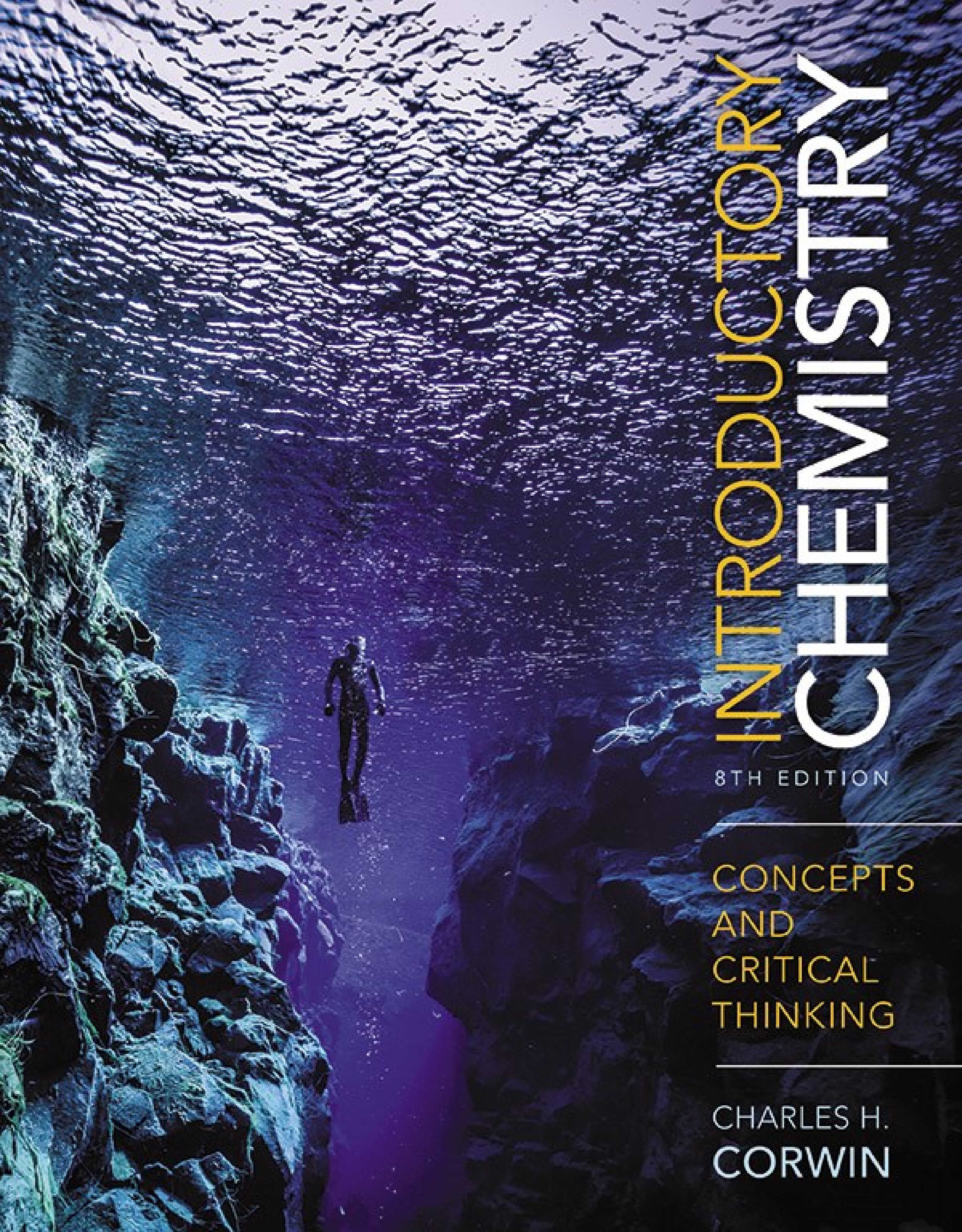

INTRODUCTORY CHEMISTRY

8TH EDITION

CONCEPTS AND CRITICAL THINKING

CHARLES H.
CORWIN

INTRODUCTORY
CHEMISTRY

INTRODUCTORY CHEMISTRY

CONCEPTS AND CRITICAL THINKING

EIGHTH EDITION

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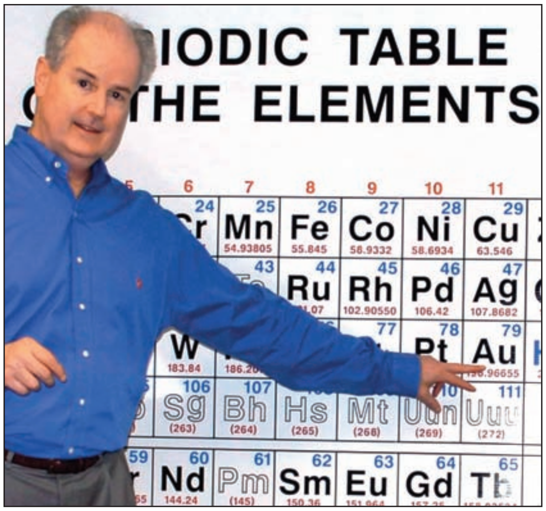
Brief Contents

About the Author



Charles H. Corwin has instructed more than 10,000 students in both lecture and laboratory of introductory and general chemistry. He acknowledges the diversity of basic chemistry students and employs a variety of classroom pedagogies based on Bloom's taxonomy, Myers-Briggs inventory, critical thinking, and algorithmic versus conceptual heuristics. His teaching strategies include learning by objective, collaborative group learning, web-based assignments, chemical demonstrations, and multimedia presentations.

The author was awarded degrees from San Jose State University, where he was a member of Tau Delta Phi honor society. He did graduate research at Stanford Research Institute on dialysis membranes, and attended the University of Akron, Institute of Polymer Science. He spent a sabbatical with the chemical education group at Purdue University studying a constructivist approach to cognitive development based on the work of Jean Piaget. Previously, he was visiting professor at Grand Valley State University on an NSF grant, and participated in a self-paced, mastery-learning study. He is currently interested in developing a hybrid online chemistry course with a laboratory component.

Professor Corwin has been recognized as instructor of the year at American River College, and has received a teaching award from Purdue University. In addition, he has been faculty mentor, department chair, academic senate representative, served on the ACS Examinations Committee and the California Chemistry Diagnostic Test Committee, has given numerous presentations to secondary schools and 2YC3, judged science projects at California state fairs, and for two decades was the examiner for the greater Sacramento region chemistry contest.

To the Student

A Personal Note I have been with you in lecture and answered your questions. I have been with you in lab and given you encouragement. Perhaps I have not spoken to you personally, but I have had countless conversations with students who are trying to juggle college life, financial aid, employment hours, relationship issues, test anxiety, and stress in general. If you are a first-time student, or a reentry student, my advice is to maintain regular student and instructor contacts, and network with others who can help guide you toward your goals.

Introductory chemistry is a subject that builds systematically and culminates in a knowledge base for the physical sciences, life sciences, health sciences, and beyond. Therefore, it is essential that you set aside time each day to study chemistry, and avoid last-minute cramming for exams. On days when you lack motivation, open the textbook to the attractive art that illustrates the topic you are studying; or, go online to the textbook website and view one of the presentations in MasteringChemistry®.

The completion of a basic chemistry course begins to open doors to a rewarding career. A rewarding career is a source of personal satisfaction that spreads to all aspects of your life and helps you avoid making poor life choices. I know you can be successful in your chemistry class. I have seen others like you with different expressions on their faces at the end of the semester. It is a look of confidence after accepting the challenge, doing their best, and knowing they are better prepared for what lies ahead.

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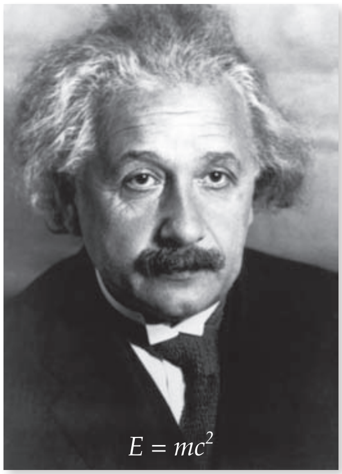

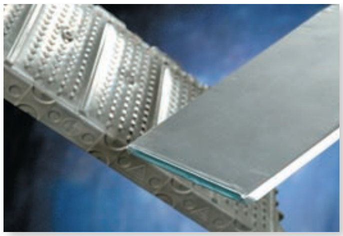

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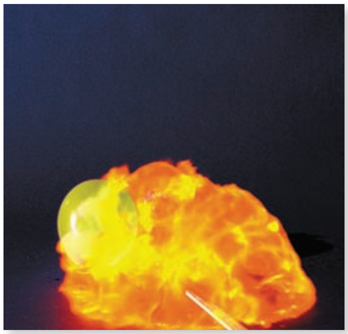

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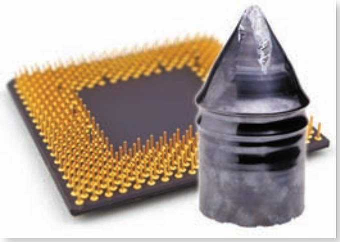

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Preface

It is my goal to hone each lecture presentation and improve my skill set each semester. Accordingly, I never leave campus without Post-It® notes for improving a presentation. I continually experiment with different topic organizations, employ multimedia and molecular models for visual concepts, perform chemical demonstrations to retain student interest, and alternate qualitative and quantitative topics to allow for topic assimilation and provide a fresh edge.

It is obvious from reviewer comments there is no consensus for the "right way" of presenting introductory chemistry. Each class situation varies with the course objectives and available facilities. In choosing an order of topic presentation in the *Eighth Edition,* I have weighed the arguments and made informed decisions.

Over a decade ago, many chemistry textbooks chose to move atomic theory and chemical bonding to later chapters. The genesis of this decision was based on Piagetian theory and the educational research that conceptual topics are a higher cognitive task, which dictates descriptive and algorithmic topics be covered early as a foundation. However, some chemistry texts prefer an "atoms first" approach, which is influenced by innate student interest in the atomic and molecular world.

Experienced instructors may have found that a mathematically "soft" approach resonates with students. In an effort to cultivate student interest, we can first discuss physical and chemical properties and assign easy tasks, such as learning names and symbols of chemical elements. The downside to this approach is that students who woefully lack basic math skills, will later experience difficulty with calculations, and may not succeed in spite of our best efforts.

The *Eighth Edition of Introductory Chemistry: Concepts and Critical Thinking* allows instructors great latitude in choosing their topic presentation. This was mandated by reviewers who argued for the early placement of certain topics, while others argued for a later placement. Suggestions for the order of topic presentation were particularly diverse in the following areas: atomic theory, chemical bonding, stoichiometry, and ionic equations.

We also asked reviewers to assess the rigor to which prerequisite science skills should be covered. The responses ranged from minimal to heavy emphasis; from relegation to an appendix to full chapter coverage. Obviously, a textbook cannot accommodate all views when there is such disparity. After weighing the alternatives, we opted for an interlude on *Prerequisite Science Skills* following Chapter 1. Students with a good preparation may find the material unnecessary; however, students with a weak preparation will find it invaluable.

A related area that has been considered at great length, is the depth of coverage for chemical calculations. In past editions, problem solving has received high marks by reviewers. In Chapter 15, *Advanced Problem Solving,* the chemical calculations are more rigorous, and the problem-solving covers a broad spectrum of techniques.

New to This Edition

In this Eighth Edition of *Introductory Chemistry: Concepts and Critical Thinking*, there have been substantial changes. In addition to a cover to cover revision, the *Eighth Edition* includes the following.

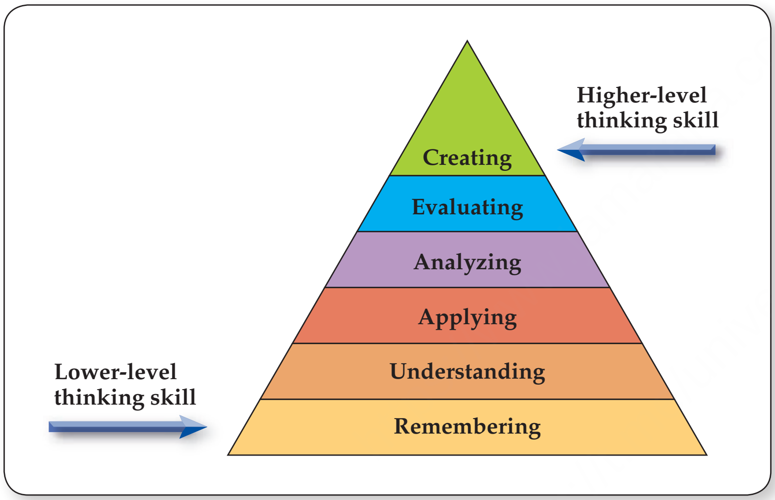
- Each **chapter opener** image is introduced with a numerical reference to an element in the periodic table. For example, Chapter 1 is introduced by the first element, hydrogen, along with a brief description.

1
H
hydrogen

Element 1: Hydrogen
Hydrogen is the simplest and most abundant element in the universe. Beginning with the Big Bang 14 billion years ago, hydrogen atoms were scattered throughout

the universe. Subsequently, two hydrogen atoms fused to produce the second element helium, which in turn fused with another hydrogen atom to give the third element, lithium. Thus, each of the elements evolved beginning with hydrogen.

- New to this edition, are numerous **Helpful Hints**, in which the author addresses the student directly, as if to give private "coaching" based on points of known difficulty and confusion.
- New to this edition, there are features termed **A Closer Look,** which offer insights and thought-provoking information designed to maintain student interest, and often provide examples of misconceptions.
- Inorganic nomenclature has become an issue in introductory general chemistry. Historically, the Stock System and Latin System of nomenclature have both been provided in textbooks. Because the Stock System is preferred by IUPAC, some textbooks have deleted the Latin System. To address this issue, the Eighth Edition of this textbook added Section 6.7, **Latin System of Nomenclature,** which provides the instructor the option to assign, or not assign, the *Latin System* based on course objectives.
- New to the Eighth Edition, are additional **Example Exercises,** each of which includes a Practice Exercise and Concept Exercise. Because introductory chemistry students learn by working examples, this is an indispensable tool.
- Approximately one third of the end-of-chapter **Exercises** have been revised, along with the **Self-Test** that follows each end-of-chapter exercise.
- New to the Eighth Edition text is a revision of the learning objectives in accord with the action verb model, set forth by Benjamin Bloom and colleagues at the University of Chicago. **Bloom's Taxonomy** describes learning that takes place at different levels, which proceeds from simple remembering to creating. The pyramid icon to the left encapsulates the Bloom learning model. The Instructor's Manual lists the learning objectives for every chapter, and assigns each to a particular level. For instructors who demand much of their students, this resource may prove invaluable for predicting the difficulty of each objective relative to its skill level.



An Integrated Learning System

More than a textbook, *Introductory Chemistry: Concepts and Critical Thinking* is a comprehensive learning system that offers print and media resources as well as an extensive website. Unlike other chemistry texts, all the materials are coherently integrated with the textbook by a single author, including the student solutions manual, laboratory manual, instructor's manual, and test item file. Moreover, the genesis of the package is based on considerable classroom experience, student feedback, instructor feedback, and multiple rounds of reviewer feedback from dozens of institutions across the country.

Students are presented with the same topics, in the same sequence, using the same vocabulary, consistently in the textbook and all the supplements. Instructors are presented with a tightly integrated package including an Instructor's Resource Manual and a 3000-question Test Bank. The media resources include an Instructor Resource Center with lecture resources, including PowerPoint presentations, animations, and interactive activities. The Mastering platform provides customizable and automatically graded assessments that motivate students to come prepared for class.

Problem Solving

An important objective of this text is to help students become effective problem solvers. This is accomplished by a walk-through discussion of each new topic, followed by an example exercise, practice exercise, concept exercise, and a problem-solving organizer at the end of each quantitative chapter.

Previous editions have received positive reviews for systematic problem solving. In as much as basic chemistry students often have weak math skills, algebra has been scrupulously avoided in the early chapters in favor of the unit analysis method of problem solving. However, since many instructors prefer an algebraic approach to the gas laws, Chapter 10 offers both an algebraic and unit analysis approach to solving gas law problems.

To assist under-prepared students with a weak math/science background, the Eighth Edition offers a brief interlude on *Prerequisite Science Skills* that instructors can assign at their discretion.

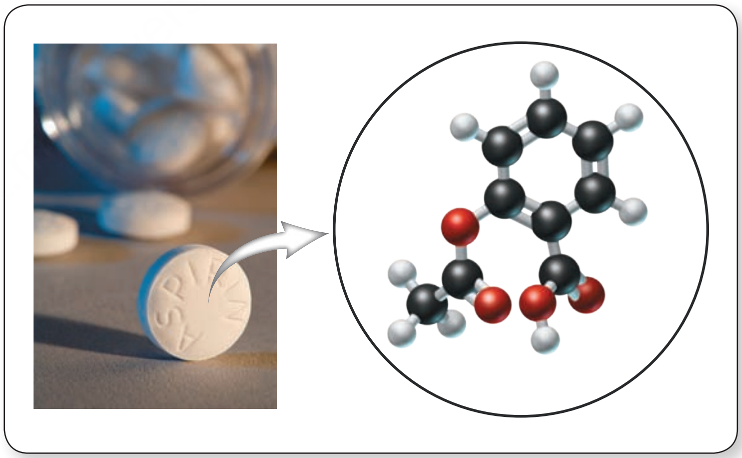
The Eighth Edition offers many end-of-chapter exercises (in matched pair format) correlated with each section of each chapter, as well as General Exercises and Challenge Exercises. Chapter 15, *Advanced Problems Solving,* discusses both the unit analysis method and algebraic method of problem solving; in addition, there is a discussion of the problem-solving techniques of visualization and drawing concept maps.

Language of Chemistry

Another important objective of this text is to help students learn the language of chemistry. To this end, each chapter has a unique matching exercise for all *key terms* that allows students to verify their definition in Appendix H. The *Study Guide and Selected Solutions Manual* has a computer-generated crossword puzzle of key terms for each chapter that provides a fun way to learn the language.

Conceptual Learning

Introductory chemistry students often require motivation to learn this subject. To this end, in the Eighth Edition we have refined the palette of colors and style of rendering to achieve a new level of sophistication in illustrations and photographs. The molecular art program has been enhanced by providing molecular structures for chemical formulas and chemical reactions. Moreover, photographs are enhanced by providing macro/micro molecular art. This visual presentation not only enhances interest in the topics, but adds an effective pedagogical tool for understanding concepts that students find difficult to grasp.



Aspirin, acetylsalicylic acid, has the molecular formula $C_9H_8O_4$.

Critical Thinking

Cognitive scientists define *"critical thinking"* as mental activity associated with three processes: applying reason, making decisions, and solving problems. In this textbook, critical thinking is especially evident in the different approaches to solving chemistry problems. Critical thinking questions are in the end-of-chapter *Self-Tests,* and misconceptions are probed in many of the vignettes entitled *A Closer Look;* for example, *Metric Labels* and *Lower Gasoline Bills.*

Flexible Chapter Sequence

For a variety of reasons, no two introductory chemistry classes present topics in the same exact order. Accordingly, the chapter sequence in *Introductory Chemistry: Concepts and Critical Thinking,* Eighth Edition is constructed in such a way so as to accommodate alternate sequences. For example, chemical bonding (Chapter 12) can immediately follow atomic theory and periodicity (Chapters 4 and 5); ionic equations can follow chemical reactions (Chapter 7); and gas stoichiometry can be deferred to the gas laws (Chapter 10).

Acknowledgments

I would like to thank the reviewers, instructors, and students who have helped make this learning package so successful. I teach chemistry at a large community college (~40,000 students) with a high enrollment for introductory chemistry. Students and colleagues continually keep me on my toes, and emails from students and instructors around the country assure broad success in the classroom.

I also want to acknowledge the "team members" for their contribution to the Eighth Edition. Scott Dustan, Senior Courseware Portfolio Manager, arrived with extensive experience, and from the outset placed a firm hand on the project. A special thank you to Brett Coker, Content Producer, who indulged me with lengthy phone conversations that ironed out issues, while keeping the project on schedule. Matt Walker, Courseware Analyst, reviewed each chapter of the previous edition and had many suggestions that were implemented. Julie Lafflin, photo researcher, provided high-quality photos within new constraints for photo usage. T. J. Mullen, accuracy reviewer, had the difficult task of ferreting out miscues introduced from a variety of sources. Nancy Moyer, proofreader, was amazingly meticulous and taught me the difference between "since" and "because" and other grammatical fine points. To each of these individuals, singly and collectively, I extend my sincere appreciation.

Reviewers of The Eighth Edition

The reviewers of this and previous editions of *Introductory Chemistry* continue to define that vague line between the simplifications that students require and the explanations that accuracy and detailed breadth of coverage demand. I thank them for working with me to make this a better resource for students learning chemistry.

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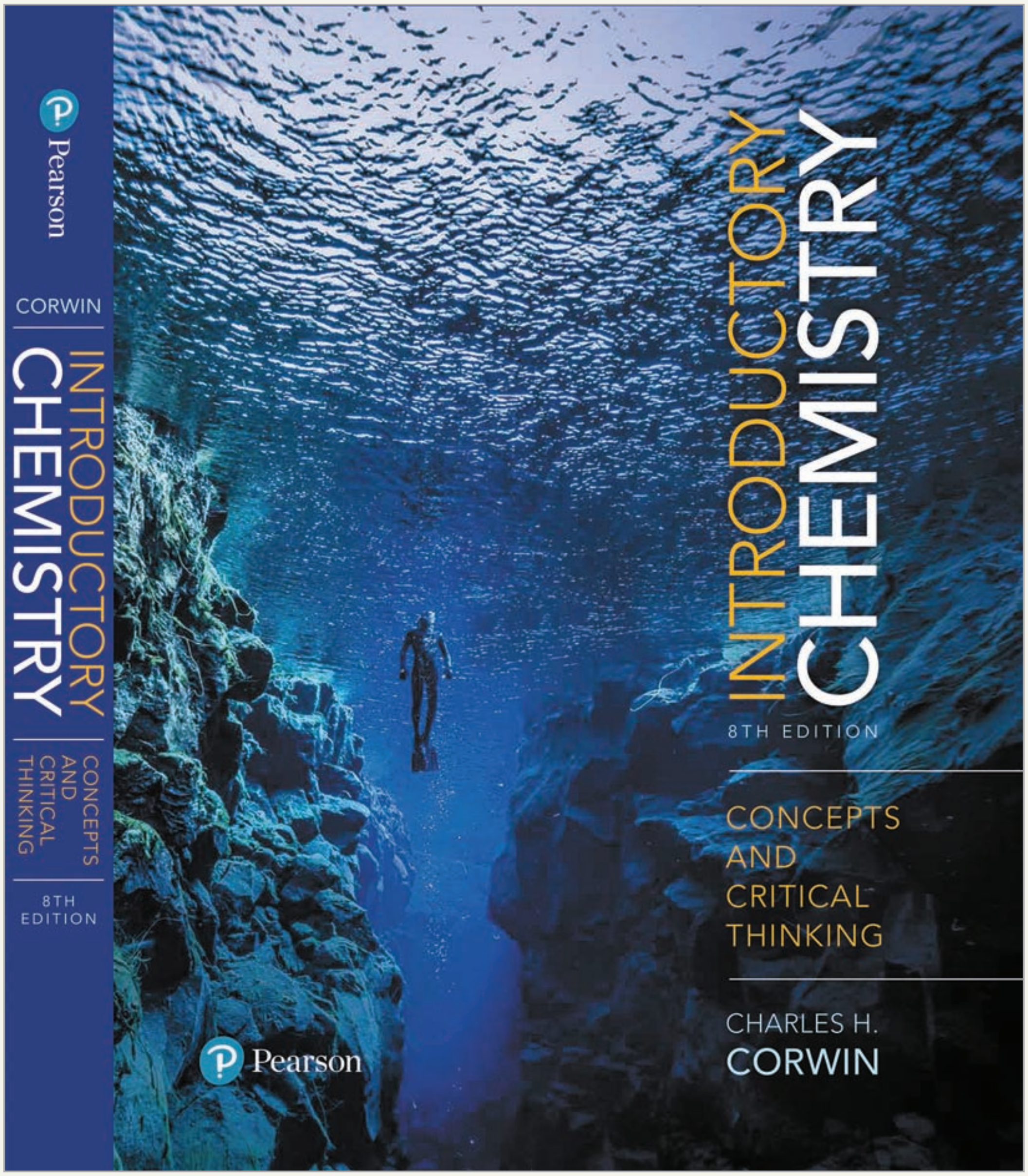
Inspire your students' curiosity around chemistry

Introductory Chemistry: Concepts and Critical Thinking will truly inspire your students to engage with the chemistry concepts presented in the text and see the world around them in a new light.

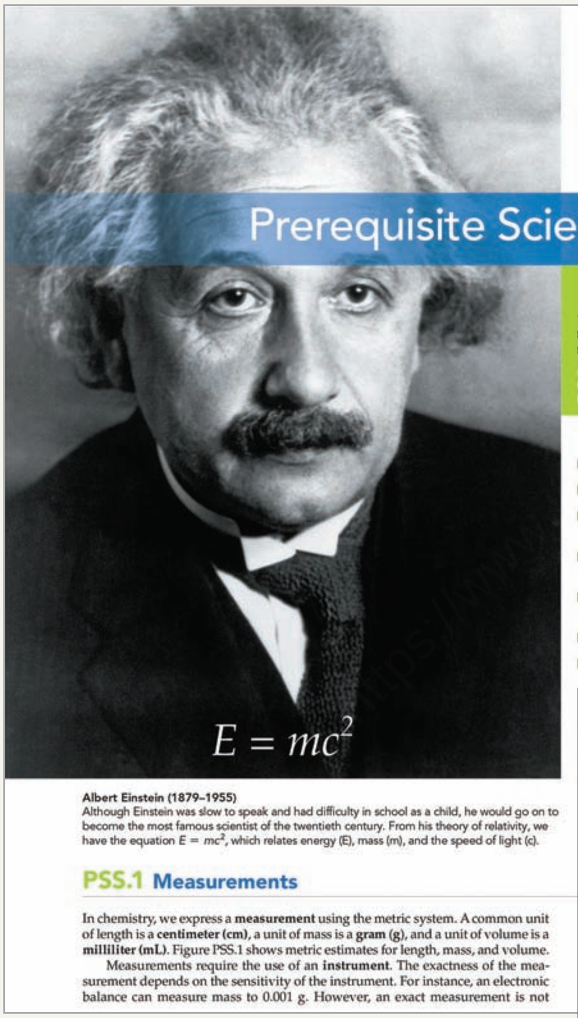
Written in an incredibly student-friendly style by experienced teacher Chuck Corwin, *Introductory Chemistry: Concepts and Critical Thinking* is a comprehensive learning system that offers print and media resources all written by the author himself, as well as an extensive MasteringChemistry course.

With this new edition, you will find two key themes woven throughout the text: addressing student needs when, where, and how they arise and engaging students through rich media and compelling applications. Turn the page to see more.



Build skills critical to student success

Problem Solving Skills



Prerequisite Science Skills

"Logic will get you from A to B. Imagination will take you everywhere."

Albert Einstein,
German-Swiss-American
Physicist (1879–1955)

$E = mc^2$

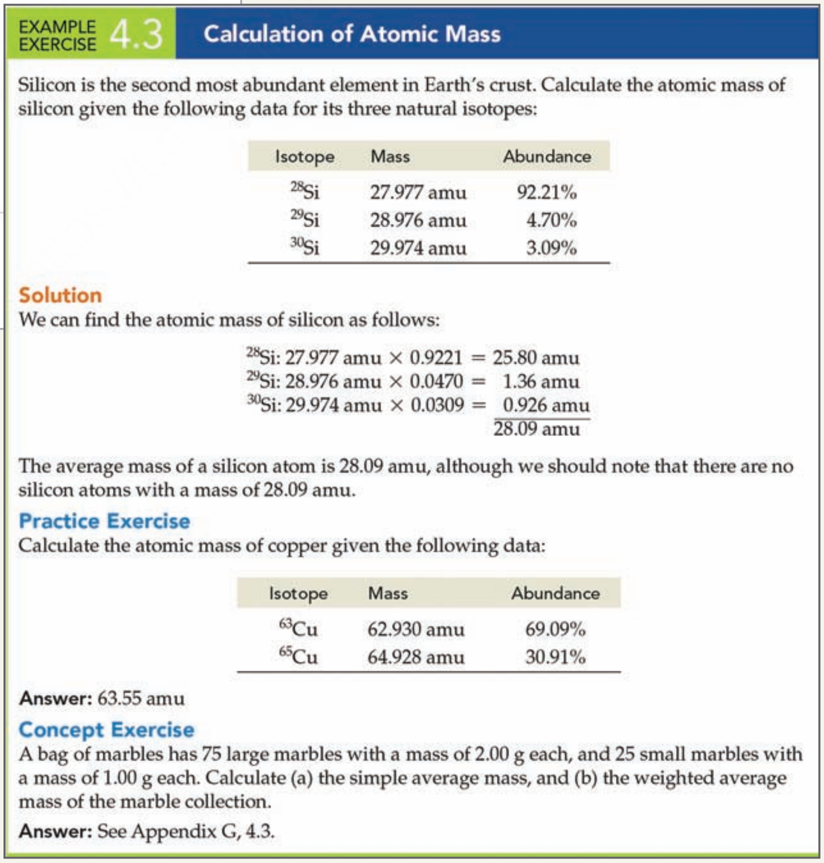
Albert Einstein (1879–1955)
Although Einstein was slow to speak and had difficulty in school as a child, he would go on to become the most famous scientist of the twentieth century. From his theory of relativity, we have the equation $E = mc^2$, which relates energy (E), mass (m), and the speed of light (c).

PSS.1 Measurements

In chemistry, we express a **measurement** using the metric system. A common unit of length is a **centimeter (cm)**, a unit of mass is a **gram (g)**, and a unit of volume is a **milliliter (mL)**. Figure PSS.1 shows metric estimates for length, mass, and volume.

Measurements require the use of an **instrument**. The exactness of the measurement depends on the sensitivity of the instrument. For instance, an electronic balance can measure mass to 0.001 g. However, an exact measurement is not

The Prerequisite Science Skills interlude helps students with weak math or science backgrounds prepare for chemical calculations. The brief interlude is an optional assignment at the discretion of the instructor and can be found between chapters one and two in the text, as well as within MasteringChemistry (many topics are available as assignable tutorials).



EXAMPLE EXERCISE 4.3 Calculation of Atomic Mass

Silicon is the second most abundant element in Earth's crust. Calculate the atomic mass of silicon given the following data for its three natural isotopes:

Isotope	Mass	Abundance
^{28}Si	27.977 amu	92.21%
^{29}Si	28.976 amu	4.70%
^{30}Si	29.974 amu	3.09%

Solution

We can find the atomic mass of silicon as follows:

$$\begin{aligned}
{}^{28}\text{Si: } 27.977\text{ amu} \times 0.9221 &= 25.80\text{ amu}\\
{}^{29}\text{Si: } 28.976\text{ amu} \times 0.0470 &= 1.36\text{ amu}\\
{}^{30}\text{Si: } 29.974\text{ amu} \times 0.0309 &= \underline{0.926\text{ amu}}\\
&\ 28.09\text{ amu}
\end{aligned}$$

The average mass of a silicon atom is 28.09 amu, although we should note that there are no silicon atoms with a mass of 28.09 amu.

Practice Exercise

Calculate the atomic mass of copper given the following data:

Isotope	Mass	Abundance
^{63}Cu	62.930 amu	69.09%
^{65}Cu	64.928 amu	30.91%

Answer: 63.55 amu

Concept Exercise

A bag of marbles has 75 large marbles with a mass of 2.00 g each, and 25 small marbles with a mass of 1.00 g each. Calculate (a) the simple average mass, and (b) the weighted average mass of the marble collection.

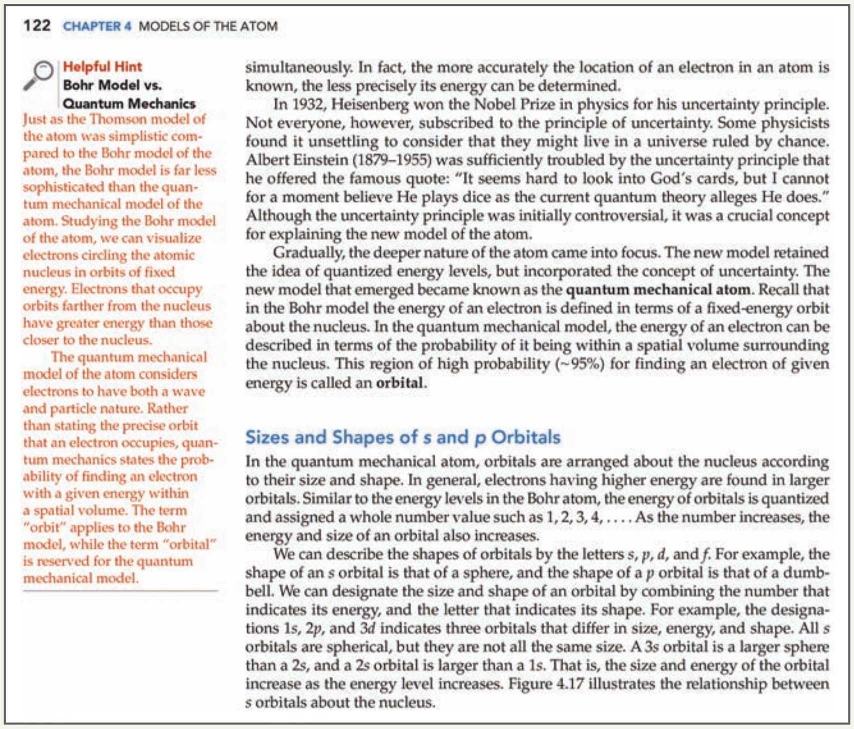
Answer: See Appendix G, 4.3.

NEW to the eighth edition are many additional **Example Exercises**, each including an accompanying **Practice Exercise and Concept Exercise**, which walk students through this important skill. This indispensable tool has been informed by the author's many years' experience of teaching thousands of introductory chemistry students.

in Future courses and careers

Critical Thinking and Conceptual Understanding

NEW to this edition, numerous **Helpful Hints** have been added; marginal notes in which the author addresses the student directly, as if giving private "coaching" based on points of known difficulty and confusion. This helps students think critically about what they are learning and gives them the confidence to move forward once they understand the concept.



122 CHAPTER 4 MODELS OF THE ATOM

Helpful Hint
Bohr Model vs. Quantum Mechanics

Just as the Thomson model of the atom was simplistic compared to the Bohr model of the atom, the Bohr model is far less sophisticated than the quantum mechanical model of the atom. Studying the Bohr model of the atom, we can visualize electrons circling the atomic nucleus in orbits of fixed energy. Electrons that occupy orbits farther from the nucleus have greater energy than those closer to the nucleus.

The quantum mechanical model of the atom considers electrons to have both a wave and particle nature. Rather than stating the precise orbit that an electron occupies, quantum mechanics states the probability of finding an electron with a given energy within a spatial volume. The term "orbit" applies to the Bohr model, while the term "orbital" is reserved for the quantum mechanical model.

simultaneously. In fact, the more accurately the location of an electron in an atom is known, the less precisely its energy can be determined.

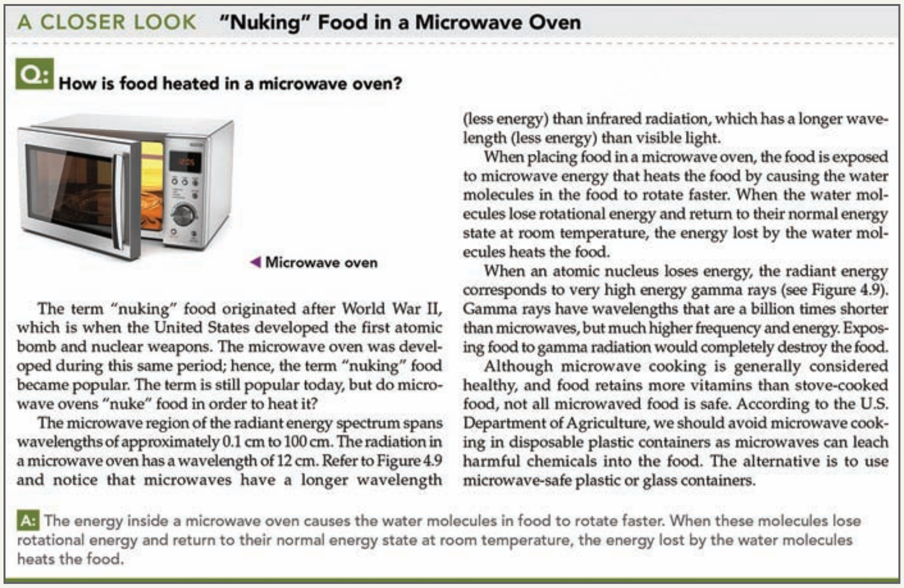
In 1932, Heisenberg won the Nobel Prize in physics for his uncertainty principle. Not everyone, however, subscribed to the principle of uncertainty. Some physicists found it unsettling to consider that they might live in a universe ruled by chance. Albert Einstein (1879–1955) was sufficiently troubled by the uncertainty principle that he offered the famous quote: "It seems hard to look into God's cards, but I cannot for a moment believe He plays dice as the current quantum theory alleges He does." Although the uncertainty principle was initially controversial, it was a crucial concept for explaining the new model of the atom.

Gradually, the deeper nature of the atom came into focus. The new model retained the idea of quantized energy levels, but incorporated the concept of uncertainty. The new model that emerged became known as the **quantum mechanical atom**. Recall that in the Bohr model the energy of an electron is defined in terms of a fixed-energy orbit about the nucleus. In the quantum mechanical model, the energy of an electron can be described in terms of the probability of it being within a spatial volume surrounding the nucleus. This region of high probability (~95%) for finding an electron of given energy is called an **orbital**.

Sizes and Shapes of *s* and *p* Orbitals

In the quantum mechanical atom, orbitals are arranged about the nucleus according to their size and shape. In general, electrons having higher energy are found in larger orbitals. Similar to the energy levels in the Bohr atom, the energy of orbitals is quantized and assigned a whole number value such as 1, 2, 3, 4, ... As the number increases, the energy and size of an orbital also increases.

We can describe the shapes of orbitals by the letters *s*, *p*, *d*, and *f*. For example, the shape of an *s* orbital is that of a sphere, and the shape of a *p* orbital is that of a dumbbell. We can designate the size and shape of an orbital by combining the number that indicates its energy, and the letter that indicates its shape. For example, the designations 1*s*, 2*p*, and 3*d* indicates three orbitals that differ in size, energy, and shape. All *s* orbitals are spherical, but they are not all the same size. A 3*s* orbital is a larger sphere than a 2*s*, and a 2*s* orbital is larger than a 1*s*. That is, the size and energy of the orbital increase as the energy level increases. Figure 4.17 illustrates the relationship between *s* orbitals about the nucleus.



A CLOSER LOOK **"Nuking" Food in a Microwave Oven**

Q: How is food heated in a microwave oven?

◀ Microwave oven

The term "nuking" food originated after World War II, which is when the United States developed the first atomic bomb and nuclear weapons. The microwave oven was developed during this same period; hence, the term "nuking" food became popular. The term is still popular today, but do microwave ovens "nuke" food in order to heat it?

The microwave region of the radiant energy spectrum spans wavelengths of approximately 0.1 cm to 100 cm. The radiation in a microwave oven has a wavelength of 12 cm. Refer to Figure 4.9 and notice that microwaves have a longer wavelength (less energy) than infrared radiation, which has a longer wavelength (less energy) than visible light.

When placing food in a microwave oven, the food is exposed to microwave energy that heats the food by causing the water molecules in the food to rotate faster. When the water molecules lose rotational energy and return to their normal energy state at room temperature, the energy lost by the water molecules heats the food.

When an atomic nucleus loses energy, the radiant energy corresponds to very high energy gamma rays (see Figure 4.9). Gamma rays have wavelengths that are a billion times shorter than microwaves, but much higher frequency and energy. Exposing food to gamma radiation would completely destroy the food.

Although microwave cooking is generally considered healthy, and food retains more vitamins than stove-cooked food, not all microwaved food is safe. According to the U.S. Department of Agriculture, we should avoid microwave cooking in disposable plastic containers as microwaves can leach harmful chemicals into the food. The alternative is to use microwave-safe plastic or glass containers.

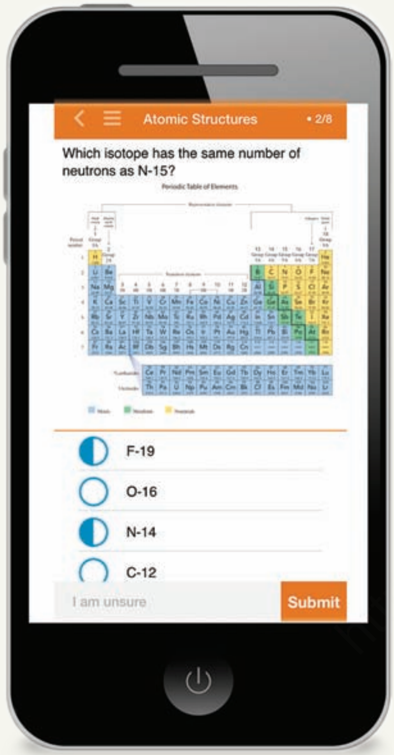
A: The energy inside a microwave oven causes the water molecules in food to rotate faster. When these molecules lose rotational energy and return to their normal energy state at room temperature, the energy lost by the water molecules heats the food.

NEW to this edition **A Closer Look** boxes offer insights and thought-provoking information on topics such as "Nuking" Food in a Microwave Oven, Nitrogen in Tires, Household Chemicals, The Ozone Hole, and Water Fluoridation, designed to connect students to the chemical world around them and, often, provide examples of common misconceptions.

Extend student learning before, during, and after class

BEFORE CLASS

Dynamic Study Modules



Atomic Structures • 2/8

Which isotope has the same number of neutrons as N-15?

◐ F-19

○ O-16

◐ N-14

○ C-12

I am unsure

Submit

NEW! 66 Dynamic Study Modules help students study effectively on their own by continuously assessing their activity and performance in real time.

- Students complete a set of questions with a unique answer format that also asks them to indicate their confidence level. Questions repeat until the student can answer them all correctly and confidently. These are available as graded assignments prior to class and are accessible on smartphones, tablets, and computers.
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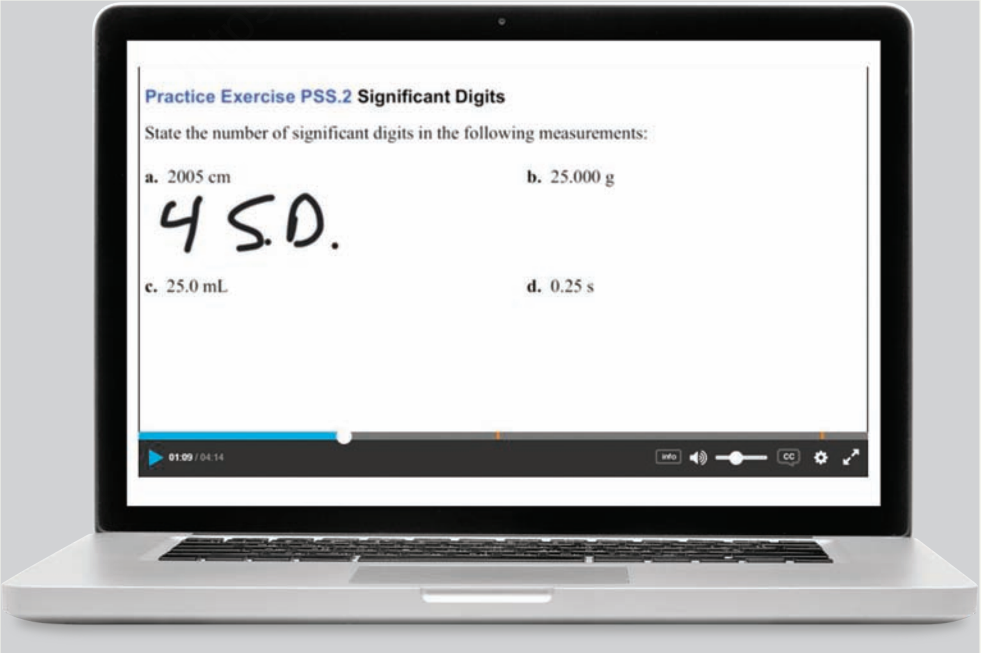
NEW! Prerequisite Science Skill Tutorials

Bringing the aforementioned interlude to life, Video Tutor Solutions cover topics of measurements, significant digits, rounding off nonsignificant digits, adding and subtracting measurements, multiplying and dividing measurements, exponential numbers, and scientific notation from the popular Prerequisite Science Skills section of the book.



Practice Exercise PSS.2 Significant Digits

State the number of significant digits in the following measurements:

a. 2005 cm

4 S.D.

b. 25.000 g

c. 25.0 mL

d. 0.25 s

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DURING CLASS

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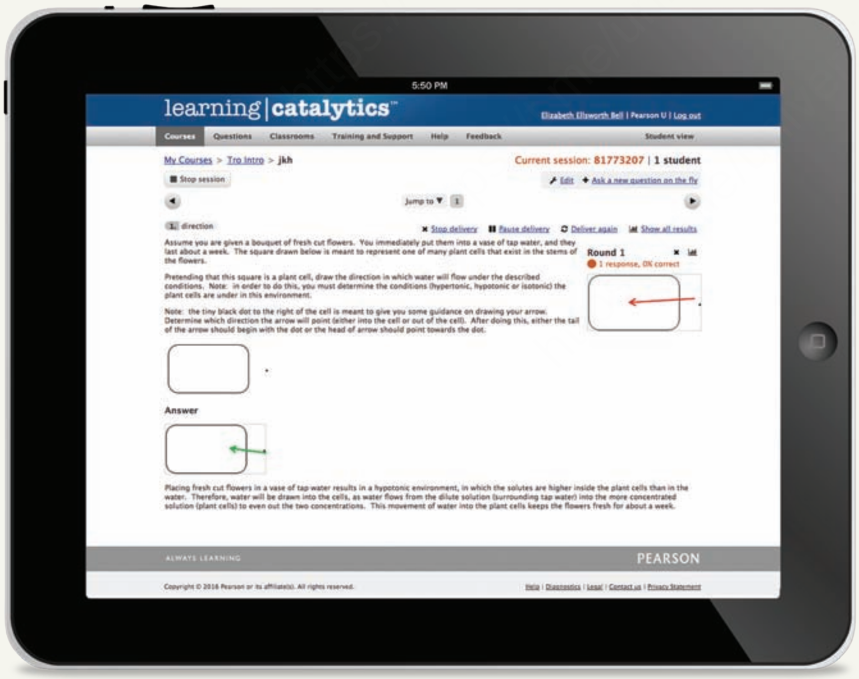

Assume you are given a bouquet of fresh cut flowers. You immediately put them into a vase of tap water, and they last about a week. The square drawn below is meant to represent one of many plant cells that exist in the stems of the flowers.

Pretending that this square is a plant cell, draw the direction in which water will flow under the described conditions. Note: in order to do this, you must determine the conditions (hypertonic, hypotonic or isotonic) the plant cells are under in this environment.

Note: the tiny black dot to the right of the cell is meant to give you some guidance on drawing your arrow. Determine which direction the arrow will point (either into the cell or out of the cell). After doing this, either the tail of the arrow should begin with the dot or the head of arrow should point towards the dot.

Placing fresh cut flowers in a vase of tap water results in a hypotonic environment, in which the solutes are higher inside the plant cells than in the water. Therefore, water will be drawn into the cells, as water flows from the dilute solution (surrounding tap water) into the more concentrated solution (plant cells) to even out the two concentrations. This movement of water into the plant cells keeps the flowers fresh for about a week.

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Close the learning loop with after-class assignments

AFTER CLASS

± Scientific Notation

Learning Goal:

To understand and be able to use scientific notation.

Scientists often make measurements that are very large or very small. Scientific notation is a way to write numbers that describe very large or very small things. There are two parts to the number in the scientific notation: the coefficient and the power of 10. (Figure 1)

Part A

Type the number 1340000 in scientific notation.

Express your answer using scientific notation.

1340000 = $1.3x10^6$

Incorrect; Try Again; 5 attempts remaining

The correct answer does not depend on: x10.

Part B

Type the number 0.000354 in scientific notation.

Express your answer using scientific notation.

0.000354 =

Hint 1. Determine the coefficient

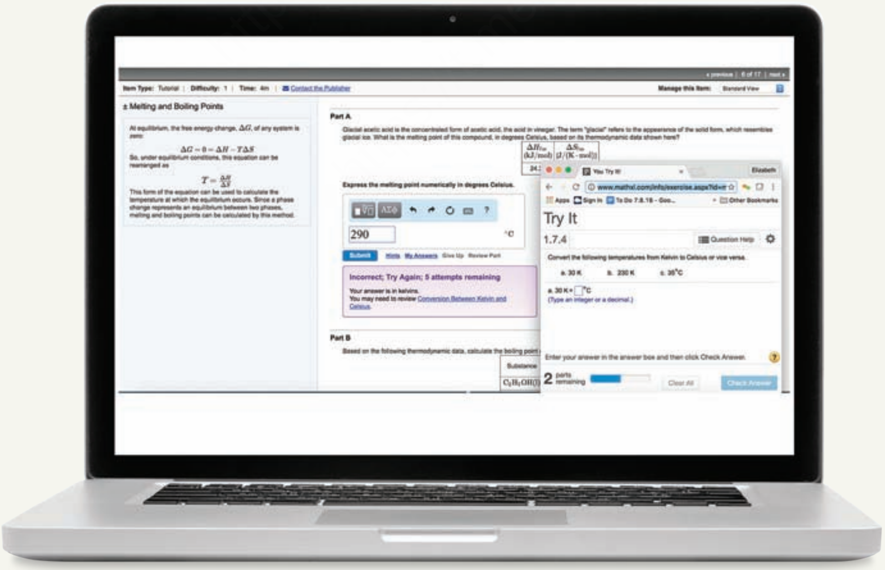
When written in scientific notation, what is the coefficient for the number 1340000?

Express the coefficient numerically.

coefficient =

Hint 2. Determine the power of 10 (click to open)

Assignable, in-depth tutorials guide students through the toughest topics in chemistry with individualized coaching. These self-paced tutorials aid students with hints and feedback specific to individual misconceptions. Tutorials respond, at any step, to a wide variety of typical wrong answers that students enter, and offer hints, allowing students to choose specific help when they need it.



± Melting and Boiling Points

At equilibrium, the free energy change, ΔG, of any system is zero:

$$\Delta G = 0 = \Delta H - T\Delta S$$

So, under equilibrium conditions, this equation can be rearranged as

$$T = \frac{\Delta H}{\Delta S}$$

This form of the equation can be used to calculate the temperature at which the equilibrium occurs. Since a phase change represents an equilibrium between two phases, melting and boiling points can be calculated by this method.

Part A

Glacial acetic acid is the concentrated form of acetic acid, the acid in vinegar. The term "glacial" refers to the appearance of the solid form, which resembles glacial ice. What is the melting point of this compound, in degrees Celsius, based on its thermodynamic data shown here?

Express the melting point numerically in degrees Celsius.

290 °C

Incorrect; Try Again; 5 attempts remaining

Your answer is in kelvins.
You may need to review Conversion Between Kelvin and Celsius.

Part B

Based on the following thermodynamic data, calculate the boiling point

Try It

1.7.4

Convert the following temperatures from Kelvin to Celsius or vice versa.

a. 30 K b. 230 K c. 35°C

a. 30 K = ☐ °C
(Type an integer or a decimal.)

Enter your answer in the answer box and then click Check Answer.

2 parts remaining

Math Remediation links in selected tutorials launch algorithmically generated math exercises that give students unlimited practice on prerequisite skills, freeing up class and office-hour time to focus on the chemistry. Exercises include guided solutions, sample problems, and learning aids for extra help, and offer helpful feedback when students enter incorrect answers.

Etext 2.0

CHAPTER 4

Models of the Atom

"Anyone who says that they can contemplate quantum mechanics without becoming dizzy has not understood the concept in the least."

Niels Bohr, Danish Physicist (1885–1962)

Element 4: Beryllium
Beryllium is a lightweight metal used in the aerospace industry for the manufacture of aircraft, missiles, and satellites. Beryllium-copper alloys are used to make tools that are uniquely spark-proof. In addition, beryllium metal is a good reflector of neutrons; thus, beryllium finds application in nuclear reactors.

4.1 Dalton Model of the Atom

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- **Seamlessly integrated videos and activities** allow students to watch and practice key concepts within the eText learning experience.
- **Accessible** (screen-reader ready).
- **Configurable reading settings**, including resizable type and night reading mode.

Instructor and student supplements

Instructor Resource Center

This lecture resource contains selected art from the textbook, three pre-built PowerPoint® presentations, animations, interactive activities, the Instructor's Manual lecture outlines in Word® format, and the Test Bank in Word® format.

Instructor Manual and Test Bank

Written by Charles H. Corwin, the Instructor Manual provides complete solutions to all even-numbered exercises in the textbook, lists learning objectives for course planning, recommends media resources, and suggests chemical demonstrations. It also includes the Test Bank with over 3000 class-tested questions that have undergone item analysis and address each topic in the textbook.

Student Study Guide & Selected Solutions Manual

Also written by the author, this printed study aid includes diagnostic test questions for each topic covered in the text, crossword puzzles using key terms, and complete solutions to all odd-numbered exercises.

Resources in Print and Online

Name of Supplement	Available in Print	Available Online	Instructor or Student Supplement	Description
MasteringChemistry®— for Introductory Chemistry: Concepts and Critical Thinking, 8/e 0134555015 / 9780134555010		✓	Supplement for Instructors and Students	The Mastering platform is the most effective and widely used online homework, tutorial, and assessment system for the sciences. It delivers self-paced tutorials that focus on your course objectives, provide individualized coaching, and respond to each student's progress. The Mastering system helps instructors maximize class time with easy-to-assign, customizable, and automatically graded assessments that motivate students to learn outside of class and arrive prepared for lecture or lab.
Study Guide & Selected Solutions Manual for Introductory Chemistry: Concepts and Critical Thinking, 8/e 0134580281 / 9780134580289	✓		Supplement for Students	Written by the author, Charles H. Corwin, this study aid includes diagnostic test questions for each topic covered in the text, crossword puzzles using key terms, and complete solutions to all odd-numbered exercises.
Instructor Manual and Test Bank for Introductory Chemistry: Concepts and Critical Thinking, 8/e 0134580265 / 9780134580265		✓	Supplement for Instructors	The manual features a list of all chapter learning objectives and complete solutions to the even-numbered chapter exercises. This has been updated to reflect the revisions in this text and contains questions in a bank of more than 4,000 multiple-choice questions.
Instructor Resource Materials for Introductory Chemistry: Concepts and Critical Thinking, 8/e 0134580257 / 9780134580258		✓	Supplement for Instructors	This lecture resource contains all art and images from the textbook, three pre built PowerPoint® presentations, animations, interactive activities, the Instructor's Manual in PDF format, and the test item file in Word® format.
Laboratory Manual for Introductory Chemistry: Concepts and Critical Thinking, 6/e 0321750942/9780321750945	✓		Supplement for Laboratory	Emphasizing environmental considerations, Corwin's acclaimed lab manual offers a proven format of a prelaboratory assignment, a stepwise procedure, and a postlaboratory assignment. More than 300,000 students to date in introductory chemistry, preparatory chemistry, and allied health chemistry have used these "bulletproof" experiments successfully. The *Sixth Edition* features environmental icons to alert students to recycle chemical waste, updated prelabs and postlabs, new experimental procedures, a new experiment (Experiment 25), and a new appendix on how to keep a laboratory notebook. Corwin's lab manual can be packaged with any Pearson intro prep chemistry book.

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CHAPTER 1

Introduction to Chemistry

> *"All our dreams can come true—if we have the courage to pursue them."*
>
> **Walt Disney, Disneyland Creator (1901–1966)**

1
H
hydrogen

Element 1: Hydrogen
Hydrogen is the simplest and most abundant element in the universe. Beginning with the Big Bang 14 billion years ago, hydrogen atoms were scattered throughout the universe. Subsequently, two hydrogen atoms fused to produce the second element, helium, which in turn fused with another hydrogen atom to give the third element, lithium. Thus, each of the elements evolved beginning with hydrogen.

In the United States, Canada, and other developed countries, we enjoy a standard of living that could not have been imagined a century ago. Owing to the evolution of science and technology, we have abundant harvests; live in comfortable, climate-controlled buildings; and travel the world via automobiles and airplanes. We also have extended life spans free of many diseases that previously ravaged humanity.

The development of technology has provided machinery and equipment to perform tedious tasks, which gives us time for more interesting activities. The arrival of the computer chip has given us electronic appliances that afford ready convenience and dazzling entertainment. We can select from a multitude of audio and video resources that offer remarkable sound and brilliant color. We can access these audio and video resources from the Internet, satellite, a compact disc, or a smartphone that can communicate wirelessly while surfing the Internet (Figure 1.1).

Our present standard of living requires scientists and technicians with educational training in chemistry. The health sciences as well as the life sciences, physical sciences, and earth sciences demand an understanding of chemical principles. In fact, chemistry is sometimes referred to as the central science because it stands at the crossroads of biology, physics, geology, and medicine. Just as personal computers and smartphones are indispensable in our everyday activities, chemistry plays an essential role in our daily lives.

▲ **Figure 1.1 Smartphone**
A typical smartphone has Internet access and cloud connectivity.

1.1 Evolution of Chemistry

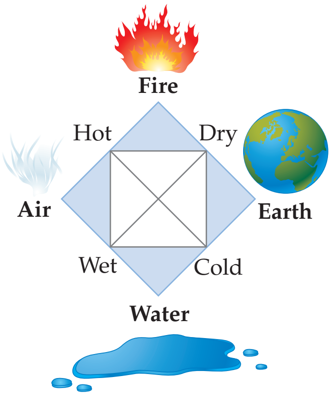
LEARNING OBJECTIVES

- Describe the early practice of chemistry.
- Identify the three steps in the scientific method.

The earliest concept of science began with the ancient Chinese, Egyptian, and Greek civilizations. The Chinese believed that the universe was created from the interaction of two forces. Yin, the feminine force, was manifested in darkness, cold, and wetness. Yang, the masculine force, was manifested in light, heat, and dryness. When the yin and yang forces interacted, they brought the earthly world into existence and were responsible for everything in nature.

As early as 600 B.C., the Greeks began to speculate that the universe was composed of a single element. Thales, the founder of Greek science, mathematics, and philosophy, suggested that water was the single element. He claimed that Earth was a dense, flat disc floating in a universe of water. He also believed that air and space were less dense forms of water.

A few years later, another Greek philosopher proposed that air was the basic element. This theory was followed by the proposals that fire, and later earth, was the basic element. About 450 B.C., the Greek philosopher Empedocles observed that when wood burned, smoke was released (air), followed by a flame (fire). He also noticed that a cool surface held over a fire collected moisture (water) and that the only remains were ashes (earth). Empedocles interpreted his observations as evidence for air, fire, water, and earth as basic elements. The conclusion was logical based on his observations and he further speculated other substances were examples of these four elements combined in varying proportions, as illustrated in Figure 1.2.



▲ **Figure 1.2 The Four Greek Elements** The four elements proposed by the Greeks: air, earth, fire, and water. Notice the properties hot, cold, wet, and dry associated with each element.

In about 350 B.C., Aristotle adopted the idea that air, earth, fire, and water were basic elements. In addition, he added a fifth element, ether, that he believed filled all space. Aristotle's influence was so great that his opinions dominated other Greek philosophers and shaped our understanding of nature for nearly 2,000 years.

The Scientific Method

In 1661, the English scientist Robert Boyle (1627–1691) published *The Sceptical Chymist*. In his classic book, Boyle stated that theoretical speculation was worthless unless it was supported by experimental evidence. This principle led to the development of the scientific method, which marked a turning point in scientific inquiry and the beginning of modern science.

Science can be defined as the methodical exploration of nature followed by a logical explanation of the observations. The practice of science entails planning an investigation, carefully recording observations, gathering data, and analyzing the results. In an **experiment**, scientists explore nature according to a planned strategy and make observations under controlled conditions.

The **scientific method** is a systematic investigation of nature and requires proposing an explanation for the results of an experiment in the form of a general principle. The initial, tentative proposal of a scientific principle is called a **hypothesis**.

After further experimentation, the initial hypothesis may be rejected, modified, or elevated to the status of a scientific principle. However, for a hypothesis to become a scientific principle, many additional experiments must support and verify the original proposal. Only after there is sufficient evidence does a hypothesis rise to the level of a **scientific theory**. We can summarize the three steps in the scientific method as follows:

▲ **Robert Boyle** Boyle performed experiments with the vacuum pump, and wrote the classic textbook, *The Skeptical Chymist*, which laid the foundation for the scientific method.

Applying the Scientific Method

Step 1: Perform a planned experiment, make observations, and record data.

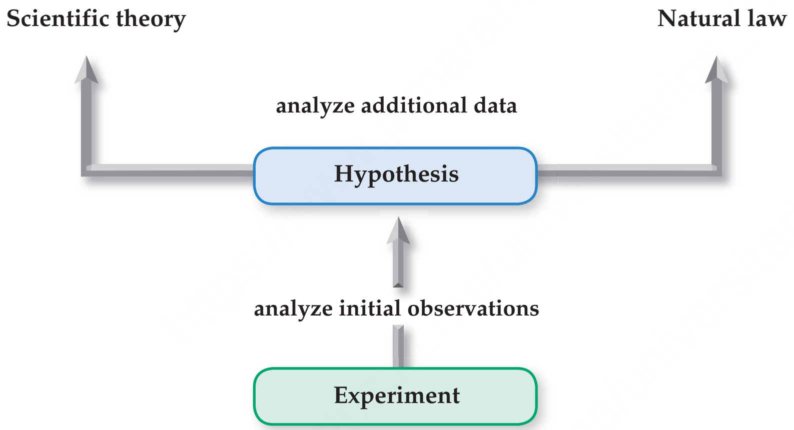
Step 2: Analyze the data and propose a tentative hypothesis to explain the experimental observations.

Step 3: Conduct additional experiments to test the hypothesis. If the evidence supports the initial proposal, the hypothesis may become a scientific theory.

We should note that scientists exercise caution before accepting a theory. Experience has shown that nature reveals its secrets slowly and only after considerable probing. A scientific theory is not accepted until rigorous testing has established that the hypothesis is a valid interpretation of the evidence. For example, in 1803, John Dalton (1766–1844) proposed that all matter was composed of small, indivisible particles called atoms. However, it took nearly 100 years of gathering additional evidence before his proposal was universally accepted and elevated to the status of the atomic theory.

Although the terms *theory* and *law* are related, there is a distinction between the two terms. A theory is a model that explains the behavior of nature. A **natural law** does not explain behavior, but rather states a measurable relationship. To illustrate, it is a *law* that heat flows from a hotter object to a cooler one because we can measure experimentally the change in temperature if we drop an ice cube into water. It is a *theory* that the transfer of heat is due to changes in the motion of molecules in the ice and water.

We can distinguish between a theory and a law by simply asking the question, "Is the proposal measurable?" If the answer is yes, the statement is a law; otherwise, the statement is a theory. Figure 1.3 summarizes the relationship of a hypothesis, a scientific theory, and a natural law.



◀ **Figure 1.3 The Scientific Method** The initial observations from an experiment are analyzed and formulated into a hypothesis. Next, additional data is collected from experiments conducted under various conditions and the data is analyzed. If the additional data supports the initial proposal, the hypothesis may be elevated to a scientific theory or a natural law.

1.2 Modern Chemistry

LEARNING OBJECTIVE

- Describe the modern practice of chemistry.

In the A.D. eighth century, the Arabs introduced the pseudoscience of **alchemy**. Alchemists conducted simple experiments and believed in the existence of a magic potion that had miraculous healing powers and could transmute lead into gold. Although alchemy did not withstand the test of time, it preceded the planned, systematic, scientific experiments that are the cornerstone of modern chemical research.

In the late eighteenth century, the French chemist Antoine Lavoisier (1743–1794) organized chemistry and wrote two important textbooks. Lavoisier also built a magnificent laboratory and invited scientists from around the world to view it; his many visitors included Benjamin Franklin and Thomas Jefferson. Lavoisier was a prolific experimenter and published his work in several languages. For his numerous contributions, he is considered the founder of modern chemistry.

Today, we define **chemistry** as the science that studies the composition of matter and its properties. Chemists have accumulated so much information during the past two centuries that we now divide the subject into several branches or specialties. The branch of chemistry that studies substances containing the element carbon is called **organic chemistry**. The study of all other substances, those that do not contain the element carbon, is called **inorganic chemistry**.

▲ **Antoine Lavoisier** In addition to writing two textbooks on chemistry, Lavoisier established a magnificent eighteenth-century laboratory that attracted scientists from around the world.

The branch of chemistry that studies substances derived from plants and animals is **biochemistry**. Another branch, analytical chemistry, includes qualitative analysis (what substances are present in a sample) and quantitative analysis (how much of each substance is present). Physical chemistry is a specialty that proposes theoretical and mathematical explanations for chemical behavior. Recently, environmental chemistry has become an important specialty that focuses on the safe disposal of chemical waste. **Green chemistry**, also termed *sustainable chemistry*, refers to the design of chemical products and processes that reduce or eliminate hazardous substances.

Chemistry plays a meaningful role in medicine, especially in the dispensing of pharmaceutical prescriptions. Chemists help ensure agricultural harvests by formulating fertilizers and pesticides. Chemistry is indispensable to many industries including the manufacture of automobiles, electronic components, aluminum, steel, paper, and plastics. One of the largest industries is the petrochemical industry. Petrochemicals are chemicals derived from petroleum and natural gas. They can be used to manufacture a wide assortment of consumer products including paints, plastics, rubber, textiles, dyes, and detergents.

In every chapter you will have example exercises that put learning into action. Each example exercise poses a question and shows the solution. There is also a practice exercise and a concept exercise to further your understanding. Example Exercise 1.1 illustrates a question, practice exercise, and concept exercise.

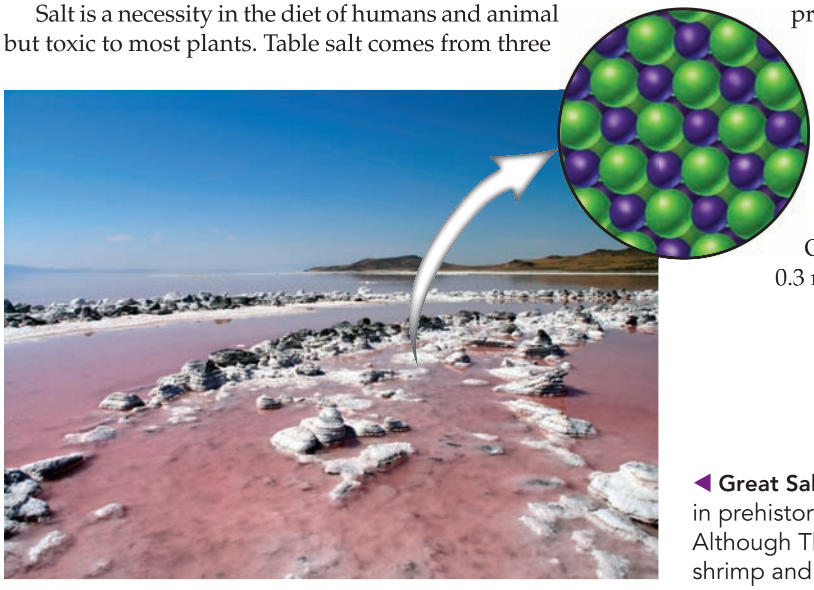
A CLOSER LOOK Watch Your Salt Intake!

Q: How much salt is too much?

The many uses for salt predate modern history. In the ancient world, towns and settlements were near salt reservoirs, because salt was a dietary necessity and a food preservative. Hippocrates, the Greek founder of medicine, urged physicians to soak their patients in salt water as treatment for various ailments. Because most natural salt was not suitable for ingestion, pure salt was a rare and valuable commodity. So-called "salt roads" were used by caravans of camels to transport salt long distances in trade for gold and textiles.

Salt is a necessity in the diet of humans and animal but toxic to most plants. Table salt comes from three sources: salt mining, solution mining, and evaporation of salt water. The United States and Canada have extensive deposits of salt, and the Great Salt Lake in Utah is so concentrated and dense that humans can easily float in the salt water.

Table salt (sodium chloride) is necessary in the human diet; however, too much sodium has been linked to high blood pressure that can lead to diabetes and heart problems. The recommended daily allowance (RDA) of table salt is a teaspoon, which contains approximately 2300 mg of sodium. Surprisingly, most salt in the human diet does not come from table salt, but from processed foods, for example, ketchup, pickles, and snack foods. Table salt contains iodine in the form of potassium iodide. Humans require iodine in small quantities for proper function of the thyroid gland. The hormone thyroxine, which contains iodine, is largely responsible for maintaining our metabolic rate. One teaspoon of iodized table salt contains about 0.3 milligram of iodine, which is twice the RDA.



◀ **Great Salt Lake** The Great Salt Lake in Utah was created in prehistoric times and contains more salt than seawater. Although The Great Salt Lake provides habitat for brine shrimp and aquatic birds, it is called "America's Dead Sea."

A: The RDA for sodium is 2300 mg; about 1 teaspoon of table salt.

EXAMPLE EXERCISE 1.1 Introduction to Chemistry

What is the difference between ancient chemistry and modern chemistry?

Solution

The principal difference is that modern chemistry is founded on the scientific method. Ancient chemistry was based on speculation, whereas modern chemistry is based on planned experiments.

Practice Exercise

What question can we ask to distinguish a scientific theory from a natural law?

Answer: We can distinguish a theory from a law by asking the question, "Is the proposed statement measurable?" If we take measurements and verify a relationship by a mathematical equation, the statement is a law; if not, it is a theory.

Concept Exercise

Alchemists believed in a magic potion that had what miraculous power?

Answer: See Appendix G, 1.1.

CHEMISTRY CONNECTION A Student Success Story

Q: Which common inexpensive metal was more valuable than gold in the nineteenth century?

In 1885, Charles Martin Hall (1863–1914) was a 22-year-old student at Oberlin College in Ohio. One day his chemistry teacher told the class that anyone who could discover an inexpensive way to produce aluminum metal would become rich and benefit humanity. At the time, aluminum was a rare and expensive metal. In fact, Napoleon III, a nephew of Napoleon Bonaparte, entertained his most honored guests with utensils made from aluminum while other guests dined with utensils of silver and gold. Although aluminum is the most abundant metal in Earth's crust, it is not found free in nature; it is usually found combined with oxygen in minerals such as bauxite.

After graduation, Charles Hall set up a laboratory in a woodshed behind his father's church in Oberlin, Ohio. Using homemade batteries, he devised a simple method for producing aluminum by passing electricity through a molten mixture of minerals. After only 8 months of experimenting, he invented a successful method for reducing an aluminum mineral to aluminum metal. In February 1886, Charles Hall walked into his former teacher's office with a handful of metallic aluminum globules.

Just as his chemistry teacher had predicted, within a short period of time, Hall became rich and famous. In 1911, he received the Perkin Medal for achievement in chemistry, and in his will, he donated $5 million to Oberlin College. He also helped to establish the Aluminum Company of America (ALCOA), and the process for making aluminum metal gave rise to a huge industry. Aluminum is now second only to steel as a construction metal.

It is an interesting coincidence that the French chemist Paul Hèroult, without knowledge of Hall's work, made a similar discovery at the same time. Thus, the industrial method for obtaining aluminum metal is referred to as the Hall–Hèroult process. In 1886, owing to the discovery of this process, the price of aluminum plummeted dramatically. Today, the price of aluminum is less than $1 a pound.

▲ **Aluminum Globules** The notebook of Charles Hall along with globules of aluminum.

A: Before 1886, aluminum was an extremely rare and expensive metal.

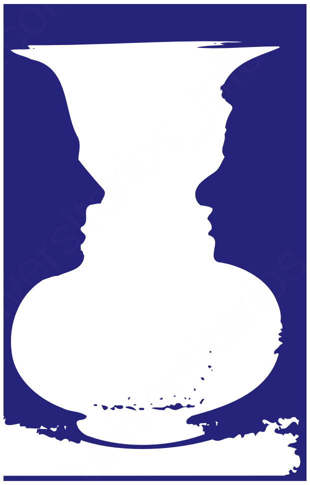
LEARNING OBJECTIVE

- Conclude that chemistry is very relevant in our daily life.

1.3 Learning Chemistry

In a survey published by the American Chemical Society, entering college students were asked to express their attitudes about science courses. The students rated chemistry as the most relevant science course, and as highly relevant to their daily lives. Unfortunately, 83% of the students thought chemistry was a difficult subject. In view of the results of the student survey, perhaps we should take a moment to consider perceptions in general, and attitudes about chemistry in particular.

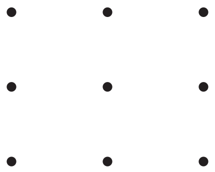
You are probably familiar with the expression that some people see a glass of water as half full, while others see the same glass as half empty. This expression implies that different people can respond to the same experience with optimism or pessimism. Moreover, experimental psychologists have found that they can use abstract visual images to discover underlying attitudes regarding a particular perception. A practical lesson involving two perceptions obtained from the same image is revealed by the following picture.



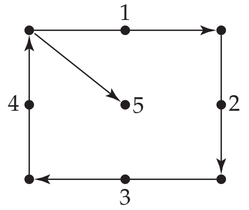
What do you see? Some students see a white vase on a dark background; others see two dark profiles facing each other. After a short period of time, one image switches to the other. If you concentrate, can you view only one of the images? Can you choose to switch the images back and forth? This exercise is an example of our brains registering dual perceptions from the same image.

Your experience of learning chemistry may be somewhat like the preceding exercise that tests your perspective. Sometimes your perception may be that chemistry is challenging, whereas a short time later your attitude may be that chemistry is easy and fun.

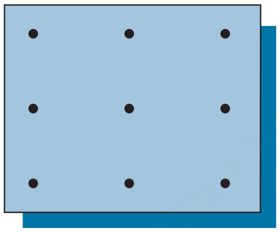
Perception is often affected by unconscious assumptions. Let's consider a type of problem that is slightly different from the vase perception. In the following problem try to connect each of the nine dots using only *four* straight, continuous lines.



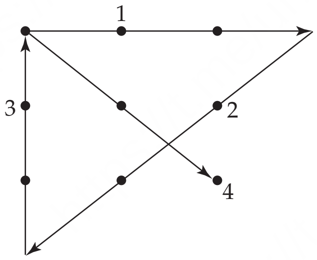
We can begin to solve the problem by experimenting. For example, let's start with the upper-left dot and draw a line to the upper-right dot. We can continue to draw straight lines as follows:



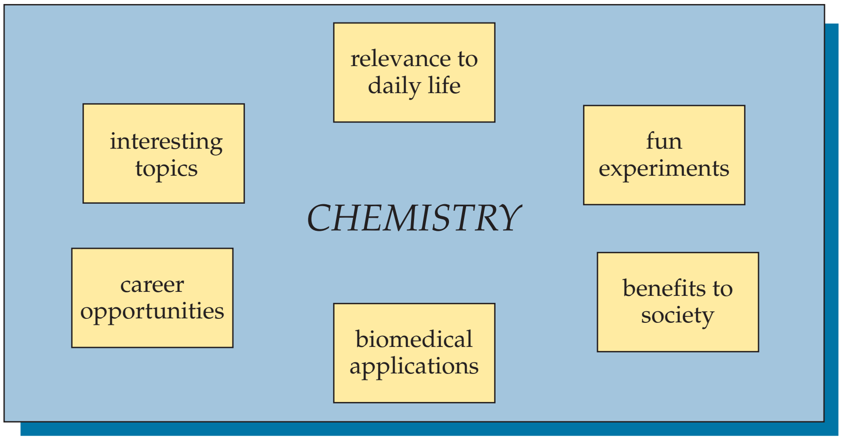
Notice that we connected the nine dots, but that it was necessary to use *five* straight, continuous lines. If we start with a different dot, we find that *five* lines are required no matter where we start. Perhaps we are bringing an underlying assumption to the nine-dot problem. That is, we may be unconsciously framing the nine dots, thus limiting the length of the four straight lines.



What will happen if we start with the upper-left dot and draw a line through the upper-right dot? If we continue, we can complete the problem with *four* straight, continuous lines as follows:



The "secret" to solving this nine-dot problem is to recognize that we may be unconsciously confining our thinking and making it impossible to solve. Similarly, we should not confine our concept of chemistry to a preconceived attitude that learning chemistry will be difficult. Or better yet, we should choose positive associations for our concept of chemistry. The following diagram illustrates a few of the positive associations for chemistry.



Helpful Hint
Studying Chemistry

It is essential to study on a regular basis. On days you lack motivation, look at the attractive art that illustrates each topic. Next, study the Example Exercises. Finally, try the odd-numbered end-of-chapter Exercises, and check your answers in Appendix I.

To prepare for a test, take the end-of-chapter Self-Test, and check your answers in Appendix J.

EXAMPLE EXERCISE 1.2 Learning Chemistry

Is it possible to have different perceptions of a single visual image?

Solution

Yes, a single image can suggest two different observations. That is, an observer may view a half glass of water as "half full" or "half empty."

Practice Exercise

Why would a student at first find a chemistry problem difficult to solve when later the solution seems easy and obvious?

Answer: If the student has an unconscious assumption that chemistry is a difficult subject, the student may assume a chemistry problem is difficult to solve.

Concept Exercise

What can a student do to eliminate an unconscious assumption regarding chemistry being a difficult subject.

Answer: See Appendix G, 1.2.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
1.1 Evolution of Chemistry This chapter traces the development of chemistry from a historical point of view. Beginning in the period 600–350 B.C., the early Greeks used reason and thoughtful mental exercises to understand the laws of nature. Although they often arrived at conclusions based on speculation, they did unveil some of nature's secrets and had a profound influence on Western civilization that lasted for 20 centuries. The term **science** implies a rigorous, systematic investigation of nature. Moreover, a scientist must accumulate significant evidence before attempting to explain the results. In the seventeenth century, Robert Boyle founded the **scientific method**, and laboratory experimentation became essential to an investigation. After an **experiment**, scientists use their observations to formulate an initial proposal, which is called a **hypothesis**. However, a hypothesis must be tested repeatedly before it is accepted as valid. After a hypothesis has withstood extensive testing, it becomes either a **scientific theory** or a **natural law**. A scientific theory is an accepted explanation for the behavior of nature, whereas a natural law states a relationship under different experimental conditions and is often expressed as a mathematical equation.	• Describe the early practice of chemistry. *Related Exercises: 1–4* • Identify the three steps in the scientific method. *Related Exercises: 5–12*
1.2 Modern Chemistry The pseudoscience of **alchemy** introduced the practice of laboratory experimentation and was the forerunner of modern **chemistry**. Today, chemistry is quite diverse and has several branches, including **inorganic chemistry**, **organic chemistry**, **green chemistry**, and **biochemistry**. The impact of chemistry is felt in medicine and agriculture, as well as in the electronics, pharmaceutical, petrochemical, and other industries.	• Describe the modern practice of chemistry. *Related Exercises: 13–16*
1.3 Learning Chemistry In this section we examined some dual perceptions and pointed out that our brains have the ability to respond to the same image in two ways. Before beginning to learn chemistry, most students have already made associations with the subject. With this knowledge, hopefully you will be able to focus on chemistry as being an interesting and relevant subject and put aside any preconceived limiting attitudes.	• Conclude that chemistry is very relevant in our daily life. *Related Exercises: 17–18*

Key Terms

Answers to Key Terms are in Appendix H.

Select the key term that corresponds to each of the following definitions.

______ **1.** the methodical exploration of nature and the logical explanation of the observations

______ **2.** a scientific procedure for gathering data and recording observations under controlled conditions

______ **3.** a systematic investigation that entails performing an experiment, proposing a hypothesis, testing the hypothesis, and stating a theory or law

______ **4.** a tentative proposal of a scientific principle that attempts to explain the meaning of a set of data collected in an experiment

______ **5.** an extensively tested proposal of a scientific principle that explains the behavior of nature

______ **6.** an extensively tested proposal of a scientific principle that states a measurable relationship under different experimental conditions

______ **7.** a pseudoscience that attempted to convert a base metal, such as lead, to gold; a medieval science that sought to discover a universal cure for disease and a magic potion for immortality

______ **8.** the branch of science that studies the composition and properties of matter

______ **9.** the study of chemical substances that contain the element carbon

______ **10.** the study of chemical substances that do not contain the element carbon

______ **11.** the study of chemical substances derived from plants and animals

______ **12.** the design of products and processes that reduce or eliminate hazardous chemical substances

(a) alchemy (*Sec. 1.2*)
(b) biochemistry (*Sec. 1.2*)
(c) chemistry (*Sec. 1.2*)
(d) experiment (*Sec. 1.1*)
(e) green chemistry (*Sec. 1.2*)
(f) hypothesis (*Sec. 1.1*)
(g) inorganic chemistry (*Sec. 1.2*)
(h) natural law (*Sec. 1.1*)
(i) organic chemistry (*Sec. 1.2*)
(j) science (*Sec. 1.1*)
(k) scientific method (*Sec. 1.1*)
(l) scientific theory (*Sec. 1.1*)

Exercises

Answers to odd-numbered Exercises are in Appendix I.

Evolution of Chemistry (Sec. 1.1)

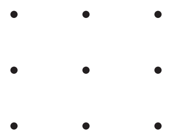
1. According to ancient Chinese beliefs, what two forces were responsible for bringing the earthly world into existence?
2. According to the Greek philosopher Thales in 600 B.C., what single element composed earth, air, and space?
3. According to the Greek philosopher Empedocles in 450 B.C., what four basic elements composed everything in nature?
4. According to the Greek philosopher Aristotle in 350 B.C., what five basic elements composed everything in nature?
5. What is the first step in the scientific method?
6. What is the second step in the scientific method?
7. What is the third step in the scientific method?
8. Who is the founder of the scientific method?
9. What is the difference between a scientific theory and a natural law?
10. What is the difference between a hypothesis and a theory?
11. Which of the following statements is a natural law?
    (a) The total mass of reacting substances remains constant after reaction.
    (b) The total energy of two gas molecules remains constant after colliding.
    (c) The volumes of two gases combine in the ratio of small whole numbers.
    (d) The nucleus of an atom contains positive charges.
12. Which of the following statements is a scientific theory?
    (a) The energy in an atomic nucleus is found by $E = mc^2$.
    (b) There is the same number of molecules in equal volumes of gases.
    (c) If the temperature of a gas doubles, the pressure doubles.
    (d) The region surrounding the nucleus has negative charges.

Modern Chemistry (Sec. 1.2)

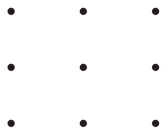
13. Why is chemistry referred to as the central science?
14. Who is the founder of modern chemistry?
15. State five professions that require a training in chemistry.
16. State five industries in which chemistry plays an important role.

Learning Chemistry (Sec. 1.3)

17. It is possible to solve the nine-dot problem with *one* straight, continuous line. Solve the problem and identify the unconscious assumption.

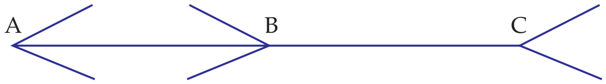


18. It is possible to solve the nine-dot problem with *three* straight, continuous lines. Solve the problem and identify the unconscious assumption.

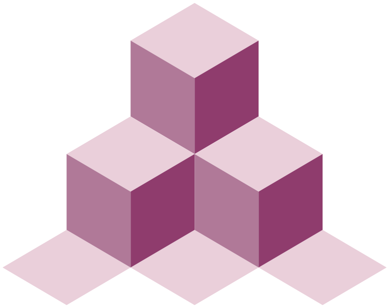


Challenge Exercises

19. Which of the line segments in the image below appears to be longer, AB or BC?



20. Stare at the image below and attempt to "flip" the stack of blocks upside down. (*Hint:* If you cannot see a second perception, stare at the point where the three blocks come together and mentally pull the point toward you.)



Chapter 1 Self-Test

Answers to Self-Test are in Appendix J.

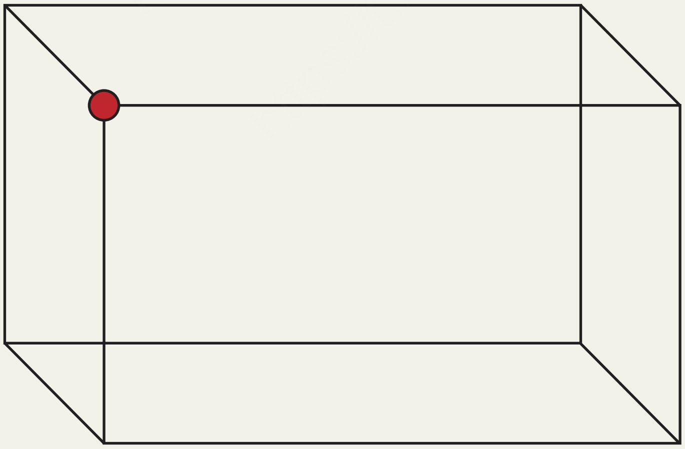
1. What is the difference between a scientific theory and a natural law? (Sec. 1.1)
- **(a)** a theory is a tentative proposal and a law is a tested proposal
- **(b)** a law is a tentative proposal and a theory is a tested proposal
- **(c)** a theory explains behavior and a law states a measurable relationship
- **(d)** a law explains behavior and a theory states a measurable relationship
- **(e)** none of the above

2. Petrochemicals are derived from which of the following resources? (Sec. 1.2)
- **(a)** atmosphere
- **(b)** petroleum
- **(c)** seawater
- **(d)** trees
- **(e)** none of the above

3. Which of the following is a positive reason for learning chemistry? (Sec. 1.3)
- **(a)** chemistry is relevant to daily life
- **(b)** chemistry offers career opportunities
- **(c)** chemistry studies interesting topics
- **(d)** chemistry benefits society
- **(e)** all of the above

Key Concept

4. Stare at the box until your first perception changes to a second view. (*Hint:* If you cannot see the second perception, stare at the red dot and mentally "push and pull" the dot.)

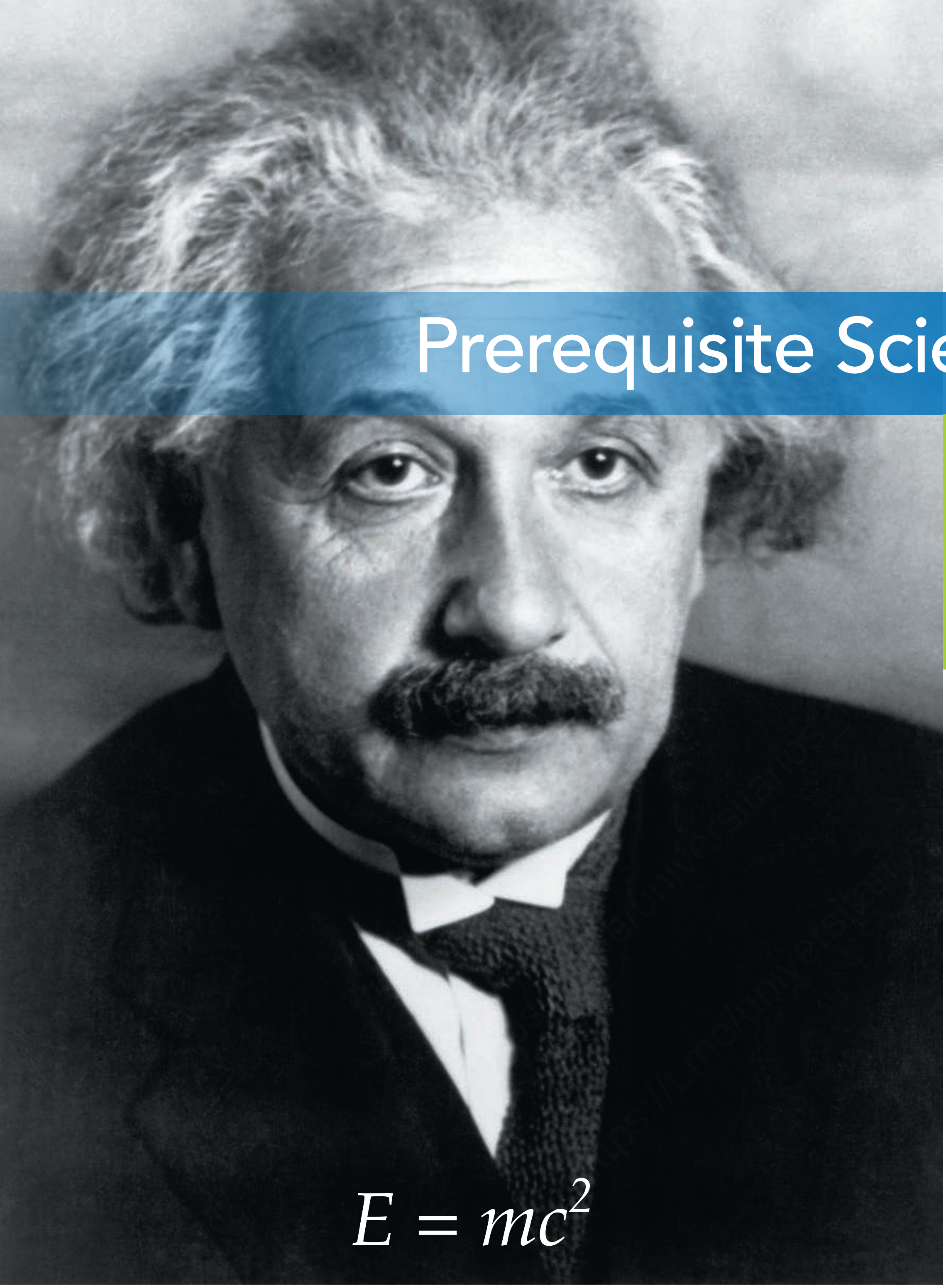


Critical Thinking

5. Chemists note the periodic table is arranged by the number of positive charges in an atom (that is, hydrogen (1), helium (2), lithium (3), etc.) and propose the chemical evolution of elements. Is chemical evolution an example of a scientific theory or a natural law?

6. Biologists note Darwin's principle of natural selection ("survival of the fittest") and propose the biological evolution of species. Is biological evolution an example of a scientific theory or a natural law?

Prerequisite Science Skills



"Logic will get you from A to B. Imagination will take you everywhere."

Albert Einstein, German-Swiss-American Physicist (1879–1955)

Albert Einstein (1879–1955)
Although Einstein was slow to speak and had difficulty in school as a child, he would go on to become the most famous scientist of the twentieth century. From his theory of relativity, we have the equation $E = mc^2$, which relates energy (E), mass (m), and the speed of light (c).

PSS.1 Measurements

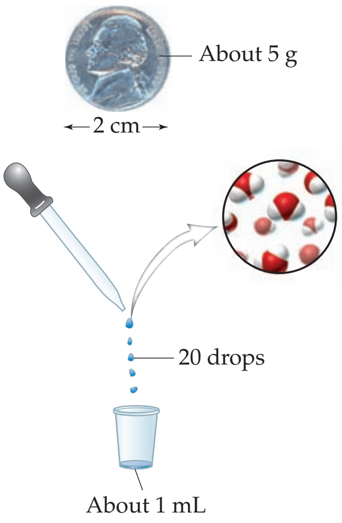
In chemistry, we express a **measurement** using the metric system. A common unit of length is a **centimeter (cm)**, a unit of mass is a **gram (g)**, and a unit of volume is a **milliliter (mL)**. Figure PSS.1 shows metric estimates for length, mass, and volume.

Measurements require the use of an **instrument**. The exactness of the measurement depends on the sensitivity of the instrument. For instance, an electronic balance can measure mass to 0.001 g. However, an exact measurement is not

LEARNING OBJECTIVES

- Explain why measurements are never exact.
- Identify instruments for taking measurements.

possible because no instrument measures exactly. An instrument may give a very sensitive reading, but every measurement has a degree of inexactness termed **uncertainty**.



▲ **Figure PSS.1 Estimates for Length, Mass, and Volume** The diameter of a 5¢ coin is about 2 cm, and its mass is about 5 g. The volume of 20 drops from an eye dropper is about 1 mL.

Length Measurements

To help you understand uncertainty, suppose we measure a metallic rectangular solid. We have two metric rulers available that differ as shown in Figure PSS.2. In this example Ruler B provides a more exact measurement than Ruler A.

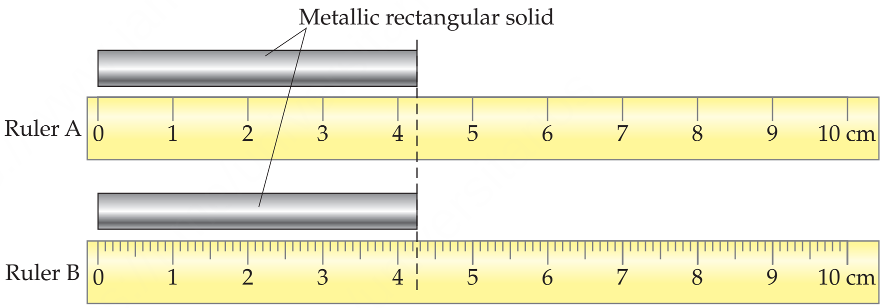
Notice that Ruler A has ten 1-cm divisions. Because the divisions are large, we can imagine 10 subdivisions. Thus, we can estimate to one-tenth of a division, that is, ±0.1 cm. On Ruler A, we see that the rectangular solid measures about 4.2 cm. Because the uncertainty is ±0.1 cm, a reading of 4.1 cm or 4.3 cm is also acceptable.

Notice that Ruler B has ten 1-cm divisions and ten 0.1-cm subdivisions. On Ruler B, the subdivisions are smaller. So, with Ruler B we can estimate to one-half of a subdivision, that is, we can estimate to ±0.05 cm. On Ruler B, we see that the rectangular solid measures about 4.25 cm. Because the uncertainty is ±0.05 cm, a reading of 4.20 cm or 4.30 cm is also acceptable.

We can compare the length of the rectangular solid measured with Rulers A and B as follows:

Ruler A: 4.2 ± 0.1 cm **Ruler B:** 4.25 ± 0.05 cm

In summary, Ruler A has more uncertainty and gives less precise measurements. Conversely, Ruler B has less uncertainty and gives more precise measurements. Example Exercise PSS.1 further illustrates the uncertainty in recorded measurements.



▶ **Figure PSS.2 Metric Rulers** On Ruler A, each division is 1 cm. On Ruler B, each division is 1 cm and each subdivision is 0.1 cm.

EXAMPLE EXERCISE PSS.1 Uncertainty in Measurement

Which measurements are consistent with the metric rulers shown in Figure PSS.2?

(a) Ruler A: 2 cm, 2.0 cm, 2.05 cm, 2.5 cm, 2.50 cm
(b) Ruler B: 3.0 cm, 3.3 cm, 3.33 cm, 3.35 cm, 3.50 cm

Solution

Ruler A has an uncertainty of ±0.1 cm, and Ruler B has an uncertainty of ±0.05 cm. Thus,

(a) Ruler A can give the measurements 2.0 cm and 2.5 cm.
(b) Ruler B can give the measurements 3.35 cm and 3.50 cm.

Practice Exercise

Which measurements are consistent with the metric rulers shown in Figure PSS.2?

(a) Ruler A: 1.5 cm, 1.50 cm, 1.55 cm, 1.6 cm, 2.00 cm
(b) Ruler B: 0.5 cm, 0.50 cm, 0.055 cm, 0.75 cm, 0.100 cm

Answers:

(a) 1.5 cm, 1.6 cm
(b) 0.50 cm, 0.75 cm

Concept Exercise

What high-tech instrument is capable of making an exact measurement?

Answer: See Appendix G, PSS.1.

Mass Measurements

The **mass** of an object is a measure of the amount of matter it contains. Although the term *weight* is often used instead of *mass*, the two terms are different, as weight is affected by gravity and mass is not. The weight of an astronaut on Earth may be 150 pounds, whereas in space the astronaut is weightless. The mass of the astronaut is the same on Earth and in space.

The measurement of mass always has uncertainty and varies with the balance. A typical balance in a laboratory may weigh a sample to 0.1 of a gram. Thus, the mass has an uncertainty of ±0.1 g. An electronic balance may weigh a sample to 0.001 of a gram. Thus, its mass has an uncertainty of ±0.001 g. Figure PSS.3 shows three common laboratory balances.

Volume Measurements

The amount of space occupied by a solid, gas, or liquid is its volume. Many pieces of laboratory equipment are available for measuring the volume of a liquid. Three of the most common are a graduated cylinder, a pipet, and a buret. Figure PSS.4 shows common laboratory equipment used for measuring volume.

A graduated cylinder is routinely used to measure a volume of liquid. The most common sizes of graduated cylinders are 10 mL, 50 mL, and 100 mL. The uncertainty of a graduated cylinder measurement varies, but usually ranges from 1/10 to 1/2 of a milliliter (±0.1 mL to ±0.5 mL).

There are many types of pipets. The volumetric pipet shown in Figure PSS.4 is used to deliver a fixed volume of liquid. The liquid is drawn up until it reaches a calibration line etched on the pipet. The tip of the pipet is then placed in a container, and the liquid is allowed to drain from the pipet. The volume delivered varies, but the uncertainty usually

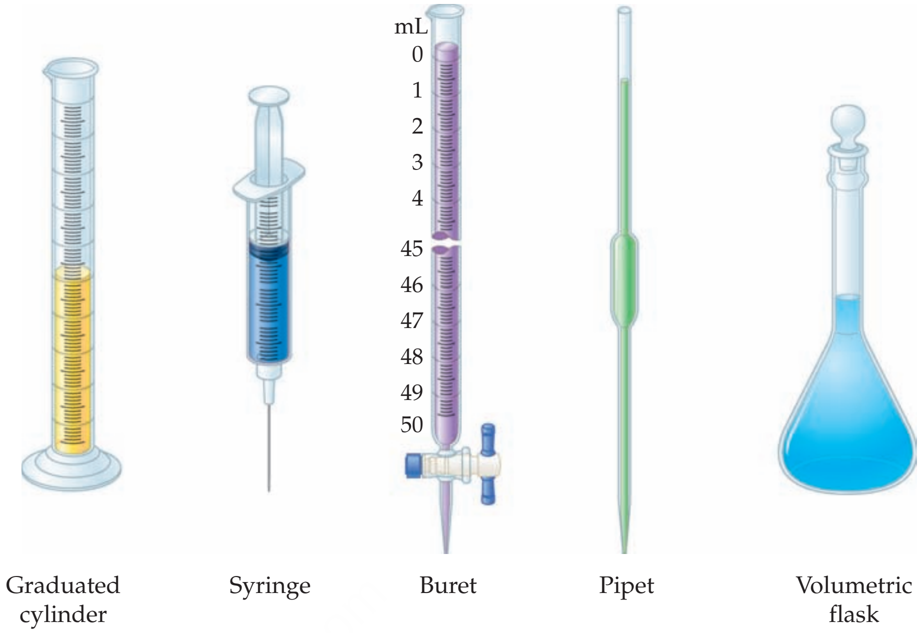
(a)

(b)

(c)

◀ **Figure PSS.3 Laboratory Balances** (a) A balance having an uncertainty of ±0.1 g. (b) A balance having an uncertainty of ±0.01 g. (c) An electronic balance having an uncertainty of ±0.001 g.

▶ Figure PSS.4 Laboratory Instruments for Volume A graduated cylinder, a syringe, and a buret deliver a variable amount of liquid, whereas a volumetric pipet and a volumetric flask contain fixed amounts of liquid, for example, 10 mL and 250 mL.



ranges from 1/10 to 1/100 of a milliliter. For instance, a 10-mL pipet can deliver 10.0 mL (±0.1 mL) or 10.00 mL (±0.01 mL), depending on the uncertainty of the instrument.

A buret is a long, narrow piece of calibrated glass tubing with a valve called a "stopcock" near the tip. The flow of liquid is regulated by opening and closing the stopcock, and the initial and final liquid levels in the buret are observed and recorded. The volume delivered is found by subtracting the initial buret reading from the final buret reading. Burets usually have uncertainties ranging from 1/10 to 1/100 of a milliliter. For instance, the liquid level in a buret can read 22.5 mL (±0.1 mL) or 22.55 mL (±0.01 mL), depending on the uncertainty of the instrument.

PSS.2 Significant Digits

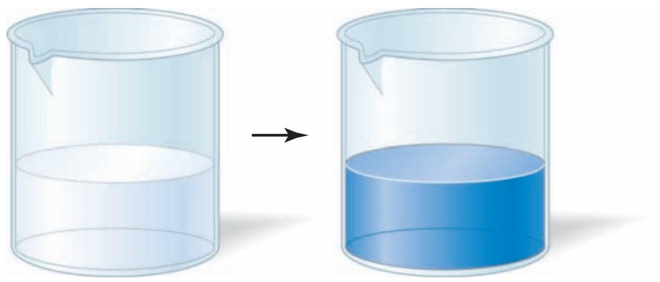
LEARNING OBJECTIVE

- Identify significant digits in a measurement.

In a recorded measurement, all numbers are **significant digits**, which can be referred to as a "significant figures" or "sig. figs." For instance, if we weigh a 5¢ nickel coin on different balances, we may record the mass of the coin as 5.0 g, 5.00 g, or 5.000 g. Although the uncertainty varies for the three balances, every digit is significant in all three measurements. Removing the last digit changes the uncertainty of the measurement. In this example, the measurements of mass have two, three, and four significant digits, respectively.

In every measurement, the significant digits express the uncertainty of the instrument. By way of example, refer to Figure PSS.5, which shows a chemical reaction that requires about 35 seconds (s).

▶ Figure PSS.5 Significant Digits and a Timed Reaction The data demonstrates uncertainty of three different stopwatches. Although each of the measurements is correct, Stopwatch A has the most uncertainty, and Stopwatch C has the least uncertainty.



	Colorless Solution	**Blue Solution**
Stopwatch A:	0 s	35 s
Stopwatch B:	0.0 s	35.1 s
Stopwatch C:	0.00 s	35.08 s

To determine the number of significant digits in a measurement, we simply count the number of digits from left to right, starting with the first nonzero digit. Therefore, 35 s has two significant digits, 35.1 s has three significant digits, and 35.08 s has four significant digits. We can summarize the rules for determining the number of significant digits with the following rules.

Determining Significant Digits

Rule 1: Count the number of digits in a measurement from left to right.

(a) Start with the first nonzero digit.
(b) Do not count placeholder zeros.
(0.11 g, 0.011 g, and 110 g each have two significant digits.)

Rule 2: The rules for significant digits apply only to measurements and not to exact numbers. Exact numbers can be derived from the following:

(a) counting items, such as 6 test tubes.
(b) exact relationships, such as 1 meter = 100 centimeters.
(c) an equation such as 1 diameter = 2 radii.

Example Exercise PSS.2 illustrates the number of significant digits in a measurement.

EXAMPLE EXERCISE PSS.2 Significant Digits

State the number of significant digits in the following measurements:

(a) 12,345 cm
(b) 0.123 g
(c) 0.5 mL
(d) 102.0 s

Solution

In each example, we simply count the number of digits. Thus,

(a) 5
(b) 3
(c) 1
(d) 4

Notice that the leading zero in (b) and (c) is not part of the measurement but is inserted to call attention to the decimal point that follows.

Practice Exercise

State the number of significant digits in the following measurements:

(a) 2005 cm
(b) 25.000 g
(c) 25.0 mL
(d) 0.25 s

Answers:

(a) 4
(b) 5
(c) 3
(d) 2

Concept Exercise

What type of measurement is exact?

Answer: See Appendix G, PSS.2.

Significant Digits and Placeholder Zeros

A measurement may contain placeholder zeros to properly locate the decimal point, for example, 500 cm and 0.005 cm. If the number is less than 1, a placeholder zero is never significant. Thus, 0.5 cm, 0.05 cm, and 0.005 cm each contain only one significant digit.

If the number is greater than 1, we will assume that placeholder zeros are not significant. Thus, 50 cm, 500 cm, and 5000 cm each contain only one significant digit. Example Exercise PSS.3 illustrates how to determine the number of significant digits in a measurement with placeholder zeros.

EXAMPLE EXERCISE PSS.3 Significant Digits

State the number of significant digits in the following measurements:

(a) 0.0025 cm	(b) 0.2050 g
(c) 250 mL	(d) 2500 s

Solution

In each example, we count the number of significant digits and disregard placeholder zeros. Thus,

(a) 2	(b) 4
(c) 2	(d) 2

Practice Exercise

State the number of significant digits in the following measurements:

(a) 0.050 cm	(b) 0.055 g
(c) 50.00 mL	(d) 1000 s

Answers:

(a) 2	(b) 2
(c) 4	(d) 1

Concept Exercise

What type of measurement has infinite significant digits?

Answer: See Appendix G, PSS.3.

Note If a placeholder zero is significant, we can express the number using a power of 10. For example, if one zero is significant in 100 cm, we can express the measurement as 1.0×10^2 cm. If both zeros are significant, we can write 1.00×10^2 cm. If neither zero is significant, we can write 1×10^2 cm. The power of 10 does not affect the number of significant digits; thus, 1.1×10^5 cm has two significant digits, and 1.11×10^{-5} cm has three significant digits.

▲ **Exact Numbers** We count seven coins in the photo, which is an exact number. This is not a measurement; thus, the concept of significant digits does not apply.

Significant Digits and Exact Numbers

All measurements have uncertainty, so a measurement is never an exact number. However, we can obtain exact numbers when counting items. For instance, a chemistry laboratory may have 24 pipets. Because we have simply counted items, 24 is an exact number. Significant digits do not apply to exact numbers, only to measurements.

LEARNING OBJECTIVE

- Apply rules for rounding off nonsignificant digits from a calculation.

PSS.3 Rounding Off Nonsignificant Digits

All digits in a correctly recorded measurement, except placeholder zeros, are significant. However, we often generate **nonsignificant digits** when using a calculator. These nonsignificant digits should not be reported, but they frequently appear in the calculator display and must be eliminated. We eliminate nonsignificant digits through a process of **rounding off**. We round off nonsignificant digits by following three simple rules.

Rounding Off Nonsignificant Digits

Rule 1: If the first nonsignificant digit is less than 5, drop all nonsignificant digits.

Rule 2: If the first nonsignificant digit is greater than 5, or equal to 5, increase the last significant digit by 1 and drop all nonsignificant digits.*

Rule 3: If a calculation has several multiplication or division operations, retain nonsignificant digits in your calculator display until the last operation. Not only is it more convenient, it is also more accurate.

*If the nonsignificant digit is 5, or 5 followed by zeros, an odd–even rule can be applied. That is, if the last significant digit is odd, round up; if it is even, drop the nonsignificant digits.

If a calculator displays 12.846239 and three significant digits are justified, we must round off. Because the first nonsignificant digit is 4 in 12.8**4**6239, we follow Rule 1, drop the nonsignificant digits, and round to 12.8. If a calculator displays 12.856239 and three significant digits are justified, we follow Rule 2. In this case, the first nonsignificant digit is 5 in 12.8**5**6239, so we round to 12.9.

▲ Scientific Calculator

Rounding Off and Placeholder Zeros

On occasion, rounding off can create a problem. For example, if we round off 151 to two significant digits, we obtain 15. Because 15 is only a fraction of the original value, we must insert a placeholder zero; thus, rounding off 151 to two significant digits gives 150. Similarly, rounding off 1514 to two significant digits gives 1500 or 1.5×10^3. Example Exercise PSS.4 further illustrates how to round off numbers.

EXAMPLE EXERCISE PSS.4 Rounding Off

Round off the following numbers to three significant digits:

(a) 22.250	(b) 0.34548
(c) 0.072038	(d) 12,267

Solution

To locate the first nonsignificant digit, count three digits from left to right. If the first nonsignificant digit is less than 5, drop all nonsignificant digits. If the first nonsignificant digit is 5 or greater, add 1 to the last significant digit.

(a) 22.3 (Rule 2)	(b) 0.345 (Rule 1)
(c) 0.0720 (Rule 1)	(d) 12,300 (Rule 2)

In (d), notice that two placeholder zeros must be added to 123 to obtain the correct decimal place.

Practice Exercise

Round off the following numbers to three significant digits:

(a) 12.514748	(b) 0.6015261
(c) 192.49032	(d) 14652.832

Answers:

(a) 12.5 (Rule 1)	(b) 0.602 (Rule 2)
(c) 192 (Rule 1)	(d) 14,700 (Rule 2)

Concept Exercise

How many significant digits are in the exact number 155?

Answer: See Appendix G, PSS.4.

PSS.4 Adding and Subtracting Measurements

LEARNING OBJECTIVE

- Apply significant digit rule for adding and subtracting measurements.

When adding or subtracting measurements, *the answer is limited by the value with the most uncertainty;* that is, the answer is limited by the decimal place. Note the decimal place in the following examples:

$$\begin{array}{r} 5\ \text{g} \\ 5.0\ \text{g} \\ +5.00\ \text{g} \\ \hline 15.00\ \text{g} \end{array}$$

The mass of 5 g has the most uncertainty because it measures only ±1 g. Thus, the sum should be limited to the nearest gram. If we round off the answer to the proper significant digit, the correct answer is **15 g**. In addition and subtraction, the unit (cm, g, mL) in the answer is the same as the unit in each piece of data. Example Exercise PSS.5 illustrates the addition and subtraction of measurements.

EXAMPLE EXERCISE PSS.5 Addition/Subtraction and Rounding Off

Add or subtract the following measurements and round off your answer:

(a) 106.7 g + 0.25 g + 0.195 g
(b) 35.45 mL − 30.5 mL

Solution

In addition or subtraction operations, the answer is limited by the measurement with the most uncertainty.

(a) Let's align the decimal places and perform the addition.

$$\begin{array}{r} 106.7\ \text{g} \\ 0.25\ \text{g} \\ \underline{+0.195\ \text{g}} \\ 107.145\ \text{g} \end{array}$$

Because 106.7 g has the most uncertainty (±0.1 g) the answer rounds off to one decimal place. The correct answer is **107.1 g** and is read "**one hundred and seven point one grams**."

(b) Let's align the decimal places and perform the subtraction.

$$\begin{array}{r} 35.45\ \text{mL} \\ \underline{-30.5\ \text{mL}} \\ 4.95\ \text{mL} \end{array}$$

Because 30.5 mL has the most uncertainty (±0.1 mL), we round off to one decimal place. The answer is **5.0 mL** and is read "**five point zero milliliters**."

Practice Exercise

Add or subtract the following measurements and round off your answer:

(a) 8.6 cm + 50.05 cm (b) 34.1 s − 0.55 s

Answers:

(a) 58.7 cm (b) 33.6 s

Concept Exercise

When adding or subtracting measurements, which measurement in a set of data limits the answer?

Answer: See Appendix G, PSS.5.

Helpful Hint
Significant Digits

The rule for rounding off addition and subtraction of measurements is different from the rule for rounding off multiplication and division of measurements.

- When adding and subtracting measurements, the answer is limited by the measurement value having the least certainty. That is, when adding or subtracting measurements, the measurement with the fewest decimal places limits the digits in the answer.
- When multiplying and dividing measurements, the answer is limited by the measurement having the least number of significant digits. That is, when multiplying and dividing measurements, it is the measurement with the fewest number of significant digits that limits the digits in the answer.

PSS.5 Multiplying and Dividing Measurements

LEARNING OBJECTIVE

- Apply significant digit rule for multiplying and dividing measurements.

Significant digits are treated differently in multiplication and division than in addition and subtraction. In multiplication and division, *the answer is limited by the measurement with the least number of significant digits.* Let's multiply the following length measurements:

$$5.15\ \text{cm} \times 2.3\ \text{cm} = 11.845\ \text{cm}^2$$

The measurement of 5.15 cm has three significant digits, and 2.3 cm has two. Thus, the product should be limited to two digits. When we round off to the proper number of significant digits, the correct answer is **12 cm^2**. Notice that the units must also be multiplied together, which we have indicated by the superscript 2. Example Exercise PSS.6 illustrates the multiplication and division of measurements.

EXAMPLE EXERCISE PSS.6 Multiplication/Division and Rounding Off

Multiply or divide the following measurements and round off your answer:

(a) 50.5 cm × 12 cm
(b) 103.37 g/20.5 mL

Solution

In multiplication and division operations, the answer is limited by the measurement with the least number of significant digits.

(a) In this example, 50.5 cm has three significant digits and 12 cm has two.

$$50.5 \text{ cm} \times 12 \text{ cm} = 606 \text{ cm}^2$$

The answer is limited to two significant digits and rounds off to **610 cm^2** after inserting a placeholder zero. The answer is read "**six hundred ten square centimeters**."

(b) In this example, 103.37 g has five significant digits and 20.5 mL has three.

$$\frac{103.37 \text{ g}}{20.5 \text{ mL}} = 5.0424 \text{ g/mL}$$

The answer is limited to three significant digits and rounds off to **5.04 g/mL**. Notice that the unit is a ratio; the answer is read as "**five point zero four grams per milliliter**."

Practice Exercise

Multiply or divide the following measurements and round off your answer.

(a) 359 cm × 0.20 cm (b) 73.950 g/25.5 mL

Answers:

(a) 72 cm^2 (b) 2.90 g/mL

Concept Exercise

When multiplying or dividing measurements, which measurement in a set of data limits the answer?

Answer: See Appendix G, PSS.6.

PSS.6 Exponential Numbers

LEARNING OBJECTIVES

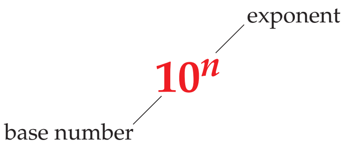
- Express an ordinary number as a power of 10.
- Express a power of 10 as an ordinary number.

When a value is multiplied times itself, the process is indicated by a number written as a superscript. The superscript indicates the number of times the process is repeated. For example, if the number 2 is multiplied two times, the product is expressed as 2^2. Thus, $(2)(2) = 2^2$. If the number 2 is multiplied three times, the product is expressed as 2^3. Thus, $(2)(2)(2) = 2^3$.

A superscript number indicating that a value is multiplied times itself is called an **exponent**. If 2 has the exponent 2, the value 2^2 is read as "2 to the second power" or "2 squared." The value 2^3 is read as "2 to the third power" or "2 cubed."

Powers of 10

A **power of 10** is a number that results when 10 is raised to an exponential power. You know that an exponent raises any number to a higher power, but we are most interested in the base number 10. A power of 10 has the general form



The number 10 raised to the *n* power is equal to 10 multiplied times itself *n* times. For instance, 10 to the second power (10^2) is equal to 10 times 10. When we write 10^2 as an ordinary number, we have 100. Notice that the exponent 2 corresponds to the number of zeros in 100. Similarly, 10^3 has three zeros (1000) and 10^6 has six zeros (1,000,000).

The exponent is positive for all numbers greater than 1. Conversely, the exponent is negative for numbers less than 1. For example, 10 to the negative first power (10^{-1}) is equal to 0.1, 10 to the negative second power (10^{-2}) is equal to 0.01, and 10 to the negative third power (10^{-3}) is equal to 0.001. Table PSS.1 lists some powers of 10, along with the equivalent ordinary number.

TABLE PSS.1 Powers of 10

Exponential Number	Ordinary Number
$1 \times 10^6 = 10 \times 10 \times 10 \times 10 \times 10 \times 10$	1,000,000
$1 \times 10^3 = 10 \times 10 \times 10$	1000
$1 \times 10^2 = 10 \times 10$	100
$1 \times 10^1 = 10$	10
$1 \times 10^0 = 1$	1
$1 \times 10^{-1} = \frac{1}{10}$	0.1
$1 \times 10^{-2} = \frac{1}{10} \times \frac{1}{10}$	0.01
$1 \times 10^{-3} = \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}$	0.001
$1 \times 10^{-6} = \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}$	0.000 001

Although you can easily carry out operations with exponents using an inexpensive scientific calculator, you will have greater confidence if you understand exponents. Example Exercises PSS.7 and PSS.8 illustrate the relationship between ordinary numbers and exponential numbers.

EXAMPLE EXERCISE PSS.7 Converting to Powers of 10

Express each of the following ordinary numbers as a power of 10:

(a) 100,000 (b) 0.000 000 01

Solution

The power of 10 indicates the number of places the decimal point has been moved.

(a) We must move the decimal five places to the left; thus, 1×10^5.

(b) We must move the decimal eight places to the right; thus, 1×10^{-8}.

▲ **Earth** A positive power of 10 can be used to express very large numbers; for example, the diameter of Earth is about 1×10^7 meters.

Practice Exercise

Express each of the following ordinary numbers as a power of 10:

(a) 10,000,000 (b) 0.000 000 000 001

Answers

(a) 1×10^7 (b) 1×10^{-12}

Concept Exercise

Which of the following lengths is less: 1×10^3 cm or 1×10^{-3} cm?

Answer: See Appendix G, PSS.7.

EXAMPLE EXERCISE PSS.8 Converting to Ordinary Numbers

Express each of the following powers of 10 as an ordinary number:

(a) 1×10^4
(b) 1×10^{-9}

Solution

The power of 10 indicates the number of places the decimal point has been moved.

(a) The exponent in 1×10^4 is positive 4, and so we must move the decimal point four places to the right of 1, thus, 10,000.

(b) The exponent in 1×10^{-9} is negative 9, and so we must move the decimal point nine places to the left of 1, thus, 0.000 000 001.

Practice Exercise

Express each of the following powers of 10 as an ordinary number:

(a) 1×10^{10} (b) 1×10^{-5}

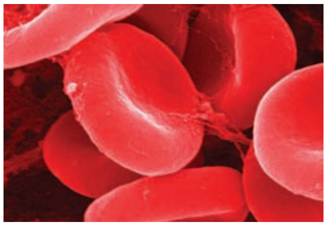
Answers

(a) 10,000,000,000 (b) 0.000 01

Concept Exercise

Which of the following masses is less: 0.000 001 g or 0.000 01 g?

Answer: See Appendix G, PSS.8.



▲ **Red Blood Cells** A negative power of 10 can be used to express very small numbers; for example, the diameter of a red blood cell is about 1×10^{-6} meter.

PSS.7 Scientific Notation

LEARNING OBJECTIVES

- Show an ordinary number in scientific notation.
- Show scientific notation as an ordinary number.

Science often deals with numbers that are very large or very small. These numbers may be awkward and incomprehensible when written with many zeros. The mass of an iron atom, for example, is 0.000 000 000 000 000 000 000 093 g. To overcome this problem, a standard notation has been devised that places the decimal after the first significant digit and sets the size of the number using a power of 10; this method is called **scientific notation**. The scientific notation format is

power of 10

$$D\,.\,D\,D \times 10^n$$

significant digits

To use scientific notation, write down the significant digits (D) in the number. Then, move the decimal point to follow the first nonzero digit. Indicate the number of places the decimal is moved using power-of-10 notation. For instance, you can write 555,000 in scientific notation by first moving the decimal five places to the left to give 5.55 and then adding the appropriate power of 10, which is 10^5. Thus, you can express 555,000 in scientific notation as 5.55×10^5.

You can also express numbers smaller than 1 in scientific notation. For example, you can write 0.000 888 in scientific notation by first moving the decimal four places to the right to give 8.88 and then adding the appropriate power of 10, which is 10^{-4}. Thus, 0.000 888 is expressed in scientific notation as 8.88×10^{-4}.

Regardless of the size of the number, in scientific notation the decimal point is always placed after the first nonzero digit. The size of the number is indicated by a power of 10. A positive exponent indicates a very large number, whereas a negative exponent indicates a very small number. To express a number in scientific notation, follow these two steps.

Applying Scientific Notation

Step 1: Place the decimal point after the first nonzero digit in the number, followed by the remaining significant digits.

Step 2: Indicate how many places the decimal is moved by the power of 10. When the decimal is moved to the left, the power of 10 is positive. When the decimal is moved to the right, the power of 10 is negative.

Example Exercises PSS.9 and PSS.10 illustrate the relationship between ordinary numbers and scientific notation.

EXAMPLE EXERCISE PSS.9 Converting to Scientific Notation

Express each of the following values in scientific notation:

(a) There are 26,800,000,000,000,000,000,000 helium atoms in a 1-liter balloon filled with helium gas.

(b) The mass of one helium atom is 0.000 000 000 000 000 000 000 006 65 g.

Solution

We can write each value in scientific notation as follows:

(a) Place the decimal after the 2, followed by the other significant digits (2.68). Next, count the number of places the decimal has moved. The decimal is moved to the left 22 places, so the exponent is +22. Finally, we have the number of helium atoms in 1.00 L of gas: 2.68×10^{22} atoms.

(b) Place the decimal after the 6, followed by the other significant digits (6.65). Next, count the number of places the decimal has shifted. The decimal has shifted 24 places to the right, so the exponent is −24. Finally, we have the mass of a helium atom: 6.65×10^{-24} g.

Practice Exercise

Express each of the following values as ordinary numbers:

(a) The mass of one mercury atom is 3.33×10^{-22} g.

(b) The number of atoms in 1 mL of liquid mercury is 4.08×10^{22}.

Answers

(a) 0.000 000 000 000 000 000 000 333 g;

(b) 40,800,000,000,000,000,000,000 atoms

Concept Exercise

Which of the following masses is greater: 1×10^{-6} g or 0.000 01 g?

Answer: See Appendix G, PSS.9.

EXAMPLE EXERCISE PSS.10 Converting to Ordinary Numbers

Express each of the following scientific notation values as an ordinary number.

(a) 1.75×10^{3} (b) 2.88×10^{-2}

Solution

The power of 10 indicates the number of places the decimal point has been moved.

(a) The exponent in 1.75×10^{3} is positive 3, so we move the decimal point three places to the right of 1.75, thus, 1750.

(b) The exponent in 2.88×10^{-2} is negative 2, so we move the decimal point two places to the left of 2.88, thus, 0.0288.

Practice Exercise

Express each of the following scientific notation values as an ordinary number.

(a) 7.11×10^5 (b) 8.23×10^{-4}

Answers:

(a) 711,000 (b) 0.000 823

Concept Exercise

Express the following as an ordinary number: 1.55×10^0 cm.

Answer: See Appendix G, PSS.10.

A CLOSER LOOK Calculators

Q: What type of calculator do I need for introductory chemistry?

A student may have a personal calculator for routine purposes. However, a personal calculator, or smartphone calculator, is not usually appropriate for a chemistry course. A scientific calculator is required for the calculations encountered in chemistry. Before purchasing a calculator, check with your instructor for a calculator recommendation for your course.

At a minimum, a chemistry course will require a calculator having a key for entering exponents (e.g., EE, EXP, EEX), a key for storing and recalling from memory (STO and RCL), and finding logarithms (LOG). Unless your college major is chemistry, a basic, inexpensive calculator is probably your best choice. The Texas Instruments TI-30 series is a popular, inexpensive, and appropriate calculator for introductory chemistry. Casio, Sharp, and Hewlett Packard offer similar inexpensive calculators (less than $10) although each operates differently.

All too often, students choose a sophisticated calculator thinking that it will be helpful in performing the calculations encountered in chemistry. In practice, the opposite is true as a sophisticated scientific calculator has so many functions that there is a steep learning curve to master even basic operations such as entering exponents and using the memory function.

The following exercise demonstrates the minimum calculator operations required for introductory chemistry.

(a) enter 6.02×10^{23} in the calculator display (Enter: 6.02 EXP23)
(b) store 6.02×10^{23} in calculator memory
(c) recall 6.02×10^{23} from calculator memory
(d) find the log of 2.55×10^{-3} (Answer: −2.59)
(e) find the inverse log of −4.57 (Answer: 0.0000269)

▲ **Scientific Calculator** A scientific calculator has an exponent key (EXP) for expressing powers of 10.

A: A scientific calculator is required for the calculations encountered in chemistry. The best choice of a calculator will vary for the individual student. For example, a science major may choose a *scientific graphing calculator*. However, a graphing calculator requires much effort to use and may actually be counterproductive to the student who is only required to perform routine introductory chemistry calculations.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
PSS.1 Measurements A **measurement** is obtained using a laboratory **instrument**. A balance is the instrument used to measure **mass**. We can record length, mass, and volume measurements in metric units of **centimeter (cm)**, **gram (g)**, and **milliliter (mL)**. No measurement is exact because every instrument has **uncertainty**.	• Explain why measurements are never exact. ***Related Exercises: 1–2*** • Identify instruments for taking measurements. ***Related Exercises: 3–10***
PSS.2 Significant Digits The **significant digits** in a measurement are those digits known with certainty, plus one digit that is estimated.	• Identify significant digits in a measurement ***Related Exercises: 11–16***
PSS.3 Rounding Off Nonsignificant Digits **Nonsignificant digits** are obtained from a calculator. The process of eliminating nonsignificant digits is called **rounding off**.	• Round off nonsignificant digits. ***Related Exercises: 17–20***
PSS.4 Adding and Subtracting Measurements When adding or subtracting measurements, the answer is limited by the measurement having the most uncertainty. If the uncertainties of three mass measurements are ±0.1 g, ±0.01 g, and ±0.001 g, the measurement with the most uncertainty, ±0.1 g, limits the answer.	• Add and subtract measurements and round off correctly. ***Related Exercises: 21–24***
PSS.5 Multiplying and Dividing Measurements When multiplying or dividing measurements, the answer is limited by the least number of significant digits in the data. Every calculated answer must include the correct number of significant digits, as well as the proper units.	• Multiply and divide measurements and round off correctly. ***Related Exercises: 25–28***
PSS.6 Exponential Numbers A number written as a superscript indicating that a value is multiplied by itself is called an **exponent**. If the value is 10, the exponent is a **power of 10**. A power of 10 can be positive or negative. If the exponent is positive (10^3), the number is greater than 1; if the exponent is negative (10^{-3}), the number is less than 1.	• Express an ordinary number as a power of 10. ***Related Exercises: 29–32*** • Express a power of 10 as an ordinary number. ***Related Exercises: 33–38***
PSS.7 Scientific Notation In science we use **scientific notation** to express numbers that are very large or very small. Scientific notation uses the following format to express a number with three significant digits: D.DD × 10^n. The three significant digits (D.DD) are multiplied by a power of 10 (10^n) to set the decimal point.	• Show an ordinary number in scientific notation. ***Related Exercises: 39–40*** • Show scientific notation as an ordinary number. ***Related Exercises: 41–44***

Problem–Solving Organizer

Topic	Procedure	Example
Measurements PSS.1	State the measurement and indicate the uncertainty.	11.50 cm ± 0.05 cm 0.545 g ± 0.001 g 10.0 mL ± 0.5 mL
Significant Digits PSS.2	Count the significant digits in a measurement from left to right: (a) start with the first nonzero digit, (b) do not count placeholder zeros.	0.05 g (1 significant digit) 50.0 g (3 significant digits) 505 g (3 significant digits) 50,000 g (1 significant digit)

Problem–Solving Organizer *(Continued)*

Topic	Procedure	Example
Rounding Off Nonsignificant Digits PSS.3	1. If the first nonsignificant digit is less than 5, drop all nonsignificant digits. 2. If the first nonsignificant digit is 5 or greater, add 1 to the last significant digit. 3. If a calculation involves several steps, retain nonsignificant digits until the final step.	505.05 g rounds to 505 g (3 significant digits) 505.50 g rounds to 506 g (3 significant digits)
Adding and Subtracting Measurements PSS.4	The answer is limited by the least certain measurement in a set of data.	5.05 g + 5.005 g = 10.06 g 5.05 mL − 2.1 mL = 3.0 mL
Multiplying and Dividing Measurements PSS.5	The answer is limited by the least number of significant digits in the given data.	5.05 cm × 5.005 cm = 25.3 cm^2 5.0 g/10.0 mL = 0.50 g/mL
Exponential Numbers PSS.6	(a) Numbers greater than 1 have a positive exponent. (b) Numbers less than 1 have a negative exponent.	1,000,000 = 1×10^6 0.000 001 = 1×10^{-6}
Scientific Notation PSS.7	Place the decimal point after the first nonzero digit. Using an exponent, indicate how many places the decimal is moved.	55,500 g = 5.55×10^4 g 0.000 555 g = 5.55×10^{-4} g

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** a numerical value and unit that expresses length, mass, or volume
______ **2.** a common metric unit of length
______ **3.** a common metric unit of mass
______ **4.** a common metric unit of volume
______ **5.** a device for recording a measurement such as length, mass, or volume
______ **6.** the degree of inexactness in an instrumental measurement
______ **7.** the quantity of matter in an object
______ **8.** the certain digits in a measurement plus one estimated digit
______ **9.** the digits in a measurement that exceed the certainty of the instrument
______ **10.** the process of eliminating digits that are not significant
______ **11.** a number written as a superscript that indicates a value is multiplied by itself, for example, $10^4 = 10 \times 10 \times 10 \times 10$, and $\text{cm}^3 = \text{cm} \times \text{cm} \times \text{cm}$
______ **12.** a positive or negative exponent of 10
______ **13.** a method for expressing numbers by moving the decimal place after the first significant digit and indicating the number of decimal moves by a power of 10

(a) centimeter (cm) *(PSS.1)*
(b) exponent *(PSS.6)*
(c) gram (g) *(PSS.1)*
(d) instrument *(PSS.1)*
(e) mass *(PSS.1)*
(f) measurement *(PSS.1)*
(g) milliliter (mL) *(PSS.1)*
(h) nonsignificant digits *(PSS.3)*
(i) power of 10 *(PSS.6)*
(j) rounding off *(PSS.3)*
(k) scientific notation *(PSS.7)*
(l) significant digits *(PSS.2)*
(m) uncertainty *(PSS.1)*

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Uncertainty in Measurements (Sec. PSS.1)

1. What is the approximate mass of a 5¢ nickel coin?
2. What is the approximate diameter of a 5¢ nickel coin?
3. What quantity (length, mass, volume, time) is expressed by the following units?
   (a) centimeter **(b)** gram
   (c) milliliter **(d)** second
4. What quantity (length, mass, volume, time) is expressed by the following units?
   (a) meter **(b)** kilogram
   (c) liter **(d)** millisecond
5. Which of the following measurements corresponds to metric Ruler A shown in Figure PSS.2?
   (a) 5.0 cm **(b)** 5.00 cm
   (c) 5.05 cm **(d)** 0.5 cm
6. Which of the following measurements corresponds to metric Ruler B shown in Figure PSS.2?
   (a) 2.0 cm **(b)** 2.00 cm
   (c) 2.05 cm **(d)** 0.5 cm

7. Which of the following measurements corresponds to the electronic balance shown in Figure PSS.2?
(a) 75 g **(b)** 75.0 g
(c) 75.22 g **(d)** 75.518 g

8. Which of the following measurements corresponds to the electronic balance shown in Figure PSS.2?
(a) 25 g **(b)** 25.0 g
(c) 25.00 g **(d)** 25.000 g

9. Which of the following measurements corresponds to the buret shown in Figure PSS.3 having an uncertainty of ±0.05 mL?
(a) 32.0 mL **(b)** 32.00 mL
(c) 32.05 mL **(d)** 32.58 mL

10. Which of the following measurements corresponds to the graduated cylinder shown in Figure PSS.3 having an uncertainty of ±0.5 mL?
(a) 25 mL **(b)** 25.0 mL
(c) 25.5 mL **(d)** 25.50 mL

Significant Digits (Sec. PSS.2)

11. State the number of significant digits in each of the following:
(a) 20 cm **(b)** 200 g
(c) 0.002 mL **(d)** 0.0002 s

12. State the number of significant digits in each of the following:
(a) 500 cm **(b)** 5000 g
(c) 0.05 mL **(d)** 0.005 s

13. State the number of significant digits in each of the following:
(a) 0.5 cm **(b)** 1055 g
(c) 1.2×10^0 mL **(d)** 4.50×10^1 s

14. State the number of significant digits in each of the following:
(a) 5.0 cm **(b)** 5.05 g
(c) 5.02×10^{-1} mL **(d)** 1.0×10^{-2} s

15. State the number of significant digits in each of the following:
(a) 5.00 cm **(b)** 5005 g
(c) 1×10^{-1} mL **(d)** 1.0×10^{-2} s

16. State the number of significant digits in each of the following:
(a) 0.5 cm **(b)** 0.50 g
(c) 1.00×10^1 mL **(d)** 1.000×10^3 s

Rounding Off Nonsignificant Digits (Sec. PSS.3)

17. Round off the following to three significant digits:
(a) 10.25 **(b)** 10.20
(c) 0.01029 **(d)** 10,248

18. Round off the following to three significant digits:
(a) 20.155 **(b)** 0.204 500
(c) 2055 **(d)** 0.2065

19. Round off the following to three significant digits:
(a) 9.123×10^5 **(b)** 9.456×10^{10}
(c) 9.000×10^{-7} **(d)** 9.075×10^{-12}

20. Round off the following to three significant digits:
(a) 1.454×10^1 **(b)** 1.455×10^2
(c) 1.508×10^{-3} **(d)** 1.503×10^{-4}

Adding and Subtracting Measurements (Sec. PSS.4)

21. Add the following measurements and round off the answer:

(a)
$$\begin{array}{r} 1.55\text{ cm} \\ 36.15\text{ cm} \\ \underline{+\ 17.3\text{ cm}} \end{array}$$

(b)
$$\begin{array}{r} 5.0\text{ cm} \\ 16.3\text{ cm} \\ \underline{+\ 0.95\text{ cm}} \end{array}$$

22. Add the following measurements and round off the answer:

(a)
$$\begin{array}{r} 0.4\text{ g} \\ 0.44\text{ g} \\ \underline{+\ 0.444\text{ g}} \end{array}$$

(b)
$$\begin{array}{r} 15.5\text{ g} \\ 7.50\text{ g} \\ \underline{+\ 0.050\text{ g}} \end{array}$$

23. Subtract the following measurements and round off the answer:

(a)
$$\begin{array}{r} 22.10\text{ cm} \\ \underline{-\ 10.5\text{ cm}} \end{array}$$

(b)
$$\begin{array}{r} 10.0\text{ cm} \\ \underline{-\ 0.15\text{ cm}} \end{array}$$

24. Subtract the following measurements and round off the answer:

(a)
$$\begin{array}{r} 242.167\text{ g} \\ \underline{-\ 175\text{ g}} \end{array}$$

(b)
$$\begin{array}{r} 27.55\text{ g} \\ \underline{-\ 14.545\text{ g}} \end{array}$$

Multiplying and Dividing Measurements (Sec. PSS.5)

25. Multiply the following measurements and round off the answer:
(a) 1.25 cm × 0.5 cm **(b)** 2.55 cm × 1.1 cm
(c) 12.0 cm^2 × 1.00 cm **(d)** 22.1 cm^2 × 0.75 cm

26. Multiply the following measurements and round off the answer:
(a) 3.65 cm × 2.10 cm **(b)** 8.75 cm × 1.15 cm
(c) 16.5 cm × 1.7 cm **(d)** 21.1 cm × 20 cm

27. Divide the following measurements and round off the answer:
(a) 66.3 g/7.5 mL **(b)** 12.5 g/4.1 mL
(c) 42.620 g/10.0 mL **(d)** 91.235 g/10.00 mL

28. Divide the following measurements and round off the answer:
(a) 26.0 cm^2/10.1 cm **(b)** 9.95 cm^3/0.15 cm^2
(c) 131.78 cm^3/19.25 cm **(d)** 131.78 cm^3/19.2 cm

Exponential Numbers (Sec. PSS.6)

29. Express each of the following products as a power of 10:
(a) 10 × 10 × 10 **(b)** 1/10 × 1/10 × 1/10

30. Express each of the following products as a power of 10:
(a) 10 × 10 × 10 × 10 **(b)** 1/10 × 1/10 × 1/10 × 1/10

31. Express each of the following products in exponential form:
(a) 2 × 2 × 2 **(b)** 1/2 × 1/2 × 1/2

32. Express each of the following products in exponential form:
(a) 3 × 3 × 3 × 3 **(b)** 1/3 × 1/3 × 1/3 × 1/3

33. Express each of the following powers of 10 as an ordinary number:
(a) 1×10^3 **(b)** 1×10^{-7}

34. Express each of the following powers of 10 as an ordinary number:
(a) 1×10^{12} **(b)** 1×10^{-22}

35. Express each of the following ordinary numbers as a power of 10:
(a) 1,000,000,000 **(b)** 0.000 000 01

36. Express each of the following ordinary numbers as a power of 10:
(a) 100,000,000,000,000,000 **(b)** 0.000 000 000 000 001

37. Express each of the following powers of 10 as an ordinary number:
(a) 1×10^1 **(b)** 1×10^{-1}

38. Express each of the following powers of 10 as an ordinary number:
(a) 1×10^0 **(b)** 1×10^{-10}

Scientific Notation (Sec. PSS.7)

39. Express the following ordinary numbers in scientific notation:
(a) 80,916,000 **(b)** 0.000 000 015
(c) 335,600,000,000,000 **(d)** 0.000 000 000 000 927

40. Express the following ordinary numbers in scientific notation:
(a) 1,010,000,000,000,000 **(b)** 0.000 000 000 000 456
(c) 94,500,000,000,000,000 **(d)** 0.000 000 000 000 000 019 50

41. There are 26,900,000,000,000,000,000,000 oxygen molecules in 1 liter of oxygen gas at standard temperature and pressure. Express this number in scientific notation.

42. There are 26,900,000,000,000,000,000 helium atoms in 1 milliliter of helium gas at standard temperature and pressure. Express this number in scientific notation.

43. The mass of an oxygen molecule is 0.000 000 000 000 000 000 000 053 2 g. Express the mass in scientific notation.

44. The mass of a helium atom is 0.000 000 000 000 000 000 000 006 64 g. Express the mass in scientific notation.

General Exercises

45. A 10-mL pipet has an uncertainty of ±0.1 mL. Express the volume of the pipet as an ordinary number.

46. A 10-mL pipet has an uncertainty of ±0.01 mL. Express the volume of the pipet as an ordinary number.

47. The velocity of light is 186,282.397 miles per second. Round off this value to three significant digits.

48. The velocity of light is 299,792,458 meters per second. Round off this value to three significant digits.

49. Find the length of magnesium ribbon that remains after two 25.0-cm strips are cut from 255 cm of the metal ribbon.

50. Find the total mass of a 126.457-g brass cylinder and a 131.60-g bronze cylinder.

51. Convert the following exponential numbers to scientific notation.
(a) 352×10^4 **(b)** 0.191×10^{-5}

52. Convert the following exponential numbers to scientific notation.
(a) 0.170×10^2 **(b)** 0.00350×10^{-1}

Challenge Exercises

53. The mass of a neutron is 1.6749×10^{-24} g, and the mass of a proton is 1.6726×10^{-24} g. What is the total mass of a neutron and a proton?

54. The mass of a neutron is 1.6749×10^{-24} g, and the mass of a proton is 1.6726×10^{-24} g. What is the mass difference of a neutron and a proton?

55. A metric ton is defined as 1000 kg, or 2.200×10^3 lb. An English ton is defined as 2000 lb, or 2.000×10^3 lb. What is the total mass of a metric ton and English ton expressed in pounds?

56. A metric ton is defined as 1000 kg, or 2.200×10^3 lb. An English ton is defined as 2000 lb, or 2.000×10^3 lb. What is the mass difference between a metric ton and English ton expressed in pounds?

PSS Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following balances has the least uncertainty? (PSS.1)
- **(a)** platform balance, ±0.1 g
- **(b)** triple-beam balance, ±0.01 g
- **(c)** electronic balance, ±0.001 g
- **(d)** analytical balance, ±0.0001 g
- **(e)** all balances have the same uncertainty

2. State the number of significant digits in a mass of 0.050 g. (PSS.2)
- **(a)** 1
- **(b)** 2
- **(c)** 3
- **(d)** 4
- **(e)** none of the above

3. Round off the following measurement to three significant digits: 7268 mL. (PSS.3)
- **(a)** 720 mL
- **(b)** 726 mL
- **(c)** 727 mL
- **(d)** 7260 mL
- **(e)** 7270 mL

4. Add 55.5 cm to 0.50 cm and round off the answer. (PSS.4)
- **(a)** 56 cm
- **(b)** 56.0 cm
- **(c)** 56.05 cm
- **(d)** 56.1 cm
- **(e)** 57 cm

5. Multiply 7.2 cm by 3.75 cm by 1.555 cm and round off the answer. (PSS.5)
- **(a)** 41.9 cm^3
- **(b)** 41.99 cm^3
- **(c)** 42 cm^3
- **(d)** 42.0 cm^3
- **(e)** 42.00 cm^3

6. Express 5.55×10^{-2} as an ordinary number. (PSS.6)
- **(a)** 0.055 5
- **(b)** 0.005 55
- **(c)** 0.000 555
- **(d)** 5,550
- **(e)** none of the above

7. Express 55,500,000,000,000,000,000,000 in scientific notation. (PSS.7)
- **(a)** 5.55×10^{-22}
- **(b)** 5.55×10^{-23}
- **(c)** 5.55×10^{22}
- **(d)** 5.55×10^{23}
- **(e)** none of the above

Key Concept

8. When using a scientific calculator to perform multistep multiplication or division, is it better to round off after each step, or round off the final answer in the calculator display?

Critical Thinking

9. The mass of an alpha particle is 6.64×10^{-24} g, and the mass of a beta particle is 9.11×10^{-28} g. Which is heavier: an alpha particle or a beta particle?

10. An ounce of aluminum has a mass of 28.4 g, and an ounce of silver weighs 31.1 g. Which is heavier: an ounce of aluminum or an ounce of silver?

CHAPTER 2

The Metric System

> *"Never has anything more grand and more simple, more coherent in all of its parts, issued from the hand of man."*
>
> **Antoine Lavoisier, French Chemist (1743–1794)**

2
He
helium

Element 2: Helium
Helium is the second most abundant element in the universe. Helium gas is more dense than hydrogen gas, but is used in blimps and balloons because it is not flammable.

Since the beginning of civilization there has been the need to convey measurement. When exchanging goods, traders had to agree on standards for length, weight, and volume. Historically, length was often defined in terms of hands and feet because it was convenient. For instance, a digit was defined as the thickness of the index finger, a span was the width of four fingers, and a cubit was the distance from the elbow to the end of the middle finger.

The Greeks proposed the foot as a unit of length and divided it into 12 parts. The Romans in turn adopted the unit and referred to 1/12 of a foot as an *uncia*. The Anglo-Saxons pronounced *uncia* as "inch," which is how it entered into the English language.

The English system of measurement originated in 1215 with the signing of the Magna Carta. The Magna Carta attempted to bring uniform measurements to world trade and decreed, "Throughout the kingdom there shall be standard measures of wine, ale, and corn." To ensure uniform measurements, the English used reference standards. An inch was defined as the length of three barleycorns, and a pound and a gallon were similarly referenced to barley.

2.1 Basic Units and Symbols

LEARNING OBJECTIVES

- List the basic units and symbols of the metric system.
- List the prefixes for multiples and fractions of basic units.

Discontent over the lack of worldwide standards in measurement reached a peak in the late 1700s. Although the **English system** was common, it was used primarily within the British Empire. In 1790, the French government appointed a committee of scientists to investigate the possibility of a universal measuring system.

The committee spent nearly 10 years devising and agreeing on a new system. The resulting **metric system** offers simplicity and basic units. The basic unit of length is the **meter** (symbol **m**), the basic unit of mass is the **gram** (symbol **g**), the basic unit of volume is the **liter** (symbol **L**), and the basic unit of time is the **second** (symbol **s**). The basic units and symbols of the metric system are shown in Table 2.1.

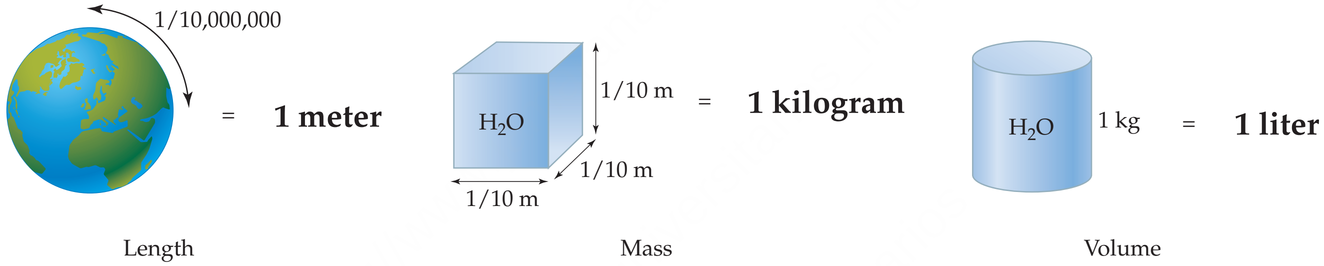
TABLE 2.1 The Metric System

Physical Quantity	Basic Unit	Symbol
length	meter*	m
mass	gram	g
volume	liter*	L
time	second	s

*The U.S. Metric Association recommends the spellings “meter” and “liter.” All other English-speaking nations use the spellings “metre” and “litre.”

The metric system is simple and coherent for two reasons: First, it uses a single basic unit for each quantity measured, that is, meter, gram, and liter. Second, it is a decimal system that uses prefixes to enlarge or reduce a basic unit. For instance, a kilometer is 1000 meters and a centimeter is 1/100 meter.

The metric committee defined a meter, kilogram, and liter as follows. A meter was defined as one ten-millionth of the distance from the North Pole to the equator. A kilogram was defined as the mass of a cube of water one-tenth of a meter on a side. A liter was equal to the volume occupied by a kilogram of water (at 4 °C).



▲ **The Original Metric References** A meter was originally defined as one ten-millionth of the distance from the North Pole to the equator. A kilogram was equal to the mass of a cube of water one-tenth of a meter on each side. A liter was equal to the volume of one kilogram of water at 4 °C.

In addition, the metric committee members made reference standards for the meter, kilogram, and liter. They cast a platinum bar (alloyed with 10% iridium metal) as the reference standard for one meter. They cast a solid platinum cylinder as the reference for one kilogram and a platinum container as the standard for one liter. These platinum castings became the reference standards for world measurement. In 1875 all the major nations of the world, including the United States, signed the Treaty of the Meter. Duplicates of the original castings were made and distributed to all member nations. The original casting of the standard reference meter is shown in Figure 2.1.

▶ **Figure 2.1** **The Meter** The official reference meter was cast from platinum and iridium metals in 1790. The original casting is kept at the International Bureau of Weights and Measures outside Paris, France.

Metric Prefixes

The United States uses the dollar ($1.00) as the basic unit of currency. In our monetary system we move the decimal to express a multiple or fraction of a dollar (for example, a 10-dollar bill equals 10 one-dollar bills, 100 dimes, or 1000 pennies). Similarly, the metric system is a decimal system, but it uses a prefix to express a multiple or a fraction of a basic unit. A metric prefix increases or decreases a basic unit by a factor of 10. For instance, the prefix *kilo-* increases a basic unit by 1000, whereas the prefix *centi-* decreases a basic unit by 100. Thus, a kilometer is 1000 m, and a centimeter is 0.01 m. Table 2.2 lists some common prefixes for basic metric units.

TABLE 2.2 Metric Prefixes

Prefix	Symbol	Multiple/Fraction
exa-	E	$1{,}000{,}000{,}000{,}000{,}000{,}000 = 1 \times 10^{18}$
peta-	P	$1{,}000{,}000{,}000{,}000{,}000 = 1 \times 10^{15}$
tera-	T	$1{,}000{,}000{,}000{,}000 = 1 \times 10^{12}$
giga-	G	$1{,}000{,}000{,}000 = 1 \times 10^{9}$
mega-	M	$1{,}000{,}000 = 1 \times 10^{6}$
kilo-	k	$1000 = 1 \times 10^{3}$
Basic unit: meter, gram, liter, second		
deci-	d	$0.1 = 1 \times 10^{-1}$
centi-	c	$0.01 = 1 \times 10^{-2}$
milli-	m	$0.001 = 1 \times 10^{-3}$
micro-	μ*	$0.000\ 001 = 1 \times 10^{-6}$
nano-	n	$0.000\ 000\ 001 = 1 \times 10^{-9}$
pico-	p	$0.000\ 000\ 000\ 001 = 1 \times 10^{-12}$
femto-	f	$0.000{,}000{,}000{,}000{,}001 = 1 \times 10^{-15}$
atto-	a	$0.000{,}000{,}000{,}000{,}000{,}001 = 1 \times 10^{-18}$

*The Greek letter mu (μ) is the symbol for micro.

The names of metric units can be abbreviated using symbols. We can illustrate this by writing symbols for the following metric units: kilometer (km), milligram (mg), microliter (μL), and nanosecond (ns). Example Exercise 2.1 further illustrates how to write names of metric units and symbols.

EXAMPLE EXERCISE 2.1 Metric Basic Units and Prefixes

Give the symbol for each of the following metric units and state the quantity measured by each unit:

(a) gigameter (b) kilogram
(c) centiliter (d) microsecond

Solution

We compose the symbol for each unit by combining the prefix symbol and the basic unit symbol. If we refer to Tables 2.1 and 2.2, we have the following:

(a) Gm, length (b) kg, mass
(c) cL, volume (d) μs, time

Practice Exercise

Give the symbol for each of the following metric units and state the quantity measured by each unit:

(a) nanosecond (b) milliliter
(c) decigram (d) megameter

Answers:

(a) ns, time (b) mL, volume
(c) dg, mass (d) Mm, length

Concept Exercise

What is the basic unit of length, mass, and volume in the metric system?

Answers: See Appendix G, 2.1.

Note The metric system is sometimes referred to as the International System of measurement. Although the **International System** (symbol **SI**) is similar to the metric system, SI is much more comprehensive. In this text, we will use the metric system and utilize SI units only in special instances.

2.2 Metric Conversion Factors

LEARNING OBJECTIVES

- Show the unit equation for a basic metric unit and a prefix unit.
- Show the two unit factors derived from a metric unit equation.

In the next section we will tackle our single most important task: problem solving. We will learn a simple but powerful method called unit analysis. To use the unit analysis method of problem solving, we must understand equivalent relationships and unit equations.

A **unit equation** relates two quantities that are equal. As an example, let's find a unit equation for kilometers and meters. Since the metric prefix *kilo-* means 1000 basic units, 1 kilometer equals 1000 meters. We can write the **exact equivalent** as 1 kilometer = 1000 meters, or simply 1 km = 1000 m.

A centimeter is 1/100 of a meter, and we can write the relationship as 1 cm = 0.01 m. However, working with fractions is more confusing than working with multiples of basic units. Fortunately, we can avoid decimal fractions by simply restating the relationship. Because there are 100 cm in a meter, we can write the exact equivalent as 1 m = 100 cm. Example Exercise 2.2 further illustrates how to write unit equations for metric relationships.

EXAMPLE EXERCISE 2.2 Metric Unit Equations

Complete the unit equation for each of the following exact metric equivalents:

(a) 1 Mm = ? m (b) 1 kg = ? g
(c) 1 L = ? dL (d) 1 s = ? ns

Solution

We can refer to Table 2.2 as necessary.

(a) The prefix *mega-* means 1,000,000 basic units; thus, 1 Mm = 1,000,000 m.
(b) The prefix *kilo-* means 1000 basic units; thus, 1 kg = 1000 g.
(c) The prefix *deci-* means 0.1 of a basic unit; thus, 1 L = 10 dL.
(d) The prefix *nano-* means 0.000 000 001 of a basic unit; thus, 1 s = 1,000,000,000 ns.

Alternatively, we can express the unit equation using exponential numbers:

(a) 1 Mm = 1×10^6 m (b) 1 kg = 1×10^3 g
(c) 1 L = 1×10^1 dL (d) 1 s = 1×10^9 ns

Practice Exercise

Complete the unit equation for each of the following exact metric equivalents:

(a) 1 Gm = ? m (b) 1 g = ? cg
(c) 1 L = ?μL (d) 1 s = ? ms

Answers:

(a) 1 Gm $= 1 \times 10^9$ m (b) 1 g $= 1 \times 10^2$ cg

(c) 1 L $= 1 \times 10^6 \mu$ L (d) 1 s $= 1 \times 10^3$ ms

Concept Exercise

How many significant digits are in the following unit equation: 1 m = 1000 mm?

Answers: See Appendix G, 2.2.

Writing Metric Unit Factors

Once we are able to write unit equations, we can easily generate two unit factors. A **unit factor** is the ratio of the two equivalent quantities. That is, the quantity in the numerator is equal to the quantity in the denominator.

It is helpful to consistently write the larger metric unit first and the smaller metric unit second in a unit equation; for example, 1 m = 100 cm, 1 kg = 1000 g, and 1 L = 1000 mL.

The numerator and the denominator are equivalent, so we can generate a second unit factor by inverting the ratio and writing the **reciprocal**. As an example, consider the unit equation 1 m = 100 cm. The two associated unit factors are

$$\frac{1\ \text{m}}{100\ \text{cm}} \quad \text{and} \quad \frac{100\ \text{cm}}{1\ \text{m}}$$

Example Exercise 2.3 further illustrates how to write unit factors corresponding to a given metric relationship:

Helpful Hint
Unit Equations

A consistent approach in applying the unit analysis method eliminates mistakes. It is helpful to always write the larger metric unit first, and the smaller unit second, in the unit equation. For example, write 1 m = 100 cm, 1 kg = 1000 g, and 1 L = 1000 mL.

EXAMPLE EXERCISE 2.3 Metric Unit Factors

Write two unit factors for each of the following metric relationships:

(a) kilometers and meters (b) grams and decigrams

Solution

We start by writing the unit equation to generate the two unit factors.

(a) The prefix *kilo*- means 1000 basic units; thus, 1 km = 1000 m. The two unit factors are

$$\frac{1\ \text{km}}{1000\ \text{m}} \quad \text{and} \quad \frac{1000\ \text{m}}{1\ \text{km}}$$

(b) The prefix *deci*- means 0.1 basic unit; thus, 1 g = 10 dg. The two unit factors are

$$\frac{1\ \text{g}}{10\ \text{dg}} \quad \text{and} \quad \frac{10\ \text{dg}}{1\ \text{g}}$$

Practice Exercise

Write two unit factors for each of the following metric relationships:

(a) liters and milliliters (b) megaseconds and seconds

Answers:

(a) 1 L/1000 mL and 1000 mL/1 L (b) 1 Ms/1,000,000 s and 1,000,000 s/1 Ms

Concept Exercise

How many significant digits are in the following unit factor: 1 L/1000 mL?

Answers: See Appendix G, 2.3.

2.3 Metric–Metric Conversions

LEARNING OBJECTIVE

- Express a given metric measurement using a different metric prefix.

Science has the reputation for tackling difficult problems. Fortunately, you are about to learn a powerful technique for solving problems. It is called the **unit analysis method**. This approach to problem solving is also referred to as *dimensional analysis*, or the *factor-label method*. Unit analysis cannot solve every problem, but it is very effective for the problems we encounter in introductory chemistry.

Solving problems by unit analysis is a simple three-step process. Step 1: Read the problem and determine the unit required in the answer. Step 2: Analyze the problem and determine the given value that is related to the answer. Step 3: Write one or more unit factors to convert the unit in the given value to the unit in the answer. You should follow these three steps consistently and systematically when using the unit analysis method of problem solving.

Applying the Unit Analysis Method

Step 1: Write down the unit asked for in the answer.

Step 2: Write down the given value related to the answer.

Step 3: Apply unit factor(s) to convert the unit in the given value to the unit in the answer.

Let's begin with a simple problem and find the mass in grams of a 325 mg aspirin tablet. *First*, we write down the unit asked for in the answer (g). *Second*, we write down the related given value (325 mg). *Third*, we apply a unit factor using the following format:

$$\begin{array}{ccccc} 325\text{ mg} & \times & \dfrac{\text{unit}}{\text{factor}} & = & \text{g} \\ \text{(Step 2)} & & \text{(Step 3)} & & \text{(Step 1)} \end{array}$$

The unit equation that relates Steps 1 and 2 is $1\text{ g} = 1000\text{ mg}$. Therefore, the two associated unit factors are

$$\frac{1\text{ g}}{1000\text{ mg}} \quad \text{and} \quad \frac{1000\text{ mg}}{1\text{ g}}$$

To cancel the unit in the given value (325 mg), we should select the unit factor 1 g/1000 mg. After substituting, we have

$$325\,\cancel{\text{mg}} \times \frac{1\text{ g}}{1000\,\cancel{\text{mg}}} = 0.325\text{ g}$$

The given value has three significant digits, and the unit factor is an exact equivalent. Thus, the answer has three significant digits.

Example Exercises 2.4 and 2.5 further illustrate the unit analysis method of problem solving for metric–metric conversions.

EXAMPLE EXERCISE 2.4 Metric–Metric Conversion

A hospital has 125 deciliter bags of blood plasma. What is the volume of plasma expressed in milliliters?

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: mL

STEP 2: What given value is related to the answer?

Step 2: 125 dL

STEP 3: What unit factor(s) should we apply? Given that 1 L = 10 dL, and 1 L = 1000 mL, the two pairs of unit factors are

Step 3: $\frac{1\text{ L}}{10\text{ dL}}$ or $\frac{10\text{ dL}}{1\text{ L}}$

$\frac{1\text{ L}}{1000\text{ mL}}$ or $\frac{1000\text{ mL}}{1\text{ L}}$

Unit Analysis Map

given value

125 dL

unit factor 1

$\times \frac{1\text{ L}}{10\text{ dL}}$ or $\frac{10\text{ dL}}{1\text{ L}}$

unit factor 2

$\times \frac{1\text{ L}}{1000\text{ mL}}$ or $\frac{1000\text{ mL}}{1\text{ L}}$

unit in answer

= ? mL

Solution

We apply the unit factor 1 L/10 dL to cancel deciliters ($\cancel{\text{dL}}$), and 1000 mL/1 L to cancel liters ($\cancel{\text{L}}$).

$$125\,\cancel{\text{dL}} \times \frac{1\,\cancel{\text{L}}}{10\,\cancel{\text{dL}}} \times \frac{1000\text{ mL}}{1\,\cancel{\text{L}}} = 1.25 \times 10^{4}\text{ mL}$$

The given value, 125 dL, limits the answer to three significant digits. Each unit factor is derived from an exact equivalent, so neither affects the number of significant digits in the answer.

Practice Exercise

A dermatology patient is treated with ultraviolet light having a wavelength of 375 nm. What is the wavelength expressed in centimeters?

Answers: 0.0000375 cm (3.75×10^{-5} cm)

Concept Exercise

Express the volume of a cube 1 cm on a side in milliliters.

Answers: See Appendix G, 2.4.

▶ **Blood Plasma** A blood plasma bag has a volume of 1 deciliter, that is, 0.1 liter.

EXAMPLE EXERCISE 2.5 Metric–Metric Conversion

The mass of Earth is 5.98×10^{24} kg. What is the mass expressed in megagrams?

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: Mg

STEP 2: What given value is related to the answer?

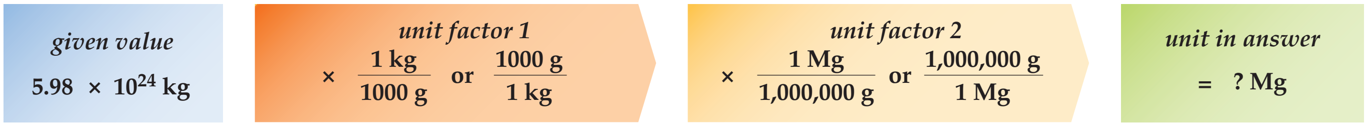
Step 2: 5.98×10^{24} kg

STEP 3: What unit factor(s) should we apply? Given that 1 kg = 1000 g, and 1 Mg = 1,000,000 g, the two pairs of unit factors are

Step 3: $\frac{1\text{ kg}}{1000\text{ g}}$ or $\frac{1000\text{ g}}{1\text{ kg}}$

$\frac{1\text{ Mg}}{1{,}000{,}000\text{ g}}$ or $\frac{1{,}000{,}000\text{ g}}{1\text{ Mg}}$

Unit Analysis Map



Solution

We apply the unit factor 1000 g/1 kg to cancel kilograms ($\cancel{\text{kg}}$), and 1 Mg/1,000,000 g to cancel grams ($\cancel{\text{g}}$).

$$5.98 \times 10^{24}\ \cancel{\text{kg}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\text{ Mg}}{1{,}000{,}000\ \cancel{\text{g}}} = 5.98 \times 10^{21}\text{ Mg}$$

The given value limits the answer to three significant digits. Because each unit factor is derived from an exact equivalent, neither affects the number of significant digits in the answer.

Earth

◀ **Earth** The mass of Earth is about 6,000,000,000,000,000,000,000,000 kilograms.

Practice Exercise

Light travels through the universe at a velocity of 3.00×10^{10} cm/s. How many gigameters does light travel in one second?

Answers: 0.300 Gm (3.00×10^{-1} Gm)

Concept Exercise

How many significant digits are in the following unit factor: 1 g/1000 mg?

Answers: See Appendix G, 2.5.

LEARNING OBJECTIVES

- Memorize the metric equivalent for inch, pound, quart, and second.
- Express a given measurement in metric units or English units.

2.4 Metric–English Conversions

Despite the superiority of the metric system, its adoption has been resisted from its inception. When the French asked the Americans and the British for their cooperation in creating a worldwide system of measurement, both countries declined. The initial objection, interestingly, was political rather than scientific. The meter was defined by a longitude passing through Dunkirk, France, and the original metric standards were kept in a special vault outside of Paris.

In 1975, President Gerald Ford signed an official metric conversion act. Thus, the United States became the last major nation in the world to formally adopt the metric system. Although the metric system is now taught in public schools, the United States is making slow progress in achieving full compliance.

In science, we regularly record measurements in metric units. In everyday activities, you may be more familiar with inches, pounds, and quarts. Therefore, let's compare the two systems to gain a practical awareness of metric sizes. Table 2.3 contrasts some common English and metric units.

Which race do you think is run faster: 100 meters or 100 yards? Obviously, it is the shorter race distance that is run faster. Given that 1 yd = 0.914 m, we can calculate the distance in yards for a 100.0-meter race.

TABLE 2.3 English–Metric Equivalents

Quantity	English Unit	Metric Equivalent
length	1 inch (in.)	1 in. = 2.54 cm
mass	1 pound (lb)	1 lb = 454 g
volume	1 quart (qt)	1 qt = 946 mL
time	1 second (sec)	1 sec = 1.00 s

When applying these metric equivalents, assume three significant digits. Because the metric and English systems have different reference standards, these are not exact equivalents; however, the U.S. Bureau of Weights and Measures has redefined 1 in. as exactly equal to 2.54 cm.

First, we write down the unit asked for in the answer (yd). *Second*, we write down the related given value (100.0 m). *Third*, we apply a unit factor using the following format:

$$\begin{array}{ccccc} 100.0\text{ m} & \times & \dfrac{\text{unit}}{\text{factor}} & = & \text{yd} \\ \text{(Step 2)} & & \text{(Step 3)} & & \text{(Step 1)} \end{array}$$

The unit equation that relates Steps 1 and 2 is 1 yd = 0.914 m. Therefore, the two unit factors are

$$\frac{1\text{ yd}}{0.914\text{ m}} \quad \text{and} \quad \frac{0.914\text{ m}}{1\text{ yd}}$$

To cancel the unit in the given value (100.0 m), we select the unit factor (1 yd/0.914 m). After substituting, we have

$$100.0\ \cancel{\text{m}} \times \frac{1\text{ yd}}{0.914\ \cancel{\text{m}}} = 109\text{ yd}$$

Although the given value (100.0 m) has four significant digits, the unit factor has only three because 1 yd = 0.914 m is not an exact equivalent. Thus, the answer is rounded to three significant digits. Because 100 meters is a longer distance than 100 yards, the 100-yard race is run faster. In fact, the world record for 100 meters is more than 9 seconds, whereas the world record for 100 yards is less than 9 seconds!

Example Exercise 2.6 and 2.7 further illustrate the unit analysis method of problem solving.

EXAMPLE EXERCISE 2.6 Metric–English Conversion

A can of soda contains 12.0 fl oz (1 quart = 32 fluid ounces). What is the volume of soda in milliliters?

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**mL**
STEP 2: What given value is related to the answer?	**Step 2:**	**12.0 fl oz**
STEP 3: What unit factor(s) should we apply? Given that 1 qt = 32 fl oz; and in Table 2.3 we find that 1 qt = 946 mL. The two pairs of unit factors are	**Step 3:**	$\dfrac{32\text{ fl oz}}{1\text{ qt}}$ **or** $\dfrac{1\text{ qt}}{32\text{ fl oz}}$ $\dfrac{1\text{ qt}}{946\text{ mL}}$ **or** $\dfrac{946\text{ mL}}{1\text{ qt}}$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit in answer*
12 fl oz	$\times\ \dfrac{32\text{ fl oz}}{1\text{ qt}}$ **or** $\dfrac{1\text{ qt}}{32\text{ fl oz}}$	$\times\ \dfrac{1\text{ qt}}{946\text{ mL}}$ **or** $\dfrac{946\text{ mL}}{1\text{ qt}}$	**= ? mL**

Solution

We apply the unit factor 1 qt/32 fl oz to cancel ounces ($\cancel{\text{fl oz}}$), and 946 mL/1 qt to cancel quarts ($\cancel{\text{qt}}$).

$$12.0\ \cancel{\text{fl oz}} \times \frac{1\ \cancel{\text{qt}}}{32\ \cancel{\text{fl oz}}} \times \frac{946\text{ mL}}{1\ \cancel{\text{qt}}} = 355\text{ mL}$$

The given value, 12.0 fl oz, limits the answer to three significant digits. Unit factor 1 has no effect as it is an exact equivalent, and unit factor 2 has three significant digits.

Practice Exercise

A can of soda contains 12.0 fl oz. What is the volume in liters (given that 1 L = 1000 mL)?

Answers: 0.355 L

Concept Exercise

How many significant digits are in the following unit equation: 1 L = 1000 mL?

Answers: See Appendix G, 2.6.

A CLOSER LOOK Metric Labels

Q: Does a 12-ounce can of soda have a volume of 354 mL or 355 mL?

The soft drink industry is a billion dollar business. Since the 1980s, Coca-Cola and Pepsi-Cola have waged the so-called "Cola Wars" to acquire more market share. Originally, soft drink cans were labeled in English units; for example, 12 fluid ounces. After the United States adopted the metric system, products were given dual labels; for example, 1 lb (454 g), and 12 fl oz (355 mL).

Initially, Pepsi-Cola labeled the volume of 12-oz cans as 354 ml, and Coca-Cola labeled their cans as 355 mL. Pepsi-Cola realized they were potentially losing market share because consumers would notice that a Coke can contained more cola than a Pepsi can. Soon thereafter, Pepsi changed their labels to read 12 FL OZ (355 mL).

So how did the label discrepancy arise? In performing an English-metric conversion, the results are as follows:

$$12\ \cancel{\text{fl oz}} \times \frac{1\ \cancel{\text{qt}}}{32\ \cancel{\text{fl oz}}} \times \frac{946\ \text{mL}}{1\ \cancel{\text{qt}}} = 354.75\ \text{mL}$$

Apparently, Pepsi rounded off the value for the volume of soda to 354 ml. Assuming three significant digits, the correct value for volume is 355 mL.

▲ **Pepsi** 12 fl oz 354 ml.

▲ **Pepsi** 12 fl oz 355 mL.

A: Rounded correctly to three significant figures, 12-fluid ounce cans of Coke and Pepsi both contain the same amount of cola, 355 mL.

EXAMPLE EXERCISE 2.7 Metric–English Conversion

If a tennis ball weighs 2.0 oz, what is the mass of the tennis ball in grams?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	g
STEP 2:	What given value is related to the answer?	**Step 2:**	2.0 oz

STEP 3: What unit factor(s) should we apply? Given that 1 lb = 16 oz, and that in Table 2.3 we find that 1 lb = 454 g, the two pairs of unit factors are as follows:

Step 3:

$$\frac{1\ \text{lb}}{16\ \text{oz}} \quad \text{or} \quad \frac{16\ \text{oz}}{1\ \text{lb}}$$

$$\frac{1\ \text{lb}}{454\ \text{g}} \quad \text{or} \quad \frac{454\ \text{g}}{1\ \text{lb}}$$

Unit Analysis Map

given value
2.0 oz

unit factor 1
$$\times \frac{1\ \text{lb}}{16\ \text{oz}} \quad \text{or} \quad \frac{16\ \text{oz}}{1\ \text{lb}}$$

unit factor 2
$$\times \frac{1\ \text{lb}}{454\ \text{g}} \quad \text{or} \quad \frac{454\ \text{g}}{1\ \text{lb}}$$

unit in answer
= ? g

Solution

We apply the unit factor 1 lb/16 oz to cancel ounces ($\cancel{\text{oz}}$), and 454 g/1 lb to cancel pounds ($\cancel{\text{lb}}$).

$$2.0\ \cancel{\text{oz}} \times \frac{1\ \cancel{\text{lb}}}{16\ \cancel{\text{oz}}} \times \frac{454\ \text{g}}{1\ \cancel{\text{lb}}} = 57\ \text{g}$$

The given value, 2.0 oz, limits the answer to two significant digits. Unit factor 1 has no effect as it is derived from an exact equivalent, and unit factor 2 has three significant digits.

Practice Exercise

If a tennis ball has a diameter of 2.5 inches, what is the diameter in millimeters?

Answers: 64 mm

Concept Exercise

How many significant digits are in the following unit factor: 1 kg/2.20 lb?

Answers: See Appendix G, 2.7.

CHEMISTRY CONNECTION The Olympics

Q: Which Olympic race is nearly equal in length to a quarter mile, that is, 440 yards?

The Olympics began in 776 B.C. with a single event, a 200-yard dash called a *stadion*. In time, a 2-stadia dash and a 24-stadia event were added, and wrestling appeared in 708 B.C. Eventually, participation in the games declined, and the ancient Olympics were discontinued in A.D. 394.

In the summer of 1896, the modern Olympic Games were initiated and held in Greece. The games attracted about 500 athletes from 13 nations. Since then, the games have been held every four years in various cities around the world.

The games have steadily increased in number of participants, as well as in number of events. In 1896, 311 men from 13 nations competed, and the United States won gold medals in 9 of the 12 events. In 2004, over 10,000 men and women athletes from 197 nations competed.

U.S. track and field competitions have been conducted using the English system of measurement. That is, American runners compete at distances of 100 yards, 440 yards, and 1 mile. In Olympic competitions, events are conducted using the metric system owing to its international acceptance.

▲ **Olympic Marathon** Marathon runners at the 2012 Olympic Summer Games in London.

Olympic runners compete at comparable distances such as 100, 400, and 1500 meters. The marathon, however, is an exception. Instituted in 1896, a marathon was to correspond to the distance run by the Greek messenger Pheidippides, who carried news of the Athenian victory over the Persians from the plains of Marathon to Athens in 490 B.C., after which, he promptly died. Today, a marathon is defined as a distance of 26 miles.

In 1924, the Olympic Winter Games were introduced at Chamonix, France, and have been held four times in the United States, twice at Lake Placid, once at Squaw Valley, and in 2002 at Salt Lake City. The winter games feature skating, skiing, bobsledding, luge, and ice hockey. In Nordic cross-country skiing, the races are 10, 20, 30, and 50 kilometers; in ski jumping, the ramps are 70 and 90 meters. Thus, Olympic competitions—winter and summer—employ the metric system of measurement.

A: The 400-meter Olympic race is nearly identical to 440 yards, and the world records for the two races are within tenths of a second.

2.5 The Percent Concept

LEARNING OBJECTIVES

- Express a quantity in a sample as a percent.
- Apply percent as a unit factor.

A **percent** (symbol **%**) expresses the amount a single quantity compared to an entire sample. For example, a dime is 10% of a dollar, and a quarter is 25% of a dollar. The percent symbol (%) is derived from the ratio sign (/) and the two zeros in 100.

A percent expresses a ratio and can be referred to as parts per 100 parts. To calculate a percent, divide a single quantity by the total sample, and multiply that ratio by 100%.

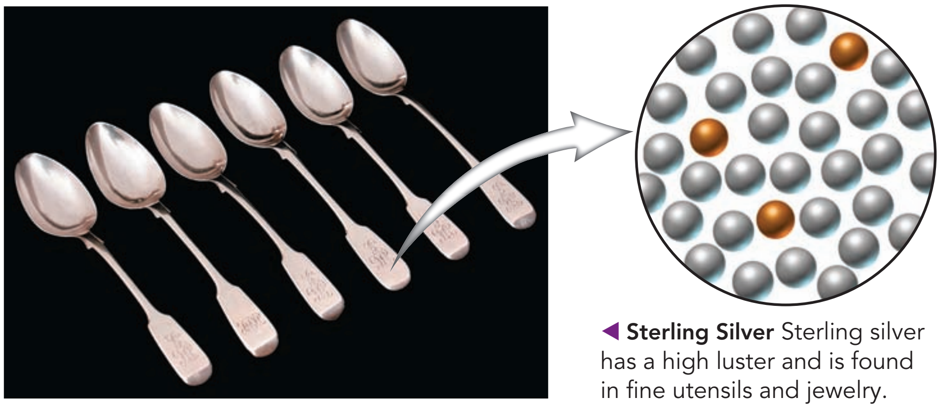
$$\frac{\text{one quantity}}{\text{total sample}} \times 100\% = \%$$

As an example, consider that 5¢ coins minted after 1971 are composed of the two metals copper and nickel. If a 5¢ coin contains 3.80 g copper and 1.27 g nickel, what is the percent of copper in the coin? We can calculate the percent of copper as follows:

$$\frac{3.80\text{ g}}{(3.80 + 1.27)\text{ g}} \times 100\% = 75.0\%$$

EXAMPLE EXERCISE 2.8 The Percent Concept

Sterling silver contains silver and copper metals. If a sterling silver chain contains 18.5 g of silver and 1.5 g of copper, what is the percent of silver?



Sterling Silver Sterling silver has a high luster and is found in fine utensils and jewelry.

Strategy Plan

STEP 1:	What is asked for in the answer?	**Step 1:**	%
STEP 2:	What given value is related to the answer?	**Step 2:**	**18.5 g and 1.5 g**
STEP 3:	What unit factor should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

To find percent, we compare the mass of silver metal to the total mass of the silver and copper in the chain, and multiply by 100%.

$$\frac{18.5\ \cancel{g}}{(18.5 + 1.5)\ \cancel{g}} \times 100\% = 92.5\%$$

Genuine sterling silver is cast from 92.5% silver and 7.5% copper. If you carefully examine a piece of sterling silver, you may see the jeweler's notation .925, which indicates the item is genuine sterling silver.

Practice Exercise

A 14-karat gold ring contains 7.45 g of gold, 2.66 g of silver, and 2.66 g of copper. Calculate the percent of gold in the 14-karat ring.

Answers: 58.3%

Concept Exercise

If a gold alloy contains 20% silver and 5% copper, what is the percent of gold in the alloy?

Answers: See Appendix G, 2.8.

Percent Unit Factors

A percent can be expressed as parts per 100 parts. For instance, 10% and 25% correspond to 10 and 25 parts per 100 parts, respectively. Thus, we can write a percent as a ratio fraction; that is, we can write 10% as 10/100, and 25% as 25/100.

Let's try a type of problem in which the percent is given. *Apollo 11* traveled 394,000 km from Earth to land the first astronaut on the Moon. How many kilometers had the astronauts traveled when they had gone 15.5% of the distance? By definition, 15.5% means 15.5 parts per 100 parts. In this example, 15.5% is 15.5 km per 100 km. Thus, we can write the ratio as a unit factor and apply the unit analysis method.

$$394{,}000\ \cancel{km} \times \frac{15.5\ km}{100\ \cancel{km}} = 61{,}100\ km$$

In this problem, notice that we used percent as a unit factor. Example Exercises 2.9 and 2.10 further illustrate the application of percent as a unit factor.

▲ **Moon** The Moon and Earth have a similar composition.

EXAMPLE EXERCISE 2.9 Percent as a Unit Factor

The Moon and Earth have a similar composition and each contains 4.70% iron, which supports the theory that the Moon and Earth were originally a single planet. What is the mass of iron in a lunar sample that weighs 235 g?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**g iron**
STEP 2:	What given value is related to the answer?	**Step 2:**	**235 g sample**
STEP 3:	What unit factor should we apply? From the definition of percent, 4.70 g iron = 100 g sample; the two unit factors are 4.70 g iron/100 g sample, and its reciprocal 100 g sample/4.70 g iron.	**Step 3:**	$\frac{4.70\ g\ iron}{100\ g\ sample}$ **or** $\frac{100\ g\ sample}{4.70\ g\ iron}$

Unit Analysis Map

given value
235 g sample

unit factor
$$\times \frac{4.70 \text{ g iron}}{100 \text{ g sample}} \text{ or } \frac{100 \text{ g sample}}{4.70 \text{ g iron}}$$

unit in answer
= ? g iron

Solution

We should apply the unit factor 4.70 g iron/100 g sample to cancel grams sample (g sample), which appears in the denominator.

$$235 \cancel{\text{g sample}} \times \frac{4.70 \text{ g iron}}{100 \cancel{\text{g sample}}} = 11.0 \text{ g iron}$$

The given value and unit factor each have three digits and limit the answer to three significant digits.

Practice Exercise

A Moon sample is found to contain 7.50% aluminum. What is the mass of the lunar sample if the amount of aluminum is 5.25 g?

Answers: 70.0 g sample

Concept Exercise

Water is 11.2% hydrogen by mass. What two unit factors express the percent hydrogen in water?

Answers: See Appendix G, 2.9.

EXAMPLE EXERCISE 2.10 Percent as a Unit Factor

A 5¢ coin minted in 2016 is an alloy, 75.0% copper and 25.0% nickel. If the coin has a mass of 5.000 g, what is the mass of nickel in the coin?

Strategy Plan

STEP 1: What unit is asked for in the answer? **Step 1:** g nickel

STEP 2: What given value is related to the answer? **Step 2:** 5.000 g coin

STEP 3: What unit factor should we apply? From the definition of percent, 25.0 g nickel = 100 g coin; the two unit factors are 25.0 g nickel/100 g coin, and its reciprocal 100 g coin/25.0 g nickel. **Step 3:**

$$\frac{25.0 \text{ g nickel}}{100 \text{ g coin}} \text{ or } \frac{100 \text{ g coin}}{25.0 \text{ g nickel}}$$

Unit Analysis Map

given value
5.000 g coin

unit factor
$$\times \frac{25.0 \text{ g nickel}}{100 \text{ g coin}} \text{ or } \frac{100 \text{ g coin}}{25.0 \text{ g nickel}}$$

unit in answer
= ? g nickel

Solution

We can cancel grams coin (g coin), by applying the unit factor 25.0 g nickel/100 g coin.

$$5.000 \cancel{\text{g coin}} \times \frac{25.0 \text{ g nickel}}{100 \cancel{\text{g coin}}} = 1.25 \text{ g nickel}$$

The given value has four significant digits and the unit factor has only three digits, thus we round off the answer to three significant digits.

Practice Exercise

A 25¢ coin minted in 2015 is an alloy, which is 91.67% copper and 8.33% nickel. If the quarter contains 5.20 g of copper, what is the mass of the coin?

Answers: 5.67 g sample

Concept Exercise

Water is 11.2% hydrogen by mass. What two unit factors express the percent hydrogen in water?

Answers: See Appendix G, 2.10.

2.6 Volume by Calculation

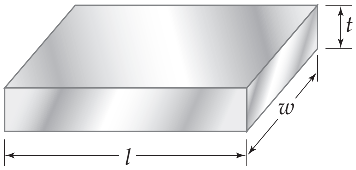
LEARNING OBJECTIVES

- Solve problems that relate length, width, thickness, and volume of a rectangular solid.
- Express a given volume in milliliters, cubic centimeters, or cubic inches.

Have you ever wondered how a building contractor arrives at the costs for a new home? For example, how does one estimate the cost of pouring a foundation? The cost is based on the volume of concrete that is poured. If the foundation is rectangular, the volume of concrete is found by multiplying length (l) times width (w) times thickness (t). We can express this relationship by the following equation:

$$l \times w \times t = \text{volume}$$

Before calculating a volume, we must express each of the measurements in the same units. If we express length, width, and thickness in centimeters, the volume of the solid has the units of **cubic centimeters** (symbol cm^3). For example, if a rectangular solid measures 3 cm by 2 cm by 1 cm, we find the volume is 6 cm^3. Example Exercise 2.11 further illustrates how to determine the volume of a rectangular solid by calculation.



▲ **Volume by Calculation** The volume of a rectangular solid is equal to its length (l) times width (w) times thickness (t) or height.

EXAMPLE EXERCISE 2.11 Volume Calculation for a Rectangular Solid

If a stainless steel rectangular solid measures 5.55 cm long, 3.75 cm wide, and 2.25 cm thick, what is the volume in cubic centimeters?

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**cm^3**
STEP 2: What given value is related to the answer?	**Step 2:**	**5.55 cm, 3.75 cm, and 2.25 cm**
STEP 3: What unit factor(s) should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

We can calculate the volume of the rectangular solid by multiplying length times width times thickness: $l \times w \times t$.

$$5.55 \text{ cm} \times 3.75 \text{ cm} \times 2.25 \text{ cm} = 46.8 \text{ cm}^3$$

The answer is rounded off to three significant digits because each given value has three significant digits.

Practice Exercise

If a rectangular brass solid measures 52.0 mm by 25.0 mm by 15.0 mm, what is the volume in cubic millimeters?

Answers: 19,500 mm^3 ($1.95 \times 10^4 \text{ mm}^3$)

Concept Exercise

Express the volume of a cube 10 cm on a side in liters.

Answers: See Appendix G, 2.11.

Let's now consider a rectangular solid with a length of 4 cm, a width of 3 cm, and a volume of 24 cm^3. Given the length, width, and volume, we can calculate the thickness of the rectangular solid by dividing the volume by its length and its width; that is, we divide 24 cm^3 by 4 cm and 3 cm. In this example, the thickness of the solid is 2 cm. Example Exercise 2.12 illustrates finding the thickness of rectangular solid.

EXAMPLE EXERCISE 2.12 Thickness Calculation for a Rectangular Solid

A sheet of aluminum foil measures 25.0 mm by 10.0 mm, and the volume is 3.75 mm^3. What is the thickness of the foil in millimeters?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**mm**
STEP 2:	What given value is related to the answer?	**Step 2:**	**3.75 mm^3, 25.0 mm, and 10.0 mm**
STEP 3:	What unit factor(s) should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

We can calculate the thickness of the foil by dividing the volume by length and width. Because the unit of volume is mm^3, we obtain the thickness in mm by unit cancellation.

$$3.75\ \text{mm}^3 \times \frac{1}{25.0\ \text{mm}} \times \frac{1}{10.0\ \text{mm}} = \frac{3.75\ \text{mm} \times \cancel{\text{mm}} \times \cancel{\text{mm}}}{25.0\ \cancel{\text{mm}} \times 10.0\ \cancel{\text{mm}}} = 0.0150\ \text{mm}$$

The answer is rounded off to three significant digits because each given value has three significant digits.

Practice Exercise

A sheet of aluminum foil measures 35.0 cm by 25.0 cm, and the volume is 1.36 cm^3. What is the thickness of the foil in centimeters?

Answers: 0.00155 cm (1.55×10^{-3} cm)

Concept Exercise

Which of the following is the greatest thickness: 1 mm, 0.1 cm, or 0.001 m?

Answers: See Appendix G, 2.12.

▲ **Aluminum Foil** A thin sheet of aluminum foil.

Volumes of Solids, Liquids, and Gases

In the metric system, the basic unit of liquid volume is the liter. A liter is equivalent to the volume occupied by a cube exactly 10 cm on a side.

A cube is a regular solid with its length, width, and thickness being equal. We calculate the volume of the cube by multiplying length by width by thickness.

$$10\ \text{cm} \times 10\ \text{cm} \times 10\ \text{cm} = 1000\ \text{cm}^3$$

The calculated volume of 1 L is 1000 cm^3. Moreover, we recall the exact equivalent: 1 L = 1000 mL. We can combine the two equivalents as follows.

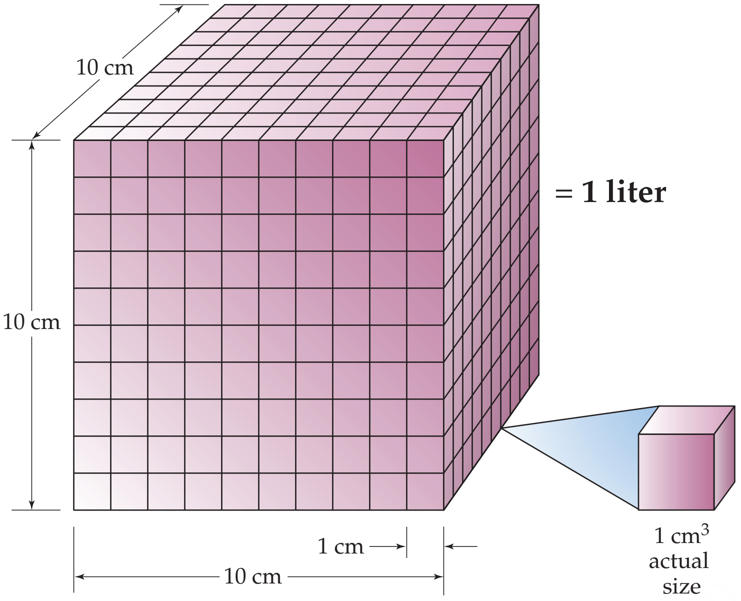
$$1000\ \text{cm}^3 = 1\ \text{L} = 1000\ \text{mL}$$

Simplifying, we have

$$1\ \text{cm}^3 = 1\ \text{mL}$$

We see that 1 cm^3 is exactly equivalent to 1 mL. Example Exercise 2.13 further illustrates the calculation of volume.

Note A cubic centimeter is sometimes abbreviated "cc," as in a 1 cc medical injection. It is not obvious that cc stands for cubic centimeters, so use of this abbreviation is discouraged in science, even though it is common in medicine.



▲ **One Liter** A liter equals the volume of a cube 10 cm on a side, that is, 1000 cm^3.

▲ **1-Liter Bottle** A 1-liter bottle contains 1000 mL of liquid.

EXAMPLE EXERCISE 2.13 Metric–English Volume Conversion

Given that an automobile engine has a volume displacement of 498 cm^3 in each cylinder, express the volume in cubic inches.

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**in.^{3}**
STEP 2:	What given value is related to the answer?	**Step 2:**	**498 cm^3**
STEP 3:	What unit factor(s) should we apply?	**Step 3:**	$\frac{\text{1 in.}}{\text{2.54 cm}}$ **or** $\frac{\text{2.54 cm}}{\text{1 in.}}$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit factor 3*	*unit in answer*
498 cm^3	$\times\ \frac{\text{1 in.}}{\text{2.54 cm}}$	$\times\ \frac{\text{1 in.}}{\text{2.54 cm}}$	$\times\ \frac{\text{1 in.}}{\text{2.54 cm}}$	**= ? in.^{3}**

Solution

We can calculate the volume in cubic inches by converting the volume given in cubic centimeters. Notice that cm units do not cancel cm^3. To obtain cubic inches, we must apply the unit factor, 1 in. /2.54 cm, three times.
Thus,

$$498\ \cancel{\text{cm}^3} \times \frac{1\ \text{in.}}{2.54\ \cancel{\text{cm}}} \times \frac{1\ \text{in.}}{2.54\ \cancel{\text{cm}}} \times \frac{1\ \text{in.}}{2.54\ \cancel{\text{cm}}} = 30.4\ \text{in.}^3$$

The given value, and unit factors, each limits the answer to three significant digits.

Practice Exercise

Given that an SUV has a 244 in.^{3} engine, express the engine volume in liters.

Answers: 4.00 L

Concept Exercise

Which of the following is the greater volume: 500 mL or 500 cm^3?

Answers: See Appendix G, 2.13.

▲ **A Typical SUV.**

LEARNING OBJECTIVE

- Estimate the volume of a solid and gas by the displacement of water.

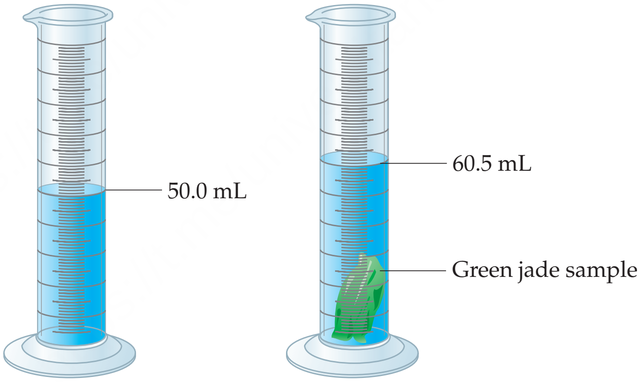
2.7 Volume by Displacement

The volumes of liquids and solids can be determined by several methods. The volumes of liquids can be determined directly with calibrated glassware, such as graduated cylinders, pipets, and burets. The volumes of regular solids (cubes, cylinders, spheres) can be calculated from measurements.

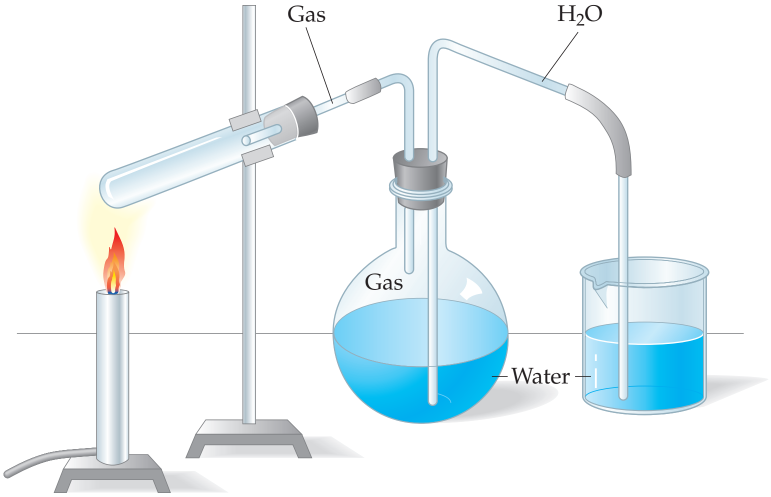
The volume of an irregular solid cannot be determined directly by multiplying its measurements. However, the volume can be determined *indirectly* by measuring the amount of water the solid displaces. This technique is called determination of **volume by displacement**.

Suppose we wish to determine the volume of a piece of green jade. Because it is an irregular solid object, we cannot measure its volume directly. We have to use the technique of determining volume by displacement. We first fill a graduated cylinder halfway with water (Figure 2.2) and record the water level. Next, we carefully slip the piece of green jade into the graduated cylinder and record the new water level. The difference between the final and the initial water levels represents the volume of the piece of jade.

We can also use the volume by displacement method to determine the volume of a gas. When a gas is produced from a chemical reaction, its volume is equal to the amount of water it displaces. This technique provides good results for gases that do not dissolve, or only slightly dissolve, in water. For instance, heating a solid sample may release oxygen gas, which displaces water from a closed container. Because oxygen dissolves only slightly in water, the volume of water displaced is equal to the volume of gas released. Determining the volume of a gas by displacement is referred to as *collecting a gas over water* and is illustrated in Figure 2.3.



▶ **Figure 2.2 Volume of a Solid by Displacement** The difference between the final water level and the initial water level in the graduated cylinder (10.5 mL) is equal to the volume of the solid piece of jade.



▶ **Figure 2.3 Volume of a Gas by Displacement** A flask is filled with water, and a sample is placed in the test tube. After heating, the sample decomposes and produces oxygen gas, which in turn displaces water from the flask. The volume of water displaced into the beaker equals the volume of oxygen gas produced by the sample.

EXAMPLE EXERCISE 2.14 Volume by Displacement

A quartz stone weighing 30.475 g is carefully placed into a graduated cylinder. If the water level increases from 25.0 mL to 36.5 mL, what is the volume of the quartz stone?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**mL**
STEP 2:	What given value is related to the answer?	**Step 2:**	**25.0 mL and 36.5 mL**
STEP 3:	What unit factor(s) should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

We can calculate the displaced volume in milliliters by subtracting the initial volume from the final volume.

$$36.5 \text{ mL} - 25.0 \text{ mL} = 11.5 \text{ mL}$$

Practice Exercise

Hydrogen peroxide decomposes to give oxygen gas, which displaces a volume of water into a beaker. If the water level in the beaker increases from 50.0 mL to 105.5 mL, what is the volume of oxygen gas?

Answers: 55.5 mL

Concept Exercise

Which of the following has the greater volume: 1 mL or 1 cm^3?

Answers: See Appendix G, 2.14.

2.8 The Density Concept

LEARNING OBJECTIVES

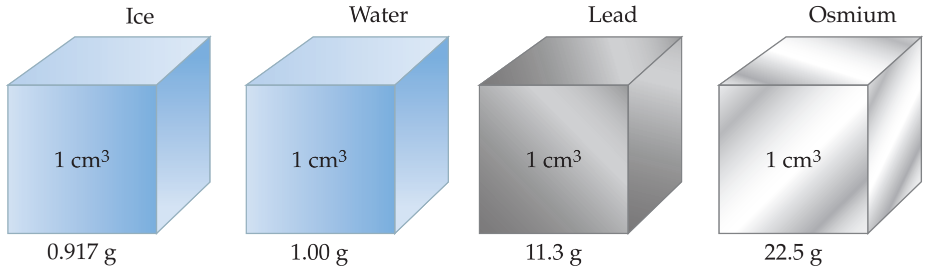
- Apply the concept of density.
- Memorize the value for the density of water in grams per milliliter.
- Solve problems that relate density, mass, and volume.

If we hold a piece of wood in one hand and a piece of metal in the other (each having the same mass), the piece of metal may feel heavier. This contradiction is explained by the term **density** (symbol ***d***), which expresses the concentration of mass. The more concentrated the mass is in a sample, the greater its density. We sometimes say an object is heavy when we are trying to convey density. A 1 kg block of lead and a 1 kg block of ice have the same mass, but the lead is much more dense.

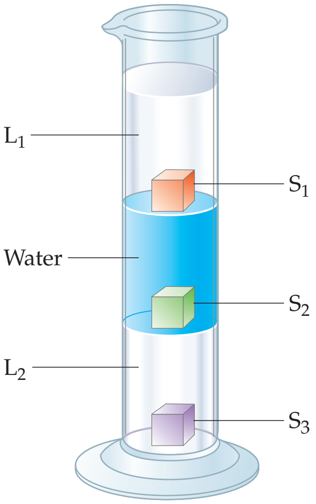
We define density as the amount of mass per unit volume. The term *per* indicates a ratio, as in "miles per hour." Thus, mass divided by volume equals density. We can write this relationship mathematically as

$$\frac{\text{mass}}{\text{volume}} = \text{density}$$

Density is expressed in different units. For solids and liquids, it is usually expressed in grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL). For gases, density is usually expressed in grams per liter (g/L). Figure 2.4 illustrates the concept of density.



◀ **Figure 2.4 An Illustration of Density** Each cube represents a volume of 1 cm^3. Notice the mass of each cube increases as its density becomes greater. Osmium has the greatest density of any metal, 22.5 g/cm^3.



▲ **Figure 2.5 The Concept of Density** In a tall glass cylinder, liquid L_1 is shown floating on water, which is floating on liquid L_2. Solid S_1 is floating on water, solid S_2 is floating on liquid L_2, and solid S_2 has sunk to the bottom of the glass cylinder.

Recall that the original metric standards for mass and volume were based on water. A mass of 1 kg was set equal to a 1 L volume of water. Therefore, the density of water is 1 kg/L. A density of 1 kg/L is exactly equivalent to 1 g/mL.

Although the density of solids and liquids varies slightly with temperature, we will ignore this small effect. For our purposes, we will consider the density of water to be 1.00 g/mL. The density of gases, however, is an exception. The density of a gas varies greatly with changes in temperature and pressure. Table 2.4 lists the density for selected solids, liquids, and gases.

We can drop a solid object into water and estimate its density. An object that sinks in water has a density greater than 1.00 g/mL. An object that floats on water has a density less than 1.00 g/mL.

Consider the following experiment. Water, chloroform, and ethyl ether are poured into a tall glass cylinder. The liquids form three layers as shown in Figure 2.5. Identify liquids L1 and L2 by referring to the density of the liquids listed in Table 2.4. Next, ice, rubber, and aluminum cubes are dropped into the cylinder. Identify solids S_1, S_2, and S_3 by referring to the density of the solids listed in Table 2.4.

We know that liquid L_1 must be less dense than water ($d = 1.00$ g/mL) because it floats; liquid L_2 must be more dense than water because it sinks. Thus, liquid L_1 is ethyl ether ($d = 0.714$ g/mL), which is less dense than water, and liquid L_2 is chloroform ($d = 1.48$ g/mL), which is more dense than water.

We see in Figure 2.5 that S_1 floats on water, S_2 floats on chloroform, and S_3 sinks in chloroform. Table 2.4 indicates the densities of ice, rubber, and aluminum are 0.917, 1.19, and 2.70 g/cm^3, respectively. Thus, S_1 must be ice because it sinks in ether and floats on water, S_2 must be rubber because it sinks in water and floats on chloroform, and S_3 must be aluminum because it sinks in chloroform.

Now that we understand the concept of density, let's try some calculations. We can find the density of a sample by simply dividing its mass by its volume. Example Exercise 2.15 illustrates the calculation of density.

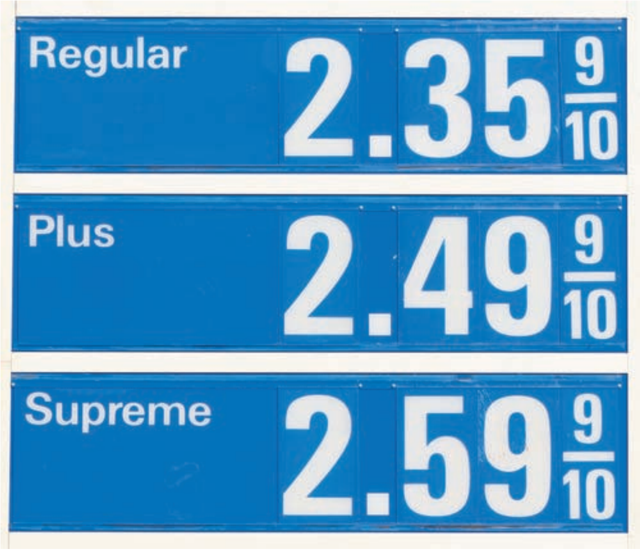
A CLOSER LOOK Lower Gasoline Bills

Q: In terms of expense, is it better to fill a gas tank in the cool morning, or in the warm afternoon?

Interestingly, you can reduce your automobile fuel bills by applying the density concept discussed in this chapter. The density of most liquids decreases as the temperature warms, and increases as the temperature cools. Thus, the density of gasoline decreases in the warm afternoon, and increases in the cool morning.

Gasoline is sold at the pump by the volume delivered (gallons), and not by the mass, which is greater when the density increases. Although $50.00 of fuel contains the *same volume* at different temperatures, $50.00 of fuel contains a *greater mass* of gasoline at 40 °F than at 70 °F because the density increases. Thus, filling up with gas in the cool morning saves money.

Two liquids violate the principle of greater density at a lower temperature. The liquids are water and ammonia. At temperatures near their freezing points (0 °C and −77 °C respectively), each of these liquids expands to form a crystalline solid. Consequently, the density of water decreases and solid ice demonstrates the unusual property of floating in water. Similarly, solid ammonia floats in liquid ammonia. All other solid substances are more dense and sink in their corresponding liquids.



▲ **Gas Prices Displayed at a Gas Station.**

A: Although it is more economical to fill a gas tank in the cool morning, the density change is small and the mass of gasoline is only about 1% greater at 40 °F than at 70 °F.

TABLE 2.4 Density of Selected Solids, Liquids, and Gases

Example	Density (*d*)
Solids	
ice	0.917 g/cm^3 or g/mL
rubber	1.19
magnesium	1.74
aluminum	2.70
iron	7.87
lead	11.3
gold	19.3
Liquids	
ethyl ether	0.714 g/cm^3 or g/mL
ethyl alcohol	0.789
water	1.000
chloroform	1.48
mercury	13.6
Gases*	
hydrogen	0.090 g/L
helium	0.179
ammonia	0.760
air	1.29
oxygen	1.43

*The density value for each gas is given at 0 °C and a pressure equal to normal atmospheric pressure.

EXAMPLE EXERCISE 2.15 Density Calculation

If a platinum nugget has a mass of 214.50 g and a volume of 10.0 cm^3, what is the density of the metal?

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**g/cm^3**
STEP 2: What given value is related to the answer?	**Step 2:**	**214.50 g and 10.0 cm^3**
STEP 3: What unit factor(s) should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

We can calculate the density of the platinum nugget by comparing the mass of metal, 214.50 g, to its volume, 10.0 cm^3.

$$\frac{214.50\ \text{g}}{10.0\ \text{cm}^3} = 21.5\ \text{g/cm}^3$$

The given volume has three significant digits, so the answer is rounded off to three digits. It is interesting to note that platinum metal is more dense than lead ($d = 11.3$ g/cm^3), and more valuable than gold.

Practice Exercise

Carbon tetrachloride is a solvent used for degreasing electronic parts. If 25.0 mL of carbon tetrachloride has a mass of 39.75 g, what is the density of the liquid?

▲ **Platinum Nugget** A nugget of precious platinum metal.

Answers: 1.59 g/mL

Concept Exercise

Which of the following has the greater density: ice or water?

Answers: See Appendix G, 2.15.

Applying Density as a Unit Factor

We can also solve problems that ask for the mass or volume of a sample by applying the unit analysis method of problem solving. The secret to solving these problems is to realize that density can be used as a unit factor because it relates mass and volume. For example, the density of mercury is 13.6 g/mL, and so we can write

$$\frac{13.6\ \text{g}}{1\ \text{mL}} \quad \text{and} \quad \frac{1\ \text{mL}}{13.6\ \text{g}}$$

Let's calculate the milliliter volume of 75.5 g of liquid mercury. *First*, we write down the unit asked for in the answer (mL). *Second*, we write down the related given value (75.5 g). *Third*, we apply a unit factor using the following format:

$$\underset{\text{(Step 2)}}{75.5\ \text{g}} \quad \times \quad \underset{\text{(Step 3)}}{\frac{\text{unit}}{\text{factor}}} \quad = \quad \underset{\text{(Step 1)}}{\text{mL}}$$

To cancel the units in the given value (75.5 g), we select the unit factor 1 mL/13.6 g. After substituting, we have

$$75.5\ \cancel{\text{g}} \times \frac{1\ \text{mL}}{13.6\ \cancel{\text{g}}} = 5.55\ \text{mL}$$

The given value and the unit factor each have three significant digits; thus, the answer is rounded to three digits. Example Exercise 2.16 further illustrates the use of density as a unit factor.

EXAMPLE EXERCISE 2.16 Density as a Unit Factor

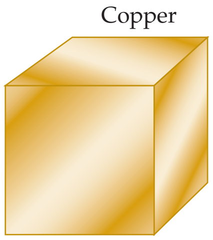
A 1.00-in. cube of copper measures 2.54 cm on a side. What is the mass of the copper cube (given that *d* of copper = 8.96 g/cm^3)?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**g**
STEP 2:	What given value is related to the answer?	**Step 2:**	**2.54 cm cube**
STEP 3:	What unit factor(s) should we apply? From the definition of density, 8.96 g = 1 cm^3; thus, the two unit factors are 8.96 g/1 cm^3, and its reciprocal 1 cm^3/8.96 g.	**Step 3:**	$\frac{8.96\ \text{g}}{1\ \text{cm}^3}$ **or** $\frac{1\ \text{cm}^3}{8.96\ \text{g}}$

Unit Analysis Map

given value	*unit factor*	*unit in answer*
16.4 cm^3	$\times\ \frac{8.96\ \text{g}}{1\ \text{cm}^3}$ **or** $\frac{1\ \text{cm}^3}{8.96\ \text{g}}$	= **? g**



▲ **Copper Metal** A 1.00-inch cube of copper metal.

Solution

First, we find the volume of the copper cube. We obtain the volume of the cube, 16.4 cm^3, by multiplying (2.54 cm) (2.54 cm) (2.54 cm). We use the given density, 8.96 g/1 cm^3, as a unit factor to cancel cubic centimeters ($\cancel{cm^3}$), which appears in the denominator.

$$16.4\ \cancel{cm^3} \times \frac{8.96\ g}{1\ \cancel{cm^3}} = 147\ g$$

The given value and unit factor each has three significant digits, so the answer is rounded off to three significant digits.

Practice Exercise

A cube of silver is 5.00 cm on a side and has a mass of 1312.5 g. What is the density of silver?

Answers: 10.5 g/cm^3

Concept Exercise

If some humans float in water and others sink, what is the approximate density of the human body?

Answers: See Appendix G, 2.16.

Specific Gravity

The ratio of the density of a liquid to the density of water at 4 °C is called **specific gravity** (symbol **sp gr**). Because specific gravity is a ratio of two densities, the units cancel and specific gravity is a unitless quantity. The density of water is 1.00 g/mL, and the specific gravity of water is 1.00.

Diagnostic medical testing often includes the specific gravity of body fluids. For instance, the specific gravity of urine may be 1.02, and the specific gravity of blood may be 1.06. Both of these values are considered to be in the normal range.

2.9 Temperature

LEARNING OBJECTIVES

- Memorize the freezing point and boiling point of water on the Fahrenheit, Celsius, and Kelvin scales.
- Express a given temperature in degrees Fahrenheit, degrees Celsius, and Kelvin (K).

The "hotness" or "coolness" of the atmosphere is determined by how fast the tiny air molecules are moving. If the temperature is warmer, molecules move faster and have more energy. If the temperature is cooler, molecules move slower and have less energy. **Temperature** is the average kinetic energy of individual particles in motion; for example, the average energy of air molecules in the atmosphere. We measure temperature using a thermometer.

In 1724, Daniel Gabriel Fahrenheit (1686–1736), a German physicist, invented the mercury thermometer. In attempting to produce as cold a temperature as possible, Fahrenheit prepared an ice bath to which he added salt to lower the temperature further. He then assigned a value of zero to that temperature and marked his Fahrenheit scale accordingly.

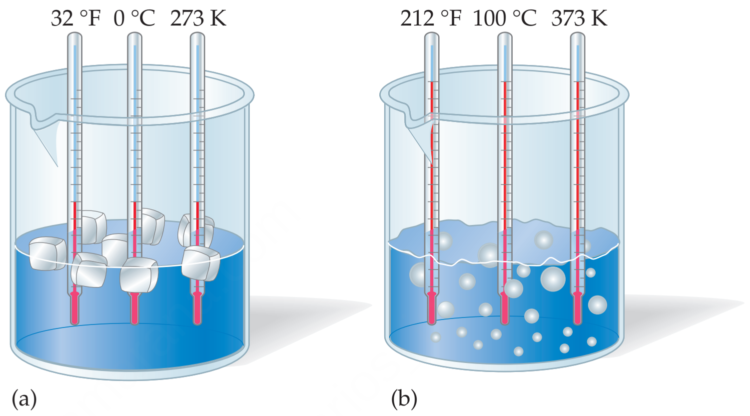
Fahrenheit obtained a second reference point by recording his underarm body temperature; he assigned that temperature a value of 96. He divided the distance between the two reference points into 96 equal units, and each division was termed a **Fahrenheit degree** (symbol **°F**). Later, the freezing and boiling points of water were selected as the standard reference points. The freezing point of water was assigned a value of 32 °F, and the boiling point a value of 212 °F. The Fahrenheit degree eventually became a basic unit in the English system of measurement.

In 1742, Anders Celsius (1701–1744), a Swedish astronomer, proposed a scale similar in principle to the Fahrenheit scale. On the Celsius scale, the freezing point of water was assigned a value of 0 °C and the boiling point of water a value of 100 °C. The scale was then divided into 100 equal divisions, each division representing a **Celsius degree** (symbol **°C**). The Celsius degree is sometimes referred to as a centigrade degree, but the use of degrees centigrade is discouraged.

In 1848, William Thomson (1824–1907), an English physicist also known as Lord Kelvin, proposed a scale based on the lowest possible temperature. The unit of temperature was the **Kelvin unit** (symbol **K**), which is a basic unit in the SI system. On the Kelvin scale, the coldest temperature was assigned a value of 0 K, and each division on the scale is equal to one Celsius degree. The lowest temperature is referred to as absolute zero and corresponds to −273.15 °C. Although there is no highest temperature on the Kelvin scale, the interior of the Sun reaches about 10,000,000 K.

Fahrenheit and Celsius Temperature Conversions

One division on the Kelvin scale is equivalent to 1° on the Celsius scale. Because 0 K is equivalent to −273 °C, the freezing point of water is 273 K, and the boiling point of water is 373 K. Figure 2.6 illustrates how the three temperature scales are related.



▶ **Figure 2.6 Temperature Scales** A Fahrenheit, a Celsius, and a Kelvin thermometer are placed in (a) ice water and (b) boiling water. Notice the freezing point and boiling point on each scale. The number of divisions is 180 units on the Fahrenheit scale, 100 units on the Celsius scale, and 100 units on the Kelvin scale.

Notice that 180 Fahrenheit units are equivalent to 100 Celsius units. Therefore, to convert from °F to °C, we first subtract 32 (the difference between the freezing point of water on the two scales) and then multiply by 100 °C/180 °F. That is,

$$(\text{°F} - 32\ \text{°F}) \times \frac{100\ \text{°C}}{180\ \text{°F}} = \text{°C}$$

To convert from °C to °F, we reverse the procedure of converting °F to °C. We first multiply the Celsius temperature by the ratio 180 °F/100 °C, and then add 32 °F. That is,

$$\left(\text{°C} \times \frac{180\ \text{°F}}{100\ \text{°C}}\right) + 32\ \text{°F} = \text{°F}$$

Example Exercise 2.17 illustrates the conversion of Fahrenheit and Celsius temperatures.

EXAMPLE EXERCISE 2.17 °F and °C Temperature Conversions

Normal body temperature is 98.6 °F. What is normal body temperature in degrees Celsius?

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**°C**
STEP 2: What given value is related to the answer?	**Step 2:**	**98.6 °F**
STEP 3: What unit factor(s) should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

To calculate °C, we refer to Figure 2.6 and compare the Celsius and Fahrenheit temperature scales. The conversion from °F to °C is as follows.

$$(98.6\ ^\circ\text{F} - 32\ ^\circ\text{F}) \times \frac{100\ ^\circ\text{C}}{180\ ^\circ\text{F}} = ^\circ\text{C}$$

Simplifying and canceling units gives

$$66.6\ \cancel{^\circ\text{F}} \times \frac{100\ ^\circ\text{C}}{180\ \cancel{^\circ\text{F}}} = 37.0\ ^\circ\text{C}$$

The given value, 98.6 °F, has three significant digits, so the answer is rounded off to three digits. Because 32 °F and 100 °C/180 °F are exact numbers, neither affects the significant digits in the answer.

Practice Exercise

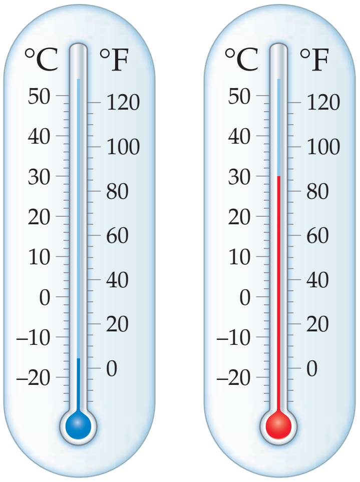
The average surface temperature of Mars is −55 °C. What is the average temperature in degrees Fahrenheit?

Answers: −67 °F

Concept Exercise

What is the relationship between the Celsius and centigrade temperature scales?

Answers: See Appendix G, 2.17.



▲ **Celsius and Fahrenheit Thermometers** Note that a temperature of –15 °C corresponds to 5 °F; a temperature of 30 °C equals 86 °F.

Note We can convert Fahrenheit and Celsius temperatures by understanding the relationship of the two temperature scales (see Figure 2.6). In practice, nurses and other professionals may have a reference chart showing equivalent Fahrenheit and Celsius temperatures. Moreover, students can convert Fahrenheit and Celsius temperatures with a single keystroke using an inexpensive scientific calculator.

Celsius and Kelvin Temperature Conversions

Let's reexamine the temperature scales in Figure 2.6. Notice that the Kelvin scale is 273 units above the Celsius scale. Therefore, to convert from °C to K, we must add 273 units to the Celsius temperature.

$$^\circ\text{C} + 273 = \text{K}$$

Conversely, to convert from K to °C, we subtract 273 units from the Kelvin temperature. It is helpful to remember that negative Kelvin temperatures are nonexistent. By definition, the lowest possible temperature is assigned a value of 0 K. Example Exercise 2.18 illustrates the conversion of Celsius and Kelvin temperatures.

EXAMPLE EXERCISE 2.18 °C and K Temperature Conversions

Dermatologists use liquid nitrogen to freeze skin tissue. If the Celsius temperature of liquid nitrogen is −196 °C, what is the Kelvin temperature?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**K**
STEP 2:	What given value is related to the answer?	**Step 2:**	**– 196 °C**
STEP 3:	What unit factor(s) should we apply? No unit factor is required.	**Step 3:**	**none**

Solution

Given the Celsius temperature, we add 273 units to find the corresponding Kelvin temperature.

$$-196\ ^\circ\text{C} + 273 = 77\ \text{K}$$

Practice Exercise

The secret to "fire-walking" is to first walk barefoot through damp grass and then step lively on the red-hot coals. If the bed of coals is 1475 K, what is the Celsius temperature?

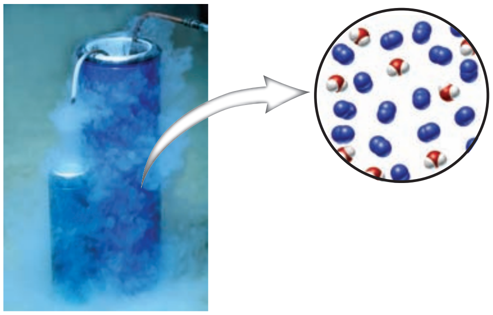
Answers: 1202 °C

Concept Exercise

Which of the following temperatures does not exist?

$$-100\ ^\circ\text{F},\ -100\ ^\circ\text{C},\ -100\ \text{K}$$

Answers: See Appendix G, 2.18.



▲ **Liquid Nitrogen** Although nitrogen is normally a gas, it liquefies at −196 °C. When liquid nitrogen is poured from a Thermos, it is cold enough to freeze the moisture in air and form a white mist.

LEARNING OBJECTIVES

- Apply the concept of specific heat to good and poor conductors of heat.
- Express heat energy in calories, kilocalories, joules, and kilojoules.
- Solve problems that relate specific heat, mass, and temperature change.

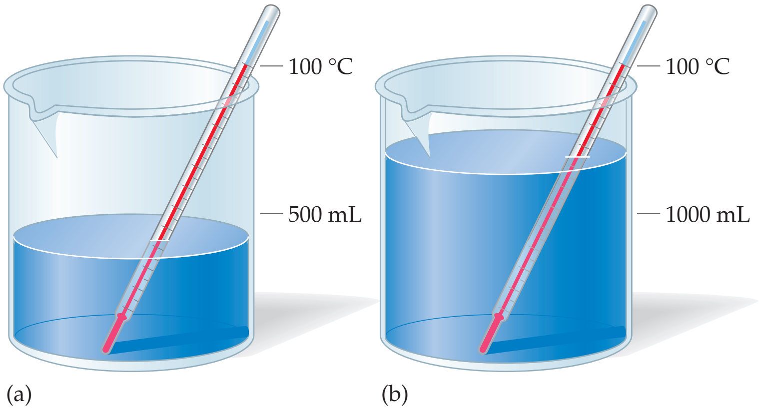
2.10 The Heat Concept

Scientists define *heat* differently depending on context. In this text, we will define heat as the flow of energy from an object at a higher temperature to an object at a lower temperature. For example, if you hold an ice cube, the heat from your hand (higher temperature) flows into the ice cube (lower temperature) and your hand feels cold.

Heat and temperature are both a measure of energy. However, **heat** is a measure of total energy, and temperature is a measure of average energy. To distinguish heat from temperature, consider a cup of hot tea and a teaspoon of hot tea, each of which is at 100 °C. Which feels hotter: quickly drinking a cup of hot tea, or sipping a teaspoon of hot tea? Obviously, the cup of hot tea is more likely to burn your mouth than the teaspoon of hot tea. The cup of tea has more heat than the teaspoon of tea, even though each is at the same temperature, 100 °C. Figure 2.7 further illustrates the concepts of heat and temperature.

Heat energy is often expressed in units of calories or kilocalories. A **calorie** (symbol **cal**) is the amount of heat necessary to raise 1 gram of water 1 degree on the Celsius scale. A kilocalorie (symbol kcal) is the amount of heat necessary to raise 1000 grams of water 1 degree on the Celsius scale. A nutritional Calorie (symbol Cal) is spelled with a capital letter to distinguish it from a metric calorie. One nutritional Calorie is equal to 1 kilocalorie, that is, 1000 calories.

The SI unit of energy is the **joule** (symbol **J**), where 1 cal = 4.184 J. The heat produced by chemical reactions is often expressed in kilocalories, as well as in kilojoules (kJ), where 1 kcal = 4.184 kJ. Example Exercise 2.19 illustrates the conversion of calories, kilocalories, and joules.



▶ **Figure 2.7 Heat versus Temperature** In (a) 500 mL of water is heated to 100 °C, and in (b) 1000 mL is heated to 100 °C. Although the temperatures are the same, the second beaker has twice the amount of heat.

EXAMPLE EXERCISE 2.19 Energy Conversion

Burning one liter of natural gas produces 9.46 kcal of heat energy. Express the energy in kilojoules (given that 1 kcal = 4.184 kJ).

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: kJ

STEP 2: What given value is related to the answer?

Step 2: 9.46 kcal

STEP 3: What unit factor(s) should we apply? The unit equation is 1 kcal = 4.184 kJ, so the two unit factors are 1 kcal/4.184 kJ, and its reciprocal 4.184 kJ/1 kcal.

Step 3: $\frac{1\ \text{kcal}}{4.184\ \text{kJ}}$ or $\frac{4.184\ \text{kJ}}{1\ \text{kcal}}$

▲ **Bunsen Burner** A laboratory burner that uses natural gas for fuel.

Unit Analysis Map

given value
9.46 kcal

unit factor
× $\frac{1\ \text{kcal}}{4.184\ \text{kJ}}$ or $\frac{4.184\ \text{kJ}}{1\ \text{kcal}}$

unit in answer
= ? kJ

Solution

We apply the unit factor 4.184 kJ/1 kcal to cancel kilocalories (kcal), which appears in the denominator.

$$9.46\ \cancel{\text{kcal}} \times \frac{4.184\ \text{kJ}}{1\ \cancel{\text{kcal}}} = 39.6\ \text{kJ}$$

The given value has three significant digits, and the unit factor has four digits. Thus, we round off the answer to three significant digits.

Practice Exercise

Burning one gram of gasoline produces 47.9 kJ of energy. Express the heat energy in kilocalories (given that 1 kcal = 4.184 kJ).

Answers: 11.4 kcal

Concept Exercise

If an aerosol can feels cold after releasing the spray, is heat flowing from the can or from your hand?

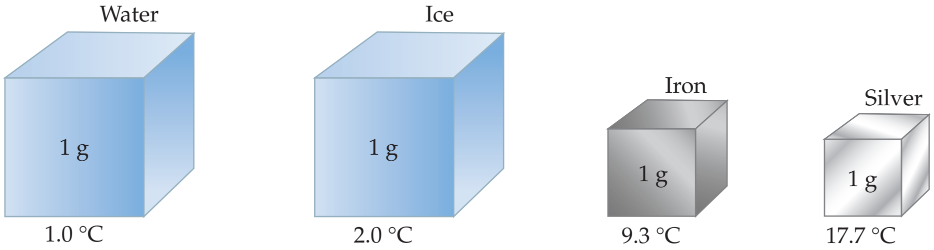
Answers: See Appendix G, 2.19.

Specific Heat

The specific heat of a substance is the amount of heat required to bring about a given change in temperature. It is observed that the amount of heat necessary is unique for each substance. The specific heat of water is relatively high, and the change in temperature is minimal as water gains or loses heat. The surface of Earth is covered with water, and fortunately its high specific heat helps to regulate climates and maintain moderate temperatures. Figure 2.8 illustrates the increase in temperature for four substances each receiving 1 cal of heat.

We can define **specific heat** as the amount of heat required to raise the temperature of one gram of substance one degree Celsius; the units of specific heat are often given in calories per gram per degree Celsius. For reference, the specific heat of water has a value of 1.00 cal/(g × °C).

Water has an unusually high specific heat and requires more energy to raise its temperature 1 °C than ice or steam. The high specific heat of water is responsible for maintaining Earth's climate as our oceans resist wide swings in temperature. Table 2.5 lists the specific heat for selected solids, liquids, and gases.



▶ **Figure 2.8 An Illustration of Specific Heat** Each cube represents a volume of 1 gram of substance receiving 1 calorie of heat. Notice the temperature change increases as the specific heat decreases. Silver has a low specific heat and is a good conductor of heat.

TABLE 2.5 Specific Heat for Selected Solids, Liquids, and Gases

Example	Specific Heat
Solids	
ice	0.492 cal/(g × °C)
aluminum	0.215
carbon (graphite)	0.170
carbon (diamond)	0.124
iron	0.108
copper	0.0920
silver	0.0566
gold	0.0305
Liquids	
water	1.00 cal/(g × °C)
ethyl alcohol (ethanol)	0.587
methyl alcohol (methanol)	0.424
freon (CFC refrigerant)	0.232
mercury	0.0331
Gases	
steam	0.485 cal/(g × °C)
nitrogen	0.249
oxygen	0.219
argon	0.124
radon	0.0224

EXAMPLE EXERCISE 2.20 Specific Heat

A 25.0-g sample of rock at 250 °C is dropped into 100.0 g of water at 19 °C. If the water temperature rises to 31 °C as the rock cools, what is the specific heat of the rock?

Strategy Plan

STEP 1: What unit is asked for in the answer?

STEP 2: What given value is related to the answer?

STEP 3: What unit factor should we apply? From the definition of specific heat, 1.00 cal = 1 g × °C; the two unit factors are 1.00 cal = 1 g × °C, and its reciprocal 1 g × °C = 1.00 cal.

Step 1: **cal/(g × °C)**

Step 2: **100.0 g water**

Step 3: $\frac{1\text{ g} \times {}^\circ\text{C}}{1.00\text{ cal}}$ **or** $\frac{1.00\text{ cal}}{1\text{ g} \times {}^\circ\text{C}}$

Unit Analysis Map

given value 100.0 g $\times$ *unit factor* $\dfrac{1\text{ g} \times {}^\circ\text{C}}{1.00\text{ cal}}$ or $\dfrac{1.00\text{ cal}}{1\text{ g} \times {}^\circ\text{C}}$ *unit in answer* = ? cal/g × °C

Solution

We can cancel grams and degrees Celsius by applying the unit factor 1.00 cal/(1 g × °C) and multiplying by the temperature change.

$$100.0\ \cancel{\text{g}} \times \frac{1.00\text{ cal}}{1\ \cancel{\text{g}} \times \cancel{{}^\circ\text{C}}} \times (31 - 19)\ \cancel{{}^\circ\text{C}} = 1200\text{ cal}$$

If the water gains 1200 cal, the rock loses 1200 cal; that is, heat gain = heat loss. Because the rock cools from 250 °C to 31 °C, we can calculate the specific heat of the rock,

$$\frac{1200\text{ cal}}{25.0\text{ g }(250 - 31){}^\circ\text{C}} = 0.22\text{ cal/g} \times {}^\circ\text{C}$$

The principle of heat transfer is demonstrated in a sauna. Pouring cool water on hot rocks causes the water to first heat, then vaporize to produce hot steam.

Practice Exercise

A 1025–g horseshoe at 425 °C is dropped into 1550 g of water at 20 °C. If the water temperature rises to 47 °C as the horseshoe cools, what is the specific heat of the horseshoe?

Answers: 0.11 cal/g × °C

Concept Exercise

What type of substance is better for storing heat energy: a substance with a high specific heat or a low specific heat?

Answers: See Appendix G, 2.20.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
2.1 Basic Units and Symbols The **English system** of measurement has many unrelated units. On the other hand, the **metric system** of measurement is a decimal system with basic units: **meter (m)**, **gram (g)**, **liter (L)**, and **second (s)**. Metric prefixes provide multiples and fractions of basic units. The common prefixes include *giga-* (G), *mega-* (M), kilo- (k), *deci-* (d), *centi-* (c), *milli-* (m), *micro-* (μ), and *nano-* (n). The **International System** (**SI**) is based on the metric system but is more comprehensive.	• List the base units and symbols of the metric system. *Related Exercises: 1–6* • List the prefixes for multiples and fractions of base units. *Related Exercises: 7–10*
2.2 Metric Conversion Factors Metric conversion problems are solved by systematically writing a **unit equation** or an **exact equivalent** (1 m = 100 cm). An equal or equivalent relationship generates a **unit factor** and its **reciprocal** (1 m/100 cm and 100 cm/1 m).	• Show the unit equation for a base metric unit and a prefix unit. *Related Exercises: 11–12* • Show the two unit factors derived from a metric unit equation. *Related Exercises: 13–14*
2.3 Metric–Metric Conversions Metric problems are solved systematically by applying the **unit analysis method.** Step 1: Write down the units of the unknown. Step 2: Write down a given value that is related to the unknown. Step 3: Apply one or more unit factors to convert the units of the given value to the units asked for in the answer.	• Express a given metric measurement using a different metric prefix. *Related Exercises: 15–18*

Key Concepts	Learning Objectives and Related Exercises
2.4 Metric–English Conversions Every nation in the world uses the metric system of measurement; however, the English system is still common in the United States. Therefore, we will relate the two systems and memorize the following: 2.54 cm = 1 in., 454 g = 1 lb, and 946 mL = 1 qt. These equivalents will be used for metric–English conversions.	• Memorize the metric equivalent for inch, pound, quart, and second. ***Related Exercises: 19–20*** • Express a given measurement in metric units or English units. ***Related Exercises: 21–28***
2.5 The Percent Concept A **percent** (%) expresses the amount of a single quantity compared to an entire sample. Percent is a ratio, so we can apply percent unit factors to perform calculations using the unit analysis method.	• Express a given quantity in a sample as a percent-by-mass. ***Related Exercises: 29–32*** • Apply percent as a unit factor. ***Related Exercises: 33–34***
2.6 Volume by Calculation The volume of a rectangular solid is equal to length times width times thickness. The calculated volume is reported in cubic units, such as **cubic centimeters** (**cm^3**). You should memorize the exact equivalent 1 mL = 1 cm^3.	• Solve problems that relate length, width, thickness, and volume of a rectangular solid. ***Related Exercises: 35–38*** • Express a given volume in milliliters, cubic centimeters, or cubic inches. ***Related Exercises: 39–42***
2.7 Volume by Displacement For irregularly shaped objects, as well as gases, we must determine volume indirectly. We can find the volume of an object or a gas from the amount of water it displaces; this technique involves **volume by displacement**.	• Estimate the volume of a solid by the displacement of water. ***Related Exercises: 43–46***
2.8 The Density Concept The **density** of water is 1.00 g/mL. A liquid or solid that floats on water is less dense than water; if it sinks, it is more dense. Density can be used as a unit factor, for example, 1.00 g/1 mL or 1 mL/1.00 g. **Specific gravity (sp gr)** is unitless and expresses the ratio of the density of a liquid to the density of water.	• Apply the concept of density. ***Related Exercises: 47–50*** • Memorize the value for the density of water in grams per milliliter. ***Related Exercises: 51–52*** • Solve problems that relate density, mass, and volume. ***Related Exercises: 53–58***
2.9 Temperature **Temperature** is the average energy of molecules in motion. The **Fahrenheit**, **Celsius**, and **Kelvin** scales have related reference points. The freezing point of water is 32 °F, 0 °C, or 273 K; and the boiling point is 212 °F, 100 °C, or 373 K.	• Memorize the freezing point and boiling point of water on the Fahrenheit, Celsius, and Kelvin scales. ***Related Exercises 59–60*** • Express a given temperature in degrees Fahrenheit (°F), degrees Celsius (°C), and Kelvin (K). ***Related Exercises: 61–66***
2.10 The Heat Concept **Heat** is the total energy of molecules in motion. Heat changes are expressed in **calories (cal)** or kilocalories (kcal). Heat changes can also be expressed in the SI unit **joule (J)**, where 1 cal = 4.184 J. **Specific heat** is the amount of heat necessary to raise the temperature of 1 g of substance 1 °C; for water, the value is 1.00 cal/(g × °C).	• Apply the concept of specific heat to good and poor conductors of heat. ***Related Exercises 67–68*** • Express heat energy in units of calories, kilocalories, joules, and kilojoules. ***Related Exercises 69–70*** • Solve problems that relate specific heat, mass, and temperature change. ***Related Exercises: 71–76***

Problem–Solving Organizer

Topic	Procedure	Example
Basic Units and Symbols Sec. 2.1	Combine basic metric units and prefixes using symbols.	centimeter, cm (length); kilogram, kg (mass); milliliter, mL (volume)
Metric Conversion Factors Sec. 2.2	(a) Write a unit equation involving basic metric units and prefix units. (b) Write two unit factors for a metric relationship. 1 m = 100 cm	$\frac{1\ \text{m}}{100\ \text{cm}}$ and $\frac{100\ \text{cm}}{1\ \text{m}}$
Metric–Metric Conversions Sec. 2.3	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert the given unit to the unit in the answer.	What is the decigram mass of a wheel of cheese that weighs 0.515 kg? $0.515\ \cancel{\text{kg}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{10\ \text{dg}}{1\ \cancel{\text{g}}} = 5150\ \text{dg}$
Metric–English Conversions Sec. 2.4	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert the given unit to the unit in the answer.	What is the pound mass of a wheel of cheese that weighs 0.515 kg? $0.515\ \cancel{\text{kg}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \text{lb}}{454\ \cancel{\text{g}}} = 1.13\ \text{lb}$
The Percent Concept Sec. 2.5	Compare the amount of a single quantity to the total sample, and multiply by 100%.	If a 14-karat ring contains 20.0 g gold and 14.3 g silver, what is the percent of gold? $\frac{20.0\ \cancel{\text{g}}}{(20.0 + 14.3)\cancel{\text{g}}} \times 100\% = 58.3\%$
Volume by Calculation Sec. 2.6	To calculate the volume of a rectangular solid, multiply length by width by thickness: $V = l \times w \times t$.	What is the volume of a rectangular solid measuring 65 mm by 35 mm by 12 mm? $65\ \text{mm} \times 35\ \text{mm} \times 12\ \text{mm} = 27{,}000\ \text{mm}^3$
Volume by Displacement Sec. 2.7	The volume by displacement is the difference between the initial and final readings in a calibrated container.	What is the volume of a gold nugget if the water level in a graduated cylinder increases from 25.0 mL to 42.5 mL? $42.5\ \text{mL} - 25.0\ \text{mL} = 17.5\ \text{mL}$
The Density Concept Sec. 2.8	The density of a sample is equal to its mass divided by its volume: density = mass/volume.	What is the density of 10.0 mL of ether if its mass is 7.142 g? $\frac{7.142\ \text{g}}{10.0\ \text{mL}} = 0.714\ \text{g/mL}$
Temperature Sec. 2.9	(a) To convert from Celsius to Kelvin, add 273 units. (b) To convert from Kelvin to Celsius, subtract 273 units.	$-88\ ^\circ\text{C} + 273 = 185\ \text{K}$ $1265\ \text{K} - 273 = 992\ ^\circ\text{C}$
The Heat Concept Sec. 2.10	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert units.	What is the heat energy of 10.0 cal expressed in joules? $10.0\ \cancel{\text{cal}} \times \frac{4.184\ \text{J}}{1\ \cancel{\text{cal}}} = 41.8\ \text{J}$

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** a nondecimal system of measurement without basic units
______ **2.** a decimal system of measurement with basic units
______ **3.** the basic unit of length in the metric system
______ **4.** the basic unit of mass in the metric system
______ **5.** the basic unit of volume in the metric system
______ **6.** the basic unit of time in the metric system
______ **7.** a comprehensive system of measurement with seven base units
______ **8.** a statement of two equivalent values, for example, 1 m = 39.4 in.
______ **9.** a statement of two exactly equal values, for example, 1 m = 100 cm
______**10.** a ratio of two quantities that are equivalent, for example, 1 lb/454 g
______**11.** the relationship between a fraction and its inverse, for example, 1 qt/946 mLand 946 mL/1 qt
______**12.** a procedure for solving problems that proceeds from a given value to a related answer by the conversion of units
______ **13.** an expression for the amount of a single quantity compared to an entire sample
______**14.** the volume occupied by a cube 1 cm on a side
______**15.** a technique for determining volume from the amount of water displaced
______**16.** the amount of mass in one unit of volume
______**17.** the ratio of the density of a liquid compared to the density of water at 4 °C
______**18.** the flow of energy from an object at a higher temperature to an object at a lower temperature
______**19.** the average energy of molecules in motion
______**20.** the basic unit of temperature in the English system
______**21.** the basic unit of temperature in the metric system
______**22.** the basic unit of temperature in the SI system
______**23.** the amount of heat required to raise 1 g of substance 1 °C
______**24.** the amount of heat required to raise 1 g of water 1 °C
______**25.** a unit of energy in the SI system

(a) calorie (cal) (*Sec. 2.10*)
(b) Celsius degree (°C) (*Sec. 2.9*)
(c) cubic centimeter (cm^3) (*Sec. 2.6*)
(d) density (*d*) (*Sec. 2.8*)
(e) English system (*Sec. 2.1*)
(f) exact equivalent (*Sec. 2.2*)
(g) Fahrenheit degree (°F) (*Sec. 2.9*)
(h) gram (g) (*Sec. 2.1*)
(i) heat (*Sec. 2.10*)
(j) International System (SI) (*Sec. 2.1*)
(k) joule (J) (*Sec. 2.10*)
(l) Kelvin unit (K) (*Sec. 2.9*)
(m) liter (L) (*Sec. 2.1*)
(n) meter (m) (*Sec. 2.1*)
(o) metric system (*Sec. 2.1*)
(p) percent (%) (*Sec. 2.5*)
(q) reciprocal (*Sec. 2.2*)
(r) second (s) (*Sec. 2.1*)
(s) specific gravity (sp gr) (*Sec. 2.8*)
(t) specific heat (*Sec. 2.10*)
(u) temperature (*Sec. 2.9*)
(v) unit analysis method (*Sec. 2.3*)
(w) unit equation (*Sec. 2.2*)
(x) unit factor (*Sec. 2.2*)
(y) volume by displacement (*Sec. 2.7*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Basic Units and Symbols (Sec. 2.1)

1. What is the basic unit for each of the following physical quantities in the metric system?
(a) length **(b)** mass
(c) volume **(d)** time

2. What is the physical quantity corresponding to each of the following metric basic units?
(a) meter **(b)** gram
(c) liter **(d)** second

3. Write the symbol for the following metric units.
(a) exameter **(b)** teragram
(c) attoliter **(d)** picosecond

4. Write the symbol for the following metric units.
(a) petameter **(b)** gigagram
(c) femtoliter **(d)** nanosecond

5. Write the name of the metric unit indicated by the following symbols.
(a) km **(b)** Mg
(c) nL **(d)** ds

6. Write the name of the metric unit indicated by the following symbols.
(a) cm **(b)** mg
(c) dL **(d)** μs

7. State the quantity that is measured using the following instruments.
(a) metric ruler **(b)** platform balance
(c) graduated cylinder **(d)** stopwatch

8. State the quantity that is measured using the following instruments.
(a) micrometer **(b)** electronic balance
(c) buret **(d)** atomic clock

9. State the name and symbol for each of the following metric quantities.
(a) 1×10^{9} m **(b)** 1×10^{3} g
(c) 1×10^{-3} L **(d)** 1×10^{-9} s

10. State the name and symbol for each of the following metric quantities.
(a) 1×10^{12} m **(b)** 1×10^{6} g
(c) 1×10^{-6} L **(d)** 1×10^{-12} s

Metric Conversion Factors (Sec. 2.2)

11. Write a unit equation for each of the following metric equivalents.
(a) km and m **(b)** Mg and g
(c) L and mL **(d)** s and μs

12. Write a unit equation for each of the following metric equivalents.
(a) Gm and m **(b)** Tg and g
(c) L and dL **(d)** s and cs

13. Write two unit factors for each of the following metric relationships.
(a) km and m **(b)** Mg and g
(c) L and mL **(d)** s and μs

14. Write two unit factors for each of the following metric relationships.
(a) Gm and m **(b)** Tg and g
(c) L and dL **(d)** s and cs

Metric–Metric Conversions (Sec. 2.3)

15. Perform the following metric–metric conversions.
(a) 5.00 m to km **(b)** 5.00 g to cg
(c) 5.00 L to dL **(d)** 5.00 s to ns

16. Perform the following metric–metric conversions.
(a) 5.00 Mm to m **(b)** 5.00 mg to g
(c) 5.00 mL to L **(d)** 5.00 ds to s

17. Perform the following metric–metric conversions.
(a) 6.50 Tm to Mm **(b)** 650 Gg to kg
(c) 0.650 cL to dL **(d)** 0.000 650 ns to ps

18. Perform the following metric–metric conversions.
(a) 7.50 km to Gm **(b)** 750 Mg to Tg
(c) 0.750 pL to μL **(d)** 0.000 750 ms to ns

Metric–English Conversions (Sec. 2.4)

19. State the following metric–English equivalents.
(a) 1 in. = ? cm **(b)** 1 lb = ? g
(c) 1 qt = ? mL **(d)** 1 sec = ? s

20. State the following metric–English equivalents.
(a) 1 m = ? yd **(b)** 1 kg = ? lb
(c) 1 L = ? qt **(d)** 1 s = ? sec

21. Perform the following metric–English conversions.
(a) 25 cm to in. **(b)** 25 g to lb
(c) 2.50 qt to mL **(d)** 2.50 sec to s

22. Perform the following metric–English conversions.
(a) 65 in. to cm **(b)** 65 lb to g
(c) 65.0 mL to qt **(d)** 6.50×10^{-3} s to sec

23. Perform the following metric–English conversions.
(a) 72 in. to m **(b)** 175 lb to kg
(c) 0.500 qt to L **(d)** 4.55×10^{-4} min to ds

24. Perform the following metric–English conversions.
(a) 800.0 m to yards **(b)** 0.375 kg to lb
(c) 1250 mL to gallons **(d)** 1.52×10^{3} ds to min

25. The tallest office building in the world is the "Freedom Tower" in New York City, which measures 1776 feet. What is the height expressed in meters?

26. The tallest structure in the world is the "Khalifa Tower" in Dubai, which measures 2722 feet. What is the height expressed in meters?

27. A Honda Accord hybrid gets 19 kilometers per liter city driving. What is the mileage in miles per gallon? (Given: 1 km = 0.621 mi and 1 L = 0.264 gal)

28. A Honda Accord hybrid gets 21 kilometers per liter highway driving. What is the mileage in miles per gallon? (Given: 1 km = 0.621 mi and 1 L = 0.264 gal)

The Percent Concept (Sec. 2.5)

29. A sterling silver spoon has a mass of 65.5 g and contains 7.50% copper. Find the mass of copper in the spoon.

30. A stainless steel spoon has a mass of 55.5 g and contains 10.5% chromium. Find the mass of chromium in the spoon.

31. If a solution contains 255 mL of ethanol and 375 mL of water, what is the percent ethanol in the solution?

32. If 20.0 gallons of gasohol contains 2.40 gal of ethanol, what is the percent ethanol in the gasohol?

33. In 1982 the U.S. Mint began to cast penny coins made of zinc plated with copper. If a 1990 penny has a mass of 2.554 g and contains 2.490 g zinc, what is the percent copper in the coin?

34. Before 1982 the U.S. Mint cast penny coins made of a copper and zinc mixture. If a 1980 penny has a mass of 3.051 g and contains 0.153 g zinc, what is the percent copper in the coin?

Volume by Calculation (Sec. 2.6)

35. Complete the following volume equivalents.
(a) 1 mL = ? cm^3 **(b)** 1 in.^{3} = ? cm^3

36. Complete the following volume equivalents.
(a) 1 L = ? cm^3 **(b)** 1 in.^{3} = ? mL

37. A piece of green jade is a rectangular solid measuring 2.50 cm by 1.25 cm by 3.50 mm. What is the volume in cubic centimeters?

38. A piece of black onyx is a rectangular solid measuring 5.00 cm by 5.00 cm by 2.50 mm. What is the volume in cubic centimeters?

39. A sheet of aluminum foil measures 75.0 cm by 35.0 cm. If the volume is 5.00 cm^3, what is the thickness of the foil?

40. A sheet of aluminum measures 4.95 cm by 2.45 cm. If the volume is 2.75 cm^3, what is the thickness of the foil?

41. If an automobile engine is 325-in.^{3}, what is the volume in liters?

42. If a motorcycle engine is 0.500-L, what is the volume in cubic inches?

Volume by Displacement (Sec. 2.7)

43. The initial water level in a 10-mL graduated cylinder reads 4.5 mL. After a ruby gemstone is dropped into the cylinder, the water level reads 7.0 mL. What is the volume of the ruby?

44. The initial water level in a 10-mL graduated cylinder is 5.0 mL. After a sapphire gemstone is added into the cylinder, the water level is 6.5 mL. What is the volume of the sapphire?

45. Magnesium metal reacts with acid to produce hydrogen gas. The gas displaces water from a 100-mL graduated cylinder, and the water level decreases from 95.0 mL to 32.5 mL. What is the volume of gas produced by the reaction?

46. Calcium metal reacts with water to produce hydrogen gas. The gas displaces water from a 100-mL graduated cylinder, and the water level decreases from 75.5 mL to 43.0 mL. What is the volume of gas produced by the reaction?

The Density Concept (Sec. 2.8)

47. State whether the following will sink or float when dropped into water.
(a) wax ($d = 0.90$ g/cm^3) **(b)** marble ($d = 3.5$ g/cm^3)

48. State whether the following will sink or float when dropped into water.
(a) redwood ($d = 1.2$ g/mL)
(b) bamboo ($d = 0.40$ g/mL)

49. Will a balloon filled with the given gas rise in the air or drop to the ground? (Assume the mass of the balloon is negligible and the density of air is 1.29 g/L.)
(a) helium ($d = 0.178$ g/L) **(b)** argon ($d = 1.78$ g/L)

50. Will a balloon filled with the given gas rise in the air or drop to the ground? (Assume the mass of the balloon is negligible and the density of air is 1.29 g/L.)
(a) laughing gas (d = 1.96 g/L)
(b) ammonia (d = 0.759 g/L)

51. State the density of water in grams per milliliter.

52. State the density of water in kilograms per liter.

53. Calculate the density in grams per milliliter for each of the following.
(a) 25.0 mL of ethyl alcohol having a mass of 19.7 g
(b) 10.0 g of ethyl ether having a volume of 14.0 mL

54. Calculate the density in grams per milliliter for each of the following.
(a) 25.5-g solid whose volume is found by displacement to be 4.5 mL
(b) 95.5-g rectangular solid measuring 3.55 cm × 2.50 cm × 1.75 cm

55. Calculate the mass in grams for each of the following liquids.
(a) 250 mL of gasoline ($d = 0.69$ g/mL)
(b) 150 mL of ethanol ($d = 0.79$ g/mL)

56. Calculate the mass in grams for each of the following solids.
(a) 5.00 cm^3 of table salt ($d = 2.18\ \text{g/cm}^3$)
(b) 2.50 cm^3 of table sugar ($d = 1.59\ \text{g/cm}^3$)

57. Calculate the volume in milliliters for each of the following liquids.
(a) 25.0 g of ether ($d = 0.714$ g/mL)
(b) 15.0 g of acetone ($d = 0.792$ g/mL)

58. Calculate the volume in milliliters for each of the following solids.
(a) 1.00 kg of silicon ($d = 2.33\ \text{g/cm}^3$)
(b) 1.00 kg of titanium ($d = 4.51\ \text{g/cm}^3$)

Temperature (Sec. 2.9)

59. State the freezing point of water on the following temperature scales.
(a) Fahrenheit **(b)** Celsius **(c)** Kelvin

60. State the boiling point of water on the following temperature scales.
(a) Fahrenheit **(b)** Celsius **(c)** Kelvin

61. Express the following Fahrenheit temperatures in degrees Celsius.
(a) 101 °F **(b)** −215 °F

62. Express the following Celsius temperatures in degrees Fahrenheit.
(a) 19 °C **(b)** −175 °C

63. Express the following Celsius temperatures in Kelvin units.
(a) 495 °C **(b)** −185 °C

64. Express the following Kelvin temperatures in degrees Celsius.
(a) 273 K **(b)** 100 K

65. The lowest recorded climatic temperature is 184 K at Vostok, Antarctica. Express the temperature in degrees Fahrenheit.

66. The highest recorded climatic temperature is 329 K in Death Valley, California. Express the temperature in degrees Fahrenheit.

The Heat Concept (Sec. 2.10)

67. State the solid listed in Table 2.5 that is the best conductor of heat.

68. State the liquid listed in Table 2.5 that is the worst conductor of heat.

69. If a gas furnace releases 450 kcal of heat energy, what is the energy in kilojoules? (1 cal = 4.184 J)

70. If a gas stove releases 120 kJ of heat energy, what is the energy in kilocalories? (1 cal = 4.184 J)

71. Calculate the heat required to raise 25.0 g of iron (0.108 cal/g × °C) from 25.0 °C to 50.0 °C.

72. Calculate the heat released as 35.5 g of copper (0.0920 cal/g × °C) cools from 50.0 °C to 25.0 °C.

73. Find the specific heat of gold if 25.5 cal are required to heat 30.0 g of gold from 27.5 °C to 55.0 °C.

74. Find the specific heat of platinum if 35.7 cal are lost as 75.0 g of the metal cools from 43.5 °C to 29.0 °C.

75. Calculate the mass of titanium (0.125 cal/g × °C) that requires 75.5 cal to heat the metal from 20.5 °C to 32.0 °C.

76. Calculate the mass of lead (0.0382 cal/g × °C) that releases 52.5 cal as the metal cools from 35.5 °C to 25.0 °C.

General Exercises

77. If a computer hard disk has 1.5 terabytes (TB) of memory, what is the storage capacity in megabytes (MB)?

78. If a computer hard disk has 850 gigabytes (GB) of memory, what is the storage capacity in megabytes (MB)?

79. If a DVD can store 4.7 gigabytes (GB) of data, and a computer hard drive has a capacity of 1.5 terabytes (TB), how many DVDs of data can be loaded onto the hard drive?

80. If a USB flash drive can store 16 gigabytes (GB) of data, and a computer hard drive has a capacity of 2.0 terabytes (TB), how many flash drives of data can be loaded onto the hard drive?

81. An automobile airbag inflates in 35 ms. How many microseconds are required for the airbag to inflate?

82. An automobile antilock brake system (ABS) operates at 30 pulses per second. How many times do the brakes pulse in 1.00 ds?

83. A light year is the distance light travels in 1.00 year. Given the velocity of light, 1.86×10^5 mi/s, how many miles does light travel in a light year?

84. A parsec is the distance light travels in 3.26 years. Given the velocity of light, 3.00×10^8 m/s, how many kilometers does light travel in a parsec?

85. A basketball court measures 94.0 feet by 50.0 feet. Calculate the area in square meters. (Given: 1 yd = 0.914 m)

86. A football field measures 100.0 yards by 160.0 feet. Calculate the area in square meters. (Given: 1 yd = 0.914 m)

87. Olympic athletes compete in a 400-meter event, but not in a quarter mile event. Which race is longer: 400 meters or 0.250 mile? (Given: 1 mi = 1.61 km)

88. Olympic athletes compete in a 1500-meter event, but not in a mile event. Which race is longer: 1500 meters or 1 mile? (Given: 1 mi = 1.61 km)

89. How many 325-mg tablets can be produced from 2.50 kg of powdered aspirin?

90. How many molecules are in one drop of water if 1.00 g of water contains 3.34×10^{22} molecules? (Given: 1 mL = 20 drops)

91. What is the specific gravity of gasohol if the density is 0.801 g/mL?

92. What is the specific gravity of jet fuel if the density is 0.775 g/mL?

Challenge Exercises

93. Express the density of water in the English units of pounds per gallon.

94. Express the density of water in the English units of pounds per cubic foot.

95. The radius (r) of the international reference kilogram cylinder is 1.95 cm. Assuming the density of the kilogram is 21.50 g/cm^3, calculate its height (h). The volume of a cylinder equals $\pi r^2 h$, where π is the constant 3.14.

96. The density of mercury is 13.6 g/mL. Express the density in SI units (kg/m^3).

Chapter 2 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following is a base unit and symbol in the metric system? (Sec. 2.1)
(a) centimeter (cm) **(b)** kilogram (kg)
(c) milliliter (mL) **(d)** all of the above
(e) none of the above

2. What are the two unit factors that relate meters and millimeters? (Sec. 2.2)
(a) 1000 mm/1 m and 1 m/1000 mm
(b) 1000 mm/1 m and 1000 m/1 mm
(c) 1 mm/1000 m and 1 m/1000 mm
(d) 1 mm/1000 m and 1000 m/1 mm
(e) none of the above

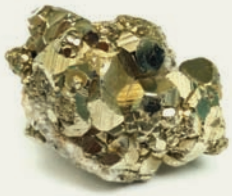
3. What is the three-step sequence in applying the unit analysis method of problem solving? (Sec. 2.3)
(a) 1–unknown unit, 2–unit factor, 3–relevant given value
(b) 1–unknown unit, 2–relevant given value, 3–unit factor
(c) 1–relevant given value, 2–unknown unit, 3–unit factor
(d) 1–unit factor, 2–unknown unit, 3–relevant given value
(e) 1–unit factor, 2–relevant given value, 3–unknown unit

4. If a chemistry student weighs 155 pounds, what is the mass in kilograms? (Sec. 2.4)
(a) 0.341 kg **(b)** 0.394 kg
(c) 70.4 kg **(d)** 341 kg
(e) 70,400 kg

5. A sample of 18K gold contains the following by mass: 18.0 grams gold, 3.0 grams silver, and 3.0 grams copper. What is the percent gold? (Sec. 2.5)
(a) 18% **(b)** 25%
(c) 33% **(d)** 75%
(e) 90%

6. If a rectangular aluminum block measures 3.80 cm by 2.55 cm by 1.25 cm, what is the volume of the rectangular solid? (Sec. 2.6)
(a) 0.0826 cm^3 **(b)** 1.19 cm^3
(c) 1.86 cm^3 **(d)** 7.75 cm^3
(e) 12.1 cm^3

7. What is the volume of a sample of fool's gold that displaces the water level from 25.5 mL to 35.0 mL in a graduated cylinder? (Sec. 2.7)
(a) 9.5 mL **(b)** 11.5 mL
(c) 25.5 mL **(d)** 35.0 mL
(e) 60.5 mL



▶ **Fool's Gold** Iron pyrite, FeS_2, is commonly referred to as fool's gold because of its yellow metallic luster.

8. If 50.0 mL of gasohol has a mass of 37.5 g, what is the density of the gasohol in grams per cubic centimeter? (Sec. 2.8)
(a) 0.00750 g/cm^3 **(b)** 0.0750 g/cm^3
(c) 0.750 g/cm^3 **(d)** 7.50 g/cm^3
(e) 10.0 g/cm^3

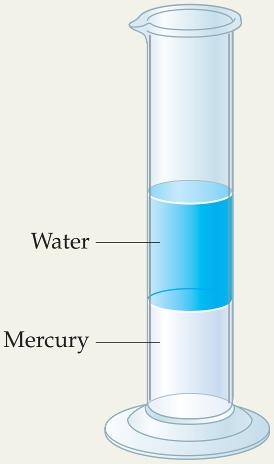
9. A rare metal alloy is a superconductor at 55 K. What is the temperature on the Celsius scale? (Sec. 2.9)
(a) −328 °C **(b)** −218 °C
(c) −55 °C **(d)** 218 °C
(e) 328 °C

10. Which of the following expresses the total heat energy in a sealed capsule? (Sec. 2.10)
(a) 20 °C **(b)** 68 °F
(c) 293 K **(d)** 20 kcal
(e) all of the above

Key Concepts

11. A tall glass cylinder contains water ($d = 1.00$ g/mL) floating on liquid mercury ($d = 13.6$ g/mL). If a piece of fool's gold ($d = 5.00$ g/mL) and a gold nugget ($d = 18.3$ g/mL) are dropped into the cylinder, where does each come to rest?

12. Which of the following temperatures is the coldest: 0 °F, 0 °C, or 0 K?



◀ **The Density Concept** A glass cylinder is shown containing water and liquid mercury.

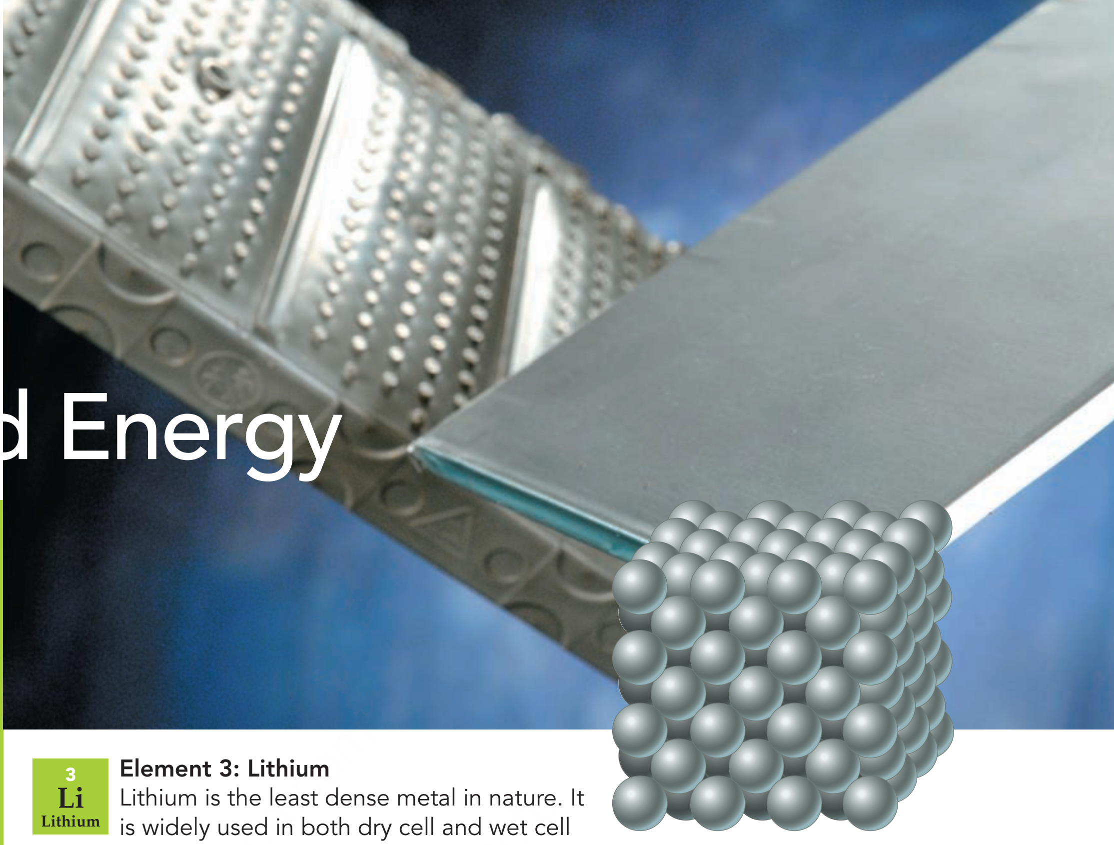
Critical Thinking

13. A silver dime is smaller and thinner than a zinc penny. Why is the mass of the 10¢ coin greater than the 1¢ coin?

14. Given the two liquids, gasoline ($d = 0.70$ g/mL) and chloroform ($d = 1.5$ g/mL), how could you identify each liquid given a beaker of water and the two liquids?

CHAPTER 3

d Energy



"Matter is discontinuous and cannot be infinitely divided."

Democritus, Greek Philosopher (450–370 B.C.)

"Matter is continuous and can be infinitely divided."

Aristotle, Greek Philosopher (384–322 B.C.)

3 **Li** Lithium

Element 3: Lithium

Lithium is the least dense metal in nature. It is widely used in both dry cell and wet cell batteries. A lithium-ion battery is the latest technology used in hybrid vehicles such as the Chevy Volt.

In the seventeenth century, British scientist Robert Boyle (1627–1691) established the importance of experiments in the study of science. Boyle realized the value of laboratory research and described his work thoroughly so that other scientists could repeat his procedures and confirm his observations. In 1661, he published *The Sceptical Chymist*, which marked a turning point in science by suggesting that theories were no better than the experimental methods that supported them.

Boyle rejected the Greek notion that air, earth, fire, and water were basic elements. Instead, he proposed that an element was a substance that could not be broken down further. Thus, an element must gain weight when it undergoes a chemical change and combines with another element. This was a practical proposal that could be tested in the laboratory.

Following the lead of Boyle, others established scientific principles by performing experiments that could be verified by fellow scientists. Although scientists continued to propose theories about the behavior of matter, their theories were now based on laboratory evidence.

3.1 Physical States of Matter

LEARNING OBJECTIVES

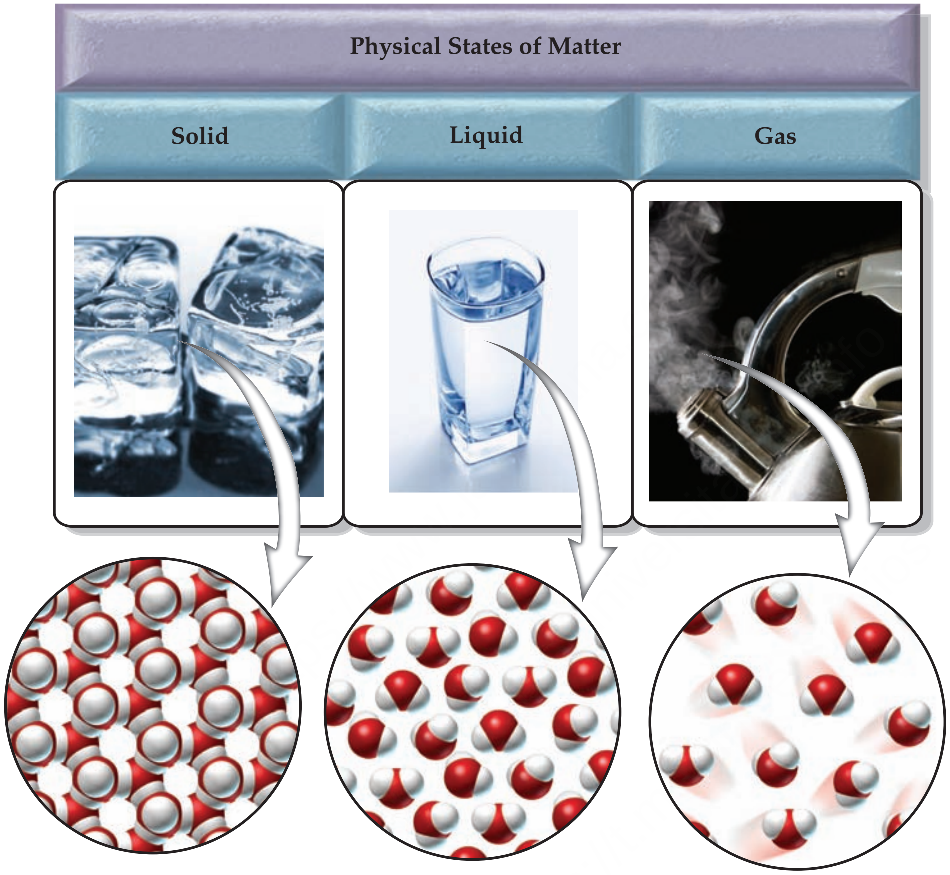
- Describe the motion of particles in the solid, liquid, and gaseous states of matter.
- Describe the effect of temperature on the solid, liquid, and gaseous states of matter.

Matter is defined as a substance that has mass and occupies volume. It exists in one of three **physical states**: solid, liquid, or gas. In the solid state, matter has a fixed shape and a fixed volume. The particles in a solid are tightly packed, held in rigid positions, and cannot be compressed together. In the liquid state, matter has a variable shape, but a fixed volume. The particles in a liquid are loosely packed, free to move around, and compress only slightly.

In the gaseous state, tiny particles of matter are widely spaced apart and uniformly distributed throughout its container. If the volume increases, the gas expands and the particles move farther away from each other. If the volume decreases, the gas compresses and the particles move closer to one another. Table 3.1 summarizes the properties of the three states of matter.

TABLE 3.1 Physical States of Matter

Property	Solid	Liquid	Gas
shape	fixed	variable	variable
volume	fixed	fixed	variable
compressibility	not significant	not significant	significant



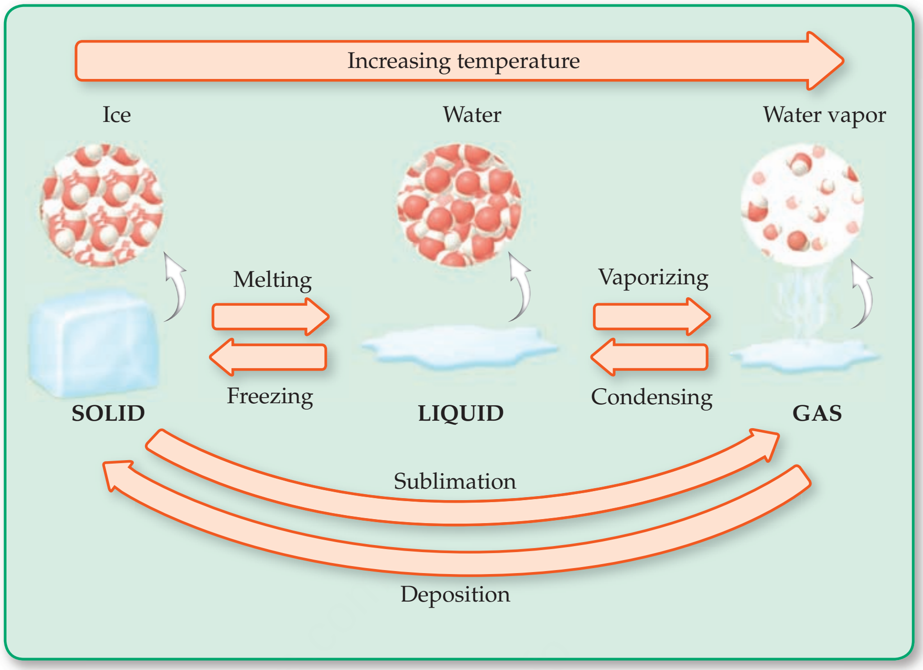
◀ **Physical States of Matter** Ice, water, and steam represent the three states of matter: solid, liquid, and gas.

Even though we may describe a substance as being a solid, every substance can exist as a solid, liquid, or gas. We can change the physical state of a substance by changing the temperature. Water, for example, can be changed from a liquid to solid ice at 0 °C, and from a liquid to gaseous steam at 100 °C. Other substances behave similarly at different temperatures. Iron, for example, can be changed to a molten liquid at 1535 °C and to a gas at 2750 °C.

We can describe changes of physical state as follows. As temperature increases, a solid *melts* to a liquid, and then the liquid *vaporizes* to a gas. A direct change of state from a solid to a gas is called **sublimation**. Dry ice, for example, undergoes sublimation and disappears as it changes directly from a solid to a gas.

Conversely, as temperature decreases, a gas *condenses* to a liquid, and the liquid then *freezes* to a solid. A direct change of state from a gas to a solid is called **deposition**. The freezer compartment in a refrigerator may demonstrate deposition by collecting ice. Opening the refrigerator door allows moist air to enter, which deposits frost without a trace of liquid water. Figure 3.1 shows the relationship of temperature and physical state.

▲ **Sublimation and Deposition** Iodine crystals in a heated beaker undergo sublimation to a purple gas. As the gas cools, it undergoes deposition to a solid on the upper rim of the beaker.



▶ **Figure 3.1 Changes in Physical State** As temperature increases, a solid melts to a liquid and then vaporizes into a gas. As temperature decreases, a gas condenses to a liquid and then freezes to a solid.

EXAMPLE EXERCISE 3.1 Change of Physical State

State the term that applies to each of the following changes of physical state:

(a) Snow changes from a solid to a liquid.
(b) Gasoline changes from a liquid to a gas.
(c) Dry ice changes from a solid to a gas.

Solution

Refer to Figure 3.1 for the changes of physical state.

(a) The change from solid to liquid is called *melting*.
(b) The change from liquid to gas is called *vaporizing*.
(c) The change from solid to gas is called *sublimation*.

Practice Exercise

State the term that applies to each of the following changes of physical state:

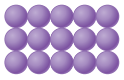
(a) A refrigerant changes from a gas to a liquid.
(b) Water changes from a liquid to a solid.
(c) Iodine vapor changes from a gas to a solid.

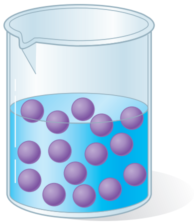
Answers:

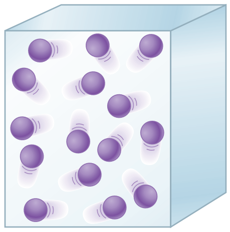
(a) condensing (b) freezing (c) deposition

Concept Exercise

Identify the physical state (solid, liquid, gas) that corresponds to each of the following pictorial representations:







Answer: See Appendix G, 3.1.

3.2 Elements, Compounds, and Mixtures

LEARNING OBJECTIVE

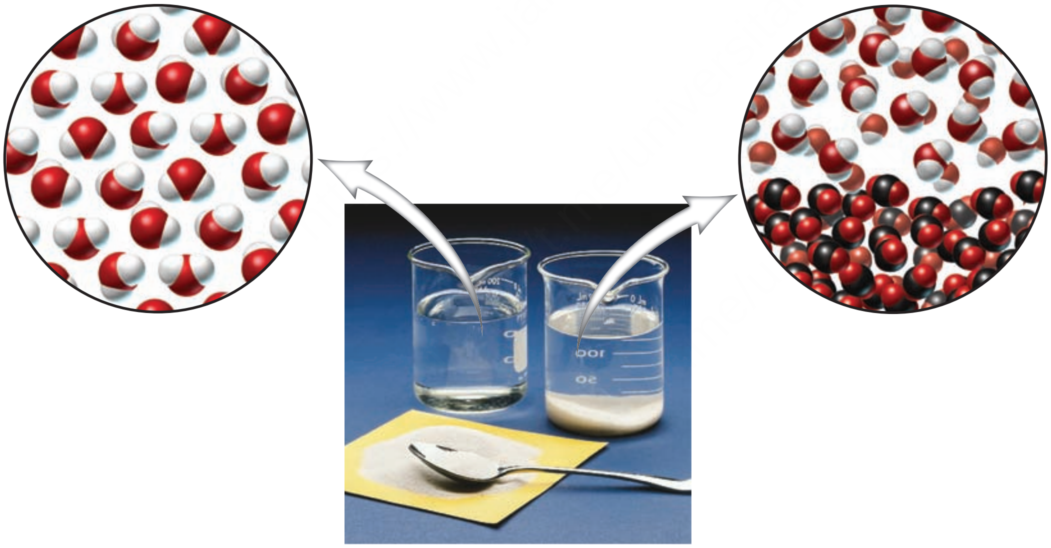
- Classify a sample of matter as an element, compound, or mixture.

A sample of matter may have properties that are consistent throughout or that vary. One way to tell if the properties of matter are consistent is to melt the sample. A sample of pure gold always melts at 1064 °C and does not vary. It makes no difference whether it is a large gold nugget or a small gold flake; the yellow metal consistently melts at 1064 °C.

Now consider a sample of quartz rock that contains a vein of pure gold. Although the quartz sample is observed to melt over a broad range from 1000 °C to 1600 °C, pure gold always melts precisely at 1064 °C. Scientists classify the quartz rock as heterogeneous and gold as homogeneous. A sample of matter is said to be *heterogeneous* if its properties are indefinite and vary. A sample of pure gold is said to be *homogeneous* because its properties are definite and constant.

A **heterogeneous mixture** can be separated into pure substances by physical methods. For example, panning for gold uses the property of density to separate pure gold from sand and rock. Gold is quite dense, 19.3 g/cm^3, and remains at the bottom of the pan, whereas the less dense sand and rock are swirled away with water.

Air is not a heterogeneous mixture, but rather a homogeneous mixture of nitrogen, oxygen, and other gases. Similarly, salt water is a homogeneous mixture of salt and water. Unlike those of a heterogeneous mixture, the properties of a **homogeneous mixture** are constant for a given sample. However, a homogeneous mixture may have properties that vary from sample to sample. For example, samples of salt water from the Pacific Ocean and from the Dead Sea have different properties. Salt water from the Dead Sea has a higher density and a greater concentration of dissolved minerals.



▲ **Heterogeneous Mixture** Sand and water are each pure substances. However, they form a heterogeneous mixture when added together, and the properties vary throughout the sample.

An **alloy** is a homogeneous mixture of two or more metals. Examples of alloys include 10K, 14K, and 18K gold jewelry. Although 10K, 14K, and 18K jewelry may contain only gold, silver, and copper metals, the amount of gold varies from 42% to 75%. Gold alloy is a homogeneous mixture; therefore the properties from different samples can vary. For instance, 10K gold is a harder alloy and is more scratch resistant than 18K gold.

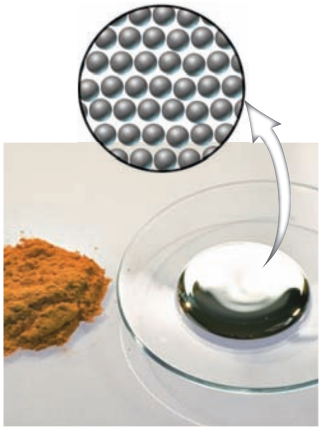
A **substance** is matter that has definite composition and constant properties. A substance is either a compound or an element. A **compound** has predictable properties, but can be broken down into elements by an ordinary chemical reaction. Table sugar is an example of a compound that can be broken down into carbon, hydrogen, and oxygen.

An **element** is a substance that cannot be broken down further by a chemical reaction. For example, table salt is a compound composed of the elements sodium and chlorine, which cannot be broken down any further. Figure 3.2 shows the overall classification of matter.

Helpful Hint
Homogeneous vs. Heterogeneous Matter

Students often confuse the terms homogeneous mixture and heterogeneous mixture. A *homogeneous mixture* is matter having a definite composition although the amount of each component can vary. For example, brass is a homogeneous mixture, an alloy of copper and zinc. The proportions of copper and zinc are the same throughout a given brass sample, and the physical properties are uniform throughout the sample.

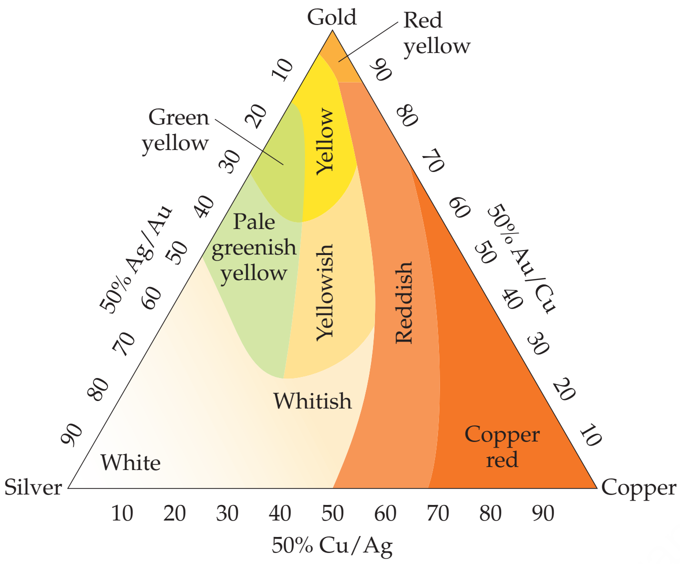
A *heterogeneous mixture* is matter having a variable composition within the sample, and the physical properties vary within the sample. For example, when gold occurs naturally it is often found in the mineral quartz. Because the gold occurs as veins of yellow metal within the quartz, the amount of gold varies within the sample. Gold in a quartz sample is a heterogeneous mixture, and physical properties such as density will vary throughout the sample.



▲ **Compound, Element, Mixture** The orange compound, mercury oxide, decomposes to give the element mercury, which is a silver liquid. Adding the orange powder and the metallic liquid together produces a heterogeneous mixture.

A CLOSER LOOK 18K Gold

Q: **Is all 18K gold equally valuable?**



Pure gold is assigned a value of 24 karat (24K). However, gold is a relatively soft metal and is easily scratched. Gold jewelry is usually alloyed with another metal, such as silver and copper, to impart color and resistance to scratching. The following figure indicates the color of a gold based on the percentages of silver and copper in the gold alloy.

▲ **18K Gold Necklace**

Sometimes, nickel, manganese, and other transition metals are used to create white gold. Zinc can be used to produce a dark yellow gold, or reddish gold. Gray gold alloys are usually made from gold and palladium. Because palladium is an expensive metal, a cheaper method to produce gray gold is to add silver, manganese, and copper to the alloy.

A: Not all 18K gold jewelry appears the same, or is equally valuable! However, all 18K gold jewelry must be 75% gold by weight. That is, pure gold is 24K, and 18K gold is 75% gold. Thus, 18K gold contains 75% gold, and 25% other metals. Depending on the other metals in the 18K gold jewelry, the color and value will vary.

EXAMPLE EXERCISE 3.2 Element, Compound, or Mixture

Consider the following properties of the element copper:

(a) Copper metal cannot be broken down by a chemical change.
(b) Copper reacts with oxygen in air to give copper oxide.
(c) Copper, in the form of malachite ore, is found worldwide.
(d) Copper and tin compose bronze alloy.

Classify each of the following copper samples as an element, a compound, a homogeneous mixture, or a heterogeneous mixture:

(a) copper wire
(b) copper oxide
(c) malachite ore
(d) bronze alloy

Solution

Refer to Figure 3.2 to classify each sample.

(a) Copper wire is a metallic *element*.
(b) Copper oxide is a *compound* of the elements copper and oxygen.
(c) Malachite ore is a *heterogeneous mixture* of copper and other substances.
(d) Bronze alloy is a *homogeneous mixture* of copper and tin.

Practice Exercise

Consider the following properties of the element mercury:

(a) Mercury liquid cannot be broken down by a chemical change.
(b) Mercury oxide can be heated to give mercury and oxygen gas.
(c) Mercury, in form of cinnabar ore, is found in Spain and Italy.
(d) Mercury and silver compose the alloy used for dental fillings.

Classify each of the following mercury samples as an element, a compound, a homogeneous mixture, or a heterogeneous mixture:

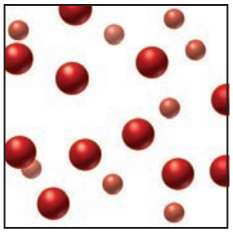
(a) mercury liquid
(b) mercury oxide
(c) cinnabar ore
(d) dental alloy

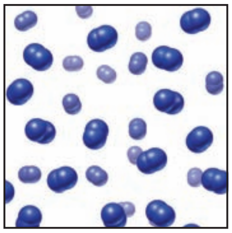
Answers:

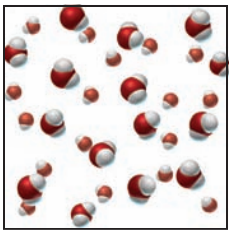
(a) element	(b) compound
(c) heterogeneous mixture	(d) homogeneous mixture

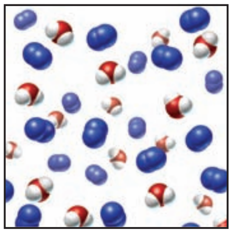
Concept Exercise

Classify each of the following gases as an element, a compound, or a mixture of gases, as shown in the illustration:

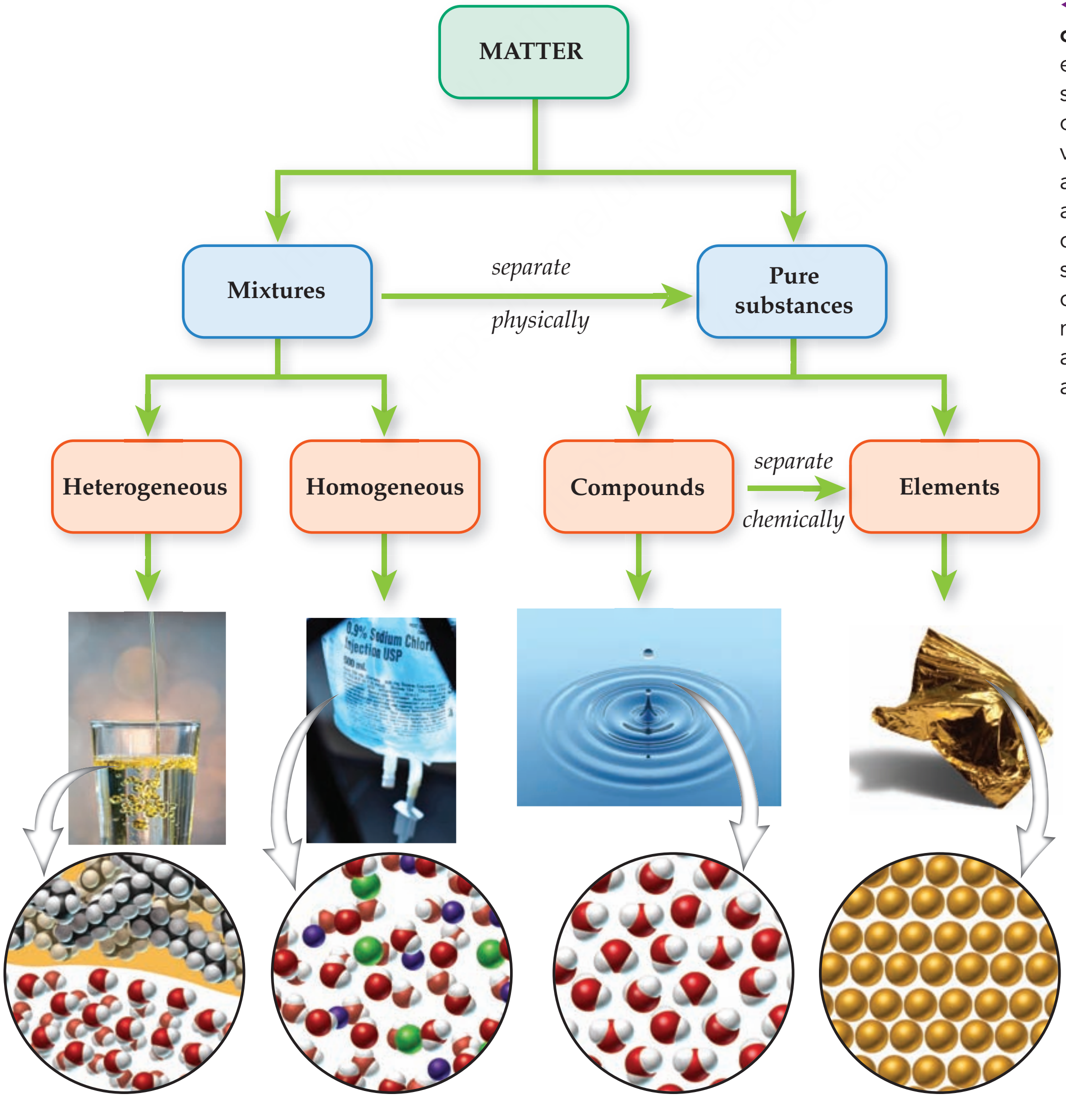

(a)


(b)


(c)


(d)

Answer: See Appendix G, 3.2.



◀ **Figure 3.2 Classification of Matter** Matter may be either a mixture or a pure substance. The properties of a heterogeneous mixture vary within the sample (oil and water). The properties of a homogeneous mixture are constant (salt solution). A pure substance may be either a compound (water) or an element (gold). Left to right: oil and water; NaCl solution; H_2O; and gold foil.

LEARNING OBJECTIVE

- Memorize the names and symbols of 48 common elements.

3.3 Names and Symbols of the Elements

There are 81 stable elements that occur in nature. In addition, there are a few other naturally occurring elements, such as uranium, which are unstable. In total, there are over 100 elements, but only 10 account for 95% of the mass of Earth's crust, water, and atmosphere.

Oxygen is the most abundant element in Earth's crust, water, and atmosphere. It is found combined with hydrogen in water and with silicon in sand and rocks. Oxygen is found in air and constitutes about 21% of our atmosphere. The mass of oxygen in Earth's crust, water, and atmosphere is about equal to the total mass of all other elements. Table 3.2 lists the 10 most abundant elements.

Oxygen, silicon, and aluminum are the three most abundant elements on Earth. The elements oxygen, carbon, hydrogen, nitrogen, calcium, and phosphorus account for over 99% of the mass of the human body. The remaining 1% consists of trace elements, many of which are essential to human life. For example, a trace of iron is necessary to bind oxygen to hemoglobin in blood. Figure 3.3 compares the distribution of elements in Earth's crust and in the human body.

TABLE 3.2 Elements in Earth's Crust, Water, and Atmosphere

Element	Mass Percent	Element Mass	Percent
oxygen	49.5%	sodium	2.6%
silicon	25.7%	potassium	2.4%
aluminum	7.5%	magnesium	1.9%
iron	4.7%	hydrogen	0.9%
calcium	3.4%	titanium	0.6%
		all other elements	0.5%

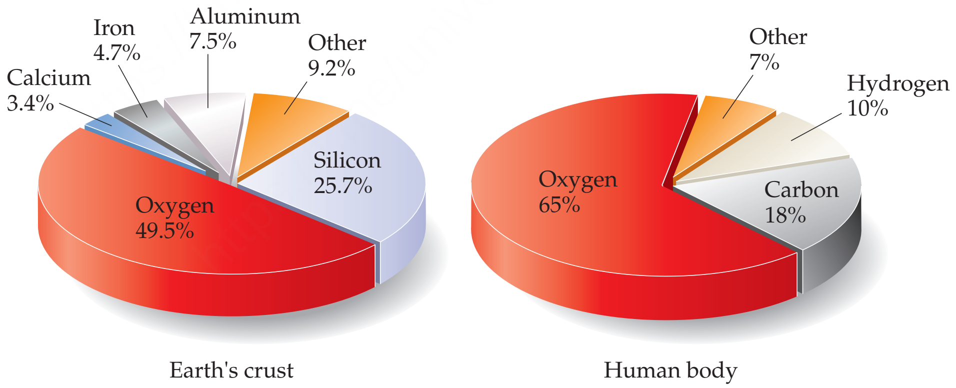
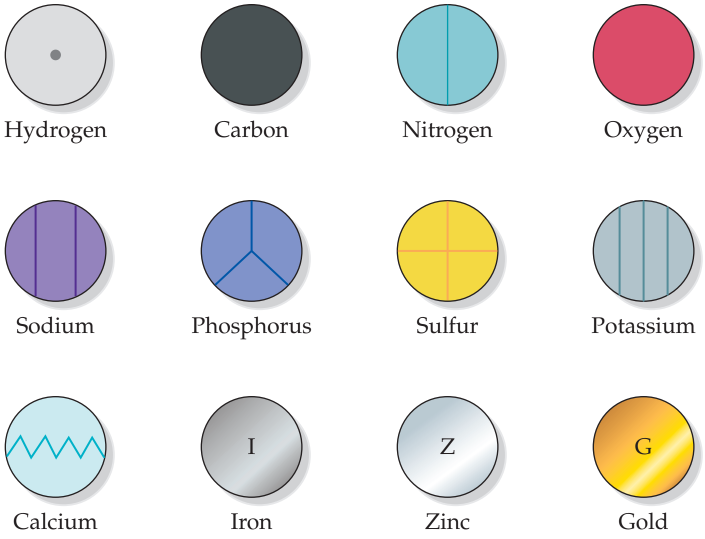


Figure 3.3 Abundance of Elements Notice that oxygen is the most abundant element in both Earth's crust and the human body. Although aluminum is abundant in Earth, there is evidence that this element is toxic to the human body.

The names of elements are derived from various sources. For example, hydrogen is derived from the Greek word *hydro* meaning "water former." Carbon is derived from the Latin word *carbo*, meaning "coal." Calcium is derived from the Latin word *calcis*, which translates as "lime," a mineral source of calcium. Some elements are named for their region of discovery. For example, germanium comes from Germany, and scandium from Scandinavia. Several elements are named for famous scientists, such as curium (Marie Curie) and nobelium (Alfred Nobel).

The name of each element is abbreviated using a **chemical symbol**. In 1803, the English chemist John Dalton (1766–1844) proposed that elements are composed of indivisible, spherical particles. Dalton referred to each of these individual particles as an **atom**, from the Greek word *atomos* meaning "indivisible." He suggested the use of circles with enclosed markings as symbols for the elements. Figure 3.4 shows selected symbols chosen by Dalton to represent elements.



◀ **Figure 3.4 Dalton's Symbols for Selected Elements** Dalton symbolized each element using a circle with an inscribed marking. Dalton's system was not practical, but it did convey the idea that elements were composed of atomic spheres.

A CLOSER LOOK Aluminum or Aluminium?

Q: What suffix is commonly attached to the name of an element to indicate the element is a metal?

At the turn of the nineteenth century, many new elements were isolated and added to a growing list. For example, sodium, potassium, magnesium, calcium, strontium, and barium were all discovered in the period from 1807 to 1808 by the English chemist Humphry Davy (1778–1829). Metallic elements were systematically named by adding an *–ium* suffix to its root name. The nonmetal selenium was discovered a few years later in 1817, and named for the Greek word *selene,* which means "moon." The semimetal tellurium had been previously discovered in 1782, and named for the Latin word *tellus,* which means "earth."

◀ **Aluminum Cap on the Washington Monument** The Washington Monument, in Washington, D.C., is capped by a shiny metal pyramid. The thin metal pyramid weighing 2.85 kilograms is composed of aluminum, which was an expensive metal at the time of its construction in 1884.

In 1808, Davy proposed the name *alumium* for an element that he was attempting to isolate from the mineral alumina. A couple of years later he changed the name to *aluminium* for the element that was not isolated until Charles Martin Hall accomplished the feat in 1885 (see Chemistry Connection in Chapter 1).

Today, the official International Union of Pure and Applied Chemistry (IUPAC) name for the element is *aluminium*; however, both Americans and Canadians refer to it as aluminum. The name *aluminum* has been explained in a couple of ways. According to Oberlin College in Ohio, the name of the element was misspelled on a notice that publicized Charles Martin Hall's discovery of the process that separated the metal from its ore. Another account is that a spelling error on the stationery of the Aluminum Company of America (ALCOA), a company that Charles Hall helped to found, was instrumental in dictating the name. In either case, it appears that a spelling error was responsible for the irregular ending for the name of the element that is extremely important today, and second only to steel as a construction metal.

For most of the nineteenth century, Americans used the *–ium* suffix and *aluminium* appeared in Webster's dictionary in 1828. However, the misspelling of the metal was common, and in 1926 the American Chemical Society decided officially to call the metal *aluminum.* Meantime, in Europe the *–ium* ending was the accepted spelling. In 1990, IUPAC adopted aluminium as the official international name for the metal, which is used by French, German, Danish, Dutch, Norwegian, and Swedish scientists. In 1993, IUPAC officially approved the alternate spelling aluminum, but stated a preference for the name *aluminium.*

A: An *–ium* suffix is the most common ending for metallic elements, although there are a few exceptions such as platinum (Pt), molybdenum (Mo), and tantalum (Ta). In addition, the Latin names of elements known since ancient times such as cuprum (Cu), argentum (Ag), and aurum (Au) are exceptions.

In 1813, the Swedish chemist J. J. Berzelius (1779–1848) proposed our current system of symbols for the elements. He suggested that a symbol correspond to the first letter of the name of the element; for example, H for hydrogen, O for oxygen, and C for carbon. Furthermore, when elements start with the same letter, he recommended using two letters in the name, for example, Ca for calcium, Cd for cadmium, and Cl for chlorine.

In some instances, the chemical symbol is derived from the original Latin name of the element. For example, the symbol Pb for lead is derived from the Latin *plumbum*. Interestingly, the Romans used lead pipe to transport water, which explains the derivation of our word "plumbing." The symbol Hg for mercury is derived from the Greek *hydrargyro*, meaning "liquid silver." Similarly, the symbols for silver (Ag), gold (Au), copper (Cu), iron (Fe), potassium (K), sodium (Na), antimony (Sb), and tin (Sn), are derived from the original Latin and Greek names.

When writing symbols for the names of elements, it is important to follow convention to avoid confusion. The first letter is always capitalized, and the second is lowercase. For example, the symbol for the metallic element cobalt is Co. In contrast, notice that the formula for the deadly gas carbon monoxide is CO. Table 3.3 lists the names and symbols of selected elements.

TABLE 3.3 Names and Symbols of Selected Elements

Name of Element	Symbol	Name of Element	Symbol	Name of Element	Symbol	Name of Element	Symbol
aluminum	Al	chlorine	Cl	lead	Pb	radium	Ra
antimony	Sb	chromium	Cr	lithium	Li	selenium	Se
argon	Ar	cobalt	Co	magnesium	Mg	silicon	Si
arsenic	As	copper	Cu	manganese	Mn	silver	Ag
barium	Ba	fluorine	F	mercury	Hg	sodium	Na
beryllium	Be	germanium	Ge	neon	Ne	strontium	Sr
bismuth	Bi	gold	Au	nickel	Ni	sulfur	S
boron	B	helium	He	nitrogen	N	tellurium	Te
bromine	Br	hydrogen	H	oxygen	O	tin	Sn
cadmium	Cd	iodine	I	phosphorus	P	titanium	Ti
calcium	Ca	iron	Fe	platinum	Pt	xenon	Xe
carbon	C	krypton	Kr	potassium	K	zinc	Zn

3.4 Metals, Nonmetals, and Semimetals

LEARNING OBJECTIVES

- List the properties of metals and nonmetals.
- Predict whether an element is a metal, nonmetal, or semimetal given its position in the periodic table.
- Predict whether an element is a solid, liquid, or gas at 25 °C and normal atmospheric pressure.

An element is a pure substance that cannot be broken down further and still maintain its unique properties. A **metal** is an element that typically is a solid, has a bright metallic luster, a high density, a high melting point, and is a good conductor of heat and electricity. A metal can usually be hammered into a thin sheet of foil and is said to be **malleable**. If it can be drawn into a fine wire, it is said to be **ductile**. Aluminum, copper, and silver are familiar examples of metals.

A **nonmetal** is an element that usually has a low density, a low melting point, and is a poor conductor of heat and electricity. Many nonmetals occur naturally in the solid state and have a dull appearance, for example, phosphorus and sulfur. These solid nonmetals are neither malleable nor ductile and crush to a powder if hammered. Eleven nonmetals occur naturally in the gaseous state; hydrogen and oxygen are familiar examples of colorless, gaseous nonmetals. Table 3.4 summarizes the properties of metals and nonmetals.

A **semimetal** (or ***metalloid***) is an element that typically has properties midway between those of metals and nonmetals. Silicon is a familiar semimetal used in the semiconductor industry for making transistors and integrated circuits. Example Exercise 3.3 further illustrates the identification of properties of metals and nonmetals.

TABLE 3.4 General Characteristics of Metals and Nonmetals*

Property	Metals	Nonmetals
physical state	solid	solid, gas
appearance	metallic luster	dull
pliability	malleable, ductile	brittle
conductivity	heat, electricity	nonconductor
density	usually high	usually low
melting point	usually high	usually low
chemical reactivity	react with nonmetals	react with metals and nonmetals

*There are numerous exceptions to these general characteristics. For example, mercury and bromine are liquids, magnesium metal has a low density, and gallium metal has a melting point below 30 °C.

EXAMPLE EXERCISE 3.3 Properties of Metals

Which of the following properties is *not* characteristic of a metal?

(a) good conductor of heat	(b) malleable
(c) high melting point	(d) reacts with other metals

Solution

Refer to Table 3.4 to classify each of the following properties:

(a) Metals are good conductors of heat.
(b) Metals are malleable.
(c) Metals usually have high melting points.
(d) Metals do *not* react with other metals; they mix to form alloys.

Practice Exercise

Which of the following properties is *not* characteristic of a nonmetal?

(a) insulator of electricity	(b) ductile
(c) low density	(d) reacts with nonmetals

Answer:

(b) Nonmetals crush to a powder and are not malleable or ductile.

Concept Exercise

Which of the following is a solid metal under normal conditions: calcium, phosphorus, mercury, or silicon?

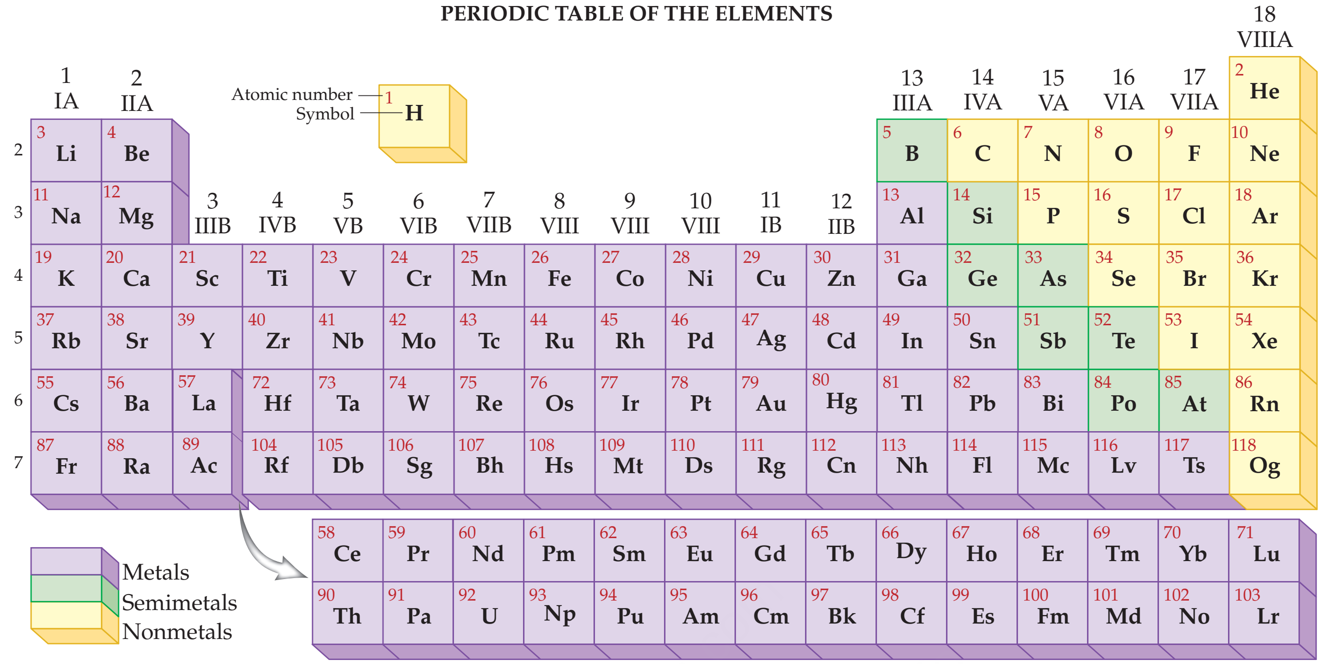
Answer: See Appendix G, 3.3.

Periodic Table of the Elements

In Chapter 5, we will devote our entire discussion to the chemical elements. For now, realize that each element is assigned a number. The number that identifies a particular element is called the **atomic number**. For example, the atomic number of hydrogen is 1, helium is 2, lithium is 3, and so on, and the atomic number of uranium is 92.

All the elements have been arranged by atomic number and placed in a special chart. This chart is called the *periodic table of the elements*, or simply the **periodic table**. Metals are placed on the left side of the table, and nonmetals are placed on the right side. Hydrogen is an exception. Although it is a nonmetal, hydrogen has unusual properties and has been placed by itself at the top center of the periodic tables in this text.

Metals and nonmetals are separated by semimetals that include boron, silicon, germanium, arsenic, antimony, and tellurium. The unstable elements polonium and astatine are also considered semimetals. Figure 3.5 shows the overall arrangement of metals, nonmetals, and semimetals in the periodic table of the elements.



	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
																		2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

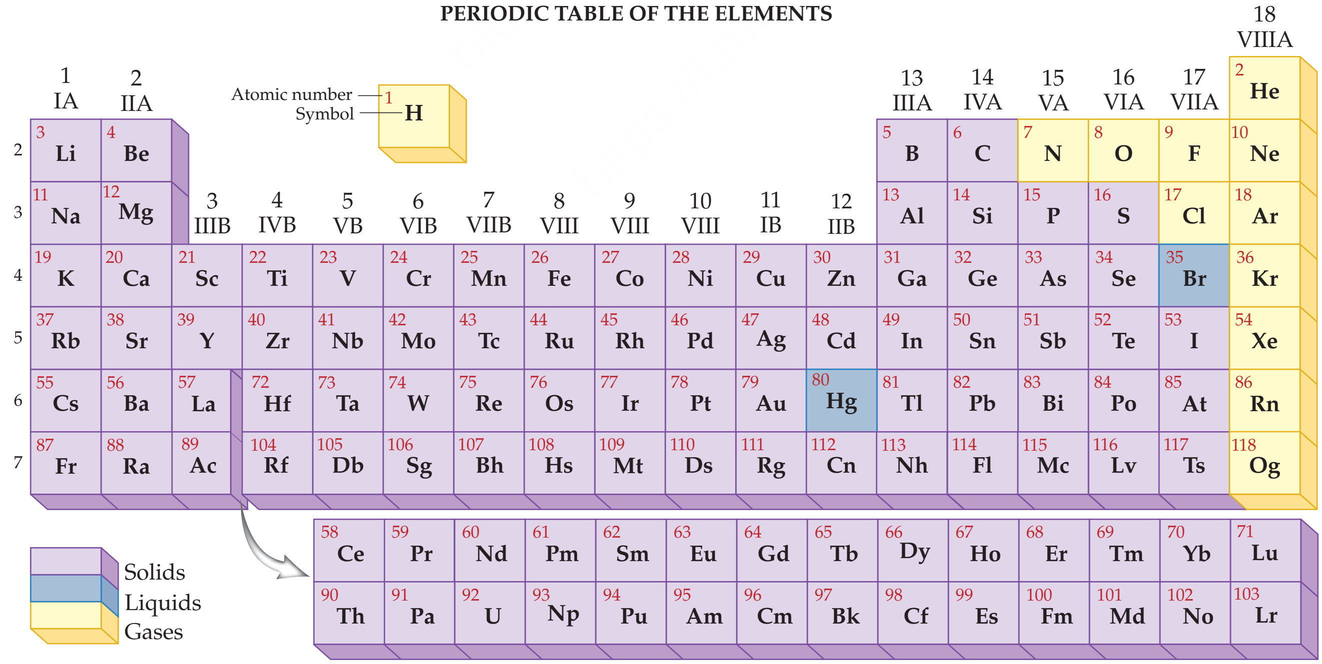
58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

▲ **Figure 3.5** **Metals, Nonmetals, and Semimetals** The symbols of elements having metallic properties are on the left side of the periodic table, nonmetallic are on the right side, and semimetallic are in between. Notice the special placement of hydrogen, a nonmetallic element.

Physical States of the Elements

The periodic table can help you master a large amount of information. With little effort, you can correctly predict the physical state of most elements. Excluding mercury, all the metals are in the solid state at normal conditions of 25 °C and normal atmospheric pressure. All semimetals are also in the solid state at normal temperature and pressure.

Nonmetals, on the other hand, show great diversity in physical state. At 25 °C and normal pressure, 5 nonmetals are solids, 1 is a liquid, and 11 are gases. The solids are carbon, phosphorus, sulfur, selenium, and iodine. The only nonmetal liquid is



	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
																		2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

▲ **Figure 3.6** **Physical States of the Elements** At 25 °C and normal atmospheric pressure, all metals are in the solid state except Hg. Most nonmetals are gases except C, P, S, Se, and I, which are solids. The only elements in the liquid state at normal conditions are Hg and Br.

reddish-brown bromine. The 11 gases are hydrogen, nitrogen, oxygen, fluorine, chlorine, helium, neon, argon, krypton, xenon, and radon. All these gases are colorless except fluorine and chlorine, which are greenish-yellow. Figure 3.6 illustrates the normal physical state of each element.

CHEMISTRY CONNECTION Elements 104 and Beyond

Q: Element 112 was recently officially named for which sixteenth-century astronomer?

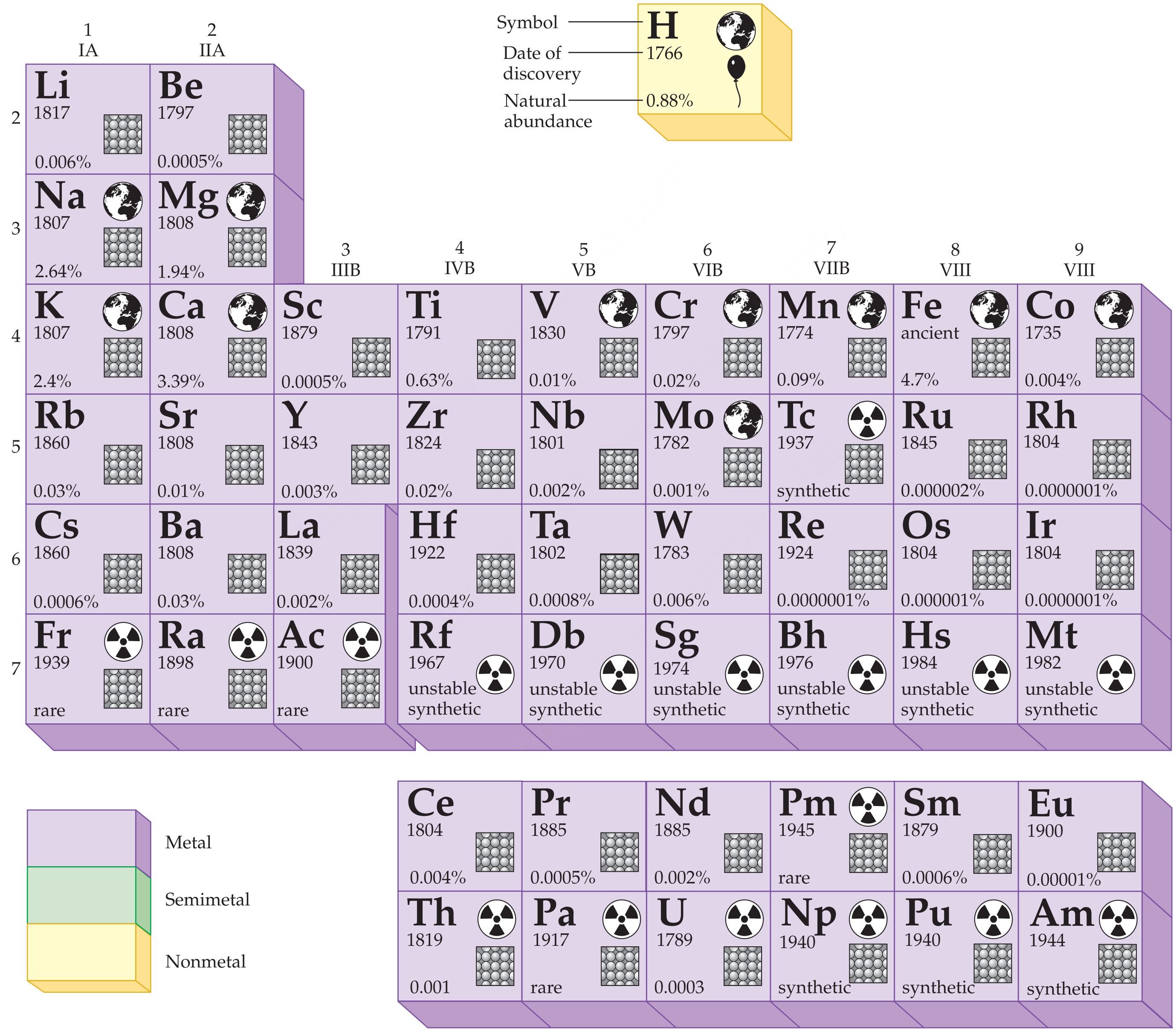
In 1964, a team of Russian scientists reported the first synthesis of element 104. The Russians proposed the name "kurchatovium" for Igor Kurchatov, a Soviet physicist. American scientists could not confirm the results, however, and denied the Russian claim of discovery. In 1969, a heavy-element team at the University of California, Berkeley, synthesized element 104 and proposed the name "rutherfordium" after the English physicist Ernest Rutherford. A few years later, a similar disagreement arose over element 105. Russian and American scientists each claimed they had discovered element 105, and each team proposed different names and symbols.

In 1985, the International Union of Pure and Applied Chemistry (IUPAC) attempted to resolve the controversy by recommending systematic names based on the atomic numbers of the elements. According to IUPAC, the names for elements 104 and beyond were to be formed from Latin prefixes plus the suffix *–ium*. The name for element 104, for example, was to be unnilquadium. Interestingly, there was a strong objection to the IUPAC recommendation because it did not honor scientists who had made significant contributions to our understanding of heavy elements.

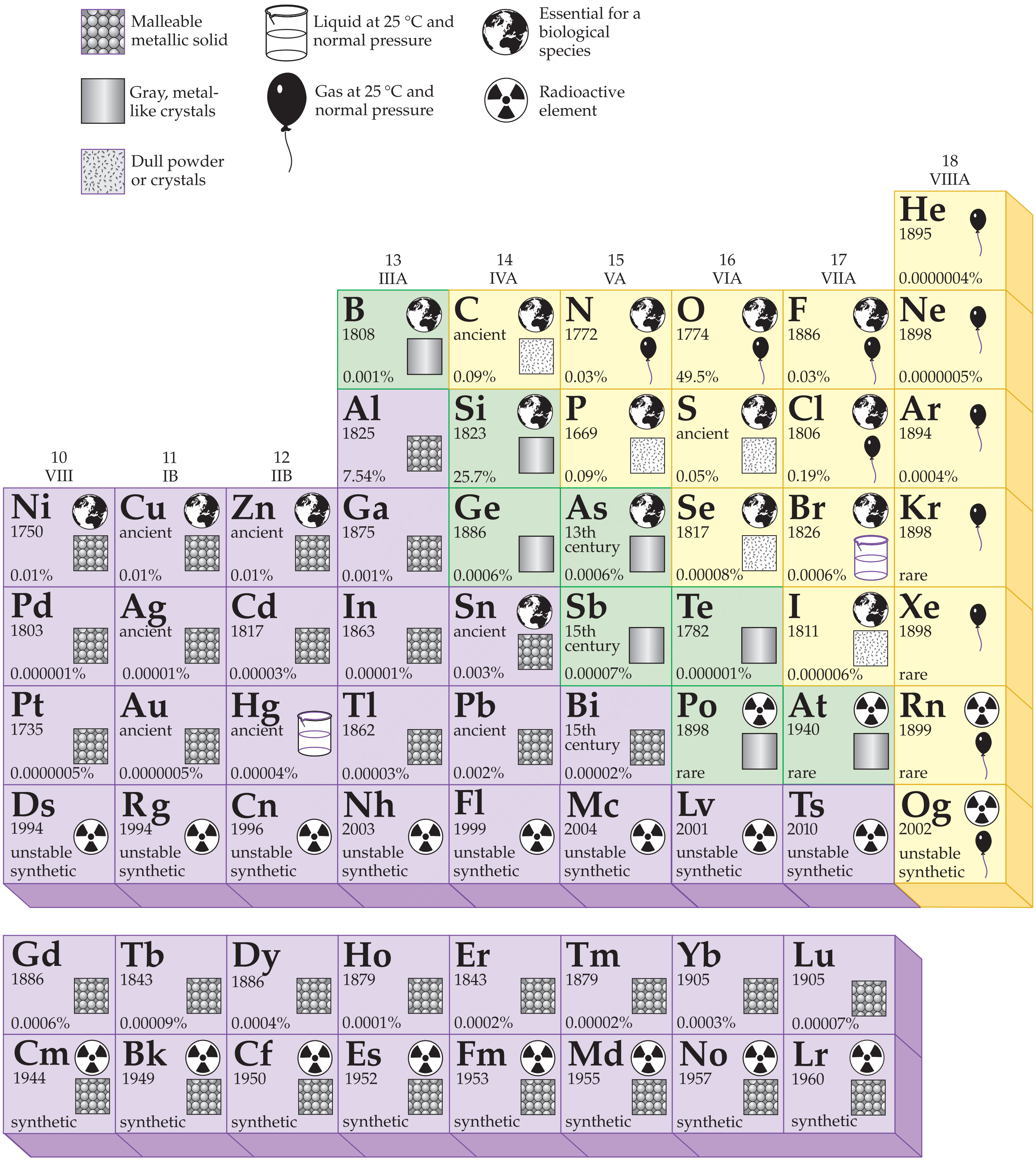
In 1997, IUPAC issued a new set of official names. Element 104 is named *rutherfordium* for Ernest Rutherford. Element 105 is named *dubnium* for the Russian nuclear research facility at Dubna. Element 106 is named *seaborgium* for the American chemist, Glenn Seaborg. Element 107 is named *bohrium* in honor of the Danish physicist Niels Bohr. Element 108 is named *hassium* for the province of Hesse in Germany. Element 109 is named *meitnerium* in honor of the Austrian physicist Lise Meitner. Element 110 is named *darmstadtium* for the German nuclear research facility at Darmstadt. Element 111 is named *roentgenium* for William Roentgen, the German physicist who discovered X-rays. Element 112 is named *copernicium* for Nicolas Copernicus, the sixteenth-century astronomer who first proposed a heliocentric model with the Sun (not Earth) at the center of our solar system. Recently, element 114 was officially named *flerovium* for Georgy Flyorov, a Russian physicist; element 116 was named *livermorium* for the Livermore Radiation Laboratory in California. Elements 113, 115, 117, and 118 have been given proposed names, which are awaiting IUPAC approval.

ELEMENT	DISCOVERY	PROPOSED NAME	IUPAC NAME
104	Russia, 1964 United States, 1969	kurchatovium, Ku rutherfordium, Rf	rutherfordium, Rf
105	Russia, 1967 United States, 1970	nielsbohrium, Ns hahnium, Ha	dubnium, Db
106	United States, 1974	seaborgium, Sg	seaborgium, Sg
107	Germany, 1974	bohrium, Bh	bohrium, Bh
108	Germany, 1984	hassium, Hs	hassium, Hs
109	Germany, 1982	meitnerium, Mt	meitnerium, Mt
110	Germany, 1994	darmstadtium, Ds	darmstadtium, Ds
111	Germany, 1994	roentgenium, Rg	roentgenium, Rg
112	Germany, 1996	copernicium, Cn	copernicium, Cn
113	Russia, 2003	nihonium, Nh	
114	Russia, 1999	flerovium, Fl	flerovium, Fl
115	Russia, 2004	moscovium, Mc	
116	Russia, 2001	livermorium, Lv	livermorium, Lv
117	Russia, 2010	tennessine, Ts	
118	Russia, 2002	oganesson, Og	

A: Element 112 was officially named *copernicium* for Nicolas Copernicus, the sixteenth-century astronomer who first proposed that the Sun, not Earth, was at the center of our solar system.



▲ **Figure 3.7** **Pictorial Periodic Table of the Elements** The natural abundance is the percent by mass of an element in Earth's crust, oceans, and atmosphere. The natural abundance of an element listed as rare is less than 1 mg per metric ton (1000 kg). An element listed as synthetic is made artificially and does not occur naturally. An element listed as unstable often disintegrates in a fraction of a second.



▦	Malleable metallic solid
▮	Gray, metal-like crystals
░	Dull powder or crystals
🥛	Liquid at 25 °C and normal pressure
🎈	Gas at 25 °C and normal pressure
🌍	Essential for a biological species
☢	Radioactive element

10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
								He 🎈 1895 0.0000004%
			B 🌍 ▮ 1808 0.001%	**C** 🌍 ░ ancient 0.09%	**N** 🌍 🎈 1772 0.03%	**O** 🌍 🎈 1774 49.5%	**F** 🌍 🎈 1886 0.03%	**Ne** 🎈 1898 0.0000005%
			Al ▦ 1825 7.54%	**Si** 🌍 ▮ 1823 25.7%	**P** 🌍 ░ 1669 0.09%	**S** 🌍 ░ ancient 0.05%	**Cl** 🌍 🎈 1806 0.19%	**Ar** 🎈 1894 0.0004%
Ni 🌍 ▦ 1750 0.01%	**Cu** 🌍 ▦ ancient 0.01%	**Zn** 🌍 ▦ ancient 0.01%	**Ga** ▦ 1875 0.001%	**Ge** ▮ 1886 0.0006%	**As** 🌍 ▮ 13th century 0.0006%	**Se** 🌍 ░ 1817 0.00008%	**Br** 🌍 🥛 1826 0.0006%	**Kr** 🎈 1898 rare
Pd ▦ 1803 0.000001%	**Ag** ▦ ancient 0.00001%	**Cd** ▦ 1817 0.00003%	**In** ▦ 1863 0.00001%	**Sn** 🌍 ▦ ancient 0.003%	**Sb** ▮ 15th century 0.00007%	**Te** ▮ 1782 0.000001%	**I** 🌍 ░ 1811 0.000006%	**Xe** 🎈 1898 rare
Pt ▦ 1735 0.0000005%	**Au** ▦ ancient 0.0000005%	**Hg** 🥛 ancient 0.00004%	**Tl** ▦ 1862 0.00003%	**Pb** ▦ ancient 0.002%	**Bi** ▦ 15th century 0.00002%	**Po** ☢ ▮ 1898 rare	**At** ☢ ▮ 1940 rare	**Rn** ☢ 🎈 1899 rare
Ds ☢ 1994 unstable synthetic	**Rg** ☢ 1994 unstable synthetic	**Cn** ☢ 1996 unstable synthetic	**Nh** ☢ 2003 unstable synthetic	**Fl** ☢ 1999 unstable synthetic	**Mc** ☢ 2004 unstable synthetic	**Lv** ☢ 2001 unstable synthetic	**Ts** ☢ 2010 unstable synthetic	**Og** ☢ 🎈 2002 unstable synthetic

Gd ▦ 1886 0.0006%	**Tb** ▦ 1843 0.00009%	**Dy** ▦ 1886 0.0004%	**Ho** ▦ 1879 0.0001%	**Er** ▦ 1843 0.0002%	**Tm** ▦ 1879 0.00002%	**Yb** ▦ 1905 0.0003%	**Lu** ▦ 1905 0.00007%
Cm ☢ ▦ 1944 synthetic	**Bk** ☢ ▦ 1949 synthetic	**Cf** ☢ ▦ 1950 synthetic	**Es** ☢ ▦ 1952 synthetic	**Fm** ☢ ▦ 1953 synthetic	**Md** ☢ ▦ 1955 synthetic	**No** ☢ ▦ 1957 synthetic	**Lr** ☢ ▦ 1960 synthetic

EXAMPLE EXERCISE 3.4 Physical States of the Elements

Indicate the physical state for each of the following elements at 25 °C and normal pressure. Classify each element as a metal, nonmetal, or semimetal:

(a) barium (b) boron
(c) bismuth (d) bromine

Solution

Referring to Figures 3.5 and 3.6, we observe the following:

(a) Barium (Ba) is on the left side of the periodic table; it is a *solid metal* under normal conditions.
(b) Boron (B) is in the middle of the periodic table; it is a *solid semimetal*.
(c) Bismuth (Bi) is to the right, but below the semimetals in the periodic table, it is a *solid metal*.
(d) Bromine (Br) is on the right side of the periodic table; it is a *liquid nonmetal* at normal conditions.

Practice Exercise

Indicate the physical state for each of the following elements at 25 °C and normal pressure. Classify each element as a metal, nonmetal, or semimetal:

(a) aluminum (b) hydrogen
(c) helium (d) radium

Answers:

(a) solid metal (b) gaseous nonmetal
(c) gaseous nonmetal (d) solid metal

Concept Exercise

Which of the following is a liquid nonmetal under normal conditions: sodium, mercury, bromine, or sulfur?

Answer: See Appendix G, 3.4.

LEARNING OBJECTIVES

- Calculate the number of atoms of each element in a compound given the chemical formula.
- Apply the law of definite composition to a compound.

3.5 Compounds and Chemical Formulas

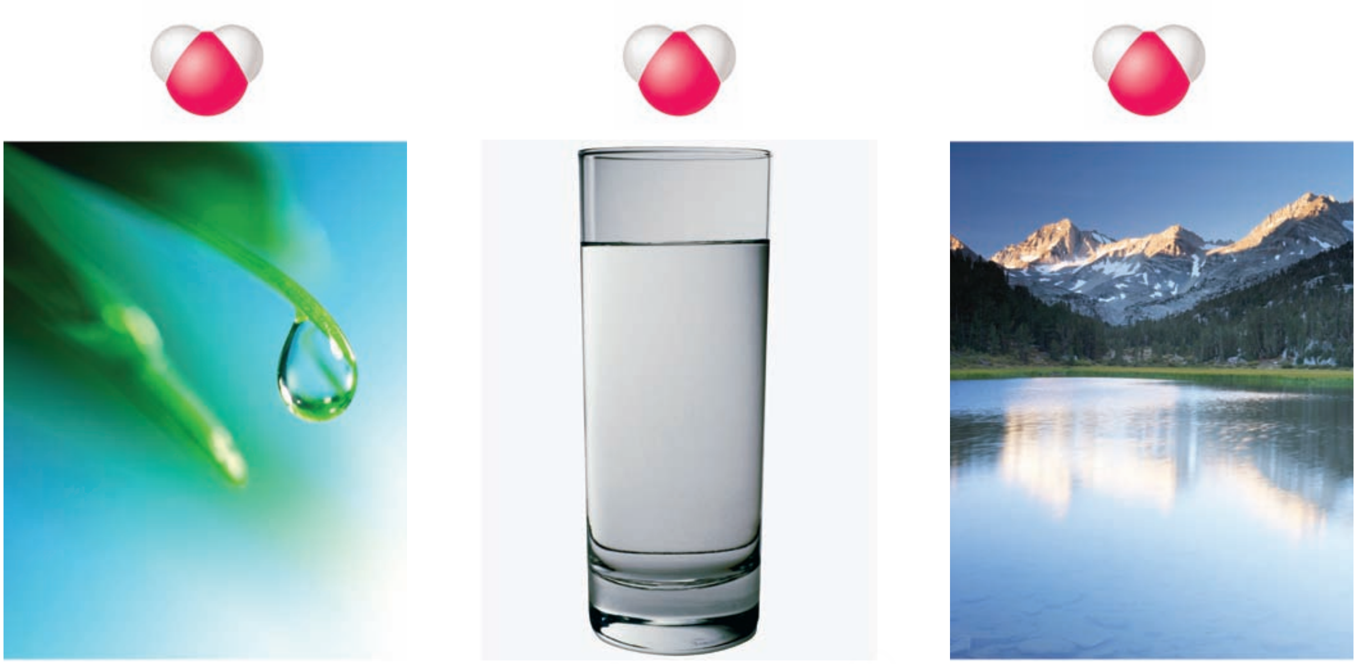
In the late 1700s, the French chemist Joseph Louis Proust (1754–1826) painstakingly analyzed the compound copper carbonate. No matter how he prepared the compound, he found that the elements copper, carbon, and oxygen were always present in the same proportion by mass. Proust studied many other compounds and obtained similar constant proportions for the elements. In 1799, Proust stated that "compounds always contain the same elements in a constant proportion by mass." This statement is now called the **law of definite composition** or the law of constant proportion.

Ordinary table salt is the compound sodium chloride, NaCl. According to the law of definite composition, salt contains sodium and chlorine in a constant proportion by mass. For NaCl, the ratio is 39.3% sodium and 60.7% chlorine. Moreover, the proportion of sodium and chlorine is the same whether we have a tiny crystal of salt, a block of salt, or a mountain of salt.

Similarly, the compound water contains 11.2% hydrogen and 88.8% oxygen by mass. Regardless of the amount, water always contains hydrogen and oxygen in a constant proportion by mass. Figure 3.8 illustrates the law of definite composition with water as an example.

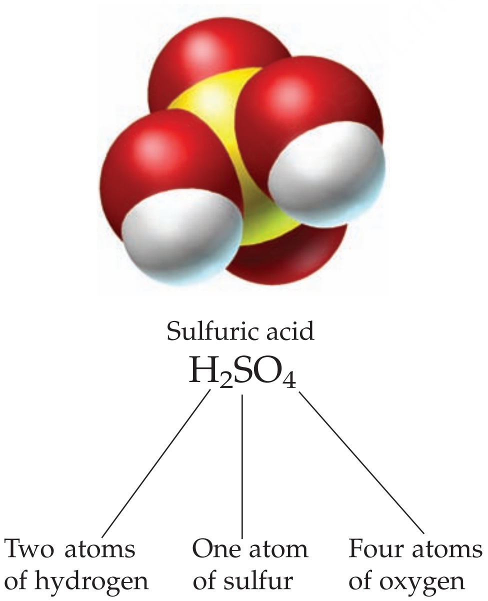
Chemical Formulas

Most elements occur naturally as a collection of individual atoms. A few nonmetal elements, such as hydrogen and oxygen, occur naturally as two atoms of the element in a single unit. A unit of matter composed of two or more nonmetal atoms is called a **molecule**. Hydrogen, H_2, and oxygen, O_2, occur naturally as molecules containing two atoms.



▲ **Figure 3.8 Law of Definite Composition** A drop of water, a glass of water, and a lake of water all contain hydrogen and oxygen in the same percent by mass, that is, 11.2% hydrogen and 88.8% oxygen.

We have learned to use chemical symbols for the names of elements. In a similar fashion, we use chemical formulas for the names of compounds. A **chemical formula** expresses the number of atoms of each element in a compound. The number of atoms is indicated with a subscript, unless the number is 1, in which case the subscript is omitted. A molecule of water contains two atoms of hydrogen and one atom of oxygen. Therefore, its chemical formula is H_2O. A molecule of ammonia has one atom of nitrogen and three atoms of hydrogen; its chemical formula is NH_3. The chemical formula of sulfuric acid, H_2SO_4, is interpreted in Figure 3.9.



◀ **Figure 3.9 Interpretation of a Chemical Formula** The chemical formula of sulfuric acid, H_2SO_4,indicates 2 hydrogen atoms, 1 sulfur atom, and 4 oxygen atoms. Thus, sulfuric acid has a total of 7 atoms.

EXAMPLE EXERCISE 3.5 Composition of Chemical Formulas

State the total number of atoms in a molecule of vitamin B_3, $C_6H_6N_2O$.

Solution

The chemical formula for vitamin B_3 indicates 6 carbon atoms, 6 hydrogen atoms, 2 nitrogen atoms, and 1 oxygen atom. Thus, $C_6H_6N_2O$ has a total of 15 atoms.

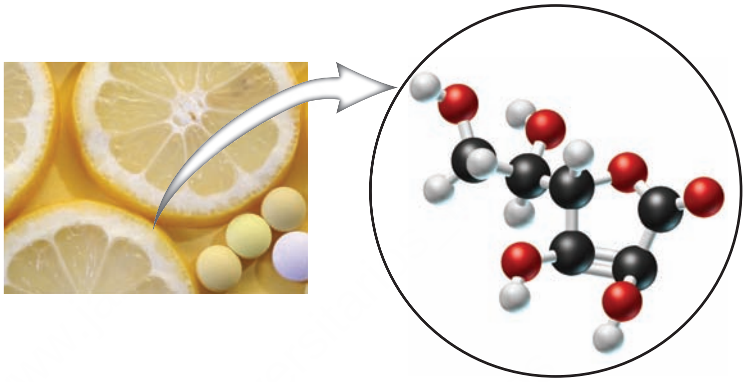
Practice Exercise

Write the chemical formula for vitamin B_6, which is composed of 8 carbon atoms, 11 hydrogen atoms, 1 nitrogen atom, and 3 oxygen atoms.

Answer: $C_8H_{11}NO_3$ (total of 23 atoms)

Concept Exercise

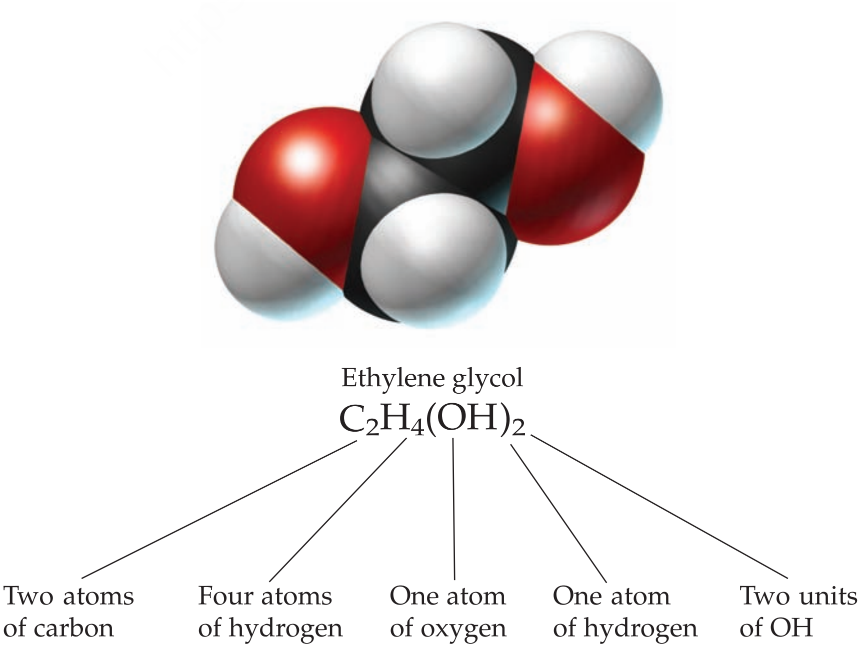
Examine the model for vitamin C and determine the molecular formula. In the model shown, the black sphere = C atom, white = H atom, and red = O atom.



Answer: See Appendix G, 3.5.

Some chemical formulas use parentheses to clarify the chemical composition. For example, antifreeze is ethylene glycol and has the chemical formula $C_2H_4(OH)_2$. A molecule of ethylene glycol is composed of 2 atoms of carbon, 4 atoms of hydrogen, and 2 units of OH. The parentheses around (OH) in the chemical formula indicate that OH is a two-atom unit. The chemical formula of ethylene glycol, $C_2H_4(OH)_2$, is interpreted in Figure 3.10.

▶ **Figure 3.10 Interpretation of a Chemical Formula** The chemical formula of ethylene glycol, $C_2H_4(OH)_2$, indicates 2 carbon atoms, 4 hydrogen atoms, and 2 OH units. The oxygen atom and hydrogen atom are a single unit that appear twice in a single molecule. Thus, ethylene glycol has a total of 10 atoms.



EXAMPLE EXERCISE 3.6 Composition of Chemical Formulas

State the total number of atoms in a molecule of glycerin, $C_3H_5(OH)_3$.

Solution

The chemical formula for glycerin indicates 3 carbon atoms, 5 hydrogen atoms, and 3 OH units. Thus, $C_3H_5(OH)_3$ has a total of 14 atoms.

Practice Exercise

Write the chemical formula for nitroglycerin if a molecule is composed of 3 carbon atoms, 5 hydrogen atoms, 3 oxygen atoms, and 3 NO_2 units.

Answer: $C_3H_5O_3(NO_2)_3$

Concept Exercise

State the total number of atoms in a molecule of nitroglycerin, $C_3H_5O_3(NO_2)_3$.

Answer: See Appendix G, 3.6.

Note In Chapter 6 we will learn how to systematically name compounds. It is helpful to begin referring to compounds by pronouncing their chemical formulas. Even professional chemists pronounce chemical formulas of compounds; for instance, water is H_2O and is referred to as "H two O." Sulfuric acid is H_2SO_4 and is referred to as "H two, S, O four." We can pronounce the chemical formula for milk of magnesia, $Mg(OH)_2$, as "Mg, OH taken twice."

3.6 Physical and Chemical Properties

LEARNING OBJECTIVE

- Classify a property of a substance as physical or chemical.

Previously, we classified matter as either a mixture or a pure substance. Recall that the physical and chemical properties of a pure substance are consistent throughout a sample. Moreover, no two substances have the same physical and chemical properties.

Furthermore, we said that a pure substance is either a compound or an element. However, the physical and chemical properties of a compound are never the same as the properties of its constituent elements. For example, sodium is a soft, silver metal, and chlorine is a poisonous, yellow gas. In a reaction, these two elements combine to form white crystals of table salt, NaCl (Figure 3.11).

A **physical property** refers to those characteristics of a pure substance that we can observe without changing the composition of the substance. The list of physical properties is extensive, but those considered important are appearance, melting and boiling points, density, heat and electrical conductivity, solubility, and physical state under normal conditions. Table 3.5 lists some of the physical properties of sodium, chlorine, and sodium chloride at 25 °C and normal atmospheric pressure.

A **chemical property** of a pure substance describes its chemical reactions with other substances. The chemical properties of sodium, for example, include its reaction with oxygen to form sodium oxide and its reaction with water to produce hydrogen gas. There are many reactions for the elements, more than we can possibly consider. We can, however, use the periodic table of elements to predict chemical behavior.

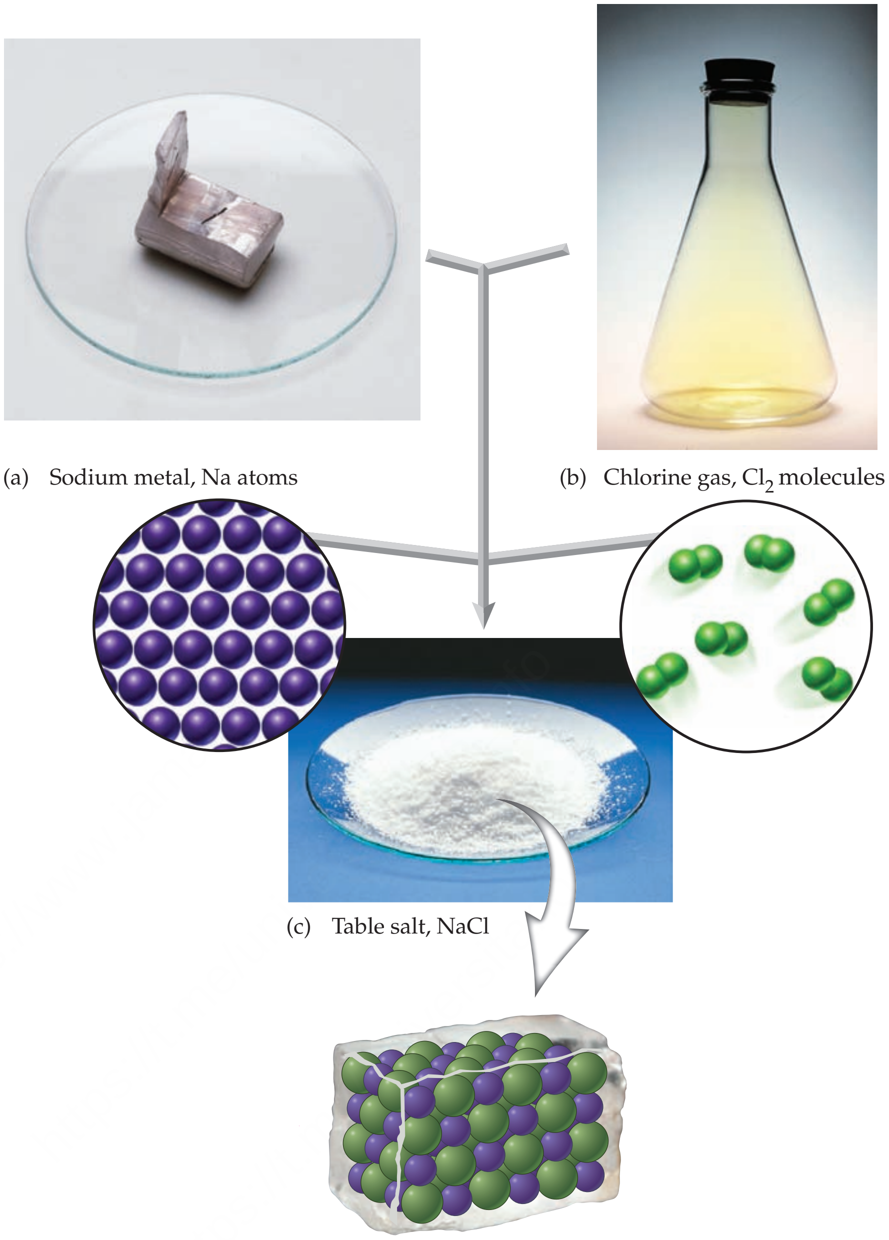
The periodic table is arranged according to families of elements that in general give similar reactions. For instance, the metals in Group IA/1 on the far left side of the periodic table react with oxygen gas to give products with similar chemical formulas, namely, Li_2O, Na_2O, K_2O, Rb_2O, and Cs_2O. Notice that in each case, the number of metal to oxygen atoms is in the ratio of 2 to 1.

In Chapter 5 we will study families of elements. For now, review Table 3.6, which illustrates that elements in the same family have similar chemical properties and give products with similar chemical formulas.

Classifying Physical and Chemical Properties

We can distinguish between a physical and a chemical property as follows. Physical properties are observed without altering the composition of the substance, whereas

▶ **Figure 3.11 Sodium, Chlorine, and Sodium Chloride** (a) Sodium is a shiny metal, (b) chlorine is a yellow gas, and (c) table salt is white crystals. (d) A salt crystal, NaCl is composed of sodium (purple) and chloride (green).



(a) Sodium metal, Na atoms

(b) Chlorine gas, Cl_2 molecules

(c) Table salt, NaCl

(d) Salt crystal, NaCl

TABLE 3.5 Physical Properties of Sodium, Chlorine, and Sodium Chloride at Normal Conditions (25 °C and 1.00 ATM)

Property	Sodium	Chlorine	Sodium Chloride
appearance	silver metal	yellowish gas	white crystals
melting point	98 °C	−101 °C	801 °C
boiling point	883 °C	−35 °C	1413 °C
density	0.97 g/cm^3	2.90 g/L	2.17 g/cm^3
solubility in 100 g water	reacts with water	0.51 g at 30 °C	35.7 g at 0 °C

chemical properties always involve a chemical change. When a chemical change occurs, the composition of a substance changes and another substance is formed. Example Exercise 3.7 illustrates the distinction between physical and chemical properties.

TABLE 3.6 Chemical Properties for Families of Elements

Family and Element	Reaction with Oxygen	Reaction with Water	Reaction with HCL Acid
Group IA/1			
lithium	Li_2O	$H_2 + LiOH$	$H_2 + LiCl$
sodium	Na_2O	$H_2 + NaOH$	$H_2 + NaCl$
potassium	K_2O	$H_2 + KOH$	$H_2 + KCl$
Group IIA/2			
calcium	CaO	$H_2 + Ca(OH)_2$	$H_2 + CaCl_2$
strontium	SrO	$H_2 + Sr(OH)_2$	$H_2 + SrCl_2$
barium	BaO	$H_2 + Ba(OH)_2$	$H_2 + BaCl_2$
Group IVA/14			
carbon	CO_2	no reaction	no reaction
silicon	SiO_2	no reaction	no reaction
germanium	GeO_2	no reaction	no reaction
Group VIIIA/18			
helium	no reaction	no reaction	no reaction
neon	no reaction	no reaction	no reaction
argon	no reaction	no reaction	no reaction

EXAMPLE EXERCISE 3.7 Physical and Chemical Properties

Classify each of the following as a physical property or a chemical property.

(a) Water appears colorless and odorless at 20 °C.
(b) Water dissolves sugar crystals.
(c) Water produces a gas with calcium metal.
(d) Water exists as ice at −10 °C.

Solution

If a reaction occurs, there is a change in composition and the property is chemical. Otherwise, the property is physical.

(a) Color and odor are *physical properties*.
(b) Solubility is a *physical property*.
(c) A reaction producing a gas is a *chemical property*.
(d) Physical state is a *physical property*.

Practice Exercise

Classify each of the following as a physical property or a chemical property.

(a) Water appears hard and crystalline at 0 °C.
(b) Water is insoluble in gasoline.
(c) Water is a very weak conductor of electricity.
(d) Water produces a gas with sodium metal.

Answers:

(a) physical property (b) physical property
(c) physical property (d) chemical property

Concept Exercise

Copper melts at 1083 °C, has a density of 8.92 g/mL, is a good conductor of electricity, and turns black when heated. Which of these is an example of a chemical property?

Answer: See Appendix G, 3.7.

3.7 Physical and Chemical Changes

LEARNING OBJECTIVE

- Classify a change in a substance as physical or chemical.

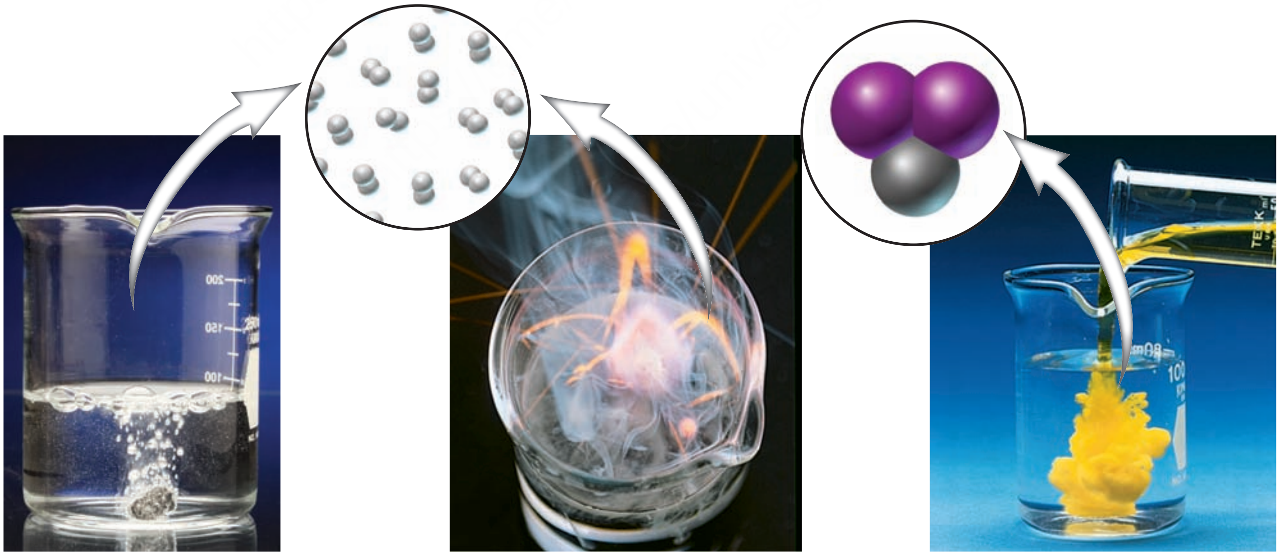
When a pure substance is altered, the alteration is classified as a physical or a chemical change. In a **physical change**, the chemical composition of the sample does not change. If we melt ice to water, heat alcohol to a vapor, or recycle aluminum cans into aluminum foil, we are performing a physical change. The chemical composition of the ice, alcohol, and aluminum is the same before and after the change. In other words, the formula of each substance (H_2O, C_2H_5OH, and Al) remains constant. Thus, altering the shape or physical state of a pure substance is an indication of a physical change.

In a **chemical change**, there is a chemical reaction. The composition of the sample changes, and we observe a new set of properties. If a banana ripens from green to yellow, an antacid tablet in water evolves gas bubbles, or fireworks display a shower of colorful lights, we are observing a chemical change. The composition of the banana, antacid tablet, and fireworks are different before and after the change.

Classifying Physical and Chemical Changes

We can distinguish between a physical and a chemical change as follows: a physical change can be observed without altering the composition of the substance, whereas a chemical change involves the formation of a new substance. When a new substance is formed, we usually observe one of the following: a permanent color change, an odor or bubbles from the release of a gas, or light or heat from the release of energy.

In the laboratory, we may observe that two solutions added together produce an insoluble substance. The formation of a solid substance in a solution is a practical observation that indicates a chemical change. Figure 3.12 illustrates three examples of chemical changes.



(a) (b) (c)

▲ **Figure 3.12** **Evidence of Chemical Change** The evidence for a chemical change includes (a) gas bubbles from the reaction of calcium metal in water, (b) heat and light energy from the reaction of potassium metal in water, and (c) a colorful, insoluble substance from the reaction of two solutions.

EXAMPLE EXERCISE 3.8 Physical and Chemical Changes

Classify each of the following as a physical change or a chemical change.

(a) Touching a lit candle to hydrogen soap bubbles gives an explosion.
(b) Heating water in a flask produces moisture on the glass.
(c) Combining two colorless solutions gives a yellow solid.
(d) Pouring vinegar on baking soda produces gas bubbles.

Solution

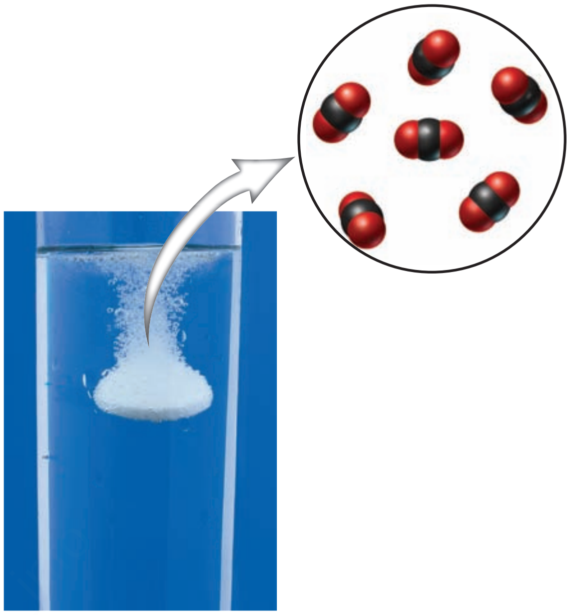
An observation that indicates a physical change is a change of physical state. The observations that suggest a chemical change include burning, fizzing, changing color, or forming an insoluble substance in solution.

(a) Hydrogen explodes; thus, it is a *chemical change*.
(b) Water is boiled; thus, it is a *physical change*.
(c) Two solutions give an insoluble substance; thus, it is a *chemical change*.
(d) Baking soda fizzes; thus, it is a *chemical change*.

Practice Exercise

Classify each of the following as a physical change or a chemical change.

(a) Freezing water in a refrigerator makes cubes of ice.
(b) Adding silver nitrate to tap water gives a cloudy solution.
(b) Burning sulfur gives a light blue flame.
(d) Grinding aspirin tablets produces a powder.



▲ **Alka-Seltzer** An Alka-Seltzer tablet in water releases carbon dioxide gas bubbles.

Answers:

(a) physical change
(b) chemical change
(c) chemical change
(d) physical change

Concept Exercise

An Alka-Seltzer tablet dissolves in water and produces gas bubbles. Is this an example of a physical change or a chemical change?

Answers: See Appendix G, 3.8.

3.8 Conservation of Mass

LEARNING OBJECTIVE

- Apply the conservation of mass law to chemical changes.

In the late 1700s, the French chemist Antoine Lavoisier (1743–1794) performed exacting laboratory experiments and found that the mass of reactants before a chemical change was always equal to the mass of products after the change. He concluded that matter was neither created nor destroyed during a chemical reaction. In 1789, he announced the conservation of mass principle. This principle has become known as the **law of conservation of mass**.

As an example of conservation of mass, consider the reaction of hydrogen and oxygen. Hydrogen and oxygen always combine in the same ratio by mass to form water. That is, if 1 g of hydrogen combines with 8 g of oxygen, then 2 g of hydrogen combines with 16 g of oxygen, 3 g of hydrogen reacts with 24 g of oxygen, and so on. We can predict the mass of water produced by each of these reactions. From the conservation of mass law, we know that the mass of hydrogen plus the mass of oxygen must equal the mass of water produced. That is,

$$1.0 \text{ g hydrogen} + 8.0 \text{ g oxygen} = 9.0 \text{ g water}$$

Conversely, we can predict the masses of hydrogen and oxygen produced by the decomposition of water. Passing an electric current through water produces hydrogen gas and oxygen gas. This process is called *electrolysis*. If the electrolysis of 45.0 g of water produces 5.0 g of hydrogen, how many grams of oxygen are evolved?

The conservation of mass law states that the mass of the hydrogen and oxygen is equal to the mass of the water, 45.0 g. Therefore, the mass of oxygen must equal the mass of the water minus the mass of the hydrogen. That is,

$$45.0 \text{ g water} - 5.0 \text{ g hydrogen} = 40.0 \text{ g oxygen}$$

The conservation of mass law is one of the most important principles in chemistry. Historically, it was used to determine the formulas of compounds and to develop the first periodic table of elements. As we will discuss in Chapter 9, it is the cornerstone of all quantitative relationships in chemistry. Example Exercise 3.9 further illustrates the conservation of mass law.

EXAMPLE EXERCISE 3.9 Conservation of Mass Law

In an experiment, 2.430 g of magnesium metal was ignited and burned with oxygen in the air. If 4.030 g of white magnesium oxide powder, MgO, was collected, what was the mass of oxygen gas that reacted?

Solution

Applying the conservation of mass law, we find that the mass of the magnesium metal plus the mass of the oxygen gas equals the mass of the magnesium oxide powder. That is,

$$\begin{aligned} 2.430 \text{ g Mg} + \text{mass of oxygen} &= 4.030 \text{ g MgO} \\ \text{mass of oxygen} &= 4.030 \text{ g MgO} - 2.430 \text{ g Mg} \\ \text{mass of oxygen} &= 1.600 \text{ g} \end{aligned}$$

Practice Exercise

If 0.654 g of zinc metal reacts with 0.321 g of yellow powdered sulfur, what is the mass of the zinc sulfide produced?

Answer: 0.975 g

Concept Exercise

Heating 1.000 g copper metal with yellow sulfur produces 1.252 g of black copper sulfide. What is the mass of sulfur that reacted with the copper metal?

Answer: See Appendix G, 3.9.

3.9 Potential and Kinetic Energy

LEARNING OBJECTIVES

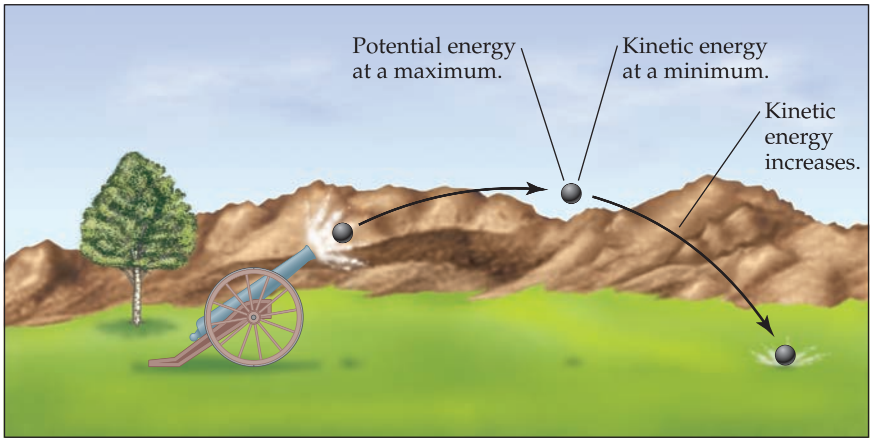
- Distinguish between potential and kinetic energy.
- Summarize the relationship of temperature and kinetic energy.
- Summarize the relationship of kinetic energy and physical state.

Energy makes things happen and is often defined as the ability to do work. There are two types of energy: potential and kinetic. **Potential energy (PE)** is stored energy that matter possesses as a result of its position or composition. **Kinetic energy (KE)** is the energy matter has as a result of its motion. A boulder perched at the top of a mountain has potential energy. If the boulder rolls down the mountain, it loses potential energy, but gains kinetic energy. In the process of rolling down the mountain, the potential energy of the boulder is converted to kinetic energy. Figure 3.13 illustrates the conversion from potential to kinetic energy.

In our study of chemistry, we are usually interested in the potential energy associated with the chemical composition of a pure substance. For example, gasoline has potential energy that is converted to kinetic energy when it undergoes combustion. As gasoline burns, the potential energy of the cool liquid is converted to the kinetic energy of the hot, expanding gas mixture.

Kinetic Energy, Temperature, and Physical State

We can view a sample of matter as a large collection of very small particles. The attraction between particles is one of the criteria that influences the physical state. In solids, this force of attraction holds particles in fixed positions and allows little or no movement.



◀ **Figure 3.13 Potential Energy and Kinetic Energy** A cannonball shot into the air loses kinetic energy while gaining potential energy. At the top of its flight, the potential energy is at a maximum and the kinetic energy is at a minimum. As the cannonball falls to Earth, the potential energy decreases as the kinetic energy increases.

When a substance in the *solid state* is heated, the temperature increases and particles begin to randomly vibrate in fixed positions. As a solid acquires more heat energy, particles vibrate so intensely that the force of attraction is overcome and the particles break free from each other. This is a description of the *liquid state*.

A liquid is, therefore, a state of matter in which particles possess sufficient energy to overcome their mutual attraction and have limited movement. When a liquid has enough energy to completely overcome the force of attraction, the particles are free to fly about and strike the walls of the container. At this point, the liquid has changed to the *gaseous state.* Table 3.7 summarizes the relationship between the properties of particles and the physical state.

TABLE 3.7 Properties and Physical States

Property of Particles	Solid	Liquid	Gas
kinetic energy	very low	high	very high
movement	insignificant	restricted	unrestricted

We can review the relationship between kinetic energy and physical state as follows. If we heat a solid, it acquires enough energy to overcome the attraction between particles and melts to a liquid. As the temperature increases, particles gradually gain more energy and demonstrate increasing motion. If we continue to heat the liquid, it acquires enough energy to vaporize to a gas. Figure 3.14 illustrates the relationship between increasing energy and physical state.

The kinetic energy of a gas is directly related to its temperature. If the temperature increases, the kinetic energy increases; if the temperature decreases, the kinetic energy decreases. Moreover, if the temperature increases, the motion of the particles increases; if the temperature decreases, the motion decreases.

Example Exercise 3.10 further illustrates how temperature, kinetic energy, and the motion of gas particles are related.

EXAMPLE EXERCISE 3.10 Kinetic Energy and Molecular Motion

A balloon filled with helium gas is cooled from 25 °C to −25 °C. State the change in each of the following:

(a) kinetic energy of the gas (b) motion of helium atoms

Solution

Temperature, kinetic energy, and velocity are related as follows:

(a) As the temperature cools from 25 °C to −25 °C, the kinetic energy of helium atoms *decreases.*

(b) Because a drop in temperature produces a decrease in kinetic energy, the motion of helium atoms *decreases.*

Practice Exercise

A steel cylinder containing air is heated from 25 °C to 50 °C. State the change in each of the following:

(a) kinetic energy of the gas
(b) motion of air molecules

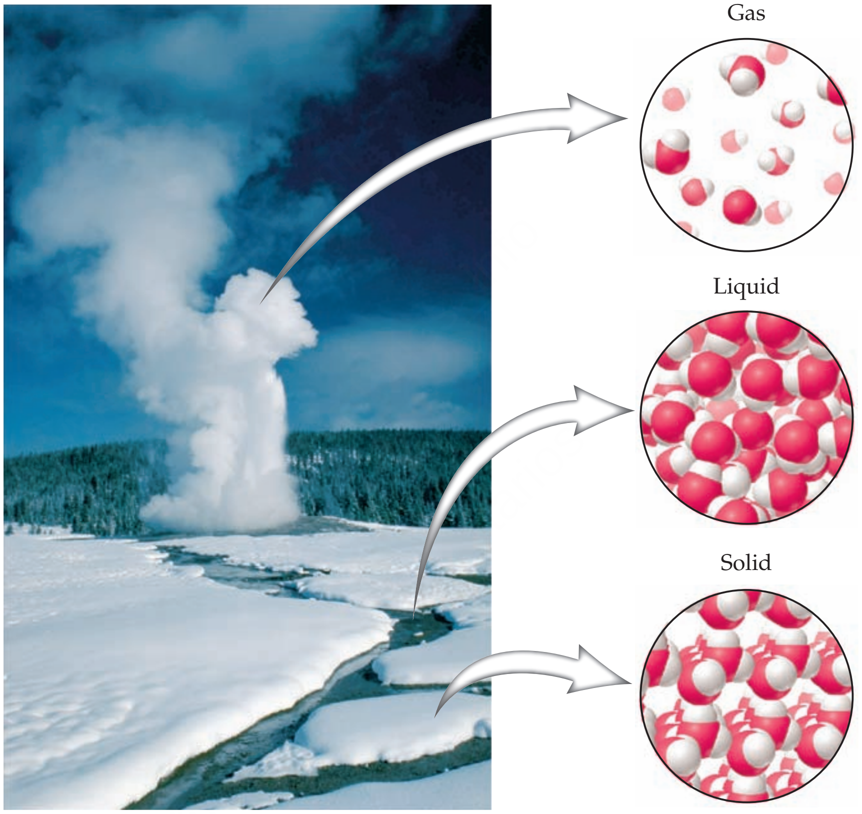
Answers:

(a) increases (b) increases

Concept Exercise

What happens to the kinetic energy and velocity of air molecules when the temperature increases?

Answer: See Appendix G, 3.10.



▶ Figure 3.14 **Kinetic Energy and Physical State** The energy of individual H_2O molecules increases as solid ice is heated. As the attraction between molecules is overcome, solid ice changes to liquid water and then to gaseous steam.

LEARNING OBJECTIVES

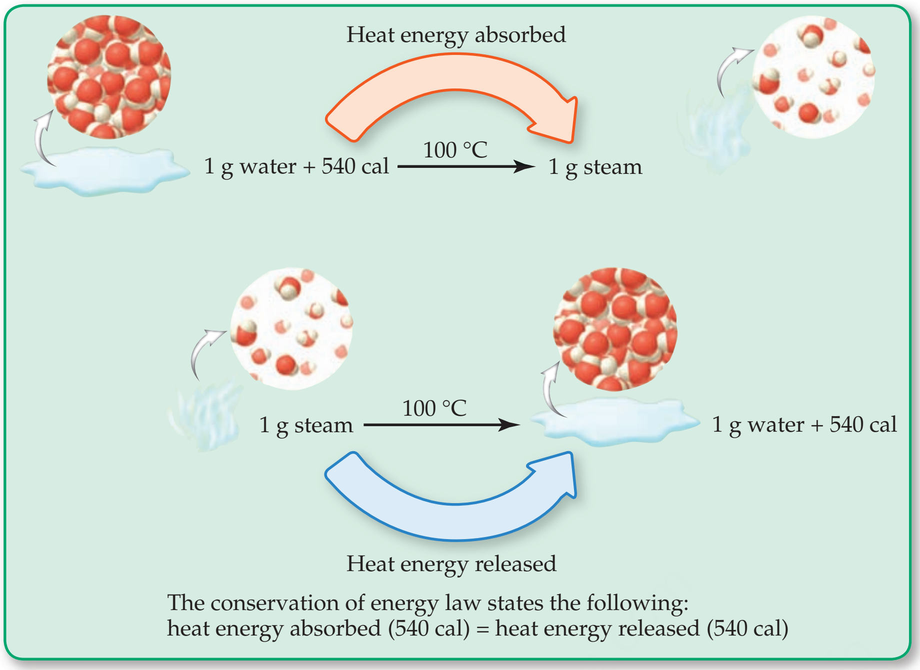
- Apply the conservation of energy law to physical and chemical changes.
- Identify the following forms of energy: chemical, electrical, mechanical, nuclear, heat, and light.

3.10 Conservation of Energy

In the middle of the nineteenth century, several English and German scientists proposed that the total energy in the universe is constant. This proposal was worded in different ways, but it essentially stated that energy cannot be created nor destroyed. Energy can, however, be converted from one form to another. This principle is known as the **law of conservation of energy**. This principle applies experimentally, and we can make energy predictions for physical and chemical changes.

Energy and Physical Changes

Let's consider the conservation of energy law applied to a change of physical state. When water is converted to steam, a given amount of heat energy is required. Energy cannot be destroyed, so the same amount of heat must be released if the steam condenses to a liquid. In practice, it requires 540 cal of heat energy to convert 1 g of water at 100 °C to steam at 100 °C. It therefore follows that 540 cal of heat are released when 1 g of steam at 100 °C condenses to water. Figure 3.15 illustrates the conservation of energy for a physical change.



◀ **Figure 3.15 Conservation of Energy for a Physical Change** The physical change of state from water to steam, and the reverse process, involve the same amount of heat energy.

Energy and Chemical Changes

The conservation of energy law applied to a chemical change is analogous to the conservation of mass law. That is, the total energy before a reaction is the same as the total energy after the reaction. It is important to understand, however, that some chemical changes release energy, whereas others absorb energy. Yet, the total energy before and after a reaction must always be constant.

How can the energy be constant for a reaction that involves a change in heat? The explanation is that a reaction involves the two types of energy discussed previously, *potential* chemical energy and *kinetic* heat energy. In a reaction, the initial reactants may lose potential energy, but the resulting products gain kinetic energy. In reactions that produce heat, the initial reactants have more potential energy than the products. Conversely, in reactions that absorb heat, the initial reactants have less potential energy than the products.

If we consider the conservation of energy law, we can say that the energy released from a given change is equal to the energy required for the reverse change. For example, we know from experiments that 3.20 kcal of heat is released when hydrogen gas and

EXAMPLE EXERCISE 3.11 Energy and Physical Changes

If 5.50 kcal of heat are required to melt an ice cube, how much energy is released when freezing the water to ice, at 0 °C?

Solution

From the law of conservation of energy, 5.50 kcal of heat energy is released when the water freezes to ice, at 0 °C.

Practice Exercise

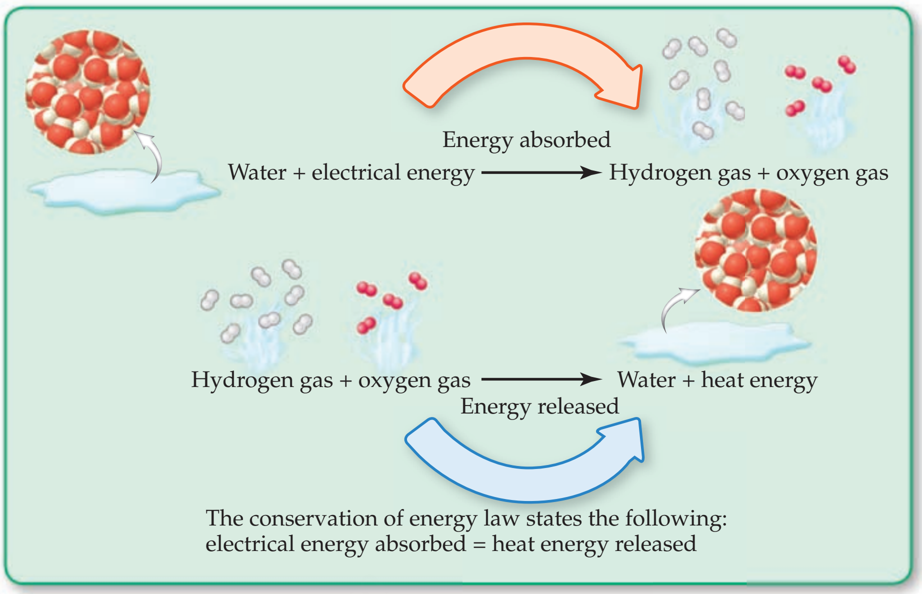
If 1250 kcal of heat are required to vaporize a sample of water, how much energy is released when the steam condenses to water, at 100 °C?

Answer: 1250 kcal

Concept Exercise

Which of the following involves more heat energy: (a) heating 10.0 g of alcohol from 20 °C to 50 °C; or, (b) cooling 10.0 g of alcohol from 50 °C to 20 °C?

Answer: See Appendix G, 3.11.



▶ Figure 3.16 **Conservation of Energy for a Chemical Change** The chemical change from water to hydrogen and oxygen gases, and the reverse process, involve the same amount of energy.

oxygen gas react to produce 1 g of water. Thus, 3.20 kcal of heat is required for the reverse process of decomposing 1 g of water into hydrogen gas and oxygen gas. If we decompose water into the two gases using an electric current, the electrical energy required is equivalent to 3.20 kcal. Figure 3.16 illustrates the conservation of energy for a chemical change.

Forms of Energy

There are six forms of energy: heat, light, chemical, electrical, mechanical, and nuclear. Each of these can be converted from one form of energy to another. For example, we can use solar energy panels to convert light energy from the Sun to heat energy, which we can store as hot water.

We can also convert mechanical energy to electrical energy, which is a more efficient process. For example, the mechanical energy of water spilling over a dam is used to spin a turbine that in turn drives an electrical generator. This hydroelectric process operates at nearly 90% efficiency. Example Exercise 3.12 illustrates the conversion of various forms of energy.

EXAMPLE EXERCISE 3.12 Forms of Energy

Identify two forms of energy that are involved in each of the following conversions:

(a) Radioactive emissions vaporize water to steam.
(b) Steam drives a turbine.
(c) A turbine spins and drives an electrical generator.

Solution

We can refer to the six forms of energy listed above. It follows that

(a) *Nuclear* energy is converted to *heat* energy.
(b) *Heat* energy is converted to *mechanical* energy.
(c) *Mechanical* energy is converted to *electrical* energy.

Practice Exercise

Identify two forms of energy that are involved in each of the following devices:

(a) flashlight
(b) solar calculator
(c) lead–acid battery

Answers:

(a) chemical and light (b) light and electrical
(c) chemical and electrical

Concept Exercise

Which of the following is *not* a basic form of energy: chemical, electrical, heat, light, mechanical, nuclear, solar?

Answers: See Appendix G, 3.12.

Conservation of Mass and Energy

In 1905, Albert Einstein (1879–1955) presented a paper on the special theory of relativity that proposed that matter and energy are interrelated, stating this relationship in the form $E = mc^2$. In this famous equation, Einstein showed that energy (E) and mass (m) are related by the velocity of light (c) squared.

Because mass and energy are related, it naturally follows that the conservation of mass and the conservation of energy are related. Thus, a more accurate statement combines these two principles into the **law of conservation of mass and energy**. The combined law states that the total mass and energy in the universe is constant. When applied to a chemical reaction, the law indicates that the total mass and energy before and after a chemical change is constant.

By applying the combined law, we realize that any decrease in mass corresponds to an increase in energy. Even a slight decrease in mass, however, corresponds to an enormous amount of energy. For instance, a mass loss of 0.001 g produces enough heat energy to raise the temperature of a large swimming pool about 30 °C. Because the mass loss for ordinary chemical reactions is undetectable from a practical laboratory point of view, the original conservation of mass law is still valid.

▲ **Albert Einstein** The relationship between mass and energy is expressed by Einstein's famous equation: $E = mc^2$.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
3.1 Physical States of Matter Matter exists in three **physical states**: solid, liquid, and gas. A *solid* is characterized by a fixed shape and fixed volume. A *liquid* has a variable shape, but a fixed volume. A *gas* has a variable shape and a variable volume. As temperature increases, a solid *melts* to a liquid and then *vaporizes* to a gas. A direct change of state from a solid to a gas is called **sublimation**. Conversely, as temperature decreases, a gas *condenses* to a liquid and then *freezes* to a solid. A direct change of state from a gas to a solid is called **deposition**.	• Describe the motion of particles in the solid, liquid, and gaseous states of matter. ***Related Exercises: 1–6*** • Describe the effect of temperature on the solid, liquid, and gaseous states of matter ***Related Exercises: 7–12***
3.2 Elements, Compounds, and Mixtures We can classify matter according to its properties. A **heterogeneous mixture** has an indefinite composition and variable properties. A **homogeneous mixture** has a definite composition, but variable properties from sample to sample; examples include **alloys**, solutions, and mixtures of gases. A **substance** can be either an **element** or a **compound**. A substance has a definite composition and constant properties.	• Classify a sample of matter as an element, compound, or mixture. ***Related Exercises: 13–16***
3.3 Names and Symbols of the Elements In 1803, the English chemist John Dalton proposed that elements are composed of indivisible spherical particles called **atoms**. In 1813, the Swedish chemist J. J. Berzelius suggested that the name of an element be abbreviated with a **chemical symbol** corresponding to the first letter of the name of the element, for instance, N for nitrogen. If two or more names began with the same letter, he recommended using two letters in the name, for example, Ni for nickel.	• Memorize the names and symbols of 48 common elements. ***Related Exercises: 17–22***

Key Concepts	Learning Objectives and Related Exercises
3.4 Metals, Nonmetals, and Semimetals An element can be classified as a **metal** or a **nonmetal**. If an element can be hammered into a thin sheet, it is said to be **malleable**. If it can be drawn into a fine wire, it is said to be **ductile**. An element having properties between those of a metal and a nonmetal is termed a **semimetal**, or ***metalloid***. All elements can be ordered by **atomic number** and arranged in a **periodic table**. The periodic table shows a general trend in properties, with metals on the left side and nonmetals on the right side.	• List the properties of metals and nonmetals. ***Related Exercises: 23–26*** • Predict whether an element is a metal, nonmetal, or semimetal given its position in the periodic table ***Related Exercises: 27–30*** • Predict whether an element is a solid, liquid, or gas at 25 °C and normal atmospheric pressure ***Related Exercises: 31–34***
3.5 Compounds and Chemical Formulas The **law of definite composition** states that a compound always contains the same elements in the same proportion by mass. The composition of a compound is given by its **chemical formula**. Vitamin C has the formula $C_6H_8O_6$; one **molecule** contains 6 carbon atoms, 8 hydrogen atoms, and 6 oxygen atoms. Thus, the total number of atoms in the chemical formula of vitamin C is 20.	• Calculate the number of atoms of each element in a compound given the chemical formula. ***Related Exercises: 35–40*** • Apply the law of definite composition to a compound. ***Related Exercises: 41–42***
3.6 Physical and Chemical Properties The properties of substances are classified as physical or chemical. A **physical property** refers to a characteristic such as color, density, melting point, conductivity of heat and electricity, solubility, and physical state. A **chemical property** describes a chemical reaction of a substance with another substance.	• Classify a property of a substance as physical or chemical. ***Related Exercises: 43–50***
3.7 Physical and Chemical Changes The changes in substances are classified as physical or chemical. A **physical change** does not alter the chemical composition and involves a change in mass, volume, or physical state. A **chemical change** involves a change in chemical composition. Evidence for a chemical change includes burning, producing a gas, forming a solid substance in solution, and releasing heat or light energy.	• Classify a change in a substance as physical or chemical. ***Related Exercises: 51–58***
3.8 Conservation of Mass The **law of conservation of mass** states that matter is neither created nor destroyed. This fundamental principle has practical applications in the chemistry laboratory; that is, the total mass of substances produced from a reaction can be calculated because the mass of the products must equal the mass of the reactants.	• Apply the conservation of mass law to chemical changes. ***Related Exercises: 59–62***
3.9 Potential and Kinetic Energy The **potential energy (PE)** of a substance is related to its chemical composition, whereas the **kinetic energy (KE)** of a substance is related to the motion of its particles. Kinetic energy is directly related to temperature. That is, when we heat a solid substance and increase its temperature, we also increase its kinetic energy. As the temperature increases, the particles gain more energy and begin to overcome their attraction for each other. If we continue heating, we will eventually produce a change of state from a solid to a liquid, and from a liquid to a gas.	• Distinguish between potential and kinetic energy. ***Related Exercises: 63–64*** • Summarize the relationship of temperature and kinetic energy. ***Related Exercises: 65–66*** • Summarize the relationship of kinetic energy and physical state. ***Related Exercises: 67–68***
3.10 Conservation of Energy The six forms of energy are heat, light, chemical, electrical, mechanical, and nuclear. The **law of conservation of energy** states that energy is neither created nor destroyed. Energy may be converted, however, from one form to another. In practice, chemical energy is used to provide heat and electrical energy. The laws of conservation of mass and conservation of energy can be treated as a single principle—the **law of conservation of mass and energy**. For ordinary chemical reactions, energy changes are sufficiently small that we can ignore this principle and apply conservation of mass and conservation of energy individually.	• Apply the conservation of energy law to physical and chemical changes. ***Related Exercises: 69–74*** • Identify the following forms of energy: chemical, electrical, mechanical, nuclear, heat, and light. ***Related Exercises: 75–78***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** the condition of matter existing as a solid, liquid, or gas
______ **2.** a direct change of state from a solid to a gas
______ **3.** a direct change of state from a gas to a solid
______ **4.** matter having an indefinite composition and variable properties
______ **5.** matter having a definite composition, but variable properties from sample to sample; for example, alloys and solutions
______ **6.** matter having a definite composition and constant properties; for example, an element or compound
______ **7.** a homogeneous mixture of two or more metals
______ **8.** a substance that can be broken down into simpler substances
______ **9.** a substance that cannot be broken down by chemical reaction
______ **10.** the smallest particle that represents an element
______ **11.** an abbreviation for the name of a chemical element
______ **12.** an element that is generally shiny in appearance, has a high density, high melting point, and is a good conductor of heat and electricity
______ **13.** the property of a metal that allows it to be machined into a foil
______ **14.** the property of a metal that allows it to be drawn into a wire
______ **15.** an element that is generally dull in appearance, has a low density, low melting point, and is not a good conductor of heat and electricity
______ **16.** an element that is generally metallic in appearance and has properties midway between those of a metal and a nonmetal
______ **17.** a number that identifies a particular element
______ **18.** a chart that arranges elements according to their properties
______ **19.** the principle that states a compound always contains the same elements in the same proportion by mass
______ **20.** a single particle composed of two or more nonmetal atoms
______ **21.** an abbreviation for the name of a chemical compound that indicates the number of atoms of each element
______ **22.** a characteristic of a substance that can be observed without changing the chemical composition
______ **23.** a characteristic of a substance that cannot be observed without changing the chemical formula
______ **24.** a modification of a substance that alters its chemical composition
______ **25.** a modification of a substance that does not alter its chemical composition
______ **26.** the stored energy that matter possesses owing to its position or composition
______ **27.** the energy associated with the mass and velocity of a particle
______ **28.** the law that states mass cannot be created nor destroyed
______ **29.** the law that states energy cannot be created nor destroyed
______ **30.** the law that states the total mass and energy in the universe is constant

(a) alloy (*Sec. 3.2*)
(b) atom (*Sec. 3.3*)
(c) atomic number (*Sec. 3.4*)
(d) chemical change (*Sec. 3.7*)
(e) chemical formula (*Sec. 3.5*)
(f) chemical property (*Sec. 3.6*)
(g) chemical symbol (*Sec. 3.3*)
(h) compound (*Sec. 3.2*)
(i) deposition (*Sec. 3.1*)
(j) ductile (*Sec. 3.4*)
(k) element (*Sec. 3.2*)
(l) heterogeneous mixture (*Sec. 3.2*)
(m) homogeneous mixture (*Sec. 3.2*)
(n) kinetic energy (KE) (*Sec. 3.9*)
(o) law of conservation of energy (*Sec. 3.10*)
(p) law of conservation of mass (*Sec. 3.8*)
(q) law of conservation of mass and energy (*Sec. 3.10*)
(r) law of definite composition (*Sec. 3.5*)
(s) malleable (*Sec. 3.4*)
(t) metal (*Sec. 3.4*)
(u) molecule (*Sec. 3.5*)
(v) nonmetal (*Sec. 3.4*)
(w) periodic table (*Sec. 3.4*)
(x) physical change (*Sec. 3.7*)
(y) physical property (*Sec. 3.6*)
(z) physical state (*Sec. 3.1*)
(aa) potential energy (PE) (*Sec. 3.9*)
(bb) semimetal (*Sec. 3.4*)
(cc) sublimation (*Sec. 3.1*)
(dd) substance (*Sec. 3.2*)

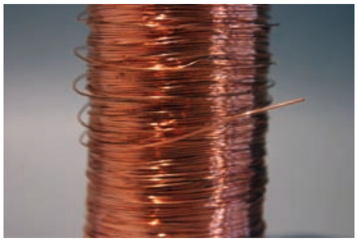
Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Physical States of Matter (Sec. 3.1)

1. Which physical state is described as having a fixed shape and fixed volume?
2. Which physical state is described as having a variable shape and variable volume?
3. Which physical state is described as having a variable shape and compresses significantly?
4. Which physical state is described as having a variable shape and does not compress significantly?
5. Which physical state is described as having particles closely packed and held rigidly in position?
6. Which physical state is described as having particles closely packed and free to move around?
7. Supply the term that describes each of the following changes of physical state.
   (a) gas to liquid **(b)** liquid to solid
8. Supply the term that describes each of the following changes of physical state.
   (a) solid to liquid **(b)** liquid to gas
9. Indicate whether the following changes of physical state require heating or cooling.
   (a) gas to liquid **(b)** liquid to solid
10. Indicate whether the following changes of physical state require heating or cooling.
    (a) solid to liquid **(b)** liquid to gas
11. State whether the process of deposition requires heating or cooling.
12. State whether the process of sublimation requires heating or cooling.

Elements, Compounds, and Mixtures (Sec. 3.2)

13. Classify each of the following as an element, compound, or mixture.
    (a) sodium metal **(b)** chlorine gas
    (c) sodium chloride **(d)** salt water
14. Classify each of the following as an element, compound, or mixture.
    (a) iron metal **(b)** oxygen gas
    (c) iron oxide **(d)** iron ore
15. Classify each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.
    (a) copper wire **(b)** copper ore
    (c) copper oxide **(d)** brass alloy



◀ **Copper Wire** Copper metal is ductile and can be drawn into a thin wire.

16. Classify each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.
    (a) silver bar **(b)** silver ore
    (c) silver oxide **(d)** silver alloy

Names and Symbols of the Elements (Sec. 3.3)

17. Write the chemical symbol for each of the following elements.
    (a) lithium **(b)** argon
    (c) magnesium **(d)** manganese
    (e) beryllium **(f)** silicon
    (g) mercury **(h)** titanium
18. Write the chemical symbol for each of the following elements.
    (a) strontium **(b)** xenon
    (c) oxygen **(d)** zinc
    (e) fluorine **(f)** sodium
    (g) copper **(h)** nickel
19. Write the name of the element for each of the following chemical symbols.
    (a) P **(b)** S
    (c) Kr **(d)** Fe
    (e) Sb **(f)** N
    (g) Pt **(h)** Ag
20. Write the name of the element for each of the following chemical symbols.
    (a) Cl **(b)** Ne
    (c) Cd **(d)** Ge
    (e) Co **(f)** Ra
    (g) Cr **(h)** Te
21. Refer to the periodic table and find the atomic number for each of the following.
    (a) hydrogen **(b)** boron
    (c) aluminum **(d)** titanium
    (e) arsenic **(f)** strontium
    (g) tin **(h)** bismuth
22. Refer to the periodic table and find the atomic number for each of the following.
    (a) helium **(b)** carbon
    (c) potassium **(d)** copper
    (e) selenium **(f)** silver
    (g) lead **(h)** radon

Metals, Nonmetals, and Semimetals (Sec. 3.4)

23. State whether each of the following properties is more typical of a metal or a nonmetal.
    (a) shiny luster
    (b) low melting point
    (c) malleable
    (d) reacts with metals and nonmetals
24. State whether each of the following properties is more typical of a metal or a nonmetal.
    (a) dense solid **(b)** low boiling point
    (c) ductile **(d)** reacts with metals
25. State whether each of the following properties is more typical of a metal or a nonmetal.
    (a) dull powder **(b)** colorless gas
    (c) high density **(d)** reacts with metals
26. State whether each of the following properties is more typical of a metal or a nonmetal.
    (a) shiny solid **(b)** brittle solid
    (c) low density **(d)** forms alloys

27. Refer to the periodic table and classify each of the following elements as a metal, nonmetal, or semimetal.
(a) H **(b)** He
(c) Hf **(d)** Ho

28. Refer to the periodic table and classify each of the following elements as a metal, nonmetal, or semimetal.
(a) C **(b)** Ca
(c) Cl **(d)** Cu

29. Refer to the periodic table and classify each of the following elements as a metal, nonmetal, or semimetal.
(a) boron **(b)** barium
(c) beryllium **(d)** bismuth

30. Refer to the periodic table and classify each of the following elements as a metal, nonmetal, or semimetal.
(a) argon **(b)** antimony
(c) arsenic **(d)** astatine

31. Refer to the periodic table and indicate the physical state for each of the following elements at 25 °C and normal pressure.
(a) N **(b)** Na
(c) Ne **(d)** Ni

32. Refer to the periodic table and indicate the physical state for each of the following elements at 25 °C and normal pressure.
(a) S **(b)** Se
(c) Si **(d)** Sn

33. Refer to the periodic table and indicate the physical state for each of the following elements at 25 °C and normal pressure.
(a) hydrogen **(b)** titanium
(c) fluorine **(d)** bromine

34. Refer to the periodic table and indicate the physical state for each of the following elements at 25 °C and normal pressure.
(a) helium **(b)** manganese
(c) iodine **(d)** mercury

Compounds and Chemical Formulas (Sec. 3.5)

35. State the composition for a molecule of each of the following compounds.
(a) cholesterol, $C_{27}H_{46}O$ **(b)** quinine, $C_{20}H_{24}N_2O_2$

36. State the composition for a molecule of each of the following vitamins.
(a) vitamin B_1, $C_{12}H_{17}Cl_2N_4OS$
(b) vitamin B_2, $C_{17}H_{20}N_4O_6$

37. Write the chemical formula for each of the following compounds.
(a) lysine (an amino acid): 6 carbon, 14 hydrogen, 2 nitrogen, and 2 oxygen atoms
(b) methionine (an amino acid): 5 carbon, 11 hydrogen, 1 nitrogen, 2 oxygen, and 1 sulfur atom

38. Write the chemical formula for each of the following compounds.
(a) vitamin A: 20 carbon, 30 hydrogen, and 1 oxygen atom
(b) vitamin K: 31 carbon, 46 hydrogen, and 2 oxygen atoms

39. What is the total number of atoms in a molecule of each of the following compounds?
(a) citric acid (in citrus fruit), $C_3H_4OH(COOH)_3$
(b) lactic acid (in sore muscles), $C_2H_4(OH)COOH$

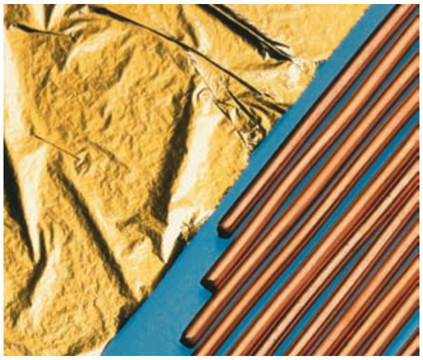
40. What is the total number of atoms in a molecule of each of the following compounds?
(a) stearic acid (in animal fat), $CH_3(CH_2)_{16}COOH$
(b) triolein (in vegetable oil), $C_3H_5(C_{17}H_{33}COO)_3$

41. The mass ratio of H:O in a bottle of water is 1:8. Predict the mass ratio of water in ten bottles of water.

42. The mass ratio of C:H:O in natural vitamin C found in rose hips is 9:1:12. Predict the mass ratio of synthetic vitamin C synthesized in the laboratory.

Physical and Chemical Properties (Sec. 3.6)

43. State whether each of the following is a physical or a chemical property of copper.
(a) shiny metallic luster **(b)** high melting point
(c) conducts heat **(d)** reacts with oxygen



◀ **Gold Foil and Copper Rods** Gold and copper are both malleable and ductile metals; that is, each can be hammered into a foil and drawn into thin rods.

44. State whether each of the following is a physical or a chemical property of gold.
(a) yellow color **(b)** high density
(c) conducts electricity **(d)** reacts with nitric acid

45. State whether each of the following is a physical or a chemical property.
(a) Manganese bubbles in acid.
(b) Silver is malleable.
(c) Calcium turns dark in air.
(d) Carbon grinds to a powder.

46. State whether each of the following is a physical or a chemical property.
(a) Titanium is ductile.
(b) Zinc generates a gas in acid.
(c) Yellow phosphorus ignites in air.
(d) Chromium conducts heat.

47. State whether each of the following is a physical or a chemical property.
(a) Electricity decomposes water.
(b) Ice floats on water.
(c) Potassium metal fizzes in water.
(d) Acetone dissolves in water.

48. State whether each of the following is a physical or a chemical property.
(a) Ether is flammable.
(b) Bromine is a reddish-orange liquid.
(c) Ethanol vaporizes at 78 °C.
(d) Steam condenses to water at 100 °C.

49. State whether each of the following is a physical or a chemical property.
(a) Iron oxidizes to rust.
(b) Dry ice undergoes sublimation.
(c) Lithium metal turns gray in air.
(d) Plutonium radioactively decays.

50. State whether each of the following is a physical or a chemical property.
- **(a)** Sugar ferments to ethanol.
- **(b)** Ethanol dissolves in gasoline.
- **(c)** Ethylene glycol freezes at −12 °C.
- **(d)** Ethanol burns with a light blue flame.

Physical and Chemical Changes (Sec. 3.7)

51. State whether each of the following is a physical or a chemical change.
- **(a)** burning
- **(b)** condensing
- **(c)** freezing
- **(d)** rusting

52. State whether each of the following is a physical or a chemical change.
- **(a)** combustion
- **(b)** vaporization
- **(c)** sublimation
- **(d)** oxidation

53. State whether each of the following is a physical or a chemical change.
- **(a)** cutting aluminum foil
- **(b)** caramelizing sugar
- **(c)** adding air to a tire
- **(d)** igniting a sparkler

54. State whether each of the following is a physical or a chemical change.
- **(a)** slicing an orange into wedges
- **(b)** fermenting apples into apple cider
- **(c)** grinding sugar crystals into powder
- **(d)** digesting carbohydrates for energy

55. State whether each of the following is a physical or a chemical change.
- **(a)** Zinc metal produces a colorless gas in acid.
- **(b)** Nickel alloys with copper in a 5¢ coin.
- **(c)** Liquid helium boils at −269 °C.
- **(d)** Sodium metal reacts with chlorine gas.

56. State whether each of the following is a physical or a chemical change.
- **(a)** Steam condenses and freezes tos ice.
- **(b)** Hydrogen and oxygen gases explode.
- **(c)** Bromine vaporizes to a brown gas.
- **(d)** Sodium metal changes to a white powder.

57. State whether each of the following is a physical or a chemical change.
- **(a)** Mercury liquid cools to a silvery solid.
- **(b)** Natural gas burns with a blue flame.
- **(c)** Ammonia and copper solution turns deep-blue.
- **(d)** Sugar crystals dissolve in water.

58. State whether each of the following is a physical or a chemical change.
- **(a)** Baking soda fizzes in vinegar.
- **(b)** Vinegar and oil separate into two layers.
- **(c)** A flaming wood log releases heat.
- **(d)** A helium balloon decreases in size.

Conservation of Mass (Sec. 3.8)

59. If 1.00 g of gray zinc powder reacts with 0.49 g of yellow sulfur powder, what is the mass of the zinc sulfide product?

60. If 5.00 g of gray iron powder reacts with 2.77 g of red phosphorus powder, what is the mass of the iron phosphide product?

61. If 10.11 g of limestone decomposes by heat to give 8.51 g of solid calcium oxide and carbon dioxide gas, what is the mass of carbon dioxide produced?

62. If 0.750 g of orange mercury oxide decomposes to 0.695 g of liquid mercury and oxygen gas, what is the mass of oxygen produced?

Potential and Kinetic Energy (Sec. 3.9)

63. Which physical state has particles with the lowest kinetic energy?

64. Which physical state has particles with the most movement?

65. If the temperature increases, does the kinetic energy increase or decrease?

66. If the temperature increases, does the motion of particles increase or decrease?

67. If a steel cylinder of helium gas is heated from 50 °C to 100 °C, does the motion of helium atoms increase or decrease?

68. If a steel cylinder of hydrogen gas is cooled from 50 °C to 0 °C, does the motion of hydrogen molecules increase or decrease?

Conservation of Energy (Sec. 3.10)

69. An ice cube at 0 °C melts to water by absorbing heat. If 10.5 kcal of heat are required to melt the ice, how much energy must be lost to freeze the water to ice, at 0 °C?

70. Steam at 100 °C passing through a radiator condenses to water and releases heat. If 1500 kcal of heat are released, how much energy is required to vaporize the water to steam, at 100 °C?

71. A 10.0-g sample of mercury absorbs 110 cal as it is heated from 25 °C to its boiling point at 356 °C. It then requires an additional 697 cal to vaporize. How much energy is released as the mercury vapor cools from 356 °C to 25 °C?

72. A 25.0-g sample of molten iron releases 1230 cal as it cools from 2000 °C to its freezing point, 1535 °C. It then releases an additional 1590 cal as it solidifies. How much energy is required to heat the solid iron from 1535 °C to 2000 °C?

73. The reaction of hydrogen and oxygen gases produces 18.0 g of water and releases 68.2 kcal of heat. How much energy is required to decompose 18.0 g of water into hydrogen and oxygen gases?

74. The decomposition of 25.0 g of water to produce hydrogen and oxygen gases requires 94.8 cal of energy. How much energy is released when hydrogen and oxygen gases react to produce 25.0 g of water?

75. Identify two forms of energy that are involved in each of the following energy conversions relating to a fossil fuel power plant.
- **(a)** Burning coal converts water to steam.
- **(b)** Steam drives a turbine.
- **(c)** A turbine turns a generator.
- **(d)** A generator makes electricity.

76. Identify two forms of energy that are involved in each of the following energy conversions relating to a nuclear power plant.
- **(a)** Uranium converts water to steam.
- **(b)** Steam drives a turbine.
- **(c)** A turbine turns a generator.
- **(d)** A generator makes electricity.

77. Identify two forms of energy that are involved in each of the following energy conversions relating to an automobile.
   (a) hot gases moving pistons
   (b) pistons powering the crankshaft
   (c) crankshaft turning the generator
   (d) generator charging the battery

78. Identify two forms of energy that are involved in each of the following energy conversions relating to an automobile.
   (a) lead–acid battery making electricity
   (b) electricity turning a starter motor
   (c) starter motor turning a flywheel
   (d) gasoline exploding into hot gases

General Exercises

79. A pre-1982 penny was minted from an alloy of copper and zinc. Is a pre-1982 penny an example of a homogeneous or heterogeneous mixture?

80. A post-1982 penny is minted by plating copper onto a zinc disk. Is a post-1982 penny an example of a homogeneous or heterogeneous mixture?

81. Gasoline contains a complex mixture of hydrocarbons. Is gasoline an example of a heterogeneous or a homogeneous mixture?

82. A dry cleaning solvent contains chlorofluorocarbons (CFCs). Is this dry cleaning solvent an example of a homogeneous mixture or a pure substance?

83. The electronics industry manufactures semiconductor chips from silicon. Refer to the periodic table and predict an element that may substitute for silicon.

84. The electronics industry manufactures transistors using arsenic diffusion. Refer to the periodic table and predict an element that may substitute for arsenic.

85. Name the three most abundant elements in Earth's crust.

86. Name the three most abundant elements in the human body.

87. State the atomic number for each of the following elements.
   (a) ferrum **(b)** cuprum
   (c) stannum **(d)** plumbum

88. State the atomic number for each of the following elements.
   (a) roentgenium **(b)** copernicium
   (c) flerovium **(d)** livermorium

89. The reaction of hydrogen and nitrogen gases produces 5.00 g of ammonia and releases 3250 cal of heat. How much energy is required to decompose 5.00 g of ammonia into hydrogen and nitrogen gases?

90. The decomposition of 10.0 g of ammonia requires 27,200 J of energy to give hydrogen and nitrogen gases. How much energy is released when hydrogen and nitrogen gases react to produce 10.0 g of ammonia?

91. Iron and sulfur react to produce iron sulfide and heat energy. An experiment shows that the mass of iron and sulfur is equal to the mass of the iron sulfide. In *theory*, should the products weigh slightly more or slightly less than the reactants?

92. Hydrogen and iodine react to give hydrogen iodide while absorbing heat energy. An experiment shows that the total mass of the reactants is equal to the product. In *theory*, should the product weigh slightly more or slightly less than the reactants?

93. Define each symbol in Einstein's equation: $E = mc^2$.

94. State the law of conservation of mass and energy as it applies to a chemical reaction.

Challenge Exercises

95. State whether the following describes a physical or a chemical change: changing physical state, but not chemical formula.

96. State whether the following describes a physical or a chemical change: changing chemical formula, but not physical state.

Chapter 3 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following describes a substance in the gaseous physical state? (Sec. 3.1)
   (a) The substance has a variable shape.
   (b) The substance has a variable volume.
   (c) The substance compresses significantly.
   (d) all of the above
   (e) none of the above

2. Water contains hydrogen and oxygen. Which of the following describes water? (Sec. 3.2)
   (a) compound **(b)** element
   (c) heterogeneous mixture **(d)** homogeneous mixture
   (e) none of the above

3. Which of the following elements is *not* one of the 10 most abundant in Earth's crust, oceans, and atmosphere? (Sec. 3.3)
   (a) aluminum **(b)** iron
   (c) oxygen **(d)** silicon
   (e) silver

4. Refer to the periodic table and predict which of the following elements is an example of a semimetal. (Sec. 3.4)
   (a) Al **(b)** Ca
   (c) Te **(d)** Ti
   (e) Zn

5. Alum, $Al_2(SO_4)_3$, is used in styptic pencils to stop minor bleeding. What is the number of atoms in the chemical formula of alum? (Sec. 3.5)
   (a) 10 **(b)** 12
   (c) 14 **(d)** 17
   (e) 21

6. Which of the following properties of gasoline is a chemical property? (Sec. 3.6)
   (a) colorless liquid **(b)** burns with air
   (c) density, 0.70 g/mL **(d)** boiling point, 97 °C
   (e) insoluble in water

7. Which of the following is an example of a chemical change? (Sec. 3.7)
   (a) condensation **(b)** crystallization
   (c) oxidation **(d)** sublimation
   (e) none of the above

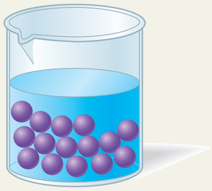
8. If 2.43 g of magnesium burns in oxygen to produce 4.03 g of magnesium oxide, what is the mass of reacting oxygen gas? (Sec. 3.8)
   (a) 1.60 g
   (b) 2.43 g
   (c) 4.03 g
   (d) 6.46 g
   (e) impossible to predict from the given information

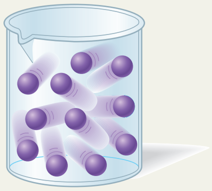
9. What happens to the kinetic energy of gas molecules when a gas is cooled? (Sec. 3.9)
- **(a)** kinetic energy increases
- **(b)** kinetic energy decreases
- **(c)** kinetic energy remains constant
- **(d)** kinetic energy depends on the gas
- **(e)** none of the above

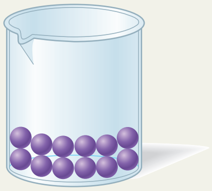
10. A nuclear power plant uses radioactive uranium to convert water to steam. The steam drives a turbine that turns a generator to produce electricity. What two forms of energy are represented by the *uranium* and the *steam*? (Sec. 3.10)
- **(a)** chemical energy and mechanical energy
- **(b)** chemical energy and heat energy
- **(c)** electrical energy and heat energy
- **(d)** nuclear energy and heat energy
- **(e)** nuclear energy and mechanical energy

Key Concepts

11. Identify the physical state (solid, liquid, or gas) that corresponds to each of the following illustrations.

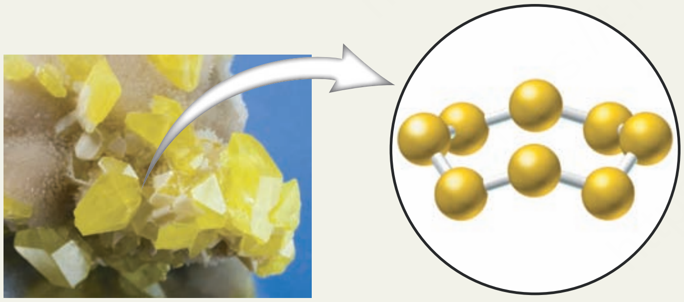

(a)


(b)

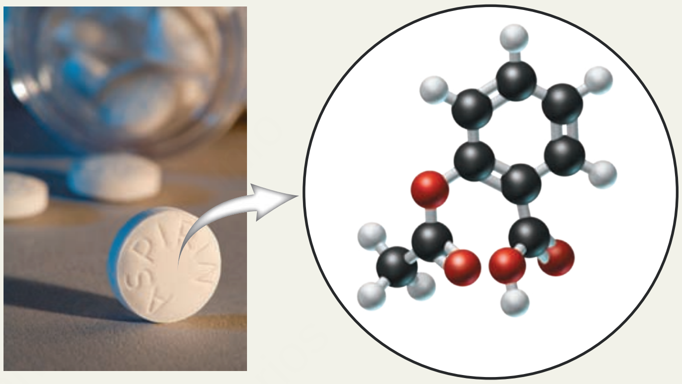

(c)

12. Classify each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.
- **(a)** iron ore
- **(b)** iron oxide
- **(c)** iron metal
- **(d)** iron alloy

13. Classify the mineral (photo) and molecule (model) of sulfur as an element, compound, homogeneous mixture, or heterogeneous mixture.



14. Examine the model for aspirin and determine the molecular formula. In the model shown, the black sphere = C atom, white = H atom, and red = O atom.



Critical Thinking

15. Is ice floating in water an example of a heterogeneous or a homogeneous mixture?

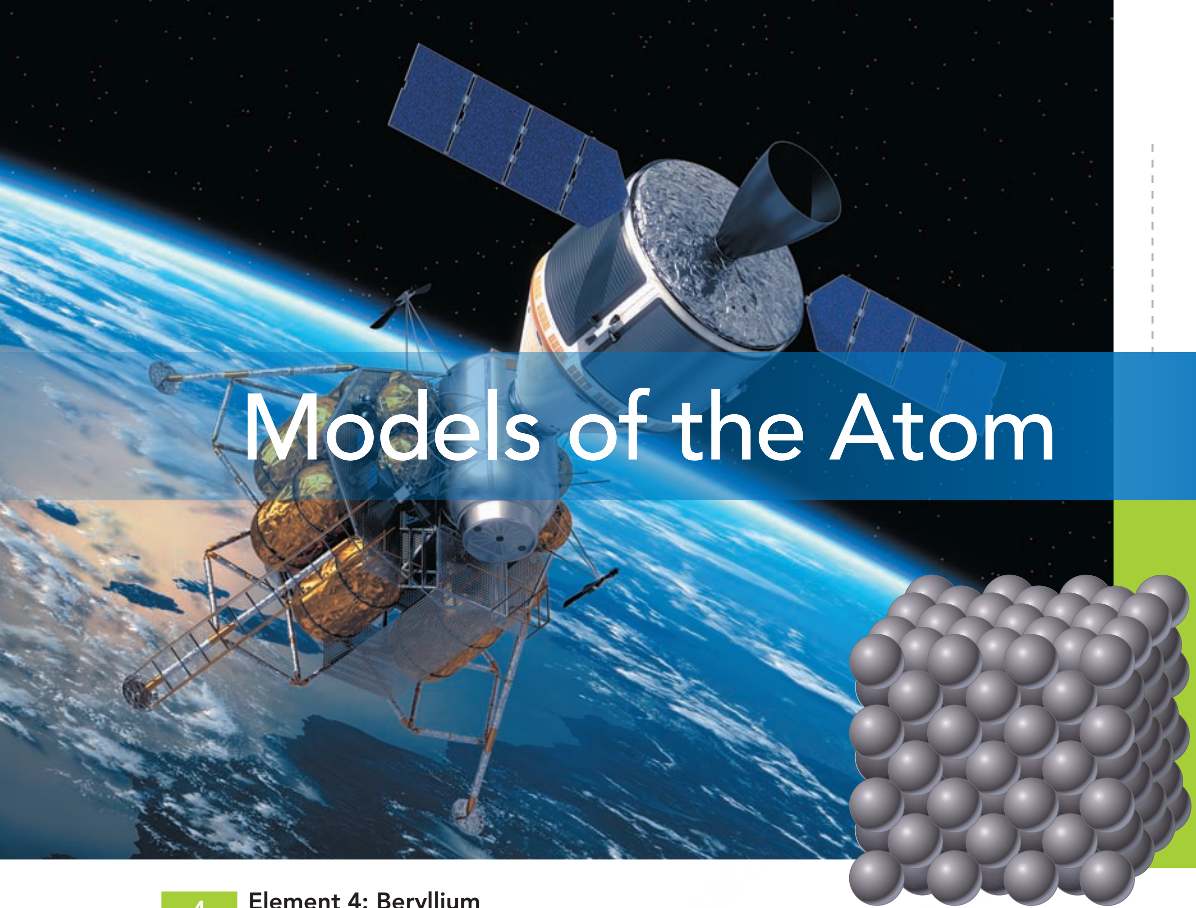
16. Which of the following compounds has physical and chemical properties similar to cobalt metal: CoO, Co_2O_3, or $Co(NO_3)_3$?

17. Using a playground swing as an example, when is the potential energy minimum and the kinetic energy maximum?

18. Using a roller coaster as an example, when is the potential energy maximum and the kinetic energy minimum?

19. Is the law of conservation of mass violated if the Sun converts 7.00 tons of hydrogen into only 6.95 tons of helium?

20. Which conservation law explains the mass difference for the conversion reaction of hydrogen to helium?



CHAPTER 4

Models of the Atom

> *"Anyone who says that they can contemplate quantum mechanics without becoming dizzy has not understood the concept in the least."*
>
> **Niels Bohr, Danish Physicist (1885–1962)**

4 **Be** Beryllium

Element 4: Beryllium
Beryllium is a lightweight metal used in the aerospace industry for the manufacture of aircraft, missiles, and satellites. Beryllium-copper alloys are used to make tools that are uniquely spark-proof. In addition, beryllium metal is a good reflector of neutrons; thus, beryllium finds application in nuclear reactors.

The concept of an atom was born in Greece about 450 B.C. One of the great issues of the day was whether matter and motion were continuous or discontinuous. Democritus, a Greek philosopher, argued that *matter is discontinuous* and could not be infinitely divided. He believed that at some point a fundamental indivisible particle would emerge. He called this ultimate particle an atom, from the Greek word *atomos*, which means "indivisible."

About 350 B.C., Aristotle—the most influential philosopher of his time—argued that both *matter and motion are continuous*. He believed that matter could be divided an infinite number of times, and that even the smallest particle could be further divided. So powerful was Aristotle's influence that the existence of atoms was not believable, and his point of view prevailed for 21 centuries.

4.1 Dalton Model of the Atom

LEARNING OBJECTIVE

- Describe the Dalton model of the atom.

In 1803, John Dalton (1766–1844), a modest Quaker schoolteacher, proposed that matter was discontinuous, not continuous as Aristotle had argued over 20 centuries before. But unlike Aristotle, Dalton offered experimental evidence for the particle nature of matter. Although he was not a distinguished scientist, Dalton presented convincing evidence based on the work of the famous French scientists Antoine Lavoisier and Joseph Proust, and the English physicist Robert Boyle.

Dalton proposed that matter consisted of atoms based on the laws of conservation of mass and definite proportion. Citing Proust's experiments in which two elements combine in more than one way, he formulated the law of multiple proportions; for example, carbon and oxygen can combine as CO and CO_2.

In the seventeenth century, Robert Boyle, the noted English physicist, had concluded from experiments that a gas was made up of tiny particles. Dalton expanded on Boyle's conclusion and proposed that all matter was composed of particles. In addition to Boyle's experiments, Dalton relied on two other scientific principles, the law of conservation of mass and the law of definite proportion.

In 1789, Antoine Lavoisier established the law of conservation of mass. By carefully weighing substances before and after a chemical reaction, Lavoisier demonstrated that matter was neither created nor destroyed. Although atoms are much too small to weigh directly, this principle allows us to predict the masses of substances involved in a chemical change. In 1799, Joseph Proust established the law of definite proportion (definite composition), which states that a compound always contains the same elements in the same proportion by mass.

Dalton initially presented his evidence for the existence of atoms to the *Literary and Philosophical Society* of Manchester, England. He proposed that an element is composed of tiny, indivisible, indestructible particles. Furthermore, he argued that compounds are simply combinations of two or more atoms of different elements. In 1808, Dalton published the atomic theory in his classic textbook *A New System of Chemical Philosophy*.

FEATURED SCIENTIST John Dalton

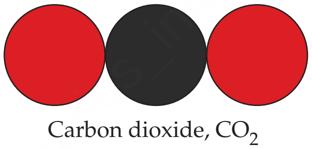
Q: How was John Dalton able to accept an honorary degree from Oxford University dressed in a scarlet robe when wearing scarlet was forbidden by his Quaker faith?

John Dalton was born the son of a weaver into a devoutly religious family. At the age of 12, he began teaching at a Quaker school and developed an interest in science. He was most interested in meteorology and kept a lifelong daily journal of atmospheric conditions for his hometown of Manchester, England.

◀ John Dalton (1766–1844)

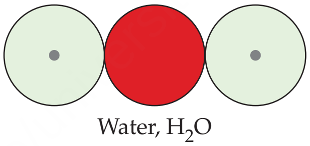
By all accounts, Dalton was not an inspiring lecturer, and he was hindered as a researcher by being color-blind. He had only a minimal education and limited finances, but he compensated for this with persistent and meticulous work habits. Dalton's daily study of the weather led him to conclude, like Robert Boyle and Isaac Newton before him, that the air was made up of gas particles. Over time, he began to construct his atomic theory as follows.

According to the law of definite composition, carbon and oxygen always react in the same mass ratio to produce carbon dioxide. Dalton proposed that one atom of carbon combines with two atoms of oxygen to produce a molecule of CO_2, as shown using Dalton's symbolism.



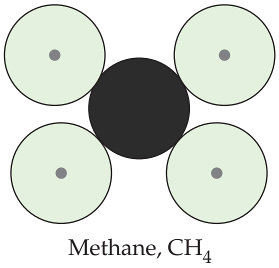
◀ Carbon dioxide, CO_2

Similarly, he proposed that two atoms of hydrogen combine with one atom of oxygen to give a molecule of H_2O.



◀ Water, H_2O

Finally, Dalton reasoned that two atoms of hydrogen in water could substitute for each of the oxygen atoms in carbon dioxide. This would result in a molecule with one carbon atom and four hydrogen atoms, that is, CH_4. At the time, this compound was known as marsh gas, but we now refer to it as methane.



◀ Methane, CH_4

Through experimentation, Dalton indeed found that the combining ratio of carbon to hydrogen in methane was 1 to 4; this agreed perfectly with his prediction. Thus, Dalton had experimental evidence to support the atomic theory.

A: Dalton rationalized that because he was color-blind and saw the ceremonial robe as gray, he was not violating Quaker doctrine.

In a surprisingly short period of time, the atomic theory was generally accepted by the scientific community. The theory can be summarized as follows:

1. An element is composed of tiny, indivisible, indestructible particles called atoms.
2. All atoms of an element are identical and have the same properties.
3. Atoms of different elements combine to form compounds.
4. Compounds contain atoms in small whole number ratios.
5. Atoms can combine in more than one ratio to form different compounds.

As we will learn, Dalton's first two proposals were incorrect, but the atomic theory was nonetheless an important step toward understanding the nature of matter.

4.2 Thomson Model of the Atom

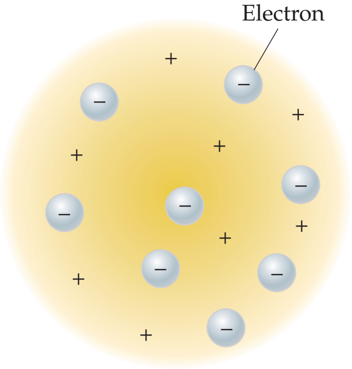
LEARNING OBJECTIVES

- Describe the Thomson raisin pudding model of the atom.
- Specify the relative charge on an electron and a proton.

Toward the end of the 1800s, there was disturbing evidence that the atom was divisible after all. Experiments involving the passage of electricity through hydrogen gas led to the discovery of a negatively charged particle called an **electron** ($\mathbf{e^-}$), and later a positively charged particle called a **proton** ($\mathbf{p^+}$). The electron was assigned a relative charge of −1, and the proton had an opposite, but equal charge of +1.

In 1903, J. J. Thomson (1856–1940) proposed a subatomic model of the atom. The subatomic model pictured a positively charged atom containing negatively charged electrons. Thomson visualized electrons in homogeneous spheres of positive charge in a way that was analogous to raisins in English plum pudding (Figure 4.1). Thus, the Thomson proposal became popularly known as the *plum pudding model* or *raisin pudding model of the atom*.

Originally, Thomson was only able to determine the relative charge-to-mass ratio of the electron and the proton. However in 1911, after 5 years of tedious observations, the American physicist Robert Millikan (1868–1953) determined the actual charge on the electron. This allowed Thomson to calculate the actual mass of the electron and the proton. Thomson calculated that the mass of the electron is 9.11×10^{-28} g, and that the mass of the proton is 1.67×10^{-24} g.



▲ **Figure 4.1 Thomson Model of the Atom** Atoms are pictured as spheres of positive charge. The small negative particles in the sphere represent electrons.

Note J. J. Thomson spent his academic career at Cambridge University and, at the age of 27, became director of the prestigious Cavendish Laboratory. In 1906, he won the Nobel Prize in physics for his work on the electron, and 2 years later he was knighted. Thomson's contribution was far-reaching—seven of his former students and assistants went on to receive Nobel Prizes.

4.3 Rutherford Model of the Atom

LEARNING OBJECTIVES

- Describe the Rutherford nuclear model of the atom.
- Specify the relative charge and approximate mass of the electron, proton, and neutron.

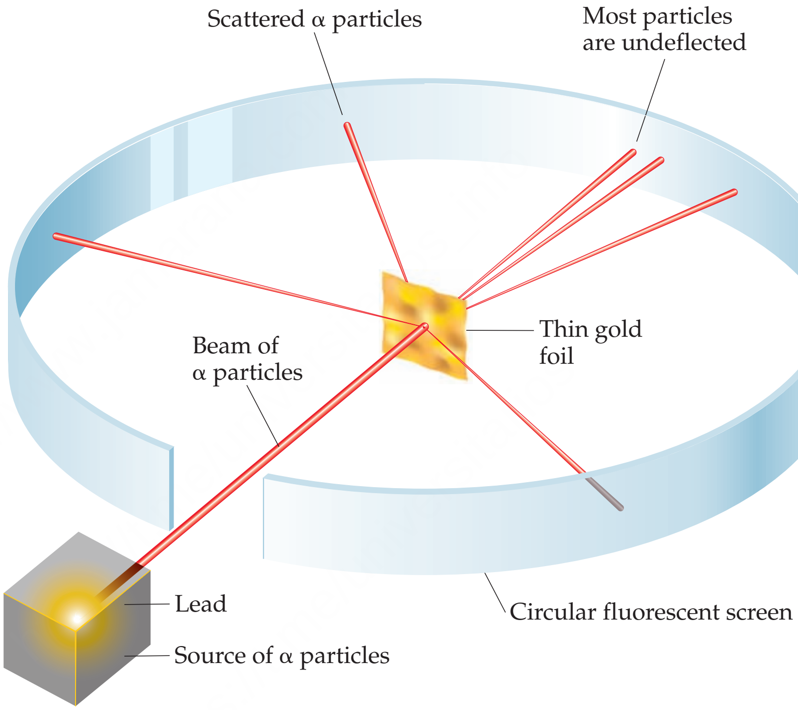
Ernest Rutherford (1871–1937) was digging potatoes on his father's farm in New Zealand when he received the news that he had won a scholarship to Cambridge University. He had actually come in second, but the winner had declined the scholarship to get married. The 24-year-old Rutherford postponed his own wedding plans and immediately set out for England.

At Cambridge, Ernest Rutherford studied subatomic particles and earned his doctorate working for J. J. Thomson. Upon graduation, Rutherford went to McGill University in Canada and began to work in the field of radioactivity, which studies unstable atoms that radiate subatomic particles and energy. After a year he went back to New Zealand, married, and returned to Manchester University in England. There he continued to study radioactivity and coined the terms *alpha ray* and *beta ray* for two types of radiation. He discovered a third type of radiation that was not affected by a magnetic field and gave it the name *gamma ray*. In 1908, Rutherford received the Nobel Prize in chemistry for his work on radioactivity.

In 1906, Rutherford found that alpha rays contained particles identical to those of helium atoms stripped of electrons. He experimented with alpha rays by firing them at thin gold foils. As expected, the particles passed straight through the foil or, on occasion, deflected slightly (Figure 4.2). This observation seemed reasonable because the plum pudding model of the atom pictured homogeneous spheres.

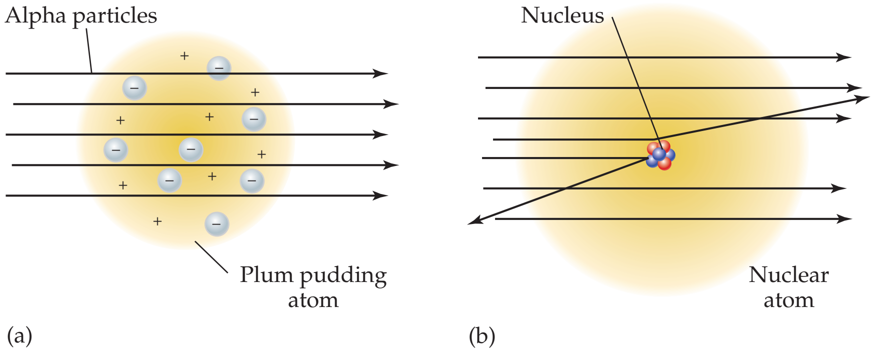
A few years later the true picture of the atom was unveiled. The person who did the experiment was Hans Geiger, Rutherford's assistant and the inventor of the Geiger counter. Here is a description of the experiment in Rutherford's own words:

> *One day Geiger came to me and said, "Don't you think that young Marsden, whom I am training in radioactive methods, ought to begin a small research?" Now I had thought that too, so I said, "Why not let him see if any alpha particles can be scattered through a large angle?" I may tell you in confidence that I did not believe that they would be, since we knew that the alpha particle was a very fast massive particle, with a great deal of energy. Then I remember two or three days later Geiger coming to me in great excitement and saying, "We have been able to get some of the alpha particles coming backwards." It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.*



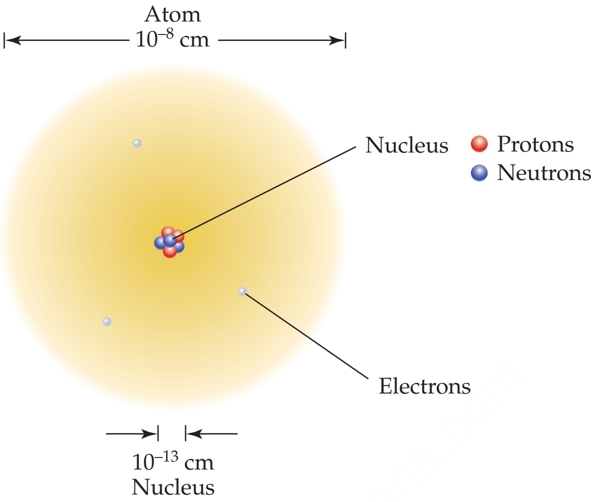
▶ **Figure 4.2 Alpha-Scattering Experiment** The diagram shows the deflection of alpha particles by a thin gold foil. Although the foil was only 0.5 μm thick, a few alpha particles actually rebounded backward.

Rutherford interpreted the alpha-scattering results as follows. He believed that most of the alpha particles passed directly through the foil because an atom is largely empty space with electrons moving about. But in the center of the atom is the **atomic nucleus** containing protons. Rutherford reasoned that, compared to the atom, the nucleus is tiny and has a very high density. The alpha particles that bounced backward recoiled after striking the dense nucleus. The scattering of alpha particles by the atomic nuclei in the gold foil is illustrated in Figure 4.3.



▶ **Figure 4.3 Explanation of Alpha-Scattering** (a) Powerful alpha particles should pass through a homogeneous plum pudding gold atom. (b) Because positively charged alpha particles are deflected, Rutherford reasoned correctly that a gold atom has a dense, positively charged nucleus.

In 1911, Rutherford proposed a new model of the atom. He suggested that negatively charged electrons were distributed about a positively charged nucleus. Rutherford was able to estimate the size of the atom and its nucleus. He calculated that an atom has a diameter of about 1×10^{-8} cm, and that the nucleus has a diameter of about 1×10^{-13} cm. The Rutherford model of the atom is illustrated in Figure 4.4.



◀ **Figure 4.4 Rutherford Model of the Atom** Most of the mass of an atom is found in the nucleus. Note that the atoms and nuclei are not to scale, as the size of an atom is about 100,000 times larger than the size of the nucleus.

The diameters of atoms and nuclei are extremely small and difficult to comprehend, so consider the following analogy to visualize an atom. The size of an atom compared to the size of its nucleus is similar to a kilometer compared to a centimeter. By analogy, if the atom were as large as the New Orleans Superdome, the nucleus would be the size of a small marble. This analogy clarifies why so few alpha particles were scattered by gold nuclei in the thin foil.

Owing to the heaviness of the nucleus, Rutherford predicted that it contained neutral particles in addition to protons. Twenty years later, one of Rutherford's former students found the elusive neutral particle. In 1932, James Chadwick (1891–1974) discovered the **neutron** (**n^0**) and 3 years later was awarded the Nobel Prize in physics. We can summarize the data for the electron, proton, and neutron as shown in Table 4.1.

TABLE 4.1 Subatomic Particles

Subatomic Particle	Symbol	Location	Relative Charge	Relative Mass
electron	e^-	outside nucleus	−1	1/1836
proton	p^+	inside nucleus	+1	1
neutron	n^0	inside nucleus	0	1

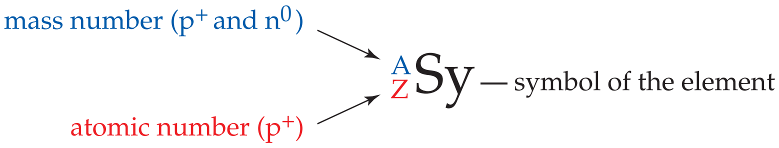
4.4 Atomic Notation

Each element has a characteristic number of protons in its atomic nucleus. This value is called the **atomic number**. The total number of protons and neutrons in the nucleus of an atom is called the **mass number**. A shorthand method for keeping track of protons and neutrons in the nucleus of an atom is called **atomic notation**. By convention, the symbol of the element (Sy) is preceded by a subscript and superscript. The subscript,

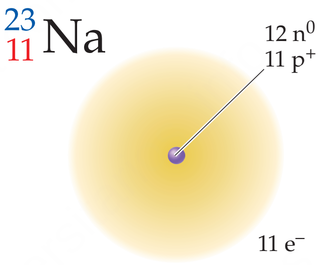
LEARNING OBJECTIVES

- Analyze atomic notation and indicate protons, neutrons, and electrons.
- Analyze a given isotope and indicate protons, neutrons, and electrons.

designated the *Z* value, represents the atomic number. The superscript, designated the *A* value, represents the mass number. Thus,



As an example, an atom of sodium can be written as $^{23}_{11}\text{Na}$. Here the atomic number (number of protons) is 11, and the mass number (number of protons + neutrons) is 23. From this information, we can determine the number of neutrons. The nucleus of this sodium atom contains 12 neutrons ($A - Z = 23 - 11 = 12$). Because atoms are neutral, the number of negative electrons must equal the number of positive protons. Thus, there are 11 electrons surrounding the nucleus. We can show a simple diagram of the sodium atom as follows:



Example Exercise 4.1 further illustrates the interpretation of atomic notation to state the composition of an atom.

EXAMPLE EXERCISE 4.1 Atomic Notation

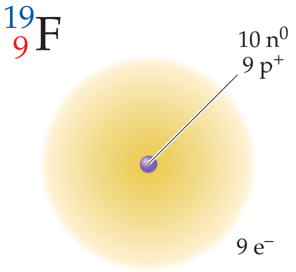
Given the atomic notation for the following atoms, draw a diagram showing the arrangement of protons, neutrons, and electrons.

(a) $^{19}_{9}\text{F}$
(b) $^{109}_{47}\text{Ag}$

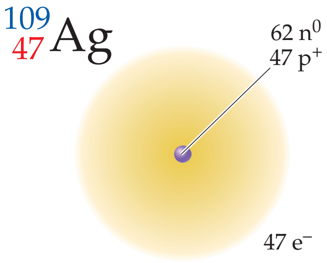
Solution

We can draw a diagram of an atom by showing protons and neutrons in the nucleus surrounded by electrons.

(a) Because the atomic number is 9 and the mass number is 19, the number of neutrons is 10 (19 − 9). If there are 9 protons, there must be 9 electrons.

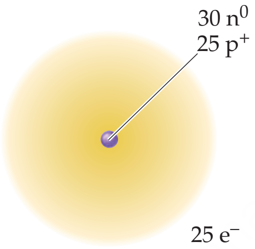


(b) Because the atomic number is 47 and the mass number is 109, the number of neutrons is 62 (109 − 47). If there are 47 protons, there must be 47 electrons.



Practice Exercise

Given the following diagram, indicate the nucleus using atomic notation.



Answer: ${}^{55}_{25}Mn$

Concept Exercise

Can atoms of different elements have the same atomic number?

Answer: See Appendix G, 4.1.

Isotopes

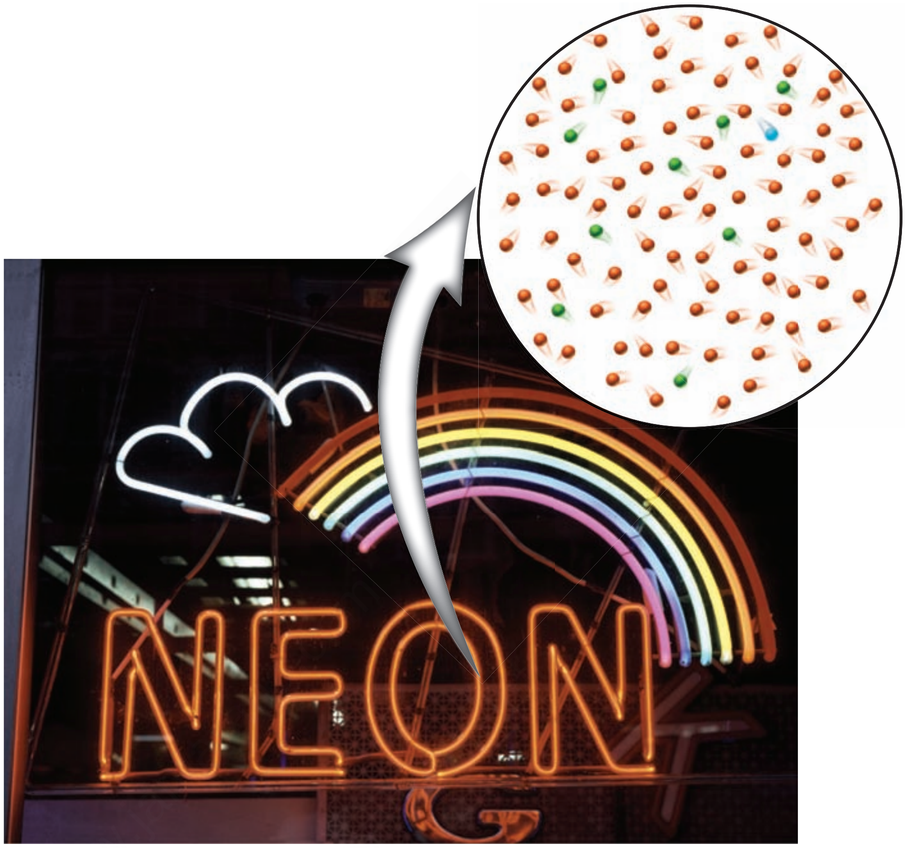
Approximately 20 elements occur naturally with a fixed number of neutrons. For most elements, the number of neutrons in the nucleus varies. Atoms of the same element that have a different number of neutrons in the nucleus are called **isotopes**.

For isotopes, as the number of neutrons varies, so does the mass number. For example, hydrogen occurs naturally as two stable isotopes, protium (${}^{1}_{1}H$) and deuterium (${}^{2}_{1}H$). Protium has only one proton in its nucleus, whereas deuterium has a proton and a neutron. A third isotope of hydrogen, tritium (${}^{3}_{1}H$), is unstable and radioactive. Tritium has one proton and two neutrons in its nucleus. Table 4.2 lists the naturally occurring stable isotopes of the first 10 elements.

TABLE 4.2 Naturally Occurring Stable Isotopes of the First 10 Elements

Atomic Number	Mass Number	Atomic Notation	Atomic Number	Mass Number	Atomic Notation
1	1	${}^{1}_{1}H$	6	13	${}^{13}_{6}C$
1	2	${}^{2}_{1}H$	7	14	${}^{14}_{7}N$
2	3	${}^{3}_{2}He$	7	15	${}^{15}_{7}N$
2	4	${}^{4}_{2}He$	8	16	${}^{16}_{8}O$
3	6	${}^{6}_{3}Li$	8	17	${}^{17}_{8}O$
3	7	${}^{7}_{3}Li$	8	18	${}^{18}_{8}O$
4	9	${}^{9}_{4}Be$	9	19	${}^{19}_{9}F$
5	10	${}^{10}_{5}B$	10	20	${}^{20}_{10}Ne$
5	11	${}^{11}_{5}B$	10	21	${}^{21}_{10}Ne$
6	12	${}^{12}_{6}C$	10	22	${}^{22}_{10}Ne$

▶ **Neon Isotopes** The element neon consists of three naturally occurring isotopes: Ne-20, Ne-21, and Ne-22. We can deduce by the atomic mass of neon (20.18 amu) that Ne-20 is the most abundant isotope.



We often refer to an isotope by stating the name of the element followed by its mass number, for example, carbon-12. To write the atomic notation, we must also know the atomic number. In this textbook, the atomic number is found above the symbol of the element in the periodic table. If we refer to the inside front cover of this textbook, we find that the atomic number of C is 6. Thus, we can write the atomic notation for carbon-12 as ${}^{12}_{6}\mathrm{C}$.

Example Exercise 4.2 further illustrates the relationship between atomic notation and the composition of an atomic nucleus.

EXAMPLE EXERCISE 4.2 Nuclear Composition of Isotopes

State the number of protons and the number of neutrons in an atom of each of the following isotopes.

(a) ${}^{37}_{17}\mathrm{Cl}$

(b) mercury-202

Solution

The subscript value refers to the atomic number (p^+), and the superscript value refers to the mass number (p^+ and n^0).

(a) Thus, ${}^{37}_{17}\mathrm{Cl}$ has 17 p^+ and 20 n^0 ($37 - 17 = 20$).

(b) In the periodic table, we find that the atomic number of mercury is 80. Thus, the atomic notation, ${}^{202}_{80}\mathrm{Hg}$, indicates 80 p^+ and 122 n^0 ($202 - 80 = 122$).

Practice Exercise

State the number of protons and the number of neutrons in an atom of each of the following isotopes.

(a) ${}^{120}_{50}\mathrm{Sn}$

(b) uranium-238

Answers:

(a) 50 p^+ and 70 n^0 (b) 92 p^+ and 146 n^0

Concept Exercise

Can atoms of different elements have the same mass number?

Answer: See Appendix G, 4.2.

Note In general, the properties of isotopes of an element are similar. Consider a carbon-12 atom, which has six protons and six neutrons in the nucleus. If the mass number increases by one neutron, we have a carbon-13 atom. Other than a slight change in mass, the properties of carbon-12 and carbon-13 are nearly identical.

On the other hand, when the atomic number of carbon-12 increases by one proton, we have a nitrogen-13 atom. The additional proton changes the properties dramatically. Carbon occurs naturally as coal, diamond, and graphite, whereas nitrogen is a colorless, odorless gas in the atmosphere.

4.5 Atomic Mass

LEARNING OBJECTIVES

- Explain the concept of relative atomic mass.
- Calculate the atomic mass for an element given the mass and abundance of the naturally occurring isotopes.

We realize, of course, that atoms are much too small to weigh directly on a balance. A carbon atom, for instance, has a mass of only 1.99×10^{-23} g. Instead of weighing atoms, scientists determine the mass of atoms relative to each other. With a special magnetic-field instrument, the mass of an atom can be compared to the mass of a carbon-12 atom. The carbon-12 isotope has been chosen as a reference standard and is assigned a value of exactly 12 atomic mass units. Stated differently, an **atomic mass unit (amu)** is equal to 1/12 the mass of a carbon-12 atom.

Simple and Weighted Averages

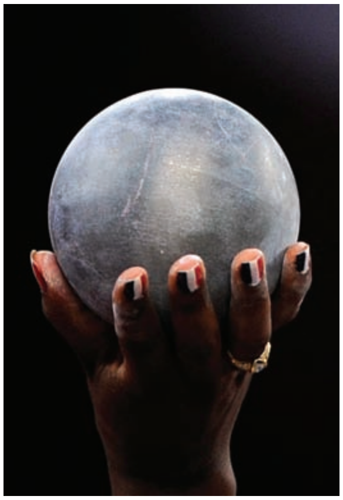
Before proceeding, let's contrast two kinds of averages. Let's suppose a company manufactures 16-lb and 8-lb shot put balls for Olympic men and women athletes. Of all the shot put balls manufactured, 75% are 16 lb and 25% are 8 lb. What is the average mass of a shot put ball? If we add the two masses together and divide by 2, we obtain the simple average mass. That is,

$$\frac{16\ \text{lb} + 8\ \text{lb}}{2} = 12\ \text{lb}$$

The simple average mass of a shot put ball is 12 lb, but the true average mass must be heavier, and weighted in favor of the higher percentage of 16-lb balls. Because the percentages are 75% and 25%, we will write the decimal fractions 0.75 and 0.25, and proceed as follows:

$$\begin{aligned} \text{16-lb balls: } 16\ \text{lb} \times 0.75 &= 12\ \text{lb} \\ \text{8-lb balls: } 8\ \text{lb} \times 0.25 &= \underline{\ \ 2\ \text{lb}} \\ &\ 14\ \text{lb} \end{aligned}$$

Although the weighted average mass is 14 lb, we should note that no shot put ball weighs 14 lb. An actual shot put ball weighs either 16 lb or 8 lb. A mass of 14 lb represents a theoretical average shot put ball.



▲ **Olympic Shot Put** In the women's Olympic shot put event, female athletes toss an 8 lb steel ball.

Atomic Mass of an Element

The **atomic mass** of an element is the weighted average mass of all naturally occurring isotopes. To calculate the atomic mass, we will use a method similar to the one used in the shot put example. That is, to calculate the weighted average mass of an atom, we must consider the mass as well as the percent abundance of each isotope.

Carbon has two naturally occurring stable isotopes: carbon-12 and carbon-13. We can calculate the atomic mass of carbon given the mass and natural abundance of each isotope.

Isotope	Mass	Abundance
^{12}C	12.000 amu	98.89%
^{13}C	13.003 amu	1.11%

The natural abundance of each isotope, expressed as a decimal, is 0.9889 and 0.0111. To calculate the true weighted average mass of a carbon atom, we must consider the mass and percentage of each isotope. We proceed as follows:

$$\begin{aligned} ^{12}\mathrm{C}: 12.000\text{ amu} \times 0.9889 &= 11.87\text{ amu} \\ ^{13}\mathrm{C}: 13.003\text{ amu} \times 0.0111 &= \underline{\ 0.144\text{ amu}} \\ & 12.01\text{ amu} \end{aligned}$$

Even though the atomic mass of carbon is 12.01 amu, we should note that no carbon atom weighs 12.01 amu. An actual carbon atom weighs either 12.000 amu or 13.003 amu. A mass of 12.01 amu represents the mass of a hypothetical average carbon atom. Example Exercise 4.3 further illustrates the calculation of atomic mass from isotopic mass and abundance data.

EXAMPLE EXERCISE 4.3 Calculation of Atomic Mass

Silicon is the second most abundant element in Earth's crust. Calculate the atomic mass of silicon given the following data for its three natural isotopes:

Isotope	Mass	Abundance
^{28}Si	27.977 amu	92.21%
^{29}Si	28.976 amu	4.70%
^{30}Si	29.974 amu	3.09%

Solution

We can find the atomic mass of silicon as follows:

$$\begin{aligned} ^{28}\mathrm{Si}: 27.977\text{ amu} \times 0.9221 &= 25.80\text{ amu} \\ ^{29}\mathrm{Si}: 28.976\text{ amu} \times 0.0470 &= 1.36\text{ amu} \\ ^{30}\mathrm{Si}: 29.974\text{ amu} \times 0.0309 &= \underline{0.926\text{ amu}} \\ & 28.09\text{ amu} \end{aligned}$$

The average mass of a silicon atom is 28.09 amu, although we should note that there are no silicon atoms with a mass of 28.09 amu.

Practice Exercise

Calculate the atomic mass of copper given the following data:

Isotope	Mass	Abundance
^{63}Cu	62.930 amu	69.09%
^{65}Cu	64.928 amu	30.91%

Answer: 63.55 amu

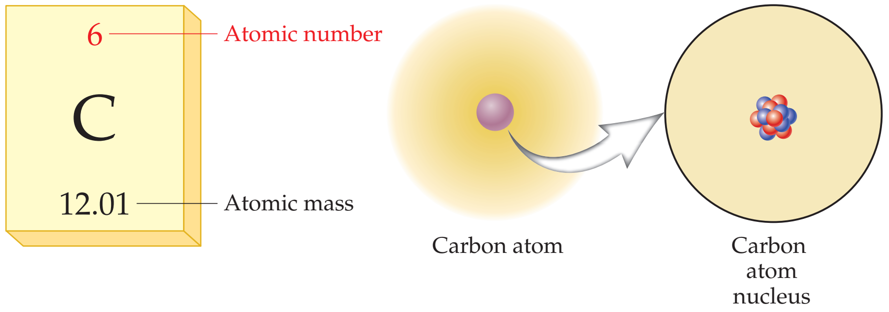
Concept Exercise

A bag of marbles has 75 large marbles with a mass of 2.00 g each, and 25 small marbles with a mass of 1.00 g each. Calculate (a) the simple average mass, and (b) the weighted average mass of the marble collection.

Answer: See Appendix G, 4.3.

The Periodic Table

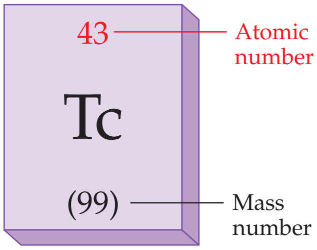
We often need to refer to the atomic number and atomic mass of an element. For convenience, this information is listed in the periodic table. In this textbook the atomic number is indicated above the symbol, and the atomic mass is given below the symbol. Figure 4.5 illustrates carbon as it appears in the periodic table.



◀ **Figure 4.5 Carbon in the Periodic Table** The atomic number designates the number of protons (6). The atomic mass designates the weighted average mass of the naturally occurring stable isotopes of the element.

When we refer to the periodic table, it is not possible to determine the number of naturally occurring isotopes for a given element. Of the first 83 elements, however, 81 elements have one or more stable isotopes. Only technetium (element 43) and promethium (element 61) are radioactive and unstable. In the periodic table on the inside front cover of this textbook, the mass number of unstable elements is given in parentheses as shown in Figure 4.6.

A value in parentheses below the symbol of an element indicates the mass number of the most stable or best known radioactive isotope. The mass number refers to the total number of protons and neutrons in the unstable nucleus. From the periodic table you can easily distinguish the elements that are stable from those that are unstable and radioactive. Example Exercises 4.4 and 4.5 further illustrate how to use the periodic table as a reference.



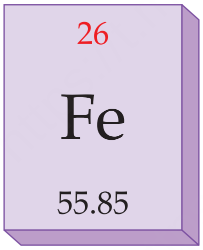
▲ **Figure 4.6 Technetium in the Periodic Table** The atomic number designates the number of protons (43). The mass number for this radioactive isotope is 99. Technetium-99 has a total of 43 protons and 56 neutrons in its nucleus.

EXAMPLE EXERCISE 4.4 Nuclear Composition from the Periodic Table

Refer to the periodic table on the inside cover of this text and determine the atomic number and atomic mass for iron.

Solution

In the periodic table we observe



The atomic number of iron is 26, and the atomic mass is 55.85 amu. It is not possible to determine the number of iron isotopes, or their masses, from the periodic table.

Practice Exercise

Refer to the periodic table and determine the atomic number and atomic mass for sulfur.

Answers: 16 and 32.07 amu

Concept Exercise

Sodium has only one stable isotope. What is the mass of the Na-23 stable isotope?

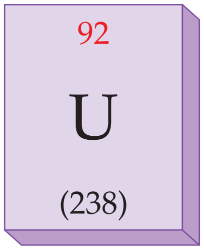
Answer: See Appendix G, 4.4.

EXAMPLE EXERCISE 4.5 Nuclear Composition from the Periodic Table

Refer to the periodic table on the inside cover of this text and determine the atomic number and mass number for element 92, uranium.

Solution

In the periodic table we observe



The atomic number of uranium is 92. The mass number shown is 238, which is indicated by the use of parentheses. U-233 and U-235 are other radioactive isotopes of uranium, which are not stable.

Practice Exercise

Refer to the periodic table and determine the atomic number and mass number listed for element 61, promethium.

Answers: 61 and (147)

Concept Exercise

Which of the following is *never* a whole number value: atomic number, atomic mass, or mass number?

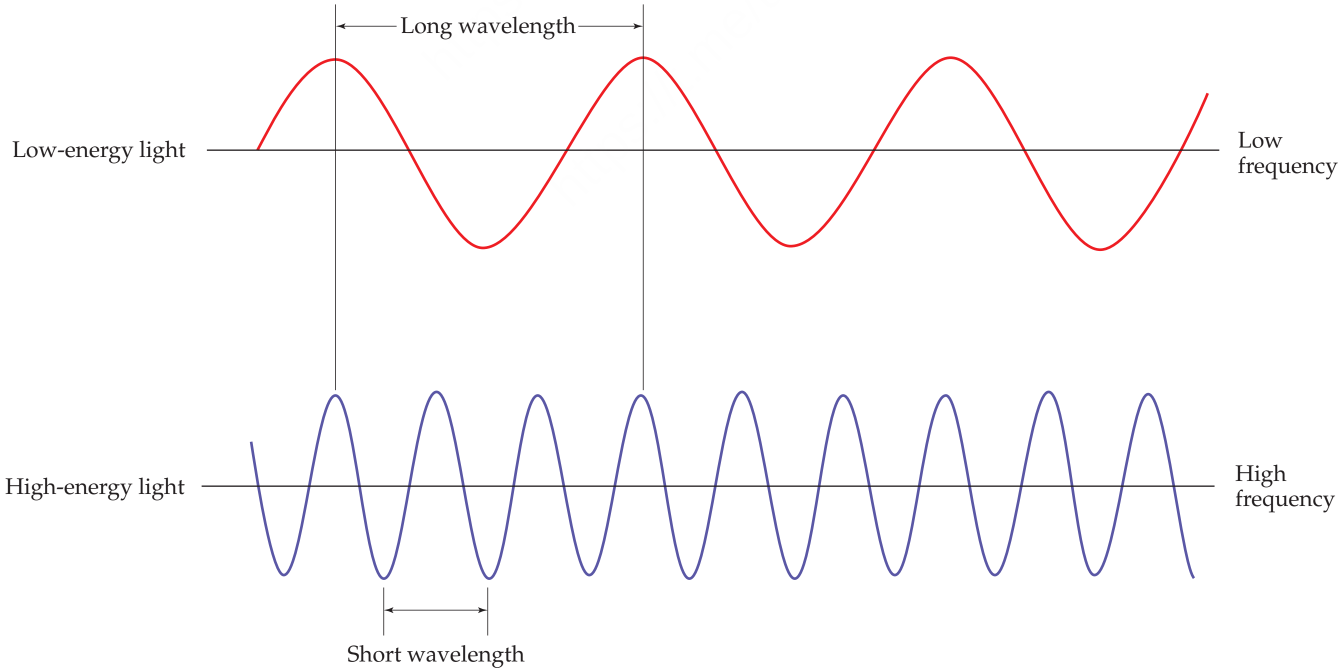
Answer: See Appendix G, 4.5.

LEARNING OBJECTIVE

- Analyze the relationship of wavelength, frequency, and energy of light.

4.6 The Wave Nature of Light

We can use our imagination to visualize light traveling through space in a fashion similar to an ocean wave. **Wavelength** refers to the distance the light wave travels to complete one cycle. **Frequency** refers to the number of wave cycles completed in each second. The velocity of light is constant, as all wavelengths and frequencies travel at 3.00×10^8 m/s. Figure 4.7 illustrates the wave nature of light.



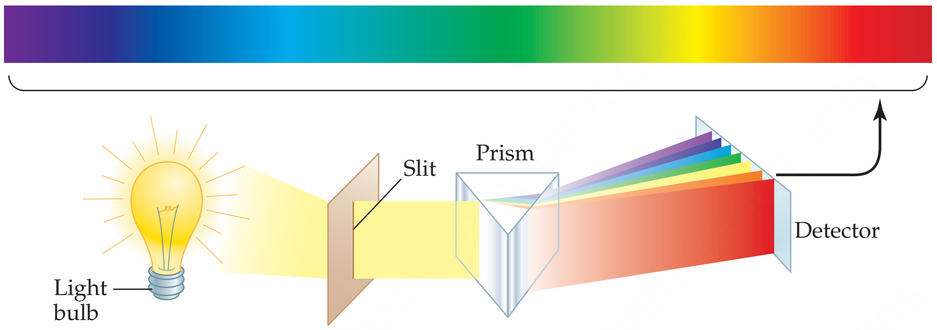
▲ **Figure 4.7 Wavelength and Frequency of Light** Notice that the wavelength is longer for low-energy light than for high-energy light. Also notice that the frequency is greater for high-energy light than for low-energy light.

As the wavelength of light decreases, the frequency of a light wave increases. Conversely, as the wavelength increases, the frequency decreases. As an example, the wavelength decreases from red to violet as the frequency increases. We can visualize a high-frequency light wave moving rapidly up and down while completing many cycles per second. Consequently, high-frequency light is more energetic than low-frequency light, and short-wavelength light is more energetic than long-wavelength light.

Light—A Continuous Spectrum

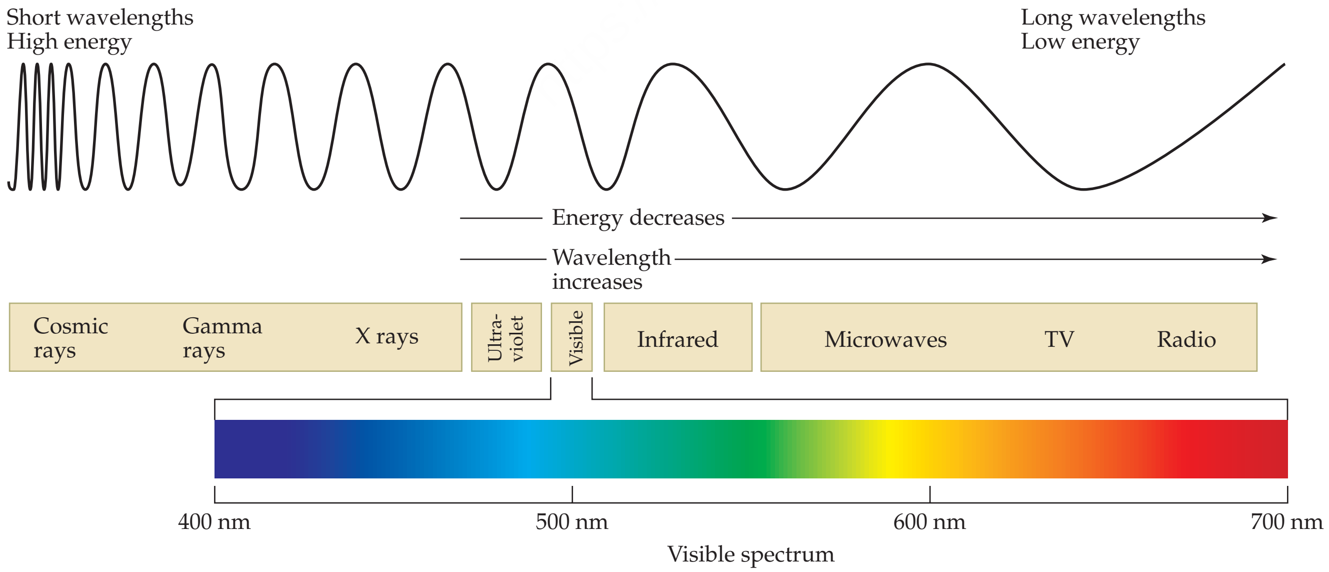
The term **light** usually refers to radiant energy that is visible. Our eyes can see light in the **visible spectrum** (400–700 nm), but not in the ultraviolet spectrum or the infrared spectrum. The wavelength of ultraviolet radiation is too short to be visible (below 400 nm), and infrared radiation is too long (above 700 nm).

A rainbow is created when sunlight passes through raindrops. Each raindrop acts as a miniature prism and separates sunlight into various bands of color. What we ordinarily observe as white light is actually several colors of light mixed together. When white light passes through a glass prism, it separates into all the colors of the rainbow (Figure 4.8).



◀ **Figure 4.8 White Light Passing through a Prism** Notice that white light separates into a rainbow of colors when it passes through a glass prism. Similarly, sunlight produces a rainbow when it passes through raindrops that act as miniature prisms.

We also, on occasion, use the term *light* when referring to radiant energy that is not visible. The complete **radiant energy spectrum** is an uninterrupted band, or **continuous spectrum**, of visible and invisible light that ranges from short to long wavelengths. The radiant energy spectrum includes invisible gamma rays, X rays, and microwaves, as well as visible light (Figure 4.9).



▲ **Figure 4.9 The Radiant Energy Spectrum** The complete radiant energy spectrum includes short-wavelength gamma rays through long-wavelength microwaves. Notice that the visible spectrum is only a narrow window in a broad band of radiant energy.

Example Exercise 4.6 illustrates how the wavelength, frequency, and energy of light are related.

EXAMPLE EXERCISE 4.6 Properties of Light

Considering blue light and yellow light, which has the

(a) longer wavelength?
(b) higher frequency?
(c) higher energy?

Solution

Referring to Figure 4.9, we notice that the wavelength of yellow light is about 600 nm and that of blue light is about 500 nm. Thus,

(a) *yellow light* has a longer wavelength than blue light.
(b) *blue light* has a higher frequency because it has a shorter wavelength.
(c) *blue light* has a higher energy because it has a higher frequency.

Practice Exercise

Considering infrared light and ultraviolet light, which has the

(a) longer wavelength?
(b) higher frequency?
(c) higher energy?

Answers:

(a) infrared
(b) ultraviolet
(c) ultraviolet

Concept Exercise

The energy of light (increases/decreases) as the wavelength increases. The energy of light (increases/decreases) as the frequency increases.

Answer: See Appendix G, 4.6.

A CLOSER LOOK "Nuking" Food in a Microwave Oven

Q: How is food heated in a microwave oven?

◀ Microwave oven

The term "nuking" food originated after World War II, which is when the United States developed the first atomic bomb and nuclear weapons. The microwave oven was developed during this same period; hence, the term "nuking" food became popular. The term is still popular today, but do microwave ovens "nuke" food in order to heat it?

The microwave region of the radiant energy spectrum spans wavelengths of approximately 0.1 cm to 100 cm. The radiation in a microwave oven has a wavelength of 12 cm. Refer to Figure 4.9 and notice that microwaves have a longer wavelength (less energy) than infrared radiation, which has a longer wavelength (less energy) than visible light.

When placing food in a microwave oven, the food is exposed to microwave energy that heats the food by causing the water molecules in the food to rotate faster. When the water molecules lose rotational energy and return to their normal energy state at room temperature, the energy lost by the water molecules heats the food.

When an atomic nucleus loses energy, the radiant energy corresponds to very high energy gamma rays (see Figure 4.9). Gamma rays have wavelengths that are a billion times shorter than microwaves, but much higher frequency and energy. Exposing food to gamma radiation would completely destroy the food.

Although microwave cooking is generally considered healthy, and food retains more vitamins than stove-cooked food, not all microwaved food is safe. According to the U.S. Department of Agriculture, we should avoid microwave cooking in disposable plastic containers as microwaves can leach harmful chemicals into the food. The alternative is to use microwave-safe plastic or glass containers.

A: The energy inside a microwave oven causes the water molecules in food to rotate faster. When these molecules lose rotational energy and return to their normal energy state at room temperature, the energy lost by the water molecules heats the food.

4.7 The Quantum Concept

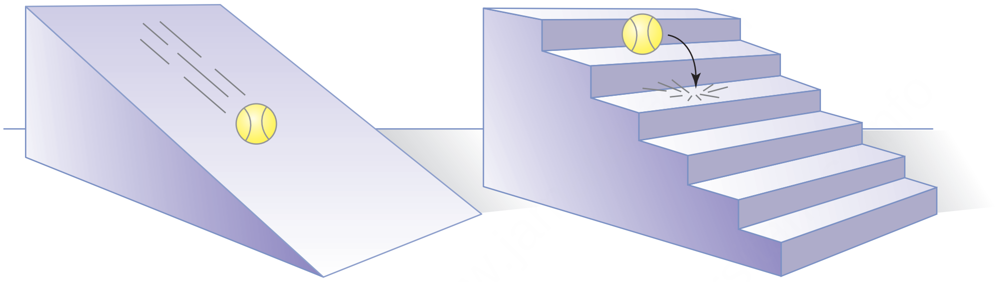
LEARNING OBJECTIVE

- Illustrate the quantum concept applied to matter and energy.

In 1900, Max Planck (1858–1947), a German physicist, introduced a revolutionary idea. Planck proposed that the energy radiated by a heated object is not continuous, but rather that the radiation is emitted in small bundles. The idea that energy is composed of discrete units is referred to as the quantum concept.

Although radiant energy has a wave nature, it also has a particle nature, and an individual unit of light energy is referred to as a quantum, or a **photon**. According to the quantum theory, an ordinary light bulb radiates light which has both a wave and particle nature.

We can illustrate the quantum concept with the following analogy. A ball rolling down a ramp loses potential energy continuously. Conversely, a ball rolling down a flight of stairs loses potential energy in discrete units each time it drops from one step to another. In this example, the loss of potential energy is continuous as the ball rolls down the ramp, and the loss is quantized as the ball rolls down the stairs. Figure 4.10 contrasts a continuous change in energy versus a quantized energy change.



◀ **Figure 4.10 Stair Analogy for the Quantum Principle** A ball rolling down a ramp loses potential energy continuously. Conversely, a ball rolling down a flight of stairs loses potential energy in quantized amounts each time it drops from one step to another one.

The following Example Exercise 4.7 further illustrates practical applications of the quantum theory.

EXAMPLE EXERCISE 4.7 Quantum Concept

State whether each of the following scientific instruments gives a continuous or a quantized measurement of mass:

(a) triple-beam balance
(b) digital electronic balance

Solution

Refer to Figure PSS.3 if you have not used these balances in the laboratory.

(a) On a triple-beam balance a small metal rider is moved along a beam. Because the metal rider can be moved to any position on the beam, a triple-beam balance gives a *continuous* mass measurement.

(b) On a digital electronic balance the display indicates the mass of an object to a particular decimal place, for example, 5.015 g. Because the last digit in the display must be a whole number, a digital balance gives a *quantized* mass measurement.

Practice Exercise

State whether each of the following musical instruments produces continuous or quantized musical notes:

(a) acoustic guitar
(b) electronic keyboard

Answers:

(a) continuous
(b) quantized

Concept Exercise

Complete the following quantum analogy: a water wave is to a drop of water, as a light wave is to a ______.

Answer: See Appendix G, 4.7.

LEARNING OBJECTIVES

- Describe the Bohr model of the atom.
- Explain the relationship between energy levels in an atom and lines in an emission spectrum.

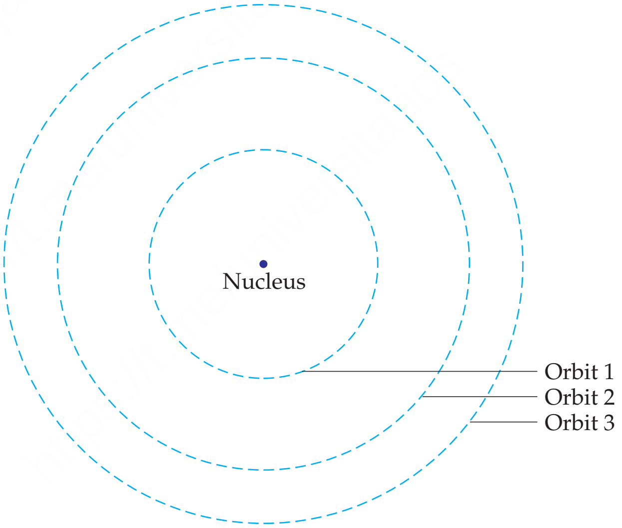
4.8 Bohr Model of the Atom

In 1911, a brilliant young Danish physicist, Niels Bohr (1885–1962), completed his doctorate and left for England to begin postdoctoral work under J. J. Thomson. After a few months, he left Cambridge to join Rutherford at Manchester where the atomic nucleus had been discovered. It took only 2 years for the young Dane to raise our understanding of the atom to yet another level.

In 1913, Bohr speculated that electrons orbit around the atomic nucleus just as planets orbit around the Sun. He further suggested that electron orbits were at a fixed distance from the nucleus and had a definite energy. The electron was said to travel in a fixed-energy orbit referred to as an **energy level**. Moreover, electrons could be found only in specific energy levels and nowhere else. Figure 4.11 illustrates the model of the hydrogen atom proposed by Niels Bohr. This model is referred to as the Bohr model of the atom, or simply, the **Bohr atom**.

Evidence for Energy Levels

The Bohr model was a beautiful mental picture of electrons in atoms. However, no one knew if the model was right or wrong because there was no experimental evidence to support the theory. Coincidentally, about this time, Bohr received an article on the emission of light from hydrogen gas. The article showed that excited hydrogen gas emits separate lines of color, rather than a continuous band of color. The three most prominent lines are violet, blue-green, and red.

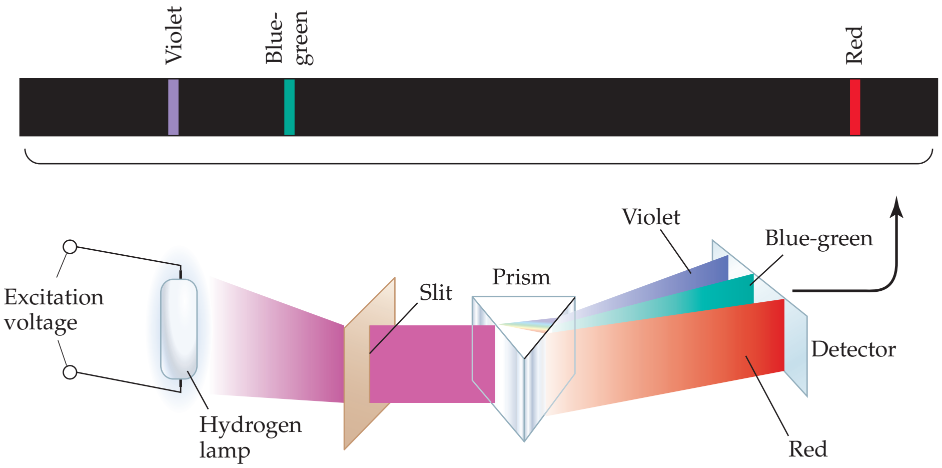


▶ **Figure 4.11 Bohr Model of the Hydrogen Atom** The electron is a specific distance from the nucleus and occupies an orbit of discrete energy. According to the Bohr model, the electron is found only in a given energy level.

If hydrogen gas in a glass tube is subjected to electricity, the excited hydrogen gas releases energy as reddish-purple light. When the reddish-purple light passes through a glass prism, it separates into narrow bands of color. This collection of narrow bands of light is called an **emission line spectrum**, and the individual bands of light are called *spectral lines*. The emission spectrum of hydrogen gas is shown in Figure 4.12.

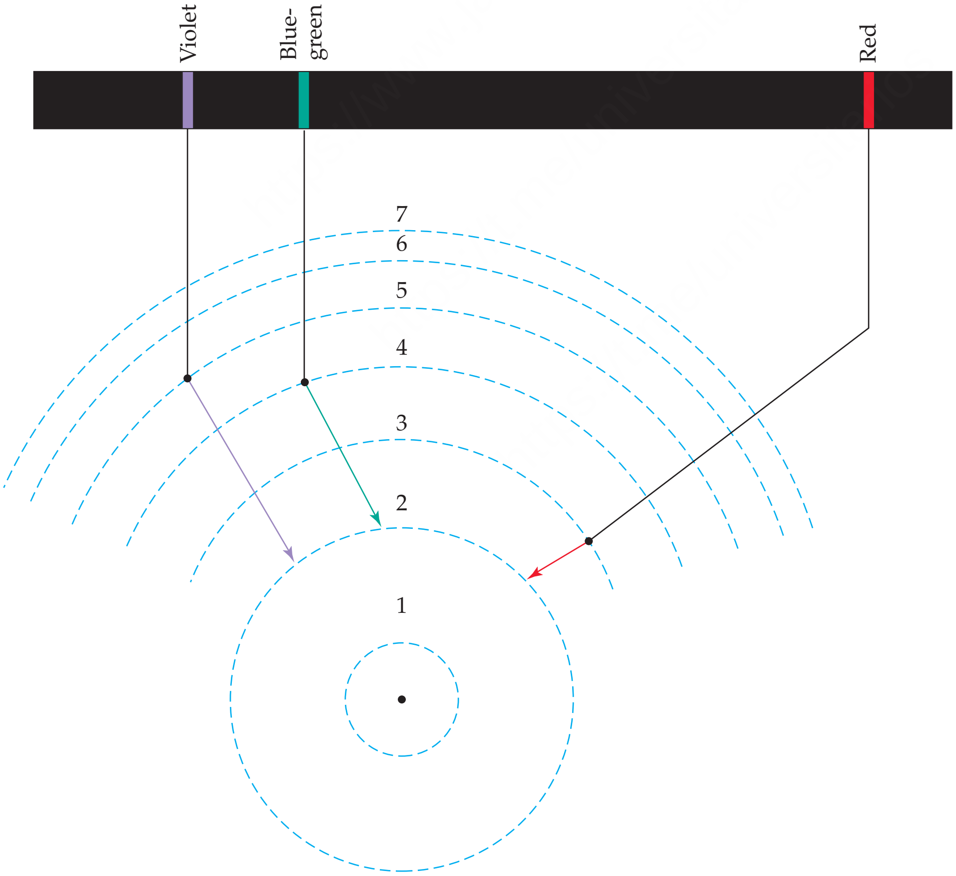
After examining the emission spectrum of hydrogen, Bohr realized that he had experimental evidence to support his model of the atom. The concept of electron energy levels was supported by the line spectrum of hydrogen. In a gas discharge tube, excited atoms of hydrogen have electrons in a high-energy orbit; for example, the electron may temporarily occupy the second orbit. Because this state is unstable the electron quickly drops from the higher level back to a lower level, that is, from level 2 to level 1. In the process, the electron loses a discrete amount of energy. This discrete energy loss corresponds to a photon of light energy. The energy of the photon of light equals the amount of energy lost by the electron as it drops from the higher to the lower energy level.

▲ **Niels Bohr** The planetary model of the atom was proposed by Niels Bohr.



◀ **Figure 4.12 Hydrogen Emission Spectrum** A reddish-purple light is observed when a voltage is applied to a hydrogen lamp. When the reddish-purple light passes through a glass prism, three lines are observed: violet, blue-green, and red.

In the hydrogen spectrum, there are three bright lines: red, blue-green, and violet. The red line has the longest wavelength of the three and the lowest energy. Bohr found that the red line corresponds to an excited electron dropping from energy level 3 to 2. The blue-green line is more energetic than the red line and corresponds to an excited electron dropping from energy level 4 to 2. The violet line is the most energetic of the three and is produced when an electron drops from energy level 5 to 2. Figure 4.13 illustrates the correlation between Bohr's electron energy levels and the observed lines in the hydrogen spectrum. When the electron drops to energy level 1, the energy of the photon is in the ultraviolet region, and therefore not visible.



◀ **Figure 4.13 Spectral Lines and Energy Levels in Hydrogen** When an electron drops from energy level **5** to **2**, a violet photon is released. When an electron drops from level **4** to **2**, a blue-green photon is released; from level **3** to **2**, a red photon is released. A visible spectral line corresponds to a large number of identical photons.

Because a single photon is emitted each time an electron drops to a lower level, it follows that several photons are emitted when several electrons change levels. For example, if the electron drops from energy level 5 to 2 in 10 hydrogen atoms, 10 photons would be emitted. Each of the photons would have the same energy, and collectively they would be observed as a violet line in the emission spectrum. Example Exercise 4.8 further illustrates the relationship between energy levels and emission lines.

EXAMPLE EXERCISE 4.8 Emission Spectra and Energy Levels

Explain the relationship between an observed emission line in a spectrum and electron energy levels.

Solution

When an electron drops from a higher to a lower energy level, light is emitted. For each electron that drops, a single photon of light energy is emitted. The energy lost by the electron that drops equals the energy of the photon that is emitted. Several photons of light having the same energy are observed as a spectral line.

Practice Exercise

Indicate the number and color of the photons emitted for each of the following electron transitions in hydrogen atoms:

(a) 1 e^- dropping from energy level 3 to 2
(b) 10 e^- dropping from energy level 3 to 2
(c) 100 e^- dropping from energy level 4 to 2
(d) 500 e^- dropping from energy level 5 to 2

Answers:

(a) 1 red photon
(b) 10 red photons
(c) 100 blue-green photons
(d) 500 violet photons

Concept Exercise

Which of the following statements are true according to the Bohr model of the atom?

(a) Electrons are attracted to the atomic nucleus.
(b) Electrons have fixed energy as they circle the nucleus.
(c) Electrons lose energy as they drop to an orbit closer to the nucleus.

Answer: See Appendix G, 4.8.

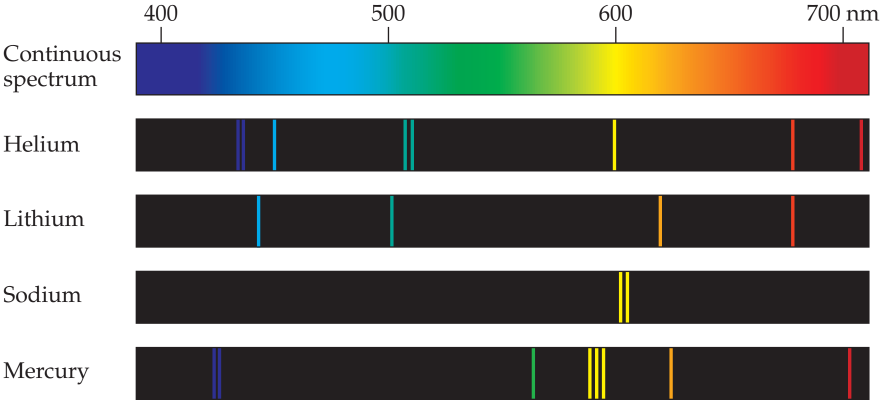
▲ **Mercury Gas Discharge Tube** The emission lines from mercury vapor illuminates the glass tube with a blue glow.

Emission Spectra as "Atomic Fingerprints"

Further study of emission spectra revealed that each element produced a unique set of spectral lines. This observation indicated that the energy levels must be unique for atoms of each element. Therefore, a line spectrum is sometimes referred to as an *atomic fingerprint.*

Atomic fingerprints are useful in the identification of elements. For instance, in 1868, the atomic fingerprint of a new element was observed in the spectrum from the Sun. This element was named helium, after *helios*, the Greek word for "sun." In 1895, an element was discovered in uranium ore with an atomic fingerprint identical to that observed for helium in the Sun's spectrum. Thus, helium was discovered on Earth 27 years after it had first been observed in the solar spectrum. Figure 4.14 compares a continuous spectrum to the emission line spectra of four elements.

▶ **Figure 4.14 A Continuous Spectrum versus Line Spectra** A continuous spectrum is produced from an ordinary light bulb. The emission line spectra are produced by excited atoms of elements in the gaseous state.



A CLOSER LOOK Neon Lights

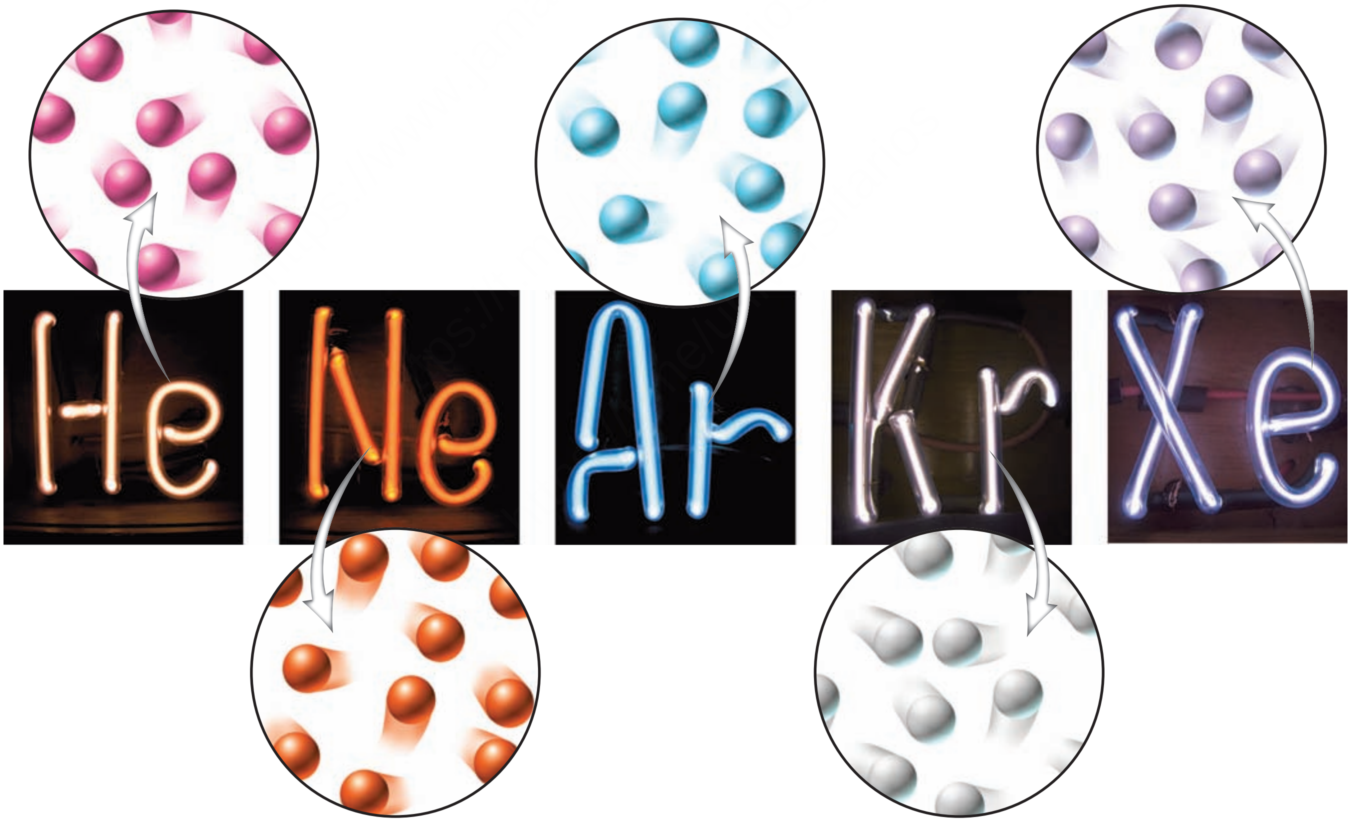
Q: Do all "neon" advertising lights contain neon gas?

At the end of the nineteenth century, J. J. Thomson discovered the electron using a gas discharge tube. Thomson constructed his gas discharge tube out of thin glass, placed a metal electrode in each end, and introduced a small amount of gas. When the gas in the tube was electrically excited, he noticed that the tube glowed. He identified the glowing rays as a stream of small negatively charged particles. These particles acquired the name electrons, and Thomson is given credit for their discovery. The glass tubes used by Thomson were similar to cathode-ray tubes, originally used in television sets and computer screens.

In 1913, Niels Bohr explained that exciting gases with electricity caused electrons to be temporarily promoted to higher-energy states within the atom. However, the excited electrons quickly lose energy by emitting light and return to their original state. The light emitted by different gases varies because the energy levels within atoms vary for each element. For example, a gas discharge tube containing hydrogen gas produces a reddish-purple glow, whereas mercury vapor gives a blue glow.

In 1898, the Scottish chemist William Ramsay discovered the gas neon. Unlike argon, which comprises about 1% of air, neon is much more rare. When neon gas is placed in a narrow glass tube and electrically excited, it produces a reddish-orange light that is very eye-catching. The fact that gas discharge tubes produce an attractive array of colors led naturally to their use as advertising lights. Light from excited neon gas is very intense, and the term neon light has become a generic term for all lighted advertising displays.

Obviously, not all lights used in advertising are the same color. That is, "neon lights" can be red, green, blue, and so on. To produce a given color, a gas discharge tube must be filled with a specific gas. For a pink light, helium gas is used, and for a blue light, argon gas can be used. Only when we wish to produce a reddish-orange light is an advertising display actually filled with neon gas.



▲ **Light Emission** Notice the color of each light varies depending on the gas contained in the glass tube.

A: Not all "neon lights" contain neon gas; only "neon lights" that have a reddish-orange glow contain neon gas.

4.9 Energy Levels and Sublevels

LEARNING OBJECTIVES

- Indicate each sublevel within a given energy level.
- Indicate the number of electrons that can occupy a given sublevel or energy level.

As we learned in the previous section, in 1913 Niels Bohr proposed a model for the atom that pictured electrons circling around the nucleus in fixed-energy levels. His proposal was supported experimentally by the lines in the emission spectrum of hydrogen. The emission spectra of other elements, however, had far too many lines to interpret. Although Bohr could not explain the spectra of other elements, he did suggest the idea of sublevels within a main energy level. The model that eventually emerged had electrons occupying an **energy sublevel** within a main energy level. These energy sublevels were designated *s*, *p*, *d*, and *f*—in reference to the *sharp*, *principal*, *diffuse*, and *fine* lines, respectively, in the emission spectra of the elements.

The number of sublevels in each level corresponds to the number of the main energy level. That is, the first energy level (1) has one sublevel designated 1*s*. The second energy level (2) has two sublevels designated 2*s* and 2*p*. The third energy level (3) has three sublevels designated 3*s*, 3*p*, and 3*d*. The fourth energy level (4) is composed of 4*s*, 4*p*, 4*d*, and 4*f* sublevels (Figure 4.15).

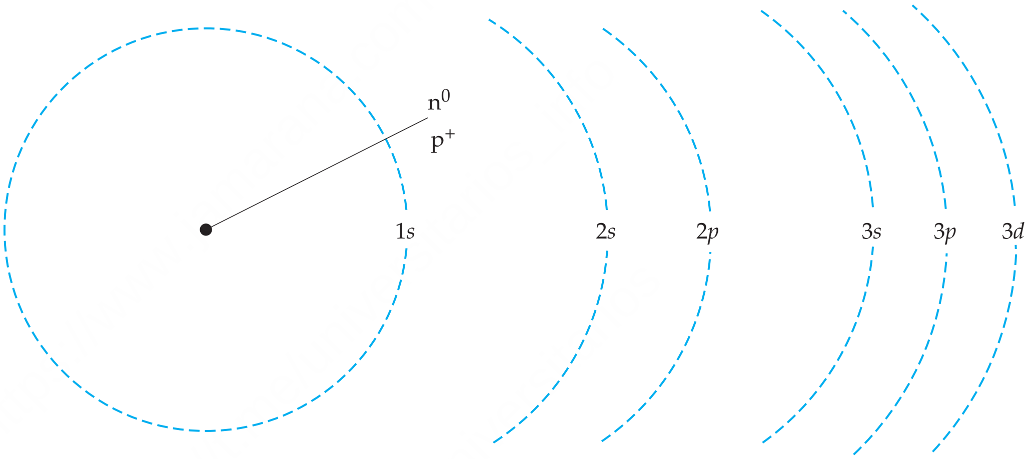


Figure 4.15 A Cross Section of an Atom The first energy level has one sublevel (1*s*). The second energy level has two sublevels (2*s* and 2*p*). The third energy level has three sublevels (3*s*, 3*p*, and 3*d*). Although the diagram suggests electrons travel in circular orbits, this is a simplification and is not actually the case.

The maximum number of electrons in each of the energy sublevels depends on the type of sublevel. That is, an *s* sublevel can hold a maximum of 2 electrons. A *p* sublevel can have a maximum of 6 electrons. A *d* sublevel can have 10 electrons, and an *f* sublevel can hold a maximum of 14 electrons.

To find the maximum number of electrons in a main energy level, we add up the electrons in each sublevel. The first energy level has one *s* sublevel; it can contain only 2 e^-. The second major energy level has two sublevels, 2*s* and 2*p*. The 2*s* can hold 2 e^- and the 2*p* can hold 6 e^-. Thus, the second energy level can hold a maximum of 8 e^-. Example Exercise 4.9 further illustrates how energy levels, sublevels, and number of electrons are related.

EXAMPLE EXERCISE 4.9 Energy Levels, Sublevels, and Electrons

What is the maximum number of electrons that can occupy the third energy level?

Solution

The third energy level is split into three sublevels: 3*s*, 3*p*, and 3*d*. The maximum number of electrons that can occupy each sublevel is as follows:

$$\begin{aligned} s \text{ sublevel} &= 2\,e^- \\ p \text{ sublevel} &= 6\,e^- \\ d \text{ sublevel} &= 10\,e^- \end{aligned}$$

The maximum number of electrons in the third energy level is found by adding the three sublevels together:

$$\begin{aligned} 3s + 3p + 3d &= \text{total electrons} \\ 2\,e^- + 6\,e^- + 10\,e^- &= 18\,e^- \end{aligned}$$

The third energy level can hold a maximum of 18 electrons. Of course, in elements where the third energy level of an atom is not filled, there are fewer than 18 electrons.

Practice Exercise

What is the maximum number of electrons that can occupy the fourth energy level?

Answer: $32\,e^-$ $(2\,e^- + 6\,e^- + 10\,e^- + 14\,e^- = 32\,e^-)$

Concept Exercise

What is the theoretical number of sublevels in the tenth energy level?

Answer: See Appendix G, 4.9.

In summary, electrons are arranged around the nucleus in sublevels that have a specific fixed energy. Electrons farther from the nucleus occupy higher energy levels than those closer to the nucleus. Table 4.3 summarizes how main energy levels, sublevels, and number of electrons are related.

TABLE 4.3 Distribution of Electrons by Energy Level

Energy Level	Energy Sublevel	Electrons/Sublevel	Electrons/Energy Level
1	$1s$	$2\,e^-$	$2\,e^-$
2	$2s$	$2\,e^-$	
	$2p$	$6\,e^-$	$8\,e^-$
3	$3s$	$2\,e^-$	
	$3p$	$6\,e^-$	
	$3d$	$10\,e^-$	$18\,e^-$
4	$4s$	$2\,e^-$	
	$4p$	$6\,e^-$	
	$4d$	$10\,e^-$	
	$4f$	$14\,e^-$	$32\,e^-$

4.10 Electron Configuration

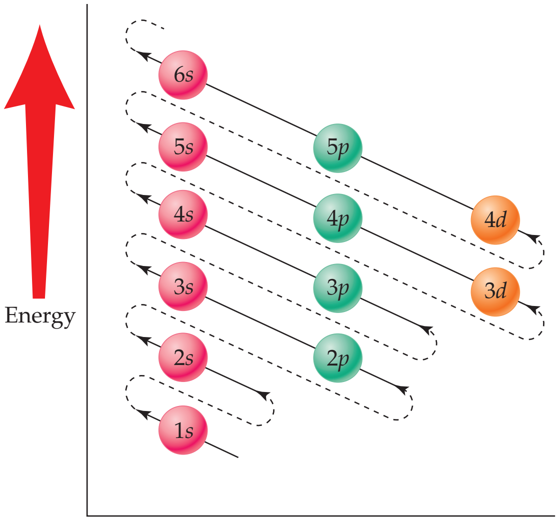
LEARNING OBJECTIVES

- Diagram the order of sublevels according to increasing energy.
- Diagram the predicted electron configurations for selected elements.

Electrons are arranged about the nucleus in a systematic order. The first electrons fill the energy sublevel closest to the nucleus. Additional electrons fill energy sublevels farther and farther from the nucleus. In other words, each energy level is filled sublevel by sublevel. The *s* sublevel is filled before a *p* sublevel, a *p* sublevel is filled before a *d* sublevel, and a *d* sublevel is filled before an *f* sublevel.

In general, sublevels are higher in energy as the energy level increases. Therefore, we would expect the order of sublevel filling to be $1s$, $2s$, $2p$, $3s$, $3p$, $3d$, $4s$, and so on. This is not quite accurate because there are exceptions. For instance, the $4s$ sublevel is lower in energy than the $3d$ sublevel, and the $5s$ is lower than the $4d$ sublevel. A partial list of sublevels in order of increasing energy is $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$.

In Chapter 6 we will learn how to predict the order in which sublevels are filled from the position of the element in the periodic table. In fact, the unusual shape of the periodic table reflects the order of sublevels according to increasing energy. For now, you should memorize the order of sublevel filling or refer to Figure 4.16.



▶ **Figure 4.16 Filling Diagram for Energy Sublevels** The order of sublevel filling is arranged according to increasing energy. Electrons first fill the 1*s* sublevel followed by the 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, and 6*s* sublevels.

EXAMPLE EXERCISE 4.10 Order of Sublevels

According to increasing energy, what is the next energy sublevel after each of the following sublevels?

(a) 3*p* (b) 4*d*

Solution

If you do not know the order of sublevels, refer to the filling diagram in Figure 4.16.

(a) Although the third energy level has 3*s*, 3*p*, and 3*d* sublevels, the 3*d* sublevel does not immediately follow the 3*p*. Instead, the 4*s* sublevel follows the 3*p* and precedes the 3*d*. Thus,

3*s*, 3*p*, 4*s*

(b) Although the fourth energy level has 4*s*, 4*p*, 4*d*, and 4*f* sublevels, the 4*f* sublevel does not immediately follow the 4*d*. Instead, the 5*p* sublevel begins accepting electrons after the 4*d* is filled. Thus,

4*p*, 5*s*, 4*d*, 5*p*

Practice Exercise

Which sublevel gains electrons after each of the following sublevels is filled?

(a) 2*s* (b) 5*p*

Answers:

(a) 2*p* (b) 6*s*

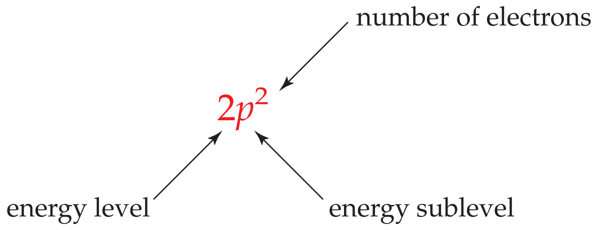
Concept Exercise

The energy difference between sublevels (increases/decreases) moving away from the nucleus.

Answer: See Appendix G, 4.10.

Writing Electron Configurations

The **electron configuration** of an atom is a shorthand statement describing the location of electrons by sublevel. First, the sublevel is written, followed by a superscript indicating the number of electrons. For example, if the 2*p* sublevel contains two electrons, the standard notation is $2p^2$. Thus,



Writing the electron configuration for an atom is a straightforward procedure. First, find the atomic number for the element in the periodic table. The atomic number corresponds to the number of protons in the nucleus, which equals the number of electrons in a neutral atom. Next, write the sublevels according to increasing energy. Each sublevel is filled with electrons in sequence until the total of all the superscripts equals the atomic number of the element.

As an example, let's write the electron configuration for iron. If we refer to the periodic table, we find that the atomic number of iron is 26. Given the atomic number, we know that an iron nucleus must have 26 protons and is surrounded by 26 electrons. The order in which the electron sublevels are arranged according to increasing energy is as follows:

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ \ldots$$

The energy sublevels in iron are filled beginning with the 1*s* sublevel and ending when there is a total of 26 electrons. The electron configuration for iron is as follows:

$$\text{Fe:}\quad 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^6$$

Notice that the sum of the superscripts equals the atomic number of iron (26). Example Exercise 4.11 illustrates the electron configuration for other elements.

EXAMPLE EXERCISE 4.11 Electron Configuration

Write the predicted electron configuration for each of the following elements:

(a) F (b) Sr

Solution

Refer to the periodic table to find the atomic number of an element.

(a) The atomic number of fluorine is 9; therefore, the number of electrons is 9. We can fill sublevels with 9 electrons as follows.

$$\text{F:}\quad 1s^2\,2s^2\,2p^5$$

(b) The atomic number of strontium is 38; therefore, the number of electrons is 38. We can fill sublevels with 38 electrons as follows:

$$\text{Sr:}\quad 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2$$

To check your answer, find the total number of electrons by adding up the superscripts. The total is 38 e^-; this agrees with the atomic number for Sr.

Practice Exercise

Write the predicted electron configuration for each of the following elements:

(a) argon (b) cadmium

Answers:

(a) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$ (b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}$

Concept Exercise

Refer to the periodic table and state whether Cr or Mn has more electrons in the outermost *d* sublevel.

Answer: See Appendix G, 4.11.

4.11 Quantum Mechanical Model of the Atom

In the mid-1920s, a new model of the atom began to emerge. A more powerful theory evolved because the behavior of electrons could not be fully explained by the Bohr model of the atom. The German physicist Werner Heisenberg (1901–1976) concluded that it was not possible to accurately determine both the location and energy of an electron. In his **uncertainty principle**, Heisenberg stated that it is impossible to precisely measure both the location and momentum (or energy) of a small particle

LEARNING OBJECTIVES

- Describe the relative size, shape, and energy of *s* and *p* orbitals.
- Specify the number of electrons that can occupy a given orbital.

simultaneously. In fact, the more accurately the location of an electron in an atom is known, the less precisely its energy can be determined.

In 1932, Heisenberg won the Nobel Prize in physics for his uncertainty principle. Not everyone, however, subscribed to the principle of uncertainty. Some physicists found it unsettling to consider that they might live in a universe ruled by chance. Albert Einstein (1879–1955) was sufficiently troubled by the uncertainty principle that he offered the famous quote: "It seems hard to look into God's cards, but I cannot for a moment believe He plays dice as the current quantum theory alleges He does." Although the uncertainty principle was initially controversial, it was a crucial concept for explaining the new model of the atom.

Gradually, the deeper nature of the atom came into focus. The new model retained the idea of quantized energy levels, but incorporated the concept of uncertainty. The new model that emerged became known as the **quantum mechanical atom**. Recall that in the Bohr model the energy of an electron is defined in terms of a fixed-energy orbit about the nucleus. In the quantum mechanical model, the energy of an electron can be described in terms of the probability of it being within a spatial volume surrounding the nucleus. This region of high probability (~95%) for finding an electron of given energy is called an **orbital**.

Helpful Hint

Bohr Model vs. Quantum Mechanics

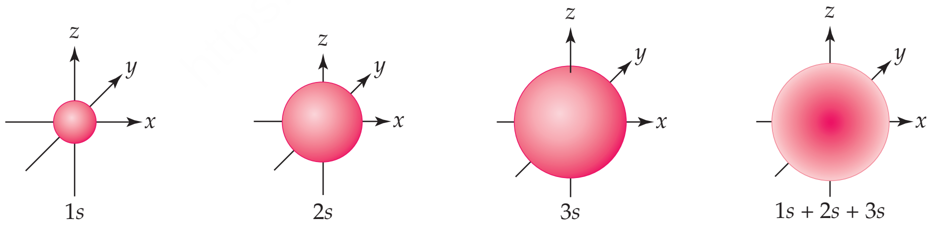
Just as the Thomson model of the atom was simplistic compared to the Bohr model of the atom, the Bohr model is far less sophisticated than the quantum mechanical model of the atom. Studying the Bohr model of the atom, we can visualize electrons circling the atomic nucleus in orbits of fixed energy. Electrons that occupy orbits farther from the nucleus have greater energy than those closer to the nucleus.

The quantum mechanical model of the atom considers electrons to have both a wave and particle nature. Rather than stating the precise orbit that an electron occupies, quantum mechanics states the probability of finding an electron with a given energy within a spatial volume. The term "orbit" applies to the Bohr model, while the term "orbital" is reserved for the quantum mechanical model.

Sizes and Shapes of *s* and *p* Orbitals

In the quantum mechanical atom, orbitals are arranged about the nucleus according to their size and shape. In general, electrons having higher energy are found in larger orbitals. Similar to the energy levels in the Bohr atom, the energy of orbitals is quantized and assigned a whole number value such as 1, 2, 3, 4, ... As the number increases, the energy and size of an orbital also increases.

We can describe the shapes of orbitals by the letters *s*, *p*, *d*, and *f*. For example, the shape of an *s* orbital is that of a sphere, and the shape of a *p* orbital is that of a dumbbell. We can designate the size and shape of an orbital by combining the number that indicates its energy, and the letter that indicates its shape. For example, the designations 1*s*, 2*p*, and 3*d* indicates three orbitals that differ in size, energy, and shape. All *s* orbitals are spherical, but they are not all the same size. A 3*s* orbital is a larger sphere than a 2*s*, and a 2*s* orbital is larger than a 1*s*. That is, the size and energy of the orbital increase as the energy level increases. Figure 4.17 illustrates the relationship between *s* orbitals about the nucleus.

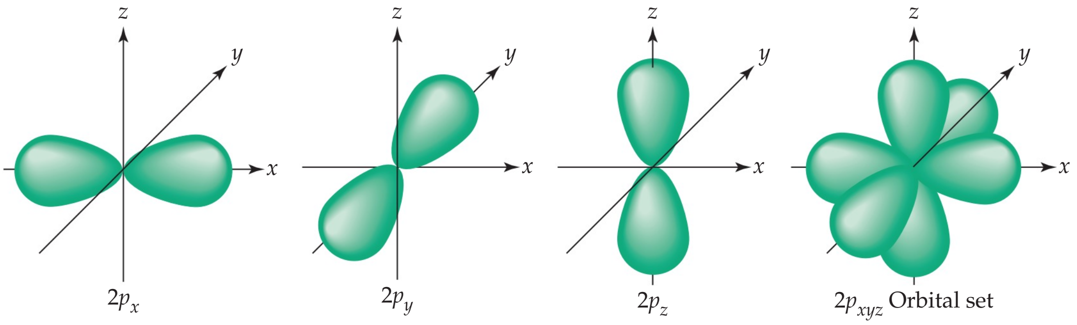


▲ **Figure 4.17** **Relative Sizes of *s* Orbitals** The relative sizes of 1*s*, 2*s*, and 3*s* orbitals are shown. As the main energy level increases, the size and energy of the orbital increases. The nucleus of the atom is located in the center where the three axes intersect. The sketch on the far right illustrates the relationship of the 1*s*, 2*s*, and 3*s* orbitals.

All *p* orbitals have the shape of a dumbbell, but they are not all equal in size or energy. A 3*p* orbital is larger than a 2*p* orbital and is at a higher-energy level. A *p* orbital is said to resemble a dumbbell because it has two lobes. Electrons in a *p* orbital can occupy either of the lobes.

There are three different 2*p* orbitals. Although these three orbitals are identical in size and shape, they differ in their orientation to each other. That is, the three 2*p* orbitals intersect at the nucleus, but they are oriented at right angles to each other.

Figure 4.18 illustrates the relationships between the $2p_x$, $2p_y$, and $2p_z$ orbitals. The p_x orbital is oriented along the x-axis of a three-dimensional system, and the p_y and p_z orbitals are oriented along the y-axis and z-axis, respectively.

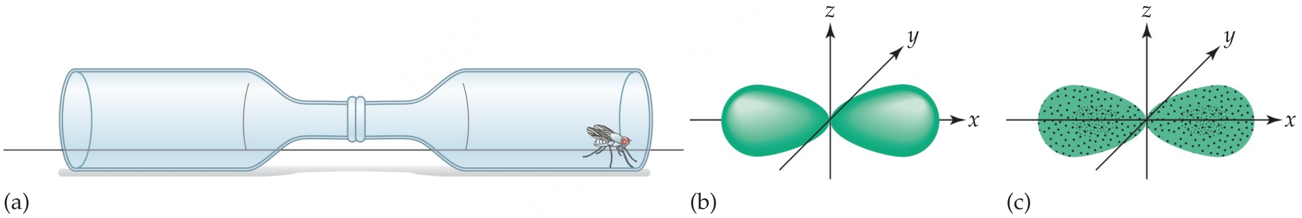


▲ **Figure 4.18 Orientation of 2*p* Orbitals** The size and shape of the three $2p$ orbitals are identical. The $2p$ orbitals do not have a fixed orientation, but rather are perpendicular to each other. An electron in a $2p_x$ orbital has the same probability of occupying the $2p_y$ or $2p_z$ orbital.

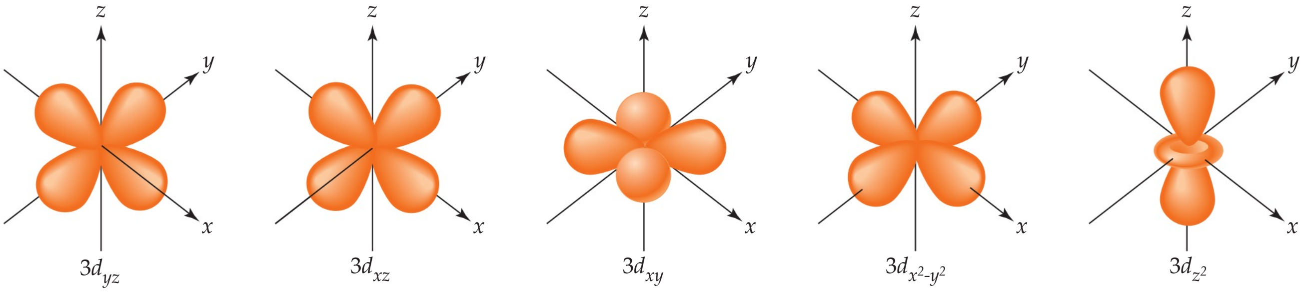
By analogy, try to visualize a flying insect trapped inside two bottles with the open ends held together. The insect is free to fly about the entire volume of the two bottles. In this analogy, the insect represents an electron, and the two bottles represent the two lobes of a *p* orbital. Thus, there is a high probability of finding the electron anywhere within the volume of the entire *p* orbital. Figure 4.19 illustrates this analogy.

So far, we have shown the shapes of *s* and *p* orbitals. However, there are many other orbitals; for example, in the third energy level there are 3*s*, 3*p*, and 3*d* orbitals, and in the fourth energy level there are 4*s*, 4*p*, 4*d*, and 4*f* orbitals.

The 3*s* sublevel contains a single spherical 3*s* orbital. The 3*p* sublevel contains three dumbbell-shaped orbitals: $3p_x$, $3p_y$, and $3p_z$. The 3*p* orbitals are similar in shape, but are larger than 2*p* orbitals. The 3*d* sublevel contains five 3*d* orbitals. Although it is not necessary to know the shapes of *d* orbitals, the five different 3*d* orbitals are shown in Figure 4.20.



▲ **Figure 4.19 Analogy for a *p* Orbital** (a) Notice the insect trapped within the bottles held end to end. The insect can be in the left bottle, or the right bottle. (b) Similarly, an electron has a probability of being found anywhere within the two lobes of the p orbital. (c) The concentration of black dots indicates the most probable location of the electrons.



▲ **Figure 4.20 Shapes of 3*d* Orbitals** The shapes of the five 3*d* orbitals are not the same. Although four of the 3*d* orbitals are similar in shape, they have a different orientation to each other.

EXAMPLE EXERCISE 4.12 Atomic Orbitals

Describe the relative size, energy, and shape for each of the following orbitals:

(a) 4*s* versus 3*s* and 5*s*

(b) 4*p* versus 3*p* and 5*p*

Solution

The size and energy of an orbital is indicated by the number; the shape of the orbital is designated by the letter.

(a) Size and energy are greater for a 4*s* orbital than for a 3*s* orbital, but less than for a 5*s* orbital. The shape of a 4*s* orbital—and all *s* orbitals—is similar to the shape of a sphere.

(b) Size and energy are greater for a 4*p* orbital than for a 3*p* orbital, but less than for a 5*p* orbital. The shape of a 4*p* orbital—and all *p* orbitals—is similar to the shape of a dumbbell.

Practice Exercise

Select the orbital in each of the following pairs that fits the description:

(a) the higher-energy orbital: 3*p* or 4*p*

(b) the larger size orbital: 4*d* or 5*d*

Answers:

(a) 4*p*

(b) 5*d*

Concept Exercise

Which of the following statements are true according to the quantum mechanical model of the atom?

(a) Orbitals represent quantum energy levels for electrons.

(b) Orbitals represent probability boundaries for electrons.

(c) Orbitals can have different shapes.

Answer: See Appendix G, 4.12.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
4.1 Dalton Model of the Atom In 1803, John Dalton proposed that matter consisted of atoms based on the laws of conservation of mass and definite proportion. Citing experiments whereby two elements combine in more than one way, he formulated the law of multiple proportions, that is, CO and CO_2.	• Describe the Dalton model of the atom. • *Related Exercises: 1–4*
4.2 Thomson Model of the Atom Toward the end of the 1800s, there was evidence that the atom was divisible. When electricity was applied to a gas in a sealed glass tube, negative and positive rays were observed in the tube. Scientists discovered the rays from hydrogen gas were composed of tiny, charged particles they named **electrons** and **protons**.	• Describe the Thomson raisin pudding model of the atom. • *Related Exercises: 5–6* • Specify the relative charge on an electron and a proton. *Related Exercises: 7–10*
4.3 Rutherford Model of the Atom In 1911, Ernest Rutherford supervised a classic experiment in which alpha particles were fired at a thin sheet of gold foil. Much to his astonishment, some of the alpha particles bounced backward. He interpreted the results as evidence for a tiny, dense **atomic nucleus** at the center of the atom. The nucleus contains positively charged protons surrounded by negatively charged electrons. Twenty years later, the nucleus was also found to contain a neutral particle called a **neutron**.	• Describe the Rutherford nuclear model of the atom. • *Related Exercises: 11–12* • Specify the relative charge and approximate mass of an electron, proton, and neutron. *Related Exercises: 13–16*
4.4 Atomic Notation Chemists use a symbolic shorthand called **atomic notation** to designate the composition of a nucleus. Atoms of an element always contain the same number of protons, but the number of neutrons in the nucleus can vary. The number of protons is called the **atomic number** (*Z*), and the sum of the protons and neutrons is called the **mass number** (*A*). Atoms with the same atomic number, but a different mass number are called **isotopes**.	• Analyze atomic notation and indicate protons, neutrons, and electrons. • *Related Exercises: 17–18* • Analyze a given isotope and indicate the number of neutrons. *Related Exercises: 19–24*

Key Concepts	Learning Objectives and Related Exercises
4.5 Atomic Mass Although the mass of an atom is much too small to measure directly, we can determine its relative mass. Carbon-12 is used as the reference isotope and is assigned a mass of exactly 12 **atomic mass units (amu)**. The masses of all other atoms are related to the mass of carbon-12. To find a representative value for the atomic mass of an element, we average the mass of each isotope. The weighted average mass of all isotopes is termed the **atomic mass** of the element.	• Explain the concept of relative atomic mass. ***Related Exercises: 25–32*** • Calculate the atomic mass of an element given the mass and abundance of the naturally occurring isotopes. ***Related Exercises: 33–36***
4.6 The Wave Nature of Light **Light** travels through space as a wave of radiant energy. The crest-to-crest distance between waves is the **wavelength**, and the number of cycles completed in a second is the **frequency**. As the wavelength decreases, the frequency and energy of light increase. The **visible spectrum** extends from 400–700 nm, but the entire **radiant energy spectrum** also includes gamma rays, X rays, and microwaves. Thus, radiant energy is a **continuous spectrum** of visible and invisible light.	• Analyze the relationship of wavelength, frequency, and energy of light. ***Related Exercises: 37–44***
4.7 The Quantum Concept In 1900, Max Planck introduced the quantum concept. Planck stated that the energy radiated by an object is not continuous, but rather that the radiation is emitted in small bundles. When an object radiates light, it releases a unit of radiant energy called a **photon**.	• Illustrate the quantum concept applied to matter and energy. ***Related Exercises: 45–50***
4.8 Bohr Model of the Atom In 1913, Niels Bohr suggested that electrons travel in circular orbits about the nucleus. The electron possesses a specific energy and is said to occupy an **energy level**. If an electron changes orbits in the **Bohr atom**, there is a quantum energy change. Bohr argued that an **emission line spectrum** results from electrons dropping from higher energy levels to lower levels. When an electron drops, a photon of light is released having an energy that corresponds to the difference in energy between the two levels.	• Describe the Bohr model of the atom. ***Related Exercises: 51–52*** • Explain the relationship between energy levels in an atom and lines in an emission spectrum. ***Related Exercises: 53–66***
4.9 Energy Levels and Sublevels A closer examination of emission line spectra from gases revealed **energy sublevels** within main levels. The number of sublevels corresponds to the number of the energy level. For example, the fourth energy level has four sublevels (*s*, *p*, *d*, *f*). An *s* sublevel can hold 2 electrons, a *p* sublevel 6 electrons, a *d* sublevel 10 electrons, and an *f* sublevel a maximum of 14 electrons.	• Indicate each sublevel within a given energy level. ***Related Exercises: 67–70*** • Indicate the number of electrons that can occupy a given sublevel or energy level. ***Related Exercises: 71–72***
4.10 Electron Configuration Electrons fill sublevels in order of increasing energy as follows: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$. Notice that the $4s$ sublevel fills before the $3d$ and the $5s$ sublevel before the $4d$. A description of sublevel filling for an element is given by the **electron configuration**. A superscript following each sublevel indicates the number of electrons in a sublevel; for instance, $1s^22s^22p^63s^1$ is the electron configuration for sodium (atomic number 11).	• Diagram the order of sublevels according to increasing energy. ***Related Exercises: 73–74*** • Diagram the predicted electron configurations for selected elements. ***Related Exercises: 75–78***
4.11 Quantum Mechanical Model In the 1920s our understanding of electrons in atoms became very sophisticated. In 1925 Werner Heisenberg suggested the **uncertainty principle**; that is, it is impossible to simultaneously know both the precise location and the energy of an electron. Instead, the energy of an electron can be known only in terms of its probability of being located somewhere within the atom. This description gave rise to the **quantum mechanical atom**. A location within the atom where there is a high probability of finding an electron having a certain energy is called an **orbital**. An orbital is a region about the nucleus having a given energy, size, shape, and orientation. The shape of an *s* orbital is spherical, and a *p* orbital resembles the shape of a dumbbell.	• Describe the relative size, shape, and energy of *s* and *p* orbitals. ***Related Exercises: 79–84*** • Specify the number of electrons that can occupy a given orbital. ***Related Exercises: 85–86***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** a negatively charged subatomic particle having a tiny mass
_____ **2.** a positively charged subatomic particle having an approximate mass of 1 amu
_____ **3.** a neutral subatomic particle having an approximate mass of 1 amu
_____ **4.** a region in the center of an atom containing protons and neutrons
_____ **5.** a value indicating the number of protons in the nucleus of an atom
_____ **6.** a value indicating the number of protons and neutrons in the nucleus of an atom
_____ **7.** a symbolic method for expressing the composition of an atomic nucleus
_____ **8.** atoms of the same element that have a different number of neutrons
_____ **9.** a unit of mass exactly equal to 1/12 the mass of a C-12 atom
_____ **10.** the average mass of all the naturally occurring isotopes of an element
_____ **11.** the distance a light wave travels to complete one cycle
_____ **12.** the number of times a light wave completes a cycle in one second
_____ **13.** a general term that can refer to either visible or invisible radiant energy
_____ **14.** a range of light energy extending from violet to red, that is, approximately 400–700 nm
_____ **15.** a range of light energy extending from gamma rays to microwaves
_____ **16.** a band of light energy that is uninterrupted
_____ **17.** a particle of radiant energy
_____ **18.** a model of the atom that describes electrons circling the nucleus in orbits
_____ **19.** an orbit of specific energy that electrons occupy as they circle the nucleus
_____ **20.** a collection of narrow bands of light produced by atoms of a given element releasing energy
_____ **21.** an electron energy level that results from splitting a main energy level
_____ **22.** a shorthand description of the arrangement of electrons by sublevels according to increasing energy
_____ **23.** the statement that it is impossible to precisely measure the location and energy of a particle at the same time
_____ **24.** a sophisticated model of the atom that describes the energy of an electron in terms of its probability of being found in a particular location about the nucleus
_____ **25.** a region about the nucleus in which there is a high probability of finding an electron with a given energy

(a) atomic mass (*Sec. 4.5*)
(b) atomic mass unit (amu) (*Sec. 4.5*)
(c) atomic notation (*Sec. 4.4*)
(d) atomic nucleus (*Sec. 4.3*)
(e) atomic number (*Z*) (*Sec. 4.4*)
(f) Bohr atom (*Sec. 4.8*)
(g) continuous spectrum (*Sec. 4.6*)
(h) electron (e^-) (*Sec. 4.2*)
(i) electron configuration (*Sec. 4.10*)
(j) emission line spectrum (*Sec. 4.8*)
(k) energy level (*Sec. 4.8*)
(l) energy sublevel (*Sec. 4.9*)
(m) frequency (*Sec. 4.6*)
(n) isotope (*Sec. 4.4*)
(o) light (*Sec. 4.6*)
(p) mass number (*A*) (*Sec. 4.4*)
(q) neutron (n^0) (*Sec. 4.3*)
(r) orbital (*Sec. 4.11*)
(s) photon (*Sec. 4.7*)
(t) proton (p^+) (*Sec. 4.2*)
(u) quantum mechanical atom (*Sec. 4.11*)
(v) radiant energy spectrum (*Sec. 4.6*)
(w) uncertainty principle (*Sec. 4.11*)
(x) visible spectrum (*Sec. 4.6*)
(y) wavelength (*Sec. 4.6*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Dalton Model of the Atom (Sec. 4.1)

1. Which Greek philosopher suggested the concept of an atom in 486 B.C.?
2. Which Greek philosopher argued that matter is continuous in 350 B.C.?
3. State three scientists whose work Dalton used to support the atomic theory.
4. State two experimental laws Dalton used to support the atomic theory.
5. Which of the following of Dalton's proposals proved to be invalid?
   (a) Atoms of different elements combine to form compounds.
   (b) Atoms can combine in small whole number ratios.
   (c) Atoms are indivisible.
6. Which of the following of Dalton's proposals proved to be invalid?
   (a) Atoms are indestructible.
   (b) Atoms of the same element are identical.
   (c) Atoms can combine in more than one whole number ratio.

Thomson Model of the Atom (Sec. 4.2)

7. What do the raisins represent in the raisin pudding model of the atom?
8. What does the pudding represent in the raisin pudding model of the atom?
9. What is the simplest negative particle in an atom?
10. What is the simplest positive particle in an atom?
11. What is the relative charge on an electron?
12. What is the relative charge on a proton?

Rutherford Model of the Atom (Sec. 4.3)

13. Where are protons and neutrons located in an atom?
14. Where are electrons located in an atom?
15. What is the approximate size (cm) of an atom?
16. What is the approximate size (cm) of a nucleus?
17. What is the relative mass of an electron and proton?
18. What is the relative charge of an electron and proton?

Atomic Notation (Sec. 4.4)

19. State the number of neutrons in an atom of each of the following isotopes.
    - **(a)** $^{1}_{1}\text{H}$
    - **(b)** $^{4}_{2}\text{He}$
    - **(c)** $^{16}_{8}\text{O}$
    - **(d)** $^{20}_{10}\text{Ne}$
20. State the number of neutrons in an atom of each of the following isotopes.
    - **(a)** $^{2}_{1}\text{H}$
    - **(b)** $^{3}_{2}\text{He}$
    - **(c)** $^{18}_{8}\text{O}$
    - **(d)** $^{22}_{10}\text{Ne}$
21. State the number of neutrons in an atom of each of the following isotopes.
    - **(a)** hydrogen-1
    - **(b)** carbon-13
    - **(c)** cobalt-59
    - **(d)** iodine-127
22. State the number of neutrons in an atom of each of the following isotopes.
    - **(a)** hydrogen-2
    - **(b)** carbon-14
    - **(c)** cobalt-60
    - **(d)** iodine-131
23. Complete the following table and provide the missing information.

Atomic Notation	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
$^{11}_{5}\text{B}$					
$^{15}_{7}\text{N}$					
$^{40}_{20}\text{Ca}$					
$^{200}_{80}\text{Hg}$					

24. Complete the following table and provide the missing information.

Atomic Notation	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
	9			10	
		31	15		
	28			31	
		210	82		

25. Draw a diagram of the arrangement of protons, neutrons, and electrons in an atom of each of the following isotopes.
    - **(a)** $^{7}_{3}\text{Li}$
    - **(b)** $^{13}_{6}\text{C}$
    - **(c)** $^{16}_{8}\text{O}$
    - **(d)** $^{20}_{10}\text{Ne}$
26. Draw a diagram of the arrangement of protons, neutrons, and electrons in an atom of each of the following isotopes.
    - **(a)** $^{31}_{15}\text{P}$
    - **(b)** $^{35}_{17}\text{Cl}$
    - **(c)** $^{40}_{18}\text{Ar}$
    - **(d)** $^{131}_{53}\text{I}$

Atomic Mass (Sec. 4.5)

27. Distinguish between atomic number and mass number.
28. Distinguish between atomic mass and isotopic mass.
29. What is the assigned mass for the current reference isotope?
30. What is the current reference isotope for the atomic mass scale?
31. Given the only naturally occurring isotope of sodium is Na-23, determine its mass from the periodic table.
32. Given the only naturally occurring isotope of aluminum is Al-27, determine its mass from the periodic table.
33. Given the only naturally occurring isotope of fluorine is F-19, determine its mass from the periodic table.
34. Given the only naturally occurring isotope of phosphorus is P-31, determine its mass from the periodic table.
35. Calculate the atomic mass for lithium given the following data for its natural isotopes.

^{6}Li:	6.015 amu	7.42%
^{7}Li:	7.016 amu	92.58%

36. Calculate the atomic mass for magnesium given the following data for its natural isotopes.

^{24}Mg:	23.985 amu	78.70%
^{25}Mg:	24.986 amu	10.13%
^{26}Mg:	25.983 amu	11.17%

37. Calculate the atomic mass for iron given the following data for its natural isotopes.

^{54}Fe:	53.940 amu	5.82%
^{56}Fe:	55.935 amu	91.66%
^{57}Fe:	56.935 amu	2.19%
^{58}Fe:	57.933 amu	0.33%

38. Calculate the atomic mass for zinc given the following data for its natural isotopes.

^{64}Zn	63.929 amu	48.89%
^{66}Zn	65.926 amu	27.81%
^{67}Zn	66.927 amu	4.11%
^{68}Zn	67.925 amu	18.57%
^{70}Zn	69.925 amu	0.62%

The Wave Nature of Light (Sec. 4.6)

39. Which has the longer wavelength: red light or yellow light?
40. Which has the shorter wavelength: blue light or green light?
41. Which has the higher frequency: red light or yellow light?
42. Which has the lower frequency: blue light or green light?
43. Which has the higher energy: red light or yellow light?
44. Which has the lower energy: blue light or green light?
45. Which wavelength has the higher energy: 400 nm or 450 nm?
46. Which wavelength has the lower energy: 1100 nm or 1250 nm?

The Quantum Concept (Sec. 4.7)

47. What particle represents the quantum nature of light energy?
48. What particle represents the quantum nature of electrical energy?
49. State whether each of the following is continuous or quantized.
    - **(a)** a rainbow
    - **(b)** a line spectrum
50. State whether each of the following is continuous or quantized.
    - **(a)** a spiral staircase
    - **(b)** an elevated ramp

51. State whether each of the following instruments gives a continuous or a quantized measurement of length.
    (a) a metric ruler **(b)** a digital laser

52. State whether each of the following instruments gives a continuous or a quantized measurement of volume.
    (a) 10-mL volumetric pipet
    (b) 10-mL graduated cylinder

Bohr Model of the Atom (Sec. 4.8)

53. Which of the following energy level changes for an electron is most energetic: 5 → 2, 4 → 2, or 3 → 2?

54. Which of the following energy level changes for an electron is least energetic: 5 → 2, 4 → 2, or 3 → 2?

55. Which of the following energy level changes for an electron is most energetic: 4 → 1, 3 → 1, or 2 → 1?

56. Which of the following energy level changes for an electron is least energetic: 4 → 1, 3 → 1, or 2 → 1?

◀ **Hydrogen Gas Discharge Tube** The emission lines from hydrogen gas illuminate the glass tube with a reddish-purple glow.

57. In a hydrogen atom, what color is the emission line when electrons drop from the third to the second energy level?

58. In a hydrogen atom, what color is the emission line observed when electrons drop from the fourth to the second energy level?

59. In the emission spectrum from hydrogen gas, what type of energy is released when electrons drop from the fourth to the first energy level?

60. In the emission spectrum from hydrogen gas, what type of energy is released when electrons drop from the fourth to the third energy level?

61. Which of the following lines in the emission spectrum of hydrogen has the least energy: red, blue-green, or violet?

62. Which of the following lines in the emission spectrum of hydrogen has the shortest wavelength: red, blue-green, or violet?

63. How many photons of light are emitted for each of the following?
    (a) 1 e^- drops from energy level 3 to 1.
    (b) 1 e^- drops from energy level 3 to 2.

64. How many photons of light are emitted for each of the following?
    (a) 100 e^- drop from energy level 3 to 2.
    (b) 100 e^- drop from energy level 4 to 2.

65. What is the color of the spectral line emitted for each of the following electron energy changes in excited hydrogen gas?
    (a) Electrons drop from energy level 2 to 1.
    (b) Electrons drop from energy level 3 to 2.
    (c) Electrons drop from energy level 4 to 3.

66. What is the color of the spectral line emitted for each of the following electron energy changes in excited hydrogen gas?
    (a) Electrons drop from energy level 5 to 1.
    (b) Electrons drop from energy level 5 to 2.
    (c) Electrons drop from energy level 5 to 4.

Energy Levels and Sublevels (Sec. 4.9)

67. What experimental evidence suggests the concept of electrons in energy levels?

68. What experimental evidence suggests main energy levels split into sublevels?

69. State the number of sublevels in each of the following main energy levels.
    (a) 1st **(b)** 2nd
    (c) 3rd **(d)** 4th

70. Designate all the sublevels within each of the following main energy levels.
    (a) 1st **(b)** 2nd
    (c) 3rd **(d)** 4th

71. What is the maximum number of electrons in each of the following sublevels?
    (a) 2*s* **(b)** 4*p*
    (c) 3*d* **(d)** 5*f*

72. What is the maximum number of electrons in each of the following?
    (a) an *s* sublevel **(b)** a *p* sublevel
    (c) a *d* sublevel **(d)** an *f* sublevel

73. What is the maximum number of electrons in the second energy level?

74. What is the maximum number of electrons in the fourth energy level?

Electron Configuration (Sec. 4.10)

75. Draw a filling diagram and predict the sublevel that follows 4*s*.

76. Draw a filling diagram and predict the sublevel that follows 5*s*.

77. Write the predicted electron configuration for each of the following elements.
    (a) He **(b)** Be
    (c) Co **(d)** Cd

78. Write the predicted electron configuration for each of the following elements.
    (a) boron **(b)** argon
    (c) manganese **(d)** nickel

79. Which element corresponds to each of the following electron configurations?
    (a) $1s^2\ 2s^1$
    (b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
    (c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$
    (d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$

80. Which element corresponds to each of the following electron configurations?
(a) $1s^2\,2s^2\,2p^5$
(b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$
(c) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^5$
(d) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^5$

Quantum Mechanical Model of the Atom (Sec. 4.11)

81. Sketch a three-dimensional representation for each of the following orbitals. Label the *x*-axis, *y*-axis, and *z*-axis.
(a) $1s$ **(b)** $2p_x$
(c) $3p_y$ **(d)** $4p_z$

82. Sketch a three-dimensional representation for each of the following orbital sets. Label the *x*-axis, *y*-axis, and *z*-axis.
(a) $1s$, $2s$, $2p_x$ **(b)** $3p_x$, $3p_y$, $3p_z$

83. Which orbital in each of the following pairs has the higher energy?
(a) $2s$ or $3s$ **(b)** $2p_x$ or $3p_x$
(c) $2p_x$ or $2p_y$ **(d)** $4p_y$ or $4p_z$

84. Which orbital in each of the following pairs has the larger size?
(a) $2s$ or $3s$ **(b)** $2p_x$ or $3p_x$
(c) $2p_x$ or $2p_y$ **(d)** $4p_y$ or $4p_z$

85. Designate the orbital that fits each of the following descriptions.
(a) a spherical orbital in the fifth energy level
(b) a dumbbell-shaped orbital in the fourth energy level

86. Designate the orbital that fits each of the following descriptions.
(a) a spherical orbital in the sixth energy level
(b) a dumbbell-shaped orbital in the third energy level

87. State the maximum number of electrons that can occupy each of the following orbitals.
(a) $1s$ **(b)** $2p$
(c) $3d$ **(d)** $4f$

88. State the maximum number of electrons that can occupy each of the following sublevels.
(a) $1s$ **(b)** $2p$
(c) $3d$ **(d)** $4f$

General Exercises

89. Silver occurs naturally as ^{107}Ag and ^{109}Ag. Given the mass and abundance of silver-109 (108.91 amu and 48.16%), what is the isotopic mass of silver-107?

90. Gallium occurs naturally as ^{69}Ga and ^{71}Ga. Given the mass and abundance of gallium-69 (68.92 amu and 60.10%), what is the isotopic mass of gallium-71?

91. Element 43 is used in medical radiology to locate tumors. Refer to the periodic table and state whether Tc has any stable isotopes.

92. Element 61 was named for the mythological Greek Prometheus who stole fire from the gods. Refer to the periodic table and state whether Pm has any stable isotopes.

93. Indicate the region of the spectrum for each of the following wavelengths of light.
(a) 300 nm **(b)** 1000 nm

94. Indicate the region of the spectrum for each of the following wavelengths of light.
(a) 280 nm **(b)** 950 nm

Challenge Exercises

95. Silver has only two natural isotopes, and they occur in about equal abundance. If ^{109}Ag is one of the isotopes, what is the other isotope given the atomic mass of silver is approximately 108 amu?

96. Bromine has only two natural isotopes, and they occur in about equal abundance. If ^{79}Br is one of the isotopes, what is the other isotope given the atomic mass of bromine is approximately 80 amu?

97. Explain why the electron configuration for silver is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1\,4d^{10}$ rather than the predicted $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^9$.

98. Explain why the electron configuration for copper is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^{10}$ rather than the predicted $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^9$.

Chapter 4 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following of Dalton's proposals is still valid? (Sec. 4.1)
(a) An element is composed of tiny particles called atoms.
(b) Atoms of different elements combine to form compounds.
(c) Compounds contain atoms in small whole number ratios.
(d) all of the above
(e) none of the above

2. What do the raisins represent in the plum pudding model of the atom? (Sec. 4.2)
(a) electrons **(b)** neutrons
(c) protons **(d)** nuclei
(e) none of the above

3. Which subatomic particle has a relative charge of +1 and a mass of 1 amu? (Sec. 4.3)
(a) alpha **(b)** electron
(c) neutron **(d)** proton
(e) none of the above

4. How many neutrons are in the nucleus of an atom of cobalt-60? (Sec. 4.4)
(a) 27 **(b)** 33
(c) 60 **(d)** 87
(e) none of the above

5. Element X has two natural isotopes: X-63 (62.940 amu) and X-65 (64.928 amu). Calculate the atomic mass of element X given the abundance of X-63 is 69.17%. (Sec. 4.5)
(a) 63.55 amu **(b)** 64.00 amu
(c) 64.32 amu **(d)** 107.85 amu
(e) 108.46 amu

6. Which of the following colors of light has the highest energy? (Sec. 4.6)
(a) blue **(b)** green
(c) red **(d)** violet
(e) yellow

7. Which of the following is an example of a quantized change in energy? (Sec. 4.7)
   (a) a ball rolling across a playground
   (b) a ball rolling down a ramp
   (c) a ball rolling down a flight of stairs
   (d) all of the above
   (e) none of the above
8. How many photons of light are emitted when the electron in a hydrogen atom drops from energy level 4 to 1? (Sec. 4.8)
   (a) 1 **(b)** 2
   (c) 3 **(d)** 4
   (e) none of the above
9. What is the maximum number of electrons that can occupy the 3*d* energy sublevel? (Sec. 4.9)
   (a) 2 **(b)** 6
   (c) 8 **(d)** 14
   (e) none of the above
10. What is the electron configuration for an atom of manganese? (Sec. 4.10)
    (a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$
    (b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$
    (c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 4p^5$
    (d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 4d^5$
    (e) none of the above
11. Which of the following is true for both an orbit and an orbital? (Sec. 4.11)
    (a) An orbit and an orbital always contain two electrons.
    (b) An orbit and an orbital can be described as elliptical.
    (c) An orbit and an orbital represent an electron boundary.
    (d) An orbit and an orbital represent a fixed energy state.
    (e) all of the above

Key Concepts

12. An atomic nucleus can be described by the analogy "like a marble in the Dallas Cowboys Stadium." If a marble represents the atomic nucleus, what does the stadium represent?

▲ **Dallas Cowboys Stadium** By analogy, imagine a marble on the 50-yard line compared to the size of the entire Dallas Cowboys Stadium.

13. The scattering of alpha particles by a thin gold foil can be described by the analogy "like a missile shot through our solar system." If the missile represents an alpha particle, what do the planets in our solar system represent?
14. Complete the following quantum analogy: an *electron* is to electricity, as a _____ is to light.
15. What is the experimental evidence for electrons occupying fixed energy states surrounding the nucleus?

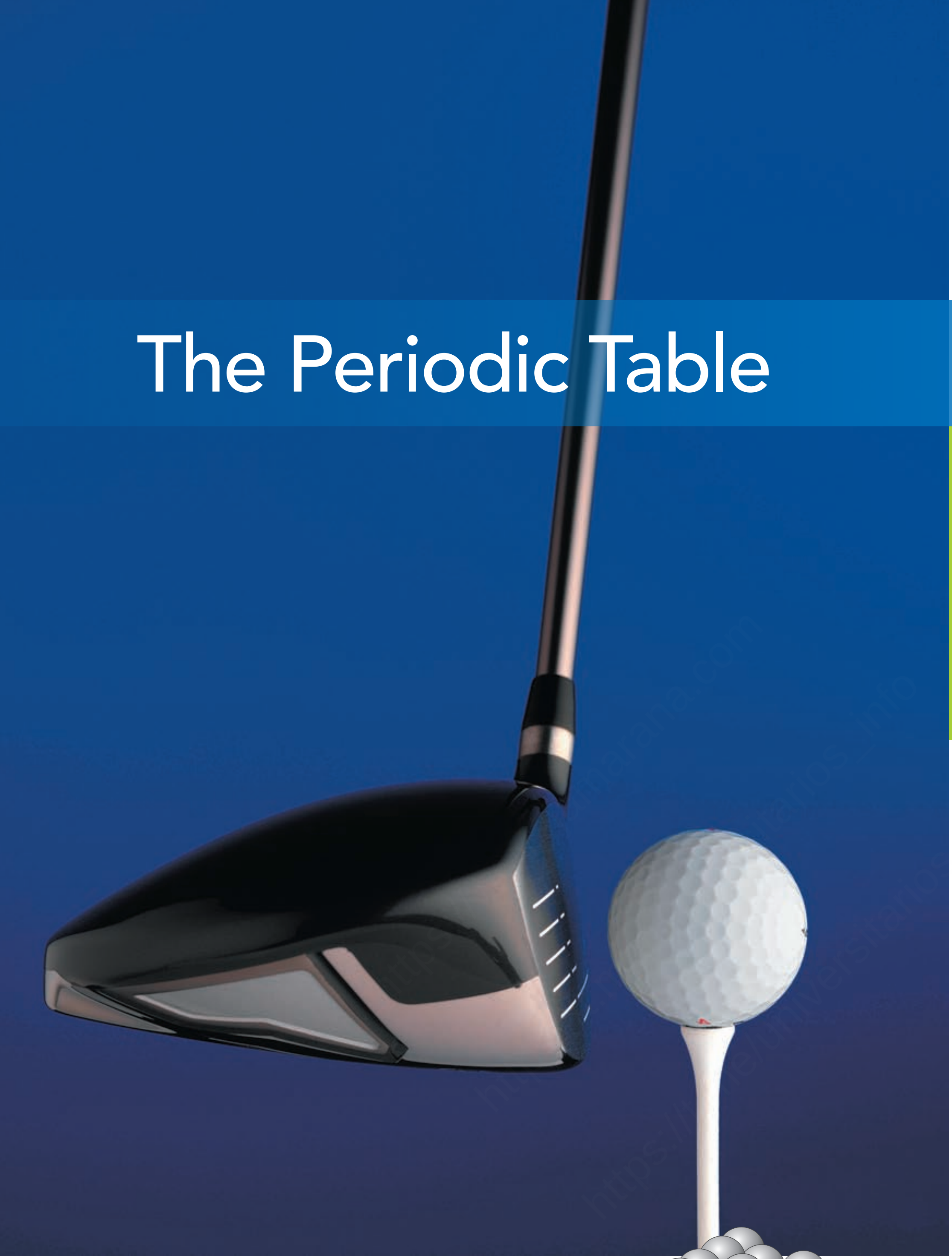
◀ **Neon Gas Discharge Tube** When the reddish-orange light from neon gas passes through a prism, a line spectrum is observed, which is the "atomic fingerprint" for neon.

Critical Thinking

16. Copper has two natural isotopes: Cu-63 and Cu-65. Which isotope is more abundant? (*Hint:* Refer to the periodic table.)
17. Can atoms of different elements have the same number of protons?
18. Can atoms of different elements have the same number of neutrons?

CHAPTER 5

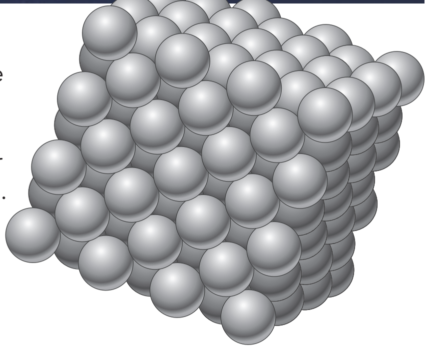
The Periodic Table



> *"It is the function of science to discover the existence of a general reign of order in nature and to find the causes governing this order."*
>
> **Dmitri Mendeleev, Russian Chemist (1834–1907)**

5
B
Boron

Element 5: Boron
Boron is a semimetal element that can be drawn into high-strength fibers for the manufacture of golf clubs and fishing rods. Boron occurs naturally in borax, which is a white powder used in detergents, cosmetics, and fire retardants.



Copper, silver, and gold have been used since prehistoric times. However, not until the 1700s were these metals defined as elements by Robert Boyle (1627–1691). In 1800, 30 elements had been isolated and identified, and by 1870, there were more than 60 known elements.

In 1829, the German chemist J. W. Döbereiner (1780–1849) observed that several elements could be classified into groups of three, or *triads*. All three elements in a triad showed remarkable similarity in their chemical properties. Each triad also revealed an orderly trend in physical properties such as density, melting point, and especially atomic mass.

In 1865, the English chemist J. A. R. Newlands (1837–1898) suggested that the 62 known elements be arranged into groups of seven according to increasing atomic mass. He proposed that an eighth element would then repeat the properties of the first element in the previous group. Interestingly, his theory, called the *law of octaves*, was received with ridicule and was not accepted for publication. Although Newlands's insight into the periodic relationships of the elements was essentially correct, it took 20 years for him to receive professional recognition. In 1887, Newlands graciously accepted the Davy Medal from the Royal Society of Great Britain.

LEARNING OBJECTIVE

- Describe the original periodic law proposed by Mendeleev.

5.1 Classification of Elements

In the 1860s, the Russian chemist Dmitri Mendeleev proposed that properties of different elements repeat at regular intervals when they are arranged in order of *increasing atomic mass*. Today, Mendeleev is regarded as the architect of the modern periodic table, which depicts a systematic arrangement of the elements.

Mendeleev's achievement was the result of his patient and systematic study of the physical and chemical properties of the elements. Furthermore, he had the insight and courage to predict the existence and properties of three elements before their discovery. In 1869, Mendeleev predicted the existence of an element he called ekasilicon. In 1886, ekasilicon was discovered in Germany and given the name germanium. The properties of ekasilicon predicted by Mendeleev and the observed properties of germanium are compared in Table 5.1.

TABLE 5.1 Mendeleev's Predictions of the Properties for Ekasilicon (Ek)

Property	Ekasilicon Predicted (1869)	Germanium Discovered (1886)
color	gray	gray
atomic mass	72 amu	72.6 amu
density	5.5 g/mL	5.32 g/mL
melting point	very high	937 °C
formula of oxide	EkO_2	GeO_2
density of oxide	4.7 g/mL	4.70 g/mL
formula of chloride	$EkCl_4$	$GeCl_4$
boiling point of chloride	100 °C	86 °C

In 1869, Mendeleev published his original periodic table with the elements arranged in vertical columns. Two years later, he created an updated version showing the elements in horizontal rows. He arranged the elements by increasing atomic mass and began a new row with an element that repeated the properties of a previous element. Figure 5.1 illustrates a portion of Mendeleev's 1871 periodic table of elements.

Group	**I**	**II**	**III**	**IV**	**V**	**VI**	**VII**	**VIII**
Formula of Oxide	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7	RO_4
	H							
	Li	Be	B	C	N	O	F	
	Na	Mg	Al	Si	P	S	Cl	
	K	Ca	**eka-**	Ti	V	Cr	Mn	Fe, Co
	Cu	Zn	**eka-**	**eka-**	As	Se	Br	& Ni
	Rb	Sr	Y	Zr	Nb	Mo	—	Ru, Rh
	Ag	Cd	In	Sn	Sb	Te	I	& Pd
	Cs	Ba	Di	Ce	—	—	—	
	—	—	—	—	—	—	—	
	—	—	Er	La	Ta	W	—	Os, Ir
	Au	Hg	Tl	Pb	Bi	—	—	& Pt
	—	—	—	Th	—	U	—	

◀ **Figure 5.1 Mendeleev's Classification of the Elements (1871)** The formula for the oxide of an element is indicated by the notation R_2O, RO, and so on. For example, Group I oxides include Li_2O, Na_2O, and K_2O; and Group II oxides include BeO, MgO, and CaO. In Groups III and IV, Mendeleev indicated the existence of three undiscovered elements, which he named ekaboron, ekaaluminum, and ekasilicon.

FEATURED SCIENTIST **Dmitri Mendeleev**

Q: Which element in the periodic table is named in honor of Dmitri Mendeleev?

Dmitri Ivanovich Mendeleev (1834–1907) was born in Siberia, the youngest of 14 to 17 children (records vary). Mendeleev's father was a high school principal, but blindness ended his career when he was quite young. Mendeleev's mother, a woman of remarkable energy and determination, subsequently started a glass factory to support the family. About the time Mendeleev graduated from high school, his father died and his mother's factory burned down. Although destitute, his mother used her influence to get him into college just a few months before her own death.

Mendeleev enrolled at the University of St. Petersburg where he graduated first in his class. Following graduate work in Europe, he returned to St. Petersburg as a professor of chemistry. Mendeleev's consuming interest was finding a common thread that linked the rapidly growing number of elements. In 1869, he published a table that related elements according to increasing atomic mass. He even predicted the existence and properties of three undiscovered elements. In 1875, the predicted element gallium was discovered in France (Gallia); in 1879, scandium was discovered in Scandinavia; and in 1886, germanium was discovered in Germany.

Gradually, scientists recognized the importance of the periodic table, and Mendeleev became the most famous chemist in the world. His professional reputation thrived, and he was invited to major universities throughout Europe and gave lectures in the United States and Canada.

His personal reputation suffered when he divorced his wife and married a young art student. In the eyes of the Russian Orthodox Church he was a bigamist because he did not wait the required 7 years before he remarried. When Czar Alexander II was questioned about Mendeleev's bigamy, he replied, "Yes, Mendeleev has two wives, but I have only one Mendeleev."

In the late nineteenth century, Russia was engulfed in political turmoil. Mendeleev was a liberal and voiced his concern about the rights of students in general and women in particular. His outspoken speech led to his forced resignation from the university in 1890. Fortunately, Mendeleev still had influential friends, and he was named director of the Bureau of Weights and Measures.

In 1906, he missed winning the Nobel Prize by a single vote. Historians speculate that Mendeleev was more deserving of the recognition, but owing to his controversial personality, the prize was awarded instead to Henri Moissan, the French chemist who discovered the element fluorine.

▲ **Dmitri Mendeleev (1834–1907)**
A spiral form of the periodic table conveys the concept of chemical evolution. Starting with H, all the elements evolved step by step. Mendeleev's likeness is commemorated in the center of the spiral of elements. Notice that there are 18 rows of elements corresponding to the 18 groups of elements in the periodic table.

A: Element 101 is named mendelevium, symbol Md, in honor of Mendeleev.

The Noble Gases

The arrangement of the elements expanded significantly with the discovery of the group of similar elements that appear on the far right side of the periodic table. Argon was isolated from air in 1894, and helium, neon, krypton, xenon, and radon were discovered in the ensuing 5 years. Originally, this collection of gaseous elements was referred to as the *inert gases* because they showed no chemical reactivity.

In 1962 a compound containing xenon gas was synthesized at the University of British Columbia. After the first compound was made, several other compounds containing inert gases quickly followed. Although compounds of xenon and krypton have been prepared, the elements argon, neon, and helium have yet to be combined. More recently, the term *noble gas* has been substituted for *inert gas* to convey the unreactive nature of the gases in this group. Similarly, copper, silver, and gold are referred to as noble metals because of their resistance to chemical reaction.

5.2 The Periodic Law Concept

LEARNING OBJECTIVE

- Describe the modern periodic law proposed by Moseley.

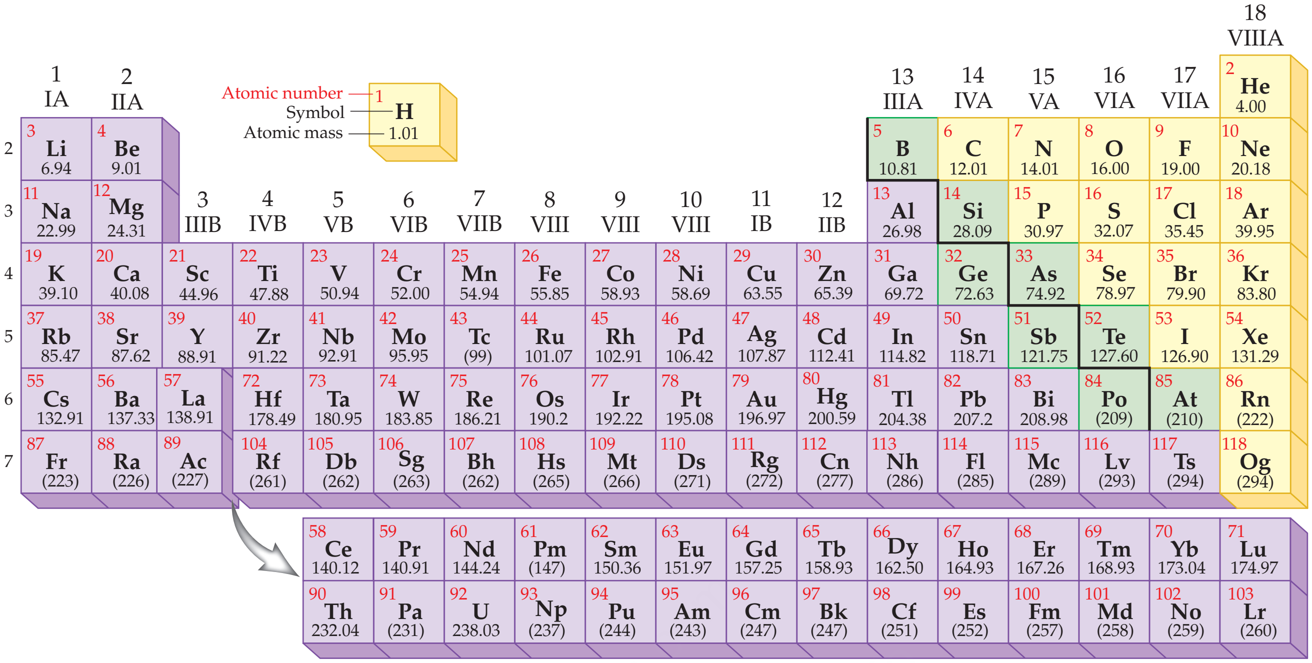
In 1869, Mendeleev proposed that elements showed recurring properties according to increasing atomic mass. In 1913, H. G. J. Moseley (1887–1915), a postdoctoral student, bombarded atomic nuclei with high-energy radiation. By studying the X rays that were subsequently emitted, Moseley discovered that the nuclear charge increased by 1 for each element in the periodic table. Thus,

H = 1 He = 2
Li = 3 Be = 4 B = 5 C = 6 N = 7 O = 8 F = 9 Ne = 10
Na = 11 Mg = 12 Al = 13 Si = 14 P = 15 S = 16 Cl = 17 Ar = 18

Moseley correctly concluded that arranging elements according to increasing nuclear charge, rather than atomic mass, more clearly explained their repeating properties. That is, arranging elements according to atomic number better explained the trends found in the periodic table. Elements, therefore, should be arranged according to the number of protons in their nucleus and not their atomic mass. The **periodic law** states that the properties of elements recur in a repeating pattern when arranged according to *increasing atomic number.* As it so happens, with only a few exceptions, these trends are identical.

In 1913, Niels Bohr introduced the concept of electron energy levels and Mendeleev's periodic table soon took on a new shape. The new shape resembled the arrangement used today, as shown in Figure 5.2.

If you closely examine the fourth row in the periodic table, you will notice that the sequence of atomic masses does not increase from Co to Ni, even though Co precedes Ni in the periodic table. But notice that the atomic number of Co (27) is less than the atomic number of Ni (28). It is also true that the properties of cobalt resemble those of the elements in Group 9, and that the properties of Ni are similar to those of the elements in Group 10.



	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
																		2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (99)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Ds (271)	111 Rg (272)	112 Cn (277)	113 Nh (286)	114 Fl (285)	115 Mc (289)	116 Lv (293)	117 Ts (294)	118 Og (294)

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (147)	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 Th 232.04	91 Pa (231)	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

▲ **Figure 5.2 The Modern Periodic Table of Elements** The atomic numbers increase stepwise throughout the periodic table. Note that atomic masses, with few exceptions, also increase. However, the nuclear charge (that is, the atomic number) more clearly explains the recurring properties of elements.

EXAMPLE EXERCISE 5.1 Periodic Law

Find the two elements in the fifth row of the periodic table that violate the original periodic law proposed by Mendeleev.

Solution

Mendeleev proposed that elements be arranged according to increasing atomic mass. Beginning with Rb, each of the elements in the fifth row increases in atomic mass until iodine. Although the atomic numbers of Te (52) and I (53) increase, the atomic masses of Te (127.60) and I (126.90) do not. Experimentally, it is I that has properties similar to F, Cl, and Br; thus, I comes after Te in the periodic table.

Practice Exercise

Find a pair of elements in the periodic table with atomic numbers less than 20 that are an exception to the original periodic law.

Answer: Ar and K.

Concept Exercise

The modern periodic law states that elements are arranged according to increasing (atomic mass/atomic number).

Answer: See Appendix G, 5.1.

5.3 Groups and Periods of Elements

A vertical column in the periodic table is called a **group** or family of elements, and a horizontal row is called a **period** or series. When we examine the periodic table on the inside front cover of this textbook, we notice there are seven horizontal rows of elements. These periods of elements are numbered 1–7. The first period has only two elements, H and He. The second and third periods each have eight elements, Li through Ne, and Na through Ar. The fourth and fifth periods have 18 elements, K through Kr and Rb through Xe.

LEARNING OBJECTIVES

- Classify the elements according to their groups (families) and periods (series) in the periodic table.
- Designate a group of elements in the periodic table using both the American convention (IA–VIIIA) and the IUPAC convention (1–18).

When Mendeleev constructed his original periodic table, he did not know where to place hydrogen as its properties were unlike any group of elements. Because hydrogen is a gas, it has physical properties similar to nonmetals. On the other hand, hydrogen reacts by losing an electron—and losing electrons is a chemical property of metals. Today, H appears at the top of the far left column in most periodic tables.

There are 18 vertical columns in the periodic table. In the past, American chemists have used a Roman numeral followed by the letter A or B to designate a group of elements, for example IA, IIA, IIB. In the periodic table, you will notice that Group IA contains the elements Li to Fr, Group IIA has the elements Be to Ra, and Group IIB has the elements Zn, Cd, Hg.

Since the 1920s, a controversy has existed over the numbering of the groups in the periodic table. Recently, the International Union of Pure and Applied Chemistry (IUPAC) resolved the dispute by proposing to number the groups using the numerals 1 to 18. For example, IUPAC recommends that Group IA be designated Group 1, Group IIB be designated Group 12, and Group VIIIA be designated Group 18. Currently, chemists are in the process of adopting the IUPAC convention. Hence, the periodic tables in this textbook will list both conventions, for example, IA/1, IIB/12, and VIIIA/18.

In the periodic table, groups of elements can also be referred to by their family names. The family name for Group IA/1 is **alkali metals**. The elements in Group IIA/2 are called **alkaline earth metals**. The elements in Group VIIA/17 are called **halogens**. Group VIIIA/18 elements are all gases that are usually unreactive under normal conditions; they are called **noble gases**.

Representative Elements and Transition Elements

Groups of elements are classified as either representative elements or transition elements. **Representative elements** (also called *main-group elements*) are found in the A groups, on the left and right sides of the periodic table. As a rule, the chemical behavior of representative elements is predictable. For example, magnesium always reacts with oxygen to produce MgO.

Transition elements are found in the B groups in the middle of the periodic table. The chemical behavior of transition elements is not as predictable as that of representative elements. For instance, in the presence of limited oxygen, iron reacts to form FeO; if excess oxygen gas is available, the product is Fe_2O_3.

Inner transition elements are found beneath the main portion of the periodic table. These two series are placed below the main portion to avoid an unduly wide periodic table. The first series of elements, Ce through Lu, lie between La and Hf in the periodic table and are considered part of Period 6. This series is called the **lanthanide series** because it follows lanthanum. The elements Ce through Lu have similar properties, and they occur together in nature along with Sc, Y, and La. The natural abundance of these elements in Earth's crust is less than 0.005%, and they are collectively referred to as **rare earth elements**.

The second series of inner transition elements, Th through Lr, is called the **actinide series** because it follows element 89, actinium. The actinide series is considered part of Period 7. All the elements in this series are radioactive; in fact, except for trace amounts, none of the elements past uranium is naturally occurring. Element 93 and beyond are created in particle accelerators called "atom smashers." The isotopes of these elements have very short lifetimes, often less than a millisecond. The elements following uranium, Np through element 118, are called **transuranium elements**. The periodic table shown in Figure 5.3 summarizes the names of the important groups and periods.

EXAMPLE EXERCISE 5.2 Groups and Periods of Elements

Select the symbol of the element that fits each of the following descriptions:

(a) the alkali metal in the fourth period
(b) the halogen in the third period
(c) the rare earth with the lowest atomic mass
(d) the metal in Group VIIB/7 and Period 4

Solution

Referring to the periodic table in Figure 5.3, we have

(a) K
(b) Cl
(c) Sc
(d) Mn

Practice Exercise

Select the symbol of the element that fits each of the following descriptions:

(a) the alkaline earth metal in the sixth period
(b) the noble gas in the third period
(c) the actinide with the highest atomic mass
(d) the semimetal in Group IIIA/13

Answers:

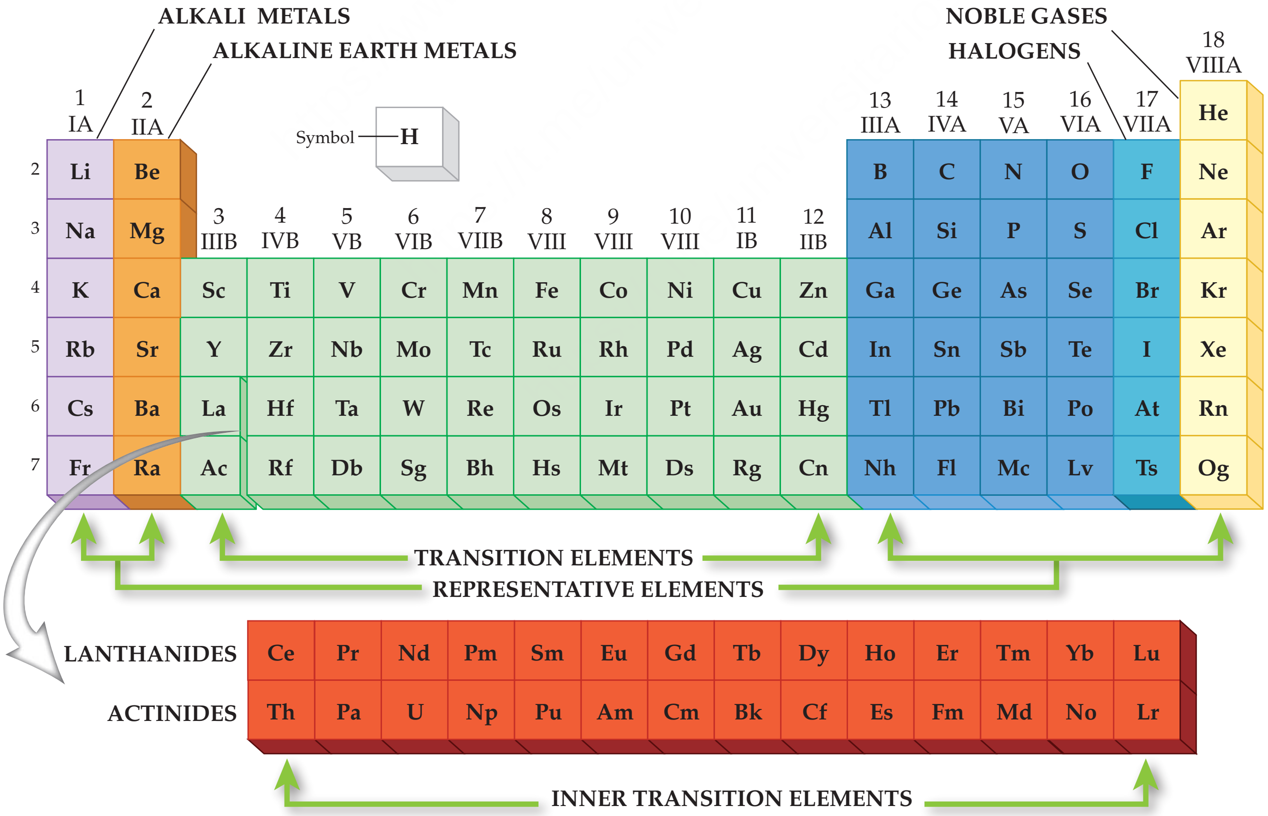
(a) Ba
(b) Ar
(c) Lr
(d) B

Concept Exercise

What is the group number for each of the following families of elements?

(a) alkali metals
(b) alkaline earth metals
(c) halogens
(d) noble gases

Answer: See Appendix G, 5.2.



	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
																		He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

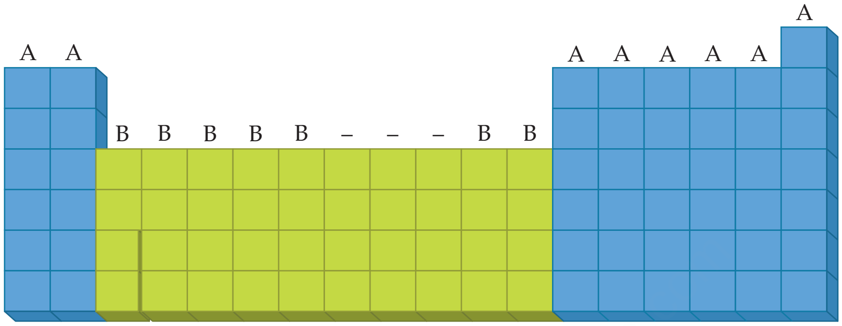
LANTHANIDES	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
ACTINIDES	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

▲ **Figure 5.3 Names of Groups and Periods** The common names of groups and periods are shown for selected families and series.

A CLOSER LOOK IUPAC Group Numbers

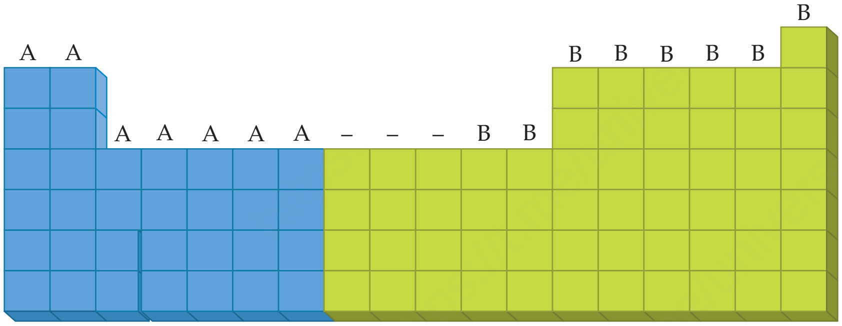
Q: Which two groups of transition elements in the periodic table have the same designation according to both the American convention and European convention?

Recently, the designation of groups of elements in the periodic table has been a topic of interest. In the 1970s a lingering controversy surfaced that was addressed by IUPAC. Previously, in the United States, elements on the left side and right side of the periodic table had been designated A groups. The elements in the middle of the periodic table were labeled B groups. This so-called American convention is shown next.



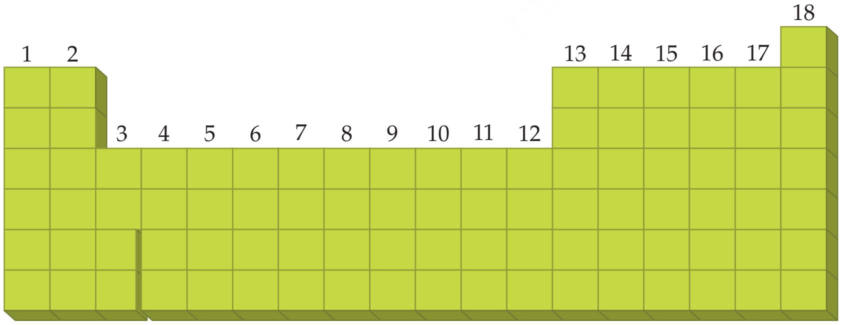
◀ **Periodic Table of Elements—American Convention** Representative elements are placed in A groups and transition elements are in B groups.

In Europe, elements on the left side of the periodic table had been previously designated A groups, and those on the right side had been labeled B groups. The so-called European convention is shown in the following illustration.



◀ **Periodic Table of Elements—European Convention** The elements on the left are A groups, and the elements on the right are B groups.

After much discussion in 1985 IUPAC proposed that groups of elements be designated by the numerals 1–18. Currently, IUPAC recommends that groups of elements be numbered as shown in the following periodic table.



◀ **Periodic Table of Elements—IUPAC Convention** Groups of elements are indicated by the sequential numerals 1–18.

One of the reasons for the IUPAC recommendation is that the last digit in the group number corresponds to the Roman numeral in both the American and European conventions. For example, in group 13, the number 3 corresponds to the Roman numeral III in both the American (IIIA) and European (IIIB) versions.

A: Groups IB and IIB have the same designation in both the American and European versions of the periodic table. These two groups are labeled 11 and 12 using IUPAC convention.

5.4 Periodic Trends

LEARNING OBJECTIVES

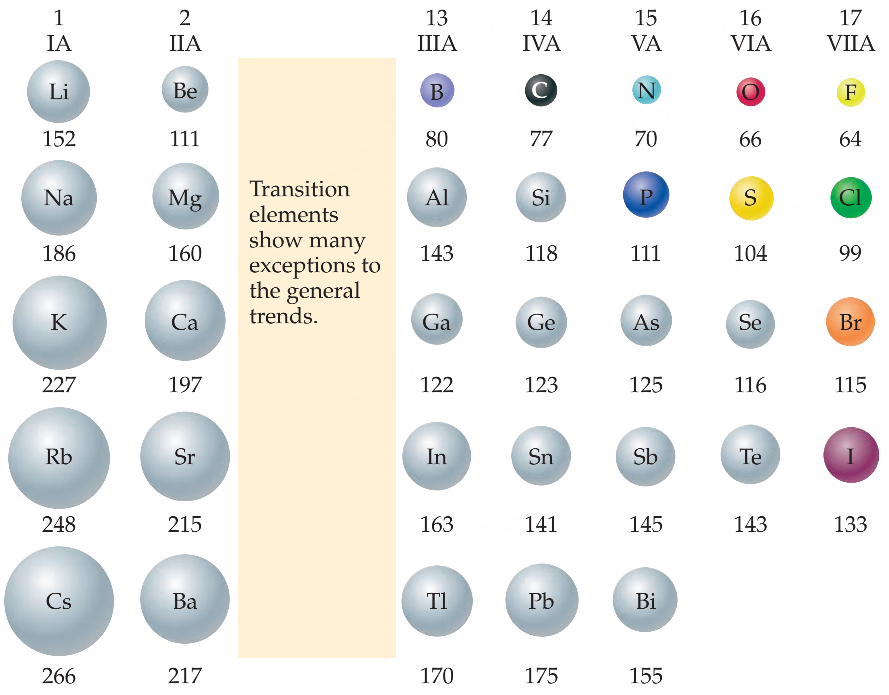
- Explain the trend in metallic character within a group or period.
- Explain the trend in atomic size within a group or period.

We can visualize atoms as spheres and express their size in terms of atomic radius. The atomic radius is the distance from the nucleus to the outermost electrons. As this distance is very small, we will express the atomic radius in units of nanometers.

In the periodic table there are two general trends in atomic size. First, as we move through a group of elements from bottom to top, the radii of the atoms decrease. We can explain the decrease in radius because there are fewer energy levels of electrons surrounding the nucleus. As the number of energy levels decreases, the distance from the nucleus to the outermost electrons decreases. Therefore, the trend in *atomic radius decreases up a group.*

Second, as we move left to right within a period of elements, the radii of the atoms decrease. The explanation for this decrease is as follows: The atomic number increases from left to right, and therefore, the number of protons increases. As the number of protons in the nucleus increases across a period, the positive nuclear charge increases. This has the effect of pulling the negative electrons closer to the nucleus and reducing the size of the atom. Thus, the *atomic radius decreases across a period* from left to right.

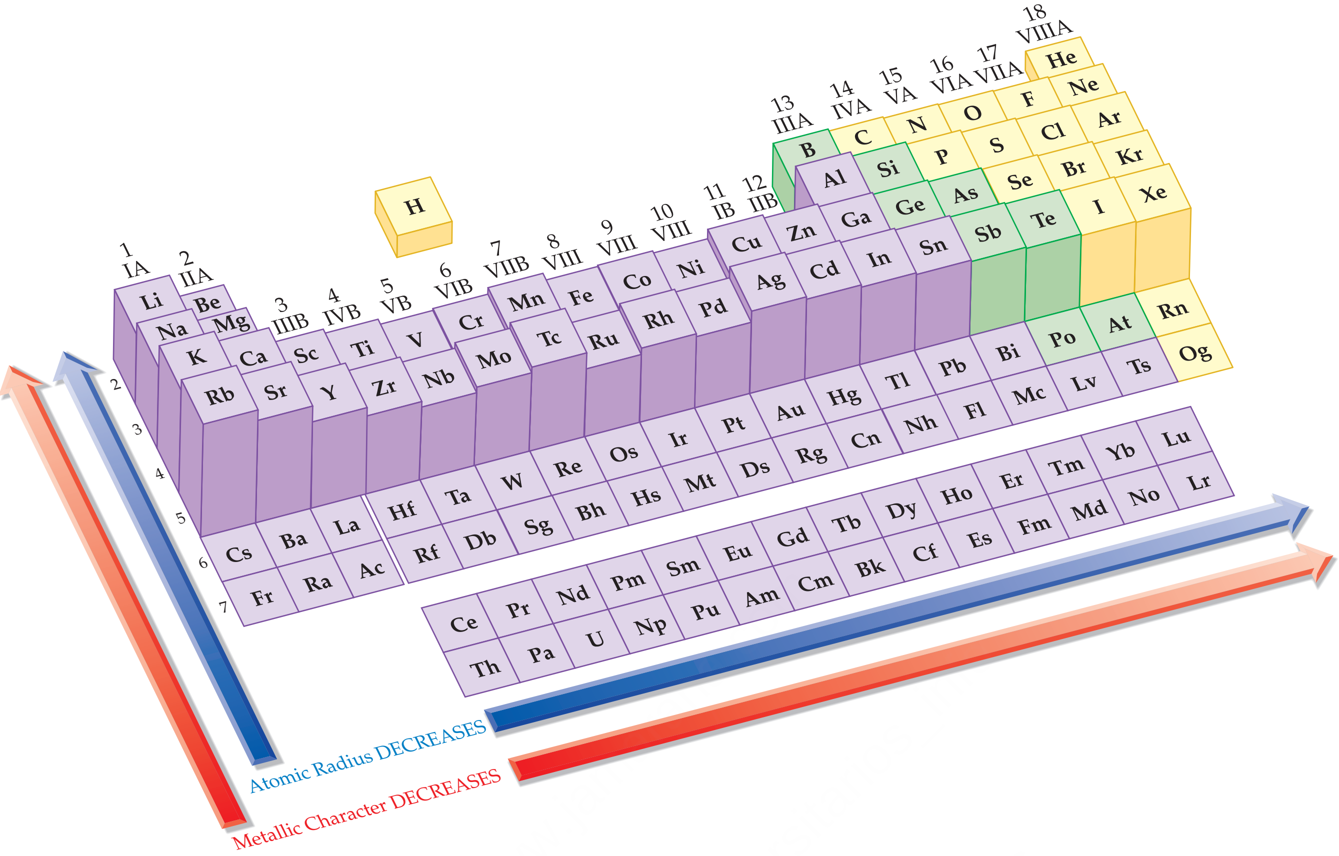
Figure 5.4 illustrates trends in the size of atomic radii for a portion of the periodic table. Transition elements and inner transition elements have been omitted as there are numerous exceptions to the general trends. Moreover, the atomic radii of transition elements in a given sublevel are reasonably similar in size.



◀ **Figure 5.4 Atomic Radii of Selected Representative Elements**
The values for atomic radii are given in picometers (pm). Note the atomic radius decreases up a group, and from left to right in the periodic table.

The periodic table shows a horizontal trend in metallic character of the elements. First, recall that metals are on the left side of the periodic table and nonmetals are on the right side. The most reactive metals are on the far left side, and the most reactive nonmetals are on the far right side of the table. Thus, as we move from left to right, *metallic character decreases across a period.*

The periodic table also shows a vertical trend in metallic character of the elements. A metal reacts by losing its outermost valence electrons. As we move up a group, the atomic radius is smaller and the outermost electrons are closer to the nucleus. When the negatively charged electrons and the positively charged nucleus are closer together, it is more difficult for an atom to lose an electron. Therefore, *metallic character decreases up a group.* Figure 5.5 illustrates the general trends in the periodic table for atomic radius and metallic character.



▲ **Figure 5.5** **General Trends in Atomic Radii and Metallic Character**

The general trend for the atomic radius is to decrease both up a group and across a period from left to right. The metallic character trend in the periodic table is similar.

EXAMPLE EXERCISE 5.3 Periodic Table Predictions

Based on the general trends in the periodic table, predict which element in each of the following pairs has the smaller atomic radius:

(a) Na or K (b) P or N (c) Ca or Ni (d) Si or S

Solution

The general trend in atomic radii decreases up a group and across a period from left to right. Referring to the periodic table,

(a) Na is above K in Group IA/1; the atomic radius of Na is smaller.
(b) N is above P in Group VA/15; the atomic radius of N is smaller.
(c) Ni is to the right of Ca in Period 4; the atomic radius of Ni is smaller.
(d) S is to the right of Si in Period 3; the atomic radius of S is smaller.

Practice Exercise

Based on the general trends in the periodic table, predict which element in each of the following pairs has the most metallic character:

(a) Sn or Pb (b) Ag or Sr (c) Al or B (d) Ga or Ge

Answers:

(a) Pb (b) Sr (c) Al (d) Ga

Concept Exercise

Which period represents the smallest atomic radius? Which group represents the most metallic character?

Answers: See Appendix G, 5.3.

A CLOSER LOOK Pencil Lead

Q: Why is the term "pencil lead" a misconception?

In the sixteenth century, a large deposit of graphite was discovered in England. At the time, scientists believed that graphite was a form of lead, because graphite looks similar to lead. However, it was immediately discovered that graphite, unlike lead, was an excellent tool for writing on paper. This unique quality of graphite made it a valuable commodity; so much so that the English mines had to be guarded to prevent theft of the graphite.

▲ Graphite sticks

Graphite is actually a crystalline form of carbon and is very brittle. To prevent breakage and make it suitable for writing, graphite was embedded in wood. The term pencil originated from the Latin word "pencillus" that translates as "little tail." The graphite in a pencil was referred to as "black lead" at the time, and the term "pencil lead" has persisted to this day.

In the nineteenth century, England and France were at war and France did not have access to pure English graphite. However, a French military officer discovered that mixing graphite powder with clay, followed by firing in a high-temperature kiln, produced a substance suitable for writing on paper. Moreover, by varying the ratio of graphite powder to clay, he could adjust the hardness of the "lead," and thus the darkness of the writing. Subsequently, numbers were assigned to indicate the hardness of the graphite/clay mixture. That is, #1 corresponded to a soft, dark "lead," #2 corresponded to a medium, dark "lead," #3 corresponded to a medium, light "lead," and #4 corresponded to a hard, light "lead." If you examine a yellow pencil, you may see the notation HB, where the H is an abbreviation for "hard," and B is an abbreviation for "black."

A: The term "pencil lead" is a misconception because pencils do not contain lead. Pencils contain graphite, which, like diamond, is a crystalline form of carbon.

5.5 Properties of Elements

LEARNING OBJECTIVES

- Predict a physical property for an element given the values of other elements in the same group.
- Predict a chemical formula for a compound given the formulas of other compounds containing an element in the same group.

Mendeleev designed the periodic table so that elements with similar properties were placed in the same group. He was then able to predict yet undiscovered elements from the gaps in his periodic table. He successfully predicted the properties of ekasilicon (i.e., germanium), for example, after noticing a trend in properties such as color, density, melting point, and atomic mass.

Predicting Physical Properties

Today, chemists have access to information regarding properties of elements cataloged in reference books such as the *Handbook of Chemistry and Physics* and online sources such as Wikipedia. However, we can use the same procedure that Mendeleev used to estimate the numerical value for a physical property. For example, if we know the density of two elements in the same group, we can make a reasonable estimate of the density for another element in the same group. Table 5.2 lists selected physical properties of alkali metals, which appear in Group IA/1 in the periodic table.

TABLE 5.2 Physical Properties of Alkali Metals

Element	Atomic Radius (pm)	Density (g/mL at 20 °C)	Melting Point (°C)	Atomic Mass (amu)
Li	152	0.53	180.5	6.94
Na	186	0.97	97.8	22.99
K	227	0.86	63.3	39.10
Rb	248	1.53	38.9	85.47
Cs	266	1.87	28.4	132.91

*Note the density of K is slightly less than that of Na. Small irregularities in group trends are not unusual.

Notice that the properties of the radioactive element francium, Fr, are not included with those of the other alkali metals in Table 5.2. We can, however, make some reasonable predictions about its physical properties based on the trends shown by the other elements. Because Fr is below Cs in Group IA/1, we can predict that its atomic radius is greater than 266 pm, its density is greater than 1.87 g/mL, its melting point is less than 28.4 °C, and its atomic mass is greater than 132.91 amu.

EXAMPLE EXERCISE 5.4 Predicting Physical Properties

Predict the missing value (?) for each physical property listed below. The (a) atomic radius, (b) density, and (c) melting point are given for two of three alkaline earth metals in Group IIA/2.

Element	Atomic Radius	Density at 20 °C	Melting Point
Ca	197 pm	1.54 g/mL	(?) °C
Sr	215 pm	(?) g/mL	769 °C
Ba	(?) pm	3.65 g/mL	725 °C

Solution

We can estimate a value for the physical property of an element by observing the trend in values for other elements within the same group.

(a) To estimate the atomic radius for Ba, we first find the atomic radius increase from Ca to Sr, that is, 215 pm − 197 pm = 18 pm. If we add the difference (18 pm) to the atomic radius of Sr (215 pm), we obtain 233 pm. This estimated value of 233 pm is reasonable as the measured value for the atomic radius of Ba is slightly less, *217 pm*.

(b) Notice that Sr lies between Ca and Ba in Group IIA/2. Thus, we can estimate that the density of Sr lies midway between Ca and Ba. To find the density of Sr, we calculate the average value for Ca and Ba; that is, (1.54 g/mL + 3.65 g/mL)/2 = 2.60 g/mL. (The actual measured value is *2.63 g/mL*.)

(c) From the general trend, we can predict that the melting point of Ca is greater than that of Sr. To determine the value, let's find the increase in melting point from Ba to Sr. It is 769 °C − 725 °C = 44 °C. Now we add 44 °C to the value of Sr: 769 °C + 44 °C = 813 °C. Therefore, we predict the melting point of Ca as 813 °C. (The actual value is *839* °C.)

Practice Exercise

Predict the missing value (?) for each physical property listed below. The (a) atomic radius, (b) density, and (c) melting point are given for two of the metals in Group VIII/10.

Element	Atomic Radius	Density at 20 °C	Melting Point
Ni	125 pm	8.91 g/cm^3	(?) °C
Pd	138 pm	(?) g/cm^3	1554 °C
Pt	(?) pm	21.5 g/cm^3	1772 °C

Answers: (a) 151 pm (b) 15.2 g/cm^3 (c) 1336 °C

Concept Exercise

Given the density of silver (10.5 g/cm^3) and gold (19.3 g/cm^3) in Group IB/11, estimate the density for the synthetic element roentgenium, which is below gold in the periodic table.

Answers: See Appendix G, 5.4.

Predicting Chemical Properties

In Chapter 7, we will study chemical reactions systematically. However, it is possible to predict the products of chemical reactions by understanding the organization of elements in the periodic table. For example, if we know that magnesium and oxygen react to give magnesium oxide (MgO), we can predict that the other Group IIA/2 elements react in a similar fashion. That is, calcium, strontium, and barium should react with oxygen to give similar oxides (CaO, SrO, and BaO), and they do. There are exceptions to the general rule, but the principle of using the periodic table is helpful to our understanding of chemical properties.

▲ **Gallium Metal**
Gallium has chemical properties similar to aluminum. However, gallium has a much lower melting point (30 °C), and will melt in the hand.

EXAMPLE EXERCISE 5.5 Predicting Chemical Properties

Metallic sodium reacts with chlorine gas to give sodium chloride, NaCl. Predict the products formed when (a) lithium and (b) potassium react with chlorine gas.

Solution

Because Li and K are in the same group as Na (Group IA/1), we can predict that the products are similar to NaCl. Thus,

(a) lithium metal should react with chlorine gas to give LiCl.
(b) potassium metal should react with chlorine gas to give KCl.

Practice Exercise

The chemical formulas for the oxides of potassium, calcium, gallium, and germanium are, respectively, K_2O, CaO, Ga_2O_3, and GeO_2. Refer to the periodic table and predict the chemical formula for each of the following compounds:

(a) rubidium oxide
(b) strontium oxide
(c) indium oxide
(d) lead oxide

Answers: (a) Rb_2O (b) SrO (c) In_2O_3 (d) PbO_2

Concept Exercise

Predict which of the following metals has chemical properties most similar to zinc: Fe, Cu, or Cd.

Answers: See Appendix G, 5.5.

5.6 Blocks of Elements

LEARNING OBJECTIVES

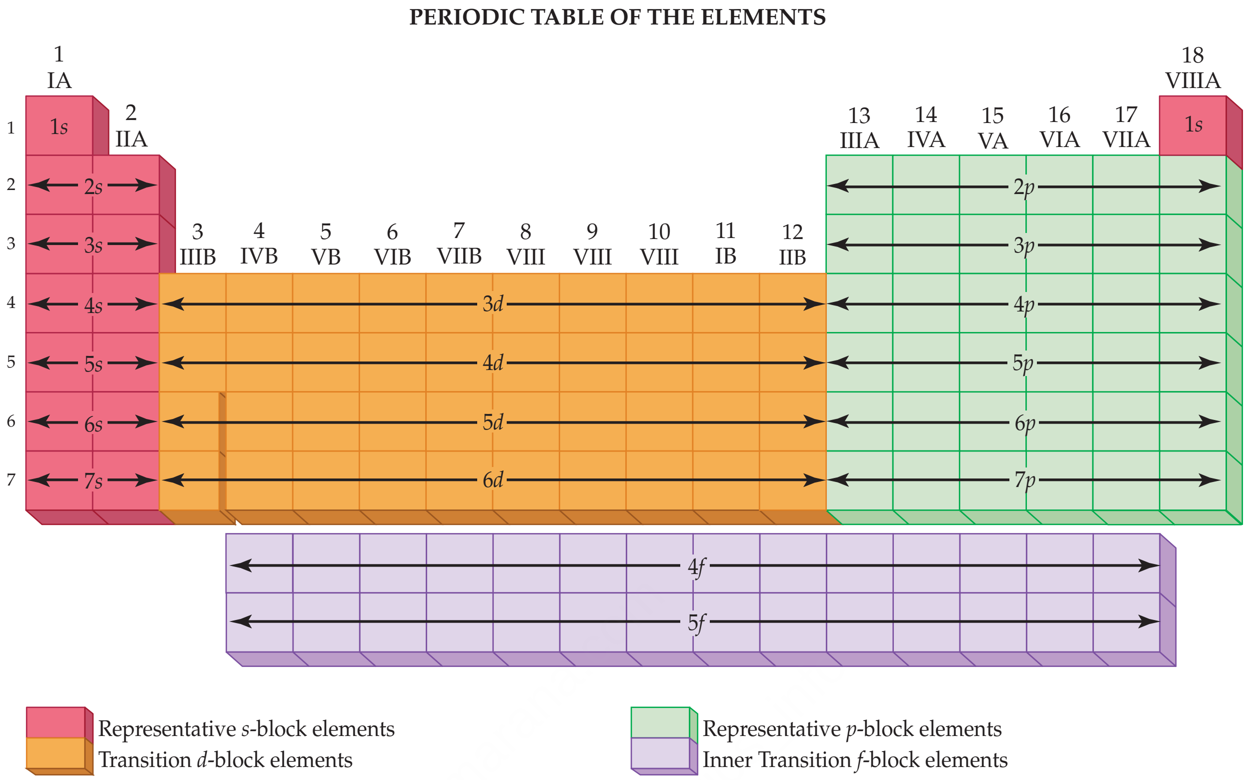
- Predict the highest energy sublevel for an element given its position in the periodic table.
- Predict the electron configuration for an element given its position in the periodic table.

In Section 4.9, we learned that the order of electron energy sublevels according to increasing energy is $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$. In this section, we will see that the unusual shape of the periodic table is the result of the ordering of energy sublevels. That is, the order of energy sublevels follows the systematic arrangement of elements by groups.

Elements in Groups IA/1 and IIA/2 fill *s* sublevels. The *s* sublevels are to the left in the periodic table and are collectively called the *s block* of elements. Elements in Groups IIIA/13 through VIIIA/18 fill *p* sublevels, and are referred to as the *p block* of elements. Transition elements in Groups IIIB/3 through IIB/12 fill *d* sublevels, and likewise make up the *d block* of elements. Finally, the inner transition elements (Ce–Lu and Th–Lr) fill *f* sublevels, and comprise the *f block* of elements. Figure 5.6 illustrates the *s*, *p*, *d*, and *f* block elements in the periodic table.

In Chapter 4, we had to memorize that the 3d energy sublevel comes after the 4s energy sublevel. Now, we can simply refer to the periodic table and see that *s* block elements precede each *d* sublevel. Also, recall that the number of sublevels corresponds to the number of the main energy level. For example, in the third main level, there are three sublevels (3*s*, 3*p*, and 3*d*). In the fourth main level, there are four sublevels (4*s*, 4*p*, 4*d*, and 4*f*).

We can see in Figure 5.6 that members of the lanthanide and actinide series are *f* block elements. The lanthanides are filling the 4*f* energy sublevel, and the actinides are filling the 5*f* sublevel.



▲ **Figure 5.6 Blocks of Elements in the Periodic Table** The relationship between energy sublevels and the *s*, *p*, *d*, and *f* blocks of elements in the periodic table are shown.

Helpful Hint
Predicting Chemical Formulas

In Chapter 5, we notice that the periodic table is a powerful tool for predicting physical and chemical properties. In Chapter 6, we will learn how to name chemical compounds and write chemical formulas using the periodic table; that is, we will learn the “language of chemistry.”

It can be time-consuming to learn the “language of chemistry.” However, if you understand periodic table relationships, you can simplify naming chemical compounds. For example, if you know that table salt is sodium chloride (NaCl), you can predict the chemical formula of lithium chloride (LiCl) and potassium chloride (KCl), because Li and K are in the same group as Na. Similarly, you can predict the chemical formula of sodium fluoride (NaF), sodium bromide (NaBr), and sodium iodide (NaI), because F, Br, and I are in the same family as Cl in the periodic table.

EXAMPLE EXERCISE 5.6 Energy Sublevels and the Periodic Table

State the highest energy sublevel in each of the following elements:

(a) H (b) S (c) Ni (d) U

Solution

Refer to the periodic table and determine the energy sublevel based on the period and block of the elements.

(a) Hydrogen has only one electron; thus, H is filling a 1*s* sublevel.
(b) Sulfur is in the third period and is a *p* block element; S is filling a 3*p* sublevel.
(c) Nickel is in the first series of *d* block elements; Ni is filling a 3*d* sublevel.
(d) Uranium is in the second series of *f* block elements; U is filling a 5*f* sublevel.

Practice Exercise

State the energy sublevel being filled in each of the following series of elements:

(a) Cs − Ba (b) Y − Cd (c) In − Xe (d) Ce − Lu

Answer: (a) 6*s* (b) 4*d* (c) 5*p* (d) 4*f*

Concept Exercise

Refer to the periodic table and state the highest energy sublevel in a silver atom.

Answer: See Appendix G, 5.6.

Electron Configuration of Elements

In Section 4.10, we learned to write the **electron configuration** for an element after memorizing the order of sublevel filling. Now that we understand the relationship of sublevels in the periodic table, we can write electron configurations by referring to blocks of elements and counting the number of electrons in each block.

The electron configuration for Na (atomic number 11) is $1s^2\,2s^2\,2p^6\,3s^1$. As a shorthand method, we can abbreviate the electron configuration by indicating the innermost electrons with the symbol of the preceding noble gas. The preceding noble gas with an atomic number less than 11 is neon (atomic number 10). The symbol for neon is placed in brackets, [Ne], followed by the outermost electrons. That is, the electron configuration for Na can be written as [Ne] $3s^1$.

EXAMPLE EXERCISE 5.7 Electron Configuration and the Periodic Table

Refer to a periodic table and write the predicted electron configuration for each of the following elements by counting the number of electrons in each block:

(a) P (b) Co

Solution

Now that you understand blocks of elements in the periodic table, you can predict the order of sublevels according to increasing energy.

(a) Phosphorus is the third element in the $3p$ sublevel. The electron configuration for P is $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$, or [Ne] $3s^2\,3p^3$.

(b) Cobalt is the seventh element in the $3d$ sublevel. The electron configuration for Co is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^7$ or [Ar] $4s^2\,3d^7$.

Practice Exercise

Refer to a periodic table and write the predicted electron configuration for each of the following elements.

(a) Zn (b) Se

Answers:

(a) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}$ or [Ar] $4s^2\,3d^{10}$

(b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^4$ [Ar] $4s^2\,3d^{10}\,4p^4$

Concept Exercise

Refer to the periodic table and predict the number of $5d$ electrons in a Pt atom.

Answer: See Appendix G, 5.7.

Note There are a few exceptions in the predicted electron configurations for some of the elements. That is, some elements have actual electron configurations that vary from their predicted arrangements. On occasion, electrons in a lower energy sublevel jump to an unfilled sublevel of slightly higher energy. Some of these exceptions are easy to explain, whereas others are not. For our purposes, we will disregard these exceptions and write electron configurations based on the general predictions for the order of sublevel filling.

5.7 Valence Electrons

LEARNING OBJECTIVE

- Predict the number of valence electrons for any representative element.

When an atom undergoes a chemical reaction, only the outermost electrons are involved. These electrons are farthest from the nucleus, and more available to react with another atom. These outermost electrons are called **valence electrons**. Valence electrons form chemical bonds between atoms and are responsible for the chemical behavior of the element. For our purposes, we will ignore transition elements and consider only the representative elements. In these cases, the number of valence electrons is equal to the total number of electrons in the outermost *s* and *p* sublevels.

We can predict the number of valence electrons by noting the group number in the periodic table. Using the Roman numeral convention, we see that the group number is identical to the total number of valence electrons. That is, elements in Group IA have one valence electron, and elements in Group IIA have two valence electrons. Group IIIA elements have two electrons from an *s* sublevel plus one electron from a *p* sublevel, for a total of three valence electrons. Group IVA elements have four valence electrons, Group VA elements have five valence electrons, and so on.

With the IUPAC designations for group numbers, the last digit (for example, the "8" in 18) indicates the number of valence electrons. Group 13 elements have three valence electrons, Group 14 elements have four valence electrons, Group 15 elements have five valence electrons, and so on. Note that the *s* and *p* sublevels can have a maximum of two and six electrons, respectively. Thus, the maximum number of valence electrons (s sublevel + p sublevel) is equal to $8\,e^-$ ($2\,e^- + 6\,e^- = 8$).

Lithium is in Group IA/1, and so a lithium atom has only one valence electron. Oxygen is in Group VIA/16 and has six valence electrons. The following example further illustrates the relationship between group number and number of valence electrons:

EXAMPLE EXERCISE 5.8 Valence Electrons and the Periodic Table

Refer to the periodic table and predict the number of valence electrons for an atom of each of the following representative elements:

(a) Na
(b) Al
(c) S
(d) Xe

Solution

Find the element in the periodic table, note the group number, and indicate the number of valence electrons.

(a) Sodium is in Group IA/1, and so Na has 1 valence electron.
(b) Aluminum is in Group IIIA/13, and so Al has 3 valence electrons.
(c) Sulfur is in Group VIA/16, and so S has 6 valence electrons.
(d) Xenon is in Group VIIIA/18, and so Xe has 8 valence electrons.

Practice Exercise

Refer to the periodic table and state the number of valence electrons for any element in each of the following groups:

(a) Group IIA
(b) Group VA
(c) Group 14
(d) Group 17

Answers:

(a) 2
(b) 5
(c) 4
(d) 7

Concept Exercise

Why are predictions for the number of valence electrons only for representative elements?

Answer: See Appendix G, 5.8.

Note We can quickly predict the number of valence electrons for any main-group element by referring to its group number in the periodic table. However, the transition elements are filling *d* sublevels, not an *s* or *p* sublevel, which complicates the discussion. Therefore, we will not predict the number of valence electrons for a transition element.

5.8 Electron Dot Formulas

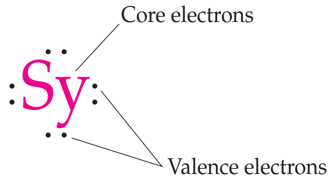
LEARNING OBJECTIVE

- Diagram the electron dot formula for any representative element.

We said previously that only valence electrons are involved when an element undergoes a chemical reaction. That's because valence electrons are farthest from the nucleus and are the most accessible. To keep track of valence electrons, chemists have devised a notation called the **electron dot formula**. Electron dot formulas are also referred to as *Lewis structures* in honor of G. N. Lewis (1875–1946), a famous American chemist who conceived the notion of valence electrons.

An electron dot formula shows the symbol of an element surrounded by its valence electrons. The symbol of the element represents the **core electrons** in an atom. That is, the symbol represents the inner electrons, which are not available for bonding. Dots are placed around the symbol to represent the valence electrons. Figure 5.7 illustrates the general symbol for an electron dot formula.

In practice, we will use the following guidelines for drawing electron dot formulas:



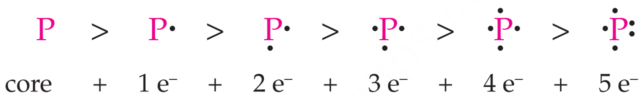
▲ **Figure 5.7 Electron Dot Formulas** The electron dot formula for an atom of an element uses the symbol for the element to represent the core electrons. Dots represent valence electrons, which are arranged clockwise around the symbol.

Guidelines for Drawing Electron Dot Formulas of Atoms

1. Write the symbol of the element to represent the core electrons.
2. We then draw a maximum of two dots on each side of the symbol. A dot represents 1 valence electron, and the maximum number is 8 electrons (s sublevel + p sublevel = 2 e^- + 6 e^- = 8 e^-).
3. Determine the number of valence electrons from the group number of the element in the periodic table. Draw one dot around the symbol for each valence electron.

Although there is no absolute rule for placing dots, the first dot is usually placed to the right of the symbol, the second dot beneath the symbol, the third dot to the left of the symbol, and the fourth dot above the symbol. The fifth dot pairs with the first dot. The sixth, seventh, and eighth dots pair with the second, third, and fourth dots, respectively, moving clockwise around the symbol.

As an example, consider the element phosphorus. Phosphorus is in Group VA/15 and, therefore, has five valence electrons. We can draw the electron dot formula for phosphorus as follows. First, write the symbol for phosphorus, then add five dots, one at a time, as follows:



Keep in mind that the placement of dots is not intended to show the actual positions of valence electrons. In point of fact, electrons are in constant motion, moving about the nucleus. Electron dot formulas are intended only to help us keep track of valence electrons. The following example exercise further illustrates electron dot formulas.

EXAMPLE EXERCISE 5.9 Electron Dot Formulas

Draw the electron dot formula for each of the following elements:

(a) Si (b) Xe

Solution

Let's find the group number of the element in the periodic table and note the number of valence electrons. We write the symbol of the element and place the same number of dots around the symbol as there are valence electrons. In these examples, Si has 4 valence electrons and Xe has 8. The electron dot formulas are as follows:

(a)


(b)


Practice Exercise

Draw the electron dot formula for each of the following elements:

(a) K (b) I

Answers:

(a) $\mathrm{K}\cdot$ (b) $:\overset{\cdot}{\underset{\cdot\cdot}{\mathrm{I}}}:$

Concept Exercise

Propose two ways of drawing the electron dot formula for Mg.

Answer: See Appendix G, 5.9.

LEARNING OBJECTIVES

- Describe the general trends of ionization energy in the periodic table.
- Predict the group with the highest and the lowest ionization energy.
- Predict the element in a pair having the higher ionization energy.

5.9 Ionization Energy

Although electrons can be removed from all elements, metals lose electrons more easily than nonmetals. In fact, metals undergo chemical reactions by losing one or more valence electrons. Because electrons are negatively charged, metals become positively charged after losing an electron. Any atom bearing a charge is called an **ion**.

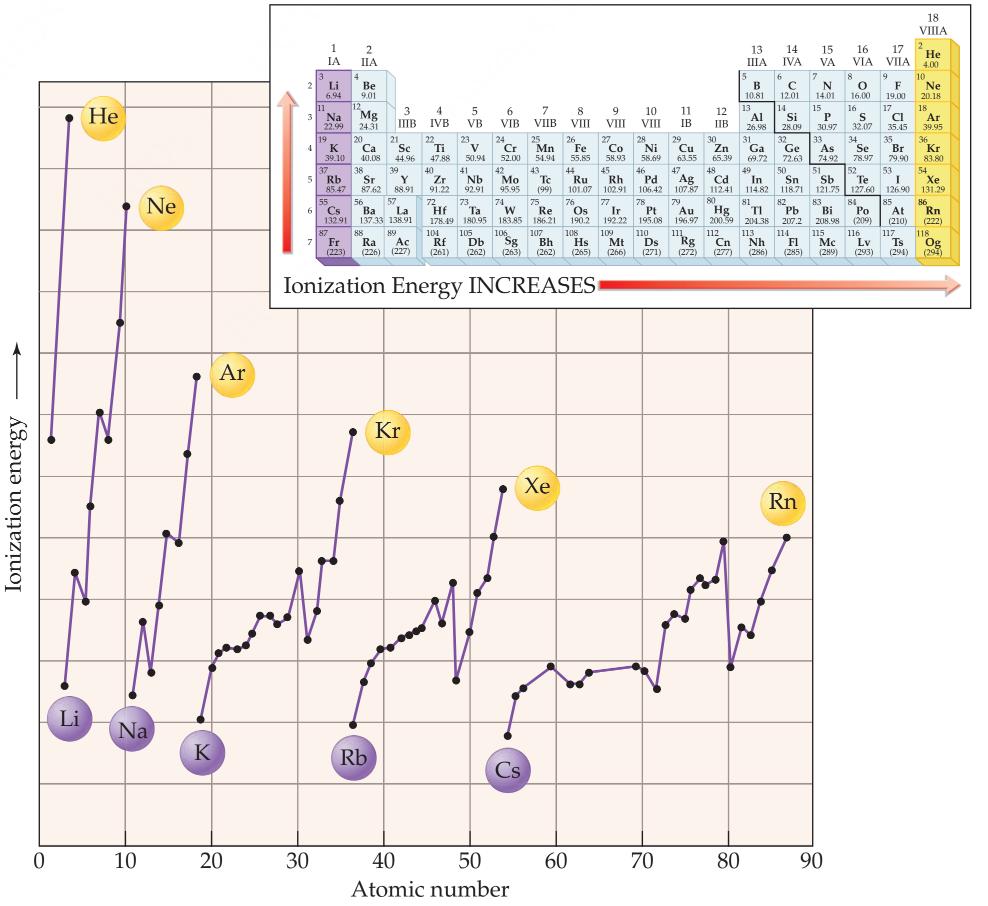
Energy is always required to remove an electron from an atom. By definition, the amount of energy necessary to remove an electron from a neutral atom in the gaseous state is called the **ionization energy**, or ionization potential. We can illustrate ionization of a sodium atom as follows:

$$\mathrm{Na} \xrightarrow{\text{ionization energy}} \mathrm{Na}^{+} + e^{-}$$

Figure 5.8 shows the relative energy required to remove a single electron from an atom of elements 1–86.

Notice in Figure 5.8 that the elements having the highest ionization energy belong to the same group—the noble gases. One reason this group shows little tendency to

Figure 5.8 Ionization Energy The diagram shows the energy required to remove a single electron from a neutral atom of an element in the gaseous state. Note the alkali metals have the lowest ionization energy, and the noble gases have the highest ionization energy.



undergo reaction is the difficulty with which an electron is removed. We can also reason that noble gases have high ionization energies because their valence shells are completely filled, and so they do not need to gain or lose electrons to become stable.

On the other hand, notice that the elements having the lowest ionization energy are alkali metals. Because they are in the same group, all alkali metals have a similar electron configuration. That is, they all have one electron in an *s* sublevel. If we remove an electron from an alkali metal, the resulting ion will have the same number of electrons as the preceding noble gas. For example, a lithium ion has 2 electrons (the same as He), a sodium ion has 10 electrons (the same as Ne), and a potassium ion has 18 electrons (the same as Ar). When an alkali metal loses an electron, it achieves a noble gas electron configuration, which is exceedingly stable.

There are general trends in the periodic table regarding ionization energy. Within a group of elements, the energy required to remove an electron increases as the atomic radius decreases. As we proceed up a group, the atomic radius decreases because the valence electrons are closer to the nucleus and are more tightly held. Thus, the ionization energy *increases up a group of elements.* As we move from left to right in the periodic table, the nuclear charge becomes greater. The energy required to remove an electron, therefore, increases. In general, the ionization energy *increases from left to right across a period of elements*.

EXAMPLE EXERCISE 5.10 Ionization Energy and the Periodic Table

Based on the general trends in the periodic table, predict which element in each of the following pairs has the higher ionization energy:

(a) Li or Na (b) O or F

Solution

Let's refer to a periodic table and apply the general trends in ionization energy, which increases up a group and across a period.

(a) Li is above Na in Group IA/1, and so Li has the higher ionization energy.

(b) F is right of O in Period 2, and so F has the higher ionization energy.

Practice Exercise

Based on the general trends in the periodic table, predict which element in each of the following pairs has the higher ionization energy:

(a) Na or Mg (b) O or S

Answers:

(a) Mg (b) O

Concept Exercise

Which group in the periodic table has the highest ionization energy? Which group has the lowest ionization energy?

Answers: See Appendix G, 5.10.

Note The amount of energy involved when a neutral atom gains an electron is called *electron affinity*. Metals show a weak tendency to gain an electron and have a low electron affinity. Nonmetals have a strong tendency to gain electrons and have a high electron affinity. The nonmetals closest to the noble gases have the highest electron affinity, so we can predict the electron affinity for F is greater than for O; and Cl is greater than S.

LEARNING OBJECTIVES

- Predict an ionic charge for a representative element.
- Diagram the predicted electron configuration for selected ions.

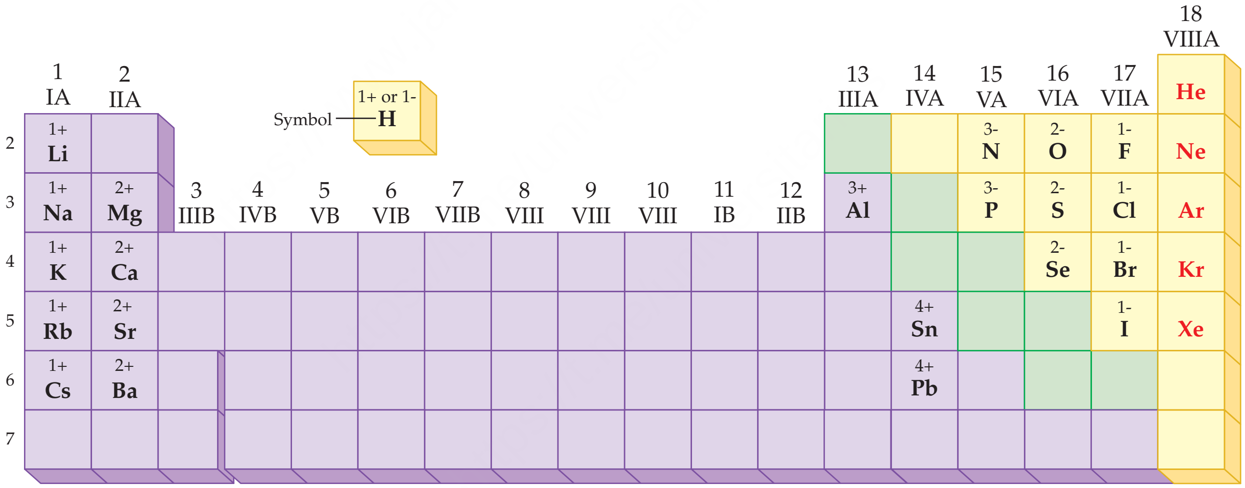
5.10 Ionic Charges

In general, when metals and nonmetals react, metals tend to lose electrons and nonmetals tend to gain electrons. More specifically, metals lose electrons from their valence shell, and nonmetals add electrons to their valence shell. Recall that an atom that bears a charge as a result of gaining or losing electrons is called an *ion*. Thus, metals become positive ions, and nonmetals become negative ions; in other words, they acquire an **ionic charge**.

The positive ionic charge on a metal ion is related to its number of valence electrons. Metals in Group IA/1 give up one valence electron to produce a positive ionic charge of 1+. Elements in Group IIA/2 give up their two valence electrons to produce a positive ionic charge of 2+. Metals in Group IIIA/13 usually lose three electrons and have an ionic charge of 3+. Group IVA/14 metals can lose four electrons, producing a charge of 4+.

The negative charge on a nonmetal ion is also governed by its number of valence electrons. Nonmetals in Group VIIA/17 have seven valence electrons and tend to add one electron to assume a noble gas configuration with eight valence electrons. After gaining one electron, the nonmetal has a negative ionic charge of 1−. Elements in Group VIA/16 gain two valence electrons, which produces an ionic charge of 2−. Nonmetals in Group VA/15 add three valence electrons, which give an ionic charge of 3−.

Figure 5.9 shows several common elements and their ionic charges as predicted by the position of the element in the periodic table.



▲ **Figure 5.9 Periodic Table of Selected Ions** Notice that metals have a positive ionic charge equal to their group number. Nonmetals have a negative ionic charge equal to 8 minus their A group number in the periodic table. Both metals and nonmetals attempt to assume a stable noble gas electron configuration.

Lithium is in Group IA/1, and so a lithium ion has an ionic charge of 1+. The lithium ion is written Li^{+}, with the 1 (in 1+) understood. Oxygen is in Group VIA/16, has an ionic charge of 2−, and is shown as O^{2-}. The following examples further illustrate the relationship between group number and ionic charge.

EXAMPLE EXERCISE 5.11 Predicting Ionic Charges of Cations

Predict the ionic charge for a sodium ion, $Na^{?+}$, based on the position of the element in the periodic table.

Solution

We refer to the periodic table to find the group number of the element. Sodium is in Group IA/1; after Na loses 1 valence electron the ionic charge is 1+, that is, Na^{+}.

Practice Exercise

Predict the ionic charge for an aluminum ion, $Al^{?+}$, based on the position of the element in the periodic table.

Answer: Al^{3+}

Concept Exercise

Predict the common ionic charge for Group IA/1 elements; Group IIA/2 elements; Group III/13 elements.

Answers: See Appendix G, 5.11.

EXAMPLE EXERCISE 5.12 Predicting Ionic Charges of Anions

Predict the ionic charge for a chlorine ion, $Cl^{?-}$, based on the position of the element in the periodic table.

Solution

We refer to the periodic table to find the group number of the element. Chlorine is in Group VIIA/17; after Cl gains 1 valence electron the ionic charge is 1−, that is, Cl^{-}.

Practice Exercise

Predict the ionic charge for a phosphorus ion, $P^{?-}$, based on the position of the element in the periodic table.

Answer: P^{3-}

Concept Exercise

Predict the common ionic charge for Group VIIA/17 elements; Group VIA/16 elements; Group VA/15 elements.

Answers: See Appendix G, 5.12.

In Section 5.9 we learned that Group IA/1 elements form ions by losing one electron. We showed that each Group IA/1 ion has an electron configuration identical to that of the previous noble gas element. Similarly, Group VIIA/17 elements gain one electron to assume an electron configuration identical to that of the next noble gas element.

For example, a sodium ion (Na^{+}) and a fluoride ion (F^{-}) each have 10 electrons, as does a neutral neon atom. Although their properties are not related, their electron configurations are identical. By definition, two or more ions having the same number of electrons are said to be **isoelectronic**. The sodium ion, the fluoride ion, and the neon atom are said to be members of an isoelectronic series.

EXAMPLE EXERCISE 5.13 Predicting Isoelectronic Ions

Refer to the periodic table and predict which of the following ions are isoelectronic with the noble gas argon:

(a) K^{+}	(b) Br^{-}
(c) Ca^{2+}	(d) O^{2-}

Solution

We refer to the periodic table to find the atomic number of the element.

(a) K^+ has 18 electrons ($19 - 1 = 18$); it is isoelectronic with argon (18 e^-).

(b) Br^- has 36 electrons ($35 + 1 = 36$); it is isoelectronic with krypton (36 e^-).

(c) Ca^{2+} has 18 electrons ($20 - 2 = 18$); it is isoelectronic with argon (18 e^-).

(d) O^{2-} has 10 electrons ($8 + 2 = 10$); it is isoelectronic with neon (10 e^-).

Practice Exercise

Refer to the periodic table and predict which of the following ions are isoelectronic with the noble gas xenon:

(a) Cs^+ (b) Cl^-

(c) La^{3+} (d) Se^{2-}

Answers:

(a) Cs^+ is isoelectronic with xenon (b) Cl^- is isoelectronic with argon

(c) La^{3+} is isoelectronic with xenon (d) Se^{2-} is isoelectronic with krypton

Concept Exercise

Predict the next ion in the isoelectronic series: Se^{2-}, Br^-, Kr, Rb^+, Sr^{2+}.

Answer: See Appendix G, 5.13.

Electron Configuration of Ions

In Section 5.6 we learned how to predict the electron configuration of an element from the blocks of elements in the periodic table. Now we will learn how to write the electron configuration of an ion. First, we find the element in the periodic table, and then we write out the electron configuration as we did in Section 5.6.

When writing the electron configuration for positive ions, we remove the number of electrons that corresponds to the positive charge on the ion. For example, the electron configuration for a sodium atom is $1s^2\,2s^2\,2p^6\,3s^1$. Sodium ion is Na^+, so it loses one electron. That is,

$$\begin{array}{ccc} \text{Na atom} & \xrightarrow{\text{loses } 1\ e^-} & Na^+ \\ 1s^2\,2s^2\,2p^6\,3s^1 & & 1s^2\,2s^2\,2^6 \end{array}$$

We should note that transition metals lose two electrons from the highest *s* sublevel first before losing electrons from their outer *d* sublevel. In the fourth-period elements, for example, 4*s* electrons are lost before 3*d* electrons. The electron configuration for a manganese(II) ion is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5$. Note that manganese forms a Mn^{2+} ion by losing two electrons from the 4*s* sublevel rather than the 3*d* sublevel. That is,

$$\begin{array}{ccc} \text{Mn atom} & \xrightarrow{\text{loses } 2\ e^-} & Mn^{2+} \\ 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^5 & & 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5 \end{array}$$

As shown in Section 5.6, electron configurations can be simplified by using a noble gas symbol to represent the core electrons. This method of showing electron configuration is called **core notation**. In the preceding example with a Mn atom, we can write the electron configuration as $[Ar]\,4s^2\,3d^5$ using core notation. For Mn^{2+}, we can write the electron configuration as $[Ar]\,3d^5$.

When writing the electron configuration for negative nonmetal ions, we must add the number of electrons that corresponds to the negative charge on the ion. For example, the electron configuration for chlorine is $[Ne]\,3s^2\,3p^5$. Because chloride ion is Cl^-, it gains one electron. That is,

$$\begin{array}{ccc} \text{Cl atom} & \xrightarrow{\text{gains } 1\ e^-} & Cl^- \\ [Ne]\,3s^2\,3p^5 & & [Ne]\,3s^2\,3p^6 \end{array}$$

We can write the electron configuration of Cl^- as [Ar]. The following example exercise further illustrates how to write the electron configuration of ions.

EXAMPLE EXERCISE 5.14 Electron Configuration of Ions

Refer to the periodic table and write the predicted electron configuration for each of the following ions using core notation:

(a) Fe^{3+} (b) Se^{2-}

Solution

We refer to the periodic table to recall the blocks of elements so that we can write the electron configuration for the element.

(a) The electron configuration for an Fe atom is [Ar] $4s^2 3d^6$. The core electron configuration for an Fe^{3+} ion is [Ar] $3d^5$.

(b) The electron configuration for a Se atom is [Ar] $4s^2 3d^{10} 4p^4$. The electron configuration for a Se^{2-} ion is [Ar] $4s^2 3d^{10} 4p^6$ or simply [Kr].

Practice Exercise

Refer to the periodic table and write the predicted electron configuration for each of the following ions using core notation:

(a) Cd^{2+} (b) P^{3-}

Answers:

(a) [Kr] $4d^{10}$

(b) [Ne] $3s^2 3p^6$, or [Ar]

Concept Exercise

What is the electron configuration for radon gas using core notation?

Answer: See Appendix G, 5.14.

CHEMISTRY CONNECTION Evolution of Chemical Elements

Q: All the chemical elements may be derived from which single element?

Although biologists are keenly interested in the "origin of species," chemists are equally curious about the origin of elements. The scientific community is not in total agreement, but in general, scientists believe that the chemical elements evolved from the simplest element hydrogen. This theory that the chemical elements evolved from hydrogen was first hypothesized in the nineteenth century, and coincided with Darwin's *On the Origin of Species* published in 1859, which set forth the principle of biological evolution.

The universe is calculated to be about 14 billion years old beginning with an event referred to as the "Big Bang." The Big Bang signaled the creation of our universe and was characterized by inconceivably high temperatures. Subsequently, the universe cooled rapidly and subatomic particles including protons, neutrons, and electrons began to form. There is evidence for both hydrogen and helium being formed in the Big Bang, but we also know that helium is currently being formed by our Sun through the process of nuclear fusion. That is, two hydrogen nuclei ($_1H$) fuse together to form one helium nucleus ($_2He$). It is reasonable to speculate that heavier nuclei can form through further nuclear fusion by a process termed nucleosynthesis.

There are a number of mechanisms that have been put forth to explain the creation of elements heavier than helium. For example, the fusion of hydrogen and helium can produce lithium ($_3Li$), which in turn can yield beryllium ($_4Be$), and eventually carbon ($_6C$). There is also evidence for the direct fusion of three helium nuclei ($_2He$) to give carbon ($_6C$). Although the study of chemical elements is clearly within the scope of chemists, the ultimate answer(s) to the evolution of the elements may come from astrophysics, the branch of astronomy that studies the physical nature of stars and other planetary bodies. Astrophysicists apply the laws of physics and chemistry to interpret astronomical observations.

◀ **Milky Way Galaxy** Gravity acting on debris from the Big Bang produced regions of high density including galaxies having numerous stars. Some of these stars simply burned out, whereas others reached high core temperatures capable of nucleosynthesis of heavier elements. Eventually, these stars collapsed and exploded creating supernovas and sending chemical elements throughout the universe.

A: Hydrogen was first formed from the Big Bang and, directly or indirectly, gave rise to the evolution of heavier chemical elements.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
5.1 Classification of Elements In 1869, Dmitri Mendeleev explained the recurring properties of more than 60 elements by arranging them according to increasing atomic mass. In 1871, Mendeleev published a revised periodic table of elements, which even included undiscovered elements and their predicted properties.	• Describe the original periodic law proposed by Mendeleev. ***Related Exercises 1–4***
5.2 The Periodic Law Concept In 1913, H. G. J. Moseley showed that the positive charge on the nucleus of each element in the periodic table increases progressively. As a result, the **periodic law** was rewritten, and it is now stated that physical and chemical properties repeat periodically when the elements are arranged in order of increasing atomic number.	• Describe the modern periodic law proposed by Moseley. ***Related Exercises 5–8***
5.3 Groups and Periods of Elements The periodic table is organized by vertical columns called **groups** and horizontal rows called **periods**. The main-group elements in Groups 1, 2, and 13–18 are referred to as **representative elements**, and the elements in Groups 3–12 are called **transition elements**. The elements in Groups 1, 2, 17, and 18 are referred to as the **alkali metals**, **alkaline earth metals**, **halogens**, and **noble gases**, respectively. The elements following lanthanum (Ce–Lu) are referred to as **lanthanides** and the elements following actinium (Th–Lr) are called **actinides**. Collectively, lanthanides and actinides are known as **inner transition elements**. The elements Sc, Y, La, and Ce–Lu are referred to as **rare earth elements**. The elements beyond uranium are called **transuranium elements**.	• Classify the elements according to their groups (families) and periods (series) in the periodic table. ***Related Exercises 9–20*** • Designate a group of elements in the periodic table using both the American convention (IA–VIIIA) and the IUPAC convention (1–18). ***Related Exercises 21–28***
5.4 Periodic Trends The periodic table is divided into metals on the left side and nonmetals on the right. Semimetals have intermediate properties and are found between the metals and the nonmetals. The atomic radius decreases up a group and from left to right. The metallic character also decreases up a group and from left to right.	• Explain the trend in metallic character within a group or period. ***Related Exercises 29–30, 33–34*** • Explain the trend in atomic size within a group or period. ***Related Exercises 31–32, 35–36***
5.5 Properties of Elements The trends in the periodic table enable us to predict physical and chemical properties. If we are given the atomic radius and oxide formula of two elements in a group, we can make a reasonable prediction for the atomic radius and oxide formula of another element in the same group.	• Predict a physical property for an element given the values of other elements in the same group. ***Related Exercises 37–40*** • Predict a chemical formula for a compound given the formulas of other compounds containing an element in the same group. ***Related Exercises 41–46***
5.6 Blocks of Elements The elements in the periodic table are arranged by sublevels according to increasing energy. We can easily predict the energy sublevel containing the outermost electrons by referring to the periodic table. We can write the **electron configuration** by noting the sublevels in the *s*, *p*, *d*, and *f* blocks of elements.	• Predict the highest energy sublevel for an element given its position in the periodic table. ***Related Exercises 47–54*** • Predict the electron configuration for an element given its position in the periodic table. ***Related Exercises 55–56***
5.7 Valence Electrons We can use the periodic table to quickly predict the number of electrons in the outermost *s* and *p* sublevels of an element, that is, the number of **valence electrons**. The number of valence electrons corresponds to the group number in the periodic table. For example, a Group IA/1 element has 1 valence electron, and a Group VIIA/17 element has 7 valence electrons.	• Predict the number of valence electrons for a representative element. ***Related Exercises 57–60***

Key Concepts	Learning Objectives and Related Exercises
5.8 Electron Dot Formulas We can diagram an atom by letting the symbol of the element represent the **core electrons** and drawing dots to represent the valence electrons. This diagram is referred to as the **electron dot formula** of the element.	- Diagram the electron dot formula for any representative element. ***Related Exercises 61–62***
5.9 Ionization Energy When an element loses or gains electrons, the resulting positively or negatively charged atoms are called **ions**. The amount of energy required to remove an electron from an atom is called the **ionization energy**. Because each noble gas has its valence shell filled, its ionization energy is extremely high. Conversely, each alkali metal has a low ionization energy because it has to lose only one valence electron to acquire a noble gas electron structure. Ionization energy increases up a group and from left to right in the periodic table.	- Describe the general trends of ionization energy in the periodic table. ***Related Exercises 63–64*** - Predict the group with the highest and the lowest ionization energy. ***Related Exercises 65–66*** - Predict the element in a pair having a higher ionization energy. ***Related Exercises 67–70***
5.10 Ionic Charges We can predict the **ionic charge** for a representative element from its group number. Metals lose valence electrons and nonmetals gain electrons to obtain a noble gas structure. For example, potassium is in Group IA/1 and loses 1 valence electron to form K^+. Sulfur is in Group VIA/16, has 6 valence electrons, and gains 2 electrons to form S^{2-}. The resulting K^+ and S^{2-} ions each have 18 electrons and are said to be **isoelectronic**. To write the electron configuration for an ion, we add or subtract electrons corresponding to the ionic charge. For the magnesium ion, Mg^{2+}, we subtract 2 electrons, and for the fluoride ion, F^-, we add 1 electron. Because Mg^{2+} and F^- each have 10 electrons, they are isoelectronic, and we can write their electron configurations as $1s^2\ 2s^2\ 2p^6$ or simply in **core notation** as [Ne].	- Predict an ionic charge for a representative element. ***Related Exercises 71–76*** - Diagram the predicted electron configuration for selected ions. ***Related Exercises 77–80***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** the properties of elements recur in a repeating pattern when arranged by increasing atomic number
_____ **2.** a vertical column of elements in the periodic table
_____ **3.** a horizontal row of elements in the periodic table
_____ **4.** the Group IA/1 elements, excluding hydrogen
_____ **5.** the Group IIA/2 elements
_____ **6.** the Group VIIA/17 elements
_____ **7.** the relatively unreactive Group VIIIA/18 elements
_____ **8.** the Group A (1, 2, and 13–18) elements in the periodic table
_____ **9.** the Group B (3–12) elements in the periodic table
_____ **10.** the elements with atomic numbers 58–71
_____ **11.** the elements with atomic numbers 90–103
_____ **12.** the elements in the lanthanide and actinide series
_____ **13.** the elements with atomic numbers 21, 39, 57, and 58–71
_____ **14.** the elements beyond atomic number 92
_____ **15.** a shorthand description of the arrangement of electrons by sublevels according to increasing energy
_____ **16.** the inner electrons in an atom which are not available for bonding
_____ **17.** the electrons that occupy the outermost *s* and *p* sublevels of an atom

(a) actinide series (*Sec. 5.3*)
(b) alkali metals (*Sec. 5.3*)
(c) alkaline earth metals (*Sec. 5.3*)
(d) core electrons (*Sec. 5.8*)
(e) core notation (*Sec. 5.10*)
(f) electron configuration (*Sec. 5.6*)
(g) electron dot formula (*Sec. 5.8*)
(h) group (*Sec. 5.3*)
(i) halogens (*Sec. 5.3*)
(j) inner transition elements (*Sec. 5.3*)
(k) ion (*Sec. 5.9*)
(l) ionic charge (*Sec. 5.10*)
(m) ionization energy (*Sec. 5.9*)
(n) isoelectronic (*Sec. 5.10*)
(o) lanthanide series (*Sec. 5.3*)
(p) noble gases (*Sec. 5.3*)

_____ **18.** the chemical symbol of an element surrounded by a dot representing each valence electron

_____ **19.** an atom that bears a charge as a result of gaining or losing valence electrons

_____ **20.** the amount of energy necessary to remove an electron from an atom

_____ **21.** refers to the positive charge on a metal atom that has lost electrons or to the negative charge on a nonmetal atom that has gained electrons

_____ **22.** refers to ions having the same electron configuration; for example, Mg^{2+} and O^{2-} each have 10 electrons

_____ **23.** a method of writing an electron configuration in which core electrons are represented by a noble gas symbol in brackets followed by the valence electrons, for example, $[Ne]\,3s^2$

(q) period (*Sec. 5.3*)
(r) periodic law (*Sec. 5.2*)
(s) rare earth elements (*Sec. 5.3*)
(t) representative elements (*Sec. 5.3*)
(u) transition elements (*Sec. 5.3*)
(v) transuranium elements (*Sec. 5.3*)
(w) valence electrons (*Sec. 5.7*)

Exercises

Answers to odd-numbered Exercises are in Appendix I.

Classification of Elements (Sec. 5.1)

1. Why did Mendeleev not include neon in his periodic table of 1871?
2. Why did Mendeleev not include argon in his periodic table of 1871?
3. Why did Mendeleev not include mendelevium in his periodic table of 1871?
4. Why did Mendeleev not include einsteinium in his periodic table of 1871?

▲ This Russian stamp depicts Mendeleev predicting the existence of gallium and indium in 1869. Aluminum had been discovered in 1825, but the existence of Ga and In were not discovered until a few years later.

The Periodic Law Concept (Sec. 5.2)

5. Before Moseley's discovery in 1913, the periodic law stated that physical and chemical properties tend to repeat periodically when elements are arranged according to what trend?
6. After Moseley's discovery in 1913, the periodic law stated that physical and chemical properties tend to repeat periodically when elements are arranged according to what trend?
7. Which two elements in the fourth period violate the original periodic law as stated by Mendeleev?
8. Which two elements in the fifth period violate the original periodic law as stated by Mendeleev?

Groups and Periods of Elements (Sec. 5.3)

9. Vertical columns in the periodic table are referred to by what two terms?
10. Horizontal rows in the periodic table are referred to by what two terms?
11. What is the term for the elements that belong to Groups IA–VIIIA (Groups 1, 2, and 13–18)?
12. What is the term for the elements that belong to Groups IIIB–IIB (Groups 3–12)?
13. What is the term for the elements in the series that follows element 57?
14. What is the term for the elements in the series that follows element 89?
15. What is the term for the series of elements that include Ce–Lu?
16. What is the term for the series of elements that include Th–Lr?
17. What type of elements (metal, nonmetal) are on the left side of the periodic table?
18. What type of elements (metal, nonmetal) are on the right side of the periodic table?
19. Which two elements are rare earth elements, but not lanthanides?
20. Which two elements are radioactive semimetals?
21. Identify the group number corresponding to each of the following families of elements.
    (a) alkali metals **(b)** alkaline earth metals
    (c) halogens **(d)** noble gases
22. Identify the group number corresponding to each of the following families of elements.
    (a) boron group **(b)** oxygen group
    (c) nickel group **(d)** copper group
23. According to IUPAC convention, what is the designation for each of the following groups of elements?
    (a) Group IA **(b)** Group IIIA
    (c) Group VA **(d)** Group VIIA
    (e) Group IB **(f)** Group IIIB
    (g) Group VB **(h)** Group VIIB

24. According to the American convention, what is the designation for each of the following groups of elements?

(a) Group 2 **(b)** Group 4
(c) Group 6 **(d)** Group 8
(e) Group 11 **(f)** Group 12
(g) Group 16 **(h)** Group 18

25. Refer to the periodic table and select the symbol of the element that fits each of the following descriptions.

(a) the semimetal in the second period
(b) the semimetal in the fourth period and Group IVA/14
(c) the halogen that exists as yellow gas
(d) the halogen that is radioactive

26. Refer to the periodic table and select the symbol of the element that fits each of the following descriptions.

(a) the semimetal in the third period
(b) the semimetal in the fourth period and Group VA/15
(c) the halogen that exists as a reddish-brown liquid
(d) the noble gas that is radioactive

27. Refer to the periodic table and select the symbol of the element that fits each of the following descriptions.

(a) the third–period alkali metal
(b) the third–period alkaline earth metal
(c) the rare earth element with the highest atomic mass
(d) the lanthanide that is radioactive

28. Refer to the periodic table and select the symbol of the element that fits each of the following descriptions.

(a) the fourth–period alkali metal
(b) the fourth–period alkaline earth metal
(c) the rare earth element with the lowest atomic mass
(d) the lanthanide that does not occur naturally

Periodic Trends (Sec. 5.4)

29. According to the general trend, metallic character (increases/decreases) proceeding up a group of elements in the periodic table.

30. According to the general trend, metallic character (increases/decreases) proceeding from left to right in the periodic table.

31. According to the general trend, the atomic radius (increases/decreases) proceeding up a group of elements in the periodic table.

32. According to the general trend, the atomic radius (increases/decreases) proceeding from left to right in the periodic table.

33. According to general trends in the periodic table, predict which element in each of the following pairs has the greater metallic character.

(a) K or Ca **(b)** Mg or Al
(c) Fe or Cu **(d)** S or Ar

34. According to general trends in the periodic table, predict which element in each of the following pairs has the greater metallic character.

(a) B or Al **(b)** Na or K
(c) Mg or Ba **(d)** H or Fe

35. According to general trends in the periodic table, predict which element in each of the following pairs has the larger atomic radius.

(a) Li or Na **(b)** N or P
(c) Mg or Ca **(d)** Ar or Kr

36. According to general trends in the periodic table, predict which element in each of the following pairs has the larger atomic radius.

(a) Rb or Sr **(b)** As or Se
(c) Pb or Bi **(d)** I or Xe

Properties of Elements (Sec. 5.5)

37. Predict the missing value (?) for radioactive francium (Fr). The atomic radius, density, and melting point are given for elements in Group IA/1.

Element	Atomic Radius	Density at 20 °C	Melting Point
Rb	248 pm	1.53 g/mL	39.3 °C
Cs	266 pm	1.93 g/mL	28.4 °C
Fr	(?) pm	(?) g/mL	(?) °C

38. Predict the missing value (?) for radioactive radium (Ra). The atomic radius, density, and melting point are given for elements in Group IIA/2.

Element	Atomic Radius	Density at 20 °C	Melting Point
Sr	215 pm	2.63 g/mL	769 °C
Ba	217 pm	3.65 g/mL	725 °C
Ra	(?) pm	(?) g/mL	(?) °C

39. Predict the missing value (?) for radioactive astatine (At). The atomic radius, density, and boiling point are given for elements in Group VIIIA/17.

Element	Atomic Radius	Density at 20 °C	Boiling Point
Br	115 pm	3.10 g/L	59 °C
I	133 pm	4.93 g/L	184 °C
At	(?) pm	(?) g/L	(?) °C

40. Predict the missing value (?) for radioactive radon (Rn). The atomic radius, density, and boiling point are given for elements in Group VIIIA/18.

Element	Atomic Radius	Density at STP	Boiling Point
Kr	202 pm	3.74 g/L	−152 °C
Xe	210 pm	5.86 g/L	−107 °C
Rn	(?) pm	(?) g/L	(?) °C

41. The formulas for the chlorides of potassium, calcium, boron, and germanium are, respectively, KCl, $CaCl_2$, BCl_3, and $GeCl_4$. Using the periodic table, predict the chemical formulas for each of the following similar compounds.

(a) potassium fluoride **(b)** calcium fluoride
(c) boron bromide **(d)** germanium iodide

42. The formulas for the oxides of sodium, magnesium, aluminum, and silicon are, respectively, Na_2O, MgO, Al_2O_3, and SiO_2. Using the periodic table, predict the chemical formulas for each of the following similar compounds.

(a) lithium oxide **(b)** barium oxide
(c) gallium oxide **(d)** tin oxide

43. The chemical formula for barium chloride is $BaCl_2$. Predict the formulas for each of the following similar compounds.

(a) calcium fluoride **(b)** calcium chloride
(c) calcium bromide **(d)** calcium iodide

44. The chemical formula for zinc oxide is ZnO. Predict the formulas for each of the following similar compounds.
(a) cadmium oxide **(b)** cadmium sulfide
(c) cadmium selenide **(d)** cadmium telluride

45. Depending upon reaction conditions, nitrogen and oxygen can produce either N_2O_3 or N_2O_5. Predict two formulas for each of the following compounds.
(a) arsenic oxide **(b)** antimony oxide

46. Depending upon reaction conditions, phosphorus and sulfur can produce either P_2S_3 or P_2S_5. Predict two formulas for each of the following compounds.
(a) arsenic sulfide **(b)** antimony sulfide

Blocks of Elements (Sec. 5.6)

47. What type of energy sublevel (*s*, *p*, *d*, *f*) is being filled by the elements in Groups IA/1 and IIA/2?

48. What type of energy sublevel (*s*, *p*, *d*, *f*) is being filled by the elements in Groups IIIA/13 through VIIIA/18?

49. What type of energy sublevel (*s*, *p*, *d*, *f*) is being filled by the transition elements?

50. What type of energy sublevel (*s*, *p*, *d*, *f*) is being filled by the inner transition elements?

51. Which energy sublevel is being filled by the lanthanide series?

52. Which energy sublevel is being filled by the actinide series?

53. Refer to the periodic table and state the highest energy sublevel for each of the following elements.
(a) H **(b)** Na
(c) Sm **(d)** Br
(e) Sr **(f)** C
(g) Sn **(h)** Cs

54. Refer to the periodic table and state the highest energy sublevel for each of the following elements.
(a) He **(b)** K
(c) U **(d)** Pd
(e) Be **(f)** Co
(g) Si **(h)** Pt

55. Refer to the periodic table and write the predicted electron configuration for each of the following elements.
(a) Li **(b)** F
(c) Mg **(d)** P
(e) Ca **(f)** Mn
(g) Ga **(h)** Rb

56. Refer to the periodic table and write the predicted electron configuration for each of the following elements.
(a) B **(b)** Ti
(c) Na **(d)** O
(e) Ge **(f)** Ba
(g) Pd **(h)** Kr

Valence Electrons (Sec. 5.7)

57. State the number of valence electrons in each of the following groups as predicted from the periodic table.
(a) Group IA/1 **(b)** Group IIIA/13
(c) Group VA/15 **(d)** Group VIIA/17

58. State the number of valence electrons in each of the following groups as predicted from the periodic table.
(a) Group IIA/2 **(b)** Group IVA/14
(c) Group VIA/16 **(d)** Group VIIIA/18

59. State the number of valence electrons for each of the following elements.
(a) H **(b)** B
(c) N **(d)** F
(e) Ca **(f)** Si
(g) O **(h)** Ar

60. State the number of valence electrons for each of the following elements.
(a) He **(b)** Pb
(c) Se **(d)** Ne
(e) Cs **(f)** Ga
(g) Sb **(h)** Br

Electron Dot Formulas (Sec. 5.8)

61. Draw the electron dot formula for each of the following elements.
(a) H **(b)** B
(c) N **(d)** F
(e) Ca **(f)** Si
(g) O **(h)** Ar

62. Draw the electron dot formula for each of the following elements.
(a) He **(b)** Pb
(c) Se **(d)** Ne
(e) Cs **(f)** Ga
(g) Sb **(h)** Br

Ionization Energy (Sec. 5.9)

63. According to the general trend, the ionization energy for a group of elements (increases/decreases) proceeding up a group in the periodic table.

64. According to the general trend, the ionization energy for a period of elements (increases/decreases) proceeding from left to right in the periodic table.

65. Which group of elements has the lowest ionization energy?

66. Which group of elements has the highest ionization energy?

67. Refer to the periodic table and predict which element in each of the following pairs has the higher ionization energy.
(a) Mg or Ca **(b)** S or Se
(c) Sn or Pb **(d)** N or P

68. Refer to the periodic table and predict which element in each of the following pairs has the higher ionization energy.
(a) Ga or Ge **(b)** Si or P
(c) Br or Cl **(d)** As or Sb

69. Refer to the periodic table and predict which element in each of the following pairs has the lower ionization energy.
(a) Rb or Cs **(b)** He or Ar
(c) B or Al **(d)** F or I

70. Refer to the periodic table and predict which element in each of the following pairs has the lower ionization energy.
(a) Mg or Si **(b)** Pb or Bi
(c) Ca or Ga **(d)** P or Cl

Ionic Charges (Sec. 5.10)

71. State the predicted ionic charge of metal ions in each of the following groups of elements.
(a) Group IA/1 **(b)** Group IIA/2
(c) Group IIIA/13 **(d)** Group IVA/14

72. State the predicted ionic charge of nonmetal ions in each of the following groups of elements.
(a) Group IVA/14
(b) Group VA/15
(c) Group VIA/16
(d) Group VIIA/17

73. Write the ionic charge for each of the following ions as predicted from the group number in the periodic table.
(a) Be ion
(b) Sn ion
(c) P ion
(d) S ion

74. Write the ionic charge for each of the following ions as predicted from the group number in the periodic table.
(a) Cs ion
(b) Ga ion
(c) O ion
(d) I ion

75. Refer to the periodic table and predict which of the following ions are isoelectronic with the noble gas krypton.
(a) K^+
(b) Sr^{2+}
(c) Cl^-
(d) Se^{2-}

76. Refer to the periodic table and predict which of the following ions are isoelectronic with the noble gas argon.
(a) Al^{3+}
(b) Ca^{2+}
(c) S^{2-}
(d) N^{3-}

77. Refer to the periodic table and write the predicted electron configuration for each of the following positive ions using core notation.
(a) Mg^{2+}
(b) K^+
(c) Fe^{2+}
(d) Zr^{2+}

78. Refer to the periodic table and write the predicted electron configuration for each of the following positive ions using core notation.
(a) Ti^{2+}
(b) Zn^{2+}
(c) Y^{3+}
(d) Cs^+

79. Refer to the periodic table and write the predicted electron configuration for each of the following negative ions using core notation.
(a) F^-
(b) S^{2-}
(c) N^{3-}
(d) I^-

80. Refer to the periodic table and write the predicted electron configuration for each of the following negative ions using core notation.
(a) Br^-
(b) Te^{2-}
(c) As^{3-}
(d) O^{2-}

General Exercises

81. Use the American convention to designate the group number corresponding to each of the following groups listed by the European convention.
(a) Group IA
(b) Group IB
(c) Group IIIA
(d) Group IIIB

82. Use the IUPAC convention to designate the group number corresponding to each of the following groups listed by the European convention.
(a) Group IIA
(b) Group IIB
(c) Group IVA
(d) Group IVB

83. Predict the missing value (?) for each property listed below. The atomic radius, density, and boiling point are given for elements in Group VIB/6.

Element	Atomic Radius	Density at 20 °C	Melting Point
Cr	125 pm	7.14 g/mL	(?) °C
Mo	(?) pm	10.28 g/mL	2617 °C
W	137 pm	(?) g/mL	3410 °C

84. Predict the missing value (?) for each property listed below. The atomic radius, density, and melting point are given for elements in Group VIII/8.

Element	Atomic Radius	Density at 20 °C	Melting Point
Fe	126 pm	7.87 g/mL	(?) °C
Ru	(?) pm	12.45 g/mL	2334 °C
Os	135 pm	(?) g/mL	3033 °C

85. Refer to the periodic table and write the predicted electron configuration for each of the following elements using core notation.
(a) Sr
(b) Ru
(c) Sb
(d) Cs

86. Refer to the periodic table and write the predicted electron configuration for each of the following elements using core notation.
(a) W
(b) Bi
(c) Ra
(d) Ac

87. Explain why the ionization energy for hydrogen is much higher than that of other Group IA/1 elements.

88. Predict two ionic charges for hydrogen. Write the formulas of the two ions and explain the ionic charges.

Challenge Exercises

89. Examine Figure 5.1 and determine the name of the element that Mendeleev predicted before its discovery and called ekaboron.

90. Examine Figure 5.1 and determine the name of the element that Mendeleev predicted before its discovery and called ekaaluminum.

91. Explain why the ionization energy for the alkali metals is less than the ionization energy for the alkaline earth metals.

92. Explain why the ionization energy for aluminum, contrary to the general trend, is less than the ionization energy for magnesium.

Chapter 5 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following scientists contributed to the development of the periodic table of elements? (Sec. 5.1)
(a) J. W. Döbereiner
(b) Dmitri Mendeleev
(c) H. G. J. Moseley
(d) J. A. R. Newlands
(e) all of the above

2. The modern periodic law states that the properties of the elements repeat when the periodic table is arranged according to which of the following? (Sec. 5.2)
(a) increasing atomic number
(b) increasing mass number
(c) increasing atomic mass
(d) increasing isotopic mass
(e) none of the above

3. Which of the following is a third-period alkaline earth metal? (Sec. 5.3)
(a) Ca **(b)** K
(c) Mg **(d)** Na
(e) Sr

4. Which of the following elements has the smallest atomic radius? (Sec. 5.4)
(a) Li **(b)** Na
(c) Be **(d)** Mg
(e) Al

5. Calculate the predicted atomic radius for selenium, Se, given the atomic radius of sulfur, S, (104 pm) and tellurium, Te, (143 pm). (Sec. 5.5)
(a) 39 pm **(b)** 62 pm
(c) 124 pm **(d)** 163 pm
(e) 182 pm

6. Which energy sublevel is being filled by the elements Sc to Zn? (Sec. 5.6)
(a) $3d$ **(b)** $4s$
(c) $4p$ **(d)** $4d$
(e) $4f$

7. Predict the number of valence electrons for a Group VIIA/17 element. (Sec. 5.7)
(a) 1 **(b)** 2
(c) 3 **(d)** 7
(e) 17

8. What is the electron dot formula for an atom of neon? (Sec. 5.8)
(a) $\mathrm{Ne}\cdot$ **(b)** $\cdot\underset{\cdot}{\mathrm{Ne}}\cdot$
(c) $\cdot\overset{\cdot}{\underset{\cdot}{\mathrm{Ne}}}\cdot$ **(d)** $:\overset{\cdot}{\underset{\cdot\cdot}{\mathrm{Ne}}}:$
(e) $:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Ne}}}:$

9. Which of the following elements has the highest ionization energy? (Sec. 5.9)
(a) H **(b)** He
(c) Ne **(d)** Cl
(e) F

10. Which of the following groups has a predictable ionic charge of three negative? (Sec. 5.10)
(a) Group IIIA/13
(b) Group IIIB/3
(c) Group VA/15
(d) Group VB/5
(e) Group VIIIA/18

Key Concepts

11. State the trends in the periodic table for decreasing metallic character.

12. State the trends in the periodic table for decreasing atomic radii.

13. State the trends in the periodic table for increasing ionization energy.

14. Given the chemical formulas for Al_2O_3, SiO_2, and P_2O_5, predict the formula for arsenic oxide, $As_?O_?$.

Critical Thinking

15. Can ions of different elements have the same number of electrons?

16. Can ions of different elements have the same number of protons?

17. Can ions of different elements have the same number of neutrons?

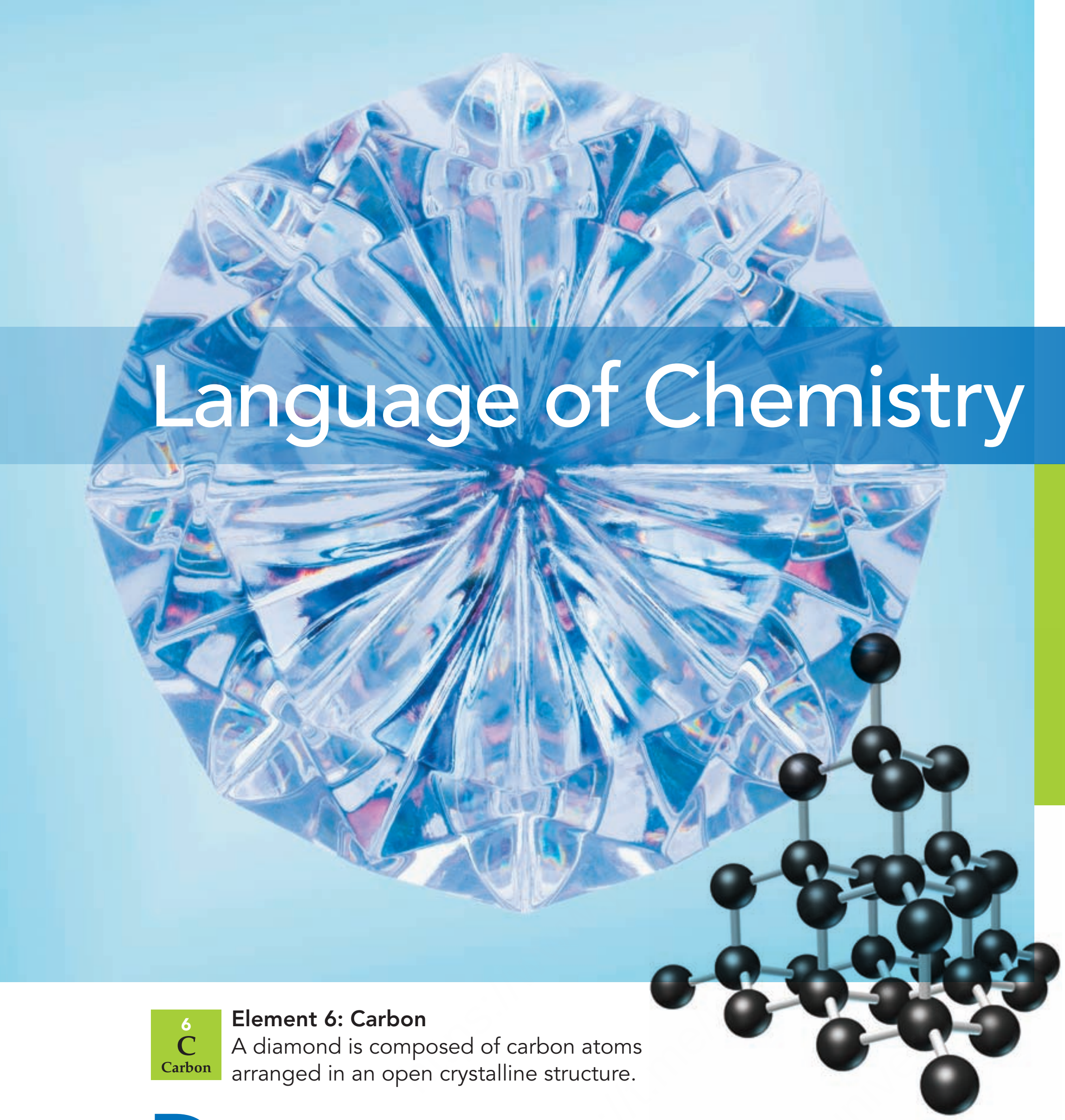
18. The following photograph shows a semimetal (Sb), nonmetal (S), and compound, (Sb_2S_5). Identify the (a) yellow, (b) gray, and (c) orange substances.

CHAPTER 6

Language of Chemistry

"Science is organized knowledge. Wisdom is organized life."

Immanuel Kant, German Philosopher (1724–1804)



6 C Carbon

Element 6: Carbon A diamond is composed of carbon atoms arranged in an open crystalline structure.

During the Middle Ages, alchemists identified many substances that they named in a haphazard fashion. There were only a few dozen substances, so this did not pose a great problem. Toward the end of the 1700s, however, they had identified more than 10,000 substances, and the number of compounds was growing rapidly. Chemists then faced the staggering task of providing names for all these substances. The problem of assigning names was eventually solved by using a set of systematic rules.

The French chemist Antoine Lavoisier developed the first systematic method of naming substances. He proposed that chemical names refer to the composition of the compound and be derived from Latin or Greek. In 1787, Lavoisier, with the aid of others, published *Methods of Chemical Nomenclature*. The naming system he proposed was so clear and logical that it was universally accepted in a short time. In fact, it was so simple and insightful that it became the basis for our current system of naming.

6.1 Classification of Compounds

In 1921, the International Union of Pure and Applied Chemistry (IUPAC) formed the Commission on the Nomenclature of Inorganic Chemistry. In 1938, the IUPAC Committee on the Reform of Inorganic Nomenclature met in Berlin. Two years later it released a comprehensive set of rules. Although the rules have been expanded and revised, the 1940 rules remain as the official international system for naming chemical compounds. These rules are referred to as **IUPAC nomenclature**.

LEARNING OBJECTIVES

- Classify a compound as a binary ionic compound, a ternary ionic compound, or a binary molecular compound.
- Classify an acid as a binary acid or a ternary oxyacid.
- Classify an ion as a monoatomic cation, a monoatomic anion, a polyatomic cation, or a polyatomic anion.

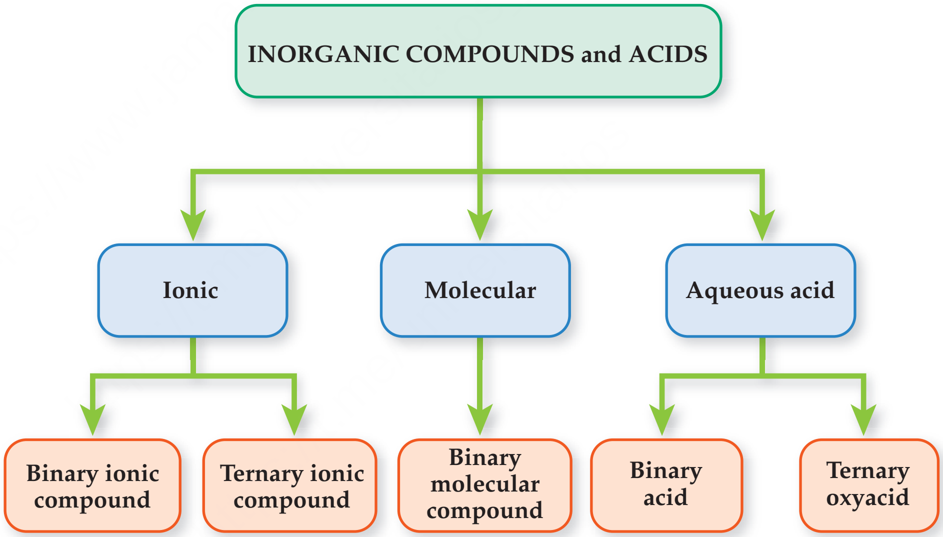
Classifying Inorganic Compounds

An **inorganic compound** does not contain the element carbon. However, carbon dioxide (CO_2), and compounds containing carbonate (CO_3^{2-}) or hydrogen carbonate (HCO_3^-) are exceptions and are considered inorganic. According to the 1940 IUPAC rules, inorganic compounds can be placed in one of several categories. Five common classes are binary ionic, ternary ionic, binary molecular, binary acid, and ternary oxyacid.

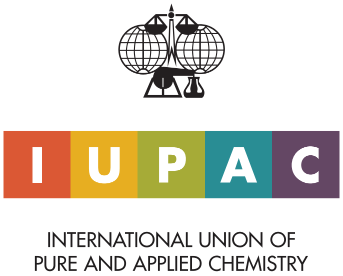
A **binary ionic compound** contains two elements, a metal and a nonmetal. Examples of binary ionic compounds are KCl and $AlCl_3$. A **ternary ionic compound** contains three elements with at least one metal and one nonmetal. Examples of ternary ionic compounds are KNO_3 and $Al(NO_3)_3$. A **binary molecular compound** contains two elements that are both nonmetals. Water, H_2O, is a common example, as is ammonia, NH_3.

An **aqueous solution** is produced when a compound dissolves in water. Aqueous solutions are indicated by the symbol (*aq*). A **binary acid** is an aqueous solution of a compound containing hydrogen and one other nonmetal. Formulas of acids begin with H, and examples of binary acids are HCl(*aq*) and H_2S(*aq*).

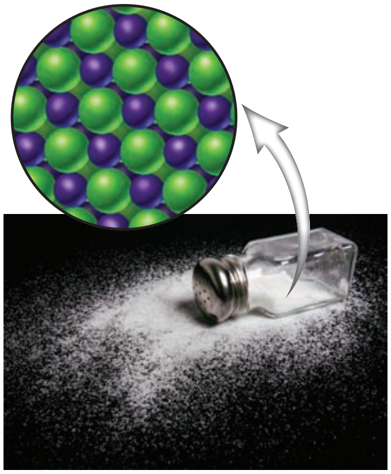
A **ternary oxyacid** is an aqueous solution of a compound containing hydrogen, a nonmetal, and oxygen. Examples of ternary oxyacids are HNO_3(*aq*) and H_2SO_4(*aq*). Figure 6.1 illustrates the relationships among the five types of compounds we will classify according to IUPAC rules.



▶ **Figure 6.1 Classification of Inorganic Compounds and Acids** According to IUPAC nomenclature, inorganic compounds and acids are divided into five categories: binary ionic, ternary ionic, binary molecular, binary acid, and ternary oxyacid.



▲ Logo for the International Union of Pure and Applied Chemistry.



◀ **Table Salt** Sodium chloride, NaCl, is an example of a binary ionic compound composed of metal and nonmetal ions.

The following example exercise illustrates the classification of different types of compounds and aqueous acids.

EXAMPLE EXERCISE 6.1 Classifying Compounds and Acids

Classify each of the following as a binary ionic compound, ternary ionic compound, binary molecular compound, binary acid, or ternary oxyacid:

(a) calcium oxide, CaO
(b) sulfur dioxide, SO_2
(c) silver chromate, Ag_2CrO_4
(d) hydrofluoric acid, $HF(aq)$
(e) carbonic acid, $H_2CO_3(aq)$

Solution

We can refer to the periodic table and classify each compound or solution as follows:

(a) CaO contains two elements, a metal and nonmetal. Thus, CaO is a *binary ionic compound.*
(b) SO_2 contains two elements, both nonmetals. Thus, SO_2 is a *binary molecular compound.*
(c) Ag_2CrO_4 contains three elements, two metals and a nonmetal. Thus, Ag_2CrO_4 is a *ternary ionic compound*.
(d) $HF(aq)$ is a compound of hydrogen and a nonmetal dissolved in water. Thus, $HF(aq)$ is a *binary acid.*
(e) $H_2CO_3(aq)$ is a compound containing three elements, including hydrogen and oxygen, dissolved in water. Thus, $H_2CO_3(aq)$ is a *ternary oxyacid.*

Practice Exercise

Classify each of the following as a binary ionic compound, ternary ionic compound, binary molecular compound, binary acid, or ternary oxyacid:

(a) carbon disulfide, CS_2
(b) lithium dichromate, $Li_2Cr_2O_7$
(c) magnesium iodide, MgI_2
(d) nitric acid, $HNO_3(aq)$
(e) hydrochloric acid, $HCl(aq)$

Answers:

(a) binary molecular compound
(b) ternary ionic compound
(c) binary ionic compound
(d) ternary oxyacid
(e) binary acid

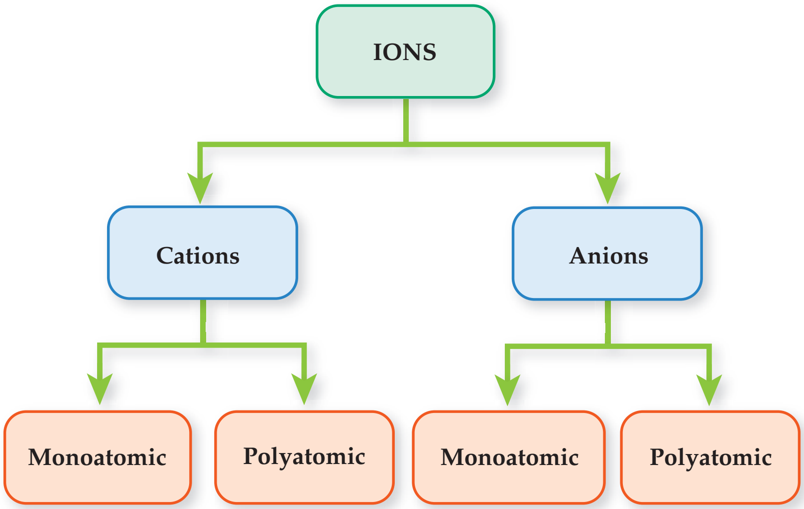
Concept Exercise

How do you distinguish a binary molecular compound from a binary acid?

Answer: See Appendix G, 6.1.

Classifying Ions

According to IUPAC nomenclature, ions are named systematically, depending on the category in which they are placed. A positive ion is referred to as a **cation**, and a negative ion is referred to as an **anion**. A single atom of a metal with a positive charge, or a single atom of a nonmetal with a negative charge, is called a **monoatomic ion**. A group of atoms bound together and having an overall positive or negative charge is called a **polyatomic ion**. Figure 6.2 illustrates the classification of and relationship of these ions.



▶ **Figure 6.2 Classification of Ions** According to the IUPAC nomenclature, ions are divided into four categories: monoatomic cations, polyatomic cations, monoatomic anions, and polyatomic anions.

The following example exercise illustrates the classification of different types of ions:

EXAMPLE EXERCISE 6.2 Classifying Cations and Anions

Classify each of the following ions as a monoatomic cation, monoatomic anion, polyatomic cation, or polyatomic anion:

(a) barium ion, Ba^{2+}
(b) chloride ion, Cl^-
(c) ammonium ion, NH_4^+
(d) carbonate ion, CO_3^{2-}

Solution

We can classify each ion as follows:

(a) Ba^{2+} is a single atom with a positive charge. Thus, Ba^{2+} is a *monoatomic cation.*
(b) Cl^- is a single atom with a negative charge. Thus, Cl^- is a *monoatomic anion.*
(c) NH_4^+ has five atoms and a positive charge. Thus, NH_4^+ is a *polyatomic cation.*
(d) carbonate ion, CO_3^{2-} has four atoms and a negative charge. Thus, carbonate ion, CO_3^{2-} is a *polyatomic anion.*

Practice Exercise

Classify each of the following ions as a monoatomic cation, monoatomic anion, polyatomic cation, or polyatomic anion:

(a) hydronium ion, H_3O^+
(b) sulfide ion, S^{2-}
(c) permanganate ion, MnO_4^-
(d) stannous ion, Sn^{2+}

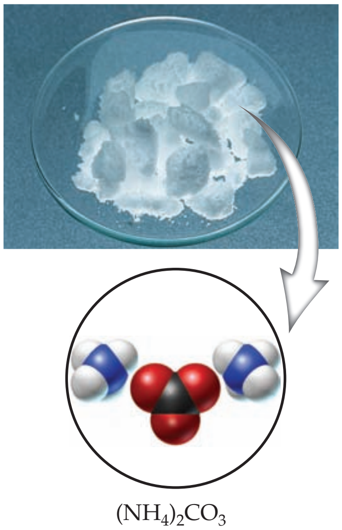
Answers:

(a) polyatomic cation
(b) monoatomic anion
(c) polyatomic anion
(d) monoatomic cation

Concept Exercise

What is the distinction between Hg^{2+} and Hg_2^{2+}?

Answer: See Appendix G, 6.2.



▲ **Ammonium Carbonate, $(NH_4)_2CO_3$** Ammonium carbonate decomposes in air to give ammonia, NH_3, and carbon dioxide, CO_2; it is used in smelling salts and is found in baking powder.

FEATURED SCIENTIST Antoine and Marie-Anne Lavoisier

Q: Who provided invaluable laboratory assistance to Lavoisier, translated the experimental results into other languages, and helped spread the enormous body of work throughout Europe and America?

Antoine Lavoisier (1743–1794) was born into an affluent French family and received an excellent education. He initially became a lawyer like his father, but later turned to science. Lavoisier was a gifted experimenter and is generally considered the founder of modern chemistry. He established a magnificent laboratory that attracted scientists from around the world.

In 1787, Lavoisier published a book entitled *Methods of Chemical Nomenclature*. The book designated systematic principles for naming chemical substances and was so logical that it became the basis for our present rules of nomenclature. In 1789, he published the landmark textbook *Elementary Treatise on Chemistry*, which offered the first modern view of chemistry.

In 1771, Lavoisier married Marie-Anne Pierrette Paulze (1758–1836). Marie was only 13 years old and she had been raised and educated in a convent after her mother died. Marie's father arranged the marriage in order to thwart a proposal by an influential nobleman who was more than three times her age. Despite her youth, she proved invaluable to Lavoisier's work and joined him in the laboratory. Marie recorded the experimental work through her drawings and notes, and translated the results from French into other languages. Owing to their close working relationship, it is impossible to assign individual credit and separate their joint contribution to the advancement of chemistry.

Unfortunately, Marie's father was an important official in a firm that collected taxes. Although Lavoisier did not collect taxes, he received royalties from the same firm. When the French Revolution broke out, revolutionaries threw tax collectors into jail. Lavoisier was first barred from his laboratory and then arrested. After a farce trial, he was sentenced to the guillotine, along with his father-in-law and other tax collectors.

On May 8, 1794, Lavoisier was executed and buried in an unmarked grave. The French court declared that France had no need for scientists, but the famous mathematician Lagrange lamented, "Only a moment was required to strike off his head, and a hundred years may not be enough to produce another like it again." Within two years, the French people regretted their hasty decision and began unveiling statues of Lavoisier.

◀ **Antoine Lavoisier** Antoine Lavoisier is shown with Marie-Anne in their laboratory.

A: Lavoisier's multilingual wife, Marie-Anne, is recognized for keeping laboratory records and translating scientific papers into other languages.

6.2 Monoatomic Ions

LEARNING OBJECTIVES

- Provide systematic names and formulas for common monoatomic ions.
- Predict the ionic charge for ions of representative elements.

Metal atoms can lose valence electrons and become positively charged cations; examples include Na^+, Ca^{2+}, and Al^{3+}. Conversely, nonmetal atoms can gain valence electrons and become negatively charged anions; examples include, Cl^-, S^{2-}, and P^{3-}.

Naming Metal Ions

According to IUPAC nomenclature rules, cations are named for the parent metal followed by the word *ion*. For example, Na^+ is named sodium ion, Ca^{2+} is named calcium ion, and Al^{3+} is named aluminum ion.

Main-group metals usually form one cation; tin and lead are exceptions. Transition metals, however, often form more than one cation. Iron, for example, can form Fe^{2+} and Fe^{3+}. To name a metal cation having more than one ionic charge, it is necessary to specify the charge. IUPAC recommends that the cation be named for the parent metal followed by its charge specified by Roman numerals in parentheses. Thus, Fe^{2+} is named iron(II) ion, and Fe^{3+} is named iron(III) ion. Similarly, Cu^{+} is named copper(I) ion and Cu^{2+} is named copper(II) ion. This method of naming transition metal cations is called the **Stock system**. We should note that the names of Ag^{+}, Zn^{2+}, and Cd^{2+} are transition metal exceptions, which do not require Roman numerals.

The formula for the mercury(I) ion is an exception. When mercury loses an electron, the resulting ion, Hg^{+}, becomes more stable by combining with another Hg^{+} ion. Thus, the mercury(I) ion is written Hg_2^{2+}. Table 6.1 lists the systematic names for common cations using the Stock system of naming.

TABLE 6.1 Common Monoatomic Cations

Cation	Stock System
Al^{3+}	aluminum ion
Ba^{2+}	barium ion
Cd^{2+}	cadmium ion
Ca^{2+}	calcium ion
Co^{2+}	cobalt(II) ion
Co^{3+}	cobalt(III) ion
Cu^{+}	copper(I) ion
Cu^{2+}	copper(II) ion
Cr^{3+}	chromium(III) ion
H^{+}	hydrogen ion
Fe^{2+}	iron(II) ion
Fe^{3+}	iron(III) ion
Pb^{2+}	lead(II) ion
Pb^{4+}	lead(IV) ion
Li^{+}	lithium ion
Mg^{2+}	magnesium ion
Mn^{2+}	manganese(II) ion
Hg_2^{2+}	mercury(I) ion*
Hg^{2+}	mercury(II) ion
Ni^{2+}	nickel(II) ion
K^{+}	potassium ion
Ag^{+}	silver ion
Na^{+}	sodium ion
Sr^{2+}	strontium ion
Sn^{2+}	tin(II) ion
Sn^{4+}	tin(IV) ion
Zn^{2+}	zinc ion

*Note that the mercury(I) ion is diatomic and is written Hg_2^{2+}.

Naming Nonmetal Ions

As we now know, nonmetal atoms can gain valence electrons and become negatively charged ions called anions. According to IUPAC rules, nonmetal ions are named by using the *nonmetal stem* plus the suffix *–ide*. Examples of nonmetal ions are Cl^-, chloride ion; S^{2-}, sulfide ion; and P^{3-}, phosphide ion. Table 6.2 lists the systematic names for common anions.

TABLE 6.2 Common Monoatomic Anions

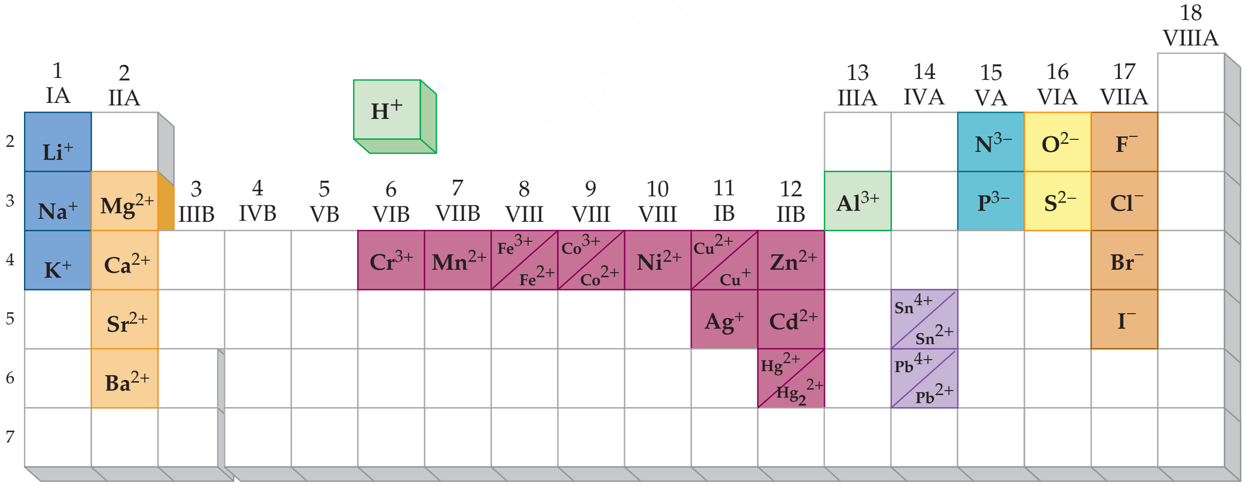
Anion	IUPAC Name	Anion	IUPAC Name
Br^-	bromide ion	N^{3-}	nitride ion
Cl^-	chloride ion	O^{2-}	oxide ion
F^-	fluoride ion	P^{3-}	phosphide ion
I^-	iodide ion	S^{2-}	sulfide ion

Predicting Formulas of Monoatomic Cations

We can use the periodic table to help learn the names and formulas of the ions and predict their charge based on the group number of the element. Recall that Group IA/1 metals always form 1+ ions. Thus, we have Li^+, Na^+, and K^+. Also recall that the Group IIA/2 elements always form 2+ ions. Thus, we have Mg^{2+}, Ca^{2+}, Sr^{2+}, and Ba^{2+}. Similarly, we can predict that aluminum in Group IIIA/13 forms a 3+ ion, that is, Al^{3+}.

Not all metal ions are predictable from the periodic table. The metals tin and lead and lead in Group IVA/14 can each form two ions: Sn^{2+} or Sn^{4+} and Pb^{2+} or Pb^{4+}. Therefore, you will have to memorize the formulas for these ions.

With few exceptions all transition elements have two *s* electrons as well as a variable number of *d* electrons. In general, a transition metal loses its two *s* electrons to form an ion with a 2+ charge, for example, Ni^{2+}. Many transition metals form additional ions having charges that are unpredictable. You will have to memorize the names of these ions. Figure 6.3 shows the relationship of ionic charge to the position of the element in the periodic table.



▲ Figure 6.3 **Periodic Table of Selected Ions** Note the correlation of ionic charge and group number. Although transition elements can exhibit more than one ionic charge, nearly all these metals exhibit a charge of 2+.

Although we should memorize the names and formulas of ions, the periodic table is a valuable resource for verifying ionic charges. The following example exercise illustrates the names and formulas of monoatomic cations:

EXAMPLE EXERCISE 6.3 Names and Formulas of Monoatomic Cations

Provide the formula for the following monoatomic cations:

(a) barium ion	(b) cobalt(II) ion

Solution

We can use the periodic table to predict the charge on a cation.

(a) Barium is found in Group IIA/2 and can lose two valence electrons. We predict the formula of the ion is Ba^{2+}.

(b) Cobalt is a transition metal and can lose two or more valence electrons. The Roman numeral (II) indicates the loss of two electrons, and so the formula of the ion is Co^{2+}.

Practice Exercise

Supply a systematic name for the following monoatomic cations:

(a) Al^{3+}	(b) Co^{3+}

Answers:

(a) aluminum ion	(b) cobalt(III) ion

Concept Exercise

How does the Stock system specify a variable charge on a cation?

Answer: See Appendix G, 6.3.

Predicting Formulas of Monoatomic Anions

Nonmetals gain electrons to form negative ions in a predictable fashion. The halogens in Group VIIA/17 need only one electron to become isoelectronic with a noble gas. Thus, the halogens have a 1− charge and the formulas F^-, Cl^-, Br^-, and I^-. The elements in Group VIA/16 gain two electrons to become isoelectronic with a noble gas. These elements have a 2− charge and the formulas O^{2-}, S^{2-}, and Se^{2-}. Nitrogen and phosphorus are in Group VA/15 and gain three electrons. These elements have a 3− charge and the formulas N^{3-} and P^{3-}.

EXAMPLE EXERCISE 6.4 Names and Formulas of Monoatomic Anions

Provide the formula for each of the following monoatomic anions:

(a) fluoride ion	(b) oxide ion

Solution

Recall that nonmetals gain electrons so as to acquire a noble gas electron configuration.

(a) Fluorine is found in Group VIIA/17 and can gain one electron to become isoelectronic with neon. We predict the formula for the fluoride ion is F^-.

(b) Oxygen is found in Group VIA/16 and can gain two electrons to become isoelectronic with neon. We predict the formula for the oxide ion is O^{2-}.

Practice Exercise

Supply a systematic name for each of the following monoatomic anions:

(a) Br^-	(b) N^{3-}

Answers:

(a) bromide ion	(b) nitride ion

Concept Exercise

Which one of the following is a monoatomic anion: iodide, iodate, or iodite?

Answer: See Appendix G, 6.4.

6.3 Polyatomic Ions

LEARNING OBJECTIVE

- Provide systematic names and formulas for common polyatomic ions.

Polyatomic anions generally contain one or more elements combined with oxygen. Such an anion is referred to as an **oxyanion**, a term that reflects its composition. Most oxyanions have names ending in the suffix *–ate*. Examples include the nitrate ion, NO_3^-, and the sulfate ion, SO_3^{2-}.

A few oxyanions have names ending in the suffix *–ite*. Examples include the nitrite ion, NO_2^-, and the sulfite ion, SO_3^{2-}. Notice that in each case the formula for ions with the suffix *–ite* has one less oxygen than that for ions with the suffix *–ate*.

This suffix pattern can provide clues when naming some polyatomic ions. Given the formula for the chlorate ion, ClO_3^-, we can predict the formula for the chlorite ion. Because the suffix *–ate* has changed to the suffix *–ite*, the formula has one less oxygen. Therefore, the formula for the chlorite ion is ClO_2^-. The general relationship between the suffixes *–ate* and *–ite* allows us to simplify our task of memorizing formulas of ions.

There are two important polyatomic ions that have the suffix *–ide*. The cyanide ion, CN^-, and the hydroxide ion, OH^-, are exceptions, and their names should be memorized. In addition to these two polyatomic anions, there is the ammonium cation, NH_4^+, whose formula is derived from that for ammonia gas, NH_3. Table 6.3 lists the common polyatomic ions whose formulas should be memorized.

TABLE 6.3 Common Polyatomic Ions

Cation		IUPAC Name	
NH_4^+		ammonium ion	
Anion	**IUPAC Name**	**Anion**	**IUPAC Name**
$C_2H_3O_2^-$	acetate ion	OH^-	hydroxide ion*
CO_3^{2-}	carbonate ion	ClO^-	hypochlorite ion
ClO_3^-	chlorate ion	NO_3^-	nitrate ion
ClO_2^-	chlorite ion	NO_2^-	nitrite ion
CrO_4^{2-}	chromate ion	ClO_4^-	perchlorate ion
CN^-	cyanide ion*	MnO_4^-	permanganate ion
$Cr_2O_7^{2-}$	dichromate ion	PO_4^{3-}	phosphate ion
HCO_3^-	hydrogen carbonate ion	SO_4^{2-}	sulfate ion
HSO_4^-	hydrogen sulfate ion	SO_3^{2-}	sulfite ion

*Note that the suffix –ide is an exception to the general –ate and –ite rule.

The following example exercise illustrates how general principles can be used to assist you in memorizing the names and formulas of polyatomic ions:

EXAMPLE EXERCISE 6.5 Names and Formulas of Polyatomic Ions

Provide a systematic name for each of the following polyatomic oxyanions:

(a) CO_3^{2-} (b) CrO_4^{2-} (c) ClO_2^- (d) HSO_4^-

Solution

We can make reasonable predictions for the names of many polyatomic ions. This makes the task of memorization much easier.

(a) CO_3^{2-} contains carbon, and we predict the name has the suffix *–ate*. Thus, we predict CO_3^{2-} is named the *carbonate ion*.

(b) CrO_4^{2-} contains chromium, and we predict the name has the suffix *–ate*. Thus, we predict CrO_4^{2-} is named the *chromate ion*.

(c) ClO_2^- is related to ClO_3^-, which is named the chlorate ion. Because ClO_2^- has one less oxygen atom, the suffix changes to *–ite*. Thus, we predict ClO_2^- is named the *chlorite ion*.

(d) HSO_4^- is related to the sulfate ion, SO_4^{2-}. With the addition of hydrogen, the name becomes the *hydrogen sulfate ion*.

Practice Exercise

Provide the formula for each of the following polyatomic oxyanions:

(a) acetate ion
(b) dichromate ion
(c) perchlorate ion
(d) hydrogen carbonate ion

Answers:

(a) $C_2H_3O_2^-$ (b) $Cr_2O_7^{2-}$ (c) ClO_4^- (d) HCO_3^-

Concept Exercise

How does systematic naming distinguish between a monoatomic anion and a polyatomic anion?

Answer: See Appendix G, 6.5.

Note You may find it useful to make flashcards to help you memorize ions. Write the name of the ion on one side of a card and the formula of the ion on the other side. Your task will be easier if you recall the following:

1. There is one common polyatomic cation, NH_4^+.
2. Most oxyanions have the suffix *–ate.*
3. The suffix *–ate* changes to the suffix *–ite* for an oxyanion with one less oxygen atom; for example, nitrate, NO_3^-, changes to nitrite, NO_2^-.
4. There are two common polyatomic anions that are exceptions: OH^- and CN^- each have the suffix *–ide.*

LEARNING OBJECTIVE

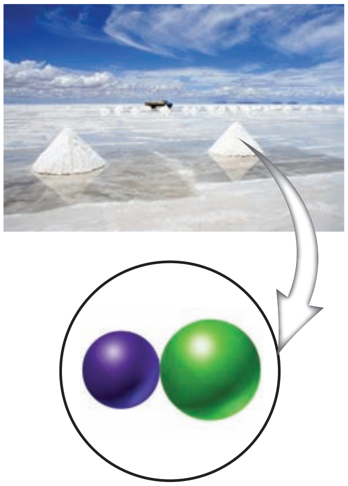
- Write chemical formulas for ionic compounds composed of monoatomic ions and polyatomic ions.

6.4 Writing Chemical Formulas

An ionic compound is composed of positive and negative ions. A **formula unit** is the simplest representative particle in an ionic compound. Because a formula unit is neutral, the total positive charge must equal the total negative charge in both the formula unit and the ionic compound. That is, the total positive charge from the metal ions must be the same as the total negative charge from the nonmetal ions.

Ordinary sodium chloride contains the ions Na^+ and Cl^-. The positive and negative ions have equal but opposite charges, so the formula for the compound is NaCl. Calcium chloride contains the ions Ca^{2+} and Cl^-. Because the charges are not the same, it is necessary to have two Cl^- ions for each Ca^{2+} ion; therefore, the formula is $CaCl_2$. Aluminum chloride contains the ions Al^{3+} and Cl^-, and it is necessary to have three Cl^- ions for each Al^{3+} ion. The formula for the compound is $AlCl_3$.

The following example exercise provides additional illustrations of how to write formulas for ionic compounds:



▲ **Salt, NaCl** Table salt is an example of a binary ionic compound composed of sodium ions and chloride ions.

EXAMPLE EXERCISE 6.6 Writing Formulas of Binary Ionic Compounds

Write the chemical formula for the following binary compounds given their constituent ions:

(a) copper(I) oxide, Cu^+ and O^{2-}
(b) cadmium oxide, Cd^{2+} and O^{2-}
(c) cobalt(III) oxide, Co^{3+} and O^{2-}

Solution

(a) The copper(I) ion has a charge of 1+, and the oxide ion has a charge of 2−. Thus, two positive ions are required for each negative ion in a neutral formula unit. The formula of copper(I) oxide is written Cu_2O.

(b) Because the cadmium ion and oxide ion each have a charge of 2, the ratio is 1:1, that is, Cd_1O_1. It is not necessary to write the subscript 1, and so the formula of cadmium oxide is simply CdO.

(c) This example is more difficult. The cobalt(III) ion has a charge of 3+ and the oxide ion has a charge of 2−. Because the lowest common multiple is 6, two 3+ ions are required to cancel the charge of three 2− ions. The ratio is 2:3, and the formula of cobalt(III) oxide is written Co_2O_3.

Practice Exercise

Write the chemical formula for the following binary compounds given their constituent ions:

(a) iron(II) sulfide, Fe^{2+} and S^{2-}
(b) mercury(I) fluoride, Hg_2^{2+} and F^-
(c) lead(IV) oxide, Pb^{4+} and O^{2-}

Answers:

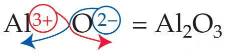
(a) FeS
(b) Hg_2F_2
(c) PbO_2

Concept Exercise

What is the formula for a metal oxide if the charge on the metal (M) is 3+?

Answer: See Appendix G, 6.6.

Note You can quickly verify that the chemical formula is written correctly by crossing over the charge on each ion. Consider aluminum oxide, which contains Al^{3+} and O^{2-}. The 3+ charge on the aluminum ion becomes the subscript 3 for the oxygen, and the 2− charge on the oxide ion becomes the subscript 2 for the aluminum ion. That is,



$$Al^{3+}\,O^{2-} = Al_2O_3$$

Formula Units Containing Polyatomic Ions

As mentioned earlier, formula units are the simplest particles representing an ionic compound, and the total positive charge is equal to the total negative charge. Previously, we learned to write neutral formula units for binary ionic compounds. Similarly, in a formula unit containing K^+ and SO_4^{2-}, two K^+ ions are required to balance the charge for each SO_4^{2-}. Thus, a neutral formula unit is written K_2SO_4.

Magnesium sulfate contains Mg^{2+} and SO_4^{2-}. Because the magnitude of the charge is the same on each ion, the ratio of positive ion to negative ion is 1:1. The formula unit is written $MgSO_4$.

Ammonium sulfate is a nitrogen-supplying component in fertilizer and contains NH_4^+ and SO_4^{2-} ions. Because the negative charge is greater for SO_4^{2-}, two NH_4^+ ions are necessary to give a neutral formula unit. To avoid misunderstanding, parentheses are placed around the NH_4^+ ion. The correct formula is written $(NH_4)_2SO_4$. The following example exercise provides additional illustrations of how to write formulas for ionic compounds:

EXAMPLE EXERCISE 6.7 Writing Formulas of Ternary Ionic Compounds

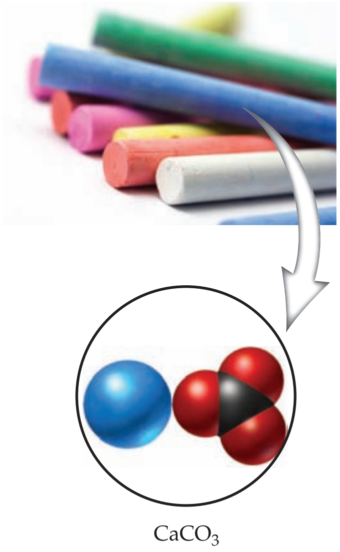
Write the chemical formula for each of the following ternary compounds given their constituent ions:

(a) calcium carbonate, Ca^{2+} and CO_3^{2-}
(b) calcium hydroxide, Ca^{2+} and OH^-
(c) calcium phosphate, Ca^{2+} and PO_4^{3-}

Solution

(a) Because the positive and negative ions each have a charge of 2, one positive ion and one negative ion are required to produce a neutral formula unit, and the formula is $CaCO_3$. Calcium carbonate occurs naturally as ordinary chalk.

(b) The positive ion has a charge of 2+, and the negative ion has a charge of 1−. Therefore, one positive ion and two negative ions are required to produce a neutral formula unit. Because OH^- is a polyatomic ion, parentheses are required, and the formula is written $Ca(OH)_2$. Calcium hydroxide is known as "slaked lime" and is sometimes used to mark the boundaries of an athletic field.



▲ **Chalk** White chalk is traditionally composed of calcium carbonate, $CaCO_3$; pastel chalk is composed of calcium sulfate, $CaSO_4$.

(c) The positive ion has a charge of 2+, and the negative ion has a charge of 3−. The lowest common multiple of the charges is 6. Three positive ions are required for every two negative ions to produce a neutral formula unit. A calcium phosphate formula unit is written $Ca_3(PO_4)_2$. Calcium phosphate is found in tooth enamel.

Practice Exercise

Write the chemical formula for each of the following ternary compounds given their constituent ions:

(a) copper(II) permanganate, Cu^{2+} and MnO_4^-
(b) iron(III) carbonate, Fe^{3+} and CO_3^{2-}
(c) potassium dichromate, K^+ and $Cr_2O_7^{2-}$

Answers:

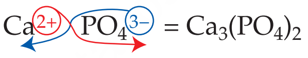
(a) $Cu(MnO_4)_2$ (b) $Fe_2(CO_3)_3$ (c) $K_2Cr_2O_7$

Concept Exercise

What is the formula for a metal carbonate if the charge on the metal (M) is 3+?

Answer: See Appendix G, 6.7.

Note As before, we can verify that the formula is correct by simply crossing over the charge on each ion. Consider calcium phosphate, which contains Ca^{2+} and PO_4^{3-}. The 2+ charge on the calcium ion becomes the subscript 2 for the phosphate ion. Phosphate is a polyatomic ion, so we must use parentheses, that is, (PO_4). Conversely, the 3− charge on the phosphate ion becomes the subscript 3 for the calcium ion.



$$Ca^{2+}\ PO_4^{3-} = Ca_3(PO_4)_2$$

A CLOSER LOOK Nutritional Potassium

Q: How can potassium be an essential mineral in the human diet when it reacts violently in water?

The human body requires many minerals such as calcium, sodium, and potassium, as well as trace amounts of iron, zinc, and selenium. Minerals perform numerous functions in the body, such as building strong bones, transmitting nerve impulses, making hormones, and maintaining a regular heartbeat.

One role of potassium is to maintain normal amounts of water in body tissues and avoid excessive water retention. Potassium is also important for building muscle and metabolizing proteins and carbohydrates. Although there is no recommended daily allowance (RDA) for potassium, the National Library of Medicine suggests about 2 grams of potassium a day.

The best dietary sources of potassium include fish (especially cod and sardines), meat, peas, tomatoes, potatoes, leafy green vegetables, citrus fruits, bananas, melons, dried fruit, nuts, and chocolate. People with high blood pressure are often put on a diet low in sodium, and a "salt substitute" high in potassium.

Pregnant women and people with chronic diseases are at risk to develop potassium deficiency. Depleted potassium levels may also be caused by vomiting, diarrhea, or use of diuretic or cortisone medications. Symptoms of potassium deficiency include fatigue, slow reflexes, muscle weakness, and dry skin. If potassium levels are depleted quickly, a person can experience heart problems. If potassium levels are not replenished quickly, a person can suffer heart failure and death.

▲ **Potassium in Water** Potassium metal reacts violently in water as shown; potassium chloride, KCl, simply dissolves in water.

Conversely, ingesting extreme amounts of potassium or other minerals can be unhealthy. In some cases, an excess mineral can make it difficult for the body to absorb other minerals and vitamins. The best remedy to avoid vitamin and mineral deficiencies is a balanced diet that includes a wide variety of fruits, green vegetables, protein sources, whole grains, and dairy products daily.

A: Nutritionists state the need for potassium in the diet. To be more precise, humans require potassium salt (not potassium metal) in their diet. In fact, individuals on low-salt diets may use a "salt substitute" containing KCl, in lieu of table salt, NaCl.

6.5 Binary Ionic Compounds

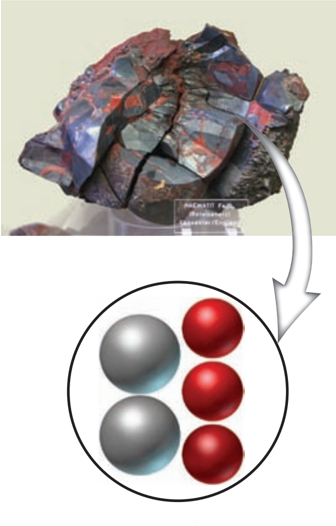
LEARNING OBJECTIVES

- Write systematic names and formulas for binary ionic compounds.
- Determine the ionic charge on a cation in a binary ionic compound.

Positively charged metal ions and negatively charged nonmetal ions compose neutral binary ionic compounds. The oppositely charged cation and anion are attracted to each other and create compounds in which the overall charge is zero. That is, an ionic compound has an overall charge of zero even though it contains charged ions.

Binary Compounds Containing a Transition Metal

Some binary ionic compounds contain transition metal ions—which have a variable ionic charge, for instance, Fe^{2+} and Fe^{3+}. To name such a compound correctly, we first have to determine the ionic charge of the metal cation. The following example exercise illustrates the determination of ionic charge:



▲ **Hematite, Fe_2O_3** Hematite is a mineral found in nature that contains iron.

EXAMPLE EXERCISE 6.8 Determining Ionic Charge in a Compound

Determine the ionic charge for iron in the mineral hematite, Fe_2O_3.

Solution

The charge on an oxide ion is 2−, and there are three oxide ions. The total negative charge must be equal to six negative:

$$O^{2-} + O^{2-} + O^{2-} = 6 \text{ negative}$$

Because all compounds are electrically neutral, the total positive charge must equal the total negative charge: 6 negative = 6 positive. Thus, the two iron ions have a total charge of six positive:

$$Fe^{x+} + Fe^{x+} = 6 \text{ positive}$$
$$Fe^{x+} = 3 \text{ positive}$$

The iron ion is therefore Fe^{3+}. The name of Fe_2O_3 is iron(III) oxide according to the Stock system.

Practice Exercise

Determine the ionic charge for each transition metal in the following compounds:

(a) Cu_3P (b) CoN

Answers:

(a) Cu^+ (b) Co^{3+}

Concept Exercise

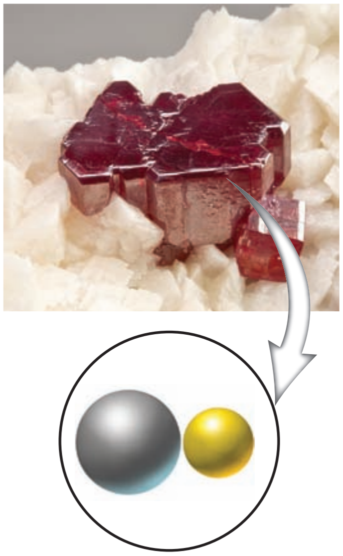
What is the ionic charge on a metal (M), given the formula of the oxide, MO_2?

Answer: See Appendix G, 6.8.

Naming Binary Ionic Compounds

When writing a chemical formula, the positive cation is always written before the negative anion. Therefore, in naming binary ionic compounds, the metal cation is named first followed by the nonmetal anion. We learned in Section 6.2 that monoatomic anions end with the suffix *–ide*. For example, MgO is composed of the magnesium ion and the oxide ion and is named magnesium oxide.

The mineral cinnabar, HgS, is composed of the mercury ion and the sulfide ion. Mercury has two possible ions, Hg_2^{2+} and Hg^{2+}. Because the ionic charge on a sulfide ion is 2−, the mercury ion is Hg^{2+}. Thus, we can name HgS mercury(II) sulfide according to the Stock system.



▲ **Cinnabar, HgS** Cinnabar is a mineral found in nature that contains mercury.

EXAMPLE EXERCISE 6.9 Names of Binary Ionic Compounds

Supply a systematic name for each of the following binary ionic compounds:

(a) ZnO (b) SnF_2

Solution

We can name an ionic compound by designating the two ions.

(a) ZnO contains the zinc ion and the oxide ion; ZnO is named zinc oxide.

(b) SnF_2 contains the tin(II) ion and the fluoride ion. Thus, SnF_2 is named tin(II) fluoride.

Practice Exercise

Supply a systematic name for each of the following binary ionic compounds:

(a) Mn_3P_2 (b) Fe_2S_3

Answers:

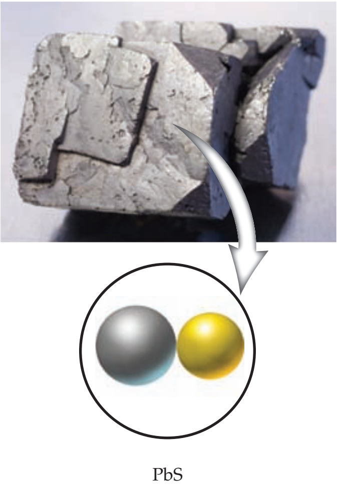
(a) manganese(II) phosphide (b) iron(III) sulfide

Concept Exercise

Which of the following is a binary ionic compound: sodium chloride, sodium chlorate, or sodium chlorite?

Answer: See Appendix G, 6.9.

EXAMPLE EXERCISE 6.10 Formulas of Binary Ionic Compounds



▲ **Galena, PbS** Galena is a mineral found in nature that contains lead.

Provide the formula for each of the following binary ionic compounds:

(a) lithium fluoride (b) lead(II) sulfide

Solution

We write the formula for a binary ionic compound by combining ions so that the total positive charge equals the total negative charge.

(a) Lithium fluoride is composed of Li^+ and F^-; thus, the formula of the compound is written LiF.

(b) Lead(II) sulfide is composed of Pb^{2+} and S^{2-}; thus, the formula of the compound is written PbS.

Practice Exercise

Provide the formula for the following binary ionic compounds:

(a) copper(II) iodide (b) mercury(II) oxide

Answers:

(a) CuI_2 (b) HgO

Concept Exercise

Which of the following is a binary ionic compound: LiCl, $LiClO_3$, or $LiClO_2$?

Answer: See Appendix G, 6.10.

Predicting Formulas of Binary Ionic Compounds

We have learned the value of the periodic table in mastering the formulas of ions. Now we are going to use the periodic table to predict the chemical formulas of binary ionic compounds.

We can predict the formulas of compounds based on the formula of a similar compound. Let's begin with alkali metal chlorides. Sodium is in Group IA/1. The formula of sodium chloride is NaCl. All alkali metal chlorides, in fact, have a similar chemical formula. For the alkali metal family, the chloride compounds have the following formulas: LiCl, NaCl, and KCl.

We can predict the formulas for alkaline earth metal chlorides in a similar manner. Barium is in Group IIA/2, and the formula for barium chloride is $BaCl_2$. Therefore, the formulas for the other alkaline earth metal chlorides are $MgCl_2$, $CaCl_2$, and $SrCl_2$.

EXAMPLE EXERCISE 6.11 Predicting Formulas of Binary Ionic Compounds

Predict the chemical formula for each of the following binary compounds given the formula of aluminum oxide, Al_2O_3:

(a) gallium oxide (b) aluminum sulfide

Solution

To predict the chemical formula, we compare the elements that are different in the similar compounds.

(a) The elements Ga and Al are both in Group IIIA/3, and so the formula is Ga_2O_3.
(b) The elements S and O are both in Group VIA/16, and so the formula is Al_2S_3.

Practice Exercise

Predict the chemical formula for each of the following binary compounds given the formula of magnesium chloride, $MgCl_2$:

(a) radium chloride (b) magnesium fluoride

Answers:

(a) $RaCl_2$ (b) MgF_2

Concept Exercise

Predict the formula for potassium iodide, given the formula of sodium fluoride, NaF.

Answer: See Appendix G, 6.11.

6.6 Ternary Ionic Compounds

LEARNING OBJECTIVES

- Write systematic names and formulas for ternary ionic compounds.
- Determine the ionic charge on a cation in a ternary ionic compound.

Compounds containing a metal and two other elements are classified as ternary ionic compounds. Typically, ternary ionic compounds have a monoatomic metal cation and a polyatomic oxyanion, for example, $NaNO_3$.

Like all ionic compounds, the cation is written first in the formula. The names of ternary ionic compounds usually have the suffix *–ate* or *–ite*. For example, $KClO_3$ is potassium chlorate, and $KClO_2$ is potassium chlorite.

Ternary Compounds Containing a Transition Metal

Some ternary compounds contain transition metal ions that have a variable ionic charge, for instance, Fe^{2+} and Fe^{3+}. To name such a compound correctly, we first have to determine the ionic charge of the transition metal cation.

EXAMPLE EXERCISE 6.12 Determining Ionic Charge in a Compound

Determine the ionic charge for iron in $Fe_3(PO_4)_2$.

Solution

The charge on a phosphate ion is 3− and there are two phosphate ions. Therefore, the total negative charge must be equal to six negative:

$$PO_4^{3-} + PO_4^{3-} = 6 \text{ negative}$$

Because all compounds are electrically neutral, the total positive charge must equal the total negative charge: 6 negative = 6 positive. Thus, the three iron ions have a total charge of six positive.

$$\mathrm{Fe}^{x+} + \mathrm{Fe}^{x+} + \mathrm{Fe}^{x+} = 6 \text{ positive}$$

$$\mathrm{Fe}^{x+} = 2 \text{ positive}$$

The iron ion is therefore Fe^{2+}. The name of $\mathrm{Fe_3(PO_4)_2}$ is iron(II) phosphate according to the Stock system.

Practice Exercise

Determine the ionic charge for the metal cation in each of the following compounds:

(a) $\mathrm{Hg(OH)_2}$ (b) $\mathrm{Co(ClO_3)_3}$

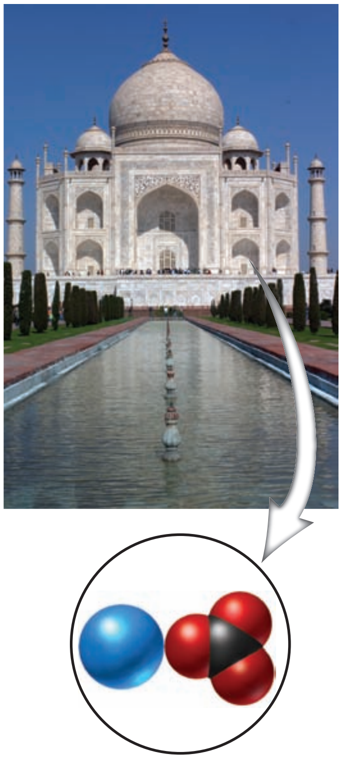
Answers:

(a) Hg^{2+} (b) Co^{3+}

Concept Exercise

What is the ionic charge on a metal (M), given the formula of the carbonate, $\mathrm{MCO_3}$?

Answer: See Appendix G, 6.12.



▲ **Marble, $\mathrm{CaCO_3}$** Marble is a mineral found in nature that was used to construct the Taj Mahal in India.

Naming Ternary Ionic Compounds

Like all ionic compounds, ternary ionic compounds are named by designating the cation followed by the anion. The mineral marble, $\mathrm{CaCO_3}$, is composed of the calcium ion and the carbonate ion and is named calcium carbonate. $\mathrm{Ca(NO_3)_2}$ is composed of the calcium ion and two nitrate ions and is named calcium nitrate.

The compound $\mathrm{CuSO_4}$ is composed of a copper ion and a sulfate ion. Copper, however, has two possible ions, Cu^{+} and Cu^{2+}. Because the sulfate charge is 2−, the charge on copper must be 2+ because ionic compounds are neutral. The copper ion is Cu^{2+}, so the name of the compound is copper(II) sulfate.

EXAMPLE EXERCISE 6.13 Names of Ternary Ionic Compounds

Supply a systematic name for the following ternary ionic compounds:

(a) $\mathrm{KMnO_4}$ (b) $\mathrm{Hg(NO_3)_2}$

Solution

We can name an ionic compound by designating the two ions.

(a) $\mathrm{KMnO_4}$ is composed of the potassium ion and the permanganate ion. Thus, the compound is named potassium permanganate.
(b) $\mathrm{Hg(NO_3)_2}$ contains the mercury(II) ion and the nitrate ion. Therefore, it is named mercury(II) nitrate.

Practice Exercise

Supply a systematic name for each of the following ternary ionic compounds.

(a) $\mathrm{BaCrO_4}$ (b) $\mathrm{Cu(NO_2)_2}$

Answers:

(a) barium chromate
(b) copper(II) nitrite

Concept Exercise

Which of the following is not a ternary ionic compound: potassium nitride, potassium nitrate, or potassium nitrite?

Answer: See Appendix G, 6.13.

EXAMPLE EXERCISE 6.14 Formulas of Ternary Ionic Compounds

Provide the formula for each of the following ternary ionic compounds:

(a) nickel(II) acetate (b) iron(III) sulfate

Solution

We can write the formula by combining the cation and polyatomic anion into a neutral formula unit.

(a) Nickel(II) acetate is composed of Ni^{2+} and $C_2H_3O_2^-$. The formula of the compound is written $Ni(C_2H_3O_2)_2$.

(b) Iron(III) sulfate contains Fe^{3+} and SO_4^{2-}; the formula is written $Fe_2(SO_4)_3$.

Practice Exercise

Provide the formula for each of the following ternary ionic compounds.

(a) mercury(I) nitrite

(b) tin(IV) permanganate

Answers:

(a) $Hg_2(NO_2)_2$ (b) $Sn(MnO_4)_4$

Concept Exercise

Which of the following is not a ternary ionic compound: LiCl, LiClO, or $LiClO_2$?

Answer: See Appendix G, 6.14.

Predicting Formulas of Ternary Ionic Compounds

The periodic table can be used to predict the formulas of ternary ionic compounds. For example, if we know that the formula of potassium nitrate is KNO_3, we can predict the formulas of sodium nitrate and lithium nitrate. Because Li, Na, and K are all in Group IA/1, we can predict that the compounds have similar chemical formulas, that is, $LiNO_3$, $NaNO_3$, and KNO_3. The following example exercise illustrates how to predict formulas of ternary ionic compounds.

EXAMPLE EXERCISE 6.15 Predicting Formulas of Ternary Ionic Compounds

Predict the chemical formula for each of the following ternary ionic compounds given the formula of calcium carbonate, $CaCO_3$:

(a) radium carbonate (b) calcium silicate

Solution

To predict the formula, we compare the elements that are different in the similar compounds.

(a) The elements Ra and Ca are both in Group IIA/2. Thus, the formula for radium carbonate is $RaCO_3$.

(b) The elements Si and C are both in Group IVA/14. Therefore, the formula for calcium silicate is $CaSiO_3$.

Practice Exercise

Predict the chemical formula for each of the following ternary compounds given the formula of potassium chlorate, $KClO_3$:

(a) lithium chlorate (b) potassium bromate

Answers:

(a) $LiClO_3$ (b) $KBrO_3$

Concept Exercise

Predict the formula for strontium sulfate, given the formula of calcium sulfate, $CaSO_4$.

Answer: See Appendix G, 6.15.

LEARNING OBJECTIVE

- Write names and formulas for ionic compounds using the Latin system.

6.7 Latin System of Nomenclature

In addition to the Stock system, IUPAC allows, but discourages, an older method for naming metal cations having two common ionic charges. It is called the **Latin system**, or the suffix system. This system uses the Latin name of the metal and applies the suffix *–ous* or *–ic*. The lower of the two ionic charges applies an –ous suffix, and the higher charge applies an *–ic* suffix. For example, the Latin name for iron is *ferrum*. To name an iron ion, we take the stem *ferr–* and add the suffix *–ous* or *–ic*. Thus, Fe^{2+} is named ferrous ion, and Fe^{3+} is named ferric ion. The Latin name for copper is *cuprum*. Similarly, we add the suffix –ous or *–ic* to the stem *cupr–*. Thus, Cu^{+} is named cuprous ion, and Cu^{2+} is named cupric ion.

The Latin name for mercury is hydrargyrum, but IUPAC recommends applying the suffix to the English name. Thus, Hg_2^{2+} is named mercurous ion, and Hg^{2+} is named mercuric ion. Table 6.4 compares systematic names for common cations using both the Stock system and Latin system of nomenclature.

TABLE 6.4 Common Monoatomic Cations

Cation	Stock System	Latin System
Co^{2+}	cobalt(II) ion	cobaltous ion
Co^{3+}	cobalt(III) ion	cobaltic ion
Cu^{+}	copper(I) ion	cuprous ion
Cu^{2+}	copper(II) ion	cupric ion
Fe^{2+}	iron(II) ion	ferrous ion
Fe^{3+}	iron(III) ion	ferric ion
Pb^{2+}	lead(II) ion	plumbous ion
Pb^{4+}	lead(IV) ion	plumbic ion
Hg_2^{2+}	mercury(I) ion*	mercurous ion
Hg^{2+}	mercury(II) ion	mercuric ion
Sn^{2+}	tin(II) ion	stannous ion
Sn^{4+}	tin(IV) ion	stannic ion

*Note that the mercury(I) ion is diatomic and is written Hg_2^{2+}.

EXAMPLE EXERCISE 6.16 Latin System Nomenclature of Binary Ionic Compounds

Supply a systematic name for the following binary ionic compounds.

(a) SnF_2 (b) $FeCl_3$

Solution

(a) The binary ionic compound SnF_2 contains the stannous ion, Sn^{2+}, and two fluoride ions, F^{-}. According to the Latin system, the name of SnF_2 is stannous fluoride, which is a dental preventative in a popular toothpaste.

(b) The binary ionic compound $FeCl_3$ contains the ferric ion, Fe^{3+}, and three chloride ions, Cl^{-}. According to the Latin system, the name of $FeCl_3$ is ferric chloride.

Practice Exercise

Provide the formula for each of the following binary ionic compounds.

(a) cuprous phosphide (b) cobaltic sulfide

Answers:

(a) Cu_3P (b) Co_2S_3

Concept Exercise

According to the anion suffix, which of the following is a binary ionic compound: sodium chloride, sodium chlorate, or sodium chlorite?

Answer: See Appendix G, 6.16.

EXAMPLE EXERCISE 6.17 Latin System Nomenclature of Ternary Ionic Compounds

Supply a systematic name for the following ternary ionic compounds.

(a) $Cu(NO_3)_2$ (b) $Co(MnO_4)_2$

Solution

We can name an ionic compound by designating the two ions.

(a) The ternary ionic compound $Cu(NO_3)_2$ contains the cupric ion, Cu^{2+}, and two nitrate ions, NO_3^-. According to the Latin system, the name of $Cu(NO_3)_2$ is cupric nitrate.

(b) The ternary ionic compound $Co(MnO_4)_2$ contains the cobaltous ion, Co^{2+}, and two permanganate ions, MnO_4^-. According to the Latin system, the name of $Co(MnO_4)_2$ is cobaltous nitrate.

Practice Exercise

Provide the formula for each of the following binary ionic compounds.

(a) stannic nitrite (b) ferrous acetate

Answers:

(a) $Sn(NO_2)_4$ (b) $Fe(C_2H_3O_2)_2$

Concept Exercise

According to the anion suffix, which of the following is a ternary ionic compound: potassium nitride, potassium nitrate, or potassium nitrite?

Answer: See Appendix G, 6.17.

Helpful Hint

In Chapter 2 we were introduced to both the metric system and English system of measurement. Although the metric system is the international system, and more convenient to use, the English system is still widely used in the United States. Similarly, both the Stock system and Latin system of nomenclature are still used extensively.

- The Stock system of naming uses Roman numerals in parentheses to indicate the ionic charge on a multivalent cation; for example, Fe(II) for the Fe^{2+} ion, and Fe(III) for the Fe^{3+} ion.
- The Latin system of naming applies an *-ous* or *-ic* suffix to the Latin name of the element; for example, ferrous for the Fe^{2+} ion, and ferric for the Fe^{3+} ion.

6.8 Binary Molecular Compounds

LEARNING OBJECTIVE

- Write systematic names and formulas for binary molecular compounds.

Recall that a binary molecular compound is composed of two nonmetal elements. The simplest representative particle in a binary molecular compound is a **molecule**. In general, the more nonmetallic element is written second in the chemical formula. IUPAC prescribes the following order for writing the elements in a compound: C, P, N, H, S, I, Br, Cl, O, F. Notice that hydrogen is in the middle of the series. Thus, the binary compounds of hydrogen are written as follows: CH_4, PH_3, NH_3, H_2S, HI, HBr, HCl, H_2O, and HF.

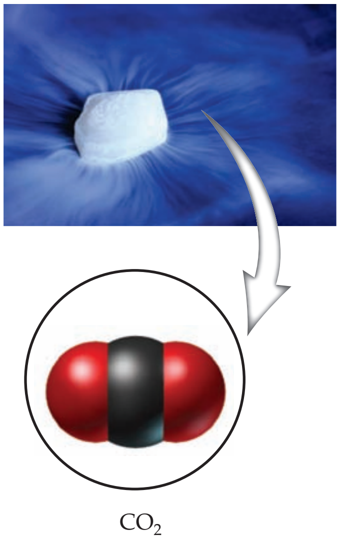
Naming Binary Molecular Compounds

In the naming of binary molecular compounds, IUPAC specifies that the first element in the compound be named followed by the second element bearing the suffix *–ide*. Thus, binary ionic and binary molecular compounds both end in the suffix *–ide*. It also specifies that the number of atoms of each element be indicated by Greek prefixes. The Greek prefixes in Table 6.5 indicate the number of atoms of each element that correspond to the subscripts in the chemical formula.

TABLE 6.5 Greek Prefixes for Binary Molecular Compounds

Atoms	Prefix	Atoms	Prefix
1	mono–	6	hexa–
2	di–	7	hepta–
3	tri–	8	octa–
4	tetra–	9	nona–*
5	penta–	10	deca–

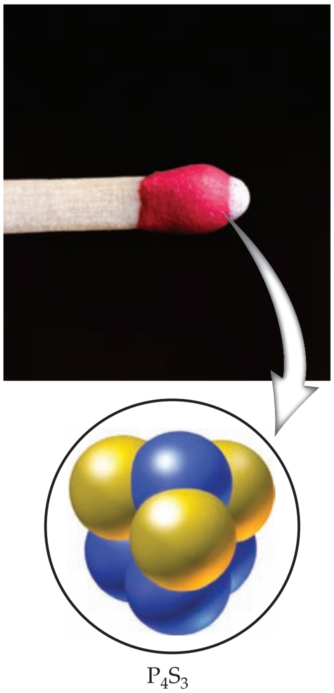
*Although the Latin prefix *nona–* is commonly used, IUPAC prefers the Greek prefix *ennea–*.



▲ **Carbon Dioxide, CO_2** Dry ice is an example of a binary molecular compound.

According to the official IUPAC recommendations, the prefix *mono–* is always omitted unless its presence is necessary to avoid confusion. The only common exceptions are CO and NO. The name of CO includes the prefix *mono–* in front of oxygen and is written carbon monoxide. Similarly, NO is named nitrogen monoxide.

Let's consider the binary molecular compound P_4S_3. This compound is found on match tips and ignites in air when struck on a rough surface (Figure 6.4). The ratio is four phosphorus atoms to three sulfur atoms, so the Greek prefixes are *tetra–* and *tri–*, respectively. Thus, the name of P_4S_3 is tetraphosphorus trisulfide. A different compound, P_4S_7, has the same elements but is named tetraphosphorus heptasulfide.



▲ **Figure 6.4 Chemistry of Matches** The substances P_4S_3 and $KClO_3$ are present on the tip of a match. When the match is struck on a rough surface, the two chemicals ignite and produce a flame. The products from the reaction are P_2O_5, KCl, and SO_2, the last of which is responsible for the characteristic sulfur smell.

EXAMPLE EXERCISE 6.18 Names of Binary Molecular Compounds

Give the IUPAC systematic name for each of the following binary molecular compounds:

(a) IF_6 (b) Br_3O_8

Solution

We name binary molecular compounds by attaching the suffix *–ide* to the second nonmetal and indicate the atomic ratios by Greek prefixes.

(a) IF_6 is first named iodine fluoride. After supplying the Greek prefixes for the atomic ratios, we have *iodine hexafluoride*.

(b) Br_3O_8 is first named bromine oxide. After supplying the Greek prefixes for the atomic ratios, we have *tribromine octaoxide*.

Practice Exercise

Give the IUPAC systematic name for each of the following binary molecular compounds:

(a) Cl_2O_5 (b) P_4S_{10}

Answers:

(a) dichlorine pentaoxide (b) tetraphosphorus decasulfide

Concept Exercise

Which of the following is named using Greek prefixes to specify the number of atoms of each element: Fe_2O_3 or P_2O_3?

Answer: See Appendix G, 6.18.

EXAMPLE EXERCISE 6.19 Formulas of Binary Molecular Compounds

Provide the formula for each of the following binary molecular compounds:

(a) diphosphorus pentasulfide (b) tetraiodine nonaoxide

Solution

To write the formula, we give the symbol for each element followed by a subscript indicating the number of atoms.

(a) Diphosphorus pentasulfide is composed of two phosphorus atoms and five sulfur atoms. The formula of the compound is written P_2S_5.
(b) Tetraiodine nonaoxide is composed of four iodine atoms and nine oxygen atoms. The formula of the compound is written I_4O_9.

Practice Exercise

Provide the formula for each of the following binary molecular compounds:

(a) diphosphorus tetraiodide (b) sulfur hexafluoride

Answers:

(a) P_2I_4 (b) SF_6

Concept Exercise

Which of the following is a binary molecular compound: magnesium oxide or nitrogen oxide?

Answer: See Appendix G, 6.19.

Note For ease of pronunciation, double vowels are sometimes avoided in naming compounds with Greek prefixes. For example, if the Greek prefix ends in "a" or "o" and the nonmetal is oxygen, two pronunciations are common. That is, the name may be pronounced tetraoxide or tetroxide, pentaoxide or pentoxide, etc. However, according to the official IUPAC recommendations, vowels are not to be dropped and tetraoxide is preferred. The one exception is monoxide, not monooxide.

6.9 Binary Acids

LEARNING OBJECTIVE

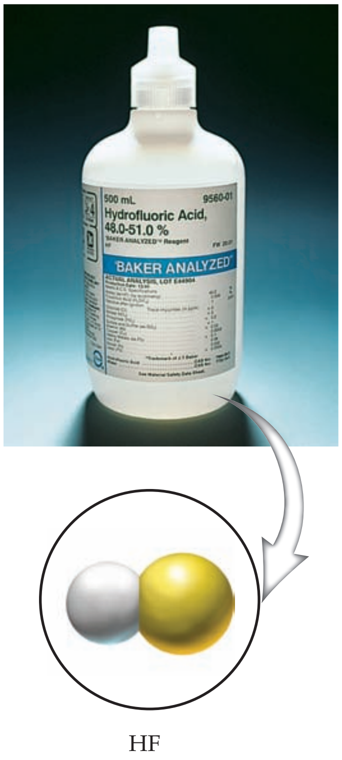
- Write systematic names and formulas for binary acids.

A binary acid is an aqueous solution of a compound containing hydrogen and a nonmetal. The chemical formulas of acids always begin with H, for example, $HCl(aq)$. A binary acid is produced by dissolving a binary molecular compound, such as HCl, in water. The resulting aqueous solution, $HCl(aq)$, is a binary acid.

Binary acids are systematically named by using the prefix *hydro–* before the nonmetal stem and adding the suffix *–ic acid*. As an example, consider the acid used to treat swimming pools and sold in supermarkets under the common name muriatic acid. Muriatic acid is aqueous hydrogen chloride, that is, $HCl(aq)$. The IUPAC name for muriatic acid is formed as follows: hydro + chlor + ic acid. The systematic name for $HCl(aq)$ is hydrochloric acid. We can illustrate the binary acid relationship to its compounds as shown in Table 6.6.

TABLE 6.6 Binary Acids and Related Compounds

Binary	Acid	Binary	Compound
HF	hydrofluoric acid	NaF	sodium fluoride
HCl	hydrochloric acid	NaCl	sodium chloride
HBr	hydrobromic acid	NaBr	sodium bromide
HI	hydroiodic acid	NaI	sodium iodide



▲ **Hydrofluoric Acid, HF**

EXAMPLE EXERCISE 6.20 Names of Binary Acids

Give the IUPAC systematic name for $HF(aq)$, a binary acid.

Solution

Binary acids are named as *hydro–* plus nonmetal stem plus *–ic acid*. Because $HF(aq)$ contains the nonmetal fluorine, we construct the systematic name as follows: hydro + fluor + ic acid gives *hydrofluoric acid*.

Practice Exercise

Give the IUPAC systematic name for $H_2S(aq)$.

Answer:

hydrosulfuric acid

Concept Exercise

Which of the following acids is named using a *hydro–* prefix: $HBr(aq)$, $HBrO_2(aq)$, $HBrO_3(aq)$?

Answer: See Appendix G, 6.20.

Note Be careful not to confuse the names of a binary acid and a binary molecular compound. A binary acid is a compound of hydrogen and a nonmetal dissolved in water. For example, $HCl(aq)$ is a binary acid and is named hydrochloric acid. On the other hand, $HCl(g)$ is a gaseous binary molecular compound and is named hydrogen chloride.

6.10 Ternary Oxyacids

LEARNING OBJECTIVE

- Write systematic names and formulas for ternary oxyacids.

A ternary oxyacid is an aqueous solution of a compound containing hydrogen and an oxyanion. Most ternary oxyacids are named by attaching the suffix *–ic acid* to the *nonmetal stem* of the oxyanion. For example, the name of $HNO_3(aq)$ is systematically formed as follows: *nonmetal stem* plus *ic acid*. The name of $HNO_3(aq)$ is nitr + ic acid, or *nitric acid*.

Some ternary oxyacids are named by attaching the suffix *–ous acid* to the *nonmetal stem* of the oxyanion. For example, the name of $HNO_2(aq)$ is systematically formed as follows: *nonmetal stem* + *ous acid*. The name of $HNO_2(aq)$ is nitr + ous acid, or *nitrous acid*.

A ternary oxyacid with the suffix *–ic acid* contains an oxyanion ending in *–ate*. Thus, the name of a compound containing the oxyanion will bear the suffix *–ate*. Consider sulfuric acid, $H_2SO_4(aq)$, which contains the sulfate oxyanion, SO_4^{2-}. The name of the compound Na_2SO_4 is sodium sulfate.

Similarly, a ternary oxyacid with the suffix *–ous acid* contains an oxyanion ending in *–ite*. Consider sulfurous acid, $H_2SO_3(aq)$, which contains the sulfite oxyanion, SO_3^{2-}. The name of the compound Na_2SO_3 is sodium sulfite. We can illustrate the ternary oxyacid relationship to its compounds as shown in Table 6.7.

H_2SO_4

▲ **Sulfuric Acid, H_2SO_4** Aqueous sulfuric acid, $H_2SO_4(aq)$, is found in automobile batteries; hence, it is commonly termed *battery acid*.

TABLE 6.7 Ternary Oxyacids and Related Compounds

Ternary	Oxyacid	Ternary	Compound
$HClO$	**hypo**chlor**ous** acid	$NaClO$	sodium **hypo**chlor**ite**
$HClO_2$	chlor**ous** acid	$NaClO_2$	sodium chlor**ite**
$HClO_3$	chlor**ic** acid	$NaClO_3$	sodium chlor**ate**
$HClO_4$	**per**chlor**ic** acid	$NaClO_4$	sodium **per**chlor**ate**

Now that we have provided an overview of ternary oxyacids, let's try the following example exercise that illustrates systematic naming:

EXAMPLE EXERCISE 6.21 Names of Ternary Oxyacids

Give the IUPAC systematic name for $H_3PO_4(aq)$, a ternary oxyacid.

Solution

Ternary oxyacids are named as *–ic acids* or *–ous acids*. Because $H_3PO_4(aq)$ contains the phosphate oxyanion, it is an *–ic acid*. We construct the systematic name as follows: phosphor + ic acid gives *phosphoric acid*.

Practice Exercise

Give the IUPAC systematic name for $H_3PO_3(aq)$, a ternary oxyacid.

Answer: phosphorous acid

Concept Exercise

Which of the following acids is named *nonmetal stem* plus *–ic acid:* $HBr(aq)$, $HBrO_2(aq)$, $HBrO_3(aq)$?

Answer: See Appendix G, 6.21.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
6.1 Classification of Compounds **Inorganic compounds** do not contain carbon and are named according to a systematic set of rules referred to as **IUPAC nomenclature**. The IUPAC rules classify substances as belonging to one of five categories: **binary ionic compound**, **ternary ionic compound**, **binary molecular compound**, **binary acid**, or **ternary oxyacid**. A substance dissolved in water is referred to as an **aqueous solution**. Ionic compounds are composed of positive **cations** and negative **anions**. A **monoatomic ion** is a single atom having a charge; whereas a **polyatomic ion** is a group of two or more atoms bearing a charge.	• Classify a compound as a binary ionic compound, a ternary ionic compound, or a binary molecular compound. ***Related Exercises 1–2*** • Classify an acid as a binary acid or a ternary oxyacid. ***Related Exercises 1–2*** • Classify an ion as a monoatomic cation, a monoatomic anion, a polyatomic cation, or a polyatomic anion. ***Related Exercises 3–4***
6.2 Monoatomic Ions The name of a monoatomic cation is derived from a metal, for example, the name of Na^+ is sodium ion. Cations having two possible ionic charges require further identification. The **Stock system** indicates the charge on the metal ion using Roman numerals in parentheses. Thus, Cu^+ is named copper(I) ion, and Cu^{2+} is named copper(II) ion, according to the Stock system. The name of a monoatomic anion is derived from a nonmetal. A monoatomic anion is named using the nonmetal stem and adding an *-ide* suffix. Examples of nonmetal ions include the chloride ion, Cl^-, and the oxide ion, O^{2-}.	• Provide systematic names and formulas for common monoatomic ions. ***Related Exercises 5–8*** • Predict the ionic charge for ions of representative elements. ***Related Exercises 9–10***
6.3 Polyatomic Ions A polyatomic anion that contains one or more elements combined with oxygen is called an **oxyanion**. The names of most polyatomic oxyanions have the suffix *-ate*, for example, the nitrate ion, NO_3^-. A few oxyanions have the suffix *-ite*, for example, the nitrite ion, NO_2^-. The nitrite ion has one less oxygen atom than the nitrate ion. Oxyanions having one less oxygen atom have the suffix *-ite*, such as the sulfite ion, SO_3^{2-}, which has one less oxygen than the sulfate ion, SO_4^{2-}.	• Provide systematic names and formulas for common polyatomic ions. ***Related Exercises 11–14***
6.4 Writing Chemical Formulas The simplest representative particle in an ionic compound is called a **formula unit**. The net charge for a formula unit is zero. That is, the total positive charge from the cations in the compound must be the same as the total negative charge from the anions.	• Write the chemical formulas for ionic compounds composed of monoatomic ions and polyatomic ions. ***Related Exercises 15–20***
6.5 Binary Ionic Compounds Binary ionic compounds have names ending with the suffix *-ide*. For example, NaCl is named sodium chloride, $CaBr_2$ is named calcium bromide, and AlF_3 is named aluminum fluoride.	• Determine the ionic charge on a cation in a binary ionic compound. ***Related Exercises 21–24*** • Write systematic names and formulas for binary ionic compounds. ***Related Exercises 25–30***

Key Concepts	Learning Objectives and Related Exercises
6.6 Ternary Ionic Compounds Ternary ionic compounds have names ending with the suffix *-ate* or *-ite*. For example, $CaSO_4$ is named calcium sulfate, and $CaSO_3$ is named calcium sulfite. A ternary compound containing CN^- or OH^- is an exception and each ends with the suffix *-ide*. For example, NaOH is named sodium hydroxide, and KCN is named potassium cyanide.	• Determine the ionic charge on a cation in a ternary ionic compound. ***Related Exercises 31–34*** • Write systematic names and formulas for ternary ionic compounds. ***Related Exercises 35–40***
6.7 Latin System of Nomenclature The naming of compounds that contain cations having two possible ionic charges can be named by either the Stock system or the Latin system. The **Latin system** adds the suffix *-ous* or *-ic* to the Latin name of the element. The ionic charge of iron in $FeCl_2$ is 2+ and the compound is named ferrous chloride; the ionic charge of iron in $FeCl_3$ is 3+ and is named ferric chloride.	• Write names and formulas for ionic compounds using the Latin system. ***Related Exercises 41–48***
6.8 Binary Molecular Compounds Binary molecular compounds contain two nonmetals, and the simplest representative particle is called a **molecule**. Binary molecular compounds are named using a Greek prefix for the number of atoms and the suffix *-ide;* for example, SO_2 and SO_3 are named sulfur dioxide and sulfur trioxide, respectively.	• Write systematic names and formulas for binary molecular compounds. ***Related Exercises 49–52***
6.9 Binary Acids A binary acid is an aqueous solution of a hydrogen containing compound. Binary acids are named *hydro* + *nonmetal stem* + *-ic acid*. For example, hydrogen chloride dissolved in water, $HCl(aq)$, is named hydrochloric acid.	• Write systematic names and formulas for binary acids. ***Related Exercises 53–54***
6.10 Ternary Oxyacids Ternary oxyacids are usually named *nonmetal stem* + *-ic acid*. For example, an aqueous solution of $HNO_3(aq)$ is named nitric acid. A few ternary oxyacids are named *nonmetal stem* + *-ous acid*. An aqueous solution of $HNO_2(aq)$ is named nitrous acid.	• Write systematic names and formulas for ternary oxyacids. ***Related Exercises 55–58***

Nomenclature Organizer

Topic	Procedure	Example
Classification of Compounds Sec. 6.1	When a formula has two elements, a metal and a nonmetal, it is a **binary ionic compound**. When a formula has three elements and at least one metal, it is a **ternary ionic compound**. When a formula has two nonmetal elements, it is a **binary molecular compound**. When a formula has H and a nonmetal in an aqueous solution, it is a **binary acid**. When a formula has H, a nonmetal, and oxygen in an aqueous solution, it is a **ternary oxyacid**.	binary ionic compound, KCl ternary ionic compound, $KClO_3$ binary molecular compound, HCl binary acid, $HCl(aq)$ ternary oxyacid, $HClO_3(aq)$
Monoatomic Ions Sec. 6.2	Representative metal ions are named for the element. Transition metal ions are named for the element, followed by the ionic charge in parentheses; in other words, the Stock system. Nonmetals are named for the element plus the **suffix *–ide***.	K^+, potassium ion Fe^{2+}, iron(II) ion or ferrous ion Cl^-, chloride ion

Nomenclature Organizer *(Continued)*

Topic	Procedure	Example
Polyatomic Ions Sec. 6.3	Polyatomic oxyanions are named with (1) the **suffix *–ate***, or (2) the **suffix *–ite*** when there is one less oxygen atom in the oxyanion. There are two common polyatomic anion exceptions that have the **suffix *–ide***.	ClO_3^-, chlorate ion ClO_2^-, chlorite ion OH^-, hydroxide ion CN^-, cyanide ion
Writing Chemical Formulas Sec. 6.4	In writing formulas of ionic compounds, the ratio of cations and anions must provide a total positive charge equal to the total negative charge so that the compound is electrically neutral.	Fe^{2+} and O^{2-} = FeO Fe^{3+} and O^{2-} = Fe_2O_3 Fe^{2+} and ClO_3^- = $Fe(ClO_3)_2$ Fe^{3+} and ClO_3^- = $Fe(ClO_3)_3$
Binary Ionic Compounds Sec. 6.5	Binary ionic compounds end in the **suffix *–ide***. The metal cation is named first followed by the name of the anion.	KCl, potassium chloride $FeCl_2$ iron(II) chloride, or ferrous chloride
Ternary Ionic Compounds Sec. 6.6	Ternary ionic compounds end in the **suffix *–ate*** or the **suffix *–ite***. The metal cation is named first followed by the name of the oxyanion.	$KClO_3$, potassium chlorate $KClO_2$, potassium chlorite
Latin System of Nomenclature Sec. 6.7	Transition metal ions can be named using the Latin system name of the element, and adding an -ous suffix, or an -ic suffix.	Fe^{2+}, ferrous ion Fe^{3+}, ferric ion
Binary Molecular Compounds Sec. 6.8	Binary molecular compounds end in the **suffix *–ide***. The number of nonmetal atoms is indicated by a Greek prefix.	CO, carbon monoxide CO_2, carbon dioxide
Binary Acids Sec. 6.9	Binary acids are named hydro- plus nonmetal stem plus ***–ic* acid**.	HCl(*aq*), hydrochloric acid
Ternary Oxyacids Sec. 6.10	Ternary oxyacids are named with (1) a **nonmetal stem** plus ***–ic* acid,** or (2) a **nonmetal stem** plus ***–ous* acid.**	$HClO_3$(*aq*), chloric acid $HClO_2$(*aq*), chlorous acid

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** the international system of rules for naming chemical compounds
_____ **2.** a compound that does not contain the element carbon
_____ **3.** a compound that contains one metal and one nonmetal
_____ **4.** a compound that contains three elements, including at least one metal
_____ **5.** a compound that contains two nonmetals
_____ **6.** a substance dissolved in water
_____ **7.** an aqueous compound that contains hydrogen and a nonmetal
_____ **8.** an aqueous compound that contains hydrogen, a nonmetal, and oxygen
_____ **9.** any positively charged ion
_____ **10.** any negatively charged ion
_____ **11.** a single atom that has a positive or negative charge
_____ **12.** a group of atoms bound together and having an overall positive or negative charge
_____ **13.** a system for naming ionic compounds that designates the charge on a cation with Roman numerals

(a) anion (*Sec. 6.1*)
(b) aqueous solution (*Sec. 6.1*)
(c) binary acid (*Sec. 6.1*)
(d) binary ionic compound (*Sec. 6.1*)
(e) binary molecular compound (*Sec. 6.1*)
(f) cation (*Sec. 6.1*)
(g) formula unit (*Sec. 6.4*)
(h) inorganic compound (*Sec. 6.1*)
(i) IUPAC nomenclature (*Sec. 6.1*)
(j) Latin system (*Sec. 6.7*)
(k) molecule (*Sec. 6.8*)

_____ **14.** a system for naming ionic compounds that designates the charge on a cation with an *–ic* or *–ous* suffix

_____ **15.** a polyatomic anion that contains one or more elements combined with oxygen, for example, CO_3^{2-}

_____ **16.** the simplest representative particle in a compound composed of ions

_____ **17.** the simplest representative particle in a compound composed of nonmetal atoms

(l) monoatomic ion (*Sec. 6.1*)
(m) oxyanion (*Sec. 6.3*)
(n) polyatomic ion (*Sec. 6.1*)
(o) Stock system (*Sec. 6.2*)
(p) ternary ionic compound (*Sec. 6.1*)
(q) ternary oxyacid (*Sec. 6.1*)

Exercises

Answers to odd-numbered Exercises are in Appendix I.

Classification of Compounds (Sec. 6.1)

1. Classify each of the following as a binary ionic compound, ternary ionic compound, binary molecular compound, binary acid, or ternary oxyacid.
   (a) water, H_2O **(b)** rust, Fe_2O_3
   (c) marble, $CaCO_3$ **(d)** stomach acid, $HCl(aq)$

2. Classify each of the following as a binary ionic compound, ternary ionic compound, binary molecular compound, binary acid, or ternary oxyacid.
   (a) ammonia, NH_3 **(b)** iodized salt, KI
   (c) bleach, NaClO **(d)** battery acid, $H_2SO_4(aq)$

3. Classify each of the following as a monoatomic cation, monoatomic anion, polyatomic cation, or polyatomic anion.
   (a) ammonium ion, NH_4^+ **(b)** aluminum ion, Al^{3+}
   (c) chloride ion, Cl^- **(d)** chlorate ion, ClO_3^-

4. Classify each of the following as a monoatomic cation, monoatomic anion, polyatomic cation, or polyatomic anion.
   (a) hydronium ion, H_3O^+ **(b)** strontium ion, Sr^{2+}
   (c) sulfide ion, S^{2-} **(d)** sulfate ion, SO_4^{2-}

Monoatomic Ions (Sec. 6.2)

5. Supply the Stock system name for each of the following monoatomic cations.
   (a) Ca^{2+} **(b)** Cd^{2+}
   (c) Co^{2+} **(d)** Cu^{2+}

6. Supply the Stock system name for each of the following monoatomic cations.
   (a) Ba^{2+} **(b)** Zn^{2+}
   (c) Co^{3+} **(d)** Cu^+

7. Supply a systematic name for each of the following monoatomic anions.
   (a) F^- **(b)** Cl^-
   (c) Br^- **(d)** I^-

8. Supply a systematic name for each of the following monoatomic anions.
   (a) O^{2-} **(b)** S^{2-}
   (c) N^{3-} **(d)** P^{3-}

9. Provide the formula for each of the following monoatomic cations.
   (a) potassium ion **(b)** barium ion
   (c) iron(II) ion **(d)** tin(II) ion

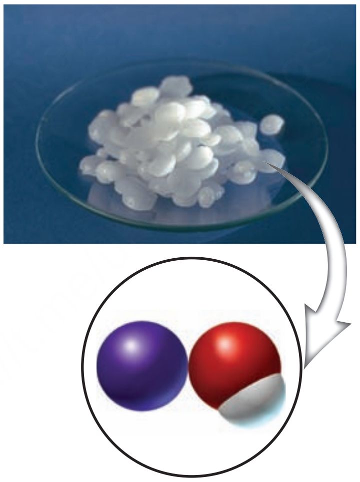
10. Provide the formula for each of the following monoatomic cations.
   (a) lithium ion **(b)** silver ion
   (c) iron(III) ion **(d)** tin(IV) ion

Polyatomic Ions (Sec. 6.3)

11. Supply a systematic name for each of the following oxyanions.
   (a) NO_3^- **(b)** NO_2^-
   (c) ClO_3^- **(d)** ClO_2^-

12. Supply a systematic name for each of the following oxyanions.
   (a) CO_3^{2-} **(b)** HCO_3^-
   (c) SO_4^{2-} **(d)** HSO_4^-

13. Provide the formula for each of the following polyatomic ions.
   (a) chromate ion **(b)** dichromate ion
   (c) hydroxide ion **(d)** cyanide ion



◀ **Sodium Hydroxide, NaOH** Sodium Hydroxide, NaOH, is used in oven cleaner and drain cleaner. Because it is strongly caustic and contains sodium, it is commonly referred to as "caustic soda."

14. Provide the formula for each of the following polyatomic ions.
   (a) chlorate ion **(b)** perchlorate ion
   (c) permanganate ion **(d)** acetate ion

Writing Chemical Formulas (Sec. 6.4)

15. Write the chemical formula for the following binary compounds given their constituent ions.
   (a) aluminum bromide, Al^{3+} and Br^-
   (b) cadmium sulfide, Cd^{2+} and S^{2-}
   (c) manganese(II) phosphide, Mn^{2+} and P^{3-}
   (d) lead(IV) fluoride, Pb^{4+} and F^-

16. Write the chemical formula for the following binary compounds given their constituent ions.
(a) lithium chloride, Li^+ and Cl^-
(b) silver oxide, Ag^+ and O^{2-}
(c) chromium(III) iodide, Cr^{3+} and I^-
(d) tin(II) nitride, Sn^{2+} and N^{3-}

17. Write the chemical formula for the following ternary compounds given their constituent ions.
(a) potassium nitrate, K^+ and NO_3^-
(b) ammonium dichromate, NH_4^+ and $Cr_2O_7^{2-}$
(c) copper (I) carbonate, Cu^+ and CO_3^{2-}
(d) manganese(II) sulfite, Mn^{2+} and SO_3^{2-}

18. Write the chemical formula for the following ternary compounds given their constituent ions.
(a) sodium acetate, Na^+ and $C_2H_3O_2^-$
(b) aluminum sulfite, Al^{3+} and SO_3^{2-}
(c) mercury(II) cyanide, Hg^{2+} and CN^-
(d) chromium(III) hypochlorite, Cr^{3+} and ClO^-

19. Write the chemical formula for the following ternary compounds given their constituent ions.
(a) strontium nitrite, Sr^{2+} and NO_2^-
(b) zinc permanganate, Zn^{2+} and MnO_4^-
(c) calcium chromate, Ca^{2+} and CrO_4^{2-}
(d) nickel(II) perchlorate, Ni^{2+} and ClO_4^-

20. Write the chemical formula for the following ternary compounds given their constituent ions.
(a) lead(IV) sulfate, Pb^{4+} and SO_4^{2-}
(b) tin(II) chlorite, Sn^{2+} and ClO_2^-
(c) cobalt(II) hydroxide, Co^{2+} and OH^-
(d) mercury(I) phosphate, Hg_2^{2+} and PO_4^{3-}

Binary Ionic Compounds (Sec. 6.5)

21. Supply a systematic name for each of the following binary ionic compounds.
(a) MgO
(b) ZnO
(c) CdO
(d) BaO

22. Supply a systematic name for each of the following binary ionic compounds.
(a) $LiBr$
(b) SrI_2
(c) Na_3N
(d) AlF_3

23. Supply a Stock system name for each of the following binary ionic compounds.
(a) CoO
(b) FeO
(c) HgO
(d) SnO

24. Supply a Stock system name for each of the following binary ionic compounds.
(a) $MnCl_2$
(b) NiF_2
(c) $SnBr_2$
(d) CrP

25. Provide the formula for each of the following binary ionic compounds.
(a) copper(I) oxide
(b) iron(II) nitride
(c) mercury(II) chloride
(d) lead(IV) sulfide

◀ **Cuprite, Cu_2O** Cuprite, copper(I) oxide, is a mineral found in nature which contains copper.

26. Provide the formula for each of the following binary ionic compounds.
(a) copper(II) sulfide
(b) iron(III) phosphide
(c) mercury(I) iodide
(d) lead(II) oxide

27. Predict the chemical formula for each of the following binary ionic compounds given the formula of sodium chloride, $NaCl$.
(a) rubidium chloride
(b) sodium bromide

28. Predict the chemical formula for each of the following binary ionic compounds given the formula of calcium oxide, CaO.
(a) beryllium oxide
(b) calcium selenide

29. Predict the chemical formula for each of the following binary ionic compounds given the formula of aluminum nitride, AlN.
(a) gallium nitride
(b) aluminum arsenide

30. Predict the chemical formula for each of the following binary ionic compounds given the formula of titanium oxide, TiO_2.
(a) zirconium oxide
(b) titanium sulfide

Ternary Ionic Compounds (Sec. 6.6)

31. Supply a systematic name for each of the following ternary ionic compounds.
(a) $LiMnO_4$
(b) $Sr(ClO_2)_2$
(c) $BaCO_3$
(d) $(NH_4)_2Cr_2O_7$

32. Supply a systematic name for each of the following ternary ionic compounds.
(a) $KClO_3$
(b) $Mg(NO_3)_2$
(c) Ag_2SO_4
(d) $Al_2(CrO_4)_3$

33. Supply a Stock system name for each of the following ternary ionic compounds.
(a) $CuSO_4$
(b) $FeCrO_4$
(c) $Hg(NO_2)_2$
(d) $Pb(C_2H_3O_2)_2$

34. Supply a Stock system name for each of the following ternary ionic compounds.
(a) Cu_2SO_4
(b) $Fe_2(CrO_4)_3$
(c) $Hg_2(NO_2)_2$
(d) $Pb(C_2H_3O_2)_4$

35. Provide the formula for each of the following ternary ionic compounds.
(a) manganese(II) acetate
(b) lead(IV) chlorite
(c) tin(II) phosphate
(d) iron(III) hypochlorite

36. Provide the formula for each of the following ternary ionic compounds.
(a) chromium(III) chlorate
(b) lead(II) sulfite
(c) tin(IV) carbonate
(d) iron(II) chromate

37. Predict the chemical formula for each of the following ternary ionic compounds given the formula of sodium sulfate, Na_2SO_4.
(a) francium sulfate
(b) sodium sulfite

38. Predict the chemical formula for each of the following ternary ionic compounds given the formula of scandium nitrate, $Sc(NO_3)_3$.
(a) lanthanum nitrate
(b) scandium nitrite

39. Predict the chemical formula for each of the following ternary ionic compounds given the formula of barium chlorate, $Ba(ClO_3)_2$.
(a) radium chlorate
(b) barium bromate

40. Predict the chemical formula for each of the following ternary ionic compounds given the formula of iron(III) sulfate, $Fe_2(SO_4)_3$.
(a) cobalt(III) sulfate
(b) iron(III) selenate

Latin System of Nomenclature (Sec. 6.7)

41. Supply the Latin system name for each of the following monoatomic cations.
(a) Co^{2+} **(b)** Co^{3+}
(c) Sn^{2+} **(d)** Sn^{4+}

42. Supply the Latin system name for each of the following monoatomic cations.
(a) Fe^{2+} **(b)** Fe^{3+}
(c) Pb^{2+} **(d)** Pb^{4+}

43. Provide the formula for each of the following monoatomic cations.
(a) cuprous ion **(b)** cupric ion
(c) plumbous ion **(d)** plumbic ion

44. Provide the formula for each of the following monoatomic cations.
(a) mercurous ion **(b)** mercuric ion
(c) stannous ion **(d)** stannic ion

45. Supply a Latin system name for each of the following ionic compounds.
(a) CuS **(b)** FeO
(c) $Hg(NO_2)_2$ **(d)** $Pb(C_2H_3O_2)_2$

46. Supply a Latin system name for each of the following ionic compounds.
(a) Cu_2S **(b)** Fe_2O_3
(c) $Hg_2(NO_2)_2$ **(d)** $Pb(C_2H_3O_2)_4$

47. Provide the formula for each of the following ionic compounds.
(a) cupric chloride **(b)** ferrous sulfide
(c) cobaltic acetate **(d)** stannous chromate

48. Provide the formula for each of the following ionic compounds.
(a) cuprous iodide **(b)** ferric nitride
(c) cobaltous phosphate **(d)** stannic hypochlorite

Binary Molecular Compounds (Sec. 6.8)

49. Provide the formula for each of the following binary molecular compounds.
(a) dinitrogen trioxide **(b)** carbon tetrachloride
(c) iodine monobromide **(d)** dihydrogen sulfide

50. Provide the formula for each of the following binary molecular compounds.
(a) chlorine dioxide **(b)** sulfur tetrafluoride
(c) iodine monochloride **(d)** nitrogen monoxide

51. Supply a systematic name for each of the following binary molecular compounds.
(a) BrF **(b)** CF_4
(c) I_2O_4 **(d)** Cl_2O_5

52. Supply a systematic name for each of the following binary molecular compounds.
(a) $BrCl$ **(b)** SF_6
(c) I_4O_9 **(d)** Cl_2O_7

Binary Acids (Sec. 6.9)

53. Provide the formula for each of the following binary acids.
(a) hydrosulfuric acid **(b)** hydrobromic acid

54. Supply the IUPAC systematic name for each of the following binary acids.
(a) $HF(aq)$ **(b)** $HI(aq)$

Ternary Oxyacids (Sec. 6.10)

55. Provide the formula for each of the following ternary oxyacids.
(a) acetic acid **(b)** nitric acid

56. Provide the formula for each of the following ternary oxyacids.
(a) carbonic acid **(b)** nitrous acid

57. Supply the IUPAC systematic name for each of the following ternary oxyacids.
(a) $HClO_2(aq)$ **(b)** $H_3PO_4(aq)$

58. Supply the IUPAC systematic name for each of the following ternary oxyacids.
(a) $HClO_4(aq)$ **(b)** $H_2SO_3(aq)$

General Exercises

59. State the ionic charge for each of the following substances.
(a) iron metal atoms **(b)** ferrous ions
(c) iron(III) ions **(d)** iron compounds

60. State the ionic charge for each of the following substances.
(a) chlorine gas molecules **(b)** chloride ions
(c) hypochlorite ion **(d)** chlorine compounds

61. Predict which of the following polyatomic anions has an ionic charge of 2−. (*Hint:* The total number of valence electrons must be an even number.)
(a) periodate ion, $IO_4^{?-}$ **(b)** silicate ion, $SiO_3^{?-}$

62. Predict which of the following polyatomic anions has an ionic charge of 1−. (*Hint:* The total number of valence electrons must be an even number.)
(a) thiocyanate ion, $CNS^{?-}$ **(b)** thiosulfate ion, $S_2O_3^{?-}$

63. Complete the table below by combining cations and anions into chemical formulas. Give the Stock system name for each of the compounds.

IONS	chloride ion	sulfide ion	phosphide ion
copper(I) ion	$CuCl$ copper(I) chloride		
cobalt(III) ion			
lead(IV) ion			

64. Complete the table below by combining cations and anions into chemical formulas. Give the Stock system name for each of the compounds.

IONS	fluoride ion	oxide ion	nitride ion
copper(II) ion	CuF_2 copper(II) fluoride		
cobalt(II) ion			
lead(II) ion			

65. Complete the table below by combining cations and oxyanions into chemical formulas. Give the Latin name for each of the compounds.

IONS	hydroxide ion	sulfate ion	phosphate ion
mercurous ion	$Hg_2(OH)_2$ mercurous hydroxide		
ferrous ion			
stannous ion			

66. Complete the table below by combining cations and oxyanions into chemical formulas. Give the Latin system name for each of the compounds.

IONS	cyanide ion	sulfite ion	permanganate ion
mercuric ion	$Hg(CN)_2$ mercuric cyanide		
ferric ion			
stannic ion			

67. State the suffix in the name for each of the following.
(a) Na_2S **(b)** $H_2S(aq)$

68. State the suffix in the name for each of the following.
(a) NaI **(b)** $HI(aq)$

69. State the suffix in the name for each of the following.
(a) Na_2SO_3 **(b)** $H_2SO_3(aq)$

70. State the suffix in the name for each of the following.
(a) $NaIO_2$ **(b)** $HIO_2(aq)$

71. State the suffix in the name for each of the following.
(a) Na_2SO_4 **(b)** $H_2SO_4(aq)$

72. State the suffix in the name for each of the following.
(a) $NaIO_3$ **(b)** $HIO_3(aq)$

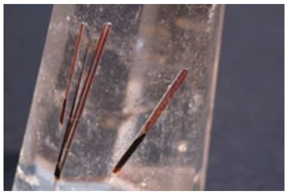
73. Write the chemical formula for each of the following household chemicals.
(a) dihydrogen oxide (common liquid)
(b) sodium hypochlorite (bleach)
(c) sodium hydroxide (caustic soda)
(d) sodium bicarbonate (baking soda)

74. Write the chemical formula for each of the following household chemicals.
(a) acetic acid (vinegar solution)
(b) aqueous nitrogen trihydride (ammonia solution)
(c) aqueous magnesium hydroxide (milk of magnesia)
(d) aqueous sodium hydrogen sulfate (porcelain cleaner)

75. Give the name for each of the following compounds containing a semimetal. (*Hint:* Name the formula as a binary molecular compound.)
(a) BF_3 **(b)** $SiCl_4$
(c) As_2O_5 **(d)** Sb_2O_3

76. Write the formula for each of the following compounds containing a semimetal.
(a) boron tribromide **(b)** trisilicon tetranitride
(c) diarsenic trioxide **(d)** diantimony pentaoxide

77. The compound titanium oxide occurs in the mineral rutile. Use the periodic table to predict the formula for titanium oxide, given the formula of zirconium oxide, ZrO_2.



◀ **Rutile, TiO_2** Rutile, titanium dioxide, is a mineral found in nature that contains titanium.

78. The compound zirconium silicate occurs in cubic zirconia. Use the periodic table to predict the formula for zirconium silicate, given the formula of titanium silicate, $TiSiO_4$.

Challenge Exercises

79. The element actinium is unstable and radioactive. Use the periodic table to predict the formula for actinium chloride, given the formula of lanthanum chloride, $LaCl_3$.

80. The element lawrencium is unstable and radioactive. Use the periodic table to predict the chemical formula for lawrencium chloride, given the formula of lutetium chloride, $LuCl_3$.

Chapter 6 Self-Test

Answers to Self-Test are in Appendix J.

1. Gaseous HCl is classified as which of the following? (Sec. 6.1)
(a) binary ionic compound
(b) ternary ionic compound
(c) binary molecular compound
(d) binary acid
(e) ternary oxyacid

2. What is the Stock system name for Hg^{2+}? (Sec. 6.2)
(a) mercury(I) ion **(b)** mercury(II) ion
(c) mercuric ion **(d)** mercurous ion
(e) none of the above

3. What is the chemical formula for the hypochlorite ion? (Sec. 6.3)
(a) ClO^- **(b)** ClO_2^-
(c) ClO_3^- **(d)** ClO_4^-
(e) none of the above

4. What is the chemical formula for the ternary compound composed of Cr^{3+} and OH^- ions? (Sec. 6.4)
(a) CrOH **(b)** $CrOH_3$
(c) $Cr(OH)_3$ **(d)** Cr_3OH
(e) none of the above

5. What is the Stock system name for CuS? (Sec. 6.5)
(a) cuprous sulfide **(b)** cupric sulfide
(c) copper(I) sulfide **(d)** copper(II) sulfide
(e) none of the above

6. What is the Stock system name for $CuSO_4$? (Sec. 6.6)
(a) cuprous sulfate
(b) cupric sulfate
(c) copper(I) sulfate
(d) copper(II) sulfate
(e) none of the above

7. What is the Latin system name for $CuSO_3$? (Sec. 6.7)
(a) cuprous sulfite **(b)** cupric sulfite
(c) copper(I) sulfite **(d)** copper(II) sulfite
(e) none of the above

8. What is the systematic name for gaseous H_2S? (Sec. 6.8)
(a) hydrogen sulfide **(b)** hydrosulfuric acid
(c) sulfurous acid **(d)** sulfuric acid
(e) none of the above

9. What is the systematic name for aqueous H_2S? (Sec. 6.9)
(a) hydrogen sulfide
(b) hydrosulfuric acid
(c) sulfurous acid
(d) sulfuric acid
(e) none of the above

10. What is the systematic name for aqueous H_2SO_4? (Sec. 6.10)
(a) hydrogen sulfide
(b) hydrosulfuric acid
(c) sulfurous acid
(d) sulfuric acid
(e) none of the above

Key Concepts

11. Classify each of the following as binary ionic, ternary ionic, binary molecular, binary acid, or ternary oxyacid: NaCl, HCl, HCl(*aq*), $NaClO_3$, $HClO_3$(*aq*).

Hydrochloric Acid, HCl Aqueous hydrochloric acid, HCl, is found in swimming pools and is produced in the stomach to aid digestion; hence, it is commonly referred to as "stomach acid."

12. Which of the following compounds is named using the suffix *-ide:* NaCl, $NaClO_2$, $NaClO_3$?

13. Which of the following compounds is named using the suffix *-ate:* NaCl, $NaClO_2$, $NaClO_3$?

14. Which of the following compounds is named using the suffix *-ite:* NaCl, $NaClO_2$, $NaClO_3$?

15. Which of the following acids is named *hydro* + *nonmetal stem* + *-ic acid:* HCl(*aq*), $HClO_3$(*aq*), $HClO_2$(*aq*)?

16. Which of the following acids is named *nonmetal stem* + *-ic acid:* HCl(*aq*), $HClO_3$(*aq*), $HClO_2$(*aq*)?

17. Which of the following acids is named *nonmetal stem* + *-ous acid:* HCl(*aq*), $HClO_3$(*aq*), $HClO_2$(*aq*)?

Critical Thinking

18. Examine the periodic table and propose an explanation why copper ions can have a charge of either 1+ or 2+.

19. Examine the periodic table and propose an explanation why iron ions can have a charge of either 2+ or 3+.

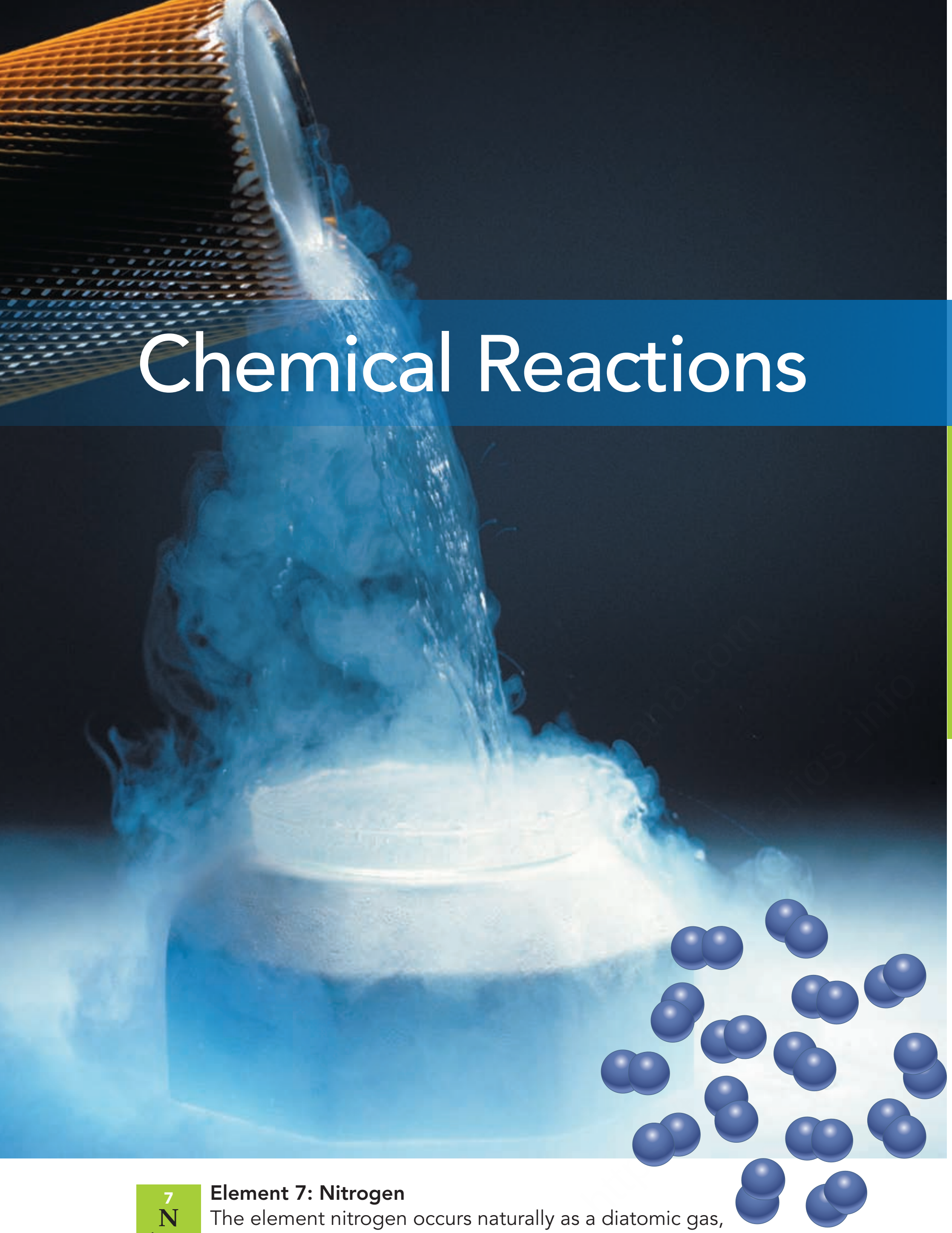
20. Chalk and marble are both examples of limestone, $CaCO_3$. Why is chalk a soft substance, while marble is a hard mineral?

CHAPTER 7

Chemical Reactions

"Science is built of facts the way a house is built of bricks; but an accumulation of facts is no more science than a pile of bricks is a house."

Henri Poincare, French Physicist (1854–1912)



7
N
nitrogen

Element 7: Nitrogen
The element nitrogen occurs naturally as a diatomic gas, N_2, which comprises about 80% of Earth's atmosphere. Nitrogen gas is quite unreactive and can be liquefied at a temperature of −196 °C. Liquid nitrogen is used in a dermatologist's office to freeze skin and remove skin lesions.

Every day substances are constantly being altered in some way. These alterations can be classified as a physical or a chemical change. In a *physical change*, the chemical composition of the substance remains constant. For instance, when we melt ice or boil water, we cause a physical change. The chemical composition of the ice and the water, H_2O, is the same before and after the change.

In a *chemical change*, the chemical composition of the substance does not remain constant. For example, when we drop an antacid tablet into a glass of water, we produce a chemical change. The chemical composition of the antacid tablet is not the same before and after the change. The evidence for the formation of a new substance is the gas bubbles that are released in the glass of water.

We may observe a *chemical reaction* when a substance undergoes a chemical change and forms a new substance. When we form a new substance, we usually observe one of the following: a change in color, an odor or bubbles indicating the

A CLOSER LOOK Nitrogen in Tires

Q: **Why is nitrogen preferable to air for the inflation of tires?**

Air contains about 78% nitrogen, 21% oxygen, and 1% other gases including water vapor. In an automobile tire, air molecules permeate the sidewalls of a rubber tire and escape six times faster than nitrogen molecule; thus causing decreased pressure and shorter tire life. In addition, oxygen gas degrades the inner liner of the tire, and can cause corrosion of steel, aluminum, and alloy tire rims.

For some time, the Federal Aviation Administration (FAA) has mandated aircraft tires be filled with nitrogen gas for safety reasons. Under the harsh conditions of landing an airplane, the tires get very hot and oxygen in the tire can undergo combustion with the rubber tire. Recently, several automobile tire manufacturers have recommended nitrogen gas for increased tire life, as well as for safety reasons.

Another reason to inflate tires with nitrogen is that compressed air contains water vapor, which can freeze under cold conditions, and expand under heated conditions. Nitrogen has significantly less moisture than compressed air. By eliminating moisture, tire pressure is more stable and predictable, which extends the life of the tire.

A: There are several reasons nitrogen is preferable to air in tires, including increased tire life and greater safety.

liberation of a gas, or light or heat from the release of energy. For instance, we can mix flour, yeast, and water and put the mixture into an oven, and in a short time we can smell the odor of bread baking. A chemical reaction has occurred in which the flour, yeast, and water have combined to produce a new substance, evidenced by a change in color and smell.

LEARNING OBJECTIVE

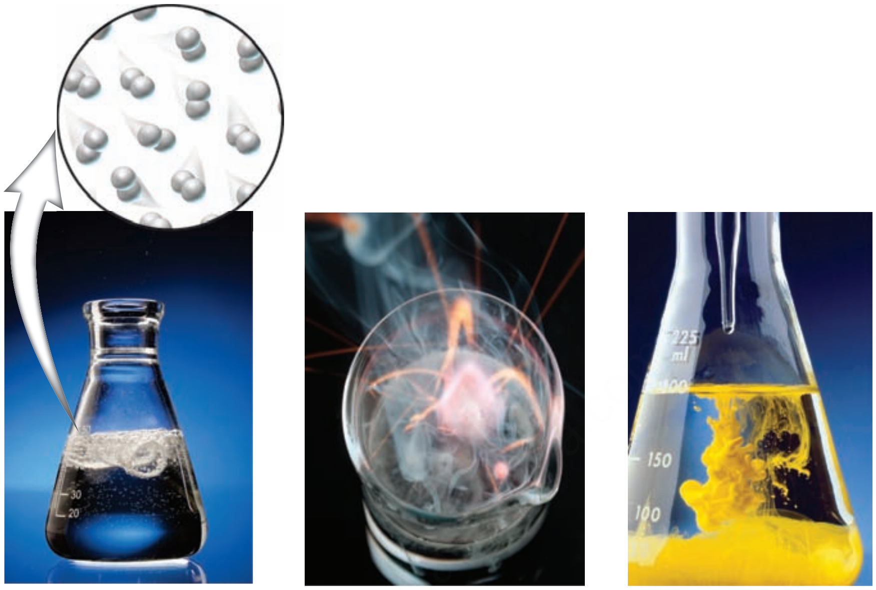
- Describe four observations that are evidence for a chemical reaction.

7.1 Evidence of Chemical Reactions

If we combine substances, the resulting change can be physical or chemical. We can distinguish between a physical and a chemical change by observing the change. Each of the following indicates a chemical change; that is, a **chemical reaction** has taken place.

1. **A gas is produced.** We can observe a gas in a number of ways; for example, we may see bubbles in a solution. When we add an Alka-Seltzer tablet to water, it begins to fizz. The fizzing is evidence of a chemical reaction and the release of carbon dioxide gas, which is produced by the reaction of citric acid and baking soda in the tablet.
2. **An insoluble solid is produced in a solution.** A substance dissolves in water to give an **aqueous solution**. If we add two aqueous solutions together, we may observe solid particles in the solution. This insoluble solid is called a **precipitate**. The formation of a precipitate is evidence of a reaction.
3. **A permanent color change is observed.** Many chemical reactions involve a permanent change in color. For example, adding aqueous ammonia to a copper(II) sulfate solution changes the color from light blue to deep blue. To observe an acid reacting with a base, we use an indicator, a substance that changes color. The indicator enables us to indirectly follow a reaction that would otherwise not be visible.
4. **An energy change is observed.** In chemical reactions there is always a change in energy. A reaction that releases heat is said to be an **exothermic reaction**; conversely, a reaction that absorbs heat is said to be an **endothermic reaction**. For example, when an emergency flare is ignited, heat and light are observed. Heat and light are two forms of energy indicating that an exothermic chemical reaction has taken place.

In summary, each of the above four criteria—the production of a gas, the formation of a precipitate, a change in color, or a change in energy—is a signal that a chemical reaction has occurred. This is shown in Figure 7.1. A chemical reaction, however, may have occurred even though no change is obvious. In some reactions, for example, acid–base reactions, the energy change may be too subtle to notice and an external indicator may be necessary.



(a) (b) (c)

◀ **Figure 7.1 Evidence of a Chemical Reaction** Each of these images is evidence of a chemical reaction: (a) the gas bubbles released from magnesium metal in acid; (b) the heat and light released from potassium metal in water; and (c) the color change and the precipitate produced by two aqueous solutions.

EXAMPLE EXERCISE 7.1 Evidence of a Chemical Reaction

Which of the following is experimental evidence of a chemical reaction?

(a) Pouring vinegar on baking soda gives foamy bubbles.
(b) Mixing two solutions produces insoluble particles.
(c) Mixing two colorless solutions gives a yellow solution.
(d) Mixing two solutions produces a temperature increase.

Solution

We can analyze each of these observations based on the criteria for a chemical reaction.

(a) The bubbles produced indicate a chemical reaction is occurring.
(b) The insoluble particles formed indicate a chemical reaction.
(c) The yellow color produced indicates a chemical reaction.
(d) The temperature increase indicates heat energy is being released from a chemical reaction.

Practice Exercise

What are four observations that a chemical reaction has occurred?

Answers:

(a) a gas is released
(b) a precipitate is produced
(c) a permanent color change is observed
(d) an energy change is noted

Concept Exercise

If two aqueous solutions are added together, how could you determine if there is an exothermic or endothermic reaction?

Answer: See Appendix G, 7.1.

CHEMISTRY CONNECTION Fireworks

Q: **What element is responsible for the red-orange color in the fireworks display?**

Have you ever watched a Fourth of July fireworks display and wondered how the colors were produced? The color of fireworks is produced by the chemicals in a rocket shell, which is fitted with a fuse and fired into the air. When the fuse burns, gunpowder ignites and sets off an explosion that shoots chemicals throughout the sky. If the rocket shell is packed with a sodium compound, an orange-yellow color is observed. If the shell is packed with a barium compound, a yellow-green color is observed.

◀ **Fourth of July Fireworks**

However, not all chemicals produce a colored display. You may recall having seen fireworks that simply produce a shower of white sparks. Powdered aluminum or magnesium metals are used to produce this effect. The following table lists the chemicals used to produce various colors in fireworks:

CHEMICAL	FIREWORKS COLOR
Na compounds	orange-yellow
K compounds	violet
Ba compounds	yellow-green
Ca compounds	red-orange
Sr compounds	bright red
Li compounds	scarlet-red
Cu compounds	blue-green
Al or Mg metals	white sparks

Interestingly, the colors of fireworks and the colors of neon signs are based on the same principle. In a fireworks display an element is energized by heat, whereas in a "neon light" an element is energized by electricity. In either case, electrons are first excited and then immediately lose energy by emitting light. The observed color corresponds to the wavelength of the light emitted. For example, if the emitted light has a wavelength near 650 nm, it is observed as red. If the light has a wavelength near 450 nm, it appears blue.

Chemists can identify an element by the characteristic color of its flame test. In the laboratory, a flame test is performed by placing a small amount of chemical on the tip of a wire and holding the wire in a hot flame. For example, sodium gives an orange-yellow flame test, and barium a yellow-green flame test. Flame-test colors are identical to fireworks colors containing the same element. The red-orange flame-test color of calcium is illustrated.

◀ **Calcium Flame Test**
The flame test for calcium is red-orange.

A: The red-orange color suggests that the element is calcium.

LEARNING OBJECTIVES

- Identify seven elements that occur naturally as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.
- Write a chemical equation from the description of a chemical reaction.

7.2 Writing Chemical Equations

In this chapter we will use formulas and symbols to describe a chemical reaction; that is, we will write a **chemical equation**. Consider the following description of a chemical reaction: substance A and substance B undergo a chemical change that produces substances C and D. We can state this description in symbols as follows:

$$A + B \longrightarrow C + D$$

In this general chemical equation, A and B are each a **reactant**, and C and D are each a **product**.

We can provide more information about the reaction if we specify the physical state of each substance, that is, solid, liquid, or gas. The physical state is indicated by the abbreviation (*s*), (*l*), or (*g*). By convention, we indicate an aqueous solution by the symbol (*aq*) and a precipitate by the symbol (*s*). For example, an aqueous solution of A reacts with a gaseous substance B to yield a precipitate C and an aqueous solution D. This statement can be written as follows:

$$A(aq) + B(g) \rightarrow C(s) + D(aq)$$

A **catalyst** is a substance that speeds up a reaction without being consumed or permanently altered. For example, the catalytic converter in an automobile contains a metallic catalyst that speeds up the conversion of unburned fuel to carbon dioxide and water. The presence of a catalyst is indicated by placing its formula above the arrow. Table 7.1 lists the symbols that are used to describe chemical reactions.

Let's interpret the symbols in the following chemical equation for the reaction of acetic acid and baking soda.

$$HC_2H_3O_2(aq) + NaHCO_3(s) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) + CO_2(g)$$

The formulas and symbols are read as follows: aqueous acetic acid is added to solid sodium hydrogen carbonate, which produces aqueous sodium acetate, water, and carbon dioxide gas.

TABLE 7.1 Chemical Equation Symbols

Symbol	Interpretation of Chemical Equation Symbol
$\rightarrow$	produces, yields, gives (separates reactants from products)
+	reacts with, added to, plus (separates two or more reactants or products)
$\xrightarrow{\Delta}$	heat is a catalyst for the reaction
$\xrightarrow{Fe}$	iron is a catalyst for the reaction
NR	no reaction
(*s*)	solid substance or precipitate
(*l*)	liquid substance
(*g*)	gaseous substance
(*aq*)	aqueous solution

Chemical reactions often involve hydrogen, nitrogen, or oxygen gases. These nonmetals occur naturally as **diatomic molecules**, that is, H_2, N_2, and O_2. Moreover, the halogens also occur naturally as diatomic molecules and are therefore written as F_2, Cl_2, Br_2, and I_2. If one of these elements appears in a chemical equation, it is written as a diatomic molecule.

The following example exercise further illustrates how to write a chemical equation from the description of a chemical reaction.

Helpful Hint
Diatomic Molecules
The following elements occur naturally as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2. If oxygen gas is a reactant or product in a chemical reaction, it is written as O_2.

EXAMPLE EXERCISE 7.2 Writing Chemical Equations

Write a chemical equation for each of the following chemical reactions:

(a) Mercury liquid and fluorine gas react to give solid mercury(II) fluoride.
(b) Zinc metal reacts with sulfuric acid to give aqueous zinc sulfate and hydrogen gas.

Solution

To write the chemical equation, we must provide formulas and symbols for each substance. We can describe each of the preceding chemical reactions as follows:

(a) $Hg(l) + F_2(g) \rightarrow HgF_2(s)$
(b) $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Practice Exercise

Write a chemical equation for each of the following chemical reactions:

(a) Aqueous solutions of sodium iodide and silver nitrate yield silver iodide precipitate and aqueous sodium nitrate.

(b) Acetic acid reacts with aqueous potassium hydroxide to give aqueous potassium acetate plus water.

Answers:

(a) $NaI(aq) + AgNO_3(aq) \rightarrow AgI(s) + NaNO_3(aq)$

(b) $HC_2H_3O_2(aq) + KOH(aq) \rightarrow KC_2H_3O_2(aq) + H_2O(l)$

Concept Exercise

Which seven nonmetals occur naturally as diatomic molecules?

Answers: See Appendix G, 7.2.

7.3 Balancing Chemical Equations

LEARNING OBJECTIVE

- Convert a chemical reaction into a balanced chemical equation.

In the previous section we translated the description of a chemical reaction into a chemical equation. In each case, the number of atoms of each element was equal in the reactants and products. Because the number of atoms of each element is the same on each side of the arrow, the chemical reaction is said to be a balanced chemical equation.

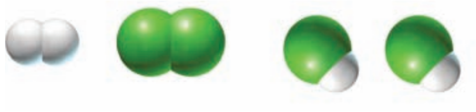
More often, the number of atoms of each element in the reactants and products is not the same. Therefore, it is necessary to balance the number of atoms of each element on each side of the chemical reaction. We balance a chemical reaction by placing a whole number **coefficient** in front of each substance. It is important to note that a coefficient multiplies all subscripts in the chemical formula that follows. That is, 3 H_2O indicates six hydrogen atoms and three oxygen atoms.

Let's consider the reaction of hydrogen gas with chlorine gas to give hydrogen chloride, a gas with a sharp odor. We can write the chemical reaction as follows:

$$H_2(g) + Cl_2(g) \longrightarrow HCl(g)$$

Notice that the **subscript** for H_2 and Cl_2 is 2, but that only one H atom and one Cl atom appear in the product. Thus, the equation is not balanced. To balance the number of H atoms on both sides of the equation, we place the coefficient 2 in front of the HCl. This gives us

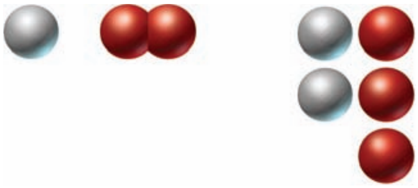
$$H_2(g) + Cl_2(g) \longrightarrow 2\ HCl(g)$$



After adding the reaction coefficient 2, we have two molecules of HCl. That is, we have two H atoms and two Cl atoms on each side of the equation. The equation is now balanced.

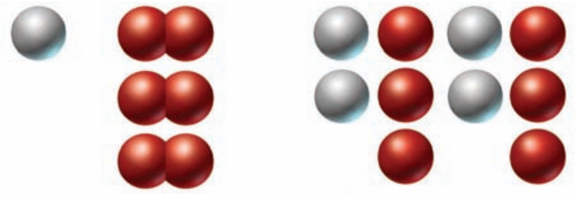
Let's try a more difficult example. Aluminum metal is heated with oxygen gas to give solid aluminum oxide. The formula for aluminum oxide is Al_2O_3. The chemical equation is as follows:

$$\mathrm{Al}(s) + \mathrm{O_2}(g) \longrightarrow \mathrm{Al_2O_3}(s)$$



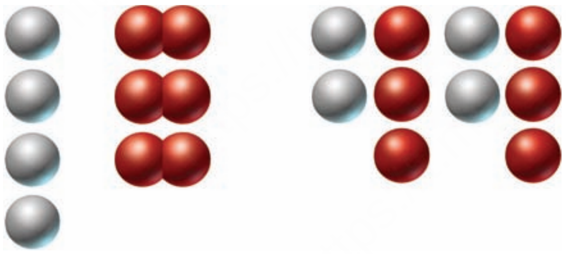
Notice that two O atoms appear in the reactants but that three O atoms appear in the product. This reaction is not balanced. To balance the numbers of O atoms, we will use the lowest common multiple of 2 and 3; the number is 6. We place the coefficient 3 in front of the O_2 to give six oxygen atoms in the reactants and the coefficient 2 in front of Al_2O_3 to give six oxygen atoms in the products. That gives us

$$\mathrm{Al}(s) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{Al_2O_3}(s)$$



The numbers of Al atoms, however, are not balanced. On the reactant side we have one Al atom. On the product side, we have two units of Al_2O_3, for a total of four Al atoms. We can place the coefficient 4 in front of the reacting Al metal. This gives a balanced chemical equation:

$$4\,\mathrm{Al}(s) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{Al_2O_3}(s)$$



This method of placing coefficients by systematically analyzing each side of an equation is called *balancing by inspection*. Although there is no formal prescription for balancing an equation by inspection, we can list some general guidelines.

General Guidelines for Balancing a Chemical Equation

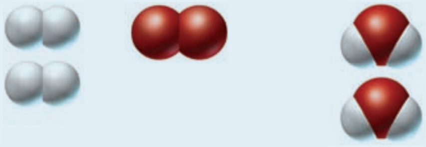
1. Before placing a coefficient in an equation, verify the formula of each substance for correct subscripts. Note that subscripts of formulas are never changed to balance an equation. In fact, if an equation is difficult to balance, it is usually because of an incorrect subscript in the formula.
2. Balance each element in the equation by placing a coefficient in front of each substance. Coefficients of 1 are assumed and do not appear in the balanced chemical equation.
    (a) Begin balancing the equation starting with the most complex formula.
    (b) Balance polyatomic ions as a single unit. However, if a polyatomic ion breaks down, you must balance each atom separately.

(c) The coefficients in an equation must be whole numbers. Occasionally, it is helpful to use a fractional coefficient to balance an element in a diatomic molecule; for example,

$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l)$$

If a fraction is used, such as 1/2, we must then multiply the equation by 2 to obtain whole number coefficients. Thus, the balanced equation is

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$$



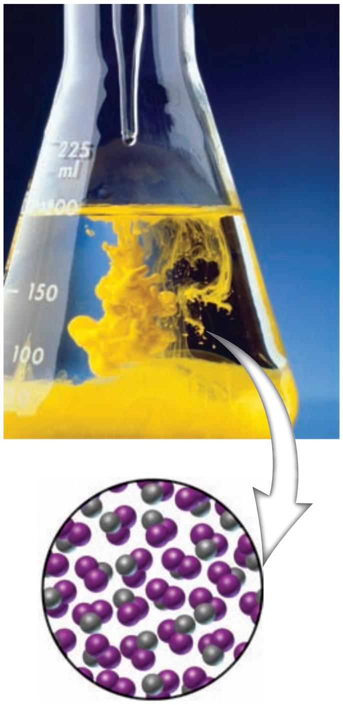
3. After balancing the equation, check (√) each symbol of every element (or polyatomic ion) to verify that the coefficients are correct. Proceed back and forth between reactants and products. The procedure for verification is to multiply the coefficient times the subscript of each element; the totals should be the same on both sides of the equation.

Finally, check the coefficients to make sure they represent the smallest whole number ratio. It may be possible to obtain a smaller set of coefficients. Although the following equation is balanced,

$$2\,\overset{\surd}{N}_2(g) + 6\,\overset{\surd}{H}_2(g) \rightarrow 4\,\overset{\surd}{N}H_3(g)$$

you must reduce the coefficients by a factor of 2 to obtain the correct balanced equation:

$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$



▲ Insoluble Lead(II) Iodide Lead(II) iodide, PbI_2, forms a yellow precipitate in an aqueous solution.

We can write an equation for the formation of the yellow precipitate shown in Figure 7.1c. In the reaction, aqueous solutions of lead(II) nitrate and potassium iodide produce a yellow precipitate of lead(II) iodide and an aqueous solution of potassium nitrate. We can write a chemical equation for the reaction as follows:

$$Pb(NO_3)_2(aq) + KI(aq) \rightarrow PbI_2(s) + KNO_3(aq)$$

Let's begin balancing with $Pb(NO_3)_2$ because it is the most complex formula. Notice that one Pb atom appears on each side of the equation. However, there are two NO_3^- ions on the reactant side and only one NO_3^- ion on the product side. We can balance the NO_3^- polyatomic ion by placing a 2 in front of KNO_3. Thus,

$$Pb(NO_3)_2(aq) + KI(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq)$$

We now have two K on the right side of the equation and only one K on the left. Thus, we need the coefficient 2 in front of KI:

$$Pb(NO_3)_2(aq) + 2\,KI(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq)$$

Notice the coefficient 2 in front of KI gives two I, which is equal to the two I in PbI_2. As a final inspection, let's check off each element and polyatomic ion to verify that we indeed have a balanced equation:

$$\overset{\surd}{Pb}(\overset{\surd}{NO})_3)_2(aq) + 2\,\overset{\surd\surd}{KI}(aq) \rightarrow \overset{\surd\surd}{PbI_2}(s) + 2\,\overset{\surd\surd}{KNO_3}(aq)$$

The following example exercises further illustrate the general guidelines for balancing a chemical equation:

EXAMPLE EXERCISE 7.3 Balancing Chemical Equations

Aqueous solutions of calcium nitrate and sodium carbonate react to give a white precipitate of calcium carbonate and aqueous sodium nitrate. Write a balanced chemical equation given

$$Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + NaNO_3(aq)$$

Solution

We see that $Ca(NO_3)_2$ has two NO_3^- ions in the formula. There is only one NO_3^- ion on the right side, so we insert the coefficient 2 as follows:

$$Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ NaNO_3(aq)$$

The coefficient 2 in front of $NaNO_3$ generates two Na on the right side, and we have two Na on the left side. Notice we have one CO_3 on each side, and one Ca on each side. Finally, we can check off each element and polyatomic ion to verify that the chemical equation is balanced.

$$\overset{\surd}{Ca}\,(\overset{\surd}{NO_3})_2(aq) + \overset{\surd}{Na_2}\overset{\surd}{CO_3}(aq) \rightarrow \overset{\surd}{Ca}\overset{\surd}{CO_3}(s) + 2\ \overset{\surd}{Na}\overset{\surd}{NO_3}(aq)$$

Practice Exercise

Aqueous solutions of aluminum sulfate and barium nitrate react to yield a white precipitate of barium sulfate and aqueous aluminum nitrate. Write a balanced chemical equation given

$$Al_2(SO_4)_3(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + Al(NO_3)_3(aq)$$

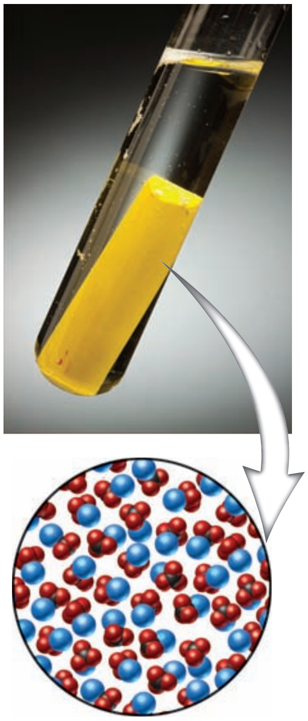
Answer:

$$\overset{\surd}{Al_2}(\overset{\surd}{SO_4})_3(aq) + 3\ \overset{\surd}{Ba}(\overset{\surd}{NO_3})_2(aq) \rightarrow 3\ \overset{\surd\surd}{BaSO_4}(s) + 2\ \overset{\surd}{Al}(\overset{\surd}{NO_3})_3(aq)$$

Concept Exercise

Do you change a subscript in a chemical formula to balance an equation?

Answer: See Appendix G, 7.3.



Insoluble Calcium Carbonate Chalk, calcium carbonate, $CaCO_3$, is insoluble in an aqueous solution.

EXAMPLE EXERCISE 7.4 Balancing Chemical Equations

Sulfuric acid reacts with aqueous sodium hydroxide to give aqueous sodium sulfate and water. Write a balanced chemical equation given

$$H_2SO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

Solution

Let's start with Na_2SO_4, which contains the same number of atoms as H_2SO_4. There is one SO_4 on each side of the equation, and so SO_4 is balanced. However, there are two Na on the right side of the equation and one Na on the left. We must place the coefficient 2 in front of NaOH.

$$H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + HOH(l)$$

To balance this reaction more easily, we can write the formula for water as *HOH*. Notice that we now have two OH on the left side of the equation, and one OH on the right side. By placing a 2 in front of the HOH, we can balance the OH. The two H on the right side are balanced by the two H in H_2SO_4. Finally, we check off each atom and polyatomic ion to verify that we have a balanced chemical equation.

$$\overset{\surd}{H_2}\overset{\surd}{SO_4}(aq) + 2\ \overset{\surd\surd}{NaOH}(aq) \rightarrow \overset{\surd}{Na_2}\overset{\surd}{SO_4}(aq) + 2\ \overset{\surd\surd}{HOH}(l)$$

Practice Exercise

Nitric acid reacts with aqueous barium hydroxide to give aqueous barium nitrate and water. Write a balanced chemical equation given

$$HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + HOH(l)$$

Answer:

$$2\ \overset{\surd}{H}\overset{\surd}{N}O_3(aq) + \overset{\surd}{Ba}(\overset{\surd}{OH})_2(aq) \rightarrow \overset{\surd}{Ba}(\overset{\surd}{NO_3})_2(aq) + 2\ \overset{\surd\surd}{HOH}(l)$$

Concept Exercise

Why write water as HOH for the product of a reaction that produces water?

Answer: See Appendix G, 7.4.

Note Balancing an equation is a straightforward task. Start with the most complex formula and balance systematically, back and forth, between reactants and products. If you encounter difficulty in the final step, it is usually because the original equation had an incorrect subscript in a chemical formula.

LEARNING OBJECTIVE

- Classify a chemical reaction as one of the following types: combination, decomposition, single-replacement, double-replacement, or neutralization.

7.4 Classifying Chemical Reactions

Because there are millions of chemical reactions, how can we attempt to master such a large topic? The answer lies in the same way that we learned previously to master chemical formulas. We can *classify chemical reactions* and put them in categories.

In this section, we will begin with five simple types of reactions: combination, decomposition, single-replacement, double-replacement, and neutralization. We can briefly describe these five types as follows:

1. A **combination reaction** involves simpler substances being combined into a more complex compound. It is also referred to as a *synthesis reaction*. In the general case, substance A combines with substance Z to produce compound AZ. The chemical equation is

$$A + Z \rightarrow AZ$$

2. In a **decomposition reaction** a single compound is broken down into two or more simpler substances. In this case, heat is usually applied to decompose the compound. In the general case, compound AZ decomposes into substances A and Z. The chemical equation is

$$AZ \xrightarrow{\Delta} A + Z$$

3. In a **single-replacement reaction** a more reactive element displaces another element from a compound or aqueous solution. The substance that is displaced shows less tendency to undergo reaction; that is, it is less active. In one case, metal A replaces metal B in BZ to give AZ and B. The chemical equation is

$$A + BZ \rightarrow AZ + B$$

4. In a **double-replacement reaction** two compounds exchange anions. Compound AX reacts with compound BZ to yield the products AZ and BX. The general form of the reaction is

$$AX + BZ \rightarrow AZ + BX$$

5. An acid and a base react with each other in a **neutralization reaction**. Acid HX reacts with base BOH to give the compound BX and water. We can write the general form of the reaction as

$$HX + BOH \rightarrow BX + HOH$$

When we examine a neutralization reaction, we see that it is actually a special type of double-replacement reaction. An acid and a base simply switch anions. Notice that we wrote water as HOH rather than H_2O. Using the formula HOH shows the double-replacement nature of the neutralization reaction and makes it easier to balance the equation.

EXAMPLE EXERCISE 7.5 Classifying Chemical Reactions

Classify each of the following reactions as combination, decomposition, single-replacement, double-replacement, or neutralization.

(a) Copper metal heated with oxygen gas produces solid copper(II) oxide.

$$2\,\mathrm{Cu}(s) + \mathrm{O_2}(g) \xrightarrow{\Delta} 2\,\mathrm{CuO}(s)$$

(b) Heating powdered iron(III) carbonate produces solid iron(III) oxide and carbon dioxide gas.

$$\mathrm{Fe_2(CO_3)_3}(s) \xrightarrow{\Delta} \mathrm{Fe_2O_3}(s) + 3\,\mathrm{CO_2}(g)$$

(c) Aluminum metal reacts with aqueous manganese(II) sulfate to give aqueous aluminum sulfate and manganese metal.

$$2\,\mathrm{Al}(s) + 3\,\mathrm{MnSO_4}(aq) \rightarrow \mathrm{Al_2(SO_4)_3}(aq) + 3\,\mathrm{Mn}(s)$$

(d) Aqueous sodium chromate reacts with aqueous barium chloride to give insoluble barium chromate and aqueous sodium chloride.

$$\mathrm{Na_2CrO_4}(aq) + \mathrm{BaCl_2}(aq) \rightarrow \mathrm{BaCrO_4}(s) + 2\,\mathrm{NaCl}(aq)$$

(e) Nitric acid reacts with aqueous potassium hydroxide to give aqueous potassium nitrate and water.

$$\mathrm{HNO_3}(aq) + \mathrm{KOH}(aq) \rightarrow \mathrm{KNO_3}(aq) + \mathrm{H_2O}(l)$$

Solution

We can classify the type of each reaction as follows:

(a) The two elements Cu and O_2 synthesize a single compound; this is an example of a *combination reaction*.

(b) The compound $Fe_2(CO_3)_3$ is heated and breaks down into a simpler compound and a gas; this is an example of a *decomposition reaction*.

(c) The metal Al displaces the metal Mn from aqueous $MnSO_4$; this is an example of a *single-replacement reaction*.

(d) The two compounds Na_2CrO_4 and $BaCl_2$ exchange anions; this is an example of a *double-replacement reaction*.

(e) The acid HNO_3 reacts with the base KOH to form KNO_3 and water; this is an example of a *neutralization reaction*.

Practice Exercise

Classify the following reactions as combination, decomposition, single-replacement, double-replacement, or neutralization:

(a) $\mathrm{Zn}(s) + \mathrm{CuSO_4}(aq) \rightarrow \mathrm{ZnSO_4}(aq) + \mathrm{Cu}(s)$

(b) $2\,\mathrm{Sr}(s) + \mathrm{O_2}(g) \rightarrow 2\,\mathrm{SrO}(s)$

(c) $\mathrm{Cd(HCO_3)_2}(s) \rightarrow \mathrm{CdCO_3}(s) + \mathrm{H_2O}(g) + \mathrm{CO_2}(g)$

(d) $\mathrm{HC_2H_3O_2}(aq) + \mathrm{NaOH}(aq) \rightarrow \mathrm{NaC_2H_3O_2}(aq) + \mathrm{H_2O}(l)$

(e) $\mathrm{AgNO_3}(aq) + \mathrm{KCl}(aq) \rightarrow \mathrm{AgCl}(s) + \mathrm{KNO_3}(aq)$

Answers:

(a) single-replacement
(b) combination
(c) decomposition
(d) neutralization
(e) double-replacement

Concept Exercise

How can a decomposition reaction be easily distinguished from other types of chemical reactions?

Answer: See Appendix G, 7.5.

LEARNING OBJECTIVES

- Propose a balanced chemical equation for the reaction of a metal and oxygen gas.
- Propose a balanced chemical equation for the reaction of a nonmetal and oxygen gas.
- Propose a balanced chemical equation for the reaction of a metal and a nonmetal.

7.5 Combination Reactions

There are many examples of combination reactions. For instance, heating a metal in air can cause the metal to combine with oxygen and form a compound. In this section, we will study three important kinds of combination reactions.

Metal and Oxygen Gas

In one kind of combination reaction, a metal and oxygen gas react to give a metal oxide. In the following example, magnesium metal and oxygen gas combine to give magnesium oxide:

$$\textbf{metal} + \textbf{oxygen gas} \rightarrow \textbf{metal oxide}$$

$$2\,\mathrm{Mg}(s) + \mathrm{O_2}(g) \xrightarrow{\Delta} 2\,\mathrm{MgO}(s)$$

▲ **Igniting Magnesium Metal** A strip of magnesium metal reacts with oxygen in air to give white smoke, MgO, and a bright white light.

A metal and oxygen react to give a metal oxide compound. We can usually predict the formulas of metal oxides containing a main-group metal. On the other hand, we cannot predict the formulas of most metal oxides containing a transition metal. For transition metal oxides, the ionic charge of the metal must be given to write the chemical formula of the metal oxide.

The following example exercise further illustrates how to write balanced equations for a combination reaction of a metal and oxygen gas:

EXAMPLE EXERCISE 7.6 Combination Reaction of a Metal and Oxygen Gas

Write a balanced chemical equation for each of the following combination reactions:

(a) Zinc metal is heated with oxygen gas in air to yield solid zinc oxide.
(b) Chromium metal is heated with oxygen gas to give solid chromium(III) oxide.

Solution

A metal and oxygen react to produce a metal oxide.

(a) Zinc is a metal with a predictable charge, that is, $\mathrm{Zn^{2+}}$. The formula of zinc oxide is ZnO. The balanced equation for the reaction is

$$2\,\mathrm{Zn}(s) + \mathrm{O_2}(g) \xrightarrow{\Delta} 2\,\mathrm{ZnO}(s)$$

(b) Chromium is a metal with a variable charge. From the name chromium(III) oxide, we know the ion is $\mathrm{Cr^{3+}}$. The formula of the oxide is, therefore, $\mathrm{Cr_2O_3}$, and the balanced equation for the reaction is

$$4\,\mathrm{Cr}(s) + 3\,\mathrm{O_2}(g) \xrightarrow{\Delta} 2\,\mathrm{Cr_2O_3}(s)$$

Practice Exercise

Write a balanced chemical equation for each of the following combination reactions:

(a) Lead metal is heated with oxygen in air to yield solid lead(IV) oxide.
(b) Cobalt metal is heated with oxygen gas to give solid cobalt(III) oxide.

Answers:

(a) $\mathrm{Pb}(s) + \mathrm{O_2}(g) \xrightarrow{\Delta} \mathrm{PbO_2}(s)$ (b) $4\,\mathrm{Co}(s) + 3\,\mathrm{O_2}(g) \xrightarrow{\Delta} 2\,\mathrm{Co_2O_3}(s)$

Concept Exercise

What is the product from heating a metal and oxygen gas?

Answer: See Appendix G, 7.6.

Nonmetal and Oxygen Gas

In this kind of combination reaction, a nonmetal and oxygen gas react to give a nonmetal oxide. In the following example, yellow sulfur powder and oxygen gas combine to give sulfur dioxide gas:

$$\textbf{nonmetal} + \textbf{oxygen gas} \quad \rightarrow \quad \textbf{nonmetal oxide}$$

$$S(s) + O_2(g) \xrightarrow{\Delta} SO_2(g)$$

Nonmetal oxides demonstrate multiple combining capacities, for example, CO and CO_2. In general, the formula of a nonmetal oxide product is unpredictable and varies with temperature and pressure. For example, nitrogen and oxygen gases can combine to give all of the following: NO, NO_2, N_2O, N_2O_3, N_2O_4, and N_2O_5. Therefore, to complete and balance an equation of a nonmetal and oxygen gas, we must be given the name or formula of the nonmetal oxide.

The following example exercise further illustrates how to write balanced equations for a combination reaction of a nonmetal and oxygen gas:

▲ **Igniting Sulfur Powder**
A small portion of yellow sulfur burns with the oxygen in air to give an intense blue flame and sulfur dioxide gas, SO_2, which has a sharp odor.

EXAMPLE EXERCISE 7.7 Combination Reaction of a Nonmetal and Oxygen Gas

Write a balanced chemical equation for each of the following combination reactions:

(a) Carbon is heated with oxygen gas to produce carbon dioxide gas.
(b) Phosphorus and oxygen gas react to give solid diphosphorus pentaoxide.

Solution

A nonmetal and oxygen combine to produce a nonmetal oxide.

(a) The formula of the nonmetal oxide is unpredictable. We are given that the product is carbon dioxide and not carbon monoxide. The balanced equation for the reaction is

$$C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)$$

(b) The formula for the oxide of phosphorus is not predictable, but we have the name of the nonmetal oxide product. The formula for diphosphorus pentaoxide is P_2O_5. The balanced equation is

$$4\ P(s) + 5\ O_2(g) \rightarrow 2\ P_2O_5(s)$$

Practice Exercise

Write a balanced chemical equation for each of the following combination reactions:

(a) Nitrogen gas is heated with oxygen to give dinitrogen trioxide gas.
(b) Chlorine gas is heated with oxygen to give dichlorine monoxide gas.

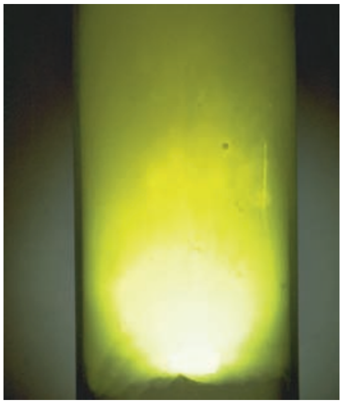
Answers:

(a) $2\ N_2(g) + 3\ O_2(g) \xrightarrow{\Delta} 2\ N_2O_3(g)$
(b) $2\ Cl_2(g) + O_2(g) \xrightarrow{\Delta} 2\ Cl_2O(g)$

Concept Exercise

What is the product from heating a nonmetal and oxygen gas?

Answer: See Appendix G, 7.7.



▲ **Igniting Sodium and Chlorine** Sodium metal reacts strongly with chlorine gas to give a bright yellow light and sodium chloride, NaCl, powder.

Metal and Nonmetal

In this kind of combination reaction, a metal and a nonmetal react to give an ionic compound. In the following example, silvery sodium metal and yellow chlorine gas combine to give solid sodium chloride:

$$\textbf{metal} + \textbf{nonmetal} \longrightarrow \textbf{ionic compound}$$

$$2\ Na(s) + Cl_2(g) \xrightarrow{\Delta} 2\ NaCl(s)$$

When a metal and a nonmetal react, they produce a binary ionic compound. If the compound contains a main-group metal, the formula is usually predictable. If the compound contains a transition metal, the formula is not predictable.

The following example exercise further illustrates how to write a balanced equation for a combination reaction of a metal and a nonmetal:

EXAMPLE EXERCISE 7.8 Combination Reaction of a Metal and a Nonmetal

Write a balanced chemical equation for each of the following combination reactions:

(a) Aluminum metal is heated with sulfur powder to yield solid aluminum sulfide.
(b) Chromium metal is heated with iodine crystals to give solid chromium(III) iodide.

Solution

A metal and a nonmetal react to produce a binary ionic compound.

(a) The formula of the product is predictable. Aluminum combines with sulfur to give aluminum sulfide, Al_2S_3. The balanced equation for the reaction is

$$2\ Al(s) + 3\ S(s) \xrightarrow{\Delta} Al_2S_3(s)$$

(b) Chromium is a transition metal, and so we cannot predict the formula for the product. We are given the name of the compound, chromium(III) iodide, and so the formula is CrI_3. The equation is

$$2\ Cr(s) + 3\ I_2(s) \xrightarrow{\Delta} 2\ CrI_3(s)$$

Practice Exercise

Write a balanced chemical equation for each of the following combination reactions:

(a) Calcium metal is heated with fluorine gas to yield solid calcium fluoride.
(b) Manganese metal reacts with bromine vapor to give crystalline manganese(II) bromide.

Answers:

(a) $Ca(s) + F_2(g) \xrightarrow{\Delta} CaF_2(s)$ (b) $Mn(s) + Br_2(g) \rightarrow MnBr_2(s)$

Concept Exercise

What is the product from heating a metal and nonmetal?

Answer: See Appendix G, 7.8.

7.6 Decomposition Reactions

LEARNING OBJECTIVES

- Propose a balanced equation for the decomposition of a metal hydrogen carbonate.
- Propose a balanced equation for the decomposition of a metal carbonate.
- Propose a balanced equation for the decomposition of a compound that releases oxygen gas.

There are many examples of decomposition reactions. For instance, heating a compound may cause it to decompose into two or more simpler substances. In this section, we will study three important kinds of decomposition reactions.

Metal Hydrogen Carbonates

A metal hydrogen carbonate undergoes a decomposition reaction when it is heated. For example, you may be aware that baking soda is a natural fire extinguisher. That is, heat from a fire decomposes baking soda (sodium hydrogen carbonate) and releases carbon dioxide gas. Because carbon dioxide is more dense than air, it can smother a fire by excluding oxygen from the flame.

The decomposition of a metal hydrogen carbonate yields a metal carbonate, water, and carbon dioxide. In the following example, sodium hydrogen carbonate is decomposed by heating to give sodium carbonate, water, and carbon dioxide:

$$\textbf{metal hydrogen carbonate} \xrightarrow{\Delta} \textbf{metal carbonate} + \textbf{water} + \textbf{carbon dioxide}$$

$$2\ NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

During the decomposition reaction of a metal hydrogen carbonate, the ionic charge of the metal does not change. Therefore, the formula of the metal carbonate is predictable. Even transition metal hydrogen carbonates are predictable. If nickel(II) hydrogen carbonate decomposes, one of the products is nickel(II) carbonate. That is,

$$Ni(HCO_3)_2(s) \xrightarrow{\Delta} NiCO_3(s) + H_2O(g) + CO_2(g)$$

The following example exercise further illustrates how to write balanced equations for a decomposition reaction of a metal hydrogen carbonate:

EXAMPLE EXERCISE 7.9 Decomposition Reaction of a Hydrogen Carbonate

Write a balanced chemical equation for each of the following decomposition reactions:

(a) Lithium hydrogen carbonate decomposes on heating.
(b) Lead(II) hydrogen carbonate decomposes on heating.

Solution

A metal hydrogen carbonate decomposes with heat to give a metal carbonate, water, and carbon dioxide gas.

(a) All the formulas are predictable, including that of the product, Li_2CO_3. The balanced equation for the reaction is

$$2\ LiHCO_3(s) \xrightarrow{\Delta} Li_2CO_3(s) + H_2O(g) + CO_2(g)$$

(b) Although the ionic charge of lead is variable, we are given that lead(II) hydrogen carbonate is the reactant. Therefore, the product is lead(II) carbonate, $PbCO_3$. The balanced equation for the reaction is

$$Pb(HCO_3)_2(s) \xrightarrow{\Delta} PbCO_3(s) + H_2O(g) + CO_2(g)$$

Practice Exercise

Write a balanced chemical equation for each of the following decomposition reactions:

(a) Barium hydrogen carbonate is decomposed with heat.
(b) Copper(I) hydrogen carbonate is decomposed with heat.

Answers:

(a) $Ba(HCO_3)_2(s) \xrightarrow{\Delta} BaCO_3(s) + H_2O(g) + CO_2(g)$
(b) $2\ CuHCO_3(s) \xrightarrow{\Delta} Cu_2CO_3(s) + H_2O(g) + CO_2(g)$

Concept Exercise

What is the product from heating a metal hydrogen carbonate?

Answer: See Appendix G, 7.9.

Metal Carbonates

After a metal hydrogen carbonate decomposes into a metal carbonate, the metal carbonate can further decompose with prolonged heating. The carbonate can decompose into a metal oxide while releasing carbon dioxide gas.

In the following example, calcium carbonate is decomposed by heating to give calcium oxide and carbon dioxide:

$$\textbf{metal carbonate} \xrightarrow{\Delta} \textbf{metal oxide} + \textbf{carbon dioxide}$$

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

The ionic charge of the metal does not change during the decomposition reaction of a metal carbonate. Therefore, we can predict the formula for the metal oxide product from the decomposition of a metal carbonate. If nickel(II) carbonate is decomposed, one of the products is nickel(II) oxide. That is,

$$NiCO_3 \xrightarrow{\Delta} NiO(s) + CO_2(g)$$

The following example exercise further illustrates how to write balanced equations for a decomposition reaction of a metal carbonate:

EXAMPLE EXERCISE 7.10 Decomposition Reaction of a Carbonate

Write a balanced chemical equation for each of the following decomposition reactions:

(a) Magnesium carbonate decomposes on heating.
(b) Copper(I) carbonate decomposes on heating.

Solution

A metal carbonate decomposes with heat to a metal oxide and carbon dioxide gas.

(a) All the formulas are predictable including that of the metal oxide, MgO. The balanced equation for the reaction is

$$MgCO_3(s) \xrightarrow{\Delta} MgO(s) + CO_2(g)$$

(b) The ionic charge of copper can be either 1+ or 2+. Because the reactant is copper(I) carbonate, the product is copper(I) oxide, that is, Cu_2O. Thus, the balanced equation for the reaction is

$$Cu_2CO_3(s) \xrightarrow{\Delta} Cu_2O(s) + CO_2(g)$$

Practice Exercise

Write a balanced chemical equation for each of the following decomposition reactions:

(a) Aluminum carbonate is decomposed on heating.
(b) Iron(II) carbonate is decomposed on heating.

Answers:

(a) $Al_2(CO_3)_3(s) \xrightarrow{\Delta} Al_2O_3(s) + 3\,CO_2(g)$
(b) $FeCO_3(s) \xrightarrow{\Delta} FeO(s) + CO_2(g)$

Concept Exercise

What is the product from heating a metal carbonate?

Answer: See Appendix G, 7.10.

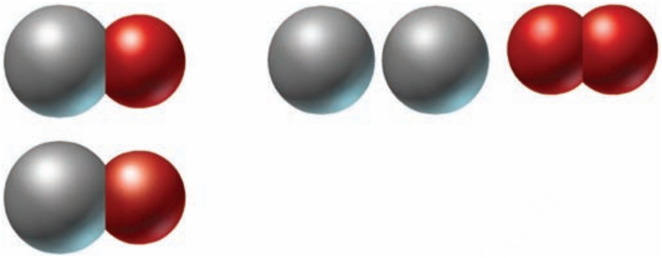
Miscellaneous Oxygen-Containing Compounds

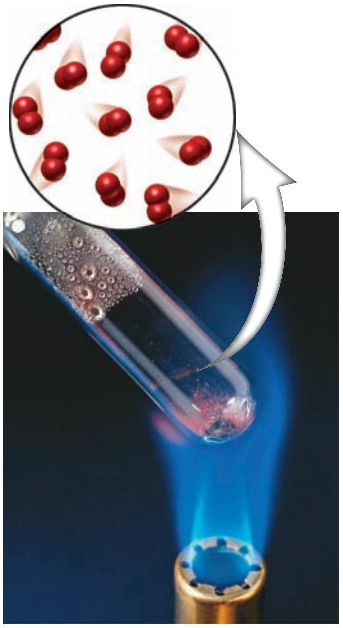
We can often decompose an oxygen-containing ionic compound, which releases oxygen gas. However, we cannot always predict the products. In this kind of decomposition reaction, we must be given the formulas of the products.

In the following example, orange mercury(II) oxide powder is decomposed by heating to give droplets of shiny liquid mercury and oxygen gas:

$$\textbf{oxygen-containing compound} \overset{\Delta}{\rightarrow} \textbf{substance} + \textbf{oxygen gas}$$

$$2\ HgO(s) \xrightarrow{\Delta} 2\ Hg(l) + O_2(g)$$





▲ **Decomposing Mercury(II) Oxide** A small portion of orange HgO powder decomposes with heat to give oxygen gas, O_2, and beads of silvery metallic mercury.

The following example exercise illustrates how to write balanced equations for miscellaneous decomposition reactions:

EXAMPLE EXERCISE 7.11 Decomposition Reaction Producing Oxygen Gas

Write a balanced chemical equation for each of the following decomposition reactions:

(a) Solid potassium chlorate decomposes upon heating to give solid potassium chloride and oxygen gas.

(b) Solid sodium nitrate decomposes upon heating to give solid sodium nitrite and oxygen gas.

Solution

There is no general format for these reactions; however, we are given the names of the reactants and products.

(a) We can write the reaction for the decomposition of potassium chlorate and balance the chemical equation as follows:

$$2\ KClO_3(s) \overset{\Delta}{\rightarrow} 2\ KCl(s) + 3\ O_2(g)$$

(b) We can write the reaction for the decomposition of sodium nitrate and balance the chemical equation as follows:

$$2\ NaNO_3(s) \overset{\Delta}{\rightarrow} 2\ NaNO_2(s) + O_2(g)$$

Practice Exercise

Write a balanced chemical equation for each of the following decomposition reactions:

(a) Solid lead(IV) oxide decomposes upon heating to give solid lead(II) oxide and oxygen gas.

(b) Aqueous hydrogen peroxide, H_2O_2, decomposes upon heating to give water and oxygen gas.

Answers:

(a) $2\ PbO_2(s) \overset{\Delta}{\rightarrow} 2\ PbO(s) + O_2(g)$ (b) $2\ H_2O_2(aq) \overset{\Delta}{\rightarrow} 2\ H_2O(l) + O_2(g)$

Concept Exercise

What is the usual product from heating an oxygen-containing compound?

Answer: See Appendix G, 7.11.

7.7 The Activity Series Concept

LEARNING OBJECTIVES

- Predict whether a metal reacts with an aqueous salt solution.
- Predict whether a metal reacts with an aqueous acid.
- Predict whether a metal reacts with water at 25 °C.

Helpful Hint
Activity Series

The relative order of metals arranged by their ability to undergo reaction is as follows: Li > K > Ba > Sr > Ca > Na > Mg > Al > Mn > Zn > Fe > Cd > Co > Ni > Sn > Pb > (H) > Cu > Ag > Hg > Au. If zinc metal is placed in an aqueous iron solution, Zn will displace Fe from the solution.

When a metal undergoes a single replacement reaction, it displaces another metal from a compound or aqueous solution. The metal that displaces the other metal does so because it has a greater tendency to undergo a reaction; in other words, it is more active. If metal A is more active than metal B, metal A will displace metal B from compound BZ. We can show the reaction as

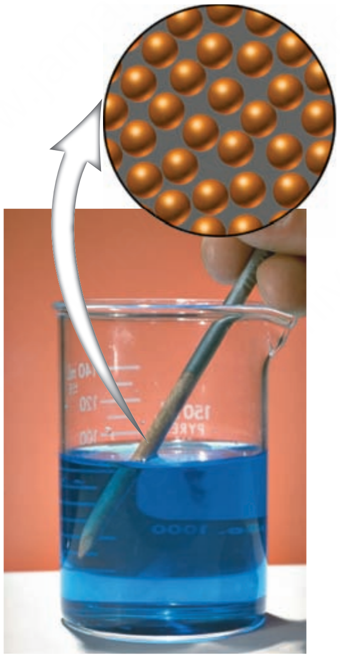
$$\text{A} + \text{BZ} \longrightarrow \text{AZ} + \text{B}$$

The activity of a metal is a measure of its ability to compete in a single replacement reaction. In an **activity series** (or electromotive series), a sequence of metals is arranged according to their ability to undergo reaction. Metals that are most reactive appear first in the activity series. Metals that are least reactive appear last in the series. The relative activity of several metals is listed in the Helpful Hint.

We can see that lithium precedes potassium in the series, which in turn precedes barium, and so on. Although hydrogen (H) is not a metal, it is included in the series as a reference. Metals that precede (H) in the series react with an aqueous acid; metals that follow (H) in the series do not react with acids. Notice that the metals Cu, Ag, Hg, and Au do not react with acids.

Consider the reaction of iron metal in an aqueous solution of copper(II) sulfate:

$$\text{Fe}(s) + \text{CuSO}_4(aq) \longrightarrow \text{FeSO}_4(aq) + \text{Cu}(s)$$



Iron in a Copper Solution An iron nail reacts in an aqueous blue $CuSO_4$ solution to give a reddish-brown deposit of Cu metal on the nail.

Iron precedes copper in the series, and so Fe displaces Cu from the aqueous solution. If we perform the reaction in the laboratory, a reddish-brown copper deposit appears on the iron metal. It is this type of experiment that establishes the order of metals in the activity series; that is, Fe > > > Cu.

Conversely, there is no reaction if we place copper metal in an aqueous solution of iron(II) sulfate. That is, Cu cannot displace Fe from the solution because copper follows iron in the activity series. Thus, if we put copper wire into a solution of iron(II) sulfate, we observe

$$\text{Cu}(s) + \text{FeSO}_4(aq) \longrightarrow \text{NR}$$

Now let's consider the reaction of iron wire with sulfuric acid. The reaction releases gas bubbles in the solution, and the equation is

$$\text{Fe}(s) + \text{H}_2\text{SO}_4(aq) \longrightarrow \text{FeSO}_4(aq) + \text{H}_2(g)$$

Iron precedes (H) in the activity series, and so Fe displaces H_2 gas from a sulfuric acid solution. When we observe the reaction, we see tiny bubbles of hydrogen gas on the iron wire. Now consider the reaction of copper metal and sulfuric acid:

$$\mathrm{Cu}(s) + \mathrm{H_2SO_4}(aq) \rightarrow \mathrm{NR}$$

Copper is after (H) in the activity series, so Cu cannot displace H_2 gas from sulfuric acid. When we observe the copper wire in sulfuric acid, there is no evidence for reaction.

Active Metals

There are a few metals that are so reactive they react directly with water at room temperature. These metals are called the **active metals** and include most of the metals in Groups IA/1 and IIA/2. Specifically, Li, Na, K, Rb, Cs, Ca, Sr, and Ba react with water. The most common active metals are listed in the Helpful Hint.

Consider an experiment in which sodium metal is dropped into water. Sodium is an active metal and reacts with water according to the following equation:

$$2\,\mathrm{Na}(s) + 2\,\mathrm{H_2O}(l) \rightarrow 2\,\mathrm{NaOH}(aq) + \mathrm{H_2}(g)$$

The following example exercise further illustrates how to predict the reactivity of metals with aqueous solutions, acids, and water:

Helpful Hint
Active Metals
The relative order of metals arranged by their ability to undergo reaction in water is as follows: Li > K > Ba > Sr > Ca > Na. If lithium metal is dropped into water, Li will displace H_2 gas from the solution.

EXAMPLE EXERCISE 7.12 Predictions Based on the Activity Series

Predict whether or not a reaction occurs for each of the following:

(a) Aluminum foil is added to an iron(II) sulfate solution.
(b) Iron wire is added to an aluminum sulfate solution.
(c) Manganese metal is added to acetic acid.
(d) Magnesium metal is added to water.

Solution

We refer to the activity series for the reactivity of each of the metals.

(a) Aluminum precedes iron in the activity series: Al > Fe. Thus, a reaction occurs, and Al displaces Fe from the solution.
(b) Conversely, iron follows aluminum in the activity series: Al > Fe. Thus, there is *no reaction.*
(c) Manganese precedes hydrogen in the activity series: Mn > (H). Thus, a reaction occurs, and Mn produces H_2 gas that bubbles from the solution.
(d) Magnesium is not an active metal. Therefore, magnesium does not react with water, and there is *no reaction.*

Practice Exercise

Predict whether or not a reaction occurs for each of the following:

(a) A gold ring is dropped into sulfuric acid.
(b) A zinc granule is dropped into hydrochloric acid.
(c) A cadmium foil is put into a lead(II) nitrate solution.
(d) A chromium strip is put into water.

Answers:

(a) There is *no reaction* because Au follows (H) in the activity series.
(b) There is a reaction because Zn precedes (H) in the activity series.
(c) There is a reaction because Cd precedes Pb in the activity series.
(d) There is *no reaction* because Cr is not an active metal.

Concept Exercise

Which of the following metals reacts with aqueous silver nitrate: Zn, Cu, Au?

Answer: See Appendix G, 7.12.

▲ **Gold in Acid** A gold ring gives no reaction in an aqueous sulfuric acid, H_2SO_4, solution.

Note In this introduction to the concept of an activity series, we have considered only the activity of metals. In addition, the ability of nonmetals to undergo a reaction also gives rise to an activity series. For example, the halogens demonstrate the following activity series: F > Cl > Br > I. Consider the reaction

$$Cl_2(g) + 2\,NaBr(aq) \quad \rightarrow \quad 2\,NaCl(aq) + Br_2(l)$$

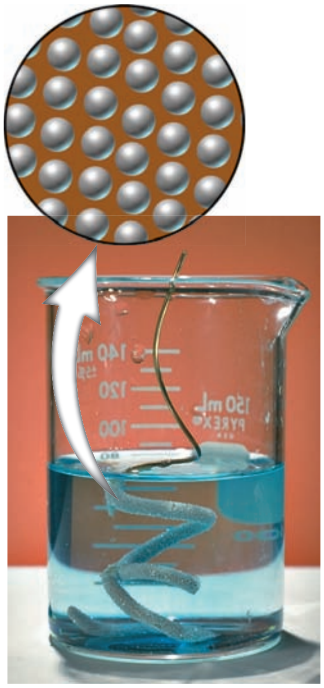
A reaction occurs because Cl precedes Br in the activity series; thus, the more active Cl displaces the less active Br from the solution. Conversely, Cl follows F in the series. Therefore, Cl is less reactive than F and cannot displace it from the solution. Thus,

$$Cl_2(g) + NaF(aq) \quad \rightarrow \quad NR$$

7.8 Single-Replacement Reactions

LEARNING OBJECTIVES

- Propose a balanced chemical equation for the reaction of a metal in an aqueous solution.
- Propose a balanced chemical equation for the reaction of a metal in an aqueous acid.
- Propose a balanced chemical equation for the reaction of an active metal in water.



▲ **Copper in a Silver Solution** A copper wire reacts in a colorless $AgNO_3$ solution to give white crystals of Ag and a light-blue aqueous solution.

There are many examples of single-replacement reactions in which a more active metal displaces a less active metal from a solution. In general, the activity of a metal follows the metallic character trend in the periodic table. To determine the relative reactivity of a specific metal, we refer to the activity series.

Metal and Aqueous Salt Solution

One kind of single-replacement reaction involves a metal reacting with an ionic compound (salt) in aqueous solution. The general equation for the reaction is

$$\mathbf{metal_1}(s) + \mathbf{salt_1}(aq) \quad \rightarrow \quad \mathbf{metal_2}(s) + \mathbf{salt_2}(aq)$$

Consider the reaction of metallic copper wire in an aqueous solution of silver nitrate. The balanced equation for the reaction is

$$Cu(s) + 2\,AgNO_3(aq) \quad \rightarrow \quad 2\,Ag(s) + Cu(NO_3)_2(aq)$$

If we refer to the activity series, we note that copper is the more active metal. That is, copper lies above silver in the activity series. Therefore, Cu metal can displace Ag metal from an aqueous solution.

Now let's consider the reaction of metallic silver wire in an aqueous solution of copper nitrate. Because Ag metal is less active than Cu metal, we can predict that there is no reaction. That is,

$$Ag(s) + Cu(NO_3)_2(aq) \rightarrow NR$$

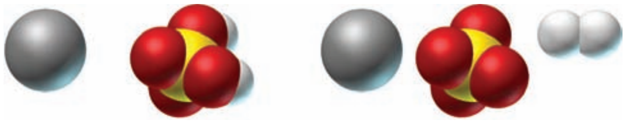
Metal and Aqueous Acid Solution

Another kind of single-replacement reaction is that of a metal with an aqueous acid. A metal and an aqueous acid react to produce hydrogen gas and an aqueous solution of an ionic compound (salt). The general equation for the reaction is

$$\mathbf{metal}(s) + \mathbf{acid}(aq) \quad \rightarrow \quad \mathbf{salt}(aq) + \mathbf{hydrogen}(g)$$

Consider the reaction of iron metal and sulfuric acid. First, refer to the activity series to check the position of iron. Iron is above hydrogen in the series. Therefore, iron displaces hydrogen gas from the acid. If we assume that the reaction produces iron(II) sulfate, we can write the balanced chemical equation.

$$Fe(s) \; + \; H_2SO_4(aq) \; \longrightarrow \; FeSO_4(aq) \; + \; H_2(g)$$



Active Metal and Water

In this type of single-replacement reaction, an active metal and water react to give a metal hydroxide and hydrogen gas. The general equation for the reaction is

$$\textbf{metal}(s) + \textbf{water}(l) \rightarrow \textbf{metal hydroxide}(aq) + \textbf{hydrogen}(g)$$

Consider the reaction of calcium metal in water to give calcium hydroxide and hydrogen gas. The balanced equation for the reaction is

$$Ca(s) + 2\,H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$$

The following example exercise further illustrates how to write balanced chemical equations for the three types of single-replacement reactions:

▲ **Calcium in Water** A piece of calcium metal in water reacts to give bubbles of hydrogen gas.

EXAMPLE EXERCISE 7.13 Single-Replacement Reactions

Write a balanced chemical equation for each of the following single-replacement reactions:

(a) Nickel metal is placed in a tin(II) sulfate solution.
(b) Cobalt metal is put in a cadmium nitrate solution.
(c) Manganese chips are added to sulfuric acid.
(d) A chunk of potassium is dropped into water.

Solution

First, we will refer to the activity series for the relative positions of each of the metals. Then we will write the equations.

(a) Nickel is above tin in the activity series: Ni > Sn. Therefore, a reaction occurs and Ni displaces Sn from the solution.

$$Ni(s) + SnSO_4(aq) \rightarrow NiSO_4(aq) + Sn(s)$$

(b) Cobalt is below cadmium in the series: Cd > Co. Therefore, there is no reaction.

$$Co(s) + Cd(NO_3)_2(aq) \rightarrow NR$$

(c) Manganese is above hydrogen in the series: Mn > (H). Therefore, a reaction occurs, and Mn releases H_2 gas bubbles from the solution.

$$Mn(s) + H_2SO_4(aq) \rightarrow MnSO_4(aq) + H_2(g)$$

(d) Potassium is a Group IA/1 active metal. Therefore, a reaction occurs, and K releases H_2 gas from the solution.

$$2\,K(s) + 2\,H_2O(l) \rightarrow 2\,KOH(aq) + H_2(g)$$

Practice Exercise

Write a balanced chemical equation for each of the following single-replacement reactions:

(a) Nickel metal is placed in a silver nitrate solution.
(b) Gold metal is placed in a silver nitrate solution.
(c) A chunk of cadmium metal is dropped into hydrochloric acid.
(d) A small piece of strontium metal is dropped into water.

Answers:

(a) $Ni(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Ni(NO_3)_2(aq)$
(b) $Au(s) + AgNO_3(aq) \rightarrow NR$
(c) $Cd(s) + 2\,HCl(aq) \rightarrow CdCl_2(aq) + H_2(g)$
(d) $Sr(s) + 2\,H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$

Concept Exercise

Which of the following metals reacts with an aqueous acid: Zn, Cu, Au?

Answer: See Appendix G, 7.13.

Note In a single-replacement reaction, it may not be possible to predict the ionic charge of the resulting metal cation. For example, when copper metal reacts with silver nitrate solution, the resulting compound may contain either Cu^+ or Cu^{2+}.

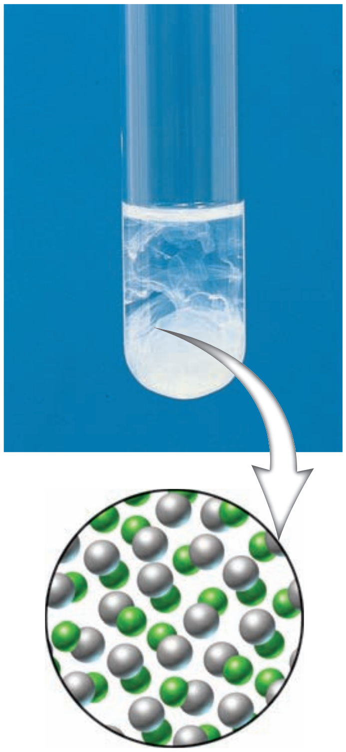
7.9 Solubility Rules

LEARNING OBJECTIVE

- Predict whether an ionic compound dissolves in water given the general rules for solubility.

In a double-replacement reaction, which we will discuss in Section 7.10, two aqueous solutions may react to form a compound that is not soluble, that is, a precipitate. When we observe that a precipitate is produced in a solution, we have evidence for a chemical reaction.

It is important to be able to identify an insoluble compound in an aqueous solution. Table 7.2 presents general rules for predicting the solubility of ionic compounds in water. However, there are exceptions to the general rules, and a few of these exceptions are indicated; for example, Na_2CO_3 is soluble in water (Rule 6).



▲ Insoluble Silver Chloride Aqueous solutions of silver nitrate and sodium chloride react to give white, insoluble AgCl.

TABLE 7.2 Solubility Rules for Ionic Compounds

Ionic compounds containing the following ions are generally *soluble* in water:

1. alkali metal ions and the ammonium ion, Li^+, Na^+, K^+, NH_4^+
2. acetate ion, $C_2H_3O_2^-$
3. nitrate ion, NO_3^-
4. halide ions ($X = Cl^-, Br^-, I^-$) (AgX, Hg_2X_2, and PbX_2 are exceptions and *insoluble*)
5. sulfate ion, SO_4^{2-} ($SrSO_4$, $BaSO_4$, and $PbSO_4$ are exceptions and *insoluble*)

Ionic compounds containing the following ions are generally *insoluble* in water:*

6. carbonate ion, CO_3^{2-} (see Rule 1 exceptions, which are *soluble*)
7. chromate ion, CrO_4^{2-} (see Rule 1 exceptions, which are *soluble*)
8. phosphate ion, PO_4^{3-} (see Rule 1 exceptions, which are *soluble*)
9. sulfide ion, S^{2-} (CaS, SrS, BaS, and Rule 1 exceptions are *soluble*)
10. hydroxide ion, OH^- [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, and Rule 1 exceptions are *soluble*]

*These compounds are actually slightly soluble, or very slightly soluble, in water.

Now let's apply the solubility rules to selected compounds and determine whether or not the compound is soluble in water.

EXAMPLE EXERCISE 7.14 Soluble Ionic Compounds

State whether each of the following compounds is soluble or insoluble in water.

(a) ammonium sulfide, $(NH_4)_2S$
(b) barium sulfate, $BaSO_4$

Solution

Refer to the solubility rules for ionic compounds in Table 7.2.

(a) Ammonium sulfide contains the ammonium ion, NH_4^+. According to Rule 1, $(NH_4)_2S$ is *soluble*.
(b) Barium sulfate contains the sulfate ion, SO_4^{2-}. According to the Rule 5 exception, $BaSO_4$ is *insoluble*.

Practice Exercise

State whether each of the following compounds is soluble or insoluble in water.

(a) magnesium carbonate, $MgCO_3$
(b) calcium hydroxide, $Ca(OH)_2$

Answers:

(a) insoluble (Rule 6) (b) soluble (Rule 10 exception)

Concept Exercise

Which of the following compounds is soluble in water: $HgCl_2$, or Hg_2Cl_2?

Answer: See Appendix G, 7.14.

EXAMPLE EXERCISE 7.15 Insoluble Ionic Compounds

State whether each of the following compounds is soluble or insoluble in water.

(a) aluminum nitrate, $Al(NO_3)_3$ (b) lead(II) chromate, $PbCrO_4$

Solution

Refer to the solubility rules for ionic compounds in Table 7.2.

(a) Aluminum nitrate contains the nitrate ion, NO_3^-. According to Rule 3, $Al(NO_3)_3$ is *soluble*.

(b) Lead(II) chromate contains the chromate ion, CrO_4^{2-}. According to Rule 7, $PbCrO_4$ is *insoluble*.

Practice Exercise

State whether each of the following compounds is soluble or insoluble in water.

(a) lead(II) acetate or $Pb(C_2H_3O_2)_2$ (b) zinc phosphate or $Zn_3(PO_4)_2$

Answers:

(a) soluble (Rule 2) (b) insoluble (Rule 8)

Concept Exercise

Which of the following compounds is insoluble in water: $HgBr_2$ or Hg_2Br_2?

Answer: See Appendix G, 7.15.

7.10 Double-Replacement Reactions

LEARNING OBJECTIVE

- Propose a balanced chemical equation for the reaction of two salts in aqueous solution.

In a double-replacement precipitation reaction, two ionic compounds in an aqueous solution switch anions and produce an insoluble substance. The general form of the reaction is

$$AX + BZ \longrightarrow AZ + BX$$

In the reaction of AX and BZ, we can assume that either AZ or BX is an insoluble precipitate. If both AZ and BX are soluble, there is no precipitate; and hence, there is no chemical reaction. Consider the double-replacement reaction of two salts in aqueous solution:

$$\textbf{salt}_1(aq) + \textbf{salt}_2(aq) \longrightarrow \textbf{salt}_3(s) + \textbf{salt}_4(aq)$$

As an example, when we add an aqueous solution of silver nitrate to an aqueous solution of sodium carbonate, we observe a white precipitate. In this reaction, the two compounds switch anions to give insoluble silver carbonate and sodium nitrate. We can write the balanced chemical equation as follows:

$$2\,AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3(s) + 2\,NaNO_3(aq)$$

By referring to solubility Rule 6, we find that most compounds containing the carbonate ion, CO_3^{2-}, are insoluble. Because silver carbonate is insoluble, it is written as $Ag_2CO_3(s)$ in the above equation. The (*s*) indicates that Ag_2CO_3 is an insoluble precipitate. The following example exercise further illustrates how to write balanced equations for double-replacement reactions:

▲ **Insoluble Silver Carbonate** Aqueous solutions of silver nitrate and sodium carbonate react to give an insoluble precipitate of Ag_2CO_3.

EXAMPLE EXERCISE 7.16 **Double-Replacement Reactions**

Write a balanced chemical equation for each of the following double-replacement reactions.

(a) Aqueous barium chloride is added to a potassium chromate solution.
(b) Aqueous strontium acetate is added to a lithium hydroxide solution.

Solution

For double-replacement reactions, we switch anions for the two compounds and check the solubility rules for an insoluble compound.

(a) Barium chloride and potassium chromate give barium chromate and potassium chloride. According to solubility Rule 7, we find that barium chromate is insoluble. The balanced equation is

$$\mathrm{BaCl_2}(aq) + \mathrm{K_2CrO_4}(aq) \rightarrow \mathrm{BaCrO_4}(s) + 2\,\mathrm{KCl}(aq)$$

(b) Strontium acetate and lithium hydroxide react as follows:

$$\mathrm{Sr(C_2H_3O_2)_2}(aq) + 2\,\mathrm{LiOH}(aq) \rightarrow \mathrm{Sr(OH)_2}(aq) + 2\,\mathrm{LiC_2H_3O_2}(aq)$$

However, the solubility rules indicate that $Sr(OH)_2$ and $LiC_2H_3O_2$ are soluble. Therefore, the equation is written

$$\mathrm{Sr(C_2H_3O_2)_2}(aq) + \mathrm{LiOH}(aq) \rightarrow \mathrm{NR}$$

Practice Exercise

Write a balanced chemical equation for each of the following double-replacement reactions:

(a) Aqueous zinc sulfate is added to a sodium carbonate solution.
(b) Aqueous manganese(II) nitrate is added to a potassium hydroxide solution.

Answers:

(a) $\mathrm{ZnSO_4}(aq) + \mathrm{Na_2CO_3}(aq) \rightarrow \mathrm{ZnCO_3}(s) + \mathrm{Na_2SO_4}(aq)$
(b) $\mathrm{Mn(NO_3)_2}(aq) + 2\,\mathrm{KOH}(aq) \rightarrow \mathrm{Mn(OH)_2}(s) + 2\,\mathrm{KNO_3}(aq)$

Concept Exercise

Identify the insoluble product from the reaction of silver nitrate and sodium chloride.

Answer: See Appendix G, 7.16.

LEARNING OBJECTIVE

- Propose a balanced chemical equation for the reaction of an acid and a base.

7.11 Neutralization Reactions

In a neutralization reaction, an acid and a base react to produce an aqueous ionic compound and water. An **acid** is a substance that releases hydrogen ions, H^+, in water; a **base** is a substance that releases hydroxide ions, OH^-, in water. The resulting ionic compound is called a **salt** and is composed of the cation from the base and the anion from the acid. The general form of the reaction is

$$\mathbf{H}\mathrm{X} + \mathrm{BO}\mathbf{H} \rightarrow \mathrm{BX} + \mathbf{HOH}$$

A neutralization reaction is a special case of a double-replacement reaction. Notice that we have written water as HOH to emphasize the switching of ions. You will also discover that writing water as HOH is helpful in balancing equations for neutralization reactions. In a neutralization reaction, an acid and a base react to produce a salt and water:

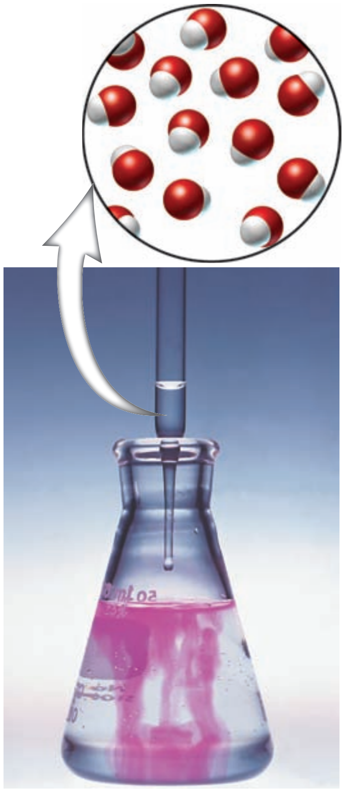
$$\textbf{acid}(aq) + \textbf{base}(aq) \rightarrow \textbf{salt}(aq) + \textbf{water}(l)$$

As an example, when we add hydrochloric acid to sodium hydroxide solution, we obtain sodium chloride and water. We can write the balanced chemical equation as follows:

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \rightarrow \mathrm{NaCl}(aq) + \mathrm{HOH}(l)$$

Other than a slight warming, there is no evidence of reaction. To provide evidence for this neutralization reaction, we can add a drop of the acid–base indicator phenolphthalein to the HCl. Phenolphthalein is colorless in acid and pink in base. Initially, the HCl acid solution is colorless; however, as we add NaOH base, we see flashes of color and eventually the solution turns pink. With an acid–base indicator we can observe a color change, which is evidence for a chemical reaction.

The following example exercise provides additional practice in writing balanced equations for neutralization reactions:



▲ **Phenolphthalein Acid–Base Indicator** In an acid–base neutralization reaction, the phenolphthalein indicator signals a change. If we have acid in a flask and add base, the indicator changes from colorless to pink.

EXAMPLE EXERCISE 7.17 Neutralization Reactions

Write a balanced chemical equation for each of the following neutralization reactions:

(a) Nitric acid neutralizes an ammonium hydroxide solution.
(b) Sulfuric acid neutralizes a potassium hydroxide solution.

Solution

A neutralization reaction produces a salt and water.

(a) Nitric acid and ammonium hydroxide produce ammonium nitrate and water. The balanced equation is

$$HNO_3(aq) + NH_4OH(aq) \rightarrow NH_4NO_3(aq) + HOH(l)$$

(b) Sulfuric acid and potassium hydroxide produce potassium sulfate and water. The balanced equation is

$$H_2SO_4(aq) + 2\,KOH(aq) \rightarrow K_2SO_4(aq) + 2\,HOH(l)$$

Practice Exercise

Write a balanced chemical equation for each of the following neutralization reactions:

(a) Chloric acid neutralizes a strontium hydroxide solution.
(b) Phosphoric acid neutralizes a sodium hydroxide solution.

Answers:

(a) $2\,HClO_3(aq) + Sr(OH)_2(aq) \rightarrow Sr(ClO_3)_2(aq) + 2\,HOH(l)$
(b) $H_3PO_4(aq) + 3\,NaOH(aq) \rightarrow Na_3PO_4(aq) + 3\,HOH(l)$

Concept Exercise

Identify the salt product from the reaction of hydrochloric acid and sodium hydroxide.

Answer: See Appendix G, 7.17.

A CLOSER LOOK Household Chemicals

Q: Is it possible that mixing two common household chemicals can produce a deadly gas?

Salt and sugar are two familiar chemicals on your kitchen table. Vinegar, another kitchen chemical, is a solution of acetic acid. Citrus fruit contains citric acid and is responsible for the sour taste of lemons and limes. Aspirin contains acetylsalicylic acid, which can irritate the lining of the stomach if taken in excess.

Hydrochloric acid, used to acidify swimming pools and clean concrete, is sold in the supermarket as muriatic acid. Sulfuric acid, found in automobile batteries, is a dangerous chemical that should be handled with great caution.

Perhaps the most dangerous chemical in the home is caustic soda, NaOH, which is sold as drain cleaner and as oven cleaner. If it contacts your skin, it gives a slippery sensation and causes redness and irritation. If taken internally, caustic soda could be fatal, especially to an infant. Household ammonia is also potentially dangerous and should not be used with bleach. Together, these two chemicals produce a poisonous gas.

▲ **Common Household Chemicals**

Common Household Chemicals

NAME	FORMULA	PRODUCT/USE	SAFETY
Acids			
acetic acid	$HC_2H_3O_2(aq)$	vinegar	
carbonic acid	$H_2CO_3(aq)$	carbonated drinks	
hydrochloric acid	$HCl(aq)$	swimming pools	avoid contact[*]
sulfuric acid	$H_2SO_4(aq)$	battery acid	avoid contact[*†]
Bases			
ammonia	$NH_3(aq)$	window cleaner	
sodium hydrogen carbonate	$NaHCO_3(s)$	baking soda, antacid	
magnesium hydroxide	$Mg(OH)_2(aq)$	milk of magnesia	
sodium hydroxide	$NaOH(aq)$	drain and oven cleaner	avoid contact[*†]
Miscellaneous			
aluminum hydroxide	$Al(OH)_3(s)$	antacid tablets	
dry ice	$CO_2(s)$	keep items cold	avoid contact
Epsom salts	$MgSO_4 \cdot 7H_2O(s)$	salt bath, laxative	
sodium hypochlorite	$NaClO(aq)$	bleach	avoid contact[*]
Organic			
ethylene glycol	$C_2H_4(OH)_2(l)$	antifreeze	avoid ingestion[†]
isopropyl alcohol	$CH_3CHOHCH_3(l)$	rubbing alcohol	avoid ingestion[†]
naphthalene	$C_{10}H_8(s)$	moth repellent	avoid ingestion[†]
trichloroethane	$C_2H_3Cl_3(l)$	spot remover	avoid ingestion[†]

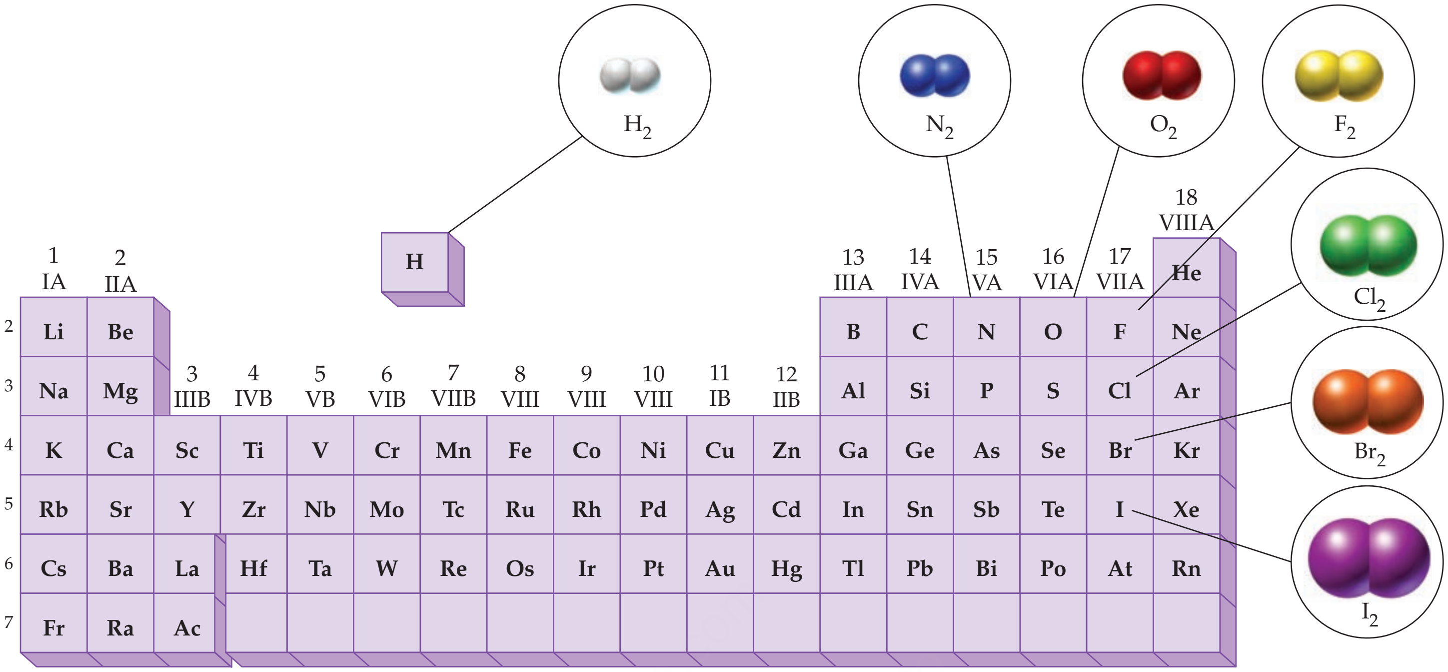
[*] In the event of contact, flush with water.

[†] In the event of ingestion, seek medical attention immediately.

A: Caution must be exercised when using any chemicals; for example, household ammonia and bleach, when mixed, release the poisonous gas chloramine, NH_2Cl.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
7.1 Evidence of Chemical Reactions In this section we learned that the evidence of a **chemical reaction** is any of the following: a gas is detected, a **precipitate** is formed in an **aqueous solution**, or a color change or energy change is observed. An **exothermic reaction** releases heat energy, and an **endothermic reaction** absorbs heat energy.	• Describe four observations that are evidence of a chemical reaction. ***Related Exercises 1–6***
7.2 Writing Chemical Equations A **reactant** substance undergoes a chemical change to become a **product**. A **catalyst** is a substance that speeds up a reaction without being permanently altered. A **chemical equation** describes a reaction using formulas and symbols. Seven nonmetal elements occur naturally as **diatomic molecules**, and their formulas are written as H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.	• Identify seven elements that occur naturally as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2. ***Related Exercises 7–8*** • Write a chemical equation given the description of a chemical reaction. ***Related Exercises 7–16***



▲ **Diatomic Molecules** Seven elements occur naturally as diatomic molecules: H_2, N_2, O_2, F_2,and Cl_2 are gases; bromine, Br_2,occurs as a liquid; and iodine, I_2,as solid crystals.

Key Concepts

Learning Objectives and Related Exercises

7.3 Balancing Chemical Equations

To balance a chemical equation, **coefficients** are systematically placed in front of each substance until there is the same number of atoms or ions of each element on both sides of the equation. If the reaction coefficient is 1, it is omitted. Similarly, if a formula **subscript** is 1, such as oxygen in H_2O, it is omitted.

- Convert a chemical reaction into a balanced chemical equation.
  Related Exercises 17–20

7.4 Classifying Chemical Reactions

Although there are thousands of chemical reactions, most can be classified as one of the following: combination, decomposition, single replacement, double replacement, or neutralization. Table 7.3 lists these five basic types.

TABLE 7.3 Types of Chemical Reactions

REACTION TYPE	GENERAL FORMAT
Combination reaction	$A + Z \rightarrow AZ$
Decomposition reaction	$AZ \rightarrow A + Z$
Single-replacement reaction	$A + BZ \rightarrow AZ + B$
Double-replacement reaction	$AX + BZ \rightarrow AZ + BX$
Neutralization reaction	$HX + BOH \rightarrow BX + HOH$

- Classify a chemical reaction as one of the following types: combination, decomposition, single-replacement, double-replacement, or neutralization reaction.
  Related Exercises 21–24

7.5 Combination Reactions

We have examined three kinds of combination reactions: a metal and oxygen, a nonmetal and oxygen, and a metal and a nonmetal. Often, we can predict the products from the reaction. Nonmetals, however, combine with oxygen gas to give a variety of oxides, and the product is not predictable. In summary,

1. **metal** + **oxygen**(g) → **metal oxide**
2. **nonmetal** + **oxygen**(g) → **nonmetal oxide**
3. **metal** + **nonmetal** → **ionic compound**

- Propose a balanced equation for the combination reaction of a metal and oxygen gas.
  Related Exercises 25–26
- Propose a balanced equation for the combination reaction of a nonmetal and oxygen gas.
  Related Exercises 27–28
- Propose a balanced equation for the combination reaction of a metal and a nonmetal.
  Related Exercises 29–36

Key Concepts	Learning Objectives and Related Exercises
7.6 Decomposition Reactions We have examined three kinds of decomposition reactions. In the first type, a metal hydrogen carbonate is decomposed by heating to give a metal carbonate product. In the second, a metal carbonate is decomposed to a metal oxide. Most oxygen-containing polyatomic ions, such as chlorate and nitrate, decompose to give oxygen gas. In summary, 1. $\textbf{metal hydrogen carbonate} \xrightarrow{\Delta} \textbf{metal carbonate} + \textbf{water} + \textbf{carbon dioxide}(g)$ 2. $\textbf{metal carbonate} \xrightarrow{\Delta} \textbf{metal oxide} + \textbf{carbon dioxide}(g)$ 3. $\textbf{oxygen-containing compound} \xrightarrow{\Delta} \textbf{substance} + \textbf{oxygen}(g)$	• Propose a balanced equation for the decomposition reaction of a metal hydrogen carbonate. ***Related Exercises 37, 38, and 41*** • Propose a balanced equation for the decomposition reaction of a metal carbonate. ***Related Exercises 39, 40, and 42*** • Propose a balanced equation for the decomposition reaction of a compound that releases oxygen gas. ***Related Exercises 43–46***
7.7 The Activity Series Concept The **activity series** lists the relative ability of a metal to undergo reaction. A metal higher in the series always displaces a metal lower in the series. Metals that are more active than (H) displace hydrogen gas from an acid solution. **Active metals** include Li, Na, K, Ba, Sr, and Ca; each of these metals displaces hydrogen gas from water.	• Predict if a metal reacts in a given aqueous salt solution. ***Related Exercises 47–48*** • Predict if a given metal reacts in an aqueous acid solution. ***Related Exercises 49–50*** • Predict if a given metal reacts in water at 25 °C. ***Related Exercises 51–52***
7.8 Single-Replacement Reactions We have examined three kinds of single-replacement reactions: a metal and aqueous solution, a metal and a dilute acid, and an active metal and water. In summary, 1. $\textbf{metal}_1 + \textbf{salt}_1(aq) \rightarrow \textbf{metal}_2 + \textbf{salt}_2(aq)$ 2. $\textbf{metal} + \textbf{acid}(aq) \rightarrow \textbf{salt}(aq) + \textbf{hydrogen}(g)$ 3. $\textbf{metal} + \textbf{water} \rightarrow \textbf{metal hydroxide}(aq) + \textbf{hydrogen}(g)$	• Propose a balanced equation for the single-replacement reaction of a given metal in an aqueous solution. ***Related Exercises 53–56 and 61–62*** • Propose a balanced equation for the single-replacement reaction of a given metal in an aqueous acid. ***Related Exercises 57–58 and 63–64*** • Propose a balanced equation for the single-replacement reaction of an active metal in water. ***Related Exercises 59–60 and 65–66***
7.9 Solubility Rules An insoluble substance does not dissolve completely in water although it may dissolve slightly. To determine whether a substance is soluble or insoluble, we will simply refer to the solubility rules in Table 7.2.	• Predict if an ionic compound dissolves in water given the general rules for solubility. ***Related Exercises 67–70***
7.10 Double-Replacement Reactions In a double-replacement precipitation reaction, two ionic compounds (salts) react in an aqueous solution. The two compounds react by switching anions and forming an insoluble substance, $\text{salt}_3(s)$. In general, $\textbf{salt}_1(aq) + \textbf{salt}_2(aq) \rightarrow \textbf{salt}_3(s) + \textbf{salt}_4(aq)$	• Propose a balanced equation for the double-replacement reaction of two salts in aqueous solution. ***Related Exercises 71–74***
7.11 Neutralization Reactions In a neutralization reaction an **acid** and a **base** react to yield a **salt** and water. A neutralization reaction is a special type of double-replacement reaction. Writing the formula for water as HOH is helpful when balancing neutralization reactions: $\textbf{acid}(aq) + \textbf{base}(aq) \rightarrow \textbf{salt}(aq) + \textbf{water}(l)$	• Propose a balanced equation for the neutralization reaction of an acid and a base. ***Related Exercises 75–78***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** the process of undergoing a chemical change
_____ **2.** a chemical reaction that releases heat energy
_____ **3.** a chemical reaction that absorbs heat energy
_____ **4.** a substance dissolved in water
_____ **5.** an insoluble solid substance produced from a reaction in aqueous solution
_____ **6.** a representation using formulas and symbols to describe a chemical change
_____ 7. a substance undergoing a chemical reaction
_____ **8.** a substance resulting from a chemical reaction
_____ **9.** a substance that speeds up a reaction without being permanently changed
_____ **10.** a particle composed of two nonmetal atoms
_____ **11.** a digit in front of a chemical formula that helps to balance a chemical equation
_____ **12.** a digit in a chemical formula that represents the number of atoms or ions appearing in the substance
_____ **13.** a type of reaction in which two substances produce a single compound
_____ **14.** a type of reaction in which a single compound produces two or more substances
_____ **15.** a type of reaction in which a more active metal displaces a less active metal from a solution
_____ **16.** a type of reaction in which two cations in different compounds exchange anions
_____ **17.** a type of reaction in which an acid and a base produce a salt and water
_____ **18.** a relative order of metals arranged by their ability to undergo a reaction
_____ **19.** a metal that reacts with water at room temperature
_____ **20.** a substance that releases hydrogen ions in water
_____ **21.** a substance that releases hydroxide ions in water
_____ **22.** an ionic compound produced by an acid–base reaction

(a) acid (*Sec. 7.11*)
(b) active metal (*Sec. 7.7*)
(c) activity series (*Sec. 7.7*)
(d) aqueous solution (*aq*) (*Sec. 7.1*)
(e) base (*Sec. 7.11*)
(f) catalyst (*Sec. 7.2*)
(g) chemical equation (*Sec. 7.2*)
(h) chemical reaction (*Sec. 7.1*)
(i) coefficient (*Sec. 7.3*)
(j) combination reaction (*Sec. 7.4*)
(k) decomposition reaction (*Sec. 7.4*)
(l) diatomic molecule (*Sec. 7.2*)
(m) double-replacement reaction (*Sec. 7.4*)
(n) endothermic reaction (*Sec. 7.1*)
(o) exothermic reaction (*Sec. 7.1*)
(p) neutralization reaction (*Sec. 7.4*)
(q) precipitate (*Sec. 7.1*)
(r) product (*Sec. 7.2*)
(s) reactant (*Sec. 7.2*)
(t) salt (*Sec. 7.11*)
(u) single-replacement reaction (*Sec. 7.4*)
(v) subscript (*Sec. 7.3*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Evidence for Chemical Reactions (Sec. 7.1)

1. Which of the following is evidence of a chemical reaction?
   (a) Dropping potassium metal into water gives a violet flame.
   (b) Dropping chromium metal into acid produces gas bubbles.
2. Which of the following is evidence of a chemical reaction?
   (a) Dropping sodium metal into water gives a yellow flame.
   (b) Dropping cobalt metal into acid produces gas bubbles.
3. Which of the following is evidence of a chemical reaction?
   (a) Heating iron in air produces a dark red solid.
   (b) Heating iron in a vacuum gives a gray film.
4. Which of the following is evidence of a chemical reaction?
   (a) Heating baking soda releases water vapor.
   (b) Heating soda ash gives a flame-extinguishing gas.
5. Which of the following is evidence of a chemical reaction?
   (a) Mixing two aqueous solutions gives gas bubbles.
   (b) Mixing two solutions produces an insoluble solid.
6. Which of the following is evidence of a chemical reaction?
   (a) Mixing two aqueous solutions produces a precipitate.
   (b) Mixing two colorless solutions gives a blue solution.

Writing Chemical Equations (Sec. 7.2)

7. Write a chemical equation for magnesium metal combining with bromine liquid to give solid magnesium bromide.
8. Write a chemical equation for iron metal combining with chlorine gas to give yellow ferrous chloride solid.

◀ **Iron in Chlorine Gas** Iron wool reacts in chlorine gas to give sparks and a yellow-brown $FeCl_3$ residue.

9. Write a chemical equation for solid zinc hydrogen carbonate decomposing to yield solid zinc carbonate, water, and carbon dioxide gas.
10. Write a chemical equation for solid cadmium hydrogen carbonate decomposing to yield solid cadmium carbonate, water, and carbon dioxide gas.
11. Write a chemical equation for the reaction of cadmium metal with aqueous cobalt(II) nitrate to produce aqueous cadmium nitrate and cobalt metal.
12. Write a chemical equation for the reaction of manganese metal with aqueous nickel(II) nitrate to produce aqueous manganese(II) nitrate and nickel metal.
13. Write a chemical equation for the reaction of aqueous solutions of lithium iodide and silver nitrate to give silver iodide precipitate and aqueous lithium nitrate.
14. Write a chemical equation for the reaction of aqueous solutions of potassium chromate and calcium sulfate to give calcium chromate precipitate and aqueous potassium sulfate.
15. Write a chemical equation for the reaction of acetic acid with aqueous ammonium hydroxide to give aqueous ammonium acetate and water.
16. Write a chemical equation for the reaction of nitric acid with aqueous lithium hydroxide to give aqueous lithium nitrate and water.

Balancing Chemical Equations (Sec. 7.3)

17. Balance each of the following chemical equations by inspection
    - **(a)** $H_2(g) + N_2(g) \rightarrow NH_3(g)$
    - **(b)** $Al_2(CO_3)_3(s) \rightarrow Al_2O_3(s) + CO_2(g)$
    - **(c)** $Sr(s) + H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$
    - **(d)** $K_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + KOH(aq)$
    - **(e)** $H_3PO_4(aq) + Mn(OH)_2(s) \rightarrow Mn_3(PO_4)_2(s) + H_2O(l)$
18. Balance each of the following chemical equations by inspection.
    - **(a)** $Co(s) + O_2(g) \rightarrow Co_2O_3(s)$
    - **(b)** $LiClO_3(s) \rightarrow LiCl(s) + O_2(g)$
    - **(c)** $Cu(s) + AgC_2H_3O_2(aq) \rightarrow Cu(C_2H_3O_2)_2(aq) + Ag(s)$
    - **(d)** $Pb(NO_3)_2(aq) + LiCl(aq) \rightarrow PbCl_2(s) + LiNO_3(aq)$
    - **(e)** $H_2SO_4(aq) + Al(OH)_3(aq) \rightarrow Al_2(SO_4)_3(aq) + H_2O(l)$
19. Balance each of the following chemical equations by inspection
    - **(a)** $Pb(s) + O_2(g) \rightarrow PbO(s)$
    - **(b)** $LiNO_3(s) \rightarrow LiNO_2(s) + O_2(g)$
    - **(c)** $Mg(s) + HC_2H_3O_2(aq) \rightarrow Mg(C_2H_3O_2)_2(aq) + H_2(g)$
    - **(d)** $Hg_2(NO_3)_2(aq) + NaBr(aq) \rightarrow Hg_2Br_2(s) + NaNO_3(aq)$
    - **(e)** $H_2CO_3(aq) + NH_4OH(aq) \rightarrow (NH_4)_2CO_3(aq) + H_2O(l)$
20. Balance each of the following chemical equations by inspection.
    - **(a)** $Sn(s) + P(s) \rightarrow Sn_3P_2(s)$
    - **(b)** $Fe_2(CO_3)_3(s) \rightarrow Fe_2O_3(s) + CO_2(g)$
    - **(c)** $Fe(s) + Cd(NO_3)_2(aq) \rightarrow Fe(NO_3)_3(aq) + Cd(s)$
    - **(d)** $Co(NO_3)_2(aq) + H_2S(g) \rightarrow CoS(s) + HNO_3(aq)$
    - **(e)** $HClO_4(aq) + Ba(OH)_2(s) \rightarrow Ba(ClO_4)_2(s) + H_2O(l)$

Classifying Chemical Reactions (Sec. 7.4)

21. Classify each reaction in Exercise 17 as one of the following: combination, decomposition, single replacement, double replacement, or neutralization.
22. Classify each reaction in Exercise 18 as one of the following: combination, decomposition, single replacement, double replacement, or neutralization.
23. Classify each reaction in Exercise 19 as one of the following: combination, decomposition, single replacement, double replacement, or neutralization.
24. Classify each reaction in Exercise 20 as one of the following: combination, decomposition, single replacement, double replacement, or neutralization.

Combination Reactions (Sec. 7.5)

25. Write a balanced equation for each of the following combination reactions.
    - **(a)** Nickel metal is heated with oxygen gas to produce solid nickel(II) oxide.
    - **(b)** Iron wool is heated with oxygen gas to produce solid iron(III) oxide.

◀ **Iron in Oxygen Gas** Iron wool reacts in oxygen gas to give bright white light and a gray Fe_2O_3 residue.

26. Write a balanced equation for each of the following combination reactions.
    - **(a)** Tin metal is heated with oxygen gas to yield solid tin(IV) oxide.
    - **(b)** Lead metal is heated with oxygen gas to yield solid lead(II) oxide.
27. Write a balanced equation for each of the following combination reactions.
    - **(a)** Carbon is heated with oxygen to give carbon monoxide gas.
    - **(b)** Phosphorus is heated with oxygen to give solid diphosphorus pentaoxide.
28. Write a balanced equation for each of the following combination reactions.
    - **(a)** Nitrogen is heated with oxygen to form dinitrogen pentaoxide gas.
    - **(b)** Chlorine is heated with oxygen to form dichlorine trioxide gas.
29. Write a balanced equation for each of the following combination reactions.
    - **(a)** Copper metal is heated with chlorine gas to give solid copper(I) chloride.
    - **(b)** Cobalt metal is heated with sulfur powder to give solid cobalt(II) sulfide.
30. Write a balanced equation for each of the following combination reactions.
    - **(a)** Iron is heated with fluorine gas to yield solid iron(III) fluoride.
    - **(b)** Lead is heated with phosphorus powder to yield solid lead(IV) phosphide.

31. Write a balanced equation for each of the following combination reactions.
    (a) Chromium is heated with oxygen to give solid chromium(III) oxide.
    (b) Chromium is heated with nitrogen to give solid chromium(III) nitride.

32. Write a balanced equation for each of the following combination reactions.
    (a) Sulfur is heated with oxygen to form sulfur dioxide gas.
    (b) Sulfur is heated with oxygen and Pt catalyst to form sulfur trioxide gas.

33. Complete and balance each of the following combination reactions.
    (a) $Li + O_2 \rightarrow$ **(b)** $Ca + O_2 \rightarrow$

34. Complete and balance each of the following combination reactions.
    (a) $Sr + O_2 \rightarrow$ **(b)** $Al + O_2 \rightarrow$

35. Complete and balance each of the following combination reactions.
    (a) $Na + I_2 \rightarrow$ **(b)** $Ba + N_2 \rightarrow$

36. Complete and balance each of the following combination reactions.
    (a) $Zn + P \rightarrow$ **(b)** $Al + S \rightarrow$

Decomposition Reactions (Sec. 7.6)

37. Write a balanced equation for each of the following decomposition reactions.
    (a) Solid potassium hydrogen carbonate decomposes with heat to give solid potassium carbonate, water, and carbon dioxide gas.
    (b) Solid strontium hydrogen carbonate decomposes with heat to give solid strontium carbonate, water, and carbon dioxide gas.

38. Write a balanced equation for each of the following decomposition reactions:
    (a) Solid silver hydrogen carbonate decomposes with heat to give solid silver carbonate, water, and carbon dioxide gas.
    (b) Solid nickel(II) hydrogen carbonate decomposes with heat to give solid nickel(II) carbonate, water, and carbon dioxide gas.

39. Write a balanced equation for each of the following decomposition reactions.
    (a) Solid cobalt(III) carbonate decomposes with heat to give solid cobalt(III) oxide, and carbon dioxide gas.
    (b) Solid tin(IV) carbonate decomposes with heat to give solid tin(IV) oxide, and carbon dioxide gas.

40. Write a balanced equation for each of the following decomposition reactions.
    (a) Solid chromium(III) carbonate decomposes with heat to give solid chromium(III) oxide and carbon dioxide gas.
    (b) Solid lead(IV) carbonate decomposes with heat to give solid lead(IV) oxide and carbon dioxide gas.

41. Complete and balance each of the following decomposition reactions.
    (a) $AgHCO_3(s) \rightarrow$
    (b) $Zn(HCO_3)_2(s) \rightarrow$

42. Complete and balance each of the following decomposition reactions.
    (a) $Li_2CO_3(s) \rightarrow$
    (b) $CdCO_3(s) \rightarrow$

43. Write a balanced equation for each of the following decomposition reactions.
    (a) Solid calcium nitrate decomposes with heat to give solid calcium nitrite and oxygen gas.
    (b) Solid silver sulfate decomposes with heat to give solid silver sulfite and oxygen gas.

44. Write a balanced equation for each of each of the following decomposition reactions.
    (a) Solid stannous chlorate decomposes with heat to give solid stannous chloride and oxygen gas.
    (b) Solid plumbic oxide decomposes with heat to give solid plumbous oxide and oxygen gas.

45. Balance each of the following decomposition reactions.
    (a) $NaClO_3(s) \rightarrow NaCl(s) + O_2(g)$
    (b) $Ca(NO_3)_2(s) \rightarrow Ca(NO_2)_2(s) + O_2(g)$

46. Balance each of the following decomposition reactions:
    (a) $AlPO_4(s) \rightarrow AlPO_3(s) + O_2(g)$
    (b) $SnSO_4(s) \rightarrow SnSO_3(s) + O_2(g)$

The Activity Series Concept (Sec. 7.7)

(Refer to the activity series in Section 7.7, or refer to Appendix C, to answer the following questions.)

47. Predict which of the following metals reacts with aqueous copper(II) nitrate.
    (a) Ag **(b)** Sn
    (c) Co **(d)** Mn

48. Predict which of the following metals reacts with aqueous cadmium nitrate.
    (a) Hg **(b)** Pb
    (c) Ni **(d)** Cr

49. Predict which of the following metals reacts with sulfuric acid.
    (a) Ag **(b)** Sn
    (c) Co **(d)** Mn

50. Predict which of the following metals reacts with hydrochloric acid.
    (a) Hg **(b)** Pb
    (c) Ni **(d)** Cr

51. Predict which of the following metals reacts with water at room temperature.
    (a) Li **(b)** Mg
    (c) Ca **(d)** Al

52. Predict which of the following metals reacts with water at room temperature.
    (a) Ba **(b)** Mn
    (c) Sn **(d)** K

Single-Replacement Reactions (Sec. 7.8)

53. Write a balanced equation for each of the following single-replacement reactions.
    (a) Aluminum wire is added to a copper(II) nitrate solution.
    (b) Copper wire is added to an aluminum nitrate solution.

54. Write a balanced equation for each of the following single-replacement reactions.
    (a) Cadmium wire is added to a silver nitrate solution.
    (b) Silver wire is added to a cadmium nitrate solution.

55. Write a balanced equation for each of the following single-replacement reactions.
    (a) Nickel metal is added to a lead(II) acetate solution.
    (b) Lead metal is added to a nickel(II) acetate solution.

56. Write a balanced equation for each of the following single-replacement reactions.
(a) Iron is added to a mercury(II) sulfate solution.
(b) Mercury is added to an iron(II) sulfate solution.

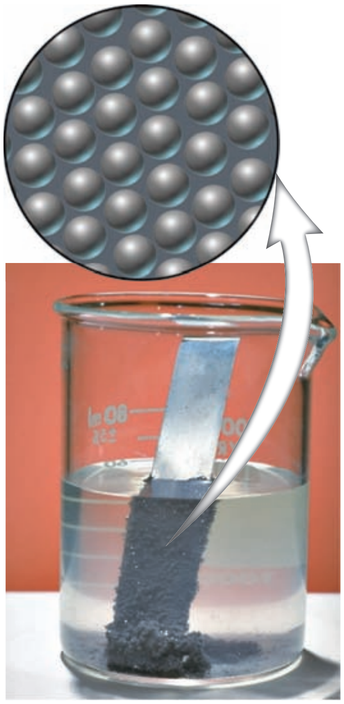
57. Write a balanced equation for each of the following single-replacement reactions.
(a) Magnesium metal is added to nitric acid.
(b) Manganese metal is added to hydrochloric acid.

58. Write a balanced equation for each of the following single-replacement reactions.
(a) Zinc granules are added to carbonic acid.
(b) Cadmium metal is added to acetic acid.

59. Write a balanced equation for each of the following single-replacement reactions.
(a) A soft, gray piece of lithium metal is added to water.
(b) A small piece of barium metal is added to water.

60. Write a balanced equation for each of the following single-replacement reactions.
(a) A gray chunk of cesium is added to water.
(b) A piece of radioactive radium is added to water.

61. Complete and balance each of the following single-replacement reactions.
(a) $Zn(s) + Pb(NO_3)_2(aq) \rightarrow$
(b) $Cd(s) + Fe(NO_3)_2(aq) \rightarrow$



Zinc in a Lead Solution Zinc metal reacts in an aqueous $Pb(NO_3)_2$ solution to give a gray deposit of spongy Pb metal.

62. Complete and balance each of the following single-replacement reactions.
(a) $Mg(s) + NiSO_4(aq) \rightarrow$
(b) $Al(s) + SnSO_4(aq) \rightarrow$

63. Complete and balance each of the following single-replacement reactions.
(a) $Zn(s) + HNO_3(aq) \rightarrow$
(b) $Cd(s) + HNO_3(aq) \rightarrow$

64. Complete and balance each of the following single-replacement reactions.
(a) $Mg(s) + H_2SO_4(aq) \rightarrow$
(b) $Al(s) + H_2SO_4(aq) \rightarrow$

Magnesium in Acid Magnesium metal reacts in aqueous sulfuric acid, H_2SO_4, to give bubbles of hydrogen gas.

65. Complete and balance each of the following single-replacement reactions.
(a) $Zn(s) + H_2O(l) \rightarrow$
(b) $Ba(s) + H_2O(l) \rightarrow$

66. Complete and balance each of the following single-replacement reactions.
(a) $Mg(s) + H_2O(l) \rightarrow$
(b) $Sr(s) + H_2O(l) \rightarrow$

Solubility Rules (Sec. 7.9)

(Refer to the solubility rules in Table 7.2, or Appendix D, to answer the following questions.)

67. Predict which of the following compounds are soluble in water:
(a) ammonium sulfide, $(NH_4)_2S$
(b) zinc carbonate, $ZnCO_3$
(c) tin(II) chromate, $SnCrO_4$
(d) lead(II) acetate, $Pb(C_2H_3O_2)_2$

68. Predict which of the following compounds are soluble in water:
(a) aluminum nitrate, $Al(NO_3)_3$
(b) sodium sulfate, Na_2SO_4
(c) cobalt(II) hydroxide, $Co(OH)_2$
(d) iron(III) phosphate, $FePO_4$

69. Predict which of the following compounds are insoluble in water:
(a) mercury(I) chloride, Hg_2Cl_2
(b) mercury(II) chloride, $HgCl_2$
(c) silver bromide, $AgBr$
(d) lead(II) iodide, PbI_2

70. Predict which of the following compounds are insoluble in water:
(a) strontium carbonate, $SrCO_3$
(b) calcium hydroxide, $Ca(OH)_2$
(c) nickel(II) sulfide, NiS
(d) mercury(II) bromide, $HgBr_2$

Double-Replacement Reactions (Sec. 7.10)

71. Write a balanced chemical equation for each of the following double-replacement reactions.
    - **(a)** Aqueous solutions of tin(II) chloride and potassium sulfide react to give a tin(II) sulfide precipitate and aqueous potassium chloride.
    - **(b)** Aqueous solutions of cobalt(II) nitrate and sodium chromate react to give a cobalt(II) chromate precipitate and aqueous sodium nitrate.
72. Write a balanced chemical equation for each of the following double-replacement reactions.
    - **(a)** Aqueous solutions of chromium(III) iodide and sodium hydroxide react to give a chromium(III) hydroxide precipitate and aqueous sodium iodide.
    - **(b)** Aqueous solutions of nickel(II) sulfate and mercury(I) nitrate react to give a mercury(I) sulfate precipitate and aqueous nickel(II) nitrate.
73. Complete and balance each of the following double-replacement reactions.
    - **(a)** $MgSO_4(aq) + BaCl_2(aq) \rightarrow$
    - **(b)** $FeSO_4(aq) + Ca(OH)_2(aq) \rightarrow$
74. Complete and balance each of the following double-replacement reactions.
    - **(a)** $AgC_2H_3O_2(aq) + SrI_2(aq) \rightarrow$
    - **(b)** $AlBr_3(aq) + Na_2CO_3(aq) \rightarrow$

Neutralization Reactions (Sec. 7.11)

75. Write a balanced equation for each of the following neutralization reactions.
    - **(a)** Lithium hydroxide solution is added to carbonic acid.
    - **(b)** Calcium hydroxide solution is added to nitric acid.
76. Write a balanced equation for each of the following neutralization reactions.
    - **(a)** Potassium hydroxide solution is added to phosphoric acid.
    - **(b)** Strontium hydroxide solution is added to acetic acid.
77. Complete and balance each of the following neutralization reactions.
    - **(a)** $HNO_3(aq) + Ca(OH)_2(aq) \rightarrow$
    - **(b)** $HNO_2(aq) + Ca(OH)_2(aq) \rightarrow$
78. Complete and balance each of the following neutralization reactions.
    - **(a)** $H_2SO_4(aq) + LiOH(aq) \rightarrow$
    - **(b)** $H_2SO_3(aq) + LiOH(aq) \rightarrow$

General Exercises

79. If the sum of the coefficients of reactants is equal to 5, is the sum of the coefficents of products equal to 5 in a balanced chemical equation?
80. If the sum of the subscripts of reactants is equal to 5, is the sum of the subscripts of products equal to 5 in a balanced chemical equation?
81. Balance each of the following chemical equations by inspection.
    - **(a)** $Fe(s) + H_2O(g) \rightarrow Fe_3O_4(s) + H_2(g)$
    - **(b)** $FeS(s) + O_2(g) \rightarrow Fe_2O_3(s) + SO_2(g)$
82. Balance each of the following chemical equations by inspection.
    - **(a)** $FeO(l) + Al(l) \rightarrow Al_2O_3(l) + Fe(l)$
    - **(b)** $MnO_2(l) + Al(l) \rightarrow Al_2O_3(l) + Mn(l)$
83. Balance each of the following chemical equations by inspection.
    - **(a)** $F_2(g) + NaBr(aq) \rightarrow Br_2(l) + NaF(aq)$
    - **(b)** $Sb_2S_3(s) + HCl(aq) \rightarrow SbCl_3(aq) + H_2S(g)$
84. Balance each of the following chemical equations by inspection.
    - **(a)** $PCl_5(s) + H_2O(l) \rightarrow H_3PO_4(aq) + HCl(aq)$
    - **(b)** $TiCl_4(s) + H_2O(g) \rightarrow TiO_2(s) + HCl(g)$
85. Balance each of the following combustion reactions by inspection.
    - **(a)** $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
    - **(b)** $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
86. Balance each of the following combustion reactions by inspection.
    - **(a)** $CH_4O(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
    - **(b)** $C_3H_8O(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Challenge Exercises

87. Iron is prepared industrially by passing carbon monoxide gas through molten iron ore, Fe_2O_3, in a blast furnace at 1500 °C. Assuming the only products are molten iron and carbon dioxide gas, write a balanced chemical equation for the manufacture of iron.
88. Chlorine is prepared industrially by heating hydrogen chloride gas with oxygen gas. Assuming the only products are water and chlorine gas, write a balanced chemical equation for the manufacture of chlorine.

Chapter 7 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following is experimental evidence of a gas being produced from a chemical reaction? (Sec. 7.1)
    - **(a)** A flaming wooden splint is extinguished.
    - **(b)** A glowing wooden splint bursts into flame.
    - **(c)** An Alka-Seltzer tablet fizzes in water.
    - **(d)** all of the above
    - **(e)** none of the above
2. Which of the following occurs naturally as diatomic molecules? (Sec. 7.2)
    - **(a)** fluorine gas
    - **(b)** helium gas
    - **(c)** krypton gas
    - **(d)** all of the above
    - **(e)** none of the above
3. What is the coefficient of oxygen gas after balancing the following equation? (Sec. 7.3)

$$\_\_SO_2(g) + \_\_O_2(g) \rightarrow \_\_SO_3(g)$$

    - **(a)** 1
    - **(b)** 2
    - **(c)** 3
    - **(d)** 5
    - **(e)** none of the above

4. Which type of chemical reaction is illustrated in question 3? (Sec. 7.4)
- **(a)** combination reaction
- **(b)** decomposition reaction
- **(c)** single-replacement reaction
- **(d)** double-replacement reaction
- **(e)** neutralization reaction

5. Predict the product from the following reaction. (Sec. 7.5)

$$Zn(s) + O_2(g) \xrightarrow{\Delta}$$

- **(a)** ZnO
- **(b)** Zn_2O
- **(c)** ZnO_2
- **(d)** Zn_2O_3
- **(e)** Zn_3O_2

6. Predict the products from the following reaction. (Sec. 7.6)

$$LiHCO_3(s) \xrightarrow{\Delta}$$

- **(a)** Li, H_2, and CO_2
- **(b)** Li, H_2O, and CO_2
- **(c)** Li_2CO_3, H_2, and CO_2
- **(d)** Li_2CO_3, H_2O, and CO_2
- **(e)** Li_2CO_3 and H_2O

7. Which of the following metals reacts with sulfuric acid? (Sec. 7.7)
- **(a)** Ag
- **(b)** Cu
- **(c)** Fe
- **(d)** all of the above
- **(e)** none of the above

8. What are the products from the following reaction? (Sec. 7.8)

$$Mg(s) + H_2SO_4(aq) \rightarrow$$

- **(a)** Mgo and H_2SO_3
- **(b)** Mgo and H_2S
- **(c)** $MgSO_4$ and H_2
- **(d)** $MgSO_4$ and H_2O
- **(e)** no reaction

9. Which of the following compounds is *insoluble* in water? (Sec. 7.9)
- **(a)** $(NH_4)_2CO_3$
- **(b)** $BaSO_4$
- **(c)** K_2CrO_4
- **(d)** Na_2S
- **(e)** $Sr(OH)_2$

10. What are the products from the following reaction? (Sec. 7.10)

$$Fe(NO_3)_3(aq) + Na_3PO_4(aq) \rightarrow$$

- **(a)** FeP and $NaNO_3$
- **(b)** $FePO_3$ and $NaNO_2$
- **(c)** $FePO_3$ and $NaNO_3$
- **(d)** $FePO_4$ and $NaNO_2$
- **(e)** $FePO_4$ and $NaNO_3$

11. What are the predicted products from the following neutralization reaction? (Sec. 7.11)

$$H_2SO_4(aq) + NH_4OH(aq) \rightarrow$$

- **(a)** NH_4SO_3 and H_2O
- **(b)** NH_4SO_4 and O_2
- **(c)** $(NH_4)_2SO_4$ and O_2
- **(d)** $(NH_4)_2SO_4$ and H_2O
- **(e)** no reaction

Key Concepts

12. Write a balanced equation for the combination of iron metal with oxygen gas to give solid iron(III) oxide.

13. Write a balanced equation for the decomposition of solid calcium hydrogen carbonate with heat to give calcium carbonate powder, steam, and carbon dioxide gas.

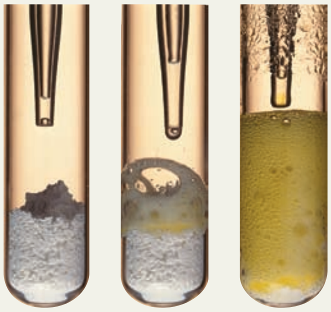
14. Write a balanced equation for the reaction of zinc metal with aqueous nickel(II) nitrate to produce aqueous zinc nitrate and nickel metal.

15. Write a balanced equation for the neutralization of nitric acid with aqueous ammonium hydroxide to give aqueous ammonium nitrate and water.

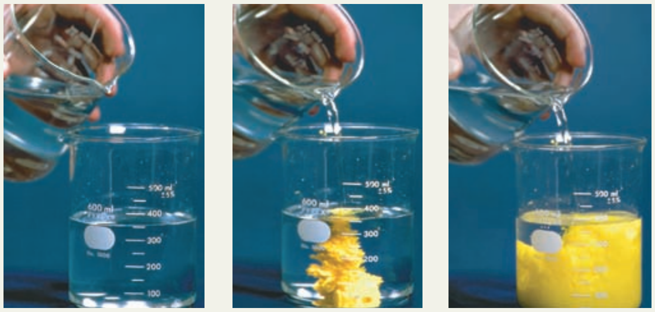
16. Write a balanced equation for the combustion of ethanol, C_2H_5OH, with oxygen to give carbon dioxide and water.

Critical Thinking

17. Write a balanced equation for the reaction of acetic acid added into a test tube containing baking soda, as shown in the illustration.

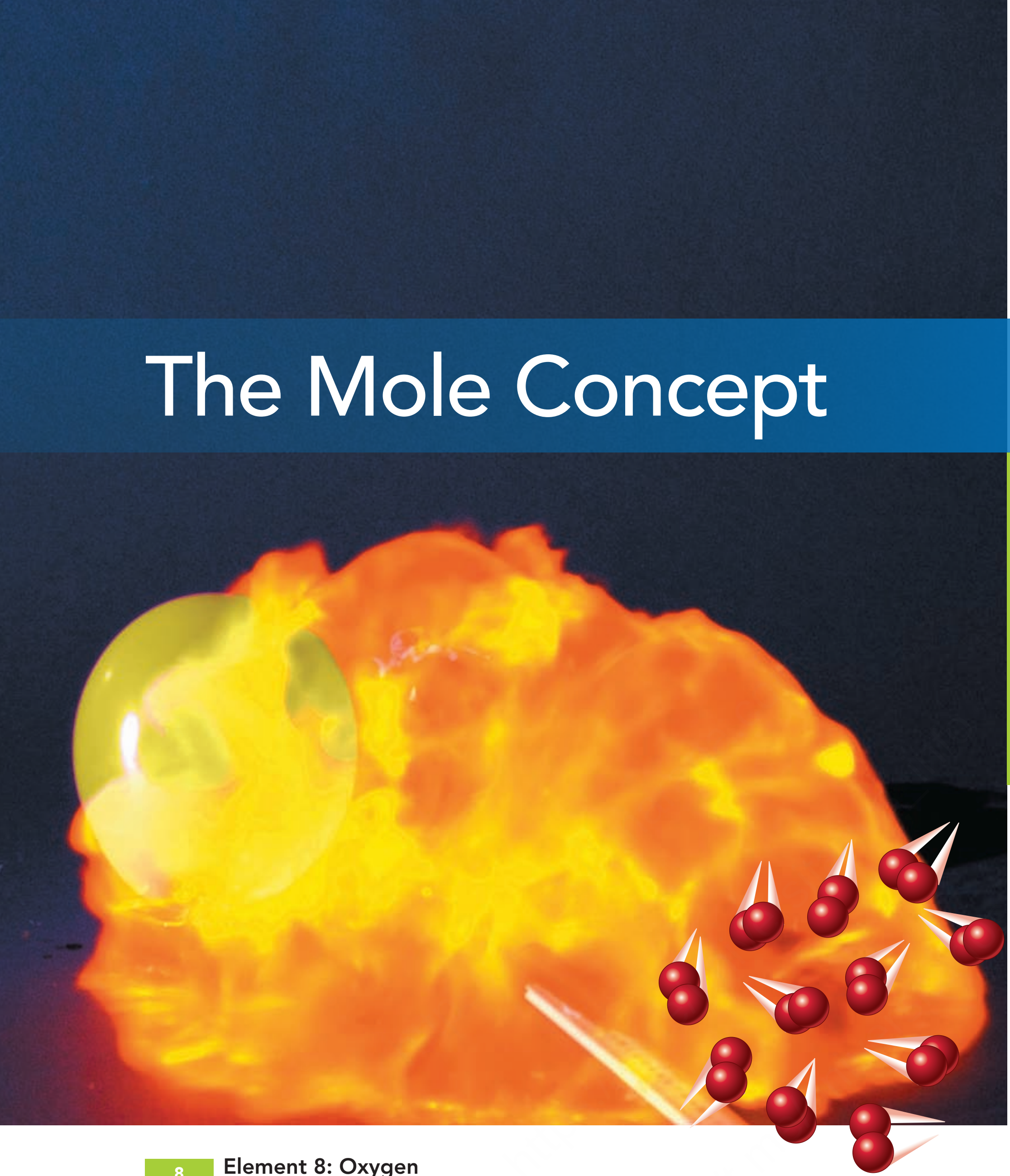


18. Write a balanced equation for the reaction of lead(II) acetate solution poured into a beaker containing sodium iodide solution, as shown in the illustration.



CHAPTER 8

The Mole Concept



8
O
oxygen

Element 8: Oxygen
The Earth's atmosphere contains about 80% nitrogen and 20% oxygen. A flame causes hydrogen gas in the yellow balloon to explode because hydrogen reacts energetically with oxygen molecules in air.

"Chance favors the prepared mind."

Louis Pasteur, French Chemist (1822–1895)

Is it possible to keep track of atoms by counting them? The answer is *"yes,"* but it is impractical to count them directly. We can, however, keep track of atoms by counting large groups of atoms. We count atoms in the same way we count eggs by the dozen (12), pencils by the gross (144), and sheets of paper by the ream (500). Because atoms are so tiny, we must consider a very large group of atoms.

When considering a very large number, it may be difficult to comprehend its magnitude. In this chapter, we will encounter an extremely large number—602,000,000,000,000,000,000,000 (6.02×10^{23}). This number refers to a group of atoms and is called *Avogadro's number.* To appreciate the enormity of this number, consider the following analogy. If you had Avogadro's number of marbles, the volume of the group of marbles would be the size of the Moon.

8.1 Avogadro’s Number

LEARNING OBJECTIVES

- Memorize the value of Avogadro’s number: 6.02×10^{23}.
- Determine the mass of Avogadro’s number of atoms for any element by referring to the periodic table.

To count the number of atoms in a sample, consider the following experiment. Let’s weigh enough atoms so that the mass in grams has the same numerical value as the atomic mass expressed in atomic mass units. For example, let’s weigh enough carbon atoms to equal a mass of 12.01 g (atomic mass = 12.01 amu).

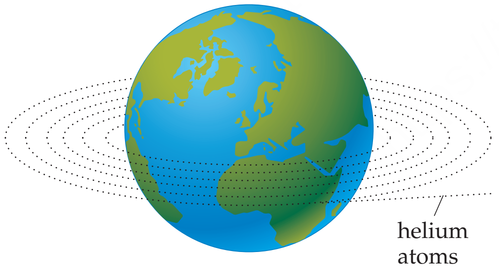
But how many atoms are in 12.01 g of carbon? Because atoms are so tiny, the number of carbon atoms in 12.01 g must be extremely large. Experiments have shown that there are 6.02×10^{23} carbon atoms in 12.01 g. This very large number is referred to as **Avogadro’s number** (symbol **N**) in honor of the Italian scientist Amedeo Avogadro (1776–1856). Avogadro’s number is sometimes referred to as the “chemist’s dozen.”

To count Avogadro’s number of atoms we only have to know the atomic mass of the element. Because one carbon atom has a mass of 12.01 amu, then 6.02×10^{23} carbon atoms have a mass of 12.01 g. The following example exercises further illustrate how to count Avogadro’s number of atoms.

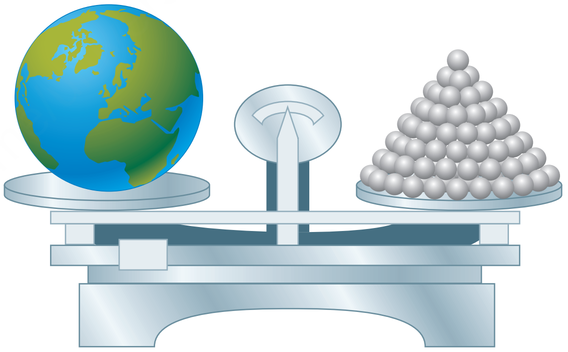
CHEMISTRY CONNECTION Analogies for Avogadro’s Number

Q: If the Moon was composed of a collection of small marbles, how many marbles would be in the collection that equals the size of the Moon?

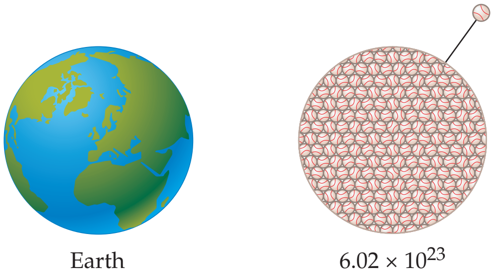
Length Analogy: If 6.02×10^{23} helium atoms were laid side by side, the total length would be long enough to encircle Earth about a million times.



Mass Analogy: The mass of 6.02×10^{23} Olympic shotput balls would be about equal to the mass of Earth.



Volume Analogy: The volume occupied by 6.02×10^{23} softballs would be about the size of Earth.



A: It would require 6.02×10^{23} ordinary small marbles to equal the size of the Moon.

EXAMPLE EXERCISE 8.1 Atomic Mass and the Periodic Table

Refer to the periodic table inside the front cover of this textbook. State the mass of Avogadro's number of atoms for each of the following metals.

(a) silver (b) mercury

Solution

The atomic mass of each element is listed below its symbol in the periodic table. The mass of Avogadro's number of atoms is the atomic mass expressed in grams. Therefore,

(a) Ag = 107.87 g (b) Hg = 200.59 g

Practice Exercise

Refer to the periodic table and state the mass of Avogadro's number of atoms for each of the following nonmetals.

(a) sulfur (b) helium

Answers:

(a) 32.07 g (b) 4.00 g

Concept Exercise

Refer to the periodic table and state the mass for each of the following number of atoms.

(a) 1 atom of carbon (b) 6.02×10^{23} atoms of carbon

Answers: See Appendix G, 8.1.

EXAMPLE EXERCISE 8.2 Avogadro's Number and the Periodic Table

State the number of atoms for each of the following metals.

(a) 63.55 g Cu (b) 40.08 g Ca

Solution

The atomic mass of each element is listed below its symbol in the periodic table. There are Avogadro's number of atoms if we express the atomic mass in grams. Therefore,

(a) 6.02×10^{23} atoms Cu (b) 6.02×10^{23} atoms Ca

Practice Exercise

Refer to the periodic table and state the number of atoms for each of the following nonmetals.

(a) 30.97 g P (b) 20.18 g Ne

Answers:

(a) 6.02×10^{23} atoms P (b) 6.02×10^{23} atoms Ne

Concept Exercise

Refer to the periodic table and state the number of atoms in each of the following masses.

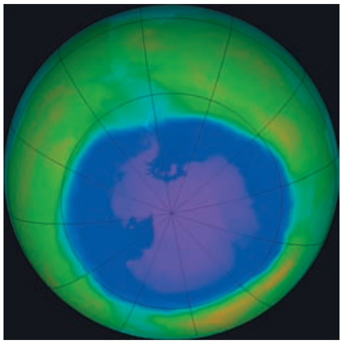
(a) 195.08 amu Pt (b) 195.08 g Pt

Answers: See Appendix G, 8.2.

CHEMISTRY CONNECTION The Ozone Hole

Q: What animal population is threatened as a result of ozone depletion over Antarctica?

In the stratosphere, about 50 miles above Earth, ozone molecules absorb high-energy ultraviolet radiation from the Sun. A reduction of ozone is of great concern because the ozone molecules protect us from the Sun's harmful UV rays. It is well known that ultraviolet radiation can cause skin cancer and eye cataracts. With less ozone in the stratosphere over Antarctica, the Sun's powerful UV rays have caused genetic damage to penguins and reduced their population.



◀ **Ozone Depletion over Antarctica** The amount of ozone in the stratosphere is indicated over Antarctica. The ozone concentration is low in the blue region and lowest in the violet area.

Chlorofluorocarbons (CFCs) have been used as propellants in aerosol cans and air-conditioning units. CFC compounds, such as Freon, are quite inert. However, when released into the environment, they diffuse into the upper atmosphere and destroy the ozone layer. CFCs react with ozone molecules in steps. First, ultraviolet radiation from the Sun strips a chlorine atom away from a CFC molecule. The resulting chlorine atom (Cl•) is referred to as a *free radical*.

$$\text{Step 1:} \qquad \text{CFC} \xrightarrow{\text{UV}} \text{Cl}\bullet$$

A chlorine free radical is very reactive and acts as a catalyst for the decomposition of ozone gas molecules, O_3, into oxygen gas molecules.

$$\text{Step 2:} \qquad 2\ O_3 \xrightarrow{\text{CL}\bullet} 3\ O_2$$

Because the chlorine free radical is not consumed in the reaction, this process can occur repeatedly, and a single chlorine free radical can destroy thousands of ozone molecules.

Ozone depletion is a worldwide problem. Many nations have halted the production and use of CFCs. The United States has resolved to phase out CFCs, but it is expensive to retrofit existing refrigeration units in automobiles and homes. However, in new homes and automobiles, air-conditioning units are being equipped with non-CFC refrigerant gases. With these preventative measures, ozone levels in the stratosphere are increasing, and scientists predict ozone levels will be restored by the middle of the twenty-first century.

A: The penguin population has been threatened by the hole in the ozone layer over Antarctica.

LEARNING OBJECTIVE

- Calculate the moles of a substance given the number of particles.

8.2 Mole Calculations I

The **mole** (symbol **mol**) is a unit of measure for an amount of a chemical substance. We define a mole as the amount of substance that contains Avogadro's number of particles, that is, 6.02×10^{23} particles. The individual particles may be atoms, molecules, formula units, or any other particles. Thus,

$$1 \text{ mol} = \text{Avogadro's number } (N) = 6.02 \times 10^{23} \text{ atoms, molecules, or formula units}$$

The mole relationship allows us to convert between the number of particles and the mass of substance. Now we are ready to try some mole calculations. Remember to apply the following three steps in the unit analysis method of problem solving.

Applying the Unit Analysis Method

Step 1: Write down the unit asked for in the answer.

Step 2: Write down the given value related to the answer.

Step 3: Apply unit factor(s) to convert the unit in the given value to the unit in the answer. To illustrate,

given value	*unit factor(s)*	*unit in answer*
(Step 2)	× (Step 3)	= (Step 1)

We can perform calculations that relate number of moles and number of particles. For instance, we can determine how many molecules of chlorine are in 0.250 mol of the yellow gas. By applying the unit analysis method of problem solving, we have

$$0.250\ \cancel{\text{mol Cl}_2} \times \frac{6.02 \times 10^{23}\ \text{molecules Cl}_2}{1\ \cancel{\text{mol Cl}_2}} = 1.51 \times 10^{23}\ \text{molecules Cl}_2$$

The following example exercises will help to reinforce the unit analysis method of problem solving for calculations involving the mole concept:

EXAMPLE EXERCISE 8.3 Mole Calculations I

Calculate the number of sodium atoms in 0.120 mol Na.

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: **atoms Na**

STEP 2: What given value is related to the answer?

Step 2: **0.120 mol Na**

STEP 3: What unit factor(s) should we apply? Since 1 mol Na = 6.02×10^{23} atoms Na, the two unit factors are 1 mol Na/6.02×10^{23} atoms Na, and its reciprocal 6.02×10^{23} atoms Na/1 mol Na.

Step 3: $\frac{\textbf{1 mol Na}}{\textbf{6.02} \times \textbf{10}^{23}\ \textbf{atoms Na}}$ **or** $\frac{\textbf{6.02} \times \textbf{10}^{23}\ \textbf{atoms Na}}{\textbf{1 mol Na}}$

Unit Analysis Map

given value	*unit factor*	*unit in answer*
0.120 mol Na	× $\frac{\textbf{6.02} \times \textbf{10}^{23}\ \textbf{atoms Na}}{\textbf{1 mol Na}}$	**= ? atoms Na**

Solution

We apply the unit factor 6.02×10^{23} atoms Na/1 mol Na to cancel moles Na (~~mol Na~~), which appears in the denominator.

$$0.120\ \cancel{\text{mol Na}} \times \frac{6.02 \times 10^{23}\ \text{atoms Na}}{1\ \cancel{\text{mol Na}}} = 7.22 \times 10^{22}\ \text{atoms Na}$$

The answer is rounded to three digits because the given value and unit factor each have three significant digits.

Practice Exercise

Calculate the number of molecules in 0.0763 mol of chlorine gas, Cl_2.

Answer:

4.59×10^{22} molecules Cl_2

Concept Exercise

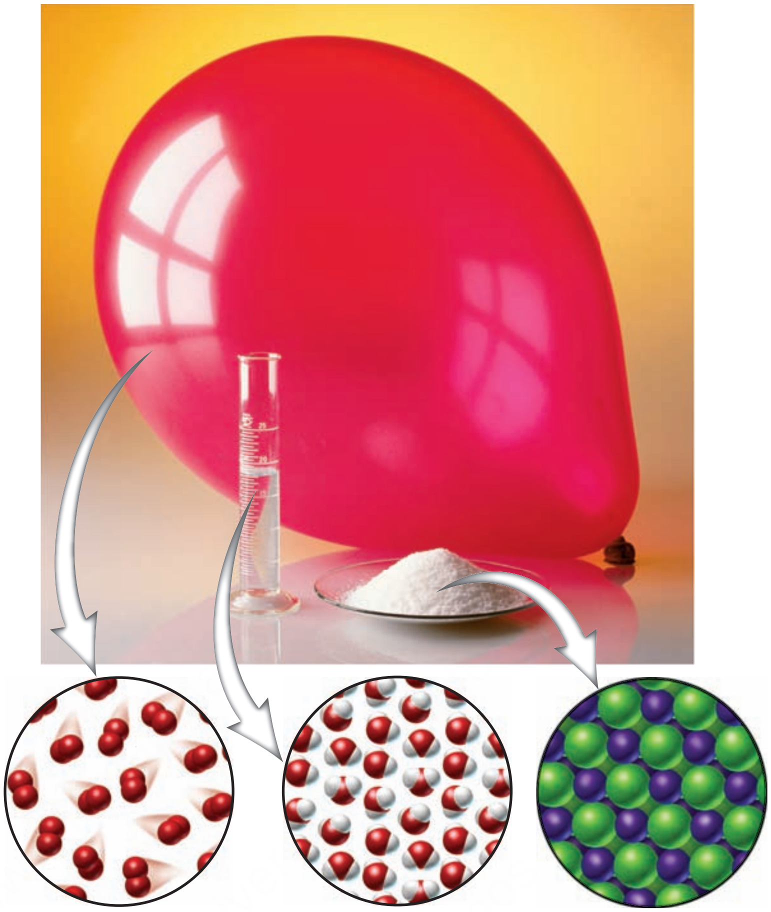
What is the number of atoms in 1.00 mol of a metallic element?

Answer: See Appendix G, 8.3.

We can also perform the reverse procedure and determine the number of moles of substance given the number of particles. For instance, we can calculate the number of moles of violet iodine crystals corresponding to 2.50×10^{23} molecules of I_2. Applying the unit analysis method of problem solving, we have

$$2.50 \times 10^{23}\ \cancel{\text{molecules I}_2} \times \frac{1\ \text{mol I}_2}{6.02 \times 10^{23}\ \cancel{\text{molecules I}_2}} = 0.415\ \text{mol I}_2$$

▶ The Mole Concept—Avogadro's Number The photograph shows a mole of oxygen, a mole of water, and a mole of table salt. The balloon contains 6.02×10^{23} molecules of O_2 gas, the graduated cylinder holds 6.02×10^{23} molecules of H_2O, and the watchglass has 6.02×10^{23} formula units of NaCl.



EXAMPLE EXERCISE 8.4 Mole Calculations I

Calculate the number of moles of potassium in 1.25×10^{21} atoms K.

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: mol K

STEP 2: What given value is related to the answer?

Step 2: 1.25×10^{21} atoms K

STEP 3: What unit factor(s) should we apply? Because 1 mol K = 6.02×10^{23} atoms K, the two unit factors are 1 mol K/6.02×10^{23} atoms K, and its reciprocal 6.02×10^{23} atoms K/1 mol K.

Step 3: $\frac{1 \text{ mol K}}{6.02 \times 10^{23} \text{ atoms K}}$ or $\frac{6.02 \times 10^{23} \text{ atoms K}}{1 \text{ mol K}}$

Unit Analysis Map

given value
1.25×10^{21} atoms K

unit factor
$\times \frac{1 \text{ mol K}}{6.02 \times 10^{23} \text{ atoms K}}$

unit in answer
= ? mol K

Solution

We apply the unit factor 1 mol K/6.02×10^{23} atoms K to cancel atoms K ($\cancel{\text{atoms K}}$), which appears in the denominator.

$$1.25 \times 10^{21}\ \cancel{\text{atoms K}} \times \frac{1\ \text{mol K}}{6.02 \times 10^{23}\ \cancel{\text{atoms K}}} = 2.08 \times 10^{-3}\ \text{mol K}$$

The answer is rounded to three digits because the given value and unit factor each have three significant digits.

Practice Exercise

Calculate the number of moles of iodine in 5.34×10^{25} molecules of I_2.

Answer:

88.7 mol I_2

Concept Exercise

What is the number of molecules in 1.00 mol of a diatomic nonmetal?

Answer: See Appendix G, 8.4.

8.3 Molar Mass

LEARNING OBJECTIVE

- Calculate the molar mass of a substance given its chemical formula.

A mole is an amount of substance that indicates the number of particles in a sample. That is, 1 mol of an element contains 6.02×10^{23} atoms. Since the atomic mass of carbon is 12.01 amu, we know that 1 mol of carbon has a mass of 12.01 g. In fact, the atomic mass of any substance expressed in grams corresponds to 1 mol of the substance. In our discussion, we will refer to the atomic mass of a substance expressed in grams as the **molar mass** (symbol **MM**).

The molar mass of an element is equal to its atomic mass expressed in grams. By referring to the periodic table, we find that the atomic mass of iron is 55.85 amu. Thus, the molar mass of iron is 55.85 g/mol. Naturally occurring oxygen is O_2, so the molecular mass is 32.00 amu, and the molar mass is 32.00 g/mol.

We can calculate the molar mass of a compound by adding the molar masses of each element. For example, we can find the molar mass of iron(III) oxide, Fe_2O_3, by summing the masses as follows:

$$Fe_2O_3\text{: } 2(55.85\ \text{g/mol Fe}) + 3(16.00\ \text{g/mol O}) = 159.70\ \text{g/mol}$$
$$Fe_2O_3\text{: } 111.70\ \text{g/mol Fe} + 48.00\ \text{g/mol O} = 159.70\ \text{g/mol}$$

The calculated molar mass of Fe_2O_3 is 159.70 g/mol. Notice that we used 16.00 g as the molar mass of oxygen. Even though the element occurs naturally as molecules of oxygen, it is atoms of oxygen that are combined in compounds. The following example exercise further illustrates the calculation of molar mass:

EXAMPLE EXERCISE 8.5 Molar Mass Calculations

Calculate the molar mass for each of the following substances:

(a) silver metal, Ag
(b) ammonia gas, NH_3
(c) magnesium nitrate, $Mg(NO_3)_2$

Solution

We begin by finding the atomic mass of each element in the periodic table. The molar mass equals the sum of the atomic masses expressed in g/mol.

(a) The atomic mass of Ag is 107.87 amu, and the molar mass of silver equals 107.87 g/mol.
(b) The sum of the atomic masses for NH_3 is 14.01 amu + 3(1.01) amu = 17.04 amu. The molar mass of ammonia equals 17.04 g/mol.
(c) The sum of the atomic masses for $Mg(NO_3)_2$ is 24.31 amu + 2(14.01 + 16.00 + 16.00 + 16.00) amu = 148.33 amu. The molar mass of magnesium nitrate equals 148.33 g/mol.

Practice Exercise

Calculate the molar mass for each of the following substances:

(a) manganese metal, Mn
(b) sulfur hexafluoride, SF_6
(b) strontium acetate, $Sr(C_2H_3O_2)_2$

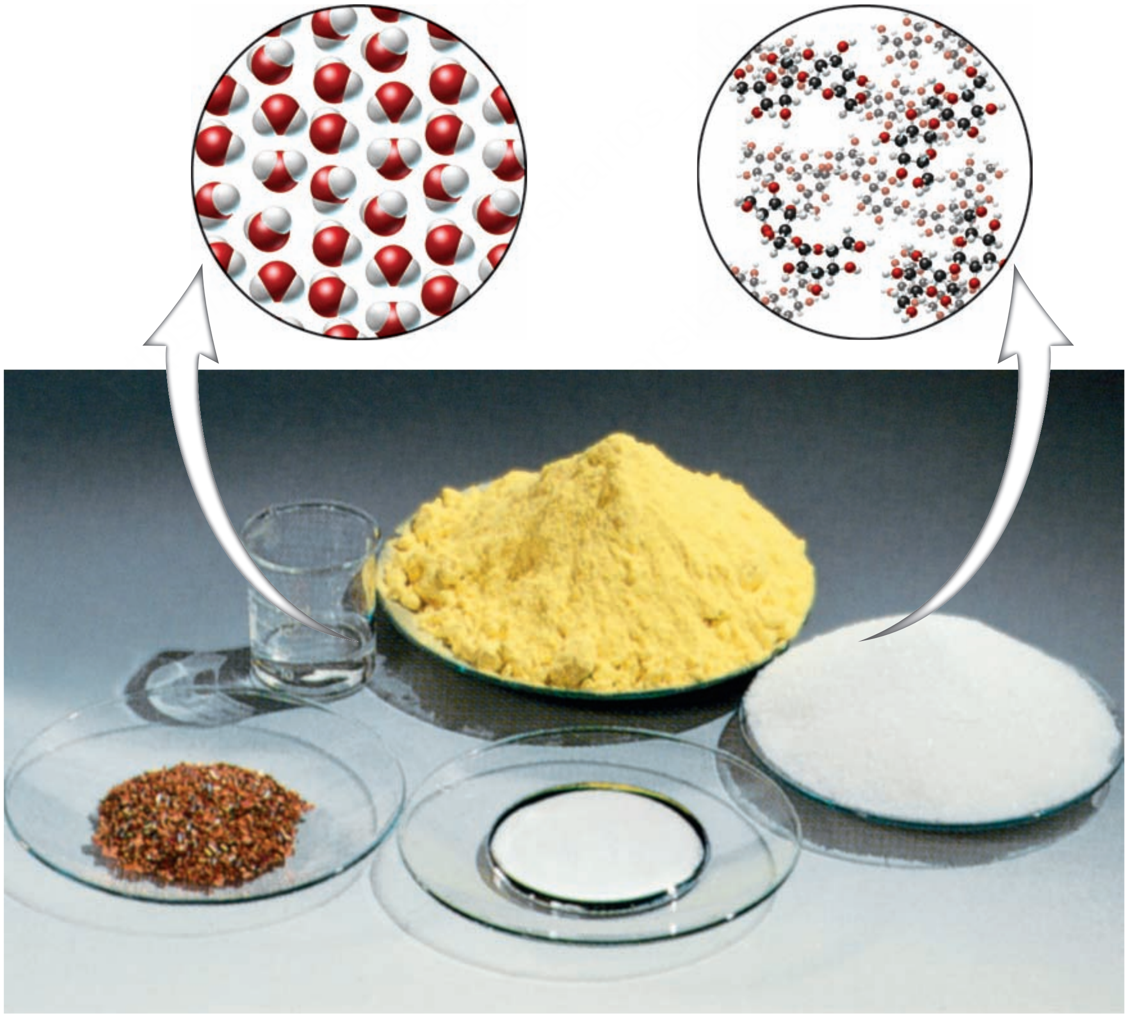
Answers:

(a) 54.94 g/mol
(b) 146.07 g/mol
(c) 205.72 g/mol

Concept Exercise

The molecular mass of water is 18.02 amu. What is the mass of Avogadro's number of water molecules?

Answer: See Appendix G, 8.5.



▲ **The Mole Concept—Molar Mass** The photograph shows a mole of water, sulfur, table sugar, mercury, and copper (clockwise). The molar masses are as follows: 18.02 g/mol H_2O, 32.07 g/mol S, 342.34 g/mol $C_{11}H_{22}O_{11}$, 200.59 g/mol Hg, and 63.55 g/mol Cu.

8.4 Mole Calculations II

LEARNING OBJECTIVE

- Calculate the mass of a substance given the number of particles.

The mole is the central unit in chemistry. It is an amount of a substance and relates the number of particles to the mass of the substance, that is,

$$6.02 \times 10^{23} \text{ particles} = 1 \text{ mol} = \text{molar mass of substance}$$

We can perform calculations that relate mass to the number of particles. For instance, we can find the mass of 2.55×10^{23} atoms of lead. By applying the unit analysis method of problem solving, we first find the number of moles:

$$2.55 \times 10^{23} \cancel{\text{atoms Pb}} \times \frac{1 \text{ mol Pb}}{6.02 \times 10^{23} \cancel{\text{atoms Pb}}} = 0.424 \text{ mol Pb}$$

To calculate the mass of lead, we multiply the number of moles of lead by its molar mass. From the periodic table we find that the molar mass of Pb is 207.2 g/mol:

$$0.424 \cancel{\text{mol Pb}} \times \frac{207.2 \text{ g Pb}}{1 \cancel{\text{mol Pb}}} = 87.8 \text{ g Pb}$$

The following example exercises reinforce mole calculations involving Avogadro's number and molar mass.

EXAMPLE EXERCISE 8.6 Mole Calculations II

What is the mass in grams of 2.01×10^{22} atoms of sulfur?

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**g S**
STEP 2:	What given value is related to the answer?	**Step 2:**	**2.01×10^{22} atoms S**
STEP 3:	What unit factor(s) should we apply? By definition, 1 mol S = 6.02×10^{23} atoms S, and the molar mass from the periodic table is 32.07 g S = 1 mol S; the two pairs of unit factors are shown in Step 3.	**Step 3:**	$\frac{1 \text{ mol S}}{6.02 \times 10^{23} \text{ atoms S}}$ **or** $\frac{6.02 \times 10^{23} \text{ atoms S}}{1 \text{ mol S}}$ $\frac{1 \text{ mol S}}{32.07 \text{ g S}}$ **or** $\frac{32.07 \text{ g S}}{1 \text{ mol S}}$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit in answer*
2.01×10^{22} atoms S	$\times \frac{1 \text{ mol S}}{6.02 \times 10^{23} \text{ atoms S}}$	$\times \frac{32.07 \text{ g S}}{1 \text{ mol S}}$	**= ? g S**

Solution

We apply the unit factor 1 mol S/6.02×10^{23} atoms S to cancel atoms S ($\cancel{\text{atoms S}}$), and 32.07 g S/1 mol S to cancel mol S ($\cancel{\text{mol S}}$):

$$2.01 \times 10^{22} \cancel{\text{atoms S}} \times \frac{1 \cancel{\text{mol S}}}{6.02 \times 10^{23} \cancel{\text{atoms S}}} \times \frac{32.07 \text{ g S}}{1 \cancel{\text{mol S}}} = 1.07 \text{ g S}$$

Practice Exercise

What is the mass of 7.75×10^{22} formula units of lead(II) sulfide, PbS?

Answer:

30.8 g PbS

Concept Exercise

What is the mass of Avogadro's number of sulfur molecules, S_8?

Answer: See Appendix G, 8.6.

EXAMPLE EXERCISE 8.7 Mole Calculations II

How many O_2 molecules are present in 0.470 g of oxygen gas?

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: **molecules O_2**

STEP 2: What given value is related to the answer?

Step 2: **0.470 g O_2**

STEP 3: What unit factor(s) should we apply? The molar mass from the periodic table is 32.00 g O_2 = 1 mol O_2, and by definition, 1 mol O_2 = 6.02×10^{23} molecules O_2. Thus, the two pairs of unit factors are shown in Step 3.

Step 3:

$$\frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \quad \text{or} \quad \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}$$

$$\frac{1 \text{ mol } O_2}{6.02 \times 10^{23} \text{ molecules } O_2} \quad \text{or} \quad \frac{6.02 \times 10^{23} \text{ molecules } O_2}{1 \text{ mol } O_2}$$

Unit Analysis Map

given value 0.470 g O_2 | *unit factor 1* $\times \dfrac{1 \text{ mol } O_2}{32.00 \text{ g } O_2}$ | *unit factor 2* $\times \dfrac{6.02 \times 10^{23} \text{ molecules } O_2}{1 \text{ mol } O_2}$ | *unit in answer* = ? molecules O_2

Solution

We apply the unit factor 1 mol O_2/32.00 g O_2 to cancel g O_2 ($\cancel{\text{g } O_2}$), and 6.02×10^{23} molecules O_2/1 mol O_2 to cancel moles O_2 ($\cancel{\text{mol } O_2}$). Thus,

$$0.470 \, \cancel{\text{g } O_2} \times \frac{1 \, \cancel{\text{mol } O_2}}{32.00 \, \cancel{\text{g } O_2}} \times \frac{6.02 \times 10^{23} \text{ molecules } O_2}{1 \, \cancel{\text{mol } O_2}} = 8.84 \times 10^{21} \text{ molecules } O_2$$

Practice Exercise

How many molecules are found in 0.175 g of fluorine gas, F_2?

Answer:

2.77×10^{21} molecules F_2

Concept Exercise

What is the mass of Avogadro's number of ozone, O_3, molecules?

Answer: See Appendix G, 8.7.

Mass of an Atom or Molecule

▲ Water molecule—H_2O

Now, let's try a different type of problem. What is the mass of a single molecule of water, given that the molar mass of water is 18.02 g/mol? This problem requires a ratio of units in the answer, that is, g/molecule:

$$\text{given value} \times \frac{\text{unit}}{\text{factor}} = \frac{\text{g}}{\text{molecule}}$$

To arrive at an answer having compound units, we start with a ratio of two units. Because the molar mass of water is 18.02 g/mol, we can write

$$\frac{18.02 \text{ g}}{1 \text{ mol } H_2O} \times \frac{\text{unit}}{\text{factor}} = \frac{\text{g}}{\text{molecule}}$$

Avogadro's number provides the unit factor 1 mol $H_2O/6.02 \times 10^{23}$ molecules. We can apply the unit factor and cancel units as follows:

$$\frac{18.02 \text{ g}}{1 \, \cancel{\text{mol } H_2O}} \times \frac{1 \, \cancel{\text{mol } H_2O}}{6.02 \times 10^{23} \text{ molecules}} = 2.99 \times 10^{-23} \text{ g/molecule}$$

The following example exercise further illustrates how to apply the mole concept to determine the mass of an individual atom or molecule.

EXAMPLE EXERCISE 8.8 Mass of a Molecule

Calculate the mass in grams for a single molecule of carbon dioxide, CO_2 (given that MM = 44.01 g/mol).

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: $\text{g } CO_2/\text{molecule } CO_2$

STEP 2: What given value is related to the answer?

Step 2: $44.01 \text{ g } CO_2/1 \text{ mol } CO_2$

STEP 3: What unit factor(s) should we apply? By definition, 1 mol CO_2 = 6.02×10^{23} molecules. Thus, the unit factors are shown in Step 3.

Step 3: $$\frac{1 \text{ mol } CO_2}{6.02 \times 10^{23} \text{ molecules } CO_2} \quad \text{or} \quad \frac{6.02 \times 10^{23} \text{ molecules } CO_2}{1 \text{ mol } CO_2}$$

Unit Analysis Map

$$\underset{\textit{given value}}{\frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}} \times \underset{\textit{unit factor}}{\frac{1 \text{ mol } CO_2}{6.02 \times 10^{23} \text{ molecules } CO_2}} = \frac{\text{g } CO_2}{\text{molecule } CO_2}$$

▲ Carbon Dioxide molecule—CO_2

Solution

We apply the unit factor 1 mol CO_2/6.02×10^{23} molecules CO_2 to cancel mol CO_2 ($\cancel{\text{mol } CO_2}$).

$$\frac{44.01 \text{ g}}{1 \cancel{\text{mol } CO_2}} \times \frac{1 \cancel{\text{mol } CO_2}}{6.02 \times 10^{23} \text{ molecules}} = 7.31 \times 10^{-23} \text{ g/molecule}$$

Practice Exercise

Calculate the mass in grams for a single molecule of carbon monoxide, CO.

Answer:

4.65×10^{-23} g/molecule

Concept Exercise

What is the mass of a carbon monoxide molecule expressed in amu?

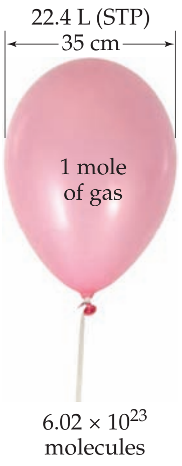
Answer: See Appendix G, 8.8.

8.5 Molar Volume

LEARNING OBJECTIVES

- Memorize the value for the molar volume of any gas at STP: 22.4 L/mol.
- Calculate the density of a gas at STP from its molar mass and molar volume.

Recall that 1 mol of any gas contains 6.02×10^{23} molecules. The gas can be hydrogen, oxygen, or any other gas. In 1811, Avogadro proposed that two gases containing equal numbers of molecules occupy equal volumes under similar conditions. This statement is known as **Avogadro's theory**, and is also referred to as Avogadro's law or Avogadro's hypothesis. It naturally follows, therefore, that 1 mol of hydrogen gas and 1 mol of oxygen gas occupy the same volume. Stated differently, 6.02×10^{23} molecules of H_2 occupy the same volume as 6.02×10^{23} molecules of O_2. In fact, 6.02×10^{23} molecules of any gas occupy the same volume as 6.02×10^{23} molecules of hydrogen or oxygen.



▲ **Figure 8.1 Molar Volume of a Gas** A balloon containing 1 mol of gas at STP has a diameter of about 35 cm. A mole of any gas contains Avogadro's number of molecules and occupies a volume of 22.4 L at STP.

What volume of gas contains Avogadro's number of molecules? At **standard temperature and pressure** (symbol **STP**) the volume is 22.4 L. Standard temperature is 0 °C. Standard pressure is 1 atm, which is the atmospheric pressure exerted by air at sea level. The volume occupied by 1 mol of any gas at STP is called the **molar volume** (Figure 8.1).

Table 8.1 compares 1 mol of five different gases. Notice that the number of molecules and the volume are constant, whereas the molar mass varies.

TABLE 8.1 Mole Relationships for Selected Gases

Gas	No. of Moles	No. of Molecules	Molar Mass	Molar Volume at STP
hydrogen, H_2	1.00	6.02×10^{23}	2.02 g/mol	22.4 L/mol
oxygen, O_2	1.00	6.02×10^{23}	32.00 g/mol	22.4 L/mol
carbon dioxide, CO_2	1.00	6.02×10^{23}	44.01 g/mol	22.4 L/mol
ammonia, NH_3	1.00	6.02×10^{23}	17.04 g/mol	22.4 L/mol
argon, Ar*	1.00	6.02×10^{23}	39.95 g/mol	22.4 L/mol

*Argon gas is composed of atoms rather than molecules.

Gas Density

The density of a gas is much less than the density of a liquid or a solid. In Section 2.8, we found that the density of water at 4 °C is 1.00 g/mL. However, the density of air at 4 °C is only 0.00129 g/mL (1.29 g/L). Therefore, the density of water is almost a 1000 times greater than that of air.

We also learned in Section 2.8 that the definition of density is mass divided by volume. The density of a gas, such as hydrogen or oxygen, is its mass divided by its volume. We can easily calculate the density of a gas at STP because 1 mol of any gas has a mass equal to its molar mass and a volume equal to its molar volume. The formula for **gas density** is

$$\frac{\text{molar mass in grams}}{\text{molar volume in liters}} = \text{density, g/L (at STP)}$$

The following example exercises illustrate how density, molar mass, and molar volume of a gas are related:

▶ **The Mole Concept—Molar Volume** The photograph depicts a molar volume of gas. For reference, a basketball contains about 7 liters of air at STP.

EXAMPLE EXERCISE 8.9 Density of a Gas at STP

Calculate the density of ammonia gas, NH_3, at STP. (MM = 17.04 g/mol)

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: g/L

STEP 2: What given value is related to the answer?

Step 2: 17.04 g/mol

STEP 3: What unit factor(s) should we apply? From the molar volume of a gas at STP, 1 mol = 22.4 L; thus, the unit factors are 1 mol/22.4 L, and 22.4 L/1 mol.

Step 3: $\frac{1\ \text{mol}}{22.4\ \text{L}}$ or $\frac{22.4\ \text{L}}{1\ \text{mol}}$

Unit Analysis Map

given value $\frac{17.04\ \text{g}}{1\ \text{mol}}$ × *unit factor* $\frac{1\ \text{mol}}{22.4\ \text{L}}$ = *unit in answer* $\frac{\text{g}}{\text{L}}$

Solution

We apply the unit factor 1 mol/22.4 L to cancel moles ($\cancel{\text{mol}}$), which appears in the denominator.

$$\frac{17.04\ \text{g}}{1\ \cancel{\text{mol}}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \text{L}} = 0.761\ \text{g/L}$$

Practice Exercise

Calculate the density of hydrogen sulfide gas, H_2S, at STP.

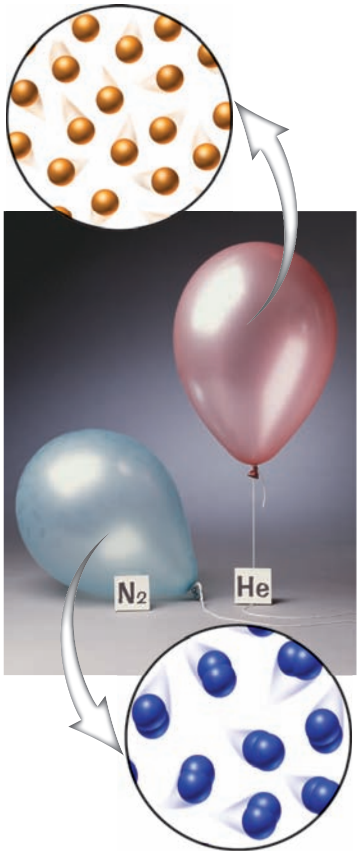
Answer:

1.52 g/L

Concept Exercise

Why does the balloon filled with helium gas rise, whereas the balloon filled with nitrogen gas does not?

Answer: See Appendix G, 8.9.



▲ Gas Balloons

EXAMPLE EXERCISE 8.10 Molar Mass of a Gas

A fire extinguisher releases 1.96 g of an unknown gas that occupies 1.00 L at STP. What is the molar mass (g/mol) of the unknown gas?

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: g/mol

STEP 2: What given value is related to the answer?

Step 2: 1.96 g/1.00 L

STEP 3: What unit factor(s) should we apply? From the molar volume of a gas at STP, 1 mol = 22.4 L; thus, the unit factors are 1 mol/22.4 L, and 22.4 L/1 mol.

Step 3: $\frac{1\ \text{mol}}{22.4\ \text{L}}$ or $\frac{22.4\ \text{L}}{1\ \text{mol}}$

Unit Analysis Map

given value		*unit factor*		*unit in answer*
$\frac{1.96\text{ g}}{1.00\text{ L}}$	×	$\frac{22.4\text{ L}}{1\text{ mol}}$	=	$\frac{\text{g}}{\text{mol}}$

Solution

We apply the unit factor 1 mol/22.4 L to cancel liters ($\cancel{\text{L}}$), which appears in the denominator.

$$\frac{1.96\text{ g}}{1.00\ \cancel{\text{L}}} \times \frac{22.4\ \cancel{\text{L}}}{1\text{ mol}} = 43.9\text{ g/mol}$$

Because the unknown gas is from a fire extinguisher, we suspect that it is carbon dioxide. By summing the molar mass for CO_2(44.01 g/mol), we help confirm that the unknown gas is carbon dioxide.

Practice Exercise

Boron trifluoride gas is used in the manufacture of computer chips. Given that 1.505 g of the gas occupies 497 mL at STP, what is the molar mass of boron fluoride gas?

Answer:

67.8 g/mol

Concept Exercise

What is the volume of one mole of any gas at STP?

Answer: See Appendix G, 8.10.

8.6 Mole Calculations III

LEARNING OBJECTIVE

- Calculate the volume of a gas at STP given its mass, or number of particles.

As mentioned previously, the mole is a central unit in chemical calculations. A mole has three interpretations:

1. A mole is Avogadro's number of particles.
2. A mole of a substance has a mass equal to its atomic mass in grams.
3. A mole of any gas at STP occupies a volume of 22.4 L.

We can state the three interpretations in the form of equations as follows:

$$1\text{ mol} = 6.02 \times 10^{23}\text{ particles}$$
$$1\text{ mol} = \text{molar mass (g/mol)}$$
$$1\text{ mol} = 22.4\text{ L at STP}$$

From the preceding mole equations, we can relate the number of particles to the mass or volume of a gaseous substance at STP. The process of interchanging these three quantities involves mole calculations.

The following example exercises illustrate calculations involving the number of moles of a substance. In these calculations, we will use the concepts of Avogadro's number, molar mass, and molar volume.

Helpful Hint
The Mole

The mole concept is invaluable in performing the calculations encountered in introductory chemistry. If we want to relate number of particles in a given substance to its mass or gaseous volume, we apply the mole concept. The relationship of number of particles to mole is found using Avogadro's number; the relationship of mole to mass employs molar mass; the relationship of mole to volume of a gas at STP employs molar volume. The molar volume of any gas is 22.4 L/mol at STP.

EXAMPLE EXERCISE 8.11 Mole Calculations III

What is the mass of 3.36 L of ozone gas, O_3, at STP?

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**$g\ O_3$**
STEP 2: What given value is related to the answer?	**Step 2:**	**$3.36\ L\ O_3$**
STEP 3: What unit factor(s) should we apply? From the molar volume of a gas at STP, $1\ mol\ O_3 = 22.4\ L\ O_3$. The molar mass from the periodic table is $48.00\ g\ O_3 = 1\ mol\ O_3$. Thus, the two pairs of unit factors are shown in Step 3.	**Step 3:**	$\dfrac{1\ mol\ O_3}{22.4\ L\ O_3}$ **or** $\dfrac{22.4\ L\ O_3}{1\ mol\ O_3}$ $\dfrac{1\ mol\ O_3}{48.00\ g\ O_3}$ **or** $\dfrac{48.00\ g\ O_3}{1\ mol\ O_3}$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit in answer*
$3.36\ L\ O_3$	$\times \dfrac{1\ mol\ O_3}{22.4\ L\ O_3}$	$\times \dfrac{48.00\ g\ O_3}{1\ mol\ O_3}$	$= ?\ g\ O_3$

Solution

We apply the unit factor $1\ mol\ O_3/22.4\ L\ O_3$ to cancel $L\ O_3$ ($\cancel{L\ O_3}$), and $48.00\ g\ O_3/1\ mol\ O_3$ to cancel moles O_3 ($\cancel{mol\ O_3}$)

$$3.36\ \cancel{L\ O_3} \times \frac{1\ \cancel{mol\ O_3}}{22.4\ \cancel{L\ O_3}} \times \frac{48.00\ g\ O_3}{1\ \cancel{mol\ O_3}} = 7.20\ g\ O_3$$

Practice Exercise

What volume is occupied by 0.125 g of methane gas, CH_4, at STP?

Answer:

$0.174\ L\ CH_4$ (174 mL)

Concept Exercise

How many molecules of ozone are in one mole of gas?

Answer: See Appendix G, 8.11.

EXAMPLE EXERCISE 8.12 Mole Calculations III

How many molecules of hydrogen gas, H_2, occupy 0.500 L at STP?

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**molecules H_2**
STEP 2: What given value is related to the answer?	**Step 2:**	**$0.500\ L\ H_2$**
STEP 3: What unit factor(s) should we apply? From the molar volume of a gas at STP, $1\ mol\ H_2 = 22.4\ L\ H_2$ and by definition, $1\ mol\ H_2 = 6.02 \times 10^{23}$ molecules H_2. Thus, the two pairs of unit factors are shown in Step 3.	**Step 3:**	$\dfrac{1\ mol\ H_2}{22.4\ L\ H_2}$ **or** $\dfrac{22.4\ L\ H_2}{1\ mol\ H_2}$ $\dfrac{1\ mol\ H_2}{6.02 \times 10^{23}\ \text{molecules}\ H_2}$ **or** $\dfrac{6.02 \times 10^{23}\ \text{molecules}\ H_2}{1\ mol\ H_2}$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit in answer*
$0.500\ \text{L H}_2$	$\times\ \dfrac{1\ \text{mol H}_2}{22.4\ \text{L H}_2}$	$\times\ \dfrac{6.02 \times 10^{23}\ \text{molecules H}_2}{1\ \text{mol H}_2}$	$=\ ?\ \text{molecules H}_2$

Solution

We apply the unit factor 1 mol H_2/22.4 L H_2 to cancel L H_2 ($\cancel{\text{L H}_2}$), and 6.02×10^{23} molecules H_2/1 mol H_2 to cancel moles H_2 ($\cancel{\text{mol H}_2}$). Thus,

$$0.500\ \cancel{\text{L H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{22.4\ \cancel{\text{L H}_2}} \times \frac{6.02 \times 10^{23}\ \text{molecules H}_2}{1\ \cancel{\text{mol H}_2}} = 1.34 \times 10^{22}\ \text{molecules H}_2$$

Practice Exercise

What is the volume occupied by 3.33×10^{21} atoms of helium gas, He, at STP?

Answer:

0.124 L He (124 mL)

Concept Exercise

What is the mass of one mole of hydrogen gas?

Answer: See Appendix G, 8.12.

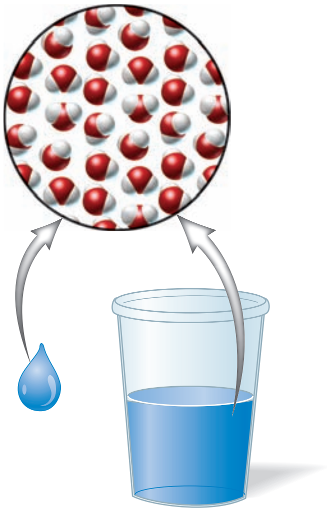
LEARNING OBJECTIVE

- Calculate the percent composition of a compound given its chemical formula.

8.7 Percent Composition

Previously, we defined percent as parts per 100 parts. Another way to define percent is to say that it is the amount of a given quantity compared to an entire sample. For example, a dime is 10% of a dollar. Now, let's apply the percent concept to the composition of a compound.

The **percent composition** of a compound lists the mass percent of each element. The percent composition of water, H_2O, for example, is 11% hydrogen and 89% oxygen. According to the law of definite composition discussed in Section 3.5, the elements in a compound are always present in the same proportion by mass. Therefore, water always contains 11% hydrogen and 89% oxygen regardless of the amount. A drop of water, a milliliter of water, and a pool of water all contain 11% hydrogen and 89% oxygen.



◀ **Percent Composition** The percent composition of H_2O is the same for a drop of water and a glass of water, that is, 11% hydrogen and 89% oxygen.

The percent of each element in water is given, but how are the values obtained? We can calculate the percent composition of water as follows. Let's begin by assuming that we have 1 mol of H_2O. A mole of H_2O contains 2 mol of hydrogen and 1 mol of oxygen. Thus,

$$\begin{aligned} 2(\text{mol H}) + 1\text{ mol O} &= 1\text{ mol H}_2\text{O} \\ 2(\text{molar mass H}) + 1\text{ molar mass O} &= 1\text{ molar mass H}_2\text{O} \\ 2(1.01\text{ g H}) + 16.00\text{ g O} &= \text{g H}_2\text{O} \\ 2.02\text{ g H} + 16.00\text{ g O} &= 18.02\text{ g H}_2\text{O} \end{aligned}$$

Next, we find the percent composition of water by comparing the molar masses of hydrogen and oxygen to the molar mass of the whole compound:

$$\frac{2.02\text{ g H}}{18.02\text{ g H}_2\text{O}} \times 100\% = 11.2\%\text{ H}$$

$$\frac{16.00\text{ g O}}{18.02\text{ g H}_2\text{O}} \times 100\% = 88.79\%\text{ O}$$

The following example exercise further illustrates the calculation of the percent composition of a compound.

EXAMPLE EXERCISE 8.13 Percent Composition of a Substance

TNT (trinitrotoluene) is a white crystalline substance that explodes at 240 °C. Calculate the percent composition of TNT, $C_7H_5(NO_2)_3$.

Solution

Let's calculate the percent composition assuming there is 1 mol of TNT. For compounds with parentheses, it is necessary to count the number of atoms of each element carefully. That is, 1 mol of $C_7H_5(NO_2)_3$ contains 7 mol of C atoms, 5 mol of H atoms, 3 mol of NO_2 (that is, 3 mol N atoms and 6 mol of O atoms).

We begin the calculation by finding the molar mass of $C_7H_5(NO_2)_3$ as follows:

$$7(12.01\text{ g C}) + 5(1.01\text{ g H}) + 3(14.01\text{ g N} + 32.00\text{ g O}) = \text{g C}_7\text{H}_5(\text{NO}_2)_3$$

$$84.07\text{ g C} + 5.05\text{ g H} + 42.03\text{ g N} + 96.00\text{ g O} = 227.15\text{ g C}_7\text{H}_5(\text{NO}_2)_3$$

Now, let's compare the mass of each element to the total molar mass of the compound, that is, 227.15 g.

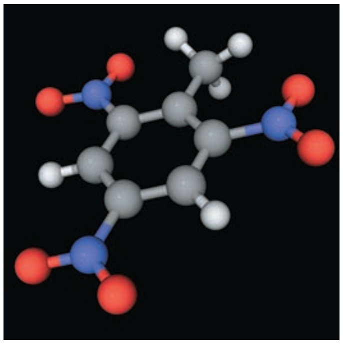
$$\frac{84.07\text{ g C}}{227.15\text{ g C}_7\text{H}_5(\text{NO}_2)_3} \times 100\% = 37.01\%\text{ C}$$

$$\frac{5.05\text{ g H}}{227.15\text{ g C}_7\text{H}_5(\text{NO}_2)_3} \times 100\% = 2.22\%\text{ H}$$

$$\frac{42.03\text{ g N}}{227.15\text{ g C}_7\text{H}_5(\text{NO}_2)_3} \times 100\% = 18.50\%\text{ N}$$

$$\frac{96.00\text{ g O}}{227.15\text{ g C}_7\text{H}_5(\text{NO}_2)_3} \times 100\% = 42.26\%\text{ O}$$

The percent composition of TNT reveals that the explosive is mostly oxygen by mass. When we sum the individual percentages (37.01 + 2.22 + 18.50 + 42.26 = 99.99%), we verify our calculation and find that the total is approximately 100%.



▲ **TNT** Trinitrotoluene is a white crystalline compound that is incorporated into a stick of dynamite.

Practice Exercise

EDTA (ethylenediaminetetraacetic acid) is used as a food preservative and in the treatment of lead poisoning. Calculate the percent composition of EDTA, $C_{10}H_{16}N_2O_8$.

Answer:

41.09% C, 5.53% H, 9.587% N, and 43.79% O

Concept Exercise

If an analysis of sugar, $C_xH_yO_z$, gave 40.0% C and 6.7% H, what is the percent oxygen?

Answer: See Appendix G, 8.13.

LEARNING OBJECTIVES

- Calculate the empirical formula of a compound given its mass composition.
- Calculate the empirical formula of a compound given its percent composition.

8.8 Empirical Formula

During the late 1700s, chemists experimented to see how elements reacted to form compounds. In particular, they were interested in the reactions of elements with oxygen to form oxides. By measuring the mass of an element and the mass of the oxide after reaction, chemists could determine the formula of the compound.

The **empirical formula** of a compound corresponds to the simplest whole-number ratio of atoms of each element in a molecule, or ions in an ionic compound. The following example illustrates the determination of an empirical formula. A 1.640 g sample of radioactive radium was heated with oxygen to produce 1.755 g of radium oxide. By subtracting the mass of radium from the mass of the oxide (1.755 g − 1.640 g), we find that the mass gain was 0.115 g. Thus, the radium reacted with 0.115 g of oxygen gas.

Now, let's find the empirical formula for radium oxide, $Ra_?O_?$. The empirical formula is the simplest whole-number ratio of radium and oxide ions in an ionic compound. We can determine this ratio from the moles of each reactant. If the molar mass of radium is given as 226.03 g/mol, we can proceed as follows:

$$1.640\ \cancel{\text{g Ra}} \times \frac{1\ \text{mol Ra}}{226.03\ \cancel{\text{g Ra}}} = 0.00726\ \text{mol Ra}$$

We can calculate the moles of oxygen (16.00 g/mol) as follows:

$$0.115\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.00719\ \text{mol O}$$

The mole ratio of the elements in radium oxide is $Ra_{0.00726}O_{0.00719}$. We can simplify the mole ratio by dividing by the smaller number:

$$\text{Ra}_{\frac{0.00726}{0.00719}}\text{O}_{\frac{0.00719}{0.00719}} = \text{Ra}_{1.01}\text{O}_{1.00}$$

Because the empirical formula must be a small whole-number ratio, we will round the calculated chemical formula of radium oxide from $Ra_{1.01}O_{1.00}$ to RaO.

Notice that the empirical formula for radium oxide is similar to other Group IIA/2 elements. Magnesium, calcium, and barium are in Group IIA/2 because their compounds have similar empirical formulas. The empirical formulas of the metal oxides all have a metal-to-oxygen ratio of 1:1, that is, MgO, CaO, and BaO. Accordingly, the empirical formula of radium oxide, which has a Group IIA/2 metal, is RaO.

▲ **Radioactivity** An element, such as radium, which is radioactive is indicated by a "propeller" symbol.

EXAMPLE EXERCISE 8.14 Empirical Formula from Mass Composition

In a laboratory experiment, 0.500 g of scandium was heated and allowed to react with oxygen from the air. The resulting product oxide had a mass of 0.767 g. Now, let's find the empirical formula for scandium oxide, $Sc_?O_?$.

Solution

The empirical formula is the simplest whole-number ratio of scandium and oxygen in the compound scandium oxide. This ratio is experimentally determined from the moles of each reactant. The moles of scandium are calculated as follows:

$$0.500\ \cancel{\text{g Sc}} \times \frac{1\ \text{mol Sc}}{44.96\ \cancel{\text{g Sc}}} = 0.0111\ \text{mol Sc}$$

The moles of oxygen are calculated after first subtracting the mass of Sc from the product, $Sc_?O_?$:

$$0.767\ \text{g Sc}_x\text{O}_y - 0.500\ \text{g Sc} = 0.267\ \text{g O}$$

The moles of oxygen are calculated from the mass of oxygen that reacted:

$$0.267\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.0167\ \text{mol O}$$

The mole ratio in scandium oxide is $Sc_{0.0111}O_{0.0167}$. To simplify the ratio and obtain small whole numbers, we divide by the smaller number:

$$Sc_{\frac{0.0111}{0.0111}}O_{\frac{0.0167}{0.0111}} = Sc_{1.00}O_{1.50}$$

Because the calculated ratio is not close to whole numbers, we cannot round off the experimental ratio $Sc_{1.00}O_{1.50}$. However, we can double the ratio to obtain whole numbers, that is, $Sc_{2.00}O_{3.00}$. Thus, the empirical formula is Sc_2O_3.

Practice Exercise

Iron can react with chlorine gas to give two different compounds, $FeCl_2$ and $FeCl_3$. If 0.558 g of metallic iron reacts with chlorine gas to yield 1.621 g of iron chloride, which iron compound is produced in the experiment?

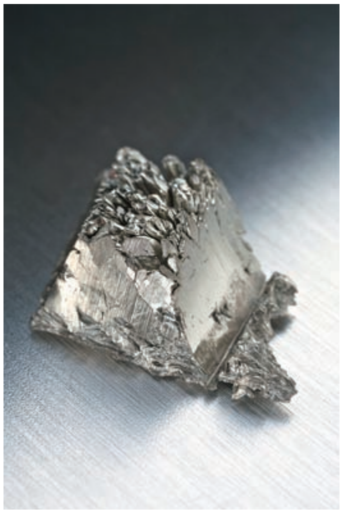
Answer:

$FeCl_3$

Concept Exercise

If 0.500 mol of yellow powder sulfur reacts with 0.500 mol of oxygen gas, what is the empirical formula of the sulfur oxide?

Answer: See Appendix G, 8.14.



▲ **Scandium** Scandium is a soft, silvery rare earth metal discovered in Scandinavia in 1879.

Empirical Formulas from Percent Composition

Benzene was a common liquid solvent until the Environmental Protection Agency (EPA) discovered that it was a carcinogen. Let's calculate the formula for benzene if its percent composition is 92.2% carbon and 7.83% hydrogen.

The empirical formula expresses the simplest whole-number ratio of carbon to hydrogen atoms in a molecule of benzene. To calculate the moles of each element, let's assume we have 100 g of sample. In 100 g of benzene there are 92.2 g of carbon and 7.83 g of hydrogen. That is, the percentage of each element corresponds to its mass in 100 g of the compound. We find the moles of C and H as we did in the previous examples:

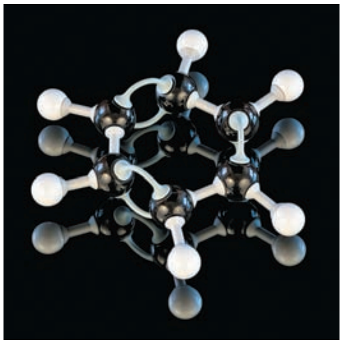
$$92.2\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 7.68\ \text{mol C}$$

$$7.83\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 7.75\ \text{mol H}$$

The mole ratio of the elements in benzene is $C_{7.68}H_{7.75}$. We can simplify the mole ratio by dividing both values by the smaller number, 7.68:

$$C_{\frac{7.68}{7.68}}H_{\frac{7.75}{7.68}} = C_{1.00}H_{1.01}$$

Because the ratio $C_{1.00}H_{1.01}$ is close to whole numbers, we can round off to CH to obtain the empirical formula for benzene. The following example exercise further illustrates the calculation of an empirical formula from percent composition data.



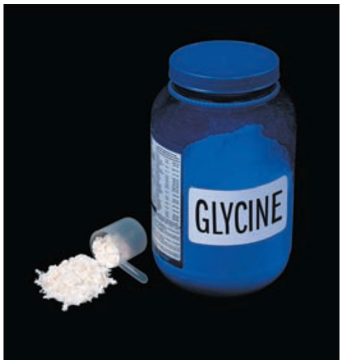
▲ **Benzene** The benzene molecule contains six carbon atoms and six hydrogen atoms joined in a circle with alternating single and double bonds.

EXAMPLE EXERCISE 8.15 Empirical Formula from Percent Composition

Glycine is an amino acid found in protein. An analysis of glycine gave the following data: 32.0% carbon, 6.7% hydrogen, 18.7% nitrogen, and 42.6% oxygen. Calculate the empirical formula of the amino acid.

Solution

If we assume a 100-g sample, then the percentage of each element equals its mass in 100 g of glycine, that is, 32.0 g C, 6.7 g H, 18.7 g N, and 42.6 g O. We can determine the empirical formula as follows:



▲ **Glycine** Glycine is one of the most common amino acids found in protein.

$$32.0\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 2.66\ \text{mol C}$$

$$6.7\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 6.6\ \text{mol H}$$

$$18.7\ \cancel{\text{g N}} \times \frac{1\ \text{mol N}}{14.01\ \cancel{\text{g N}}} = 1.33\ \text{mol N}$$

$$42.6\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 2.66\ \text{mol O}$$

The mole ratio of the elements in the amino acid is $C_{2.66}H_{6.6}N_{1.33}O_{2.66}$. We can find a small whole-number ratio by dividing by the smallest number:

$$\text{C}\,\frac{2.66}{1.33}\,\text{H}\,\frac{6.6}{1.33}\,\text{N}\,\frac{1.33}{1.33}\,\text{O}\,\frac{2.66}{1.33} = C_{2.00}H_{5.0}N_{1.00}O_{2.00}$$

Simplifying, we find that the empirical formula for the amino acid glycine is $C_2H_5NO_2$.

Practice Exercise

Calculate the empirical formula for caffeine given the following percent composition: 49.5% C, 5.15% H, 28.9% N, and 16.5% O.

Answer:

$C_4H_5N_2O$

Concept Exercise

If 0.500 mol of yellow powder sulfur reacts with 0.750 mol of oxygen gas, what is the empirical formula of the sulfur oxide?

Answer: See Appendix G, 8.15.

8.9 Molecular Formula

LEARNING OBJECTIVE

- Calculate the molecular formula for a compound given its empirical formula and molar mass.

Molecular compounds are represented by individual molecules. In the previous section, we found the empirical formula for benzene is CH. However, this does not represent the actual formula for benzene. The actual formula, or **molecular formula**, is a multiple of the empirical formula. We can represent the molecular formula as $(CH)_n$, where n is some multiple of CH, such as 2, 3, 4, 5, or 6.

Now consider acetylene, used in oxyacetylene welding, which also has the empirical formula CH. Even though acetylene and benzene have unrelated properties, each shares the same empirical formula, CH. In addition, the compound styrene, used in manufacturing styrofoam cups, also shares the same empirical formula, CH.

Because benzene, acetylene, and styrene are different compounds, their molecular formulas must be different. Experiments have provided the molar mass for each of these compounds. For benzene, the molar mass is 78 g/mol; for acetylene, it is 26 g/mol; and for styrene, it is 104 g/mol. We can indicate the number of multiples of the empirical formula for each compound as follows:

$$\begin{aligned} \text{Benzene: } (CH)_n &= 78\ \text{g/mol} \\ \text{Acetylene: } (CH)_n &= 26\ \text{g/mol} \\ \text{Styrene: } (CH)_n &= 104\ \text{g/mol} \end{aligned}$$

The mass of the empirical formula CH is found by adding the molar masses of C and H, that is, 12 g + 1 g = 13 g/mol. Now we can determine how many multiples of the empirical formula are in benzene. We have

$$\text{Benzene: } \frac{(CH)_n}{CH} = \frac{78\ \text{g/mol}}{13\ \text{g/mol}}$$

$$n = 6$$

Therefore, the molecular formula of benzene is $(CH)_6$, which we can write as C_6H_6. Similarly, we can find the molecular formula for acetylene:

$$\text{Acetylene: } \frac{(CH)_n}{CH} = \frac{26\ g/mol}{13\ g/mol}$$
$$n = 2$$

The molecular formula of acetylene is $(CH)_2$ or C_2H_2. In a similar fashion, we can find the number of multiples of the empirical formula in styrene:

$$\text{Styrene: } \frac{(CH)_n}{CH} = \frac{104\ g/mol}{13\ g/mol}$$
$$n = 8$$

Thus, the molecular formula of styrene is $(CH)_8$ or C_8H_8.

Even though benzene, acetylene, and styrene have the same empirical formula, their molecular formulas are different. Thus, benzene, acetylene, and styrene are different compounds. The following example exercise further illustrates how to determine a molecular formula from an empirical formula:

EXAMPLE EXERCISE 8.16 Molecular Formula from Empirical Formula

The empirical formula for fructose, or fruit sugar, is CH_2O. If the molar mass of fructose is 180 g/mol, find the actual molecular formula for the sugar.

Solution

We can indicate the molecular formula of fructose as $(CH_2O)_n$. The molar mass of the empirical formula CH_2O is 12 g C + 2(1 g H) + 16 g O = 30 g/mol. Thus, the number of multiples of the empirical formula is

$$\text{Fructose: } \frac{(CH_2O)_n}{CH_2O} = \frac{180\ g/mol}{30\ g/mol}$$
$$n = 6$$

Thus, the molecular formula of fructose is $(CH_2O)_6$ or $C_6H_{12}O_6$.

Practice Exercise

Ethylene dibromide was banned as a grain pesticide in 1983 because it was found to accumulate in groundwater. Laboratory tests showed EDB to be a carcinogen that causes reproductive disorders in test animals. Calculate (a) the empirical formula and (b) the molecular formula for ethylene dibromide given its approximate molar mass of 190 g/mol and its percent composition:12.7% C, 2.1% H, and 85.1% Br.

Answers:

(a) CH_2Br (b) $C_2H_4Br_2$

Concept Exercise

If the molecular formula of hydrogen peroxide is H_2O_2, what is the empirical formula?

Answer: See Appendix G, 8.16.

▲ **Honey** Glucose and fructose are sugars in honey, and have the same molecular formula, $C_6H_{12}O_6$. However, the structural formulas for glucose and fructose are different.

Note For molecular compounds, the empirical and molecular formulas are usually different. For ionic compounds, the actual formulas are almost always identical to the empirical formulas. For example, the actual formula for sodium chloride is NaCl, not Na_2Cl_2 or some other multiple. The common exception, however, is the mercury(I) compounds. Notice that the actual formula of mercury(I) chloride, Hg_2Cl_2, is twice the empirical formula.

A CLOSER LOOK Avogadro's Number

Q: What is the mass of Avogadro's number of silicon atoms?

In 1911, Ernest Rutherford determined a value for Avogadro's number. Rutherford counted the alpha particles emitted from radioactive radium using a Geiger counter, which had been invented by his assistant Hans Geiger. As the alpha particles decayed into helium gas, he then measured the volume of the gas. From the data, Rutherford calculated a value of 6.11×10^{23} for Avogadro's number.

The most recent experimental value for Avogadro's number was found using an X-ray method. X-rays were used to determine the spacing of atoms in an ultrapure crystal of silicon. After striking the nuclei of silicon atoms, the path of the X-ray is altered. The volume occupied by a single atom is calculated from the spacing between two X-rays. A precise value for Avogadro's number is obtained by dividing the volume of an individual silicon atom into the volume of a molar mass of pure silicon.

We can illustrate how to calculate a value for Avogadro's number as follows. In the figure notice that the volume occupied by each silicon atom is 0.020116 nm^3 and that the volume of the crystal is 12.114 cm^3. To find Avogadro's number we simply divide the volume of one atom of silicon into the volume of the silicon crystal. However, we must use the same units for the volumes of the atom and the crystal. For example, we can convert the volume of the crystal from cm^3 to nm^3:

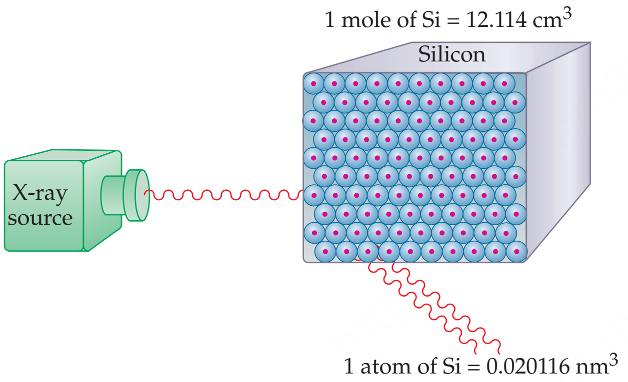
$$12.114\ \text{cm}^3 \times \left(\frac{1\ \text{m}}{10^2\text{cm}}\right)^3 \times \left(\frac{10^9\ \text{nm}}{1\ \text{m}}\right)^3 = \text{nm}^3$$

$$12.114\ \cancel{\text{cm}^3} \times \frac{1\ \cancel{\text{m}^3}}{10^6\ \cancel{\text{cm}^3}} \times \frac{10^{27}\text{nm}^3}{1\ \cancel{\text{m}^3}} = 1.2114 \times 10^{22}\text{nm}^3$$

When we divide the volume of the silicon atom into the volume of the crystal, we find the number of silicon atoms:

$$1.2114 \times 10^{22}\ \cancel{\text{nm}^3} \times \frac{1\ \text{atom Si}}{0.020116\ \cancel{\text{nm}^3}} = 6.0221 \times 10^{23}\ \text{atoms Si}$$

In this example, we obtained an experimental value of 6.0221×10^{23} atoms for Avogadro's number. Currently, the most accurate value for Avogadro's number is 6.0221415×10^{23}.



◀ **X-Ray Diffraction of a Silicon Crystal** We can calculate the volume of a Si atom after using an X-ray method to find the spacing of atoms in a silicon crystal.

A: The mass of Avogadro's number of silicon atoms equals its molar mass—28.09 g.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
8.1 Avogadro's Number **Avogadro's number** (*N*) is the value that corresponds to the number of atoms in 12.01 g of carbon. Moreover, the number of atoms in the atomic mass of any element expressed in grams is 6.02×10^{23}.	• Memorize the value of Avogadro's number: 6.02×10^{23}. ***Related Exercises: 1–4*** • Determine the mass of Avogadro's number of atoms for any element by referring to the periodic table. ***Related Exercises: 1–4***
8.2 Mole Calculations I **A mole (mol)** is the amount of substance that contains Avogadro's number of particles, that is, 6.02×10^{23} particles. A mole is the central unit in performing chemical formula calculations. These computations, based on moles of substance, are known as mole calculations.	• Calculate the moles of a substance given the number of particles. ***Related Exercises: 5–12***

Key Concepts	Learning Objectives and Related Exercises
8.3 Molar Mass The mass of 1 mol of any substance is called the **molar mass (MM)**. The molar mass of an element corresponds to the atomic mass of the element. For instance, the molar mass of carbon is 12.01 g/mol, and the molar mass of oxygen is 16.00 g/mol. The molar mass of a compound corresponds to the sum of the atomic masses for all the atoms of each element in the compound. For example, the molar mass of carbon monoxide, CO, equals 28.01 g/mol (12.01 g/mol + 16.00 g/mol).	• Calculate the molar mass of a substance given its chemical formula. ***Related Exercises: 13–16***
8.4 Mole Calculations II The mole is perhaps the most important unit in chemistry. A mole relates the number of particles (atoms, molecules, or formula units) to the amount of a substance. Given the moles of a substance, we can calculate the mass of the substance as well as the number of particles.	• Calculate the mass of a substance given the number of particles. ***Related Exercises: 17–22***
8.5 Molar Volume **Avogadro's theory** states that equal volumes of gases, at the same conditions, contain equal numbers of molecules. The volume of 1 mol of a gaseous substance at standard conditions is called the **molar volume.** The molar volume of any gas at **standard temperature and pressure (STP)** is 22.4 L/mol. Standard conditions are 0 °C and 1 atm pressure. **Gas density** is found by dividing the molar mass of a gas by the molar volume; that is, we divide the molar mass by 22.4 L/mol.	• Memorize the value for the molar volume of any gas at STP: 22.4 L/mol. ***Related Exercises: 23–24*** • Calculate the density of a gas at STP from its molar mass and molar volume. ***Related Exercises: 25–30***
8.6 Mole Calculations III The mole concept relates the number of particles, molar mass, and molar volume of a gas. Given the number of liters of a gaseous substance at STP, we can apply the mole concept to find the mass of the substance as well as the number of particles.	• Calculate the volume of a gas at STP given its mass, or number of particles. ***Related Exercises: 31–38***
8.7 Percent Composition The **percent composition** of a compound is found by comparing the mass contributed by each element to the molar mass of the substance. A list of the resulting percentages of each element is the percent composition of the compound.	• Calculate the percent composition of a compound given its chemical formula. ***Related Exercises: 39–46***
8.8 Empirical Formula The **empirical formula** is the simplest whole-number ratio of the atoms of the elements in a compound. The empirical formula can be calculated from (1) experimental synthesis data, or (2) the percent composition of the compound. After calculating the moles of each element in a compound, the mole ratio is simplified to small whole numbers.	• Calculate the empirical formula of a compound given its mass composition. ***Related Exercises: 47–52*** • Calculate the empirical formula of a compound given its percent composition. ***Related Exercises: 53–58***
8.9 Molecular Formula The **molecular formula** is the actual ratio of the atoms in a molecular compound. For example, lactic acid and glucose each have the empirical formula CH_2O. The molecular formula of lactic acid, $C_3H_6O_3$, is three times the empirical formula. The actual formula of glucose, $C_6H_{12}O_6$, is six times the empirical formula.	• Calculate the molecular formula of a compound given its empirical formula and molar mass. ***Related Exercises: 59–68***

Problem-Solving Organizer

Topic	Procedure	Example
Avogadro's Number Sec. 8.1	Express the atomic mass of an element in grams to find the mass of Avogadro's number of atoms.	Avogadro's number of Fe atoms has a mass of 55.85 g.
Mole Calculations I Sec. 8.2	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert the given unit to the unit in the answer.	How many iron atoms are in 0.500 mol of Fe? $0.500\ \cancel{\text{mol}} \times \frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \cancel{\text{mol}}} = 3.01 \times 10^{23}\ \text{atoms}$
Molar Mass Sec. 8.3	Sum the masses of each atom of each element in a compound to calculate the molar mass.	What is the molar mass of iron(III) oxide, Fe_2O_3? $2(55.85\ \text{g/mol}) + 3(16.00\ \text{g/mol}) = 159.70\ \text{g/mol}$
Mole Calculations II Sec. 8.4	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert the given unit to the unit in the answer.	What is the mass of 5.00×10^{23} formula units of Fe_2O_3? $5.00 \times 10^{23}\ \cancel{\text{formula units}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{formula units}}}$ $\times \frac{159.70\ \text{g}\ Fe_2O_2}{1\ \cancel{\text{mol}}} = 133\ \text{g}\ Fe_2O_3$
Molar Volume Sec. 8.5	The volume occupied by 1 mol of any gas at STP is 22.4 L.	Volume of 1.00 mol of O_2 at STP is 22.4 L.
Mole Calculations III Sec. 8.6	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert the given unit to the unit in the answer.	What is the mass of 2.50 L of O_2 gas at STP? $2.50\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{32.00\ \text{g}\ O_2}{1\ \cancel{\text{mol}}} = 3.57\ \text{g}\ O_2$
Percent Composition Sec. 8.7	Find the mass of each element compared to the total molar mass of the compound, all multiplied by 100%.	What is the percent composition of hydrogen peroxide, H_2O_2 (MM = 34.02 g/mol)? $\frac{2(1.01)\ \cancel{\text{g}}}{34.02\ \cancel{\text{g}}} \times 100\% = 5.94\%\ \text{H}$ $\frac{2(16.00)\ \text{g}}{34.02\ \text{g}} \times 100\% = 94.6\%\ \text{O}$
Empirical Formula Sec. 8.8	Find the mole ratio of each element in the compound, and simplify to whole numbers.	What is the empirical formula of hydrogen peroxide, given 2.25 mol H and 2.22 mol O? $\text{H}\,\frac{2.25}{2.22}\,\text{O}\,\frac{2.22}{2.22} = H_{1.01}O_{1.00}$ thus, HO
Molecular Formula Sec. 8.9	Find the ratio of the molar mass of the compound to the molar mass of the empirical formula unit. Use this ratio as a multiplier for the subscripts in the empirical formula unit.	What is the molecular formula of hydrogen peroxide, given the molar mass is about 34 g/mol? $\frac{(\text{HO})_n}{\text{HO}} = \frac{34\ \text{g/mol}}{17\ \text{g/mol}} = 2$ thus, H_2O_2

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_______ **1.** the value that corresponds to the number of atoms in 12.01 g of carbon, that is, 6.02×10^{23} particles
_______ **2.** the amount of substance that contains 6.02×10^{23} particles
_______ **3.** the mass of 1 mol of pure substance expressed in grams
_______ **4.** the volume occupied by 1 mol of any gas at standard conditions
_______ **5.** a temperature of 0 °C and 1 atm pressure
_______ **6.** equal volumes of gases, under the same conditions of temperature and pressure, contain equal numbers of molecules
_______ **7.** the ratio of mass per unit volume for a gas, expressed in grams per liter
_______ **8.** a list of the mass percent of each element in a compound
_______ **9.** a chemical formula that expresses the simplest whole-number ratio of atoms of each element in a molecule, or ions in an ionic compound
_______**10.** a chemical formula that expresses the actual number of atoms of each element in a molecule

(a) Avogadro's number (N) (*Sec. 8.1*)
(b) Avogadro's theory (*Sec. 8.5*)
(c) empirical formula (*Sec. 8.8*)
(d) gas density (*Sec. 8.5*)
(e) molar mass (MM) (*Sec. 8.3*)
(f) molar volume (*Sec. 8.5*)
(g) mole (mol) (*Sec. 8.2*)
(h) molecular formula (*Sec. 8.9*)
(i) percent composition (*Sec. 8.7*)
(j) standard temperature and pressure (STP) (*Sec. 8.5*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Avogadro's Number (Sec. 8.1)

1. Refer to the periodic table and state the atomic mass (in amu) of one atom for each of the following metals.
(a) sodium **(b)** strontium
(c) silicon **(d)** selenium

2. Refer to the periodic table and state the atomic mass (in amu) of one atom for each of the following nonmetals.
(a) beryllium **(b)** barium
(c) boron **(d)** bromine

3. Refer to the periodic table and state the mass of 6.02×10^{23} atoms of each of the following metals.
(a) sodium **(b)** strontium
(c) silicon **(d)** selenium

4. Refer to the periodic table and state the mass of 6.02×10^{23} atoms of each of the following nonmetals.
(a) beryllium **(b)** barium
(c) boron **(d)** bromine

Mole Calculations I (Sec. 8.2)

5. State the number of particles in each of the following.
(a) 1 mol of carbon atoms, C
(b) 1 mol of methane molecules, CH_4

6. State the number of particles in each of the following.
(a) 1 mol of silicon atoms, Si
(b) 1 mol of silane molecules, SiH_4

7. State the number of moles represented by each of the following.
(a) 6.02×10^{23} atoms of sulfur, S
(b) 6.02×10^{23} molecules of sulfur dioxide, SO_2

8. State the number of moles represented by each of the following.
(a) 6.02×10^{23} atoms of selenium, Se
(b) 6.02×10^{23} molecules of selenium dioxide, SeO_2

9. Calculate the number of particles in each of the following.
(a) 0.100 mol magnesium atoms, Mg
(b) 0.200 mol chlorine molecules, Cl_2
(c) 0.300 mol magnesium chloride formula units, $MgCl_2$

10. Calculate the number of particles in each of the following.
(a) 0.250 mol calcium atoms, Ca
(b) 0.500 mol fluorine molecules, F_2
(c) 0.300 mol calcium fluoride formula units, CaF_2

11. Calculate the number of moles containing each of the following.
(a) 2.50×10^{22} atoms of iron, Fe
(b) 5.00×10^{23} molecules of sulfur tri oxide, SO_3
(c) 7.50×10^{24} formula units of iron(II) sulfate, $FeSO_4$

12. Calculate the number of moles containing each of the following.
(a) 1.00×10^{22} atoms of cobalt, Co
(b) 2.00×10^{23} molecules of carbon dioxide, CO_2
(c) 3.00×10^{24} formula units of cobalt(II) carbonate, $CoCO_3$

Molar Mass (Sec. 8.3)

13. Refer to the periodic table and state the molar mass for each of the following.
(a) aluminum, Al **(b)** silicon, Si
(c) arsenic, As **(d)** sulfur, S

14. Refer to the periodic table and state the molar mass for each of the following.
(a) gallium, Ga **(b)** germanium, Ge
(c) antimony, Sb **(d)** selenium, P

15. Refer to the periodic table and calculate the molar mass for each of the following.
    (a) calcium sulfide, CaS
    (b) calcium sulfate, $CaSO_4$
    (c) dichlorine pentaoxide, Cl_2O_5
    (d) glycerin, $C_3H_5(OH)_3$

16. Refer to the periodic table and calculate the molar mass for each of the following.
    (a) iron(II) acetate, $Fe(C_2H_3O_2)_2$
    (b) iron(II) phosphate, $Fe_3(PO_4)_2$
    (c) tribromine octaoxide, Br_3O_8
    (d) trinitroglycerin, $C_3H_5O_3(NO_2)_3$

Mole Calculation II (Sec. 8.4)

17. Calculate the mass in grams for each of the following.
    (a) 2.95×10^{23} atoms of mercury, Hg
    (b) 1.16×10^{22} molecules of nitrogen, N_2
    (c) 5.05×10^{21} formula units of barium chloride, $BaCl_2$

18. Calculate the mass in grams for each of the following.
    (a) 1.21×10^{24} atoms krypton, Kr
    (b) 6.33×10^{22} molecules of dinitrogen oxide, N_2O
    (c) 4.17×10^{21} formula units of magnesium perchlorate, $Mg(ClO_4)_2$

19. Calculate the number of particles in each of the following.
    (a) 1.50 g potassium, K
    (b) 0.470 g oxygen, O_2
    (c) 0.555 g silver chlorate, $AgClO_3$

20. Calculate the number of particles in each of the following.
    (a) 7.57 g platinum, Pt
    (b) 3.88 g ethane, C_2H_6
    (c) 0.152 g aluminum chloride, $AlCl_3$

21. Calculate the mass in grams for a single particle of the following.
    (a) lithium, Li
    (b) silicon, Si
    (c) sulfur dioxide, SO_2
    (d) nitrous oxide, N_2O

22. Calculate the mass in grams for a single particle of the following.
    (a) sodium, Na
    (b) germanium, Ge
    (c) sulfur trioxide, SO_3
    (d) nitric oxide, NO

Molar Volume (Sec. 8.5)

23. State standard conditions for a gas in degrees Celsius and atmospheres.

24. State standard conditions for a gas in Kelvin units and atmospheres.

25. Calculate the density for each of the following gases at STP.
    (a) argon, Ar
    (b) fluorine, F_2
    (c) methane, CH_4
    (d) propane, C_3H_8

26. Calculate the density for each of the following gases at STP.
    (a) xenon, Xe
    (b) chlorine, Cl_2
    (c) ethane, C_2H_6
    (d) butane, C_4H_{10}

27. Calculate the molar mass for each of the following gases given the STP density.
    (a) oxygen, 1.43 g/L
    (b) phosphine, 1.52 g/L
    (c) nitrous oxide, 1.97 g/L
    (d) Freon-12, 5.40 g/L

28. Calculate the molar mass for each of the following gases given the STP density.
    (a) ozone, 2.14 g/L
    (b) silane, 1.43 g/L
    (c) nitric oxide, 1.34 g/L
    (d) Freon-22, 3.86 g/L

29. Given 1 mol of each gas listed, complete the following table.

Gas	Molecules	Mass	STP Volume
nitrogen, N_2			
oxygen, O_2			
nitrogen dioxide, NO_2			

30. Given 1 mol of each gas listed, complete the following table.

Gas	Molecules	Mass	STP Volume
nitrogen, N_2			
hydrogen, H_2			
nitrogen ammonia, NH_3			

Mole Calculations III (Sec. 8.6)

31. Calculate the number of molecules in each of the following gases at STP.
    (a) 0.150 g of carbon monoxide, CO
    (b) 100.0 mL of hydrogen, H_2

32. Calculate the number of molecules in each of the following gases at STP.
    (a) 2.75 g of nitrogen monoxide, NO
    (b) 70.5 mL of ammonia, NH_3

33. Calculate the mass in grams for each of the following gases at STP.
    (a) 1.05 L of hydrogen sulfide, H_2S
    (b) 5.42×10^{22} molecules of propane, C_3H_8

34. Calculate the mass in grams for each of the following gases at STP.
    (a) 5.33 L of dinitrogen trioxide, N_2O_3
    (b) 1.82×10^{23} molecules of butane, C_4H_{10}

35. Calculate the volume in liters for each of the following gases at STP.
    (a) 0.250 g of helium, He
    (b) 2.22×10^{22} molecules of methane, CH_4

36. Calculate the volume in liters for each of the following gases at STP.
    (a) 5.05 g of nitrogen, N_2
    (b) 4.18×10^{24} molecules of ethane, C_2H_6

37. Use the given quantity for each gas listed to complete the following table.

Gas	Molecules	Mass	STP Volume
methane, CH_4	1.50×10^{23}		
ethane, C_2H_6		7.50 g	
propane, C_3H_8			5.58 L

38. Use the given quantity for each gas listed to complete the following table.

Gas	Molecules	Mass	STP Volume
ozone, O_3	2.50×10^{22}		
carbon dioxide, CO_2		1.83 g	
carbon monoxide, CO			0.930 L

Percent Composition (Sec. 8.7)

39. If the percent sodium in a salt crystal, NaCl, is 39.34%, what is the percent sodium in a kilogram of salt?
40. If the percent chlorine in a salt crystal, NaCl, is 60.66%, what is the percent chlorine in a pillar of salt?
41. Lactose is a sugar found in milk. Calculate the percent composition of lactose, $C_{12}H_{22}O_{11}$.
42. Benzoyl peroxide is the active ingredient in an acne cream. Calculate the percent composition of benzoyl peroxide, $C_7H_6O_3$.
43. Methionine is an essential amino acid in protein. Calculate the percent composition of methionine, $C_5H_{11}NSO_2$.
44. Monosodium glutamate (MSG) is added to food to enhance the flavor. Calculate the percent composition of MSG, $NaC_5H_8NO_4$.
45. Mercurochrome dissolved in water is a reddish liquid formerly used as an antiseptic. Calculate the percent composition of mercurochrome, $HgNa_2C_{20}H_8Br_2O_6$.
46. Chlorophyll is a dark green plant pigment. Calculate the percent composition of chlorophyll, $C_{55}H_{70}MgN_4O_6$.

Empirical Formula (Sec. 8.8)

47. A 1.115-g sample of cobalt was heated with excess sulfur to give 1.722 g of cobalt sulfide. What is the empirical formula of the cobalt sulfide?
48. A 1.115-g sample of cobalt was heated with excess sulfur to give 2.025 g of cobalt sulfide. What is the empirical formula of the cobalt sulfide?
49. In an experiment, 2.410 g of copper oxide produced 1.925 g of copper metal after heating with hydrogen gas. What is the empirical formula of the copper oxide?
50. In an experiment, 1.550 g of mercury oxide decomposed to give oxygen gas and 1.435 g of liquid mercury. What is the empirical formula of the mercury oxide?
51. In an experiment, 0.500 g of tin reacted with nitric acid to give tin oxide. If the oxide had a mass of 0.635 g, what is the empirical formula of the tin oxide?
52. In an experiment, 0.500 g of nickel reacted with air to give 0.704 g of nickel oxide. What is the empirical formula of the nickel oxide?
53. The percent composition of copper chloride is 64.1% Cu and 35.9% Cl. Calculate the empirical formula.
54. The percent composition of vanadium oxide is 68.0% V and 32.0% O. Calculate the empirical formula.
55. The percent composition of tin bromide is 42.6% Sn and 57.4% Br. Calculate the empirical formula.
56. The percent composition of bismuth oxide is 89.7% Bi and 10.3% O. Calculate the empirical formula.
57. Trichloroethylene (TCE) is a common solvent used to degrease machine parts. Calculate the empirical formula for TCE if the percent composition is 18.25% C, 0.77% H, and 80.99% Cl.
58. Dimethyl sulfoxide (DMSO) is a liniment for horses and has been used in the treatment of arthritis for humans. Calculate the empirical formula for DMSO, given that the percent composition is 30.7% C, 7.74% H, 20.5% O, and 41.0% S.

Molecular Formula (Sec. 8.9)

59. Aspirin has a molar mass of 180 g/mol. If the empirical formula is $C_9H_8O_4$, what is the molecular formula of aspirin?
60. Quinine is used to treat malaria and has a molar mass of 325 g/mol. If the empirical formula is $C_{10}H_{12}NO$, what is the molecular formula of quinine?
61. Adipic acid is used to manufacture nylon. If the molar mass is 147 g/mol and the empirical formula is $C_3H_5O_2$, what is the molecular formula of the compound?
62. Hexamethylene diamine is used to make nylon. If the molar mass is 115 g/mol and the empirical formula is C_3H_8N, what is the molecular formula of the compound?
63. Ethylene glycol is used as permanent antifreeze. If the molar mass is 62 g/mol and the percent composition is 38.7% C, 9.74% H, and 51.6% O, what is the molecular formula of ethylene glycol?
64. Dioxane is a common solvent for plastics. If the molar mass is 88 g/mol and the percent composition is 54.5% C, 9.15% H, and 36.3% O, what is the molecular formula of dioxane?
65. Lindane is an insecticide. If the molar mass is 290 g/mol and the percent composition of 24.8% C, 2.08% H, and 73.1% Cl, what is the molecular formula of lindane?
66. Mercurous chloride is a fungicide. If the molar mass is 470 g/mol and the percent composition is 85.0% Hg and 15.0% Cl, what is the molecular formula of mercurous chloride?
67. Nicotine has a molar mass of 160 g/mol. If the percent composition is 74.0% C, 8.70% H, and 17.3% N, what is the molecular formula of nicotine?
68. Allose is a sugar with a molar mass of 180 g/mol. If the percent composition is 40.0% C, 6.72% H, and 53.3% O, what is the molecular formula of allose?

General Exercises

69. In 1871, Mendeleev predicted the undiscovered element *ekasilicon.* In 1886, the element was discovered in Germany and was given the name germanium. If 0.500 g of germanium reacts with chlorine gas to give 1.456 g of germanium chloride, what is its empirical formula?
70. In 1871, Mendeleev predicted the undiscovered element *ekaaluminum.* In 1875, the element was discovered in Gaul (France) and was given the name gallium. If 0.500 g of gallium reacts with oxygen gas to give 0.672 g of gallium oxide, what is its empirical formula?
71. A quadrillion, 1×10^{15}, is the approximate number of red blood cells in 50,000 people. Which is greater: a quadrillion red blood cells or the number of nickel atoms in a 5-g nickel coin?
72. By rubbing the Lincoln profile on a penny coin, a student removed just enough mass to be detected by an electronic balance, 0.0001 g. How many copper atoms were rubbed off the coin?
73. The mass of Earth is 6×10^{24} kg, and an average rodent mole has a mass of 100 g. Which weighs more, Earth or a mole of furry moles?
74. A light year, 9.5×10^{12} km, is the distance light travels in a year and an average rodent mole has a length of 17 cm. Which is longer, a light year or a mole of furry moles?

75. Calculate the number of carbon atoms in 1.00 g of table sugar, $C_{12}H_{22}O_{11}$.

76. Calculate the number of carbon atoms in 1.00 g of blood sugar, $C_6H_{12}O_6$.

77. What is the mass of cuprite, Cu_2O, that contains 10.0 g of copper?

78. What is the mass of rust, Fe_2O_3, that contains 10.0 g of iron?

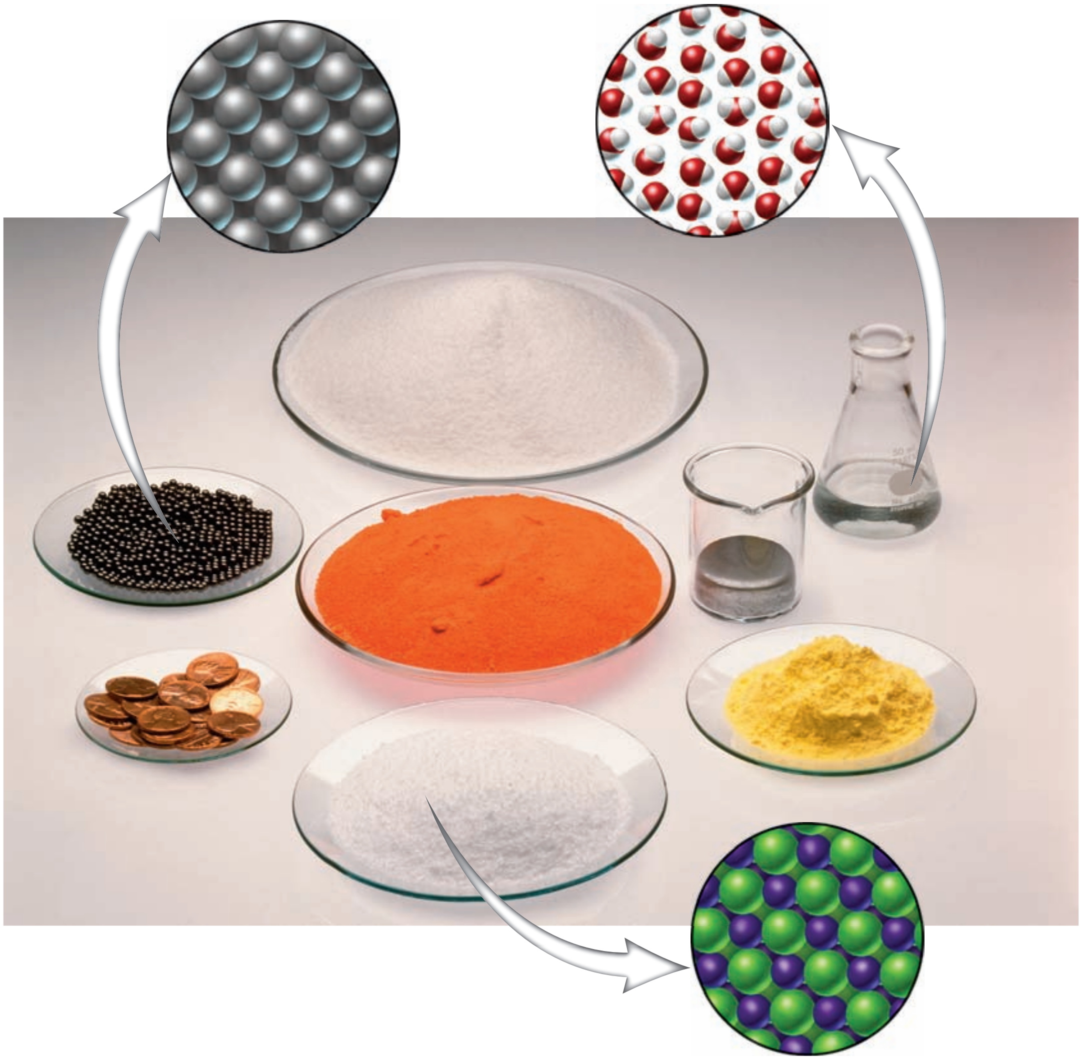
Challenge Exercises

79. Calculate the volume in cubic centimeters occupied by one molecule of water in a beaker of water. The density of water is 1.00 g/cm^3.

80. Calculate the volume in milliliters occupied by one molecule of ethyl alcohol, C_2H_5OH, in a beaker of ethyl alcohol. The density of ethyl alcohol is 0.789 g/mL.

81. The volume occupied by each copper atom in a 1-mole crystal is 0.0118 nm^3. If the density of the copper crystal is 8.92 g/cm^3, what is the experimental value of Avogadro's number?

82. Each atom in a crystal of aluminum metal occupies a theoretical cube that is 0.255 nm on a side. If the density of the aluminum crystal is 2.70 g/cm^3, what is the experimental value of Avogadro's number?



▲ **The Mole Concept** The photograph shows a mole of sugar, water, mercury, sulfur, salt, copper, lead (clockwise), and orange potassium dichromate (center). Each sample contains 6.02×10^{23} particles; however, the molar masses differ as follows: 342.34 g/mol $C_{12}H_{22}O_{11}$, 18.02 g/mol H_2O, 200.59 g/mol Hg, 32.07 g/mol S, 58.44 g/mol NaCl, 63.55 g/mol Cu, 207.2 g/mol Pb, and 294.20 g/mol $K_2Cr_2O_7$.

Chapter 8 Self-Test

Answers to Self-Test are in Appendix J.

1. How many atoms of silver equal a mass of 107.87 amu? (Sec. 8.1)
(a) 1
(b) 47
(c) 107.87
(d) 108
(e) 6.02×10^{23}

2. How many atoms of silver are in 2.50 mol of Ag metal? (Sec. 8.2)
(a) 1.51×10^{22} atoms
(b) 1.51×10^{23} atoms
(c) 1.51×10^{24} atoms
(d) 2.41×10^{23} atoms
(e) 2.41×10^{24} atoms

3. What is the molar mass of ammonium dichromate, $(NH_4)_2Cr_2O_7$? (Sec. 8.3)
(a) 238.09 g/mol
(b) 252.10 g/mol
(c) 260.18 g/mol
(d) 372.18 g/mol
(e) 386.19 g/mol

4. What is the mass of 3.30×10^{23} atoms of silver, Ag? (Sec. 8.4)
(a) 0.549 g
(b) 25.8 g
(c) 59.1 g
(d) 85.7 g
(e) 197 g

5. How many hydrogen molecules are in 22.4 liters of H_2 gas at STP? (Sec. 8.5)
(a) 1.20×10^{24}
(b) 1.35×10^{25}
(c) 1.81×10^{24}
(d) 2.69×10^{22}
(e) 6.02×10^{23}

6. What is the volume of 1.20×10^{22} molecules of hydrogen gas, H_2, at STP? (Sec. 8.6)
(a) 0.0199 L
(b) 0.447 L
(c) 2.24 L
(d) 5.02 L
(e) 1120 L

7. An emerald has the chemical formula $Be_3Al_2Si_{16}O_{18}$. What is the percentage of aluminum in the gemstone? (Sec. 8.7)
(a) 3.30%
(b) 3.37%
(c) 6.59%
(d) 35.2%
(e) 54.9%

8. If 0.500 mol of silver combines with 0.250 mol of sulfur, what is the empirical formula of the silver sulfide product? (Sec. 8.8)
(a) AgS
(b) Ag_2S
(c) AgS_2
(d) Ag_5S_5
(e) none of the above

9. A common name for fructose is *fruit sugar.* What is the molecular formula of fructose if the empirical formula is CH_2O, and the approximate molar mass is 180 g/mol? (Sec. 8.9)
(a) CHO
(b) CH_2O
(c) CH_2O_6
(d) $C_6H_{12}O_6$
(e) $C_{12}H_{22}O_{11}$

Key Concepts

10. Given that Al-27 is the only natural isotope of aluminum, what is the mass of one Al atom and Avogadro's number of Al atoms, respectively?

11. Given that I-127 is the only natural isotope of iodine, what is the mass of one I atom and Avogadro's number of I atoms, respectively?

12. Given Avogadro's number of shotput balls, which of the following is the best estimate of the total mass: an automobile, the Empire State building, or Earth?

13. Given Avogadro's number of marbles, which of the following is the best estimate of the total volume: a basketball, Dallas Cowboys Stadium, or the Moon?

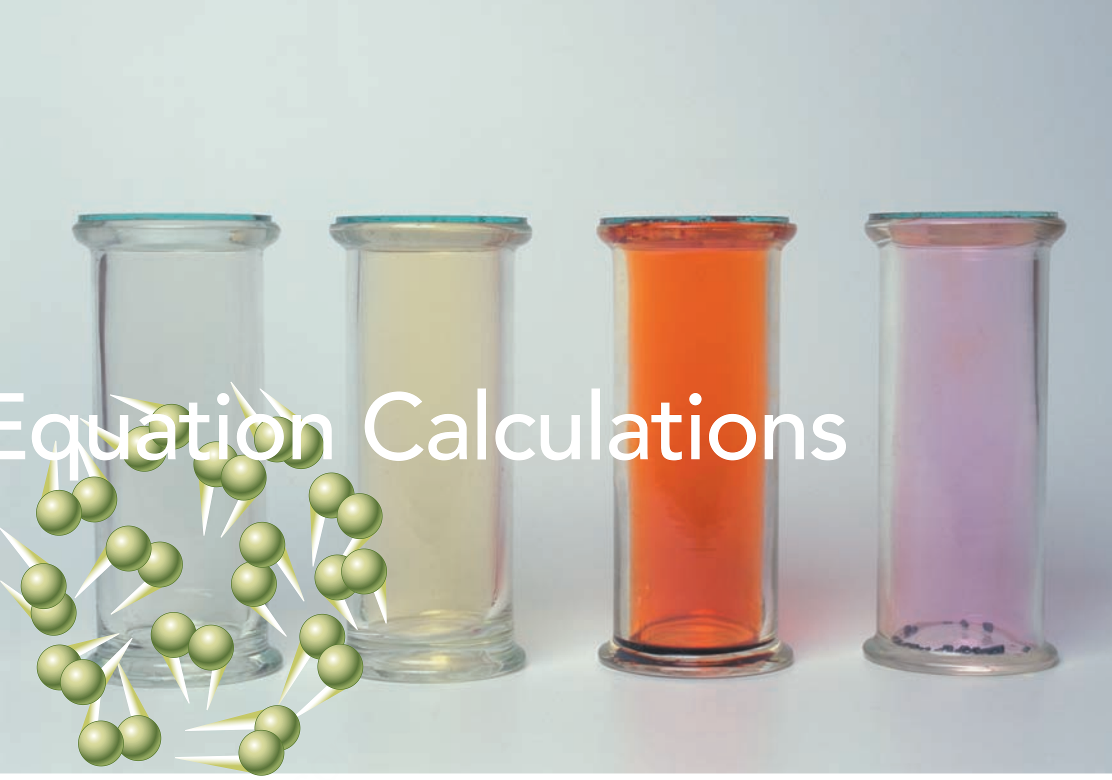
Critical Thinking

14. If an ultra-sensitive balance can weigh samples having a mass of 1 μg or greater, can it weigh 10 trillion copper atoms?

15. A 1-carat diamond is pure, crystalline carbon and has a mass of 0.200 g. Which has more atoms: a 1-carat diamond, or a googol (1×10^{100}) of atoms?

CHAPTER 9

Equation Calculations



"A single conversation with a wise teacher is worth a month's study of books."

– Chinese Proverb

9
F
fluorine

Element 9: Fluorine
Fluorine is a halogen, shown above with chlorine, bromine, and iodine vapor (left to right). Fluorine is a pale yellow gas that is extremely reactive and very dangerous. However, fluoride compounds such as Freon and Teflon are quite stable and have an important role in modern society. Fluorine was not isolated from its natural mineral form until 1886, and a French chemist received the 1906 Nobel Prize in chemistry for the feat.

The manufacturing of chemicals is one of the most important industries in the United States. Millions of tons of chemicals are produced each year for use both in the United States and in other countries. These chemicals are used to manufacture cars, computer chips, electronic instruments, prescription drugs, glass, paper, fertilizers, pesticides, plumbing supplies, synthetic fibers, and other products.

Chemists and chemical engineers routinely perform calculations based on balanced chemical equations to determine the cost of producing chemicals. These calculations are critical to many industries. The amounts of chemicals manufactured are published annually, and sulfuric acid, found in automobile batteries, consistently tops this list. The amount of sulfuric acid produced indicates the general activity of manufacturing and for that reason is referred to as the barometer of the chemical industry. However, sulfuric acid is an environmental problem that results from burning low-grade coal containing sulfur. Sulfur dioxide is released into the atmosphere and is converted to sulfuric acid. So-called "acid rain" contains sulfuric acid, which is a corrosive chemical that harms plants, birds animals, and aquatic life.

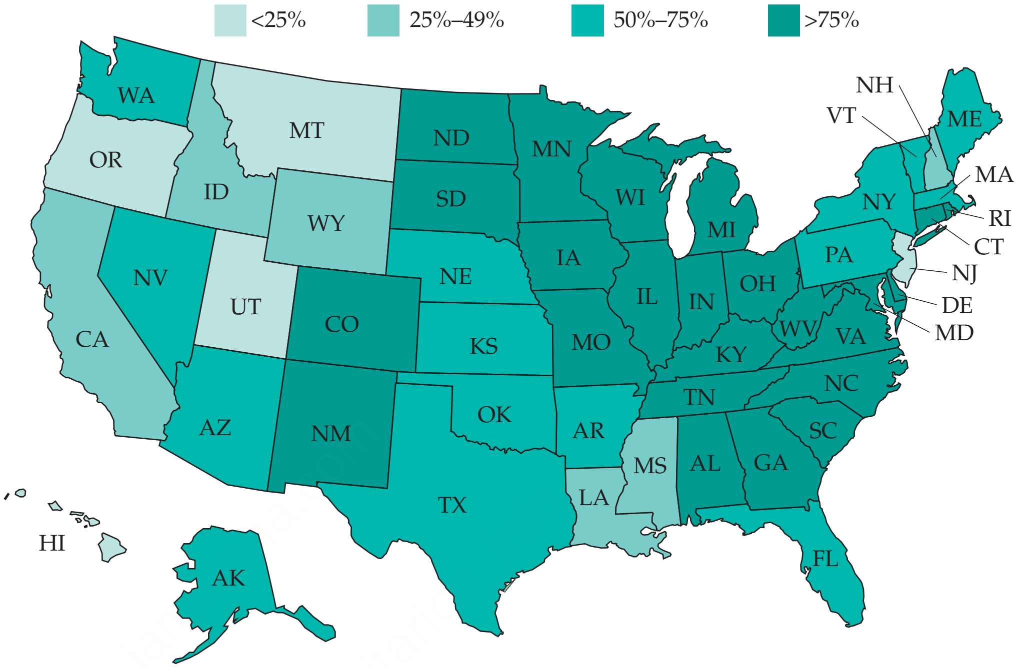
9.1 Interpreting a Chemical Equation

Let's examine the information we can obtain from a balanced chemical equation. Consider nitrogen monoxide gas, which is present in automobile emissions. In the atmosphere, ultraviolet light catalyzes the reaction of nitrogen monoxide and oxygen to produce reddish-brown nitrogen dioxide smog. The balanced chemical equation is:

$$2\,\mathrm{NO}(g) + \mathrm{O}_2(g) \xrightarrow{\mathrm{UV}} 2\,\mathrm{NO}_2(g)$$

A CLOSER LOOK Water Fluoridation

Q: How does sodium fluoride in the public water supply promote dental health?



▶ Fluoridated Water
The map of the United States indicates the availability of fluoridated water in municipalities throughout the nation.

The U.S. Public Health Service states that fluoridated water can help prevent dental cavities. The initial study in the 1950s showed that children who grew up drinking naturally fluoridated water had an average of three cavities, whereas children who grew up without fluoridated water had an average of 10 cavities.

The effect of fluoride is to make the enamel surface of teeth more resistant to decay. Tooth enamel is mainly hydroxyapatite, $Ca_{10}(PO_4)_6(OH)_2$, the hardest substance in the body. However, bacteria in the mouth can cause pitting in tooth enamel. Fluoride can prevent cavities by converting hydroxyapatite to $Ca_{10}(PO_4)_6F_2$, which is more resistant to bacteria and acids in the mouth.

Every state has drinking water with natural or controlled fluoridation. In 2000, 42 of the 50 largest cities in the United States had fluoridated drinking water, and two-thirds of Americans drink fluoridated water. However, the percentage of people who drink fluoridated water varies widely from state to state. In 2015, the Department of Health and Human Services announced a change in the fluoride level in public water; it now limits the concentration of fluoride to 0.7 milligrams per liter (0.7 ppm).

In areas where the public water supply is not fluoridated, people can take inexpensive fluoride tablets. In addition, many toothpastes and mouthwashes contain fluoride. For example, sodium fluoride and stannous fluoride have been added to many toothpastes. Another alternative is to have fluoride treatments in a dental office.

A: Drinking fluoridated water, especially while teeth are developing, helps prevent dental cavities by making the enamel surface of teeth harder.

The Mole Interpretation of Equation Coefficients

The coefficients in the balanced equation indicate that 2 molecules of NO react with 1 molecule of O_2 to produce 2 molecules of NO_2. These coefficients indicate relative numbers of reactant and product molecules. It therefore follows that multiples of these coefficients are in the same ratios. For example, 2000 molecules of NO react with 1000 molecules of O_2 to give 2000 molecules of NO_2. Let's consider a number even larger than 2000; let's consider Avogadro's number (symbol *N*) of molecules, that is, 6.02×10^{23} molecules. When we substitute this very large number of molecules for the coefficients, the equation remains balanced. That is,

$$2\,N\,\text{NO} + 1\,N\,\text{O}_2 \longrightarrow 2\,N\,\text{NO}_2$$

LEARNING OBJECTIVES

- Relate the coefficients in a balanced chemical equation to
  (a) moles of reactants and products, and
  (b) liters of gaseous reactants and products.
- Verify the coefficients of a balanced chemical equation using the conservation of mass law.

We read the preceding equation as *two times Avogadro's number of nitrogen monoxide molecules reacts with Avogadro's number of oxygen molecules to produce two times Avogadro's number of nitrogen dioxide molecules.* Because Avogadro's number is the number of molecules in 1 mol, we can write the equation in terms of moles of substance:

$$2 \text{ mol NO} + 1 \text{ mol } O_2 \longrightarrow 2 \text{ mol } NO_2$$

Similarly, this equation reads *two moles of nitrogen monoxide react with one mole of oxygen to give two moles of nitrogen dioxide.* Furthermore, the coefficients indicate the ratio of moles, or the **mole ratio**, of reactants and products in every balanced chemical equation.

▲ **Nitrogen Oxides** This photo shows a smoggy day in Los Angeles as a result of automobile emissions and the release of nitrogen oxides into the atmosphere.

The Volume Interpretation of Equation Coefficients

According to **Avogadro's theory**, there are equal numbers of molecules in equal volumes of gas at the same temperature and pressure. It therefore follows that twice the number of molecules occupies twice the volume of gas. In the preceding equation, 2 molecules of NO react with 1 molecule of O_2 to give 2 molecules of NO_2. Accordingly, 2 volumes of NO react with 1 volume of O_2 gas to give 2 volumes of NO_2 gas. If we choose the liter as our unit of volume, we can write the ratio of volumes of gas for the reaction.

$$2 \text{ L NO} + 1 \text{ L } O_2 \longrightarrow 2 \text{ L } NO_2$$

Regardless of the units, the numbers in the ratio of volumes correspond to the coefficients of the balanced chemical equation, that is, 2:1:2. For example, 2 mL of NO reacts with 1 mL of O_2 to give 2 mL of NO_2. Similarly, 20 mL of NO reacts with 10 mL of O_2 to give 20 mL of NO_2. Although the ratio of gaseous volumes is 20:10:20, the smallest whole-number ratio is 2:1:2. Table 9.1 summarizes the information that can be obtained from the coefficients of a balanced chemical equation.

TABLE 9.1 Interpretation of Chemical Equation Coefficients

For the General Equation	2 A	+	3 B	→	C	+	2 D
The ratio of molecules	2	:	3	:	1	:	2
The ratio of moles	2	:	3	:	1	:	2
The ratio of volumes of gas	2	:	3	:	1	:	2

The following example exercise illustrates the information that can be obtained by interpreting the coefficients in a balanced chemical equation:

EXAMPLE EXERCISE 9.1 Interpreting Chemical Equation Coefficients

Given the chemical equation for the combustion of methane, CH_4, balance the equation and interpret the coefficients in terms of (a) moles and (b) liters:

$$CH_4(g) + O_2(g) \xrightarrow{\text{spark}} CO_2(g) + H_2O(g)$$

Solution

The balanced chemical equation is

$$CH_4(g) + 2\,O_2(g) \xrightarrow{\text{spark}} CO_2(g) + 2\,H_2O(g)$$

(a) The coefficients in the equation (1:2:1:2) indicate the ratio of moles as well as molecules.

$$1 \text{ mol } CH_4 + 2 \text{ mol } O_2 \longrightarrow 1 \text{ mol } CO_2 + 2 \text{ mol } H_2O$$

(b) The coefficients in the equation (1:2:1:2) indicate the ratio of volumes of gases. If we express the volume in liters, we have

$$1 \text{ L } CH_4 + 2 \text{ L } O_2 \longrightarrow 1 \text{ L } CO_2 + 2 \text{ L } H_2O$$

Practice Exercise

Given the chemical equation for the combustion of propane, C_3H_8, balance the equation and interpret the coefficients in terms of (a) moles and (b) milliliters:

$$C_3H_8(g) + O_2(g) \xrightarrow{\text{spark}} CO_2(g) + H_2O(g)$$

Answers:

(a) $1\text{ mol } C_3H_8 + 5\text{ mol } O_2 \longrightarrow 3\text{ mol } CO_2 + 4\text{ mol } H_2O$

(b) $1\text{ mL } C_3H_8 + 5\text{ mL } O_2 \longrightarrow 3\text{ mL } CO_2 + 4\text{ mL } H_2O$

Concept Exercise

Which of the following is in the same ratio as the coefficients in a balanced equation: moles of gas, mass of gas, volume of gas?

Answer: See Appendix G, 9.1.

Verifying the Conservation of Mass Law

In 1787 after conducting numerous experiments, Antoine Lavoisier formulated the **law of conservation of mass**. He stated that mass is neither created nor destroyed during a chemical reaction. As a result, the combined masses of the reactants must equal the combined masses of the products.

We can verify the conservation of mass law for any balanced equation. Let's consider the previous equation for the formation of nitrogen dioxide:

$$2\,NO(g) + O_2(g) \xrightarrow{\text{UV}} 2\,NO_2(g)$$

According to the coefficients in the balanced equation, the ratio of the substances is 2:1:2. Therefore, the mole ratio of each substance is 2:1:2.

$$2\text{ mol NO} + 1\text{ mol } O_2 \longrightarrow 2\text{ mol } NO_2$$

By summing the atomic masses for each substance, we find that the **molar mass** for NO is 30.01 g/mol; for O_2, 32.00 g/mol; and for NO_2, 46.01 g/mol. Substituting the molar mass for a mole of each substance gives

$$2\,(30.01\text{ g}) + 1\,(32.00\text{ g}) \longrightarrow 2\,(46.01\text{ g})$$

$$\text{Simplifying, } 60.02\text{ g} + 32.00\text{ g} \longrightarrow 92.02\text{ g}$$

$$\text{and, } 92.02\text{ g} \longrightarrow 92.02\text{ g}$$

In this example, the total mass of the reactants is 92.02 g, which is equal to the total mass of the products. Thus from our calculation, we have verified the conservation of mass law.

9.2 Mole–Mole Relationships

LEARNING OBJECTIVE

- Relate the number of moles of two substances in a balanced chemical equation.

As we demonstrated in Section 9.1, the coefficients in a chemical equation indicate the mole ratio of the reactants and products. Consider the combination reaction of nitrogen and oxygen to give nitrogen monoxide. The equation is

$$N_2(g) + O_2(g) \xrightarrow{\Delta} 2\,NO(g)$$

From the coefficients in the balanced equation, we see that 1 mol of nitrogen reacts with 1 mol of oxygen to produce 2 mol of nitrogen monoxide. We can write several mole ratios for this equation.

Because 1 mol of N_2 reacts with 1 mol of O_2, the corresponding mole ratios are

$$\frac{1\text{ mol } N_2}{1\text{ mol } O_2} \quad \text{and} \quad \frac{1\text{ mol } O_2}{1\text{ mol } N_2}$$

Because 1 mol of N_2 produces 2 mol of NO, the corresponding mole ratios are

$$\frac{1\ \text{mol}\ N_2}{2\ \text{mol NO}} \quad \text{and} \quad \frac{2\ \text{mol NO}}{1\ \text{mol}\ N_2}$$

Because 1 mol of O_2 produces 2 mol of NO, the corresponding mole ratios are

$$\frac{1\ \text{mol}\ O_2}{2\ \text{mol NO}} \quad \text{and} \quad \frac{2\ \text{mol NO}}{1\ \text{mol}\ O_2}$$

Suppose we want to find out how many moles of oxygen react with 2.25 mol of nitrogen. As usual, we apply the unit analysis method of problem solving:

$$2.25\ \text{mol}\ N_2 \times \frac{\text{unit}}{\text{factor}} = \text{mol}\ O_2$$

To cancel units, we select the mole ratio 1 mol O_2/1 mol N_2 as our unit factor:

$$2.25\ \cancel{\text{mol}\ N_2} \times \frac{1\ \text{mol}\ O_2}{1\ \cancel{\text{mol}\ N_2}} = 2.25\ \text{mol}\ O_2$$

We can also calculate the moles of nitrogen monoxide produced by the reaction. To cancel units, we select the mole ratio 2 mol NO/1 mol N_2 as our unit factor:

$$2.25\ \cancel{\text{mol}\ N_2} \times \frac{2\ \text{mol NO}}{1\ \cancel{\text{mol}\ N_2}} = 4.50\ \text{mol NO}$$

Whenever we have a balanced chemical equation, we can always convert from moles of one substance to moles of another substance using a mole ratio as a unit factor. Example Exercise 9.2 further illustrates the relationship between moles of reactants and moles of products.

EXAMPLE EXERCISE 9.2 Mole–Mole Relationships

Carbon monoxide is produced in a blast furnace by passing oxygen gas over hot coal. How many moles of oxygen react with 2.50 mol of carbon, according to the balanced equation?

$$2\ C(s) + O_2(g) \xrightarrow{\Delta} 2\ CO(g)$$

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**mol O_2**
STEP 2:	What given value is related to the answer?	**Step 2:**	**2.50 mol C**
STEP 3:	What unit factor(s) should we apply? Using the balanced equation, we see that 2 mol C = 1 mol O_2. Thus, the mole ratio and the two unit factors are	**Step 3:**	$\frac{2\ \text{mol C}}{1\ \text{mol}\ O_2}$ **or** $\frac{1\ \text{mol}\ O_2}{2\ \text{mol C}}$

Unit Analysis Map

$$\underset{\textit{given value}}{2.50\ \text{mol C}} \quad \underset{\textit{unit factor}}{\times \frac{1\ \text{mol}\ O_2}{2\ \text{mol C}}} \quad \underset{\textit{unit in answer}}{= ?\ \text{mol}\ O_2}$$

Solution

We select the unit factor that cancels the unit in the given value (~~mol C~~). Thus,

$$2.50\ \cancel{\text{mol C}} \times \frac{1\ \text{mol}\ O_2}{2\ \cancel{\text{mol C}}} = 1.25\ \text{mol}\ O_2$$

Practice Exercise

Iron is produced from iron ore in a blast furnace by passing carbon monoxide gas through molten iron(III) oxide. The balanced equation is

$$Fe_2O_3(l) + 3\,CO(g) \xrightarrow{\Delta} 2\,Fe(l) + 3\,CO_2(g)$$

(a) How many moles of carbon monoxide react with 2.50 mol of Fe_2O_3?
(b) How many moles of iron are produced from 2.50 mol of Fe_2O_3?

Answers:

(a) 7.50 mol CO
(b) 5.00 mol Fe

Concept Exercise

How many unit factors are required to solve a mole–mole problem?

Answer: See Appendix G, 9.2.

A CLOSER LOOK Iron versus Steel

Q: What is the difference between iron and steel?

The Iron Age began about 1000 B.C. in western Asia and Europe. The Iron Age followed the Bronze Age after it was discovered that iron could be made by heating iron ore with coal. Interestingly, iron is still obtained by a similar method.

Today, an industrial blast furnace converts iron ore to iron. In the process, iron ore, coke (a form of coal), and limestone, $CaCO_3$, are added into the top of a heated furnace while a blast of hot air is blown in at the bottom.

Many reactions take place in the process, but first, oxygen in the hot air reacts with carbon in coke to produce carbon monoxide:

$$2\,C(s) + O_2(g) \xrightarrow{\Delta} 2\,CO(g)$$

▲ **Blast Furnace** Molten pig iron reacts with a "blast" of oxygen gas that converts carbon and sulfur impurities into CO_2 and SO_2 gases.

Iron ore is an impure mixture of hematite, Fe_2O_3, and magnetite, Fe_3O_4. In the blast furnace hematite and magnetite are converted to iron by reaction with carbon monoxide gas. The reactions are as follows:

$$Fe_2O_3(l) + 3\,CO(g) \longrightarrow 2\,Fe(l) + 3\,CO_2(g)$$
$$Fe_3O_4(l) + 4\,CO(g) \longrightarrow 3\,Fe(l) + 4\,CO_2(g)$$

The iron ore also contains silicon impurities that must be removed. This is done by reacting the impurities with CaO, which is produced from limestone, that is, $CaCO_3$. Silicon compounds such as $CaSiO_3$ form and are referred to as "slag." This less dense slag is easily removed, as it floats on the molten iron at the bottom of the furnace.

The molten iron obtained from the blast furnace is still not pure and is referred to as "pig iron." Pig iron contains impurities such as carbon and sulfur. In a second stage of the process, these impurities are removed by blowing oxygen gas through the molten pig iron. The impurities are converted to oxides, and are released as gases such as CO_2 and SO_2.

The iron obtained after the oxygen process contains a trace amount of carbon. Manganese is then often added for strength and flexibility, and the resulting product is referred to as "steel." A large number of alloy steels can be produced by adding other metals. For example, stainless steel contains about 10% chromium.

A: Iron is an element, whereas steel is an alloy of iron with traces of carbon and metals, such as manganese, nickel, and chromium, which impart special properties.

9.3 Types of Stoichiometry Problems

LEARNING OBJECTIVE

- Classify the three basic types of stoichiometry problems: mass–mass, mass–volume, and volume–volume.

In Chapter 8, we performed mole calculations based on chemical formulas. In this chapter, we will perform mole calculations based on chemical equations. That is, we will apply mole ratios to relate quantities of reactants and products. The term **stoichiometry** (stoy-kee-AHM-uh-tree) is used to refer to the relationship between quantities in a chemical reaction according to the balanced chemical equation. The term is derived from the Greek words *stoicheion*, meaning "element," and *metron*, meaning "measure."

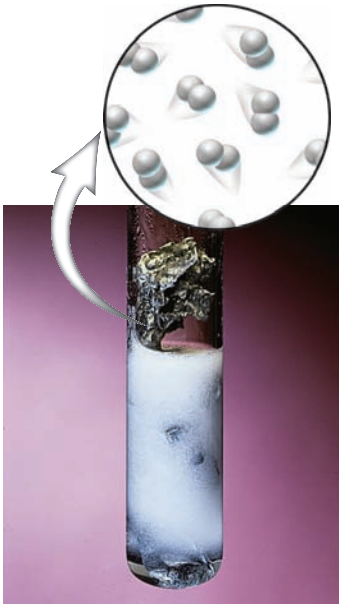
Some stoichiometry calculations relate a given amount of reactant to an unknown amount of product. Conversely, other stoichiometry calculations relate a given amount of product to an unknown amount of reactant. Whether we relate a reactant to a product, or vice versa, the calculations use the same basic method of problem solving.

We can classify three types of stoichiometry problems based on the given quantity and the quantity to be calculated. Consider the general equation for any reaction, where a, b, c, and d are the respective coefficients for the substances A, B, C, and D:

$$a\,\mathrm{A} + b\,\mathrm{B} \longrightarrow c\,\mathrm{C} + d\,\mathrm{D}$$

If we are given grams of A and asked to find how many grams of C are produced, we classify this as a **mass–mass problem**. If we are given grams of A and asked to calculate the milliliters of a gas D, we classify this as a **mass–volume problem**. If we are given liters of a gas A and asked to find the liters of a gas D, we classify this as a **volume–volume problem**. In summary, stoichiometry problems can be classified as one of these basic types: mass–mass, mass–volume, or volume–volume.

After classifying a stoichiometry problem, we can apply the unit analysis method of problem solving to its solution. Example Exercise 9.3 offers practice in classifying types of stoichiometry problems.



▲ **Reaction of Zn Metal** Zinc metal reacts with hydrochloric acid to give ZnCl and H_2 bubbles.

Helpful Hint
Stoichiometry

In Chapter 2 we introduced the unit analysis method of problem solving. Before attempting any stoichiometry problem, it is necessary to write a balanced chemical equation. The coefficients in the balanced equation provide the mole ratio of reactants and products.

EXAMPLE EXERCISE 9.3 Classifying Stoichiometry Problems

Classify the type of stoichiometry problem for each of the following:

(a) How many grams of Zn metal react with hydrochloric acid to give 0.500 g of zinc chloride?

(b) How many liters of H_2 gas react with chlorine gas to yield 50.0 cm^3 of hydrogen chloride gas?

(c) How many kilograms of Fe react with sulfuric acid to produce 50.0 mL of hydrogen gas?

Solution

After analyzing a problem for the unknown quantity and the relevant given value, we classify the type of problem.

(a) The problem asks for grams of Zn (mass) that react to give 0.500 g of $ZnCl_2$ (mass). This a *mass–mass* type of problem.

(b) The problem asks for liters of H_2 gas (volume) that react to yield 50.0 cm^3 of HCl gas (volume). This is a *volume–volume* type of problem.

(c) The problem asks for kilograms of Fe (mass) that react to produce 50.0 mL of H_2 gas (volume). This is a *mass–volume* type of problem.

Practice Exercise

Classify the type of stoichiometry problem for each of the following:

(a) How many grams of HgO decompose to give 0.500 L of oxygen gas at STP?

(b) How many grams of AgCl are produced from the reaction of 0.500 g of solid sodium chloride with silver nitrate solution?

(c) How many milliliters of H_2 gas react with nitrogen gas to yield 1.00 L of ammonia gas?

Answers:

(a) mass–volume
(b) mass–mass
(c) volume–volume

Concept Exercise

What are the three types of stoichiometry problems?

Answers: See Appendix G, 9.3.

9.4 Mass–Mass Problems

LEARNING OBJECTIVE

- Solve mass–mass stoichiometry problems.

A mass–mass stoichiometry problem is so named because an *unknown mass* of substance is calculated from a *given mass* of reactant or product in a chemical equation. After balancing the chemical equation, we proceed as follows:

(a) Convert the given mass of substance to moles using the molar mass of the substance as a unit factor.

(b) Convert the moles of the given to moles of the unknown using the mole ratio from the coefficients in the balanced equation.

(c) Convert the moles of the unknown to grams using the molar mass of the substance as a unit factor.

To see how to solve a mass–mass problem, consider the high-temperature conversion of 14.4 g of iron(II) oxide to elemental iron with aluminum metal. The balanced equation for the reaction is

$$3\,\mathrm{FeO}(l) + 2\,\mathrm{Al}(l) \xrightarrow{\Delta} 3\,\mathrm{Fe}(l) + \mathrm{Al_2O_3}(l)$$

To calculate the mass of aluminum necessary for the reaction, we must first find the moles of iron(II) oxide. Given that the molar mass of FeO is 71.85 g/mol, the unit factor conversion is as follows:

(a) $$14.4\ \cancel{\mathrm{g\ FeO}} \times \frac{1\ \mathrm{mol\ FeO}}{71.85\ \cancel{\mathrm{g\ FeO}}} = 0.200\ \mathrm{mol\ FeO}$$

Next, we convert the moles of FeO to moles of Al by applying a mole ratio using the coefficients in the balanced equation. Because 3 mol FeO = 2 mol Al, the mole ratio unit factor is

(b) $$0.200\ \cancel{\mathrm{mol\ FeO}} \times \frac{2\ \mathrm{mol\ Al}}{3\ \cancel{\mathrm{mol\ FeO}}} = 0.134\ \mathrm{mol\ Al}$$

Last, we use the molar mass of aluminum (26.98 g/mol) as a unit factor to obtain the mass of Al reacting with 14.4 g of FeO.

(c) $$0.134\ \cancel{\mathrm{mol\ Al}} \times \frac{26.98\ \mathrm{g\ Al}}{1\ \cancel{\mathrm{mol\ Al}}} = 3.60\ \mathrm{g\ Al}$$

After you gain confidence in solving stoichiometry problems, it will be more convenient to perform one continuous calculation. For example, the solution to the preceding problem can be shown as

$$14.4\ \cancel{\mathrm{g\ FeO}} \times \underbrace{\frac{1\ \cancel{\mathrm{mol\ FeO}}}{71.85\ \cancel{\mathrm{g\ FeO}}}}_{(a)} \times \underbrace{\frac{2\ \cancel{\mathrm{mol\ Al}}}{3\ \cancel{\mathrm{mol\ FeO}}}}_{(b)} \times \underbrace{\frac{26.98\ \mathrm{g\ Al}}{1\ \cancel{\mathrm{mol\ Al}}}}_{(c)} = 3.60\ \mathrm{g\ Al}$$

The following example exercises give additional illustrations of mass–mass stoichiometry calculations using unit factors.

▲ **Decomposition of HgO**
Orange mercury(II) oxide decomposes by heating to give droplets of silver Hg liquid and O_2 gas.

EXAMPLE EXERCISE 9.4 Mass–Mass Stoichiometry

Calculate the mass of mercury produced from the decomposition of 1.25 g of orange mercury(II) oxide (MM = 216.59 g/mol):

$$2\,\mathrm{HgO}(s) \xrightarrow{\Delta} 2\,\mathrm{Hg}(l) + \mathrm{O_2}(g)$$

Strategy Plan

STEP 1: What unit is asked for in the answer? **Step 1:** g Hg

STEP 2: What given value is related to the answer? **Step 2:** 1.25 g HgO

STEP 3: What unit factor(s) should we apply? We are given that 1 mol HgO = 216.59 g HgO; the mole ratio from the balanced equation is 2 mol HgO = 2 mol Hg. From the periodic table, we find that 1 mol Hg = 200.59 g Hg. Thus, the unit factors are

Step 3:

$$\frac{\textbf{1 mol HgO}}{\textbf{216.59 g HgO}} \quad \textbf{or} \quad \frac{\textbf{216.59 g HgO}}{\textbf{1 mol HgO}}$$

$$\frac{\textbf{2 mol HgO}}{\textbf{2 mol Hg}} \quad \textbf{or} \quad \frac{\textbf{2 mol Hg}}{\textbf{2 mol HgO}}$$

$$\frac{\textbf{1 mol Hg}}{\textbf{200.59 g Hg}} \quad \textbf{or} \quad \frac{\textbf{200.59 g Hg}}{\textbf{1 mol Hg}}$$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit factor 3*	*unit in answer*
1.25 g HgO	$\times \ \frac{\textbf{1 mol HgO}}{\textbf{216.59 g HgO}}$	$\times \ \frac{\textbf{2 mol Hg}}{\textbf{2 mol HgO}}$	$\times \ \frac{\textbf{200.59 g Hg}}{\textbf{1 mol Hg}}$	**= ? g Hg**

Solution

We select from each of the three pairs of ratios a unit factor that cancels the previous units. Thus,

$$1.25\ \cancel{\text{g HgO}} \times \frac{1\ \cancel{\text{mol HgO}}}{216.59\ \cancel{\text{g HgO}}} \times \frac{2\ \cancel{\text{mol Hg}}}{2\ \cancel{\text{mol HgO}}} \times \frac{200.59\ \text{g Hg}}{1\ \cancel{\text{mol Hg}}} = 1.16\ \text{g Hg}$$

Practice Exercise

Calculate the mass of carbon dioxide released from 10.0 g of cobalt(III) carbonate given the *unbalanced* equation for the reaction:

$$Co_2(CO_3)_3(s) \xrightarrow{\Delta} Co_2O_3(s) + CO_2(g)$$

Answer: 4.43 g CO_2

Concept Exercise

What are the three steps in the unit analysis method of problem solving?

Answer: See Appendix G, 9.4.

Precipitation of HgI_2 Potassium iodide and mercury(II) nitrate react to give an orange precipitate of HgI_2.

EXAMPLE EXERCISE 9.5 Mass–Mass Stoichiometry

Calculate the mass of potassium iodide (166.00 g/mol) required to yield 1.78 g of mercury(II) iodide precipitate (454.39 g/mol):

$$2\ KI(s) + Hg(NO_3)_2(aq) \longrightarrow HgI_2(s) + 2\ KNO_3(aq)$$

Strategy Plan

STEP 1: What unit is asked for in the answer? **Step 1:** **g KI**

STEP 2: What given value is related to the answer? **Step 2:** **1.78 g HgI_2**

STEP 3: What unit factor(s) should we apply? From the given molar mass, we see that 1 mol HgI_2 = 454.39 g HgI_2. The mole ratio from the balanced equation is 2 mol KI = 1 mol HgI_2, and from the given molar mass, 1 mol KI = 166.00 g KI. Thus, the three pairs of unit factors are

Step 3:

$$\frac{\textbf{1 mol HgI}_2}{\textbf{454.39 g HgI}_2} \quad \textbf{or} \quad \frac{\textbf{454.39 g HgI}_2}{\textbf{1 mol HgI}_2}$$

$$\frac{\textbf{2 mol KI}}{\textbf{1 mol HgI}_2} \quad \textbf{or} \quad \frac{\textbf{1 mol HgI}_2}{\textbf{2 mol KI}}$$

$$\frac{\textbf{1 mol KI}}{\textbf{166.00 g KI}} \quad \textbf{or} \quad \frac{\textbf{166.00 g KI}}{\textbf{1 mol KI}}$$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit factor 3*	*unit in answer*
$1.78\ \text{g HgI}_2$	$\times\ \frac{1\ \text{mol HgI}_2}{454.39\ \text{g HgI}_2}$	$\times\ \frac{2\ \text{mol KI}}{1\ \text{mol HgI}_2}$	$\times\ \frac{166.00\ \text{g KI}}{1\ \text{mol KI}}$	$=\ ?\ \text{g KI}$

Solution

We select from each of the three pairs of ratios a unit factor that cancels the previous units. Thus,

$$1.78\ \cancel{\text{g HgI}_2} \times \frac{1\ \cancel{\text{mol HgI}_2}}{454.39\ \cancel{\text{g HgI}_2}} \times \frac{2\ \cancel{\text{mol KI}}}{1\ \cancel{\text{mol HgI}_2}} \times \frac{166.00\ \text{g KI}}{1\ \cancel{\text{mol KI}}} = 1.30\ \text{g KI}$$

Practice Exercise

Calculate the mass of iron filings required to produce 0.455 g of silver metal given the *unbalanced* equation for the reaction:

$$\text{Fe}(s) + \text{AgNO}_3(aq) \longrightarrow \text{Fe(NO}_3)_3(aq) + \text{Ag}(s)$$

Answer: 0.0785 g Fe

Concept Exercise

In general, how many unit factors are required to solve a mass–mass stoichiometry problem?

Answer: See Appendix G, 9.5.

9.5 Mass–Volume Problems

LEARNING OBJECTIVE

- Solve mass–volume stoichiometry problems.

A mass–volume stoichiometry problem is so named because an *unknown volume* of substance is calculated from a *given mass* of reactant or product in a chemical equation. After balancing the chemical equation, we proceed as follows:

(a) Convert the given mass of substance to moles using the molar mass of the substance as a unit factor.

(b) Convert the moles of the given to moles of the unknown using the coefficients in the balanced equation.

(c) Convert the moles of the unknown to liters using the **molar volume** of a gas as a unit factor. At **standard temperature and pressure (STP)**, the molar volume of a gas is 22.4 L/mol.

Of course, it is also possible to perform the reverse calculation; that is, we can find the mass of an unknown substance from a given volume of gas. For simplicity, this reverse process is also referred to as a mass–volume stoichiometry problem.

To understand how to solve a mass–volume stoichiometry problem, consider the reaction of 0.165 g of aluminum metal with dilute hydrochloric acid. The balanced equation for the reaction is

$$2\ \text{Al}(s) + 6\ \text{HCl}(aq) \longrightarrow 2\ \text{AlCl}_3(aq) + 3\ \text{H}_2(g)$$

Let's calculate the volume of hydrogen gas produced by the reaction at STP. This is a mass–volume stoichiometry problem. First, we calculate the moles of aluminum whose molar mass is 26.98 g/mol:

(a)
$$0.165\ \cancel{\text{g Al}} \times \frac{1\ \text{mol Al}}{26.98\ \cancel{\text{g Al}}} = 0.00611\ \text{mol Al}$$

Next, we use the coefficients in the balanced equation to find the moles of hydrogen gas. According to the balanced equation, 2 mol Al = 3 mol H_2. Applying the mole ratio unit factor, we have

(b)
$$0.00611\ \cancel{\text{mol Al}} \times \frac{3\ \text{mol H}_2}{2\ \cancel{\text{mol Al}}} = 0.00917\ \text{mol H}_2$$

Last, we multiply by the molar volume, 22.4 L/ mol, to obtain the volume of H_2 gas produced at STP:

(c)
$$0.00917\ \cancel{\text{mol H}_2} \times \frac{22.4\ \text{L H}_2}{1\ \cancel{\text{mol H}_2}} = 0.205\ \text{L H}_2$$

The solution to the preceding problem can also be shown as one continuous calculation:

$$0.165\ \cancel{\text{g Al}} \times \underset{(a)}{\frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}}} \times \underset{(b)}{\frac{3\ \cancel{\text{mol H}_2}}{2\ \cancel{\text{mol Al}}}} \times \underset{(c)}{\frac{22.4\ \text{L H}_2}{1\ \cancel{\text{mol H}_2}}} = 0.205\ \text{L H}_2$$

EXAMPLE EXERCISE 9.6 Mass–Volume Stoichiometry

In an automobile collision, sodium azide, NaN_3, decomposes and fills an air bag with nitrogen gas. If an air bag contains 100.0 g of NaN_3(65.02 g/mol), what is the volume of nitrogen gas produced at STP? The equation for the reaction is

$$2\ \text{NaN}_3(s) \xrightarrow{\text{spark}} 2\ \text{Na}(s) + 3\ \text{N}_2(g)$$

Strategy Plan

STEP 1: What unit is asked for in the answer?	**Step 1:**	**L N_2**		
STEP 2: What given value is related to the answer?	**Step 2:**	**100.0 g NaN_3**		
STEP 3: What unit factor(s) should we apply? Given the molar mass, 1 mol NaN_3 = 65.02 g NaN_3. The mole ratio from the balanced equation is 2 mol NaN_3 = 3 mol N_2, and the molar volume of a gas is 1 mol N_2 = 22.4 L N_2. Thus, the unit factors are	**Step 3:**	$\dfrac{1\ \text{mol NaN}_3}{65.02\ \text{g NaN}_3}$	**or**	$\dfrac{65.02\ \text{g NaN}_3}{1\ \text{mol NaN}_3}$
		$\dfrac{2\ \text{mol NaN}_3}{3\ \text{mol N}_2}$	**or**	$\dfrac{3\ \text{mol N}_2}{2\ \text{mol NaN}_3}$
		$\dfrac{1\ \text{mol N}_2}{22.4\ \text{L N}_2}$	**or**	$\dfrac{22.4\ \text{L N}_2}{1\ \text{mol N}_2}$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit factor 3*	*unit in answer*
100.0 g NaN_3	$\times \dfrac{1\ \text{mol NaN}_3}{65.02\ \text{g NaN}_3}$	$\times \dfrac{3\ \text{mol N}_2}{2\ \text{mol NaN}_3}$	$\times \dfrac{22.4\ \text{L N}_2}{1\ \text{mol N}_2}$	= ? L N_2

Solution

We select from each of the three pairs of ratios a unit factor that cancels the previous units. Thus,

$$100.0\ \cancel{\text{g NaN}_3} \times \frac{1\ \cancel{\text{mol NaN}_3}}{65.02\ \cancel{\text{g NaN}_3}} \times \frac{3\ \cancel{\text{mol N}_2}}{2\ \cancel{\text{mol NaN}_3}} \times \frac{22.4\ \text{L N}_2}{1\ \cancel{\text{mol N}_2}} = 51.7\ \text{L N}_2$$

Practice Exercise

Calculate the volume of hydrogen gas produced at STP from 1.55 g of sodium metal in water, given the *unbalanced* equation for the reaction:

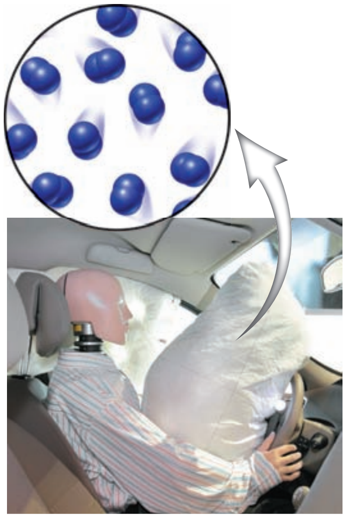
$$\text{Na}(s) + \text{H}_2\text{O}(l) \longrightarrow \text{NaOH}(aq) + \text{H}_2(g)$$

Answer: 0.755 L H_2

Concept Exercise

What are the three steps in the unit analysis method of problem solving?

Answer: See Appendix G, 9.6.



▲ **Automobile Air Bag** The decomposition of sodium azide releases N_2 gas, which inflates the air bag.

EXAMPLE EXERCISE 9.7 Mass–Volume Stoichiometry

Baking soda can be used as a fire extinguisher. When heated, it decomposes to carbon dioxide gas, which can smother a fire. If a sample of $NaHCO_3$ (84.01 g/mol) produces 0.500 L of carbon dioxide gas at STP, what is the mass of the sample?

$$2\ \text{NaHCO}_3(s) \xrightarrow{\Delta} \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(g) + \text{CO}_2(g)$$

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: g $NaHCO_3$

STEP 2: What given value is related to the answer?

Step 2: 0.500 L CO_2

STEP 3: What unit factor(s) should we apply? From molar volume, we know that 1 mol CO_2 = 22.4 L CO_2. The mole ratio from the balanced equation is 2 mol $NaHCO_3$ = 1 mol CO_2, and from the given molar mass, 1 mol $NaHCO_3$ = 84.01 g $NaHCO_3$. Thus, the unit factors are

Step 3:

$$\frac{1\ \text{mol CO}_2}{22.4\ \text{L CO}_2} \quad \text{or} \quad \frac{22.4\ \text{L CO}_2}{1\ \text{mol CO}_2}$$

$$\frac{2\ \text{mol NaHCO}_3}{1\ \text{mol CO}_2} \quad \text{or} \quad \frac{1\ \text{mol CO}_2}{2\ \text{mol NaHCO}_3}$$

$$\frac{1\ \text{mol NaHCO}_3}{84.01\ \text{g NaHCO}_3} \quad \text{or} \quad \frac{84.01\ \text{g NaHCO}_3}{1\ \text{mol NaHCO}_3}$$

Unit Analysis Map

given value	*unit factor 1*	*unit factor 2*	*unit factor 3*	*unit in answer*
0.500 L CO_2	$\times \dfrac{1\ \text{mol CO}_2}{22.4\ \text{L CO}_2}$	$\times \dfrac{2\ \text{mol NaHCO}_3}{1\ \text{mol CO}_2}$	$\times \dfrac{84.01\ \text{g NaHCO}_3}{1\ \text{mol NaHCO}_3}$	= ? g $NaHCO_3$

Solution

We select from each of the three pairs of ratios a unit factor that cancels the previous units. Thus,

$$0.500\ \cancel{\text{L CO}_2} \times \frac{1\ \cancel{\text{mol CO}_2}}{22.4\ \cancel{\text{L CO}_2}} \times \frac{2\ \cancel{\text{mol NaHCO}_3}}{1\ \cancel{\text{mol CO}_2}} \times \frac{84.01\text{ g NaHCO}_3}{1\ \cancel{\text{mol NaHCO}_3}} = 3.75\text{ g NaHCO}_3$$

Practice Exercise

Calculate the mass of aluminum metal required to release 2160 mL of hydrogen gas at STP from sulfuric acid given the *unbalanced* equation for the reaction:

$$\text{Al}(s) + \text{H}_2\text{SO}_4(aq) \longrightarrow \text{Al}_2(\text{SO}_4)_3(aq) + \text{H}_2(g)$$

Answer: 1.73 g Al

Concept Exercise

In general, how many unit factors are required to solve a mass–volume stoichiometry problem?

Answer: See Appendix G, 9.7.

▲ Baking Soda, $NaHCO_3$
Ordinary baking soda can be used to extinguish a fire. When heated, it decomposes and releases carbon dioxide gas, which is more dense than air and displaces oxygen from the flame.

LEARNING OBJECTIVE

- Solve volume–volume stoichiometry problems.

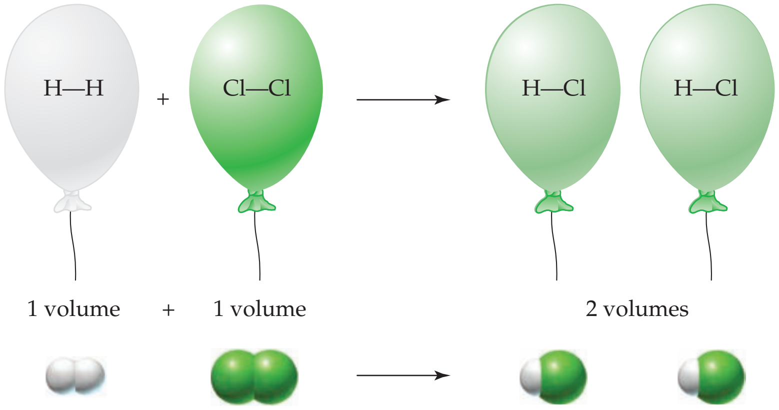
9.6 Volume–Volume Problems

In 1804, the French chemist Joseph Gay-Lussac ascended to 23,000 feet in a hydrogen-filled balloon. During his record-setting ascent, he took samples of the atmosphere and later studied their composition.

In 1808, Gay-Lussac proposed the **law of combining volumes**. This law states that volumes of gases, under similar conditions, combine in small whole-number ratios. For example, 10.0 mL of hydrogen gas combines with 10.0 mL of chlorine gas to give 20.0 mL of hydrogen chloride gas:

$$\underset{10.0\ mL}{\text{H}_2(g)} + \underset{10.0\ mL}{\text{Cl}_2(g)} \longrightarrow \underset{20.0\ mL}{2\ \text{HCl}(g)}$$

Notice that the volumes of gas (10.0 mL:10.0 mL:20.0 mL) are in the ratio of small whole numbers (1:1:2). Moreover, the ratio of the volumes is identical to the mole ratio in the balanced chemical equation (1:1:2).



The law of combining volumes states that gases combine in small whole-number ratios. The volume ratio is identical to the coefficients in the balanced equation, so we can convert from a given volume to an unknown volume of gas in a single conversion using the mole ratio. A volume–volume stoichiometry problem is so named because an *unknown volume* of gas is calculated from a *given volume* of gaseous reactant or product in a chemical equation.

We can illustrate a volume–volume problem using the industrial method for preparing sulfuric acid. Sulfuric acid is manufactured by the Contact process, which begins with the conversion of sulfur dioxide to sulfur trioxide using heat and a catalyst. The sulfur trioxide gas is then passed through water to produce sulfuric acid. The balanced equation for the conversion of SO_2 to SO_3 is

$$2\ SO_2(g) + O_2(g) \xrightarrow{\Delta} 2\ SO_3(g)$$

Let's calculate the liters of oxygen gas that react with 37.5 L of sulfur dioxide. From the balanced equation, we see that 2 volumes of SO_2 = 1 volume of O_2. When we express the volume in liters, we have 2 L SO_2 = 1 L O_2. Using the volume ratio as a unit factor, we have

$$37.5\ \cancel{L\ SO_2} \times \frac{1\ L\ O_2}{2\ \cancel{L\ SO_2}} = 18.8\ L\ O_2$$

Now let's calculate the volume, in liters, of sulfur trioxide produced. From the balanced equation we see that 2 volumes of SO_2 = 2 volumes of SO_3. Therefore,

$$37.5\ \cancel{L\ SO_2} \times \frac{2\ L\ SO_3}{2\ \cancel{L\ SO_2}} = 37.5\ L\ SO_3$$

The following example exercise further illustrates the application of stoichiometry principles to a problem of the volume–volume type:

▲ **Ammonia, NH_3** Aqueous ammonia, NH_3, is found in laboratory reagent bottles labeled "ammonium hydroxide," NH_4OH.

EXAMPLE EXERCISE 9.8 Volume–Volume Stoichiometry – Haber Process

In the Haber process, nitrogen and hydrogen gases combine to give ammonia gas. If 5.55 L of nitrogen gas is available, calculate the volume of ammonia that is produced. Assume that all volumes of gas are measured under constant conditions of 500 °C and 300 atm pressure:

$$N_2(g) + 3\ H_2(g) \xrightarrow{Fe/Al_2O_3} 2\ NH_3(g)$$

Strategy Plan

STEP 1: What unit is asked for in the answer? **Step 1:** L NH_3

STEP 2: What given value is related to the answer? **Step 2:** 5.55 L N_2

STEP 3: What unit factor(s) should we apply? Using the balanced equation, we see that 1 mol N_2 = 2 mol NH_3. The volumes of gas have the same ratio as the mole ratio, so the unit factors are

Step 3: $\frac{1\ L\ N_2}{2\ L\ NH_3}$ or $\frac{2\ L\ NH_3}{1\ L\ N_2}$

Unit Analysis Map

given value	*unit factor(s)*	*unit in answer*
5.55 L N_2	$\times \frac{2\ L\ NH_3}{1\ L\ N_2}$	= ? L NH_3

Solution

We select the unit factor that cancels the unit in the given value (L N_2). Thus,

$$5.55\ \cancel{\text{L N}_2} \times \frac{2\ \text{L NH}_3}{1\ \cancel{\text{L N}_2}} = 11.1\ \text{L NH}_3$$

Practice Exercise

Calculate the volumes of (a) hydrogen chloride gas and (b) oxygen gas that react to yield 40.00 mL of chlorine gas given the following unbalanced equation. Assume that all gases are at the same temperature and pressure.

$$\text{HCl}(g) + \text{O}_2(g) \longrightarrow \text{Cl}_2(g) + \text{H}_2\text{O}(g)$$

Answers:

(a) 80.0 mL HCl

(b) 20.0 mL O_2

Concept Exercise

In general, how many unit factors are required to solve a volume–volume stoichiometry problem?

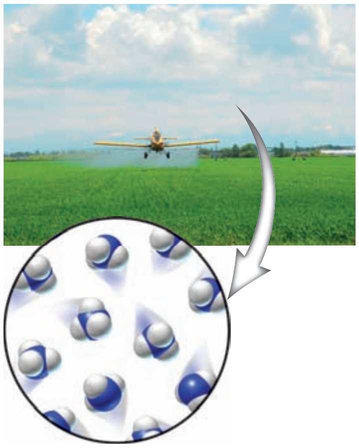
Answer: See Appendix G, 9.8.

CHEMISTRY CONNECTION Manufacturing Ammonia

Q: What are the main uses of ammonia, NH_3, which is one of the 10 most important industrial chemicals?

Over 15 million tons of ammonia are produced annually in the United States. Ammonia, NH_3, is primarily used in agriculture to fertilize the soil and replenish nitrogen depleted by crops. Although our atmosphere is a vast potential source of nitrogen, the gaseous element is unreactive and does not readily combine with other elements. Nitrogen is essentially a colorless, odorless, inert gas.

In 1905, the German chemist Fritz Haber (1868–1934) successfully prepared ammonia in the laboratory for the first time. He discovered that nitrogen and hydrogen gases combine directly at high temperatures and pressures in the presence of metal oxide catalysts. The Haber process thus provides a method for converting unreactive atmospheric nitrogen into the versatile compound ammonia, NH_3. Ammonia is easily converted to important ammonium compounds and nitric acid, HNO_3. Today, the Haber process is the main source for manufacturing nitrogen compounds throughout the world.



Soil Nitrogen Spraying anhydrous ammonia provides nitrogen to soil for optimum crop yields.

Until the beginning of the 1900s, the bulk of the world's nitrate was supplied by the rich saltpeter (KNO_3) deposits in South America. Although nitrates were used principally as fertilizer, they were also used to make explosives. Indirectly, the Haber process had a pronounced effect on World War I. Soon after the outbreak of the war, the British blockade halted the supply of Chilean saltpeter to Germany. Without a source of nitrate, Germany would have soon run out of ammunition. But by late 1914, Germany had built factories that applied the Haber process for the manufacture of ammonia and nitrate compounds.

During World War I, Haber helped to develop weapons used in chemical warfare, such as chlorine gas and even more lethal mustard gas. Haber's wife was so troubled by his crimes against humanity that she committed suicide at a dinner party honoring his role in the use of poison gas. She shot herself with his service revolver, but did not die until the following morning as Haber was leaving for the Eastern front to personally oversee the use of poison gas against the Russians. In 1918, the same year World War I ended, Haber was awarded the Nobel Prize in chemistry. Following the announcement of the Nobel Prize, he was denounced by American, English, and French scientists for his involvement in the war.

Ironically, the patriotic Haber was later condemned by his own country. When the Nazis rose to power in 1933, Haber had to resign as director of the Kaiser Wilhelm Institute because he was of Jewish descent. In distress, he fled to Cambridge University in England where he spent a few months before passing away while traveling in Switzerland.

A: The chief use of ammonia is as an agricultural fertilizer, but it is also used in explosives and household products such as ordinary window cleaner.

EXAMPLE EXERCISE 9.9 Volume–Volume Stoichiometry – Ostwald Process

In the Ostwald process, ammonia and oxygen gases combine to give nitrogen monoxide gas and steam using a platinum catalyst. The NO is then converted to nitric acid, HNO_3. If 50.0 L of NO are produced, what is the volume of NH_3 required? Assume all gaseous volumes are measured at the same conditions of temperature and pressure.

$$4\,NH_3(g) \quad + \quad 5\,O_2(g) \xrightarrow{Pt} 4\,NO(g) \quad + \quad 6\,H_2O(g)$$

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: **L NH_3**

STEP 2: What given value is related to the answer?

Step 2: **50.0 L NO**

STEP 3: What unit factor(s) should we apply? From the balanced equation, we see 4 mol NH_3 = 4 mol NO. The two unit factors are

Step 3: $\frac{4\,L\,NH_3}{4\,L\,NO}$ **or** $\frac{4\,L\,NO}{4\,L\,NH_3}$

Unit Analysis Map

given value
50.0 L NO

unit factor
$\times \frac{4\,L\,NH_3}{4\,L\,NO}$

unit in answer
= ? L NH_3

Solution

We select the unit factor that cancels the unit in the given value ($\cancel{L\,NO}$). Thus,

$$50.0\,\cancel{L\,NO} \times \frac{4\,L\,NH_3}{4\,\cancel{L\,NO}} = 50.0\,L\,NH_3$$

Practice Exercise

Calculate the volumes of (a) nitrogen monoxide gas and (b) oxygen gas that react to yield 50.0 mL of nitrogen dioxide gas given the following unbalanced equation. Assume all gases are at the same temperature and pressure.

$$NO(g) \quad + \quad O_2(g) \quad \rightarrow \quad NO_2(g)$$

Answers:

(a) 50.0 mL NO; (b) 25.0 mL O_2

Concept Exercise

Why is only one unit factor usually required to solve a volume–volume stoichiometry problem?

Answer: See Appendix G, 9.9.

9.7 Limiting Reactant Concept

LEARNING OBJECTIVES

- Explain the concept of a limiting reactant.
- Predict the limiting reactant in a chemical reaction given the number of moles of each reactant.

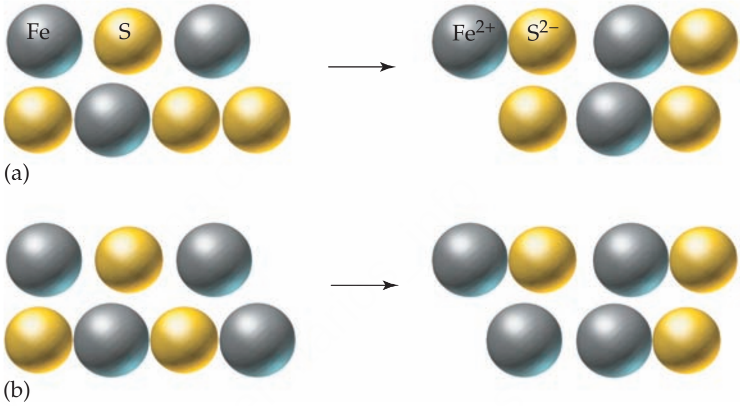
In world-class cycling, competitors carry extra bicycle frames and tires. If a cyclist participating in the Tour de France bicycle race has 5 frames and 20 tires, how many complete bicycles can be assembled? Because 5 complete frames require 10 tires, there are 10 spare tires. In this example, the frames limit the number of bicycles that can be assembled. Similarly, we will learn that one of the reactants in a chemical reaction can limit the amount of product.

▲ **Producing FeS** Heating a mixture of iron powder and sulfur powder produces iron(II) sulfide; either the iron or the sulfur can be the limiting reactant.

We have learned how to perform calculations using information from a balanced chemical equation. For example, we can calculate the amount of iron(II) sulfide produced by the reaction of metallic iron and powdered sulfur:

$$\mathrm{Fe}(s) + \mathrm{S}(s) \xrightarrow{\Delta} \mathrm{FeS}(s)$$

Previously, we assumed that there is sufficient iron and sulfur to react completely with each other. However, this assumption may not be true. What if there is not enough sulfur to react with the iron? If that were the case, the sulfur would *limit* the amount of FeS produced. In a stoichiometry problem, the reactant that controls the amount of product is called the **limiting reactant** and must be identified to correctly calculate the amount of product. Figure 9.1 is an example model to help you visualize the concept of a limiting reactant.



▲ **Figure 9.1 Limiting Reactant Concept** (a) In this reaction there is excess S; thus, Fe is the limiting reactant. Notice that an unreacted S atom remains. (b) In another reaction, there is excess Fe; thus, S is the limiting reactant. Notice that an unreacted Fe atom remains.

Determining the Limiting Reactant

Now that we have a visual model for the limiting reactant concept, let's examine data from an experiment and solve a problem. Suppose we heat 2.50 mol of iron with 3.00 mol of sulfur. How many moles of FeS are formed? We will analyze the problem by considering the amount of each substance before and after the reaction.

Before the reaction begins, there are 2.50 mol of Fe, 3.00 mol of S, and 0.00 mol of FeS. According to the balanced equation, 1 mol of Fe reacts with 1 mol of S to give 1 mol of FeS. Therefore, 2.50 mol of Fe reacts with 2.50 mol of S to give 2.50 mol of FeS. At the outset we have 3.00 mol of S, so there is not enough iron to react with all the sulfur. Therefore, iron is the limiting reactant and sulfur is the excess reactant. After the reaction, the excess sulfur is 0.50 mol (3.00 mol − 2.50 mol). Table 9.2 summarizes the amounts of each substance before and after the reaction.

TABLE 9.2 Summary of Experimental Synthesis of FeS

Experiment	mol Fe	mol S	mol FeS
before reaction:	2.50	3.00	0.00
after reaction:	0.00	0.50	2.50

EXAMPLE EXERCISE 9.10 Limiting Reactant Concept

A 1.00 mol sample of iron(II) oxide is heated with 1.00 mol of aluminum metal and converted to molten iron. Identify the limiting reactant, and calculate the moles of iron produced in the so-called Thermite reaction:

$$3\,\mathrm{FeO}(l) + 2\,\mathrm{Al}(l) \xrightarrow{\Delta} 3\,\mathrm{Fe}(l) + \mathrm{Al_2O_3}(s)$$

Strategy Plan

STEP 1:	What unit is asked for in the answer?	**Step 1:**	**mol Fe**
STEP 2:	What given value is related to the answer?	**Step 2:**	**1.00 mol FeO and 1.00 mol Al**
STEP 3:	What unit factor(s) should we apply? From the balanced equation, we have the mole ratio 3 mol FeO/3 mol Fe, and 2 mol Al/3 mol Fe.	**Step 3:**	$\frac{3\text{ mol FeO}}{3\text{ mol Fe}}$ **or** $\frac{3\text{ mol Fe}}{3\text{ mol FeO}}$ $\frac{2\text{ mol Al}}{3\text{ mol Fe}}$ **or** $\frac{3\text{ mol Fe}}{2\text{ mol Al}}$

Unit Analysis Map

given value	*unit factor*	*unit in answer*
1.00 mol FeO	× $\frac{3\text{ mol Fe}}{3\text{ mol FeO}}$	= **? mol Fe**

given value	*unit factor*	*unit in answer*
1.00 mol Al	× $\frac{3\text{ mol Fe}}{2\text{ mol Al}}$	= **? mol Fe**

Solution

We select the unit factor that cancels the unit in the given value (~~mol FeO~~). Thus,

$$1.00\,\cancel{\text{mol FeO}} \times \frac{3\text{ mol Fe}}{3\,\cancel{\text{mol FeO}}} = 1.00\text{ mol Fe}$$

Second, we select the unit factor that cancels the unit in the given value (~~mol Al~~). Thus,

$$1.00\,\cancel{\text{mol Al}} \times \frac{3\text{ mol Fe}}{2\,\cancel{\text{mol Al}}} = 1.50\text{ mol Fe}$$

Notice that FeO produces 1.00 mol Fe, whereas Al produces 1.50 mol Fe. Thus, FeO is the *limiting reactant* and the amount of product is 1.00 mol Fe.

▲ **Thermite Reaction** The thermite reaction of iron oxide and aluminum metal produces molten iron and a shower of sparks.

Practice Exercise

A 5.00 mol sample of iron(III) oxide is heated with 5.00 mol aluminum metal and converted to molten iron. Identify the limiting reactant, and calculate the moles of iron produced given the *unbalanced* equation for the reaction:

$$\mathrm{Fe_2O_3}(l) + \mathrm{Al}(l) \xrightarrow{\Delta} \mathrm{Fe}(l) + \mathrm{Al_2O_3}(s)$$

Answer: The limiting reactant is Al, which produces 5.00 mol of Fe.

Concept Exercise

A Tour de France cyclist has 10 tires and 3 frames. How many complete bicycles can be assembled?

Answer: See Appendix G, 9.10.

LEARNING OBJECTIVE

- Solve mass–mass stoichiometry problems involving a limiting reactant.
- Solve volume–volume stoichiometry problems involving a gaseous limiting reactant.

9.8 Limiting Reactant Problems

Suppose we wish to find how much molten iron is produced from the reaction of 25.0 g of FeO with 25.0 g of Al. The balanced equation for the reaction is

$$3\,\text{FeO}(l) + 2\,\text{Al}(l) \xrightarrow{\Delta} 3\,\text{Fe}(l) + \text{Al}_2\text{O}_3(s)$$

Because the mass of each reactant is given, it is not obvious whether FeO or Al limits the amount of iron produced. In problems where we are given the amounts of two reactants, we can determine the limiting reactant in three steps.

I. *Calculate the mass of product that can be produced from the first reactant.*
  (a) Calculate the moles of reactant.
  (b) Calculate the moles of product.
  (c) Calculate the mass of product.

II. *Calculate the mass of product that can be produced from the second reactant.*
  (a) Calculate the moles of reactant.
  (b) Calculate the moles of product.
  (c) Calculate the mass of product.

III. *State the limiting reactant and the corresponding mass of product formed.* The limiting reactant gives the least amount of product. The actual mass of product obtained from the chemical reaction is the lesser of the two product masses calculated in Steps I and II.

Applying I: We can calculate the amount of iron produced from iron(II) oxide in the preceding problem. Using the balanced equation for the reaction, we can find the mass of Fe produced from 25.0 g of FeO. Let's outline the solution to the problem:

$$\text{g FeO} \xrightarrow[\text{I(a)}]{} \text{mol FeO} \xrightarrow[\text{I(b)}]{} \text{mol Fe} \xrightarrow[\text{I(c)}]{} \text{g Fe}$$

From the periodic table, we find that for Fe, the molar mass is 55.85 g/mol, and for FeO, the molar mass is 71.85 g/mol. The unit analysis solution to the problem is

$$25.0\,\cancel{\text{g FeO}} \times \frac{1\,\cancel{\text{mol FeO}}}{71.85\,\cancel{\text{g FeO}}} \times \frac{3\,\cancel{\text{mol Fe}}}{3\,\cancel{\text{mol FeO}}} \times \frac{55.85\text{ g Fe}}{1\,\cancel{\text{mol Fe}}} = 19.4\text{ g Fe}$$

Applying II: We calculate the amount of iron that can be produced by the aluminum. Using the balanced equation, we can find the mass of Fe produced from 25.0 g of Al. We can outline the solution to the problem as follows:

$$\text{g Al} \xrightarrow[\text{II(a)}]{} \text{mol Al} \xrightarrow[\text{II(b)}]{} \text{mol Fe} \xrightarrow[\text{II(c)}]{} \text{g Fe}$$

From the periodic table, we find that for Al, the molar mass is 26.98 g/mol. The unit analysis solution to the problem is

$$25.0\,\cancel{\text{g Al}} \times \frac{1\,\cancel{\text{mol Al}}}{26.98\,\cancel{\text{g Al}}} \times \frac{3\,\cancel{\text{mol Fe}}}{2\,\cancel{\text{mol Al}}} \times \frac{55.85\text{ g Fe}}{1\,\cancel{\text{mol Fe}}} = 77.6\text{ g Fe}$$

Appying III: We compare the grams of Fe produced form each of the reactants:

$$25.0\text{ g FeO} \longrightarrow 19.4\text{ g Fe}$$

$$25.0\text{ g Al} \longrightarrow 77.6\text{ g Fe}$$

We see that the FeO yields less mass of product; thus, FeO is the *limiting reactant.* Conversely, aluminum metal is the excess reactant, because not all the Al is used in the reaction. Therefore, the maximum yield of metallic iron from the reaction is limited to 19.4 g of Fe.

Example Exercise 9.11 provides additional practice in solving a limiting reactant problem.

EXAMPLE EXERCISE 9.11 Mass–Mass Limiting Reactant

In a reaction, 50.0 g of manganese(IV) oxide reacts with 25.0 g of aluminum. Identify the limiting reactant and calculate the mass of manganese metal produced. The equation for the reaction is

$$3\ \mathrm{MnO_2}(l) + 4\ \mathrm{Al}(l) \xrightarrow{\Delta} 3\ \mathrm{Mn}(l) + 2\ \mathrm{Al_2O_3}(s)$$

Solution

STEP 1: We calculate the mass of Mn produced from 50.0 g of MnO_2. We can outline the solution as follows:

$$\mathrm{g\ MnO_2} \longrightarrow \mathrm{mol\ MnO_2} \longrightarrow \mathrm{mol\ Mn} \longrightarrow \mathrm{g\ Mn}$$

From the periodic table, we find the molar mass of MnO_2 is 86.94 g/mol and Mn is 54.94 g/mol. The unit analysis solution to the problem is

$$50.0\ \cancel{\mathrm{g\ MnO_2}} \times \frac{1\ \cancel{\mathrm{mol\ MnO_2}}}{86.94\ \cancel{\mathrm{g\ MnO_2}}} \times \frac{3\ \cancel{\mathrm{mol\ Mn}}}{3\ \cancel{\mathrm{mol\ MnO_2}}} \times \frac{54.94\ \mathrm{g\ Mn}}{1\ \cancel{\mathrm{mol\ Mn}}} = 31.6\ \mathrm{g\ Mn}$$

STEP 2: We calculate the mass of Mn produced from 25.0 g of Al. We can outline the solution as follows:

$$\mathrm{g\ Al} \longrightarrow \mathrm{mol\ Al} \longrightarrow \mathrm{mol\ Mn} \longrightarrow \mathrm{g\ Mn}$$

From the periodic table, we find that the molar mass of Al is 26.98 g/mol. Starting with 25.0 g of Al, the unit analysis solution is

$$25.0\ \cancel{\mathrm{g\ Al}} \times \frac{1\ \cancel{\mathrm{mol\ Al}}}{26.98\ \cancel{\mathrm{g\ Al}}} \times \frac{3\ \cancel{\mathrm{mol\ Mn}}}{4\ \cancel{\mathrm{mol\ Al}}} \times \frac{54.94\ \mathrm{g\ Mn}}{1\ \cancel{\mathrm{mol\ Mn}}} = 38.2\ \mathrm{g\ Mn}$$

STEP 3: We compare the mass of product obtained from each of the reactants:

$$50.0\ \mathrm{g\ MnO_2} \longrightarrow 31.6\ \mathrm{g\ Mn}$$

$$25.0\ \mathrm{g\ Al} \rightarrow 38.2\ \mathrm{g\ Mn}$$

In this example, MnO_2 is the *limiting reactant* because it yields less product. Al is the excess reactant and is not completely consumed in the reaction. Thus, the maximum amount of product from the reaction is 31.6 g of Mn.

Practice Exercise

In a reaction, 75.0 g of manganese(IV) oxide react with 30.0 g of aluminum. Identify the limiting reactant and calculate the mass of aluminum oxide produced:

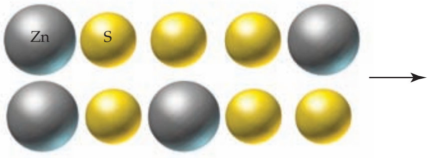
$$3\ \mathrm{MnO_2}(l) + 4\ \mathrm{Al}(l) \xrightarrow{\Delta} 3\ \mathrm{Mn}(l) + 2\ \mathrm{Al_2O_3}(s)$$

Answer: The limiting reactant is Al, which gives 56.7 g of Al_2O_3.

Concept Exercise

Indicate the limiting reactant shown in the following illustration:

$$\mathrm{Zn}(s) + \mathrm{S}(s) \xrightarrow{\Delta} \mathrm{ZnS}(s)$$



Answer: See Appendix G, 9.11.

Limiting Reactants Involving Volumes of Gas

Now let's try a limiting reactant problem that involves reactants and products in the gaseous state. For example, consider the reaction of nitrogen monoxide and oxygen to give nitrogen dioxide. The balanced equation for the reaction is

$$2\ NO(g) + O_2(g) \xrightarrow{UV} 2\ NO_2(g)$$

Suppose 5.00 L of NO gas reacts with 5.00 L of O_2 gas. The volumes of both reactants are given, so it is not obvious which gas limits the amount of product. We will assume that the temperature and pressure remain constant and use the following three steps to determine the limiting reactant:

I. *Calculate the volume of product that can be produced from the first reactant.*
   Recall that the coefficients in the balanced equation are in the same mole ratio as the volumes of gas.

II. *Calculate the volume of product that can be produced from the second reactant.*
   Again, the coefficients in the balanced equation are in the same mole ratio as the volumes of gas.

III. *State the limiting reactant and the corresponding volume of product.*
   The limiting reactant is the gas that gives the least amount of product. The actual volume of product obtained from the chemical reaction is the lesser of the two volumes calculated in Steps I and II.

▲ **Oxyacetylene Welding** The reaction of oxygen and acetylene gases releases enough heat to melt iron.

Applying I: We calculate the volume of nitrogen dioxide produced from 5.00 L of nitrogen monoxide gas. From the balanced equation, we see that 2 mol of NO produce 2 mol of NO_2. Accordingly, 2 L of NO produce 2 L of NO_2. Hence,

$$5.00\ \cancel{L\ NO} \times \frac{2\ L\ NO_2}{2\ \cancel{L\ NO}} = 5.00\ L\ NO_2$$

Applying II: We calculate the volume of nitrogen dioxide produced from 5.00 L of oxygen gas. From the balanced equation, we see that 1 mol of O_2 produces 2 mol of NO_2. Accordingly, 1 L of O_2 produces 2 L of NO_2. Thus,

$$5.00\ \cancel{L\ O_2} \times \frac{2\ L\ NO_2}{1\ \cancel{L\ O_2}} = 10.0\ L\ NO_2$$

Applying III: We compare the liters of NO_2 gas from each of the reactants.

$$5.00\ L\ NO \longrightarrow 5.00\ L\ NO_2$$

$$5.00\ L\ O_2 \longrightarrow 10.0\ L\ NO_2$$

We see that the NO yields less volume of product; thus, NO is the *limiting reactant*. Conversely, O_2 is the excess reactant, and not all the O_2 gas is used in the reaction. Therefore, the maximum yield of product from the reaction is 5.00 L of NO_2. The following example exercise provides additional practice in solving a limiting reactant problem involving gases.

EXAMPLE EXERCISE 9.12 Volume–Volume Limiting Reactant

In oxyacetylene welding, acetylene reacts with oxygen to give carbon dioxide and water. If 25.0 mL of C_2H_2 reacts with 75.0 mL of O_2, what is the limiting reactant? Assuming constant conditions, what is the volume of CO_2 produced?

$$2\,C_2H_2(g) + 5\,O_2(g) \xrightarrow{\text{spark}} 4\,CO_2(g) + 2\,H_2O(g)$$

Strategy Plan

STEP 1: What unit is asked for in the answer?

Step 1: mL CO_2

STEP 2: What given value is related to the answer?

Step 2: 25.0 mL C_2H_2 and 75.0 mL O_2

STEP 3: What unit factor(s) should we apply? From the balanced equation, we have the mole ratio so the volume ratio is 2 mL C_2H_2/4 mL CO_2, and 5 mL O_2/4 mL CO_2.

Step 3: $\dfrac{2\text{ mL } C_2H_2}{4\text{ mL } CO_2}$ or $\dfrac{4\text{ mL } CO_2}{2\text{ mL } C_2H_2}$

$\dfrac{5\text{ mL } O_2}{4\text{ mL } CO_2}$ or $\dfrac{4\text{ mL } CO_2}{5\text{ mL } O_2}$

Unit Analysis Mapw

given value	unit factor	unit in answer
25.0 mL C_2H_2	$\times\ \dfrac{4\text{ mL } CO_2}{2\text{ mL } C_2H_2}$	= ? mL CO_2
75.0 mL O_2	$\times\ \dfrac{4\text{ mL } CO_2}{5\text{ mL } O_2}$	= ? mL CO_2

Solution

First, we find the amount of product that can be produced from the first reactant (mL C_2H_2). Thus,

$$25.0\ \cancel{\text{mL } C_2H_2} \times \frac{4\text{ mL } CO_2}{2\ \cancel{\text{mL } C_2H_2}} = 50.0\text{ mL } CO_2$$

Second, we find the amount of product that can be produced from the second reactant (mL O_2):

$$75.0\ \cancel{\text{mL } O_2} \times \frac{4\text{ mL } CO_2}{5\ \cancel{\text{mL } O_2}} = 60.0\text{ mL } CO_2$$

Third, we compare the amount of product from each reactant and see that C_2H_2 produces 50.0 mL CO_2, whereas O_2 produces 60.0 mL CO_2. Thus, C_2H_2 is the *limiting reactant* and the amount of product is limited to 50.0 mL CO_2.

Practice Exercise

Ethane undergoes combustion to give carbon dioxide and water. If 10.0 L of C_2H_6 reacts with 25.0 L of O_2, what is the limiting reactant? Assuming constant conditions, what is the volume of CO_2 produced? The equation is

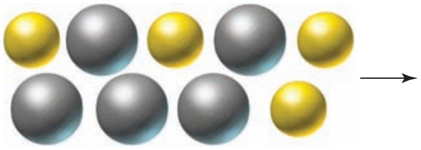
$$2\,C_2H_6(g) + 7\,O_2(g) \xrightarrow{\text{spark}} 4\,CO_2(g) + 6\,H_2O(g)$$

Answers: The limiting reactant is O_2, which gives 14.3 L of CO_2.

Concept Exercise

Indicate the limiting reactant shown in the following illustration:

$$Zn(s) + S(s) \xrightarrow{\Delta} ZnS(s)$$



Answer: See Appendix G, 9.12.

LEARNING OBJECTIVE

- Calculate the percent yield for a reaction given the actual yield and theoretical yield.

▲ **Precipitate of CdS** The aqueous solution reaction of $CdCl_2$ and Na_2S yields a yellow precipitate of CdS.

▲ **Precipitate of $CuCO_3$** The aqueous solution reaction of $Cu(NO_3)_2$ and Na_2CO_3 yields a precipitate of $CuCO_3$.

9.9 Percent Yield

To understand the concept of yield, consider the following laboratory experiment. First, a student weighs a 1.00-g sample of cadmium chloride and dissolves it in water. Then, the student adds aqueous sodium sulfide solution to obtain a precipitate of cadmium sulfide, CdS. The balanced equation for the reaction is

$$CdCl_2(aq) + Na_2S(aq) \longrightarrow CdS(s) + 2\,NaCl(aq)$$

The student collects the insoluble CdS precipitate in filter paper and determines its mass on a balance. The mass of the precipitate is called the **actual yield**.

The mass of cadmium sulfide obtained, starting with the 1.00-g sample of cadmium chloride, can also be calculated as a mass–mass problem. The calculated amount of precipitate is called the **theoretical yield**. In practice, there is always inherent experimental error. Thus, we compare the actual and theoretical yields and express the results in terms of percent yield. The **percent yield** is the amount of the actual yield compared to the theoretical yield expressed as a percent:

$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \text{percent yield}$$

To find the percent yield for this laboratory experiment, we start with the 1.00-g sample of cadmium chloride. Applying stoichiometry to the balanced equation, we calculate that the amount of cadmium sulfide precipitate is 0.788 g. Suppose the student weighs the CdS precipitate and finds that it has an actual mass of 0.775 g. To find the percent yield, we compare the actual yield (0.775 g) to the theoretical yield (0.788 g) as follows:

$$\frac{0.775\,\cancel{g}}{0.788\,\cancel{g}} \times 100\% = 98.4\%$$

Suppose that the student performs a second trial for the same experiment, using a 1.00-g sample of cadmium chloride. In this second trial, the precipitate has a mass of 0.805 g. Since the student started with 1.00 g of cadmium chloride, the theoretical yield for the two trials is the same, 0.788 g. As before, to find the percent yield we compare the actual yield to the theoretical yield.

$$\frac{0.805\,\cancel{g}}{0.788\,\cancel{g}} \times 100\% = 102\%$$

Notice that the percent yield for this second trial is greater than 100%. Although surprising, this result is possible. In an experiment, some errors lead to high results and other errors lead to low results. In the two trials, the student made different types of experimental errors. For example, high results will be obtained when impurities are trapped, or when the precipitate is not completely dry.

The following example exercise provides additional practice in calculating percent yield:

EXAMPLE EXERCISE 9.13 Percent Yield

A student dissolves 1.50 g of copper(II) nitrate in water. After adding aqueous sodium carbonate solution, the student obtains 0.875 g of $CuCO_3$ precipitate. If the theoretical yield is 0.988 g, what is the percent yield?

$$Cu(NO_3)_2(aq) + Na_2CO_3(aq) \longrightarrow CuCO_3(s) + 2\,NaNO_3(aq)$$

Solution

The percent yield is the ratio of the actual yield compared to the theoretical yield. In this experiment, the actual yield is 0.875 g and the theoretical yield is 0.988 g. The percent yield, therefore, is

$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \text{percent yield}$$

$$\frac{0.875\,\cancel{g}}{0.988\,\cancel{g}} \times 100\% = 88.6\%$$

The percent yield obtained by the student is 88.6%.

Practice Exercise

Ammonium nitrate is used in explosives and is produced by the reaction of ammonia, NH_3, and nitric acid. The equation for the reaction is

$$NH_3(g) + HNO_3(aq) \longrightarrow NH_4NO_3(s)$$

If 15.0 kg of ammonia gives an actual yield of 65.3 kg of ammonium nitrate, what is the percent yield? The calculated yield of ammonium nitrate for the experiment is 70.5 kg.

Answer: 92.6%

Concept Exercise

Is it possible to have a percent yield greater than 100%?

Answer: See Appendix G, 9.13.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
9.1 Interpreting a Chemical Equation We learned to interpret the coefficients in a balanced equation as a **mole ratio**. We also showed that by applying **Avogadro's theory** we can express the mole ratio in terms of volumes of gases. Furthermore, we illustrated the **law of conservation of mass** by substituting the **molar mass** for a mole of each reactant and product into a balanced chemical equation.	• Relate the coefficients in a balanced chemical equation to. (a) moles of reactants and products, and (b) liters of gaseous reactants and products. ***Related Exercises: 1–4*** • Verify the coefficients of a balanced chemical equation using the conservation of mass law. ***Related Exercises: 5–6***
9.2 Mole–Mole Relationships We learned to relate the moles of reactants to the moles of products by simply applying the coefficients in the balanced chemical reaction.	• Relate the number of moles of two substances in a balanced equation. ***Related Exercises: 7–14***
9.3 Types of Stoichiometry Problems **Stoichiometry** relates the quantities of substances involved in a chemical reaction. There are three basic types of stoichiometry problems. In a **mass–mass problem**, the mass of one substance is related to the mass of another substance according to a balanced chemical equation. In a **mass–volume problem**, the mass of a substance is related to the volume of a gas. In a **volume–volume problem**, the volumes of two gaseous substances are related.	• Classify the three basic types of stoichiometry problems: mass–mass, mass–volume, and volume–volume. ***Related Exercises: 15–20***
9.4 Mass–Mass Problems In a mass–mass problem, we first convert the given mass of a substance to the number of moles using the molar mass. If the molar mass is not shown in the problem, we refer to the periodic table and add up the gram atomic masses for each element in the substance. Next, we apply the mole ratio to convert the moles of the given substance to the moles of the unknown substance using the coefficients from the balanced equation. Finally, we determine the mass of the unknown by multiplying by the molar mass of the unknown.	• Perform mass–mass stoichiometry calculations. ***Related Exercises: 21–30***
9.5 Mass–Volume Problems In a mass–volume problem, we first convert the given mass of substance to the number of moles using the molar mass. Next, we apply the mole ratio to convert the moles of the given substance to the moles of the unknown substance using the coefficients from the balanced equation. Finally, we determine the volume of the unknown gas at **standard temperature and pressure (STP)** by multiplying by the **molar volume**. The volume occupied by 1 mol of gas is 22.4 L at STP, that is, 0 °C and 1 atm pressure.	• Perform mass–volume stoichiometry calculations. ***Related Exercises: 31–38***
9.6 Volume–Volume Problems In a volume-volume problem, we can take advantage of Gay-Lussac's **law of combining volumes**. Since the volumes of gases combine in the same whole-number ratio as the coefficients in the balanced equation, we can solve this type of problem in one conversion step. We simply multiply the given volume by the ratio of the coefficients in the equation.	• Perform volume–volume stoichiometry calculations. ***Related Exercises: 39–48***

Key Concepts	Learning Objectives and Related Exercises
9.7 Limiting Reactant Concept In more difficult stoichiometry problems, we cannot assume that there is a sufficient quantity of each reactant available. The amount of one reactant may not be sufficient to completely consume another reactant that is in excess. The reactant that controls the reaction and limits the amount of product formed is referred to as the **limiting reactant**.	• Explain the concept of a limiting reactant. *Related Exercises: 49–52* • Predict the limiting reactant in a chemical reaction given the number of moles of each reactant. *Related Exercises: 49–58*
9.8 Limiting Reactant Problems If we are given the amounts of two reactants, we can determine the limiting reactant by finding the amount of product yield from each of the reactants. The reactant that produces the least amount of product controls the reaction and is the limiting reactant. A limiting reactant problem may involve the masses of two reacting substances as well as the volumes of two gases.	• Perform mass–mass stoichiometry calculations involving a limiting reactant. *Related Exercises: 59–66* • Perform volume–volume stoichiometry calculations involving a gaseous limiting reactant. *Related Exercises: 67–74*
9.9 Percent Yield The percent yield provides an evaluation of the laboratory method and the skill of the analyst. A **percent yield** from an experiment expresses the amount of the **actual yield** compared to the **theoretical yield**.	• Calculate the percent yield for a reaction given the actual yield and theoretical yield. *Related Exercises: 75–78*

Problem–Solving Organizer

Topic	Procedure	Example
Interpreting a Chemical Equation Sec. 9.1	Express the mole ratio and volume ratio of gaseous substances in a balanced equation using the coefficients.	Given that $N_2(g) + O_2(g) \longrightarrow 2\,NO(g)$, interpret the mole ratio and the volume ratio in liters. $1\text{ mol } N_2 + 1\text{ mol } O_2 \longrightarrow 2\text{ mol NO}$ $1\text{ L } N_2 + 1\text{ L } O_2 \longrightarrow 2\text{ L NO}$
Mole–Mole Relationships Sec. 9.2	Convert moles of a given substance to moles of an unknown using the mole ratio.	How many moles of NO are produced from 0.500 mol N_2? $0.500\ \cancel{\text{mol } N_2} \times \frac{2\text{ mol NO}}{1\ \cancel{\text{mol } N_2}} = 1.00\text{ mol NO}$
Types of Stoichiometry Problems Sec. 9.3	Classify a stoichiometry problem as a mass–mass, mass–volume, or volume–volume type of problem.	What type of problem is the following: How many grams of NO are produced from 11.2 L N_2? This is a mass–volume type problem.
Mass–Volume Problems Sec. 9.4	(a) Convert mass of given to moles of given. (b) Convert moles of given to moles of unknown. (c) Convert moles of unknown to mass of unknown.	How many grams of NO are produced from 10.0 g N_2? $10.0\ \cancel{\text{g } N_2} \times \frac{1\ \cancel{\text{mol } N_2}}{28.02\ \cancel{\text{g } N_2}} \times \frac{2\ \cancel{\text{mol NO}}}{1\ \cancel{\text{mol } N_2}} \times \frac{30.01\text{ g NO}}{1\ \cancel{\text{mol NO}}}$ $= 21.4\text{ g NO}$
Mass–Volume Problems Sec. 9.5	(a) Convert mass of given to moles of given. (b) Convert moles of given to moles unknown. (c) Convert moles unknown to volume unknown.	How many liters of NO at STP are produced from 10.0 g N_2? $10.0\ \cancel{\text{g } N_2} \times \frac{1\ \cancel{\text{mol } N_2}}{28.02\ \cancel{\text{g } N_2}} \times \frac{2\ \cancel{\text{mol NO}}}{1\ \cancel{\text{mol } N_2}} \times \frac{22.4\text{ L NO}}{1\ \cancel{\text{mol NO}}}$ $= 16.0\text{ L NO}$
Volume–Volume Problems Sec. 9.6	Convert volume of given to volume of unknown.	How many liters of NO at STP are produced from 10.0 L N_2? $10.0\ \cancel{\text{L } N_2} \times \frac{2\text{ L NO}}{1\ \cancel{\text{L } N_2}} = 20.0\text{ L NO}$

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** the ratio of moles of reactants and products according to the coefficients in the balanced chemical equation

______ **2.** the principle that equal volumes of gases contain equal numbers of molecules, at the same temperature and pressure

______ **3.** the principle that mass is neither gained nor lost during a chemical reaction

______ **4.** the mass in grams of 1 mol of substance

______ **5.** the relationship of quantities (i.e., mass of a substance, or volume of a gas) in a chemical reaction according to the balanced equation

______ **6.** a type of stoichiometry calculation that relates the masses of two substances

______ **7.** a type of stoichiometry calculation that relates the mass of a substance to the volume of a gas

______ **8.** a type of stoichiometry calculation that relates the volumes of two gases at the same temperature and pressure

______ 9. the volume of 1 mol of gas at STP, that is, 22.4 L/mol

______ **10.** a temperature of 0 °C and a pressure of 1 atm

______ **11.** the principle that states that the volumes of gases that react in a chemical reaction are in the ratio of small whole numbers

______ **12.** the substance in a chemical reaction that controls the maximum amount of product formed

______ **13.** the amount of product experimentally obtained from a given amount of reactant

______ **14.** the amount of product calculated to be obtained from a given amount of reactant

______ **15.** the actual yield compared to the theoretical yield expressed as a percent

(a) actual yield (*Sec. 9.9*)
(b) Avogadro's theory (*Sec. 9.1*)
(c) law of combining volumes (*Sec. 9.6*)
(d) law of conservation of mass (*Sec. 9.1*)
(e) limiting reactant (*Sec. 9.7*)
(f) mass–mass problem (*Sec. 9.3*)
(g) mass–volume problem (*Sec. 9.3*)
(h) molar mass (*Sec. 9.1*)
(i) molar volume (*Sec. 9.5*)
(j) mole ratio (*Sec. 9.1*)
(k) percent yield (*Sec. 9.9*)
(l) standard temperature and pressure (STP) (*Sec. 9.5*)
(m) stoichiometry (*Sec. 9.3*)
(n) theoretical yield (*Sec. 9.9*)
(o) volume–volume problem (*Sec. 9.3*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Interpreting a Chemical Equation (Sec. 9.1)

1. Consider the general chemical equation:
$\text{A} + 2\,\text{B} \longrightarrow \text{C} + 2\,\text{D}$
(a) How many moles of C are produced from 1 mol of A?
(b) How many liters of gas B must react to give 2 L of gas D at the same temperature and pressure?

2. Consider the general chemical equation:
$2\,\text{A} + 3\,\text{B} \longrightarrow 2\,\text{C} + \text{D}$
(a) How many moles of C are produced from 2 mol of A?
(b) How many liters of gas B must react to give 1 L of gas D at the same temperature and pressure?

3. Consider the general chemical equation: $\text{A} + 3\,\text{B} \longrightarrow 2\,\text{C}$
(a) If 1.00 g of A reacts with 1.50 g of B, what is the mass of C?
(b) If 2.00 g of A reacts to produce 5.00 g of C, what is the mass of B?

4. Consider the general chemical equation: $3\,\text{A} + \text{B} \longrightarrow 2\,\text{C}$
(a) If 1.50 g of A reacts with 1.65 g of B, what is the mass of C?
(b) If 4.50 g of A reacts to produce 9.45 g of C, what is the mass of B?

5. Illustrate the conservation of mass law using the molar masses of reactants and products for each substance in the following balanced equations.
(a) $2\,N_2(s) + 3\,O_2(g) \longrightarrow 2\,N_2O_3(s)$
(b) $2\,N_2(s) + 5\,O_2(g) \longrightarrow 2\,N_2O_5(s)$

6. Illustrate the conservation of mass law using the molar masses of reactants and products for each substance in the following balanced equations.
(a) $P_4(s) + 3\,O_2(g) \longrightarrow 2\,P_2O_3(s)$
(b) $P_4(s) + 5\,O_2(g) \longrightarrow 2\,P_2O_5(s)$

Mole–Mole Relationships (Sec. 9.2)

7. Given the following decomposition reaction, calculate the moles of oxygen gas produced from 5.00 mol of H_2O_2.

$$2\,H_2O_2(l) \xrightarrow{\Delta} 2\,H_2O(l) + O_2(g)$$

8. Given the following decomposition reaction, calculate the moles of water produced from 5.00 mol of H_2O_2.

$$2\,H_2O_2(l) \xrightarrow{\Delta} 2\,H_2O(l) + O_2(g)$$

9. Given the following decomposition reaction, calculate the moles of oxygen gas produced from 5.00 mol of $LiClO_3$.

$$2\,LiClO_3(s) \xrightarrow{\Delta} 2\,LiCl(s) + O_2(g)$$

10. Given the following decomposition reaction, calculate the moles of lithium chloride produced from 5.00 mol of $LiClO_3$.

$$2\,LiClO_3(s) \xrightarrow{\Delta} 2\,LiCl(s) + O_2(g)$$

11. How many moles of chlorine gas react with 0.100 mol of metallic iron given the *unbalanced* chemical equation? How many moles of iron(III) chloride are produced?

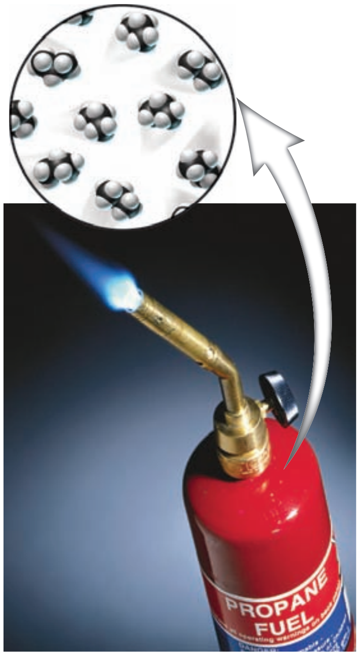
$$Fe(s) + Cl_2(g) \xrightarrow{\Delta} FeCl_3(s)$$

12. How many moles of barium metal react to produce 0.100 mol of barium nitride given the *unbalanced* chemical equation? How many moles of nitrogen gas react?

$$Ba(s) + N_2(g) \xrightarrow{\Delta} Ba_3N_2(s)$$

13. How many moles of propane gas, C_3H_8, react with 2.50 mol of oxygen gas given the *unbalanced* chemical equation? How many moles of carbon dioxide are produced?

$$C_3H_8(g) + O_2(g) \xrightarrow{\Delta} CO_2(g) + H_2O(g)$$



◀ **Propane Torch** The reaction of propane gas, C_3H_8 and oxygen in the air releases heat energy.

14. How many moles of butane gas, C_4H_{10}, react to produce 2.50 mol of water given the *unbalanced* chemical equation? How many moles of oxygen gas react?

$$C_4H_{10}(g) + O_2(g) \xrightarrow{\Delta} CO_2(g) + H_2O(g)$$

Types of Stoichiometry Problems (Sec. 9.3)

15. Classify the following type of stoichiometry problem: How many grams of carbon dioxide gas are produced from the decomposition of 1.10 g $NaHCO_3$?

16. Classify the following type of stoichiometry problem: How many grams of baking soda decompose to give 1.25 g Na_2CO_3?

17. Classify the following type of stoichiometry problem: How many grams of ferric oxide are produced from iron metal and 50.0 L of steam?

18. Classify the following type of stoichiometry problem: How many liters of steam react with iron metal to yield 1.88 g FeO?

19. Classify the following type of stoichiometry problem: How many milliliters of hydrogen chloride gas are produced from hydrogen and 25.0 mL Cl_2?

20. Classify the following type of stoichiometry problem: How many milliliters of chlorine gas react with 50.0 mL H_2?

Mass–Mass Problems (Sec. 9.4)

21. What is the mass of zinc oxide that can be prepared from 1.00 g of zinc metal?

$$2\ Zn(s) + O_2(g) \xrightarrow{\Delta} 2\ ZnO(s)$$

22. How many grams of oxygen gas react to give 1.50 g of ZnO according to the equation in Exercise 21?

23. What is the mass of product that can be prepared from 3.45 g of bismuth metal.

$$2\ Bi(s) + 3Cl_2(g) \longrightarrow 2\ BiCl_3(s)$$

24. How many grams of chlorine gas react to give 3.52 g of $BiCl_3$ according to the equation in Exercise 23?

25. What is the mass of silver that can be prepared from 1.25 g of copper metal?

$$Cu(s) + 2\ AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2\ Ag(s)$$

26. How many grams of copper react to give 1.50 g Ag according to the equation in Exercise 25?

27. What is the mass of mercury that can be prepared from 1.25 g of cobalt metal?

$$Co(s) + HgCl_2(aq) \longrightarrow CoCl_3(aq) + Hg(l)$$

28. How many grams of mercuric chloride react to give 5.11 g of Hg according to the equation in Exercise 27?

29. What is the mass of calcium phosphate that can be prepared from 1.78 g Na_3PO_4?

$$Na_3PO_4(aq) + Ca(OH)_2(aq) \longrightarrow Ca_3(PO_4)_2(s) + NaOH(aq)$$

30. How many grams of calcium hydroxide react to give 2.39 g of $Ca_3(PO_4)_2$ according to the equation in Exercise 29?

Mass–Volume Problems (Sec. 9.5)

31. How many milliliters of carbon dioxide gas at STP are produced from the decomposition of 1.75 g of sodium hydrogen carbonate?

$$2\ NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

32. How many grams of sodium hydrogen carbonate decompose to give 25.0 mL of carbon dioxide gas at STP according to the equation in Exercise 31?

33. How many milliliters of carbon dioxide gas at STP are produced from the decomposition of 5.00 g of iron(III) carbonate?

$$Fe_2(CO_3)_3(s) \xrightarrow{\Delta} Fe_2O_3(s) + 3CO_2(g)$$

34. How many grams of iron(III) carbonate decompose to give 50.0 mL of carbon dioxide gas at STP according to the equation in Exercise 33?

35. How many milliliters of oxygen gas at STP are released from the decomposition of 2.50 g of mercury(II) oxide?

$$2\ HgO(s) \xrightarrow{\Delta} 2\ Hg(l) + O_2(g)$$

36. How many grams of mercury(II) oxide decompose to give 50.0 mL of oxygen gas at STP according to the equation in Exercise 35?

37. How many milliliters of oxygen gas at STP are produced from the decomposition of 1.55 g of hydrogen peroxide, H_2O_2?

$$H_2O_2(l) \xrightarrow{\Delta} H_2O(l) + O_2(g)$$

38. How many grams of hydrogen peroxide decompose to give 50.0 mL of oxygen gas at STP according to the equation in Exercise 37?

◀ **Hydrogen Peroxide** Hydrogen peroxide is unstable and breaks down to give water and oxygen gas. When H_2O_2 is poured on a wound, the fizzing bubbles are O_2 gas.

Volume–Volume Problems (Sec. 9.6)

39. Assuming the same temperature and pressure for each gas, how many liters of nitrogen gas react with 50.0 L of hydrogen gas?

$$3\,H_2(g) + N_2(g) \xrightarrow{\Delta} 2\,NH_3(g)$$

40. Assuming the same temperature and pressure for each gas, how many liters of ammonia are produced from 50.0 L of H_2 according to the equation in Exercise 39?

41. Assuming the same temperature and pressure for each gas, how many milliliters of oxygen gas react with 10.0 mL of carbon monoxide?

$$2\,CO(g) + O_2(g) \xrightarrow{\Delta} 2\,CO_2(g)$$

42. Assuming the same temperature and pressure for each gas, how many milliliters of carbon dioxide are produced from 10.0 mL of CO according to the equation in Exercise 41?

43. Assuming constant conditions, how many milliliters of chlorine gas react to yield 1.75 L of dichlorine trioxide?

$$Cl_2(g) + O_2(g) \longrightarrow Cl_2O_3(g)$$

44. Assuming constant conditions, how many milliliters of oxygen gas react to give 1.75 L of Cl_2O_3 according to the equation in Exercise 43?

45. Assuming constant conditions, what volume of sulfur trioxide gas is produced from 25.0 L of oxygen gas?

$$SO_2(g) + O_2(g) \longrightarrow SO_3(g)$$

46. Assuming constant conditions, what volume of sulfur dioxide gas reacts with 25.0 L of O_2 according to the equation in Exercise 45?

47. Assuming constant conditions, what volume of nitrogen gas reacts to yield 500.0 cm^3 of dinitrogen pentaoxide?

$$N_2(g) + O_2(g) \longrightarrow N_2O_5(g)$$

48. Assuming constant conditions, what volume of oxygen gas reacts to yield 500.0 cm^3 of N_2O_5 according to the equation in Exercise 47?

Limiting Reactant Concept (Sec. 9.7)

▲ **Tour de France** The relationship between bicycle frames and tires illustrates the limiting reactant concept. If 5 frames and 20 tires are available, the number of frames limits the number of bicycles to 5.

49. If 5.00 mol of nitrogen gas and 5.00 mol of oxygen gas react, what is the limiting reactant and how many moles of NO are produced from the reaction?

$$N_2(g) + O_2(g) \longrightarrow 2\,NO(g)$$

50. If 0.125 mol of nitrogen gas and 0.125 mol of oxygen gas react, what is the limiting reactant and how many moles of NO are produced according to the equation in Exercise 49?

51. If 1.50 mol of nitrogen monoxide gas and 1.50 mol of oxygen gas react, what is the limiting reactant and how many moles of NO_2 are produced from the reaction?

$$2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g)$$

52. If 1.50 mol of nitrogen monoxide gas and 2.50 mol of oxygen gas react, what is the limiting reactant and how many moles of NO_2 are produced according to the equation in Exercise 51?

53. If 5.00 mol of hydrogen gas and 5.00 mol of oxygen gas react, what is the limiting reactant and how many moles of water are produced from the reaction?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$$

54. If 5.00 mol of hydrogen gas and 1.50 mol of oxygen gas react, what is the limiting reactant and how many moles of water are produced according to the equation in Exercise 53?

55. If 1.00 mol of ethane gas and 5.00 mol of oxygen gas react, what is the limiting reactant and how many moles of water are produced from the reaction?

$$2\,C_2H_6(g) + 7\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$$

56. If 1.00 mol of ethane gas and 3.00 mol of oxygen gas react, what is the limiting reactant and how many moles of water are produced according to the equation in Exercise 55?

57. The following table indicates the amounts of Co and S before reaction. Complete the table after reactions (a) and (b) according to the balanced equation:

$$\mathrm{Co}(s) + \mathrm{S}(s) \xrightarrow{\Delta} \mathrm{CoS}(s)$$

Experiment	mol Co	mol S	mol CoS
(a) before reaction:	1.00	1.00	0.00
after reaction:			
(b) before reaction:	2.00	3.00	0.00
after reaction:			

58. The following table indicates the amounts of Co and S before reaction. Complete the table after reactions (a) and (b) according to the balanced equation:

$$2\,\mathrm{Co}(s) + 3\,\mathrm{S}(s) \xrightarrow{\Delta} \mathrm{Co_2S_3}(s)$$

Experiment	mol Co	mol S	mol Co_2S_3
(a) before reaction:	1.00	1.00	0.00
after reaction:			
(b) before reaction:	2.00	3.00	0.00
after reaction:			

Limiting Reactant Problems (Sec. 9.8)

59. If 50.0 g of molten iron(II) oxide reacts with 10.0 g of magnesium, what is the mass of iron produced?

$$\mathrm{FeO}(l) + \mathrm{Mg}(l) \xrightarrow{\Delta} \mathrm{Fe}(l) + \mathrm{MgO}(s)$$

60. If 50.0 g of molten iron(II) oxide reacts with of 20.0 g magnesium, what is the mass of iron produced according to the equation in Exercise 59?

61. If 50.0 g of molten iron(III) oxide reacts with 15.0 g of aluminum, what is the mass of iron produced?

$$\mathrm{Fe_2O_3}(l) + \mathrm{Al}(l) \xrightarrow{\Delta} \mathrm{Fe}(l) + \mathrm{Al_2O_3}(s)$$

62. If 50.0 g of molten iron(III) oxide reacts with 25.0 g of aluminum, what is the mass of iron produced according to the equation in Exercise 61?

63. If 1.00 g of magnesium hydroxide reacts with 0.605 g of sulfuric acid, what is the mass of magnesium sulfate produced?

$$\mathrm{Mg(OH)_2}(s) + \mathrm{H_2SO_4}(l) \longrightarrow \mathrm{MgSO_4}(s) + \mathrm{H_2O}(l)$$

64. If 0.605 g of magnesium hydroxide reacts with 1.00 g of sulfuric acid, what is the mass of magnesium sulfate produced according to the equation in Exercise 63?

65. If 1.00 g of aluminum hydroxide reacts with 3.00 g of sulfuric acid, what is the mass of water produced?

$$\mathrm{Al(OH)_3}(s) + \mathrm{H_2SO_4}(l) \longrightarrow \mathrm{Al_2(SO_4)_3}(aq) + \mathrm{H_2O}(l)$$

66. If 3.00 g of aluminum hydroxide reacts with 1.00 g of sulfuric acid, what is the mass of water produced according to the equation in Exercise 65?

67. If 45.0 mL of nitrogen gas reacts with 95.0 mL of oxygen gas, what is the volume of NO_2 produced? Assume all gases are at the same temperature and pressure.

$$\mathrm{N_2}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{NO_2}(g)$$

68. If 95.0 mL of nitrogen gas reacts with 45.0 mL of oxygen gas, what is the volume of NO_2 produced according to the equation in Exercise 67? Assume all gases are at the same temperature and pressure.

69. If 70.0 mL of nitrogen gas reacts with 45.0 mL of oxygen gas, what is the volume of N_2O_3 produced? Assume all gases are at the same temperature and pressure.

$$\mathrm{N_2}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{N_2O_3}(g)$$

70. If 45.0 mL of nitrogen gas reacts with 70.0 mL of oxygen gas, what is the volume of N_2O_3 produced according to the equation in Exercise 69? Assume all gases are at the same temperature and pressure.

71. If 3.00 L of sulfur dioxide gas reacts with 1.25 L of oxygen gas, what is the volume of sulfur trioxide gas produced? Assume all gases are at the same temperature and pressure.

$$\mathrm{SO_2}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{SO_3}(g)$$

72. If 1.25 L of sulfur dioxide gas reacts with 3.00 L of oxygen gas, what is the volume of sulfur trioxide gas produced according to the equation in Exercise 71? Assume all gases are at the same temperature and pressure.

73. If 50.0 L of hydrogen chloride gas reacts with 10.0 L of oxygen gas, what is the volume of chlorine gas produced? Assume all gases are at the same temperature and pressure.

$$\mathrm{HCl}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{Cl_2}(g) + \mathrm{H_2O}(g)$$

74. If 10.0 L of hydrogen chloride gas reacts with 50.0 L of oxygen gas, what is the volume of chlorine gas produced according to the equation in Exercise 73? Assume all gases are at the same temperature and pressure.

Percent Yield (Sec. 9.9)

75. A chemistry student prepared sodium carbonate by heating 1.50 g of baking soda. If the student collected 0.988 g of Na_2CO_3 and the theoretical yield is 0.946 g, what is the percent yield?

76. A chemistry student prepared lead(II) iodide by precipitating 10.0 g of lead(II) nitrate with aqueous potassium iodide. If the student collected 12.5 g of PbI_2 and the theoretical yield is 13.9 g, what is the percent yield?

77. A 1.50-g sample of sodium nitrate is decomposed by heating. If the resulting sodium nitrite has a mass of 1.29 g and the calculated yield is 1.22 g, what is the percent yield?

78. A 1.00-g sample of potassium bicarbonate is decomposed by heating. If the resulting potassium carbonate weighs 0.725 g and the calculated yield is 0.690 g, what is the percent yield?

General Exercises

79. What are the units of molar mass?

80. What are the units of molar volume?

81. The "volcano reaction" produces chromium(III) oxide, water, and nitrogen gas from the decomposition of ammonium dichromate. Calculate the mass of chromium(III) oxide produced by the decomposition of 1.54 g of ammonium dichromate.

82. What STP volume of nitrogen gas is produced in the "volcano reaction" from the decomposition of 1.54 g of ammonium dichromate according to the reaction described in Exercise 81?

83. Aluminum metal and manganese(IV) oxide react to produce aluminum oxide and manganese metal. Calculate the mass of aluminum metal that yields 1.00 kg of manganese metal.

84. What mass of manganese(IV) oxide reacts to produce 1.00 kg of manganese metal according to the reaction described in Exercise 83?

85. The mineral stibnite, antimony(III) sulfide, is treated with hydrochloric acid to give antimony(III) chloride and hydrogen sulfide gas. Calculate the STP volume of H_2S produced from a 3.00-g sample of stibnite.

86. What mass of $SbCl_3$ is produced from a 3.00-g sample of stibnite according to the reaction described in Exercise 85?

Challenge Exercises

87. Recreational vehicles use propane for cooking and heating. The combustion of propane and oxygen produces carbon dioxide and water. What mass of water is produced from the combustion of 10.0 g of propane, C_3H_8?

88. What STP volume of carbon dioxide gas is produced from the combustion of 10.0 g of propane according to the reaction described in exercise 87?

89. When an electric current is passed through water, hydrogen and oxygen gases are released. What STP volume of hydrogen gas is released from the electrolysis of 100.0 mL of water? (*Hint:* The density of water is 1.00 g/mL.)

90. What STP volume of oxygen gas is released from the electrolysis of 100.0 mL of water according to the reaction described in exercise 89?

Chapter 9 Self-Test

Answers to Self-Test are in Appendix J.

1. Oxygen gas in air reacts with 0.115 g hydrogen gas to give 1.024 g of water. Use the conservation of mass law to predict the mass of reacting oxygen. (Sec. 9.1)

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$$

(a) 0.115 g **(b)** 0.909 g
(c) 1.024 g **(d)** 1.139 g
(e) none of the above

2. How many moles of water decompose to yield 0.500 mol of oxygen gas? (Sec. 9.2)

$$2\,H_2O(l) \longrightarrow 2\,H_2(g) + O_2(g)$$

(a) 0.250 mol **(b)** 0.500 mol
(c) 1.00 mol **(d)** 2.00 mol
(e) none of the above

3. Classify the following type of stoichiometry problem: *How many cubic centimeters of oxygen are produced from decomposing 5.00 mL of steam?* (Sec. 9.3)

(a) mass–mass problem **(b)** mass–volume problem
(c) volume–volume problem **(d)** mole–mole problem
(e) none of the above

4. What mass of oxygen gas is produced from decomposing 5.00 g of water (18.02 g/mol)? (Sec. 9.4)

$$2\,H_2O(l) \longrightarrow 2\,H_2(g) + O_2(g)$$

(a) 2.22 g **(b)** 2.50 g
(c) 4.44 g **(d)** 5.00 g
(e) 8.88 g

5. What STP volume of hydrogen gas is produced from decomposing 5.00 g of water (18.02 g/mol)? (Sec. 9.5)

$$2\,H_2O(l) \longrightarrow 2\,H_2(g) + O_2(g)$$

(a) 1.55 L **(b)** 3.11 L
(c) 5.00 L **(d)** 6.22 L
(e) 12.4 L

6. What volume of oxygen gas is released from decomposing water if 5.00 mL of hydrogen gas is produced (assume constant temperature and pressure)? (Sec. 9.6)

$$2\,H_2O(l) \longrightarrow 2\,H_2(g) + O_2(g)$$

(a) 2.50 mL **(b)** 5.00 mL
(c) 7.50 mL **(d)** 10.0 mL
(e) 20.0 mL

7. Consider the limiting reactant concept and predict how many moles of steam are released from the reaction of 5.00 mol of hydrogen gas and 5.00 mol of oxygen gas (assume constant temperature and pressure)? (Sec. 9.7)

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

(a) 2.50 mol **(b)** 5.00 mol
(c) 10.0 mol **(d)** 15.0 mol
(e) none of the above

8. Consider the limiting reactant concept and predict how many liters of steam are released from the reaction of 5.00 L of hydrogen gas and 5.00 L of oxygen gas (assume constant temperature and pressure)? (Sec. 9.8)

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

(a) 2.50 L **(b)** 5.00 L
(c) 10.0 L **(d)** 15.0 L
(e) none of the above

9. The decomposition of 5.00 g of water releases 4.00 g of oxygen gas. If the calculated mass of oxygen is 4.44 g, what is the percent yield? (Sec. 9.9)

$$2\,H_2O(l) \longrightarrow 2\,H_2(g) + O_2(g)$$

(a) 80.0% **(b)** 88.8%
(c) 90.1% **(d)** 111%
(e) 125%

Key Concepts

10. What is required before applying unit analysis to solve a stoichiometry problem?

11. How many unit factors are typically required to solve a mass–mass problem?

12. How many unit factors are typically required to solve a mass–volume problem?

13. How many unit factors are typically required to solve a volume–volume problem?

Critical Thinking

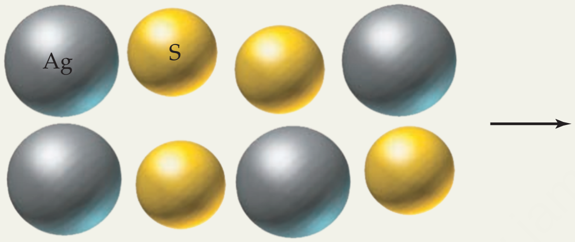
14. Gasoline and air undergo a combustion reaction in an automobile engine. What is the limiting reactant? What is the excess reactant?

15. Silver metal heated with yellow sulfur powder gives Ag_2S.

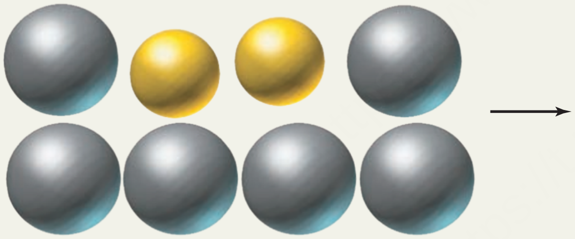
$$2\ Ag(s) + S(s) \xrightarrow{\Delta} Ag_2S(s)$$

Complete the following illustrations by drawing the product and excess reactant. Indicate the element that is the limiting reactant in each example.

(a)

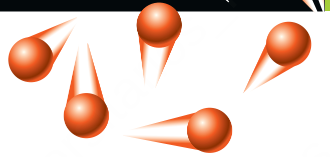


(b)



CHAPTER 10

Gases



"When we attempt to conceive the number of particles in an atmosphere, it is somewhat like attempting to conceive the number of stars in the universe."

John Dalton, English Chemist (1766–1844)

10
Ne
neon

Element 10: Neon
Neon is a colorless, odorless, inert gas, which is abundant in the universe, but rare in the Earth's atmosphere. Neon is often used in advertising lights and emits a bright, reddish-orange glow. Although advertising lights are often referred to as "neon lights," other gases are used to produce different colors of light.

English chemist Robert Boyle is best known for establishing the scientific method and his work with gases. He studied the atmosphere and found that a volume of air could be compressed. He envisioned air as being composed of invisible particles that can be squeezed closer together. Furthermore, he found that other gases could be compressed and reasoned that all gases were composed of tiny, discrete particles. His experiments were remarkable because they offered the first evidence of the particle nature of matter. In 1803, John Dalton proposed the atomic theory, which in part was supported by Boyle's experiments with gases.

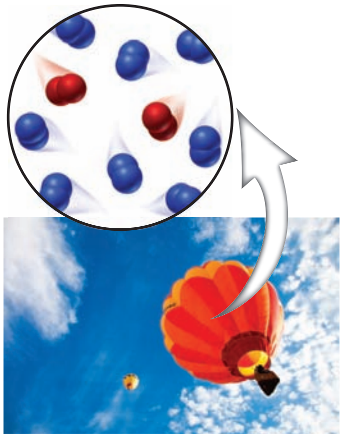
10.1 Properties of Gases

LEARNING OBJECTIVE

- List five observed properties of a gas.

In the late 1700s, high-altitude ballooning became popular in France. Even French scientists filled balloons with hydrogen gas and ascended miles into the sky. Ballooning in turn helped to create a great deal of interest in the study of gases. After numerous experiments, scientists concluded that all gases have common characteristics. We can summarize these properties as follows.

1. **Gases have a variable shape and volume.** A gas has the same shape and volume as its container. If the volume of the container increases or decreases, the volume of the gas increases or decreases.



▲ **Hot Air Balloon** The hot air inside the balloon is less dense than the air surrounding the balloon. As the less dense air rises, the balloon ascends.

2. **Gases expand uniformly.** When a container increases its volume, a gas expands, and the gas molecules distribute uniformly throughout the container.
3. **Gases compress uniformly.** When a container decreases its volume, a gas compresses, and gas molecules are closer to each other. If the volume of the container is reduced sufficiently, the gas will eventually liquefy.
4. **Gases have a low density.** The density of air is about 1 g/L (0.001 g/mL). The density of water is 1 g/mL. Thus, water is about 1000 times more dense than air.
5. **Gases mix uniformly with other gases in the same container.** Air is a common example of a gaseous mixture. The burning of fossil fuels releases carbon dioxide into the atmosphere where it mixes with the gases in air. During photosynthesis, plants consume carbon dioxide from air and release oxygen gas into the atmosphere.

Although most gases are colorless and invisible, we can observe their behavior indirectly. For example, if we decrease the volume of a gas, we observe an increase in gas pressure. Thus, it is possible to gather information about a gas by studying its properties.

10.2 Atmospheric Pressure

LEARNING OBJECTIVES

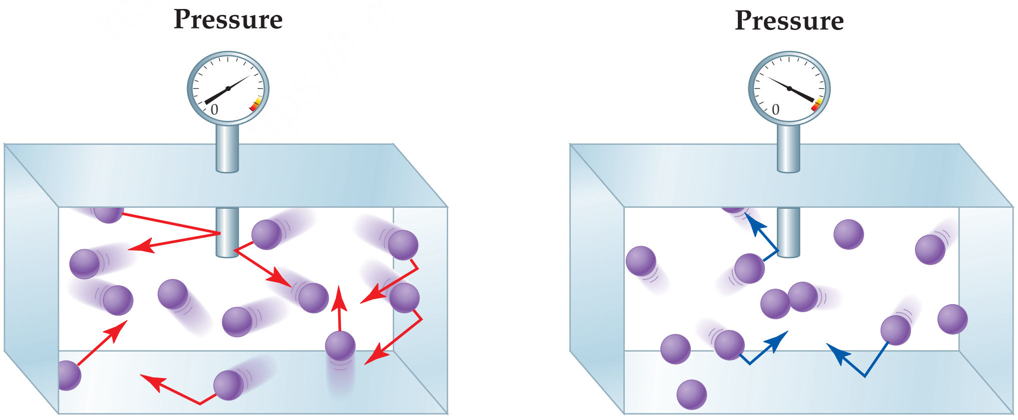
- Memorize standard atmospheric pressure in the following units: atm, mm Hg, torr, cm Hg, in. Hg, psi, and kPa.
- Convert a given gas pressure to a different unit of measurement.

Gas pressure is the result of constantly moving molecules striking the inside surface of their container. The pressure that a gas exerts depends on how often and how hard these molecules strike the walls of the container.

1. If the molecules collide more often, the gas pressure *increases*.
2. If the molecules collide with more energy, the pressure *increases*.

As the temperature increases, the molecules move faster and collide more frequently and with more energy. Thus, the gas pressure increases (Figure 10.1).

▶ **Figure 10.1 Gas Pressure and Temperature** (a) The molecules are moving faster and colliding more frequently with the container. (b) The molecules are moving slower and colliding less frequently with the container.



(a) Higher temperature
Higher pressure

(b) Lower temperature
Lower pressure

CHEMISTRY CONNECTION The Greenhouse Effect

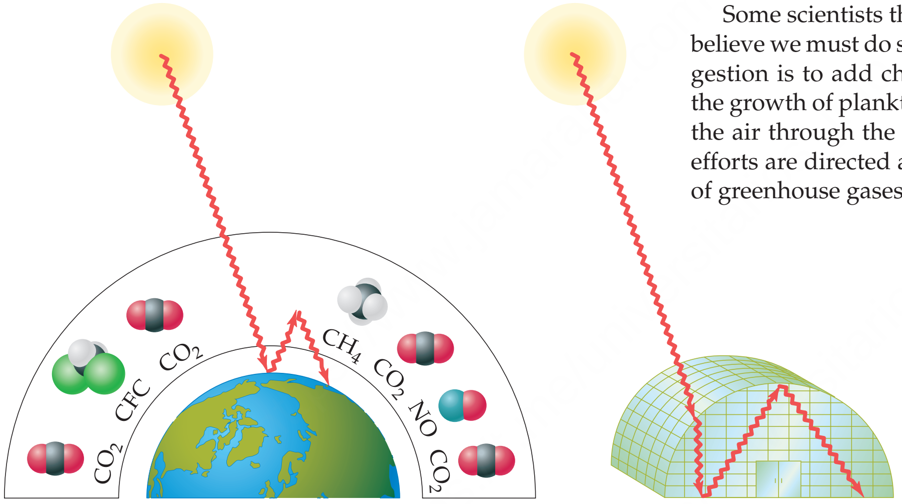
Q: Why has planting millions of trees been suggested as a means of reducing global warming?

When we read about the greenhouse effect, we may be alarmed and in fear that it is life-threatening to our planet. However, the greenhouse effect is necessary for life on Earth as it helps to maintain a moderately mild climate. The current concern is that our atmosphere has an abundance of so-called greenhouse gases that will cause drastic global warming. Although carbon dioxide is singled out as the chief offender, other gases such as chlorofluorocarbons, CFCs, methane, CH_4, and the oxides of nitrogen, NO_x, also contribute to global warming.

Scientists explain the greenhouse effect as follows: Earth's atmosphere is transparent to short-wavelength radiation from the Sun. After striking Earth's surface, the radiation from the Sun is emitted back into space as long-wavelength heat energy. Long-wavelength heat energy can be absorbed by carbon dioxide and other trace gases in the atmosphere. Subsequently, these gas molecules release heat energy in all directions, some of which is radiated back toward Earth. The net effect is that carbon dioxide and other gases in the atmosphere form an invisible canopy that traps heat energy.

Our present concern is that additional gases in the atmosphere are increasing the temperature of our planet at an alarming rate. Global warming can result from high concentrations of carbon dioxide in the atmosphere. Carbon dioxide is produced by burning fossil fuels (coal, oil, and gas) in automobiles, homes, and power plants. The problem is further accelerated by massive deforestation, especially in South America. The loss of forests is especially harmful because trees convert carbon dioxide in the atmosphere to oxygen.

Some scientists think global warming is a severe threat and believe we must do something immediately. One futuristic suggestion is to add chemicals to the oceans that will stimulate the growth of plankton. Plankton remove carbon dioxide from the air through the process of photosynthesis. Currently, our efforts are directed at reforestation and reducing the emission of greenhouse gases into the atmosphere.



(a)

(b)

◀ **Greenhouse Effect** Greenhouse gases trap heat energy from the Sun in much the same way as a plant greenhouse.

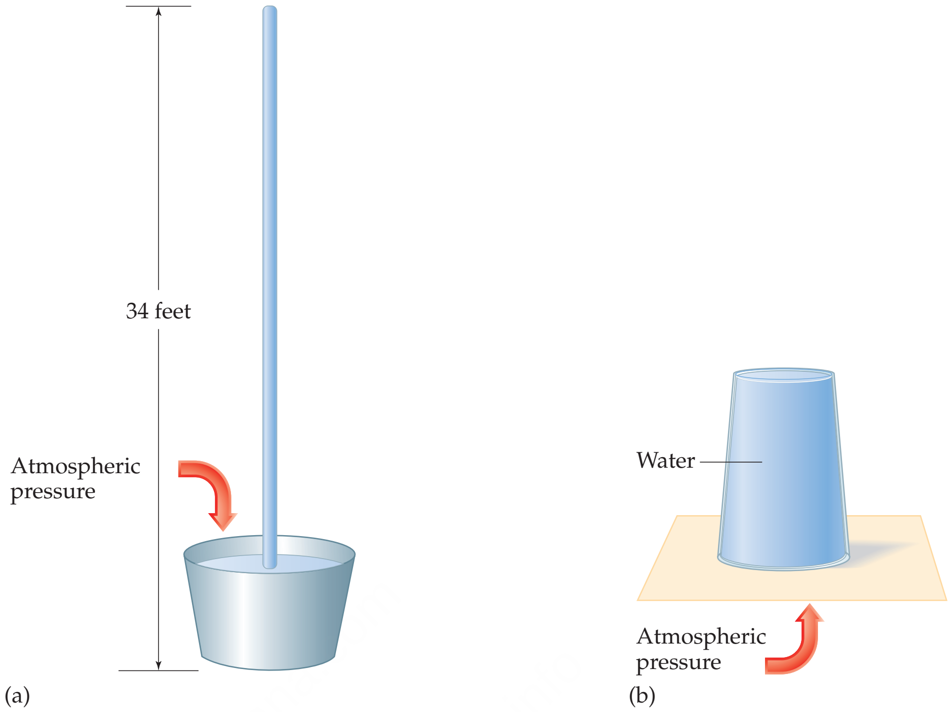
A: The greenhouse gas that contributes most to global warming is carbon dioxide. The planting of millions of trees will allow the process of photosynthesis to reduce the amount of carbon dioxide and return more oxygen to the atmosphere.

Atmospheric Pressure

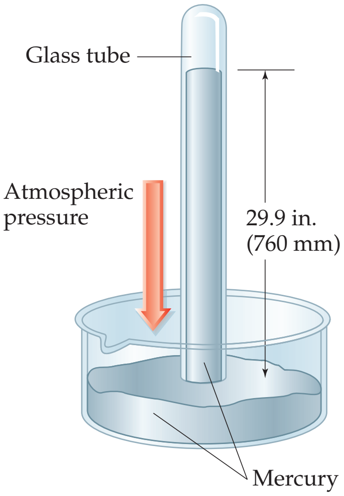
Italian scientist Evangelista Torricelli (1608–1647) published the first scientific explanation of a vacuum. He proposed that a "sea of air" surrounds Earth and argued that this sea of air exerts a pressure on everything it contacts. We now understand that there are no gas molecules in a **vacuum** and that the gas pressure is therefore zero.

By the end of the 1600s, the concept of atmospheric pressure was well established. Today, we understand that **atmospheric pressure** is the result of air molecules in the environment. The pressure of the atmosphere is considerable, about 15 pounds on every square inch of surface. In fact, the atmosphere exerts a total weight of nearly 20 tons on an average-size human body! More effects of atmospheric pressure are illustrated in Figure 10.2.

In 1643, Torricelli invented the **barometer** to measure atmospheric pressure (Figure 10.3). He showed that at sea level, the height of mercury measures 29.9 inches, which we define as standard atmospheric pressure, that is, 1 atmosphere (symbol atm). Above sea level, the height of mercury is less than 29.9 inches, indicating that atmospheric pressure is less at higher elevations.



▶ **Figure 10.2 Atmospheric Pressure** (a) The pressure of air molecules in the atmosphere supports a 34-foot column of water. (b) Because atmospheric pressure operates in all directions, air molecules strike the bottom of the card and support the water in the inverted glass.



▲ **Figure 10.3 Torricelli's Mercury Barometer** A sealed glass tube is filled with liquid mercury, inverted, and placed into a dish of mercury. No gas molecules are above the column of mercury inside the glass tube. The pressure of the atmosphere at sea level supports a column of mercury 29.9 inches (760 mm) high.

We can express standard pressure in many units besides atmospheres and inches of mercury. For example, 29.9 in. of Hg is equivalent to 760 mm Hg. In honor of Torricelli, standard pressure is stated as 760 torr, where 1 torr equals 1 mm Hg. Table 10.1 lists units that are frequently used to express pressure.

TABLE 10.1 Units of Gas Pressure

Unit	Standard Pressure
atmosphere	1 atm (exactly)
inches of mercury	29.9 in. Hg
centimeters of mercury	76 cm Hg (exactly)
millimeters of mercury	760 mm Hg (exactly)
torr[1]	760 torr (exactly)
pounds per square inch	14.7 psi
kilopascal[2]	101 kPa

[1] A torr is defined as exactly equal to 1 mm Hg pressure; thus, standard pressure is 760 torr.
[2] The kilopascal (kPa) is the official SI unit of pressure.

Each of the values in Table 10.1 is an expression for the same standard pressure, so we can relate one unit to another. For example, if the pressure of propane gas in a steel tank is 2550 torr, we can relate the pressure to atmospheres. We can perform the conversion by applying the unit analysis method. The unknown unit is atm, and the given value is 2550 torr. Thus,

$$2550 \text{ torr} \times \frac{\text{unit}}{\text{factor}} = \text{atm}$$

To convert from one unit of pressure to another, we derive a unit factor from the relationship among units of standard pressure. Because 1 atm = 760 torr, the two related unit factors are 1 atm/760 torr, and its reciprocal 760 torr/1 atm. We select the first unit factor to cancel units.

$$2550 \cancel{\text{torr}} \times \frac{1 \text{ atm}}{760 \cancel{\text{torr}}} = 3.36 \text{ atm}$$

The following example exercise further illustrates the conversion of gas pressures.

▲ Adding air to an automobile tire increases the pressure because there are more molecules, which cause an increase in the number of collisions inside the tire.

EXAMPLE EXERCISE 10.1 Gas Pressure Conversion

Meteorologists state that a "falling" barometer indicates an approaching storm. Given a barometric pressure of 27.5 in. Hg, express the pressure in each of the following units of pressure:

(a) atm (b) mm Hg (c) psi (d) kPa

Solution

For each conversion, we apply a unit conversion factor related to units of standard pressure.

(a) To express the pressure in atmospheres, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 1 atm.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{1\ \text{atm}}{29.9\ \cancel{\text{in. Hg}}} = 0.920\ \text{atm}$$

(b) To convert to millimeters of mercury, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 760 mm Hg.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{760\ \text{mm Hg}}{29.9\ \cancel{\text{in. Hg}}} = 699\ \text{mm Hg}$$

(c) To calculate the pressure in pounds per square inch, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 14.7 psi.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{14.7\ \text{psi}}{29.9\ \cancel{\text{in. Hg}}} = 13.5\ \text{psi}$$

(d) To find the pressure in kilopascals, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 101 kPa.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{101\ \text{kPa}}{29.9\ \cancel{\text{in. Hg}}} = 92.9\ \text{kPa}$$

Because 1 atm and 760 mm Hg are exact values, the answers have been rounded to three significant digits, which is consistent with the given value, 27.5 in. Hg.

Practice Exercise

Given that the air pressure inside an automobile tire is 34.0 psi, express the pressure in each of the following units:

(a) atm (b) cm Hg (c) torr (d) kPa

Answers:

(a) 2.31 atm (b) 176 cm Hg (c) 1760 torr (d) 234 kPa

Concept Exercise

Which of the following expresses the greatest gas pressure: 1 atm, 1 in. Hg, 1 torr, 1 cm Hg, 1 mm Hg, or 1 psi?

Answer: See Appendix G, 10.1.

10.3 Variables Affecting Gas Pressure

In Section 10.2, we learned that gas pressure is related to the frequency and energy of gas molecule collisions. Experimentally, there are only three ways to change the number of collisions, and hence the pressure.

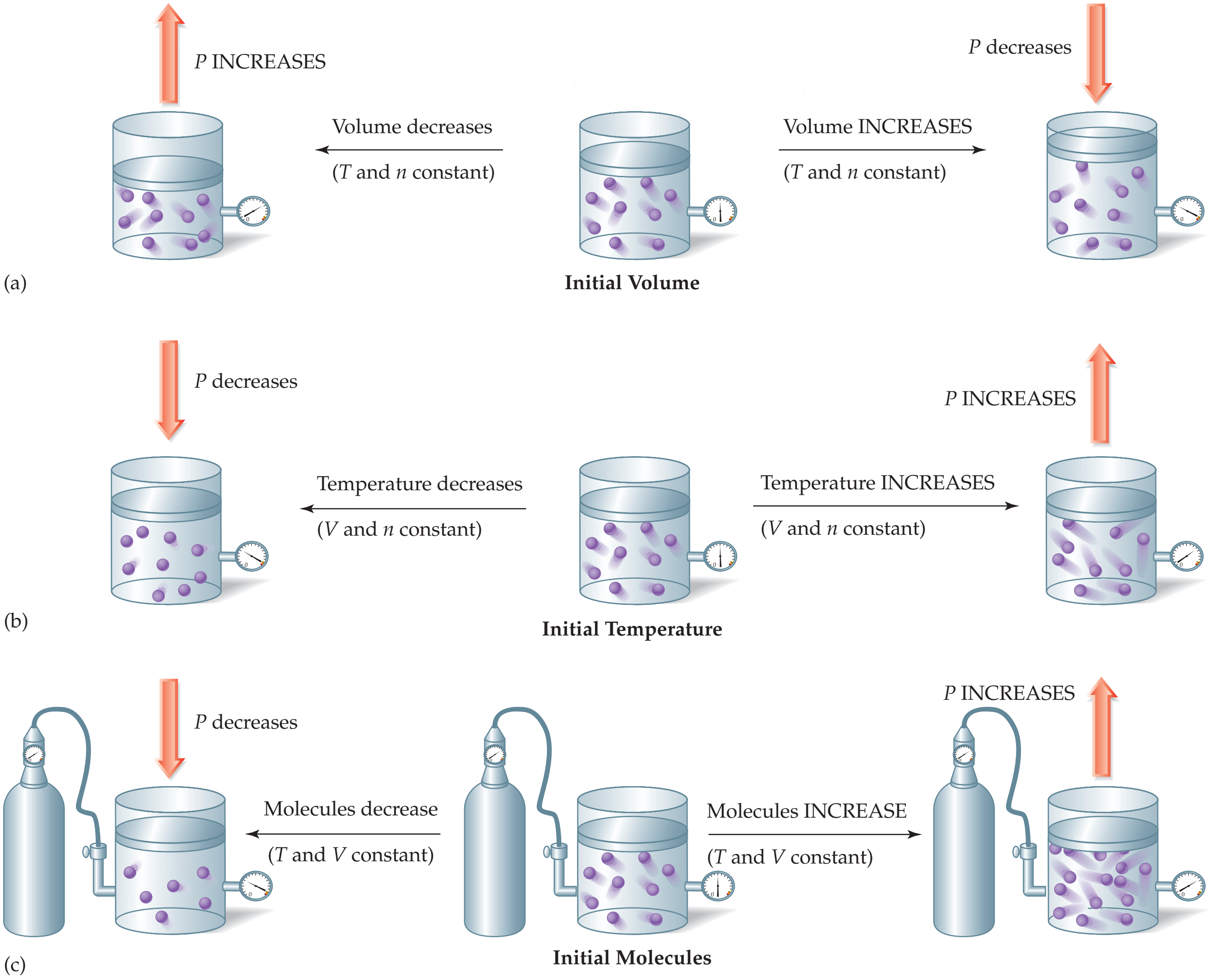
1. **Increase or decrease the volume of the container.** When we increase the volume, gas molecules are farther apart and collide less frequently, and the pressure decreases. When we decrease the volume, gas molecules are closer together and collide more frequently, and the pressure increases. The pressure is said to be *inversely* related to the volume. That is, when the volume increases, the pressure decreases. When the volume decreases, the pressure increases.

LEARNING OBJECTIVES

- Identify the three variables that affect the pressure of a gas.
- Indicate whether gas pressure increases or decreases for a given change in volume, temperature, or the number of moles of gas.

2. **Increase or decrease the temperature of the gas.** When we increase the temperature, gas molecules move faster and collide with a greater frequency and energy. When we decrease the temperature, the gas molecules move slower and collide less frequently. The pressure is said to be *directly* related to the temperature. When the temperature increases, the pressure increases. When the temperature decreases, the pressure decreases.
3. **Increase or decrease the number of molecules in the container. Avogadro's theory**, also called Avogadro's law, states that equal volumes of gases, at the same temperature and pressure, contain equal numbers of molecules. When we increase the number of gas molecules, there are more collisions and the pressure increases. When we decrease the number of gas molecules, there are fewer collisions and the pressure decreases. The pressure is said to be *directly* related to the number of moles of gas, or the number of gas molecules. When the number of gas molecules increases, the pressure increases. When the number of molecules decreases, the pressure decreases.

The effects of these three variables (volume, temperature, and number of molecules) on pressure are illustrated in Figure 10.4.



▲ **Figure 10.4 Variables Affecting Gas Pressure** A gas is contained in a cylinder with a moving piston. The gas pressure, P, is affected by (a) changing volume, V; (b) changing temperature, T; and (c) changing the number of gas molecules in the cylinder.

EXAMPLE EXERCISE 10.2 Gas Pressure Changes

State whether the pressure of a gas in a sealed container increases or decreases with each of the following changes:

(a) The volume changes from 250 mL to 500 mL.
(b) The temperature changes from 20 °C to −80 °C.
(c) The moles of gas change from 1.00 mol to 1.50 mol.

Solution

For each of the changes we must consider whether the number of molecular collisions increases or decreases.

(a) The volume increases, and so the number of collisions decreases; thus the *pressure decreases.*
(b) The temperature decreases, and so the number of collisions decreases; thus, the *pressure decreases.*
(c) When the moles of gas increase, the number of molecules increases. With more molecules, there are more collisions, and the *pressure increases*.

Practice Exercise

Indicate whether gas pressure increases or decreases with each of the following changes in a sealed container:

(a) increasing the temperature
(b) increasing the volume
(c) increasing the number of gas moalecules

Answers:

(a) pressure increases
(b) pressure decreases
(c) pressure increases

Concept Exercise

Which of the following decreases gas pressure: increasing volume, increasing temperature, or increasing the number of molecules?

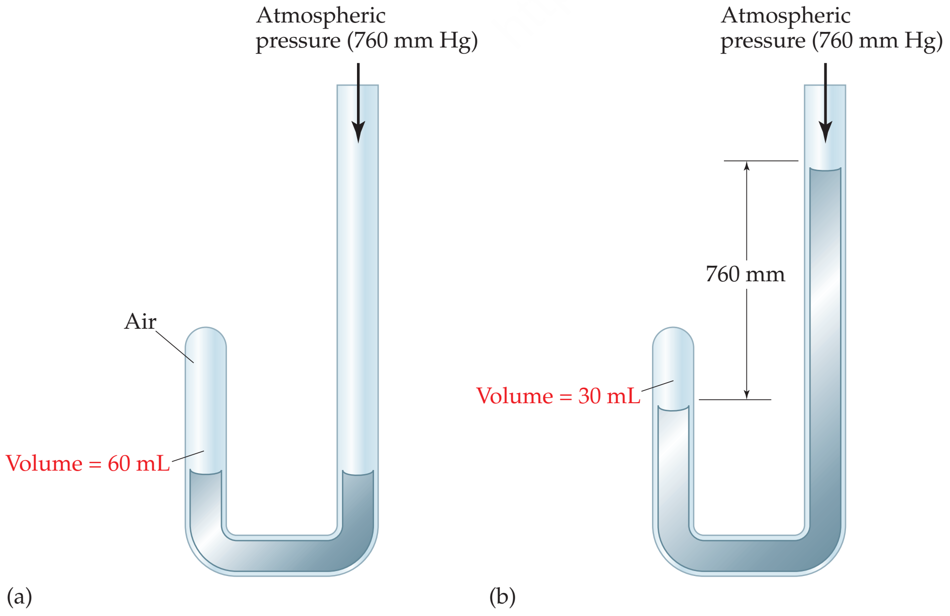
Answer: See Appendix G, 10.2.

10.4 Boyle's Law: Pressure–Volume Relationships

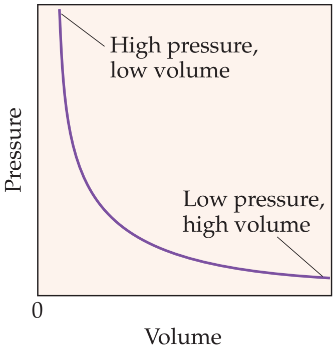
LEARNING OBJECTIVES

- Illustrate a graph of the pressure–volume relationship for a gas.
- Calculate the pressure or volume of a gas after a change in conditions.

We have mentioned that Robert Boyle was the founder of the scientific method. In one of his experiments, Boyle trapped air in a J-shaped tube with liquid mercury. As he added mercury into the tube, he found that the volume of trapped air decreased (Figure 10.5). When he doubled the pressure, the volume was reduced to half. When he tripled the pressure, the volume was reduced to a third. Thus, Boyle established the relationship between the pressure and the volume of a gas.



◀ **Figure 10.5 Boyle's Law** (a) The initial volume of gas at atmospheric pressure is 60 mL. (b) After adding mercury, the volume of gas is 30 mL. Note that the volume is halved when the pressure is doubled. Initially, the pressure is 760 mm Hg; after adding mercury, the pressure is 1520 mm Hg (760 mm Hg + 760 mm Hg).



▲ **Figure 10.6 Gas Pressure versus Volume** As the volume of a gas increases, the pressure decreases. Conversely, as the volume decreases, the pressure increases. The volume and the pressure of a gas are inversely proportional.

The results of Boyle's experiments are known as **Boyle's law**, which states that the volume of a gas is inversely proportional to the pressure when the temperature remains constant. **Inversely proportional** means that two variables have a reciprocal relationship. As one variable gets larger, the other variable gets smaller. We can express Boyle's law mathematically, where the symbol $\propto$ is read as "proportional to."

$$V \propto \frac{1}{P} \qquad (T \text{ remains constant})$$

That is, the volume (V) is proportional to the reciprocal pressure ($1/P$) when the temperature (T) remains constant. Using experimental data, we can plot pressure versus volume and obtain the curve shown in Figure 10.6.

The relationship between pressure and volume can be written as an equation by using a proportionality constant (k):

$$V = k \times \frac{1}{P}$$

Multiplying both sides of the equation by P, we see that the product of pressure and volume is equal to the constant (k). The equation becomes

$$PV = k$$

We can consider a sample of a gas under changing conditions but with temperature remaining the same. Let the initial conditions of pressure and volume be P_1 and V_1, respectively. After a change in conditions, we can indicate the final pressure and volume as P_2 and V_2. Because the product of pressure and volume is a constant, we write

$$P_1 V_1 = k = P_2 V_2$$

Solving Boyle's Law Problems

In this textbook, we will present two methods for solving gas law problems. First, we will solve problems using a conceptual approach and a reasoning method. Second, we will solve problems using an algebraic method.

Because the pressure and the volume of a gas are inversely related, we can apply a proportionality factor to find an unknown variable. To find the gas pressure after a change in volume, we apply a volume factor to the initial pressure (P_1).

$$\boxed{P_1 \times V_{\text{factor}} = P_2}$$

Similarly, to find the volume after a change in gas pressure, we apply a pressure factor to the initial volume (V_1).

$$\boxed{V_1 \times P_{\text{factor}} = V_2}$$

Let's consider the following Boyle's law gas problem. In an experiment, 5.00 L of propane gas is compressed and the pressure increases from 1.00 atm to 1.50 atm. To calculate the final volume (V_2), we write

$$5.00 \text{ L} \times P_{\text{factor}} = V_2$$

The pressure of the propane gas increases from 1.00 atm to 1.50 atm, so the final volume must decrease. Therefore, the P_{factor} must be less than 1. That is, the smaller pressure value must appear in the numerator.

$$5.00 \text{ L} \times \frac{1.00 \cancel{\text{atm}}}{1.50 \cancel{\text{atm}}} = 3.33 \text{ L}$$

Notice that the cancellation of units takes place in the unit factor (P_{factor}). Although the numerator and the denominator are not equal, this method uses a format very much like that of the unit analysis method of problem solving. Now, let's reinforce our understanding of Boyle's law with an example exercise by applying either a conceptual solution or algebraic solution.

EXAMPLE EXERCISE 10.3 Boyle's Law

A 1.50-L sample of methane gas exerts a pressure of 1650 mm Hg. Calculate the final pressure if the volume changes to 7.00 L. Assume the temperature remains constant.

Conceptual Solution

We can find the final pressure (P_2) by applying Boyle's law and using the relationship

$$P_1 \times V_{\text{factor}} = P_2$$

The volume increases from 1.50 L to 7.00 L. Thus, the pressure decreases. The V_{factor} must be less than 1. Hence,

$$1650 \text{ mm Hg} \times \frac{1.50 \cancel{\text{L}}}{7.00 \cancel{\text{L}}} = 354 \text{ mm Hg}$$

We can visually summarize the Boyle's law solution as follows:

Algebraic Solution

Alternatively, we can solve this problem using the equation

$$P_1V_1 = P_2V_2$$

Solving for P_2 gives

$$\frac{P_1V_1}{V_2} = P_2$$

Substituting for each variable and simplifying, we obtain

$$\frac{1650 \text{ mm Hg} \times 1.50 \cancel{\text{L}}}{7.00 \cancel{\text{L}}} = 354 \text{ mm Hg}$$

Practice Exercise

A sample of ethane gas has a volume of 125 mL at 20 °C and 725 torr. What is the volume of the gas at 20 °C when the pressure decreases to 475 torr?

Answer:

191 mL

Concept Exercise

When air in a steel cylinder is compressed from 10 L to 5 L, and temperature remains constant, the gas pressure inside the cylinder (increases/decreases).

Answer: See Appendix G, 10.3.

Note A conceptual solution for solving gas law problems is valuable because it emphasizes reason and a deeper understanding of gaseous behavior. A conceptual solution requires that we understand how changes in pressure, volume, and temperature of a gas affect each other. On the other hand, mathematics is a powerful tool for solving problems, and an algebraic solution is valuable where the concepts are initially difficult to grasp.

FEATURED SCIENTIST Robert Boyle

Q: What instrument did Robert Boyle invent that demonstrated a feather and a lump of lead are affected identically by gravity?

Robert Boyle (1627–1691) was a child prodigy who was studying Latin and Greek by the age of 8. He was the son of wealthy English aristocrats and traveled throughout Europe with a private tutor to gain a broad education. When he was 18 years old, his father died and left him with a lifetime income. At 27, he enrolled in Oxford University to continue his studies. There he became interested in experimentation, although laboratory work was considered to be of minor importance at the time. Most influential scientists believed, like the Greeks, that reason was far superior to experimentation.

In 1657, Boyle designed a vacuum pump that was an improvement over the first one developed a few years earlier. In one experiment, he evacuated most of the air from a sealed chamber and showed that a ticking clock could not be heard in a vacuum. He then correctly concluded that sound does not exist in the absence of air. Boyle was fascinated by the behavior of gases and formulated the law stating that pressure and volume of a gas are inversely related.

In 1661, Boyle published *The Sceptical Chymist* and argued that theories were no better than the experiments on which they were based. Gradually, this point of view was accepted, and his text marked a turning point for recognition of the importance of experimentation. Because of his numerous contributions to chemistry and physics, Boyle is generally regarded as the founder of the modern scientific method.

Boyle felt strongly about keeping meticulous notes and reporting experimental results. He thought that everyone should be able to learn from published accounts of laboratory inquiry. Boyle felt that experiments should be held up to scientific scrutiny and that others should have the opportunity to confirm or disprove the results. This practice has become a cornerstone of science and, except for information involving military and industrial secrets, research experiments are published and made available to the scientific community and to the public at large.

◀ **Robert Boyle (1627–1691)**

A: Boyle invented a vacuum pump to remove air from a cylinder. When he released a feather and a lump of lead from the same height in the cylinder, he found that they landed simultaneously at the bottom of the cylinder.

10.5 Charles's Law: Volume–Temperature Relationships

LEARNING OBJECTIVES

- Illustrate a graph of the volume–temperature relationship for a gas.
- Calculate the volume or temperature of a gas after a change in conditions.

In 1783, the French scientist Jacques Charles (1746–1823) filled a balloon with hydrogen gas and ascended to 10,000 feet—nearly two miles! Although hot air ballooning had been pioneered by the Montgolfier brothers, Jacques Charles was the first to make an ascent in a hydrogen-filled balloon.

In 1787, Charles discovered the effect of temperature on the volume of a gas. From his experiments, he correctly deduced that the volume of a gas is **directly proportional** to the Kelvin temperature. That is, when we double the Kelvin temperature of a gas, we double its volume. When we halve the Kelvin temperature of a gas, we halve its volume. We can state **Charles's law** as follows: The volume of a gas is directly proportional to the Kelvin temperature if the pressure remains constant. Using experimental data, we can plot volume versus Kelvin temperature and obtain the straight-line graph shown in Figure 10.7.

We can express the direct relationship between volume and Kelvin temperature as

$$V \propto T \qquad (P \text{ remains constant})$$

The relationship of volume and Kelvin temperature can also be written as an equation using a proportionality constant (k).

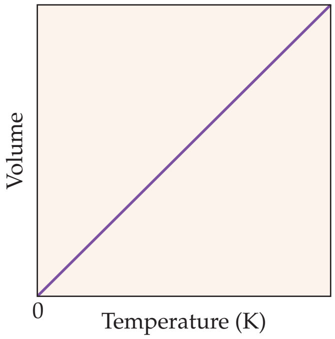
$$V = kT$$

Dividing both sides of the equation by T, we see that the ratio of volume to temperature equals the constant k. The equation becomes

$$\frac{V}{T} = k$$

Let's consider a sample of gas under changing conditions, but with pressure remaining the same. We indicate the initial conditions of volume and temperature as V_1 and T_1, respectively. After a change in conditions, we indicate the final volume and temperature as V_2 and T_2. Because the ratio of volume to temperature equals the constant k, we write

$$\frac{V_1}{T_1} = k = \frac{V_2}{T_2}$$



▲ **Figure 10.7 Gas Volume versus Temperature** As the Kelvin temperature decreases, the volume of a gas decreases. The Kelvin temperature and the volume of a gas are directly proportional.

Solving Charles's Law Problems

The volume and temperature of a gas are directly related, so we can apply a proportionality factor to find an unknown variable. To find the volume after a change in temperature, we apply a temperature factor to the initial volume (V_1):

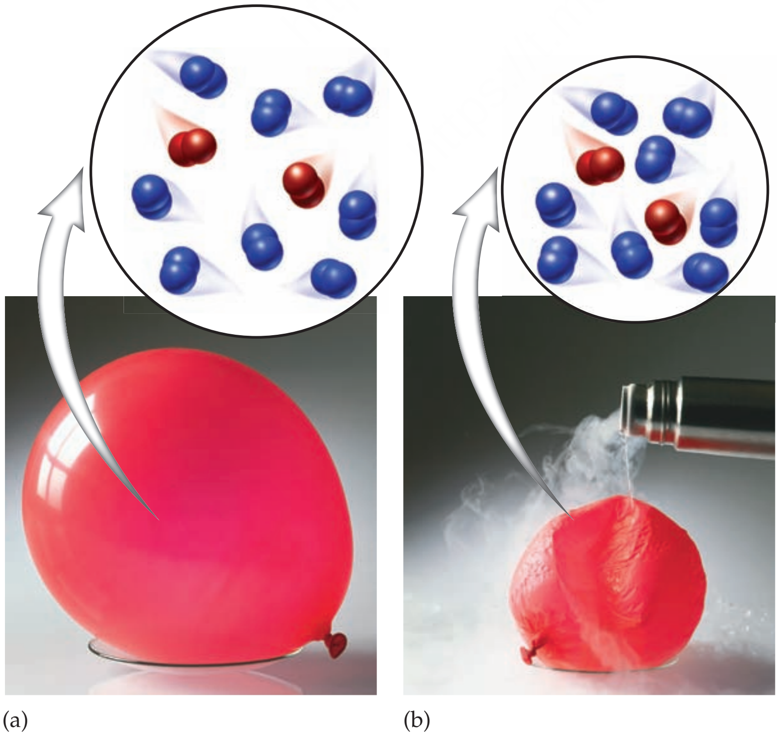
$$\boxed{V_1 \times T_{\text{factor}} = V_2}$$

Similarly, to find the temperature after a change in volume, we apply a volume factor to the initial temperature (T_1):

$$\boxed{T_1 \times V_{\text{factor}} = T_2}$$

Let's consider the following Charles's law gas problem. In an experiment, a sample of argon gas at 225 K is heated and the volume increases from 3.50 L to 12.5 L. To calculate the final temperature (T_2), we write

$$225\text{ K} \times V_{\text{factor}} = T_2$$



◀ **Illustration of Charles's Law** (a) A balloon contains air at room temperature. (b) As liquid nitrogen at −196 °C (77 K) is poured on the balloon, the air inside the balloon cools, causing it to decrease in volume.

Because the volume of argon gas increases from 3.50 L to 12.5 L, the final temperature must increase. Therefore, the V_{factor} must be greater than 1. That is, the larger value must appear in the numerator.

$$225\ \text{K} \times \frac{12.5\ \cancel{\text{L}}}{3.50\ \cancel{\text{L}}} = 804\ \text{K}$$

The final temperature is 804 K, or 531 °C. Now, let's reinforce our understanding of Charles's law with an example exercise by applying either a conceptual solution, or an algebraic solution.

EXAMPLE EXERCISE 10.4 Charles's Law

A 275 L helium balloon is heated from 20 °C to 40 °C. Calculate the final volume assuming the pressure remains constant.

Conceptual Solution

We first convert the Celsius temperatures to Kelvin by adding 273 units.

$$20\ ^\circ\text{C} + 273 = 293\ \text{K}$$
$$40\ ^\circ\text{C} + 273 = 313\ \text{K}$$

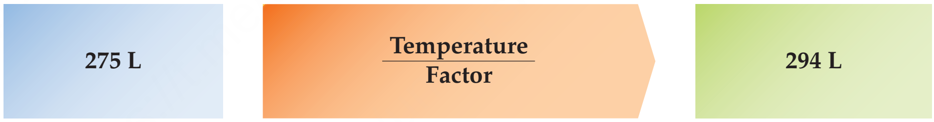
We can find the final volume, V_2, by applying Charles's law and using the relationship

$$V_1 \times T_{\text{factor}} = V_2$$

The temperature increases from 293 K to 313 K. It follows that the volume increases and the T_{factor} must be greater than 1.

$$275\ \text{L} \times \frac{313\ \cancel{\text{K}}}{293\ \cancel{\text{K}}} = 294\ \text{L}$$

We can visually summarize the Charles's law solution as follows:



Algebraic Solution

Alternatively, we can solve this problem using the equation

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Solving for V_2 gives

$$\frac{V_1 T_2}{T_1} = V_2$$

Substituting for each variable and simplifying, we obtain

$$\frac{275\ \text{L} \times 313\ \cancel{\text{K}}}{293\ \cancel{\text{K}}} = 294\ \text{L}$$

Practice Exercise

A krypton balloon has a volume of 555 mL at 21 °C. If the balloon is cooled and the volume decreases to 475 mL, what is the final temperature? Assume that the pressure remains constant.

Answer:
−21 °C (252 K)

Concept Exercise

When air in an elastic balloon cools from 25 °C to 20 °C, the volume of the balloon (increases/decreases).

Answer: See Appendix G, 10.4.

Note In Charles's law problems we can state the volume in any units, but all volumes must be expressed in the same units. For example, we can express the volume in liters or cubic centimeters. However, we must *always* express the temperature in Kelvin units.

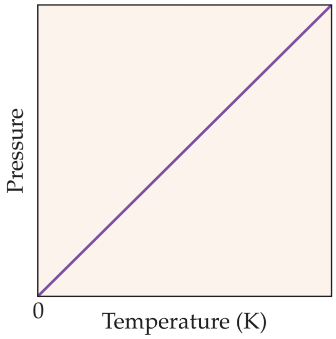
10.6 Gay-Lussac's Law: Pressure–Temperature Relationships

LEARNING OBJECTIVES

- Illustrate a graph of the pressure–temperature relationship for a gas.
- Calculate the pressure or temperature of a gas after a change in conditions.

Like Jacques Charles, French scientist Joseph Gay-Lussac (1778–1850) was also interested in ballooning and made a record ascent to 23,000 feet in a hydrogen-filled balloon. During his ascent, he collected samples of the atmosphere for study in his laboratory.

In 1802, Gay-Lussac established the relationship between pressure and temperature. He correctly interpreted his experiments to mean that the pressure of a gas is directly proportional to the Kelvin temperature. That is, when we double the Kelvin temperature of a gas, we double its pressure. When we halve the Kelvin temperature of a gas, we halve its pressure. We can state **Gay-Lussac's law** as follows: The pressure of a gas is directly proportional to the Kelvin temperature if the volume remains constant. Using experimental data, we can plot pressure versus Kelvin temperature and obtain the straight-line graph shown in Figure 10.8.



▲ **Figure 10.8 Gas Pressure versus Temperature** As the Kelvin temperature decreases, the pressure of a gas decreases. The Kelvin temperature and the pressure exerted by a gas are directly proportional.

We can express the relationship between pressure and Kelvin temperature as

$$P \propto T \qquad (V \text{ remains constant})$$

The relationship between pressure and temperature can also be written as an equation using the proportionality constant (k):

$$P = kT$$

Dividing both sides of the equation by T, we see that the ratio of pressure to temperature equals the constant k. The equation becomes

$$\frac{P}{T} = k$$

Let's consider a sample of a gas under changing conditions, but with volume remaining the same. We indicate the initial conditions of pressure and temperature as P_1 and T_1, respectively. After a change in conditions, we indicate the final pressure and temperature as P_2 and T_2. Because the ratio of pressure to temperature equals the constant (k), we can write

$$\frac{P_1}{T_1} = k = \frac{P_2}{T_2}$$

Solving Gay-Lussac's Law Problems

The pressure and temperature of a gas are directly related, so we can apply a proportionality factor to find an unknown variable. To find the pressure after a change in temperature, we apply a temperature factor to the initial pressure (P_1).

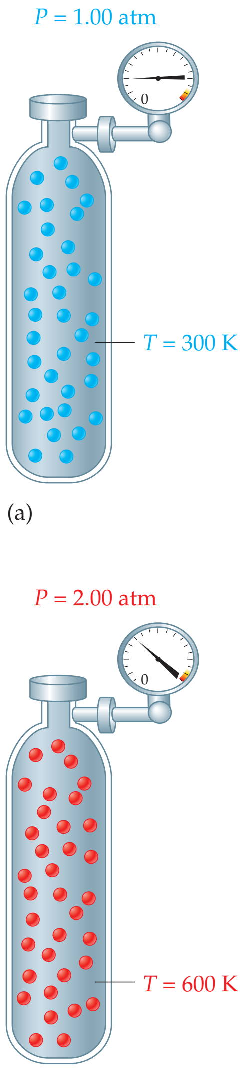
$$\boxed{P_1 \times T_{\text{factor}} = P_2}$$

Similarly, to find the temperature after a change in pressure, we apply a pressure factor to the initial temperature (T_1).

$$\boxed{T_1 \times P_{\text{factor}} = T_2}$$

Let's consider the following problem involving Gay-Lussac's law. An automobile tire is inflated to 28.0 psi at a temperature of 20 °C. After it has traveled at high speed, the tire pressure is 36.0 psi. To calculate the final temperature (T_2), we first convert the initial temperature to Kelvin (20 °C + 273 = 293 K), and then write

$$293\text{ K} \times P_{\text{factor}} = T_2$$



(a)

(b)

▲ **Illustration of Gay-Lussac's Law** (a) A cylinder contains helium gas at 300 K and a pressure of 1.00 atm. (b) After doubling the temperature of the helium gas to 600 K, the pressure doubles to 2.00 atm.

Because the volume of the tire remains constant and the pressure increases from 28.0 psi to 36.0 psi, the final temperature must increase. Therefore, the P_{factor} must be greater than 1. That is, the larger pressure value must appear in the numerator.

$$293\text{ K} \times \frac{36.0\ \cancel{\text{psi}}}{28.0\ \cancel{\text{psi}}} = 377\text{ K}$$

The final temperature is 377 K, or 104 °C. Now, let's reinforce our understanding of Gay-Lussac's law with an example exercise by applying either a conceptual solution or algebraic solution.

EXAMPLE EXERCISE 10.5 Gay-Lussac's Law

A steel container filled with nitrous oxide at 15.0 atm is cooled from 25 °C to −40 °C. Calculate the final pressure, assuming the volume remains constant.

Conceptual Solution

We must first convert the Celsius temperatures to Kelvin by adding 273.

$$25\text{ °C} + 273 = 298\text{ K}$$
$$-40\text{ °C} + 273 = 233\text{ K}$$

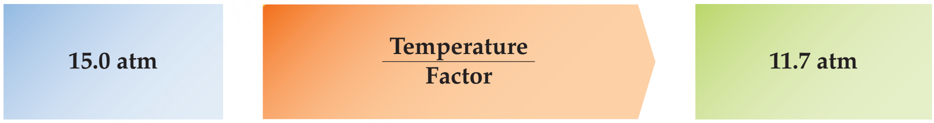
We can find the final pressure (P_2) by applying Gay-Lussac's law and using the relationship

$$P_1 \times T_{\text{factor}} = P_2$$

The volume of the container remains constant, but the temperature decreases from 298 K to 233 K. Therefore, the pressure decreases. The T_{factor} must be less than 1. Hence,

$$15.0\text{ atm} \times \frac{233\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.7\text{ atm}$$

We can visually summarize the Gay-Lussac's law solution as follows:



Algebraic Solution

Alternatively, we can solve this problem using the equation

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Solving for P_2 gives

$$\frac{P_1 T_2}{T_1} = P_2$$

Substituting for each variable and simplifying, we have

$$\frac{15.0\text{ atm} \times 233\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.7\text{ atm}$$

Practice Exercise

A copper container has a volume of 555 mL and is filled with air at 25 °C. The container is immersed in dry ice, and the pressure of the gas drops from 761 torr to 495 torr. What is the final temperature of the air in the copper container?

Answer:

194 K (−79 °C)

Concept Exercise

When air in a rigid steel tank cools from 25 °C to 20 °C, the pressure inside the tank (increases/decreases).

Answer: See Appendix G, 10.5.

10.7 Combined Gas Law

LEARNING OBJECTIVE

- Calculate the pressure, volume, or temperature of a gas after a change in conditions.

We began our discussion of gas laws with a simplifying assumption. In discussing Boyle's, Charles's, and Gay-Lussac's laws, we assumed that we could limit our treatment of gases to two variables. Experimentally, all three variables (pressure, volume, and temperature) usually change simultaneously. Now, we will bring all three variables together in a single expression. The resulting expression is called the **combined gas law**, and the equation is

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Because the variables P, V, and T are proportional, we can solve combined gas law problems using the unit analysis approach. To calculate a final pressure (P_2), we apply a volume factor and a temperature factor to the initial pressure (P_1).

$$\boxed{P_1 \times V_{\text{factor}} \times T_{\text{factor}} = P_2}$$

To calculate a final volume (V_2), we apply a pressure factor and a temperature factor to the initial volume (V_1).

$$\boxed{V_1 \times P_{\text{factor}} \times T_{\text{factor}} = V_2}$$

To calculate a final temperature (T_2), we apply a pressure factor and a volume factor to the initial temperature (T_1).

$$\boxed{T_1 \times P_{\text{factor}} \times V_{\text{factor}} = T_2}$$

In a combined gas law problem, we have three variables at initial and final conditions. That gives us six pieces of data, and so it is helpful to use a table. We place the data in a table and organize the information as follows:

CONDITIONS	P	V	T
initial	P_1	V_1	T_1
final	P_2	V_2	T_2

Let's apply the combined gas law to 10.0 L of carbon dioxide gas at 300 K and 1.00 atm. When both the volume and the Kelvin temperature double, what is the final pressure in atmospheres? The initial conditions are 300 K and 1.00 atm. Doubling the temperature gives 600 K; doubling the volume gives 20.0 L. Now let's enter the data in a table:

CONDITIONS	P	V	T
initial	1.00 atm	10.0 L	300 K
final	P_2	20.0 L	600 K

We calculate the final pressure (P_2) by applying a V_{factor} and a T_{factor} to the initial pressure, 1.00 atm.

$$1.00 \text{ atm} \times V_{\text{factor}} \times T_{\text{factor}} = P_2$$

The volume of the gas increases from 10.0 L to 20.0 L, so P_2 must decrease. The V_{factor} is less than 1 and the smaller value is placed in the numerator.

$$1.00 \text{ atm} \times \frac{10.0 \cancel{\text{L}}}{20.0 \cancel{\text{L}}} \times T_{\text{factor}} = P_2$$

The temperature of the gas increases from 300 K to 600 K; therefore, P_2 increases. The T_{factor} is greater than 1 and the larger value is placed in the numerator.

Helpful Hint
Gas Law Calculations

Previously, we have used the unit analysis method of problem solving. In this chapter we can solve gas law problems conceptually using a variation of the same method. Alternatively, we can solve gas law problems algebraically by solving an equation that relates pressure, volume, and temperature.

The conceptual method of solving gas law problems is preferred because it offers a better understanding of the concepts, although some students have difficulty applying the conceptual method and prefer to solve gas law problems algebraically.

$$1.00\ \text{atm} \times \frac{10.0\ \cancel{\text{L}}}{20.0\ \cancel{\text{L}}} \times \frac{600\ \cancel{\text{K}}}{300\ \cancel{\text{K}}} = 1.00\ \text{atm}$$

In this problem, the pressure did not change even though both the volume and the temperature each doubled. The pressure remained unchanged because pressure and volume are inversely proportional, whereas pressure and temperature are directly proportional. Thus, in effect the volume and temperature changes cancel and the pressure remains constant.

Standard Conditions of Temperature and Pressure

The **standard temperature and pressure** (symbol **STP**) for a gas are 0 °C and 1 atm. We could also express standard temperature as 273 K and standard pressure in other units such as 760 mm Hg, 760 torr, or 76 cm Hg. The following example exercise illustrates how to convert a set of conditions to STP conditions by applying either a conceptual solution or algebraic solution.

EXAMPLE EXERCISE 10.6 Combined Gas Law

A nitrogen gas sample occupies 50.5 mL at −80 °C and 1250 torr. What is the volume at STP?

Conceptual Solution

Although the final conditions are not given, we know that STP conditions are 273 K and 760 torr. We can summarize the information as follows:

CONDITIONS	P	V	T
initial	1250 torr	50.5 mL	−80 °C + 273 = 193 K
final	760 torr	V_2	273 K

We can calculate the final volume by applying a P_{factor} and a T_{factor} to the initial volume.

$$V_1 \times P_{\text{factor}} \times T_{\text{factor}} = V_2$$

The pressure decreases, and so the volume increases; thus, the P_{factor} is greater than 1. The temperature increases, and so the volume increases; thus, the T_{factor} is also greater than 1.

$$50.5\ \text{mL} \times \frac{1250\ \cancel{\text{torr}}}{760\ \cancel{\text{torr}}} \times \frac{273\ \cancel{\text{K}}}{193\ \cancel{\text{K}}} = 117\ \text{mL}$$

We can visually summarize the combined gas law solution as follows:

Algebraic Solution

Alternatively, we can solve this problem using the equation

$$\frac{P_1\,V_1}{T_1} = \frac{P_2\,V_2}{T_2}$$

Rearranging variables and solving for V_2,

$$\frac{P_1\,V_1\,T_2}{T_1\,P_2} = V_2$$

Substituting for each variable and simplifying, we obtain

$$\frac{1250\ \cancel{\text{torr}} \times 50.5\ \text{mL} \times 273\ \cancel{\text{K}}}{193\ \cancel{\text{K}} \times 760\ \cancel{\text{torr}}} = 117\ \text{mL}$$

Practice Exercise

An oxygen gas sample occupies 50.0 mL at 27 °C and 765 mm Hg. What is the final temperature if the gas is cooled to a volume of 35.5 mL and a pressure of 455 mm Hg?

Answer:

127 K (−146°C)

Concept Exercise

A sample of air is at an initial pressure of 1.00 atm. If both the volume and Kelvin temperature each double, what is the final pressure?

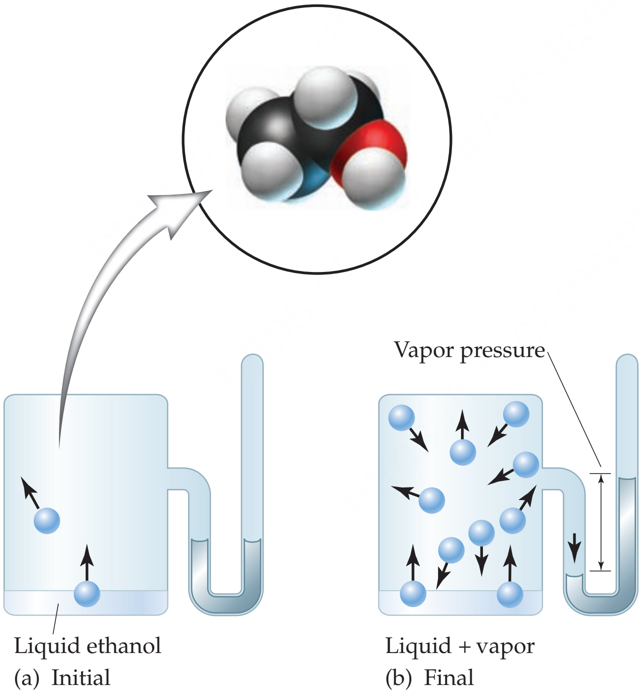
Answer: See Appendix G, 10.6.

10.8 The Vapor Pressure Concept

LEARNING OBJECTIVE

- Describe the relationship between vapor pressure and temperature.

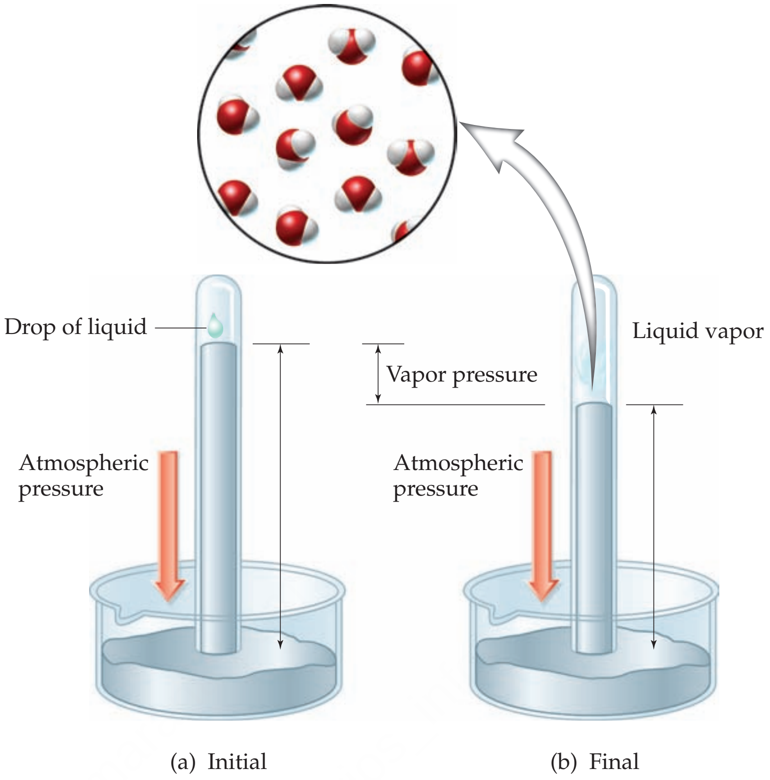
Everyone has observed that water in an open container evaporates. Vaporization occurs when molecules have enough energy to escape from the liquid. When the container is enclosed, the vapor molecules are trapped above the liquid. We should note that molecules continuously escape from the liquid into the vapor. Simultaneously, other molecules return from the vapor to the liquid. The **vapor pressure** is the pressure exerted by the gaseous vapor above a liquid when the rates of evaporation and condensation are equal (Figure 10.9).



◀ **Figure 10.9 Vapor Pressure of a Liquid** (a) Initially, ethanol is placed in a sealed container and molecules begin to escape from the liquid. After awhile, some of the ethanol molecules in the vapor return to the liquid. (b) Eventually, the ethanol molecules are evaporating and condensing at the same rate. The vapor pressure is measured by the difference in the height of mercury in the side tube.

The vapor pressure of a liquid can be determined with a mercury barometer (Figure 10.10). A drop of liquid is introduced at the bottom of the barometer. The less dense liquid droplet floats to the top of the mercury and partially vaporizes. The vaporized liquid exerts a gas pressure, thus driving the column of mercury down. *The decrease in the mercury level corresponds to the increase in vapor pressure.* If the mercury level drops from 760 mm to 740 mm, the vapor pressure of the liquid is recorded as 20 mm Hg.

We can observe that vapor pressure increases as the temperature increases. This is because molecules in the liquid evaporate faster at a higher temperature and the vapor molecules have more kinetic energy. At 25 °C the vapor pressure of water is 23.8 mm Hg. At 50 °C the vapor pressure increases to 92.5 mm Hg. Table 10.2 lists selected values for the vapor pressure of water.



▶ **Figure 10.10 Measuring Vapor Pressure** (a) Initially, a single drop of water is introduced into the tube of mercury. As molecules of the liquid evaporate, the gaseous vapor molecules exert a pressure on the column of mercury. (b) Eventually, molecules are evaporating and condensing at the same rate. The difference in the height of the mercury before and after a drop of liquid is added is the vapor pressure of the liquid at the given temperature.

TABLE 10.2 Vapor Pressure of Water

Temperature (°C)	Pressure (mm Hg)	Temperature (°C)	Pressure (mm Hg)
5	6.5	55	118.0
10	9.2	60	149.4
15	12.8	65	187.5
20	17.5	70	233.7
25	23.8	75	289.1
30	31.8	80	355.1
35	41.2	85	433.6
40	55.3	90	525.8
45	71.9	95	633.9
50	92.5	100	760.0

10.9 Dalton's Law of Partial Pressures

LEARNING OBJECTIVE

- Apply Dalton's law of partial pressures to a mixture of gases.

When John Dalton proposed the atomic theory, his evidence was largely based on experiments with gases. His knowledge of gases was related to his interest in meteorology, and in fact he kept daily records of the weather throughout his life. In 1801, Dalton proposed that each gas in a mixture of gases contributes to the total pressure of the mixture. He then formulated a law, now called **Dalton's law of partial pressures**, that states that the total pressure of a gaseous mixture is equal to the sum of the individual pressures of each gas. We can express this law as follows:

$$P_1 + P_2 + P_3 + \ldots = P_{\text{total}}$$

The pressure exerted by each gas in a mixture of gases is called the **partial pressure**. In the equation, P_1, P_2, and P_3 represent the partial pressures of each of the gases in the mixture; P_{total} symbolizes the total of all the partial pressures.

In 1952, Stanley Miller conducted a historic experiment at the University of Chicago. In an attempt to prove the chemical evolution of life, Miller attempted to simulate Earth's primordial atmosphere. He placed hydrogen, ammonia, methane, and water vapor in a glass container and energized the mixture with ultraviolet light and electric sparks. After a week, Miller analyzed the contents of the glass container. In the "primordial soup," he found amino acids and nucleic acids—the basic building blocks of life! Biologists frequently cite this famous experiment as evidence for the chemical evolution of life. Miller won a Nobel Prize for his work.

We can use Stanley Miller's gas mixture to explore Dalton's law in more detail. Using a mixture of gases, we can write an expression for the total pressure of the gaseous mixture:

$$P_{\text{hydrogen}} + P_{\text{ammonia}} + P_{\text{methane}} + P_{\text{water vapor}} = P_{\text{total}}$$

If the partial pressures of hydrogen, ammonia, and methane were 275 torr, 125 torr, and 340 torr, respectively, what was the partial pressure of the water vapor? Assuming the pressure in the container was standard pressure, 760 torr, we have

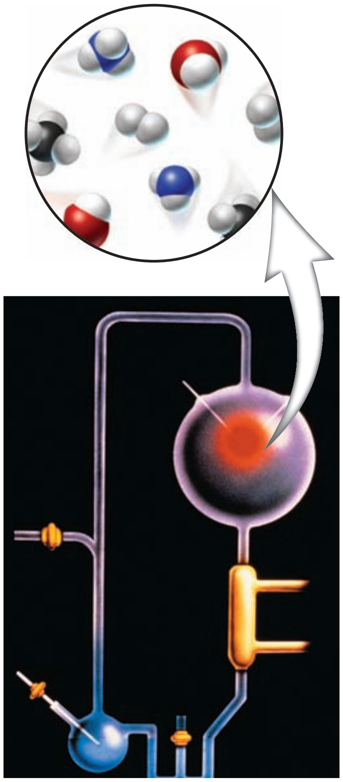
$$275 \text{ torr} + 125 \text{ torr} + 340 \text{ torr} + P_{\text{water vapor}} = 760 \text{ torr}$$

Subtracting the partial pressures of hydrogen, ammonia, and methane from each side of the equation, we have

$$P_{\text{water vapor}} = 760 \text{ torr} - (275 \text{ torr} + 125 \text{ torr} + 340 \text{ torr})$$

$$P_{\text{water vapor}} = 760 \text{ torr} - 740 \text{ torr} = 20 \text{ torr}$$

The following example exercise illustrates an application of Dalton's law.



▲ **Primordial Soup** Stanley Miller placed H_2, NH_3, CH_4, and H_2O vapor in his apparatus and energized the gaseous mixture with ultraviolet light and electric sparks.

EXAMPLE EXERCISE 10.7 Dalton's Law of Partial Pressures

An atmospheric sample contains nitrogen, oxygen, argon, and traces of other gases. If the partial pressure of nitrogen is 587 mm Hg, oxygen is 158 mm Hg, and argon is 7 mm Hg, what is the observed pressure as read on the barometer?

Solution

The sum of the individual partial pressures equals the total atmospheric pressure; therefore,

$$P_{\text{nitrogen}} + P_{\text{oxygen}} + P_{\text{argon}} = P_{\text{total}}$$

Substituting the values for the partial gas pressures, we have

$$587 \text{ mm Hg} + 158 \text{ mm Hg} + 7 \text{ mm Hg} = 752 \text{ mm Hg}$$

Thus, the atmospheric pressure as read on the barometer is 752 mm Hg.

Practice Exercise

The regulator on a steel scuba tank containing compressed air indicates that the pressure is 2250 psi. If the partial pressure of nitrogen is 1755 psi and that of argon is 22 psi, what is the partial pressure of oxygen in the tank?

Answer:

473 psi

Concept Exercise

A rigid steel cylinder contains N_2, O_2, and NO at a total pressure of 2.50 atm. What is the partial pressure of NO gas if N_2 and O_2 are each 1.00 atm?

Answer: See Appendix G, 10.7.

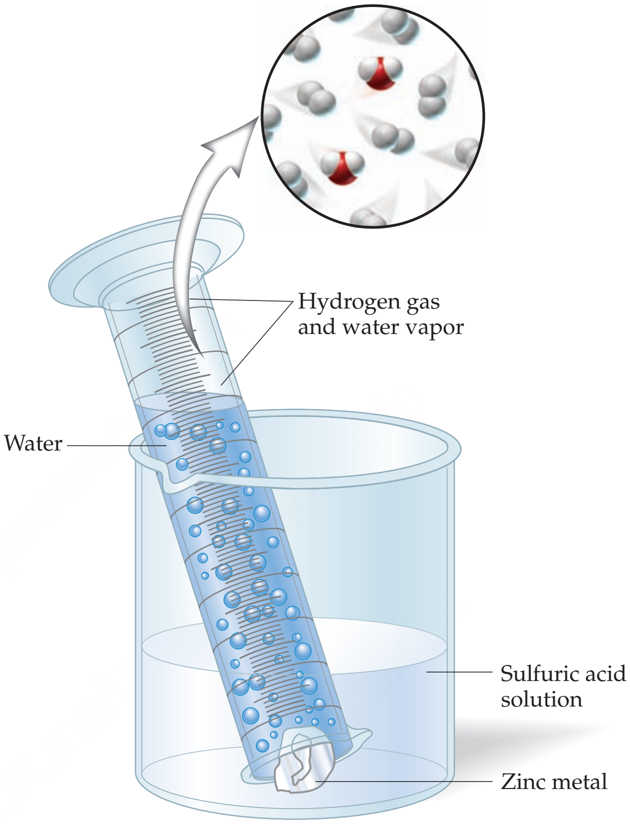
Collecting a Gas over Water

One way of measuring the volume of a gas is to measure the amount of water it displaces. This method is called determining **volume by displacement**. The volume of gas equals the volume of water displaced.

Let's consider a laboratory experiment that produces hydrogen gas from the reaction of zinc metal and sulfuric acid.

$$\mathrm{Zn}(s) + \mathrm{H_2SO_4}(aq) \rightarrow \mathrm{ZnSO_4}(aq) + \mathrm{H_2}(g)$$

To collect the hydrogen gas produced in a laboratory experiment, we can invert a graduated cylinder filled with water over the zinc metal. Figure 10.11 shows the experimental apparatus.



▶ **Figure 10.11 Collecting a Gas over Water** Zinc metal reacts with sulfuric acid to give bubbles of hydrogen gas. A graduated cylinder full of water is placed over the metal to collect the gas bubbles. The volume of water displaced from the graduated cylinder equals the volume of hydrogen gas liberated from the acid.

When the reaction is complete, the pressure exerted by the hydrogen gas inside the graduated cylinder is balanced by the atmospheric pressure outside. That is, the hydrogen gas pressure on the water inside the graduated cylinder is equal to the atmospheric pressure on the surface of the solution in the beaker.

However, there are actually two gases inside the graduated cylinder. Because the hydrogen gas is collected over an aqueous solution, water vapor is also present. We can use Dalton's law to determine the partial pressure of the hydrogen gas.

$$P_{\text{hydrogen}} + P_{\text{water vapor}} = P_{\text{atmosphere}}$$

If the reaction takes place at 25 °C and the barometer reads 767 mm Hg, what is the partial pressure of the hydrogen gas? We refer to Table 10.2 and find that the vapor pressure of water is 23.8 mm Hg at 25 °C. Then we calculate the pressure of the hydrogen gas.

$$P_{\text{hydrogen}} + 23.8 \text{ mm Hg} = 767 \text{ mm Hg}$$

After rearranging, we have

$$P_{\text{hydrogen}} = 767 \text{ mm Hg} - 23.8 \text{ mm Hg} = 743 \text{ mm Hg}$$

Note When a gas is collected over water it is often called a "wet" gas; that is, the collected gas contains water vapor. A gas that does not contain water vapor is sometimes referred to as a "dry" gas.

EXAMPLE EXERCISE 10.8 Dalton's Law and the Combined Gas Law

A 92.5 mL sample of hydrogen gas is collected over water at 25 °C and 767 mm Hg. If the vapor pressure of water is 23.8 mm Hg, what is the volume at STP?

Conceptual Solution

The partial pressure of hydrogen is: 767 mm Hg − 23.8 mm Hg = 743 mm Hg. We can summarize the information in this problem as follows.

Conditions	P	V	T
initial	743 mm Hg	92.5 mL	25 °C + 273 = 298 K
final	760 mm Hg	V_2	273 K

We can calculate the final volume by applying a P_{factor} and a T_{factor} to the initial volume.

$$V_1 \quad \times \quad P_{\text{factor}} \quad \times \quad T_{\text{factor}} \quad = \quad V_2$$

The pressure increases, and so the volume decreases; thus, the P_{factor} is less than 1. The temperature decreases, and so the volume decreases; thus, the T_{factor} is less than 1.

$$92.5\ \text{mL} \times \frac{743\ \cancel{\text{mm Hg}}}{760\ \cancel{\text{mm Hg}}} \times \frac{273\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 82.8\ \text{mL}$$

We can show the combined gas law solution visually as follows.

$$92.5\ \text{mL} \times \frac{\text{pressure}}{\text{factor}} \times \frac{\text{temperature}}{\text{factor}} = 82.8\ \text{mL}$$

Algebraic Solution

Alternatively, we can solve this problem using the equation

$$\frac{P_1\,V_1}{T_1} = \frac{P_2\,V_2}{T_2}$$

Rearranging variables and solving for V_2,

$$\frac{P_1\,V_1\,T_2}{T_1\,P_2} = V_2$$

Substituting for each variable and simplifying, we obtain

$$\frac{743\ \cancel{\text{mm Hg}} \times 92.5\ \text{mL} \times 273\ \cancel{\text{K}}}{298\ \cancel{\text{K}} \times 760\ \cancel{\text{mm Hg}}} = 82.8\ \text{mL}$$

Practice Exercise

If 75.0 mL of oxygen gas is collected over water at 20 °C and 762 mm Hg, what is the pressure at 72.5 mL and 10 °C? The vapor pressure of water at 10 °C is 9.2 mm Hg.

Answer:

752 mm Hg

Concept Exercise

A sample of air is at an initial pressure of 1.00 atm. If both the volume and Kelvin temperature each halve, what is the final pressure?

Answer: See Appendix G, 10.8.

CHEMISTRY CONNECTION Top 10 Industrial Chemicals

Q: What is the most manufactured chemical each year, and is referred to as the "barometer" of the chemical industry?

Each year a list is published of the most produced chemicals worldwide; and each year, sulfuric acid tops the list. The leading producer of sulfuric acid is the United States, followed by China, Russia, Canada, Germany, and Japan. In 2015 over 200,000,000 tons of sulfuric acid was produced. The top chemicals on the list were as follows:

1. sulfuric acid, $H_2SO_4(aq)$
2. nitrogen, $N_2(g)$
3. oxygen, $O_2(g)$
4. ethylene, $C_2H_4(g)$
5. propylene, $C_3H_6(g)$
6. chlorine, $Cl_2(g)$
7. ethylene dichloride, $C_2H_4Cl_2(g)$
8. phosphoric acid, $H_3PO_4(aq)$
9. ammonia, $NH_3(g)$
10. sodium hydroxide, $NaOH(s)$

Concentrated sulfuric acid (18 *M*) is used as a dehydrating agent to remove water vapor from air and other gases that contain moisture. Sulfuric acid is such a strong dehydrating agent that it can remove water from powdered sugar, $C_{12}H_{22}O_{11}$. Pouring concentrated sulfuric acid onto powdered sugar produces a mass of black carbon and steam. The steam contains small particles of powdered sugar that appear as white smoke. The chemical equation for the reaction is as follows:

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2SO_4} 12\ C(s) + 11\ H_2O(g)$$

Because sulfuric acid is the most manufactured chemical each year, sulfuric acid production is a general indicator of chemical manufacturing for any given year. The principle use of sulfuric acid is to make phosphoric acid, which in turn is used to make phosphate fertilizers. Sulfuric acid is found in lead-acid automobile batteries, and has a variety of other uses. For example, sulfuric acid is used in the production of nitroglycerin, which is an explosive. However, nitroglycerin has a medical application as a vasodilator, a substance that dilates blood vessels. Accordingly, nitroglycerin in the form of tablets, sprays, and patches is given to patients for the treatment of heart disease.

A: Sulfuric acid, $H_2SO_4(aq)$, is the most manufactured chemical each year. The amount of acid produced is a "barometer" of chemical productivity worldwide.

LEARNING OBJECTIVES

- List five characteristics of an ideal gas according to the kinetic theory of gases.
- Determine the value of absolute zero from a graph of volume or pressure versus temperature.

10.10 Ideal Gas Behavior

By the early 1800s, experiments had provided a lot of information about the gaseous state. There was, however, no clear way to interpret this information. By 1850, scientists had enough data to begin to construct a theory to explain the behavior of gases. British physicist James Joule (1818–1889) suggested that the temperature of a gas is related to the energy of its molecules. Joule proposed that temperature and the motion of molecules in a gas are related. At higher temperatures, gas molecules move faster, and at lower temperatures, gas molecules move slower.

Kinetic Theory of Gases

In the period from 1850 to 1870, scientists attempted to construct a model that would explain the behavior of individual gas molecules. They reasoned that given a theoretical model, they would be able to explain the behavior of real gases. This model would also allow predictions of ideal gas behavior. An **ideal gas** is a gas that always behaves in a consistent and predictable manner. Experimentally, a **real gas**, such as hydrogen or oxygen, does not behave ideally under all conditions. Specifically, a real gas does not behave ideally at low temperatures and high pressures.

In due time, the behavior of an ideal gas was described by a model called the *kinetic molecular theory*, or simply the **kinetic theory** of gases. According to the kinetic theory, an ideal gas has the following characteristics:

1. **Gases are made up of tiny molecules.** The size of gas molecules is tiny, and the distance between molecules is quite large. Therefore, gases are mostly empty space. Gas molecules occupy zero volume in an ideal gas.
2. **Gas molecules demonstrate rapid motion, move in straight lines, and travel in random directions.**
3. **Gas molecules have no attraction for one another.** After colliding with each other, molecules simply bounce off in different directions.
4. **Gas molecules collide without losing energy.** Two gas molecules are said to have an **elastic collision**. If a high-energy molecule strikes a less energetic molecule, part of the energy can be transferred. However, the total kinetic energy of both molecules, before and after the collision, does not change.
5. **The average kinetic energy of gas molecules is proportional to the Kelvin temperature that is, $KE \propto T$.** At the same temperature, all gas molecules have equal kinetic energy.

Even for different gases, the average kinetic energy is equal at the same temperature. For example, the kinetic energy of hydrogen and oxygen gases is equal at 25 °C. Although the kinetic energy is equal, hydrogen molecules move faster than oxygen molecules because they are lighter. Given the equation for kinetic energy, $KE = \frac{1}{2}mv^2$, we can explain this mathematically. If the kinetic energy of the two gases is equal, then the lighter hydrogen molecules ($m = 2$ amu) must have a greater velocity (v) than the heavier oxygen molecules ($m = 32$ amu).

At higher temperatures, molecules have more kinetic energy, move faster, and collide more frequently. At lower temperatures, molecules have less kinetic energy, move slower, and collide less frequently. The following example exercise further illustrates the characteristics of an ideal gas according to the kinetic theory.

EXAMPLE EXERCISE 10.9 Ideal Gas Behavior

Suppose we have two 5.00-L samples of gas at 25 °C. One sample is ammonia, NH_3, and the other nitrogen dioxide, NO_2. Which gas has the greater kinetic energy? Which gas has the faster molecules?

Solution

Because the temperature of each gas is 25 °C, we know that the kinetic energy is the same for NH_3 and NO_2. At the same temperature, we know that lighter molecules move faster than heavier molecules. The molecular mass of NH_3 is 17 amu and NO_2 is 46 amu. Because NH_3 is lighter than NO_2, the ammonia molecules have a higher velocity than the nitrogen dioxide molecules.

Practice Exercise

Which of the following statements is *not* true according to the kinetic theory of gases?

(a) Molecules occupy a negligible volume.
(b) Molecules move in straight-line paths.
(c) Molecules are attracted to each other.
(d) Molecules undergo elastic collisions.
(e) Molecules of different gases at the same temperature have the same average kinetic energy.

Answer:

All these statements are true except (c). Molecules of an ideal gas are not attracted to each other and behave as independent particles.

Concept Exercise

Which of the following is an example of an ideal gas: H_2, O_2, or He?

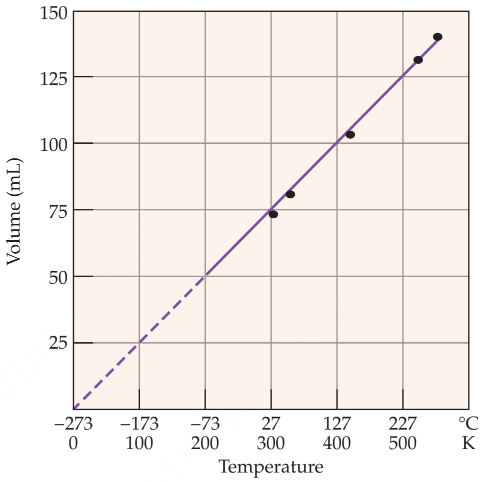
Answer: See Appendix G, 10.9.

Absolute Zero

The temperature at which the pressure and volume of a gas theoretically reach zero is referred to as **absolute zero**. Absolute zero is the coldest possible temperature and corresponds to −273 °C, or 0 K. An ideal gas at absolute zero has no kinetic energy, and therefore no molecular motion.

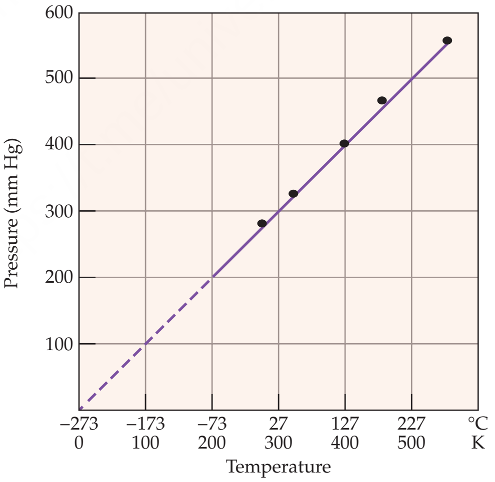
How do we determine the value of absolute zero if −273 °C is impossible to attain experimentally? We determine absolute zero from the data for the volume of a gas over a range of temperatures above 0 K. When we plot the volume and temperature data points and extend the graph to zero volume, we have the temperature corresponding to absolute zero. Figure 10.12 illustrates an extrapolated value for absolute zero.

▶ **Figure 10.12 Absolute Zero at Zero Volume** As the temperature of the gas decreases, the volume decreases. As an ideal gas approaches zero volume, the temperature approaches absolute zero.



We can also determine absolute zero by obtaining data for the pressure and temperature of a gas. We can plot the pressure and temperature data and extrapolate the graph to absolute zero. Figure 10.13 illustrates the determination of absolute zero.

▶ **Figure 10.13 Absolute Zero at Zero Pressure** As the temperature of the gas decreases, the pressure decreases. As an ideal gas approaches zero pressure, the temperature approaches absolute zero.



LEARNING OBJECTIVE

- Calculate the pressure, volume, temperature, or moles of gas from the ideal gas equation.

10.11 Ideal Gas Law

In Section 10.3, we learned that the pressure (P) of a gas is inversely proportional to the volume (V). We also learned that pressure is directly proportional to the number of molecules, or moles of gas (n), and directly proportional to the Kelvin temperature (T). We can therefore write a relationship that pressure is proportional to the moles of gas, multiplied by the temperature, divided by the volume:

$$P \propto \frac{nT}{V}$$

By introducing the proportionality constant, R, we can write this relationship as an equation:

$$P = \frac{RnT}{V}$$

After rearranging,

$$PV = nRT$$

This equation is called the **ideal gas law**. The constant R is the **ideal gas constant** and can be expressed as 0.0821 L · atm/mol · K. To use this value of R in a gas law calculation, the pressure must be expressed in atmospheres, the volume in liters, and the temperature in Kelvin. The ideal gas law is a powerful equation that summarizes the behavior of gases. That is, any change in the pressure, volume, temperature, or amount of a gas (moles of gas) is related through this single equation.

Let's try a calculation using the ideal gas law. If a sealed container holds 1.10 mol of nitrogen gas at 25 °C and 3.75 atm, what is the volume in liters? We need to find volume, so we begin by rearranging the ideal gas equation to solve for volume, V. If $PV = nRT$, then

$$V = \frac{nRT}{P}$$

To convert the temperature to Kelvin, we add 273 to 25 °C which gives 298 K. Now we can substitute for each variable in the ideal gas equation.

$$V = \frac{1.10\ \cancel{\text{mol}} \times \dfrac{0.0821\ \text{L} \cdot \cancel{\text{atm}}}{1\ \cancel{\text{mol}} \cdot \cancel{\text{K}}} \times 298\ \cancel{\text{K}}}{3.75\ \cancel{\text{atm}}} = 7.18\ \text{L}$$

A CLOSER LOOK Conceptualizing Gases

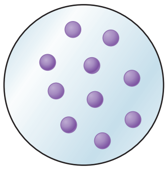
Q: **Describe the distribution of helium atoms in a steel tank containing helium gas at 200 °C, −200 °C, and −273 °C.**

Conceptual Questions

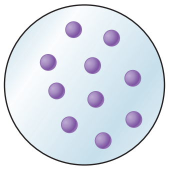
The following diagram represents a steel tank filled with helium gas at 20 °C and 1 atm pressure. The dots represent helium atoms in the steel container.

Initial Conditions:

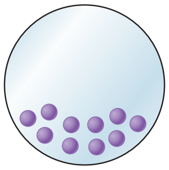
Diagram of a steel tank with helium gas at 20 °C and 1 atm pressure.



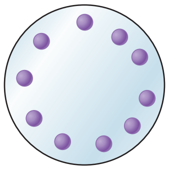
(1) Which of the following diagrams represents the steel container filled with helium gas heated to 200 °C?



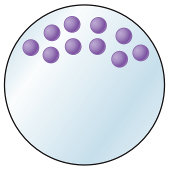
(a)



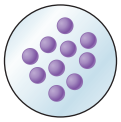
(b)



(c)

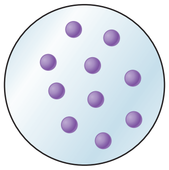


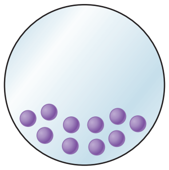
(d)

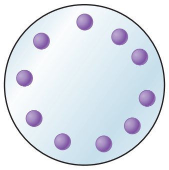


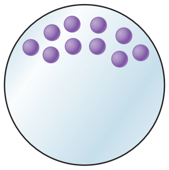
(e)

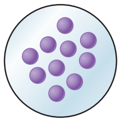
(2) Which of the following diagrams represents the steel container filled with helium gas cooled to −200 °C? (The boiling point of helium is −269 °C.)


(a)

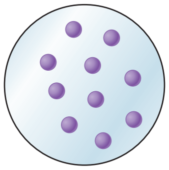

(b)

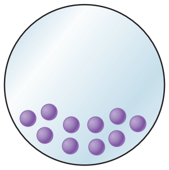

(c)

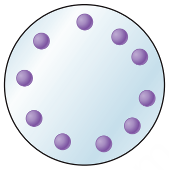

(d)

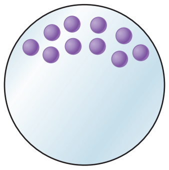

(e)

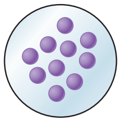
(3) Which of the following diagrams represents the steel container filled with helium gas cooled to −273 °C? (Recall the concept of absolute zero.)


(a)


(b)


(c)


(d)


(e)

A: (1) At 200 °C, the helium atoms are moving faster, but still distributed randomly as shown in (a). (2) At −200 °C, the atoms are moving slower, but still distributed randomly as shown in (a). (3) At absolute zero, −273 °C, the helium atoms are not moving and the gas condenses as shown in (b).

The following example exercise further illustrates the ideal gas law.

EXAMPLE EXERCISE 10.10 Ideal Gas Law

How many moles of hydrogen gas occupy a volume of 0.500 L at STP?

Solution

We begin by rearranging the ideal gas equation, $PV = nRT$, and solving for n:

$$n = \frac{PV}{RT}$$

The temperature at STP is 273 K, and the pressure is 1 atm. Then we substitute for each variable in the ideal gas equation:

$$n = \frac{1\ \cancel{\text{atm}} \times 0.500\ \cancel{\text{L}}}{\dfrac{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}}{1\ \text{mol} \cdot \cancel{\text{K}}} \times 273\ \cancel{\text{K}}} = 0.0223\ \text{mol}$$

Alternatively, we can display the value for the ideal gas constant separately to more clearly show the cancellation of units:

$$n = \frac{1\ \cancel{\text{atm}} \times 0.500\ \cancel{\text{L}}}{273\ \cancel{\text{K}}} \times \frac{1\ \text{mol} \cdot \cancel{\text{K}}}{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}} = 0.0223\ \text{mol}$$

Practice Exercise

What is the temperature of 0.250 mol of chlorine gas at 655 torr if the volume is 3.50 L?

Answer:

147 K (−126°C)

Concept Exercise

Solve for the proportionality constant, R, in the ideal gas law.

Answer: See Appendix G, 10.10.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises

10.1 Properties of Gases

The shape and volume of a gas are variable and occupy the entire volume of a sealed container. A gas can expand and compress, and the gaseous state has a density much less than that of the liquid state. A gas mixes completely with other gases in the same container.

- List five observed properties of a gas.
  Related Exercises: 1–2

10.2 Atmospheric Pressure

Gas pressure results from gas molecules striking the inside surface of their container. The pressure exerted by a gas depends on how often and how hard the molecules strike the container. In a **vacuum**, there are no gas molecules and the pressure is zero. **Atmospheric pressure** results from molecules in the air and is measured with a **barometer**. Standard pressure can be expressed as 1 atm, 760 mm Hg, 760 **torr**, 76 cm Hg, 29.9 in. Hg, 14.7 psi, or 101 kPa.

- Memorize standard atmospheric pressure in the following units: atm, mm Hg, torr, cm Hg, in. Hg, psi, and kPa.
  Related Exercises: 3–4
- Convert a given gas pressure to a different unit of measurement.
  Related Exercises: 5–8

10.3 Variables Affecting Gas Pressure

We can increase gas pressure by changing one of three variables. First, if we *decrease the volume*, molecules will collide more frequently with the inner surface of the container and the pressure will increase. Second, if we *increase the temperature*, molecules will collide more frequently with the walls of the container and the pressure will increase. Third, if we *increase the number of molecules*, there will be more collisions against the walls of the container and the pressure will increase. **Avogadro's theory** states that equal volumes of gases, at the same temperature and pressure, contain equal numbers of molecules.

- Identify the three variables that affect the pressure of a gas.
  Related Exercises: 9–10
- Indicate whether gas pressure increases or decreases for a given change in volume, temperature, or number of moles of gas.
  Related Exercises: 11–14

10.4 Boyle's Law: Pressure–Volume Relationships

The relationship between the pressure and the volume of a gas is **inversely proportional**. In solving problems involving **Boyle's law**, the pressure increases when the volume decreases, while the temperature remains constant.

- Illustrate a graph of the pressure–volume relationship for a gas.
  Related Exercises: 15–16
- Calculate the pressure or volume of a gas after a change in conditions.
  Related Exercises: 17–20

10.5 Charles's Law: Volume–Temperature Relationships

The relationship between the volume and the temperature of a gas is **directly proportional**. In solving problems involving **Charles's law**, the volume increases when the temperature increases, while the pressure remains constant.

- Illustrate a graph of the volume–temperature relationship for a gas.
  Related Exercises: 21–22
- Calculate the volume or temperature of a gas after a change in conditions.
  Related Exercises: 23–26

10.6 Gay-Lussac's Law: Pressure–Temperature Relationships

The relationship between the pressure and the temperature of a gas is directly proportional. In solving problems involving **Gay-Lussac's law**, the pressure increases when the temperature increases, while the volume remains constant. We can summarize all the relationships for the gas laws as shown in Table 10.3.

- Illustrate a graph of the pressure–temperature relationship for a gas.
  Related Exercises: 27–28
- Calculate the pressure or temperature of a gas after a change in conditions.
  Related Exercises: 29–32

TABLE 10.3 Summary of Gas Law Variables

Gas Law	Pressure	Volume	Temperature
Boyle's	increases	decreases	*constant*
	decreases	increases	*constant*
Charles's	*constant*	increases	increases
	constant	decreases	decreases
Gay-Lussac's	increases	*constant*	increases
	decreases	*constant*	decreases

Key Concepts	Learning Objectives and Related Exercises
10.7 Combined Gas Law **Standard temperature and pressure** (**STP**) is 273 K and exactly 1 atm. In solving problems involving the **combined gas law**, simultaneous changes occur in pressure, volume, and temperature.	• Calculate the pressure, volume, or temperature of a gas after a change in conditions. *Related Exercises: 33–42*
10.8 The Vapor Pressure Concept **Vapor pressure** is the pressure exerted by gaseous vapor above a liquid in a sealed container, when the rates of evaporation and condensation are equal. A familiar example is the pressure of H_2O molecules in the vapor above water in a sealed bottle. When the temperature increases, the vapor pressure increases. Conversely, when the temperature decreases, the vapor pressure decreases.	• Describe the relationship between vapor pressure and temperature. *Related Exercises: 43–46*
10.9 Dalton's Law of Partial Pressures **Dalton's law of partial pressures** states that the total pressure of gases in a sealed container is equal to the sum of the **partial pressure** of each gas in the mixture. We can collect a gas over water to determine the **volume by displacement**. A "wet" gas collected over water contains water vapor. The partial pressure of the gas is found by subtracting the vapor pressure of water from the total pressure.	• Apply Dalton's law of partial pressures to a mixture of gases. *Related Exercises: 47–54*
10.10 Ideal Gas Behavior According to the **kinetic theory**, a gas is made up of individual molecules behaving ideally. Molecules in an **ideal gas** move about rapidly and randomly in straight-line paths. Gas molecules show no attraction for one another and undergo **elastic collisions**. The kinetic energy is directly proportional to the Kelvin temperature. At **absolute zero** (0 K), the pressure of an ideal gas is zero. The velocity of gas molecules increases with the temperature and decreases with the molecular mass. The behavior of a **real gas** can deviate substantially at low temperatures and high pressures from that predicted by the kinetic theory.	• List five characteristics of an ideal gas according to the kinetic theory of gases. *Related Exercises: 55–56* • Determine the value of absolute zero from a graph of volume or pressure versus temperature. *Related Exercises: 57–62*
10.11 Ideal Gas Law A more mathematical method for solving gas problems involves the **ideal gas law**. In the equation $PV = nRT$, the **ideal gas constant** (symbol $\boldsymbol{R}$) has a value of $0.0821\ \mathrm{L \cdot atm/mol \cdot K}$.	• Calculate the pressure, volume, temperature, or moles of gas from the ideal gas equation. *Related Exercises: 63–66*

Problem–Solving Organizer

Topic	Procedure	Example
Properties of Gases Sec. 10.1	Predict the shape and volume of a gas.	The shape and volume of argon gas are variable.
Atmospheric Pressure Sec. 10.2	Express a given gas pressure in units of atm, mm Hg, torr, cm Hg, in. Hg, psi, or kPa.	If argon gas is at 655 mm Hg, what is the pressure in atm? $655\ \cancel{\text{mm Hg}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{mm Hg}}} = 0.862\ \text{atm}$
Variables Affecting Gas Pressure Sec. 10.3	Predict an increase or decrease in gas pressure for a change in volume, temperature, or number of molecules.	An increase in which variable causes a decrease in pressure? Answer: *volume*
Boyle's Law Sec. 10.4	Calculate a final pressure after a change in volume (temperature constant). Pressure and volume are inversely proportional.	If a gas is at 1.00 atm, what is the pressure when the volume decreases from 5.00 L to 2.00 L? $1.00\ \text{atm} \times \frac{5.00\ \cancel{\text{L}}}{2.00\ \cancel{\text{L}}} = 2.50\ \text{atm}$

Problem–Solving Organizer *(Continued)*

Topic	Procedure	Example
Charles's Law Sec. 10.5	Calculate a final volume after a change in temperature (pressure constant). Volume and temperature are directly proportional.	If a gas occupies 50.0 mL, what is the volume when the gas is heated from 300 K to 450 K? $50.0\text{ mL} \times \frac{450\ \cancel{\text{K}}}{300\ \cancel{\text{K}}} = 75.0\text{ mL}$
Gay-Lussac's Law Sec. 10.6	Calculate a final pressure after a change in temperature (volume constant). Pressure and temperature are directly proportional.	If a gas is at 1.00 atm, what is the pressure when the gas is heated from 300 K to 450 K? $1.00\text{ atm} \times \frac{450\ \cancel{\text{K}}}{300\ \cancel{\text{K}}} = 1.50\text{ atm}$
Combined Gas Law Sec. 10.7	Calculate a final volume after a change in pressure and temperature. Volume and pressure are inversely proportional; volume and temperature are directly proportional.	If 50.0 mL of gas is at 1.50 atm and 450 K, what is the volume at 1.00 atm and 273 K? $50.0\text{ mL} \times \frac{1.50\ \cancel{\text{atm}}}{1.00\ \cancel{\text{atm}}} \times \frac{273\ \cancel{\text{K}}}{450\ \cancel{\text{K}}} = 45.5\text{ mL}$
Dalton's Law Sec. 10.9	Calculate the total pressure of a gas mixture by summing the individual partial pressures.	If a gas sample contains He at 0.25 atm and Ar at 1.50 atm, what is the total pressure? $P_1 + P_2 + P_3 + \ldots = P_{\text{total}}$ $0.25\text{ atm} + 1.50\text{ atm} = 1.75\text{ atm}$

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** a measure of the frequency and energy of molecules colliding against the walls of a container

______ **2.** a volume of space that does not contain gas molecules

______ **3.** the pressure exerted by the molecules in air

______ **4.** an instrument for measuring atmospheric pressure

______ **5.** a unit of pressure equal to 1 mm Hg

______ **6.** the statement that equal volumes of gases, under the same conditions of temperature and pressure, contain equal numbers of molecules

______ **7.** a relationship between two variables such that when one variable doubles, the other variable halves

______ **8.** a relationship between two variables such that when one variable doubles, the other variable doubles

______ **9.** the statement that the pressure and volume of a gas are inversely proportional at constant temperature

______ 10. the statement that the volume and the Kelvin temperature of a gas are directly proportional at constant pressure

______ **11.** the statement that the pressure and the Kelvin temperature of a gas are directly proportional at constant volume

______ **12.** the statement that the pressure exerted by a gas is inversely proportional to its volume and directly proportional to its Kelvin temperature

______ **13.** the conditions of a gas at 273 K and 760 mm Hg

______ **14.** the pressure exerted by gaseous molecules above a liquid in a sealed container when the rates of evaporation and condensation are equal

(a) absolute zero (*Sec. 10.10*)
(b) atmospheric pressure (*Sec. 10.2*)
(c) Avogadro's theory (*Sec. 10.3*)
(d) barometer (*Sec. 10.2*)
(e) Boyle's law (*Sec. 10.4*)
(f) Charles's law (*Sec. 10.5*)
(g) combined gas law (*Sec. 10.7*)
(h) Dalton's law of partial pressures (*Sec. 10.9*)
(i) directly proportional (*Sec. 10.5*)
(j) elastic collision (*Sec. 10.10*)
(k) gas pressure (*Sec. 10.2*)
(l) Gay-Lussac's law (*Sec. 10.6*)
(m) ideal gas (*Sec. 10.10*)
(n) ideal gas constant (*Sec. 10.11*)
(o) ideal gas law (*Sec. 10.11*)
(p) inversely proportional (*Sec. 10.4*)

______ **15.** the statement that the pressure exerted by a mixture of gases is equal to the sum of the individual pressures exerted by each gas
______ **16.** the pressure exerted by an individual gas in a mixture of two or more gases
______ **17.** a technique for determining the volume of gas from the volume of water it displaces
______ **18.** a theoretical description of gas molecules demonstrating ideal behavior
______ **19.** a theoretical gas that obeys the kinetic theory under all conditions
______ **20.** a gas that deviates from ideal behavior at low temperature and high pressure
______ **21.** an impact between gas molecules with no change in total energy
______ **22.** the theoretical temperature at which the kinetic energy of a gas is zero
______ **23.** the principle stated by the relationship $PV = nRT$
______ **24.** the proportionality constant R in the equation $PV = nRT$

(q) kinetic theory (*Sec. 10.10*)
(r) partial pressure (*Sec. 10.9*)
(s) real gas (*Sec. 10.10*)
(t) standard temperature and pressure (*Sec. 10.7*)
(u) torr (Sec. 10.2)
(v) vacuum (*Sec. 10.2*)
(w) vapor pressure (*Sec. 10.8*)
(x) volume by displacement (*Sec. 10.9*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Properties of Gases (Sec. 10.1)

1. Which of the following are observed properties of gases?
   (a) variable shape **(b)** expand uniformly
   (c) compress infinitely **(d)** mix with other gases
2. Which of the following are observed properties of gases?
   (a) variable volume **(b)** expand infinitely
   (c) compress uniformly **(d)** low density

Atmospheric Pressure (Sec. 10.2)

3. Give the value for standard atmospheric pressure in each of the following units.
   (a) millimeters of mercury **(b)** pounds per square inch
4. Give the value for standard atmospheric pressure in each of the following units.
   (a) centimeters of mercury **(b)** inches of mercury

◀ **The Atmosphere** The atmospheric pressure decreases at high altitudes. The atmospheric pressure on Mt. Everest is about one-third that at sea level.

5. If oxygen gas in a steel container is at a pressure of 2.50 atm, what is the pressure expressed in each of the following units?
   (a) mm Hg **(b)** psi
6. If oxygen gas in a steel container is at a pressure of 5.00 atm, what is the pressure expressed in each of the following units?
   (a) cm Hg **(b)** in. Hg
7. An American newscast states that the barometer reads 30.8 in. Hg. Express the atmospheric pressure in each of the following units.
   (a) kPa **(b)** torr
8. A Canadian newscast states that the barometer reads 104 kPa. Express the atmospheric pressure in each of the following units.
   (a) cm Hg **(b)** psi

Variables Affecting Gas Pressure (Sec. 10.3)

9. Indicate what happens to the pressure of a gas with the following changes:
   (a) The volume increases.
   (b) The temperature increases.
   (c) The moles of gas increase.
10. Indicate what happens to the pressure of a gas with the following changes:
    (a) The volume decreases.
    (b) The temperature decreases.
    (c) The moles of gas decrease.
11. State whether the pressure of a gas in a sealed container increases or decreases with the following changes.
    (a) The volume changes from 25.0 mL to 50.0 mL.
    (b) The temperature changes from 25 °C to 50 °C.
    (c) The moles of gas change from 0.250 mol to 0.500 mol.
12. State whether the pressure of a gas in a sealed container increases or decreases with the following changes.
    (a) The volume changes from 75.0 mL to 50.0 mL.
    (b) The temperature changes from 0 °C to −80 °C.
    (c) The moles of gas change from 5.00 mol to 2.50 mol.
13. Explain how increasing the volume of a gas decreases its pressure.

14. Explain how increasing the temperature of a gas increases its pressure.

◀ **Tire Pressure** The pressure inside the bicycle tire increases as the number of air molecules increases.

Boyle's Law: Pressure–Volume Relationships (Sec. 10.4)

15. Sketch a graph of pressure versus volume, assuming temperature is constant. Label the vertical axis P and the horizontal axis V.

16. Sketch a graph of pressure versus inverse volume, assuming temperature is constant. Label the vertical axis P and the horizontal axis 1/V.

17. A sample of air at 5.00 atm expands from 1.75 L to 2.50 L. If the temperature remains constant, what is the final pressure in atm?

18. A sample of air at 15.0 psi compresses from 555 mL to 275 mL. If the temperature remains constant, what is the final pressure in psi?

19. A 5.00 L sample of argon gas changes pressure from 1.55 atm to 6.50 atm. If the temperature remains constant, what is the final volume in liters?

20. A 125 mL sample of argon gas changes pressure from 705 mm Hg to 385 mm Hg. If the temperature remains constant, what is the final volume in liters?

Charles's Law: Volume–Temperature Relationships (Sec. 10.5)

21. Sketch a graph of volume versus Kelvin temperature, assuming pressure is constant. Label the vertical axis V and the horizontal axis T(K).

22. Sketch a graph of volume versus Celsius temperature, assuming pressure is constant. Label the vertical axis V and the horizontal axis t(°C). Assume the Celsius temperature approaches zero at the origin.

23. A 555 mL sample of nitrous oxide at 25 °C is heated to 50 °C. If the pressure remains constant, what is the final volume?

24. A 2.50 L sample of nitric oxide gas at 100 °C is cooled to 20 °C. If the pressure remains constant, what is the final volume?

25. A 5.00 L sample of fluorine gas at 50 °C is cooled to a volume of 3.75 L. If the pressure remains constant, what is the final Celsius temperature?

26. A 125 mL sample of chlorine gas at 25 °C is heated to a volume of 175 mL. If the pressure remains constant, what is the final Celsius temperature?

Gay-Lussac's Law: Pressure–Temperature Relationships (Sec. 10.6)

27. Sketch a graph of pressure versus Kelvin temperature, assuming volume is constant. Label the vertical axis P and the horizontal axis T (K).

28. Sketch a graph of pressure versus Celsius temperature, assuming volume is constant. Label the vertical axis P and the horizontal axis t(°C). Assume the Celsius temperature approaches zero at the origin.

29. A 1.00-L sample of argon gas at 1.00 atm is heated from 20 °C to 110 °C. If the volume remains constant, what is the final pressure?

30. A 25.0-mL sample of neon gas at 455 mm Hg is cooled from 100 °C to 10 °C. If the volume remains constant, what is the final pressure?

31. A 10.0 L sample of oxygen gas at 20 °C and 0.500 atm is heated to a pressure of 0.750 atm. If the volume remains constant, what is the final Celsius temperature?

32. A 25.0 mL sample of ozone gas at 25 °C and 0.500 atm is cooled to a pressure of 0.250 atm. If the volume remains constant, what is the final Celsius temperature?

Combined Gas Law (Sec. 10.7)

33. A 5.00-L sample of air is collected at 500 °C and 5.00 atm. What is the volume of air at STP?

34. A 100.0-mL sample of air is collected at 21 °C and 772 mm Hg. What is the volume of air at STP?

35. If a sample of hydrogen gas occupies 2.00 L at −50 °C and 155 mm Hg, what is the volume at 75 °C and 365 mm Hg?

36. If a sample of oxygen gas occupies 25.0 mL at −25 °C and 650 mm Hg, what is the volume at 25 °C and 350 mm Hg?

37. A sample of hydrogen chloride gas occupies 0.750 L at STP. What is the pressure in atm if the volume is 0.100 L at 25 °C?

38. A sample of hydrogen fluoride gas occupies 1250 mL at STP. What is the pressure in mm Hg if the volume is 255 mL at 300 °C?

39. A sample of nitrogen gas has a volume of 1.00 L at STP. What is the Celsius temperature if the volume is 10.0 L at 2.00 atm?

40. A sample of ammonia gas has a volume of 50.0 mL at STP. What is the Celsius temperature if the volume is 350.0 mL at 350 mm Hg?

41. A sample of argon gas occupies 2.75 L at STP. What is the pressure in atm, if the gas expands to 4.55 L and the temperature is 45 °C?

42. A sample of xenon gas occupies 500.0 mL at STP. What is the Celsius temperature, if the gas contracts to 250.0 mL and the pressure is 655 mm Hg?

The Vapor Pressure Concept (Sec. 10.8)

43. Which of the following liquids has the higher vapor pressure at 25 °C: water or mercury?

44. Which of the following liquids has the higher vapor pressure at 50 °C: alcohol or mercury?

45. Refer to Table 10.2 and state the vapor pressure for water in mm Hg at each of the following temperatures:
(a) 25 °C **(b)** 50 °C

46. Refer to Table 10.2 and state the vapor pressure for water in atm at each of the following temperatures:
(a) 75 °C **(b)** 100 °C

Dalton's Law of Partial Pressures (Sec. 10.9)

47. Air contains nitrogen, oxygen, argon, and trace gases. If the partial pressure of nitrogen is 592 mm Hg, oxygen is 160 mm Hg, argon is 7 mm Hg, and trace gas is 1 mm Hg, what is the atmospheric pressure?

48. Air contains nitrogen, oxygen, argon, and trace gases. If the partial pressure of nitrogen is 587 mm Hg, oxygen is 158 mm Hg, argon is 7 mm Hg, and trace gas pressure is negligible, what is the atmospheric pressure?

49. A steel cylinder with sulfur dioxide, sulfur trioxide, and oxygen gases is at 825 °C and 1.00 atm. If the partial pressure of sulfur dioxide is 150 mm Hg and sulfur trioxide is 475 mm Hg, what is the partial pressure of oxygen in mm Hg?

50. A steel cylinder with nitrogen, hydrogen, and ammonia gases is at 500 °C and 5.00 atm. If the partial pressure of nitrogen is 1850 mm Hg and hydrogen is 1150 mm Hg, what is the partial pressure of ammonia in mm Hg?

51. Describe the meaning of the expression "collecting a gas over water."

52. Distinguish between a "wet" gas and a "dry" gas.

53. If oxygen is collected over water at 20 °C and 766 mm Hg, what is the partial pressure of the oxygen? Refer to Table 10.2 for the vapor pressure of water.

54. If ozone is collected over water at 30 °C and 775 mm Hg, what is the partial pressure of the ozone? Refer to Table 10.2 for the vapor pressure of water.

Ideal Gas Behavior (Sec. 10.10)

55. What is the pressure exerted by an ideal gas at absolute zero?

56. What is the volume occupied by an ideal gas at absolute zero?

57. What is the kinetic energy of an ideal gas at absolute zero?

58. What are the conditions of temperature and pressure for a real gas to behave most like an ideal gas?

59. State five characteristics of an ideal gas according to the kinetic theory.

60. Distinguish between a real gas and an ideal gas.

61. A sealed container holds the gases He, Ne, and Ar. Which of these gases corresponds to each of the following descriptions?
(a) highest kinetic energy **(b)** lowest kinetic energy
(c) fastest atoms **(d)** slowest atoms

62. A sealed container holds the gases H_2, N_2, and O_2. Which of these gases corresponds to each of the following descriptions?
(a) highest kinetic energy **(b)** lowest kinetic energy
(c) fastest molecules **(d)** slowest molecules

Ideal Gas Law (Sec. 10.11)

63. How many moles of helium gas occupy 5.00 L at 373 K and 1.75 atm?

64. How many moles of helium gas occupy 10.0 L at 100 K and 3.50 atm?

65. If 0.500 mol of helium gas occupies 5.00 L at 25 °C, what is the pressure in atmospheres?

66. If 1.25 mol of helium gas occupies 5.00 L at 2.25 atm, what is the temperature in degrees Celsius?

General Exercises

67. What is standard atmospheric pressure (in cm) measured by a barometer filled with water? (Given: Mercury is 13.6 times more dense than water.)

68. What is standard atmospheric pressure (in cm) measured by a barometer filled with alcohol? (Given: Mercury is 17.2 times more dense than alcohol.)

69. A deep-sea scuba tank contains oxygen and helium gases. If the partial pressures of the gases are 35 atm and 125 atm, what is the total pressure (in psi)?

70. A bicycle tire contains nitrogen, oxygen, and argon gases. If the partial pressures are 51.0 psi, 13.5 psi, and 0.5 psi, what is the total pressure (in atm)?

71. The decomposition of baking soda produces carbon dioxide gas. If 42.5 mL of wet gas is collected over water at 22 °C and 764 mm Hg, what is the volume of dry gas at STP conditions? (The vapor pressure of water at 22 °C is 19.8 mm Hg.)

72. Calcium metal and water produce hydrogen gas. If the volume of wet gas collected over water is 95.0 mL at 20 °C and 758 mm Hg, what is the volume of dry gas at STP conditions? (The vapor pressure of water at 20 °C is 17.5 mm Hg.)

73. What pressure (in atm) is exerted by 1.51×10^{23} oxygen molecules at 25 °C in a 5.00 L container?

74. What volume (in liters) is occupied by 3.38×10^{22} nitrogen molecules at 100 °C and 255 mm Hg?

75. Natural gas contains methane, CH_4, ethane, C_2H_6, and hydrogen sulfide, H_2S. Which molecules have the fastest velocity?

76. Bottled gas for recreational vehicles contains propane, C_3H_8, and butane, C_4H_{10}. Which molecules have the faster velocity?

77. A sample of chlorine gas occupies 1550 mL at 0.945 atm and 50 °C. What is the mass of the sample?

78. A sample of fluorine gas occupies 855 mL at 710 mm Hg and 155 °C. What is the mass of the sample?

Challenge Exercises

79. If the surface area of a human body is 255 square inches, what is the total weight of the atmosphere on a human body expressed in pounds?

80. If the surface area of a human hand is 165 square centimeters, what is the total weight of the atmosphere on a human hand expressed in pounds?

81. Which of the following gases has a density of 2.14 g/L at STP: H_2, O_2, O_3?

82. Which of the following gases has a density of 5.40 g/L at STP: H_2S, SO_3, CF_2Cl_2?

83. A sample of unknown gas has a mass of 1.95 g and occupies 3.00 L at 1.25 atm and 20 °C. What is the molar mass of the unknown gas?

84. A sample of unknown gas has a mass of 2.85 g and occupies 750 mL at 760 mm Hg and 100 °C. What is the molar mass of the unknown gas?

Chapter 10 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following is *not* an observed property of gases? (Sec. 10.1)
(a) gases vary in shape and volume
(b) gases expand infinitely
(c) gases compress infinitely
(d) gases have low density
(e) gases mix completely

2. Which of the following expresses standard atmospheric pressure? (Sec. 10.2)
(a) 29.9 in. Hg **(b)** 76.0 cm Hg
(c) 760 mm Hg **(d)** 14.7 psi
(e) all of the above

3. Which of the following changes increases the pressure of a gas? (Sec. 10.3)
(a) increasing the volume
(b) decreasing temperature
(c) increasing the number of molecules
(d) all of the above
(e) none of the above

4. A volume of hydrogen gas at 1.00 atm decreases from 0.250 L to 0.125 L. If the temperature remains constant, what is the final pressure? (Sec. 10.4)
(a) 0.250 atm **(b)** 0.500 atm
(c) 1.00 atm **(d)** 2.00 atm
(e) none of the above

5. A 1.00 L volume of oxygen gas is cooled from 50.0 °C to 25.0 °C. If the pressure remains constant, what is the final volume? (Sec. 10.5)
(a) 0.500 L **(b)** 0.923 L
(c) 1.00 L **(d)** 1.08 L
(e) 2.00 L

6. A sample of nitrogen gas at 1.00 atm is heated from 250 K to 500 K. If the volume remains constant, what is the final pressure? (Sec. 10.6)
(a) 0.250 atm **(b)** 0.500 atm
(c) 1.00 atm **(d)** 2.00 atm
(e) none of the above

7. A sample of air occupies a volume of 10.0 L. If both the pressure and Kelvin temperature each double, what is the final volume? (Sec. 10.7)
(a) 2.50 L **(b)** 5.00 L
(c) 10.0 L **(d)** 20.0 L
(e) 40.0 L

8. If the temperature of a liquid increases from 20.0 °C to 50.0 °C, what happens to the vapor pressure? (Sec. 10.8)
(a) increases
(b) decreases
(c) remains constant
(d) unpredictable
(e) none of the above

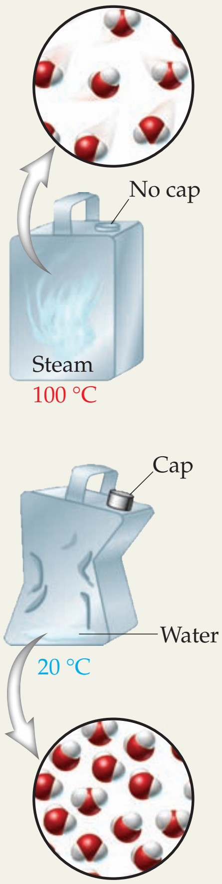
9. What is the partial pressure of a gas collected over water at 25 °C and 775 mm Hg? The vapor pressure of water at 25 °C is 23.8 mm Hg. (Sec. 10.9)
(a) 750 torr
(b) 751 torr
(c) 760 torr
(d) 775 torr
(e) 799 torr

10. Which of the following is true for an ideal gas? (Sec. 10.10)
(a) all molecules have the same velocity
(b) all molecules have the same kinetic energy
(c) all molecular collisions have the same energy
(d) all of the above
(e) none of the above

11. How many moles of helium occupy a volume of 5.00 L at 227 °C and 5.00 atm (given that $R = 0.0821\ \text{atm} \cdot \text{L/mol} \cdot \text{K}$)? (Sec. 10.11)
(a) 0.609 mol
(b) 1.64 mol
(c) 5.00 mol
(d) 25.0 mol
(e) 6090 mol

Key Concepts

12. A few drops of water in a gallon can are heated to steam, and the can is sealed with a cap. As the steam cools from 100 °C to 20 °C, the can is crushed, as shown in the diagram. Explain the observation.



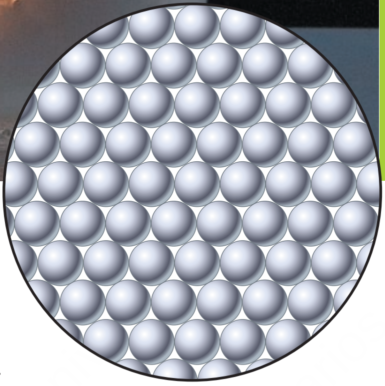
13. Methyl acetate, $C_3H_6O_2$, smells fragrant, and hydrogen sulfide gas, H_2S, smells like rotten eggs. Would you be attracted or repulsed first, if both gases escaped from the same container at the same time?
14. If you check automobile tire pressure on a cold morning before leaving home, and check it again after driving a few miles, would you expect the tire pressure to increase or decrease?
15. A half-full plastic water bottle is taken on a mountain hike. Describe the water bottle at the top of the mountain.

Critical Thinking

16. When the air inside a hot-air balloon is heated, why does the balloon rise? (Assume the air pressure inside the balloon equals the surrounding air pressure.)
17. When a helium balloon is released into the air, why does the volume increase as the balloon ascends? (Assume constant temperature.)
18. A half-full plastic water bottle in a car on a hot day bulges after an hour. Explain the observation.
19. A half-full plastic water bottle in a refrigerator is crushed after an hour. Explain the observation.

CHAPTER 11

Liquids and Solids



"If I have ever made any valuable discoveries it has been owing more to patient attention, than to any other talent."

Sir Isaac Newton, English Scientist (1643–1727)

11
Na
sodium

Element 11: Sodium
Sodium is the sixth most abundant element in Earth's crust and comprises about 2.6%. Sodium was discovered in 1807 by the English chemist Humphry Davy who isolated the element by passing an electric current through sodium hydroxide. Sodium only occurs naturally as a single, stable isotope, Na–23. Sodium is an essential element for all animals and some plants.

In Chapter 10, we discussed the gaseous state of matter, where individual molecules are relatively distant from one another. In this chapter, we will discuss the liquid and solid states of matter in which particles contact one another. We will use water in much of our discussion to reinforce the topics of nomenclature, mole calculations, and stoichiometry that we discussed in earlier chapters.

Water is the most important liquid on Earth and covers about three-fourths of the surface of our planet. It is necessary for all the chemical reactions that support plant and animal life, and about two-thirds of the mass of the human body is water. With the exception of oxygen, water is our most critical substance. We can survive quite some time without food, but only a few days without water.

The physical properties of water are unusual in several ways. One unusual property is its density. The density of a substance is generally greater in the solid state than in the liquid state. The two exceptions to this generalization are water and ammonia. Because the density of solid ice is less than that of liquid water, ice floats on water. If ice were more dense than water, marine life would not survive. Rivers and lakes would freeze into solid chunks of ice in wintry climates. Initially, ice would form on the surface and then sink, and eventually the entire body of water would change to solid ice. This, of course, would destroy aquatic life.

11.1 Properties of Liquids

LEARNING OBJECTIVE

- Identify five observed properties of a liquid.

Unlike gases, liquids do not respond significantly to temperature and pressure changes. Also, the mathematical relationships that apply to gases, such as the combined gas law, do not apply to liquids. Rather, when we study the liquid state, we observe the following general properties:

1. **Liquids have a variable shape, but a fixed volume.** The shape of a liquid conforms to the shape of its container.
2. **Liquids usually flow readily.** Liquids flow at different rates; for example, petroleum flows more slowly than water.
3. **Liquids do not compress or expand significantly.** The volume of a liquid varies very little with changes in temperature or pressure.
4. **Liquids have a high density compared to gases.** Gases and liquids are both fluids; that is, the shape is not fixed and individual particles are free to move throughout the container. However, the liquid state is about 1000 times more dense than the gaseous state because particles in a liquid are closer together. For example, the density of water is 1.00 g/mL, whereas the density of air at sea level is about 0.001 g/mL.
5. **Liquids that are soluble mix homogeneously.** Liquids diffuse more slowly than gases. However, liquids that are soluble eventually form a homogeneous mixture. For example, when food coloring is added to water, the liquids diffuse slowly and mix uniformly.

▲ **Scientific Misconception** It is a misconception that stained glass windows are thicker at the bottom due to liquid flow; the thickness varies owing to the glassmaking process.

A CLOSER LOOK **Soda Ash**

Q: **What are some common uses and properties of soda ash?**

Sodium carbonate, Na_2CO_3, is also known as soda ash and washing soda. It commonly occurs as a crystalline hydrate, which readily effloresces forming a white powder. Sodium carbonate is a white, odorless powder that absorbs moisture from the air, and is used as a water softener.

Sodium carbonate, Na_2CO_3, crystallizes from aqueous solution to form three different hydrates: sodium carbonate monohydrate ($Na_2CO_3 \cdot H_2O$), sodium carbonate heptahydrate ($Na_2CO_3 \cdot 7\,H_2O$), and sodium carbonate decahydrate ($Na_2CO_3 \cdot 10\,H_2O$). Each hydrate can be decomposed by heating with predictable results, which make them suitable for routine chemistry laboratory experiments.

Powdered Soda Ash

The manufacture of glass is one of the most important uses of sodium carbonate. This type of glass is known as “soda lime” glass. The “soda” refers to sodium carbonate, and “lime” refers to calcium carbonate, which are used in its manufacture. Soda lime glass has been the most common form of glass dating back to the Egyptians, circa 1000 B.C.

Sodium carbonate is also used in the process of developing photographic film. Sodium carbonate in solution has the ability to react with metals such as zinc and aluminum while releasing hydrogen gas. It is frequently added to swimming pools in order to reduce the corrosive effects of chlorine and raise the pH.

In the last 10 years, the Department of Transportation has reported over 100,000 hazardous waste spills. The hazardous waste includes flammable liquids, corrosive chemicals, and poisonous materials. Hazardous waste spills of acids or bases are neutralized with soda ash and baking soda, which can then be washed down a sewer drain.

A: The main use of soda ash is the manufacture of glass. Soda ash is also used to clean up hazardous spills, and hydrates of soda ash are used in chemistry laboratory experiments.

11.2 The Intermolecular Bond Concept

LEARNING OBJECTIVES

- Explain the concept of an intermolecular bond.
- Describe three types of attraction between molecules in a liquid.

The Internet is an international network for computers, whereas an Intranet is a localized network; and highways are designated interstate (between states) or intrastate (in state). Similarly, an intermolecular bond indicates an attraction between molecules, whereas an intramolecular bond is between atoms in a molecule.

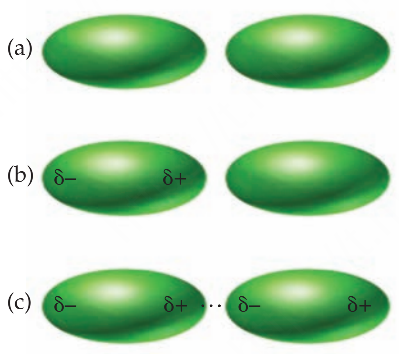
Some properties of liquids, such as vapor pressure, are determined by the strength of attraction between molecules. This attraction is called an *intermolecular bond*, and it is much weaker than an intramolecular bond between atoms. The three types of intermolecular attraction are based on temporary dipoles, permanent dipoles, and hydrogen bonds.

Temporary Dipole Attraction

In molecules, positive and negative charges are concentrated in different regions of individual molecules. These two charged regions are created by an uneven distribution of electrons in the molecule. A molecule having regions of positive and negative charge is said to possess a *dipole.*

Let's first consider temporary dipole attraction. Even though atoms in a molecule may share electrons equally, the electrons are constantly shifting about. This shifting about of electrons produces regions in the molecule that are temporarily "electron rich" and slightly negative. Simultaneously, another region of the molecule is temporarily "electron poor" and slightly positive. A negative region in one molecule has a weak attraction for a positive region in another molecule. This temporary attraction between molecules is referred to as a **dispersion force**.

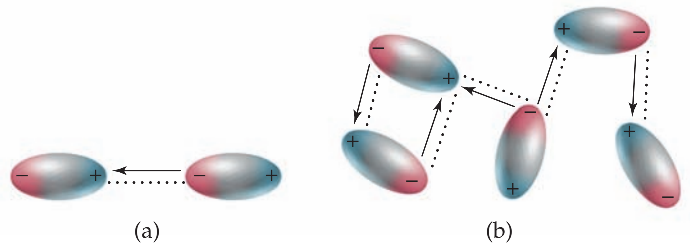
Although dispersion forces last for only brief periods of time, they occur frequently between molecules. Figure 11.1 illustrates temporary intermolecular dispersion forces between the positive end of a molecule (∂+), and the negative end of a molecule (∂–).



◀ **Figure 11.1 Intermolecular Dispersion Forces** (a) Two molecules are shown. (b) The molecule on the left forms a temporary dipole, which can induce a dipole in the molecule on the right. (c) The two molecules are temporarily attracted to each other. In turn, these molecules induce temporary dipole attractions in the surrounding molecules in the liquid.

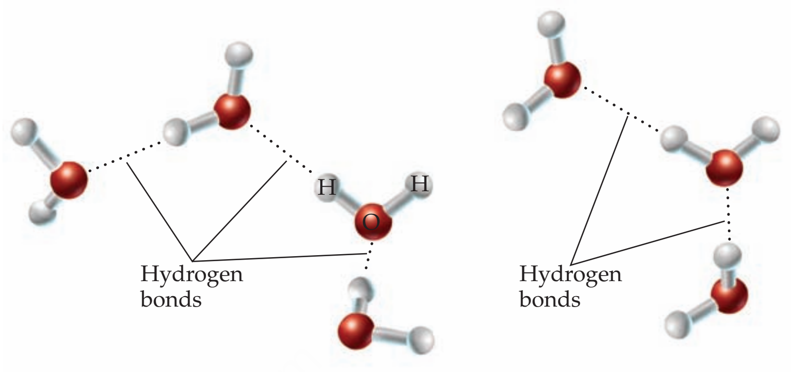
Permanent Dipoles and Hydrogen Bonds

Atoms of different elements vary in their attraction for electrons. This gives rise to regions of positive and negative charge in a molecule. As a result, similar to magnets, individual molecules are attracted to each other. In contrast to temporary dipole attraction, permanent dipole attraction operates continuously. That is, a permanent **dipole force** operates between molecules as shown in Figure 11.2.



◀ **Figure 11.2 Intermolecular Dipole Forces** (a) A dipole force results from a dipole–dipole attraction between two molecules. (b) Dipole forces between several molecules in a liquid are shown.

When a hydrogen atom is bonded to an oxygen atom or a nitrogen atom in a molecule, an especially strong dipole results. This strong dipole produces an attraction between molecules that is typically about 10% of the strength of an intramolecular bond. This special type of permanent dipole attraction is called a **hydrogen bond** (Figure 11.3). Hydrogen bonds are important in living organisms where they maintain the structure of protein molecules, and the double-helix structure of DNA molecules.



▲ **Figure 11.3 Intermolecular Hydrogen Bonds** Water molecules form intermolecular hydrogen bonds. Notice that the hydrogen atom is attracted to the nonbonding electrons on the highly electronegative oxygen atom.

The following example exercise illustrates the characteristics of intermolecular bonds:

Helpful Hint
Covalent Bonds vs. Intermolecular Bonds

In the next chapter (Chapter 12), we will discuss the types of bonds which hold molecules together. A covalent bond is formed between two nonmetal atoms by the sharing of valence electrons. Covalent bonds are strong, permanent, and require much energy to break.

In this chapter, we see that molecules can be temporarily attracted to each other. However, the strength of attraction for an intermolecular bond is much less than that of a covalent bond.

- A molecule may have a temporary dipole that weakly attracts another molecule by dipole–dipole intermolecular attraction.
- A molecule may have a permanent dipole that strongly attracts another molecule by dipole–dipole intermolecular attraction.
- A molecule may contain an H—O bond, or an H—N bond, that very strongly attracts another molecule by forming an intermolecular hydrogen bond.

EXAMPLE EXERCISE 11.1 Intermolecular Attraction

If molecules in a liquid have a permanent dipole, what is the strongest type of intermolecular attraction?

Solution

In a liquid, intermolecular attraction is the result of both permanent and temporary dipoles. Molecules that contain either H—O or H—N bonds have the strongest type of intermolecular attraction, that is, hydrogen bonds. Water, H_2O, and ammonia, NH_3, are examples of liquids with hydrogen bonds.

Practice Exercise

If molecules in a liquid do not have a permanent dipole, what is the strongest type of intermolecular attraction?

Answer:

The intermolecular attraction between molecules in a liquid not having a permanent dipole is a dispersion force resulting from temporary dipoles.

Concept Exercise

Which of the following liquids has the stronger intermolecular attraction between molecules: $CH_3—CH_2—OH$ or $CH_3—O—CH_3$?

Answer: See Appendix G, 11.1.

LEARNING OBJECTIVES

- Describe the relationship between intermolecular attraction in a liquid and the properties of vapor pressure, boiling point, viscosity, and surface tension.
- Predict which liquid in a pair has the higher vapor pressure, boiling point, viscosity, and surface tension.

11.3 Vapor Pressure, Boiling Point, Viscosity, and Surface Tension

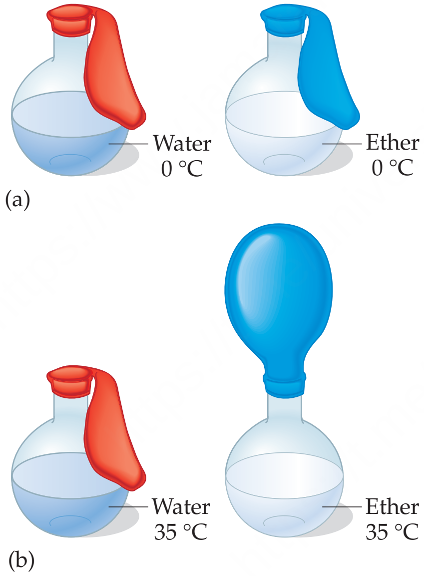
In the previous chapter, we introduced the kinetic theory of gases to explain the behavior of gases. Now, we will extend this theory to explain the properties of liquids.

According to the kinetic theory the attraction between gas molecules is negligible. This is not the case for liquids. Molecules in the liquid state are in contact with each other, and their attraction restricts their movement. Individual molecules do have enough energy to move about one another; that is, a liquid is free to flow. The liquid state and the gaseous state are analogous to honeybees swarming in a hive (liquid state), and individual bees flying off to gather pollen (gaseous state). Now let's consider some properties of liquids.

Vapor Pressure

The kinetic theory relates the average energy of molecules to temperature. The higher the temperature, the greater the kinetic molecular energy. However, not all molecules at the same temperature have identical energies. Some are more energetic than others. At the surface of a liquid, for example, some molecules have enough energy to completely escape the attraction of neighboring molecules. The molecules that escape enter the gaseous state, which is referred to as *vapor*. This process is called *vaporization.* In the reverse process, some molecules in the vapor return to the liquid. This process is called *condensation.*

When the rates of vaporization and condensation are equal, the pressure exerted by the gas molecules above a liquid is called the **vapor pressure**. However, the vapor pressure above a liquid depends on the attraction between molecules in the liquid. Water molecules have a strong attraction for each other, and ethyl ether molecules have a relatively weak attraction. Because of the lesser attraction between molecules, ether molecules escape from the liquid state more readily than water molecules. Therefore, the vapor pressure of ether is greater than that of water at the same temperature (Figure 11.4).



◀ **Figure 11.4 Vapor Pressure of Water and Ether** (a) At 0 °C neither water nor ether has sufficient vapor pressure to affect the balloons. (b) At 35 °C the vapor pressure of water is still low; however, the vapor pressure of ether is considerable and inflates the balloon.

Table 11.1 compares the attraction between molecules, that is, *intermolecular attraction,* and the vapor pressure for a few selected liquids. In general, as the attraction between molecules increases, the vapor pressure decreases. In Table 11.1 we see that propionic acid and butyl alcohol both have a strong intermolecular attraction. Thus, their vapor pressures at 20 °C are quite low. Conversely, propyl chloride and ethyl ether each has a weak intermolecular attraction. Thus, their vapor pressures are considerably

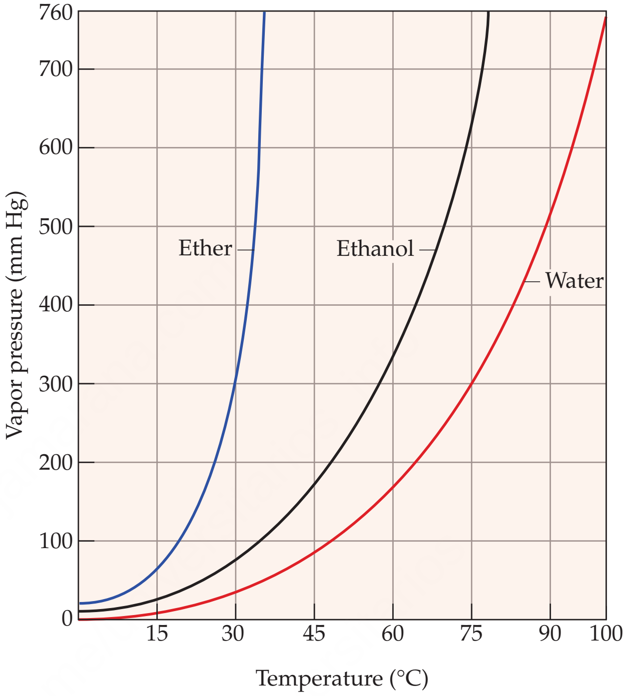
TABLE 11.1 Vapor Pressure of Selected Liquids

Liquid	Approximate Molar Mass	Intermolecular Attraction	Vapor Pressure at 20 °C
water	18 g/mol	strong	18 mm Hg
propionic acid, C_2H_5COOH	74 g/mol	strong	5 mm Hg
butyl alcohol, C_4H_9OH	74 g/mol	strong	6 mm Hg
propyl chloride, C_3H_7Cl	79 g/mol	weak	300 mm Hg
ethyl ether, $C_2H_5OC_2H_5$	74 g/mol	weak	450 mm Hg

higher. When comparing vapor pressure values, it is necessary to choose two liquids that have approximately the same molar mass. Note that in these comparisons, each liquid, except water, has an approximate molar mass of 74 g/mol.

Boiling Point

As the temperature of a liquid increases, its vapor pressure increases. At 30 °C the vapor pressure of water is about 30 mm Hg, and at 60 °C it is 150 mm Hg. Figure 11.5 shows the relationship between vapor pressure and temperature for ether, ethanol, and water.



▶ **Figure 11.5 Vapor Pressure versus Temperature** The vapor pressure of a liquid increases as the temperature increases. The vapor pressure of water equals 760 mm Hg at 100 °C. Thus, at 100 °C the vapor pressure of water equals the standard atmospheric pressure (760 mm Hg), and water begins to boil.

A liquid begins to boil when the pressure of the vapor above the liquid equals the pressure of the atmosphere. Thus, we define the normal **boiling point** (symbol **Bp**) as the temperature at which the vapor pressure equals standard atmospheric pressure. Accordingly, liquids that have *high* boiling points must have *low* vapor pressures. The following example exercise illustrates the determination of boiling points from vapor pressure data.

EXAMPLE EXERCISE 11.2 Boiling Point Predictions

Refer to Figure 11.5 and determine the approximate boiling point of ether.

Solution

The normal boiling point of a liquid is the temperature at which the vapor pressure equals standard atmospheric pressure, that is, 760 mm Hg. From the graph, we find that the vapor pressure of ether is 760 mm Hg at about 35 °C. Thus, the normal boiling point of ether is about 35 °C. The actual observed value for the boiling point of ether is 36 °C.

Practice Exercise

Refer to Figure 11.5 and determine the approximate boiling point of ethanol.

Answer:

~80 °C (The actual boiling point of ethanol is 78 °C.)

Concept Exercise

Which of the following liquids has the higher boiling point: $CH_3—CH_2—OH$ or $CH_3—O—CH_3$?

Answer: See Appendix G, 11.2.

Viscosity

Some liquids are easier to pour than others. Water pours easily, whereas honey does not. The resistance of a liquid to flow is a property called **viscosity**. Viscosity is the result of an attraction between molecules. It is also affected by factors such as the size and shape of the molecules. In principle, the greater the attraction between molecules, the higher the viscosity. Table 11.2 compares the attraction between molecules and the viscosity for a few selected liquids.

▲ **Viscosity** The motor oil on the left is more viscous and flows more slowly than the motor oil on the right.

TABLE 11.2 Viscosity of Selected Liquids

Liquid	Approximate Molar Mass	Intermolecular Attraction	Viscosity* at 20 °C
water	18 g/mol	strong	1.00
propionic acid, C_2H_5COOH	74 g/mol	strong	1.10
butyl alcohol, C_4H_9OH	74 g/mol	strong	2.95
propyl chloride, C_3H_7Cl	79 g/mol	weak	0.35
ethyl ether, $C_2H_5OC_2H_5$	74 g/mol	weak	0.23

*Values are expressed in centipoise, a common unit of viscosity.

In Table 11.2, we see that propionic acid and butyl alcohol have a strong intermolecular attraction and that the viscosity of both is greater than that of water. We also see that propyl chloride and ethyl ether have a weak attraction between molecules and that the viscosity of both is much lower than that of water. Honey is an example of a viscous liquid. We can therefore predict that there is a strong attraction between molecules in honey.

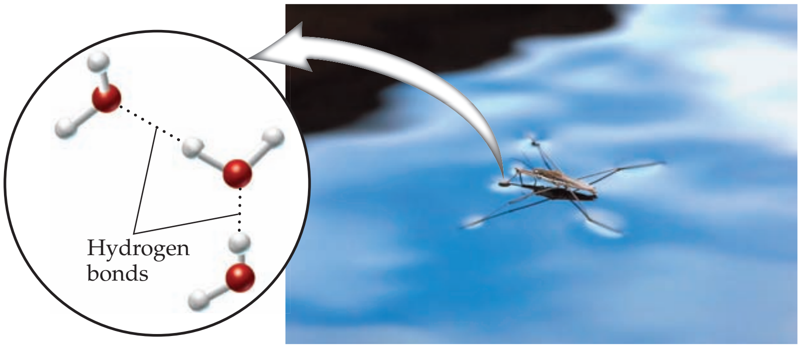
Surface Tension

At some time you have probably noticed a small insect floating on water. For an insect or any object to sink in a liquid, it has to break through the surface. But the molecules on the surface of a liquid resist being pushed apart. The attraction between the surface molecules in a liquid is called **surface tension**. There are other factors to consider, but in general the greater the intermolecular attraction, the higher the surface tension. Table 11.3 compares the intermolecular attraction and surface tension for a few selected liquids.

TABLE 11.3 Surface Tension of Selected Liquids

Liquid	Approximate Molar Mass	Intermolecular Attraction	Surface Tension* at 20 °C
water	18 g/mol	strong	70
propionic acid, C_2H_5COOH	74 g/mol	strong	27
butyl alcohol, C_4H_9OH	74 g/mol	strong	25
propyl chloride, C_3H_7Cl	79 g/mol	weak	18
ethyl ether, $C_2H_5OC_2H_5$	74 g/mol	weak	17

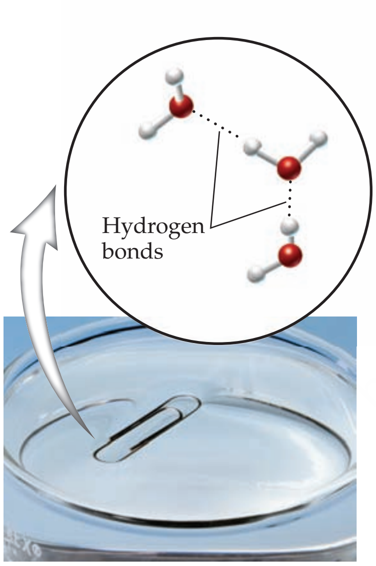
*Values are expressed in dynes per square centimeter, a common unit of surface tension.



◀ **Surface Tension** Water has a high surface tension owing to hydrogen bonds, and this allows the insect to "walk" on water.

We all know that rain forms drops as it falls. In fact, when we spray any liquid, it forms drops, and each drop has the shape of a small sphere. Drops of liquid are spherical because surface tension causes them to have the smallest possible surface area, and that area corresponds to a spherical droplet. Table 11.3 shows that water has an unusually high surface tension.

Example Exercise 11.3 illustrates the relationships between the properties of liquids and intermolecular attraction.



▲ **Surface Tension** A paper clip can float on water.

EXAMPLE EXERCISE 11.3 Physical Property Predictions

Consider the following properties of liquids. State whether the value for each property is high or low for a liquid with a strong intermolecular attraction:

(a) vapor pressure
(b) boiling point
(c) viscosity
(d) surface tension

Solution

For a liquid having a strong attraction between molecules, properties (b), (c), and (d) are generally high; property (a) is low.

(a) Molecular attraction slows vaporization. Therefore, vapor pressure is *low* for liquids with a strong intermolecular attraction.
(b) Attraction between molecules inhibits boiling. Thus, the boiling point is *high* for liquids with a strong intermolecular attraction.
(c) Molecular attraction increases the resistance of a liquid to flow. Viscosity is *high* for liquids with a strong intermolecular attraction.
(d) Attraction between molecules causes a drop of liquid to form a sphere. Surface tension is *high* for liquids with a strong intermolecular attraction.

Practice Exercise

The intermolecular attraction is greater in isopropyl alcohol, C_3H_7OH, than in pentane, C_5H_{12}. Predict which liquid has the higher value for each of the following:

(a) vapor pressure
(b) boiling point
(c) viscosity
(d) surface tension

Answers:

(a) C_5H_{12}
(b) C_3H_7OH
(c) C_3H_7OH
(d) C_3H_7OH

Concept Exercise

Explain why the paper clip can float on water even though its density is much greater than the density of water.

Answer: See Appendix G, 11.3.

LEARNING OBJECTIVE

▸ Identify five observed properties of a solid.

11.4 Properties of Solids

Unlike liquids, solids have a fixed shape. The reason for this is that individual particles of the solid are not free to move. Unlike gases, the volume of a solid shows very little response to changes in temperature or pressure. The observed properties of the solid state are as follows:

1. **Solids have a fixed shape and a fixed volume.** Unlike liquids, solids are rigid and their shape is fixed.
2. **Solids are either crystalline or noncrystalline.** A **crystalline solid** contains particles arranged in a regular repeating pattern. Each particle occupies a fixed position in the crystal. The high degree of order of the molecules can produce a beautiful clear crystal, for example, a diamond or a ruby. An ordinary stone is a noncrystalline solid and is not transparent.
3. **Solids do not compress or expand to any degree.** Assuming no change in physical state, temperature and pressure have a negligible effect on the volume of a solid.

4. **Solids have a slightly higher density than their corresponding liquids.** For example, solid chunks of iron sink in a high-temperature furnace containing molten iron. One important exception to this rule is water; ice is less dense than liquid water. As a result, ice floats on water.
5. **Solids do not mix by diffusion.** In a solid heterogeneous mixture, the particles are not free to diffuse and they cannot mix uniformly. In an alloy, which is a homogeneous mixture, the metals mix uniformly in the molten liquid state before cooling to a solid.

11.5 Crystalline Solids

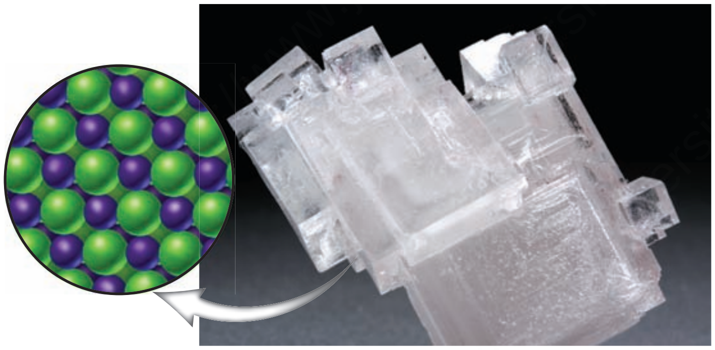
LEARNING OBJECTIVE

- Describe three types of crystalline solids: ionic, molecular, and metallic.

Chemists know that the particles of a crystalline solid are arranged in a regular geometric pattern. However, the particles can be of different types. They can be ionic, molecular, or metallic. In ionic solids the crystals are composed of regular patterns of ions. In molecular solids the molecules, not ions, form repeating patterns. In metallic solids, individual metal atoms are arranged geometrically.

Ionic Solids

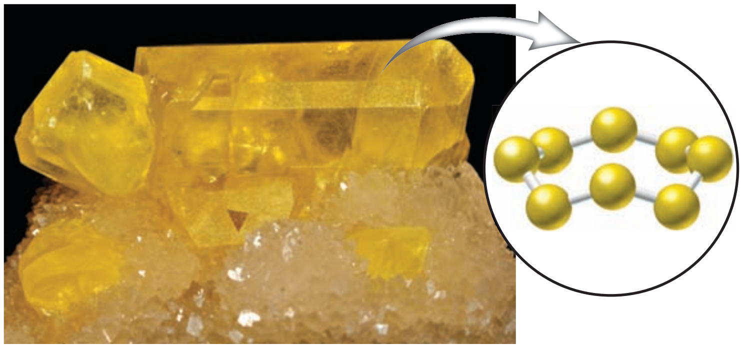
A crystalline **ionic solid** is an ionic compound composed of positive and negative ions. Salt, for example, is a crystalline solid of NaCl. Here, sodium ions, Na^+, and chloride ions, Cl^-, are arranged in a regular three-dimensional structure referred to as a *crystal lattice*. Notice the clear crystal of table salt shown in Figure 11.6 has the shape of a cube. Other ionic crystalline compounds, such as NaF, CaF_2, and $CaCO_3$, have different geometric shapes.



◀ **Figure 11.6 A Crystalline Ionic Solid** A sodium chloride crystal, NaCl, is an example of a crystalline ionic solid.

Molecular Solids

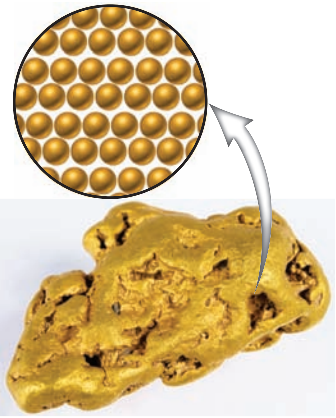
A crystalline **molecular solid** has molecules arranged in a particular configuration. Crystalline sucrose (table sugar), for example, is composed of $C_{12}H_{22}O_{11}$ molecules. The sucrose molecules are arranged in a regular order that allows light to pass through the crystal. Therefore, a large crystal of sucrose appears transparent. Other molecular solids are sulfur and phosphorus. Sulfur crystals (Figure 11.7) are made from S_8 molecules held together by intramolecular bonds. Phosphorus powder is made from P_4 molecules.



◀ **Figure 11.7 A Crystalline Molecular Solid** A sulfur crystal, S_8, is an example of a crystalline molecular solid.

Metallic Solids

A crystalline **metallic solid** has atoms of metals arranged in a definite pattern. That is, a metallic crystal is made up of positive metal ions surrounded by valence electrons. Metals are good conductors of electricity because the valence electrons are free to move about the crystal. This arrangement of atoms and freely moving electrons is referred to as the "electron sea" model. The flow of electricity is associated with the movement of electrons through a metal. Figure 11.8 shows a crystal of gold.



▲ **Figure 11.8 A Crystalline Metallic Solid** A gold nugget is an example of a crystalline metallic solid.

Classifying Crystalline Solids

It is helpful to classify crystalline solids to predict their properties. In general, the properties of ionic solids such as melting point, hardness, electrical conductivity, and water solubility are similar. The properties of molecular solids are usually similar as well. Metals show a range of physical properties, but all are malleable (form thin foils) and ductile (drawn into wire) and are good conductors of electricity. Table 11.4 lists the general properties for each of the three types of crystalline solids.

TABLE 11.4 General Properties of Crystalline Solids

Type of Solid	General Properties	Examples
ionic	high melting point, hard, brittle, at least slightly soluble in water, conductor of electricity when melted or in solution	NaCl, $CaCO_3$, $MgSO_4$
molecular	low melting point, generally insoluble in water, nonconductor of electricity	S_8, $C_{10}H_8$, $C_6H_{12}O_6$
metallic	low to high melting point, malleable, ductile, conductor of electricity, insoluble in most solvents	Cu, Ag, Au

The following example exercise further illustrates the classification of ionic, molecular, and metallic types of crystalline solids.

EXAMPLE EXERCISE 11.4 Classifying Crystalline Solids

Classify each of the following crystalline solids as ionic, molecular, or metallic:

(a) nickel, Ni
(b) nickel(II) oxide, NiO

Solution

The type of crystalline solid is dictated by the type of particle in the solid.

(a) Nickel is a metal composed of atoms; thus, Ni is a *metallic solid.*
(b) Nickel(II) oxide contains ions and is therefore an *ionic solid.*

Practice Exercise

Classify each of the following crystalline solids as ionic, molecular, or metallic:

(a) iodine, I_2
(b) silver iodide, AgI

Answers:

(a) molecular
(b) ionic

Concept Exercise

Which type of crystalline solid has a high melting point and is a nonconductor of electricity in the solid state, but a good conductor of electricity when melted?

Answer: See Appendix G, 11.4.

11.6 Changes of Physical State

LEARNING OBJECTIVES

- Explain the concepts of specific heat, heat of fusion, and heat of vaporization.
- Calculate heat changes that involve the specific heat, heat of fusion, and heat of vaporization.

Heat is necessary to change the physical state of a substance. **Specific heat** is the amount of heat required to raise one gram of substance one degree Celsius. Every substance has a unique value for its specific heat. Water is considered a reference, and its specific heat is 1.00 calorie per gram per degree Celsius, that is, 1.00 cal/(g × °C). The specific heats for ice and steam are approximately half that of liquid water.

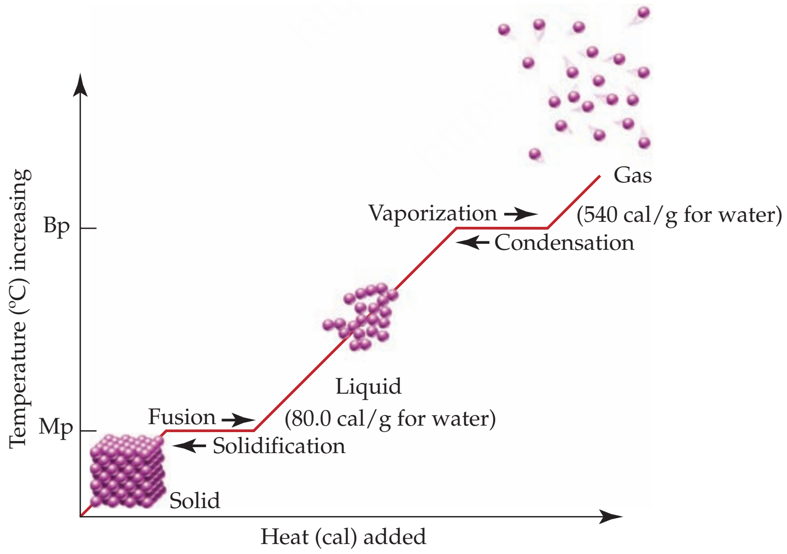
Next, let's consider a substance changing state from a solid to a liquid at its melting point. The amount of heat required to melt 1.00 g of substance is called the **heat of fusion** (H_{fusion}). For water, the heat of fusion is 80.0 cal/g. Water releases the same amount of heat energy, 80 cal/g, when it changes from the liquid to the solid state. This heat change is called the *heat of solidification* (H_{solid}). For any substance, the heats of fusion and solidification are equal. That is, the amount of heat to melt a substance equals the amount of heat released when the substance solidifies.

A substance rapidly changes state from a liquid to a vapor at its boiling point. The amount of heat required to vaporize 1.00 g of a substance is called the **heat of vaporization** (H_{vapor}). For water, it is 540 cal/g. Conversely, water releases the same amount of heat energy, 540 cal/g, when it condenses from a gas to a liquid. This heat change is called the *heat of condensation* (H_{cond}). For any substance, the heats of vaporization and condensation are equal. Table 11.5 lists the heat values for water, ice, and steam.

TABLE 11.5 Heat Values for Water

Substance	Specific Heat [cal/(g × °C)]	H_{fusion} (cal/g)	H_{solid} (cal/g)	H_{vapor} (cal/g)	H_{cond} (cal/g)
ice, $H_2O(s)$	0.50	80.0			
water, $H_2O(l)$	1.00		80.0	540	
steam, $H_2O(g)$	0.48				540

To see the change in temperature with a constant application of heat, we draw a temperature–energy graph, sometimes called a *heating curve*. The heating curve for water is shown in Figure 11.9.



◀ **Figure 11.9 Temperature–Energy Graph** As heat is continuously added to a substance, the substance eventually changes its physical state. Notice the temperature remains constant during fusion (solid to liquid), vaporization (liquid to gas), condensation (gas to liquid), and solidification (liquid to solid).

Let's combine the concept of a heating curve with the heat values in Table 11.5. For example, let's find the amount of heat energy necessary to convert 25.5 g of ice at −5.0 °C to steam at 100.0 °C. This problem requires four steps: (1) heat the ice from −5.0 °C to 0.0 °C, (2) convert the ice to liquid water at 0.0 °C, (3) heat the water from 0.0 °C to 100.0 °C, and (4) convert the water to steam at 100.0 °C.

1. To calculate the amount of energy required to heat the ice we use its mass (25.5 g), the temperature change (−5.0 °C to 0.0 °C), and the specific heat of ice (0.50 cal/(g × °C)). Using 0.50 cal/(g × °C) as a unit factor, we have

$$25.5\ \cancel{g} \times \frac{0.50\ \text{cal}}{1\ \cancel{g} \times \cancel{^\circ\text{C}}} \times [0.0 - (-5.0)]\ \cancel{^\circ\text{C}} = 64\ \text{cal}$$

2. The heat of fusion for ice, 80.0 cal/g, is found in Table 11.5. The energy required to melt 25.5 g of ice is

$$25.5\ \cancel{g} \times \frac{80.0\ \text{cal}}{1\ \cancel{g}} = 2040\ \text{cal}$$

3. To calculate the amount of energy required to heat the liquid water we must know its mass, the temperature change (0.0 °C to 100.0 °C), and the specific heat of water; therefore, we have

$$25.5\ \cancel{g} \times \frac{1.00\ \text{cal}}{1\ \cancel{g} \times \cancel{^\circ\text{C}}} \times (100.0 - 0.0)\ \cancel{^\circ\text{C}} = 2550\ \text{cal}$$

4. The heat of vaporization, 540 cal/g, is found in Table 11.5. The energy required to vaporize the water to steam is

$$25.5\ \cancel{g} \times \frac{540\ \text{cal}}{1\ \cancel{g}} = 13{,}800\ \text{cal}$$

The total heat energy required to heat and vaporize the ice is equal to the sum of the values obtained in Steps 1–4.

$$64\ \text{cal} + 2040\ \text{cal} + 2550\ \text{cal} + 13{,}800\ \text{cal} = 18{,}500\ \text{cal}$$

The heat required to raise the temperature of the ice at −5.0 °C to steam at 100.0 °C is 18,500 cal, or 18.5 kcal. The following example exercise illustrates the heat changes associated with the cooling of water and its solidification to ice.

EXAMPLE EXERCISE 11.5 Change in Heat Problem

Calculate the amount of heat released when 15.5 g of liquid water at 22.5 °C cools to ice at −10.0 °C.

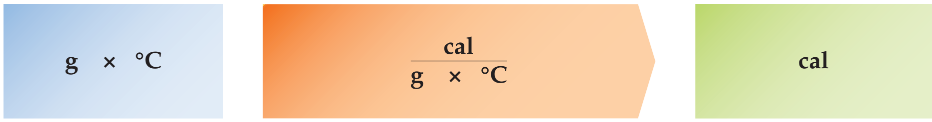
Solution

In this problem we have to consider (1) the specific heat of water, (2) the heat of solidification, and (3) the specific heat of ice.

(1) To calculate the amount of heat released when cooling the water, consider the mass, the temperature change (22.5 °C to 0.0 °C), and the specific heat of water, 1.00 cal/(g × °C).

$$15.5\ \text{g} \times \frac{1.00\ \text{cal}}{1\ \text{g} \times {^\circ\text{C}}} \times (22.5 - 0.0)\ {^\circ\text{C}} = 349\ \text{cal}$$

We can summarize the solution visually as follows:



(2) The heat of solidification, found in Table 11.5, is 80.0 cal/g. The heat released when water solidifies to ice is

$$15.5\ \cancel{g} \times \frac{80.0\ \text{cal}}{1\ \cancel{g}} = 1240\ \text{cal}$$

(3) The specific heat of ice is 0.50 cal/(g × °C). The heat released as the ice cools to −10.0 °C is found as follows:

$$15.5\ \cancel{\text{g}} \times \frac{0.50\ \text{cal}}{1\ \cancel{\text{g}} \times \cancel{^\circ\text{C}}} \times [0.0 - (-10.0)]\cancel{^\circ\text{C}} = 78\ \text{cal}$$

The total heat energy released when the water cools to ice at −10.0 °C equals the sum of the values obtained in Steps 1–3.

$$349\ \text{cal} + 1240\ \text{cal} + 78\ \text{cal} = 1670\ \text{cal}$$

Thus, the heat released when the water cools is 1670 cal, or 1.67 kcal.

Practice Exercise

Calculate the amount of heat required to convert 50.0 g of steam at 100.0 °C to ice at 0.0 °C.

Answer: 3.60×10^4 cal (36.0 kcal)

Concept Exercise

Which of the following changes involves more heat energy: (a) heating 1 g of ice at 0 °C to water at 0 °C, or (b) heating 1 g of water at 100 °C to steam at 100 °C.

Answer: See Appendix G, 11.5.

11.7 Structure of Water

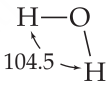
LEARNING OBJECTIVE

- Illustrate the bond angle and net dipole in a water molecule.

We learned in Section 11.3 that the boiling point and surface tension of water are unusually high. We also saw that water has a strong intermolecular attraction because of hydrogen bonding. To understand these properties more completely, let's review the water molecule. Specifically, let's examine the structural formula, bond angle, and net dipole for a water molecule.

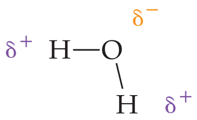
Structural Formula and Bond Angle

The structural formula for a molecule uses dashes to represent bonds between atoms. We will draw the water molecule with the two hydrogen atoms at an angle to each other. Experimental evidence shows that the angle between the two hydrogen atoms is 104.5°. The angle formed by the central atom and two attached atoms is referred to as the **bond angle**.

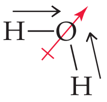


Dipoles and Net Dipole

In a water molecule, each of the two bonds is a dipole. That is, the oxygen atom is slightly negative and each hydrogen atom is slightly positive. We can indicate this using delta notation (δ^+ and δ^-).



Notice that a water molecule has two dipoles, and the central oxygen atom is pulling electrons toward the O atom. The two dipoles create a single dipole passing through the center of the molecule. The single, overall dipole for a molecule having two or more dipoles is called the **net dipole**.



The net dipole produces a negative end and a positive end in the water molecule. The negative end of the molecule is indicated by the tip of the arrow. The positive end is indicated by the plus sign on the opposite end of the net dipole arrow. The positive and negative ends of different water molecules form hydrogen bonds. This strong intermolecular attraction between water molecules explains water's unusual properties, including a high boiling point and high surface tension.

CHEMISTRY CONNECTION Water Purification

Q: What is the process for changing hard water into soft water?

In some areas of the country, the minerals dissolved in water give it a high concentration of various ions that make the water suitable neither for drinking nor for agriculture. Such water is called **hard water**. Sometimes the mineral content in hard water is so great that it causes plumbing and corrosion problems. Hard water typically contains a high concentration of the following ions: Ca^{2+}, Mg^{2+}, Fe^{3+}, Cl^-, CO_3^{2-}, SO_4^{2-}, and PO_4^{3-}.

In homes having hard water, soap rings are sometimes found in bathtubs. A soap ring is formed by the cations in hard water reacting with soap to create a compound that deposits as an insoluble film. This problem can be eliminated by removing Ca^{2+}, Mg^{2+}, and Fe^{3+} from water by passing hard water through a water softener. A water softener replaces Ca^{2+}, Mg^{2+}, and Fe^{3+} with Na^+; compounds containing Na^+ are generally soluble. Note, however, that although a water softener removes ions, **soft water** still has a high concentration of other ions. In addition to Na^+, soft water contains Cl^-, CO_3^{2-}, SO_4^{2-}, and PO_4^{3-}.

Soft water can interfere with the results of a chemical analysis. Therefore, chemists routinely use water that contains no ions. Chemists may use **distilled water** that is made by boiling water and collecting the condensed vapor. Alternatively, chemists use water with the minerals removed using an ion exchange system. Water purified by this method is called **deionized water**, or demineralized water. Deionized water is made by passing hard water through a system that exchanges both cations and anions. First, cations such as Na^+ in the water are exchanged for hydrogen ions on the resin. Second, anions such as Cl^- in the water are exchanged for hydroxide ions.

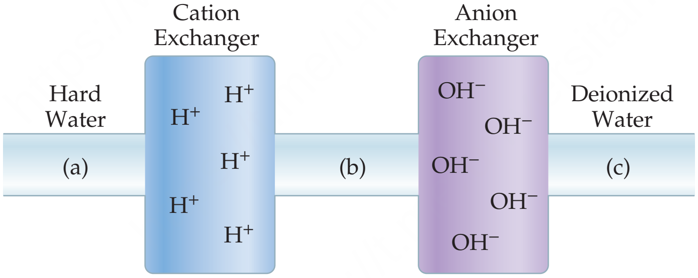
$$Na^+(aq) + H(resin) \longrightarrow Na(resin) + H^+(aq)$$

$$Cl^-(aq) + (resin)OH \longrightarrow (resin)Cl + OH^-(aq)$$

Notice that the ion exchange resin produces both hydrogen ions and hydroxide ions, which combine to give water.

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

The net result is that the resin removes all ions from water passing through the ion exchange resin.



▲ **Deionized Water** The cations in hard water are exchanged for H^+. The anions in hard water are exchanged for OH^-. The H^+ and OH^- combine to produce deionized H_2O.

A: Hard water is changed to soft water using an ion-exchange resin, which exchanges Na^+ ions for Ca^+, Mg^+, and Fe^+.

LEARNING OBJECTIVE

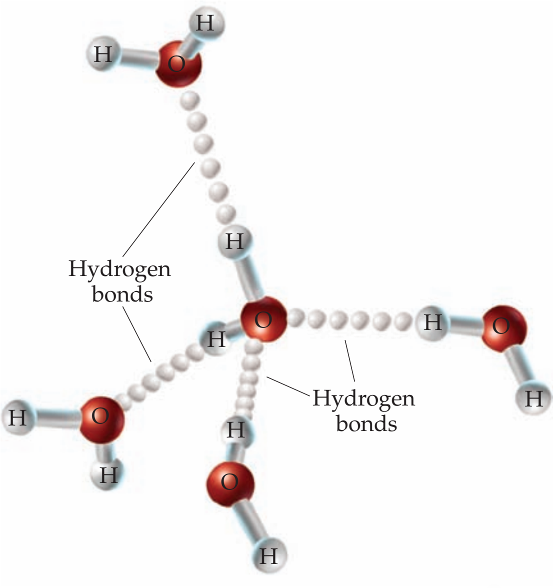
- Explain the unusual physical properties of water.

11.8 Physical Properties of Water

Water is a colorless, odorless, tasteless liquid, and a powerful solvent. At room temperature water has the highest specific heat, heat of fusion (except for ammonia), and heat of vaporization of any liquid.

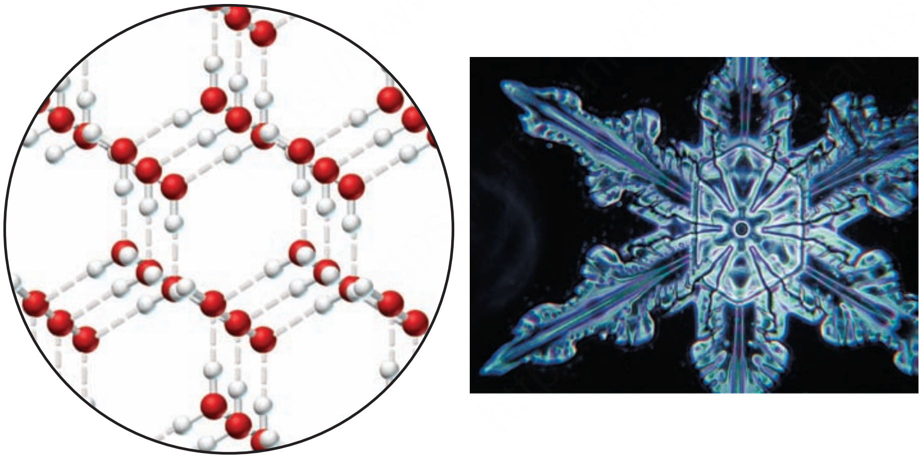
Density

Generally, a substance in the solid state has a higher density than one in the liquid state. Therefore, we would predict that the density of ice is greater than that of water. But, we know that ice floats in water. That is, the solid form of water is *less dense* than the liquid form. The reason solid ice is less dense than liquid water relates to hydrogen bonding. Figure 11.10 illustrates three-dimensional hydrogen bonding in water.



◀ **Figure 11.10 Hydrogen Bonding** Each water molecule is attracted to four other water molecules. These intermolecular hydrogen bonds are about 50% longer than an ordinary covalent bond. Because a hydrogen bond is longer, it is weaker and requires much less energy to break.

When water freezes to ice, the hydrogen bonds produce a three-dimensional crystal. Figure 11.11 illustrates the structure of an ice crystal. Because of the arrangement of water molecules, however, the crystal has holes. These holes create a volume for ice that is greater than that for an equal mass of water. Thus, the density of solid ice is less than that of liquid water. At 0 °C, the density of ice is 0.917 g/mL, and the density of water is 1.00 g/mL.



◀ **Figure 11.11 Structure of Ice Crystals** Water molecules hydrogen bond to form six-member rings. The rings in turn hydrogen bond to other rings, producing large, three-dimensional crystalline structures.

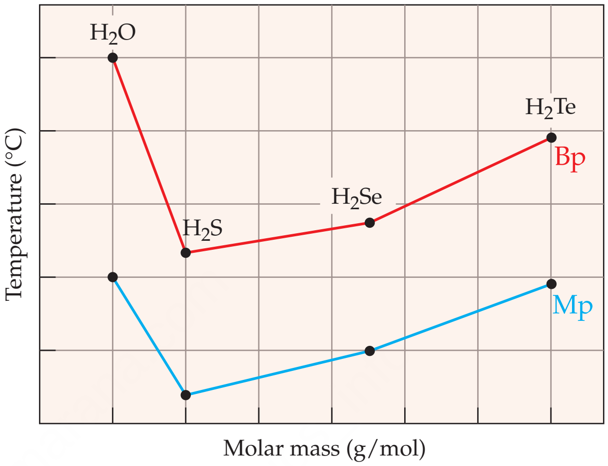
Melting and Boiling Points

Water has an unusually high melting point and boiling point for a small molecule. To see how unusual these properties are, compare them with those of some hydrogen compounds of Group VIA/16 as outlined in Table 11.6.

TABLE 11.6 Group VIA/16 Hydrogen Compounds

Compound	Molar Mass (g/mol)	Mp (°C)	Bp (°C)	H_{fusion} (cal/mol)	H_{vapor} (cal/mol)
H_2O	18	0.0	100.0	1440	9720
H_2S	34	−85.5	−60.7	568	4450
H_2Se	81	−60.4	−41.5	899	4620
H_2Te	130	−48.9	−2.2	1670	5570

If we ignore water, we see a clear trend in the melting and boiling points of Group VIA/16 hydrogen compounds. First, notice the increase in molar mass for H_2S through H_2Te. Next, notice that the values for the melting point and the boiling point increase simultaneously. As with density, the properties of water are unusual because of hydrogen bonding. Hydrogen bonding produces a strong intermolecular attraction that resists the movement of molecules. Therefore, a higher temperature is needed to melt ice and to boil water. Figure 11.12 illustrates the melting and boiling points of Group VIA/16 hydrogen compounds.



▶ **Figure 11.12 Mp and Bp of Group VIA/16 Hydrogen Compounds** Notice the systematic trend in melting point and boiling point as the molar mass increases. Water is a striking exception because of strong intermolecular hydrogen bonds.

EXAMPLE EXERCISE 11.6 Physical Property Predictions

Predict which compound in each of the following pairs has the higher boiling point:

(a) NH_3 or PH_3

(b) PH_3 or AsH_3

Solution

Ammonia, NH_3, molecules can hydrogen bond, while PH_3 and AsH_3 cannot.

(a) Ammonia, NH_3, has a higher boiling point than PH_3.

(b) AsH_3 has a higher boiling point than PH_3 because it has a greater molar mass.

Practice Exercise

Predict which compound in each of the following pairs has the higher heat of vaporization:

(a) NH_3 or PH_3

(b) PH_3 or AsH_3

Answers:

(a) NH_3;

(b) AsH_3

Concept Exercise

In order of strength of intermolecular attraction, which of the following is the strongest: temporary dipole, permanent dipole, or hydrogen bond?

Answer: See Appendix G, 11.6.

Heats of Fusion and Vaporization

Water also has surprisingly high values for heat of fusion and heat of vaporization. This is illustrated in Table 11.6. Excluding water, we notice that as the molar mass of H_2S through H_2Te increases, the values for the heat of fusion and heat of vaporization increase. The explanation for this trend is that as the molecular size increases, the attractive forces increase slightly. Therefore, more energy is required to melt a solid or vaporize a liquid. The unusually high values for water are due to hydrogen bonding.

11.9 Chemical Properties of Water

LEARNING OBJECTIVE

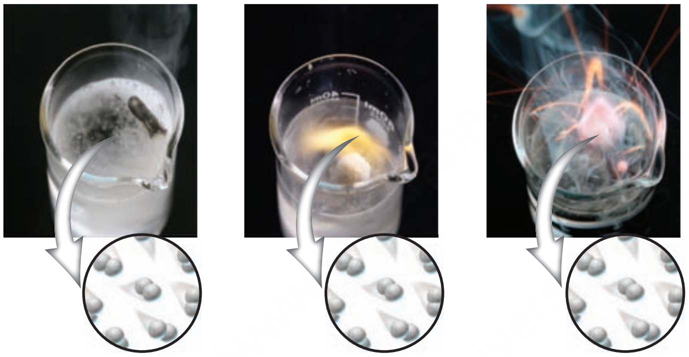
- Write chemical equations for the chemical reactions of water.

Previously, we studied five basic types of chemical reactions. Water is usually the solvent for these chemical reactions, although water can also be a reactant. One example is the passing of an electric current through water, which decomposes H_2O into hydrogen and oxygen gases. This process is called the **electrolysis** of water. From the balanced chemical equation, we notice that two volumes of hydrogen are produced for every volume of oxygen.

$$2\,H_2O(l) \xrightarrow{\text{electricity}} 2\,H_2(g) + O_2(g)$$

One of the five basic types of reactions is a replacement reaction. In this reaction an active metal (Li, Na, K, Ca, Sr, or Ba) reacts directly with water to give a metal hydroxide and hydrogen gas. These reactions occur rapidly at room temperature. For example, potassium metal reacts violently with water as follows:

$$2\,K(s) + 2\,H_2O(l) \longrightarrow 2\,KOH(aq) + H_2(g)$$



Alkali Metals in Water The reaction of Li, Na, and K in water (left to right) produces hydrogen gas, which is flammable.

The oxides of many metals can react with water to yield a metal hydroxide. Hydroxide compounds are said to be basic, or alkaline. A **metal oxide** reacts with water to yield a basic solution, so a metal oxide is referred to as a *basic oxide*. For example, calcium oxide reacts with water as follows:

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq) \qquad (\textit{a base})$$

The oxides of most nonmetals react with water to yield an acidic solution. A **nonmetal oxide** reacts with water to yield an acid, so a nonmetal oxide is referred to as an *acidic oxide*. For example, carbon dioxide reacts with water as follows:

$$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(aq) \qquad (\textit{an acid})$$

Reactions That Produce Water

Water is produced by several types of reactions. The simplest reaction is the formation of water directly from hydrogen and oxygen. In this reaction, hydrogen and oxygen gases react to give H_2O. The reaction takes place very slowly at room temperature but explosively if exposed to a flame or spark. From the balanced chemical equation, we note that two volumes of hydrogen react with one volume of oxygen.

$$2\,H_2(g) + O_2(g) \xrightarrow{\text{spark}} 2\,H_2O(l)$$

Another reaction that produces water is the combustion of hydrocarbons. Hydrocarbons are organic compounds containing hydrogen and carbon. They burn in oxygen to give carbon dioxide and water. For example, propane, C_3H_8, undergoes combustion as follows:

$$C_3H_8(g) + 5\,O_2(g) \xrightarrow{\text{spark}} 3\,CO_2(g) + 4\,H_2O(g)$$

Organic compounds containing hydrogen and oxygen also undergo combustion to give carbon dioxide and water. Ethanol, C_2H_5OH, for example, is currently blended with gasoline to produce gasohol. It undergoes combustion to give carbon dioxide and water as follows:

$$C_2H_5OH(g) + 3\,O_2(g) \xrightarrow{\text{spark}} 2\,CO_2(g) + 3\,H_2O(g)$$

Recall that neutralization reactions also produce water. An acid neutralizes a base to produce an aqueous salt and water. For example, sulfuric acid, H_2SO_4, reacts with aqueous sodium hydroxide, NaOH, to produce sodium sulfate and water.

$$H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2\,H_2O(l)$$

The decomposition of a hydrate compound also produces water. A **hydrate** is a crystalline compound that contains a specific number of water molecules. Gypsum is a hydrate of calcium sulfate, $CaSO_4$. The formula $CaSO_4 \cdot 2\,H_2O$ indicates that two water molecules are attached to each $CaSO_4$ formula unit. Heating a hydrate releases water from the compound. For example, heat decomposes gypsum to give $CaSO_4$ and two molecules of water. Thus,

$$CaSO_4 \cdot 2\,H_2O(s) \xrightarrow{\Delta} CaSO_4(s) + 2\,H_2O(g)$$

11.10 Hydrates

LEARNING OBJECTIVES

- Calculate the percentage of water in a hydrate.
- Determine the water of hydration for a hydrate.

A hydrate is a crystalline compound containing a specific number of water molecules. Common examples of hydrates include borax, $Na_2B_4O_7 \cdot 10\,H_2O$, and Epsom salts, $MgSO_4 \cdot 7\,H_2O$. The dot ($\cdot$) in the hydrate formula indicates that water molecules are attached to each unit of hydrate. In Epsom salts, $MgSO_4 \cdot 7\,H_2O$, seven molecules of water are attached to each $MgSO_4$ formula unit.

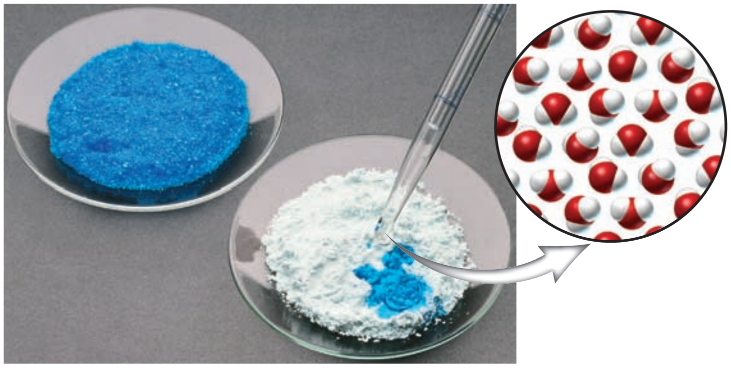
Heating a hydrate drives off the water and produces an **anhydrous compound** (without water). For example, when we heat copper(II) sulfate pentahydrate, it decomposes to give anhydrous copper(II) sulfate and water. The equation for the reaction is

$$CuSO_4 \cdot 5\,H_2O(s) \xrightarrow{\Delta} CuSO_4(s) + 5\,H_2O(g)$$

The water molecules in the hydrate are referred to as the **water of hydration**, or *water of crystallization*. Thus, the water of hydration for copper(II) sulfate pentahydrate is 5. Figure 11.13 shows the hydrate and anhydrous forms of copper(II) sulfate.

IUPAC prescribes rules for the nomenclature of hydrate compounds. According to these rules, the name of the anhydrous compound appears first, followed by the number for the water of hydration (as indicated by a Greek prefix) and the word "hydrate." For example, gypsum, $CaSO_4 \cdot 2\,H_2O$, is systematically named calcium sulfate dihydrate; the *di*–indicates the water of hydration is 2. The following example exercise further illustrates how to name hydrate compounds:

▲ **Mineral Turquoise** The mineral turquoise is an example of a tetrahydrate that contains copper, $CuAl_6(PO_4)_4(OH)_8 \cdot 4\,H_2O$.



▶ **Figure 11.13 Hydrate and Anhydrous Copper(II) Sulfate** Copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5\,H_2O$, crystals are deep blue, whereas anhydrous copper(II) sulfate, $CuSO_4$, is a white powder. A few drops of water produce the blue color of the hydrate.

EXAMPLE EXERCISE 11.7 Naming Hydrates

Supply a systematic name for each of the following hydrate compounds:

(a) $Co(CN)_3 \cdot 3\ H_2O$ (b) $FeSO_4 \cdot H_2O$

Solution

First, name the anhydrous compound and then indicate the water of hydration by a Greek prefix. (Refer to Table 7.4 if necessary.)

(a) $Co(CN)_3$ is named cobalt(III) cyanide according to the Stock system, and cobaltic cyanide according to the Latin system. The Greek prefix for 3 is *tri–*, so the hydrate is named *cobalt(III) cyanide trihydrate*, or *cobaltic cyanide trihydrate*.

(b) $FeSO_4$ is named iron(II) sulfate according to the Stock system, and ferrous sulfate according to the Latin system. The Greek prefix for 1 is *mono–*, so the hydrate is named *iron(II) sulfate monohydrate*, or *ferrous sulfate monohydrate*.

Practice Exercise

Provide the chemical formula for each of the following hydrate compounds:

(a) zinc sulfate heptahydrate
(b) sodium chromate tetrahydrate

Answers:

(a) $ZnSO_4 \cdot 7\ H_2O$ (b) $Na_2CrO_4 \cdot 4\ H_2O$

Concept Exercise

What are the systematic names for the following hydrates: $CaCl_2 \cdot H_2O$, $CaCl_2 \cdot 2\ H_2O$, and $CaCl_2 \cdot 6\ H_2O$?

Answers: See Appendix G, 11.7.

Percent Water in a Hydrate

In Section 8.7 we studied the percent composition of compounds. Here, we see that the percent water in a hydrate is the ratio of the mass of water in the hydrate to the mass of the hydrate, all multiplied by 100%.

$$\frac{\text{mass of water}}{\text{mass of hydrate}} \times 100\% = \%\ H_2O$$

As an example, we'll find the percentage of water in gypsum, $CaSO_4 \cdot 2\ H_2O$. Using the periodic table, we find the molar mass of H_2O to be 18.02 g, and that of $CaSO_4$ to be 136.15 g. We find the percentage of water from the following ratio:

$$\frac{2(18.02)\ \cancel{g}}{136.15\ \cancel{g} + 2(18.02)\ \cancel{g}} \times 100\% = 20.93\%$$

Notice that in the numerator we multiplied the molar mass of water by 2 because the compound is a dihydrate. In the denominator, we added twice the molar mass of water to that of the anhydrous compound. The following example exercise provides additional practice in calculating the percent water in a hydrate.

EXAMPLE EXERCISE 11.8 Percentage of Water in Hydrates

Calculate the percentage of water in each of the following hydrates:

(a) $CuSO_4 \cdot 5\ H_2O$ (b) $Na_2B_4O_7 \cdot 10\ H_2O$

Solution

In each example, first obtain the molar mass of the anhydrous compound from the periodic table. The molar mass of water is 18.02 g.

(a) Using the periodic table, we calculate that the molar mass of $CuSO_4$ is 159.62 g. Because the water of hydration for the hydrate is 5, we have

$$\frac{5(18.02)\ \cancel{g}}{159.62\ \cancel{g} + 5(18.02)\ \cancel{g}} \times 100\% = \%\ H_2O$$
$$= 36.08\%$$

(b) Using the periodic table, we calculate that the molar mass of $Na_2B_4O_7$ is 201.22 g. The water of hydration for the hydrate is 10, so we have

$$\frac{10(18.02)\ \cancel{g}}{201.22\ \cancel{g} + 10(18.02)\ \cancel{g}} \times 100\% = \%\ H_2O$$
$$= 47.24\%$$

Practice Exercise

Calculate the percentage of water in each of the following hydrates:

(a) $NaC_2H_3O_2 \cdot 3\ H_2O$ (b) $Na_2S_2O_3 \cdot 5\ H_2O$

Answers:

(a) 39.72% (b) 36.30%

Concept Exercise

Which of the following hydrates has the highest percentage of water:
$CaCl_2 \cdot H_2O$, $CaCl_2 \cdot 2\ H_2O$, or $CaCl_2 \cdot 6\ H_2O$?

Answer: See Appendix G, 11.8.

Determining the Chemical Formula of a Hydrate

In Section 8.8, we calculated the empirical formula for a compound from its percent composition. To determine the water of hydration for a hydrate we will proceed in a similar fashion.

The empirical formula of a compound is the simplest whole-number ratio of its elements. Similarly, the formula of a hydrate is the simplest whole-number ratio of water molecules to the anhydrous compound. In the chemical formula for washing soda, $Na_3PO_4 \cdot XH_2O$, X represents the water of hydration. To determine the value of X we must know the percent water. Let's assume that $Na_3PO_4 \cdot XH_2O$ is found by experiment to contain 52.3% water. The equation for the decomposition reaction is as follows:

$$Na_3PO_4 \cdot XH_2O(s) \xrightarrow{\Delta} Na_3PO_4(s) + XH_2O(g)$$

Because the hydrate contains 52.3% water, the percentage of Na_3PO_4 is 47.7% (100.0 − 52.3 = 47.7%). As in an empirical formula, let's assume we have a 100.0-g sample of hydrate. Therefore, our sample has 52.3 g of water and 47.7 g of anhydrous compound. The next step is to calculate the moles of water.

$$52.3\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 2.90\ mol\ H_2O$$

Using the periodic table, we calculate that the molar mass of Na_3PO_4 is 163.94 g. The moles of the anhydrous compound are

$$47.7\ \cancel{g\ Na_3PO_4} \times \frac{1\ mol\ Na_3PO_4}{163.94\ \cancel{g\ Na_3PO_4}} = 0.291\ mol\ Na_3PO_4$$

We can write the mole ratio of the hydrate as $Na_3PO_4 \cdot (2.90/0.291)H_2O$, where the water of hydration is 2.90/0.291. This ratio simplifies to 10/1, meaning that the water of hydration is 10 and the formula of the hydrate is $Na_3PO_4 \cdot 10\ H_2O$.

EXAMPLE EXERCISE 11.9 Water of Hydration for a Hydrate

Determine the water of hydration for a hydrate of magnesium iodide. In an experiment, $MgI_2 \cdot XH_2O$ was found to contain 34.0% water.

Solution

We begin by writing an equation for the decomposition.

$$MgI_2 \cdot XH_2O(s) \xrightarrow{\Delta} MgI_2(s) + XH_2O(g)$$

Because the hydrate contains 34.0% water, the percentage of MgI_2 is 66.0% (100.0 − 34.0 = 66.0%). Assume that we have a 100.0-g sample of hydrate. Therefore, we have 34.0 g of water and 66.0 g of anhydrous compound. The moles of water are

$$34.0\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 1.89\ mol\ H_2O$$

Using the periodic table, we calculate the molar mass of MgI_2 as 278.11 g. We find the moles of the anhydrous compound as follows:

$$66.0\ \cancel{g\ MgI_2} \times \frac{1\ mol\ MgI_2}{278.11\ g\ MgI_2} = 0.237\ mol\ MgI_2$$

We can write the mole ratio of the hydrate as $MgI_2 \cdot (1.89/0.237)\ H_2O$. The ratio 1.89/0.237 reduces to 7.97 and rounds off to the whole number 8. The water of hydration is 8 and the chemical formula is $MgI_2 \cdot 8\ H_2O$. The name of the hydrate is *magnesium iodide octahydrate*.

Practice Exercise

Determine the water of hydration for the hydrate of copper(II) fluoride. In an experiment, $CuF_2 \cdot XH_2O$ was found to contain 26.2% water.

Answer: $CuF_2 \cdot 2\ H_2O$

Concept Exercise

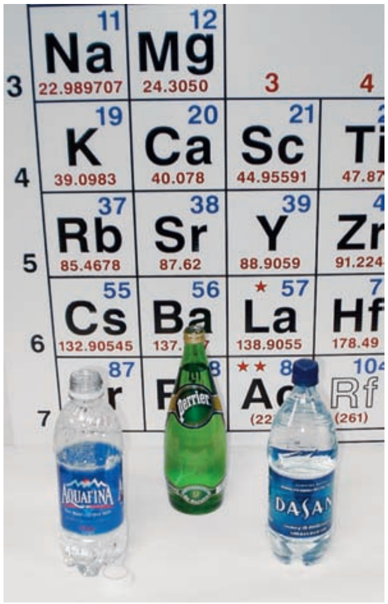
Which of the following hydrates has the highest water of hydration: $CaCl_2 \cdot H_2O$, $CaCl_2 \cdot 2\ H_2O$, or $CaCl_2 \cdot 6\ H_2O$?

Answer: See Appendix G, 11.9.

A CLOSER LOOK Bottled Water

Q: What is the difference between bottled water and tap water?

Recently, bottled water has become a huge industry and Americans consume nearly six billion gallons of bottled water each year. One survey found that more than half of all Americans drink bottled water, and about a third of the public consumes it regularly. Many different types of bottled water are now available including purified water, spring water, mineral water, and vitamin-enriched bottled water.



◀ **Bottled Water** Aquafina, Perrier, and Dasani are popular brands of bottled water.

The water we drink from a home faucet is referred to as *tap water*. It is usually chlorinated for sanitation purposes. Tap water that has been processed by distillation or deionization is referred to as *purified water*. Some brands of purified water add minerals to mimic the taste of drinking water. Water obtained from natural springs that lie deep within the Earth is referred to as *spring water*. Although spring water does not require purification, the transportation of spring water can jeopardize purity. This fact has led some spring water suppliers to bottle the water directly at the source of the natural spring.

Consumers have been led to believe that bottled water is superior to tap water in terms of purity and health benefits. Surprisingly, in most cases bottled water adheres to purity standards that are less strict than tap water. Although the Environmental Protection Agency (EPA) regulates the quality of public water supplies, the EPA has no authority over bottled water. Bottled water that crosses state lines is considered a food product and is overseen by the Food and Drug Administration (FDA), which does require that the water be bottled in sanitary conditions.

In a study by the Natural Resources Defense Council that tested 38 brands of California bottled water, eight samples were found to have measurable levels of contamination. In a study at Case Western Reserve University and Ohio State University, 57 samples of bottled water were compared to Cleveland's tap water. Although 39 of the bottled water samples were purer than the tap water, 15 of the bottles had significantly higher bacteria levels. However, the scientists stated that all of the water was safe to drink.

An example of a sparkling mineral water is Perrier®, which comes from a source in the south of France. A distinguishing characteristic of sparkling mineral water is its natural carbonation. Carbonation is the result of volcanic activity within Earth. Heat from a volcano produces steam that passes through limestone ($CaCO_3$), which in turn releases carbon dioxide gas into the mineral water.

A: The difference between bottled water and tap water depends on the source. Some bottled water originates from natural springs; whereas, other bottled water is distilled or deionized tap water, sometimes with minerals added.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
11.1 Properties of Liquids The liquid state has a variable shape but a fixed volume. Liquids usually flow readily but do not compress or expand as do gases. The densities of liquids vary but are approximately 1000 times greater than the densities of gases. Liquids mix and diffuse uniformly in a container, similar to gases.	• Identify five observed properties of a liquid. ***Related Exercises: 1–4***
11.2 The Intermolecular Bond Concept There are three basic types of intermolecular attraction: dispersion forces, dipole forces, and hydrogen bonds. A **dispersion force** has the weakest attraction and results from temporary dipoles in molecules. A **dipole force** has a stronger attraction and results from permanent dipoles in molecules. A **hydrogen bond** exerts the strongest intermolecular attraction, and occurs between molecules having H—O, or H—N bonds.	• Explain the concept of an intermolecular bond. ***Related Exercises: 5–6*** • Describe three types of attraction between molecules in a liquid: temporary dipoles, permanent dipoles, and hydrogen bonds. ***Related Exercises: 5–6***
11.3 Vapor Pressure, Boiling Point, Viscosity, and Surface Tension When the **vapor pressure** of a liquid equals the atmospheric pressure, the liquid is at its **boiling point**. The normal boiling point of a liquid is the temperature at which the vapor pressure is 760 mm Hg. Liquids have a resistance to flow, and this property is called **viscosity**. The attraction between molecules at the surface of a liquid is called **surface tension**. Each of these properties is affected by the degree of attraction between molecules in the liquid. If the attraction is low, the vapor pressure is high. Conversely, if the attraction is high, the values for the boiling point, viscosity, and surface tension are also high.	• Describe the relationship between intermolecular attraction in a liquid and the properties of vapor pressure, boiling point, viscosity, and surface tension. ***Related Exercises: 7–12*** • Predict which liquid in a pair has the higher vapor pressure, boiling point, viscosity, and surface tension. ***Related Exercises: 13–22***
11.4 Properties of Solids The solid state has a fixed shape and a fixed volume. A solid that has a highly defined structure is called a **crystalline solid**. Solids do not compress or expand to any large degree. The density of a substance in the solid state is usually higher than that in the liquid state. Water and ammonia are two interesting exceptions to this rule. Solid ice and solid ammonia float on their respective liquids. Because the solid state contains particles in fixed positions, solids do not mix or diffuse.	• Identify five observed properties of a solid. ***Related Exercises: 23–26***
11.5 Crystalline Solids There are three basic types of crystalline solids. **Ionic solids** are made up of ions. The ions are attracted to each other and form repeating geometric patterns. **Molecular solids** form crystals made up of molecules. **Metallic solids** are made up of metal atoms arranged in a definite pattern. The properties of ionic, molecular, and metallic solids differ. Metallic solids are good conductors of electricity. Ionic solids conduct electricity only when they are melted or are in an aqueous solution. As a rule, molecular solids do not conduct electricity.	• Describe three types of crystalline solids: ionic, molecular, and metallic. ***Related Exercises: 27–32***

Key Concepts

Learning Objectives and Related Exercises

11.6 Changes of Physical State

Specific heat is the heat required to raise one gram of substance one degree Celsius. The heat required to melt a substance is called the **heat of fusion**. The heat released when a substance freezes to a solid is called the *heat of solidification*. The heat of solidification of a substance has the same value as its heat of fusion. The heat required to convert a liquid to a gas at its boiling point is the **heat of vaporization**. When a vapor condenses to a liquid, it releases the same amount of heat energy that was necessary to vaporize the liquid. This energy is termed the *heat of condensation*.

- Explain the concepts of specific heat, heat of fusion, and heat of vaporization.
  Related Exercises: 33–34
- Calculate heat changes that involve specific heat, heat of fusion, and heat of vaporization.
  Related Exercises: 35–42

11.7 Structure of Water

Water has unusual properties because of the two O—H bonds separated by a **bond angle** of 104.5 °C. The oxygen atom has a partial negative charge, and each hydrogen atom has a partial positive charge. These two dipoles create an overall **net dipole** for a water molecule. Water is classified as **hard water**, or **soft water**, depending on the cations and anions present. Water can be boiled and collected to give **distilled water**, or passed through an ion-exchange resin to give pure **deionized water**.

- Illustrate the bond angle and net dipole in a water molecule.
  Related Exercises: 43–50

11.8 Physical Properties of Water

Hydrogen bonds between water molecules explain why water has extraordinarily high values for boiling point, viscosity, and surface tension. Hydrogen bonding also explains why the density of ice is less than that of water. For a small molecule, water has an unusually high melting point, boiling point, heat of fusion, and heat of vaporization.

- Explain the unusual physical properties of water.
  Related Exercises: 51–60

11.9 Chemical Properties of Water

Water undergoes the following chemical reactions: (1) An electric current decomposes water into hydrogen and oxygen gases by **electrolysis**. (2) The active metals of Groups IA/1 and IIA/2 react with water to give a metal hydroxide and hydrogen gas. (3) A **metal oxide** reacts with water to give a metal hydroxide. (4) A **nonmetal oxide** combines with water to yield an acidic solution. (See Table 11.7.)

- Write chemical equations for the chemical reactions of water.
  Related Exercises: 61–70

TABLE 11.7 Summary of the Chemical Reactions of Water

water	$\xrightarrow{\text{electricity}}$	hydrogen + oxygen
active metal + water	$\longrightarrow$	metal hydroxide + hydrogen
metal oxide + water	$\longrightarrow$	basic solution
nonmetal oxide + water	$\longrightarrow$	acidic solution

Water is produced from (1) the reaction of hydrogen and oxygen gases, (2) burning hydrocarbons, (3) neutralizing an acid with a base, and (4) heating a **hydrate**. (See Table 11.8.)

TABLE 11.8 Summary of the Chemical Reactions Producing Water

hydrogen + oxygen	$\xrightarrow{\text{spark}}$	water
hydrocarbon + oxygen	$\xrightarrow{\text{spark}}$	carbon dioxide + water
acid + base	$\longrightarrow$	salt + water
hydrate	$\xrightarrow{\Delta}$	anhydrous compound + water

Key Concepts	Learning Objectives and Related Exercises
11.10 Hydrates The name of a hydrate consists of the name of the **anhydrous compound** followed by a Greek prefix and the word "hydrate"; for example, $BaCl_2 \cdot 2H_2O$ is named barium chloride dihydrate. Determining the chemical formula of a hydrate is similar to determining an empirical formula. From the percent water in a hydrate, we can calculate the moles of water and anhydrous compound. The ratio of moles of water to moles of the anhydrous compound gives the **water of hydration**, which is also termed *water of crystallization*.	• Calculate the percentage of water in a hydrate. ***Related Exercises: 71–76*** • Determine the water of hydration for a hydrate. ***Related Exercises: 77–78***

Key Terms *Answers to Key Terms are in Appendix H.*

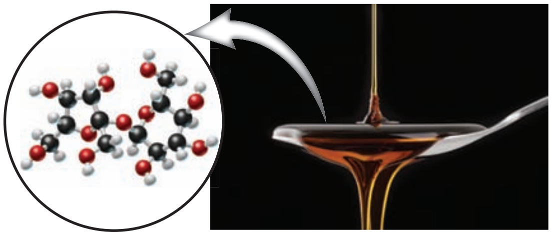
Select the key term that corresponds to each of the following definitions.

______ **1.** an intermolecular attraction based on temporary dipoles
______ **2.** an intermolecular attraction based on permanent dipoles
______ **3.** an intermolecular attraction based on permanent dipoles involving a H atom bonded to O or N
______ **4.** the pressure exerted by gaseous molecules above a liquid in a closed container when the rates of vaporization and condensation are equal
______ **5.** the temperature at which the vapor pressure of a liquid is equal to the atmospheric pressure
______ **6.** the resistance of a liquid to flow
______ **7.** the tendency of a liquid to form spherical drops
______ **8.** a solid substance composed of ions or molecules that repeat in a regular geometric pattern
______ **9.** a crystalline solid composed of ions that repeat in a regular pattern
______ **10.** a crystalline solid composed of molecules that repeat in a regular pattern
______ **11.** a crystalline solid composed of metal atoms that repeat in a regular pattern
______ **12.** the heat required to raise 1 g of any substance 1 °C
______ **13.** the heat required to convert a solid to a liquid at its melting point
______ **14.** the heat required to vaporize a liquid to a gas at its boiling point
______ **15.** the angle formed by two atoms bonded to a central atom in a molecule
______ **16.** the overall direction of partial negative charge in a molecule having two or more dipoles
______ **17.** the chemical reaction produced by the passage of electric current through an aqueous solution
______ **18.** a compound that reacts with water to form a basic solution; also termed a *basic oxide*
______ **19.** a compound that reacts with water to form an acidic solution; also termed an *acidic oxide*
______ **20.** a substance that contains a specific number of water molecules in a crystalline compound
______ **21.** refers to a compound that does not contain water
______ **22.** the number of water molecules in a hydrate
______ **23.** water containing a variety of cations and anions such as Ca^{2+}, Mg^{2+}, Fe^{3+}, CO_3^{2-}, SO_4^{2-}, and PO_4^{3-}
______ **24.** water containing sodium ions and a variety of anions
______ **25.** water purified by removing ions using an ion exchange method; also termed *demineralized water*
______ **26.** water purified by boiling and collecting the condensed vapor

(a) anhydrous compound (*Sec. 11.10*)
(b) boiling point (Bp) (*Sec. 11.3*)
(c) bond angle (*Sec. 11.7*)
(d) crystalline solid (*Sec. 11.4*)
(e) deionized water (*Sec. 11.7*)
(f) dipole force (*Sec. 11.2*)
(g) dispersion force (*Sec. 11.2*)
(h) distilled water (*Sec. 11.7*)
(i) electrolysis (*Sec. 11.9*)
(j) hard water (*Sec. 11.7*)
(k) heat of fusion (*Sec. 11.6*)
(l) heat of vaporization (*Sec. 11.6*)
(m) hydrate (*Sec. 11.9*)
(n) hydrogen bond (*Sec. 11.2*)
(o) ionic solid (*Sec. 11.5*)
(p) metal oxide (*Sec. 11.9*)
(q) metallic solid (*Sec. 11.5*)
(r) molecular solid (*Sec. 11.5*)
(s) net dipole (*Sec. 11.7*)
(t) nonmetal oxide (*Sec. 11.9*)
(u) soft water (*Sec. 11.7*)
(v) specific heat (*Sec. 11.6*)
(w) surface tension (*Sec. 11.3*)
(x) vapor pressure (*Sec. 11.3*)
(y) viscosity (*Sec. 11.3*)
(z) water of hydration (*Sec. 11.10*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Properties of Liquids (Sec. 11.1)



▲ **The Liquid State** Maple syrup can be poured and its rate of flow depends on its viscosity.

1. Which of the following are observed general properties of liquids?
   (a) fixed volume **(b)** mix homogeneously
   (c) compress significantly **(d)** more dense than gases

2. Which of the following are observed general properties of liquids?
   (a) variable shape **(b)** flow readily
   (c) expand significantly **(d)** less dense than gases

3. Indicate the physical state (solid, liquid, or gas) for each of the following at the given temperature.
   (a) H_2O at −10 °C **(b)** H_2O at 110 °C
   (c) NH_3 at −100°C **(d)** NH_3 at −50 °C

   The melting points and boiling points for water and ammonia are as follows.

Compound	Melting Point	Boiling Point
water, H_2O	0 °C	100 °C
ammonia, NH_3	−78 °C	−33 °C

4. Indicate the physical state (solid, liquid, gas) for each of the following at the given temperature.
   (a) Ne at −225 °C **(b)** Ne at −255 °C
   (c) Ar at −175 °C **(d)** Ar at −200 °C

   The melting points and boiling points for neon and argon are as follows.

Element	Melting Point	Boiling Point
neon, Ne	−249 °C	−246 °C
argon, Ar	−189 °C	−186 °C

The Intermolecular Bond Concept (Sec. 11.2)

5. Distinguish temporary dipole attraction from permanent dipole attraction between molecules in a liquid.

6. Distinguish permanent dipole attraction from hydrogen bond attraction between molecules in a liquid.

Vapor Pressure, Boiling Point, Viscosity, and Surface Tension (Sec. 11.3)

7. Describe and illustrate the concept of vapor pressure using water as an example.

8. Describe and illustrate the concept of boiling point using water as an example.

9. Describe and illustrate the concept of viscosity using water as an example.

10. Describe and illustrate the concept of surface tension using water as an example.

11. If the molecules in a liquid are strongly attracted, which of the following properties has a high value?
    (a) vapor pressure **(b)** boiling point
    (c) viscosity **(d)** surface tension

12. If the molecules in a liquid are weakly attracted, which of the following properties has a high value?
    (a) vapor pressure **(b)** boiling point
    (c) viscosity **(d)** surface tension

13. Predict which liquid in each pair has the higher vapor pressure:
    (a) CH_3COOH or C_2H_5Cl **(b)** C_2H_5OH or CH_3OCH_3

14. Predict which liquid in each pair has the higher boiling point:
    (a) CH_3COOH or C_2H_5Cl **(b)** C_2H_5OH or CH_3OCH_3

15. Predict which liquid in each pair has the higher viscosity:
    (a) CH_3COOH or C_2H_5Cl **(b)** C_2H_5OH or CH_3OCH_3

16. Predict which liquid in each pair has the higher surface tension:
    (a) CH_3COOH or C_2H_5Cl **(b)** C_2H_5OH or CH_3OCH_3

17. What is the general relationship between the vapor pressure of a liquid and its temperature?

18. What is the general relationship between the boiling point of a liquid and its vapor pressure?

19. Refer to Figure 11.5 and estimate the vapor pressure of ethanol at each of the following temperatures:
    (a) 45 °C **(b)** 60 °C

20. Refer to Figure 11.5 and estimate the vapor pressure of ether at each of the following temperatures:
    (a) 15 °C **(b)** 30 °C

21. The vapor pressure of acetone is 1 mm Hg at −59 °C, 10 mm Hg at −31 °C, 100 mm Hg at 8 °C, 400 mm Hg at 40 °C, and 760 mm Hg at 56 °C. What is the normal boiling point of acetone?

22. The vapor pressure of methanol is 1 atm at 65 °C, 2 atm at 84 °C, 5 atm at 112 °C, 10 atm at 138 °C, and 20 atm at 168 °C. What is the normal boiling point of methanol?

Properties of Solids (Sec. 11.4)

23. Which of the following are observed general properties of solids?
    (a) fixed volume **(b)** noncrystalline structure
    (c) do not expand **(d)** alloys mix by diffusion

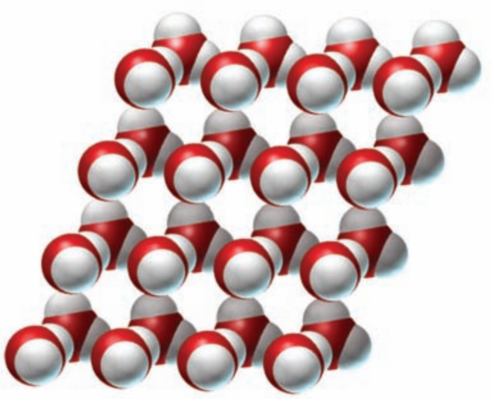
24. Which of the following are observed general properties of solids?
    (a) fixed shape **(b)** crystalline structure
    (c) do not compress **(d)** less dense than liquids

25. Indicate the physical state (solid, liquid, or gas) for gallium given its melting point (30 °C) and boiling point (2403 °C) under the following conditions:
    (a) Ga in ice water **(b)** Ga in boiling water

26. Indicate the physical state (solid, liquid, or gas) for mercury given its melting point (−40 °C) and boiling point (357 °C) under the following conditions:
    (a) Hg in ice water **(b)** Hg in boiling water

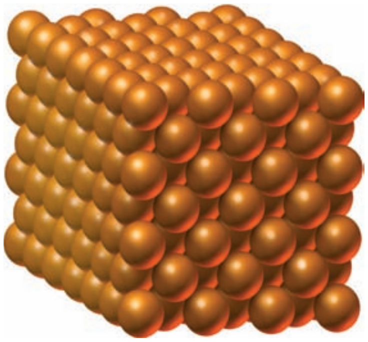
Crystalline Solids (Sec. 11.5)

27. List three examples of crystalline solids.
28. List three examples of noncrystalline solids.
29. State the type of particles that compose each of the following:
    (a) ionic solid **(b)** molecular solid
    (c) metallic solid



▲ A Molecular Solid, H_2O

30. State whether the following list of properties is most descriptive of an ionic, a molecular, or a metallic solid:
    (a) malleable, ductile, insoluble in water
    (b) high melting point, conductor of electricity when melted
    (c) low melting point, nonconductor of electrical current
31. Classify each of the following crystalline solids as ionic, molecular, or metallic.
    (a) sulfur, S_8 **(b)** sulfur dioxide, SO_2
    (c) copper, Cu **(d)** silver nitrate, $AgNO_3$



◀ A Metallic Solid, Cu

32. Classify each of the following crystalline solids as ionic, molecular, or metallic.
    (a) zinc, Zn
    (b) zinc oxide, ZnO
    (c) phosphorus, P_4
    (d) iodine monobromide, IBr

Changes of Physical State (Sec. 11.6)

33. Draw the general shape of the temperature–energy graph for the heating of ethanol from −120 °C to 120 °C. (*Given*: Mp = −117.3 °C; Bp = 78.5 °C)
34. Draw the general shape of the temperature–energy graph for the cooling of acetone from 100 °C to −100 °C. (*Given*: Mp = −95.4 °C; Bp = 56.2 °C)
35. Calculate the heat absorbed when melting 25.0 g of ice at 0 °C.
36. Calculate the heat released when freezing 25.0 g of water at 0 °C.
37. Calculate the heat required to convert 25.0 g of water at 25.0 °C to steam at 100.0 °C.
38. Calculate the heat released when 65.5 g of water at 55.5 °C cools to ice at 0.0 °C.
39. Calculate the heat required to convert 115 g of ice at 0.0 °C to steam at 100.0 °C.
40. Calculate the heat released when 155 g of steam at 100.0 °C cools to ice at 0.0 °C.
41. Calculate the heat required to convert 38.5 g of ice at −20.0 °C to steam at 100.0 °C.
42. Calculate the heat released when 90.5 g of steam at 110.0 °C cools to ice at 0.0 °C.

Structure of Water (Sec. 11.7)

43. Is the vapor pressure of water unusually high or low?
44. Is the boiling point of water unusually high or low?
45. Is the viscosity of water unusually high or low?
46. Is the surface tension of water unusually high or low?
47. What is the observed bond angle in a water molecule?
48. What type of intermolecular bond exists between water molecules?
49. Draw the two dipoles in a water molecule using delta convention.
50. Draw the net dipole for a water molecule using standard convention.

Physical Properties of Water (Sec. 11.8)

51. Which liquid has the highest specific heat?
52. Which liquid has the highest heat of fusion and heat of vaporization?
53. A solid cube of ice floats in water. Which is more dense: ice or water?
54. A solid cube of ammonia floats in liquid ammonia. Which is more dense: solid or liquid ammonia?
55. Without referring to Table 11.6, predict which compound in each of the following pairs has the higher melting point:
    (a) H_2O or H_2S **(b)** H_2S or H_2Se
56. Without referring to Table 11.6, predict which compound in each of the following pairs has the higher boiling point:
    (a) H_2O or H_2Se **(b)** H_2S or H_2Te
57. Without referring to Table 11.6, predict which compound in each of the following pairs has the higher heat of fusion (cal/mol):
    (a) H_2O or H_2S **(b)** H_2S or H_2Se
58. Without referring to Table 11.6, predict which compound in each of the following pairs has the higher heat of vaporization (cal/mol):
    (a) H_2O or H_2Se **(b)** H_2S or H_2Te
59. In general, as the molar mass of Group VIA/16 hydrogen compounds increases, does each of the following increase or decrease?
    (a) melting point **(b)** boiling point
    (c) heat of fusion **(d)** heat of vaporization
60. Refer to the trends in Table 11.6 and estimate a value for the physical properties of radioactive H_2Po, that is, predict a value for the Mp, Bp, H_{fusion}, and H_{vapor}.

Chemical Properties of Water (Sec. 11.9)

61. Write a balanced equation for the reaction of hydrogen and oxygen gases.

62. Write a balanced equation for the decomposition of water by electrolysis.

63. Complete and balance the following equations.
    - **(a)** $Li(s) + H_2O(l) \rightarrow$
    - **(b)** $Na_2O(s) + H_2O(l) \rightarrow$
    - **(c)** $CO_2(g) + H_2O(l) \rightarrow$

64. Complete and balance the following equations.
    - **(a)** $Na(s) + H_2O(l) \rightarrow$
    - **(b)** $K_2O(s) + H_2O(l) \rightarrow$
    - **(c)** $N_2O_3(g) + H_2O(l) \rightarrow$

65. Complete and balance the following equations.
    - **(a)** $Mg(s) + H_2O(l) \rightarrow$
    - **(b)** $SrO(s) + H_2O(l) \rightarrow$
    - **(c)** $N_2O_5(g) + H_2O(l) \rightarrow$

66. Complete and balance the following equations.
    - **(a)** $Ca(s) + H_2O(l) \rightarrow$
    - **(b)** $BaO(s) + H_2O(l) \rightarrow$
    - **(c)** $P_2O_5(s) + H_2O(l) \rightarrow$

67. Complete and balance the following equations:
    - **(a)** $C_3H_6(g) + O_2(g) \xrightarrow{\text{spark}}$
    - **(b)** $Na_2Cr_2O_7 \cdot 2\,H_2O(s) \xrightarrow{\Delta}$
    - **(c)** $HF(aq) + Ca(OH)_2(aq) \rightarrow$

68. Complete and balance the following equations.
    - **(a)** $C_4H_{10}(g) + O_2(g) \xrightarrow{\text{spark}}$
    - **(b)** $Co(C_2H_3O_2)_2 \cdot 4\,H_2O(s) \xrightarrow{\Delta}$
    - **(c)** $HNO_3(aq) + Ba(OH)_2(aq) \rightarrow$

69. Complete and balance the following equations.
    - **(a)** $C_3H_7OH(l) + O_2(g) \xrightarrow{\text{spark}}$
    - **(b)** $Ca(NO_3)_2 \cdot 4\,H_2O(s) \xrightarrow{\Delta}$
    - **(c)** $H_2CO_3(aq) + KOH(aq) \rightarrow$

70. Complete and balance the following equations:
    - **(a)** $C_4H_9OH(l) + O_2(g) \xrightarrow{\text{spark}}$
    - **(b)** $KAl(SO_4)_2 \cdot 12\,H_2O(s) \xrightarrow{\Delta}$
    - **(c)** $H_3PO_4(aq) + NaOH(aq) \rightarrow$

Hydrates (Sec. 11.10)

71. Supply a systematic name for each of the following hydrate compounds.
    - **(a)** $MgCl_2 \cdot 6\,H_2O$
    - **(b)** $MnSO_4 \cdot 7\,H_2O$
    - **(c)** $Co(CN)_3 \cdot 3\,H_2O$
    - **(d)** $Cr(NO_3)_3 \cdot 9\,H_2O$

72. Supply a systematic name for each of the following hydrate compounds.
    - **(a)** $MgSO_4 \cdot 7\,H_2O$
    - **(b)** $MnSO_4 \cdot H_2O$
    - **(c)** $Co(C_2H_3O_2)_2 \cdot 4\,H_2O$
    - **(d)** $Cr(NO_3)_3 \cdot 3\,H_2O$

73. Provide the chemical formula for each of the following hydrate compounds.
    - **(a)** sodium acetate trihydrate
    - **(b)** calcium sulfate dihydrate
    - **(c)** potassium chromate tetrahydrate
    - **(d)** zinc sulfate heptahydrate

74. Provide the chemical formula for each of the following hydrate compounds.
    - **(a)** sodium carbonate decahydrate
    - **(b)** nickel(II) nitrate hexahydrate
    - **(c)** cobalt(III) iodide octahydrate
    - **(d)** chromium(III) acetate monohydrate

75. Calculate the percentage of water in each of the following hydrates.
    - **(a)** $MnSO_4 \cdot H_2O$
    - **(b)** $Sr(NO_3)_2 \cdot 6\,H_2O$
    - **(c)** $Co(C_2H_3O_2)_2 \cdot 4\,H_2O$
    - **(d)** $Cr(NO_3)_3 \cdot 9\,H_2O$

76. Calculate the percentage of water in each of the following hydrates.
    - **(a)** $SrCl_2 \cdot 6\,H_2O$
    - **(b)** $K_2Cr_2O_7 \cdot 2\,H_2O$
    - **(c)** $Co(CN)_3 \cdot 3\,H_2O$
    - **(d)** $Na_2CrO_4 \cdot 4\,H_2O$

77. Determine the water of hydration for the following hydrates and write the chemical formula.
    - **(a)** $SrCl_2 \cdot XH_2O$ is found to contain 18.5% water.
    - **(b)** $Sr(NO_3)_2 \cdot XH_2O$ is found to contain 33.8% water.
    - **(c)** $Ca(NO_3)_2 \cdot XH_2O$ is found to contain 30.5% water.
    - **(d)** $Na_2B_4O_7 \cdot XH_2O$ is found to contain 30.9% water.

78. Determine the water of hydration for the following hydrates and write the chemical formula.
    - **(a)** $NiCl_2 \cdot XH_2O$ is found to contain 21.7% water.
    - **(b)** $Ni(NO_3)_2 \cdot XH_2O$ is found to contain 37.2% water.
    - **(c)** $CoSO_4 \cdot XH_2O$ is found to contain 10.4% water.
    - **(d)** $CrI_3 \cdot XH_2O$ is found to contain 27.2% water.

General Exercises

79. What is the approximate percentage of the human body that is composed of water?

80. What is the approximate percentage of the Earth's surface that is covered by water?

81. What is the strongest type of intermolecular attraction that exists in each of the following liquids?
    - **(a)** C_5H_{12}
    - **(b)** $CH_3CH_2CH_2{-}Cl$
    - **(c)** $CH_3{-}OH$
    - **(d)** $CH_3{-}O{-}CH_3$

82. What is the strongest type of intermolecular attraction that exists in each of the following liquids?
    - **(a)** C_6H_{14}
    - **(b)** $CH_3CH_2CH_2{-}F$
    - **(c)** $CH_3{-}COOH$
    - **(d)** $CH_3CH_2{-}O{-}CH_2CH_3$

83. Refer to Figure 11.5 and determine the boiling point of water at an elevation where the atmospheric pressure is 0.500 atm.

84. Refer to Figure 11.5 and determine the boiling point of ethanol at an elevation where the atmospheric pressure is 0.500 atm.

85. Venus has clouds of sulfuric acid, H_2SO_4. Why is it reasonable to predict that Venus "raindrops" are spheres?

86. Saturn has clouds of ammonia, NH_3. Why is it reasonable to predict that Saturn "raindrops" are spheres?

Challenge Exercises

87. The antifreeze in some automobiles is methyl alcohol. Calculate the heat required to vaporize 5.00 L of methyl alcohol at 25.0 °C. (Methyl alcohol: density $= 0.792$ g/mL; specific heat $= 0.610$ cal/(g × °C); Bp $= 65.0$ °C; $H_{vapor} = 263$ cal/g.)

88. The antifreeze in some automobiles is ethylene glycol. Calculate the heat required to vaporize 5.00 L of ethylene glycol at 25.0 °C. (Ethylene glycol: density $= 1.11$ g/mL; specific heat $= 0.561$ cal/(g × °C); Bp $= 197.6$ °C; $H_{vapor} = 201$ cal/g.)

Chapter 11 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following is an observed property of liquids? (Sec. 11.1)
   (a) Liquids are more dense than gases.
   (b) Liquids expand and compress significantly.
   (c) Liquids have a fixed shape and variable volume.
   (d) Liquids that are insoluble mix homogeneously.
   (e) all of the above

2. What is the strongest intermolecular force in liquid ammonia, NH_3? (Sec. 11.2)
   (a) dipole forces **(b)** dispersion forces
   (c) hydrogen bonds **(d)** all of the above
   (e) none of the above

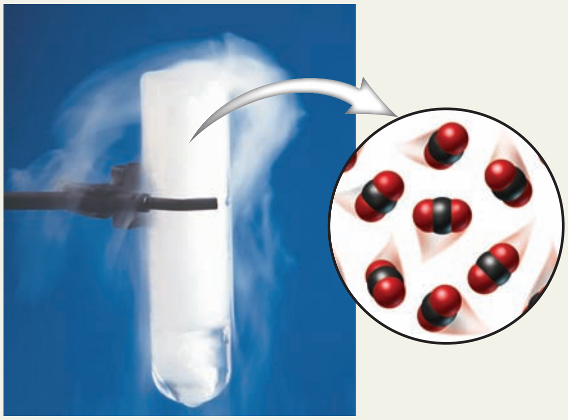
3. If the molecules in a liquid have a strong attraction for each other, which of the following properties has a relatively high value? (Sec. 11.3)
   (a) boiling point **(b)** viscosity
   (c) surface tension **(d)** all of the above
   (e) none of the above

4. Which of the following is an observed property of solids? (Sec. 11.4)
   (a) Solids have a variable shape and fixed volume.
   (b) Solids are usually less dense than liquids.
   (c) Solids can have noncrystalline structures.
   (d) Solids compress and expand significantly.
   (e) none of the above

5. Which of the following is an example of a molecular crystalline solid? (Sec. 11.5)
   (a) cuprite, Cu_2O **(b)** dry ice, CO_2
   (c) limestone, $CaCO_3$ **(d)** iron pyrite, FeS_2
   (e) none of the above

6. Calculate the heat required to convert 10.0 g of water at 100 °C to steam at 100 °C. The specific heat of water is 1.00 cal/(g × °C); the heat of fusion is 80.0 cal/g; and the heat of vaporization is 540.0 cal/g. (Sec. 11.6)
   (a) 8.0 cal **(b)** 10.0 cal
   (c) 54.0 cal **(d)** 8.00×10^2 cal
   (e) 5.40×10^3 cal

7. How many dipoles are in a water molecule? (Sec. 11.7)
   (a) 0 **(b)** 1
   (c) 2 **(d)** 3
   (e) 4



▲ Dry Ice Solid carbon dioxide, CO_2 is undergoing sublimation from a solid to a gas.

8. Which of the following explains why ice floats on water? (Sec. 11.8)
   (a) Ice has a greater density than water.
   (b) Ice has a greater heat of fusion than water.
   (c) Ice has a greater molar mass than water.
   (d) Ice has a greater specific heat than water.
   (e) Ice has a greater volume than an equal mass of water.

9. What is the product from the reaction of water and a metal oxide? (Sec. 11.9)
   (a) an acid
   (b) a base
   (c) a nonmetal hydroxide and hydrogen gas
   (d) a nonmetal hydroxide and oxygen gas
   (e) none of the above

10. An unknown hydrate of sodium dichromate, $Na_2Cr_2O_7 \cdot XH_2O$, is heated and releases 12.1% water. What is the water of crystallization (X) for the hydrate? (Sec. 11.10)
   (a) 1 **(b)** 2
   (c) 3 **(d)** 4
   (e) 12

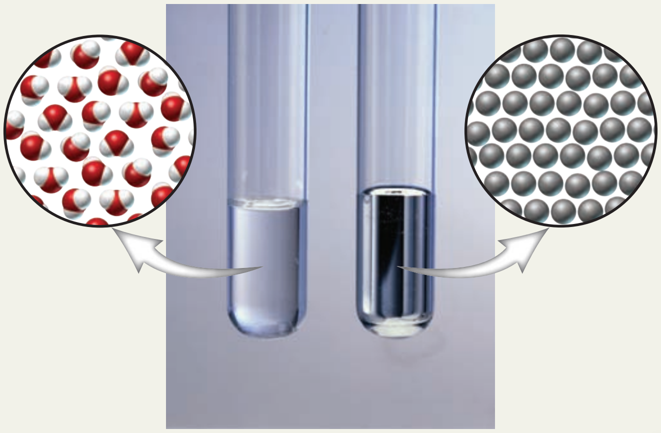
Key Concepts

11. Which physical state can be described by the following analogy? A beehive is surrounded by honey bees, which fly faster when approached by fire.

12. If a water bottle is filled completely and capped, what will happen if the water in the bottle freezes?

Critical Thinking

13. Why does it take longer to hard boil an egg in Denver, CO, than Los Angeles, CA?



▲ **Water vs. Mercury** The concave surface of water in a test tube is shown on the left, and the convex surface of mercury is shown on the right.

14. Why does water form a concave lens in a test tube, whereas mercury forms a convex lens.

CHAPTER

12

Bonding

> *"The best way to have a good idea is to have lots of ideas."*
>
> **Linus Pauling, American Chemist (1901–1994)**

12
Mg
magnesium

Element 12: Magnesium
Magnesium is a lightweight metal, exceeded only by iron and aluminum as a structural metal. Owing to its low density and high strength, magnesium metal is used in aircraft parts, racecar frames and wheels, and motorcycle engines. Magnesium metal and its alloys are flammable, and fires cannot be put out with water or carbon dioxide extinguishers.

At the beginning of the twentieth century, scientists were rapidly unraveling the mysteries of the atom. In 1897, J. J. Thomson discovered the electron and, shortly thereafter, the proton. In 1911, Ernest Rutherford unveiled the atomic nucleus. And in 1913, Niels Bohr offered evidence that electrons occupied orbits and circled the nucleus. During this same period of time, scientists began to speculate that electrons were responsible for holding molecules together. They suggested that electrons provided the "glue" that bonded one atom to another.

In 1916, American chemist G. N. Lewis (1875–1946) formulated one of the first theories of chemical bonding. He noted that the noble gases were unusually stable and that, with the exception of helium, all these gases had eight electrons in their outer shell. Lewis theorized that atoms bonded together to attain a stable noble gas electron structure. This principle, referred to as the *octet rule,* states that atoms bond in such a way that each atom attains eight electrons in its outer shell.

12.1 The Chemical Bond Concept

LEARNING OBJECTIVES

- Discuss the role of valence electrons in a chemical bond.
- Predict whether a bond is ionic or covalent given a formula unit or molecule.

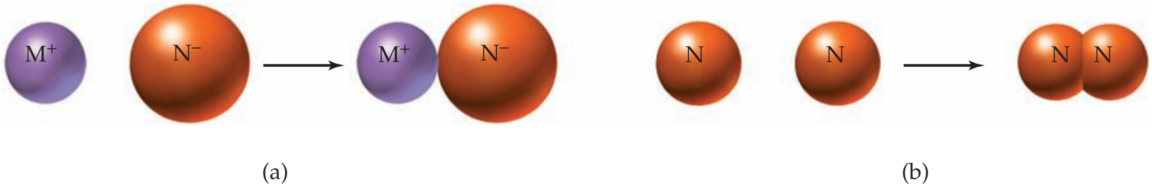
In Section 5.7, we learned that an atom has core electrons and valence electrons. Core electrons are found closer to the nucleus, whereas **valence electrons** are found in the most distant *s* and *p* energy subshells. Valence electrons are responsible for holding two or more atoms together in a **chemical bond**.

The **octet rule** states that an atom tends to bond in such a way that it acquires eight electrons in its outer shell. In the words of G. N. Lewis, the chemical bond concept can be described as follows:

> *Two atoms may conform to the rule of eight, or the octet rule, not only by the transfer of electrons from one atom to another, but also by sharing one or more pairs of electrons. These electrons which are held in common by two atoms may be considered to belong to the outer shells of both atoms.*
>
> *The simplest explanation of the predominant occurrence of an even number of electrons in the valence shells of molecules is that the electrons are definitely paired with one another. Two electrons thus coupled together, when lying between two atomic centers, and held jointly in the shells of the two atoms, I have considered to be the chemical bond.*

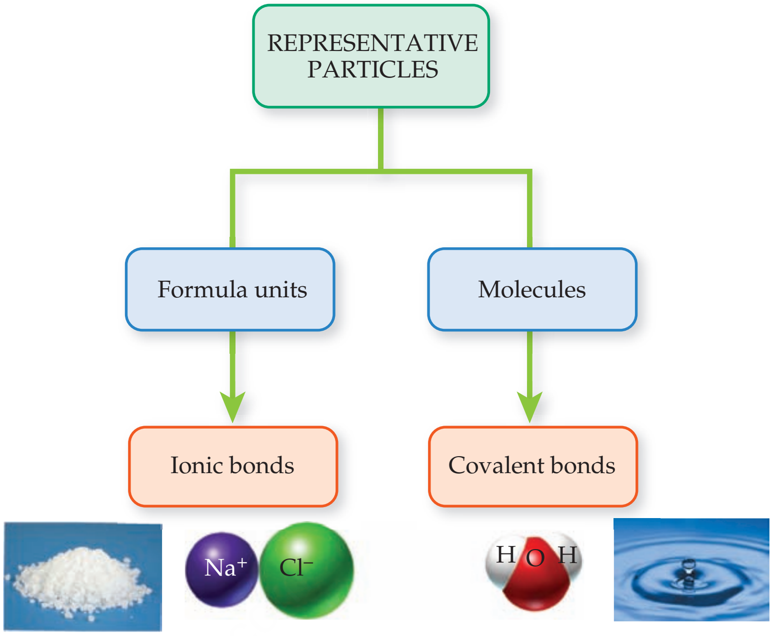
In an **ionic bond**, a metal cation is attracted to a nonmetal anion. The positively charged cation and negatively charged anion are held together by the attraction of opposite charges. A fundamental particle held together by an ionic bond is a **formula unit**. We can easily identify a formula unit by the fact that it contains metal and nonmetal ions. NaCl, MgO, and AlN all contain a metal cation and a nonmetal anion and are examples of formula units.

In a **covalent bond**, two nonmetal atoms share valence electrons. A fundamental particle held together by covalent bonds is a **molecule**. We can easily identify a molecule by the fact that it contains nonmetal atoms. H_2O, NH_3, and CH_4 all contain nonmetal atoms and are examples of molecules. The formation of an ionic bond and a covalent bond is shown in Figure 12.1.



▲ **Figure 12.1 Chemical Bonds** (a) An ionic bond forms when a metal cation (M^+) is attracted to a nonmetal anion (N^-). (b) A covalent bond forms when valence electrons are shared by two nonmetal atoms (N).

A substance can be composed of either formula units or molecules. The distinction between the types of representative particles found in a substance is shown in Figure 12.2.



▶ **Figure 12.2 Classification of Representative Particles** Particles can be formula units (for example, NaCl) held together by ionic bonds or they can be molecules (for example, H_2O) held together by covalent bonds.

The following example exercise further illustrates the relationship between bonding and the type of fundamental particles.

EXAMPLE EXERCISE 12.1 Bond Predictions

Predict whether each of the following is held together by ionic or by covalent bonds:

(a) ammonia, NH_3	(b) magnesium nitride, Mg_3N_2

Solution

A metal ion and nonmetal ion are attracted in ionic bonds; two or more nonmetal atoms are attracted in covalent bonds.

(a) Ammonia contains the nonmetals nitrogen and hydrogen. It follows that NH_3 has *covalent* bonds.

(b) Magnesium nitride contains a metal (Mg) and a nonmetal (N). It follows that Mg_3N_2 has *ionic* bonds.

Practice Exercise

Predict whether each of the following is held together by ionic or by covalent bonds:

(a) aluminum oxide, Al_2O_3

(b) sulfur dioxide, SO_2

Answers:

(a) ionic	(b) covalent

Concept Exercise

What type of chemical bond results from the attraction between a metal cation and a nonmetal anion? between two nonmetal atoms?

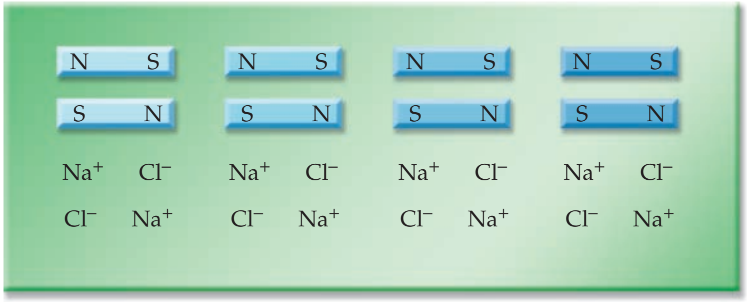
Answer: See Appendix G, 12.1.

LEARNING OBJECTIVE

▸ Illustrate the formation of an ionic bond between a metal atom and a nonmetal atom.

12.2 Ionic Bonds

An ionic bond results from the attraction between a positively charged cation and a negatively charged anion. In a crystal of table salt, NaCl, there is an attraction between Na^+ ions and Cl^- ions. This attraction between ions is similar to the attraction between the ends of two magnets. Figure 12.3 illustrates the attraction between ions, and the similar attraction between two magnets.



◀ Figure 12.3 Electrostatic Attraction Opposite ends of ordinary magnets are attracted to each other in much the same way that oppositely charged cations and anions are attracted. Notice that attraction can take place vertically as well as horizontally.

In ordinary table salt, NaCl, the ionic bonds between sodium ion and chloride ion are so strong that they can create a rigid crystalline structure (Figure 12.4). In fact, the high melting point of NaCl, 801 °C, is explained by the rigid crystal structure.

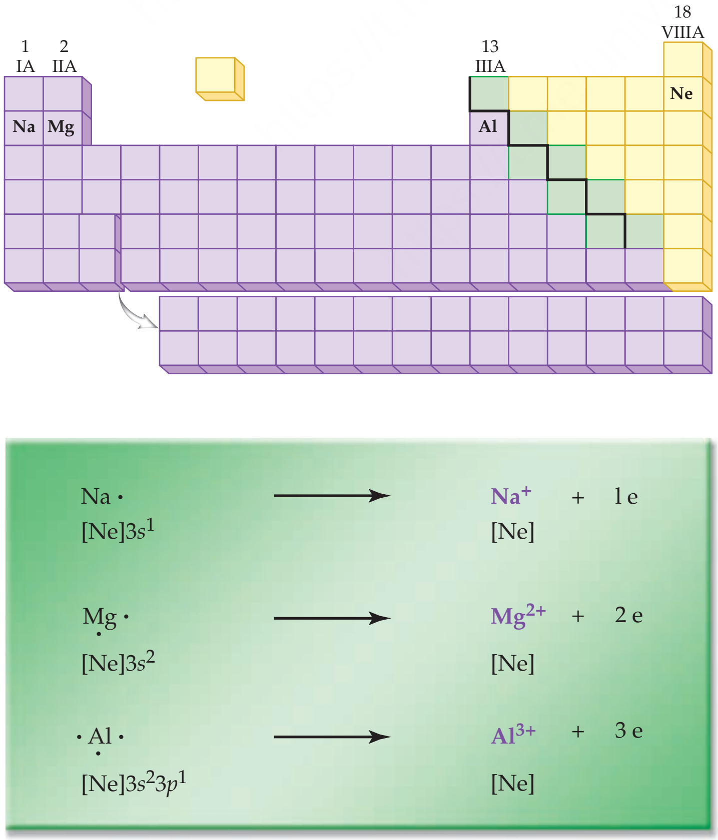
▲ Figure 12.4 Crystalline Structure of Sodium Chloride A crystal of sodium chloride is composed of sodium and chloride ions that repeat regularly. Notice that each sodium ion is attracted to several nearby chloride ions. The alternating pattern of cations and anions gives rise to a three-dimensional crystalline structure.

Formation of Cations

When a metal atom loses valence electrons to a nonmetal atom, the metal atom becomes positively charged. In Section 5.7, we learned that the number of valence electrons corresponds to the group number of the element in the periodic table. When a sodium atom (Group IA/1) loses its one valence electron, it becomes Na^+. Referring to the periodic table, we notice that Na^+ has 10 electrons ($11 - 1 = 10$ e−). This is the same number of electrons found in the preceding noble gas neon (10 e−).

Although there are exceptions, main group metals usually achieve a noble gas electron configuration after losing their valence electrons. Stated differently, main group metal ions are generally *isoelectronic* with a noble gas as explained in Section 5.10. When a magnesium atom (Group IIA/2) loses its 2 valence electrons, it becomes Mg^{2+}. A magnesium ion has 10 electrons ($12 - 2 = 10$ e−) and is also isoelectronic with neon.

In Section 5.8 we learned how to write electron dot formulas for atoms of elements. We can now use electron dot diagrams to clarify the formation of ions. Electron dot diagrams allow us to focus our attention on valence electrons involved in this process of forming ions. Figure 12.5 illustrates the formation of ions by third-period metals.



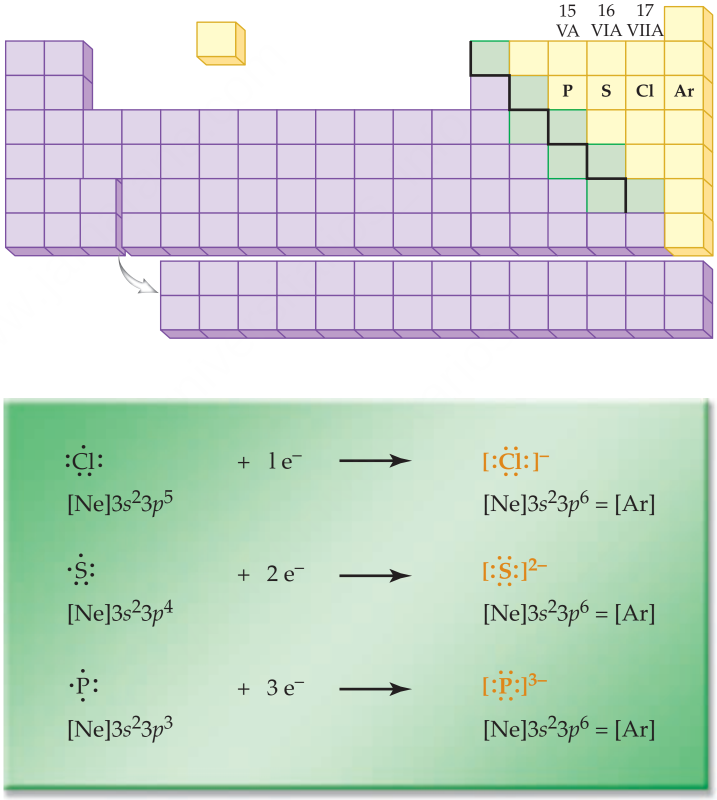
◀ Figure 12.5 Formation of Metal Ions The atoms of the metals in Period 3 form cations by losing 1, 2, and 3 electrons, respectively. In each case the metal atom becomes isoelectronic with the preceding noble gas neon, [Ne].

Note Notice that the electron dot diagrams of atoms in Figure 12.5 show electron dots, whereas the ions in Figure 12.5 do not show electron dots. Each of the ions has the electron configuration of neon. And each of these ions, like neon, has a stable electron configuration consisting of eight electrons. By convention, chemists do not draw the underlying octet of electrons because they are not valence electrons.

Formation of Anions

When a nonmetal atom gains valence electrons from a metal atom, the nonmetal atom becomes negatively charged. When a chlorine atom (Group VIIA/17) gains one valence electron, it becomes Cl^-. Referring to the periodic table, we notice that Cl^- has 18 electrons ($17 + 1 = 18$ e−). This is the same number of electrons found in the next noble gas argon (18 e−). When a sulfur atom (Group VIA/16) gains two valence electrons, it becomes S^{2-}. A sulfide ion has 18 electrons ($16 + 2 = 18$ e−) and is also isoelectronic with argon. Figure 12.6 illustrates the formation of ions by third-period nonmetals.

▶ **Figure 12.6 Formation of Nonmetal Ions** The nonmetals in Period 3 gain 1, 2, and 3 electrons, respectively. In each case the nonmetal ion becomes isoelectronic with the next noble gas argon, [Ar].



The following example exercise further illustrates the formation of cations and anions.

EXAMPLE EXERCISE 12.2 Electron Configuration of Ions

Which noble gas has an electron configuration identical to that of each of the following ions?

(a) lithium ion (b) oxide ion

(c) calcium ion (d) bromide ion

Solution

Refer to the group number in the periodic table for the number of valence electrons.

(a) Lithium (Group IA/1) forms an ion by losing one valence electron; Li^+ has only 2 electrons remaining and is isoelectronic with *He*.

(b) Oxygen (Group VIA/16) forms an ion by gaining 2 electrons; thus, O^{2-} has 10 electrons and is isoelectronic with *Ne*.

(c) Calcium (Group IIA/2) forms an ion by losing 2 electrons; Ca^{2+} has 18 electrons remaining and is isoelectronic with *Ar.*
(d) Bromine (Group VIIA/17) forms an ion by gaining one electron; Br^- has 36 electrons and is isoelectronic with *Kr.*

Practice Exercise

Which noble gas element has an electron configuration identical to that of each of the following ions?

(a) potassium ion
(b) nitride ion
(c) strontium ion
(d) iodide ion

Answers:

(a) K^+ is isoelectronic with Ar.
(b) N^{3-} is isoelectronic with Ne.
(c) Sr^{2+} is isoelectronic with Kr.
(d) I^- is isoelectronic with Xe.

Concept Exercise

Which noble gas is isoelectronic with a barium ion? with an iodide ion?

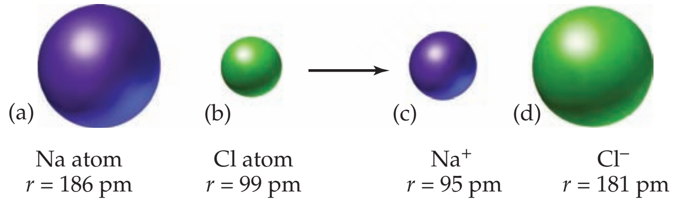
Answers: See Appendix G, 12.2.

Ionic Radii versus Atomic Radii

To visualize an ionic bond more clearly, we can picture a sodium atom becoming smaller after losing its valence electron. The radius of a sodium atom is 186 pm, whereas the radius of the sodium ion is 95 pm. The reason for this decrease in radius is that the sodium atom has lost its 3s energy sublevel. Moreover, in the sodium ion the positive nuclear charge (11+) is greater than the negative electron charge (10−) and pulls the electrons closer to the nucleus; thus, the radius decreases.

Conversely, we can picture a chlorine atom becoming larger after gaining a valence electron. The radius of a chlorine atom is 99 pm, whereas the radius of the chloride ion is 181 pm. The reason for this increase in radius is that the additional electron repels the electrons already present. In fact, the radius of the charged chloride ion is almost twice as large as that of the neutral chlorine atom.

The behavior of sodium and chlorine atoms show the same general trend observed for all metals and nonmetals with respect to atomic radius. The radius of a *cation is smaller* than that of the corresponding metal atom; the radius of an *anion is larger* than that of the corresponding nonmetal atom. The specific changes in the radius of a sodium atom and a chlorine atom are depicted in Figure 12.7.



◀ **Figure 12.7 Formation of Sodium and Chloride Ions** The radius of the sodium atom decreases (a and c) as a result of losing an electron, whereas the radius of the chlorine atom increases (b and d) after gaining an electron.

The following example exercise further illustrates the characteristics associated with the formation of an ionic bond.

EXAMPLE EXERCISE 12.3 Characteristics of Ionic Bonds

Which of the following statements are true regarding an ionic bond between an iron ion and a sulfide ion in an FeS formula unit?

(a) The iron atom loses electrons and the sulfur atom gains electrons.
(b) The iron atom has a larger radius than the iron ion.
(c) Iron and sulfide ions form a bond by electrostatic attraction.

Solution

All of these statements are true regarding an ionic bond.

(a) An iron atom loses electrons and a sulfur atom gains electrons.
(b) The radius of an iron atom is greater than its ionic radius.
(c) An ionic bond is from the attraction of positive and negative ions.

Practice Exercise

Which of the following statements are true regarding an ionic bond between a zinc ion and an oxide ion in a ZnO formula unit?

(a) Zinc and oxygen form a bond by sharing electrons.
(b) The oxygen atom is larger in radius than the oxide ion.
(c) Zinc and oxide ions form a bond by the repulsion of ions.

Answers:

(a) False, zinc and oxygen bond by transferring electrons.
(b) False, the radius of an oxygen atom is smaller than its ionic radius.
(c) False, zinc and oxygen form a bond by the attraction of ions.

Concept Exercise

Which of the following statements are true regarding the formation of an ionic bond between a metal and a nonmetal?

(a) The atomic radius of a metal atom is greater than its ionic radius.
(b) The atomic radius of a nonmetal atom is less than its ionic radius.
(c) Energy is released in the formation of an ionic bond.

Answer: See Appendix G, 12.3.

12.3 Covalent Bonds

LEARNING OBJECTIVE

- Illustrate the formation of a covalent bond between two nonmetal atoms by sharing valence electrons.

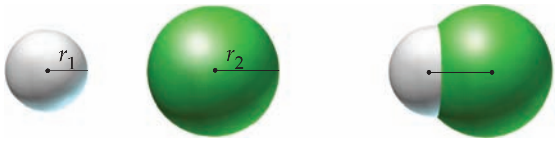
A covalent bond results from the sharing of electrons by two nonmetal atoms. To create an accurate picture, we can visualize shared electrons roaming about both atoms. The electrons belong to both nonmetal atoms, and each atom uses these bonding electrons to complete an octet. That is, each nonmetal atom uses shared electrons to complete its valence shell.

Now, let's see what happens during the formation of a covalent bond. As an example, consider the formation of hydrogen chloride, HCl, from a hydrogen atom and a chlorine atom. During bond formation, the hydrogen atom shares its one valence electron with the chlorine atom and its seven valence electrons. The additional electron gives the chlorine atom eight electrons in its valence shell, and thus fulfills the octet rule.

In the same process, the chlorine atom shares one of its valence electrons with the hydrogen atom. The additional electron gives the hydrogen atom two electrons in its valence shell. The hydrogen atom is thus similar to an atom of the noble gas helium, and is therefore stable. In this process, both the hydrogen atom and the chlorine atom become stable by sharing an electron pair and forming a covalent bond.

Bond Length

To understand bond formation more clearly, try to visualize the valence shells of the two atoms overlapping each other. In a molecule of HCl, the valence shells of the two atoms overlap as the 1s energy sublevel of the hydrogen atom mixes with the 3*p* sublevel of the chlorine atom. This mixing of sublevels draws the two nuclei closer together. The radius of the hydrogen atom is 37 pm, and that of the chlorine atom is 99 pm. If the two atoms were simply next to each other, the distance from the hydrogen nucleus to the chlorine nucleus would be $(37 + 99)$ pm or 136 pm. But experiments reveal that the distance between the two nuclei is actually 127 pm. Thus, the shells overlap, as shown in Figure 12.8. The distance between the two nuclei is called the **bond length**.



H atom
atomic radius
$r_1 = 37$ pm

Cl atom
atomic radius
$r_2 = 99$ pm

HCl molecule
bond length = 127 pm
$r_1 + r_2 = 136$ pm

◀ **Figure 12.8 Formation of a Covalent Bond** The sharing of an electron pair by a hydrogen atom and a chlorine atom produces a covalent bond. The two atoms are held together in a molecule of HCl. The bond length (127 pm) is less than the sum of the two atomic radii 37 pm + 99 pm = 136 pm.

Bond Energy

Energy is released when two ions are attracted to one another and form an ionic bond. For example, the attraction of Na^+ and Cl^- to form NaCl releases heat energy. Similarly, energy is released when two atoms are attracted and form a covalent bond. When hydrogen and chlorine atoms form an HCl molecule, heat energy is released.

$$H(g) + Cl(g) \longrightarrow HCl(g) + \textit{heat}$$

Conversely, energy is necessary to break an H—Cl bond. The amount of energy required to break a covalent bond between two atoms is referred to as the **bond energy**. The amount of energy necessary to break the H—Cl bond is identical to the amount of heat released when H and Cl form a bond in the creation of a hydrogen chloride molecule.

$$HCl(g) + \textit{heat} \longrightarrow H(g) + Cl(g)$$

The following example exercise summarizes the characteristics associated with the formation of a covalent bond.

EXAMPLE EXERCISE 12.4 Characteristics of Covalent Bonds

Which of the following statements are true regarding a covalent bond between a hydrogen atom and an oxygen atom in an H_2O molecule?

(a) Valence electrons are shared by hydrogen and oxygen atoms.
(b) The H—O bond length is less than the sum of the two atomic radii.
(c) Breaking the bond between H—O requires energy.

Solution

All of the previous statements regarding a covalent bond in a molecule are true.

(a) Valence electrons are shared by the two nonmetal atoms.
(b) The bond length is less than the sum of the two atomic radii.
(c) The breaking of a covalent bond requires energy.

Practice Exercise

Which of the following statements are true regarding a covalent bond between a hydrogen atom and a sulfur atom in an H_2S molecule?

(a) Valence electrons are transferred from hydrogen to sulfur.
(b) The H—S bond length equals the sum of the two atomic radii.
(c) Forming a bond between H—S requires energy.

Answers:

(a) False, valence electrons are shared by hydrogen and sulfur.
(b) False, the bond length is less than the sum of the two atomic radii.
(c) False, the forming of a covalent bond releases energy.

Concept Exercise

Which of the following statements are true regarding the formation of a covalent bond between two nonmetal atoms?

(a) Bonding electrons are shared between two nonmetal atoms.
(b) The bond length is less than the sum of the two atomic radii.
(c) The bond energy is the energy required to break a covalent bond.

Answer: See Appendix G, 12.4.

Note Energy is always released when two *atoms* form a covalent bond. However, energy is required to form an H—I bond from *molecules* of H_2 and I_2. Realize that energy is required to break H—H and I—I bonds in order to form H—I bonds. The formation of HI is an *endothermic* reaction because more energy is required in breaking H—H and I—I bonds than is released in forming H—I bonds.

A CLOSER LOOK Epsom Salts

Q: What are some common uses and properties of Epsom salts?

In 1618, a farmer in Epsom, England noticed that his cows would not drink water from a particular well on his property. The farmer investigated and observed that the water contained a substance with a bitter taste. The substance became known as Epsom Salts and it was eventually identified as magnesium sulfate heptahydrate, $MgSO_4 \cdot 7\ H_2O$.

Bath salts containing $MgSO_4 \cdot 7\ H_2O$ are commonplace for soothing tired and aching feet. Epsom Salts solutions have been shown to speed the healing of skin rashes and scratches. It is believed that magnesium sulfate reduces inflammation when absorbed into the skin. In addition, Epsom Salts is used in flotation therapy, where high concentrations of salt raise the density of bath water, thus, making the human body more buoyant. The salt solution soothes the skin and relaxes tired and cramping muscles.

Crystalline Epsom Salts

Epsom Salts are also commonly used by gardeners to correct a magnesium deficiency in the soil. In agriculture, magnesium sulfate is used to correct a magnesium deficiency, which is an essential element in a chlorophyll molecule. The advantage of magnesium sulfate over other magnesium compounds is its high solubility and neutral pH in water.

Anhydrous magnesium sulfate is used by chemists in the laboratory to absorb water vapor. During an experiment in which water can interfere with a reaction taking place in an organic solvent, solid magnesium sulfate is added. The magnesium sulfate absorbs moisture and the resulting solid clump can be easily removed from the organic solution by filtration.

A: Soaking the body in Epsom Salts soothes the skin and relaxes tired muscles. Epsom Salts is also used in agriculture and in the chemistry laboratory.

LEARNING OBJECTIVES

- Diagram the electron dot formula for a molecule.
- Diagram the structural formula for a molecule.

12.4 Electron Dot Formulas of Molecules

In Section 5.8, we drew the electron dot formula for an atom of a representative element. We found the element in the periodic table and noted its group number to predict the number of valence electrons. For example, in the second period, we found that Li has one valence electron (Group IA/1), Be has 2 (Group IIA/2), B has 3 (Group IIIA/13), C has 4 (Group IVA/14), N has 5 (Group VA/15), O has 6 (Group VIA/16), F has 7 (Group VIIA/17), and the noble gas Ne has 8 (Group VIIIA/18). We represented each valence electron by a dot surrounding the symbol for the element.

Now, we will draw the **electron dot formula** (also called a *Lewis structure*) for a molecule. An electron dot formula shows the valence electrons that form bonds between atoms. These electron pairs shared by atoms are called **bonding electrons**. Other electron pairs that simply complete the octet around each atom are called **nonbonding electrons**, or *lone pairs.*

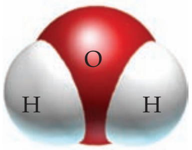
Guidelines for Drawing Electron Dot Formulas of Molecules[†]

1. Calculate the total number of valence electrons by adding all the valence electrons for each atom in the molecule. The total must be an even number to obey the octet rule. If the total is an odd number, check your calculations.
2. Divide the total number of valence electrons by 2 to find the number of electron pairs in the molecule.
3. Surround the central atom with four electron pairs. Use the remaining electron pairs to complete an octet around each of the other atoms. (Hydrogen is the sole exception as it requires only one electron pair.)
4. If there are not enough electron pairs to provide an octet for each atom, move a nonbonding electron pair between two atoms that already share an electron pair.

[†] When two or more atoms are bonded to a central atom, it is often difficult to recognize the central atom. Thus, this textbook will indicate the central atom in bold to aid in writing electron dot formulas for molecules.

Electron Dot Formula for Water, H_2O

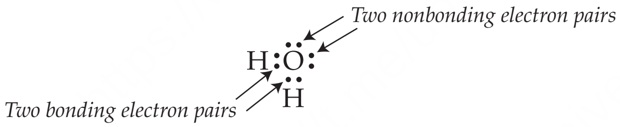
To learn how to draw electron dot formulas for molecules, we will draw the electron dot formula for water, H_2**O**. The total number of valence electrons in the molecule is one for each hydrogen atom and six for the oxygen atom. Thus, the total number of electrons is $2(1\ \text{e}-) + 6\ \text{e}- = 8\ \text{e}-$. The number of electron pairs is 4 ($8/2 = 4$).



▲ Water, H_2O

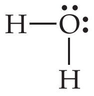
In the formula H_2**O**, oxygen is the central atom and is indicated in bold. We place the four electron pairs around the oxygen to provide the necessary octet.

We then place the two hydrogen atoms in any of the four electron pair positions.



Notice that there are two bonding and two nonbonding electron pairs. However, bonding and nonbonding electrons are identical, and the four electron pairs move about the molecule. Thus, bonding and nonbonding valence electrons spread over the entire molecule and are said to be *delocalized.*

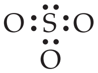
To simplify, we represent each pair of bonding electrons by a single dash, which is called a **single bond**. The resulting structure is referred to as the **structural formula** of the molecule. In a structural formula the nonbonding electrons are usually shown, but their inclusion is optional.

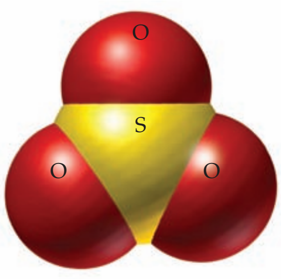


Electron Dot Formula for Sulfur Trioxide, SO_3

Now we'll try a more difficult example. Let's draw the electron dot formula for sulfur trioxide, **S**O_3. Sulfur is the central atom in the molecule and is indicated in bold. Sulfur and oxygen each have six valence electrons, so the total number of valence electrons in the molecule is $6\ \text{e}- + 3(6\ \text{e}-) = 24\ \text{e}-$. The number of electron pairs is 12 ($24/2 = 12$).

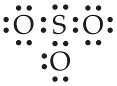
Because sulfur is the central atom, we can begin by placing four electron pairs around the sulfur and then attaching the three oxygen atoms. That gives us



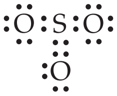


▲ Sulfur Trioxide, SO_3

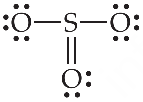
We started with 12 electron pairs, and so we have eight pairs remaining. We place the remaining pairs around the oxygen atoms to complete each octet.



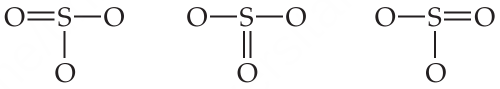
Notice that the bottom oxygen atom does not have an octet; it has only three electron pairs. According to the guidelines, we can move a nonbonding electron pair from the S atom so as to provide two bonding electron pairs between the sulfur and the oxygen atoms.



The two shared electron pairs constitute a **double bond**. All four electrons in the double bond are shared by the S and the O atoms. Therefore, the sulfur atom still has an octet and the bottom oxygen atom has gained two electrons to complete its octet. The structural formula for SO_3 has one double bond and two single bonds. We represent the double bond by a double dash and the single bond by a single dash.



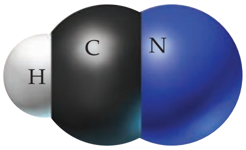
We can draw the structural formula for an SO_3 molecule with the double bond between any of the oxygen atoms. Experimentally, the double bond is found to be shared by each of the three oxygen atoms. This phenomenon is known as *resonance,* and we can draw three resonance structures for an SO_3 molecule.



$$\begin{array}{c}O{=}S{-}O\\ |\\ O\end{array}\qquad\begin{array}{c}O{-}S{-}O\\ \|\\ O\end{array}\qquad\begin{array}{c}O{-}S{=}O\\ |\\ O\end{array}$$

Although these structures may appear to be different, all three formulas are identical. If we were to construct the molecule using models, we could easily verify this statement. Each molecular model can be rotated and shown to be identical to the other two.

Electron Dot Formula for Hydrogen Cyanide, HCN



▲ **Hydrogen Cyanide, HCN**

Let's draw the electron dot formula for hydrogen cyanide, H**C**N. Carbon is the central atom in the molecule and is indicated in bold. The total number of valence electrons in the molecule is 1 e− + 4 e− + 5 e− = 10 e−. The number of electron pairs is 5 (10/2 = 5). Because carbon is the central atom, we begin by placing four electron pairs around the carbon and attaching the hydrogen and nitrogen atoms.

We started with five electron pairs, and there is only one pair remaining. The hydrogen atom has its two required electrons, so we place the remaining pair next to the nitrogen atom.

The nitrogen atom now has four electrons, which is 4 e− less than an octet. According to the guidelines, we can move nonbonding electron pairs between the C and the N atoms. For example,

The nitrogen atom has only six electrons. If we move the other nonbonding electron pair from the carbon atom, we have

$$\text{H}:\text{C}:::\text{N}:$$

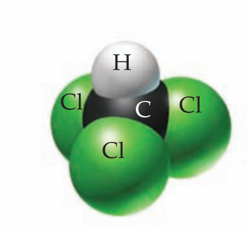
The three electron pairs produce a **triple bond**. All six electrons in the triple bond are shared by the C and the N atoms. Therefore, both the carbon atom and the nitrogen atom have obtained an octet. A triple bond is indicated by a triple dash. The structural formula for HCN is therefore written as

$$\text{H}{-}\text{C}{\equiv}\text{N}:$$

The following example exercises further illustrate how to draw electron dot and structural formulas of molecules.

EXAMPLE EXERCISE 12.5 Electron Dot Formulas—Single Bonds

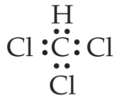
Draw the electron dot formula and the structural formula for a chloroform molecule, **C**HCl_3.



▲ **Chloroform, $CHCl_3$**

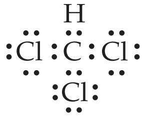
Solution

Carbon is the central atom in the chloroform molecule and is indicated in bold. The total number of valence electrons is $4\ e- + 1\ e- + 3(7\ e-) = 26\ e-$. The number of electron pairs is $13\ (26/2 = 13)$. We begin by placing four electron pairs around the carbon and adding one H and three Cl atoms.



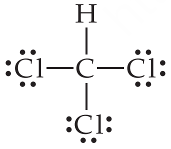
$$\begin{array}{c} \text{H} \\ \text{Cl}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{C}}}:\text{Cl} \\ \text{Cl} \end{array}$$

We have 13 electron pairs minus the four pairs we used for the carbon. Because 9 electron pairs remain, let's place three pairs around each chlorine.



$$\begin{array}{c} \text{H} \\ :\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Cl}}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{C}}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Cl}}}: \\ :\underset{\cdot\cdot}{\text{Cl}}: \end{array}$$

Each atom is surrounded by an octet of electrons, except hydrogen, which has two. This is the correct electron dot formula. In the corresponding structural formula we replace each bonding electron pair by a single dash. The structural formula is



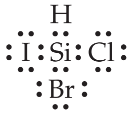
$$\begin{array}{c} \text{H} \\ | \\ :\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Cl}}}-\text{C}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Cl}}}: \\ | \\ :\underset{\cdot\cdot}{\text{Cl}}: \end{array}$$

Note that the structural formula is free to rotate. We could have also written the hydrogen atom below or at the side of the carbon atom.

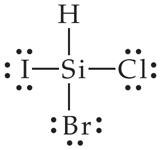
Practice Exercise

Draw the electron dot formula and the structural formula for a molecule of **Si**HClBrI.

Answers:



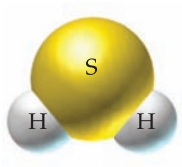
$$\begin{array}{c} \text{H} \\ :\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{I}}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Si}}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Cl}}}: \\ :\underset{\cdot\cdot}{\text{Br}}: \end{array}$$



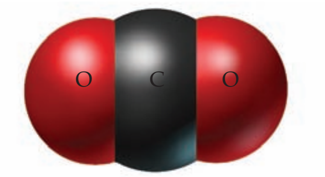
$$\begin{array}{c} \text{H} \\ | \\ :\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{I}}}-\text{Si}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Cl}}}: \\ | \\ :\underset{\cdot\cdot}{\text{Br}}: \end{array}$$

Concept Exercise

Draw the electron dot formula for H_2S. How many nonbonding electron pairs are in a hydrogen sulfide molecule?



Answers: See Appendix G, 12.5.



▲ **Carbon Dioxide, CO_2**

Helpful Hint
Octet Rule Violations

Writing electron dot formulas is a trial-and-error procedure that involves the use of valence electrons to satisfy the octet rule. However, there are exceptions to the octet rule that are beyond this introduction to chemical bonding. For example, boron trifluoride, $\mathbf{B}F_3$, has only six electrons (three electron pairs) around the central **B** atom. Moreover, there are exceptions that violate the octet rule by exceeding eight electrons around the central atom. For example, $\mathbf{P}F_5$ has 10 electrons (five electron pairs) around the central **P** atom; and in $\mathbf{S}F_6$ there are 12 electrons (six electron pairs) around the central **S** atom.

In addition, there are molecules that have an odd number of electrons; for example, nitrogen monoxide, NO, and nitrogen dioxide, NO_2. The number of valence electrons is 11 in an NO molecule, and 17 in an NO_2 molecule. Although NO_2 is a free radical, it can stabilize by joining another NO_2 molecule to give N_2O_4. In general, free radical compounds are unstable, and highly reactive.

EXAMPLE EXERCISE 12.6 Electron Dot Formulas—Double Bonds

Draw the electron dot formula and the structural formula for a carbon dioxide molecule, $\mathbf{C}O_2$.

Solution

Carbon is the central atom in the carbon dioxide molecule, as indicated in bold. The total number of valence electrons is $4\ e- + 2(6\ e-) = 16\ e-$. The number of electron pairs is 8 $(16/2 = 8)$. We begin by placing four electron pairs around the carbon and adding the two O atoms.

$$\text{O}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{C}}}:\text{O}$$

The number of remaining electron pairs is $8 - 4 = 4$ pairs. We can add 2 pairs to each oxygen.

$$\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{C}}}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}$$

Each oxygen atom shares six electrons, two less than an octet. We can use the nonbonding electron pairs around carbon to complete the octet around each oxygen. We move one pair to the oxygen on the left, and the other pair to the oxygen on the right to complete the octet around each oxygen. Each carbon–oxygen bond shares two electron pairs.

$$\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}::\text{C}::\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{O}}}$$

Now the octet rule is satisfied for each atom. This is the correct electron dot formula. There are two electron pairs between each oxygen and carbon. Thus, there are two double bonds in the carbon dioxide molecule. In the structural formula each double bond is indicated by a double dash.

$$:\ddot{\text{O}}{=}\text{C}{=}\ddot{\text{O}}:$$

Practice Exercise

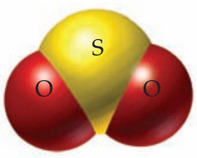
Draw the electron dot formula and the structural formula for a molecule of $\mathbf{Si}O_2$.

Answers:

$$:\ddot{\text{O}}::\text{Si}::\ddot{\text{O}}: \qquad :\ddot{\text{O}}{=}\text{Si}{=}\ddot{\text{O}}:$$

Concept Exercise

Draw the electron dot formula for $\mathbf{S}O_2$. How many nonbonding electron pairs are in a sulfur dioxide molecule?



Answer: See Appendix G, 12.6.

EXAMPLE EXERCISE 12.7 Electron Dot Formulas—Triple Bonds

Draw the electron dot formula and the structural formula for an acetylene molecule, $\mathbf{C}_2H_2$.

Solution

Carbon is the central atom in the acetylene molecule as indicated in bold. The total number of valence electrons is $2(4\ e-) + 2(1\ e-) = 10\ e-$. The number of electron pairs is $5\ (10/2 = 5)$. We begin by placing four electron pairs around a carbon atom and one electron pair between C and H.

$$\mathrm{H}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{C}}}:\mathrm{C}:\mathrm{H}$$

Because each carbon atom requires eight electrons, we must move nonbonding electron pairs to acquire an octet around each carbon. We can move two nonbonding electron pairs as follows:

$$\mathrm{H}:\mathrm{C}:::\mathrm{C}:\mathrm{H}$$

Now the octet rule is satisfied for each atom. This is the correct electron dot formula. There are three electron pairs between each carbon atom. Thus, there is one triple bond in the acetylene molecule. In the structural formula, the triple bond is indicated by a triple dash. In this example, there are no nonbonding electron pairs.

$$\mathrm{H{-}C{\equiv}C{-}H}$$

Practice Exercise

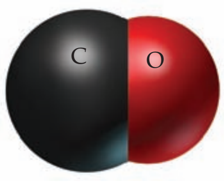
Draw the electron dot formula and the structural formula for a molecule of HO**C**N.

Answers:

$$\begin{array}{l} \mathrm{H}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{O}}}: \\ \quad\ \mathrm{C}:::\mathrm{N}: \end{array} \qquad\qquad \begin{array}{l} \mathrm{H{-}O} \\ \quad\ \ | \\ \quad\ \mathrm{C{\equiv}N} \end{array}$$

Concept Exercise

Draw the electron dot formula for **C**O. How many nonbonding electron pairs are in a carbon monoxide molecule?



Answers: See Appendix G, 12.7.

12.5 Electron Dot Formulas of Polyatomic Ions

LEARNING OBJECTIVES

- Diagram the electron dot formula for a polyatomic ion.
- Diagram the structural formula for a polyatomic ion.

So far in our discussion of ions, we have considered only single atoms that have gained or lost electrons. This type of simple ion having a positive or negative charge is called a **monoatomic ion**. An ion can also contain two or more atoms that bears an overall charge; this type of ion is called a **polyatomic ion**. Some common polyatomic ions are the ammonium ion, NH_4^+, found in fertilizer; the hydroxide ion, OH^-, found in household products like drain and oven cleaner; and the hydrogen carbonate ion, HCO_3^-, found in baking soda.

In Section 12.4, we learned to draw the electron dot formula for molecules. Now, we can draw the electron dot formula for polyatomic ions. The following directions for ions are similar to the previous guidelines for molecules.

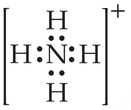
Guidelines for Drawing Electron Dot Formulas of Polyatomic Ions[†]

1. Calculate the valence electrons by adding the valence electrons for each atom in the polyatomic ion. If the ion is negatively charged, add electrons—equal to the charge—to the number of valence electrons. If the ion is positively charged, subtract electrons—equal to the charge—from the number of valence electrons. In each case, the result will be the total number of valence electrons.
2. Divide the total number of valence electrons and charges by two to find the number of electron pairs in the polyatomic ion.
3. Surround the central atom with four electron pairs. Use the remaining electron pairs to complete an octet around each of the other atoms. (Hydrogen is the sole exception, as H requires only one electron pair.)
4. If there are not enough electron pairs to provide an octet for each atom, move a nonbonding electron pair between two atoms that already share an electron pair.

[†] When two or more atoms are bonded to a central atom, it is often difficult to recognize the central atom. Thus, this textbook will indicate the central atom in bold to aid in writing electron dot formulas for polyatomic ions.

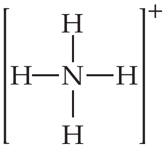
Electron Dot Formula for the Ammonium Ion, NH_4^+

We can draw the electron dot formula for the ammonium ion, $\mathbf{N}H_4^+$, as follows. The total number of valence electrons in this positive polyatomic ion is 5e− + 4(1e−) − 1e− = 8e−. (Note that we must subtract 1 e−, because the ion has a 1+ charge.) The number of electron pairs is 4(8/2 = 4). In the formula $\mathbf{N}H_4^+$, nitrogen is the central atom. We place the four electron pairs around the nitrogen atom to provide the necessary octet.



Notice that the electron dot formula is written correctly for the ammonium ion. The nitrogen atom has an octet, and each hydrogen atom shares two electrons.

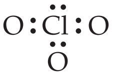
Now, let's simplify the electron dot formula and draw the structural formula of this polyatomic ion. We represent each bonding pair of electrons by a single dash and draw the structural formula for the ammonium ion as follows:



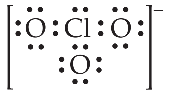
The usual way to draw structural formulas of polyatomic ions is to enclose them in brackets. The overall charge on the polyatomic ion is indicated outside the brackets. Placing the charge outside the brackets emphasizes that the positive charge is distributed over the entire ion.

Electron Dot Formula for the Chlorate Ion, ClO_3^-

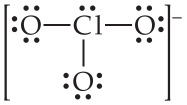
Now, let's draw the electron dot formula for the chlorate ion, $\mathbf{Cl}O_3^-$. The total number of valence electrons for this negative ion is 7 e− + 3(6 e−) + 1 e− = 26 e−. (Note that we must add 1 e−, because the ion has a 1− charge.) The number of electron pairs is 13 (26/2 = 13). Because chlorine is the central atom, we begin by placing four electron pairs around the Cl atom and then add the three O atoms.



We started with 13 electron pairs and placed four pairs around the central atom. We place the remaining 9 electron pairs around the oxygen atoms to complete each octet.

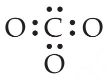


In the structural formula, each bonding electron pair is represented by a single dash. Thus, we can show the structural formula for ClO_3^- as follows:

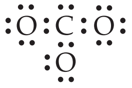


Electron Dot Formula for the Carbonate Ion, CO_3^{2-}

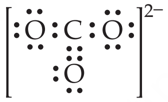
Let's try a more difficult example and draw the electron dot formula for the carbonate ion, CO_3^{2-}. The total number of valence electrons in this polyatomic ion is $4\ e- + 3(6\ e-) + 2\ e-$ (because the ion has a $2-$ charge) $= 24\ e-$. The total number of electron pairs is 12 ($24/2 = 12$). Carbon is the central atom, so we begin by placing four pairs of electrons around the C atom and then add three O atoms. We have



We started with 12 electron pairs, and so we have 8 pairs remaining. We place the remaining pairs around the oxygen atoms to give octets as follows:

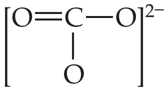


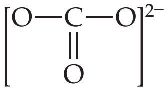
Notice that the bottom oxygen atom does not have an octet; it has only three electron pairs. According to our guidelines, we can move a nonbonding electron pair. Let's shift the nonbonding pair between the C and the O atoms.

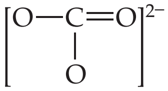


The two electron pairs constitute a double bond. All four electrons in the double bond are shared by the C and the O atoms. Therefore, the central carbon atom has an octet and the oxygen atom has gained two electrons to complete its octet.

However, the double bond can form between any one of the three oxygen atoms. We use brackets to convey the idea that valence electrons are free to move about the entire polyatomic ion. Consequently, there appear to be three different structural formulas for the carbonate ion.







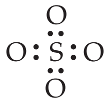
Although the structures for the carbonate ion appear to be different, they are all equivalent. No matter how we represent CO_3^{2-}, the structural formula has one double bond and two single bonds. Thus, we can draw three resonance structures for the carbonate ion. The following example exercise further illustrates how to draw electron dot and structural formulas for polyatomic ions.

EXAMPLE EXERCISE 12.8 Electron Dot Formulas of Polyatomic Ions

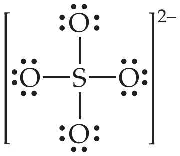
Draw the electron dot formula and the structural formula for the sulfate ion, SO_4^{2-}.

Solution

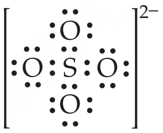
The total number of valence electrons is the sum of each atom plus 2 for the negative charge: $6\,e- + 4(6\,e-) + 2\,e- = 32\,e-$. The number of electron pairs is 16 ($32/2 = 16$). We begin by placing four electron pairs around the central sulfur atom and attaching the four oxygen atoms as follows:



We have 12 remaining electron pairs, and so we place 3 electron pairs around each oxygen atom.



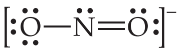
Notice that each atom is surrounded by an octet of electrons. This is the correct electron dot formula. In the structural formula each bonding electron pair is replaced by a single dash. The structural formula is



Practice Exercise

Draw the electron dot formula and the structural formula for the nitrite ion, NO_2^-.

Answers:



Concept Exercise

Draw the electron dot formula for the polyatomic sulfite ion, SO_3^{2-}. How many nonbonding electron pairs are in a sulfite ion?

Answers: See Appendix G, 12.8.

12.6 Polar Covalent Bonds

LEARNING OBJECTIVES

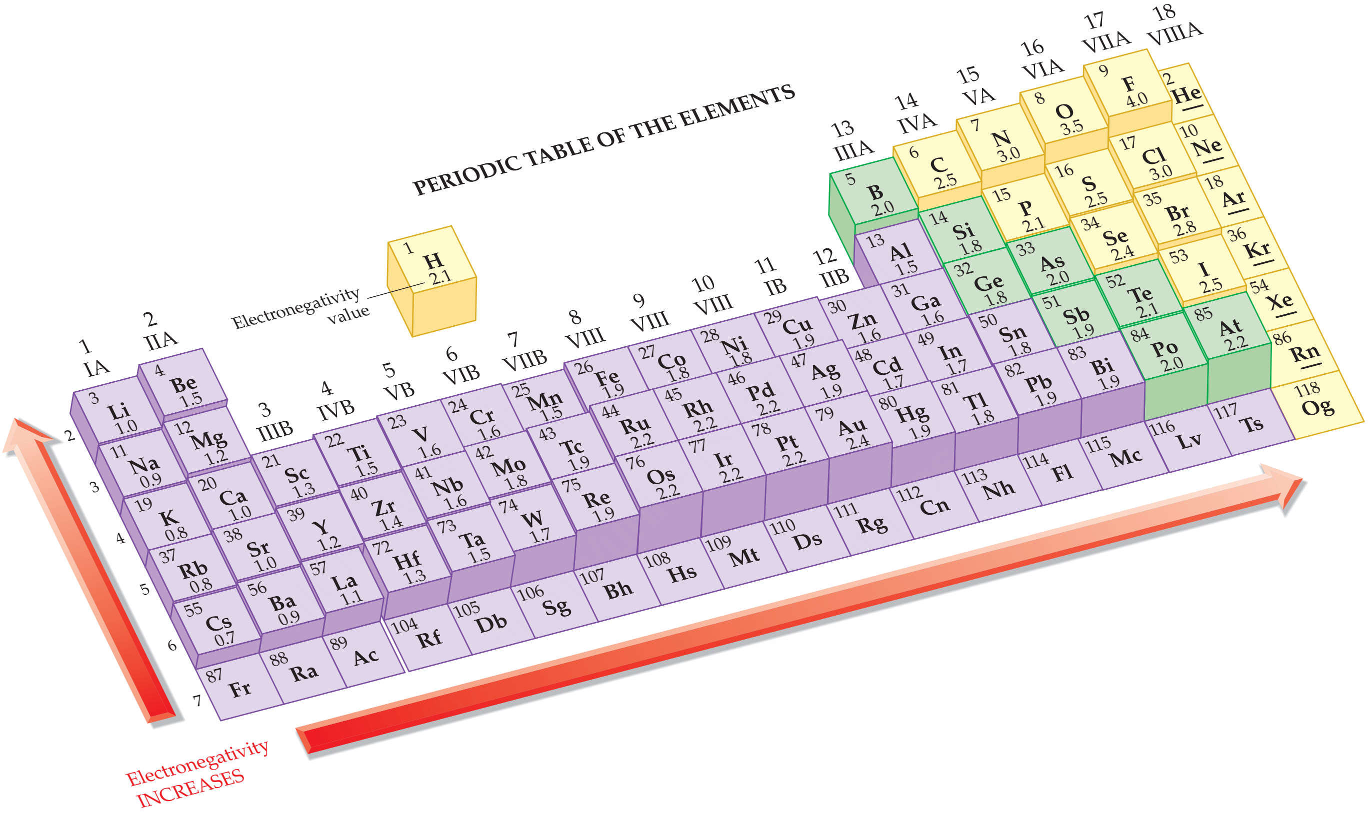
- Describe the electronegativity trends in the periodic table.
- Calculate the electronegativity difference in a polar covalent bond.
- Apply delta notation (δ^+ and δ^-) to a polar bond.

Covalent bonds result from the sharing of valence electrons. To this point we have assumed that electrons are shared equally in a covalent bond. What about a covalent bond in which the two atoms do not share electrons equally? In many instances one of the two atoms holds the electron pair more tightly. When the electrons are drawn more closely to one of the atoms, the bond is said to be polarized. This type of bond is called a *polar covalent bond,* or simply a **polar bond**.

Electronegativity Trends

Each element has an innate ability to attract valence electrons. This ability is related to the nearness of the valence shell to the nucleus. It is also related to the magnitude of the positive charge in the nucleus. The ability of an atom to attract electrons in a covalent bond is referred to as its **electronegativity**. Atoms of elements that strongly attract bonding electrons are said to be highly electronegative.

Linus Pauling, an American chemist, devised a method for measuring the electronegativity values for each of the elements. He assigned carbon a value of 2.5 and then determined the ability of other elements to attract bonding electrons relative to carbon. He found that fluorine is the most electronegative element. It has a value of 4.0 compared to carbon. Other highly electronegative elements are oxygen, 3.5; nitrogen, 3.0; and chlorine, 3.0. The most electronegative elements are the nonmetals on the far right of the periodic table. Figure 12.9 shows selected elements in the periodic table and their Pauling electronegativity value.



▲ **Figure 12.9 Electronegativity Values for the Elements** The Pauling electronegativity value for an element is shown below the symbol. In general, electronegativity increases across a period and up a group.

Note the electronegativity trends in Figure 12.9. First, the elements generally become more electronegative from *left to right* in the periodic table. Second, elements become more electronegative from *bottom to top* within a group.

We can see that the horizontal trend is consistent with the chemical properties of elements. That is, nonmetals react by gaining electrons, and metals react by losing electrons. The vertical trend follows the trend in nonmetallic character. As an example, examine Group VA/15. We see that Bi is a metal, Sb and As are semimetals, and P and N are nonmetals. Therefore, the trend in electronegativity increases corresponding to the trend in nonmetallic character. That is, nonmetals are more electronegative than semimetals, which in turn are more electronegative than metals.

The following example exercise illustrates the prediction of electronegativity values from the general trends in the periodic table.

EXAMPLE EXERCISE 12.9 Electronegativity Predictions

Predict which element in each of the following pairs is more electronegative according to the general electronegativity trends in the periodic table:

(a) N or O	(b) Br or Se
(c) F or Cl	(d) Si or C

Solution

According to the trends in the periodic table, elements that lie to the right in a series or at the top of a group are more electronegative.

(a) O is more electronegative than N.	(b) Br is more electronegative than Se.
(c) F is more electronegative than Cl.	(d) C is more electronegative than Si.

Practice Exercise

Predict which element in each of the following pairs is more electronegative according to the general trends in the periodic table:

(a) H or Cl	(b) Br or I
(c) P or S	(d) As or Sb

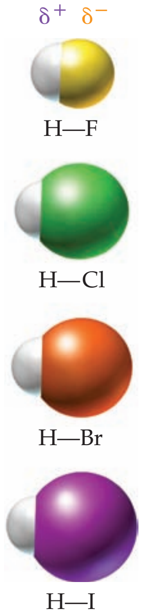
Answers:

(a) Cl	(b) Br
(c) S	(d) As

Concept Exercise

State the general trends in the periodic table for increasing electronegativity.

Answer: See Appendix G, 12.9.



▲ **Hydrogen Halides** The two atoms in each hydrogen halide molecule are joined by a polar covalent bond owing to the different electronegativity value of each halogen.

Delta Notation for Polar Bonds

In Section 12.1, we introduced the covalent bond. In Figure 12.9, we see that the electronegativity value of H is 2.1 and of Cl is 3.0. Because there is a difference in electronegativity between the two elements ($3.0 - 2.1 = 0.9$), the bond in a H—Cl molecule is polar. Moreover, Cl is the more electronegative element, so the bonding electron pair is attracted toward the Cl atom and away from the H atom. The Cl atom thus becomes slightly negatively charged, whereas the H atom becomes slightly positively charged.

We can indicate a polar bond with a special symbol, the Greek letter delta (δ). We indicate the atom having a partially negative charge by the symbol δ^-. Similarly, we indicate the atom having a partially positive charge by the symbol δ^+. These symbols, δ^- and δ^+, are referred to as **delta notation**. We can apply delta notation to illustrate the polar bond in a HCl molecule as follows:

$$\delta^{+}\ \mathbf{H{-}Cl}\ \delta^{-}$$

FEATURED SCIENTIST Linus Pauling

Q. For which of his many accomplishments was Linus Pauling awarded the Nobel Prize in chemistry?

Linus Pauling (1901–1994) is regarded by many as the most influential chemist of the twentieth century. Pauling was born in Portland, Oregon, to a family of modest means. His father died when Linus was quite young, and he had to take on a series of menial jobs to help support the family. At age 16, he enrolled at Oregon State University where he received a bachelor's degree in chemical engineering. There he met his wife and acknowledged inspiration, Ava Helen Miller, while teaching a chemistry class to home economics students.

Pauling decided on the California Institute of Technology for graduate school. In 1925, he received a Ph.D. and following graduation, was awarded a Guggenheim fellowship to study in Europe. In 1927, he returned to Caltech as a professor of chemistry and remained there for most of his career. Pauling received immediate recognition and was described by his department chairman in the following words:

> *Were all the rest of the chemistry department wiped away except Pauling, it would still be one of the most important departments of chemistry in the world.*

Among his many accomplishments, Linus Pauling invented the concept of electronegativity and provided an ingenious method for calculating the electronegativity values of elements. One of his most brilliant achievements was an explanation for the unusual stability of the benzene molecule, C_6H_6, which was discussed in Section 8.9. His most celebrated work was investigating the shapes of proteins. He found that their structures resembled a helix and further established the relationship between abnormal molecular structure and genetic disease. His work in this area earned him the 1954 Nobel Prize in chemistry.

In the 1960s, Pauling focused his energies on banning aboveground testing of nuclear weapons and pointed out the potential dangers of radiation. He spoke passionately against war and encouraged global disarmament. For his efforts, Pauling was awarded the 1962 Nobel Peace Prize. In the 1970s, he stirred controversy when he publicly advocated massive doses of vitamin C to combat the common cold. In addition to two Nobel Prizes, Linus Pauling received numerous medals and awards, accepted honorary memberships in several scientific societies, published more than 400 research papers, and wrote the classic textbook *The Nature of the Chemical Bond.*

▲ **Linus Pauling (1901–1994)**

A: Linus Pauling won the 1954 Nobel Prize in chemistry for his pioneering work on the structure of proteins.

Because the chlorine atom is more electronegative, it draws the electron pair away from the hydrogen atom. As a result, the hydrogen atom becomes slightly positively charged. The following example exercise illustrates the application of delta notation.

EXAMPLE EXERCISE 12.10 Delta Notation for Polar Bonds

Calculate the electronegativity difference and apply delta notation to the bond between carbon and oxygen, C—O.

Solution

From Figure 12.9, we find that the electronegativity value of C is 2.5 and of O is 3.5. By convention, we always subtract the lesser value from the greater. The difference between the two elements is $3.5 - 2.5 = 1.0$. The result shows that the C—O bond is slightly polarized.

Because O is the more electronegative element, the bonding electron pair is drawn away from the C atom and toward the O atom. The O atom thus becomes slightly negatively charged, whereas the C atom becomes slightly positively charged. Applying the delta convention, we have

$$\delta^+\,\mathrm{C{-}O}\,\delta^-$$

Practice Exercise

From the trends in the periodic table, apply delta notation and label each atom in the following polar covalent bonds:

(a) N—O (b) Br—F

Answers:

(a) δ^+ N—O δ^- (b) δ^+ Br—F δ^-

Concept Exercise

Refer to the electronegativity values in Figure 12.9 and predict which of the following bonds is most polar: H—N, H—O, or H—F.

Answer: See Appendix G, 12.10.

Note Delta convention indicates partially negative and partially positive atoms in a molecule. However, polar bonds in molecules differ from ionic bonds in formula units. Ionic bonds are formed between negative anions and positive cations. An electronegativity difference between two atoms that is greater than 1.8 indicates an ionic bond. For example, NaCl has an electronegativity difference of 2.1 $(3.0 - 0.9)$ and it is clearly an example of an ionic bond.

LEARNING OBJECTIVES

- Describe and identify a nonpolar covalent bond.
- Identify seven elements that occur naturally as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.

12.7 Nonpolar Covalent Bonds

In Section 12.6, we learned that polar bonds result from the unequal sharing of bonding electrons. In a polar bond, one of the two atoms has a greater tendency to attract electrons. The atom having the greater electronegativity attracts the bonded electron pair more strongly and has a partial negative charge.

But what about a covalent bond in which each atom has the same or similar electronegativity? To answer this question, let's begin by reexamining the Pauling electronegativity values in Figure 12.9. We see that electronegativity increases across a period and up a group. Notice, however, that several elements have the same Pauling electronegativity value. For example, N and Cl have a value of 3.0, and C, S, and I all have a value of 2.5.

How then would we describe an S—I bond? The electronegativity of each atom is the same, so the bond is not polarized. A covalent bond between two atoms with the same electronegativity is referred to as a nonpolar covalent bond, or simply a **nonpolar bond**. The following example exercise illustrates the classification of polar and nonpolar bonds.

EXAMPLE EXERCISE 12.11 Polar versus Nonpolar Bonds

Classify each of the following as a polar or a nonpolar bond:

(a) Cl—Cl	(b) Cl—N
(c) Cl—Br	(d) Cl—H

Solution

From Figure 12.9 we find the following electronegativity values: Cl = 3.0, N = 3.0, Br = 2.8, and H = 2.1.

(a) The Cl—Cl bond $(3.0 - 3.0 = 0)$ is *nonpolar.*
(b) The Cl—N bond $(3.0 - 3.0 = 0)$ is *nonpolar.*
(c) The Cl—Br bond $(3.0 - 2.8 = 0.2)$ is slightly *polar.*
(d) The Cl—H bond $(3.0 - 2.1 = 0.9)$ is *polar.*

Practice Exercise

Classify each of the following as a polar or a nonpolar bond:

(a) C—C	(b) C—O
(c) C—S	(d) C—H

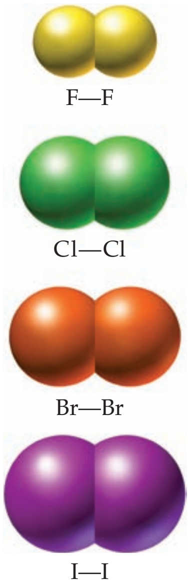
Answers:

(a) nonpolar	(b) polar
(c) nonpolar	(d) slightly polar

Concept Exercise

Refer to the electronegativity values in Figure 12.9 and predict which of the following bonds is least polar: C—N, C—O, or C—I.

Answer: See Appendix G, 12.11.



▲ **Halogens** The two atoms in each halogen molecule are joined by a nonpolar covalent bond.

Diatomic Nonpolar Molecules

A **diatomic molecule** consists of two nonmetal atoms joined by a covalent bond. The simplest example of a nonpolar bond is found between two identical atoms. For example, oxygen occurs naturally as a gas and contains O_2 molecules. Even though oxygen is quite electronegative, the bond in an O_2 molecule is nonpolar. This is because each oxygen atom attracts the electron pair with the same force.

In Section 5.3, we briefly studied the elements in Group VIIA/17. Recall that, collectively, these elements are called *halogens.* The halogens (F, Cl, Br, and I) are a family and have similar properties. All four elements occur naturally as diatomic molecules: F_2, Cl_2, Br_2, and I_2. Elemental fluorine is a yellowish-green gas, chlorine is a greenish-yellow gas, bromine is a reddish-brown liquid, and iodine is a dark-violet crystalline solid.

Seven elements exist naturally as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. Although each molecule is distinctly different, each exhibits a nonpolar covalent bond. For our discussion, we will classify a bond as nonpolar when there is no electronegativity difference between the two bonded atoms. In practice, a bond between two atoms having an electronegativity difference of 0.5 or less is usually considered a nonpolar bond.

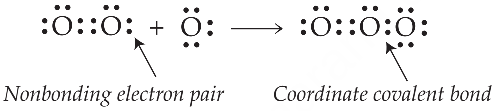
12.8 Coordinate Covalent Bonds

LEARNING OBJECTIVE

- Describe and identify a coordinate covalent bond.

We have described a covalent bond as an electron pair that is shared by two nonmetal atoms. Furthermore, each nonmetal atom is surrounded by nonbonding electron pairs to complete its octet.

Now, let's consider a special case of sharing electrons. An atom such as oxygen (six valence electrons) can attain an octet by sharing a nonbonding electron pair on another atom. A covalent bond resulting from one atom donating an electron pair to another atom is called a **coordinate covalent bond**. A good example of a coordinate covalent bond is found in a molecule of ozone, O_3. Here, an oxygen molecule donates a nonbonding electron pair to an oxygen atom to produce ozone.



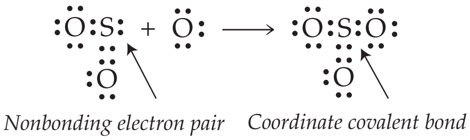
We can identify a coordinate covalent bond as follows. First, we draw the electron dot formula for the given molecule. If we can remove an atom and still have an octet around the remaining atoms, we know that the atom was joined by a coordinate covalent bond. The following example exercise further illustrates coordinate covalent bonds.

EXAMPLE EXERCISE 12.12 Coordinate Covalent Bonds

Burning yellow sulfur powder produces sulfur dioxide, SO_2. Sulfur dioxide is a colorless gas with a suffocating odor. It is used to kill insect larvae and produces the odor released when a match is ignited. Draw the electron dot formula for SO_2. Then show the formation of a coordinate covalent bond in SO_3.

Solution

The total number of valence electrons in one molecule of SO_2 is $6\ e^- + 2(6\ e^-) = 18\ e^-$. By trial and error, we find that the electron dot formula for SO_2 has a double bond. Because the sulfur atom has a nonbonding electron pair, it can bond to an additional oxygen atom. A diagram of the formation of the coordinate covalent bond is as follows:



Practice Exercise

A nitrogen molecule can form a coordinate covalent bond with an oxygen atom to give nitrous oxide, N_2O. Draw the electron dot formula for a nitrogen molecule and attach an oxygen atom.

Answers:

(a) :N:::N: (b) :N:::N:Ö:

Concept Exercise

Draw the electron dot formula for HClO and determine if the molecule contains a coordinate covalent bond.

Answers: See Appendix G, 12.12.

LEARNING OBJECTIVE

- Describe hydrogen bond attraction between two molecules.

12.9 Hydrogen Bonds

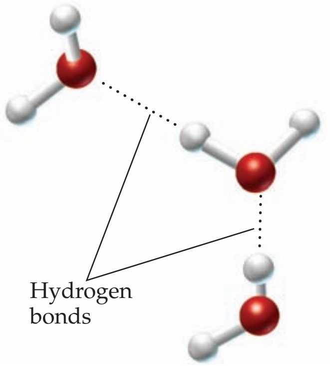
When a molecule has a polar covalent bond, it has a region that is positively charged, and a region that is negatively charged. This usually produces a molecule that has a positive end and a negative end. We can visualize this type of molecule as being similar to a miniature magnet. Like two magnets that are attracted to each other, two polar molecules are attracted to each other.

This attraction between two molecules can be significant, but is much less than the attraction between two atoms in a covalent bond. The bond length between molecules is considerably longer than the bond length between atoms in a polar bond. The most common example of this type of attraction is found between water molecules.

Recall that the bonds between H—O in water are highly polar because of the large electronegativity difference between the two elements. Thus, in a water molecule the oxygen atom is partially negative, and each hydrogen atom is partially positive. As a result, a H atom in one molecule of water is attracted to the nonbonding electron pair of the oxygen atom in another molecule of water. This intermolecular attraction between H and O in different molecules is referred to as a **hydrogen bond**.

Moreover, a hydrogen bond can result from the attraction of any two polar molecules having a hydrogen atom bonded to a highly electronegative atom (O, N, or F). Another example of hydrogen bonding occurs between ammonia molecules, NH_3. That is, the partially positive H atom in one ammonia molecule is attracted to the nonbonding electron pair of the electronegative nitrogen atom in a second molecule of ammonia.

Hydrogen bonds are especially important in living organisms where they are responsible for maintaining the structure of biochemical molecules (see Chapter 20). In large, complex biochemical molecules, hydrogen bonds are formed between H atoms and an electronegative atom (O or N) in the same molecule. In proteins, this can produce a single-stranded helix, whereas in DNA, hydrogen bonds are responsible for the double-stranded helix.



▲ **Hydrogen Bonding** This diagram shows hydrogen bonds between water molecules.

LEARNING OBJECTIVES

- Determine the shape of a molecule by applying VSEPR theory.
- Propose how a molecule with polar bonds can be nonpolar.

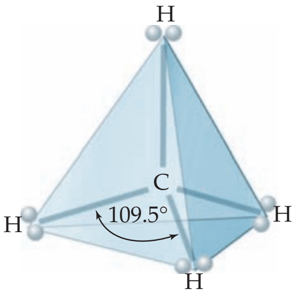
12.10 Shapes of Molecules

In the 1950s, a simple theory was proposed to explain the shapes of molecules. This theory is an extension of electron dot formulas in which pairs of electrons surround a central atom. In essence, the theory states that the electron pairs surrounding an atom tend to repel each other and that the shape of the molecule is the result of this electron pair repulsion. This model is referred to as the **valence shell electron pair repulsion** (**VSEPR**, pronounced "vesper") theory. The VSEPR model portrays bonding and nonbonding electron pairs as occupying specific positions around the central atom in a molecule.

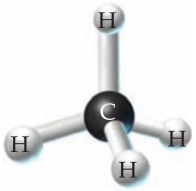
The VSEPR theory uses the term **electron pair geometry** to indicate the arrangement of bonding and nonbonding electron pairs around the central atom. VSEPR theory uses the term *molecular geometry*, or **molecular shape**, to indicate the arrangement of atoms around the central atom as a result of electron pair repulsion. The angle formed by any two atoms bonded to the central atom is referred to as the **bond angle**.

Tetrahedral Molecules

Let's apply VSEPR theory to a methane molecule, CH_4, which has four pairs of bonding electrons. Electron pairs repel each other, so they try to get away from each other. The result is that in a molecule of CH_4, the central **C** atom is surrounded by four electron pairs that are repelled to the four corners of a three-dimensional figure called a *tetrahedron.* Each electron pair is bonded to a hydrogen atom, and so the shape of the molecule is said to be *tetrahedral.* Moreover, according to VSEPR theory, any molecule with four electron pairs around the central atom has a tetrahedral electron pair geometry.



▲ The electron pair geometry in methane, CH_4, is *tetrahedral.*

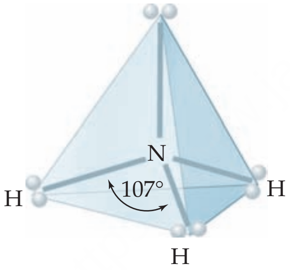


▲ The molecular shape of CH_4 is *tetrahedral;* the bond angle is 109.5°.

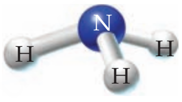
Trigonal Pyramidal Molecules

In a molecule of ammonia, $\mathbf{NH_3}$, the central **N** atom is surrounded by three bonding electron pairs and one nonbonding electron pair. The electron pair geometry is once again tetrahedral, so ammonia should theoretically have bond angles of 109.5°. Experimentally, the bond angle in $\mathbf{NH_3}$ is found to be 107°.

VSEPR theory explains the smaller bond angle by suggesting that the nonbonding electron pair exerts a stronger repelling force than the bonding pairs. Hence, the hydrogen atoms are pushed closer together, that is, from 109.5° to 107°. The molecular shape of $\mathbf{NH_3}$ is said to be *trigonal pyramidal.*



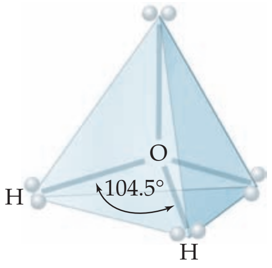
▲ The electron pair geometry in ammonia, NH_3, is *tetrahedral.*



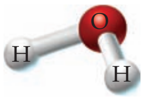
▲ The molecular shape of NH_3 is *trigonal pyramidal;* the bond angle is 107°.

Bent Molecules

In a molecule of water, H_2O, the central **O** atom is surrounded by two bonding electron pairs and two nonbonding electron pairs. Because the electron pair geometry is tetrahedral, the predicted bond angle between the two hydrogen atoms is 109.5°. The experimentally measured bond angle in $H_2\mathbf{O}$ is only 104.5°. As in the preceding example, it is proposed that the nonbonding electrons in water exert a greater repelling force than the bonding pairs. The resulting bond angle is smaller, allowing more space for the unshared electron pairs. The molecular shape of $H_2\mathbf{O}$ is said to be *bent* or V-shaped.



▲ The electron pair geometry in water, H_2O, is *tetrahedral.*



▲ The molecular shape of H_2O is *bent;* the bond angle is 104.5°.

Linear Molecules

In a molecule of carbon dioxide, CO_2, the central **C** atom is bonded to each O atom by two bonding electron pairs; that is, there is a double bond to each O atom ($O{=}\mathbf{C}{=}O$). The four pairs of electrons are on opposite sides of the **C** atom, so the three atoms lie in a straight line. The electron pair geometry in CO_2 is said to be linear, and the predicted bond angle between the two oxygen atoms is 180°. Similarly, the molecular shape is said to be *linear.*

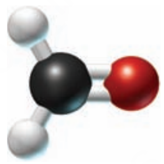
▲ The electron pair geometry in carbon dioxide, CO_2, is *linear.*

▲ The molecular shape of CO_2 is *linear*; the bond angle is 180°.

Trigonal Planar Molecules

In a molecule of formaldehyde, CH_2O, the central **C** atom is bonded to the O atom by a double bond ($\mathbf{C}{=}O$), and to each H atom by a single bond. The bonding electron pairs repel one another, so the three atoms are found at the corners of a triangle. The electron pair geometry in CH_2O is said to be trigonal planar, and the predicted bond angle between the atoms is 120°. Similarly, the molecular shape is said to be *trigonal planar.*

▲ The electron pair geometry in formaldehyde, CH_2O, is *trigonal planar.*



▲ The molecular shape of CH_2O is *trigonal planar*; the bond angle is 120°.

We can summarize the information obtained from VSEPR theory about these types of molecules, as shown in Table 12.1.

TABLE 12.1 Summary of VSEPR Theory

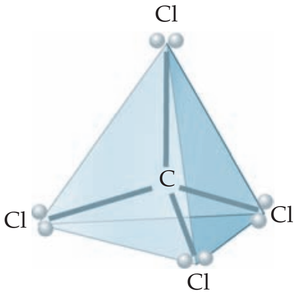
Example Molecule	Bonded Groups	Bonding Electron Pairs*	Nonbonding Electron Pairs*	Electron Pair Geometry	Molecular Shape	Bond Angle
CH_4	4	4	0	tetrahedral	tetrahedral	109.5°
NH_3	3	3	1	tetrahedral	trigonal pyramidal	107°
$H_2\mathbf{O}$	2	2	2	tetrahedral	bent	104.5°
CH_2O	3	4	0	trigonal planar	trigonal planar	120°
SO_2	2	3	1	trigonal planar	bent	120°
CO_2	2	4	0	linear	linear	180°

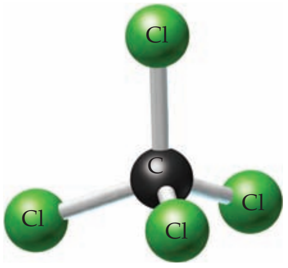
*Bonding electron pairs and nonbonding electron pairs that surround the central atom.

Nonpolar Molecules with Polar Bonds

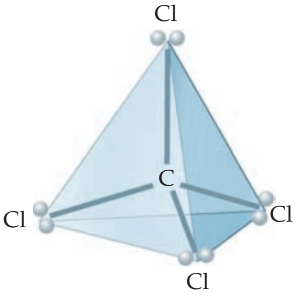
Why is carbon tetrachloride, CCl_4, a nonpolar molecule even though it has polar bonds? Recall that a chlorine atom is more electronegative than a carbon atom; hence, the $\mathbf{C}{-}Cl$ bond is polar. Furthermore, there are four $\mathbf{C}{-}Cl$ polar bonds in a CCl_4 molecule.

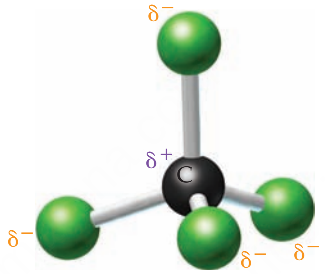
To explain why a CCl_4 molecule is nonpolar, we can apply VSEPR theory. There are four bonding electron pairs in CCl_4; thus, the electron pair geometry and the molecular shape are both tetrahedral. We illustrate the CCl_4 molecule as having four chlorine atoms at the corners of a tetrahedron.





Note that each chlorine atom exerts equal attractive forces in four opposing directions. Thus, the attractions exerted by the four polar bonds cancel each other. The net effect of the polar bonds is zero, and overall, the molecule is considered nonpolar. We can illustrate that the $\mathbf{C}Cl_4$ molecule is nonpolar as follows:



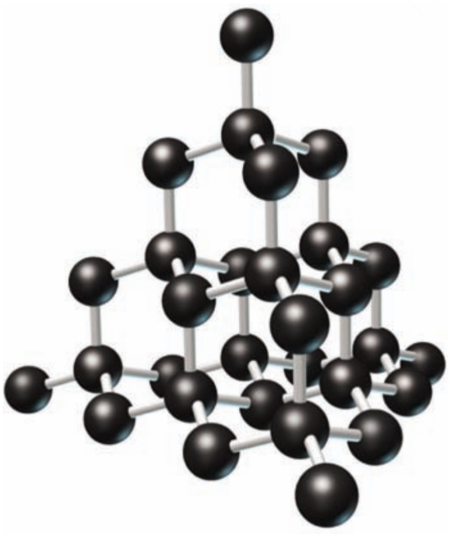


Consequently, a molecule may contain polar bonds and yet be nonpolar. In carbon dioxide, $\mathbf{C}O_2$, there are two polar bonds. An O═**C**═O molecule is nonpolar because the more electronegative oxygen atoms pull equally in opposite directions.

On the other hand, formaldehyde, $H_2\mathbf{C}{=}O$, is a polar molecule. The reason that formaldehyde is polar is that the $H_2\mathbf{C}{=}O$ molecule has only one oxygen atom and has a flat, triangular shape. The oxygen atom pulls electrons away from the central carbon atom and the two hydrogen atoms have little effect.

A CLOSER LOOK Diamond versus Graphite

Q: How is it possible that a natural diamond and the graphite in a pencil are both examples of pure carbon?

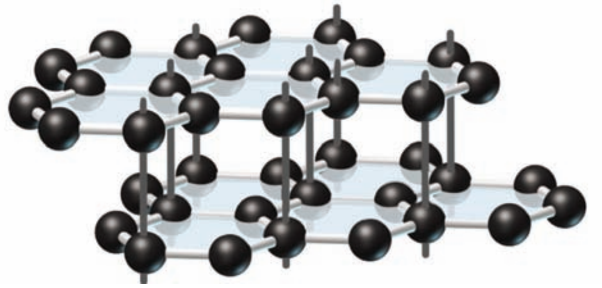


▲ **Diamond—Tetrahedral Structure**

Diamonds have been prized for centuries owing to their clarity and fiery sparkle. Diamonds originate deep within Earth, where they are formed under conditions of extreme heat and pressure. A diamond is the hardest known substance; its melting point is above 3500 °C, and it is inert to corrosive chemicals.

The hardness and crystalline structure of a diamond enables a rough stone to be cut so as to give numerous tiny facets. On average, an uncut stone loses 50% of its original weight in becoming a finished diamond. Moreover, a misdirected blow by a diamond cutter can shatter the stone into small shards.

Diamond is a crystalline solid containing carbon atoms arranged in a three-dimensional network. Carbon atoms are bonded in a repeating tetrahedral structure as shown in the illustration. The more extensive the repeating network, the larger the actual diamond.



▲ **Graphite—Planar Structure**

Like diamond, graphite is also a crystalline solid composed exclusively of carbon atoms. Unlike diamond, graphite is a soft substance that we know as the dark gray "lead" in an ordinary pencil. The structure of graphite consists of fused hexagon rings, each containing six carbon atoms. Unlike diamond, graphite is arranged in a flat, two-dimensional crystalline network. The lubricating property of graphite is due to "sheets" of carbons rings sliding past one another.

Diamonds can be synthesized from graphite by mimicking the conditions under which natural diamonds are formed. A sample of graphite is placed in a hydraulic press and converted to diamond in a few hours at high temperature and pressure. Typically, synthetic diamonds are only a few millimeters in size and are too flawed to be used as gemstones. However, synthetic diamonds are valuable industrially for use in coating drill bits and cutting tools.

▲ **Diamond and Graphite** Graphite and diamond are both examples of crystalline carbon. Powdered graphite can be converted to diamond when subjected to very high temperatures and pressures; for example, 3000 °C and 100,000 atm.

A: Although diamond and graphite are both composed of pure carbon, the structure of diamond is a repeating tetrahedron, whereas the structure of graphite is a repeating hexagon.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
12.1 The Chemical Bond Concept G. N. Lewis first described the **chemical bond** when he proposed that **valence electrons** in different atoms interact in such a way that each atom completes its valence shell. Because eight electrons are found in the valence shells of noble gas atoms, the Lewis theory became known as the *rule of eight*, or more commonly, the **octet rule**. The simplest representative particle in a substance held together by ionic bonds is a **formula unit**. When a metal atom loses valence electrons as a nonmetal atom gains electrons, an **ionic bond** results. The simplest representative particle held together by covalent bonds is a **molecule**. When valence electrons are shared by two nonmetal atoms, a **covalent bond** results.	• Discuss the role of valence electrons in a chemical bond. ***Related Exercises: 1–4*** • Predict whether a bond is ionic or covalent given a formula unit or molecule. ***Related Exercises: 5–10***
12.2 Ionic Bonds While forming ions, metal atoms lose valence electrons and are often left with a noble gas structure. Nonmetal atoms gain electrons to complete their valence shells. In the process of forming ions, the radii of metal atoms become smaller while the radii of nonmetal atoms become larger. The formation of ionic bonds between metal and nonmetal ions always releases heat energy.	• Illustrate the formation of an ionic bond between a metal atom and a nonmetal atom. ***Related Exercises: 11–28***
12.3 Covalent Bonds In the formation of covalent bonds, the outer valence shells of two nonmetal atoms overlap and share electrons. Each of the atoms in the bond shares the number of electrons necessary to complete their individual octets. The **bond length** between the two atoms is always less than the sum of their atomic radii. When atoms of different elements bond, energy is always released. The amount of energy released during the formation of a bond is exactly equal to the amount of energy required to break that bond. The amount of energy required to break a covalent bond between two atoms is called the **bond energy**.	• Illustrate the formation of a covalent bond between two nonmetal atoms by sharing valence electrons. ***Related Exercises: 29–32***

Key Concepts	Learning Objectives and Related Exercises
12.4 Electron Dot Formulas of Molecules To write an **electron dot formula**, we first calculate the total number of valence electrons from all of the atoms in the molecule. We then divide the valence electron total by two to find the number of electron pairs. The electron pairs are placed around the central atom and then the remaining atoms in the molecule so as to provide octets. One pair of electrons shared by two atoms is a **single bond**. A molecule can also contain two or three electron pairs between two atoms. These bonds are referred to as a **double bond** and a **triple bond**, respectively Double and triple bonds result from an insufficient number of valence electrons around an atom in a molecule. To provide an octet, it may be necessary to move **nonbonding electrons** between two atoms; these electrons then become **bonding electrons**. In the **structural formula** of a molecule, a single bond is shown as a dash, a double bond as two dashes, and a triple bond as three dashes.	• Diagram the electron dot formula for a molecule. ***Related Exercises: 33–38*** • Diagram the structural formula for a molecule. ***Related Exercises: 33–38***
12.5 Electron Dot Formulas of Polyatomic Ions An ionic bond can be formed by the attraction between a simple **monoatomic ion** and a complex **polyatomic ion**. Ammonium chloride, NH_4Cl, and cupric sulfate, $CuSO_4$, each contain a polyatomic ion. To draw the electron dot formula for a polyatomic ion, we must find the total number of valence electrons for each atom and then add or subtract the number of electrons equal to its ionic charge. For the ammonium cation, NH_4^+, the total number of valence electrons is 8 (9 e− minus 1 e−). For the sulfate ion, SO_4^{2-}, the total number of valence electrons is 32 e− (30 e− plus 2 e−).	• Diagram the electron dot formula for a polyatomic ion. ***Related Exercises: 39–44*** • Diagram the structural formula for a polyatomic ion. ***Related Exercises: 39–44***
12.6 Polar Covalent Bonds If one atom in a covalent bond attracts electrons more strongly than the other, that atom is said to have a higher **electronegativity**. A **polar bond** results when one of the bonded atoms has a greater attraction for the electron pair. We can indicate a polar bond using **delta notation**: the more electronegative atom is labeled δ^- and the less electronegative atom is labeled δ^+.	• Describe the electronegativity trends in the periodic table. ***Related Exercises: 45–50*** • Calculate the electronegativity difference in a polar covalent bond. ***Related Exercises: 51–52*** • Apply delta notation ($\delta+$ and $\delta-$) to a polar bond. ***Related Exercises: 53–54***
12.7 Nonpolar Covalent Bonds If each atom in a bond has an equal attraction for bonding electrons, the bond is said to be a **nonpolar bond**. A nonpolar bond always occurs in a **diatomic molecule** having two identical atoms: the bonds in naturally occurring H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2 are nonpolar bonds.	• Describe and identify a nonpolar covalent bond. ***Related Exercises: 55–56*** • Identify seven elements that occur naturally as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2. ***Related Exercises: 57–58***
12.8 Coordinate Covalent Bonds A **coordinate covalent bond** is a special type of covalent bond and is formed when one atom donates an electron pair to another atom. A simple illustration of a coordinate covalent bond occurs when the chlorine atom in HCl donates an electron pair to an oxygen atom to give HClO.	• Describe and identify a coordinate covalent bond. ***Related Exercises: 59–66***
12.9 Hydrogen Bonds A **hydrogen bond** usually refers to the attraction between a hydrogen atom in one molecule, and a nonbonding electron pair on a highly electronegative atom (O or N) in a second molecule. For example, this attraction occurs between water molecules, when the partially positive H atom in one molecule attracts the unshared electron pair on the electronegative O atom in another molecule.	• Describe hydrogen bond attraction between two molecules. ***Related Exercises: 67–70***

Key Concepts	Learning Objectives and Related Exercises
12.10 Shapes of Molecules After writing the electron dot formula for a molecule, we can draw its shape. That is, we can predict the shape of a molecule by applying valence shell electron pair repulsion (VSEPR) theory. According to **VSEPR theory**, bonding and nonbonding electron pairs surround the central atom. The arrangement of electron pairs around the central atom is termed **electron pair geometry**, and the arrangement of atoms is termed **molecular shape**. The angle formed by two atoms bonded to the central atom is referred to as the **bond angle**. If there are four electron pairs around the central atom, they are located at the corners of a three-dimensional figure called a *tetrahedron*. Examples of molecules having four electron pairs around the central atom include methane, CH_4, ammonia, NH_3, and water, H_2O. In each of these examples, the electron pair geometry is said to be *tetrahedral*.	• Determine the shape of a molecule by applying VSEPR theory. ***Related Exercises: 71–76*** • Propose how a molecule with polar bonds can be nonpolar. ***Related Exercises: 77–78***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** the attraction between positive and negative ions, or two nonmetal atoms
_____ **2.** the electrons in the outermost *s* and *p* sublevels of an atom that undergo reaction and form chemical bonds
_____ **3.** the statement that an atom tends to bond in such a way so as to acquire eight valence electrons
_____ **4.** a bond resulting from the attraction between a positively charged cation and a negatively charged anion
_____ **5.** a bond resulting from the sharing of one or more pairs of valence electrons between two nonmetal atoms
_____ **6.** the simplest representative particle in a substance held together by ionic bonds
_____ **7.** the simplest representative particle in a substance held together by covalent bonds
_____ **8.** the distance between the nuclei of two atoms joined by a covalent bond
_____ **9.** the amount of energy required to break a covalent bond between two atoms
_____ **10.** the valence electrons in a molecule that are shared
_____ **11.** the valence electrons in a molecule that are not shared
_____ **12.** a diagram of a molecule that shows the chemical symbol of each atom and a dot representing each bonding and nonbonding electron
_____ **13.** a diagram of a molecule that shows the chemical symbol of each atom and a dash representing each pair of bonding electrons
_____ **14.** a bond composed of one electron pair shared by two atoms
_____ **15.** a bond composed of two electron pairs shared by two atoms
_____ **16.** a bond composed of three electron pairs shared by two atoms
_____ **17.** a single atom that bears a negative or positive charge as the result of gaining or losing valence electrons
_____ **18.** a group of atoms held together by covalent bonds and having an overall positive or negative charge
_____ **19.** the ability of an atom to attract a shared pair of electrons in a covalent bond
_____ **20.** a method of indicating a partial positive and a partial negative charge in a chemical bond
_____ **21.** a bond in which a pair of electrons is shared unequally
_____ **22.** a bond in which a pair of electrons is shared equally
_____ **23.** a molecule composed of two nonmetal atoms held together by a covalent bond
_____ **24.** a bond in which an electron pair is shared but both electrons have been donated by a single atom

(a) bond angle (*Sec. 12.10*)
(b) bond energy (*Sec. 12.3*)
(c) bond length (*Sec. 12.3*)
(d) bonding electrons (*Sec. 12.4*)
(e) chemical bond (*Sec. 12.1*)
(f) coordinate covalent bond (*Sec. 12.8*)
(g) covalent bond (*Sec. 12.1*)
(h) delta (δ) notation (*Sec. 12.6*)
(i) diatomic molecule (*Sec. 12.7*)
(j) double bond (*Sec. 12.4*)
(k) electron dot formula (*Sec. 12.4*)
(l) electron pair geometry (*Sec. 12.10*)
(m) electronegativity (*Sec. 12.6*)
(n) formula unit (*Sec. 12.1*)
(o) hydrogen bond (*Sec. 12.9*)
(p) ionic bond (*Sec. 12.1*)
(q) molecular shape (*Sec. 12.10*)
(r) molecule (*Sec. 12.1*)
(s) monoatomic ion (*Sec. 12.5*)
(t) nonbonding electrons (*Sec. 12.4*)
(u) nonpolar bond (*Sec. 12.7*)
(v) octet rule (*Sec. 12.1*)
(w) polar bond (*Sec. 12.6*)
(x) polyatomic ion (*Sec. 12.5*)
(y) single bond (*Sec. 12.4*)

_____ **25.** the attraction between a H atom bonded to O or N, and a nonbonding electron pair on an electronegative atom in another molecule

_____ **26.** a model that explains the shapes of molecules as a result of electron pairs about the central atom repelling each other

_____ **27.** the geometric shape formed by bonding and nonbonding electron pairs about the central atom in a molecule

_____ **28.** the geometric shape formed by atoms bonded to the central atom in a molecule

_____ **29.** the angle formed by two atoms attached to the central atom in a molecule

(z) structural formula (*Sec. 12.4*)
(aa) triple bond (*Sec. 12.4*)
(bb) valence electrons (*Sec. 12.1*)
(cc) VSEPR theory (*Sec. 12.10*)

Exercises

Answers to odd-numbered Exercises are in Appendix I.

The Chemical Bond Concept (Sec. 12.1)

1. State the number of valence electrons for a sodium atom and a chlorine atom.
2. State the number of valence electrons for a hydrogen atom and a chlorine atom.
3. State the number of valence electrons for a sodium ion and a chloride ion in an ionic bond.
4. State the number of valence electrons for a hydrogen atom and a chlorine atom in a covalent bond.
5. Predict whether each of the following is held together by ionic or covalent bonds.
   (a) aluminum chloride, $AlCl_3$
   (b) water, H_2O
   (c) sulfur trioxide, SO_3
   (d) iron(II) carbonate, $FeCO_3$
6. Predict whether each of the following is held together by ionic or covalent bonds.
   (a) zinc bromide, $ZnBr_2$
   (b) ammonia, NH_3
   (c) iodine heptafluoride, IF_7
   (d) lead(II) sulfate, $PbSO_4$
7. State whether the representative particle in each of the following substances is a formula unit or a molecule.
   (a) methanol, CH_3OH
   (b) potassium chlorite, $KClO_2$
   (c) manganese(II) oxide, MnO_2
   (d) acetic acid, $HC_2H_3O_2$
8. State whether the representative particle in each of the following substances is a formula unit or a molecule.
   (a) ethanol, CH_3CH_2OH
   (b) sodium chlorate, $NaClO_3$
   (c) nickel(II) oxide, NiO
   (d) sulfuric acid, H_2SO_4
9. State whether the representative particle in each of the following substances is an atom, a formula unit, or a molecule.
   (a) propane, C_3H_8
   (b) chromium, Cr
   (c) hematite, Fe_2O_3
   (d) sulfur, S_8
10. State whether the representative particle in each of the following substances is an atom, a formula unit, or a molecule.
    (a) acetone, C_3H_6O
    (b) cobalt, Co
    (c) magnetite, Fe_3O_4
    (d) phosphorus, P_4

Ionic Bonds (Sec. 12.2)

11. Use the periodic table to predict an ionic charge for each of the following metal ions.
    (a) Li ion
    (b) Mg ion
    (c) Al ion
    (d) Sn ion
12. Use the periodic table to predict an ionic charge for each of the following metal ions.
    (a) Na ion
    (b) Ba ion
    (c) Ga ion
    (d) Pb ion
13. Use the periodic table to predict the ionic charge for each of the following nonmetal ions.
    (a) F ion
    (b) Br ion
    (c) S ion
    (d) N ion
14. Use the periodic table to predict the ionic charge for each of the following nonmetal ions.
    (a) Cl ion
    (b) I ion
    (c) Se ion
    (d) P ion
15. Write out the electron configuration for each of the following metal ions.
    (a) Na^+
    (b) Mg^{2+}
    (c) Al^{3+}
    (d) V^{4+}
16. Write out the electron configuration for each of the following metal ions.
    (a) K^+
    (b) Ca^{2+}
    (c) Sc^{3+}
    (d) Ti^{4+}
17. Write out the electron configuration for each of the following nonmetal ions.
    (a) F^-
    (b) O^{2-}
    (c) N^{3-}
    (d) C^{4-}
18. Write out the electron configuration for each of the following nonmetal ions.
    (a) Cl^-
    (b) S^{2-}
    (c) P^{3-}
    (d) Si^{4-}
19. Refer to the periodic table and state the noble gas with an electron configuration identical to each of the following ions.
    (a) S^{2-}
    (b) Cl^-
    (c) K^+
    (d) Ca^{2+}
20. Refer to the periodic table and state the noble gas with an electron configuration identical to each of the following ions?
    (a) Se^{2-}
    (b) Br^-
    (c) Rb^+
    (d) Sr^{2+}

21. Which noble gas is isoelectronic with each of the following metal ions?
(a) Li^{+}
(b) Al^{3+}
(c) Ca^{2+}
(d) Mg^{2+}

22. Which noble gas is isoelectronic with each of the following metal ions?
(a) Sc^{3+}
(b) K^{+}
(c) Ti^{4+}
(d) Ba^{2+}

23. Which noble gas is isoelectronic with each of the following nonmetal ions?
(a) Cl^{-}
(b) I^{-}
(c) S^{2-}
(d) P^{3-}

24. Which noble gas is isoelectronic with each of the following nonmetal ions?
(a) Br^{-}
(b) O^{2-}
(c) Se^{2-}
(d) N^{3-}

25. In each of the following pairs, which has the larger radius?
(a) Li atom or Li ion
(b) Mg atom or Mg ion
(c) F atom or F ion
(d) O atom or O ion

26. In each of the following pairs, which has the larger radius?
(a) Al atom or Al ion
(b) Pb atom or Pb ion
(c) Se atom or Se ion
(d) N atom or N ion

27. Which of the following statements are true regarding an ionic bond between calcium and oxygen in an CaO formula unit?
(a) Calcium ions and oxide ions bond by covalent attraction.
(b) Calcium atoms lose electrons and oxygen atoms gain electrons.
(c) The ionic radius of a calcium ion is less than its atomic radius.
(d) Breaking an ionic bond between calcium and oxygen releases energy.

28. Which of the following statements are true regarding an ionic bond between cobalt and sulfur in a CoS formula unit?
(a) Cobalt ions and sulfide ions bond by electrostatic attraction.
(b) Cobalt atoms gain electrons and sulfur atoms lose electrons.
(c) The ionic radius of a cobalt ion is greater than its atomic radius.
(d) Breaking an ionic bond between cobalt and sulfur requires energy.

Covalent Bonds (Sec. 12.3)

29. Which of the following is less?
(a) the sum of the H and Cl atomic radii, or the bond length in $H{-}Cl$
(b) the sum of the O and O atomic radii, or the bond length in $O{=}O$

30. Which of the following is less?
(a) the sum of the H and Br atomic radii, or the bond length in $H{-}Br$
(b) the sum of the S and O atomic radii, or the bond length in $S{=}O$

31. Which of the following statements are true regarding a covalent bond between nitrogen and oxygen atoms in a nitric oxide, NO, molecule?
(a) Valence electrons are shared between nitrogen and oxygen atoms.
(b) Bonding electrons are found only between the bonded atoms.
(c) The bond length is greater than the sum of the two atomic radii.
(d) Energy is required to break a covalent bond.

32. Which of the following statements are true regarding a covalent bond between carbon and oxygen atoms in a carbon monoxide, CO, molecule?
(a) Valence electrons are transferred from carbon to oxygen atoms.
(b) Bonding electrons are distributed over the entire CO molecule.
(c) The bond length is less than the sum of the two atomic radii.
(d) Energy is released when a covalent bond is broken.

Electron Dot Formulas of Molecules (Sec. 12.4)

*(Central atoms are indicated in **bold.**)*

33. Write the electron dot formula and draw the structural formula for each of the following molecules.
(a) $\mathbf{H}_2$
(b) $\mathbf{F}_2$
(c) $H\mathbf{Br}$
(d) $\mathbf{N}H_3$

34. Write the electron dot formula and draw the structural formula for each of the following molecules.
(a) $\mathbf{Cl}_2$
(b) $\mathbf{O}_2$
(c) $H\mathbf{I}$
(d) $\mathbf{P}H_3$

35. Write the electron dot formula and draw the structural formula for each of the following molecules.
(a) $\mathbf{N}_2$
(b) $\mathbf{P}I_3$
(c) $HO\mathbf{N}O$
(d) $\mathbf{C}_2H_4$

36. Write the electron dot formula and draw the structural formula for each of the following molecules.
(a) $H\mathbf{O}Cl$
(b) $\mathbf{S}O_2$
(c) $\mathbf{C}S_2$
(d) $\mathbf{C}_2H_2$

37. Write the electron dot formula and draw the structural formula for each of the following molecules.
(a) $\mathbf{C}H_4$
(b) $\mathbf{O}F_2$
(c) $H_2\mathbf{O}_2$
(d) $\mathbf{N}F_3$

38. Write the electron dot formula and draw the structural formula for each of the following molecules.
(a) $\mathbf{C}Cl_4$
(b) $HO\mathbf{N}O_2$
(c) $\mathbf{C}H_3OH$
(d) $HO\mathbf{C}N$

Electron Dot Formulas of Polyatomic Ions (Sec. 12.5)

*(Central atoms are indicated in **bold.**)*

39. Write the electron dot formula and draw the structural formula for each of the following polyatomic ions.
(a) $\mathbf{I}O^{-}$
(b) $\mathbf{I}O_2^{-}$
(c) $\mathbf{I}O_3^{-}$
(d) $\mathbf{I}O_4^{-}$

40. Write the electron dot formula and draw the structural formula for each of the following polyatomic ions.
(a) $\mathbf{Br}O^{-}$
(b) $\mathbf{Br}O_2^{-}$
(c) $\mathbf{Br}O_3^{-}$
(d) $\mathbf{Br}O_4^{-}$

41. Write the electron dot formula and draw the structural formula for each of the following polyatomic ions.
(a) $\mathbf{S}O_4^{2-}$
(b) $H\mathbf{S}O_4^{-}$
(c) $\mathbf{S}O_3^{2-}$
(d) $H\mathbf{S}O_3^{-}$

42. Write the electron dot formula and draw the structural formula for each of the following polyatomic ions.

(a) $\mathbf{P}O_4{}^{3-}$ **(b)** $H\mathbf{P}O_4{}^{2-}$

(c) $\mathbf{P}O_3{}^{3-}$ **(d)** $H\mathbf{P}O_3{}^{2-}$

43. Write the electron dot formula and draw the structural formula for each of the following polyatomic ions.

(a) $H_3\mathbf{O}^+$ **(b)** $\mathbf{O}H^-$

(c) $\mathbf{S}H^-$ **(d)** $\mathbf{C}N^-$

44. Write the electron dot formula and draw the structural formula for each of the following polyatomic ions.

(a) $\mathbf{P}H_4{}^+$ **(b)** $\mathbf{Se}O_3{}^{2-}$

(c) $\mathbf{C}O_3{}^{2-}$ **(d)** $\mathbf{B}O_3{}^{3-}$

Polar Covalent Bonds (Sec. 12.6)

45. What is the general trend in electronegativity within a period of elements in the periodic table?

46. What is the general trend in electronegativity within a group of elements in the periodic table?

47. Which elements are more electronegative: metals or nonmetals?

48. Which elements are more electronegative: semimetals or nonmetals?

49. Predict which element in each of the following pairs is more electronegative according to the general trends in the periodic table.

(a) Br or Cl **(b)** O or S

(c) Se or As **(d)** N or F

50. Predict which element in each of the following pairs is more electronegative according to the general trends in the periodic table.

(a) Se or Br **(b)** C or B

(c) Te or S **(d)** Ba or Be

51. Refer to the values in Figure 12.9 and calculate the electronegativity difference in each of the following bonds.

(a) Br—Cl **(b)** Br—F

(c) I—Cl **(d)** I—Br

52. Refer to the values in Figure 12.9 and calculate the electronegativity difference in each of the following bonds.

(a) H—Cl **(b)** H—Br

(c) N—O **(d)** C—O

53. Refer to Figure 12.9 and label each atom in the following polar covalent bonds using delta notation (δ^+ and δ^-).

(a) C—H **(b)** Se—O

(c) P—I **(d)** H—Br

54. Refer to Figure 12.9 and label each atom in the following polar covalent bonds using delta notation (δ^+ and δ^-).

(a) H—S **(b)** O—S

(c) N—F **(d)** S—Cl

Nonpolar Covalent Bonds (Sec. 12.7)

55. Refer to Figure 12.9 and indicate which of the following are nonpolar covalent bonds.

(a) Cl—Cl **(b)** Cl—N

(c) N—H **(d)** H—P

56. Refer to Figure 12.9 and indicate which of the following are nonpolar covalent bonds.

(a) I—C **(b)** C—S

(c) S—H **(d)** H—Br

57. Which of the following elements occur naturally as diatomic molecules: H, N, Cl, Br, Ar?

58. Which of the following elements occur naturally as diatomic molecules: He, O, F, I, Ne?

Coordinate Covalent Bonds (Sec. 12.8)

59. An oxygen atom can bond to a hydrogen bromide molecule to give HBrO. Draw the electron dot formula for $H\mathbf{Br}O$ and label a coordinate covalent bond.

60. An oxygen atom can bond to a hydrogen iodide molecule to give HIO. Draw the electron dot formula for $H\mathbf{I}O$ and label a coordinate covalent bond.

61. An oxygen atom can bond to a HBrO molecule to give $HBrO_2$. Draw the electron dot formula for $H\mathbf{Br}O_2$ and label a coordinate covalent bond.

62. An oxygen atom can bond to a HIO molecule to give HIO_2. Draw the electron dot formula for $H\mathbf{I}O_2$ and label a coordinate covalent bond.

63. A hydrogen ion can bond to an ammonia molecule, NH_3, forming $NH_4{}^+$. Draw the electron dot formula for $\mathbf{N}H_4{}^+$ and label a coordinate covalent bond.

64. A hydrogen ion can bond to a phosphine molecule, PH_3, forming $PH_4{}^+$. Draw the electron dot formula for $\mathbf{P}H_4{}^+$ and label a coordinate covalent bond.

65. A nitrite ion, $NO_2{}^-$, can bond to an oxygen atom to form the nitrate ion. Draw the electron dot formula for $\mathbf{N}O_3{}^-$ and label a coordinate covalent bond.

66. A phosphite ion, $PO_3{}^{3-}$, can bond to an oxygen atom to form the phosphate ion. Draw the electron dot formula for $\mathbf{P}O_4{}^{3-}$ and label a coordinate covalent bond.

Hydrogen Bonds (Sec. 12.9)

67. Diagram and label the hydrogen bond between two molecules of hydrogen fluoride, HF.

68. Diagram and label the hydrogen bond between two molecules of ammonia, NH_3.

69. Which is greater, the energy to break a hydrogen bond or a polar covalent bond?

70. Which is longer, the length of a hydrogen bond or a polar covalent bond?

Shapes of Molecules (Sec. 12.10)

71. Predict the electron pair geometry, the molecular shape, and the bond angle for a silane molecule, $\mathbf{Si}H_4$, using VSEPR theory.

72. Predict the electron pair geometry, the molecular shape, and the bond angle for a carbon tetrabromide molecule, $\mathbf{C}Br_4$, using VSEPR theory.

73. Predict the electron pair geometry, the molecular shape, and the bond angle for a nitrogen triiodide molecule, $\mathbf{N}I_3$, using VSEPR theory.

74. Predict the electron pair geometry, the molecular shape, and the bond angle for a phosphine molecule, $\mathbf{P}H_3$, using VSEPR theory.

75. Predict the electron pair geometry, the molecular shape, and the bond angle for a hydrogen sulfide molecule, $H_2\mathbf{S}$, using VSEPR theory.

76. Predict the electron pair geometry, the molecular shape, and the bond angle for a dichlorine monoxide molecule, $Cl_2\mathbf{O}$, using VSEPR theory.

77. Apply VSEPR theory to explain why $\mathbf{C}F_4$ is a nonpolar molecule even though it has four polar bonds.

78. Apply VSEPR theory to explain why $\mathbf{Si}F_4$ is a nonpolar molecule even though it has four polar bonds.

General Exercises

79. State whether the representative particle in each of the following substances is an atom, a molecule, or a formula unit.
   (a) uranium, U
   (b) fluorine, F_2
   (c) uranium hexafluoride, UF_6
   (d) hydrogen fluoride, HF

80. State whether the representative particle in each of the following substances is an atom, a molecule, or a formula unit.
   (a) plutonium, Pu
   (b) oxygen, O_2
   (c) plutonium(III) oxide, Pu_2O_3
   (d) hydrogen peroxide, H_2O_2

81. Write formula units by combining the cations and anions in each of the following pairs.
   (a) Sr^{2+} and As^{3-}
   (b) Ra^{2+} and O^{2-}
   (c) Al^{3+} and $CO_3{}^{2-}$
   (d) Cd^{2+} and OH^-

82. Write formula units by combining the cations and anions in each of the following pairs.
   (a) Sc^{3+} and N^{3-}
   (b) Ti^{4+} and O^{2-}
   (c) $NH_4{}^+$ and $CO_3{}^{2-}$
   (d) $Hg_2{}^{2+}$ and $PO_4{}^{3-}$

83. Explain why the radius of a sodium ion (95 pm) is about half that of a sodium atom (186 pm).

84. Explain why the radius of a chloride ion (181 pm) is about twice that of a chlorine atom (99 pm).

85. Refer to the values in Figure 12.9 and calculate the electronegativity difference in a B—Cl bond.

86. Refer to the values in Figure 12.9 and calculate the electronegativity difference in a Sb—Cl bond.

87. Refer to the values in Figure 12.9 and calculate the electronegativity difference in a H—P bond.

88. Refer to the values in Figure 12.9 and calculate the electronegativity difference in a S—I bond.

89. Label the polar Ge—Cl bond using delta notation (δ^+ and δ^-).

90. Label the polar As—Cl bond using delta notation (δ^+ and δ^-).

91. Write the electron dot formula and draw the structural formula for silane, $\mathbf{Si}H_4$, whose central atom is a semimetal.

92. Write the electron dot formula and draw the structural formula for stibine, $\mathbf{Sb}H_3$, whose central atom is a semimetal.

93. Write the electron dot formula and draw the structural formula for the arsenate ion, $\mathbf{As}O_3{}^{3-}$, whose central atom is a semimetal.

94. Write the electron dot formula and draw the structural formula for the silicate ion, $\mathbf{Si}O_3{}^{2-}$, whose central atom is a semimetal.

95. Predict the electron pair geometry, the molecular shape, and the bond angle for a carbon disulfide molecule, $\mathbf{C}S_2$, using VSEPR theory.

96. Predict the electron pair geometry, the molecular shape, and the bond angle for a silicon dioxide molecule, $\mathbf{Si}O_2$, using VSEPR theory.

97. Predict the electron pair geometry, the molecular shape, and the bond angle for a phosgene molecule, $\mathbf{C}Cl_2O$, using VSEPR theory.

98. Predict the electron pair geometry, the molecular shape, and the bond angle for a carbonate ion, $\mathbf{C}O_3{}^{2-}$, using VSEPR theory.

Challenge Exercises

99. Xenon dioxide is a slightly stable molecule that contains a noble gas. Write the electron dot formula and draw the structural formula for $\mathbf{Xe}O_2$.

100. Xenon trioxide is a slightly stable molecule that contains a noble gas. Write the electron dot formula and draw the structural formula for $\mathbf{Xe}O_3$.

Chapter 12 Self-Test

Answers to Self-Test are in Appendix J.

1. Predict which of the following has an ionic bond. (Sec. 12.1)
   (a) CaO
   (b) NO
   (c) CO
   (d) all of the above
   (e) none of the above

2. Which of the following statements is true regarding an ionic bond between zinc and oxygen in a ZnO formula unit? (Sec. 12.2)
   (a) Zinc atoms lose electrons and oxygen atoms gain electrons.
   (b) The ionic radius of a zinc ion is less than its atomic radius.
   (c) Forming a bond between zinc and oxide ions releases energy.
   (d) all of the above
   (e) none of the above

3. Which of the following statements is true regarding a covalent bond between hydrogen and oxygen atoms in an H_2O molecule? (Sec. 12.3)
   (a) Bonding electrons are stationary between the H and O atoms.
   (b) The bond length is greater than the sum of the two atomic radii.
   (c) Forming a bond between H and O atoms requires energy.
   (d) all of the above
   (e) none of the above

4. Draw the electron dot formula for hydrogen peroxide, H_2O_2. How many nonbonding electron pairs are in a hydrogen peroxide molecule? (Sec. 12.4)
   (a) 1
   (b) 2
   (c) 3
   (d) 4
   (e) none of the above

5. Draw the electron dot formula for the carbonate ion, CO_3^{2-}. How many nonbonding electron pairs are in a carbonate ion? (Sec. 12.5)
(a) 4 **(b)** 8
(c) 12 **(d)** 24
(e) none of the above

6. Which of the following shows a general increasing trend for electronegativity in the periodic table? (Sec. 12.6)
(a) up a group of metals
(b) up a group of nonmetals
(c) across a series of metals from left to right
(d) across a series of nonmetals from left to right
(e) all of the above

7. Given the values for H (2.1) and F (4.0), calculate the electronegativity difference in a hydrogen fluoride bond, HF. (Sec. 12.7)
(a) -1.9 **(b)** -6.1
(c) 1.9 **(d)** 6.1
(e) none of the above

8. Chlorine has an electronegativity value of 3.0. Given the electronegativity of C, N, and O (2.5, 3.0, and 3.5, respectively), which of the following molecules has nonpolar bonds? (Sec. 12.8)
(a) CCl_4 **(b)** NCl_3
(c) Cl_2O **(d)** all of the above
(e) none of the above

9. What is the electron pair geometry for a molecule of H_2S? (Sec. 12.9)
(a) bent **(b)** linear
(c) tetrahedral **(d)** trigonal pyramidal
(e) none of the above

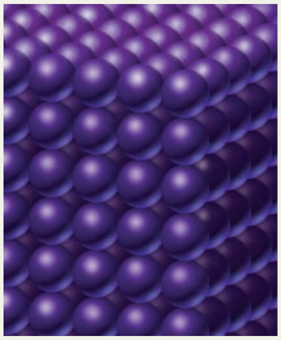
10. What is the molecular shape for a molecule of H_2S? (Sec. 12.10)
(a) bent **(b)** linear
(c) tetrahedral **(d)** trigonal pyramidal
(e) none of the above

Key Concepts

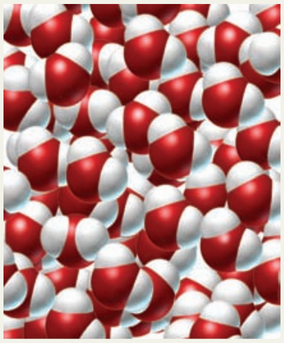
11. Classify each of the following diagrams as illustrating a substance composed of atoms, formula units, or molecules.

12. Based on electronegativity trends in the periodic table, which of the following molecules is most polar: HF, HCl, HBr, HI?

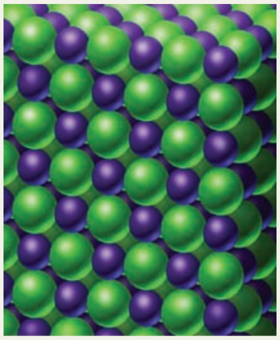
(a)



(b)



(c)



13. Given a H_2O molecule, how many additional H_2O molecules can attach directly to the given water molecule by hydrogen bonds?

14. Using VSEPR theory, predict the electron pair geometry and molecular shape for a chloroform molecule, $CHCl_3$.

Critical Thinking

15. "Acid rain" can originate from the sulfur dioxide produced from burning coal. Draw the structural formulas for two resonance structures of SO_2. Which of the two bonds in an SO_2 molecule is stronger?

16. "Acid rain" can originate from the sulfur trioxide produced from burning coal. Draw the structural formulas for two resonance structures of SO_3. Which of the three bonds in an SO_3 molecule is strongest?

17. Using VSEPR theory, contrast the molecular shape of a water molecule, H_2O, with that of a hydronium ion, H_3O^+.

18. Using VSEPR theory, contrast the molecular shape of an ammonia molecule, NH_3, with that of an ammonium ion, NH_4^+.

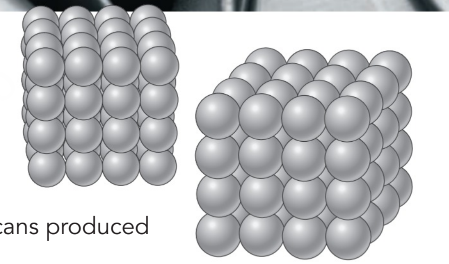
CHAPTER 13

Solutions

> *"Do not let what you cannot do interfere with what you can do."*
>
> **John Wooden, legendary UCLA basketball coach (1910–2010)**

13
Al
aluminum

Element 13: Aluminum
Aluminum is the most common metal in Earth's crust. Aluminum is a lightweight metal that resists corrosion and is used in aerospace manufacturing. It's principle use is that of making beverage containers. Typically, the United States recycles over 65% of the aluminum cans produced each year.



A solution is a homogeneous mixture. This means that a solution is the same throughout, and that every sample of it has the same properties. A solution consists of a solute dissolved in a solvent. For instance, the solute sugar dissolved in the solvent water gives an aqueous solution of sugar. Other examples of solutions are carbon dioxide dissolved in a soft drink and salt dissolved in the ocean. The proportion of solute and solvent can vary, but the solvent is always the greater quantity.

We usually think of solutions as liquids. However, solutes scattered throughout solvents exist in all three physical states: gaseous, liquid, and solid solutions. Table 13.1 lists some common examples of solutions whose physical state corresponds to that of the solvent.

TABLE 13.1 Types of Solutions

Solute	Solvent	Solution
Gaseous Solutions		
gas	gas	air (nitrogen, oxygen, argon gases)
liquid	gas	humid air (water vapor in air)
Liquid Solutions		
gas	liquid	carbonated drinks (CO_2 in water)
liquid	liquid	vinegar (acetic acid in water)
solid	liquid	salt water (NaCl in water)
Solid Solutions		
liquid	solid	dental amalgam (Hg in Ag)
solid	solid	sterling silver (Cu in Ag)

13.1 Gases in Solution

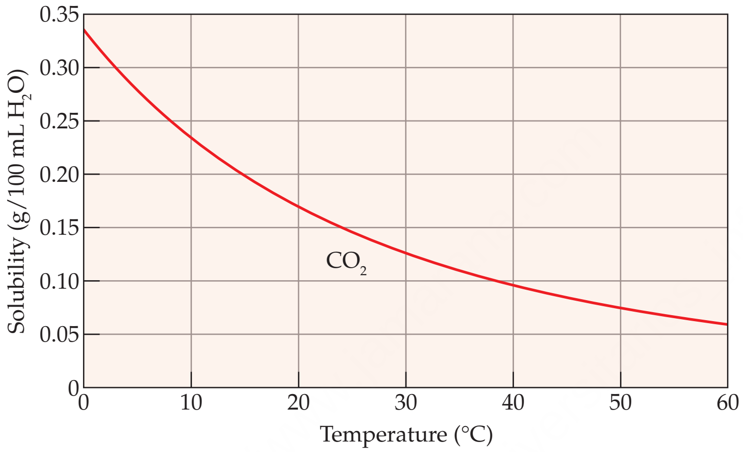
LEARNING OBJECTIVES

- Describe the effect of temperature on the solubility of a gas in a liquid.
- Describe the effect of pressure on the solubility of a gas in a liquid.

Hydrochloric acid is sold in the supermarket as muriatic acid, which is used to acidify swimming pools and clean concrete. Hydrochloric acid is HCl gas dissolved in water. Ammonia solutions have NH_3 gas dissolved in water. Carbonated drinks have CO_2 gas dissolved in water.

Solubility and Temperature

Solutions of gases in liquids are greatly affected by changes in temperature. As the temperature increases, gas molecules acquire more energy and escape from the solvent. Thus, as the temperature increases, the solubility of a gas in a liquid decreases.



Solubility of a Gas in a Liquid Notice the solubility of carbon dioxide gas in water *decreases* as the temperature *increases*.

A practical example of this principle is illustrated by a carbonated soda. When you take a cold can of soda from the refrigerator and pop the tab, you notice slight fizzing as CO_2 escapes from the solution. However, if you pop the tab on a can of warm soda, you notice intense foaming as CO_2 escapes. Table 13.2 indicates the effect of temperature on the solubility of carbon dioxide in water.

Solubility and Pressure

The solubility of a gas in a liquid is strongly influenced by pressure. In 1803, the English chemist William Henry conducted experiments on the solubility of gases in liquids. He found the solubility of a gas was proportional to the partial pressure of the gas above the liquid. This principle is known as **Henry's law**. That is, if the partial pressure of a

TABLE 13.2 Solubility of Carbon Dioxide in Water

Temperature	Pressure	Solubility of CO_2
Temperature Effect		
0 °C	1.00 atm	0.348 g/100 mL water
20 °C	1.00 atm	0.176 g/100 mL water
40 °C	1.00 atm	0.097 g/100 mL water
60 °C	1.00 atm	0.058 g/100 mL water
Pressure Effect		
0 °C	1.00 atm	0.348 g/100 mL water
0 °C	2.00 atm	0.696 g/100 mL water
0 °C	3.00 atm	1.044 g/100 mL water

Notice at higher temperatures the solubility of a gas *decreases*, but at higher pressures the solubility *increases*.

gas increases, the solubility of the gas increases proportionally. Table 13.2 indicates the effect of pressure on the solubility of CO_2 in water.

Let's apply Henry's law to calculate the solubility of CO_2 when the pressure increases from 1.00 atm to 2.00 atm. The solubility of CO_2 is given in Table 13.2 as 0.348 g/100 mL water at 0 °C and 1.00 atm.

$$0.348\text{ g}/100\text{ mL} \times \frac{2.00\ \cancel{\text{atm}}}{1.00\ \cancel{\text{atm}}} = 0.696\text{ g}/100\text{ mL}$$

Notice the solubility of the carbon dioxide in water given in Table 13.2 agrees with the calculated value, 0.696 g/100 mL. The following example exercise further illustrates Henry's law.

A CLOSER LOOK Deep-Sea Diving

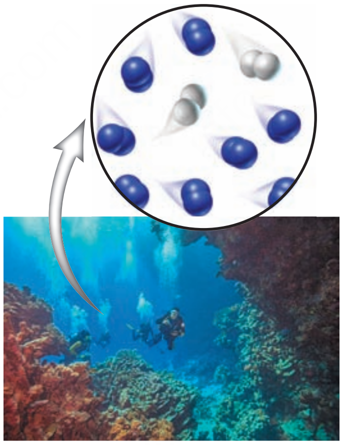
Q: Why is a gaseous mixture of helium and oxygen substituted for air when deep-sea diving?

Diving to a depth of 100 feet or more is considered a deep dive. At 100 feet the atmospheric pressure is about four times that of sea level and the amounts of dissolved gases in the blood are about four times greater according to Henry's law. At this depth, a diver may experience a phenomenon known as "nitrogen narcosis," or "rapture of the deep." A high concentration of nitrogen gas produces a disorientating effect similar to drunkenness.

Nitrogen narcosis is usually reversed in a few minutes by slowly ascending to a shallower depth. If a diver ascends too rapidly, the dissolved nitrogen does not have time to effuse from the circulatory system. Thus, the diver suffers from a painful condition referred to as decompression sickness, or the "bends." To avoid the bends, a diver can substitute a gaseous mixture of helium and oxygen, which is termed "heliox." The helium gas in a heliox mixture effuses more rapidly from the circulatory system than nitrogen gas, thus avoiding the painful consequences.

In extreme cases, when a diver is forced to ascend too rapidly, the diver can be placed in a hyperbaric chamber. A hyperbaric chamber is an enclosure where gas pressure can be regulated. The chamber can be underwater or out of the water, and in it a diver breathes oxygen or air at an elevated pressure that is slowly decreased to normal over a period of time.

For very deep dives, down to 2,000 feet, a diver can use an atmospheric diving suit. A diving suit is capable of withstanding pressures that are many times normal. Using a pressurized diving suit eliminates the problems associated with breathing gases at high pressure, and avoids nitrogen narcosis and the bends.



◀ **Deep-Sea Diving** As a diver descends into the ocean, the solubility of air increases. At a depth of 33 feet (1 atm hydrostatic pressure) the solubility of air is twice that of sea level; at 66 feet, the solubility is three times that of sea level.

An interesting medical aside is that a helium-oxygen mixture has been shown to be beneficial to asthma patients. Patients with chronic obstructive pulmonary disease (COPD), who were administered heliox, not only had fewer asthma attacks, but were able to increase their amount of exercise.

A: As a diver ascends, helium gas escapes from the blood in the circulatory system. Because a helium atom is much smaller than a nitrogen molecule, helium can escape much faster than nitrogen, thus avoiding the painful "bends."

EXAMPLE EXERCISE 13.1 Henry's Law

Calculate the solubility of carbon dioxide in water at 0 °C and a pressure of 3.00 atm. The solubility of carbon dioxide is 0.348 g/100 mL water at 0 °C and 1.00 atm.

Solution

The solubility of a gas is proportional to the partial pressure of the gas above the liquid.

$$\text{solubility} \times \text{pressure factor} = \text{solubility}$$

Since pressure increases, solubility increases, and the pressure factor is greater than 1.

$$0.348\text{ g}/100\text{ mL} \times \frac{3.00\ \cancel{\text{atm}}}{1.00\ \cancel{\text{atm}}} = 1.044\text{ g}/100\text{ mL}$$

Practice Exercise

Oxygen is much less soluble in water than carbon dioxide, 0.00412 g/100 mL at 20 °C and 760 mm Hg. Calculate the solubility of oxygen gas in water at 20 °C and a pressure of 1150 mm Hg.

Answer: 0.00623 g/100 mL

Concept Exercise

The elevation of Lake Havasu is 400 feet, and Lake Tahoe is 6200 feet. If the water temperatures are the same, which lake has a lower concentration of oxygen gas?

Answer: See Appendix G, 13.1.

13.2 Liquids in Solution

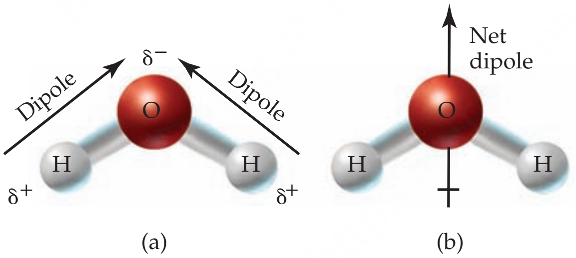
LEARNING OBJECTIVE

- Determine whether a liquid is miscible or immiscible in another liquid by applying the *like dissolves like* rule.

A **solution** is composed of a solute dissolved in a solvent. The **solute** is the lesser quantity, and the **solvent** is the greater quantity. Water is the most common solvent in the laboratory, and every living thing carries out chemical reactions in an aqueous solution.

Dipoles

In Section 12.9, we learned that in water the more electronegative oxygen atom has a partial negative charge (δ^-). The two hydrogen atoms are less electronegative than the oxygen atom and have a partial positive charge (δ^+). A molecule having regions of a positive and a negative charge possesses a **dipole**. There are two dipoles in a water molecule. Together, the two dipoles produce a **net dipole** for the molecule. The symbol for a net dipole is an arrow pointing toward the negative end of the molecule, as shown in Figure 13.1.



(a) (b)

Figure 13.1 The Polar Water Molecule (a) The more electronegative oxygen atom polarizes the O—H bond, which in turn creates two dipoles in a water molecule. (b) The two dipoles produce a net dipole for the entire water molecule.

Polar and Nonpolar Solvents

A liquid composed of polar molecules is called a **polar solvent**. Water molecules are polar, and water is the most common polar solvent. Although there are many exceptions, solvent molecules containing oxygen atoms, such as water, are usually polar. Methanol, CH_3OH, and ethanol, C_2H_5OH, are also examples of polar solvents.

A liquid composed of nonpolar molecules is a **nonpolar solvent**. Nonpolar solvents include hexane, C_6H_{14}; carbon tetrachloride, CCl_4; and ethyl ether, $C_4H_{10}O$. Table 13.3 lists some common polar and nonpolar solvents.

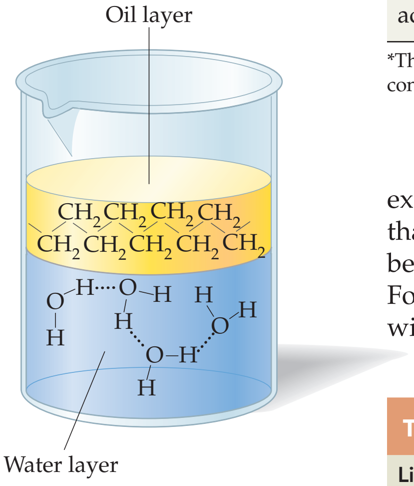
Like Dissolves Like Rule

Polar solvents such as water, H_2O, and ethanol, C_2H_5OH, dissolve in one another. Nonpolar solvents such as hexane, C_6H_{14}, and carbon tetrachloride, CCl_4, also dissolve in one another. If two solvents are both polar, or both nonpolar, they dissolve in one another. This principle is the ***like dissolves like* rule**.

TABLE 13.3 Selected Polar and Nonpolar Solvents

Polar Solvents	Nonpolar Solvents
water, H_2O	hexane, C_6H_{14}
methanol, CH_3OH	heptane, C_7H_{16}
ethanol, C_2H_5OH	toluene, C_7H_8
acetone, C_3H_6O	carbon tetrachloride, CCl_4
methyl ethyl ketone, C_4H_8O	chloroform, $CHCl_3$
formic acid, HCOOH	methylene chloride, CH_2Cl_2
acetic acid, CH_3COOH	ethyl ether, $C_4H_{10}O$*

*The general rule that oxygen-containing solvents are polar has some exceptions. For example, ethyl ether contains oxygen and is a nonpolar solvent.



▲ **Immiscible Liquids** Oil and water are immiscible because oil molecules are *nonpolar* and water molecules are *polar*. A nonpolar solvent (oil) and a polar solvent (water) obey the *like dissolves like* rule.

Two solvents that dissolve completely in one another are said to be **miscible**. For example, water and ethanol are miscible. From the *like dissolves like* rule, it also follows that a polar solvent and a nonpolar solvent are not miscible. These liquids are said to be **immiscible**. Experimentally, mixtures of immiscible liquids separate into layers. For example, oil and water are immiscible and if mixed will separate into two layers with the oil floating on the water. The *like dissolves like* rule is summarized in Table 13.4.

TABLE 13.4 Summary of the *Like Dissolves Like* Rule for Two Liquids

Liquid Solute	Polar Solvent	Nonpolar Solvent
polar	miscible	immiscible
nonpolar	immiscible	miscible

The following example exercise illustrates the use of the *like dissolves like* rule to predict the miscibility of two liquids.

EXAMPLE EXERCISE 13.2 Miscibility Predictions

Predict whether each of the following solvents is miscible or immiscible with water:

(a) methanol, CH_3OH (b) toluene, C_7H_8

Solution

Let's use the simplifying assumption that most solvents containing oxygen are polar. Thus, methanol is polar and toluene is nonpolar. Applying the *like dissolves like* rule gives the following:

(a) CH_3OH is polar and therefore is *miscible* with H_2O.
(b) C_7H_8 is nonpolar and therefore is *immiscible* with H_2O.

Practice Exercise

Predict whether each of the following solvents is miscible or immiscible with water:

(a) methylene chloride, CH_2Cl_2 (b) glycerin, $C_3H_5(OH)_3$

Answers:

(a) immiscible with H_2O (b) miscible with H_2O

Concept Exercise

The *like dissolves like* rule states that two liquids are miscible if what property of the two liquids is alike?

Answer: See Appendix G, 13.2.

▲ **Oil Spill** Petroleum oil is less dense than water; thus, the nonpolar "oil slick" floats on the ocean surface.

Note In our discussion, we have classified two solvents as either miscible or immiscible; that is, two solvents either dissolve or do not dissolve in each other. In practice, two solvents can dissolve partially in one another. For example, a nonpolar solvent can dissolve partially in a polar solvent. Although ethyl ether is generally considered a nonpolar solvent, it dissolves partially in water (~7% soluble).

13.3 Solids in Solution

LEARNING OBJECTIVE

- Determine whether a solid is soluble or insoluble in a liquid by applying the *like dissolves like* rule.

A solid substance such as sugar dissolves in a liquid, such as water, because the sugar and the water are attracted to each other. That is, the solute sugar particles are more strongly attracted to the solvent water molecules than they are to each other. The solute–solvent interaction is strongest when the polarities of the solute and the solvent are similar.

In general, a polar compound is likely to dissolve in a polar solvent such as water (Table 13.5). Sucrose, $C_{12}H_{22}O_{11}$, is ordinary table sugar and contains several oxygen atoms, and is therefore a polar compound. Applying the *like dissolves like* rule, we can correctly predict that sucrose, $C_{12}H_{22}O_{11}$, is soluble in water.

TABLE 13.5 Summary of the *Like Dissolves Like* Rule for a Solid in a Liquid

Solid Solute	Polar Solvent	Nonpolar Solvent
polar	soluble	insoluble
nonpolar	insoluble	soluble
ionic	soluble	insoluble

Similarly, molecules in a nonpolar compound are attracted by molecules in a nonpolar solvent. Thus, nonpolar compounds dissolve in nonpolar solvents. For example, grease dissolves in turpentine. This is because grease is a nonpolar compound and turpentine is a nonpolar solvent. Dissolving grease with turpentine is an illustration of the *like dissolves like* rule. Conversely, we cannot remove grease from our hands using water. Grease is a nonpolar compound and water is a polar solvent, so grease does not dissolve in water.

Because ionic compounds are composed of charged ions, they are similar to polar compounds in that they dissolve in polar solvents. That is, ionic compounds are more soluble in a polar solvent than in a nonpolar solvent. Consider ordinary table salt, NaCl, which dissolves readily in water. It does not dissolve in gasoline, which is a nonpolar solvent. This is because the solute (salt) is ionic and the solvent (gasoline) is nonpolar. It is important to note, however, that many ionic compounds are only slightly soluble in water (refer to the solubility rules in Section 7.9).

The following example exercise illustrates the *like dissolves like* rule for a solid compound in water.

Helpful Hint
Predicting Solubility

The *like dissolves like* rule is usually sufficient to predict solubility. However, some molecules contain both polar and nonpolar bonds. Thus, we must evaluate whether the molecule, overall, is more polar or nonpolar.

Because a cholesterol molecule, $C_{27}H_{46}O$, contains a polar oxygen atom, we might expect it to be polar. In practice, however, cholesterol is nonpolar. It is insoluble in water, but soluble in a nonpolar solvent such as hexane, C_6H_{14}.

We can explain this behavior by noting that $C_{27}H_{46}O$ contains only one polar bond, but many nonpolar bonds. Therefore, the molecule, overall, is more nonpolar than polar. The fact that cholesterol is nonpolar explains why cholesterol is insoluble in the body, except for fat tissue where it deposits.

EXAMPLE EXERCISE 13.3 Solubility Predictions

Predict whether each of the following solid compounds is soluble or insoluble in water.

(a) fructose, $C_6H_{12}O_6$

(b) lithium carbonate, Li_2CO_3

(c) naphthalene, $C_{10}H_8$

Solution

Generally, we can apply the *like dissolves like* rule to determine if a compound is soluble. Since water is a polar solvent, we can predict that water dissolves polar compounds and many ionic compounds.

(a) Fructose has six oxygen atoms and is a polar compound. We can predict that $C_6H_{12}O_6$ is *soluble* in water.
(b) Lithium carbonate contains lithium and carbonate ions; it is therefore an ionic compound. We can predict that Li_2CO_3 is *soluble* in water.
(c) Naphthalene does not contain an oxygen atom and is a nonpolar compound. Thus, $C_{10}H_8$ is *insoluble* in water.

Practice Exercise

Predict whether each of the following solid compounds is soluble or insoluble in water.

(a) anthracene, $C_{14}H_{10}$
(b) cupric sulfate, $CuSO_4$
(c) lactic acid, $HC_3H_5O_3$

Answers:

(a) insoluble
(b) soluble
(c) soluble

Concept Exercise

The *like dissolves like* rule states that a solid is soluble in a liquid if what property is alike?

Answer: See Appendix G, 13.3.

CHEMISTRY CONNECTION Colloids

Q: Why is the beam of light not visible in the solution (left), but observed in the colloid (right)?

A colloid is similar to a solution in that both contain particles in a liquid. Familiar examples of colloids include fat molecules in milk, and protein molecules in blood. The size of the dispersed particle distinguishes a colloid from a solution. A **colloid** contains large particles ranging in size from 1 nm to 100 nm, whereas a solution contains particles smaller than 1 nm. If we shine light on a colloid, the particles scatter light and we can see the light beam. This phenomenon is called the **Tyndall effect**.

Colloids and solutions are both stable, and their particles do not separate spontaneously from the liquid. However, we can separate colloid particles by centrifuging. For example, we can separate a colloid precipitate of AgCl by placing it in a test tube and spinning it in a centrifuge. Although colloid particles can pass through ordinary filter paper, they are too large to pass through a cell membrane, or a synthetic membrane like cellophane.

Kidney dialysis is based on the separation of colloid particles using a synthetic membrane. During kidney dialysis, blood is circulated through an artificial kidney containing a synthetic membrane. Large protein molecules and red blood cells cannot pass through the membrane, but small toxin molecules do pass through the membrane and are removed. After 2–3 hours of dialysis, a patient has blood that is mostly free of small toxic molecules. We can summarize the distinctions between a solution and a colloid in the following ways.

▲ **Tyndall Effect** The Tyndall effect is demonstrated as a beam of light passes through a red colloid (shown right).

CRITERION	SOLUTIONS	COLLOIDS
particle size	1 nm or less (small)	1–100 nm (large)
centrifuge separation	do not separate	separate
filter paper separation	do not separate	do not separate
membrane separation	do not separate	separate
Tyndall effect	light beam not visible	light beam visible

A: The beam of light is not observed as it passes through a solution (left), but the Tyndall effect is visible as light passes through the colloid (right) with large solute particles.

13.4 The Dissolving Process

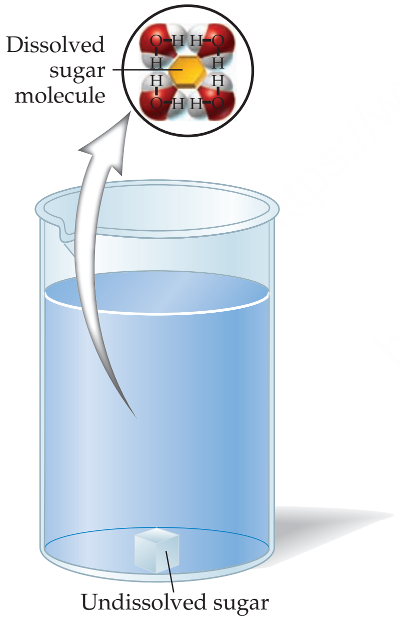
LEARNING OBJECTIVE

- Illustrate how an ionic compound and a molecular compound dissolve in water.

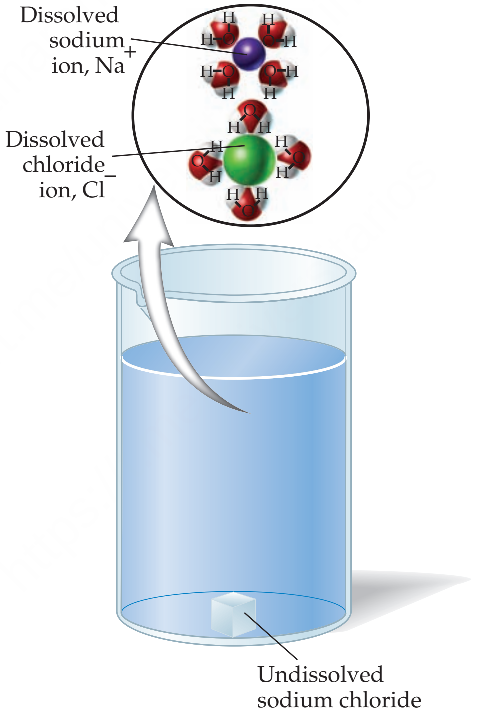
Let's try to visualize the dissolving process. When a soluble crystal is dropped into water, the crystal begins to dissolve. This is because water molecules attack the crystal and begin pulling away the solute. Specifically, water molecules attack the edges and corners of the crystal.

As an example, suppose we drop a cube of table sugar into water. Water molecules are attracted to the polar sugar molecules, $C_{12}H_{22}O_{11}$, and pull molecules of sugar into the solution. Several water molecules surround each sugar molecule in the solution. The sugar molecules, which are held within a cluster of water molecules, are said to be in a **solvent cage**. The number of water molecules in the solvent cage varies, depending on the size and concentration of the solute (Figure 13.2).

As another example, consider what happens when crystals of table salt dissolve in water. According to the *like dissolves like* rule, polar water molecules are attracted to each ionic salt crystal. Once again, water molecules attack the edges of each crystal and begin pulling away the solute. For an ionic compound such as table salt, NaCl, the water molecules pull away positive and negative ions during the dissolving process. The partially negative oxygen atom in a water molecule is attracted to the positive sodium ion, Na^+. The partially positive hydrogen atoms in the water molecule are attracted to the negative chloride ion, Cl^-. Figure 13.3 illustrates this process.



▲ **Figure 13.2 Table Sugar Dissolving in Water** Water molecules attack a sugar cube along the edges and at the corners. Each sugar molecule is pulled into the solution and surrounded by several water molecules.



▲ **Figure 13.3 Table Salt Dissolving in Water** Note the orientation of the water molecules as they surround each ion. The Na^+ is attracted to the oxygen atom in a water molecule, whereas the Cl^- is attracted to the hydrogen atoms.

EXAMPLE EXERCISE 13.4 The Dissolving Process

When potassium iodide, KI, dissolves in water, why does the oxygen atom of the water molecule attack the potassium ion, K^+?

Solution

Refer to Figure 13.1 and notice the oxygen atom bears a partial negative charge. Thus, the oxygen atom in a water molecule is attracted to the positive potassium ion.

Practice Exercise

When potassium iodide, KI, dissolves in water, why do the hydrogen atoms of the water molecule attack the iodide ion, I^-?

Answer:

Refer to Figure 13.1 and notice the hydrogen atoms bear a partial positive charge. Thus, the hydrogen atoms in a water molecule are attracted to the negative iodide ion.

Concept Exercise

When an ionic compound dissolves in water, is the oxygen atom in a water molecule attracted to the cation or anion in the compound?

Answer: See Appendix G, 13.4.

LEARNING OBJECTIVE

- Describe the effect of temperature, stirring, and particle size on the rate at which a solid compound dissolves in water.

13.5 Rate of Dissolving

The rate at which a solid compound such as table sugar dissolves in a solution depends on three factors. We can increase the rate of dissolving by any one of the following:

1. heating the solution
2. stirring the solution
3. grinding the solute

By heating and stirring the solution or by grinding the solute, we increase the rate at which solvent molecules attack the solute. *Heating the solution* increases the kinetic energy of the solution, and the solvent molecules move faster. In an aqueous solution, water molecules attack the solute more frequently. Solute molecules are pulled into the solution faster, thus increasing the rate of dissolving. *Stirring the solution* increases the interaction between water molecules and the solute. Because the solute and solvent interact more often, the rate of dissolving is faster.

Grinding the solute into smaller particles creates more surface area. As the solute crystals become smaller, the total surface area increases. We know that solvent molecules attack the surface of crystals along the edges. When smaller crystals are created, more edges are exposed to attack by water molecules. The water molecules attack the solute more frequently, thus increasing the rate of dissolving (Figure 13.4).

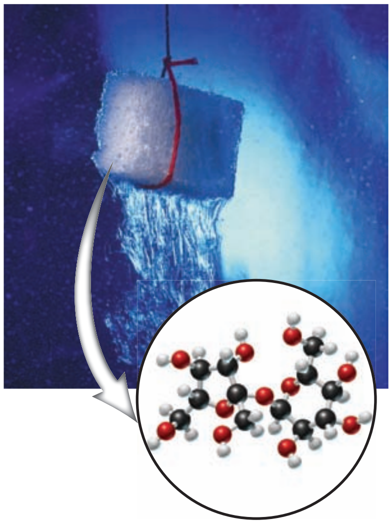


Figure 13.4 Rate of Dissolving Table Sugar A cube of sugar suspended in water at 20 °C dissolves slowly. An equal amount of powdered sugar stirred in water at 100 °C dissolves immediately.

EXAMPLE EXERCISE 13.5 Rate of Dissolving

Why does heating a solution increase the rate of dissolving of a solid solute in water?

Solution

Heating increases the motion and energy of solvent molecules; thus, water molecules attack the solid solute at a faster rate.

Practice Exercise

Why does stirring a solution increase the rate of dissolving of a solid solute in water?

Answer:

Stirring increases the rate at which solvent molecules contact the solute, thus increasing the rate of solvent pulling solute into solution.

Concept Exercise

Why does grinding solid crystals increase the rate of dissolving of a solid solute in water?

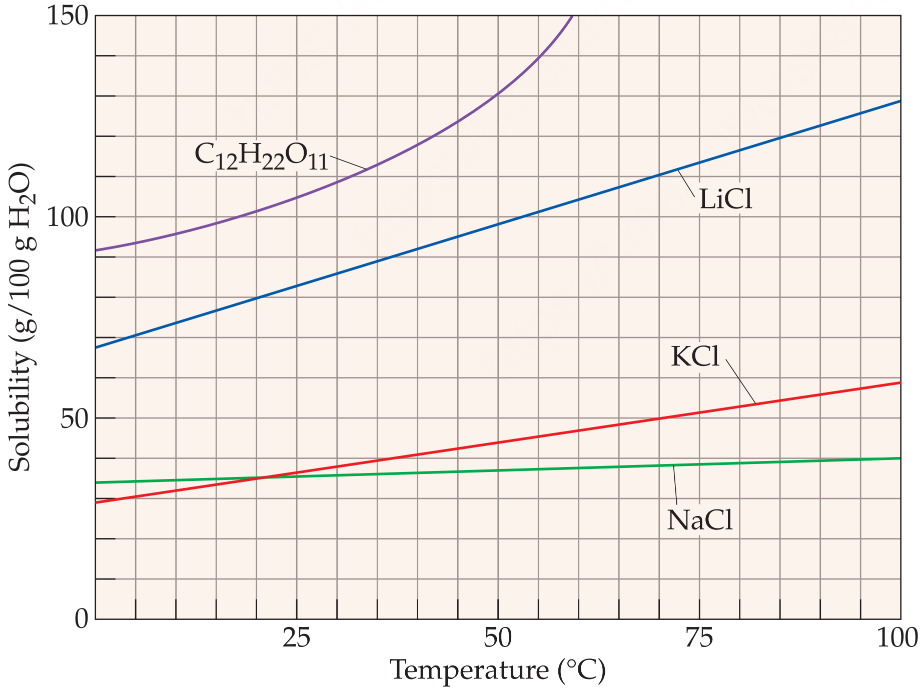
Answer: See Appendix G, 13.5.

13.6 Solubility and Temperature

LEARNING OBJECTIVE

- Interpret a graph that shows temperature versus solubility of a solid compound in water.

If we heat a solution, the interaction between the solute particles and the solvent increases. In general, the solubility of a solid compound becomes greater as the temperature increases. We will define the **solubility** of a compound as the maximum amount of solute that can be dissolved in 100 g of water at a given temperature. Figure 13.5 graphs the solubility of various compounds in 100 grams of water at different temperatures.



◀ **Figure 13.5 Solubility of Solid Compounds in Water** Although there are a few exceptions, most solid compounds become more soluble as the temperature increases.

Adding sugar to a hot beverage is an interesting example of the effect of temperature on solubility. Let's suppose you add sugar to a cup of hot tea and then allow the tea to cool. The solubility of the sugar decreases as the tea cools. If you stir the cold tea, you may notice crystals of sugar at the bottom of the cup that have crystallized from solution at the cooler temperature.

Figure 13.5 illustrates that the effect of temperature on solubility varies with the compound. The solubility of table salt, NaCl, in water is only slightly affected by temperature. Its solubility is about 35 g of NaCl per 100 g of water at 20 °C and increases to 40 g/100 g water at 100 °C.

Note the solubility of table sugar, $C_{12}H_{22}O_{11}$, is greatly affected by an increase in temperature. At 20 °C the solubility of sugar is about 100 g/100 g water. At 55 °C the solubility is 140 g/100 g water. If you dissolve 140 g of table sugar in 100 g of hot water at 55 °C, and allow it to cool to 20 °C, you would see sugar crystals, just as in the previous cold tea example.

EXAMPLE EXERCISE 13.6 Determining Solubility from a Graph

Determine the solubility of each of the following solid compounds at 50 °C as shown in Figure 13.5.

(a) NaCl
(b) KCl
(c) LiCl
(d) $C_{12}H_{22}O_{11}$

Solution

From Figure 13.5, let's find the point at which the solubility of the compound intersects 50 °C.

(a) The solubility of NaCl at 50 °C is about 38 g/100 g water.
(b) The solubility of KCl at 50 °C is about 45 g/100 g water.
(c) The solubility of LiCl at 50 °C is about 98 g/100 g water.
(d) The solubility of $C_{12}H_{22}O_{11}$ at 50 °C is about 130 g/100 g water.

Practice Exercise

Refer to the solubility behavior shown in Figure 13.5 and determine the minimum temperature required to obtain the following solutions.

(a) 35 g NaCl per 100 g of water
(b) 45 g KCl per 100 g of water
(c) 120 g LiCl per 100 g of water
(d) 109 g $C_{12}H_{22}O_{11}$ per 100 g of water

Answers:

(a) 20 °C
(b) 55 °C
(c) 85 °C
(d) 30 °C

Concept Exercise

Given a hot cup of coffee, and a warm cup of coffee, which can dissolve more sugar?

Answer: See Appendix G, 13.6.

13.7 Unsaturated, Saturated, and Supersaturated Solutions

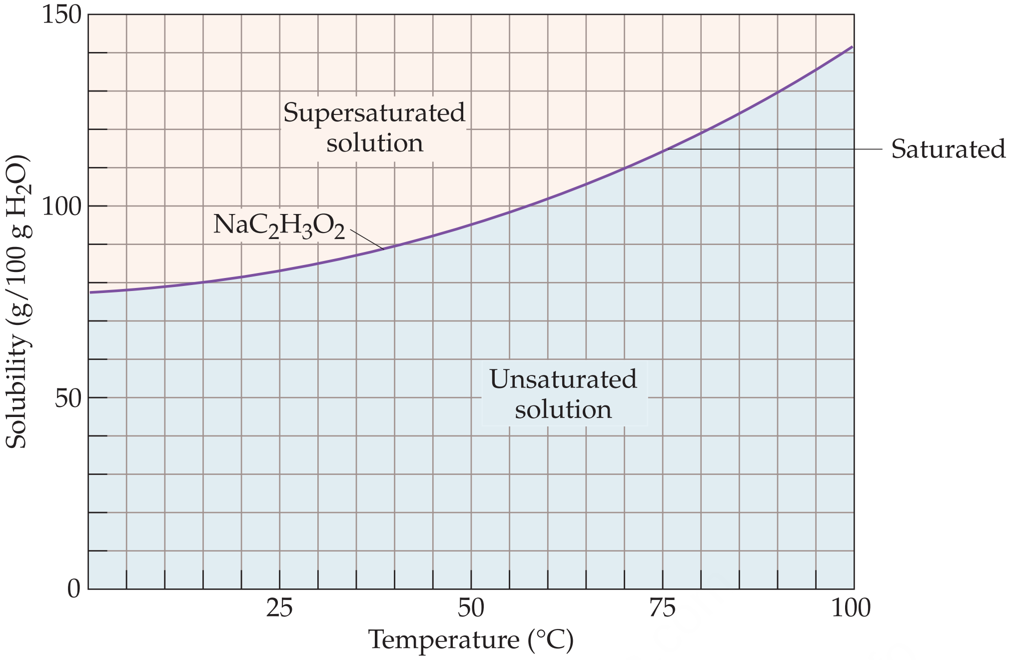
LEARNING OBJECTIVES

- Interpret a graph of temperature versus solubility and determine whether a solution is saturated, unsaturated, or supersaturated.
- Distinguish among solutions that are saturated, unsaturated, or supersaturated.

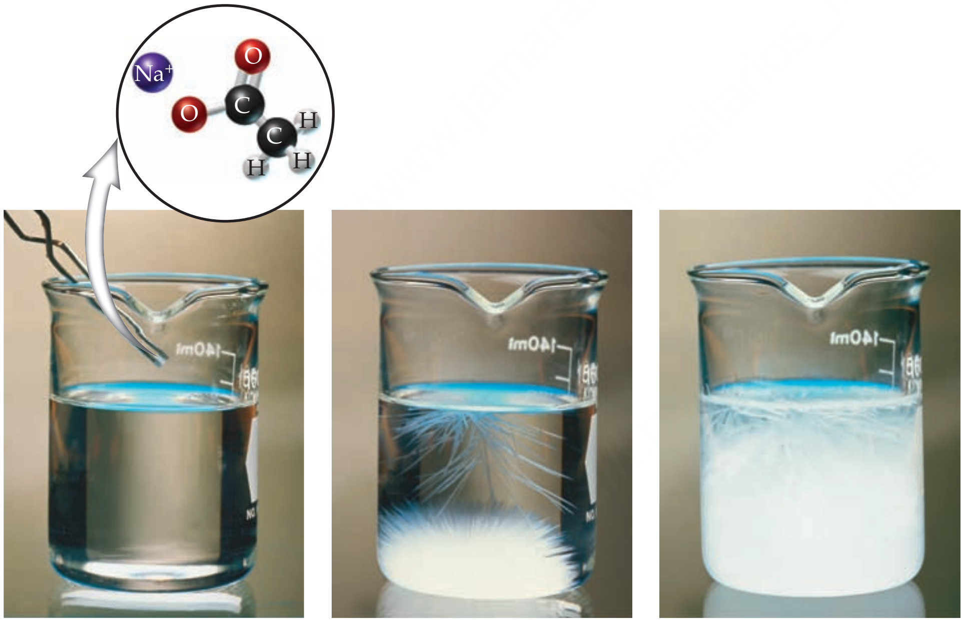
In Section 13.6, we noted that the solubility of a compound usually increases as the temperature increases. Consider the solubility of sodium acetate, $NaC_2H_3O_2$. The solubility of sodium acetate is 100 g of $NaC_2H_3O_2$ per 100 g of water at 55 °C. This is the maximum amount of solute that dissolves at 55 °C.

A solution containing the maximum amount of solute at a given temperature is a **saturated solution**. At higher temperatures, the solubility increases. A saturated solution containing 100 g of $NaC_2H_3O_2$ per 100 g of water at 55 °C is not saturated at 75 °C. Since more solute can be dissolved in the solution at 75 °C, it is an **unsaturated solution**. Figure 13.6 shows the solubility of $NaC_2H_3O_2$ at different temperatures.

A solution that contains more solute than ordinarily dissolves at a given temperature is a **supersaturated solution**. For example, if a saturated solution of $NaC_2H_3O_2$ cools from to 55 °C to 20 °C slowly, all the solute will remain in solution. At 20 °C the solubility is less, yet the excess solute remains in solution. However, supersaturated solutions are unstable. In fact, one tiny crystal can cause massive crystallization, as shown in Figure 13.7.



◀ **Figure 13.6 Solubility of Sodium Acetate in Water** The graph represents a saturated solution of sodium acetate, $NaC_2H_3O_2$, at various temperatures.



(a) (b) (c)

◀ **Figure 13.7 Illustration of Supersaturation** (a) A single crystal of sodium acetate, $NaC_2H_3O_2$, is dropped into a supersaturated solution. (b) The small crystal causes crystallization, and eventually, (c) the solution forms a solid.

▲ **Rock Candy** Rock candy is made by putting a single sugar crystal, held by a thread, into a supersaturated sugar solution.

EXAMPLE EXERCISE 13.7 Determining Saturation from a Graph

A sodium acetate solution contains 110 g of $NaC_2H_3O_2$ per 100 g of water. Refer to Figure 13.6 and determine whether the solution is unsaturated, saturated, or supersaturated at each of the following temperatures.

(a) 50 °C (b) 70 °C (c) 90 °C

Solution

(a) At 50 °C the solubility of $NaC_2H_3O_2$ is about 97 g/100 g water. Since the solution contains more solute, 110 g/100 g water, the solution is *supersaturated.*

(b) At 70 °C the solubility is about 110 g/100 g water. Since the solution contains the same amount of solute, 110 g/100 g water, the solution is *saturated.*

(c) At 90 °C the solubility is about 130 g/100 g water. Since the solution has only 110 g/100 g water, the solution is *unsaturated.*

Practice Exercise

A sodium acetate solution contains 80 g of $NaC_2H_3O_2$ per 100 g water. Refer to Figure 13.6 and determine whether the solution is unsaturated, saturated, or supersaturated at each of the following temperatures.

(a) 0 °C (b) 15 °C (c) 45 °C

Answers:

(a) supersaturated (b) saturated (c) unsaturated

Concept Exercise

How is it possible to exceed the saturation of a solution and produce a supersaturated solution?

Answer: See Appendix G, 13.7.

LEARNING OBJECTIVES

- Calculate the mass percent concentration of a solution.
- Write three pairs of unit factors given the mass percent concentration of a solution.
- Solve problems that involve the mass of solute, mass of solvent, and mass percent concentration of a solution.

13.8 Mass/Mass Percent Concentration

The concentration of a solution indicates how much solute is dissolved in a solution. Chemists often refer to a solution as *dilute* or *concentrated* to describe the approximate concentration. A concentrated solution has a large amount of dissolved solute, whereas a dilute solution does not.

Chemists may describe the precise concentration of a solution by comparing the mass of solute to the mass of solution. The **mass/mass percent** (symbol **m/m %**) concentration is the mass of solute dissolved in 100 g of solution. This concentration expression can be written as follows:

$$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% = \text{m/m}\ \%$$

Since the total mass of the solution equals the mass of the solute plus solvent, we can also write the mass/mass percent concentration expression in the following form:

$$\frac{\text{g solute}}{\text{g solute} + \text{g solvent}} \times 100\% = \text{m/m}\ \%$$

If a student prepares a solution from 5.00 g of NaCl dissolved in 95.00 g of water, what is the mass/mass percent concentration? Since NaCl is the solute and water is the solvent, we can calculate the mass/mass percent concentration as follows:

$$\frac{5.00\ \text{g NaCl}}{5.00\ \text{g NaCl} + 95.00\ \text{g H}_2\text{O}} \times 100\% = \text{m/m}\ \%$$

$$\frac{5.00\ \text{g NaCl}}{100.00\ \text{g solution}} \times 100\% = 5.00\%$$

Writing Mass/Mass Percent Concentration Unit Factors

The unit analysis method of problem solving involves three steps: unknown units, relevant given value, and application of a unit factor. We can solve mass/mass percent calculations by using the solution concentration as a unit factor. For example, the aforementioned 5.00% solution contains 5.00 g of NaCl in each 100 g of solution. We can write this pair of unit factors as

$$\frac{5.00\ \text{g NaCl}}{100\ \text{g solution}} \quad \text{and} \quad \frac{100\ \text{g solution}}{5.00\ \text{g NaCl}}$$

Since the solution is composed of solute and solvent, we can write a second pair of unit factors. The solution contains 5.00 g of NaCl solute in 95.00 g of water solvent.

$$\frac{5.00 \text{ g NaCl}}{95.00 \text{ g water}} \quad \text{and} \quad \frac{95.00 \text{ g water}}{5.00 \text{ g NaCl}}$$

Furthermore, we can express the ratio of the mass of solvent to the mass of solution in a third pair of unit factors. The solution contains 95.00 g of solvent in 100 g of solution.

$$\frac{95.00 \text{ g water}}{100 \text{ g solution}} \quad \text{and} \quad \frac{100 \text{ g solution}}{95.00 \text{ g water}}$$

Solving Mass/Mass Percent Concentration Problems

We can now apply unit factors derived from the mass/mass percent concentration to solution calculations. For example, if a glucose solution is 5.00%, what is the mass of solution that contains 1.25 g of glucose sugar?

Let's use the unit analysis method of problem solving. Step 1: The unknown quantity is g solution. Step 2: The relevant given value is 1.25 g glucose.

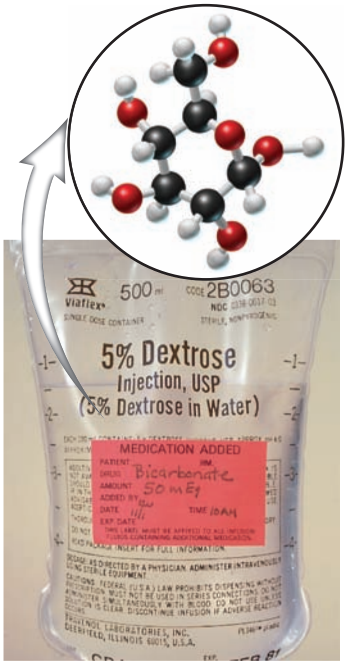
$$1.25 \text{ g glucose} \times \frac{\text{unit}}{\text{factor}} = \text{g solution}$$

Step 3: The solution concentration is 5.00%, so there is 5.00 g of solute in 100 g of solution. We can write the unit factors as

$$\frac{5.00 \text{ g glucose}}{100 \text{ g solution}} \quad \text{and} \quad \frac{100 \text{ g solution}}{5.00 \text{ g glucose}}$$

We must select the second unit factor to cancel units.

$$1.25 \cancel{\text{ g glucose}} \times \frac{100 \text{ g solution}}{5.00 \cancel{\text{ g glucose}}} = 25.0 \text{ g solution}$$



▲ **5% Dextrose** A 5% dextrose solution is often given to patients before surgery. Dextrose is the common name for glucose sugar, $C_6H_{12}O_6$.

EXAMPLE EXERCISE 13.8 Mass/Mass Percent Concentration

Intravenous dextrose injections are given to restore sugar levels in patients. What is the mass of sugar dissolved in 25.0 g of a 10.0% dextrose solution?

We will solve this problem using the unit analysis method of problem solving; that is,

given value $\times$ $\dfrac{\textbf{unit}}{\textbf{factor}}$ = **unit in answer**

Solution

STEP 1: The unit asked for in the answer is g solution.

STEP 2: The given value is 25.0 g solution.

STEP 3: We apply a conversion factor to cancel units. Because the solution concentration is 10.0%, there is 10.0 g of sugar in 100 g of solution.

$$\frac{10.0 \text{ g sugar}}{100.0 \text{ g solution}} \quad \text{and} \quad \frac{100.0 \text{ g solution}}{10.0 \text{ g sugar}}$$

In this example, we select the first unit factor to cancel the given units.

$$25.0 \cancel{\text{ g solution}} \times \frac{10.0 \text{ g sugar}}{100.0 \cancel{\text{ g solution}}} = 2.50 \text{ g sugar}$$

We can summarize the conversion as follows.

25.0 g solution $\frac{\textbf{10.0 g sugar}}{\textbf{100.0 g solution}}$ **2.50 g sugar**

Practice Exercise

Fructose is referred to as fruit sugar and accounts for the sweet taste of fruit. What is the mass of sugar dissolved in 10.0 g of a 5.00% fructose solution?

Answer: 0.500 g sugar

Concept Exercise

Explain how dextrose and fructose can have the same molecular formula, $C_6H_{12}O_6$, but different physical and chemical properties?

Answer: See Appendix G, 13.8.

EXAMPLE EXERCISE 13.9 Mass/Mass Percent Concentration

Intravenous saline injections are given to restore the mineral balance in trauma patients. What is the mass of water required to dissolve 1.50 g of NaCl for a 0.90% normal IV saline solution?

We will solve the problem using the unit analysis method of problem solving; that is,

given value $\frac{\textbf{unit}}{\textbf{factor}}$ **unit in answer**

Solution

STEP 1: The unit asked for in the answer is g water.

STEP 2: The given value is 1.50 g NaCl.

STEP 3: We apply a conversion factor to cancel units. Since the solution concentration is 0.90%, there is 0.90 g of solute in 100 g of solution. Therefore, there is 0.90 g NaCl for every 99.10 g water. We can write the unit factors

$$\frac{0.90 \text{ g NaCl}}{99.10 \text{ g water}} \quad \text{and} \quad \frac{99.10 \text{ g water}}{0.90 \text{ g NaCl}}$$

In this example, we select the second unit factor to cancel the given units.

$$1.50 \cancel{\text{ g NaCl}} \times \frac{99.10 \text{ g water}}{0.90 \cancel{\text{ g NaCl}}} = 170 \text{ g water}$$

We can summarize the conversion as follows:

1.50 g NaCl $\frac{\textbf{99.10 g water}}{\textbf{0.90 g NaCl}}$ **170 g water**

Practice Exercise

A 7.50% potassium chloride solution is prepared by dissolving enough of the salt to give 100.0 g of solution. What is the mass of water required?

Answer: 92.5 g water

Concept Exercise

Given the mass percent concentration of a solution, how many unit factors can we write?

Answer: See Appendix G, 13.9.

▲ **Blood Glucose Meter** Diabetes patients use a self-test meter for measuring glucose levels in blood. A normal fasting blood glucose reading is 70–110 mg/dL.

13.9 Molar Concentration

LEARNING OBJECTIVES

- Calculate the molar concentration of a solution.
- Write a pair of unit factors given the molar concentration of a solution.
- Solve problems that involve a mass of solute, volume of solution, and the molar concentration of a solution.

We can describe the precise concentration of a solution as mass/mass percent concentration. However, most of the time we express the precise concentration of solutions in terms of molar concentration. The molar concentration, or **molarity** (symbol ***M***), relates the amount of solute to a given volume of solution. Specifically, molarity is the number of moles of a solute dissolved in one liter of a solution. We can express the molarity as follows:

$$\frac{\text{moles solute}}{\text{liter solution}} = M$$

As an example, consider a household drain cleaner, which is a solution of sodium hydroxide, NaOH. If the drain cleaner contains 240.0 g of NaOH dissolved in 1.00 L of solution, what is the molarity? We can calculate the molarity of NaOH (40.00 g/mol) as follows:

$$\frac{240.0\ \cancel{\text{g NaOH}}}{1.00\ \text{L solution}} \times \frac{1\ \text{mol NaOH}}{40.00\ \cancel{\text{g NaOH}}} = \frac{\text{mol NaOH}}{\text{L solution}} = 6.00\ M\ \text{NaOH}$$

Notice that we started the calculation with a ratio of two units (g/L) to obtain an answer that is a ratio of two units (mol/L).

Writing Molar Concentration Unit Factors

The unit analysis method of problem solving involves the application of a unit factor. We can solve molarity calculations by using the solution concentration as a unit factor. For example, a 6.00 *M* solution of NaOH contains 6.00 mol of solute in each liter of solution. We can write this pair of unit factors as

$$\frac{6.00\ \text{mol NaOH}}{1\ \text{L solution}} \quad \text{and} \quad \frac{1\ \text{L solution}}{6.00\ \text{mol NaOH}}$$

In the laboratory, the volume of a solution is usually measured in milliliters. Since there are 1000 mL in a liter, we can substitute 1000 mL for 1 L in the unit factor.

$$\frac{6.00\ \text{mol NaOH}}{1000\ \text{mL solution}} \quad \text{and} \quad \frac{1000\ \text{mL solution}}{6.00\ \text{mol NaOH}}$$

Solving Molar Concentration Problems

We can now apply unit factors derived from the molar concentration to solution calculations. Consider a solution prepared by dissolving potassium dichromate in water. Let's calculate the mass of solute dissolved in 250.0 mL of 0.100 *M* $K_2Cr_2O_7$ solution.

We begin with the unit analysis method of problem solving. Step 1: The unknown quantity is g $K_2Cr_2O_7$. Step 2: The relevant given value is 250.0 mL solution.

$$250.0\ \text{mL solution} \times \frac{\text{unit}}{\text{factor}} = \text{g}\ K_2Cr_2O_7$$

Since the solution concentration is 0.100 *M*, there is 0.100 mol of $K_2Cr_2O_7$ solute in each liter of solution. We can write two unit factors.

$$\frac{0.100\ \text{mol}\ K_2Cr_2O_7}{1000\ \text{mL solution}} \quad \text{and} \quad \frac{1000\ \text{mL solution}}{0.100\ \text{mol}\ K_2Cr_2O_7}$$

We choose the first unit factor so as to cancel units properly. In addition, we must use the molar mass of $K_2Cr_2O_7$ (294.20 g/mol) to convert from moles to grams.

$$250.0\ \cancel{\text{mL solution}} \times \frac{0.100\ \cancel{\text{mol}\ K_2Cr_2O_7}}{1000\ \cancel{\text{mL solution}}} \times \frac{294.20\ \text{g}\ K_2Cr_2O_7}{1\ \cancel{\text{mol}\ K_2Cr_2O_7}} = 7.36\ \text{g}\ K_2Cr_2O_7$$

▲ **Potassium Dichromate** Orange crystals of potassium dichromate dissolve in water to give, $K_2Cr_2O_7(aq)$.

EXAMPLE EXERCISE 13.10 Molar Concentration

How many milliliters of 12.0 M hydrochloric acid contain 7.30 g of HCl solute (36.46 g/mol)?

In molarity calculations, we relate the given value to moles and then convert to the unit asked for in the answer.

given value $\quad \frac{\text{unit}}{\text{factor}} \quad \frac{\text{unit}}{\text{factor}} \quad$ unit in answer

Solution

STEP 1: The unit asked for in the answer is mL acid.

STEP 2: The relevant given value is 7.30 g HCl.

STEP 3: The molar mass of HCl (36.46 g/mol) and the molarity of the solution (12.0 mol/1000 mL) provide the unit factors. We select unit factors so as to cancel units.

$$7.30\ \text{g}\ \cancel{\text{HCl}} \times \frac{1\ \cancel{\text{mol HCl}}}{36.46\ \text{g}\ \cancel{\text{HCl}}} \times \frac{1000\ \text{mL acid}}{12.0\ \cancel{\text{mol HCl}}} = 16.7\ \text{mL acid}$$

We can summarize the two conversions as follows:

7.30 g HCl $\quad \frac{\text{molar}}{\text{mass}} \quad \frac{\text{molar}}{\text{concentration}} \quad$ 16.7 mL acid

Practice Exercise

What volume of 6.00 M hydrochloric acid contains 10.0 g of HCl solute (36.46 g/mol)?

Answer: 45.7 mL acid

Concept Exercise

Given the molar concentration of a solution, how many unit factors can we write?

Answer: See Appendix G, 13.10.

LEARNING OBJECTIVE

- Solve problems that involve a solution undergoing dilution.

13.10 Dilution of a Solution

In your chemistry laboratory, you may have noticed bottles labeled 0.1 M HCl and 0.1 M NaOH. These solutions, whose concentrations are only approximate, are often prepared by diluting a more concentrated solution, called a stock solution. For example, a stock solution of 6 M NaOH may be available to prepare a 0.1 M NaOH solution for an experiment.

When calculating the concentration of a diluted solution, we obey the following principle: *the amount of solute does not change during dilution.* Stated another way, the moles of solute in the concentrated solution is equal to the moles of solute in the diluted solution. Suppose we want to prepare 5.00 L of 0.1 M NaOH using a 6 M NaOH stock solution. First, we find the moles of NaOH in the diluted solution using 0.1 M NaOH as a unit factor.

$$5.00\ \cancel{\text{L solution}} \times \frac{0.1\ \text{mol NaOH}}{1\ \cancel{\text{L solution}}} = 0.5\ \text{mol NaOH}$$

Since we know the moles of solute remain constant during dilution, the amount of NaOH does not change. That is, we start with—and end up with—the same amount of NaOH. The moles of solute we start with in the concentrated solution is 0.5 mol NaOH.

Using 6 M NaOH as a unit factor, we can calculate the volume of 6 M stock solution as follows:

$$0.5\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{L solution}}{6\ \cancel{\text{mol NaOH}}} = 0.08\ \text{L NaOH}$$

In this example, we demonstrated how to prepare 5.00 L of 0.1 M NaOH from a more concentrated stock solution. That is, we simply dilute 0.08 L (~80 mL) of 6 M NaOH stock solution to a total volume of 5.00 L with distilled water.

Alternatively, we can solve this problem using a simple algebraic equation. The product of molarity (mol/L) and volume (L) gives units of mole. Since *the moles of NaOH in the concentrated and the dilute solutions are equal,* their products are also equal.

$$M_1 \times V_1 = M_2 \times V_2$$

where M_1 and V_1 represent the molarity and the volume of the concentrated solution before dilution, and M_2 and V_2 are the molarity and volume after dilution. In the preceding problem, we found the volume of a 6 M NaOH stock solution that must be diluted to give 5.00 L of 0.1 M NaOH. We can also find the volume using the previous equation. After substituting the given values into the equation, we have

$$6\ M \times V_1 = 0.1\ M \times 5.00\ \text{L}$$

Solving for V_1,

$$\frac{0.1\ M \times 5.00\ L}{6\ M} = V_1 = 0.08\ \text{L NaOH}$$

The following example exercise provides additional practice in solving dilution problems involving the concentration of a solution.

▲ Sodium Hydroxide, NaOH Drain cleaner found in a home contains the same chemical as bottles of sodium hydroxide found in a chemistry laboratory.

EXAMPLE EXERCISE 13.11 Dilution of a Solution

Concentrated hydrochloric acid is available commercially as a 12 M solution. What is the molarity of an HCl solution prepared by diluting 50.0 mL of concentrated acid with distilled water to give a total volume of 2.50 L?

Conceptual Solution

Let's find the moles of concentrated HCl before dilution.

$$50.0\ \cancel{\text{mL solutlion}} \times \frac{12\ \text{mol HCl}}{1000\ \cancel{\text{mL solution}}} = 0.60\ \text{mol HCl}$$

We know that the moles of HCl do not change during dilution; the moles of diluted HCl solution must be the same as the moles of concentrated HCl. Since the volume of diluted HCl is 2.50 L, we can calculate the diluted concentration.

$$\frac{0.60\ \text{mol HCl}}{2.50\ \text{L solution}} = \frac{0.24\ \text{mol HCl}}{1\ \text{L solution}} = 0.24\ M\ \text{HCl}$$

Algebraic Solution

Alternatively, we can solve this problem using the equation

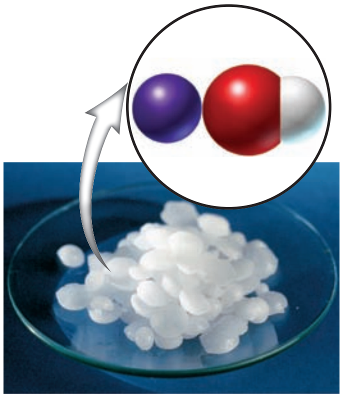
$$M_1 \times V_1 = M_2 \times V_2$$

Substituting, we have 12 $M \times 50.0$ mL $= M_2 \times 2.50$ L
We must use the same units for volume; for example, we can convert 2.50 L to mL, which equals 2500 mL. Solving for M_2, we have

$$\frac{12\ M \times 50.0\ \cancel{\text{mL}}}{2500\ \cancel{\text{mL}}} = 0.24\ M\ \text{HCl}$$

Practice Exercise

Battery acid is 18 M sulfuric acid. What volume of battery acid must be diluted with distilled water to prepare 1.00 L of 0.50 M H_2SO_4?



▲ Sodium Hydroxide, NaOH Chemistry stockrooms prepare solutions of 6 M NaOH by weighing out sodium hydroxide and dissolving the pellets in a given volume of distilled water.

Answer: 28 mL H_2SO_4 (diluted to 1.00 L with distilled H_2O)

Concept Exercise

If equal volumes of 6 *M* sulfuric acid and distilled water are added together, what is the concentration of the diluted acid?

Answer: See Appendix G, 13.11.

LEARNING OBJECTIVE

- Solve problems that involve a balanced chemical equation and the molar concentration of a solution.

13.11 Solution Stoichiometry

In Chapter 9, we performed mole calculations involving chemical equations. The relationship between quantities in a balanced equation is called *stoichiometry*. Now, let's apply stoichiometry to chemical reactions involving the molar concentration of a solution. Suppose we are given 4.24 g of sodium carbonate, Na_2CO_3, which react with 0.150 *M* hydrochloric acid, HCl. The balanced equation for the reaction is

$$Na_2CO_3(s) + 2\,HCl(aq) \rightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

Let's calculate the volume of 0.150 *M* HCl that reacts with 4.24 g of Na_2CO_3. We can solve this problem in three steps. In the first step, we calculate the moles of Na_2CO_3. The molar mass of Na_2CO_3 is 105.99 g/mol. Therefore,

$$4.24\ \text{g}\ \cancel{Na_2CO_3} \times \frac{1\ \text{mol}\ Na_2CO_3}{105.99\ \text{g}\ \cancel{Na_2CO_3}} = 0.0400\ \text{mol}\ Na_2CO_3$$

In the second step, we convert moles of Na_2CO_3 to moles of HCl by applying the coefficients from the balanced equation. The relationship between the two substances is 1 mol Na_2CO_3 = 2 mol HCl. The unit factor conversion is

$$0.0400\ \cancel{\text{mol}\ Na_2CO_3} \times \frac{2\ \text{mol HCl}}{1\ \cancel{\text{mol}\ Na_2CO_3}} = 0.0800\ \text{mol HCl}$$

In the third step, we use the molar concentration of the hydrochloric acid solution as a unit factor to obtain the volume of HCl that reacts with Na_2CO_3. In this example, 0.150 *M* HCl gives two unit factors: 0.150 mol HCl/1 L solution, and the reciprocal, 1 L solution/0.150 mol HCl.

$$0.0800\ \cancel{\text{mol HCl}} \times \frac{1\ \text{L solution}}{0.150\ \cancel{\text{mol HCl}}} = 0.533\ \text{L solution}$$

It is also possible to solve this stoichiometry problem in one continuous calculation. We can show the complete solution to the problem as

$$4.24\ \text{g}\ \cancel{Na_2CO_3} \times \frac{1\ \cancel{\text{mol}\ Na_2CO_3}}{105.99\ \text{g}\ \cancel{Na_2CO_3}} \times \frac{2\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{mol}\ Na_2CO_3}} \times \frac{1\ \text{L solution}}{0.150\ \cancel{\text{mol HCl}}}$$
$$= 0.533\ \text{L solution}$$

The following example exercise provides additional practice in solving calculations involving the concentration of a solution.

EXAMPLE EXERCISE 13.12 Solution Stoichiometry Problem

Given that 37.5 mL of 0.100 *M* aluminum bromide solution reacts with a silver nitrate solution, what is the mass of AgBr (187.77 g/mol) produced? The balanced equation is

$$AlBr_3(aq) + 3\ AgNO_3(aq) \rightarrow 3\ AgBr(s) + Al(NO_3)_3(aq)$$

Solution

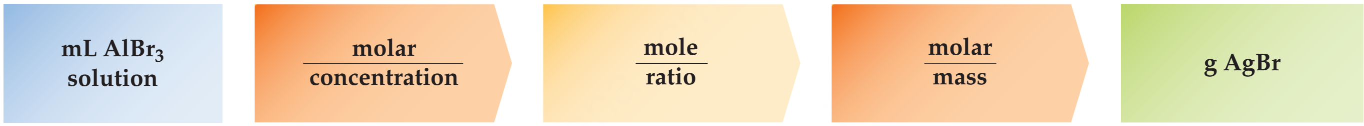
STEP 1: The unit asked for in the answer is g AgBr.

STEP 2: The given value is 37.5 mL solution.

STEP 3: We apply unit conversion factors to cancel units. In this example, we can use molar concentration (0.100 mol $AlBr_3$/1000 mL solution) and molar mass (187.77 g/mol) as unit factors. From the balanced equation, we see that 1 mol $AlBr_3$ = 3 mol AgBr; therefore,

$$37.5\ \cancel{\text{mL solution}} \times \frac{0.100\ \cancel{\text{mol AlBr}_3}}{1000\ \cancel{\text{mL solution}}} \times \frac{3\ \cancel{\text{mol AgBr}}}{1\ \cancel{\text{mol AlBr}_3}} \times \frac{187.77\ \text{g AgBr}}{1\ \cancel{\text{mol AgBr}}} = 2.11\ \text{g AgBr}$$

We can summarize the three conversions as follows:



Practice Exercise

Given that 27.5 mL of 0.210 *M* lithium iodide solution reacts with 0.133 *M* lead(II) nitrate solution, what volume of $Pb(NO_3)_2$ solution is required for complete precipitation?

$$Pb(NO_3)_2(aq) + 2\ LiI(aq) \rightarrow PbI_2(s) + 2\ LiNO_3(aq)$$

Answer: 21.7 mL $Pb(NO_3)_2$

Concept Exercise

Before applying unit analysis and solving a solution stoichiometry problem, what must always be done first?

Answer: See Appendix G, 13.12.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
13.1 Gases in Solution The most common solutions are formed from a solid, a liquid, or a gas dissolved in water. If the solute is a gas, its solubility is affected by temperature and pressure. Raising the temperature of the solution decreases the solubility of the dissolved gas. On the other hand, raising the partial pressure of the gas above the solution increases the amount of dissolved gas in the solution. In fact, **Henry's law** states that the solubility of a gas in a liquid is proportional to the partial pressure of the gas above the liquid.	• Describe the effect of temperature and pressure on the solubility of a gas in a liquid *Related Exercises 1–2* • Apply Henry's law to the solubility of a gas in a liquid. *Related Exercises 3–6*
13.2 Liquids in Solution A **solution** consists of a **solute** dissolved in a **solvent**. A solution results from the interaction of solute and solvent molecules. A polar molecule has a region of both partial positive and partial negative charge that is termed a **dipole**. In water, there are two dipoles that combine to produce a single **net dipole** for the molecule. Water is composed of polar H_2O molecules, hence it is considered a **polar solvent**. Hexane is composed of nonpolar C_6H_{14} molecules, and so it is considered a **nonpolar solvent**. Although it is a simplification, polar solvents are often small molecules that contain one or more oxygen atoms, for example, H_2O, CH_3OH, and C_2H_5OH.	• Determine whether a liquid is miscible or immiscible in another liquid by applying the *like dissolves like* rule *Related Exercises 7–14*

Key Concepts	Learning Objectives and Related Exercises
We can predict whether two liquids dissolve by applying the ***like dissolves like* rule**. This rule states that two liquids are **miscible** if they are both polar solvents or both nonpolar solvents. Furthermore, two liquids are **immiscible** if one is a polar solvent and the other a nonpolar solvent. Since CH_3OH is polar, it is miscible with water. Since C_6H_{14} is nonpolar, it is immiscible with water.	
13.3 Solids in Solution We can also usually predict whether or not a solid compound is soluble in a given solvent. According to the ***like dissolves like*** rule, a polar solvent dissolves a polar compound, and a nonpolar solvent dissolves a nonpolar compound. Ionic compounds are not soluble in nonpolar solvents, and some ionic compounds are only slightly soluble in polar solvents.	• Determine whether a solid is soluble or insoluble in a liquid by applying the *like dissolves like* rule. *Related Exercises 15–20*
13.4 The Dissolving Process In dissolving a solid compound in water, solvent molecules attack solute particles and pull them into the solution. The solute particles are then surrounded by water molecules forming a **solvent cage**. Molecular compounds release molecules in the solution, whereas ionic compounds release ions in the solution. A **colloid** is similar to a solution except that the dissolved particles are larger. These larger colloid particles demonstrate an ability to scatter a beam of light; this phenomenon is called the **Tyndall effect**.	• Illustrate an ionic compound and a molecular compound dissolving in water. *Related Exercises 21–24*
13.5 Rate of Dissolving The rate of dissolving a solid solute compound in a solution is increased by three factors: heating the solution, stirring the solution, and grinding the solute.	• Describe the effect of temperature, stirring, and particle size on the rate of dissolving for a solid compound in water. *Related Exercises 25–26*
13.6 Solubility and Temperature Generally, raising the temperature of a solution increases the **solubility** of a solid compound. By referring to a graph of solubility, we can determine the amount of solute that can dissolve in 100 g of water at a given temperature.	• Interpret a graph that shows temperature versus solubility of a solid compound in water. *Related Exercises 27–36*
13.7 Unsaturated, Saturated, and Supersaturated Solutions A solution containing the maximum amount of dissolved solute possible at a given temperature is a **saturated solution**. If the concentration of the solution is less than its maximum solubility, it is an **unsaturated solution**. Under special circumstances, it is possible to exceed the maximum solubility of a solution. Such a solution is unstable and it is said to be a **supersaturated solution**.	• Interpret a graph of temperature versus solubility and determine whether a solution is saturated, unsaturated, or supersaturated. *Related Exercises 37–40* • Distinguish among solutions that are saturated, unsaturated, or supersaturated. *Related Exercises 41–42*
13.8 Mass/Mass Percent Concentration Often, a solution concentration is expressed in terms of **mass/mass percent (m/m %)**. The mass/mass percent concentration is the mass in grams of dissolved solute in 100 g of solution.	• Calculate the mass percent concentration of a solution. *Related Exercises 43–44* • Write three pairs of unit factors given the mass percent concentration of a solution. *Related Exercises 45–46* • Solve problems that involve a mass of solute, mass of solvent, and the mass percent concentration of a solution. *Related Exercises 47–52*
13.9 Molar Concentration A solution concentration can be expressed as **molarity (*M*)**. The molar concentration, or molarity, is the number of moles of solute dissolved in one liter of solution (mol solute/L solution).	• Calculate the molar concentration of a solution. *Related Exercises 53–54* • Write a pair of unit factors given the molar concentration of a solution. *Related Exercises 55–56* • Solve problems that involve a mass of solute, volume of solution, and the molar concentration of a solution. *Related Exercises 57–64*

Key Concepts	Learning Objectives and Related Exercises
13.10 Dilution of a Solution Many dilute solutions in the chemistry laboratory are prepared from concentrated solutions. For example, a 6 *M* HCl solution can be diluted to 0.1 *M* HCl. When performing dilution calculations, it is important to realize that the amount of solute does not change when a solution is diluted.	• Solve problems that involve a solution undergoing dilution. *Related Exercises 65–68*
13.11 Solution Stoichiometry Solution stoichiometry calculations are similar to other stoichiometry calculations and require a balanced chemical equation. However, solution stoichiometry problems are different in that they involve the molar concentration of a solution.	• Solve problems that involve a balanced chemical equation and the molar concentration of a solution. *Related Exercises 69–70*

Problem–Solving Organizer

Topic	Procedure	Example
Gases in Solution Sec. 13.1	Henry's law states that the solubility of a gas is directly proportional to the partial pressure of the gas above the liquid.	If the solubility of a gas is 0.100 g/100 mL at 1.00 atm, what is it at 3.50 atm? $0.100\ \text{g/100 mL} \times \frac{3.50\ \cancel{\text{atm}}}{1.00\ \cancel{\text{atm}}} = 0.350\ \text{g/100 mL}$
Mass/Mass Percent Concentration Sec. 13.8	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply a unit factor to convert the given unit to the unit in the answer.	What mass of 10.0% table sugar solution contains 15.5 g of sugar? $15.5\ \cancel{\text{g sugar}} \times \frac{100\ \text{g solution}}{10.0\ \cancel{\text{g sugar}}} = 155\ \text{g solution}$
Molar Concentration Sec. 13.9	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply unit factors to convert the given unit to the unit in the answer.	What is the mass of sucrose, $C_{12}H_{22}O_{11}$ (342.34 g/mol), in 0.500 L of a 0.100 *M* solution? $0.500\ \cancel{\text{L solution}} \times \frac{0.100\ \cancel{\text{mol sucrose}}}{1\ \cancel{\text{L solution}}} \times \frac{342.34\ \text{g sucrose}}{1\ \cancel{\text{mol sucrose}}}$ $= 17.1\ \text{g sucrose}$
Dilution of a Solution Sec. 13.10	Solve dilution problems by unit analysis or by the algebraic method $M_1 \times V_1 = M_2 \times V_2$	If 50.0 mL of 6.0 *M* NaOH is diluted to a total volume of 250 mL, what is the molarity? $6.0\ M \times 50.0\ \text{mL} = M_2 \times 250\ \text{mL}$ Solving for M_2, $\frac{6.0\ M \times 50.0\ \cancel{\text{mL}}}{250\ \cancel{\text{ml}}} = 1.2\ M$
Solution Stoichiometry Sec. 13.11	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply unit factors to convert the given unit to the unit in the answer.	What is the mass of AgCl (143.32 g/mol) produced from 50.0 mL of 0.100 *M* KCl and excess $AgNO_3$? Given: $KCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + KNO_3(aq)$ $50.0\ \cancel{\text{mL KCl}} \times \frac{0.100\ \cancel{\text{mol KCl}}}{1000\ \cancel{\text{mL KCl}}} \times \frac{1\ \cancel{\text{mol AgCl}}}{1\ \cancel{\text{mol KCl}}}$ $\times \frac{143.32\ \text{g AgCl}}{1\ \cancel{\text{mol AgCl}}} = 0.717\ \text{g AgCl}$

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** the solubility of a gas in a liquid is proportional to the partial pressure of the gas
______ **2.** the component of a solution that is present in the lesser quantity
______ **3.** the component of a solution that is present in the greater quantity
______ **4.** the general term for a solute dissolved in a solvent
______ **5.** a region in a polar molecule with a partial negative and a partial positive charge
______ **6.** the overall direction of partial negative charge in a molecule
______ **7.** a dissolving liquid composed of polar molecules
______ **8.** a dissolving liquid composed of nonpolar molecules
______ **9.** the principle that solubility is greatest when the polarity of the solute is similar to that of the solvent
______ **10.** refers to liquids that dissolve completely in one another
______ **11.** refers to liquids that do not dissolve in one another and separate into two layers
______ **12.** a cluster of solvent molecules surrounding a solute particle in solution
______ **13.** a mixture in which the dispersed particles range from 1 to 100 nm
______ **14.** the scattering of a beam of light by colloid particles
______ **15.** the maximum amount of solute that can dissolve in a solvent at a given temperature
______ **16.** a solution containing the maximum solute that can dissolve at a given temperature
______ **17.** a solution containing less than the maximum solute that can dissolve at a given temperature
______ **18.** a solution containing more than the maximum solute that can dissolve at a given temperature
______ **19.** a solution concentration expression that relates the mass of solute in grams dissolved in each 100 g of solution
______ **20.** a solution concentration expression that relates the moles of solute dissolved in each liter of solution

(a) colloid (*Sec. 13.3*)
(b) dipole (*Sec. 13.2*)
(c) Henry's law (*Sec. 13.1*)
(d) immiscible (*Sec. 13.2*)
(e) *like dissolves like* rule (*Sec. 13.2*)
(f) mass/mass percent (m/m %) (*Sec. 13.8*)
(g) miscible (*Sec. 13.2*)
(h) molarity (*M*) (*Sec. 13.9*)
(i) net dipole (*Sec. 13.2*)
(j) nonpolar solvent (*Sec. 13.2*)
(k) polar solvent (*Sec. 13.2*)
(l) saturated solution (*Sec. 13.7*)
(m) solubility (*Sec. 13.6*)
(n) solute (*Sec. 13.2*)
(o) solution (*Sec. 13.2*)
(p) solvent (*Sec. 13.2*)
(q) solvent cage (*Sec. 13.4*)
(r) supersaturated solution (*Sec. 13.7*)
(s) Tyndall effect (*Sec. 13.3*)
(t) unsaturated solution (*Sec. 13.7*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Gases in Solution (Sec. 13.1)

1. Indicate whether the solubility of carbon dioxide gas in soda increases or decreases for each of the following changes.
   (a) The temperature of the soda decreases from 20 °C to 10 °C.
   (b) The partial pressure of CO_2 decreases from 3 atm to 2 atm.
2. Indicate whether the solubility of carbon dioxide gas in soda increases or decreases for each of the following changes.
   (a) The temperature of the soda increases from 10 °C to 20 °C.
   (b) The partial pressure of CO_2 increases from 1 atm to 2 atm.
3. The solubility of nitrogen gas is 1.90 mL/dL of blood at 1.00 atm. What is the solubility of nitrogen gas in a deep-sea diver's blood at a depth of 200 feet and a pressure of 7.00 atm?
4. The solubility of nitrogen gas is 1.90 mL/dL of blood at 1.00 atm. What is the solubility of nitrogen gas in a deep-sea diver's blood at a depth of 125 feet and a pressure of 4.50 atm?

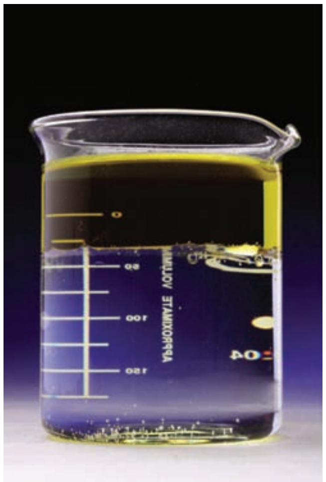
◀ **Illustration of Henry's Law** After a bottle of soda is opened, it fizzes as the carbonation escapes. The carbon dioxide gas is less soluble than when it was under pressure in the bottle.

5. The solubility of nitrous oxide is 0.12 g/100 g water at 20 °C and 1.00 atm. What is the partial pressure required to dissolve 0.25 g of the gas in 100 g of water at 20 °C?

6. The solubility of nitrous oxide is 0.12 g/100 g water at 20 °C and 1.00 atm. What is the partial pressure required to dissolve 0.55 g of the gas in 100 g of water at 20 °C?

Liquids in Solution (Sec. 13.2)

7. State whether the solutes and solvents in each of the following combinations are miscible or immiscible.
   (a) polar solute + polar solvent
   (b) polar solute + nonpolar solvent
8. State whether the solutes and solvents in each of the following combinations are miscible or immiscible.
   (a) nonpolar solute + polar solvent
   (b) nonpolar solute + nonpolar solvent
9. Predict whether each of the following solvents is polar or nonpolar.
   (a) water, H_2O **(b)** hexane, C_6H_{14}
   (c) acetone, C_3H_6O **(d)** chloroform, $CHCl_3$
10. Predict whether each of the following solvents is polar or nonpolar.
    (a) isopropyl alcohol, C_3H_7OH
    (b) pentane, C_5H_{12}
    (c) xylene, $C_6H_4(CH_3)_2$
    (d) trichloroethane, $C_2H_3Cl_3$
11. Predict whether each of the following solvents is miscible or immiscible with hexane, C_6H_{14}.
    (a) heptane, C_7H_{16}
    (b) methanol, CH_3OH
    (c) methyl ethyl ketone, C_4H_8O
    (d) chloroform, $CHCl_3$
12. Predict whether each of the following solvents is miscible or immiscible with water.
    (a) toluene, C_7H_8
    (b) ethanol, C_2H_5OH
    (c) acetic acid, CH_3COOH
    (d) trichloroethylene, C_2HCl_3
13. In the laboratory, how could you quickly determine whether an unknown liquid is polar or nonpolar?
14. Oil and water separate into two layers. Explain why the two liquids are immiscible using the *like dissolves like* rule.



◀ **Oil and Water** Salad oil is less dense than water; thus, salad oil floats on water.

Solids in Solution (Sec. 13.3)

15. State whether the following combinations of solute and solvent are generally soluble or insoluble.
    (a) polar solute + nonpolar solvent
    (b) nonpolar solute + nonpolar solvent
    (c) ionic solute + nonpolar solvent
16. State whether the following combinations of solute and solvent are generally soluble or insoluble.
    (a) polar solute + polar solvent
    (b) nonpolar solute + polar solvent
    (c) ionic solute + polar solvent
17. Predict whether each of the following compounds is soluble or insoluble in water.
    (a) naphthalene, $C_{10}H_8$
    (b) potassium hydroxide, KOH
    (c) calcium acetate, $Ca(C_2H_3O_2)_2$
    (d) trichlorotoluene, $C_7H_5Cl_3$
18. Predict whether each of the following compounds is soluble or insoluble in hexane, C_6H_{14}.
    (a) trichloroethylene, C_2HCl_3
    (b) iron(III) nitrate, $Fe(NO_3)_3$
    (c) tartaric acid, $H_2C_4H_4O_6$
    (d) dodecane, $C_{12}H_{26}$
19. Predict whether each of the following compounds is water soluble or fat soluble.
    (a) oleic acid, $C_{18}H_{34}O_2$
    (b) citric acid, $C_6H_8O_7$
    (c) lactic acid, $CH_3CH(OH)COOH$
    (d) stearic acid, $CH_3(CH_2)_{16}COOH$
20. Predict whether each of the following vitamins is water soluble or fat soluble.
    (a) vitamin B_3, $C_6H_6N_2O$ **(b)** vitamin C, $C_6H_8O_6$
    (c) vitamin D, $C_{27}H_{44}O$ **(d)** vitamin K, $C_{31}H_{46}O_2$

The Dissolving Process (Sec. 13.4)

21. Diagram a molecule of fructose, $C_6H_{12}O_6$, dissolved in water.
22. Diagram a crystal of sucrose, $C_{12}H_{22}O_{11}$, dissolving in water.
23. Diagram a formula unit of each of the following substances dissolved in water.
    (a) lithium bromide, LiBr
    (b) calcium chloride, $CaCl_2$
24. Diagram a formula unit of each of the following substances dissolved in water.
    (a) cobalt(II) sulfate, $CoSO_4$
    (b) nickel(II) nitrate, $Ni(NO_3)_2$

Rate of Dissolving (Sec. 13.5)

25. What three factors increase the rate of dissolving of a solid substance in solution?
26. Indicate whether each of the following increases, decreases, or has no effect on the rate at which 10.0 g of sugar dissolves in a liter of water.
    (a) using water from the refrigerator
    (b) shaking the sugar and water
    (c) using powdered sugar rather than crystals
    (d) using tap water rather than distilled water

Solubility and Temperature (Sec. 13.6)

27. How many grams of each of the following solutes can dissolve in 100 g of water at 20 °C? (Refer to Figure 13.5.)
(a) NaCl **(b)** KCl

28. How many grams of each of the following solutes can dissolve in 100 g of water at 30 °C? (Refer to Figure 13.5.)
(a) LiCl **(b)** $C_{12}H_{22}O_{11}$

29. Determine the maximum solubility of each of the following solid compounds at 40 °C. (Refer to Figure 13.5.)
(a) NaCl **(b)** KCl

30. Determine the maximum solubility of each of the following solid compounds at 50 °C. (Refer to Figure 13.5.)
(a) LiCl **(b)** $C_{12}H_{22}O_{11}$

31. What is the minimum temperature required to dissolve each of the following? (Refer to Figure 13.5.)
(a) 34 g NaCl in 100 g of water
(b) 50 g KCl in 100 g of water

32. What is the minimum temperature required to dissolve each of the following? (Refer to Figure 13.5.)
(a) 90 g LiCl in 100 g of water
(b) 120 g $C_{12}H_{22}O_{11}$ in 100 g of water

33. At what temperature is each of the following solutions saturated? (Refer to Figure 13.5.)
(a) 40 g NaCl/100 g water
(b) 40 g KCl/100 g water

34. At what temperature is each of the following solutions saturated? (Refer to Figure 13.5.)
(a) 105 g LiCl/100 g water
(b) 140 g $C_{12}H_{22}O_{11}$/100 g water

35. State whether each of the following solutions is saturated, unsaturated, or supersaturated. (Refer to Figure 13.5.)
(a) 110 g LiCl/100 g H_2O at 50 °C
(b) 110 g LiCl/100 g H_2O at 70 °C
(c) 110 g LiCl/100 g H_2O at 90 °C

36. State whether each of the following solutions is saturated, unsaturated, or supersaturated. (Refer to Figure 13.5.)
(a) 130 g $C_{12}H_{22}O_{11}$/100 g H_2O at 40 °C
(b) 130 g $C_{12}H_{22}O_{11}$/100 g H_2O at 50 °C
(c) 130 g $C_{12}H_{22}O_{11}$/100 g H_2O at 60 °C

Unsaturated, Saturated, and Supersaturated Solutions (Sec. 13.7)

37. Assume that 100 g of LiCl is dissolved in 100 g of water at 100 °C and the solution is allowed to cool to 20 °C. (Refer to Figure 13.5.)
(a) How much LiCl remains in solution?
(b) How much LiCl crystallizes from the solution?

38. Assume that 50 g of KCl is dissolved in 100 g of water at 100 °C and the solution is allowed to cool to 20 °C. (Refer to Figure 13.5.)
(a) How much KCl remains in solution?
(b) How much KCl crystallizes from the solution?

39. State whether each of the following solutions is saturated, unsaturated, or supersaturated. (Refer to Figure 13.6.)
(a) 120 g $NaC_2H_3O_2$ in 100 g of water at 70 °C
(b) 120 g $NaC_2H_3O_2$ in 100 g of water at 80 °C
(c) 120 g $NaC_2H_3O_2$ in 100 g of water at 90 °C

40. State whether each of the following solutions is saturated, unsaturated, or supersaturated. (Refer to Figure 13.6.)
(a) 100 g $NaC_2H_3O_2$ in 100 g of water at 25 °C
(b) 100 g $NaC_2H_3O_2$ in 100 g of water at 55 °C
(c) 100 g $NaC_2H_3O_2$ in 100 g of water at 75 °C

41. The solubility of rock salt is 40 g/100 g water at 30 °C. If a solution contains 10 g of rock salt in 25 g of water at 30 °C, is the solution saturated, unsaturated, or supersaturated?

42. The solubility of sugar is 110 g/100 g water at 30 °C. If a solution contains 25 g of sugar in 25 g of water at 30 °C, is the solution saturated, unsaturated, or supersaturated?

Mass/Mass Percent Concentration (Sec. 13.8)

43. Calculate the mass/mass percent concentration for each of the following solutions.
(a) 1.25 g NaCl in 100.0 g of solution
(b) 2.50 g $K_2Cr_2O_7$ in 95.0 g of solution
(c) 10.0 g $CaCl_2$ in 250.0 g of solution
(d) 65.0 g $C_6H_{12}O_6$ in 125.0 g of solution

44. Calculate the mass/mass percent concentration for each of the following solutions.
(a) 20.0 g KI in 100.0 g of water
(b) 2.50 g $AgC_2H_3O_2$ in 95.0 g of water
(c) 5.57 g $SrCl_2$ in 225.0 g of water
(d) 50.0 g $C_{12}H_{22}O_{11}$ in 200.0 g of water

45. Write three pairs of unit factors for each of the following aqueous solutions given the mass/mass percent concentration.
(a) 1.50% KBr **(b)** 2.50% $AlCl_3$
(c) 3.75% $AgNO_3$ **(d)** 4.25% Li_2SO_4

46. Write three pairs of unit factors for each of the following aqueous solutions given the mass/mass percent concentration.
(a) 3.35% $MgCl_2$ **(b)** 5.25% $Cd(NO_3)_2$
(c) 6.50% Na_2CrO_4 **(d)** 7.25% $ZnSO_4$

47. What mass of solution contains the following amount of solute?
(a) 5.36 g of glucose in a 10.0% solution
(b) 25.0 g of sucrose in a 12.5% solution

48. What mass of solution contains the following amount of solute?
(a) 35.0 g of hydrochloric acid in a 5.00% solution
(b) 10.5 g of acetic acid in a 4.50% solution

49. What mass of solute is dissolved in the following solutions?
(a) 85.0 g of 2.00% $FeBr_2$ solution
(b) 105.0 g of 5.00% Na_2CO_3 solution

50. What mass of solute is dissolved in the following solutions?
(a) 10.0 g of 2.50% K_2CO_3 solution
(b) 50.0 g of 5.00% Li_2SO_4 solution

51. What mass of water is necessary to prepare the following solutions?
(a) 100.0 g of 5.00% nitric acid, HNO_3
(b) 250.0 g of 10.0% NaOH solution

52. What mass of water is necessary to prepare the following solutions?
(a) 250.0 g of 0.90% saline solution
(b) 100.0 g of 5.00% sugar solution

Molar Concentration (Sec. 13.9)

53. Calculate the molar concentration for each of the following solutions.
    (a) 1.50 g NaCl in 100.0 mL of solution
    (b) 1.50 g $K_2Cr_2O_7$ in 100.0 mL of solution
    (c) 5.55 g $CaCl_2$ in 125 mL of solution
    (d) 5.55 g Na_2SO_4 in 125 mL of solution

54. Calculate the molar concentration for each of the following solutions.
    (a) 1.00 g KCl in 75.0 mL of solution
    (b) 1.00 g Na_2CrO_4 in 75.0 mL of solution
    (c) 20.0 g $MgBr_2$ in 250.0 mL of solution
    (d) 20.0 g Li_2CO_3 in 250.0 mL of solution

55. Write two pairs of unit factors for the following aqueous solutions given the molar concentration.
    (a) 0.100 *M* LiI **(b)** 0.100 *M* $NaNO_3$
    (c) 0.500 *M* K_2CrO_4 **(d)** 0.500 *M* $ZnSO_4$

56. Write two pairs of unit factors for the following aqueous solutions given the molar concentration.
    (a) 0.150 *M* KBr **(b)** 0.150 *M* $Ca(NO_3)_2$
    (c) 0.333 *M* $Sr(C_2H_3O_2)_2$ **(d)** 0.333 *M* NH_4Cl

57. What volume of solution contains the stated amount of solute?
    (a) 2.50 g solute in 0.325 *M* KNO_3
    (b) 2.50 g solute in 0.325 *M* $AlBr_3$
    (c) 2.50 g solute in 1.00 *M* $Co(C_2H_3O_2)_2$
    (d) 2.50 g solute in 1.00 *M* $(NH_4)_3PO_4$

58. What volume of solution contains the stated amount of solute?
    (a) 10.0 g solute in 0.275 *M* NaF
    (b) 10.0 g solute in 0.275 *M* $CdCl_2$
    (c) 10.0 g solute in 0.408 *M* K_2CO_3
    (d) 10.0 g solute in 0.408 *M* $Fe(ClO_3)_3$

59. What is the mass of solute in each of the following solutions?
    (a) 2.25 L of 0.200 *M* $FeCl_3$
    (b) 2.25 L of 0.200 *M* KIO_4
    (c) 50.0 mL of 0.295 *M* $ZnSO_4$
    (d) 50.0 mL of 0.295 *M* $Ni(NO_3)_2$

60. What is the mass of solute in each of the following solutions?
    (a) 1.00 L of 0.100 *M* NaOH
    (b) 1.00 L of 0.100 *M* $LiHCO_3$
    (c) 25.0 mL of 0.500 *M* $CuCl_2$
    (d) 25.0 mL of 0.500 *M* $KMnO_4$

61. What is the molar concentration of a saturated solution of calcium hydroxide that contains 0.185 g of solute in 100.0 mL of solution?

62. What is the molar concentration of a saturated solution of calcium sulfate that contains 0.209 g of solute in 100.0 mL of solution?

63. A hospital glucose solution is analyzed to verify its concentration. A 10.0 mL sample with a mass of 10.483 g is evaporated to dryness. If the solid glucose residue has a mass of 0.524 g, what are (a) the mass/mass percent concentration, and (b) the molar concentration of the glucose, $C_6H_{12}O_6$, solution?

64. A hospital saline solution is analyzed to confirm its concentration. A 50.0 mL sample with a mass of 50.320 g is evaporated to dryness. If the solid sodium chloride residue has a mass of 0.453 g, find (a) the mass/mass percent concentration , and (b) the molar concentration of the NaCl solution.

Dilution of a Solution (Sec. 13.10)

65. Concentrated hydrochloric acid is 12 *M*. What is the molarity of an HCl solution prepared by diluting 10.0 mL of concentrated acid to a total volume of 250.0 mL?

66. Concentrated acetic acid is 17 *M*. What is the molarity of an CH_3COOH solution prepared by diluting 50.0 mL of concentrated acid to a total volume of 500.0 mL?

67. Concentrated sulfuric acid is 18 *M*. What volume of H_2SO_4 is diluted with distilled water to prepare 2.50 L of 0.10 *M* H_2SO_4?

68. Concentrated nitric acid is 16 *M*. What volume of HNO_3 is diluted with distilled water to prepare 5.00 L of 0.10 *M* HNO_3?

Solution Stoichiometry (Sec. 13.11)

69. Given that 50.0 mL of 0.125 *M* silver nitrate reacts completely with 15.0 mL magnesium iodide solution according to the unbalanced equation:

$$MgI_2(aq) + AgNO_3(aq) \rightarrow AgI(s) + Mg(NO_3)_2(aq)$$

    (a) What is the molarity of the $Mg(NO_3)_2$ solution?
    (b) What is the mass of AgI precipitate?

70. Given that 25.0 mL of 0.170 *M* potassium iodide reacts completely with 0.209 *M* mercury(II) nitrate solution according to the unbalanced equation:

$$Hg(NO_3)_2(aq) + KI(aq) \rightarrow HgI_2(s) + KNO_3(aq)$$

    (a) What volume of $Hg(NO_3)_2$ is required for complete precipitation?
    (b) What is the mass of HgI_2 precipitate?

General Exercises

71. In foggy weather, automobile headlights demonstrate the Tyndall effect. What is the approximate size of water droplets in the fog?

72. In a smoky auditorium, stage lights often demonstrate the Tyndall effect. What is the approximate size of smoke particles in the auditorium?

73. Explain why scuba divers use a special gas mixture of oxygen and helium, rather than compressed air, when they dive to depths greater than 125 feet.

74. Calculate the solubility of nitrogen in blood for a scuba diver at a depth of 125 feet and a pressure of 4.68 atm. The solubility of nitrogen gas is 0.0019 g/100 g blood at normal pressure.

75. Calculate the mass of sulfur dioxide gas in 1.00 L of saturated solution at 20 °C. The solubility of SO_2 at 20 °C is 22.8 g/100 mL.

76. Calculate the mass of chlorine gas in 500 mL of saturated solution at 20 °C. The solubility of Cl_2 at 20 °C is 0.63 g/100 mL.

77. Ethyl ether, $C_4H_{10}O$, is 7% miscible in water. Explain why ether is only partially miscible even though molecules contain an electronegative oxygen atom.

78. Phenol, C_6H_5OH, is 7% soluble in water. Explain why phenol is only partially soluble even though molecules contain an electronegative oxygen atom.

79. Explain the *decrease* in solubility for the following alcohols in water.

Alcohol	Solubility
ethanol, C_2H_5OH	miscible
pentanol, $C_5H_{11}OH$	2.2 g/100 mL H_2O
hexanol, $C_6H_{13}OH$	0.59 g/100 mL H_2O
decanol, $C_{10}H_{21}OH$	immiscible

80. Explain the *decrease* in solubility for the following acids in water.

Acid	Solubility
acetic acid, CH_3COOH	miscible
pentanoic acid, C_4H_9COOH	5.0 g/100 mL H_2O
hexanoic acid, $C_5H_{11}COOH$	1.1 g/100 mL H_2O
decanoic acid, $C_9H_{19}COOH$	immiscible

81. Identify the solutes and solvents in the following solutions.
(a) 80-proof ethyl alcohol (40% ethanol in water)
(b) 190-proof ethyl alcohol (95% ethanol in water)

82. Predict whether water or chloroform, $CHCl_3$, is a better solvent for each of the following household substances.
(a) grease **(b)** maple syrup
(c) food color **(d)** gasoline

Challenge Exercises

83. A vinegar solution is 5.00% acetic acid, CH_3COOH. What is the molarity of CH_3COOH in vinegar? (Assume the density is 1.01 g/mL.)

84. A bleach solution is 5.00% sodium hypochlorite, NaClO. What is the molarity of NaClO in a bleach? (Assume the density is 1.04 g/mL.)

Chapter 13 Self-Test

Answers to Self-Test are in Appendix J.

1. Yeast and sugar are added to champagne to give the sparkle of carbonation. Under what conditions is carbon dioxide gas most soluble? (Sec. 13.1)
(a) low temperature, high pressure
(b) low temperature, low pressure
(c) high temperature, low pressure
(d) high temperature, high pressure
(e) none of the above

2. Apply the *like dissolves like* rule to predict which of the following liquids is *miscible* with water. (Sec. 13.2)
(a) benzene, C_6H_6 **(b)** chloroform, $CHCl_3$
(c) methanol, CH_3OH **(d)** all of the above
(e) none of the above

3. Apply the *like dissolves like* rule to predict which of the following is a water-soluble vitamin. (Sec. 13.3)
(a) vitamin A, $C_{20}H_{30}O$ **(b)** vitamin C, $C_6H_8O_6$
(c) vitamin D, $C_{27}H_{44}O$ **(d)** vitamin E, $C_{29}H_{50}O_2$
(e) none of the above

4. When glucose, $C_6H_{12}O_6$, dissolves in water, which of the following is formed in the solution? (Sec. 13.4)
(a) hydrated clusters of H_2O
(b) hydrated clusters of CO_2
(c) hydrated clusters of H_2CO_3
(d) hydrated clusters of $C_6H_{12}O_6$
(e) none of the above

5. Which of the following increases the rate of dissolving for a solid solute in a solvent? (Sec. 13.5)
(a) heating the solution **(b)** stirring the solution
(c) grinding the solute **(d)** all of the above
(e) none of the above

6. What is the solubility of $C_{12}H_{22}O_{11}$ at 20 °C? (Refer to Figure 13.5 in Sec. 13.6.)
(a) 20 g/100 g water **(b)** 55 g/100 g water
(c) 100 g/100 g water **(d)** 140 g/100 g water
(e) 150 g/100 g water

7. How concentrated is a solution containing 130 g of $NaC_2H_3O_2$ in 100 g water at 90 °C? (Refer to Figure 13.6 in Sec. 13.7.)
(a) unsaturated **(b)** saturated
(c) supersaturated **(d)** superunsaturated
(e) none of the above

8. What mass of solute is dissolved in 5.00 kg of 40.0% antifreeze solution? (Sec. 13.8)
(a) 1.25 kg **(b)** 2.00 kg
(c) 7.50 kg **(d)** 12.5 kg
(e) 200 kg

9. What is the volume of 3.00 *M* sulfuric acid that contains 9.809 g of H_2SO_4 solute (98.09 g/mol)? (Sec. 13.9)
(a) 0.300 mL **(b)** 30.0 mL
(c) 33.3 mL **(d)** 333 mL
(e) 3330 mL

10. What volume of 16 *M* nitric acid must be diluted with distilled water to prepare 500.0 mL of 0.50 *M* HNO_3? (Sec. 13.10)
(a) 0.016 mL **(b)** 0.16 mL
(c) 1.6 mL **(d)** 16 mL
(e) 160 mL

11. What volume of 0.115 *M* hydrochloric acid reacts completely with 0.125 g of sodium carbonate, Na_2CO_3 (105.99 g/mol)? (Sec. 13.11)

$$Na_2CO_3(s) + 2\,HCl(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l) + 2\,CO_2(g)$$

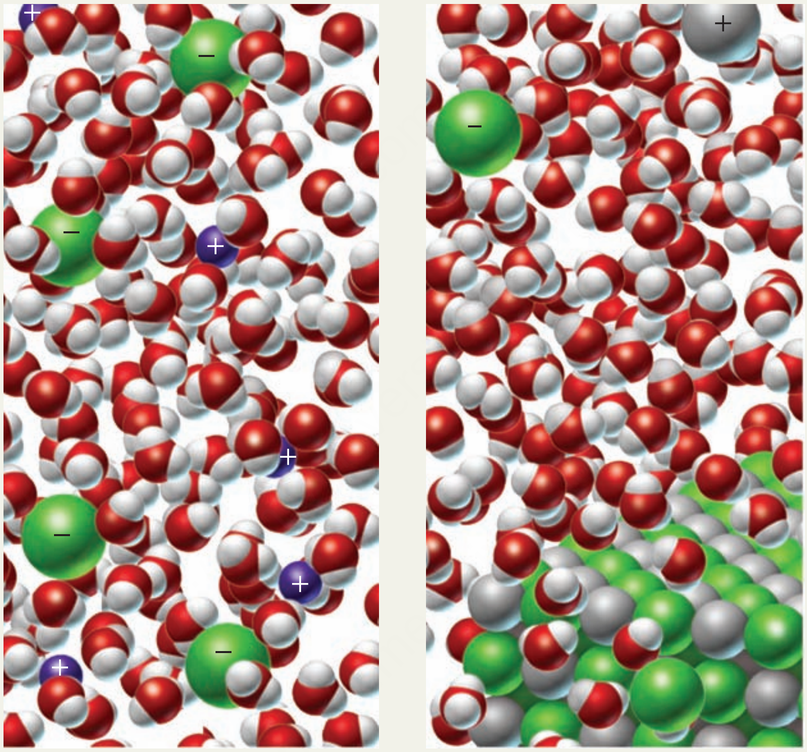
(a) 9.75 mL **(b)** 10.3 mL
(c) 20.5 mL **(d)** 41.0 mL
(e) 97.5 mL

Key Concepts

12. Explain why bubbles form on the inside surface of a pan when water is heated.
13. Explain why peanut butter can be used to remove grease from clothing.
14. Explain why powdered sugar dissolves more rapidly in water than granulated sugar.

Critical Thinking

15. Lake Nuclear and Lake Green are adjoining at the same elevation. If water from Lake Nuclear cools a nuclear power plant and Lake Green does not, which lake has more dissolved oxygen gas?
16. Examine the following illustrations of NaCl and AgCl dissolving in water. At the molecular level, explain why NaCl is more soluble than AgCl.

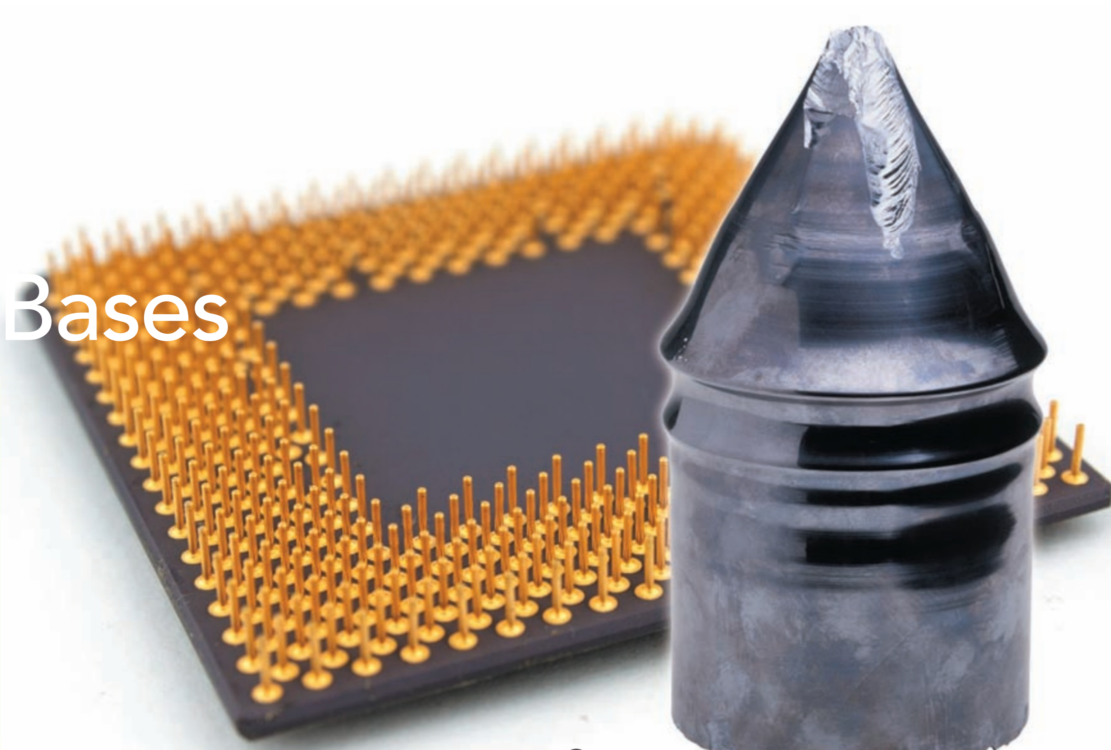


(a) NaCl very soluble

(b) AgCl slightly soluble

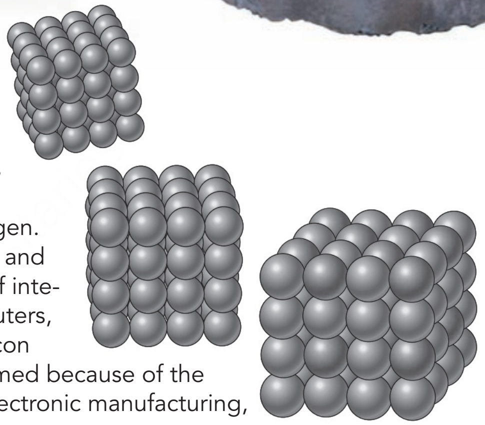
CHAPTER 14

Bases



"*Many of life's failures are people who did not realize how close they were to success when they gave up.*"

Thomas Alva Edison, Inventor and Scientist (1847–1931)



14 Si silicon

Element 14: Silicon

Silicon is a semimetal that comprises about 25% of Earth's crust. It occurs naturally in mineral form combined with oxygen. Silicon plays an important role in technology and is the starting material for the manufacture of integrated circuit chips, which are used in computers, smartphones, and electronic appliances. Silicon Valley in the San Francisco Bay area is so-named because of the large number of corporations carrying out electronic manufacturing, as well as research and development.

Acids and bases play an important role in our lives. The proper acidity of our blood and other body fluids is vital to our well-being and is carefully controlled by an elaborate buffering system. Many of the foods we eat are acidic. For instance, oranges and lemons contain citric acid, and vinegar is a solution of acetic acid. Vitamin C is ascorbic acid, and aspirin is acetylsalicylic acid. We add muriatic acid to our swimming pools and sulfuric acid to our car batteries.

Acids and bases neutralize each other. Baking soda and milk of magnesia are basic. Antacid tablets contain basic substances such as carbonates, bicarbonates, and hydroxides for neutralizing an acidic, upset stomach. We use a dilute solution of basic ammonia to clean floors and a concentrated solution of sodium hydroxide to clean drains and ovens.

LEARNING OBJECTIVES

- List the general properties of acids and bases.
- Classify a solution of given pH as one of the following: strongly acidic, weakly acidic, neutral, weakly basic, or strongly basic.

14.1 Properties of Acids and Bases

An *acid* is any substance that releases hydrogen ions, H^+, in water. This property of acids allows us to test for acidity. In the laboratory we can determine if a solution is acidic using a strip of litmus paper. Litmus paper turns red in the presence of hydrogen ions. Therefore, if we put a piece of litmus paper in a solution and the paper turns red, we know the solution is an acid. For example, litmus paper in lemon juice turns red because lemon juice is acidic.

Litmus paper is made from a plant pigment that is sensitive to changes in acid concentration. The colors of many plants are affected by the acidity of the soil in which they are grown. For example, the color of an orchid can vary from pale lavender to deep purple depending upon the acidity of the soil. The color of a rose fades from red to pink when the rose is placed in an acid solution.

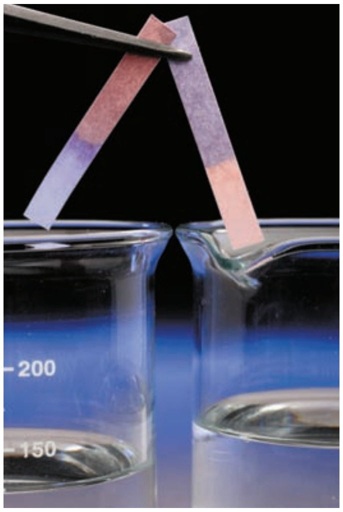
Acids have their own special properties. One property is that they have a sour taste. The tart taste of a lemon, an apple, or vinegar shows us that these are acidic foods. The taste buds that are sensitive to acids are located along the edge of our tongue, which is why we roll our tongue when we taste something that is sour.

A *base* is any substance that releases hydroxide ions, OH^-, in water. Bases also have special properties. For instance, they feel slippery or soapy to the touch and have a bitter taste. The bitter taste of milk of magnesia shows us that it is a basic substance. The taste buds that are sensitive to bases are located on the tongue toward the back of our throat, and we often react by sticking out our tongue when we taste something that is bitter.

We can use litmus paper to determine if a solution is basic. Litmus paper turns blue in the presence of hydroxide ion, which is found in drain cleaner. Another property of acids and bases is their ability to undergo a neutralization reaction, which we studied in Section 7.11. An acid and a base react to produce a salt and water. For example, hydrochloric acid neutralizes a potassium hydroxide solution to give potassium chloride and water.

$$HCl(aq) + KOH(aq) \longrightarrow KCl(aq) + H_2O(l)$$

The properties of acids and bases are summarized in Table 14.1.



▲ **Litmus Paper** Red litmus paper dipped in a basic solution turns blue (left). Blue litmus paper dipped in an acidic solution turns red (right).

TABLE 14.1 Properties of Acids and Bases

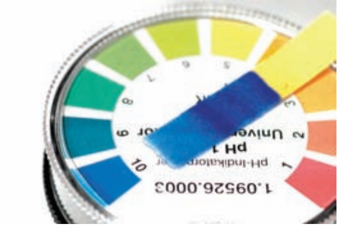
Property	Acidic Solutions	Basic Solutions
taste	sour	bitter
feel	—	slippery, soapy
litmus paper	litmus turns red	litmus turns blue
pH value	less than 7	greater than 7
neutralization reaction	react with bases to give a salt and water	react with acids to give a salt and water

pH Scale

A pH value expresses the acidity or basicity of a solution. Most solutions have a pH between 0 and 14. For example, a 1 *M* HCl solution has a pH of 0 (strongly acidic), and a 1 *M* NaOH solution has a pH of 14 (strongly basic). A pH of 7 is considered neutral; pure distilled water has a pH of 7.

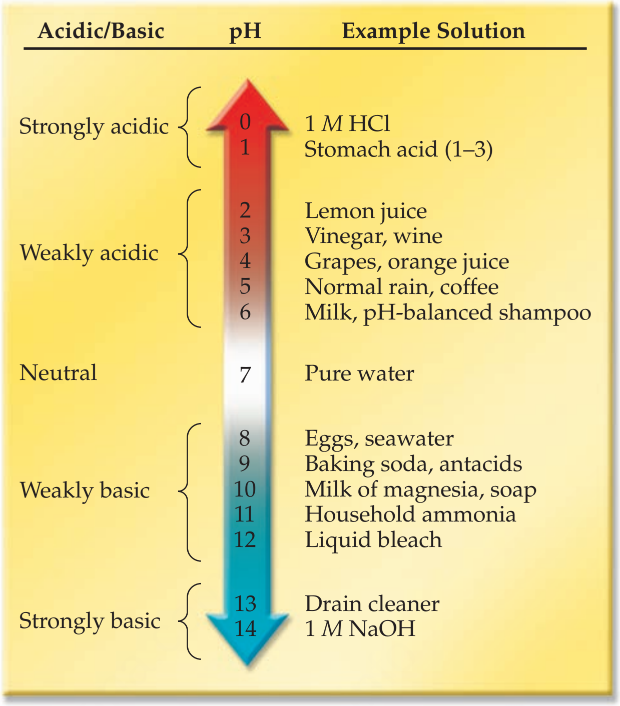
On the pH scale, acidic solutions have a pH less than 7. As the acidity increases, the pH value *decreases.* A solution having a pH of 3 is more acidic than a solution with a pH of 4. Basic solutions have a pH greater than 7. As the basicity increases, the pH value *increases.* A solution having a pH of 11 is more basic than a solution with a pH of 10.

A solution can be classified as strongly acidic, weakly acidic, neutral, weakly basic, or strongly basic on the basis of pH. A strongly acidic solution has a pH between 0 and 2. A weakly acidic solution has a pH between 2 and 7. A weakly basic solution has a pH between 7 and 12. A strongly basic solution has a pH between 12 and 14. Figure 14.1 illustrates some common substances and their approximate pH.



▲ **pH Test Strip** The indicator test strip paper uses universal indicator to test for pH. The test strip shown indicates a base at pH 10.

▶ **Figure 14.1 The pH Scale** Note that solutions may be strongly acidic (pH ~ 1), weakly acidic (pH ~ 4), neutral (pH ~ 7), weakly basic (pH ~ 10), or strongly basic (pH ~ 14).



Acidic/Basic	pH	Example Solution
Strongly acidic	0	1 *M* HCl
	1	Stomach acid (1–3)
Weakly acidic	2	Lemon juice
	3	Vinegar, wine
	4	Grapes, orange juice
	5	Normal rain, coffee
	6	Milk, pH-balanced shampoo
Neutral	7	Pure water
Weakly basic	8	Eggs, seawater
	9	Baking soda, antacids
	10	Milk of magnesia, soap
	11	Household ammonia
	12	Liquid bleach
Strongly basic	13	Drain cleaner
	14	1 *M* NaOH

The following example exercise illustrates the acid–base strength of a few common solutions:

EXAMPLE EXERCISE 14.1 The pH Scale

Indicate whether each of the following solutions is considered strongly acidic, weakly acidic, neutral, weakly basic, or strongly basic:

(a) stomach acid, pH 1.5
(b) oven cleaner, pH 13.5
(c) orange juice, pH 4.5
(d) deionized water, pH 7.0
(e) eggs, pH 7.5
(f) carbonated soda, pH 4.0

Solution

Refer to the guidelines in Figure 14.1 to classify each solution.

(a) Gastric juice has a pH between 0 and 2. It is *strongly acidic.*
(b) Oven cleaner has a pH of 13.5. It is *strongly basic.*
(c) Orange juice has a pH between 2 and 7. It is *weakly acidic.*
(d) Deionized water has a pH of 7.0 and is *neutral.*
(e) Eggs have a pH of 7.5 and are *weakly basic.*
(f) Carbonated soda has a pH of 4.0 and is *weakly acidic.*

Practice Exercise

Indicate whether each of the following properties corresponds to an acid or to a base:

(a) sour taste
(b) slippery feel
(c) turns litmus paper red
(d) pH greater than 7

Answers: (a) acid (b) base (c) acid (d) base

Concept Exercise

Which of the following has the highest acidity: pH 1, pH 5, or pH 9?

Answer: See Appendix G, 14.1.

▲ **Lemons Are Sour** Lemons, limes, oranges, and grapefruit contain citric acid, which is sour.

Buffers

Controlling acidity is important in the manufacture of food, paper, and chemicals. In agriculture, controlling the pH of the soil is necessary to obtain good crop yields.

A **buffer** is a solution that resists changes and controls the pH when an acid or a base is added. A buffer solution functions by consuming excess hydrogen ions or hydroxide ions, and thereby maintains its pH.

A buffer solution is composed of an aqueous weak acid and one of its salts, for example, citric acid and sodium citrate. If we add acid to a solution, the citrate ion combines with excess hydrogen ions to form citric acid. Conversely, if we add base to a solution, citric acid neutralizes excess hydroxide ions to form a citrate salt. If the amount of acid or base added to a buffer solution is relatively small compared to the amount of buffer, there is very little change in pH.

Other examples of buffers include hydrogen carbonate ions and carbonate ions, which are present in seawater and help to preserve the pH near 8. A similar buffer system helps to control the pH of blood, which must be maintained in a very narrow range near a pH of 7.4.

▲ **Buffer Solutions** Standard buffer solutions are available for a wide range of pH values.

14.2 Arrhenius Acids and Bases

LEARNING OBJECTIVES

- Identify an Arrhenius acid and base.
- Classify a strong or weak acid and base given the degree of ionization.
- Indicate the acid and base that react to produce a given salt.

In 1884, the Swedish chemist Svante Arrhenius suggested the first definitions for an acid and a base. He proposed that an acid is a substance that ionizes in water to produce hydrogen ions. Conversely, he proposed that a base is a substance that dissociates in water to release hydroxide ions. These simple concepts define an **Arrhenius acid** and **Arrhenius base.**

Ionization versus Dissociation

Acids and bases have varying strengths. The strength of an Arrhenius acid is measured by the degree of ionization in solution. **Ionization** is the process by which molecules of a polar compound ionize to form cations and anions. Thus, a molecule of hydrochloric acid, HCl, ionizes into H^+ and Cl^-.

The strength of an Arrhenius base is measured by the degree of dissociation in the solution. **Dissociation** is the process whereby already existing ions of an ionic compound simply separate. Thus, a formula unit of sodium hydroxide, NaOH, dissociates into Na^+ and OH^-.

Strong and Weak Arrhenius Acids

An acid is considered either strong or weak depending on how much it ionizes in water. According to the Arrhenius definition, a strong acid ionizes extensively to release hydrogen ions in solution. Hydrochloric acid is considered a strong acid because it ionizes nearly 100%. Hydrofluoric acid is a weak acid because it is only slightly ionized, about 1%. Table 14.2 lists some common Arrhenius acids.

TABLE 14.2 Common Arrhenius Acids

Aqueous Acid	Percent Ionization	Acid Strength
hydrochloric acid, $HCl(aq)$	~100%	strong
nitric acid, $HNO_3(aq)$	~100%	strong
sulfuric acid, $H_2SO_4(aq)$	~100%	strong
acetic acid, $HC_2H_3O_2(aq)$	slightly	weak
carbonic acid, $H_2CO_3(aq)$	slightly	weak
hydrofluoric acid, $HF(aq)$	slightly	weak
phosphoric acid, $H_3PO_4(aq)$	slightly	weak

▲ **Hydrofluoric Acid, HF** Although aqueous HF is a weak acid, it attacks glass. This design was created by coating glass with wax and then removing wax from the pattern areas. The glass was dipped into hydrofluoric acid, which etched the design.

All Arrhenius acids have a hydrogen atom attached to an acid molecule by a polar bond. Most acids are weak ternary oxyacids in which the H atom is attached to an O atom. When the acid molecule ionizes, it breaks this bond. In an aqueous solution, polar water molecules help the acid ionize by pulling the hydrogen ion away from the acid molecule. We can show the ionization for hydrochloric acid, HCl, and acetic acid, $HC_2H_3O_2$, as follows:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq) \qquad (\sim 100\%)$$

$$HC_2H_3O_2(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + C_2H_3O_2^-(aq) \qquad (\sim 1\%)$$

Although chemists often refer to acids as solutions containing hydrogen ions, the actual species is H_3O^+. The **hydronium ion,** H_3O^+, results from a hydrogen ion attaching to a nonbonding electron pair on a water molecule. For simplicity, we will usually refer to a hydrogen ion in aqueous solution as $H^+(aq)$.

Strong and Weak Arrhenius Bases

Recall that a base is considered either strong or weak depending on how much of it dissociates. According to the Arrhenius definition, a strong base dissociates extensively to release hydroxide ions in solution. Sodium hydroxide is a strong base and dissociates nearly 100% in an aqueous solution. A weak base dissociates very little to release few hydroxide ions in solution. Ammonium hydroxide is a weak base because it provides relatively few hydroxide ions in solution. Table 14.3 lists some common Arrhenius bases.

TABLE 14.3 Common Arrhenius Bases

Aqueous Base	Percent Dissociation	Base Strength
barium hydroxide, $Ba(OH)_2(aq)$	~100%	strong
calcium hydroxide, $Ca(OH)_2(aq)$	~100%	strong
lithium hydroxide, $LiOH(aq)$	~100%	strong
potassium hydroxide, $KOH(aq)$	~100%	strong
sodium hydroxide, $NaOH(aq)$	~100%	strong
ammonium hydroxide, $NH_4OH(aq)$	slightly	weak

When we dissolve sodium hydroxide in water, we obtain aqueous NaOH. In the dissolving process, NaOH dissociates into aqueous sodium ions and aqueous hydroxide ions. Aqueous ammonium hydroxide, NH_4OH, is a weak base and provides relatively few hydroxide ions in solution. We can show this process for the two Arrhenius bases as follows:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq) \qquad (\sim 100\%)$$

$$NH_4OH(aq) \longrightarrow NH_4^+(aq) + OH^-(aq) \qquad (\sim 1\%)$$

The following example exercise illustrates the classification of acid–base strength based on the ability to donate ions in solution.

EXAMPLE EXERCISE 14.2 Strong and Weak Acids

Classify each of the following solutions as a strong or a weak Arrhenius acid given the degree of ionization:

(a) perchloric acid, $HClO_4(aq)$	~100%
(b) hypochlorous acid, $HClO(aq)$	~1%

Solution

To classify an aqueous solution as a strong or a weak acid we must know the amount of ionization.

(a) Perchloric acid is extensively ionized and is considered a *strong acid.* An aqueous solution is primarily $H^+(aq)$ and $ClO_4^-(aq)$.

(b) Hypochlorous acid is slightly ionized and is considered a *weak acid.* An aqueous solution contains primarily molecules of HClO.

Practice Exercise

Classify each of the following solutions as a strong or a weak Arrhenius base given the degree of dissociation:

(a) magnesium hydroxide, $Mg(OH)_2(aq)$ ~1%
(b) strontium hydroxide, $Sr(OH)_2(aq)$ ~100%

Answers:

(a) weak base (b) strong base

Concept Exercise

Which of the following is a strong Arrhenius acid: $HNO_3(aq)$, $H_2CO_3(aq)$, or $H_3PO_4(aq)$?

Answer: See Appendix G, 14.2.

A CLOSER LOOK Ammonia Solution

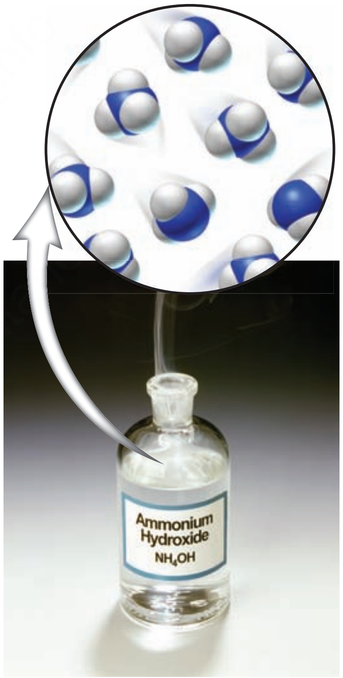
Q: Why is it incorrect to label an ammonia solution as NH_4OH?

We are all familiar with the odor of ammonia, NH_3, when using cleaning products. Ammonia is a general purpose cleaner for many surfaces. Because it does not leave streaks, it is ideal for cleaning glass and stainless steel surfaces. Ammonia is also used for cleaning ovens and removing residue on baking pans.

In an introductory chemistry laboratory, bottles of ammonia solution are routinely labeled ammonium hydroxide, NH_4OH. Ironically, the species NH_4OH does not exist. Ammonium hydroxide, NH_4OH, solutions are prepared by dissolving ammonia gas in water and are more correctly labeled aqueous ammonia, $NH_3(aq)$. However, NH_3 can act as a Brønsted-Lowry base and ionize water by forming an ammonium ion, NH_4^+, and hydroxide ion, OH^-.

Ammonia has a boiling point of −33°C and is a gas at room temperature and atmospheric pressure. Ammonia is an unusual substance, in that it is the only substance—other than water—which is less dense as a solid than it is as a liquid. Thus, solid ammonia cubes float in liquid ammonia, similar to ice cubes floating in water! The concentration of an ammonia solution in the laboratory is usually 6 *M*; household ammonia ranges from 3–6 *M* (5–10%).

Ammonia is a common farm fertilizer. In addition, ammonia vapor is used in smelling salts, and is a respiratory stimulant. Interestingly, ammonia is also a critical ingredient in making methamphetamine, an illegal stimulant. Manufacturing methamphetamine is dangerous because the ingredients, sodium metal and liquid ammonia, are extremely reactive and can explode when the two chemicals are combined.



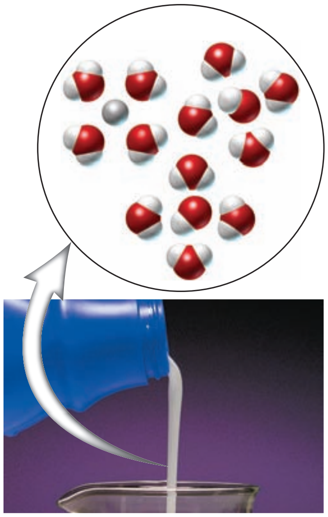
▲ **Ammonium hydroxide, NH_4OH** Ammonia gas dissolved in water is commonly referred to as an *ammonium hydroxide solution.*

A: Ammonia gas dissolved in aqueous solution is most correctly referred to as $NH_3(aq)$. However, ammonia solutions are usually labeled ammonium hydroxide, NH_4OH, because ammonium and hydroxide ions are produced in an ammonia solution.

Neutralization Reactions

An acid neutralizes a base to give a **salt** and water. For example, hydrochloric acid reacts with sodium hydroxide to give the salt sodium chloride and water. The equation for the reaction is

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$



▲ **Stomach Antacid** Milk of Magnesia, $Mg(OH)_2$, is a weak base that can neutralize excess stomach acid.

This reaction produces the aqueous salt NaCl. Different acids give other salts. When we completely neutralize sulfuric acid using aqueous sodium hydroxide, for example, we obtain sodium sulfate. The equation is

$$H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2\,H_2O(l)$$

The hydrogen ions from an acid are neutralized by the hydroxide ions from a base. The hydrogen ions and hydroxide ions combine to form water. Note from the preceding neutralization reaction that Na_2SO_4 is the salt produced.

We can identify the Arrhenius acid and base that produce any given salt because each salt is composed of the cation from the reacting base and the anion from the acid. As an example, we can predict the neutralization reaction that produces the salt potassium acetate, $KC_2H_3O_2$. Since the salt contains potassium, the neutralized base must be potassium hydroxide, KOH. The acetate ion came from acetic acid, $HC_2H_3O_2$. The equation for the reaction is

$$HC_2H_3O_2(aq) + KOH(aq) \longrightarrow KC_2H_3O_2(aq) + H_2O(l)$$

The following example exercise further illustrates the neutralization of acids and bases to give a salt and water.

EXAMPLE EXERCISE 14.3 Neutralization Reactions

Determine the acid and base that produce each of the following salts. Write a balanced equation for the neutralization reaction:

(a) lithium fluoride, $LiF(aq)$ (b) calcium sulfate, $CaSO_4(aq)$

Solution

(a) The salt LiF is produced from the neutralization of the base LiOH and the acid HF. The equation for the reaction is

$$HF(aq) + LiOH(aq) \longrightarrow LiF(aq) + H_2O(l)$$

(b) The salt $CaSO_4$ is produced from the neutralization of the base $Ca(OH)_2$ and the acid H_2SO_4. The equation for the reaction is

$$H_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow CaSO_4(aq) + 2\,H_2O(l)$$

Practice Exercise

Determine the acid and base that produce each of the following salts. Write a balanced equation for the neutralization reaction.

(a) potassium iodide, $KI(aq)$ (b) barium nitrate, $Ba(NO_3)_2(aq)$

Answers:

(a) HI is the acid and KOH is the base:

$$HI(aq) + KOH(aq) \longrightarrow KI(aq) + H_2O(l)$$

(b) HNO_3 is the acid and $Ba(OH)_2$ is the base:

$$2\,HNO_3(aq) + Ba(OH)_2(aq) \longrightarrow Ba(NO_3)_2(aq) + 2\,H_2O(l)$$

Concept Exercise

Which acid and base undergo neutralization to give sodium acetate, $NaC_2H_3O_2$?

Answer: See Appendix G, 14.3.

FEATURED SCIENTIST Svante Arrhenius

Q: What grade was Arrhenius awarded for his brilliant Ph.D. dissertation on the theory of ionic solutions?

Svante Arrhenius (1859–1927) was a child genius who taught himself to read by age three. After graduating from high school at the top of his class, he enrolled at the University of Uppsala in his native Sweden. He majored in chemistry and did research on the passage of electricity through aqueous solutions.

At 22, Arrhenius began his doctoral study in Stockholm. For his graduate thesis, he continued to pursue his interest in the behavior of solutions. In particular, he found it puzzling that sodium chloride solutions conducted electricity, whereas sugar solutions did not. After carefully considering his observations, Arrhenius boldly proposed that a solution of sodium chloride conducts an electric current because it separates into charged particles in solution. Sugar does not conduct electricity because it does not form charged particles in solution.

Arrhenius also noted that the freezing point of water was lowered twice as much for a salt solution as it was for a sugar solution. He explained that when salt dissolves in solution, NaCl separates into sodium ions and chloride ions. When sugar dissolves in solution, it remains as molecules.

Arrhenius was aware, however, that there were difficulties with the concept of charged ions in solution. At the time, the scientific community considered atoms to be indivisible, electrically neutral particles.

In 1884, Arrhenius defended his Ph.D. dissertation on the theory of ionic solutions. After a grueling four-hour defense, he was awarded the lowest possible passing grade. The dissertation committee simply was not ready to accept the idea of ions in solution.

Arrhenius' career stagnated for almost a decade, and his ideas found little support. The ionic theory was eventually championed by Wilhelm Ostwald (1853–1932), a noted German chemist, who invited Arrhenius to work with him. Gradually, evidence accumulated that supported the concept of ions. Most notably, the discovery of the electron in 1897 proved the existence of charged subatomic particles.

After the ionic theory gained credibility, Arrhenius was offered a professorship at the University of Stockholm. In 1899, he published the classic paper "On the Dissociation of Substances in Aqueous Solutions." In 1903, after struggling for 20 years to establish his ideas, Arrhenius was awarded the Nobel Prize in chemistry for his ionic theory of solutions.

◀ **Svante August Arrhenius** (1859–1927)

A: Arrhenius received the lowest possible passing grade; the examining committee simply did not believe in the ionic theory of solutions.

14.3 Brønsted–Lowry Acids and Bases

LEARNING OBJECTIVE

- Identify a Brønsted–Lowry acid and base in a neutralization reaction.

In 1923 the Danish chemist Johannes Brønsted (1879–1947) and the English chemist Thomas Lowry (1874–1936) independently proposed broader definitions than those of Arrhenius for an acid and a base. Whereas Arrhenius defined an acid as a substance that donates a hydrogen ion in water, Brønsted and Lowry defined an acid as a substance that donates a hydrogen ion to any other substance. Since a hydrogen ion is a single proton, a **Brønsted–Lowry acid** is also referred to as a **proton donor.**

Recall the Arrhenius definition of a base as a hydroxide ion donor in water. We know that hydroxide ions neutralize hydrogen ions to form water. There are, however, many substances that neutralize hydrogen ions in addition to hydroxide ions. Brønsted and Lowry proposed that a base is any substance that accepts a hydrogen ion. Thus, a **Brønsted–Lowry base** is a **proton acceptor.**

Although the Brønsted–Lowry definitions of an acid and a base are a bit different from the Arrhenius definitions, an acid and a base still neutralize each other.

Helpful Hint
Acids and Bases

A simple definition of an acid is a substance that releases hydrogen ions in aqueous solution, and a base is defined as a substance that releases hydroxide ions in aqueous solution. These two Arrhenius definitions suffice most of the time.

However, there are acid–base neutralization reactions in which the hydroxide ion is not involved. For example, sodium carbonate, Na_2CO_3, can act as a base to neutralize an acid. The term Brønsted-Lowry base is used to describe a substance in a neutralization reaction, which involves a base that is not hydroxide.

There is yet a more sophisticated definition of an acid and base. A Lewis acid is defined as a substance that can accept an electron pair, while a Lewis base can donate an electron pair. An example of a Lewis acid is $AlCl_3$, and a Lewis base is NH_3.

The following reactions illustrate the neutralization of a Brønsted–Lowry base by hydrochloric acid:

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$
$$HCl(aq) + NH_3(aq) \longrightarrow NH_4Cl(aq)$$
$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

In the first equation, aqueous NaOH accepts a proton. Therefore, according to the Brønsted–Lowry definition, it is a base. In the second equation, aqueous NH_3 accepts a proton. It is also a base. As the third equation shows, even water can act as a Brønsted–Lowry base because it can accept a hydrogen ion.

Unlike Arrhenius acids and bases that are defined by hydrogen ions or hydroxide ions dissolved in water, Brønsted–Lowry acids and bases depend on a particular reaction. For example, aqueous $NaHCO_3$ can act as a *base* by accepting a proton. But in a different reaction, aqueous $NaHCO_3$ can act as an *acid* by donating a proton. The following examples illustrate $NaHCO_3$ acting as both an acid and a base:

$$HCl(aq) + NaHCO_3(aq) \longrightarrow NaCl(aq) + H_2CO_3(aq)$$
$$NaOH(aq) + NaHCO_3(aq) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

In the first reaction, $NaHCO_3$ accepts a proton from HCl and therefore acts as a Brønsted–Lowry base. In the second reaction, $NaHCO_3$ donates a proton to NaOH and therefore acts as a Brønsted–Lowry acid. A substance that is capable of both accepting and donating a proton is said to be **amphiprotic.** According to the Brønsted–Lowry theory, water is amphiprotic. That is, H_2O can act as an acid and donate a proton; H_2O can also act as a base and accept a proton to become the hydronium ion, H_3O^+. We can show the amphiprotic nature of water as follows:

$$\underset{\substack{\text{proton donor}\\ \textit{acid}}}{H_2O(l)} + \underset{\substack{\text{proton acceptor}\\ \textit{base}}}{H_2O(l)} \longrightarrow H_3O^+(aq) + OH^-(aq)$$

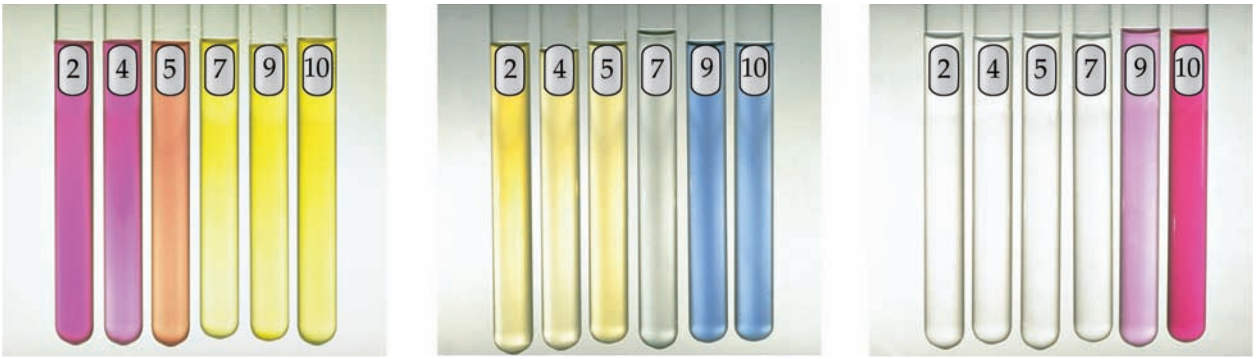
LEARNING OBJECTIVE

- Describe the color of a solution with a given pH and a drop of phenolphthalein, methyl red, or bromthymol blue indicator.

14.4 Acid–Base Indicators

We mentioned in Section 14.1 that litmus paper can be used to show whether a solution is acidic or basic. In addition, other paper test strips are available that indicate the pH of a solution. These test strips are permeated with substances that change color depending on the pH of the test solution. These strips contain substances extracted from plants. Some of the plant extracts come from grape juice, red cabbage, and flower petals such as those of roses, violets, and orchids.

A substance that is sensitive to a change in pH by exhibiting a different color is referred to as an **acid–base indicator.** Figure 14.2 illustrates the range in color of these indicators. Table 14.4 summarizes the pH range for each indicator.



▲ **Figure 14.2 pH and Acid–Base Indicator Color** The color of three acid–base indicators at different pH values. Methyl red (left) changes from red to yellow at pH 5. Bromthymol blue (middle) changes from yellow to blue at pH 7. Phenolphthalein (right) changes from colorless to pink at pH 9.

TABLE 14.4 Acid–Base Indicators

Indicator	Color Change	Color
methyl red	pH ~5	below pH 5, red above pH 5, yellow
bromthymol blue	pH ~7	below pH 7, yellow above pH 7, blue
phenolphthalein	pH ~9	below pH 9, colorless above pH 9, pink

If a solution has a pH above 5, the indicator methyl red is yellow. If a solution has a pH below 5, the indicator is red. At a pH of 5, methyl red appears orange because it contains both the yellow and red forms of the indicator.

If a solution has a pH of 9 or above, the indicator phenolphthalein is pink. If a solution has a pH below 9, the indicator is colorless. Although there are many acid–base indicators available in the laboratory, phenolphthalein is the only common indicator that changes to colorless. This practical consideration and ease of observation makes phenolphthalein the indicator of choice for many applications.

The following example exercise illustrates the relationship of the color of an acid–base indicator and the pH of a given solution.

EXAMPLE EXERCISE 14.4 Acid–Base Indicators

State the color of the acid–base indicator in each of the following solutions:

(a) A solution at pH 3 contains a drop of methyl red.
(b) A solution at pH 8 contains a drop of bromthymol blue.
(c) A solution at pH 11 contains a drop of phenolphthalein.

Solution

Refer to Table 14.4 to determine the color of each solution.

(a) A pH 3 solution containing methyl red indicator is *red.*
(b) A pH 8 solution containing bromthymol blue indicator is *blue.*
(c) A pH 11 solution containing phenolphthalein indicator is *pink.*

Practice Exercise

State the pH at which each of the following acid–base indicators changes color:

(a) methyl red
(b) bromthymol blue
(c) phenolphthalein

Answers: (a) 5 (b) 7 (c) 9

Concept Exercise

Which common acid–base indicator is colorless in acidic and neutral solutions?

Answer: See Appendix G, 14.4.

14.5 Acid–Base Titrations

Vinegar tastes sour because it contains acetic acid. To find the molar concentration of acetic acid in vinegar we analyze the vinegar sample using a base solution of known concentration. In the laboratory, we can analyze acetic acid by a titration method. We use this **titration** process to deliver a measured volume of solution. To analyze the acetic acid in a vinegar sample we can titrate a sample by delivering a measured volume of sodium hydroxide solution. We can use phenolphthalein as an indicator to signal when we have neutralized all the acetic acid.

LEARNING OBJECTIVES

- Solve problems given acid–base titration data.
- Calculate the mass percent concentration of a solution given the molarity.

After the acetic acid is neutralized, an extra drop of NaOH increases the pH dramatically and the phenolphthalein indicator changes from colorless to pink. When the indicator changes color, we stop adding NaOH base and cease the titration. This turning point in the titration is the **endpoint.** Figure 14.3 illustrates the experimental procedure.

Consider the following titration of a vinegar sample for acetic acid. A 10.0-mL sample of vinegar containing acetic acid requires 37.55 mL of 0.223 *M* NaOH. The balanced equation for the reaction is

$$\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{NaOH}(aq) \longrightarrow \text{NaC}_2\text{H}_3\text{O}_2(aq) + \text{H}_2\text{O}(l)$$

This is a solution stoichiometry problem (Section 13.11). To find the molarity of the acetic acid, we must first find the moles of NaOH. The volume of NaOH is 37.55 mL, and the molar concentration is 0.223 M. The molarity of the NaOH can be written as the unit factor 0.223 mol NaOH/1000 mL solution.

$$37.55\ \cancel{\text{mL solution}} \times \frac{0.223\ \text{mol NaOH}}{1000\ \cancel{\text{mL solution}}} = 0.00837\ \text{mol NaOH}$$

Next, we find the moles of acetic acid in the acid solution. From the balanced equation we see that 1 mol of NaOH base neutralizes 1 mol of $\text{HC}_2\text{H}_3\text{O}_2$.

$$0.00837\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{mol HC}_2\text{H}_3\text{O}_2}{1\ \cancel{\text{mol NaOH}}} = 0.00837\ \text{mol HC}_2\text{H}_3\text{O}_2$$

Finally, we can calculate the molarity of the acetic acid. The volume of the $\text{HC}_2\text{H}_3\text{O}_2$ solution is 10.0 mL.

$$\begin{aligned}\frac{0.00837\ \text{mol HC}_2\text{H}_3\text{O}_2}{10.0\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\ \text{L solution}} &= \frac{0.837\ \text{mol HC}_2\text{H}_3\text{O}_2}{1\ \text{L solution}}\\ &= 0.837\ M\ \text{HC}_2\text{H}_3\text{O}_2\end{aligned}$$

Now let's express the molar concentration of the acetic acid as a mass percent (m/m %) concentration. Given that the density of the vinegar is 1.01 g/mL and that the molar mass of acetic acid is 60.06 g/mol, we can proceed as follows:

$$\begin{aligned}&\frac{0.837\ \cancel{\text{mol HC}_2\text{H}_3\text{O}_2}}{1000\ \cancel{\text{mL solution}}} \times \frac{60.06\ \text{g HC}_2\text{H}_3\text{O}_2}{1\ \cancel{\text{mol HC}_2\text{H}_3\text{O}_2}} \times \frac{1\ \cancel{\text{mL Solution}}}{1.01\ \text{g solution}} \times 100\%\\ &= \frac{0.0498\ \text{g HC}_2\text{H}_3\text{O}_2}{1\ \text{g solution}} \times 100\% = 4.98\%\ \text{HC}_2\text{H}_3\text{O}_2\end{aligned}$$

▲ **Figure 14.3 Titration of Acetic Acid with Sodium Hydroxide**
(a) The flask contains a sample of acetic acid and a drop of phenolphthalein indicator.
(b) The buret delivers a measured volume of sodium hydroxide solution into the flask.
(c) The titration is complete when the solution attains a permanent pink color.

Therefore, the concentration of acetic acid in the vinegar solution is 0.837 *M*, which is equivalent to 4.98% $HC_2H_3O_2$. Typically, the label on a bottle of vinegar reads ”diluted in water to 5% acidity.”

Titrations can also be performed on basic solutions in the same way, but an acid is added from the buret instead of a base. We can calculate the concentration of ammonia in a household cleaning solution by titrating a sample. The following example exercise illustrates the analysis of an ammonia solution.

EXAMPLE EXERCISE 14.5 Acid–Base Titration of Ammonia

If 25.30 mL of 0.277 *M* HCl is used to titrate 10.0 mL of aqueous ammonia to a methyl red endpoint, what is the molarity of the ammonia? When we write aqueous ammonia as NH_4OH, the balanced equation is

$$HCl(aq) + NH_4OH(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)$$

Solution

To calculate the molarity of the NH_4OH, we must find the moles of NH_4OH. From the balanced equation, we see that 1 mol of HCl neutralizes 1 mol of NH_4OH. Since the acid solution is 0.277 *M* HCl, we have the unit factor 0.277 mol HCl/1000 mL solution.

$$25.30\ \cancel{\text{mL solution}} \times \frac{0.277\ \cancel{\text{mol HCl}}}{1000\ \cancel{\text{mL solution}}} \times \frac{1\ \text{mol } NH_4OH}{1\ \cancel{\text{mol HCl}}} = 0.00701\ \text{mol } NH_4OH$$

We can summarize the conversion as follows:

We obtain the molar concentration by dividing the moles of NH_4OH by the volume of aqueous ammonia titrated, 10.0 mL.

$$\frac{0.00701\ \text{mol } NH_4OH}{10.0\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\ \text{L solution}} = \frac{0.701\ \text{mol } NH_4OH}{1\ \text{L solution}}$$
$$= 0.701\ M\ NH_4OH$$

Practice Exercise

If 38.30 mL of 0.250 *M* NaOH is used to titrate 25.0 mL of phosphoric acid, what is the molarity of the acid? The balanced equation is

$$H_3PO_4(aq) + 3\ NaOH(aq) \longrightarrow Na_3PO_4(aq) + 3\ H_2O(l)$$

Answer: 0.128 *M*

Concept Exercise

Which of the following describes the titration of $H_3PO_4(aq)$ with $NaOH(aq)$: strong acid with strong base, strong acid with weak base, or weak acid with strong base?

Answer: See Appendix G, 14.5.

Let’s try another type of problem and find the volume of base required to neutralize an acid. For example, we can find the volume of base needed to neutralize sulfuric acid in a battery acid sample. The following example exercise illustrates the analysis of battery acid by a titration method.

EXAMPLE EXERCISE 14.6 Acid–Base Titration of Battery Acid

A 10.0-mL sample of 0.555 M battery acid, H_2SO_4, is titrated with 0.223 M NaOH. What volume of sodium hydroxide is required for the titration?

$$H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2\,H_2O(l)$$

Solution

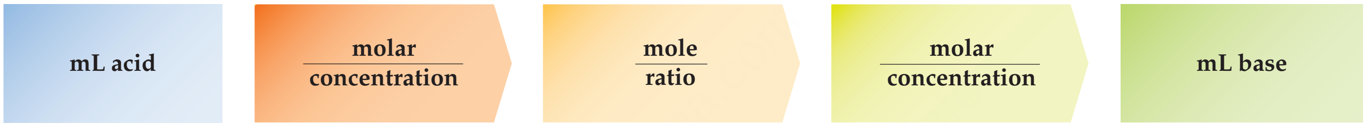
STEP 1: The unit asked for in the answer is mL NaOH.

STEP 2: The given value is 10.0 mL H_2SO_4.

STEP 3: We apply unit conversion factors to cancel units. In this example, we use the molar concentration of an acid (0.555 mol H_2SO_4 per 1000 mL H_2SO_4) and the molar concentration of a base (0.223 mol NaOH per 1000 mL NaOH) as unit factors. From the balanced equation, we see that 1 mol H_2SO_4 = 2 mol NaOH.

$$10.0\ \cancel{\text{mL solution}} \times \frac{0.555\ \cancel{\text{mol } H_2SO_4}}{1000\ \cancel{\text{mL solution}}} \times \frac{2\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol} H_2SO_4}} \times \frac{1000\ \text{mL NaOH}}{0.223\ \cancel{\text{mol NaOH}}} = 49.8\ \text{mL NaOH}$$

We can summarize the overall solution as follows:



Practice Exercise

A 25.0-mL sample of hydrochloric acid is titrated with 0.125 M $Ba(OH)_2$. If 50.0 mL of barium hydroxide is required for the titration, what is the molar concentration of the acid? The balanced equation for the reaction is

$$2\,HCl(aq) + Ba(OH)_2(aq) \longrightarrow BaCl_2(aq) + 2\,H_2O(l)$$

Answer: 0.500 M HCl

Concept Exercise

Which of the following describes the titration of $HCl(aq)$ with $Ba(OH)_2(aq)$: strong acid with strong base, strong acid with weak base, or weak acid with strong base?

Answer: See Appendix G, 14.6.

14.6 Acid–Base Standardization

LEARNING OBJECTIVES

- Describe a standard solution of acid or base.
- Solve problems that involve standardization of an acid or base.

A **standard solution** is a solution in which the concentration is known precisely, that is, three or four significant digits. For example, the concentration of NaOH used in the acetic acid titration was 0.223 M.

Chemists use standard solutions to repetitively analyze substances. For instance, a chemist may use a standard solution to analyze the neutralizing capacity of antacid tablets, or the tartness of soft drink flavors. Standard solutions are routinely used in the chemical industry to assure quality. Chemical samples are randomly selected and analyzed using a standard solution. This procedure is referred to as *quality control* (QC) or *quality assurance* (QA).

To standardize a solution of acid, chemists use a weighed quantity of a solid base. To standardize hydrochloric acid, for example, we can use solid sodium carbonate, Na_2CO_3. Let's find the molarity of hydrochloric acid if 25.50 mL of HCl solution is required to neutralize 0.375 g of Na_2CO_3. The balanced equation for the reaction of Na_2CO_3 dissolved in water is

$$2\,HCl(aq) + Na_2CO_3(aq) \longrightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

To calculate the molarity of the hydrochloric acid, we first find the number of moles of HCl. From the balanced equation we note that 2 mol of HCl react with 1 mol of Na_2CO_3 (105.99 g/mol).

$$0.375\ \cancel{\text{g Na}_2\text{CO}_3} \times \frac{1\ \cancel{\text{mol Na}_2\text{CO}_3}}{105.99\ \cancel{\text{g Na}_2\text{CO}_3}} \times \frac{2\ \text{mol HCl}}{1\ \cancel{\text{mol Na}_2\text{CO}_3}} = 0.00708\ \text{mol HCl}$$

To obtain the molarity of the acid, we divide the moles of HCl by the 25.50 mL of HCl required to neutralize the sodium carbonate.

$$\frac{0.00708\ \text{mol HCl}}{25.50\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\ \text{L solution}} = \frac{0.277\ \text{mol HCl}}{1\ \text{L solution}}$$
$$= 0.277\ M\ \text{HCl}$$

To standardize a solution of base, chemists can use a weighed quantity of a solid acid. For example, we can use crystals of oxalic acid, $H_2C_2O_4$, to standardize aqueous sodium hydroxide. We first dissolve a weighed sample of $H_2C_2O_4$ in water and then neutralize it with a measured volume of a base. The following example exercise illustrates how to determine the concentration of a standard NaOH solution.

EXAMPLE EXERCISE 14.7 Acid–Base Standardization of a Base

What is the molarity of a sodium hydroxide solution if 32.15 mL of NaOH is required to neutralize 0.424 g of oxalic acid, $H_2C_2O_4$ (90.04 g/mol)?

$$H_2C_2O_4(aq) + 2\ NaOH(aq) \longrightarrow Na_2C_2O_4(aq) + 2\ H_2O(l)$$

Solution

To calculate the molarity of the NaOH, we find the number of moles of $H_2C_2O_4$. The mass of $H_2C_2O_4$ to be neutralized is 0.424 g, and the molar mass is 90.04 g/mol. From the balanced equation we see that 2 mol of NaOH reacts with 1 mol of $H_2C_2O_4$.

$$0.424\ \cancel{\text{g H}_2\text{C}_2\text{O}_4} \times \frac{1\ \cancel{\text{mol H}_2\text{C}_2\text{O}_4}}{90.04\ \cancel{\text{g H}_2\text{C}_2\text{O}_4}} \times \frac{2\ \text{mol NaOH}}{1\ \cancel{\text{mol H}_2\text{C}_2\text{O}_4}} = 0.00942\ \text{mol NaOH}$$

We can summarize the conversion as follows:

We can obtain the molar concentration of the base from the moles of NaOH divided by the volume of base required to neutralize the oxalic acid.

$$\frac{0.00942\ \text{mol NaOH}}{32.15\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\ \text{L solution}} = \frac{0.293\ \text{mol NaOH}}{1\ \text{L solution}}$$
$$= 0.293\ M\ \text{NaOH}$$

Practice Exercise

If an unknown sample of oxalic acid, $H_2C_2O_4$, is neutralized by 33.50 mL of 0.293 M NaOH, what is the mass of the sample? Refer to the previous balanced chemical equation for the neutralization of oxalic acid with sodium hydroxide.

Answer: 0.442 g $H_2C_2O_4$

Concept Exercise

What color is the phenolphthalein indicator at the endpoint of the titration of oxalic acid with sodium hydroxide?

Answer: See Appendix G, 14.7.

Molar Mass of a Solid Acid or Solid Base

One application of a standard solution is to identify a crystalline solid acid by determining its molar mass. For example, using a standard base solution, we could identify an unknown solid acid as citric acid by determining its molar mass. We begin by dissolving crystals of the solid acid in water and then neutralizing the acid using a standard base solution. Given the mass of solid acid and the volume of standard base, we can calculate the molar mass of the acid.

The following example exercise illustrates the determination of molar mass for citric acid.

EXAMPLE EXERCISE 14.8 Molar Mass of an Acid

Citric acid, abbreviated H_3Cit, gives a soft drink its tart taste. If 36.10 mL of 0.293 *M* NaOH neutralizes a 0.677-g sample of citric acid, what is the molar mass of citric acid? The balanced equation for the reaction is

$$H_3Cit(aq) + 3\ NaOH(aq) \longrightarrow Na_3Cit(aq) + 3\ H_2O(l)$$

Solution

The given volume of NaOH solution is 36.10 mL. Since the molar concentration is 0.293 *M*, we write the unit factor as 0.293 mol NaOH/1000 mL solution.

$$36.10\ \cancel{\text{mL solution}} \times \frac{0.293\ \text{mol NaOH}}{1000\ \cancel{\text{mL solution}}} = 0.0106\ \text{mol NaOH}$$

From the balanced chemical equation, we see that 3 mol of NaOH reacts with 1 mol of H_3Cit. Thus, the number of moles of citric acid is

$$0.0106\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{mol}\ H_3Cit}{3\ \cancel{\text{mol NaOH}}} = 0.00353\ \text{mol}\ H_3Cit$$

We can summarize the conversion as follows:

Since the molar mass is defined as grams per mole, we have

$$\frac{0.677\ \text{g}\ H_3Cit}{0.00353\ \text{mol}\ H_3Cit} = 192\ \text{g/mol}$$

From titration data, we calculated the molar mass of citric acid as 192 g/mol. Given that the chemical formula of citric acid is $H_3C_6H_5O_7$, we can verify the molar mass by adding the atomic masses to obtain 192.14 g/mol.

Practice Exercise

Benzoic acid is a common food preservative, and its molecular formula is $HC_7H_5O_2$. If 22.55 mL of 0.293 *M* NaOH neutralizes a 0.806-g sample of benzoic acid, what is the molar mass of the acid? The balanced equation for the reaction is

$$HC_7H_5O_2(aq) + NaOH(aq) \longrightarrow NaC_7H_5O_2(aq) + H_2O(l)$$

Answer: 122 g/mol

Concept Exercise

What color is the phenolphthalein indicator at the endpoint of the titration of benzoic acid with sodium hydroxide?

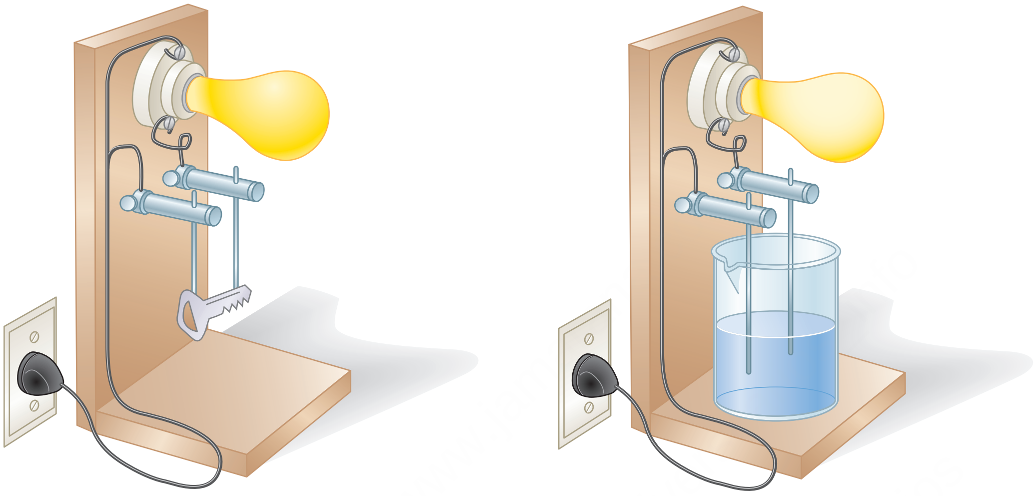
Answer: See Appendix G, 14.8.

14.7 Ionization of Water

LEARNING OBJECTIVES

- Show the ionization constant equation for water, K_w.
- Calculate the molar hydroxide ion concentration given the molar hydrogen ion concentration.

We know that metals are good conductors of electricity. What would happen if we tested the electrical conductivity of a key? We can predict that the key conducts electricity because it is made of metal. In Figure 14.4a, we see that the key completes the circuit in the conductivity apparatus and that the bulb lights up. The explanation is as follows. Electricity is defined as the flow of electrons. Recall from the electron sea model described in Section 11.5 that the valence electrons in metals are free to flow. An electric current can push electrons from one atom to another. Thus, electricity flows from one end of the metal key to the other as electrons move from atom to atom.

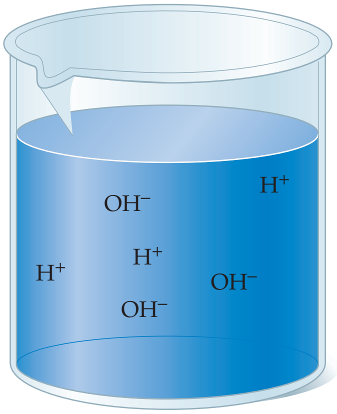

(a) Strong conductor

(b) Weak conductor

◀ **Figure 14.4 Conductivity Apparatus** (a) Metals are good conductors of electricity. A metal key completes the electrical circuit and the light bulb shows a bright light. (b) Pure water is a very poor conductor. It does, however, very weakly conduct electricity and the light bulb shows a dim light.

What happens when we test the electrical conductivity of pure water? In Figure 14.4b, we see that pure water is a very poor conductor of electricity. Evidence suggests that it is difficult to push electrons from one water molecule to another. How then do we explain the observation that some electricity flows in pure water? We explain this observation by suggesting that there are ions present in pure water.

Arrhenius found that salt solutions are good conductors of electricity. He correctly concluded that *ions* in the solution are responsible for electrical conductivity. Since water is a very poor conductor, we can conclude that only a few ions are present in pure water (Figure 14.5).



▲ **Figure 14.5 Ionization of Water** Pure water ionizes to give a few hydrogen ions and hydroxide ions. Water is a very weak conductor of electricity because only 1 molecule in about 500,000,000 forms ions.

We know that water molecules in the liquid state are free to move about and collide with one another. Although it happens only rarely, on occasion two water molecules collide with sufficient energy for their bonds to break apart. This bond breaking produces a hydronium ion, H_3O^+, and a hydroxide ion, OH^-. Moreover, this occurrence is a dynamic process. While some molecules are breaking apart, other ions are combining to form water molecules. At any given moment, only 1 water molecule in 500 million molecules is ionized. Apparently, however, these few ions in water are sufficient to conduct a very weak electric current. The chemical equation for the collision reaction is

$$H_2O(l) + H_2O(l) \longrightarrow \underset{\textit{hydronium ion}}{H_3O^+(aq)} + \underset{\textit{hydroxide ion}}{OH^-(aq)}$$

Alternatively, we can simplify the reaction by writing the ionization of water as

$$H_2O(l) \longrightarrow \underset{\textit{hydrogen ion}}{H^+(aq)} + \underset{\textit{hydroxide ion}}{OH^-(aq)}$$

The concentration of hydrogen ions in pure water is 1.0×10^{-7} mol/L at 25 °C. If we know the concentration of hydrogen ions, we also know the concentration of hydroxide ions. Since the ionization of water gives one H^+ and one OH^- for each molecule of water that ionizes, the two concentrations are equal. Therefore, the concentration of OH^- must also be 1.0×10^{-7} mol/L at 25 °C. Moreover, the molar concentration of H^+ multiplied by the molar concentration of OH^- equals a constant. This product is called the **ionization constant of water** (symbol K_w).

Let's calculate the value of K_w at 25 °C. For convenience, we use brackets to symbolize molar concentration. Hence, $[H^+]$ is the symbol for the molar concentration of hydrogen ions. We calculate a value for K_w as follows:

If $$[H^+] = 1.0 \times 10^{-7}$$

then $$[OH^-] = 1.0 \times 10^{-7}$$

and $$[H^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

Since we add exponents when multiplying exponential numbers, we have

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$
$$K_w = 1.0 \times 10^{-14} \text{ (at 25 °C)}$$

Notice that $[H^+]$ and $[OH^-]$ are inversely proportional. That is, if $[H^+]$ increases, $[OH^-]$ decreases, so that the constant K_w remains the same.

We should emphasize that water is neutral even though it contains small amounts of both H^+ and OH^-. Moreover, every aqueous solution has hydrogen ions and hydroxide ions. Even a hydrochloric acid solution contains H^+ and OH^-. Although an aqueous HCl solution has a high concentration of H^+, it has a few OH^- as well. In every aqueous solution, the product of $[H^+]$ and $[OH^-]$ is equal to the ionization constant of water.

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$

If an aqueous HCl solution is 0.1 *M*, we can calculate its hydroxide ion concentration.

$$[0.1][OH^-] = 1.0 \times 10^{-14}$$

and

$$[OH^-] = 1.0 \times 10^{-13}$$

Although $[OH^-]$ is only 0.000 000 000 0001 *M*, the hydrochloric acid solution contains a trace of hydroxide ion!

To summarize, in an acidic aqueous solution the $[H^+]$ is greater than $[OH^-]$. In a basic aqueous solution, $[H^+]$ is less than $[OH^-]$. In a neutral aqueous solution, $[H^+]$ is equal to $[OH^-]$. Example Exercise 14.9 illustrates the concept of hydrogen and hydroxide ions in aqueous solutions.

EXAMPLE EXERCISE 14.9 The Ionization Constant, K_w

Given the following hydrogen ion concentrations in an aqueous solution at 25°C, what is the molar concentration of hydroxide ion?

(a) $[H^+] = 1.4 \times 10^{-6}$ (b) $[H^+] = 5.2 \times 10^{-11}$

Solution

From the ionization constant of water, we know that $[H^+][OH^-] = 1.0 \times 10^{-14}$.

(a) If the $[H^+]$ is 1.4×10^{-6}, we can calculate $[OH^-]$ as follows:

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$
$$[1.4 \times 10^{-6}][OH^-] = 1.0 \times 10^{-14}$$

Dividing both sides of the equation by 1.4×10^{-6}, we have

$$[OH^-] = 7.1 \times 10^{-9}$$

(b) If the $[H^+]$ is 5.2×10^{-11}, we can calculate $[OH^-]$ as follows:

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$
$$[5.2 \times 10^{-11}][OH^-] = 1.0 \times 10^{-14}$$

Dividing both sides of the equation by 5.2×10^{-11}, we have

$$[OH^-] = 1.9 \times 10^{-4}$$

Practice Exercise

Given the following hydroxide ion concentrations in an aqueous solution, what is the molar concentration of hydrogen ion?

(a) $[OH^-] = 7.5 \times 10^{-4}$ (b) $[OH^-] = 2.1 \times 10^{-10}$

Answers: (a) $[H^+] = 1.3 \times 10^{-11}$ (b) $[H^+] = 4.8 \times 10^{-5}$

Concept Exercise

Explain why distilled water is neutral even though it contains hydrogen ions and hydroxide ions.

Answer: See Appendix G, 14.9.

14.8 The pH Concept

LEARNING OBJECTIVE

- Describe the relationship between pH and molar hydrogen ion concentration.

In Section 14.1, we introduced the term *pH*. Recall that pure water is neutral and has a pH of 7. We also know that an acid has a pH of less than 7. As the pH value *decreases*, the solution becomes more acidic. Thus, a pH of 3 is more acidic than a pH of 4. As the pH value *increases*, the solution becomes more basic. Thus, a pH of 11 is more basic than a pH of 10.

▶ **Digital pH Meter** The digital pH meter measures the hydrogen ion concentration in the beaker with the glass electrode (shown). Pure boiled water has a hydrogen ion concentration of 1.0×10^{-7} *M*, which corresponds to a pH of 7.00.

Converting from $[H^+]$ to pH

In an aqueous solution, the hydrogen ion concentration can vary from more than 1 *M* to less than 0.000 000 000 000 01 *M*. The pH scale is a convenient way to express this broad range of hydrogen ion concentrations. The pH scale expresses the molar hydrogen ion concentration, $[H^+]$, as a power of 10.

In other words, the pH scale uses powers of 10 to express the amount of acid in a solution. A solution with a pH of 3 is 10 times more acidic than a solution with a pH of 4. Conversely, a solution with a pH of 11 is 10 times less acidic than a solution with a pH of 10. In mathematical terms, the **pH** is the negative logarithm (log) of the molar hydrogen ion concentration.

$$\text{pH} = -\log\,[H^+]$$

▲ **Scientific Calculator**

For example, if the molar hydrogen ion concentration, $[H^+]$, is 0.1 M, then

$$pH = -\log 0.1$$

We do not need to understand logarithms to solve pH problems. All we need to know is that if we express the hydrogen ion concentration as a power of 10, the exponent equals the logarithm. Similarly, you may be familiar with the Richter scale, which expresses the severity of an earthquake. The Richter scale is a logarithmic scale and an earthquake of magnitude 6 is 10 times more severe than an earthquake of magnitude 5.

We can write 0.1 *M* hydrogen ion concentration as a power of 10 (that is, 10^{-1}) and find the pH as follows:

$$\begin{aligned} pH &= -\log 10^{-1} \\ &= -(-1) = 1 \end{aligned}$$

Thus, the pH of a solution with a hydrogen ion concentration of 0.1 *M* is 1.

The following example exercise provides practice in finding the pH of an aqueous solution given the molar hydrogen ion concentration.

▲ **Vinegar** Acetic acid is found in vinegar, which gives vinegar a sour taste.

EXAMPLE EXERCISE 14.10 Converting $[H^+]$ to pH

Calculate the pH of the following solutions given the molar hydrogen ion concentration:

(a) vinegar, $[H^+] = 0.001\ M$ (b) antacid, $[H^+] = 0.000\,000\,001\ M$

Solution

(a) The pH of vinegar is equal to $-\log [H^+]$.

$$\begin{aligned} pH &= -\log 0.001 = -\log 10^{-3} \\ &= -(-3) = 3 \end{aligned}$$

The sour taste of vinegar indicates that an acid is present. The acid in vinegar is acetic acid, $HC_2H_3O_2$ and the pH is approximately 3.

(b) The pH of the antacid is equal to $-\log [H^+]$.

$$\begin{aligned} pH &= -\log 0.000\,000\,001 = -\log 10^{-9} \\ &= -(-9) = 9 \end{aligned}$$

Practice Exercise

Calculate the pH for apple juice if the molar hydrogen ion concentration is 0.0001 M.

Answer: pH = 4

Concept Exercise

Classify apple juice as weakly acidic, neutral, or weakly basic given a pH of 4.

Answer: See Appendix G, 14.10.

Converting from pH to $[H^+]$

We can calculate the hydrogen ion concentration given the pH of a solution. Let's rearrange the definition of pH to express the molar hydrogen ion concentration. The $[H^+]$ is equal to 10 raised to a negative pH value. That is,

$$[H^+] = 10^{-pH}$$

If milk has a pH of 6, then the $[H^+]$ can be expressed as

$$[H^+] = 10^{-6} = 0.000\,001\ M$$

Example Exercise 14.11 further illustrates the conversion of a given pH to a molar hydrogen ion concentration.

EXAMPLE EXERCISE 14.11 Converting pH to $[H^+]$

Calculate the molar hydrogen ion concentration given the pH of the following solutions:

(a) lemon juice, pH 2 (b) tomato juice, pH 5

Solution

(a) The hydrogen concentration of lemon juice is equal to 10 raised to the negative value of the pH. Since the pH is 2,

$$[H^+] = 10^{-pH}$$
$$= 10^{-2} = 0.01\ M$$

(b) If the pH of tomato juice is 5, $[H^+]$ is expressed as

$$[H^+] = 10^{-pH}$$
$$= 10^{-5} = 0.000\ 01\ M$$

Notice that lemon juice and tomato juice differ by three pH units. Thus, lemon juice is 1000 times more acidic than tomato juice.

Practice Exercise

Calculate the hydrogen ion concentration for milk of magnesia if the pH is 10.

Answer: $[H^+] = 0.000\ 000\ 000\ 1\ M$ ($1 \times 10^{-10}\ M$)

Concept Exercise

Classify milk of magnesia as weakly acidic, neutral, or weakly basic given a pH of 10.

Answer: See Appendix G, 14.11.

▲ **Lemon Juice** Citric acid is found in lemon juice, which gives it a sour taste.

14.9 Advanced pH Calculations

LEARNING OBJECTIVE

- Calculate pH values and molar hydrogen ion concentrations.

We defined pH as an exponential way of expressing the hydrogen ion concentration, concentrations. but we considered only whole-number pH values. Many chemical reactions, however, must be carefully controlled to a fraction of a pH unit. Biochemical reactions, for example, are extremely sensitive to small pH changes. In fact, the pH of our blood must be maintained within the narrow range of 7.3–7.5. There is an elaborate buffering system that accomplishes this task, which involves carbonates, phosphates, and proteins.

Converting from $[H^+]$ to pH

Recall the mathematical definition of pH. That is, pH is the negative logarithm of the molar hydrogen ion concentration.

$$pH = -\log [H^+]$$

Suppose we wish to express the pH of a solution having a hydrogen ion concentration of 0.00015 M. The pH expression is

$$pH = -\log 0.00015$$
$$= -\log 1.5 \times 10^{-4}$$

We can easily obtain the logarithm for a number using a scientific calculator. If we enter 1.5×10^{-4} and touch the log key, the display shows −3.82. Since pH is the negative logarithm, we must change the sign to 3.82.

$$pH = -(-3.82)$$
$$= 3.82$$

Note that the number of significant digits in the given value equals the number of decimal places in the log expression. In this example, there are two significant digits and two decimal places in the log value. Example Exercise 14.12 provides practice in finding the pH of an aqueous solution given the molar hydrogen ion concentration.

▲ **Grape Juice** Tartaric acid is found in grape juice, which gives it a mild sour taste.

EXAMPLE EXERCISE 14.12 Converting $[H^+]$ to pH

Calculate the pH of the following solutions given the molar hydrogen ion concentration:

(a) stomach acid, $[H^+] = 0.020\ M$ (b) blood, $[H^+] = 0.000\ 000\ 048\ M$

Solution

(a) We can calculate the pH of stomach acid as follows:

$$\begin{aligned} pH &= -\log 0.020 \\ &= -(-1.70) \\ &= 1.70 \end{aligned}$$

The pH of stomach acid ranges from 1 to 3; hence, this sample is normal.

(b) We can calculate the pH of blood as follows:

$$\begin{aligned} pH &= -\log 0.000\ 000\ 048 \\ &= -\log 4.8 \times 10^{-8} \\ &= -(-7.32) \\ &= 7.32 \end{aligned}$$

The pH of blood ranges from 7.3 to 7.5; hence, this sample is normal.

Practice Exercise

Calculate the pH of grape juice if the hydrogen ion concentration is 0.000 089 M.

Answer: pH = 4.05

Concept Exercise

Classify grape juice as weakly acidic, neutral, or weakly basic given a pH of 4.05.

Answer: See Appendix G, 14.12.

Converting from pH to $[H^+]$

Previously, we expressed $[H^+]$ by raising 10 to the negative pH value.

$$[H^+] = 10^{-pH}$$

If orange juice has a pH of 2.75, then $[H^+]$ can be expressed as

$$[H^+] = 10^{-2.75}$$

To find the number that corresponds to the fractional exponent, we obtain the inverse logarithm using a scientific calculator. If we enter −2.75 and touch the inverse log key, the display shows 0.0018, or 1.8×10^{-3}.

$$[H^+] = 10^{-2.75} = 0.0018\ M$$

The following example exercise provides further practice in converting a given pH to a molar hydrogen ion concentration.

EXAMPLE EXERCISE 14.13 Converting pH to $[H^+]$

Calculate the molar hydrogen ion concentration given the pH of the following solutions:

(a) acid rain, pH 3.68 (b) seawater, pH 7.85

Solution

(a) If acid rain has a pH of 3.68, we can find the $[H^+]$ as follows:

$$[H^+] = 10^{-3.68}$$

Using the inverse log key on a scientific calculator, we have

$$\begin{aligned} [H^+] &= 2.1 \times 10^{-4} \\ &= 0.000\ 21\ M \end{aligned}$$

(b) If seawater has a pH of 7.85, we can find the $[H^+]$ as follows:

$$\begin{aligned}[H^+] &= 10^{-7.85} \\ &= 1.4 \times 10^{-8} \\ &= 0.000\,000\,014\ M\end{aligned}$$

Practice Exercise

Calculate the hydrogen ion concentration in wine if the pH is 3.25.

Answer: $[H^+] = 0.000\,56\ M$ ($5.6 \times 10^{-4}\ M$).

Concept Exercise

Classify wine as weakly acidic, neutral, or weakly basic given a pH of 3.25.

Answer: See Appendix G, 14.13.

EXAMPLE EXERCISE 14.14 K_w, $[H^+]$, $[OH^-]$, and pH

Indicate the change in hydroxide ion concentration for the following changes in hydrogen ion concentration.

(a) What happens to the $[OH^-]$, if the $[H^+]$ increases?
(b) What happens to the $[OH^-]$, if the $[H^+]$ decreases?

Solution

(a) The $[OH^-]$ decreases if the $[H^+]$ increases; $K_w = [H^+][OH^-]$
(b) The $[OH^-]$ increases if the $[H^+]$ decreases; $K_w = [H^+][OH^-]$.

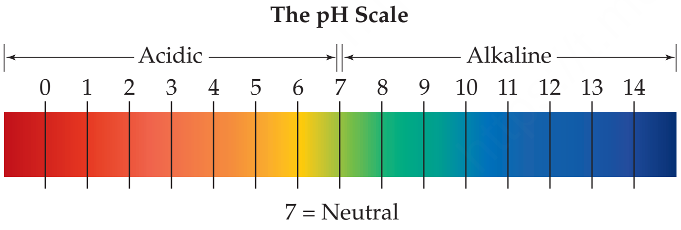
Practice Exercise

Indicate the change in pH for the following changes in molar hydrogen ion concentration.

(a) What happens to the pH, if the $[H^+]$ increases?
(b) What happens to the pH, if the $[H^+]$ decreases?

Answers:

(a) The pH decreases if the $[H^+]$ increases.
(b) The pH increases if the $[H^+]$ decreases.



◀ **The pH Scale**

Concept Exercise

As the temperature of water increases above 25 °C, the kinetic energy increases and more water molecules ionize. Does the value of K_w increase or decrease above 25 °C? Is the pH of distilled water at 50 °C greater or less than 7.0?

Answer: See Appendix G, 14.14.

A CLOSER LOOK Acid Rain

Q: What acids are found in "acid rain"?

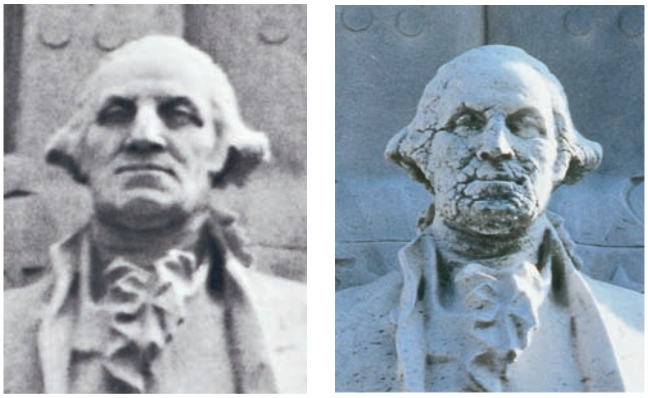
Compare a 1935 photograph of a statue of George Washington with a current photograph. The striking difference is attributed to acid rain, which attacks limestone by reacting with the calcium carbonate in the statue.

The English chemist Robert Smith coined the term *acid rain* after studying the rainfall in London. He found that the air, heavily polluted from coal burning, produced rain that was abnormally acidic. The term *acid rain* has persisted and today usually refers to rain having a pH 5 or below.

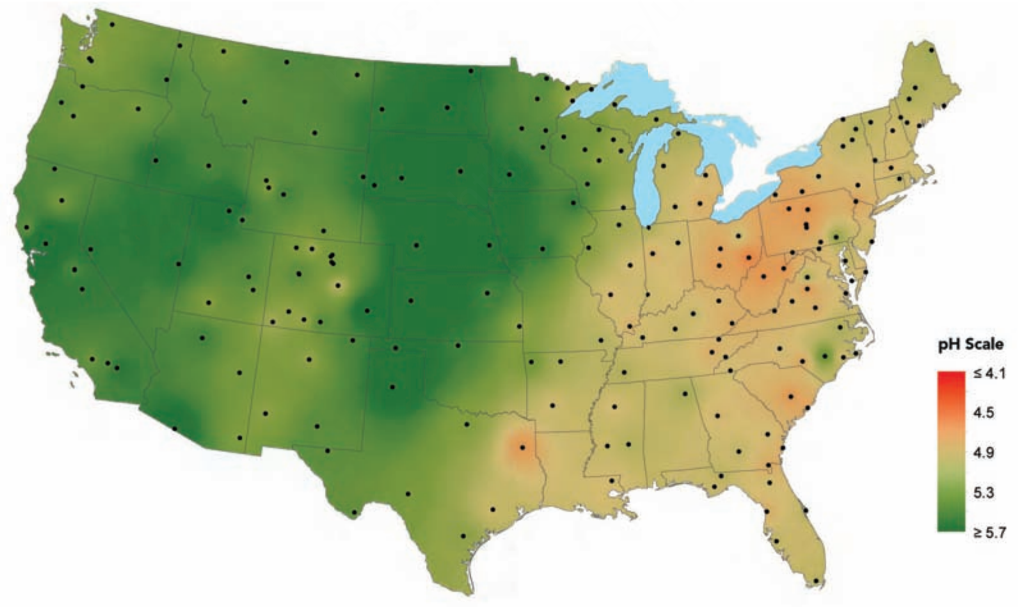
The gases that most contribute to acid rain are oxides of sulfur and nitrogen. Sulfur dioxide and sulfur trioxide in the atmosphere are released mainly by industrial steel plants and electric power plants. These plants burn low-grade coal, which has a high sulfur content. Most of the oxides of nitrogen are emitted from automobiles. When oxides of sulfur and nitrogen are released into the air, they dissolve in atmospheric water vapor. Subsequently, the rain that falls is composed of sulfuric acid and nitric acid raindrops.

Normal rain and acid rain are both acidic, but their pH values are different. Normal rain has a pH of about 5.5. It is slightly acidic because carbon dioxide in the atmosphere dissolves in raindrops to form a dilute solution of carbonic acid. Severe acid rain can be a hundred times more acidic than normal rain, and an unhealthy pH of 2.8 has been recorded. In the northeastern United States, rainfall with a pH of 4 has been blamed on sulfur oxides released from the burning of coal (see map below).

Acid rain is a global problem, and countries all over the world are attempting to reduce it. One way is to reduce the emission of environmental pollutants. Canada has begun a program to reduce its sulfur dioxide emissions. The United States, Japan, and Germany are using high-grade coal to reduce their emissions. Ultimately, the problem of acid rain may be minimized by alternative energy sources. For instance, hydroelectric power and solar power are potential sources of energy that do not pollute the atmosphere.



◀ **Acid Rain** The photograph on the left shows a limestone statue of George Washington in 1935. The recent photograph on the right shows the effect of acid rain, which attacks the statue by dissolving away calcium carbonate in the marble.



A: Acid rain contains carbonic acid from burning fossil fuels, nitric acid from automobile emissions, and sulfuric acid from burning low-grade coal. (The dots represent major U.S. cities.)

14.10 Strong and Weak Electrolytes

LEARNING OBJECTIVES

- Describe the electrical conductivity of strong and weak electrolytes.
- Show strong and weak electrolytes as ionized or nonionized.

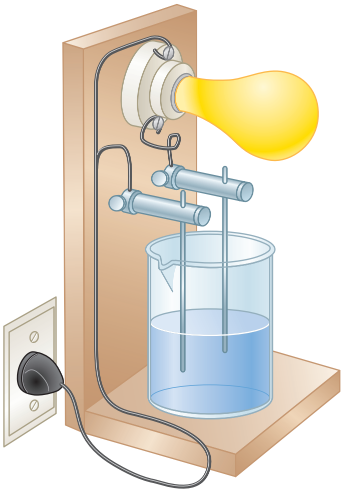
Water is a weak conductor of electricity because it is very slightly ionized. In the laboratory, we can use a conductivity apparatus to determine whether a substance in an aqueous solution is a strong or a weak conductor of electricity. If an aqueous solution is a good conductor, it is called a **strong electrolyte.** If an aqueous solution is a poor conductor, it is called a **weak electrolyte.**

We can test the electrical conductivity of hydrochloric acid and acetic acid using the apparatus in Figure 14.6. From the observations, we can conclude that hydrochloric acid is highly ionized because it is a strong electrolyte. Acetic acid is only slightly ionized because it is a weak electrolyte. Similar experiments using the conductivity apparatus demonstrate the electrolyte behavior of aqueous solutions.

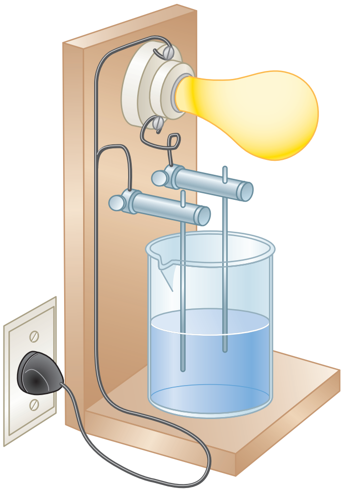
Degree of Ionization in an Aqueous Solution

We can measure the degree of ionization by testing the conductivity of aqueous solutions. Strong electrolytes are highly ionized. Strong electrolytes include strong acids, strong bases, and soluble ionic compounds. Soluble ionic compounds dissolve by dissociating into ions.

Weak electrolytes are slightly ionized. They include weak acids, weak bases, and slightly soluble ionic compounds. Although an ionic compound may be classified as insoluble, it is actually very slightly soluble. A sufficient amount of the insoluble ionic compound dissolves in a solution to act as a weak electrolyte. Table 14.5 lists examples of strong and weak electrolytes.



(a) Strong electrolyte



(b) Weak electrolyte

▲ **Figure 14.6 Conductivity of Aqueous Solutions** (a) An aqueous solution of hydrochloric acid is a strong electrolyte, and the bulb glows brightly. (b) An aqueous solution of acetic acid is a weak electrolyte, and the bulb glows dimly.

TABLE 14.5 Strong Electrolytes and Weak Electrolytes

Strong Electrolytes	Weak Electrolytes
Strong Acids—	***Weak Acids—***
hydrochloric acid, HCl (aq)	hydrofluoric acid, HF(aq)
nitric acid, $HNO_3(aq)$	nitrous acid, $HNO_2(aq)$
sulfuric acid, $H_2SO_4(aq)$	sulfurous acid, $H_2SO_3(aq)$
perchloric acid, $HClO_4(aq)$	acetic acid, $HC_2H_3O_2(aq)$
	carbonic acid, $H_2CO_3(aq)$
	phosphoric acid, $H_3PO_4(aq)$
Strong Bases—	***Weak Bases—***
sodium hydroxide, NaOH(aq)	ammonium hydroxide, $NH_4OH(aq)$
potassium hydroxide, KOH(aq)	insoluble hydroxides such as
lithium hydroxide, LiOH (aq)	$Mg(OH)_2(s)$, $Al(OH)_3(s)$, $Fe(OH)_3(s)$
calcium hydroxide, $Ca(OH)_2(aq)$	
strontium hydroxide, $Sr(OH)_2(aq)$	
barium hydroxide, $Ba(OH)_2(aq)$	
Soluble Salts—	***Insoluble Salts*—***
sodium chloride, NaCl(aq)	silver chloride, AgCl(s)
potassium carbonate, $K_2CO_3(aq)$	calcium carbonate, $CaCO_3(s)$
zinc sulfate, $ZnSO_4(aq)$	barium sulfate, $BaSO_4(s)$

*Insoluble salts are very slightly soluble in water and dissociate into a sufficient number of ions to give a very weak electrolyte solution.

We can determine the degree of ionization in a solution from conductivity testing experiments. That is, we can distinguish between a highly ionized strong electrolyte and a slightly ionized weak electrolyte. Strong electrolytes ionize almost completely in an aqueous solution. For example, sodium chloride and calcium chloride dissociate into ions as follows:

$$\mathrm{NaCl}(aq) \longrightarrow \mathrm{Na}^{+}(aq) + \mathrm{Cl}^{-}(aq)$$
$$\mathrm{CaCl_2}(aq) \longrightarrow \mathrm{Ca}^{2+}(aq) + 2\,\mathrm{Cl}^{-}(aq)$$

We can estimate the amount of ionization in an aqueous solution given the electrolyte strength. That is, the number of ions in solution is proportional to the conductivity. Strong electrolytes are highly ionized, and so we write their formulas in the ionized form. Weak electrolytes are only slightly ionized, and so we write their formulas in the nonionized form. The nonionized formula is sometimes referred to as the *molecular formula*. This is not quite accurate since ionic compounds exist as formula units, not molecules. The following example exercise illustrates how to write formulas for electrolytes in an aqueous solution.

EXAMPLE EXERCISE 14.15 Strong and Weak Electrolytes

Write the ionized or nonionized formula for each of the following aqueous solutions given the electrolyte strength:

(a) $\mathrm{HNO_3}(aq)$, strong (b) $\mathrm{NH_4OH}(aq)$, weak (c) $\mathrm{K_2CO_3}(aq)$, strong

Solution

(a) A nitric acid solution is a strong electrolyte and therefore is highly ionized. It is written in the ionized form as $\mathrm{H}^{+}(aq)$ and $\mathrm{NO_3}^{-}(aq)$.

(b) An ammonium hydroxide solution is a weak electrolyte and therefore is slightly ionized. It is written in the nonionized form as $\mathrm{NH_4OH}(aq)$.

(c) A potassium carbonate solution is a strong electrolyte and therefore is highly ionized. It is written as $2\,\mathrm{K}^{+}(aq)$ and $\mathrm{CO_3}^{2-}(aq)$.

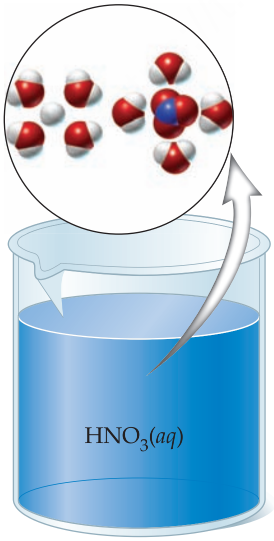
Practice Exercise

Write the ionized or nonionized formula for each of the following aqueous solutions given the electrolyte strength:

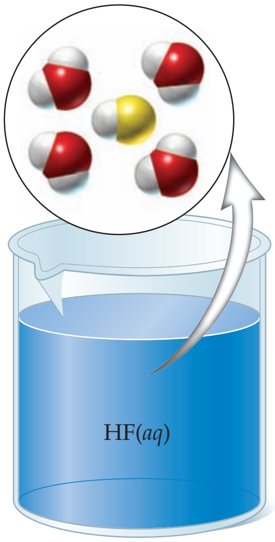
(a) $\mathrm{HF}(aq)$, weak (b) $\mathrm{Ba(OH)_2}(aq)$, strong (c) $\mathrm{CaCO_3}(s)$, weak

Answers:

(a) $\mathrm{HF}(aq)$ (b) $\mathrm{Ba}^{2+}(aq)$ and $2\,\mathrm{OH}^{-}(aq)$ (c) $\mathrm{CaCO_3}(s)$



▲ **Nitric Acid, $\mathrm{HNO_3}$** Nitric acid is a strong electrolyte, which is nearly 100% ionized in an aqueous solution.



▲ **Hydrofluoric Acid, HF** Hydrofluoric acid is a weak electrolyte, which is less than 1% ionized in an aqueous solution.

Concept Exercise

When the weak acid HF(*aq*) neutralizes the weak base $NH_4OH(aq)$, will a conductivity apparatus glow brightly or dimly?

Answer: See Appendix G, 14.15.

Note In these examples, we considered aqueous solutions to be either strong or weak electrolytes. However, some liquids are nonelectrolytes. That is, they do not conduct electricity. Examples include organic liquids such as alcohol, ether, and acetone.

14.11 NET IONIC EQUATIONS

LEARNING OBJECTIVES

- Describe the procedure for writing a net ionic equation.
- Write net ionic equations for given chemical reactions.

In Section 14.10 we learned how to write ionized formulas for strong electrolytes and nonionized formulas for weak electrolytes. The concept of ionization allows us to portray ionic solutions more accurately. That is, we can now write chemical equations by showing strong electrolytes in the ionized form.

Let's consider the neutralization reaction of hydrochloric acid and sodium hydroxide. The nonionized equation for the reaction is

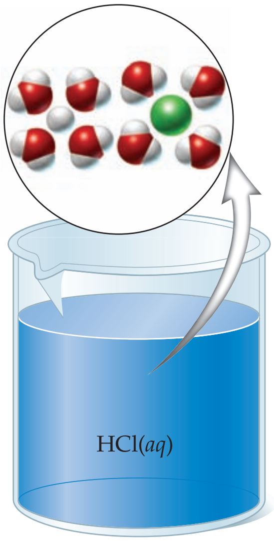
$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l)$$

By writing the ionized formula for the strong acid, strong base, and soluble salt, we can describe the reaction more accurately. Each substance in the **total ionic equation** is written as it primarily exists in the solution. Table 14.5 shows highly ionized substances. Strong electrolytes include HCl, NaOH, and NaCl. Thus, the equation is

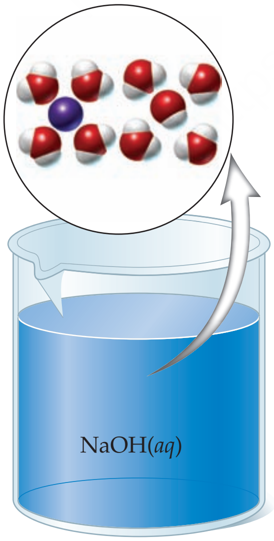
$$\mathrm{H^+}(aq) + \mathrm{Cl^-}(aq) + \mathrm{Na^+}(aq) + \mathrm{OH^-}(aq) \longrightarrow \mathrm{Na^+}(aq) + \mathrm{Cl^-}(aq) + \mathrm{H_2O}(l)$$

Note that $Na^+(aq)$ and $Cl^-(aq)$ appear on both sides of the equation. These ions are called **spectator ions.** They are in the solution but do not participate in the reaction. We can simplify the total ionic equation by eliminating spectator ions. The resulting **net ionic equation** shows only the substances undergoing reaction. The net ionic equation for the preceding reaction is

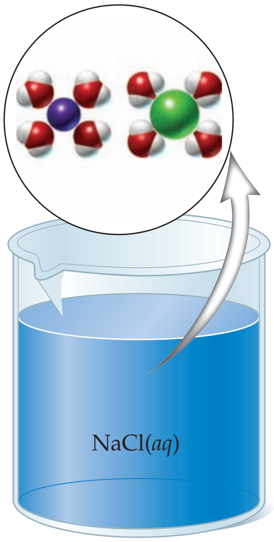
$$\mathrm{H^+}(aq) + \mathrm{OH^-}(aq) \longrightarrow \mathrm{H_2O}(l)$$



▲ **Hydrochloric Acid, HCl** Hydrochloric acid is a strong acid, which is nearly 100% ionized in an aqueous solution.



▲ **Sodium Hydroxide, NaOH** Sodium hydroxide is a strong base, which is nearly 100% dissociated in an aqueous solution.



▲ **Sodium Chloride, NaCl** Sodium chloride is a soluble salt, which is nearly 100% dissociated in an aqueous solution.

Note that we do not show the actual acid and base that were neutralized. This net ionic equation indicates that a strong acid and a strong base reacted to give water. In fact, the net ionic equation is identical for all strong acid and strong base reactions that yield a soluble salt.

Let's state a general procedure to be used when we write net ionic equations. Keep in mind that the net ionic equation gives us a good picture of substances undergoing reaction. That is, the net ionic equation helps us focus on only those substances that are undergoing reaction.

General Guidelines for Writing Net Ionic Equations

1. Complete and balance the nonionized chemical equation.
2. Convert the nonionized equation to a total ionic equation. Write strong electrolytes in the ionized form and weak electrolytes in the nonionized form. Write water and dissolved gases in the nonionized form. Refer to Table 7.2 to determine if a compound is soluble or insoluble.
3. Cancel spectator ions to obtain the net ionic equation.
   - (a) If canceling spectator ions eliminates all species, there is no reaction.
   - (b) If the coefficients can be simplified, do so to have the simplest whole number relationship.
3. Check (√) each ion or atom on both sides of the equation. The total charge (positive or negative) on the reactants side of the equation must equal the total charge on the products side of the equation.

The following example exercises illustrate the procedure for writing balanced net ionic equations.

EXAMPLE EXERCISE 14.16 Net Ionic Equations for Neutralization Reactions

Write a net ionic equation for the reaction between nitric acid and aqueous potassium hydrogen carbonate.

$$HNO_3(aq) + KHCO_3(aq) \longrightarrow KNO_3(aq) + H_2O(l) + CO_2(g)$$

Solution

STEP 1: We verify that the chemical equation is balanced. In this case, the coefficients are all 1 and the equation is balanced.

STEP 2: We determine which species are highly ionized. From Table 14.5, we find that HNO_3 is a strong electrolyte. From Table 7.2, we find that $KHCO_3$ and KNO_3 are soluble strong electrolytes; H_2O and CO_2 are weak electrolytes. The total ionic equation is

$$H^+(aq) + NO_3^-(aq) + K^+(aq) + HCO_3^-(aq) \longrightarrow K^+(aq) + NO_3^-(aq) + H_2O(l) + CO_2(g)$$

STEP 3: We obtain the net ionic equation by eliminating the spectator ions. In this example, K^+ and NO_3^- are in the aqueous solution, but do not participate in the reaction. The net ionic equation is

$$H^+(aq) + HCO_3^-(aq) \longrightarrow H_2O(l) + CO_2(g)$$

STEP 4: We check to verify that the net ionic equation is balanced. The equation is balanced because (a) the number of atoms of each element is the same on each side and (b) the total charge on both sides is identical. In this case the net charge is zero.

Practice Exercise

Write a net ionic equation for the reaction between aqueous solutions of nitric acid and barium hydroxide.

$$2\ HNO_3(aq) + Ba(OH)_2(aq) \longrightarrow Ba(NO_3)_2(aq) + 2\ H_2O(l)$$

Answer:

In this example, nitrate and barium are spectator ions and cancel.

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Concept Exercise

When the strong acid $HNO_3(aq)$ neutralizes the strong base $Ba(OH)_2(aq)$, will a conductivity apparatus glow brightly or dimly?

Answer: See Appendix G, 14.16.

EXAMPLE EXERCISE 14.17 Net Ionic Equations for Precipitation Reactions

Write a net ionic equation for the reaction between aqueous solutions of silver nitrate and aluminum chloride.

$$AgNO_3(aq) + AlCl_3(aq) \longrightarrow AgCl(s) + Al(NO_3)_3(aq)$$

Solution

STEP 1: We balance the chemical equation as follows:

$$3\ AgNO_3(aq) + AlCl_3(aq) \longrightarrow 3\ AgCl(s) + Al(NO_3)_3(aq)$$

STEP 2: We determine which of the species are strong electrolytes. From Table 7.2, we find that all the compounds are soluble except AgCl. Thus, AgCl is weakly ionized, and the total ionic equation is written as

$$3\ Ag^+(aq) + 3\ NO_3^-(aq) + Al^{3+}(aq) + 3\ Cl^-(aq) \longrightarrow 3\ AgCl(s) + Al^{3+}(aq) + 3\ NO_3^-(aq)$$

STEP 3: We cancel spectator ions. In this example, Al^{3+} and NO_3^- do not participate in the reaction.

$$3\ Ag^+(aq) + 3\ Cl^-(aq) \longrightarrow 3\ AgCl(s)$$

STEP 4: We verify that the equation is balanced. Note that each species has a coefficient of 3 in this equation. Thus, after simplifying, the net ionic equation is written as

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

Practice Exercise

Write a net ionic equation for the reaction between aqueous solutions of lithium sulfate and magnesium nitrate.

$$Li_2SO_4(aq) + Mg(NO_3)_2(aq) \longrightarrow MgSO_4(aq) + 2\ LiNO_3(aq)$$

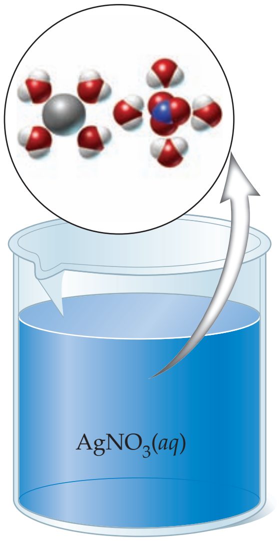
Answer:

$$Li_2SO_4(aq) + Mg(NO_3)_2(aq) \longrightarrow \text{no reaction}$$

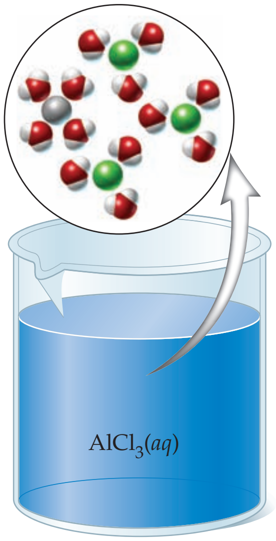
Concept Exercise

When the two strong electrolytes, $AgNO_3(aq)$ and $AlCl_3(aq)$, react to give insoluble AgCl(s), will a conductivity apparatus glow brightly or dimly?

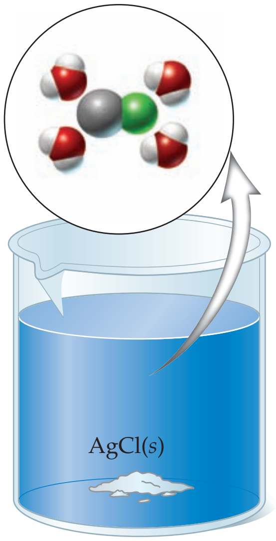
Answer: See Appendix G, 14.17.



▲ **Silver Nitrate, AgNO$_3$** Silver nitrate is a soluble salt, which is nearly 100% dissociated in an aqueous solution.



▲ **Aluminum Chloride, AlCl$_3$** Aluminum chloride is a soluble salt, which is nearly 100% dissociated in an aqueous solution.



▲ **Silver Chloride, AgCl** Silver chloride is an insoluble salt, which is less than 1% dissociated in an aqueous solution.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
14.1 Properties of Acids and Bases The properties of an acid include having a sour taste, turning litmus paper red, having a pH less than 7, and the ability to neutralize a base. The properties of a base include having a bitter taste, being slippery to the touch, turning litmus paper blue, and having a pH greater than 7. A solution that resists changes in pH when an acid or a base is added is referred to as a **buffer.**	• List the general properties of acids and bases. • *Related Exercises: 1–2* • Classify a solution of given pH as one of the following: strongly acidic, weakly acidic, neutral, weakly basic, or strongly basic. *Related Exercises: 3–4*
14.2 Arrhenius Acids and Bases An **Arrhenius acid** is a substance that undergoes **ionization** in water to give hydrogen ions. Hydrochloric acid is a strong acid because it ionizes extensively. Acetic acid is a weak acid because it does not ionize extensively. A hydrogen ion attaches to a water molecule in solution to form a **hydronium ion.** An **Arrhenius base** is a substance that undergoes **dissociation** in water to give hydroxide ions. A **salt** is obtained from the reaction of an acid and a base. Barium nitrate, Ba(NO$_3$)$_2$, is obtained from the neutralization of barium hydroxide and nitric acid.	• Identify an Arrhenius acid and base. • *Related Exercises: 5–8* • Classify a strong or weak acid and base given the degree of ionization. • *Related Exercises: 9–10* • Indicate the acid and base that react to produce a given salt. *Related Exercises: 11–14*
14.3 Brønsted–Lowry Acids and Bases A **Brønsted–Lowry acid** is a **proton donor** and a **Brønsted–Lowry base** is a **proton acceptor.** A substance capable of either donating or accepting a proton in an acid–base reaction is said to be **amphiprotic.**	• Identify a Brønsted–Lowry acid and base in a neutralization reaction. • *Related Exercises: 15–20*
14.4 Acid–Base Indicators We can use litmus paper to indicate if a solution is acidic or basic. In the laboratory, we can use an **acid–base indicator.** Three of the most common indicators are methyl red, bromthymol blue, and phenolphthalein. Methyl red changes color at a pH of 5, bromthymol blue changes at a pH of 7, and phenolphthalein at a pH of 9.	• Describe the color of a solution with a given pH and a drop of phenolphthalein, methyl red, or bromthymol blue indicator. • *Related Exercises: 21–26*

Key Concepts	Learning Objectives and Related Exercises
14.5 Acid–Base Titrations A **titration** is a laboratory procedure for analyzing the amount of acid or base present in a solution. A measured volume of acid or base solution is delivered with a buret into a flask containing a known amount of sample. An indicator is used to signal the neutralization point in the titration. When the **endpoint** is reached, the indicator changes color, and the titration is stopped.	• Perform calculations given acid–base titration data. • *Related Exercises: 27–32* • Calculate the mass percent concentration of a solution given the molarity. *Related Exercises: 33–34*
14.6 Acid–Base Standardization The concentration of a **standard solution** is known accurately to at least three significant digits. Standard solutions are used to analyze the amount of acid or base in a given sample. For example, a chemist can determine the concentration of acetic acid in vinegar or the amount of baking soda in an antacid tablet.	• Describe a standard solution of acid or base. • *Related Exercises: 35–36* • Perform calculations that involve standardization of an acid or base. *Related Exercises: 35–42*
14.7 Ionization of Water On the basis of conductivity experiments, pure water is found to be a very weak electrolyte. Water ionizes to give only a few hydrogen and hydroxide ions. $[H^+]$ and $[OH^-]$ are equal, and both have a value of 1×10^{-7}. The **ionization constant of water,** K_w, is 1×10^{-14} at 25 °C. In every aqueous solution, there is always a small number of both H^+ and OH^-.	• Show the ionization constant equation for water, *Kw*. • *Related Exercises: 43–44* • Calculate the molar hydroxide ion concentration given the molar hydrogen ion concentration. *Related Exercises: 45–48*
14.8 The pH Concept The **pH** of a solution expresses $[H^+]$ on an exponential scale. That is, a solution of pH 1 is 10 times more acidic than a solution of pH 2. To calculate the pH of a solution, we express $[H^+]$ as a power of 10 and change the sign of the exponent. For example, if $[H^+]$ is 10^{-2} *M*, the pH is 2. Conversely, to find $[H^+]$ given the pH, we raise 10 to the negative pH. For example, if the pH is 3, $[H^+]$ is 10^{-3} *M*, that is, 0.001 M.	• Describe the relationship between pH and molar hydrogen ion concentration • *Related Exercises: 49–52*
14.9 Advanced pH Calculations To calculate the pH of solutions whose hydrogen ion concentrations are not whole-number powers of 10, we find the negative logarithm values using a scientific calculator. That is, the pH is equal to $-\log [H^+]$. Conversely, to calculate the hydrogen ion concentration that corresponds to a given pH, we find the inverse logarithm of −pH using a calculator.	• Calculate pH values and molar hydrogen ion concentrations. • *Related Exercises: 53–60*
14.10 Strong and Weak Electrolytes A strong acid, a strong base, and a soluble ionic compound are each an example of a **strong electrolyte.** Therefore, these solutions are highly ionized. Hydrochloric acid is a strong acid and is written as $H^+(aq)$ and $Cl^-(aq)$. Sodium chloride is soluble in water and is represented as $Na^+(aq)$ and $Cl^-(aq)$. A weak acid, a weak base, and a slightly soluble ionic compound are each an example of a **weak electrolyte.** Therefore, these solutions are slightly ionized. Acetic acid is a weak acid and is written as $HC_2H_3O_2(aq)$. The slightly soluble ionic compound silver chloride is written as $AgCl(s)$.	• Describe the electrical conductivity of strong and weak electrolytes. • *Related Exercises: 61–66* • Show strong and weak electrolytes as ionized or nonionized. *Related Exercises: 67–70*
14.11 Net Ionic Equations A **net ionic equation** portrays a solution reaction more accurately. The first step is to balance the chemical equation. Second, write a **total ionic equation** for the reaction. Strong electrolytes are written as ions in solution and weak electrolytes as nonionized. Third, cancel **spectator ions** that are identical on both sides of the equation. Verify the net ionic equation is balanced correctly and simplify the coefficients when possible.	• Describe the procedure for writing a net ionic equation. • *Related Exercises: 71–72* • Write net ionic equations for given chemical reactions. *Related Exercises: 73–76*

Problem–Solving Organizer

Topic	Procedure	Example
Acid–Base Titrations Sec. 14.5	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply unit factors to convert the given unit to the unit in the answer.	If 10.00 mL of 0.100 M sulfuric acid neutralizes 15.50 mL of KOH, what is the molarity of the base? Given: $H_2SO_4(aq) + 2KOH(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$ $10.00\ \cancel{\text{mL solution}} \times \frac{0.100\ \cancel{\text{mol } H_2SO_4}}{1000\ \cancel{\text{mL solution}}} \times \frac{2\text{ mol KOH}}{1\ \cancel{\text{mol } H_2SO_4}}$ $= 0.00200\text{ mol KOH}$ $\frac{0.00200\text{ mol KOH}}{15.50\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\text{ L solution}} = 0.129\ M\text{ KOH}$
Acid–Base Standardization Sec. 14.6	1. Write down the unit asked for in the answer. 2. Write down the related given value. 3. Apply unit factors to convert the given unit to the unit in the answer.	If 15.50 mL of KOH neutralizes 0.995 g of KHP (204.23 g/mol), $KHC_8H_4O_4$, what is the molarity of the base? Given: $KHP(aq) + KOH(aq) \longrightarrow K_2P(aq) + H_2O(l)$ $0.995\ \cancel{\text{g KHP}} \times \frac{1\ \cancel{\text{mol KHP}}}{204.23\ \cancel{\text{g KHP}}} \times \frac{1\text{ mol KOH}}{1\ \cancel{\text{mol KHP}}}$ $= 0.00487\text{ mol KOH}$ $\frac{0.0487\text{ mol KOH}}{15.50\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\text{ L solution}} = 0.314\ M\text{ KOH}$
Ionization of Water Sec. 14.7	Solve for $[H^+]$ or $[OH^-]$ in an aqueous solution using the ionization of water equation. $K_w = [H^+][OH^-]$	If $[H^+]$ is 0.0015 M in an aqueous solution, what is the hydroxide ion concentration? $[H^+][OH^-] = 1.0 \times 10^{-14}$ $[0.0015][OH^-] = 1.0 \times 10^{-14}$ $[OH^-] = 6.7 \times 10^{-12}$
The pH Concept Sec. 14.8	Convert from $[H^+]$ to pH using the following equation: $\text{pH} = -\log [H^+]$	If $[H^+]$ in an acid is 0.001 M, what is the pH? $\text{pH} = -\log [10^{-3}] = -(-3)$ $= 3$
Advanced pH Calculations Sec. 14.9	Convert from pH to $[H^+]$ using the following equation: $[H^+] = 10^{-\text{pH}}$	If the pH of an acid is 3.20, what is the $[H^+]$? $[H^+] = 10^{-3.20} = 0.00063\ M$

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** a solution that resists changes in pH when an acid or a base is added
_____ **2.** a substance that releases hydrogen ions when dissolved in water
_____ **3.** a substance that releases hydroxide ions when dissolved in water
_____ **4.** the process of a polar compound dissolving in water and forming positive and negative ions
_____ **5.** the process of an ionic compound dissolving in water and separating into positive and negative ions
_____ **6.** the ion that best represents the hydrogen ion in an aqueous solution
_____ **7.** a product obtained from a neutralization reaction in addition to water
_____ **8.** a substance that donates a proton in an acid–base reaction
_____ **9.** a substance that accepts a proton in an acid–base reaction
_____ **10.** a synonym for hydrogen ion donor
_____ **11.** a synonym for hydrogen ion acceptor
_____ **12.** a substance capable of either donating or accepting a proton in an acid–base reaction
_____ **13.** a chemical substance that changes color according to the pH of the solution
_____ **14.** a procedure for delivering a measured volume of solution using a buret
_____ **15.** the stage in a titration when the indicator changes color
_____ **16.** a solution whose concentration has been established precisely
_____ **17.** a constant that equals the product of the molar hydrogen ion concentration and the molar hydroxide ion concentration
_____ **18.** the molar hydrogen ion concentration expressed on an exponential scale
_____ **19.** an aqueous solution that is a good conductor of electricity
_____ **20.** an aqueous solution that is a poor conductor of electricity
_____ **21.** a chemical equation that shows ionized substances in solution as ions and slightly ionized substances as nonionized
_____ **22.** ions in an aqueous solution that do not participate in a reaction, and do not appear in the net ionic equation
_____ **23.** a chemical equation that results from canceling spectator ions in a total ionic equation

(a) acid–base indicator *(Sec. 14.4)*
(b) amphiprotic *(Sec. 14.3)*
(c) Arrhenius acid *(Sec. 14.2)*
(d) Arrhenius base *(Sec. 14.2)*
(e) Brønsted–Lowry acid *(Sec. 14.3)*
(f) Brønsted–Lowry base *(Sec. 14.3)*
(g) buffer *(Sec. 14.1)*
(h) dissociation *(Sec. 14.2)*
(i) endpoint *(Sec. 14.5)*
(j) hydronium ion (H_3O^-) *(Sec. 14.2)*
(k) ionization *(Sec. 14.2)*
(l) ionization constant of water (K_w) *(Sec. 14.7)*
(m) net ionic equation *(Sec. 14.11)*
(n) pH *(Sec. 14.8)*
(o) proton acceptor *(Sec. 14.3)*
(p) proton donor *(Sec. 14.3)*
(q) salt *(Sec. 14.2)*
(r) spectator ions *(Sec. 14.11)*
(s) standard solution *(Sec. 14.6)*
(t) strong electrolyte *(Sec. 14.10)*
(u) titration *(Sec. 14.5)*
(v) total ionic equation *(Sec. 14.11)*
(w) weak electrolyte *(Sec. 14.10)*

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Properties of Acids and Bases (Sec. 14.1)

1. Which of the following is a general property of an acidic solution?
   (a) tastes sour **(b)** turns litmus paper red
   (c) pH greater than 7 **(d)** feels slippery
2. Which of the following is a general property of a basic solution?
   (a) tastes bitter **(b)** turns litmus paper blue
   (c) pH less than 7 **(d)** neutralizes a base
3. Classify the following aqueous solutions as strongly acidic, weakly acidic, neutral, weakly basic, or strongly basic.
   (a) stomach acid, pH 1.5 **(b)** vinegar, pH 3.2
   (c) distilled water, pH 7.0 **(d)** drain cleaner, pH 13.5
4. Classify the following aqueous solutions as strongly acidic, weakly acidic, neutral, weakly basic, or strongly basic.
   (a) battery acid, pH 0 **(b)** lime juice, pH 2.7
   (c) shampoo, pH 7.0 **(d)** ammonia, pH 11.1

▲ **Common Acids** Methyl red indicator is added to beakers of lime juice, vinegar, and club soda. The color of the indicator shows a different pH in each beaker.

Arrhenius Acids and Bases (Sec. 14.2)

5. Classify each of the following Arrhenius acids as strong or weak given the degree of ionization.
   - **(a)** hypochlorous acid, $HClO(aq)$, ~1%
   - **(b)** chlorous acid, $HClO_2(aq)$, ~1%
   - **(c)** chloric acid, $HClO_3(aq)$, ~100%
   - **(d)** perchloric acid, $HClO_4(aq)$, ~100%
6. Classify each of the following Arrhenius bases as strong or weak given the degree of dissociation.
   - **(a)** aluminum hydroxide, $Al(OH)_3(aq)$, ~1%
   - **(b)** calcium hydroxide, $Sr(OH)_2(aq)$, ~100%
   - **(c)** barium hydroxide, $Ba(OH)_2(aq)$, ~100%
   - **(d)** magnesium hydroxide, $Mg(OH)_2(aq)$, ~1%
7. Classify each of the following as an Arrhenius acid, an Arrhenius base, or a salt.
   - **(a)** $H_2SO_4(aq)$ **(b)** $Na_2SO_4(aq)$
   - **(c)** $NaOH(aq)$ **(d)** $Na_2CO_3(aq)$
8. Classify each of the following as an Arrhenius acid, an Arrhenius base, or a salt.
   - **(a)** $HNO_3(aq)$ **(b)** $KNO_3(aq)$
   - **(c)** $KOH(aq)$ **(d)** $K_2CO_3(aq)$
9. Identify the Arrhenius acid and Arrhenius base in each of the following neutralization reactions.
   - **(a)** $HBr(aq) + NaOH(aq) \longrightarrow NaBr(aq) + H_2O(l)$
   - **(b)** $H_2SO_4(aq) + Mg(OH)_2(aq) \longrightarrow K_2SO_4(aq) + 2\,H_2O(l)$
10. Identify the Arrhenius acid and Arrhenius base in each of the following neutralization reactions.
    - **(a)** $HC_2H_3O_2(aq) + LiOH(aq) \longrightarrow LiC_2H_3O_2(aq) + H_2O(l)$
    - **(b)** $H_3PO_4(aq) + 3\,KOH(aq) \longrightarrow K_3PO_4(aq) + 3\,H_2O(l)$
11. Determine the acid and base that were neutralized to produce each of the following salts.
    - **(a)** sodium fluoride, $NaF(aq)$
    - **(b)** magnesium iodide, $MgI_2(aq)$
    - **(c)** calcium nitrate, $Ca(NO_3)_2(aq)$
    - **(d)** lithium carbonate, $Li_2CO_3(aq)$
12. Determine the acid and base that were neutralized to produce each of the following salts.
    - **(a)** potassium bromide, $KBr(aq)$
    - **(b)** barium chloride, $BaCl_2(aq)$
    - **(c)** cobalt(II) sulfate, $CoSO_4(aq)$
    - **(d)** sodium phosphate, $Na_3PO_4(aq)$
13. Complete and balance the following neutralization reactions.
    - **(a)** $HNO_3(aq) + Ca(OH)_2(aq) \longrightarrow$
    - **(b)** $H_2CO_3(aq) + Ba(OH)_2(aq) \longrightarrow$
14. Complete and balance the following neutralization reactions.
    - **(a)** $HC_2H_3O_2(aq) + Sr(OH)_2(aq) \longrightarrow$
    - **(b)** $H_2SO_4(aq) + NH_4OH(aq) \longrightarrow$

Brønsted–Lowry Acids and Bases (Sec. 14.3)

15. Identify the Brønsted–Lowry acid and base in each of the following neutralization reactions.
    - **(a)** $HC_2H_3O_2(aq) + LiOH(aq) \longrightarrow LiC_2H_3O_2(aq) + H_2O(l)$
    - **(b)** $NH_3(aq) + HCl(aq) \longrightarrow NH_4Cl(aq)$
16. Identify the Brønsted–Lowry acid and base in each of the following neutralization reactions.
    - **(a)** $HNO_3(aq) + NaHCO_3(aq) \longrightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$
    - **(b)** $NaCN(aq) + HBr(aq) \longrightarrow NaBr(aq) + HCN(aq)$
17. Identify the Brønsted–Lowry acid and base in each of the following neutralization reactions.
    - **(a)** $HI(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + I^-(aq)$
    - **(b)** $HC_2H_3O_2(aq) + HS^-(aq) \longrightarrow H_2S(aq) + C_2H_3O_2^-(aq)$
18. Identify the Brønsted–Lowry acid and base in each of the following neutralization reactions.
    - **(a)** $HCO_3^-(aq) + OH^-(aq) \longrightarrow CO_3^{2-}(aq) + H_2O(l)$
    - **(b)** $NO_2^-(aq) + HClO_4(aq) \longrightarrow HNO_2(aq) + ClO_4^-(aq)$
19. Complete and balance the following Brønsted–Lowry neutralization reactions.
    - **(a)** $HF(aq) + NaHS(aq) \longrightarrow$
    - **(b)** $HNO_2(aq) + NaC_2H_3O_2(aq) \longrightarrow$
20. Complete and balance the following Brønsted–Lowry neutralization reactions.
    - **(a)** $H_3O^+(aq) + SO_4^{2-}(aq) \longrightarrow$
    - **(b)** $H_2PO_4^-(aq) + NH_3(aq) \longrightarrow$

Acid–Base Indicators (Sec. 14.4)

21. Given the following pH values for solutions containing a drop of phenolphthalein indicator, state the color of each solution.
    - **(a)** pH 6 **(b)** pH 12
22. What is the color of phenolphthalein indicator in deionized water?
23. Given the following pH values for solutions containing a drop of methyl red indicator, state the color of each solution.
    - **(a)** pH 3 **(b)** pH 7
24. What is the color of methyl red indicator in a solution of pH 5?
25. Given the following pH values for solutions containing a drop of bromthymol blue indicator, state the color of each solution.
    - **(a)** pH 5 **(b)** pH 9
26. What is the color of bromthymol blue indicator in a solution of pH 7?

◀ **Natural Indicator** A rose contains an acid-base indicator. After the indicator is extracted from the dark-red rose (left beaker), the rose petals are a pale pink (right beaker).

Acid–Base Titrations (Sec. 14.5)

27. If 21.50 mL of 0.115 *M* barium hydroxide is required to neutralize 25.0 mL of nitric acid, what is the molarity of the acid?

$$2\,HNO_3(aq) + Ba(OH)_2(aq) \longrightarrow Ba(NO_3)_2(aq) + 2\,H_2O(l)$$

28. If 27.55 mL of 0.101 *M* barium hydroxide is required to neutralize 25.0 mL of nitrous acid, what is the molarity of the acid?

$$2\,HNO_2(aq) + Ba(OH)_2(aq) \longrightarrow Ba(NO_2)_2(aq) + 2\,H_2O(l)$$

29. If 28.15 mL of 0.100 *M* ammonium hydroxide is required to neutralize 10.0 mL of sulfuric acid, what is the molarity of the acid?

$$H_2SO_4(aq) + 2\,NH_4OH(aq) \longrightarrow (NH_4)_2SO_4(aq) + 2\,H_2O(l)$$

30. If 27.05 mL of 0.105 *M* ammonium hydroxide is required to neutralize 10.0 mL of sulfurous acid, what is the molarity of the acid?

$$H_2SO_3(aq) + 2\,NH_4OH(aq) \longrightarrow (NH_4)_2SO_3(aq) + 2\,H_2O(l)$$

31. If 42.45 mL of 0.500 *M* lithium hydroxide is required to neutralize 50.0 mL of phosphoric acid, what is the molarity of the acid?

$$H_3PO_4(aq) + 3\,LiOH(aq) \longrightarrow Li_3PO_4(aq) + 3\,H_2O(l)$$

32. If 45.05 mL of 0.488 *M* lithium hydroxide is required to neutralize 50.0 mL of phosphorous acid, what is the molarity of the acid?

$$H_3PO_3(aq) + 3\,LiOH(aq) \longrightarrow Li_3PO_3(aq) + 3\,H_2O(l)$$

33. Given the molarity and density for each of the following acidic solutions, calculate the mass/mass percent concentration.

(a) 6.00 *M* HCl ($d = 1.10$ g/mL)
(b) 1.00 *M* $HC_2H_3O_2$ ($d = 1.01$ g/mL)
(c) 0.500 *M* HNO_3 ($d = 1.01$ g/mL)
(d) 3.00 *M* H_2SO_4 ($d = 1.18$ g/mL)

34. Given the molarity and density for each of the following basic solutions, calculate the mass/mass percent concentration.

(a) 3.00 *M* NaOH ($d = 1.12$ g/mL)
(b) 0.500 *M* KOH ($d = 1.02$ g/mL)
(c) 6.00 *M* NH_3 ($d = 0.954$ g/mL)
(d) 1.00 *M* Na_2CO_3 ($d = 1.10$ g/mL)

Acid–Base Standardization (Sec. 14.6)

35. What is the molarity of hydrochloric acid if 33.25 mL of HCl is required to neutralize 1.115 g of sodium hydrogen carbonate?

$$HCl(aq) + NaHCO_3(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

36. What is the molarity of hydrochloric acid if 29.75 mL of HCl is required to neutralize 0.515 g of sodium carbonate?

$$2\,HCl(aq) + Na_2CO_3(aq) \longrightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

37. What is the molarity of lithium hydroxide if 28.85 mL of LiOH is required to neutralize 0.506 g of oxalic acid, $H_2C_2O_4$?

$$H_2C_2O_4(aq) + 2\,LiOH(aq) \longrightarrow Li_2C_2O_4(aq) + H_2O(l)$$

38. What is the molarity of potassium hydroxide if 29.05 mL of KOH is required to neutralize 0.627 g of oxalic acid, $H_2C_2O_4$?

$$H_2C_2O_4(aq) + 2\,KOH(aq) \longrightarrow K_2C_2O_4(aq) + 2\,H_2O(l)$$

39. What volume of 0.128 *M* calcium hydroxide is required to neutralize 1.466 g of potassium hydrogen phthalate, $KHC_8H_4O_4$ (204.23 g/mol)?

$$2\,KHC_8H_4O_4(aq) + Ca(OH)_2(aq) \longrightarrow CaK_2(C_8H_4O_4)_2(aq) + 2\,H_2O(l)$$

40. What volume of 0.150 *M* barium hydroxide is required to neutralize 1.655 g of potassium hydrogen phthalate, $KHC_8H_4O_4$ (204.23 g/mol)?

$$2\,KHC_8H_4O_4(aq) + Ba(OH)_2(aq) \longrightarrow BaK_2(C_8H_4O_4)_2(aq) + 2\,H_2O(l)$$

41. Ascorbic acid, abbreviated HAsc, is the chemical name for vitamin C. If 30.95 mL of 0.176 *M* NaOH neutralizes 0.959 g of acid, what is the molar mass of vitamin C?

$$HAsc(aq) + NaOH(aq) \longrightarrow NaAsc(aq) + H_2O(l)$$

42. Tartaric acid, abbreviated H_2Tart, gives grapes a sour taste. If 28.15 mL of 0.295 *M* NaOH neutralizes 0.623 g of acid, what is the molar mass of tartaric acid?

$$H_2Tart(aq) + 2\,NaOH(aq) \longrightarrow Na_2Tart(aq) + 2\,H_2O(l)$$

Ionization of Water (Sec. 14.7)

43. Indicate each of the following for the ionization of pure water.

(a) the simplified ionization equation
(b) the ionization constant equation, K_w
(c) the ionization constant at 25 °C

44. Indicate each of the following for the ionization of pure water.

(a) the molecular collision equation
(b) the molar hydrogen ion concentration at 25 °C
(c) the molar hydroxide ion concentration at 25 °C

45. Given the molar concentration of hydrogen ion, calculate the concentration of hydroxide ion.

(a) $[H^+] = 0.025$ **(b)** $[H^+] = 0.000\,017$

46. Given the molar concentration of hydrogen ion, calculate the concentration of hydroxide ion.

(a) $[H^+] = 6.2 \times 10^{-7}$ **(b)** $[H^+] = 4.6 \times 10^{-12}$

47. Given the molar concentration of hydroxide ion, calculate the concentration of hydrogen ion.

(a) $[OH^-] = 0.0016$ **(b)** $[OH^-] = 0.000\,29$

48. Given the molar concentration of hydroxide ion, calculate the concentration of hydrogen ion.

(a) $[OH^-] = 8.8 \times 10^{-8}$ **(b)** $[OH^-] = 4.6 \times 10^{-13}$

The pH Concept (Sec. 14.8)

49. Calculate the pH of each of the following given the molar hydrogen ion concentration.

(a) soft drink, $[H^+] = 0.001$ *M*
(b) coffee, $[H^+] = 0.000\,01$ *M*

50. Calculate the pH of each of the following given the molar hydrogen ion concentration.

(a) egg white, $[H^+] = 0.000\,000\,01$ *M*
(b) sour milk, $[H^+] = 0.000\,001$ *M*

51. Calculate the molar hydrogen ion concentration of each of the following given the pH.

(a) shampoo, pH 6
(b) pH-balanced shampoo, pH 8

52. Calculate the molar hydrogen ion concentration of each of the following given the pH.

(a) phosphate detergent, pH 9
(b) nonphosphate detergent, pH 11

Advanced pH Calculations (Sec. 14.9)

53. Calculate the pH of each of the following given the molar hydrogen ion concentration.
(a) milk, $[\mathrm{H^+}] = 0.000\,000\,30\ M$
(b) eggs, $[\mathrm{H^+}] = 0.000\,000\,016\ M$

54. Calculate the pH of each of the following given the molar hydrogen ion concentration.
(a) carrots, $[\mathrm{H^+}] = 0.000\,007\,9\ M$
(b) peas, $[\mathrm{H^+}] = 0.000\,000\,39\ M$

55. Calculate the molar hydrogen ion concentration of each of the following biological solutions given the pH.
(a) saliva, pH = 6.55 **(b)** blood, pH = 7.50

56. Calculate the molar hydrogen ion concentration of each of the following biological solutions given the pH.
(a) gastric juice, pH = 1.80 **(b)** urine, pH = 4.75

57. Calculate the pH of each of the following given the molar hydroxide ion concentration.
(a) $[\mathrm{OH^-}] = 0.11\ M$ **(b)** $[\mathrm{OH^-}] = 0.000\,55\ M$

58. Calculate the pH of each of the following given the molar hydroxide ion concentration.
(a) $[\mathrm{OH^-}] = 0.000\,031\ M$
(b) $[\mathrm{OH^-}] = 0.000\,000\,000\,66\ M$

59. Calculate the molar hydroxide ion concentration of each of the following solutions given the pH.
(a) pH = 4.55 **(b)** pH = 5.20

60. Calculate the molar hydroxide ion concentration of each of the following solutions given the pH.
(a) pH = 0.90 **(b)** pH = 1.62

Strong and Weak Electrolytes (Sec. 14.10)

61. State whether each of the following substances in aqueous solution is highly ionized or slightly ionized.
(a) strong acids
(b) strong bases
(c) soluble ionic compounds

62. State whether each of the following substances in aqueous solution is highly ionized or slightly ionized.
(a) weak acids
(b) weak bases
(c) slightly soluble ionic compounds

63. Classify each of the following acids as a strong or weak electrolyte.
(a) $\mathrm{HF}(aq)$ **(b)** $\mathrm{HCl}(aq)$
(c) $\mathrm{HNO_3}(aq)$ **(d)** $\mathrm{HNO2}(aq)$

64. Classify each of the following bases as a strong or weak electrolyte.
(a) $\mathrm{KOH}(aq)$ **(b)** $\mathrm{NH_4OH}(aq)$
(c) $\mathrm{Sr(OH)_2}(aq)$ **(d)** $\mathrm{Al(OH)_3}(s)$

65. Classify each of the following aqueous solutions as a strong or weak electrolyte.
(a) $\mathrm{ZnCO_3}(s)$ **(b)** $\mathrm{Sr(NO_3)_2}(aq)$
(c) $\mathrm{K_2SO_4}(aq)$ **(d)** $\mathrm{PbI_2}(s)$

66. Classify each of the following aqueous solutions as a strong or weak electrolyte.
(a) $\mathrm{Fe(C_2H_3O_2)_3}(aq)$ **(b)** $\mathrm{AlPO_4}(s)$
(c) $\mathrm{Ag_2CrO_4}(s)$ **(d)** $\mathrm{CdSO_4}(aq)$

67. Write the following acids in either the ionized or the nonionized form to best represent an aqueous solution.
(a) $\mathrm{HF}(aq)$ **(b)** $\mathrm{HBr}(aq)$
(c) $\mathrm{HNO_3}(aq)$ **(d)** $\mathrm{HNO_2}(aq)$

68. Write the following bases in either the ionized or the nonionized form to best represent an aqueous solution.
(a) $\mathrm{NaOH}(aq)$ **(b)** $\mathrm{NH_4OH}(aq)$
(c) $\mathrm{Ba(OH)_2}(aq)$ **(d)** $\mathrm{Al(OH)_3}(s)$

69. Write the following salts in either the ionized or the nonionized form to best represent an aqueous solution.
(a) $\mathrm{AgF}(aq)$ **(b)** $\mathrm{AgI}(s)$
(c) $\mathrm{Hg_2Cl_2}(s)$ **(d)** $\mathrm{NiCl_2}(aq)$

70. Write the following salts in either the ionized or the nonionized form to best represent an aqueous solution.
(a) $\mathrm{AlPO_4}(s)$ **(b)** $\mathrm{Co(C_2H_3O_2)_3}(aq)$
(c) $\mathrm{MnSO_4}(aq)$ **(d)** $\mathrm{PbSO_4}(s)$

Net Ionic Equations (Sec. 14.11)

71. List the four steps for writing a balanced net ionic equation.

72. What term describes ions in a total ionic equation that do not react?

73. Write a balanced net ionic equation for each of the following acid–base reactions. Refer to Table 14.5 and Appendix D for electrolyte information.
(a) $\mathrm{HCl}(aq) + \mathrm{KOH}(aq) \longrightarrow \mathrm{KCl}(aq) + \mathrm{H_2O}(l)$
(b) $\mathrm{HC_2H_3O_2}(aq) + \mathrm{Ca(OH)_2}(aq) \longrightarrow \mathrm{Ca(C_2H_3O_2)_2}(aq) + \mathrm{H_2O}(l)$

74. Write a balanced net ionic equation for each of the following acid–base reactions. Refer to Table 14.5 and Appendix D for electrolyte information.
(a) $\mathrm{HF}(aq) + \mathrm{Li_2CO_3}(aq) \longrightarrow \mathrm{LiF}(aq) + \mathrm{H_2O}(l) + \mathrm{CO_2}(g)$
(b) $\mathrm{H_2SO_4}(aq) + \mathrm{Ba(OH)_2}(aq) \longrightarrow \mathrm{BaSO_4}(s) + \mathrm{H_2O}(l)$

75. Write a balanced net ionic equation for each of the following solution reactions. Refer to Table 14.5 and Appendix D for electrolyte information.
(a) $\mathrm{AgNO_3}(aq) + \mathrm{KI}(aq) \longrightarrow \mathrm{AgI}(s) + \mathrm{KNO_3}(aq)$
(b) $\mathrm{BaCl_2}(aq) + \mathrm{K_2CrO_4}(aq) \longrightarrow \mathrm{BaCrO_4}(s) + \mathrm{KCl}(aq)$

76. Write a balanced net ionic equation for each of the following solution reactions. Refer to Table 14.5 and Appendix D for electrolyte information.
(a) $\mathrm{Zn(NO_3)_2}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{Zn(OH)_2}(s) + \mathrm{NaNO_3}(aq)$
(b) $\mathrm{MgSO_4}(aq) + \mathrm{NH_4OH}(aq) \longrightarrow \mathrm{Mg(OH)_2}(s) + \mathrm{(NH_4)_2SO_4}(aq)$

General Exercises

77. A drop of methyl red and a drop of phenolphthalein are added to a beaker of distilled water. What is the resulting color of the water?

78. A drop of bromthymol blue and a drop of phenolphthalein are added to a beaker of distilled water. What is the resulting color of the water?

79. Methyl orange changes color from pH 3.2 to pH 4.4. If the indicator is red at pH 3 and yellow at pH 5, what is the color of the indicator in a pH 3.8 solution?

80. Bromcresol green changes color from pH 3.8 to pH 5.4. If the indicator is yellow at pH 3 and blue at pH 6, what is the color of the indicator in a pH 4.6 solution?

81. Identify the amphiprotic substance given the following reactions.

$$\mathrm{HCl}(aq) + \mathrm{NaH_2PO_4}(aq) \longrightarrow \mathrm{H_3PO_4}(aq) + \mathrm{NaCl}(aq)$$
$$\mathrm{NaOH}(aq) + \mathrm{NaH_2PO_4}(aq) \longrightarrow \mathrm{Na_2HPO_4}(aq) + \mathrm{H_2O}(l)$$

82. Identify the amphiprotic substance given the following reactions.

$$HI(aq) + K_2HPO_4(aq) \longrightarrow KH_2PO_4(aq) + KI(aq)$$
$$KOH(aq) + K_2HPO_4(aq) \longrightarrow K_3PO_4(aq) + H_2O(l)$$

83. What is the pH of a 0.50 *M* HCl solution?

84. What is the pH of a 0.50 *M* NaOH solution?

85. In 1 mL of water there are 3×10^{22} molecules of H_2O. How many hydrogen ions are in 1 mL of water? (*Hint:* In 1 billion water molecules, 2 are ionized.)

86. In 10 mL of water there are 3×10^{23} molecules of H_2O. How many hydroxide ions are in 10 mL of water? (*Hint:* In 1 billion water molecules, 2 are ionized.)

87. Write the net ionic equation for the reaction of the weak electrolytes acetic acid and ammonium hydroxide.

88. Write the net ionic equation for the reaction of the strong electrolytes nitric acid and potassium hydroxide.

Challenge Exercises

89. Alanine, abbreviated HAla, is an amino acid in protein. If 21.05 mL of 0.145 M NaOH neutralizes 0.272 g HAla, what is the molar mass of the amino acid?

90. Cream of tartar, abbreviated KHTart, is used in baking. If 42.10 mL of 0.100 M KOH neutralizes 0.791 g KHTart, what is the molar mass of cream of tartar?

Chapter 14 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following is a general property of an acidic solution? (Sec. 14.1)
(a) tastes sour
(b) turns litmus paper red
(c) pH less than 7
(d) neutralizes bases
(e) all of the above

2. Which of the following is a strong Arrhenius acid? (Sec. 14.2)
(a) $HCl(aq)$ (~100% ionized)
(b) $HNO_3(aq)$ (~100% ionized)
(c) $H_2SO_4(aq)$ (~100% ionized)
(d) all of the above
(e) none of the above

3. In the following reaction, which reactant is a Brønsted–Lowry base? (Sec. 14.3)

$$HCl(aq) + KHS(aq) \longrightarrow KCl(aq) + H_2S(aq)$$

(a) HCl
(b) KHS
(c) KCl
(d) H_2S
(e) none of the above

4. What is the color of phenolphthalein indicator at pH 7? (Sec. 14.4)
(a) red
(b) blue
(c) pink
(d) colorless
(e) none of the above

5. If 25.0 mL of 0.100 *M* HCl is titrated with 0.150 *M* $Ba(OH)_2$, what volume of barium hydroxide is required to neutralize the acid? (Sec. 14.5)

$$2\,HCl(aq) + Ba(OH)_2(aq) \longrightarrow BaCl_2(aq) + 2\,H_2O(l)$$

(a) 8.33 mL
(b) 16.7 mL
(c) 18.8 mL
(e) 37.5 mL
(d) 33.3 mL

6. What is the molarity of a hydrochloric acid solution if 20.00 mL of HCl is required to neutralize 0.424 g of sodium carbonate (105.99 g/mol)? (Sec. 14.6)

$$2\,HCl(aq) + Na_2CO_3(aq) \longrightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

(a) 0.100 *M*
(b) 0.200 *M*
(c) 0.300 *M*
(d) 0.400 *M*
(e) 0.500 *M*

7. Given an aqueous solution in which the $[H^+] = 2.5 \times 10^{-7}$, what is the molar hydroxide ion concentration? (Sec. 14.7)
(a) $[OH^-] = 2.5 \times 10^{-7}\,M$
(b) $[OH^-] = 2.5 \times 10^{-8}\,M$
(c) $[OH^-] = 4.0 \times 10^{-6}\,M$
(d) $[OH^-] = 4.0 \times 10^{-7}\,M$
(e) $[OH^-] = 4.0 \times 10^{-8}\,M$

8. What is the hydrogen ion concentration in a urine sample that registers a pH of 4 on a strip of pH paper? (Sec. 14.8)
(a) 0.001 *M*
(b) 0.004 *M*
(c) 0.0001 *M*
(d) 0.0004 *M*
(e) 0.00001 *M*

9. What is the hydrogen ion concentration in a blood sample that registers a pH of 7.30 using a pH meter? (Sec. 14.9)
(a) 0.000 000 020 *M*
(b) 0.000 000 050 *M*
(c) 0.000 000 20 *M*
(d) 0.000 000 50 *M*
(e) 0.000 073 *M*

10. If a drain cleaner solution is a strong electrolyte, which of the following must be true? (Sec. 14.10)
(a) Drain cleaner is highly reactive.
(b) Drain cleaner is slightly reactive.
(c) Drain cleaner is highly ionized.
(d) Drain cleaner is slightly ionized.
(e) none of the above

11. Nitric acid is a strong acid, sodium hydroxide is a strong base, and sodium nitrate is a soluble salt. Which of the following is the net ionic equation for the reaction? (Sec. 14.11)

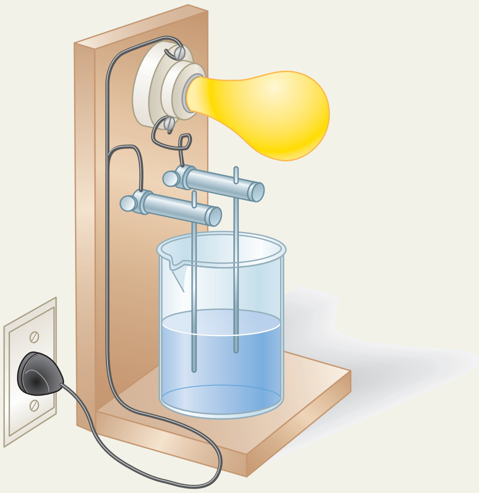
$$HNO_3(aq) + NaOH(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$$

(a) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
(b) $NO_3^-(aq) + Na^+(aq) \longrightarrow NaNO_3(s)$
(c) $H^+(aq) + NaOH(aq) \longrightarrow Na^+(aq) + H_2O(l)$
(d) $HNO_3(aq) + OH^-(aq) \longrightarrow NO_3^-(aq) + H_2O(l)$
(e) none of the above

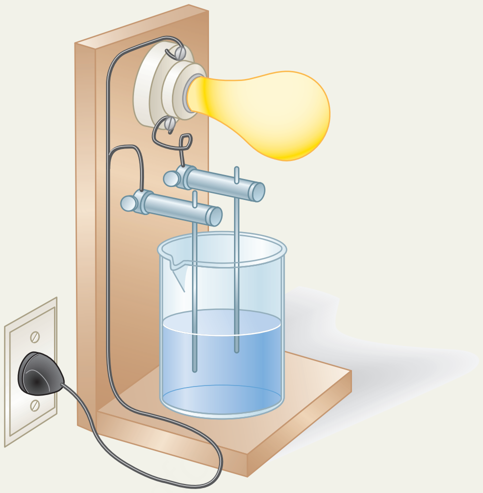
Key Concepts

12. What must be done first before applying unit analysis to solve an acid–base titration problem?

13. Beakers X and Y contain either sulfuric acid or sulfurous acid. Based on the conductivity test shown, which beaker contains H_2SO_4? Which contains H_2SO_3?



(a) Beaker X



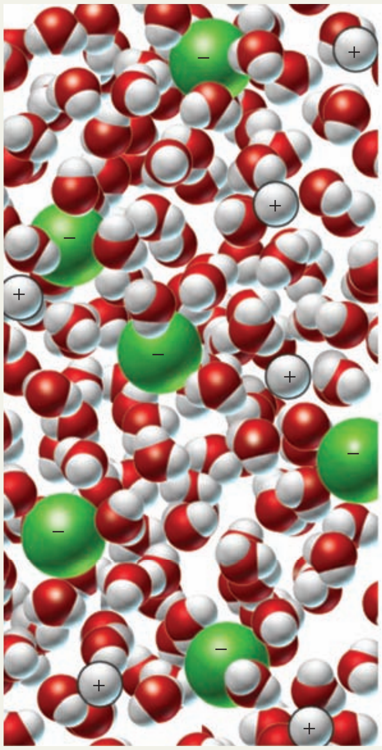
(b) Beaker Y

Critical Thinking

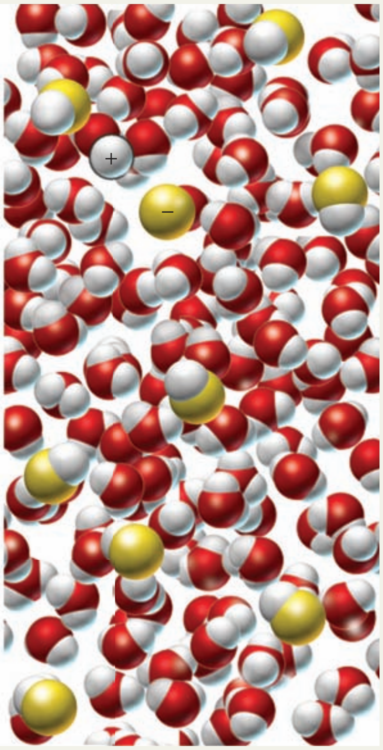
14. The smell of fish is due to a basic organic compound related to ammonia. What common kitchen chemical can be used to neutralize an unpleasant fish odor?

15. What color is the solution in a test tube that contains 10 mL of acid, 1 drop of phenolphthalein, and 1 drop of bromthymol blue indicator?

16. Examine the following illustrations of HCl and HF ionizing in solution. At the molecular level, explain why HCl is more highly ionized than HF.



(a) HCl highly ionized



(b) HF slightly ionized

CHAPTER 15

Advanced Problem Solving

"Simple can be harder than complex: You have to work hard to get your thinking clean to make it simple. But it's worth it in the end because once you get there, you can move mountains."

Steve Jobs, Founder of Apple Computer (1955–2012)

15
P
phosphorus

Element 15: Phosphorus

Phosphorus is not found free in nature, but rather combined in minerals. The element phosphorus occurs in different crystalline forms that are white, red, or black in color. White phosphorus is the most reactive, glows in the dark, and slowly oxidizes in air to red phosphorus.

In preceding chapters we have been building a foundation for solving chemistry problems. One important skill we learned was how to perform calculations using the unit analysis method of problem solving. With each new topic we applied a consistent and systematic approach to solving problems. First, we applied unit analysis to metric system conversions. Next, we applied unit analysis to chemical formula calculations and stoichiometry problems. Then, we applied the unit analysis approach for solving gas law questions and solution concentration problems.

In this chapter we will combine many of the preceding calculation topics we learned; that is, topics previously covered individually will now be considered in combination. We will incorporate two or more of topics in the same problem; for example, empirical formula and the ideal gas law.

15.1 Advanced Problem-Solving Strategies

LEARNING OBJECTIVES

- Describe several techniques for solving problems, including: unit analysis, algebraic analysis, concept maps, and visualization.
- Estimate a "ballpark" answer before using a calculator to obtain a precise answer.

Previously, we considered problem solving involving a single concept. In this chapter, we will consider problems that involve two or more concepts. We will also study alternative methods for solving problems. You will have options for solving a problem and you will choose an approach that is most appropriate.

To solve multi-concept problems successfully, you will need to analyze the given information systematically. First, examine the problem to find the unknown quantity. Next, determine which of the given information is related to the unknown. Finally, plan a strategy that links the given value to the unknown. In summary, the three steps are as follows.

1. *Examine the problem to determine the unknown quantity.*
2. *Write down the given value related to the unknown quantity.*
3. *Plan a strategy that relates the given value to the unknown quantity.*

After reading the description of a problem, you should ask: "What is the unknown quantity that is to be calculated?" The unknown quantity can be a mass, a volume, or a temperature. In most of the problems we have solved, the units of the unknown quantity have been specified. That is, the quantity you were asked to calculate is the mass *in grams*, the volume *in liters*, or a solution concentration *in moles per liter*. You will now encounter problems that may ask you to calculate an unknown quantity without specifying units. In these cases, choose a unit that is convenient. For example, if the quantity sought is a mass, the mass may be expressed in kg, g, mg, etc.

The second question to ask is, "What given value is related to the unknown quantity to be calculated?" If the unknown quantity is volume, the related given value may be in units of L, mL, or cm^3. You would therefore search the problem for a relevant volume such as 1.00 L, 10.0 mL, or 10.0 cm^3.

A third question to ask is, "What steps are necessary to proceed from the given value to the unknown?" In advanced problem solving, there is a variety of techniques available to solve a problem including: *unit analysis*, *algebraic analysis*, *concept maps*, and *visualization*.

Unit Analysis

A technique with which you are quite familiar is solving problems by *unit analysis*. The unit analysis method of problem solving is a powerful tool for solving simple chemistry problems. Recall the format for unit analysis problem solving is

$$\underset{(2)}{\text{relevant given value}} \times \underset{(3)}{\frac{\text{unit}}{\text{factor(s)}}} = \underset{(1)}{\text{units in answer}}$$

To solve the problems in this chapter, we will need a more complex approach. Consider these guidelines for using unit analysis in these more advanced problems.

(a) *Many problems cannot be solved in a single unit analysis operation.* If the problem is complex, we must break it down into two or more simpler problems. Before starting any calculations, plan an overall strategy to arrive at a final answer.

(b) *If the unknown quantity is a single unit (e.g., cm, g, mL), the given value should also be a single unit.* If the unknown quantity is a ratio (e.g., g/mL, g/mol, mol/L), the given value should also be a ratio. That is, the relevant given is a ratio with units in both the numerator and the denominator.

(c) *A problem may include given values that are not relevant.* We must then sort through the given information to determine what is essential and what can be ignored.

(d) *A problem may omit information that is necessary to obtain a solution.* You then have to find the information in a reference source. For example, you can use the periodic table to determine the molar mass of a substance. The appendices in chemistry textbooks have reference tables. A reference for data such as density, specific heat, and solubility, is the *Handbook of Chemistry and Physics.* You can also use an online reference such as Google.com or Wikipedia.org.

(e) *If the calculation is based upon a chemical reaction, first write a balanced chemical equation for the reaction.*

(f) *Before using your calculator, estimate an approximate answer. For example, consider the following calculation:*

$$1.125\ \cancel{\text{g KI}} \times \frac{1\ \cancel{\text{mol KI}}}{166.00\ \cancel{\text{g KI}}} \times \frac{1\ \cancel{\text{mol HgI}_2}}{2\ \cancel{\text{mol KI}}} \times \frac{454.39\ \text{g HgI}_2}{1\ \cancel{\text{mol HgI}_2}} = \text{g HgI}_2$$

There are no absolute rules for rounding off, but we can approximate the answer as follows:

$$\sim 1\ \cancel{\text{g KI}} \times \frac{1\ \cancel{\text{mol KI}}}{\sim 200\ \cancel{\text{g KI}}} \times \frac{1\ \cancel{\text{mol HgI}_2}}{2\ \cancel{\text{mol KI}}} \times \frac{\sim 500\ \text{g HgI}_2}{1\ \cancel{\text{mol HgI}_2}} = \sim 1\ \text{g HgI}_2$$

The estimated answer ("ballpark answer") is 1 g. Using a calculator, we can check the estimated answer. In this example, the calculator value is 1.540 g, which agrees with the estimated value of ~1 g.

▲ **HgI_2 Precipitate** Mercury(II) iodide is an orange precipitate.

Algebraic Analysis

For most problems the unit analysis method of problem solving is our first choice. However, there are some problems for which unit analysis is not appropriate. Some problems are more easily solved using *algebraic analysis*. In Section 2.8 we solved density problems using a unit analysis approach. We could have solved these problems using algebra.

Density (d) is mass (m) per unit volume (v). Therefore, we can write the equation

$$d = \frac{m}{v}$$

If we have 1550 g of recycled aluminum cans (~100 cans), we can calculate the volume after crushing the cans, given the density of aluminum is 2.70 g/cm^3.

$$2.70\ \text{g/cm}^3 = \frac{1550\ \text{g}}{v}$$

Rearranging the equation, we find the volume of crushed aluminum cans:

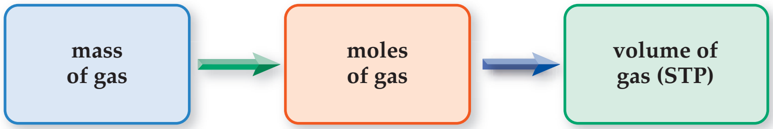
$$v = \frac{1550\ \cancel{\text{g}}}{2.70\ \cancel{\text{g}}/\text{cm}^3} = 574\ \text{cm}^3$$

In Section 10.11 we introduced the ideal gas equation: $PV = nRT$. This is another example where algebra is more appropriate for solving a chemical calculation.

▲ **Aluminum Cans for Recycling**

Concept Maps

A *concept map* is a diagram that relates one or more chemical concepts. As an example, consider the following problem. What is the volume in liters of 3.25 g of oxygen gas at standard conditions? The concept map for the problem is:



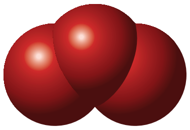
To solve this problem, we must perform two conversions. First, we convert the grams of O_2 to moles of O_2. Second, we convert the moles of O_2 to liters of O_2 at STP.

Visualization

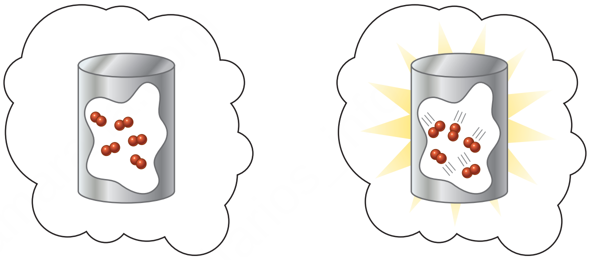
Although you have never been on the Moon, you could describe its craters and dusty surface. We have a mental picture of the Moon's surface based on photographs from the Apollo missions. Many of the concepts in chemistry are similarly removed from our actual experience. These concepts become more clear through the process of *visualization*. One example is the visualization of atoms and molecules. We cannot see an atom, but we can picture an atom or group of atoms in a molecule. This process of forming mental pictures helps us understand concepts we cannot observe directly. Through our ability to form mental pictures, we can perceive the concept of atoms more clearly.

Let's reexamine the problem of finding the volume of 3.25 g of oxygen gas at STP. To use visualization to help solve the problem, we begin by forming a mental picture of oxygen molecules moving about randomly in a stainless steel container. As we picture the shiny metal container, we can ask the question, "How big is the container?"

The ability to visualize is a skill that you can acquire and develop with practice. Refer to Figure 15.1, which illustrates the visualization process.



▲ Ozone, O_3



(a)

(b)

▶ **Figure 15.1 The Visualization Process** (a) A student can create a mental image of a gas by visualizing small spheres in a container. (b) The image becomes more vivid by picturing gas molecules moving rapidly about a shiny metal container.

Helpful Hint
Problem Solving Methods

In this chapter, we tackle more difficult problems that may incorporate two or more chemical concepts. Although unit analysis is sufficient for most chemistry problems, there are instances for which algebra is more convenient. Therefore, students may design solutions to problems using unit analysis or algebra, and may benefit from additional techniques such as concept maps and visualization.

Before calculating an answer, it is good practice to estimate a numerical value for the answer. After using a calculator, compare the calculator answer to your estimated "ballpark answer."

"Ballpark" Answers

We often perform chemical calculations correctly, although we do not know whether or not our answer is reasonable. With experience, we can begin to estimate an answer *before* using a calculator. After setting up a numerical calculation, round off each value to one significant digit, and mentally perform the calculation. The resulting estimated answer is often referred to as a *ballpark answer* because the estimated value is only approximate. The following example using ozone gas illustrates estimating a ballpark answer.

$$4.75\ \cancel{\text{g O}_3} \times \frac{1\ \cancel{\text{mol O}_3}}{48.00\ \cancel{\text{g O}_3}} \times \frac{22.4\ \text{L O}_3}{1\ \cancel{\text{mol O}_3}} = \text{L O}_3$$

Rounding off each numerical value to one digit, we have

$$5\ \cancel{\text{g O}_3} \times \frac{1\ \cancel{\text{mol O}_3}}{50\ \cancel{\text{g O}_3}} \times \frac{20\ \text{L O}_3}{1\ \cancel{\text{mol O}_3}} = {\sim}2\ \text{L O}_3$$

The estimated answer is ~2 L O_3, and the calculator answer is 2.22 L O_3. Thus, our ballpark answer agrees with the calculated answer and we have confidence that our answer is correct.

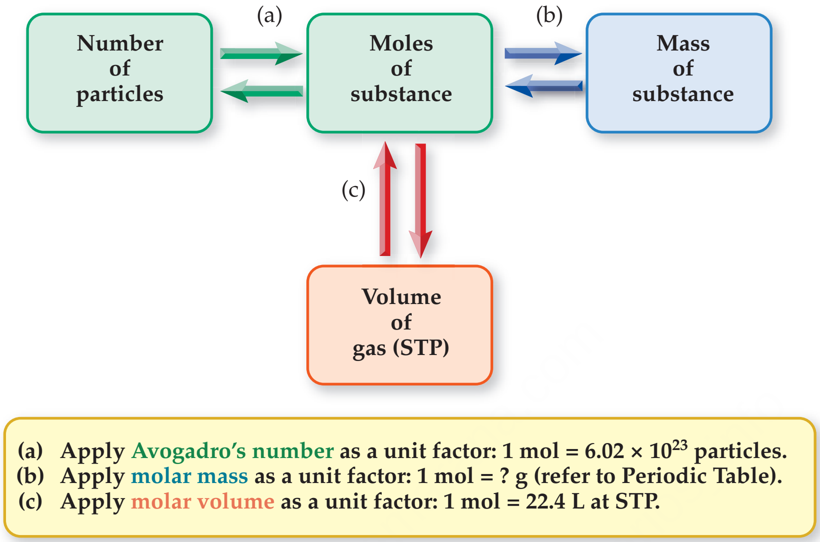
As we gain experience in solving problems, we develop an appreciation for answers that are reasonable, and those that are unreasonable. For example, given 2.28 L of ozone gas at STP, we can calculate that the number of oxygen molecules is 6.13×10^{22}. This is a reasonable answer because 2.28 L of oxygen gas is approximately one-tenth of a mole (22.4 L), and 6.13×10^{22} is approximately one-tenth of Avogadro's number (6.02×10^{23}).

15.2 Concept Maps—Mole Relationships

LEARNING OBJECTIVE

- Solve problems that relate moles of substance in the following:
  - (a) number of atoms, molecules, or formula units
  - (b) mass of substance
  - (c) volume of gas

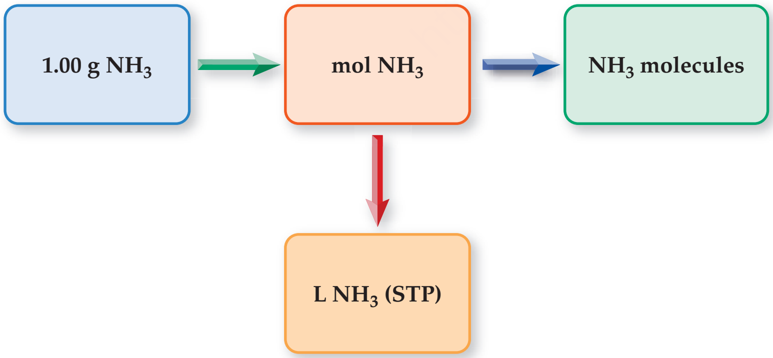
In Chapter 8 we learned about mole calculations. We learned to relate Avogadro's number, molar mass, and molar volume. To gain an overall perspective of mole relationships, we can draw a concept map, as shown in Figure 15.2.

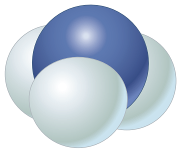


(a) Apply **Avogadro's number** as a unit factor: 1 mol = 6.02×10^{23} particles.
(b) Apply **molar mass** as a unit factor: 1 mol = ? g (refer to Periodic Table).
(c) Apply **molar volume** as a unit factor: 1 mol = 22.4 L at STP.

▲ **Figure 15.2** **Mole Concept Map** The concept map shows the various quantities related to the central mole concept. We can find the moles of a substance from the (a) number of particles, (b) mass of substance, or (c) volume of a gas. If we are given the moles of a substance, we can calculate any of the other quantities.

Given 1.00 g of ammonia gas, NH_3, let's calculate the number of ammonia molecules, and the volume of NH_3 gas at STP. We see from the concept map in Figure 15.2 that we must first calculate the moles of NH_3. We can draw a concept map for the problem as follows:





▲ Ammonia, NH_3

Using the periodic table, we can calculate the molar mass of NH_3 is 17.04 g/mol, and find the moles of NH_3 as follows:

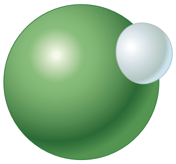
$$1.00\ \cancel{\text{g NH}_3} \times \frac{1\text{ mol NH}_3}{17.04\ \cancel{\text{g NH}_3}} = 0.0587\text{ mol NH}_3$$

Having found the number of moles of ammonia, we can calculate any of the other quantities. For instance, in one mole of ammonia there is Avogadro's number of molecules. We can find the number of molecules in our given sample as follows:

$$0.0587\ \cancel{\text{mol NH}_3} \times \frac{6.02 \times 10^{23}\text{ molecules NH}_3}{1\ \cancel{\text{mol NH}_3}} = 3.53 \times 10^{22}\text{ molecules NH}_3$$

We find the volume of ammonia gas at STP using the molar volume concept. One mole of ammonia, or any other gas at STP, occupies a volume of 22.4 liters. The volume of the given sample is

$$0.0587\ \cancel{\text{mol NH}_3} \times \frac{22.4\text{ L NH}_3}{1\ \cancel{\text{mol NH}_3}} = 1.31\text{ L NH}_3$$



▲ **Hydrogen Chloride, HCl**

EXAMPLE EXERCISE 15.1 Chemical Formula Calculations

Given 50.0 mL of 0.500 *M* hydrochloric acid, calculate each of the following:

(a) grams of HCl gas dissolved in the acid solution
(b) liters of HCl gas (at STP) dissolved in the solution
(c) molecules of HCl gas dissolved in the solution

Solution

We must first determine the number of moles of HCl before we can calculate the other quantities. Recall that 0.500 *M* corresponds to the unit factor 0.500 mol/1000 mL. Hence,

$$50.0\ \cancel{\text{mL solution}} \times \frac{0.500\text{ mol HCl}}{1000\ \cancel{\text{mL solution}}} = 0.0250\text{ mol HCl}$$

(a) From the periodic table we find the molar mass of HCl is 36.46 g/mol. The mass of HCl dissolved in the acid solution is therefore

$$0.0250\ \cancel{\text{mol HCl}} \times \frac{36.46\text{ g HCl}}{1\ \cancel{\text{mol HCl}}} = 0.912\text{ g HCl}$$

(b) The molar volume of any gas, including HCl, is 22.4 L/mol at STP. The volume of HCl gas at standard conditions is

$$0.0250\ \cancel{\text{mol HCl}} \times \frac{22.4\text{ L HCl}}{1\ \cancel{\text{mol HCl}}} = 0.560\text{ L HCl (STP)}$$

(c) The number of HCl molecules dissolved in the acid is found using Avogadro's number. To obtain the number of molecules, multiply the number of moles times 6.02×10^{23}.

$$0.0250\ \cancel{\text{mol HCl}} \times \frac{6.02 \times 10^{23}\text{ molecules HCl}}{1\ \cancel{\text{mol HCl}}} = 1.51 \times 10^{22}\text{ molecules HCl}$$

Practice Exercise

Given 100.0 mL of 0.105 *M* aqueous carbon dioxide, calculate each of the following:

(a) mass of CO_2 dissolved in the solution
(b) volume of CO_2 gas (at STP) dissolved in the solution
(c) molecules of CO_2 dissolved in the solution

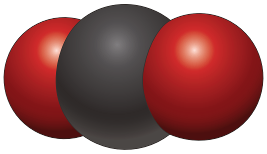
Answers:

(a) 0.462 g CO_2
(b) 0.235 L CO_2 (at STP)
(c) 6.32×10^{21} molecules of CO_2

Concept Exercise

What four quantities (e.g., mass) are related to the concept of a mole of gas?

Answers: See Appendix G, 15.1.



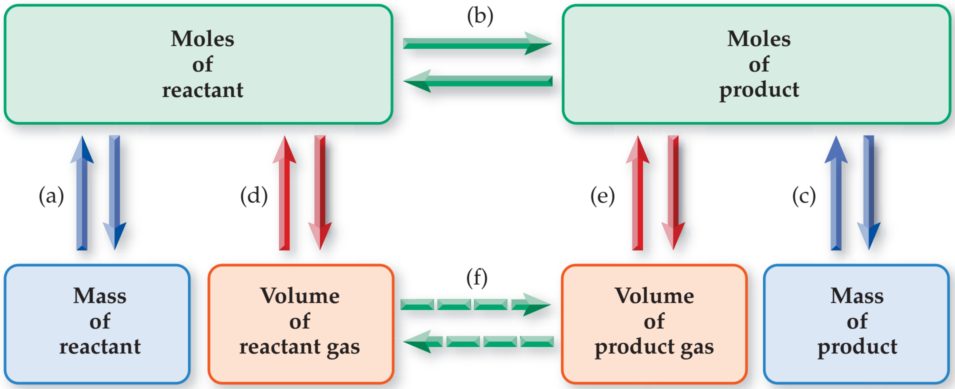
▲ **Carbon Dioxide, CO_2**

15.3 Concept Maps—Stoichiometry

LEARNING OBJECTIVE

- Solve problems that relate amounts of substance in the following:
  - (a) mass–mass stoichiometry
  - (b) mass–volume stoichiometry
  - (c) volume–volume stoichiometry
  - (d) solution stoichiometry

In Chapter 9 we learned about chemical equation calculations. We learned to relate quantities involved in mass–mass, mass–volume, and volume–volume problems. To gain an overall perspective of stoichiometry relationships, we can draw the concept map shown in Figure 15.3.



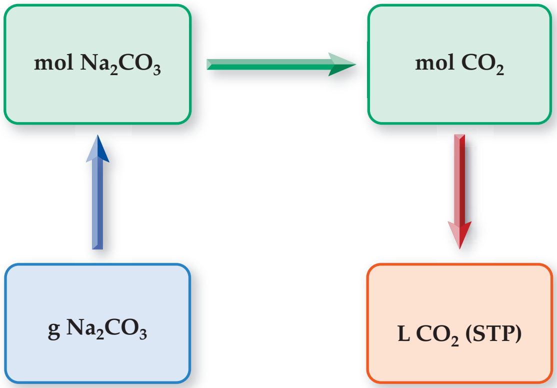
(a) Apply molar mass as a unit factor: 1 mol = ? g (refer to Periodic Table).
(b) Apply mole ratio as a unit factor (refer to balanced equation coefficients).
(c) Apply molar mass as a unit factor: 1 mol = ? g (refer to Periodic Table).
(d) Apply molar volume as a unit factor: 1 mol = 22.4 L at STP.
(e) Apply molar volume as a unit factor: 1 mol = 22.4 L at STP.
(f) Apply volume ratio as a unit factor (refer to balanced equation coefficients).

▲ Figure 15.3 **Stoichiometry Concept Map** The concept map shows the quantities that are related by a balanced chemical equation. We can relate the masses of reactants and products, as well as the volumes of gaseous reactants and products.

Now let's reinforce our understanding of chemical equation calculations. Suppose we are given 4.24 g of sodium carbonate, Na_2CO_3, which we react with 0.150 *M* HCl. The equation for the reaction is

$$Na_2CO_3(s) + 2\,HCl(aq) \longrightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

Let's calculate the STP volume of CO_2 in liters produced from the reaction. Referring to Figure 15.3, we first convert the mass of Na_2CO_3 to moles. Second, we apply the coefficients of the balanced chemical equation to find the moles of CO_2. Third, we convert the moles of gas to a volume of gas at STP. We can draw a concept map for the problem as follows:



Next, we show the unit analysis solution to the problem. By adding up the atomic masses from the periodic table, we find the molar mass of Na_2CO_3 to be 105.99 g/mol. The value for molar volume is 22.4 L/mol. The calculation is as follows:

$$4.24\ \cancel{\text{g Na}_2\text{CO}_3} \times \frac{1\ \cancel{\text{mol Na}_2\text{CO}_3}}{105.99\ \cancel{\text{g Na}_2\text{CO}_3}} \times \frac{1\ \cancel{\text{mol CO}_2}}{1\ \cancel{\text{mol Na}_2\text{CO}_3}} \times \frac{22.4\ \text{L CO}_2}{1\ \cancel{\text{mol CO}_2}} = 0.896\ \text{L CO}_2$$

Example Exercise 15.2 provides practice in solving chemical equation calculations involving the concentration of a solution.

▲ PbI_2 Precipitate

EXAMPLE EXERCISE 15.2 Chemical Equation Calculations

Given that 25.0 mL of 0.100 *M* lead(II) nitrate reacts with 37.5 mL of lithium iodide solution according to the equation

$$\text{Pb(NO}_3)_2(aq) + 2\ \text{LiI(aq)} \longrightarrow \text{PbI}_2(s) + 2\ \text{LiNO}_3(aq)$$

(a) What is the molarity of the LiI solution?
(b) What is the mass of PbI_2 precipitate?

Solution

We begin by verifying the chemical equation is balanced.

(a) We next outline a plan for finding the molarity of aqueous LiI:

$$\text{mL Pb(NO}_3)_2 \longrightarrow \text{mol Pb(NO}_3)_2 \longrightarrow \text{mol LiI} \longrightarrow M\ \text{LiI}$$

From the balanced equation, we see 1 mol $Pb(NO_3)_2$ reacts with 2 mol LiI. We calculate the moles of LiI as follows:

$$25.0\ \cancel{\text{mL Pb(NO}_3)_2} \times \frac{0.100\ \cancel{\text{mol Pb(NO}_3)_2}}{1000\ \cancel{\text{mL Pb(NO}_3)_2}} \times \frac{2\ \text{mol LiI}}{1\ \cancel{\text{mol Pb(NO}_3)_2}} = 0.00500\ \text{mol LiI}$$

Molarity is the ratio of moles to volume. Because molarity has compound units (mol/L), it is necessary to start with compound units (mol/L). We have 0.00500 mol LiI in 37.5 mL of solution; therefore,

$$\frac{0.00500\ \text{mol LiI}}{37.5\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \text{L}} = \frac{0.133\ \text{mol LiI}}{1\ \text{L solution}} = 0.133\ M\ \text{LiI}$$

(b) To find the mass of PbI_2 precipitate, we can relate the concepts:

$$\text{mL Pb(NO}_3)_2 \longrightarrow \text{mol Pb(NO}_3)_2 \longrightarrow \text{mol PbI}_2 \longrightarrow \text{g PbI}_2$$

From the balanced equation, we see 1 mol $Pb(NO_3)_2$ yields 1 mol PbI_2 The molar mass of PbI_2 is 461.0 g/mol; therefore,

$$25.0\ \cancel{\text{mL Pb(NO}_3)_2} \times \frac{0.100\ \cancel{\text{mol Pb(NO}_3)_2}}{1000\ \cancel{\text{mL Pb(NO}_3)_2}} \times \frac{1\ \cancel{\text{mol PbI}_2}}{1\ \cancel{\text{mol Pb(NO}_3)_2}} \times \frac{461.0\ \text{g PbI}_2}{1\ \cancel{\text{mol PbI}_2}} = \text{g PbI}_2$$
$$= 1.15\ \text{g PbI}_2$$

Practice Exercise

Given that 25.0 mL of 0.100 *M* aluminum bromide reacts with 0.125 *M* silver nitrate solution

$$\text{AlBr}_3(aq) + 3\ \text{AgNO}_3(aq) \longrightarrow 3\ \text{AgBr}(s) + \text{Al(NO}_3)_3(aq)$$

(a) What volume of $AgNO_3$ is required for complete precipitation?
(b) What is the mass of AgBr (187.77 g/mol) precipitate?

Answers:

(a) 60.0 mL $AgNO_3$
(b) 1.41 g AgBr

Concept Exercise

What is first required before solving a stoichiometry problem?

Answer: See Appendix G, 15.2.

15.4 Multiple-Reaction Stoichiometry

LEARNING OBJECTIVE

- Solve stoichiometry problems that involve two or more chemical reactions.

In the chemical industry, it is important to perform stoichiometry calculations to determine the cost of manufacturing a chemical. Moreover, many industrial processes involve a series of two or more chemical reactions. For example, the Ostwald Process for making nitric acid, HNO_3, and the Contact Process for making sulfuric acid, H_2SO_4, each consists of three reactions. Fortunately, the stoichiometry techniques we have learned can be extended to multiple reactions.

Sodium carbonate, Na_2CO_3, is an important industrial chemical; its common name is soda ash. Soda ash is mined from rich deposits in Wyoming and the Mojave Desert in California. It is also chemically manufactured. In 1869 Ernest Solvay invented the industrial method for producing soda ash, which bears his name. The first reaction in the Solvay Process is to convert carbon dioxide and aqueous ammonia to ammonium hydrogen carbonate:

$$CO_2(g) + NH_3(g) + H_2O(l) \longrightarrow NH_4HCO_3(aq) \qquad (1)$$

In the second reaction ammonium bicarbonate undergoes a double replacement reaction to give sodium bicarbonate, $NaHCO_3$ (baking soda):

$$NH_4HCO_3(aq) + NaCl(aq) \longrightarrow NaHCO_3(s) + NH_4Cl(aq) \qquad (2)$$

The third reaction involves heating the baking soda to give soda ash, Na_2CO_3:

$$2\ NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + CO_2(g) + H_2O(g) \qquad (3)$$

Let's suppose we want to calculate the mass of soda ash produced from 75.0 g of carbon dioxide gas in reaction (1). We can outline the relationship of CO_2 to Na_2CO_3 as follows:

$$\text{mol } CO_2 \longrightarrow \text{mol } NH_4HCO_3 \longrightarrow \text{mol } NaHCO_3 \longrightarrow \text{mol } Na_2CO_3$$

Starting with CO_2, we followed carbon as it was converted from one substance to another. In this problem we want to calculate the mass of Na_2CO_3 derived from 75.0 g of CO_2.

In a multiple-reaction problem we have to apply additional unit factors. That is, we must relate moles of substance in each reaction. In this problem, the unit analysis solution is:

$$75.0\ \cancel{g\ CO_2} \times \frac{1\ \cancel{mol\ CO_2}}{44.01\ \cancel{g\ CO_2}} \times \frac{1\ \cancel{mol\ NH_4HCO_3}}{1\ \cancel{mol\ CO_2}} \times \frac{1\ \cancel{mol\ NaHCO_3}}{1\ \cancel{mol\ NH_4HCO_3}}$$
$$\times \frac{1\ \cancel{mol\ Na_2CO_3}}{2\ \cancel{mol\ NaHCO_3}} \times \frac{105.99\ g\ Na_2CO_3}{1\ \cancel{mol\ Na_2CO_3}} = 90.3\ g\ Na_2CO_3$$

Notice in this example of multiple–reaction stoichiometry, there were *three chemical equations* to consider rather than a single equation. Example Exercise 15.3 illustrates multiple-reaction stoichiometry for the manufacture of nitric acid from ammonia.

▲ **Nitric Acid, HNO_3**

EXAMPLE EXERCISE 15.3 Multiple-Reaction Stoichiometry

Nitric acid is prepared industrially by the conversion of ammonia. There are three steps in the conversion, which is referred to as the Ostwald Process.

$$4\,NH_3(g) + 5\,O_2(g) \xrightarrow{Pt/850°C} 4\,NO(g) + 6\,H_2O(g)$$
$$2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g)$$
$$3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(aq) + NO(g)$$

What is the mass of HNO_3 (63.02 g/mol) produced, assuming 10.0 L of NH_3 at STP is completely converted to nitric acid?

Solution

We can plan our strategy for the relationship of NH_3 to HNO_3 as follows:

$$\text{L } NH_3 \longrightarrow \text{mol } NH_3 \longrightarrow \text{mol NO} \longrightarrow \text{mol } NO_2$$
$$\longrightarrow \text{mol } HNO_3 \longrightarrow g\ HNO_3$$

Starting with NH_3, we followed the conversion of nitrogen to another substance containing the nitrogen atom. We are asked to calculate the mass of HNO_3 from 10.0 L of NH_3. Therefore let's show the general format for the unit analysis calculation. It is

$$10.0\text{ L } NH_3 \times \text{unit factors} = \text{g } HNO_3$$

We see from the strategy plan that five unit factors are required for the conversion. The complete unit analysis solution to the problem is as follows:

$$10.0\ \cancel{\text{L } NH_3} \times \frac{1\ \cancel{\text{mol } NH_3}}{22.4\ \cancel{\text{L } NH_3}} \times \frac{4\ \cancel{\text{mol NO}}}{4\ \cancel{\text{mol } NH_3}} \times \frac{2\ \cancel{\text{mol } NO_2}}{2\ \cancel{\text{mol NO}}}$$
$$\times \frac{2\ \cancel{\text{mol } HNO_3}}{3\ \cancel{\text{mol } NO_2}} \times \frac{63.02\text{ g } HNO_3}{1\ \cancel{\text{mol } HNO_3}} = 18.8\text{ g } HNO_3$$

The conversion of 10.0 L of NH_3 to 18.8 g HNO_3 requires three reactions. We have assumed that in each reaction we obtain a 100% yield. For industrial processes, the actual yield is less. The industrial Ostwald Process provides an approximately 90% yield.

Practice Exercise

A large quantity of sulfur is obtained industrially from "sour" natural gas containing hydrogen sulfide. First, hydrogen sulfide is burned in air to give sulfur dioxide. Second, the SO_2 produced reacts with additional H_2S to give elemental sulfur.

$$2\,H_2S(g) + 3\,O_2(g) \longrightarrow 2\,SO_2(g) + 2\,H_2O(g) \quad (1)$$
$$2\,SO_2(g) + 2\,H_2S(g) \longrightarrow 3\,S(l) + 2\,H_2O(g) \quad (2)$$

Calculate the volume of H_2S at STP that is required to produce 1.00 kg of sulfur.

Answer: 466 L of H_2S at STP

Concept Exercise

If 2.00 L of hydrogen sulfide gas reacts with 2.00 L of oxygen gas at STP, which of the two gases is the limiting reactant?

Answer: See Appendix G, 15.3.

Note Multiple-reaction stoichiometry problems involve two or more balanced chemical equations. These stoichiometry problems can relate any of the quantities we have previously studied for single reactions, such as mass to mass, mass to volume, or volume to volume. Although these problems appear to be more complicated, multiple-reaction calculations only require an additional unit factor(s) in the unit analysis conversion.

15.5 Advanced Problem-Solving Examples

LEARNING OBJECTIVE

- Solve problems that involve two or more chemical principles.

The chemical calculations problems we have solved up to now involved only one chemical principle. Therefore, we have been able to solve them systematically by using a set of rules. Most of the time, we have applied the three basic steps in the unit analysis method of problem solving. Now it is time to tackle more challenging problems. The problems that follow involve two or more chemical principles. To solve them, we will have to analyze each problem carefully and then develop a strategy for its solution.

Although we have studied only introductory chemical principles, there is a very large number of combinations and variations of those principles. The problem examples that follow have been selected to illustrate the most important concepts we have studied. These examples, by no means, represent all the possible problems you may encounter.

Illustration: Metric System and Density

Let's consider a calculation that involves the metric system and density. We will need to perform a metric–English conversion and then apply the density concept.

EXAMPLE EXERCISE 15.4 Metric System and Density

Lithium is the least dense metallic element. Its density is 0.534 g/mL. What is the mass of lithium in a 1.00-inch cube of the metal?

Solution

We always begin a problem by determining the quantity asked for. Because the quantity is mass, we will choose units of *grams*. Next, we notice that the density is in metric units and the volume is in English units, in.^3. Let's plan the solution as two separate problems. First, we convert the volume from English to metric units.

$$\text{in.}^3\ \text{Li} \longrightarrow \text{cm}^3\ \text{Li} \longrightarrow \text{mL Li}$$

Second, we convert the volume of Li to mass of Li using density as a unit factor.

$$\text{mL Li} \longrightarrow \text{g Li}$$

The cube is 1.00 in. per side. Thus, the volume is: $1.00\ \text{in.} \times 1.00\ \text{in.} \times 1.00\ \text{in.} = 1.00\ \text{in.}^3$. Recall that $1\ \text{in.} = 2.54\ \text{cm}$, and $1\ \text{cm}^3 = 1\ \text{mL}$. The unit conversion is as follows:

$$1.00\ \cancel{\text{in.}^3} \times \left(\frac{2.54\ \cancel{\text{cm}}}{1\ \cancel{\text{in.}}}\right)^3 \times \frac{1\ \text{mL}}{1\ \cancel{\text{cm}^3}} = \text{mL Li}$$

$$1.00\ \cancel{\text{in.}^3} \times \frac{16.4\ \cancel{\text{cm}^3}}{1.00\ \cancel{\text{in.}^3}} \times \frac{1\ \text{mL}}{1\ \cancel{\text{cm}^3}} = 16.4\ \text{mL Li}$$

Because the density of lithium is 0.534 g/mL, 1 mL = 0.534 g. Converting from volume to mass, we have

$$16.4\ \cancel{\text{mL}} \times \frac{0.534\ \text{g}}{1\ \cancel{\text{mL}}} = 8.76\ \text{g Li}$$

For comparison, gold is one of the most dense metals in the periodic table. A 1.00-in. cube of gold weighs over 300 grams—more than 30 times the mass of the lithium cube!

Practice Exercise

Calculate the mass of water in a 2.00-in. cube of ice. The density of ice is $0.917\ \text{g/cm}^3$.

Answer: The strategy for solving this problem is as follows:

$$\text{in.}^3\ H_2O \longrightarrow \text{cm}^3\ H_2O \longrightarrow \text{g}\ H_2O$$

The ice cube has a volume of $8.00\ \text{in.}^3$, or $131\ \text{cm}^3$; its mass is 120 g.

Concept Exercise

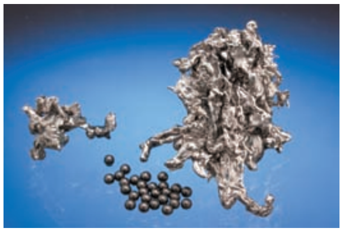
Which weighs more: a 1.00-cm cube of ice or $1.00\ \text{cm}^3$ of water?

Answer: See Appendix G, 15.4.

▲ **Lithium Metal, Li** Lithium metal is shown in mineral oil, which is floating on water.

Illustration: Measurement and the Mole Concept

Consider a problem that involves measurement and the mole concept. In the following example exercise we will use the concentration of lead in gasoline and the mole concept to calculate the number of lead atoms released into the environment.



▲ **Lead Metal, Pb**

EXAMPLE EXERCISE 15.5 Measurement and the Mole Concept

An off-road vehicle using leaded gasoline travels 125 miles, averaging 10.5 miles per gallon. How many lead atoms are released to the environment if leaded gasoline contains 0.10 g Pb/gal. (In 1996, the EPA banned leaded gasoline for on-road vehicles.)

Solution

We begin by identifying the quantity asked for. In this example, it is the number of lead atoms. With the given information, we can find the mass of lead as follows:

$$\text{mi} \longrightarrow \text{gal} \longrightarrow \text{g Pb}$$

We then convert the mass of Pb to the number of Pb atoms:

$$\text{g Pb} \longrightarrow \text{mol Pb} \longrightarrow \text{Pb atoms}$$

We are given the following information: 1 gal of gas = 10.5 mi; 0.10 g Pb = 1 gal of gas. Applying unit analysis,

$$125\ \cancel{\text{mi}} \times \frac{1\ \cancel{\text{gal}}}{10.5\ \cancel{\text{mi}}} \times \frac{0.10\ \text{g Pb}}{1\ \cancel{\text{gal}}} = 1.2\ \text{g Pb}$$

A mole of lead contains Avogadro's number of lead atoms. From the periodic table, we find the molar mass, 1 mol Pb = 207.2 g. Thus,

$$1.2\ \cancel{\text{g Pb}} \times \frac{1\ \cancel{\text{mol Pb}}}{207.2\ \cancel{\text{g Pb}}} \times \frac{6.02 \times 10^{23}\ \text{Pb atoms}}{1\ \cancel{\text{mol Pb}}} = 3.5 \times 10^{21}\ \text{Pb atoms}$$

Practice Exercise

An automobile travels 375 miles, averaging 24.0 mi/gal. If 1.00 gallon of gasoline produces 8184 g of CO_2, how many liters of carbon dioxide at STP are released during the trip?

Answer: The strategy for solving this problem is as follows:

$$\text{mi} \longrightarrow \text{gal} \longrightarrow \text{g CO}_2 \longrightarrow \text{mol CO}_2 \longrightarrow \text{L CO}_2$$

On the trip, 65,100 L (STP) of carbon dioxide are released.

Concept Exercise

Carbon dioxide is a "greenhouse" gas. What is meant by the term *greenhouse gas* and how does carbon dioxide contribute to global warming?

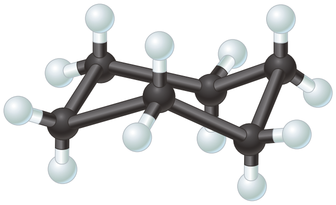
Answers: See Appendix G, 15.5.

Illustration: Empirical Formula and the Ideal Gas Law

Let's consider a calculation that involves empirical formula and the ideal gas law. We will first find the empirical formula of an organic vapor and then use a gas law calculation to find the molar mass and the actual formula of the compound.

EXAMPLE EXERCISE 15.6 Empirical Formula and the Ideal Gas Law

Cyclohexane is used as an organic solvent and fungicide. An elemental analysis of the compound showed its percent composition to be 85.7% C and 14.3% H. If the cyclohexane vapor in a 0.500-L flask at 100 °C and 0.995 atm weighed 1.37 g, what is the molar mass and molecular formula?



▲ **Cyclohexane Model**

Solution

In this problem we are asked to find the molecular formula. Recall that the molecular formula corresponds to a number of repeating empirical formula units; that is, $(C_xH_y)_n$. We must first determine the empirical formula from percentage composition. If we assume a 100-g sample of cyclohexane, there are 85.7 g C and 14.3 g H. We can calculate the empirical formula as follows:

$$85.7\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 7.14\ \text{mol C}$$

$$14.3\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{gH}}} = 14.2\ \text{mol H}$$

The empirical formula is in the ratio of $C_{7.14}H_{14.2}$. Dividing by the smaller mole value and simplifying, we have

$$C_{\frac{7.14}{7.14}}H_{\frac{14.2}{7.14}} = C_1H_{1.99} = C_1H_2$$

Next, let's calculate the molar mass of cyclohexane from the vapor density information. In this example, it is more convenient to use an algebraic approach. We can use the ideal gas law equation, $PV = nRT$. Recall that R is a constant with a value of 0.0821 atm · L/mol · K. Substituting the given information into the equation, we have

$$0.995\ \cancel{\text{atm}} \times 0.500\ \cancel{\text{L}} = n \times \frac{0.0821\ \cancel{\text{atm}} \cdot \cancel{\text{L}}}{1\ \text{mol} \cdot \cancel{\text{K}}} \times 373\ \cancel{\text{K}}$$

Rearranging and solving for n, we obtain

$$n = 0.0162\ \text{mol}$$

We are asked to find the molar mass which has the units of g/mol. The mass of vapor in the flask is given, 1.37 g. Therefore,

$$\frac{1.37\ \text{g}}{0.0162\ \text{mol}} = 84.6\ \text{g/mol}$$

Because the empirical formula is CH_2, the molecular formula is $(CH_2)_n$. The molar mass of the empirical formula unit, CH_2, is 14.03 g/mol. The molar mass of cyclohexane is 84.6 g/mol. Thus, the number of repeating empirical formula units in cyclohexane is

$$\text{cyclohexane:} \frac{(CH_2)_n}{CH_2} = \frac{84.6\ \text{g/mol}}{14.03\ \text{g/mol}}$$

$$n = 6.03 \approx 6$$

Because n equals 6, the molecular formula for cyclohexane is $(CH_2)_6$ and is written C_6H_{12}.

Practice Exercise

The percentage composition of a gaseous fuel is 92.3% C and 7.7% H. If 1.33 g of the gas occupies a volume of 1.25 L at 20 °C and 750 mm Hg, what is the (a) empirical formula and (b) molecular formula of the gas?

Answers:

(a) CH

(b) C_2H_2

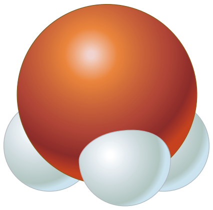
Concept Exercise

Rewrite the ideal gas law so that the molar mass (MM) of a gas is equal to the other variables in the ideal gas law equation.

Answer: See Appendix G, 15.6.

Illustration: Stoichiometry and the Combined Gas Law

Consider a problem that involves chemical equation calculation and the combined gas law. This problem is more complex than mass–volume stoichiometry because the gas produced is not at standard conditions.



▲ **Phosphine, PH_3**

EXAMPLE EXERCISE 15.7 Stoichiometry and the Combined Gas Law

Phosphine, PH_3, has a foul odor and irritates the respiratory system. The reaction of calcium phosphide and water gives phosphine gas and aqueous calcium hydroxide.

$$Ca_3P_2(s) + H_2O(l) \longrightarrow PH_3(g) + Ca(OH)_2(aq)$$

If 3.50 L of phosphine gas is collected at 749 mm Hg and 25 °C, what mass of calcium phosphide underwent reaction?

Solution

In every stoichiometry problem, we must begin with a balanced chemical equation. Balancing the above equation, we have

$$Ca_3P_2(s) + 6\ H_2O(l) \longrightarrow 2\ PH_3(g) + 3\ Ca(OH)_2(aq)$$

Stoichiometry problems involving gases must be at STP. In this example the PH_3 gas is not at standard conditions. Before starting the stoichiometry calculation, we must correct the volume of gas to 760 mm Hg and 0 °C. In Section 12.9, we solved gas law problems using a modified unit analysis approach. Let's begin with a data table showing the gas at initial and final conditions.

Conditions	*P*	*V*	*T*
initial	749 mm Hg	3.50 L	25 °C + 273 = 298K
final	760 mm Hg	V_2	273 K

The unit analysis format for gas law calculations is

$$3.50\ \text{L} \times P_{\text{factor}} \times T_{\text{factor}} = \text{L at STP}$$

Because the gas pressure increases at STP conditions, the volume decreases. Because the temperature decreases, the volume must decrease. Thus, the unit factor solution is

$$3.50\ \text{L} \times \frac{749\ \cancel{\text{mm Hg}}}{760\ \cancel{\text{mm Hg}}} \times \frac{273\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 3.16\ \text{L}$$

In this problem, we are asked to find the mass of Ca_3P_2. We notice that the volume of PH_3 is given. Thus, our strategy plan is as follows:

$$\text{L } PH_3 \longrightarrow \text{mol } PH_3 \longrightarrow \text{mol } Ca_3P_2 \longrightarrow \text{g } Ca_3P_2$$

From the balanced equation we see that 2 mol PH_3 are obtained from every 1 mol Ca_3P_2. From the periodic table we compute the molar mass of Ca_3P_2 is 182.3 g/mol. Let's calculate the mass of Ca_3P_2 as follows:

$$3.16\ \cancel{\text{L } PH_3} \times \frac{1\ \cancel{\text{mol } PH_3}}{22.4\ \cancel{\text{L } PH_3}} \times \frac{1\ \cancel{\text{mol } Ca_3P_2}}{2\ \cancel{\text{mol } PH_3}} \times \frac{182.18\ \text{g } Ca_3P_2}{1\ \cancel{\text{mol } Ca_3P_2}} = 12.9\ \text{g } Ca_3P_2$$

This problem involved several principles. However, by systematically applying the guidelines for solving simpler problems, we successfully solved a difficult problem.

Practice Exercise

The reaction of calcium metal and water gives aqueous calcium hydroxide and hydrogen gas. If 1.05 g of calcium reacts completely and the hydrogen gas is collected over water at 759 mm Hg and 20 °C, what is the STP volume of the gas? (The vapor pressure of water at 20 °C is 24 mm Hg.)

$$Ca(s) + 2\ H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$$

Answer: 529 mL (0.529 L) H_2 at STP

Concept Exercise

Is the volume of wet hydrogen gas collected greater or less than the same number of hydrogen molecules in dry hydrogen gas?

Answer: See Appendix G, 15.7.

Illustration: Molarity and Ionic Solutions

Let's consider a calculation that involves the molarity of an ionic solution. We will analyze the nitrate ion concentration in two aqueous solutions before and after mixing the two solutions.

EXAMPLE EXERCISE 15.8 Molarity and Ionic Solutions

If 25.0 mL of 0.500 *M* potassium nitrate solution are mixed with 75.0 mL of 0.125 *M* aluminum nitrate, what is the resulting molar concentration of the nitrate ion? (You can assume that the solution volumes are additive and the total volume is 100.0 mL.)

Solution

We are asked to find the molarity of nitrate ion after mixing aqueous solutions KNO_3 and $Al(NO_3)_3$. Let's begin with equations for the dissociation of ions in each solution. They are

$$KNO_3(aq) \longrightarrow K^+(aq) + NO_3^-(aq)$$
$$Al(NO_3)_3(aq) \longrightarrow Al^{3+}(aq) + 3\ NO_3^-(aq)$$

We are asked to find molarity, which has the units of mol/L. We know the volume after mixing is 100.0 mL. Therefore, we have to find the moles of NO_3^- from each solution. The calculations are as follows:

$$25.0\ \cancel{\text{mL}} \times \frac{0.500\ \cancel{\text{mol KNO}_3}}{1000\ \cancel{\text{mL solution}}} \times \frac{1\ \text{mol NO}_3^-}{1\ \cancel{\text{mol KNO}_3}} = 0.0125\ \text{mol NO}_3^-$$
$$75.0\ \cancel{\text{mL}} \times \frac{0.125\ \cancel{\text{mol Al(NO}_3)_3}}{1000\ \cancel{\text{mL solution}}} \times \frac{3\ \text{mol NO}_3^-}{1\ \cancel{\text{mol Al(NO}_3)_3}} = 0.0281\ \text{mol NO}_3^-$$

After combining the two solutions, the total moles of NO_3^- is

$$0.0125\ \text{mol NO}_3^- + 0.0281\ \text{mol NO}_3^- = 0.0406\ \text{mol NO}_3^-$$

The molar concentration of nitrate ion is found by dividing the moles of NO_3^- by the volume of solution, 100.0 mL. That is,

$$\frac{0.0406\ \text{mol NO}_3^-}{100.0\ \cancel{\text{mL solution}}} \times \frac{1000\ \cancel{\text{mL solution}}}{1\ \text{L solution}} = \frac{0.406\ \text{mol NO}_3^-}{1\ \text{L solution}} = 0.406\ M\ \text{NO}_3^-$$

Practice Exercise

If 50.0 mL of 0.100 *M* sodium chloride solution are mixed with 100.0 mL of 0.150 *M* calcium chloride, what is the resulting molar concentration of the chloride ion? (Assume the solution volumes are additive and the resulting total volume is 150.0 mL.)

Answer: 0.233 M Cl^-

Concept Exercise

If 1 mole each of NaCl, KCl, and AgCl are added into separate 1-L beakers of water, which of the solutions contains the least number of chloride ions?

Answer: See Appendix G, 15.8.

Illustration: Multiple-Reaction Stoichiometry

As a final example, let's consider a calculation involving two reactions. What makes this problem different, and somewhat more difficult, is that it does not ask for an unknown quantity. Instead, the problem asks for you to decide which of two substances is the most effective antacid tablet.

EXAMPLE EXERCISE 15.9 Multiple-Reaction Stoichiometry

Sodium bicarbonate and aluminum hydroxide are found in antacid tablets. Which antacid tablet neutralizes the most 0.100 *M* HCl stomach acid: a tablet containing 1.00 g $NaHCO_3$ or a tablet with 1.00 g $Al(OH)_3$?

Solution

In stoichiometry problems, we begin with a balanced chemical equation. There are two reactions, so we must write two balanced equations:

$$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$
$$Al(OH)_3(s) + 3\ HCl(aq) \longrightarrow AlCl_3(aq) + 3\ H_2O(l)$$

Because we are asked to find whether 1.00 g $NaHCO_3$ or 1.00 g $Al(OH)_3$ neutralizes the most acid, we have two problems. First, let's calculate the volume of acid consumed by the $NaHCO_3$. Second, let's find the volume of acid consumed by the $Al(OH)_3$.

▲ Antacid Tablet in Water

From the balanced equation we see that 1 mol $NaHCO_3$ reacts with 1 mol HCl. The molar mass of $NaHCO_3$ is found by adding the molar mass values in the periodic table, 84.01 g/mol. We can calculate the milliliters of acid as follows:

$$1.00\ \cancel{\text{g NaHCO}_3} \times \frac{1\ \cancel{\text{mol NaHCO}_3}}{84.01\ \cancel{\text{g NaHCO}_3}} \times \frac{1\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{mol NaHCO}_3}} \times \frac{1000\ \text{mL HCl}}{0.100\ \cancel{\text{mol HCl}}} = 119\ \text{mL HCl}$$

From the balanced equation we see that 1 mol $Al(OH)_3$ reacts with 3 mol HCl. The molar mass of $Al(OH)_3$ is 78.01 g/mol. The milliliters of acid that react is

$$1.00\ \cancel{\text{g Al(OH)}_3} \times \frac{1\ \cancel{\text{mol Al(OH)}_3}}{78.01\ \cancel{\text{g Al(OH)}_3}} \times \frac{3\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{mol Al(OH)}_3}} \times \frac{1000\ \text{mL HCl}}{0.100\ \cancel{\text{mol HCl}}} = 385\ \text{mL HCl}$$

Therefore, the $Al(OH)_3$ tablet neutralizes more stomach acid than the $NaHCO_3$ tablet.

Practice Exercise

Two antacid tablets are taken orally—one contains magnesium hydroxide and the other aluminum hydroxide. If each tablet contains 500 mg of antacid, what is the total volume of 0.100 *M* HCl neutralized by both tablets?

Answer: 171 mL + 192 mL = 363 mL of 0.100 *M* HCl

Concept Exercise

Which of the following antacid tablets neutralizes more acid: a 1.00-g $NaHCO_3$ tablet or a 1.00-g $KHCO_3$ tablet?

Answer: See Appendix G, 15.9.

Note In this section on advanced problem solving, we studied a few examples involving more than one chemical principle. In fact, there are thousands of possible chemical calculations based upon the principles we have learned. When we encounter more difficult problems, we can employ a combination of the strategies we learned in this chapter.

A CLOSER LOOK The International System of Units

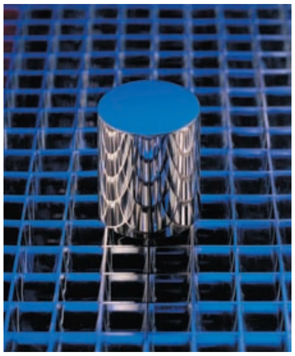
Q: Which reference standard is identical in both the metric system and SI?

In 1960, several nations, including the United States, agreed to establish an International System of Units (abbreviated SI). SI is an extension of the metric system and has seven base units. The base unit of length is the meter, and the six other SI quantities are mass, time, temperature, electric current, light intensity, and amount of substance. Volume is not a base unit, because it can be derived from cubic length (m^3).

The International System of Units (SI)

Quantity	Base Unit	Symbol
length	meter	m
mass	kilogram	kg
time	second	s
temperature	Kelvin	K
electric current	ampere	A
light intensity	candela	cd
amount of substance	mole	mol

It is important to note that SI units are defined in terms of natural phenomena and do not require physical reference standards. For example, a meter is defined as 1,650,763.73 wavelengths of the orange-red light emitted from excited atoms of krypton gas. Mass is the sole exception and still relies on the original international prototype kilogram, which serves as the reference standard for mass.



◀ **The Kilogram** The official reference kilogram was cast from platinum and iridium metals in 1790. The original casting is kept at the International Bureau of Weights and Measures outside Paris, France.

Currently, the U.S. Bureau of Weights and Measures is encouraging the use of SI units. In practice, scientists do not always follow SI conventions. It requires considerable effort to apply SI correctly, and many instruments are not calibrated in SI units. Moreover, SI disallows the familiar term "liter," and discourages use of the common prefixes *deci-* and *centi-*.

A: The original prototype kilogram is the reference standard for mass in both the metric system and SI.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
15.1 Advanced Problem-Solving Strategies We began our discussion with an overview of problem solving. Each time we approach a problem, we should ask three questions: What is the quantity asked for in the answer? What given information is relevant? What strategy can be used to convert the given information to the answer? We most often use *unit analysis* to solve a problem. However, gas law problems are better suited to an *algebraic analysis*. To design a solution to a problem we can also use a *concept map* and/or *visualization*.	• Describe several techniques for solving chemical problems, including: unit analysis, algebraic analysis, concept maps, and visualization. Estimate a "ballpark" answer before using a calculator to obtain a precise answer. ***Related Exercises: 1–10***
15.2 Concept Maps–Mole Relationships The central concept in all chemical formula calculations is the mole. Recall that a mole is related to each of the following: (1) number of entities, (2) molar mass, (3) molar volume, and (4) molar concentration. The mole concept can be used to relate the number of particles, the mass of a substance, the volume of a gas, or the concentration of an aqueous solution.	• Solve calculations that relate moles of substance to the following quantities: (a) number of atoms, molecules, or formula units (b) mass of substance (c) volume of gas ***Related Exercises: 11–22***
15.3 Concept Maps–Stoichiometry Stoichiometry relates quantities of substance based upon a balanced chemical equation. Direct stoichiometry relates a given amount of reactant to a product. Conversely, indirect stoichiometry relates a given amount of product to a reactant. Given a reactant or a product, we can relate (1) two masses of substance, (2) two volumes of gaseous substance, (3) two volumes of aqueous solution, or, various combinations thereof.	• Solve calculations that relate amounts of substance to the following quantities: (a) mass–mass stoichiometry (b) mass–volume stoichiometry (c) volume–volume stoichiometry (d) solution stoichiometry ***Related Exercises: 23–28***
15.4 Multiple-Reaction Stoichiometry Many industrial processes involve a series of two or more reactions. To perform stoichiometric calculations, we must relate a common quantity in all of the reactions. For example, sulfuric acid is manufactured from elemental sulfur in a series of three reactions. We can relate the element to the acid by keeping track of each of the intermediate products. In this case, sulfur is converted to sulfuric acid as follows: $S \longrightarrow SO_2 \longrightarrow SO_3 \longrightarrow H_2SO_4$ Quantities involving multiple reactions are related by extending the basic rules of stoichiometry. In the second step in a stoichiometry problem, moles of reactant are related to moles of product. For multiple reactions, this step is repeated for each additional reaction.	• Solve calculations that involve two or more chemical reactions. ***Related Exercises: 29–34***
15.5 Advanced Problem-Solving Examples This section presented some difficult introductory chemical calculations. For example, this section had problems involving two or more principles. Previously, problem solving was limited to a single principle and a specific problem-solving method. Now, you are asked to design a plan for finding a solution to a problem and to estimate a final answer. After calculating the final answer with a calculator, compare the calculated value to your estimated value to determine if it is reasonable.	• Solve calculations that involve two or more chemical principles. ***Related Exercises: 35–46***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** the amount of mass in one unit volume of matter

_____ **2.** the principle that states a compound always contains the same elements in the same proportion by mass

_____ **3.** a homogeneous mixture of a substance dissolved in water

_____ **4.** the value that corresponds to the number of atoms in 12.01 g of carbon

_____ **5.** the amount of substance that contains Avogadro's number of particles

_____ **6.** the mass of 1 mol of substance expressed in grams

_____ **7.** the volume occupied by 1 mol of gas at STP

_____ **8.** the principle that equal volumes of gases, at the same temperature and pressure, contain equal numbers of molecules

_____ **9.** a chemical formula that expresses the simplest whole-number ratio of atoms of each element in a molecule or of ions in a formula unit

_____ **10.** a chemical formula that expresses the actual number of atoms of each element in a molecule

_____ **11.** the ratio of moles of reactants and products according to the coefficients in the balanced chemical equation

_____ **12.** the statement that mass is neither gained or lost during a chemical reaction

_____ **13.** the relationship of quantities (i.e., mass of substance or volume of gas) in a chemical reaction according to the balanced chemical equation

_____ **14.** the substance in a chemical reaction that controls the maximum amount of product formed

_____ 15. a relationship between two variables such that when one variable doubles, the other variable doubles

_____ **16.** a relationship between two variables such that when one variable doubles, the other variable halves

_____ **17.** the statement that the pressure exerted by a gas is inversely proportional to its volume and directly proportional to its Kelvin temperature

_____ **18.** a temperature of 273 K and a pressure of 760 mm Hg for a gas

_____ **19.** the principle stated by the relationship $PV = nRT$

_____ **20.** the proportionality constant R in the equation $PV = nRT$

_____ **21.** a solution concentration expression that relates the mass of solute in grams dissolved in each 100 g of solution

_____ **22.** a solution concentration expression that relates the moles of solute dissolved in each liter of solution

(a) aqueous solution (*aq*)
(b) Avogadro's number
(c) Avogadro's theory
(d) combined gas law
(e) density
(f) directly proportional
(g) empirical formula
(h) ideal gas constant
(i) ideal gas law
(j) inversely proportional
(k) law of conservation of mass
(l) law of definite composition
(m) limiting reactant
(n) mass/mass percent (m/m %)
(o) molarity (*M*)
(p) molar mass (MM)
(q) molar volume
(r) mole
(s) mole ratio
(t) molecular formula
(u) standard temperature and pressure (STP)
(v) stoichiometry

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Advanced Problem-Solving Strategies (Sec. 15.1)

1. State the problem-solving strategy described by each of the following.
   (a) What is the term for writing a series of steps that connects the unknown quantity to a given quantity using unit factors?
   (b) What is the term for rearranging the variables in an equation in order to solve for the unknown variable?

2. State the problem-solving strategy described by each of the following.
   (a) What is the term for a drawing that shows the relationship between chemical concepts by connected boxes?
   (b) What is the term for forming a mental picture in order to understand a chemical concept more clearly?

3. State the basic metric system unit for each of the following physical quantities.
   (a) length
   (b) mass
   (c) volume
   (d) time
   (e) temperature
   (f) heat energy
4. State the basic SI unit for each of the following physical quantities.
   (a) length
   (b) mass
   (c) volume
   (d) time
   (e) temperature
   (f) heat energy
5. State the physical quantity corresponding to each of the following units.
   (a) cm
   (b) kg
   (c) cm^3
   (d) s
   (e) °C
   (f) kcal
6. State the physical quantity corresponding to each of the following units.
   (a) mm
   (b) ng
   (c) mm^3
   (d) μs
   (e) K
   (f) kJ
7. State the physical quantity corresponding to each of the following metric units.
   (a) m/s
   (b) g/mL
   (c) cal/(g × °C)
   (d) g solute/100 g solution
8. State the physical quantity corresponding to each of the following metric units.
   (a) km/h
   (b) kg/m^3
   (c) J/(kg × K)
   (d) mol solute/1 L solution
9. Estimate an approximate answer for each of the following calculations. Verify your ballpark answer using a calculator.

   (a) $$1.228 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{1 \text{ mol MgO}}{1 \text{ mol Mg}} \times \frac{40.31 \text{ g MgO}}{1 \text{ mol MgO}} = \text{? g MgO}$$

   (b) $$1.150 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} \times \frac{1 \text{ mol Ag}_2\text{O}}{2 \text{ mol Ag}} \times \frac{231.74 \text{ g Ag}_2\text{O}}{1 \text{ mol Ag}_2\text{O}} = \text{? g Ag}_2\text{O}$$

10. Estimate an approximate answer for each of the following calculations. Verify your ballpark answer using a calculator.

    (a) $$39.50 \text{ mL AlCl}_3 \times \frac{0.115 \text{ mol AlCl}_3}{1000 \text{ mL AlCl}_3} \times \frac{3 \text{ mol PbCl}_2}{2 \text{ mol AlCl}_3} \times \frac{278.1 \text{ g PbCl}_2}{1 \text{ mol PbCl}_2} = \text{? g PbCl}_2$$

    (b) $$42.05 \text{ mL HCl} \times \frac{0.195 \text{ mol HCl}}{1000 \text{ mL HCl}} \times \frac{1 \text{ mol Ba(OH)}_2}{2 \text{ mol HCl}} \times \frac{1000 \text{ mL Ba(OH)}_2}{0.105 \text{ mol Ba(OH)}_2} = \text{? mL Ba(OH)}_2$$

Concept Maps—Mole Relationships (Sec. 15.2)

11. Draw a concept map that relates each of the following mole quantities.
    (a) molecules of sulfur dioxide gas and volume SO_2 at STP
    (b) volume of nitrogen gas at STP and mass N_2
12. Draw a concept map that relates each of the following mole quantities.
    (a) mass of chloroform and molecules of chloroform, $CHCl_3$
    (b) volume of hydrochloric acid and mass HCl
13. Draw a concept map that relates each of the following stoichiometric quantities.
    (a) mass of reactant and mass of product
    (b) mass of reactant and volume of gaseous product
14. Draw a concept map that relates each of the following stoichiometric quantities.
    (a) volume of gaseous reactant and volume of gaseous product
    (b) volume of gaseous product and volume of gaseous reactant
15. Given that 0.0142 mol of hydrogen chloride gas dissolves in 47.5 mL of solution, draw a mole concept map and calculate each of the following.
    (a) grams of HCl gas (STP) dissolved in the solution
    (b) liters of HCl gas (STP) dissolved in the solution
    (c) molecules of HCl gas dissolved in the solution
    (d) molar concentration of the hydrochloric acid solution
16. Given that 0.0755 mol of ammonia gas dissolves in 0.155 L of solution, draw a mole concept map and calculate each of the following.
    (a) grams of NH_3 gas (at STP) dissolved in the solution
    (b) liters of NH_3 gas (at STP) dissolved in the solution
    (c) molecules of NH_3 gas dissolved in the solution
    (d) molar concentration of the ammonia solution
17. Given that 1.00 g of hydrogen fluoride dissolves in 100.0 mL of solution, draw a mole concept map and calculate each of the following.
    (a) liters of HF gas (STP) dissolved in the solution
    (b) molecules of HF gas dissolved in the solution
    (c) molar concentration of the hydrofluoric acid solution
18. Given that 2.00 g of hydrazine, N_2H_4, dissolves in 0.250 L of solution, draw a mole concept map and calculate each of the following.
    (a) liters of N_2H_4 gas (at STP) dissolved in the solution
    (b) molecules of N_2H_4 gas dissolved in the solution
    (c) molar concentration of the hydrazine solution
19. Given that 1.00 L of carbon dioxide at STP dissolves in 500.0 mL of solution, draw a mole concept map and calculate each of the following.
    (a) grams of CO_2 gas (at STP) dissolved in the solution
    (b) molecules of CO_2 gas dissolved in the solution
    (c) molar concentration of the carbonic acid, H_2CO_3, solution
20. Given that 555 mL of sulfur dioxide at STP dissolves in 0.250 L of solution, draw a mole concept map and calculate each of the following.
    (a) grams of SO_2 gas (at STP) dissolved in the solution
    (b) molecules of SO_2 gas dissolved in the solution
    (c) molar concentration of the sulfurous acid, H_2SO_3, solution
21. Given that 2.22×10^{22} molecules of hydrogen sulfide dissolve in 450.0 mL of solution, draw a mole concept map and calculate each of the following.
    (a) liters of H_2S gas (STP) dissolved in the solution
    (b) grams of H_2S gas dissolved in the solution
    (c) molar concentration of the hydrosulfuric acid, H_2S, solution

22. Given that 1.12×10^{23} molecules of sulfur trioxide dissolve in 0.325 L of solution, draw a mole concept map and calculate each of the following.
(a) liters of SO_3 gas (STP) dissolved in the solution
(b) grams of SO_3 gas dissolved in the solution
(c) molar concentration of the sulfuric acid, H_2SO_4, solution

Concept Maps—Stoichiometry (Sec. 15.3)

23. Methane gas, CH_4, reacts with oxygen to give 5.00 g of water according to the chemical reaction:

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

Draw a stoichiometry concept map and calculate:
(a) the mass of methane that reacted
(b) the volume of carbon dioxide that was produced at STP

24. Ammonia gas, NH_3, reacts with oxygen and platinum catalyst to give 25.0 L of nitrogen monoxide gas at STP according to the chemical reaction:

$$NH_3(g) + O_2(g) \xrightarrow{Pt/825\,^{\circ}C} NO(g) + H_2O(l)$$

Draw a stoichiometry concept map and calculate:
(a) the volume of ammonia that reacted at STP
(b) the mass of water that was produced

25. Hydrochloric acid reacts with 0.466 g aluminum metal to give hydrogen gas according to the chemical reaction:

$$Al(s) + HCl(aq) \rightarrow AlCl_3(aq) + H_2(g)$$

Draw a stoichiometry concept map and calculate:
(a) the volume of hydrogen gas produced at STP
(b) the volume of 0.100 *M* HCl required for complete reaction

26. A sample of potassium hydrogen carbonate decomposes to give 255 mL of carbon dioxide at STP according to the chemical reaction:

$$KHCO_3(s) \rightarrow K_2CO_3(aq) + H_2O(l) + CO_2(g)$$

Draw a stoichiometry concept map and calculate:
(a) the mass of $KHCO_3$ decomposed
(b) the mass of K_2CO_3 produced

27. Given that 50.0 mL of 0.100 *M* magnesium bromide reacts completely with 13.9 mL of silver nitrate solution according to the chemical reaction:

$$MgBr_2(aq) + AgNO_3(aq) \rightarrow AgBr(s) + Mg(NO_3)_2(aq)$$

Draw a stoichiometry concept map and calculate:
(a) the molarity of the $AgNO_3$ solution
(b) the mass of AgBr precipitate produced

28. Given that 24.0 mL of 0.170 *M* sodium iodide reacts completely with 0.209 *M* mercury(II) nitrate solution according to the chemical reaction:

$$Hg(NO_3)_2(aq) + NaI(aq) \rightarrow HgI_2(s) + NaNO_3(aq)$$

Draw a stoichiometry concept map and calculate:
(a) the volume of $Hg(NO_3)_2$ solution that reacted
(b) the mass of HgI_2 precipitate that was produced

Multiple-Reaction Stoichiometry (Sec. 15.4)

29. Carbon monoxide gas is used in the blast furnace process to convert iron ore to impure pig iron. Carbon dioxide is used as a fire extinguisher and to make dry ice. Starting with coal, carbon undergoes the following reactions.

$$2\,C(g) + O_2(g) \rightarrow 2\,CO(g)$$
$$2\,CO(g) + O_2(l) \rightarrow 2\,CO_2(g)$$

Starting with 25.0 g of carbon and excess oxygen gas, calculate:
(a) the mass of carbon dioxide produced
(b) the volume of carbon dioxide produced at STP
(c) the mass of oxygen consumed in the process

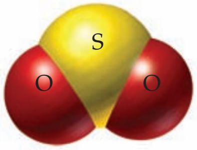
30. Ground water sometimes contains traces of hydrogen sulfide, which has the odor of rotten eggs. Chlorine gas is used to purify the water for drinking. The resulting sulfur reacts with fluorine gas to give sulfur hexafluoride. The reactions are as follows:

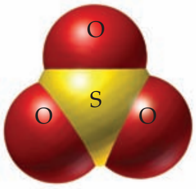
$$8\,H_2S(aq) + 8\,Cl_2(g) \rightarrow 16\,HCl(aq) + S_8(s)$$
$$S_8(s) + 24\,F_2(g) \rightarrow 8\,SF_6(g)$$

Starting with 0.950 L of Cl_2 (STP) and excess fluorine gas, calculate:
(a) the mass of sulfur hexafluoride produced
(b) the volume of sulfur hexafluoride produced at STP
(c) the volume of aqueous 0.0265 *M* H_2S that reacted

31. Sulfuric acid is the single most important industrial chemical. About 40 million tons are manufactured each year. Elemental sulfur is converted to sulfuric acid by the Contact Process as follows:

$$S(g) + O_2(g) \rightarrow SO_2(g)$$
$$2\,SO_2(g) + O_2(g) \xrightarrow{V_2O_5} 2\,SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$





Starting with 1.00 kg of sulfur and excess oxygen gas, calculate:
(a) the mass of sulfur trioxide produced
(b) the STP volume of sulfur trioxide produced
(c) the mass of sulfuric acid produced assuming a 55.0% process yield

32. Iron ore is converted to pig iron in an industrial blast furnace. The pig iron, in turn, is converted to carbon steel by high-temperature oxidation. The blast furnace process takes place in a series of three reactions.

$$3\ \mathrm{Fe_2O_3}(s) + \mathrm{CO}(g) \xrightarrow{200\ ^\circ\mathrm{C}} 2\ \mathrm{Fe_3O_4}(s) + \mathrm{CO_2}(g)$$

$$\mathrm{Fe_3O_4}(s) + \mathrm{CO}(g) \xrightarrow{700\ ^\circ\mathrm{C}} 3\ \mathrm{FeO}(s) + \mathrm{CO_2}(g)$$

$$\mathrm{FeO}(g) + \mathrm{CO}(g) \xrightarrow{1200\ ^\circ\mathrm{C}} \mathrm{Fe}\ (l) + \mathrm{CO_2}(g)$$

Starting with 1.00 kg of iron(III) oxide and excess carbon monoxide gas, calculate:

(a) the kilogram mass of iron(II) oxide produced
(b) the kilogram mass of iron produced
(c) the kilogram mass of iron produced assuming a 70.0% process yield

33. Steering rockets on the Space Shuttle are powered by the reaction of hydrazine, N_2H_4, and dinitrogen tetraoxide, N_2O_4. The reactions for making hydrazine and the reaction with dinitrogen tetraoxide are as follows.

$$2\ \mathrm{NH_3}(aq) + \mathrm{NaOCl}(aq) \rightarrow \mathrm{N_2H_4}(aq) + \mathrm{NaCl}(aq) + \mathrm{H_2O}(l)$$

$$2\ \mathrm{N_2H_4}(l) + \mathrm{N_2O_4}(l) \rightarrow 3\ \mathrm{N_2}(g) + 4\ \mathrm{H_2O}(g)$$

Starting with 50.0 mL of 6.00 *M* ammonia and excess other reactants, calculate:

(a) the mass of nitrogen gas produced
(b) the volume of nitrogen gas produced at STP
(c) the mass of water produced

34. Hydrogen peroxide, H_2O_2, is prepared industrially from the reaction of oxygen and isopropyl alcohol, C_3H_7OH. Hydrogen peroxide reacts explosively with hydrazine, N_2H_4, and is used as a rocket fuel. The reactions for manufacturing hydrazine and the rocket explosion are as follows.

$$\mathrm{C_3H_7OH}(l) + \mathrm{O_2}(g) \rightarrow \mathrm{H_2O_2}(l) + \mathrm{C_3H_6O}(l)$$

$$2\ \mathrm{H_2O_2}(l) + \mathrm{N_2H_4}(l) \rightarrow \mathrm{N_2}(g) + 4\ \mathrm{H_2O}(g)$$

Starting with 50.0 mL of isopropyl alcohol (d = 0.786 g/mL) and excess other reactants, calculate:

(a) the mass of nitrogen gas produced
(b) the volume of nitrogen gas produced at STP
(c) the mass of water produced

Advanced Problem-Solving Examples (Sec. 15.5)

35. Calculate the molar mass of an unknown gas if an experiment yielded a value of 1.45 g/L at 100 °C and 0.989 atm.

36. Calculate the mass of nitrogen dioxide gas occupying a volume of 2.50 L at 35 °C and 0.974 atm pressure.

37. A 34.5-g sample of calcium chloride is dissolved in 500.0 mL of solution. Find the molar concentration of the calcium ions and chloride ions in solution. (Assume the salt is 100% ionized.)

38. What is the molar chloride ion concentration that results from mixing 100.0 mL of 0.156 *M* lithium chloride and 150.0 mL of 0.225 *M* barium chloride?

39. How many milliliters of 0.100 *M* hydrochloric acid react with excess zinc metal in order to collect 50.0 mL of hydrogen gas over water at STP?

40. What volume of 0.150 *M* hydrochloric acid reacts with excess lead(II) nitrate solution in order to yield 1.88 g of lead(II) chloride precipitate?

41. Given 0.150 L of nitrous oxide at 25 °C and 749 mm Hg pressure, calculate the number of N_2O molecules. What is the mass of the gas?

42. Given 6.55 g of carbon dioxide gas, calculate the volume of gas at 75 °C and 0.750 atm pressure. What is the number of gas molecules?

43. Mannitol is an artificial sweetener found in sugarless gum. The percentage composition is 39.6% carbon, 7.77% hydrogen, and 52.8% oxygen and the density is 1.47 g/mL. If mannitol contains 4.93×10^{21} molecules per milliliter, what is the molecular formula of mannitol?

44. Liquid bromine can be prepared by passing chlorine gas through an aqueous solution of sodium bromide. How many milliliters of bromine are produced from the reaction of 10.0 g of Cl_2 and 10.0 g of NaBr? (Given: the density of liquid Br_2 is 3.12 g/mL.)

45. Heating solid sodium chlorate with manganese dioxide catalyst produces solid sodium chloride and oxygen gas. If 245 mL of *wet* oxygen are collected over water at 25 °C and 754 mm Hg, what mass of sodium chlorate was decomposed? The vapor pressure of water at 25 °C is 24 mm Hg.

46. Heating solid potassium chlorate with manganese dioxide catalyst produces solid potassium chloride and oxygen gas. If 455 mL of *wet* oxygen are collected over water at 23 °C and 766 mm Hg, what mass of potassium chlorate was decomposed? The vapor pressure of water at 23 °C and 21 mm Hg.

General Exercises

47. If a pencil manufacturer produces 288,125 pencils, how many gross of pencils can be sold? (A gross is 144 pencils.)

48. If a paper manufacturer produces 4,125,365 sheets of paper, how many reams of paper can be sold? (A ream is 500 sheets.)

49. Given that 45.0 mL of nitrogen gas react with 65.0 mL of oxygen gas according to the unbalanced equation:

$$\mathrm{N_2}(g) + \mathrm{O_2}(g) \rightarrow \mathrm{NO_2}(g)$$

(a) What is the limiting reactant?
(b) Assuming constant conditions, what is the volume of NO_2?

50. Given that 25.0 L of sulfur dioxide reacts with 25.0 L of oxygen gas according to the unbalanced equation:

$$\mathrm{SO_2}(g) + \mathrm{O_2}(g) \xrightarrow{\mathrm{V_2O_5}/650\ ^\circ\mathrm{C}} \mathrm{SO_3}(g)$$

(a) What is the limiting reactant?
(b) Assuming constant conditions, what is the volume of SO_3?

51. Given that 1.00 g of aluminum hydroxide react with 25.0 mL of 0.500 M sulfuric acid according to the unbalanced equation:

$$\mathrm{Al(OH)_3}(s) + \mathrm{H_2SO_4}(l) \rightarrow \mathrm{Al_2(SO_4)_3}(aq) + \mathrm{H_2O}(l)$$

(a) What is the limiting reactant?
(b) What is the mass of H_2O produced?

52. Given that 36.5 mL of 0.266 M calcium acetate react with 25.0 mL of 0.385 M sodium carbonate according to the unbalanced equation:

$$Ca(C_2H_3O_2)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + NaC_2H_3O_2(aq)$$

(a) What is the limiting reactant?
(b) What is the mass of $CaCO_3$ precipitate formed?

53. Seawater contains approximately 1.22×10^{10} atoms of gold per milliliter. How many kilograms of seawater must be evaporated to obtain 12.0 g of gold? (Assume the density of seawater is 1.05 g/mL.)

54. The volume of the oceans is 1.4×10^{21} liters. Calculate the mass of chlorine in all seawater given the mass percent concentration of chloride ion is 1.90%. (Assume the density of seawater is 1.05 g/mL.)

55. What is the molar sodium ion concentration resulting from mixing of 50.0 mL of 0.100 *M* sodium chloride and 50.0 mL of 0.200 *M* sodium sulfate?

56. What is the molar chloride ion concentration resulting from mixing of 50.0 mL of 0.100 *M* sodium chloride and 25.0 mL of 0.100 *M* potassium chloride?

57. A sample of sodium carbonate is treated with 50.0 mL of 0.345 *M* HCl. The excess hydrochloric acid is titrated with 15.9 mL of 0.155 *M* NaOH. Calculate the mass of the sodium carbonate sample.

$$Na_2CO_3(s) + 2\,HCl(aq) \rightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

58. The chloride in an aqueous sample of $BaCl_2$ is precipitated with 50.0 mL of 0.100 *M* $AgNO_3$. The excess silver nitrate is titrated with 17.0 mL of 0.125 *M* K_2CrO_4. Calculate the mass of the barium chloride in the sample.

$$BaCl_2(aq) + 2\,AgNO_3(aq) \rightarrow 2\,AgCl(s) + Ba(NO_3)_2(aq)$$

$$2\,AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2\,KNO_3(aq)$$

Challenge Exercises

59. The president of a college reports to a board of trustees, and three vice-presidents report to the college president. The chief financial officer reports directly to the board of trustees. Draw a concept map for the college administration.

60. A computer application program can transfer information to a word processing document, a data base, and a spreadsheet. Draw a concept map for the computer application.

Chapter 15 Self-Test

Answers to Self-Test are in Appendix J.

1. After reading a chemistry problem, what is always the first step? (Sec. 15.1)
(a) determine the relevant given value
(b) determine the unknown quantity
(c) write down unit factors
(d) estimate a "ballpark" answer
(e) consult a reference source

2. What are the units of the unknown quantity for the following problem? Calculate the molar mass of oxygen gas given that 1.00 g of gas occupies 0.700 L at STP. (Sec. 15.1)
(a) g/L
(b) L/g
(c) mol/g
(d) g/mol
(e) none of the above

3. Which of the following units is an expression of gas pressure? (Sec. 15.1)
(a) atm
(b) kPa
(c) mm Hg
(d) psi
(e) all of the above

4. After organizing a solution to a chemistry calculation, what should be done before using a calculator? (Sec. 15.1)
(a) decide on a problem-solving strategy
(b) determine the unknown quantity
(c) estimate a ballpark answer
(d) refer to a reference book
(e) write a balanced chemical equation

5. If the unknown quantity is g/cm^3, what is true of the relevant given value? (Sec. 15.1)
(a) the given value is a ratio of two units
(b) the given value is expressed in metric units
(c) the given value contains an exponent
(d) all of the above
(e) none of the above

6. If 1.12 L of CO_2 gas at STP is dissolved in 75.0 mL of aqueous solution, what is the number of carbon dioxide molecules dissolved in solution? (Sec. 15.2)
(a) 2.39×10^{22} molecules
(b) 3.01×10^{22} molecules
(c) 4.17×10^{23} molecules
(d) 1.20×10^{25} molecules
(e) 1.51×10^{25} molecules

7. If a sample of methane gas, CH_4, reacts completely with oxygen gas to give 1.75 g of water, what STP volume of methane reacted? (Sec. 15.3)

$$CH_4(g) + O_2(g) \xrightarrow{\text{spark}} CO_2(g) + H_2O(l)$$

(a) 0.352 L
(b) 0.703 L
(c) 1.09 L
(d) 2.18 L
(e) 4.35 L

8. Given the three-step "blast furnace" process that converts iron ore, Fe_2O_3, into molten "pig iron":

$$3\,Fe_2O_3(l) + CO(g) \rightarrow 2\,Fe_3O_4(l) + CO_2(g)$$
$$Fe_3O_4(l) + CO(g) \rightarrow 3\,FeO(l) + CO_2(g)$$
$$FeO(l) + CO(g) \rightarrow Fe(l) + CO_2(g)$$

What mass of Fe_2O_3 produces 1.00 metric ton (1000 kg) of iron? (Sec. 15.4)
(a) 175 kg
(b) 699 kg
(c) 715 kg
(d) 1430 kg
(e) 2860 kg

9. What is the mass of a 1.00-inch cube of magnesium ($d = 1.74\ \text{g/cm}^3$)? (Sec. 15.5)
(a) 0.575 g
(b) 1.74 g
(c) 9.43 g
(d) 16.4 g
(e) 28.5 g

10. How many water molecules are in a 1.00-mm^3 cube of ice ($d = 0.917\ \text{g/cm}^3$)? (Sec. 15.5)
(a) 9.17×10^{-4} molecules
(b) 3.06×10^{19} molecules
(c) 3.06×10^{22} molecules
(d) 5.09×10^{18} molecules
(e) 5.09×10^{20} molecules

Key Concepts

11. What problem-solving technique converts a given value to an unknown value by applying one or more unit factors?

12. What problem-solving technique requires rearranging variables in an equation to calculate an unknown quantity?

13. What problem-solving technique traces the relationship of chemical concepts, often shown by connected boxes?

14. What problem-solving technique helps to understand concepts that cannot be observed directly, such as the molecules in a gas?

Critical Thinking

15. If the unknown quantity in a unit analysis solution to a problem has compound units, for example, kg/L, what is true of the relevant given value?

16. State a reference for each of the following:
(a) molar mass of a substance
(b) physical properties of a substance

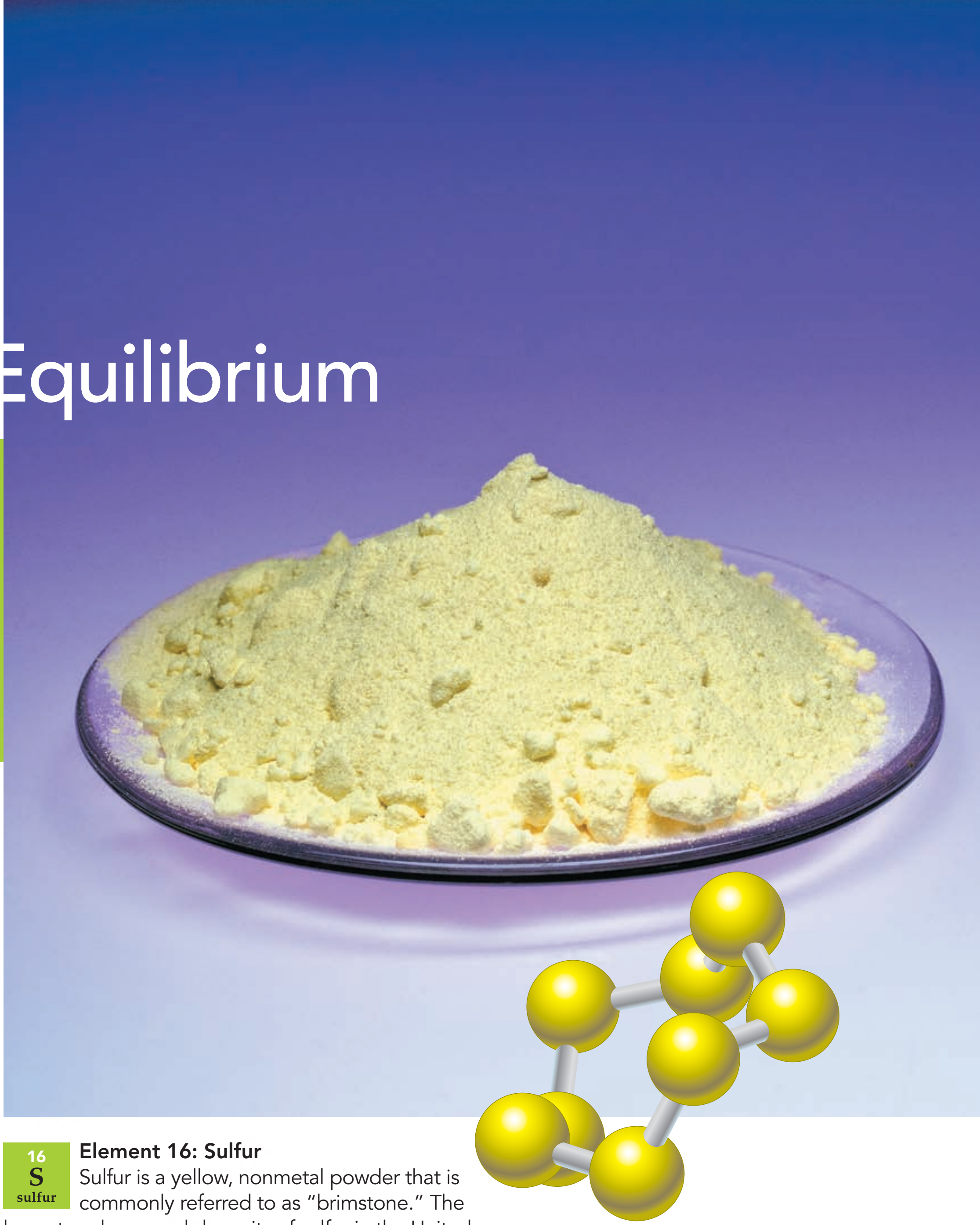
CHAPTER

16

Equilibrium

"Do not think there is any other quality so essential to success of any kind as the quality of perseverance."

John D. Rockefeller, Founder of Standard Oil



16
S
sulfur

Element 16: Sulfur
Sulfur is a yellow, nonmetal powder that is commonly referred to as "brimstone." The largest underground deposits of sulfur in the United States are in Texas and Louisiana. The most important industrial chemical, sulfuric acid, is manufactured by converting S to SO_3, and dissolving in water to give aqueous H_2SO_4.

We usually assume that a chemical reaction continues until the reactants are used up and the reaction stops. In reality, most reactions continue to occur as an ongoing, reversible process. That is, a reaction occurs in the forward direction to give products, and simultaneously a reverse reaction occurs in the opposite direction to give the original reactants.

As an example of a reversible reaction, let's consider a chemical change that gives a precipitate (Figure 16.1). Initially, insoluble particles form in the aqueous solution. Then a very small number of these particles dissolve and go back into the solution. Since some particles are precipitating while others are dissolving, the ongoing process is said to be dynamic. We can indicate this reversible chemical reaction as follows:

$$\text{soluble substance} \underset{\textit{dissolving}}{\overset{\textit{precipitating}}{\rightleftarrows}} \text{insoluble particles}$$

◀ **Figure 16.1 A Reversible Reaction** The reaction of lead(II) nitrate and potassium iodide produces insoluble particles of yellow lead(II) iodide precipitate. Simultaneously, slightly soluble particles of lead(II) iodide dissolve in the solution, while other yellow particles precipitate from the solution.

Most chemical reactions involve a dynamic, reversible change. As reactants undergo a reaction to form products, the products react with each other to give the original reactants. Initially, the rate of the forward reaction is rapid. However, as the reaction proceeds, the concentrations of the reactants decrease and the concentrations of the products increase. With this change in concentrations, the rate of the forward reaction slows down, and the rate of the reverse reaction speeds up.

When the forward reaction and the reverse reaction are taking place at the same rate, the reaction is at equilibrium. We can show these dynamic, reversible processes as follows:

$$\text{reactants} \quad \underset{\textit{reverse reaction}}{\overset{\textit{forward reaction}}{\rightleftharpoons}} \quad \text{products}$$

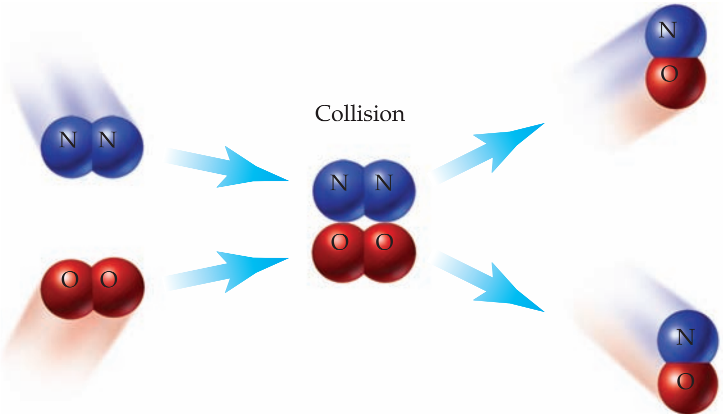
It is important to keep in mind that a chemical change is a *dynamic, reversible process.* That is, in most chemical changes both the forward and reverse reactions continue to take place at the same time.

16.1 Collision Theory

Chemists propose a theoretical model for reactions whereby molecules randomly collide and react. If a collision is successful, molecules of reactants are changed into products. In a successful collision, existing bonds in a molecule are broken and new bonds are formed. This model is referred to as the **collision theory** of reactions. In a successful collision, molecules go through a transition state in which bonds are rearranged. Figure 16.2 illustrates the collision theory by showing molecules of nitrogen and oxygen colliding to form two molecules of nitrogen monoxide.

LEARNING OBJECTIVES

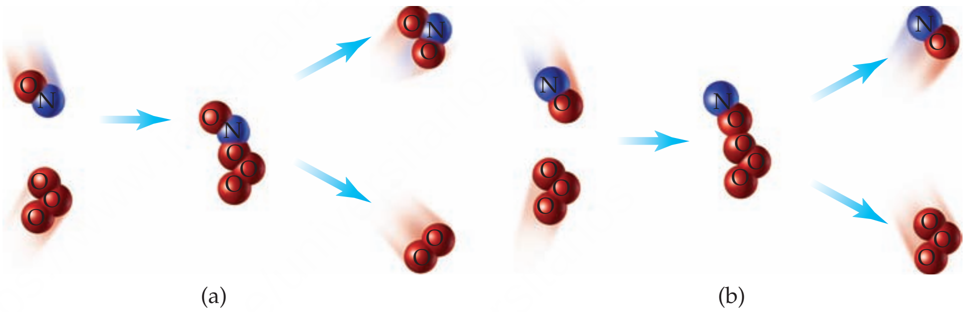
- Explain the effect of collision frequency, collision energy, and orientation of molecules on the rate of a chemical reaction.
- Explain the effect of concentration, temperature, and a catalyst on the rate of a chemical reaction.



◀ **Figure 16.2 Collision Theory** Nitrogen and oxygen gas molecules collide to form a temporary transition state. If the collision is successful, a molecule of N_2 and a molecule of O_2 produce two molecules of NO.

Collision theory proposes that three factors affect the rate of a chemical reaction. That is, three factors affect the rate of effective collisions: (1) collision frequency, (2) collision energy, and (3) orientation of the molecules. Let's discuss each factor in more detail.

1. **Collision Frequency.** When we increase the frequency of collision between molecules, we increase the rate of the reaction. One way to increase the collision frequency is to increase the concentrations of molecules. A second way is to increase the temperature. As the temperature increases, the molecules move at a faster velocity, and therefore collide more often.
2. **Collision Energy.** Even though molecules collide, they may not react. For a reaction to occur, molecules must collide with sufficient energy to break bonds. When we increase the temperature, the molecules move at a faster velocity, and the collisions between molecules are more energetic. Thus, the rate of a reaction increases as the energy of collision increases.
3. **Collision Orientation.** Even if molecules collide with a lot of energy, this does not assure that a reaction will occur. For a reaction to take place, the molecules must have a favorable orientation to have an effective collision. Otherwise, the molecules bounce off one another without reacting. Figure 16.3 illustrates (a) an effective collision, and (b) an ineffective collision.



▶ Figure 16.3 **Effective versus Ineffective Collisions** (a) An effective collision of gaseous NO and O_3 produces NO_2 and O_2. (b) An ineffective collision of NO and O_3 molecules does not give products. If molecules do not have the correct orientation, they simply bounce off each other without forming new molecules.

Effects of Concentration, Temperature, and Catalyst

So far we have described the factors that influence the rate of reaction on the conceptual level. Experimentally, three factors have been identified that affect the speed at which a chemical reaction proceeds. These factors are the concentration of reactants, the temperature of the reaction, and the presence of a catalyst. Let's consider each of these factors.

1. **Reactant Concentration.** When we increase the concentration of a reactant, the molecules are closer together and collide more frequently. It therefore follows that a chemical reaction proceeds faster as the concentrations of the reactants increase.
2. **Reaction Temperature.** If we increase the temperature of the reaction, we increase the energy of the molecules. Increasing the temperature affects the rate of reaction in two ways. First, the collision frequency increases, and the reaction speeds up. Second, the collision energy is greater, and there are more molecules with enough energy to break bonds and form new molecules. Thus, as the temperature increases, the rate of reaction increases because of increased collision frequency and greater collision energy.
3. **Catalyst.** If we add a catalyst, we increase the rate of a chemical reaction. A catalyst speeds up a reaction by increasing the number of effective collisions. For example, in reactions involving hydrogen gas the metals Pt, Ni, and Zn can be used as catalysts. The rate of reaction increases because H_2 molecules are attracted to the surface of the metal catalyst, creating a more stable and favorable *collision orientation.*

EXAMPLE EXERCISE 16.1 Rate of Reaction

According to collision theory, which of the following factors influences the rate of a chemical reaction?

(a) frequency of molecular collisions
(b) energy of molecular collisions
(c) orientation of molecular collisions

Solution

Each of the factors affects the rate of reaction as follows:

(a) Increasing collision frequency increases the rate of reaction.
(b) Increasing collision energy increases the rate of reaction.
(c) A favorable molecular orientation increases the rate of reaction.

Practice Exercise

State three experimental variables that can be changed to increase the rate of a chemical reaction.

Answers: The rate of reaction can be increased by (1) increasing the concentration of a reactant, (2) raising the reaction temperature, or (3) introducing a catalyst.

Concept Exercise

When temperature increases, the rate of effective collision increases for two reasons. Explain the two reasons.

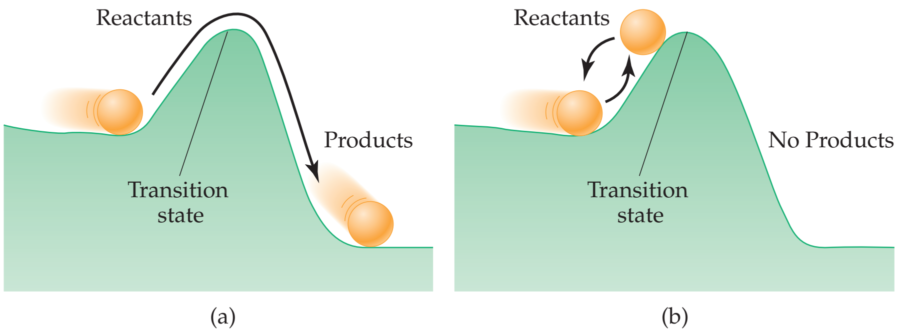
Answers: See Appendix G, 16.1.

16.2 Energy Profiles of Chemical Reactions

LEARNING OBJECTIVES

- Diagram the general energy profile for an endothermic and an exothermic reaction.
- Label the transition state, energy of activation, and heat of reaction on a given energy profile.

In air, nitrogen and oxygen gases do not react. In a hot automobile engine, however, the two gases react to produce nitrogen monoxide. To react, molecules of N_2 and O_2 must collide with sufficient energy to achieve the transition state before forming products. Since the transition state is a high-energy point, there is an energy barrier between the reactants and products. Figure 16.4 shows an analogy illustrating the energy changes that occur during a chemical reaction.

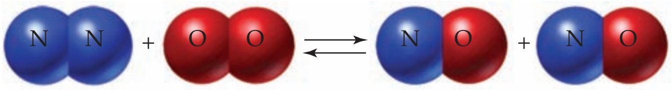


▲ **Figure 16.4 Energy Barrier Analogy for a Chemical Reaction** (a) The ball symbolizes molecules in a reaction that have sufficient energy to give products. If the ball has enough energy, it reaches the transition point at the top of the hill and rolls down the other side. (b) The ball symbolizes a reaction in which the molecules do not have sufficient energy to give products. If the ball does not have enough energy, it cannot reach the transition point and roll down the other side of the hill.

Endothermic Reaction Profiles

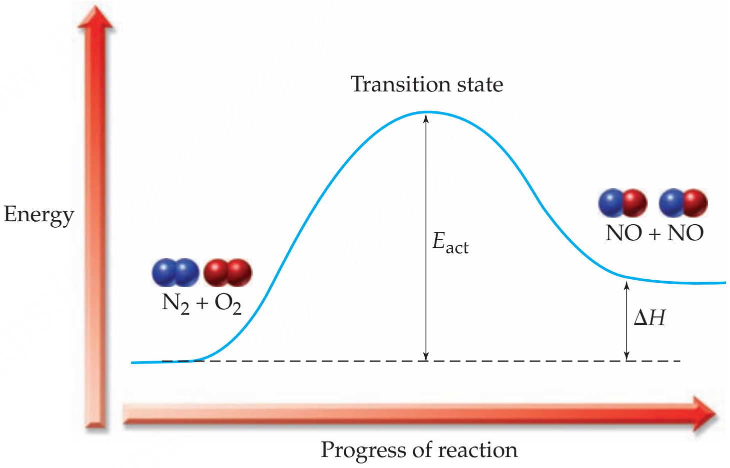
An **endothermic reaction** absorbs heat and proceeds by consuming heat energy. As an example, consider the formation of nitrogen monoxide from nitrogen and oxygen gases. The reaction is

$$N_2(g) + O_2(g) + heat \rightleftharpoons 2\,NO(g)$$



We can follow the progress of a chemical reaction by drawing an energy profile of reactants and products. A **reaction profile** shows the energy of reactants and products during the course of a reaction. The highest point on the reaction profile is the **transition state**, where bonds are breaking, and simultaneously other bonds are forming.

A reaction profile for the reaction of nitrogen and oxygen is shown in Figure 16.5. The energy required for the reactants to achieve the transition state is called the **activation energy** (symbol E_{act}). The energy difference between reactants and products is termed the **heat of reaction** (symbol ΔH). For endothermic reactions energy is absorbed, and the products end up at a higher potential energy level than the reactants.

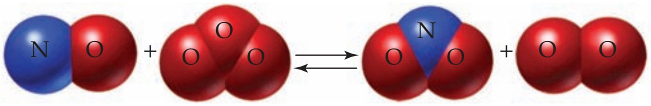


▶ **Figure 16.5 Reaction Profile for an Endothermic Reaction** Notice that the product of the reaction, NO, is at a higher energy level than the reactants, N_2 and O_2.

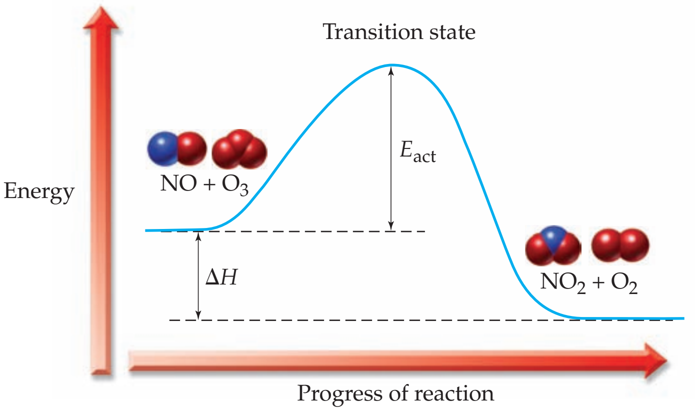
Exothermic Reaction Profiles

An **exothermic reaction** proceeds by releasing heat energy. As an example of a reaction that proceeds by liberating heat energy, let's again consider that automobiles produce nitrogen monoxide, NO. Also consider that air contains ozone, O_3. The ozone in air converts the NO to NO_2, a brown gaseous component of smog.

$$NO(g) + O_3(g) \rightleftharpoons NO_2(g) + O_2(g) + heat$$



As you can see from the chemical equation, the formation of nitrogen dioxide releases heat. As before, we can construct a reaction profile to follow the changes in energy for this reaction. Figure 16.6 illustrates the energy difference between reactants and products for an exothermic reaction. For exothermic reactions energy is released, and the products end up at a lower energy level than the reactants.

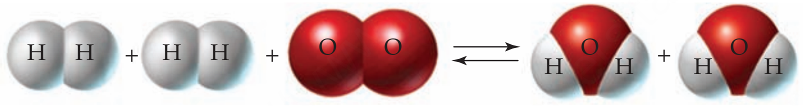


◀ **Figure 16.6 Reaction Profile for an Exothermic Reaction** Notice that the products of the reaction, NO_2 and O_2, are at a lower energy level than the reactants, NO and O_3.

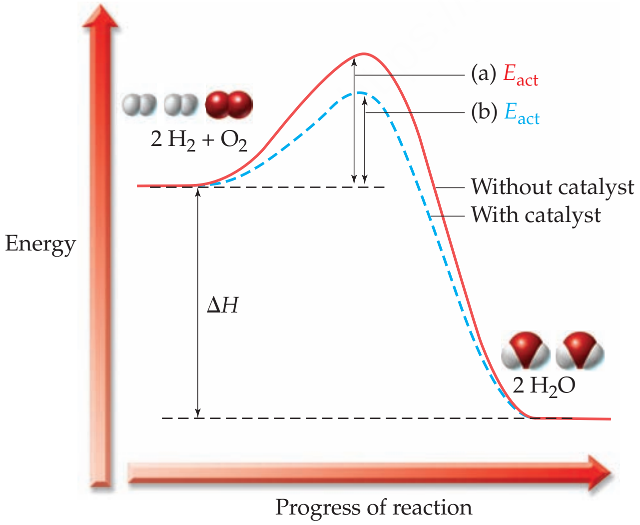
For reactants to form products, they must have sufficient energy to overcome the activation energy barrier. The higher the activation energy, the more slowly the reaction proceeds. The reaction is slower because there are fewer molecules with enough energy to achieve the transition state.

To help a reaction proceed at a faster rate, we can use a catalyst. Recall that a catalyst is a substance that speeds up the rate of a reaction without being consumed. Now, we can define a **catalyst** as a substance that allows a reaction to proceed faster by lowering the energy of activation. As an example, consider how water is produced from the exothermic reaction of hydrogen and oxygen gases. The equation for the reaction is

$$2\ H_2(g) + O_2(g) \rightleftharpoons 2\ H_2O(g) + \textit{heat}$$



Although the reaction produces heat energy, it is very slow at normal temperatures. When a mixture of hydrogen and oxygen gases is exposed to a flame or spark, the reaction is instantaneous. In fact, the instantaneous release of a large amount of energy creates an explosion. At room temperature, a powdered metal such as zinc dust acts as a catalyst and causes the reaction to occur rapidly without explosion. Figure 16.7 illustrates the reaction profile for hydrogen and oxygen gases with and without a catalyst.



◀ **Figure 16.7 Effect of a Catalyst on Activation Energy** (a) Reaction profile for the reaction of H_2 and O_2 without a catalyst (solid line). (b) Reaction profile for the reaction of H_2 and O_2 with a zinc metal catalyst (dashed line). Notice the heat of reaction, ΔH, is not affected by the catalyst.

Note that the heat of reaction is the same with and without a catalyst and that ΔH does not depend on the catalyst. Thus, the rate of a reaction is not related to the heat of reaction. Table 16.1 lists the rate of reaction and the heat of reaction for the formation of water from hydrogen and oxygen gases in the presence of various catalysts.

TABLE 16.1 Reaction of Hydrogen and Oxygen Gases

Catalyst	Rate of Reaction	Heat of Reaction*
none	very slow	57.8 kcal (242 kJ)
spark	explosive	57.8 kcal (242 kJ)
zinc dust	rapid	57.8 kcal (242 kJ)

*Energy released per mole of water produced at 25 °C.

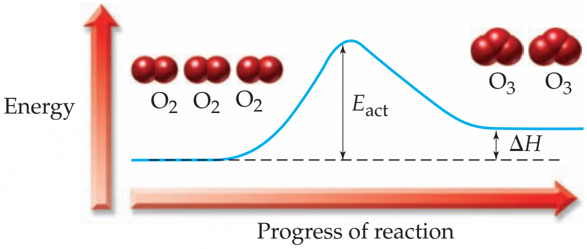
EXAMPLE EXERCISE 16.2 Energy Profiles of Chemical Reactions

Ultraviolet light from the Sun converts oxygen molecules in the upper atmosphere to ozone molecules. Draw the reaction profile for the reaction:

$$3\,O_2(g) + \textit{heat} \overset{UV}{\rightleftharpoons} 2\,O_3(g)$$

Solution

Since the reaction is endothermic, the reaction profile is

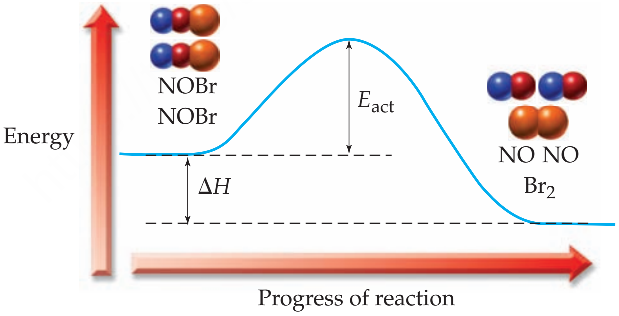


Practice Exercise

Nitrosyl bromide, NOBr, decomposes to give nitric oxide and bromine gas. Draw the reaction profile for the reaction:

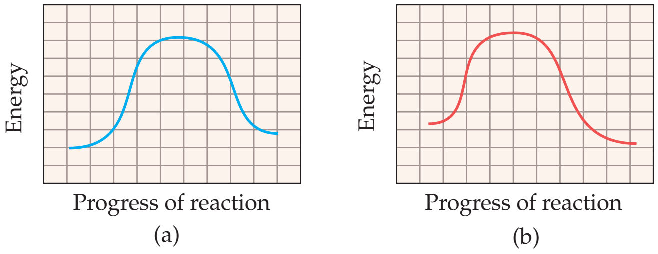
$$2\,NOBr(g) \rightleftharpoons 2\,NO(g) + Br_2(g) + \textit{heat}$$

Answer: Since the reaction is exothermic, the reaction profile is



Concept Exercise

Which of the following reaction profiles illustrates an exothermic reaction?



(a)

(b)

Answer: See Appendix G, 16.2.

16.3 The Chemical Equilibrium Concept

LEARNING OBJECTIVES

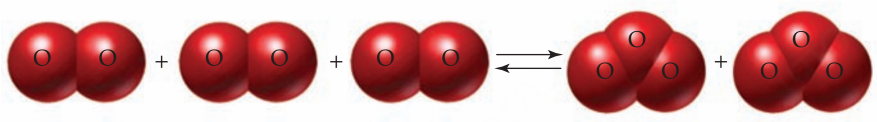
- Describe the equilibrium concept for a reversible reaction.
- Express the law of chemical equilibrium as an equation.

We have said that most chemical reactions are reversible processes. That is, a chemical change is a **reversible reaction** that can proceed simultaneously both in the forward direction toward products and in the reverse direction toward reactants. When the rate of the forward reaction is equal to the rate of the reverse reaction, a reaction is said to be in a state of **chemical equilibrium**.

It is important to understand that when a reaction is at equilibrium, it does not mean that the reaction has stopped. It also does not mean that the amounts of reactants and products are equal. Rather, a chemical reaction at equilibrium implies a *dynamic, reversible process*. That is, an ongoing forward reaction ($\longrightarrow$) and an ongoing reverse reaction ($\longleftarrow$) are taking place at the same rate.

As an example, consider the conversion of oxygen molecules, O_2, to ozone molecules, O_3, in the upper atmosphere. Ultraviolet light initiates the reaction. The equation for the reversible reaction is

$$\underset{\textit{oxygen}}{3\,O_2(g)} \rightleftharpoons \underset{\textit{ozone}}{2O_3(g)}$$



We can experimentally simulate this reversible reaction in a glass sphere. We begin by pumping oxygen gas into the glass sphere. Next, we radiate ultraviolet light on the oxygen gas in the sphere. Initially, the rate of the forward reaction (rate_f) is quite rapid and the rate of the reverse reaction (rate_r) is slow. As the reaction proceeds, the forward reaction slows down and the reverse reaction speeds up. When the gaseous reaction reaches chemical equilibrium, we have

$$\text{rate}_f\ (O_2\ \text{reaction}) = \text{rate}_r\ (O_3\ \text{reaction})$$

We can define the **rate of reaction** as the rate at which the concentrations of reactants decrease per unit time. Alternatively, we can define the rate of reaction as the rate at which the concentrations of products increase per unit time. Figure 16.8 shows the changes in the rates of the forward and reverse reactions as the reversible oxygen-ozone reaction progresses.

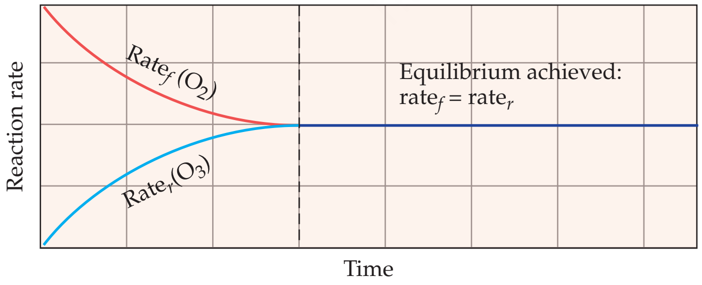
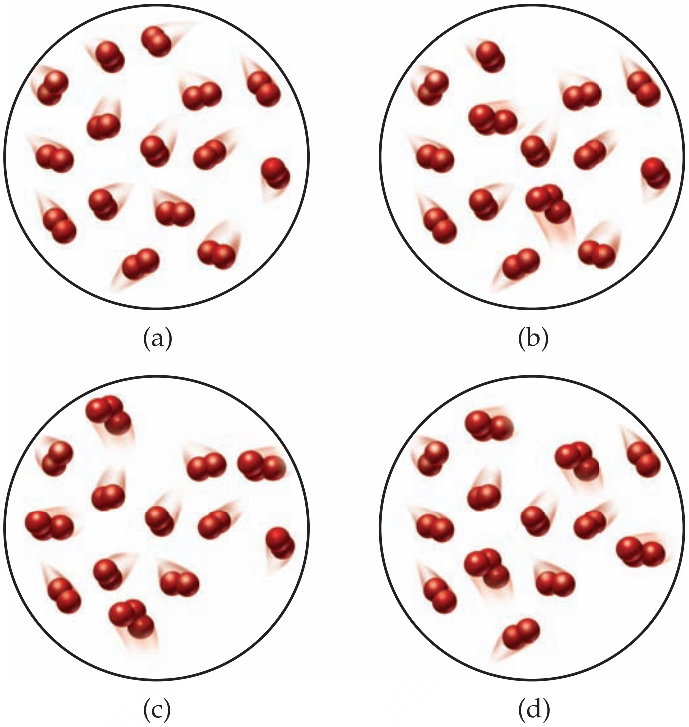


Figure 16.8 Chemical Equilibrium and Rates of Reaction The rate of reaction for O_2 decreases with time. As the reaction proceeds, the rate of reaction for O_3 increases. At equilibrium, the rates of the forward and reverse reactions are the same.

Figure 16.9 provides a model to help us understand the equilibrium process. Oxygen molecules, O_2, react in a closed container to give ozone molecules, O_3. Simultaneously, some of the O_3 molecules decompose to give O_2 molecules. Initially, the O_2 concentration is high and the forward reaction is rapid. The O_3 concentration is low and the reverse reaction is slow. As equilibrium is approached, the forward reaction slows down and the reverse reaction speeds up. When the rates of the forward and reverse reactions are equal, the reaction is at equilibrium.

▶ **Figure 16.9 A Dynamic Equilibrium Model** (a) The container has only O_2 gas molecules. (b) As the reaction proceeds, O_3 is formed. (c) At equilibrium, the forward and reverse reactions occur at the same rate. (d) Later, the amounts of O_2 and O_3 are constant even though molecules of O_2 and O_3 continue to react.



Law of Chemical Equilibrium

In 1864, the Norwegian chemists Cato Guldberg and Peter Waage observed that a change in the amount of substance participating in a reversible reaction produced a shift in the equilibrium. When they added more of a reactant substance, the reaction shifted to form more product. Conversely, when they added a product substance, the reaction shifted to form more reactant. We now know that a shift in equilibrium can be produced for any reversible reaction.

$$a\,\mathrm{A} + b\,\mathrm{B} \rightleftharpoons c\,\mathrm{C} + d\,\mathrm{D}$$

where A and B represent reactants, C and D represent products, and a, b, c, and d represent the coefficients for a general reversible reaction.

Moreover, at equilibrium we would observe that the ratio of product concentrations to reactant concentrations remains constant at a given temperature. Guldberg and Waage found that the molar concentrations of the products (raised to the powers c and d), divided by the molar concentrations of reactants (raised to the powers a and b), always gave the same value. Regardless of the initial amounts of reactants, at equilibrium the molar concentration ratio is a constant at a given temperature. This relationship, which applies to every reversible reaction, is known as the **law of chemical equilibrium**. Mathematically, we can express the law of chemical equilibrium as follows:

$$K_{\mathrm{eq}} = \frac{[\mathrm{C}]^c\,[\mathrm{D}]^d}{[\mathrm{A}]^a\,[\mathrm{B}]^b}$$

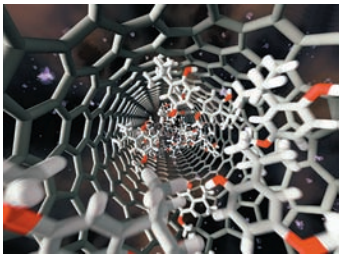
where [A], [B], [C], and [D] represent the molar concentrations of reactants and products, and a, b, c, and d correspond to the coefficients in the balanced chemical equation. The ratio is equal to the **general equilibrium constant** (symbol K_{eq}), which expresses the molar concentration of each substance participating in the reaction at a given temperature.

A CLOSER LOOK Nanotechnology

Q: What is a nanotechnology?

The transistor is considered one of the greatest inventions of the twentieth century. The transistor revolutionized the electronics industry by replacing large radio and television vacuum tubes with microscale chips. The term *microelectronics* refers to electronic devices that operate at one micrometer (10^{-6} m). The term *nanotechnology* refers to devices that operate at one nanometer (10^{-9} m). The size of this device is smaller than the wavelength of visible light (400–700 nm). Applications of nanotechnology include nanoparticles of titanium dioxide in sunscreens that absorb ultraviolet radiation; and nanoparticles of silver metal in sports clothing that act as an antibacterial agent.

Scientists have discovered that hollow nanoscale cylinders having a diameter of 100 nm or less are capable of allowing fluids to flow much faster than predicted. Despite its small inner diameter, a so-called "nanotube" allows for relatively fast flow of water and gas molecules. One reason that small molecules speed through the tiny nanotube is that the inner wall of the tube has a slippery surface.



◀ **Nanotube** A carbon nanotube has a slippery inner surface that allows the flow of water molecules and gases.

Scientists are hopeful that nanotechnology will lead to practical applications, such as the removal of minerals from salt water to produce drinking water. Scientists also foresee medical applications, such as nanoscale devices for clearing restricted arteries; and devices for kidney dialysis that purify blood by separating toxins from proteins and other biomolecules.

A: Nanotechnology refers to the science of devices in the range of 1–100 nm.

16.4 General Equilibrium Constant, K_{eq}

LEARNING OBJECTIVES

- Write the equilibrium constant expression for a reversible reaction.
- Calculate an equilibrium constant, K_{eq}, from experimental data.

Let's construct the equilibrium constant expression for the following reversible reaction.

$$2\text{ A} \rightleftharpoons \text{B}$$

To write the equilibrium expression, K_{eq}, we place the concentration of the product in the numerator and the concentration of the reactant in the denominator.

$$K_{eq} = \frac{[\text{B}]}{[\text{A}]}$$

The coefficients in the balanced equation are written as exponents in the K_{eq} expression. The coefficient of A is 2, and the coefficient of B is understood to be 1. These coefficients are written as the power of the corresponding concentration.

$$K_{eq} = \frac{[\text{B}]}{[\text{A}]^2}$$

The following example exercise provides additional practice in writing K_{eq} expressions for general cases of reversible reactions.

Helpful Hint
Equilibrium Constants
If K_{eq} is greater than 1, a reaction occurs that favors the product. Conversely, if K_{eq} is less than 1, a reaction favors the reactant.

EXAMPLE EXERCISE 16.3 General Equilibrium Constant, K_{eq}

Write the equilibrium constant expression for the following reversible reaction.

$$\text{A} + 2\text{ B} \rightleftharpoons 3\text{ C} + \text{D}$$

Solution

Let's proceed in two steps. First, let's substitute into the K_{eq} expression. The ratio of product concentration to reactant concentration is

$$K_{eq} = \frac{[\text{C}]^c\,[\text{D}]^d}{[\text{A}]^a\,[\text{B}]^b}$$

Second, let's account for the coefficients of the balanced equation. The coefficients are 1, 2, 3, and 1, respectively.

$$K_{eq} = \frac{[C]^3\,[D]}{[A]\,[B]^2}$$

Practice Exercise

Write the equilibrium constant expression for the following reversible reaction.

$$2\,A + 3\,B \rightleftharpoons C + 4\,D$$

Answer:

$$K_{eq} = \frac{[C]\,[D]^4}{[A]^2\,[B]^3}$$

Concept Exercise

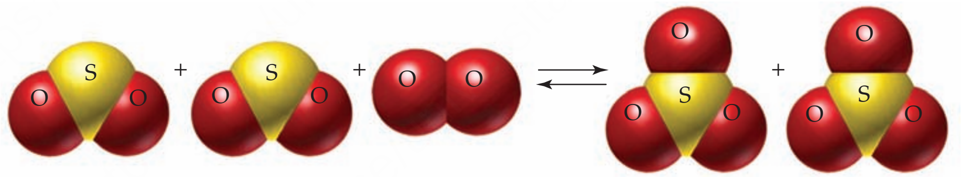
Describe the dynamic reactant and product concentrations for a reversible reaction *after* the reaction reaches equilibrium.

Answer: See Appendix G, 16.3.

Chemical Equilibria in the Same Physical State

In most reversible reactions, all the reactants and products are in the same physical state. In the ozone reaction, both the reactants and products are gases. A reversible reaction at equilibrium in which all the substances are in the same state is referred to as a **homogeneous equilibrium**. For example, low-grade coal has a high sulfur content, and burning it produces SO_2. In the atmosphere, oxygen gas converts SO_2 gas to SO_3 gas according to the following reversible reaction.

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$



If we substitute the concentrations of reactants and products into the general equilibrium expression, we have

$$K_{eq} = \frac{[SO_3]}{[SO_2]\,[O_2]}$$

The respective coefficients in the balanced equation are 2, 1, and 2. Thus, the equilibrium expression for the reaction is

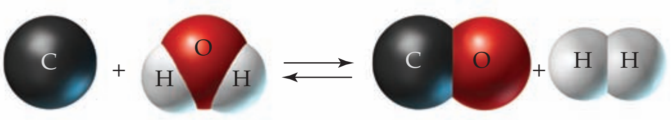
$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2\,[O_2]}$$

Chemical Equilibria in a Different Physical State

In the previous examples of gaseous equilibria, liquids and solids were not involved and did not appear in the K_{eq} expression. On occasion, a reactant or product may be in the solid or the liquid state. A reaction at equilibrium in which one of the substances is in a different physical state is referred to as a **heterogeneous equilibrium**.

As an example, consider hydrogen gas which can be manufactured by passing steam (a gas) over hot charcoal (a solid). The equation for the reaction is as follows.

$$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$$



We can write the equilibrium constant expression as

$$K_{eq} = \frac{[CO]\,[H_2]}{[C]\,[H_2O]}$$

Charcoal is a solid but all the other substances in the reaction are gases. When the reaction is studied experimentally, it is found that the amount of charcoal has no effect on the equilibrium constant. Therefore, the concentration of charcoal, [C], is set equal to unity (1) and is omitted from the equilibrium expression, giving us

$$K_{eq} = \frac{[CO]\,[H_2]}{[H_2O]}$$

In addition, it is found that liquids and solids have no effect on any gaseous equilibrium. Therefore, liquids and solids do not appear in the equilibrium expression. The following example exercise provides additional practice in writing K_{eq} expressions for homogeneous and heterogeneous equilibria.

EXAMPLE EXERCISE 16.4 Writing K_{eq} Expressions

Write the equilibrium constant expression for each of the following reversible reactions:

(a) $2\,NO(g) + 2\,H_2(g) \rightleftharpoons N_2(g) + 2\,H_2O(g)$
(b) $NH_4NO_3(s) \rightleftharpoons N_2O(g) + 2\,H_2O(g)$

Solution

We can substitute into the K_{eq} expression as follows:

(a) Since the equation is balanced, we raise the concentration of each species to the power corresponding to the coefficient in the balanced equation.

$$K_{eq} = \frac{[N_2]\,[H_2O]^2}{[NO]^2\,[H_2]^2}$$

(b) In this reaction NH_4NO_3 is a solid. Since NH_4NO_3 is not a gas, it does not appear in the equilibrium expression; thus, the equilibrium expression is

$$K_{eq} = [N_2O]\,[H_2O]^2$$

Practice Exercise

Write the equilibrium constant expression for each of the following reversible reactions:

(a) $4\,NH_3(g) + 5\,O_2(g) \rightleftharpoons 4\,NO(g) + 6\,H_2O(g)$
(b) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

Answers:

(a) $K_{eq} = \dfrac{[NO]^4\,[H_2O]^6}{[NH_3]^4\,[O_2]^5}$
(b) $K_{eq} = [CO_2]$

Concept Exercise

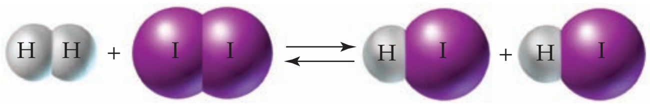
Describe the dynamic reactant and product concentrations for a reversible reaction *before* the reaction reaches equilibrium.

Answers: See Appendix G, 16.4.

Determining K_{eq} by Experiment

One of the most thoroughly investigated equilibrium reactions is involved in the formation of hydrogen iodide from hydrogen gas and iodine vapor.

$$\underset{\textit{colorless gas}}{H_2(g)} + \underset{\textit{purple vapor}}{I_2(g)} \rightleftharpoons \underset{\textit{colorless gas}}{2\ HI(g)}$$



Experimentally, we can measure the molar concentration of each gas at equilibrium. If we start with 1.000 mol of H_2 and I_2 in a 1-liter container, we find at equilibrium that $[H_2] = 0.212\ M$, $[I_2] = 0.212\ M$, and $[HI] = 1.576\ M$. To calculate an experimental value for the equilibrium constant, K_{eq}, let's write the equilibrium expression.

$$K_{eq} = \frac{[HI]^2}{[H_2]\,[I_2]}$$

Substituting, we have

$$K_{eq} = \frac{[1.576]^2}{[0.212]\,[0.212]} = 55.3$$

Interestingly, if we start with 1.000 mol of HI and approach equilibrium from the opposite direction, we obtain the same value for K_{eq}. In fact, the K_{eq} value for the reaction is 55.3 at a given temperature regardless of the initial concentrations of each gas. Table 16.2 presents data for four experiments involving the HI equilibrium.

TABLE 16.2 Experimental Determination of K_{eq}

Experiment	Initial Concentration			Equilibrium Concentration			Calculated K_{eq}* at 425°C
	$[H_2]$	$[I_2]$	$[HI]$	$[H]$	$[I_2]$	$[HI]$	$\frac{[HI]^2}{[H_2]\,[I_2]}$
1	1.000	1.000	0	0.212	0.212	1.576	55.3
2	0.500	0.500	0	0.106	0.106	0.788	55.3
3	0	0	1.000	0.106	0.106	0.788	55.3
4	0	0	2.000	0.212	0.212	1.576	55.3

*The units for an equilibrium constant are usually omitted. In this example, the units in the numerator and denominator of the K_{eq} expression cancel.

Conceptually, we can visualize a closed vessel containing the three gases H_2, I_2, and HI in equilibrium. When we change the amount of any gas, the reaction shifts forward or backward in such a way that the ratio among the concentrations of reactants and products remains constant. The following example exercise further illustrates the calculation of an equilibrium constant from experimental data.

EXAMPLE EXERCISE 16.5 Experimental Equilibrium Constant, K_{eq}

In the upper atmosphere, nitric oxide and nitrogen dioxide participate in the following equilibrium:

$$2\ NO(g) + O_2(g) \rightleftharpoons 2\ NO_2(g)$$

Given the equilibrium concentrations for each gas at 25°C, what is K_{eq}?

$$[NO] = 1.5 \times 10^{-11}\ M$$
$$[O_2] = 8.9 \times 10^{-3}\ M$$
$$[NO_2] = 2.2 \times 10^{-6}\ M$$

Solution

We first write the equilibrium expression for the reaction.

$$K_{eq} = \frac{[NO_2]^2}{[NO]^2\,[O_2]}$$

Second, we substitute the concentration values into the expression.

$$K_{eq} = \frac{[2.2 \times 10^{-6}]^2}{[1.5 \times 10^{-11}]^2\,[8.9 \times 10^{-3}]}$$
$$= 2.4 \times 10^{12}$$

The large K_{eq} for this reaction indicates that the equilibrium overwhelmingly favors the formation of the product, NO_2. However, K_{eq} does not indicate the rate of reaction. In the atmosphere, the conversion of NO to NO_2 takes place slowly.

Practice Exercise

Given the equilibrium concentrations for the gas mixture at 100°C, calculate the value of K_{eq} for the following reaction:

$$\begin{array}{ccc} N_2O_4(g) & \rightleftharpoons & 2\,NO_2(g) \\ 0.00140\,M & & 0.0172\,M \end{array}$$

Answer: $K_{eq} = 0.211$

Concept Exercise

Write the equilibrium constant expression for the following reversible reaction:

$$2\,X_2(g) + 3\,Y_2(g) \rightleftharpoons 2\,X_2Y_3(g) + \textit{heat}$$

Answer: See Appendix G, 16.5.

Note In Section 10.3, we learned that gas pressure is proportional to the concentration of molecules. Therefore, we can write an equilibrium expression for a reversible gaseous reaction on the basis of partial gas pressures. The equilibrium expression for a reversible gaseous reaction can be written as

$$K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

where K_p is the equilibrium constant, and P_A, P_B, P_C, and P_D are the partial pressures of the gases participating in the equilibrium reaction.

16.5 Equilibria Shifts for Gases

LEARNING OBJECTIVE

- Apply Le Chatelier's principle to reversible reactions in the gaseous state.

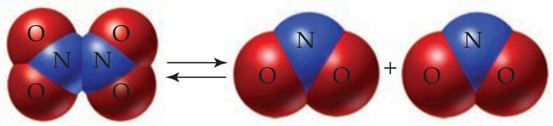
Henri Louis Le Chatelier (1850–1936) was trained in France as a mining engineer. When he graduated from college, he became a professor of chemistry at a mining school. His first interest was in the study of flames in the hope of preventing mine explosions. Eventually, he studied the effect of heat on reversible reactions and in 1888 stated the principle for which he is famous: "Every change of one of the factors of an equilibrium brings about a rearrangement of the system in such a direction as to minimize the original change."

The factors affecting a chemical equilibrium system are concentration, temperature, and pressure. A change in any of these factors shifts the equilibrium. For chemical reactions, **Le Chatelier's principle** can be stated as follows: when a reversible reaction at equilibrium is disturbed by a change in concentration, temperature, or pressure, the equilibrium shifts to relieve the stress.

Effect of Concentration

For any equilibrium system, we can experimentally shift the reaction forward or backward by changing the concentration of a reactant or product. Let's consider the following equilibrium.

$$\underset{\textit{colorless}}{N_2O_4(g)} \rightleftharpoons \underset{\textit{brown}}{2\,NO_2(g)}$$



If we increase the N_2O_4 reactant concentration, the equilibrium shifts to the right and the NO_2 product concentration increases. If we decrease the N_2O_4 concentration, the equilibrium shifts to the left and the NO_2 concentration decreases.

Conversely, if we increase the NO_2 product concentration, the equilibrium shifts to the left and the N_2O_4 reactant concentration increases. If we decrease the NO_2 concentration, the equilibrium shifts to the right and the N_2O_4 concentration decreases.

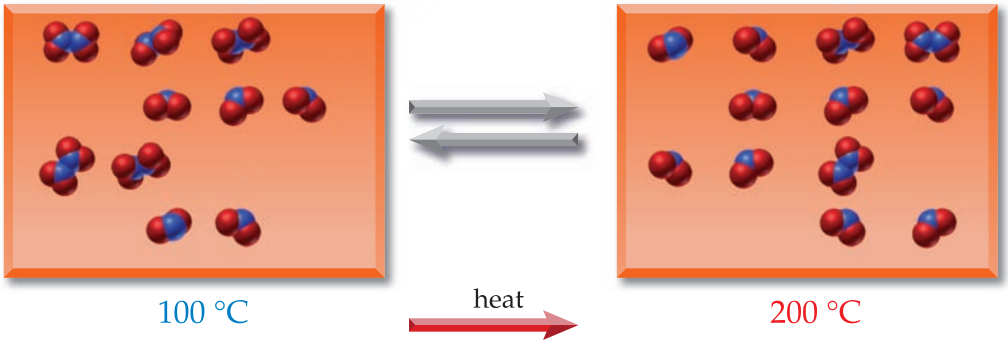
Effect of Temperature

Now, let's consider the effect of temperature on the following equilibrium:

$$\underset{\textit{colorless}}{N_2O_4(g)} + \textit{heat} \rightleftharpoons \underset{\textit{brown}}{2\,NO_2(g)}$$

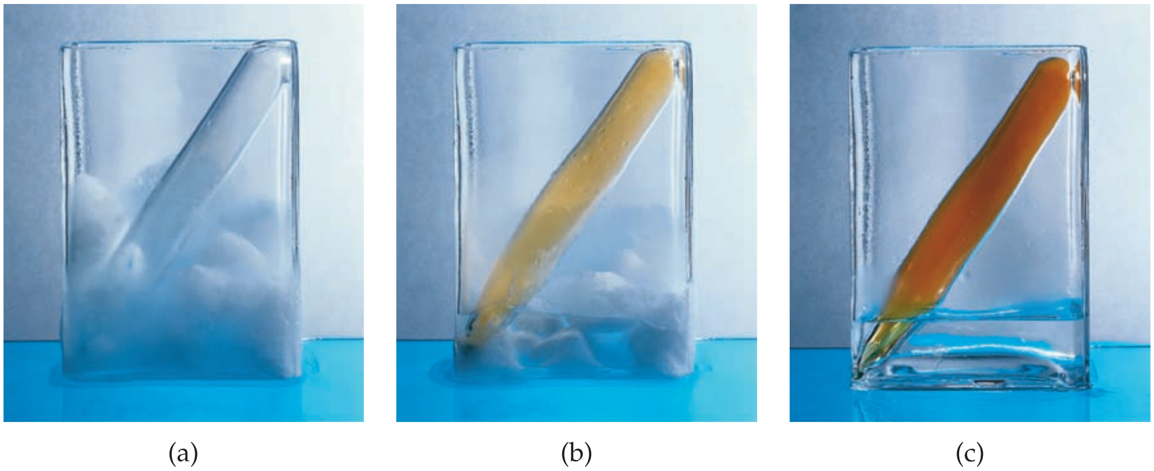
When we change the temperature for a reversible reaction, we stress the equilibrium system, and the equilibrium shifts to relieve the stress. In this example, if we heat this endothermic reaction, the equilibrium shifts to the *right;* the N_2O_4 concentration decreases and the NO_2 concentration increases. If we cool the reaction, the equilibrium shifts to the *left.* As the equilibrium shifts to the left, the NO_2 concentration decreases and the N_2O_4 concentration increases.

In an *endothermic reaction*, heat causes the equilibrium to shift to the right toward the products to relieve the stress on the reactants side of the equation. In an *exothermic reaction*, heat causes the equilibrium to shift to the left toward the reactants to relieve the stress on the products side of the equation. Figure 16.10 illustrates the shift in equilibrium that results from heating the N_2O_4 reaction.



▲ **Figure 16.10 Effect of Temperature on Equilibrium** (a) Initially, there are five molecules of N_2O_4 and five molecules of NO_2 in the equilibrium mixture. (b) If the temperature increases, the equilibrium shifts to the right to relieve the stress. As a result, the number of N_2O_4 molecules decreases and the number of NO_2 molecules increases.

Note that N_2O_4 is a colorless gas and that NO_2 is a dark-brown gas. Since each gas has a distinctive color, we can observe an equilibrium shift by noting the color of the gaseous mixture. Figure 16.11 illustrates the shift in equilibrium resulting from heating the colorless N_2O_4 gas to give dark-brown NO_2 gas.



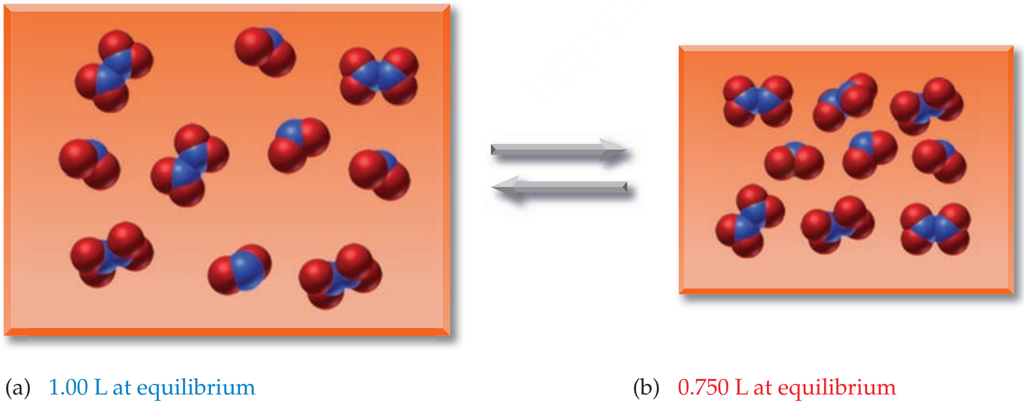
▲ **Figure 16.11** **Heating the N_2O_4–NO_2 Equilibrium** (a) At low temperatures, the N_2O_4–NO_2 equilibrium favors N_2O_4, which is a colorless gas. (b) If we heat the reaction, the equilibrium shifts to produce more NO_2, which is a brown gas. (c) Further heating produces more NO_2, and the gas mixture becomes a darker brown.

Effect of Pressure

For many gaseous equilibrium systems we can experimentally shift the reaction forward or backward by changing the volume, and therefore the pressures of the gases in the system. If we decrease the volume to increase the pressure, the equilibrium will shift to reduce the pressure change by producing fewer molecules. Once again, let's consider the following equilibrium:

$$\underset{\textit{colorless}}{N_2O_4(g)} \rightleftharpoons \underset{\textit{brown}}{2\,NO_2(g)}$$

Notice that 1 molecule of N_2O_4 produces 2 molecules of NO_2. If we decrease the volume of the container, the gas pressure increases. To relieve the stress, the equilibrium shifts toward fewer molecules. That is, the equilibrium shifts to the left because there is 1 molecule of N_2O_4 for every 2 molecules of NO_2. Figure 16.12 illustrates the equilibrium shift for a decrease in volume.



◀ **Figure 16.12** **Effect of Pressure on Equilibrium** (a) Initially, there are 10 molecules in the N_2O_4–NO_2 equilibrium. (b) When the volume decreases and the pressure increases, the equilibrium shifts to relieve the stress. As a result, NO_2 is converted to N_2O_4, and the total number of molecules decreases.

If we increase the volume of the container, the gas pressure decreases. To relieve the stress, the equilibrium shifts toward more molecules. That is, the equilibrium shifts to the right because there are 2 molecules of NO_2 for 1 molecule of N_2O_4.

Effect of an Inert Gas

Let's consider what happens to the N_2O_4–NO_2 equilibrium if we add an inert gas such as helium. Since helium does not participate in the N_2O_4–NO_2 equilibrium, the partial gas pressures of N_2O_4 and NO_2 do not change. Since the pressures of N_2O_4 and NO_2 remain constant, there is no shift in equilibrium. The following example exercise further illustrates Le Chatelier's principle as applied to gaseous systems.

EXAMPLE EXERCISE 16.6 Gaseous State Equilibria Shifts

Charcoal reacts with steam according to the equation.

$$C(s) + H_2O(g) + \textit{heat} \rightleftharpoons CO(g) + H_2(g)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[H_2O]$
(b) decrease $[H_2O]$
(c) increase [CO]
(d) decrease $[H_2]$
(e) increase temperature
(f) decrease volume
(g) add a catalyst
(h) add charcoal

Solution

Let's apply Le Chatelier's principle to each of the stresses.

(a) If $[H_2O]$ increases, the equilibrium shifts to the *right.*
(b) If $[H_2O]$ decreases, the equilibrium shifts to the *left.*
(c) If [CO] increases, the equilibrium shifts to the *left.*
(d) If $[H_2]$ decreases, the equilibrium shifts to the *right.*
(e) If the temperature increases, the equilibrium shifts to the *right.*
(f) If the volume decreases, the equilibrium shifts toward fewer gas molecules. The equilibrium shifts to the *left.*
(g) Adding a catalyst increases the rate of reaction but has no affect on the amount of reactant or product. Thus, there is *no shift* in equilibrium.
(h) Adding carbon has no effect on the gaseous equilibrium because carbon is a solid. Thus, there is *no shift* in equilibrium.

Practice Exercise

Methane, CH_4, reacts with limited oxygen according to the equation.

$$2\,CH_4(g) + O_2(g) \rightleftharpoons 2\,CO(g) + 4\,H_2(g) + \textit{heat}$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[CH_4]$
(b) increase [CO]
(c) decrease $[O_2]$
(d) decrease $[H_2]$
(e) decrease temperature
(f) decrease volume
(g) add a catalyst
(h) add gaseous neon

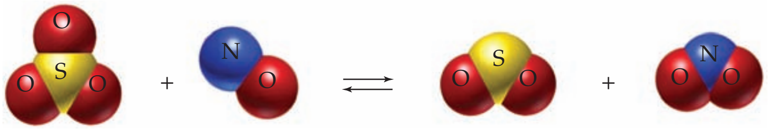
Answers:

(a) shifts right
(b) shifts left
(c) shifts left
(d) shifts right
(e) shifts right
(f) shifts left
(g) no shift
(h) no shift

Concept Exercise

What is the effect of increasing (a) temperature and (b) pressure for the following equilibrium?

$$SO_3(g) + NO(g) + \textit{heat} \rightleftharpoons SO_2(g) + NO_2(g)$$



Answers: See Appendix G, 16.6.

16.6 Ionization Equilibrium Constant, K_i

LEARNING OBJECTIVES

- Write the equilibrium constant expression for a weak acid or a weak base.
- Calculate an ionization constant, K_i, from experimental data.

In Section 16.3, we developed the concept of a general equilibrium constant expression, K_{eq}. Here we will develop a special equilibrium constant expression for a weak acid or weak base. This type of equilibrium is characterized by a weak acid or base in equilibrium with its ions.

Ionization of a Weak Acid

Let's begin with an aqueous solution of a weak acid. For example, acetic acid in aqueous solution yields the following:

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

In aqueous solutions the concentration of water is constant. We can therefore simplify the reversible reaction as follows:

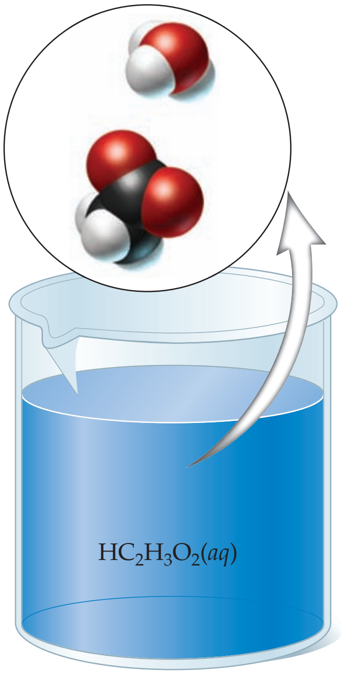
$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

In general, the equilibrium for weak acids lies overwhelmingly to the left. Typically, less than 1% of the parent molecules ionize. In the preceding reaction, very little ionization of acetic acid occurs. Although about 1% of the molecules form ions, 99% do not. Thus, the equilibrium strongly favors the reverse reaction.

When we substitute the acetic acid equilibrium into the general equilibrium constant expression, we obtain

$$K_i = \frac{[H^+]\,[C_2H_3O_2^-]}{HC_2H_3O_2}$$

where $\boldsymbol{K_i}$ is the **ionization equilibrium constant** for the weak acid. Since a weak acid ionizes only slightly, K_i is a very small number. The following example exercise provides practice in writing equilibrium constant expressions.



▲ **Acetic Acid, $HC_2H_3O_2$** Acetic acid dissolved in water is a weak acid, which is less than 1% ionized.

EXAMPLE EXERCISE 16.7 Writing K_i Expressions

Write the equilibrium constant expression for the following weak acid:

$$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$$

Solution

We substitute the product concentrations into the numerator and the reactant concentration into the denominator; thus,

$$K_i = \frac{[H^+]\,[NO_2^-]}{[HNO_2]}$$

Practice Exercise

Write the equilibrium constant expression for the following weak acid:

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

Answer: Carbonic acid, H_2CO_3, ionizes to give H^+ and HCO_3^-. Substituting into the ionization equilibrium expression, we have

$$K_i = \frac{[H^+]\,[HCO_3^-]}{[H_2CO_3]}$$

Concept Exercise

Write the equilibrium constant expression for the following weak acid:

$$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq)$$

Answer: See Appendix G, 16.7.

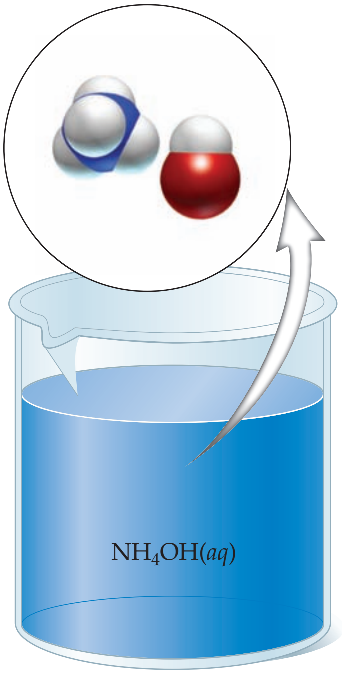
Ionization of a Weak Base

An aqueous solution of a weak base behaves similarly to an aqueous solution of a weak acid. That is, a weak base ionizes slightly and is in equilibrium with its ions. For example, aqueous ammonium hydroxide gives the following equilibrium.

$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Typically, the equilibrium for weak bases lies overwhelmingly to the left. In ammonium hydroxide, less than 1% ionization occurs. Thus, the equilibrium strongly favors the reverse reaction. Substituting into the general equilibrium constant expression, we have

$$K_i = \frac{[NH_4^+]\,[OH^-]}{[NH_4OH]}$$



▲ **Aqueous Ammonia, NH_4OH** Ammonia gas dissolved in water is a weak base, which is less than 1% ionized.

Determining K_i by Experiment

In the laboratory, we can determine the experimental equilibrium constant for a weak acid or weak base from pH measurements. For example, we can use a pH meter to measure in $[H^+]$ a solution. K_i is then calculated from the hydrogen ion concentration.

Let's calculate the ionization constant, K_i, for acetic acid. In an experiment, a 0.100 *M* solution of acetic acid is found to have a hydrogen ion concentration of 0.00134 *M*. We begin by analyzing the equilibrium concentrations. For each acetic acid molecule that ionizes, we obtain one H^+ and one $C_2H_3O_2^-$. Therefore, $[H^+]$ must equal $[C_2H_3O_2^-]$, that is, 0.00134 *M*. The equilibrium reaction is

$$\begin{array}{ccccc} HC_2H_3O_2(aq) & \rightleftharpoons & H^+(aq) & + & C_2H_3O_2^-(aq) \\ 0.100\,M & & 0.00134\,M & & 0.00134\,M \end{array}$$

Next, we write the equilibrium constant expression by placing the ion concentrations in the numerator and the molar concentration of the weak acid in the denominator. That gives

$$K_i = \frac{[H^+]\,[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Substituting into the equilibrium expression for each concentration, we have

$$K_i = \frac{[0.00134][0.00134]}{[0.100]}$$
$$= 1.80 \times 10^{-5}$$

The ionization constant for acetic acid is a small value, 1.80×10^{-5}, because acetic acid is a weak acid and ionizes only slightly in an aqueous solution.

16.7 Equilibria Shifts for Weak Acids and Bases

LEARNING OBJECTIVE

- Apply Le Chatelier's principle to solutions of weak acids and weak bases.

In Section 16.5, we discussed Le Chatelier's principle applied to equilibria in the gaseous state. This principle also applies to equilibria in aqueous solutions. That is, a change in the concentration of any species in an aqueous solution disturbs the equilibrium and causes a shift between reactants and products.

Let's reconsider acetic acid, which ionizes very little in an aqueous solution. We can indicate the weak acid equilibrium as follows:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

If we increase the molar concentration of $HC_2H_3O_2$, we stress the left side of the equilibrium. The reversible reaction responds to the stress by shifting to the right. The final result is that more $HC_2H_3O_2$ molecules ionize and the molar concentrations of H^+ and $C_2H_3O_2^-$ and increase.

If we increase the molar concentration of H^+ by adding a strong acid, we stress the right side of the equilibrium. The equilibrium responds to the stress by shifting to the left. The final result is that the molar concentration of $HC_2H_3O_2$ increases, whereas the molar concentration of $C_2H_3O_2^-$ decreases. Since additional acid was added, the molar concentration of H^+ also increases.

We can also indicate a change in $[H^+]$ that corresponds to a change in pH. Since $[H^+]$ and pH are inversely related, $[H^+]$ decreases as pH increases. Therefore, if the pH of an acetic acid solution increases, $[H^+]$ decreases. A decrease in the $[H^+]$ causes the equilibrium to shift to the right. Thus, $[HC_2H_3O_2]$ decreases, whereas $[C_2H_3O_2^-]$ increases.

If we increase the molar concentration of $C_2H_3O_2^-$ by adding solid $NaC_2H_3O_2$, we stress the right side of the equilibrium. As the $NaC_2H_3O_2$ dissolves, the equilibrium responds to the stress by shifting to the left. The final result is that the $[HC_2H_3O_2]$ increases and $[H^+]$ decreases. Since $NaC_2H_3O_2$ was added, $[C_2H_3O_2^-]$ also increases.

If we add solid NaOH, we neutralize H^+ and stress the right side of the equilibrium. The NaOH lowers the H^+ concentration, which shifts the equilibrium to the right. After the equilibrium shift, the $HC_2H_3O_2$ concentration has decreased, whereas the $C_2H_3O_2^-$ concentration has increased.

Finally, consider what happens if we add solid $NaNO_3$ to aqueous acetic acid. Since sodium nitrate is a soluble salt, it dissolves giving Na^+ and NO_3^- in solution. Neither of these ions participates in the acetic acid equilibrium. Thus, adding $NaNO_3$ has no effect, and there is no shift in the equilibrium. The following example exercise further illustrates Le Chatelier's principle.

EXAMPLE EXERCISE 16.8 Weak Acid Equilibria Shifts

Hydrofluoric acid, HF, ionizes according to the equation:

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase [HF]
(b) increase $[H^+]$
(c) decrease $[F^-]$
(d) add gaseous HCl
(e) add solid NaF
(f) add solid NaCl
(g) add solid NaOH
(h) decrease pH

Solution

Let's apply Le Chatelier's principle to each of the stresses:

(a) If [HF] increases, the equilibrium shifts to the *right*.
(b) If $[H^+]$ increases, the equilibrium shifts to the *left*.
(c) If $[F^-]$ decreases, the equilibrium shifts to the *right*.
(d) If HCl is added, $[H^+]$ increases and the equilibrium shifts to the *left*.
(e) If NaF is added, $[F^-]$ increases and the equilibrium shifts to the *left*.
(f) If NaCl is added, $[Cl^-]$ increases, which has no effect on the equilibrium. Thus, there is *no shift*.
(g) If NaOH is added, the OH^- neutralizes H^+ and lowers its concentration. Thus, the equilibrium shifts to the *right*.
(h) If the pH decreases, $[H^+]$ increases, and the equilibrium shifts to the *left*.

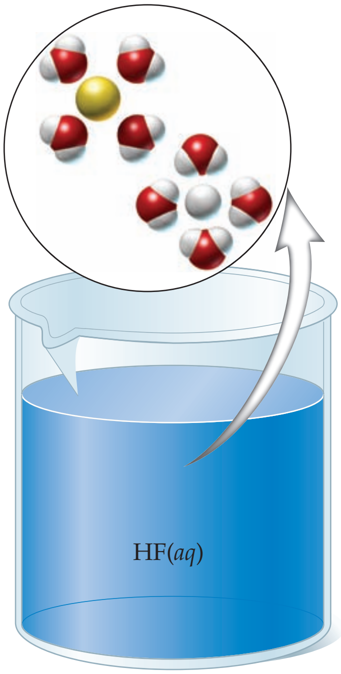
Practice Exercise

Hydrocyanic acid, HCN, ionizes according to the equation

$$HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase [HCN]
(b) decrease $[H^+]$
(c) increase $[CN^-]$
(d) add aqueous HNO_3
(e) add solid KCN
(f) add solid KCl
(g) add solid KOH
(h) increase pH



▲ **Hydrofluoric Acid, HF**
Hydrofluoric acid is a weak acid, which is less than 1% ionized in an aqueous solution.

Answers:

(a) shifts right	(b) shifts right
(c) shifts left	(d) shifts left
(e) shifts left	(f) no shift
(g) shifts right	(h) shifts right

Concept Exercise

What is the effect on the following equilibrium of adding (a) NH_4Cl and (b) NaOH?

$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Answer: See Appendix G, 16.8.

16.8 Solubility Product Equilibrium Constant, K_{sp}

LEARNING OBJECTIVES

- Write the equilibrium constant expression for a slightly soluble ionic compound.
- Calculate a solubility product constant, K_{sp}, from experimental data.

Guldberg and Waage laid the foundation for the law of chemical equilibrium based on their study of insoluble barium carbonate in an aqueous solution. After adding soluble potassium sulfate, they found that the insoluble precipitate contained both barium carbonate and barium sulfate. The Norwegian chemists proposed the following reversible reaction in an aqueous solution:

$$BaCO_3(s) + K_2SO_4(aq) \rightleftharpoons BaSO_4(s) + K_2CO_3(aq)$$

Notice that $BaCO_3$ and $BaSO_4$ are both insoluble precipitates; however, each substance is very slightly soluble. In an aqueous solution, $BaCO_3$ dissociates to give

$$BaCO_3(s) \rightleftharpoons Ba^{2+}(aq) + CO_3^{2-}(aq)$$

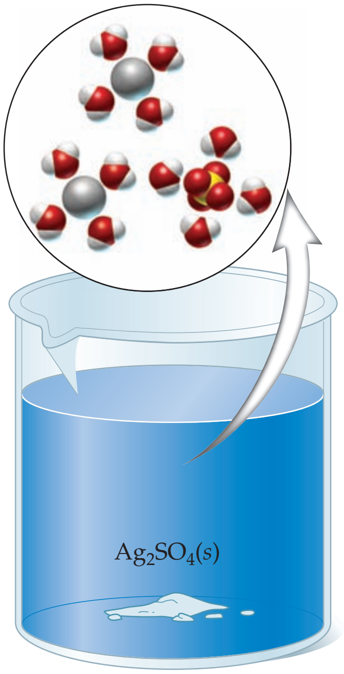
When $BaCO_3$ dissociates into ions, Ba^{2+} can recombine with SO_4^{2-} to give an insoluble $BaSO_4$ precipitate. Therefore, they concluded that seemingly insoluble precipitates in an aqueous solution are actually *very slightly soluble*. Moreover, these insoluble ionic compounds are in dynamic equilibrium with their constituent ions.

Now, let's write the equilibrium reaction for very slightly soluble Ag_2SO_4 in an aqueous solution. Ag_2SO_4 precipitate dissociates to give

$$Ag_2SO_4(s) \rightleftharpoons 2\,Ag^+(aq) + SO_4^{2-}(aq)$$

Notice that Ag_2SO_4 dissociates into 2 Ag^+ and 1 SO_4^{2-}. We can substitute this equilibrium into the general K_{eq} expression. This gives

$$K_{eq} = \frac{[Ag^+]^2\,[SO_4^{2-}]}{[Ag_2SO_4]}$$



▲ **Silver Sulfate, Ag_2SO_4**
Silver sulfate is an insoluble compound, which is less than 1% dissociated in an aqueous solution.

If an aqueous solution is saturated with insoluble Ag_2SO_4, the amount of solid precipitate has no effect on the equilibrium. As long as there is some undissolved precipitate, the concentrations of Ag^+ and SO_4^{2-} do not change. The amount of solid has no effect on the equilibrium and we can write a special equilibrium expression as

$$K_{sp} = [Ag^+]^2\,[SO_4^{2-}]$$

where K_{sp} is the **solubility product equilibrium constant**. We can define K_{sp} as a constant that expresses the molar equilibrium concentrations of ions in an aqueous solution for a slightly dissociated ionic compound. For an insoluble compound in an aqueous solution, only the ions appear in the equilibrium expression; the precipitate never appears in the K_{sp} expression. The following example exercise provides further practice in writing solubility product expressions.

EXAMPLE EXERCISE 16.9 Writing K_{sp} Expressions

Write the solubility product expression for slightly soluble calcium phosphate in an aqueous solution.

$$Ca_3(PO_4)_2(s) \rightleftharpoons 3\ Ca^{2+}(aq) + 2\ PO_4{}^{3-}(aq)$$

Solution

Calcium phosphate dissociates into 3 Ca^{2+} and 2 $PO_4{}^{3-}$. Thus, the Ca^{2+} concentration is raised to the third power and the $PO_4{}^{3-}$ concentration to the second power.

$$K_{sp} = [Ca^{2+}]^3\,[PO_4{}^{3-}]^2$$

Practice Exercise

Write the solubility product expression for slightly soluble aluminum carbonate in an aqueous solution.

$$Al_2(CO_3)_3(s) \rightleftharpoons 2\ Al^{3+}(aq) + 3\ CO_3{}^{2-}(aq)$$

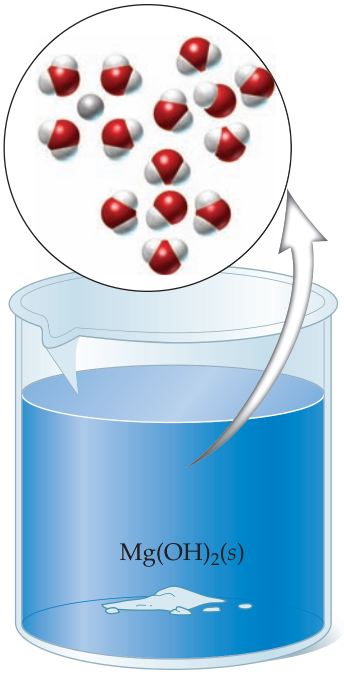
Answer: $K_{sp} = [Al^{3+}]^2\,[CO_3{}^{2-}]^3$

Concept Exercise

Write the equilibrium constant expression for the following slightly soluble ionic compound:

$$Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3\ OH^-(aq)$$

Answer: See Appendix G, 16.9.



▲ **Magnesium Hydroxide, $Mg(OH)_2$** Magnesium hydroxide is an insoluble compound, which is less than 1% dissociated in an aqueous solution.

Determining K_{sp} by Experiment

We can use several methods to find K_{sp} for a slightly soluble ionic compound. One method is to determine experimentally the ion concentrations in an aqueous solution and calculate a value for the solubility product constant. For example, in a saturated solution of milk of magnesia, $Mg(OH)_2$, the hydroxide ion concentration is found to be 0.00032 *M*. To calculate the solubility product constant, let's first analyze the equilibrium.

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq)$$

Notice that $Mg(OH)_2$ dissociates into 1 Mg^{2+} and 2 OH^-. The equilibrium constant expression for the dissociation is

$$K_{sp} = [Mg^{2+}]\,[OH^-]^2$$

Since there is 1 Mg^{2+} for every 2 OH^-, $[Mg^{2+}]$ equals half of $[OH^-]$. The value of $[OH^-]$ is given as 0.00032 *M*, and so the $[Mg^{2+}]$ is 0.00016 *M*. Substituting into the K_{sp} expression, we have

$$K_{sp} = [1.6 \times 10^{-4}]\,[3.2 \times 10^{-4}]^2$$

Simplifying,

$$K_{sp} = 1.6 \times 10^{-11}$$

The solubility product constant for magnesium hydroxide is 1.6×10^{-11}.

16.9 Equilibria Shifts for Slightly Soluble Compounds

LEARNING OBJECTIVE

- Apply Le Chatelier's principle to a saturated solution of a slightly soluble ionic compound.

In Section 16.7, we discussed Le Chatelier's principle as it applies to ionization equilibria. This principle also applies to dissociation equilibria. That is, if we change the concentration of any ions participating in a dissociation equilibrium, there will be a shift to relieve the stress.

Let's consider the dissociation of an antacid tablet in an aqueous solution. The tablet contains aluminum hydroxide, $Al(OH)_3$, which is only slightly soluble. The dissociation equilibrium is

$$Al(OH)_3(s) \rightleftarrows Al^{3+}(aq) + 3\,OH^-(aq)$$

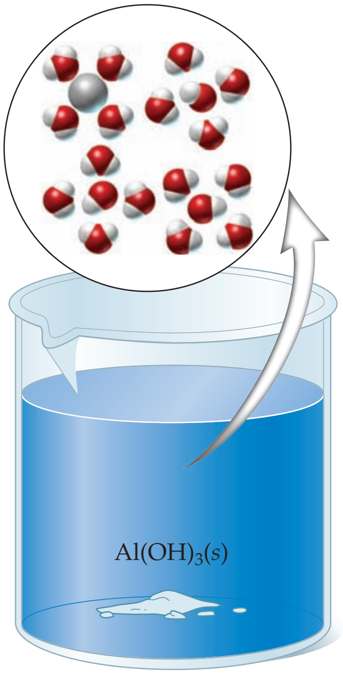
If the solution is saturated, Al^{3+} and OH^- are at maximum concentration. Therefore, if we add more aluminum hydroxide, $Al(OH)_3$, to the solution, it has no effect. Hence, there is no shift in equilibrium.

Suppose we add solid $AlCl_3$ to the solution to increase the Al^{3+} concentration. The solid $AlCl_3$ dissolves to give more Al^{3+}. The aluminum ion stresses the right side of the reaction, and the equilibrium shifts to the left. Furthermore, more hydroxide ions precipitate and the concentration of OH^- decreases.

Suppose we add solid NaCl to the aqueous solution. Sodium chloride dissolves to give sodium ions and chloride ions. Since neither Na^+ nor Cl^- participate in the equilibrium, NaCl has no effect. Thus, there is no shift in equilibrium, and the concentrations of Al^{3+} and OH^- remain constant.

Let's consider a less obvious stress on the equilibrium. Let's add aqueous HCl and increase the H^+ concentration. Hydrochloric acid is a strong acid and ionizes completely to give H^+ and Cl^-. Although neither ion is shown in the equilibrium, H^+ neutralizes OH^-. Thus, the H^+ decreases the OH^- concentration.

As the concentration of OH^- decreases, we stress the right side of the equation and shift the equilibrium to the right. As the equilibrium shifts, more $Al(OH)_3$ dissolves and the Al^{3+} concentration increases. The following example exercise further illustrates Le Chatelier's principle.



▲ Aluminum Hydroxide, $Al(OH)_3$ Aluminum hydroxide is an insoluble compound, which is less than 1% dissociated in an aqueous solution.

▲ Saturated Solution of AgCl By adding substances into the solution, we stress the solubility equilibrium. A stress may cause a Le Chatelier shift to the left (more precipitate) or to the right (less precipitate).

EXAMPLE EXERCISE 16.10 Solubility Equilibria Shifts

A saturated solution of silver chloride dissociates according to the following equation:

$$AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[Ag^+]$
(b) decrease $[Ag^+]$
(c) increase $[Cl^-]$
(d) decrease $[Cl^-]$
(e) add solid $AgNO_3$
(f) add solid $NaNO_3$
(g) add solid AgCl
(h) decrease pH

Solution

Let's apply Le Chatelier's principle to each of the stresses:

(a) If $[Ag^+]$ increases, the equilibrium shifts to the *left.*
(b) If $[Ag^+]$ decreases, the equilibrium shifts to the *right.*
(c) If $[Cl^-]$ increases, the equilibrium shifts to the *left.*
(d) If $[Cl^-]$ decreases, the equilibrium shifts to the *right.*
(e) If $AgNO_3$ is added, $[Ag^+]$ increases and the equilibrium shifts to the *left.*
(f) If $NaNO_3$ is added, $[Na^+]$ and $[NO_3^-]$ increase, which has no effect on the equilibrium. Thus, there is *no shift.*
(g) If AgCl is added, there is no effect on the equilibrium because the solution is saturated with AgCl. Thus, there is *no shift.*
(h) If the pH decreases, $[H^+]$ increases. Since the hydrogen ion has no effect on the equilibrium, there is *no shift.*

Practice Exercise

A saturated solution of magnesium hydroxide dissociates according to the following equation:

$$\mathrm{Mg(OH)_2}(s) \rightleftharpoons \mathrm{Mg^{2+}}(aq) + 2\,\mathrm{OH^-}(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[Mg^{2+}]$
(b) decrease $[Mg^{2+}]$
(c) increase $[OH^-]$
(d) decrease $[OH^-]$
(e) add solid KOH
(f) add solid KNO_3
(g) add solid $Mg(OH)_2$
(h) decrease pH

Answers:

(a) shifts left
(b) shifts right
(c) shifts left
(d) shifts right
(e) shifts left
(f) no shift
(g) no shift
(h) shifts right (neutralizes OH^-)

Concept Exercise

What is the effect on the following equilibrium of adding: (a) $Ca(NO_3)_2$, and (b) Na_2CO_3?

$$\mathrm{CaCO_3}(s) \rightleftharpoons \mathrm{Ca^{2+}}(aq) + \mathrm{CO_3^{2-}}(aq)$$

Answer: See Appendix G, 16.10.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
16.1 Collision Theory According to the **collision theory** of reaction rates, the rate of a reaction is regulated by the collision frequency, collision energy, and orientation of the molecules. We can speed up a reaction by (1) increasing the concentration of reactants, (2) raising the temperature, or (3) adding a catalyst.	• Explain the effect of collision frequency, collision energy, and orientation of molecules on the rate of a chemical reaction. ***Related Exercises: 1–4*** • Explain the effect of concentration, temperature, and a catalyst on the rate of a chemical reaction. ***Related Exercises: 5–8***
16.2 Energy Profiles of Chemical Reactions An **endothermic reaction** absorbs heat energy, and an **exothermic reaction** releases heat energy. A **reaction profile** graphs the change in energy as reactants are converted to products. For molecules to react, they must have enough energy to achieve the **transition state**. The energy necessary to reach the transition state is called the **activation energy**, E_{act}. A **catalyst** can be defined as any substance that speeds up a reaction by lowering the activation energy. The difference between the energy of the reactants and the products is called the **heat of reaction**, ΔH.	• Diagram the general energy profile for an endothermic and an exothermic reaction. ***Related Exercises: 9–10*** • Label the transition state, energy of activation, and heat of reaction on a given energy profile. ***Related Exercises: 11–14***
16.3 The Chemical Equilibrium Concept A **reversible reaction** takes place simultaneously in both the forward and reverse directions. The **rate of reaction** is the rate at which the concentration of the reactants decreases, or the concentration of the products increases, per unit time. When the rates of the forward and reverse reactions are equal, the reaction is at **chemical equilibrium**. The **law of chemical equilibrium** states that the molar concentrations of products divided by reactants (each raised to a power corresponding to a coefficient in the balanced chemical equation) equals a **general equilibrium constant**, K_{eq}. The law of chemical equilibrium applies to every reversible reaction and is illustrated in Table 16.3.	• Describe the equilibrium concept for a reversible reaction. ***Related Exercises: 15–16*** • Express the law of chemical equilibrium as an equation.

Key Concepts

Learning Objectives and Related Exercises

TABLE 16.3 Selected Types of Chemical Equilibria

Type of Equilibrium	Equilibrium Constant Expression
General Equilibrium	
$a\,\mathrm{A} + b\,\mathrm{B} \rightleftharpoons c\,\mathrm{C} + d\,\mathrm{D}$	$K_{eq} = \dfrac{[\mathrm{C}]^c\,[\mathrm{D}]^d}{[\mathrm{A}]^a\,[\mathrm{B}]^b}$
Ionization Equilibrium	
$\mathrm{HC_2H_3O_2}(aq) \rightleftharpoons \mathrm{H^+}(aq) + \mathrm{C_2H_3O_2^-}(aq)$	$K_i = \dfrac{[\mathrm{H^+}]\,[\mathrm{C_2H_3O_2^-}]}{[\mathrm{HC_2H_3O_2}]}$
Dissociation Equilibrium	
$\mathrm{Al(OH)_3}(s) \rightleftharpoons \mathrm{Al^{3+}}(aq) + 3\,\mathrm{OH^-}(aq)$	$K_{sp} = [\mathrm{Al^{3+}}]\,[\mathrm{OH^-}]^3$

16.4 General Equilibrium Constant, K_{eq}

In a **homogeneous equilibrium**, all of the participating substances are in the same physical state, for example, an equilibrium in which all the reactants and products are in the gaseous state. In a **heterogeneous equilibrium**, one of the participating substances is in a different physical state, for example, an equilibrium in which a solid substance decomposes to give a mixture of gases.

- Write the equilibrium constant expression for a reversible reaction.
  Related Exercises: 17–18
- Calculate an equilibrium constant, K_{eq}, from experimental data.
  Related Exercises: 19–20

16.5 Equilibria Shifts for Gases

A change in concentration, temperature, or pressure can cause a stress to a reversible reaction. According to **Le Chatelier's principle**, a chemical reaction at equilibrium shifts in order to relieve a stress. The shifts in equilibrium for various stresses are summarized in Table 16.4.

- Apply Le Chatelier's principle to reversible reactions in the gaseous state.
  Related Exercises: 21–24

TABLE 16.4 Summary of Chemical Equilibria Shifts

Stress on Equilibrium	Effect on Equilibrium
Concentration Increases (all systems)	
for a reactant	shifts to the right
for a product	shifts to the left
Temperature Increases (gaseous systems)	
for an endothermic reaction	shifts to the right
for an exothermic reaction	shifts to the left
Pressure Increases (gaseous systems)	
more reactant molecules	shifts to the right
more product molecules	shifts to the left
molecules of reactants and products are equal	no shift

16.6 Ionization Equilibrium Constant, K_i

The law of chemical equilibrium applies to aqueous solutions of weak acids and weak bases. The **ionization equilibrium constant, K_i**, is equal to the molar concentrations of the ions divided by the molar concentration of the weak acid or base.

- Write the equilibrium constant expression for a weak acid or a weak base.
  Related Exercises: 25–26
- Calculate an ionization constant, K_i, from experimental data.
  Related Exercises: 27–30

Key Concepts	Learning Objectives and Related Exercises
16.7 Equilibria Shifts for Weak Acids and Bases Le Chatelier's principle applies to aqueous solutions of weak acids and weak bases. If we change the concentration of one of the ions in an aqueous solution, the equilibrium will shift to relieve the stress. If we add a substance that does not effect the equilibrium, there will be no shift.	• Apply Le Chatelier's principle to solutions of weak acids and weak bases. ***Related Exercises: 31–34***
16.8 Solubility Product Equilibrium Constant, K_{sp} The law of chemical equilibrium applies to saturated solutions of slightly soluble ionic compounds. The **solubility product equilibrium constant**, K_{sp}, is equal to the product of the molar concentrations of the ions in solution. The insoluble compound is not written in the K_{sp} expression because it has no effect on the equilibrium in a saturated solution.	• Write the equilibrium constant expression for a slightly soluble ionic compound. ***Related Exercises: 35–36*** • Calculate a solubility product constant, K_{sp}, from experimental data. ***Related Exercises: 37–38***
16.9 Equilibria Shifts for Slightly Soluble Compounds Le Chatelier's principle applies to saturated solutions of slightly soluble ionic compounds. If we change the concentration of one of the ions in an aqueous solution, the equilibrium will shift to relieve the stress. If we add a substance that does not participate in the equilibrium, there will be no shift.	• Apply Le Chatelier's principle to a saturated solution of a slightly soluble ionic compound. ***Related Exercises: 39–42***

Key Terms

Answers to Key Terms are in Appendix H.

Select the key term that corresponds to each of the following definitions.

_____ **1.** the principle that the rate of a chemical reaction is controlled by the frequency and energy of molecules striking each other

_____ **2.** a chemical reaction that consumes heat energy

_____ **3.** a chemical reaction that liberates heat energy

_____ **4.** a graph of the energy of reactants and products as a reaction occurs

_____ **5.** the highest point on the reaction profile

_____ **6.** the energy required for reactants to reach the transition state

_____ **7.** the difference in energy between the reactants and the products

_____ **8.** a substance that speeds up a reaction by lowering the energy of activation

_____ **9.** a reaction that proceeds toward reactants and products simultaneously

_____ **10.** the rate at which the concentrations of reactants decrease per unit time

_____ **11.** a dynamic state for a reversible reaction in which the rates of the forward and reverse reactions are the same

_____ **12.** the principle that the molar concentrations of the products in a reversible reaction divided by the molar concentrations of the reactants (each raised to a power corresponding to a coefficient in the balanced equation) is equal to a constant

_____ **13.** the constant that expresses the molar equilibrium concentration of each substance participating in a reversible reaction at a given temperature

_____ **14.** a type of equilibrium in which all participating species are in the same physical state (e.g., gaseous state or aqueous solution)

_____ **15.** a type of equilibrium in which all participating species are not in the same physical state

_____ **16.** the statement that a chemical equilibrium stressed by a change in concentration, temperature, or pressure, shifts to relieve the stress

_____ **17.** the constant that expresses the molar equilibrium concentrations of ions in an aqueous solution for a slightly ionized acid or base

_____ **18.** the constant that expresses the molar equilibrium concentrations of ions in an aqueous solution for a slightly dissociated ionic compound

(a) activation energy (E_{act}) (*Sec. 16.2*)

(b) catalyst (*Sec. 16.2*)

(c) chemical equilibrium (*Sec. 16.3*)

(d) collision theory (*Sec. 16.1*)

(e) endothermic reaction (*Sec. 16.2*)

(f) exothermic reaction (*Sec. 16.2*)

(g) general equilibrium constant (K_{eq}) (*Sec. 16.3*)

(h) heat of reaction (ΔH) (*Sec. 16.2*)

(i) heterogeneous equilibrium (*Sec. 16.4*)

(j) homogeneous equilibrium (*Sec. 16.4*)

(k) ionization equilibrium constant (K_i) (*Sec. 16.6*)

(l) law of chemical equilibrium (*Sec. 16.3*)

(m) Le Chatelier's principle (*Sec. 16.5*)

(n) rate of reaction (*Sec. 16.3*)

(o) reaction profile (*Sec. 16.2*)

(p) reversible reaction (*Sec. 16.3*)

(q) solubility product equilibrium constant (K_{sp}) (*Sec. 16.8*)

(r) transition state (*Sec. 16.2*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Collision Theory (Sec. 16.1)

1. Which of the following theoretical factors decreases the rate of a reaction?
   (a) decrease collision frequency
   (b) decrease collision energy
   (c) ineffective collision orientation
2. Which of the following theoretical factors increases the rate of a reaction?
   (a) increase collision frequency
   (b) increase collision energy
   (c) effective collision orientation
3. Draw a diagram showing an effective collision orientation between a hydrogen molecule and a bromine molecule.
4. Draw a diagram showing an ineffective collision orientation between a hydrogen molecule and a bromine molecule.
5. State the effect on the rate of reaction for each of the following.
   (a) decrease the concentration of a reactant
   (b) decrease the temperature of the reaction
   (c) add an acid catalyst
6. State the effect on the rate of reaction for each of the following.
   (a) increase the concentration of a reactant
   (b) increase the temperature of the reaction
   (c) add a metal catalyst
7. The Contact Process uses a metal catalyst to produce sulfur trioxide gas for the manufacture of sulfuric acid. Does the catalyst increase the amount of sulfur trioxide gas? Explain.
8. The Haber Process uses a metal oxide catalyst to produce ammonia gas. Does the catalyst increase the amount of ammonia gas? Explain.

Energy Profiles of Chemical Reactions (Sec. 16.2)

9. Phosphorus pentachloride is used in the electronics industry to manufacture computer chips. Draw the energy profile for the following reaction.

$$PCl_5(g) \quad + \quad \textit{heat} \quad \rightleftarrows \quad PCl_3(g) \quad + \quad Cl_2(g)$$

10. Boron trifluoride is used in the electronics industry to manufacture computer chips. Draw the energy profile for the following reaction.

$$2\,B(s) \quad + \quad 3\,F_2(g) \quad \rightleftarrows \quad 2\,BF_3(g) \quad + \quad \textit{heat}$$

11. Draw the energy profile for the following exothermic reaction:

$$H_2(g) \quad + \quad Cl_2(g) \quad \rightleftarrows \quad 2\,HCl(g)$$

Label the axes *Energy* and *Progress of reaction* and indicate the reactants, products, transition state, activation energy, and energy of reaction.

12. Draw the energy profile for the following endothermic reaction:

$$H_2(g) \quad + \quad I_2(g) \quad \rightleftarrows \quad 2\,HI(g)$$

Label the axes *Energy* and *Progress of reaction* and indicate the reactants, products, transition state, activation energy, and energy of reaction.

13. State the effect of a catalyst on the energy of activation, E_{act}.
14. State the effect of a catalyst on the heat of reaction, ΔH.

The Chemical Equilibrium Concept (Sec. 16.3)

15. Which of the following statements is true regarding the general equilibrium expression?
   (a) K_{eq} for gaseous equilibria does not depend on temperature.
   (b) K_{eq} for gaseous equilibria does not include solid substances.
16. Which of the following statements is true regarding the general equilibrium expression?
   (a) K_{eq} can be determined experimentally.
   (b) K_{eq} can be determined theoretically.

General Equilibrium Constant, K_{eq} (Sec. 16.4)

17. Write the equilibrium constant expression for each of the following reversible reactions:
   (a) $H_2(g) + F_2(g) \rightleftarrows 2\,HF(g)$
   (b) $4\,NH_3(g) + 7\,O_2(g) \rightleftarrows 4\,NO_2(g) + 6\,H_2O(g)$
   (c) $ZnCO_3(s) \rightleftarrows ZnO(s) + CO_2(g)$
18. Write the equilibrium constant expression for each of the following reversible reactions:
   (a) $H_2(g) + Br_2(g) \rightleftarrows 2\,HBr(g)$
   (b) $4\,HCl(g) + O_2(g) \rightleftarrows 2\,Cl_2(g) + 2\,H_2O(g)$
   (c) $CO(g) + 2\,H_2(g) \rightleftarrows CH_3OH(l)$
19. Given the equilibrium concentrations for each gas at 850 °C, calculate the value of K_{eq} for the manufacture of sulfur trioxide.

$$\begin{array}{ccccc} 2\,SO_2(g) & + & O_2(g) & \rightleftarrows & 2\,SO_3(g) \\ 1.75\,M & & 1.50\,M & & 2.25\,M \end{array}$$

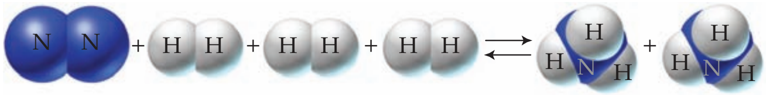
20. Given the equilibrium concentrations for each gas at 500 °C, calculate the value of K_{eq} for the manufacture of ammonia.

$$\begin{array}{ccccc} N_2(g) & + & 3\,H_2(g) & \rightleftarrows & 2\,NH_3(g) \\ 0.400\,M & & 1.20\,M & & 0.195\,M \end{array}$$

Equilibria Shifts for Gases (Sec. 16.5)

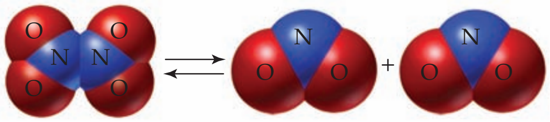
21. The conditions for producing ammonia industrially are 500 °C and 300 atm. What happens to the ammonia concentration if (a) the temperature increases and (b) the pressure increases?

$$N_2(g) \quad + \quad 3\,H_2(g) \quad \rightleftarrows \quad 2\,NH_3(g) \quad + \quad \textit{heat}$$



22. Weather conditions affect the smog equilibrium in the atmosphere. What happens to the nitrogen dioxide concentration on (a) hot, sunny days and (b) cool, overcast days?

$$N_2O_4(g) \quad + \quad \textit{heat} \quad \rightleftarrows \quad 2\,NO_2(g)$$



23. The industrial process for producing hydrogen gas involves reacting methane and steam at a high temperature.

$$CH_4(g) + H_2O(g) + \textit{heat} \rightleftarrows CO(g) + 3\,H_2(g)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[CH_4]$
(b) decrease $[H_2O]$
(c) increase [CO]
(d) decrease $[H_2]$
(e) increase temperature
(f) decrease temperature
(g) increase volume
(h) decrease volume
(i) add Zn dust
(j) add Xe inert gas

24. Smog contains formaldehyde that is responsible for an eye-burning sensation. Formaldehyde, CH_2O, is produced from the reaction of ozone and atmospheric ethylene, C_2H_4, as follows.

$$2\,C_2H_4(g) + 2\,O_3(g) \rightleftarrows 4\,CH_2O(g) + O_2(g) + \textit{heat}$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[C_2H_4]$
(b) decrease $[O_3]$
(c) increase $[CH_2O]$
(d) decrease $[O_2]$
(e) increase temperature
(f) decrease temperature
(g) increase volume
(h) decrease volume
(i) add Ne inert gas
(j) ultraviolet light

Ionization Equilibrium Constant, K_i (Sec. 16.6)

25. Write the equilibrium constant expression for each of the following weak acids:

(a) $HCHO_2(aq) \rightleftarrows H^+(aq) + CHO_2^-(aq)$
(b) $H_2C_2O_4(aq) \rightleftarrows H^+(aq) + HC_2O_4^-(aq)$
(c) $H_3C_6H_5O_7(aq) \rightleftarrows H^+(aq) + H_2C_6H_5O_7^-(aq)$

26. Write the equilibrium constant expression for each of the following weak bases:

(a) $NH_2OH(aq) + H_2O(l) \rightleftarrows NH_3OH^+(aq) + OH^-(aq)$
(b) $C_6H_5NH_2(aq) + H_2O(l) \rightleftarrows C_6H_5NH_3^+(aq) + OH^-(aq)$
(c) $(CH_3)_2NH(aq) + H_2O(l) \rightleftarrows (CH_3)_2NH_2^+(aq) + OH^-(aq)$

27. Nitrous acid, HNO_2, is used in the synthesis of organic dye compounds. If the hydrogen ion concentration of a 0.125 *M* solution is 7.5×10^{-3} *M*, what is the ionization constant for the acid?

28. Aqueous ammonium hydroxide, NH_4OH, is used as a household cleaning solution. If the hydroxide ion concentration of a 0.245 *M* solution is 2.1×10^{-3}*M*, what is the ionization constant for the base?

29. Hydrofluoric acid, HF, is used in the manufacture of silicon computer chips. If the pH of a 0.139 *M* solution is 2.00, what is the ionization constant of the acid?

30. Hydrazine, N_2H_4, is a weak base and is used as fuel in the space shuttle. If the pH of a 0.139 *M* solution is 11.00, what is the ionization constant of the base?

$$N_2H_4(aq) + H_2O(l) \rightleftarrows N_2H_5^+(aq) + OH^-(aq)$$

Equilibria Shifts for Weak Acids and Bases (Sec. 16.7)

31. Given the chemical equation for the ionization of hydrofluoric acid

$$HF(aq) \rightleftarrows H^+(aq) + F^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase [HF]
(b) increase $[H^+]$
(c) decrease [HF]
(d) decrease $[F^-]$
(e) add solid NaF
(f) add gaseous HCl
(g) add solid NaOH
(h) increase pH

32. Given the chemical equation for the ionization of nitrous acid

$$HNO_2(aq) \quad \rightleftarrows \quad H^+(aq) \quad + \quad NO_2^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) decrease $[HNO_2]$
(b) decrease $[H^+]$
(c) increase $[HNO_2]$
(d) increase $[NO_2^-]$
(e) add solid KNO_2
(f) add solid KCl
(g) add solid KOH
(h) increase pH

33. Given the chemical equation for the ionization of acetic acid

$$HC_2H_3O_2(aq) \quad \rightleftarrows \quad H^+(aq) \quad + \quad C_2H_3O_2^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[HC_2H_3O_2]$
(b) increase $[H^+]$
(c) decrease $[HC_2H_3O_2]$
(d) decrease $[C_2H_3O_2^-]$
(e) add solid $NaC_2H_3O_2$
(f) add solid NaCl
(g) add solid NaOH
(h) increase pH

34. Given the chemical equation for the ionization of ammonium hydroxide

$$NH_4OH(aq) \quad \rightleftarrows \quad NH_4^+(aq) \quad + \quad OH^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[NH_4^+]$
(b) decrease $[OH^-]$
(c) increase $[NH_4OH]$
(d) decrease pH
(e) add gaseous NH_3
(f) add solid KCl
(g) add solid KOH
(h) add solid NH_4Cl

Solubility Product Equilibrium Constant, K_{sp} (Sec. 16.8)

35. Write the solubility product expression for each of the following slightly soluble ionic compounds in a saturated aqueous solution:

(a) $AgI(s) \rightleftarrows Ag^+(aq) + I^-(aq)$
(b) $Ag_2CrO_4(s) \rightleftarrows 2\,Ag^+(aq) + CrO_4^{2-}(aq)$
(c) $Ag_3PO_4(s) \rightleftarrows 3\,Ag^+(aq) + PO_4^{3-}(aq)$

36. Write the solubility product expression for each of the following slightly soluble ionic compounds in a saturated aqueous solution:

(a) $Cu_2CO_3(s) \rightleftarrows 2\,Cu^+(aq) + CO_3^{2-}(aq)$
(b) $ZnCO_3(s) \rightleftarrows Zn^{2+}(aq) + CO_3^{2-}(aq)$
(c) $Al_2(CO_3)_3(s) \rightleftarrows 2\,Al^{3+}(aq) + 3\,CO_3^{2-}(aq)$

37. The cobalt(II) ion concentration in a saturated solution of cobalt(II) sulfide, CoS, is 7.7×10^{-11} *M*. Calculate the value for the equilibrium constant.

38. The fluoride ion concentration in a saturated solution of magnesium fluoride, MgF_2, is 2.3×10^{-3} *M*. Calculate the value for the equilibrium constant.

Equilibria Shifts for Slightly Soluble Compounds (Sec. 16.9)

39. Teeth and bones are composed mainly of calcium phosphate, which dissociates slightly in an aqueous solution as follows:

$$Ca_3(PO_4)_2(s) \rightleftarrows 3\,Ca^{2+}(aq) + 2\,PO_4^{\;3-}(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[Ca^{2+}]$
(b) increase $[PO_4^{\;3-}]$
(c) decrease $[Ca^{2+}]$
(d) decrease $[PO_4^{\;3-}]$
(e) add solid $Ca_3(PO_4)_2$
(f) add solid $Ca(NO_3)_2$
(g) add solid KNO_3
(h) decrease pH

40. Strontium carbonate dissociates slightly in an aqueous solution as follows:

$$SrCO_3(s) \rightleftarrows Sr^{2+}(aq) + CO_3^{\;2-}(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[Sr^{2+}]$
(b) increase $[CO_3^{\;2-}]$
(c) decrease $[Sr^{2+}]$
(d) decrease $[CO_3^{\;2-}]$
(e) add solid $SrCO_3$
(f) add solid $Sr(NO_3)_2$
(g) add solid KNO_3
(h) add H^+

41. Cupric hydroxide dissociates slightly in an aqueous solution as follows:

$$Cu(OH)_2(s) \rightleftarrows Cu^{2+}(aq) + 2\,OH^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[Cu^{2+}]$
(b) increase $[OH^-]$
(c) decrease $[Cu^{2+}]$
(d) decrease $[OH^-]$
(e) add solid $Cu(OH)_2$
(f) add solid NaOH
(g) add solid NaCl
(h) decrease pH

42. Cadmium sulfide dissociates slightly in an aqueous solution as follows:

$$CdS(s) \rightleftarrows Cd^{2+}(aq) + S^{2-}(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[Cd^{2+}]$
(b) increase $[S^{2-}]$
(c) decrease $[Cd^{2+}]$
(d) decrease $[S^{2-}]$
(e) add solid CdS
(f) add solid $Cd(NO_3)_2$
(g) add solid $NaNO_3$
(h) add H^+

General Exercises

43. The N_2O_4–NO_2 reversible reaction is found to have the following equilibrium concentrations at 100°C. Calculate K_{eq} for the reaction.

$$\begin{array}{ccc} N_2O_4(g) & \rightleftarrows & 2\;NO_2(g) \\ 4.5 \times 10^{-5}\,M & & 3.0 \times 10^{-3}\,M \end{array}$$

44. The N_2O_4–NO_2 reversible reaction is found to have the following equilibrium partial pressures at 100°C. Calculate K_{sp} for the reaction.

$$\begin{array}{ccc} N_2O_4(g) & \rightleftarrows & 2\;NO_2(g) \\ 0.0014\text{ atm} & & 0.092\text{ atm} \end{array}$$

45. Given the chemical equation for the ionization of water:

$$H_2O(l) \rightleftarrows H^+(aq) + OH^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) increase $[H^+]$
(b) decrease $[OH^-]$
(c) increase pH
(d) decrease pH

46. Given the chemical equation for the ionization of water:

$$H_2O(l) \rightleftarrows H^+(aq) + OH^-(aq)$$

Predict the direction of equilibrium shift for each of the following stresses:

(a) add gaseous HCl
(b) add solid NaOH
(c) add liquid H_2SO_4
(d) add solid NaF

Challenge Exercises

47. A saturated solution of calcium hydroxide, $Ca(OH)_2$, has a pH of 12.35. Find the hydroxide ion concentration and calculate the equilibrium constant.

48. A saturated solution of zinc hydroxide, $Zn(OH)_2$, has a pH of 8.44. Find the hydroxide ion concentration and calculate the equilibrium constant.

Chapter 16 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following factors influences the rate of a chemical reaction? (Sec. 16.1)
(a) concentration
(b) temperature
(c) catalyst
(d) all of the above
(e) none of the above

2. If the heat of reaction is exothermic, which of the following is true? (Sec. 16.2)
(a) The energy of the reactants is greater than the products.
(b) The energy of the reactants is less than the products.
(c) The reaction rate is fast.
(d) The reaction rate is slow.
(e) none of the above

3. Which of the following is true after a reaction reaches equilibrium? (Sec. 16.3)

(a) The amount of reactants and products are equal.
(b) The amount of reactants and products are constant.
(c) The amount of reactants is increasing.
(d) The amount of products is increasing.
(e) none of the above

4. Dinitrogen tetraoxide decomposes to produce nitrogen dioxide. Calculate the equilibrium constant for the reaction given the equilibrium concentrations at 100 °C: $[N_2O_4] = 0.800$ and $[NO_2] = 0.400$. (Sec. 16.4)

$$N_2O_4(g) \rightleftarrows 2\ NO_2(g)$$

(a) $K_{eq} = 0.200$
(b) $K_{eq} = 0.500$
(c) $K_{eq} = 0.625$
(d) $K_{eq} = 2.00$
(e) $K_{eq} = 5.00$

5. Which of the changes listed has *no effect* on the equilibrium for the following reversible reaction? (Sec. 16.5)

$$CO(g) + H_2O(g) \rightleftarrows CO_2(g) + H_2(g) + \text{heat}$$

(a) increase volume
(b) decrease volume
(c) add a catalyst
(d) add neon gas
(e) all of the above

6. If the hydrogen ion concentration of a 0.100 *M* $HC_3H_5O_3$ solution is 0.0038 *M*, what is the ionization constant for lactic acid? (Sec. 16.6)

(a) $K_i = 1.4 \times 10^{-3}$
(b) $K_i = 1.4 \times 10^{-4}$
(c) $K_i = 1.4 \times 10^{-5}$
(d) $K_i = 3.8 \times 10^{-3}$
(e) $K_i = 3.8 \times 10^{-4}$

7. Which of the changes listed has *no effect* on the equilibrium for the following reversible reaction in an aqueous solution? (Sec. 16.7)

$$HC_2H_3O_2(aq) \rightleftarrows H^+(aq) + C_2H_3O_2^-(aq)$$

(a) increase $[HC_2H_3O_2]$
(b) increase $[H^+]$
(c) add solid $NaC_2H_3O_2$
(d) add solid $NaNO_3$
(e) add solid NaOH

8. What is the K_{sp} for silver carbonate, Ag_2CO_3, if the silver ion concentration in a saturated solution is 0.00026 *M*? (Sec. 16.8)

(a) $K_{sp} = 3.4 \times 10^{-8}$
(b) $K_{sp} = 6.8 \times 10^{-8}$
(c) $K_{sp} = 3.6 \times 10^{-11}$
(d) $K_{sp} = 4.4 \times 10^{-12}$
(e) $K_{sp} = 8.8 \times 10^{-12}$

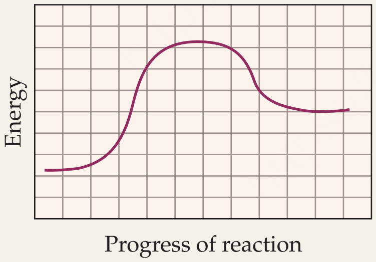
9. Which of the changes listed below has *no effect* on the equilibrium for the following reversible reaction? (Sec. 16.9)

$$Mg(OH)_2(s) \rightleftarrows Mg^{2+}(aq) + 2\ OH^-(aq)$$

(a) increase $[Mg^{2+}]$
(b) increase pH
(c) add sulfuric acid
(d) add solid $Mg(OH)_2$
(e) add solid $Mg(NO_3)_2$

Key Concepts

10. Does the following reaction profile represent an exothermic or endothermic reaction?



11. What is the equilibrium constant expression for the oxygen/ozone reaction?

$$3\ O_2(g) \quad \rightleftarrows \quad 2\ O_3(g)$$

12. What is the direction of equilibrium shift after heating the reaction?

$$3\ O_2(g) \quad \rightleftarrows \quad 2\ O_3(g)$$

13. What is the direction of equilibrium shift after reducing the volume?

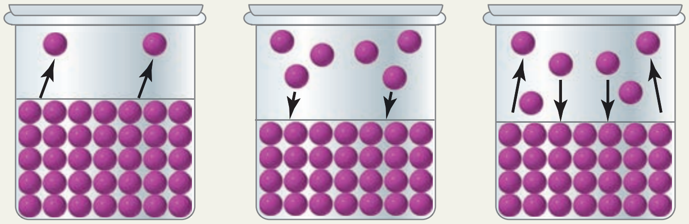
$$3\ O_2(g) \quad \rightleftarrows \quad 2\ O_3(g)$$

14. What is the direction of equilibrium shift after adding a catalyst?

$$3\ O_2(g) \quad \rightleftarrows \quad 2\ O_3(g)$$

Critical Thinking

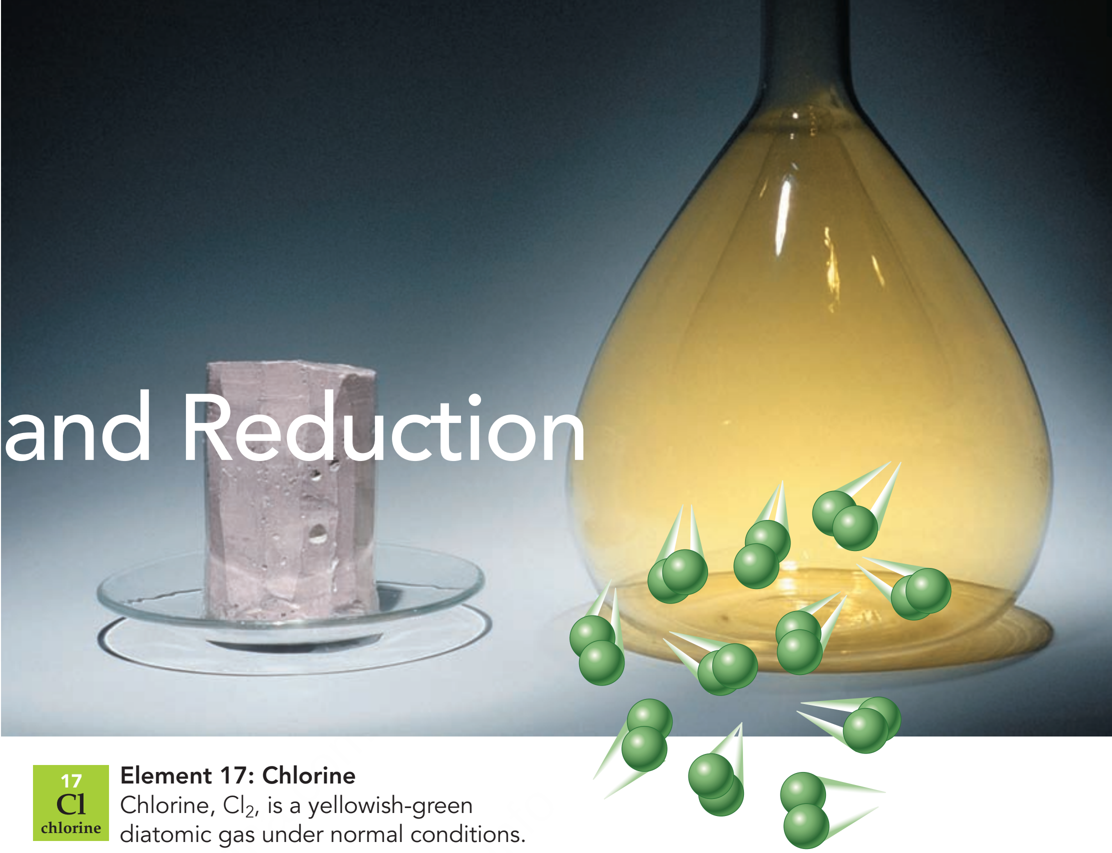
15. A liquid in a sealed beaker begins to evaporate as shown in (a). When the volume above the liquid becomes saturated with vapor, the vapor begins to condense to a liquid as shown in (b). What is true of the rates of evaporation and condensation when the system achieves equilibrium (c)?



(a) (b) (c)

CHAPTER 17

and Reduction



> *"The only person who never makes a mistake is the person who never does anything."*
>
> Theodore Roosevelt, President (1858–1919)

17
Cl
chlorine

Element 17: Chlorine
Chlorine, Cl_2, is a yellowish-green diatomic gas under normal conditions. The most common compound of chlorine is sodium chloride, which has been known since ancient times and used as payment for slave labor. The English chemist Humphry Davy described the yellowish-green gaseous element in 1774 and named it chlorine, after the Greek meaning "pale yellow." (Sodium metal shown left.)

When electrons are transferred from one substance to another, one substance is said to be *oxidized* and the other is said to be *reduced*. In this chapter we will study reactions involving the transfer of electrons from one substance to another, that is, *oxidation–reduction* reactions.

Although we have not previously used the term *oxidation–reduction*, many of the reactions we have already studied involve the transfer of electrons. In fact, whenever a metal reacts with a nonmetal, electrons are transferred. For instance, iron metal rusts when exposed to oxygen in the atmosphere (Figure 17.1b). That is, iron reacts with oxygen to give iron(III) oxide. The equation for the reaction is

$$4\ \mathrm{Fe}(s) + 3\ \mathrm{O_2}(g) \longrightarrow 2\ \mathrm{Fe_2O_3}(s)$$

Iron(III) oxide is an ionic compound, so Fe_2O_3 exists as the ions Fe^{3+} and O^{2-}. In the process of becoming an ion, an iron atom loses three electrons. We say that the iron metal is *oxidized* because it loses electrons to the oxygen gas.

$$\mathrm{Fe} \longrightarrow \mathrm{Fe^{3+}} + 3\ \mathrm{e^-}$$

As the reaction proceeds, we say that the oxygen is *reduced* because it gains electrons from the iron metal. In this example, each oxygen molecule reacts by gaining four electrons, two for each oxygen atom, and becoming two oxide ions.

$$\mathrm{O_2} + 4\ \mathrm{e^-} \longrightarrow 2\ \mathrm{O^{2-}}$$

Let's consider another example. When silver metal is exposed to traces of hydrogen sulfide in the air, the silver tarnishes. The equation for the reaction is

$$2\ \mathrm{Ag}(s) + \mathrm{H_2S}(g) \longrightarrow \mathrm{Ag_2S}(s) + \mathrm{H_2}(g)$$

Because silver sulfide is an ionic compound, we can write the ions as Ag^+ and S^{2-}. In the process of becoming an ion, a silver atom loses an electron. We say that the silver metal is *oxidized* by hydrogen sulfide gas.

$$\mathrm{Ag} \longrightarrow \mathrm{Ag^+} + \mathrm{e^-}$$

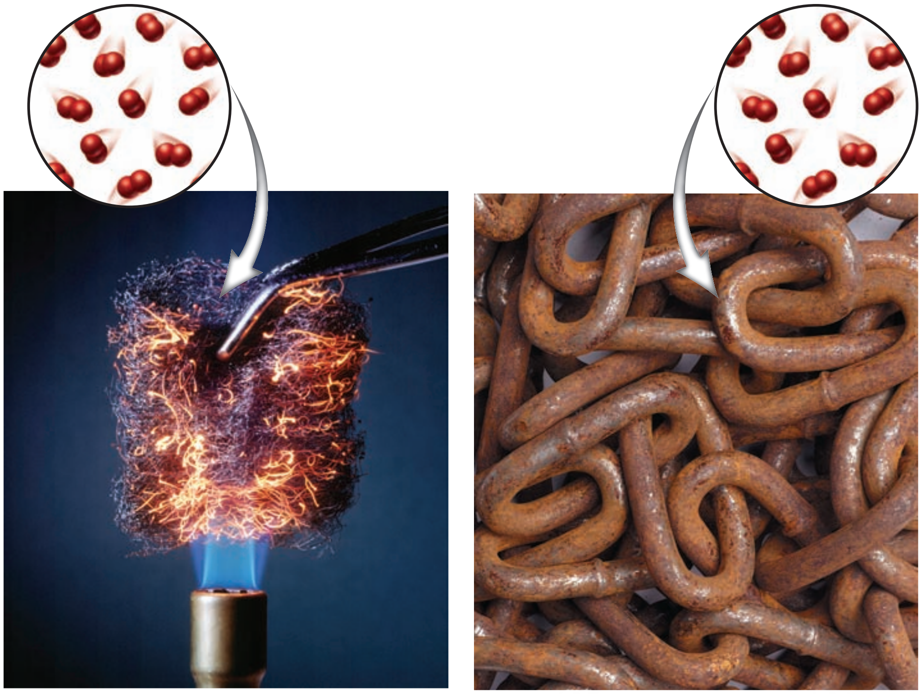

(a) (b)

Figure 17.1 Oxidation of Iron Examples of oxidation include (a) burning, as shown by the rapid reaction of oxygen and steel wool; and (b) rusting, as shown by the slow reaction of oxygen and a steel chain.

Simultaneously, hydrogen sulfide is *reduced* as silver is oxidized. Hydrogen sulfide reacts by gaining two electrons and yielding a hydrogen molecule.

$$H_2S + 2\,e^- \longrightarrow H_2 + S^{2-}$$

This overall process of one substance being oxidized, while another is being reduced, is called an oxidation–reduction reaction. Chemists refer to this process as simply a *redox reaction*.

17.1 Oxidation Numbers

The redox reactions involving silver and iron show that an atom can lose or gain different numbers of electrons. In order to describe the number of electrons lost or gained by an atom, an **oxidation number** is assigned according to a set of rules (Table 17.1). All uncombined elements are said to be in the free state, and are electrically neutral. Their oxidation numbers have a value of zero. For example, Ag, Cu, Fe, S, and P have an oxidation number of zero. Even nonmetals that exist as molecules, such as O_2 and Cl_2, have an oxidation number of zero.

LEARNING OBJECTIVES

- Explain the concept of an oxidation number.
- Solve for the oxidation number for an element in each of the following:
  - (a) metals and nonmetals
  - (b) monoatomic and polyatomic ions
  - (c) ionic and molecular compounds.

Metals in the Free State The metals shown are magnesium, copper, and gold (clockwise from top) and have an oxidation number of zero.

▶ **Nonmetals in the Free State** The nonmetals shown are yellow sulfur, violet iodine, and red phosphorus (clockwise) and have an oxidation number of zero.

The oxidation number of a monoatomic ion is equal to its ionic charge. For example, the oxidation number for Ca^{2+} is positive two (+2), and for Cl^- negative one (−1). Note the distinction between the way ionic charges and oxidation numbers are written. Ionic charges are indicated by the number followed by the sign, and oxidation numbers are indicated by the sign followed by the number. For example, the oxidation numbers for Fe^{2+} and Fe^{3+} are +2, and +3, respectively.

Let's see how to assign oxidation numbers to elements in a compound. Recall that compounds are overall electrically neutral, and so the sum of the oxidation numbers of the individual atoms in a compound is 0. In binary ionic compounds, compounds that contain one metal and one nonmetal, the oxidation numbers for the metal and the nonmetal correspond to their ionic charges. For example, in NaCl the oxidation number of Na^+ is +1, and that of Cl^- is −1. In AlF_3, the oxidation number of Al^{3+} is +3, and that of F^- is −1. In ternary ionic compounds, compounds that contain three elements including at least one metal, the situation is more complex, but the sum of the oxidation numbers is still equal to 0.

In binary molecular compounds, compounds that contain two nonmetals, the more electronegative element is assigned an oxidation number equal to the ionic charge of its free ion.

Oxygen usually has an oxidation number of −2, and hydrogen usually has an oxidation number of +1. After assigning oxidation numbers to oxygen and hydrogen, we can determine the value for the other nonmetal in a binary molecular compound. For example, in NO the oxidation number of nitrogen is +2; in NO_2 the oxidation number of nitrogen is +4. We can state general rules for assigning oxidation numbers as listed in Table 17.1.

TABLE 17.1 Rules for Assigning Oxidation Numbers

1. A metal or a nonmetal in the free state has an oxidation number of 0.
2. A monoatomic ion has an oxidation number equal to its ionic charge.
3. A hydrogen atom is usually assigned an oxidation number of +1.
4. An oxygen atom is usually assigned an oxidation number of −2.
5. For a molecular compound, the more electronegative element is assigned a negative oxidation number equal to its charge as an anion.
6. For a compound, the sum of the oxidation numbers for each of the atoms in the compound is equal to 0.
7. For a polyatomic ion, the sum of the oxidation numbers for each of the atoms in the compound is equal to the ionic charge on the polyatomic ion.

The rules for assigning oxidation numbers can be illustrated by listing some examples. Table 17.2 provides examples of oxidation numbers for elements in the free state, monoatomic ions, and polyatomic ions. Note that each example is referenced to the appropriate rule in Table 17.1.

▲ **Carbon** Diamond and coal are both examples of the element carbon.

TABLE 17.2 Assigning Oxidation Numbers

Example	Oxidation Number	
magnesium metal, Mg	Mg = 0	(Rule 1)
bromine liquid, Br_2	Br = 0	(Rule 1)
potassium ion, K^+	K = +1	(Rule 2)
sulfide ion, S^{2-}	S = −2	(Rule 2)
water, H_2O	H = +1	(Rule 3)
	O = −2	(Rule 4)
carbon tetrachloride, CCl_4	Cl = −1	(Rule 5)
	C = +4	(Rule 6)
barium chloride, $BaCl_2$	Ba = +2	(Rule 6)
	Cl = −1	(Rule 6)
nitrate ion, NO_3^-	O = −2	(Rule 4)
	N = +5	(Rule 7)

Determining Oxidation Numbers in Compounds

Now, let's determine the oxidation number for an element in a compound. Oxalic acid, $H_2C_2O_4$, is used to clean jewelry. To find the oxidation number of carbon in $H_2C_2O_4$, we proceed as follows. First, we assign hydrogen an oxidation number of +1. Second, we assign oxygen, the most electronegative element in the compound, an oxidation number of −2. Because compounds are electrically neutral, the sum of the oxidation numbers (ox no) is equal to 0.

$$2\,(\text{ox no H}) + 2\,(\text{ox no C}) + 4\,(\text{ox no O}) = 0$$

Substituting the oxidation numbers of H and O, we have

$$2\,(+1) + 2\,(\text{ox no C}) + 4\,(-2) = 0$$

Simplifying and solving for the oxidation number of carbon, we have

$$\begin{aligned} +2 + 2\,(\text{ox no C}) + (-8) &= 0 \\ 2\,(\text{ox no C}) &= +6 \\ \text{ox no C} &= +3 \end{aligned}$$

We find that the oxidation number of carbon in $H_2C_2O_4$ is +3. The following example exercise provides additional illustrations of how to determine oxidation numbers for an element in a compound.

EXAMPLE EXERCISE 17.1 Calculating Oxidation Numbers for Carbon

Calculate the oxidation number for carbon in each of the following compounds:

(a) diamond, C
(b) dry ice, CO_2
(c) marble, $CaCO_3$
(d) baking soda, $NaHCO_3$

Solution

Let's begin by recalling that uncombined elements, as well as compounds, are electrically neutral. Thus, free elements and compounds have no charge.

(a) In diamond, the oxidation number of carbon is zero.

(b) In CO_2, we assign oxygen an oxidation number of −2. We can determine the oxidation number of carbon in CO_2 as follows:

$$\begin{aligned} \text{ox no C} + 2\,(\text{ox no O}) &= 0 \\ \text{ox no C} + 2\,(-2) &= 0 \\ \text{ox no C} &= +4 \end{aligned}$$

(c) In $CaCO_3$, we assign calcium ion an oxidation number of +2, and oxygen a value of −2. We can determine the value of carbon in $CaCO_3$ as follows:

$$\begin{aligned} \text{ox no Ca} + \text{ox no C} + 3\,(\text{ox no O}) &= 0 \\ +2 + \text{ox no C} + 3\,(-2) &= 0 \\ \text{ox no C} &= +4 \end{aligned}$$

(d) In $NaHCO_3$, we assign sodium ion an oxidation number of +1, hydrogen a value of +1, and oxygen a value of −2. We can determine the oxidation number of carbon in $NaHCO_3$ as follows:

$$\begin{aligned} \text{ox no Na} + \text{ox no H} + \text{ox no C} + 3\,(\text{ox no O}) &= 0 \\ +1 + +1 + \text{ox no C} + 3\,(-2) &= 0 \\ +1 + +1 + \text{ox no C} + -6 &= 0 \\ \text{ox no C} &= +4 \end{aligned}$$

Practice Exercise

Calculate the oxidation number for iodine in each of the following compounds:

(a) iodine, I_2
(b) potassium iodide, KI
(c) silver periodate, $AgIO_4$
(d) zinc iodate, $Zn(IO_3)_2$

Answers: (a) 0 (b) −1 (c) +7 (d) +5

Concept Exercise

Calculate the oxidation number for nonmetal X in each of the following compounds:

(a) X_2 (b) X_2O (c) CaX_2 (d) HXO_4

Answers: See Appendix G, 17.1.

Determining Oxidation Numbers in Polyatomic Ions

Since many compounds contain polyatomic ions, let's determine the oxidation number of an element in a polyatomic ion. Consider the dichromate ion, $Cr_2O_7^{2-}$, which can be used as a fungicide in a fish aquarium. Unlike compounds, polyatomic ions are not neutral. The dichromate ion has a charge of 2−. To find the oxidation number of chromium in $Cr_2O_7^{2-}$, we assign oxygen an oxidation number of −2. Because $Cr_2O_7^{2-}$ has an overall charge of 2−, we can write the equation

$$2\,(\text{ox no Cr}) + 7\,(\text{ox no O}) = -2$$

Substituting −2 for the oxidation number of O, we have

$$2\,(\text{ox no Cr}) + 7\,(-2) = -2$$

Simplifying and solving for the oxidation number of Cr, we have

$$\begin{aligned} 2\,(\text{ox no Cr}) &= +12 \\ \text{ox no Cr} &= +6 \end{aligned}$$

The following example exercise provides additional practice in calculating the oxidation number for an element in a polyatomic ion.

▲ **Chromium Compounds** The oxidation number of Cr in yellow K_2CrO_4 and orange $K_2Cr_2O_7$ is +6.

EXAMPLE EXERCISE 17.2 Calculating Oxidation Numbers for Sulfur

Calculate the oxidation number for sulfur in each of the following ions:

(a) sulfide ion, S^{2-}
(b) sulfite ion, SO_3^{2-}
(c) sulfate ion, SO_4^{2-}
(d) thiosulfate ion, $S_2O_3^{2-}$

Solution

We can begin by recalling that the charge on an ion corresponds to the sum of the oxidation numbers.

(a) In S^{2-}, the oxidation number of sulfur is −2.

(b) In SO_3^{2-}, the polyatomic anion has a charge of 2−. We assign oxygen an oxidation number of −2 and write the equation

$$\begin{aligned} \text{ox no S} + 3\,(\text{ox no O}) &= -2 \\ \text{ox no S} + 3\,(-2) &= -2 \\ \text{ox no S} &= +4 \end{aligned}$$

(c) In SO_4^{2-}, the polyatomic anion has a charge of 2−. We assign oxygen an oxidation number of −2 and write the equation

$$\begin{aligned} \text{ox no S} + 4\,(\text{ox no O}) &= -2 \\ \text{ox no S} + 4\,(-2) &= -2 \\ \text{ox no S} &= +6 \end{aligned}$$

(d) In $S_2O_3^{2-}$, the polyatomic anion has a charge of 2−. We assign oxygen an oxidation number of −2 and write the equation

$$\begin{aligned} 2\,(\text{ox no S}) + 3\,(\text{ox no O}) &= -2 \\ 2\,(\text{ox no S}) + 3\,(-2) &= -2 \\ 2\,(\text{ox no S}) &= +4 \\ \text{ox no S} &= +2 \end{aligned}$$

Practice Exercise

Calculate the oxidation number for chlorine in each of the following ions:

(a) hypochlorite ion, ClO^-
(b) chlorite ion, ClO_2^-
(c) chlorate ion, ClO_3^-
(d) perchlorate ion, ClO_4^-

Answers: (a) +1 (b) +3 (c) +5 (d) +7

Concept Exercise

Calculate the oxidation number for nonmetal X in each of the following ions:

(a) X^-
(b) XO^-
(c) XO_2^-
(d) XO_3^-

Answers: See Appendix G, 17.2.

17.2 Oxidation–Reduction Reactions

LEARNING OBJECTIVES

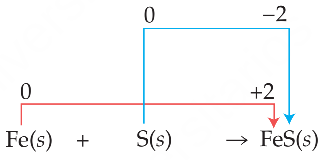
- Identify the oxidized and reduced substances in a given redox reaction.
- Identify the oxidizing and reducing agents in a given redox reaction.

What do burning charcoal and burning magnesium have in common? The answer is that both reactions involve oxygen gas and the transfer of electrons. Charcoal is oxidized to carbon dioxide, and magnesium metal is oxidized to magnesium oxide. As we have said, a chemical change that involves the transfer of electrons is an oxidation–reduction reaction, or simply a **redox reaction**.

As an example, we can heat iron metal with sulfur powder to give the following redox reaction:

$$Fe(s) + S(s) \xrightarrow{\Delta} FeS(s)$$

The oxidation number of S changes from 0 to −2 (see Table 17.2). Simultaneously, Fe changes from 0 to +2. Therefore, each iron atom is oxidized and loses 2 electrons, and each sulfur atom is reduced and gains 2 electrons. We can diagram the process as follows:

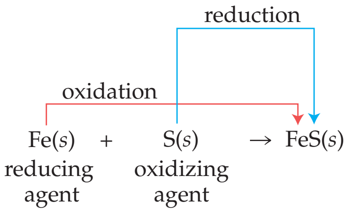


Because electrons are transferred during the chemical change, this is an example of a redox reaction. Notice that both iron and sulfur undergo a change in oxidation number.

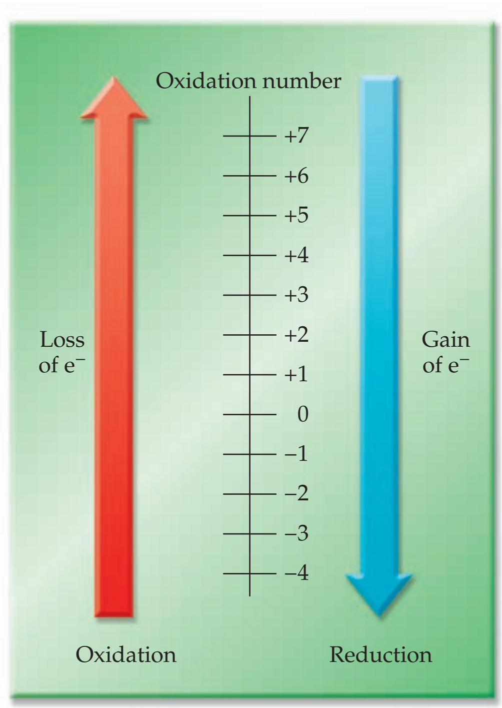
Oxidizing and Reducing Agents

By definition, the process of **oxidation** is characterized by the loss of electrons. Conversely, the process of **reduction** is characterized by the gain of electrons. In the preceding example, iron undergoes oxidation because it loses 2 electrons; sulfur undergoes reduction because it gains 2 electrons.

An **oxidizing agent** is a substance that causes the oxidation in a redox reaction by accepting the transferred electrons. Thus, an oxidizing agent gains electrons. When sulfur reacts with iron, S causes Fe to be oxidized from 0 to +2. Hence, S is the oxidizing agent and gains electrons. A **reducing agent** is any substance that causes reduction by providing the transferred electrons. Thus, a reducing agent loses electrons. When iron reacts with sulfur, Fe causes S to be reduced from 0 to −2. Hence, Fe is the reducing agent. We can illustrate the redox process as follows:



In a redox reaction one substance increases its oxidation number, while another substance decreases its oxidation number. In most instances, the oxidation numbers range from +7 to −4. Figure 17.2 illustrates the relationship of oxidation numbers in a redox reaction.



◀ **Figure 17.2 Oxidation and Reduction** The process of oxidation involves the loss of electrons (the oxidation number increases). The process of reduction involves the gain of electrons (the oxidation number decreases).

The following example exercise further illustrates redox reactions and the designation of oxidizing agents and reducing agents.

EXAMPLE EXERCISE 17.3 Identifying Oxidizing and Reducing Agents

An oxidation–reduction reaction occurs when a stream of hydrogen gas is passed over hot copper(II) sulfide.

$$CuS(s) + H_2(g) \xrightarrow{\Delta} Cu(s) + H_2S(g)$$

Indicate each of the following for the above redox reaction:

(a) substance oxidized

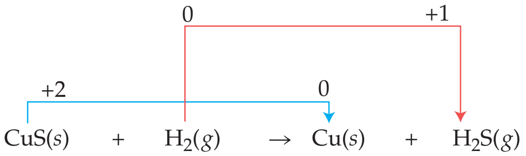
(b) substance reduced

(c) oxidizing agent

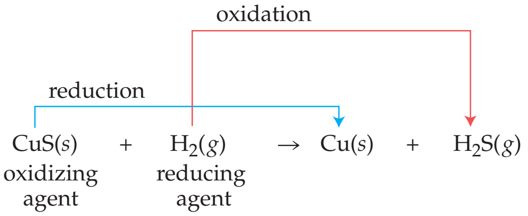
(d) reducing agent

Solution

By definition, the substance oxidized loses electrons, and its oxidation number increases. The substance reduced gains electrons, and its oxidation number decreases. After assigning oxidation numbers to each atom, we have the following:



The oxidation number of hydrogen increases from 0 to +1. Thus, H_2 is oxidized. The oxidation number of copper decreases from +2 to 0. Thus, the Cu in CuS is reduced. Note that the oxidation number of S remains constant (−2).



The oxidizing agent is CuS because it causes hydrogen to be oxidized from 0 to +1. The reducing agent is H_2 because it causes copper to be reduced from +2 to 0.

Practice Exercise

A redox reaction occurs when molten aluminum reacts with iron(III) oxide.

$$Fe_2O_3(l) + 2\,Al(l) \xrightarrow{\Delta} 2\,Fe(l) + Al_2O_3(l)$$

Indicate each of the following for the above redox reaction:

(a) substance oxidized
(b) substance reduced
(c) oxidizing agent
(d) reducing agent

Answers: (a) Al (b) Fe_2O_3 (c) Fe_2O_3 (d) Al

Concept Exercise

Identify the oxidizing agent and reducing agent in the following redox reaction:

$$M(s) + 2\,HX(aq) \longrightarrow MX_2(aq) + H_2(g)$$

Answer: See Appendix G, 17.3.

Ionic Equations in an Aqueous Solution

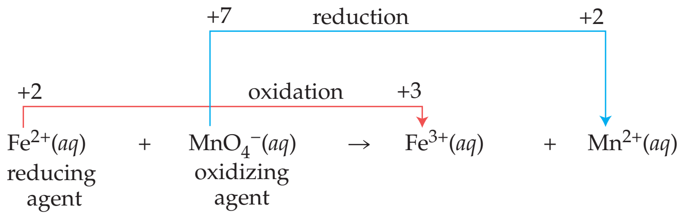
Recall that we learned to write ionic equations in Section 14.11. Redox reactions in an aqueous solution are most often shown in the ionic form. That is because ionic equations readily show us the change in oxidation number. As an example, the amount of iron in a sample is routinely analyzed by a redox method. A typical reaction uses a potassium permanganate solution as follows:

$$5\,Fe^{2+}(aq) + MnO_4^{-}(aq) + 8\,H^{+}(aq) \longrightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l)$$

Although the reaction appears complex, we can systematically study the equation for the oxidized and reduced substances. First, notice from the ionic charges that the oxidation number of iron changes from +2 to +3. Because Fe^{2+} loses an electron, it is oxidized. Because Fe^{2+} is oxidized, it is the reducing agent in this reaction.

Next, let's identify the substance being reduced. When we calculate the oxidation number of manganese in permanganate ion, MnO_4^{-}, we find that the value is +7. On the product side of the equation, we find that the oxidation number of Mn^{2+} is +2. Because permanganate ion gains electrons, it is the substance reduced. Because MnO_4^{-} is reduced, it is the oxidizing agent in this reaction.

We can illustrate the oxidation and reduction processes for the redox reaction as follows:



The following example exercise further illustrates a determination of an oxidizing agent and a reducing agent in a redox reaction.

EXAMPLE EXERCISE 17.4 Identifying Oxidizing and Reducing Agents

The amount of iodine in a solution can be determined by a redox method using a sulfite solution.

$$I_2(s) + SO_3^{2-}(aq) + H_2O(l) \longrightarrow 2\,I^-(aq) + SO_4^{2-}(aq) + 2\,H^+(aq)$$

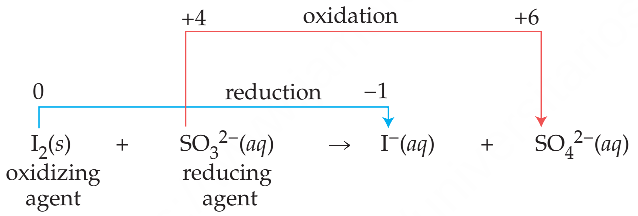
Indicate each of the following for the preceding reaction:

(a) substance oxidized
(b) substance reduced
(c) oxidizing agent
(d) reducing agent

Solution

Notice that iodine solid is converted to iodide ion in an aqueous solution. Since I_2 gains electrons, it is *reduced* and I_2 is the *oxidizing agent.*

The reducing agent is not as obvious. If we calculate the oxidation number for sulfur in SO_3^{2-} and SO_4^{2-}, we find sulfur changes from +4 to +6 and loses electrons. Thus, the sulfur in SO_3^{2-} is *oxidized*, and SO_3^{2-} is the *reducing agent.* We can illustrate the oxidation and reduction processes for the redox reaction as follows:



Practice Exercise

A redox reaction occurs when the tin(II) ion reacts with the iodate ion as follows:

$$6\,H^+(aq) + 3\,Sn^{2+}(aq) + IO_3^-(aq) \longrightarrow 3\,Sn^{4+}(aq) + I^-(aq) + 3\,H_2O(l)$$

Indicate each of the following for the preceding redox reaction:

(a) substance oxidized
(b) substance reduced
(c) oxidizing agent
(d) reducing agent

Answers: (a) Sn^{2+} (b) IO_3^- (c) IO_3^- (d) Sn^{2+}

Concept Exercise

Identify the oxidizing agent and reducing agent in the following redox reaction:

$$Y_2(g) + 2\,X^-(aq) \longrightarrow 2\,Y^-(aq) + X_2(g)$$

Answer: See Appendix G, 17.4.

17.3 Balancing Redox Equations: Oxidation Number Method

LEARNING OBJECTIVE

- Write a balanced chemical equation for a redox reaction using the oxidation number method.

In Section 7.3, we learned to balance chemical equations by inspection. We can also balance simple redox reactions by inspection. It is often difficult, however, to balance complex redox reactions by this method. Therefore, we will use the following general guidelines for balancing redox equations with an *oxidation number method.*

Guidelines for Balancing Redox Equations: Oxidation Number Method

1. Inspect the reactants and products to determine the substances undergoing a change in oxidation number. In some cases, it is not obvious and you must calculate the oxidation number.
  (a) Write the oxidation number above the element that is oxidized and above the element that is reduced.
  (b) Diagram the number of electrons lost by the oxidized substance and gained by the reduced substance.

2. Balance each element in the equation using a coefficient. Keep in mind that the total electrons lost by oxidation must equal the total electrons gained by reduction.
  (a) In front of the oxidized substance, place a coefficient equal to the number of electrons gained by the reduced substance.*
  (b) In front of the reduced substance, place a coefficient equal to the number of electrons lost by the oxidized substance.*
  (c) Balance the remaining elements by inspection.

3. After balancing the equation, verify that the coefficients are correct.
  (a) Place a check mark (√) above the symbol for each element to indicate the number of atoms is the same for reactants and products.
  (b) For ionic equations, verify that the total charge on the left side of the equation equals the total charge on the right side.

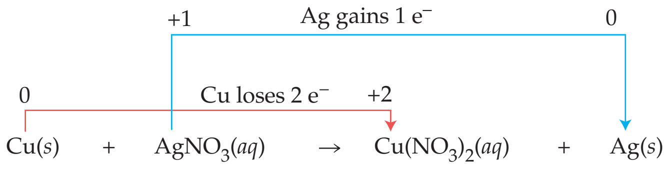
*If the number of electrons gained equals the number of electrons lost, a coefficient is not required.

Balancing Redox Equations by Oxidation Number

Let's apply the guidelines for balancing a redox reaction using the oxidation number method. Specifically, consider the reaction of a metal in an aqueous solution. For example, copper metal reacts with aqueous silver nitrate as follows:

$$\text{Cu}(s) + \text{AgNO}_3(aq) \longrightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag}(s)$$

In this reaction, the oxidation number of copper increases from 0 to +2. Simultaneously, the oxidation number of silver decreases from +1 to 0. Thus, copper loses 2 electrons and silver gains 1 electron. We can diagram the loss and gain of electrons as follows:



Because copper loses 2 electrons and silver gains only 1 electron, two silver atoms are required to balance the transfer of electrons. Therefore, we place the coefficient 2 in front of each silver substance. The balanced equation for the reaction is

$$\text{Cu}(s) + 2\,\text{AgNO}_3(aq) \longrightarrow \text{Cu(NO}_3)_2(aq) + 2\,\text{Ag}(s)$$

After placing the coefficients, we verify that the equation is balanced by checking (√) each element in the reaction. Because the nitrate ion did not change, we can check off the polyatomic ion as a single unit.

$$\overset{\surd}{\text{Cu}}(s) + 2\,\overset{\surd}{\text{Ag}}\overset{\surd}{\text{NO}_3}(aq) \longrightarrow \overset{\surd}{\text{Cu}}(\overset{\surd}{\text{NO}_3})_2(aq) + 2\,\overset{\surd}{\text{Ag}}(s)$$

Because all the elements are balanced, we have verified that this is a balanced equation. The following example exercise provides additional practice in balancing redox equations.

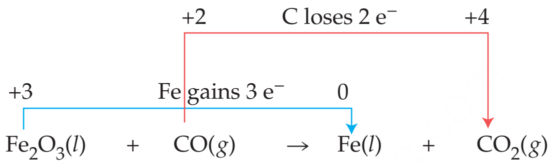
EXAMPLE EXERCISE 17.5 Balancing Redox Equations by Oxidation Number

An industrial blast furnace reduces iron ore, Fe_2O_3, to molten iron. Balance the following redox equation using the oxidation number method:

$$Fe_2O_3(l) + CO(g) \longrightarrow Fe(l) + CO_2(g)$$

Solution

In this reaction, the oxidation number of iron decreases from +3 in Fe_2O_3 to 0 in Fe. Simultaneously, the oxidation number of carbon increases from +2 to +4. Thus, each Fe gains 3 electrons while each C loses 2 electrons. We can diagram the redox process as follows:



Because the number of electrons gained and lost must be equal, we find the lowest common multiple. In this case, it is 6. Each Fe gains 3 electrons, so we place the coefficient 3 in front of CO and CO_2.

$$Fe_2O_3(l) + 3\,CO(g) \longrightarrow Fe(l) + 3\,CO_2(g)$$

Each carbon atom loses 2 electrons, and so we place the coefficient 2 in front of each iron atom. Because Fe_2O_3 has two iron atoms, it does not require a coefficient.

$$Fe_2O_3(l) + 3\,CO(g) \longrightarrow 2\,Fe(l) + 3\,CO_2(g)$$

Finally, we verify that the equation is balanced. We check (√) each element in the equation:

$$\overset{\surd}{Fe_2}\overset{\surd}{O_3}(l) + 3\,\overset{\surd}{C}\overset{\surd}{O}(g) \longrightarrow 2\,\overset{\surd}{Fe}(l) + 3\,\overset{\surd}{C}\overset{\surd}{O_2}(g)$$

Because all the elements are balanced, we have a balanced redox equation.

Practice Exercise

Balance the following redox equation by the oxidation number method:

$$Cl_2O_5(g) + CO(g) \longrightarrow Cl_2(g) + CO_2(g)$$

Answer: $Cl_2O_5(g) + 5\,CO(g) \longrightarrow Cl_2(g) + 5\,CO_2(g)$

Concept Exercise

Balance the following redox reaction using the oxidation number method:

$$M_2O_3(s) + X_2(g) \longrightarrow MX_3(aq) + O_2(g)$$

Answer: See Appendix G, 17.5.

Balancing Ionic Redox Equations by Oxidation Number

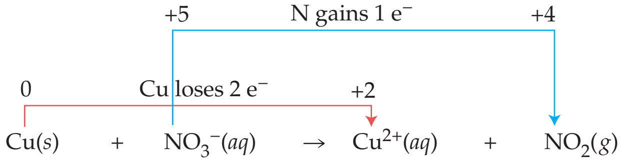
Now, let's apply the general guidelines for balancing a redox reaction to an ionic equation. Figure 17.3 shows a copper penny reacting with nitric acid, according to the following equation:

$$Cu(s) + H^+(aq) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g) + H_2O(l)$$

In this ionic equation, we see that the oxidation number of copper increases from 0 to +2. Although it is not as evident, the oxidation number of nitrogen decreases from

▲ **Figure 17.3 A Penny in Nitric Acid** The copper penny reacts with nitric acid, HNO_3, to give a green solution and brown NO_2 gas.

+5 to +4. Thus, Cu loses 2 electrons while N gains 1 electron. We can diagram the loss and gain of electrons as follows:



Helpful Hint
Balance Mass and Charge

For redox reactions, it is necessary to verify both the mass and the charge. That is, there must be the same number of atoms of each element in the reactants and products, and the total charge on the reactants must equal the total charge on the products.

We can balance the loss and gain of electrons by the coefficient 2 in front of each nitrogen substance. It is understood that the coefficient of each copper is 1.

$$\mathrm{Cu}(s) + \mathrm{H^+}(aq) + 2\,\mathrm{NO_3^-}(aq) \longrightarrow \mathrm{Cu^{2+}}(aq) + 2\,\mathrm{NO_2}(g) + \mathrm{H_2O}(l)$$

In this example, H^+ and H_2O take part in the reaction. Thus, they also must be balanced. We complete balancing the equation by inspection. Because there is a total of six oxygen atoms as reactants, there must be six oxygen atoms in the products. Therefore, we need the coefficient 2 in front of H_2O. After balancing the remaining hydrogen atoms in the reaction, we have

$$\mathrm{Cu}(s) + 4\,\mathrm{H^+}(aq) + 2\,\mathrm{NO_3^-}(aq) \longrightarrow \mathrm{Cu^{2+}}(aq) + 2\,\mathrm{NO_2}(g) + 2\,\mathrm{H_2O}(l)$$

Now, we verify that the equation is balanced by checking (√) each element in the redox reaction.

$$\overset{\surd}{\mathrm{Cu}}(s) + 4\,\overset{\surd}{\mathrm{H}}{}^+(aq) + 2\,\overset{\surd}{\mathrm{N}}\overset{\surd}{\mathrm{O}}{}_3^-(aq) \longrightarrow \overset{\surd}{\mathrm{Cu}}{}^{2+}(aq) + 2\,\overset{\surd}{\mathrm{N}}\overset{\surd}{\mathrm{O}}{}_2(g) + 2\,\overset{\surd}{\mathrm{H}}{}_2\overset{\surd}{\mathrm{O}}(l)$$

Finally, we check to verify that the total charge on the reactants equals the total charge on the products. On the left side of the equation, we have +4−2 = +2. On the right side of the equation, we have +2. Because the ionic charge on each side is +2, the equation is balanced. The following example exercise provides additional practice in balancing ionic redox equations.

EXAMPLE EXERCISE 17.6 Balancing Redox Equations by Oxidation Number

Aqueous sodium iodide reacts with a potassium dichromate solution. Write a balanced equation for the following redox reaction:

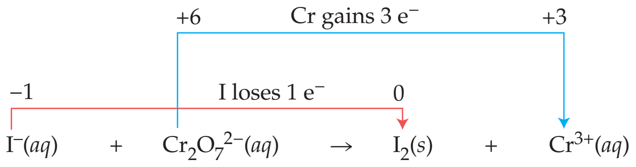
$$\mathrm{H^+}(aq) + \mathrm{I^-}(aq) + \mathrm{Cr_2O_7^{2-}}(aq) \longrightarrow \mathrm{I_2}(s) + \mathrm{Cr^{3+}}(aq) + \mathrm{H_2O}(l)$$

Solution

In this special example, $Cr_2O_7^{2-}$ and I_2 each contain a subscript that affects electron transfer. Note that there are 2 atoms of chromium in the reactant and 2 atoms of iodine in the product. Let's balance the chromium and iodine atoms first.

$$\mathrm{H^+}(aq) + 2\,\mathrm{I^-}(aq) + \mathrm{Cr_2O_7^{2-}}(aq) \longrightarrow \mathrm{I_2}(s) + 2\,\mathrm{Cr^{3+}}(aq) + \mathrm{H_2O}(l)$$

In this reaction, the oxidation number of each iodine atom increases from −1 to 0, and the oxidation number of each chromium atom decreases from +6 to +3. We can show the loss and gain of electrons as



There are 2 chromium atoms, and so the total electron gain is 6 electrons. Thus, the total electron loss must also be 6 electrons. Because there are 2 iodine atoms, and 1 iodide ion, we place the coefficients as follows:

$$H^+(aq) + 6\,I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow 3\,I_2(s) + 2\,Cr^{3+}(aq) + H_2O(l)$$

Next, we balance the oxygen and hydrogen atoms. Because there are 7 oxygen atoms as reactants, we place the coefficient 7 in front of H_2O. This gives 14 hydrogen atoms, and so we place the coefficient 14 in front of H^+.

$$14\,H^+(aq) + 6\,I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow 3\,I_2(s) + 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

We can verify that the equation is balanced by checking (√) each element.

$$14\,\overset{\surd}{H}{}^+(aq) + 6\,\overset{\surd}{I}{}^-(aq) + \overset{\surd}{Cr}_2\overset{\surd}{O}_7{}^{2-}(aq) \longrightarrow 3\,\overset{\surd}{I}_2(s) + 2\,\overset{\surd}{Cr}{}^{3+}(aq) + 7\,\overset{\surd}{H}_2\overset{\surd}{O}(l)$$

Last, we verify that the ionic charges are balanced. On the left side of the equation, we have $+14 - 6 - 2 = +6$. On the right side of the equation, we have $2\,(+3) = +6$. Because the ionic charge on each side is +6, the equation is balanced.

Practice Exercise

Write a balanced equation for the following redox reaction:

$$H^+(aq) + MnO_4^-(aq) + NO_2^-(aq) \longrightarrow Mn^{2+}(aq) + NO_3^-(aq) + H_2O(l)$$

Answer: $6\,H^+(aq) + 2\,MnO_4^-(aq) + 5\,NO_2^-(aq) \longrightarrow 2\,Mn^{2+}(aq) + 5\,NO_3^-(aq) + 3\,H_2O(l)$

Concept Exercise

Balance the following redox reaction in an acidic solution using the oxidation number method:

$$MnO_4^-(aq) + SO_3^{2-}(aq) \longrightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$$

Answer: See Appendix G, 17.6.

Helpful Hint
Balancing Redox Equations

In this chapter we have learned to balance redox equations by noting the decrease in oxidation number for the oxidizing agent, and the simultaneous increase in oxidation number for the reducing agent. This method is effective for redox reactions that occur in neutral or acidic solution.

A second method for balancing redox equations is to balance the oxidation and reduction half-reactions separately. First, balance the mass in each half-reaction, and then balance the charge in each half-reaction using electrons. This method is effective for balancing redox reactions that occur in acidic, neutral, or basic solutions.

17.4 Balancing Redox Equations: Half-Reaction Method

LEARNING OBJECTIVES

- Explain the concept of a half-reaction.
- Write a balanced chemical equation for a redox reaction using the half-reaction method:
  (a) in an acidic solution.
  (b) in a basic solution.

In the preceding section, we learned how to balance redox equations using oxidation numbers. In this section, we are going to balance redox equations with a different procedure. This alternative method is referred to as the *half-reaction method*.

A **half-reaction** shows the oxidation and reduction processes separately. The half-reaction method systematically balances the oxidation half-reaction and then balances the reduction half-reaction. We will use the following general guidelines for balancing redox equations with the half-reaction method.

Guidelines for Balancing Redox Equations: Half-Reaction Method

1. Write the half-reaction for both oxidation and reduction.
   (a) Identify the reactant that is oxidized and its product.
   (b) Identify the reactant that is reduced and its product.
2. Balance the atoms in each half-reaction using coefficients.
   (a) Balance all elements except oxygen and hydrogen.
   (b) Balance oxygen using H_2O.
   (c) Balance hydrogen using H^+.
   (**Note:** For reactions in a basic solution, add OH^- to neutralize the added H^+. For example, $2\,OH^-$ neutralizes $2\,H^+$ to give $2\,H_2O$.)
   (d) Balance the ionic charges using e^-.

3. Multiply each half-reaction by a whole number so the number of electrons lost by oxidation equals the number of electrons gained by reduction.
4. Add the two half-reactions together and cancel identical species, including electrons, on each side of the equation.
5. After balancing the equation, verify the coefficients are correct; and verify the total charge on the left side of the equation equals the total charge on the right side.

Helpful Hint
Balancing Half-Reactions

- Electrons are added to the *product* side of *the oxidation* half-reaction.
- Electrons are added to the *reactant* side of *the reduction* half-reaction.

Balancing Redox Equations in an Acidic Solution

Let's apply the general guidelines for balancing a redox reaction in an acidic aqueous solution. As an example, zinc reacts in nitric acid to give nitrogen monoxide gas. The ionic equation for the redox reaction is

$$\mathrm{Zn}(s) + \mathrm{NO_3}^-(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{NO}(g)$$

STEP 1: Write the oxidation and reduction half-reactions. Because the oxidation number of zinc increases, Zn is oxidized. Because the oxidation number of nitrogen decreases, $\mathrm{NO_3}^-$ is reduced.

$$\begin{array}{lrcl} \text{Oxidation:} & \mathrm{Zn} & \longrightarrow & \mathrm{Zn}^{2+} \\ \text{Reduction:} & \mathrm{NO_3}^- & \longrightarrow & \mathrm{NO} \end{array}$$

STEP 2: Balance the atoms and the charges for each half-reaction. The balanced half-reactions are

$$\begin{array}{rcl} \mathrm{Zn} & \longrightarrow & \mathrm{Zn}^{2+} + 2\,e^- \\ 3\,e^- + 4\,\mathrm{H}^+ + \mathrm{NO_3}^- & \longrightarrow & \mathrm{NO} + 2\,\mathrm{H_2O} \end{array}$$

STEP 3: Note that Zn loses 2 electrons and $\mathrm{NO_3}^-$ gains 3 electrons. So that the electron loss and gain will be equal, we multiply the Zn half-reaction by 3 and the $\mathrm{NO_3}^-$ half-reaction by 2.

$$\begin{array}{rcl} 3\,\mathrm{Zn} & \longrightarrow & 3\,\mathrm{Zn}^{2+} + 6\,e^- \\ 6\,e^- + 8\,\mathrm{H}^+ + 2\,\mathrm{NO_3}^- & \longrightarrow & 2\,\mathrm{NO} + 4\,\mathrm{H_2O} \end{array}$$

STEP 4: Now, we add the two half-reactions together and cancel the 6 electrons.

$$\begin{array}{rcl} 3\,\mathrm{Zn} & \longrightarrow & 3\,\mathrm{Zn}^{2+} + \cancel{6}\,e^- \\ \cancel{6}\,e^- + 8\,\mathrm{H}^+ + 2\,\mathrm{NO_3}^- & \longrightarrow & 2\,\mathrm{NO} + 4\,\mathrm{H_2O} \\ \hline 3\,\mathrm{Zn} + 8\,\mathrm{H}^+ + 2\,\mathrm{NO_3}^- & \longrightarrow & 3\,\mathrm{Zn}^{2+} + 2\,\mathrm{NO} + 4\,\mathrm{H_2O} \end{array}$$

STEP 5: Finally, let's check the atoms and ionic charges to verify that the equation is balanced. We have

$$\begin{array}{ll} \text{Atoms:} & 3\,\mathrm{Zn}, 8\,\mathrm{H}, 2\,\mathrm{N}, 6\,\mathrm{O} = 3\,\mathrm{Zn}, 2\,\mathrm{N}, 6\,\mathrm{O}, 8\,\mathrm{H} \\ \text{Charges:} & 8\,(+1) + 2\,(-1) = 3\,(+2) \\ & +6 = +6 \end{array}$$

Because the atoms and ionic charges are equal, the redox equation is balanced.

Balancing Redox Equations in a Basic Solution

Now, let's balance a redox reaction that takes place in a basic aqueous solution. For example, sodium iodide and potassium permanganate react in a basic solution to produce iodine and manganese(IV) oxide. The ionic equation is

$$\mathrm{I}^-(aq) + \mathrm{MnO_4}^-(aq) \longrightarrow \mathrm{I_2}(s) + \mathrm{MnO_2}(s)$$

The procedure for balancing a redox equation in a base is similar to the procedure given in the general guidelines for balancing a redox reaction in an acid.

STEP 1: First, we write the oxidation and reduction half-reactions. We note that iodine is oxidized from −1 to 0. Simultaneously, manganese is reduced from +7 to +4.

$$\begin{array}{lrcl} \text{Oxidation:} & I^- & \longrightarrow & I_2 \\ \text{Reduction:} & MnO_4^- & \longrightarrow & MnO_2 \end{array}$$

STEP 2: Second, we balance each half-reaction for both atoms and charges. In the oxidation half-reaction, we place the coefficient 2 in front of I^- and add 2 electrons to the product side. We balance the reduction half reaction as follows: add H_2O to balance O atoms, add H^+ to balance H atoms, and add 3 e^- to balance charge. The balanced half-reactions are

$$\begin{array}{rcl} 2\,I^- & \longrightarrow & I_2 + 2\,e^- \\ 3\,e^- + 4\,H^+ + MnO_4^- & \longrightarrow & MnO_2 + 2\,H_2O \end{array}$$

Note that this redox reaction takes place in a basic solution, not in an acidic solution. Therefore, we must follow guideline 2(c), which states that the H^+ should be neutralized with OH^-. In this example, we add 4 OH^- to each side of the reduction half-reaction and simplify the equation.

$$\begin{array}{lcl} 3\,e^- + 4\,H^+ + 4\,OH^- + MnO_4^- & \longrightarrow & MnO_2 + 2\,H_2O + 4\,OH^- \\ 3\,e^- + 4\,H_2O + MnO_4^- & \longrightarrow & MnO_2 + 2\,H_2O + 4\,OH^- \\ 3\,e^- + 2\,H_2O + MnO_4^- & \longrightarrow & MnO_2 + 4\,OH^- \end{array}$$

STEP 3: Note that 2 I^- lose 2 electrons and MnO_4^- gains 3 electrons. To equalize electron loss and gain, we multiply the I^- oxidation half-reaction by 3, and the MnO_4^- reduction half-reaction by 2.

$$\begin{array}{rcl} 6\,I^- & \longrightarrow & 3\,I_2 + 6\,e^- \\ 6\,e^- + 4\,H_2O + 2\,MnO_4^- & \longrightarrow & 2\,MnO_2 + 8\,OH^- \end{array}$$

STEP 4: We add the two half-reactions together and cancel the 6 electrons.

$$\begin{array}{rcl} 6\,I^- & \longrightarrow & 3\,I_2 + \cancel{6}e^- \\ \cancel{6}e^- + 4\,H_2O + 2\,MnO_4^- & \longrightarrow & 2\,MnO_2 + 8\,OH^- \\ \hline 6\,I^- + 4\,H_2O + 2\,MnO_4^- & \longrightarrow & 3\,I_2 + 2\,MnO_2 + 8\,OH^- \end{array}$$

STEP 5: Finally, we check to verify that the equation is balanced.

Atoms: 6 I, 8 H, 12 O, 2 Mn = 6 I, 2 Mn, 12 O, 8 H

Charges: $6\,(-1) + 2\,(-1) = 8\,(-1)$

$-8 = -8$

Because the atoms and ionic charges are equal for the reactants and products, this redox equation is balanced. Example Exercise 17.7 illustrates balancing redox reactions by the half-reaction method.

EXAMPLE EXERCISE 17.7 Balancing Redox Equations by Half-Reaction

Write a balanced ionic equation for the reaction of iron(II) sulfate and potassium permanganate in acidic solution. The ionic equation is

$$MnO_4^-(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + Mn^{2+}(aq)$$

Solution

We can balance the redox reaction by the half-reaction method as follows:

STEP 1: Because Fe^{2+} is oxidized from +2 to +3, MnO_4^- must be reduced. The two half-reactions are as follows:

$$\begin{array}{ll} \text{Oxidation:} & Fe^{2+} \longrightarrow Fe^{3+} \\ \text{Reduction:} & MnO_4^- \longrightarrow Mn^{2+} \end{array}$$

STEP 2: We can balance each half-reaction as follows:

$$\begin{array}{rcl} Fe^{2+} & \longrightarrow & Fe^{3+} + e^- \\ 5\,e^- + 8\,H^+ + MnO_4^- & \longrightarrow & Mn^{2+} + 4\,H_2O \end{array}$$

STEP 3: Because Fe^{2+} loses 1 e^- and MnO_4^- gains 5 e^-, we multiply the Fe^{2+} half-reaction by 5.

$$\begin{array}{rcl} 5\,Fe^{2+} & \longrightarrow & 5\,Fe^{3+} + 5\,e^- \\ 5\,e^- + 8\,H^+ + MnO_4^- & \longrightarrow & Mn^{2+} + 4\,H_2O \end{array}$$

STEP 4: Then, we add the two half-reactions together and cancel the 5 e^-.

$$\begin{array}{rcl} 5\,Fe^{2+} & \longrightarrow & 5\,Fe^{3+} + \cancel{5\,e^-} \\ \cancel{5\,e^-} + 8\,H^+ + MnO_4^- & \longrightarrow & Mn^{2+} + 4\,H_2O \\ \hline 5\,Fe^{2+} + 8\,H^+ + MnO_4^- & \longrightarrow & 5\,Fe^{3+} + Mn^{2+} + 4\,H_2O \end{array}$$

STEP 5: Finally, let's check the atoms and ionic charges to verify that the equation is balanced. We have

$$\begin{array}{lrcl} \text{Atoms:} & 5\,\text{Fe}, 8\,\text{H}, 1\,\text{Mn}, 4\,\text{O} & = & 5\,\text{Fe}, 1\,\text{Mn}, 8\,\text{H}, 4\,\text{O} \\ \text{Charges:} & 5\,(+2) + 8\,(+1) + (-1) & = & 5\,(+3) + (+2) \\ & +17 & = & +17 \end{array}$$

Because the atoms and ionic charges are equal for reactants and products, the redox equation is balanced.

Practice Exercise

Write a balanced ionic equation for the reaction of sodium nitrite and potassium permanganate in acidic solution. The ionic equation is

$$MnO_4^-(aq) + NO_2^-(aq) \longrightarrow Mn^{2+}(aq) + NO_3^-(aq)$$

Answer: $6\,H^+ + 2\,MnO_4^- + 5\,NO_2^- \longrightarrow 2\,Mn^{2+} + 5\,NO_3^- + 3\,H_2O$

Concept Exercise

Balance the following redox reaction in a basic solution using the half-reaction method.

$$MnO_4^-(aq) + SO_3^{2-}(aq) \longrightarrow MnO_2(s) + SO_4^{2-}(aq)$$

Answer: See Appendix G, 17.7.

17.5 Predicting Spontaneous Redox Reactions

LEARNING OBJECTIVES

- Predict the stronger oxidizing agent and reducing agent given a list of reduction potentials.
- Predict whether a redox reaction is spontaneous or nonspontaneous given a list of reduction potentials.

A redox reaction that occurs without any input of electrical energy is referred to as a *spontaneous reaction*. Let's consider the spontaneous reaction of a metal reacting in an aqueous solution. For example, consider the reaction of zinc metal in aqueous copper(II) sulfate. The net ionic equation for the reaction is

$$\mathrm{Zn}(s) + \mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Cu}(s)$$

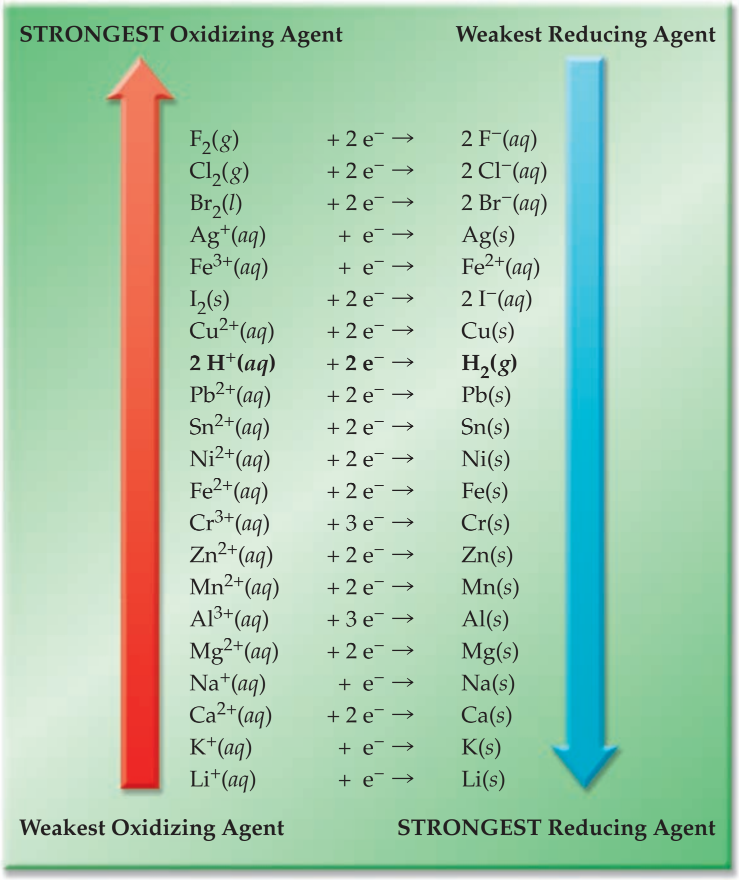
It is experimentally observed that the reaction of zinc metal in aqueous copper(II) sulfate proceeds spontaneously. Because copper and zinc ions can both be reduced, let's compare the two reduction half-reactions of copper and zinc:

$$\mathrm{Cu}^{2+}(aq) + 2\,\mathrm{e}^- \longrightarrow \mathrm{Cu}(s)$$
$$\mathrm{Zn}^{2+}(aq) + 2\,\mathrm{e}^- \longrightarrow \mathrm{Zn}(s)$$

According to the overall redox reaction of zinc in copper(II) sulfate solution, $\mathrm{Zn}(s)$ is oxidized and $\mathrm{Cu}^{2+}(aq)$ is reduced. Therefore, we can conclude that the reduction of $\mathrm{Cu}^{2+}(aq)$ has a greater tendency to occur than the reduction of $\mathrm{Zn}^{2+}(aq)$.

Stated differently, Cu^{2+} has a greater tendency to gain electrons than Zn^{2+}. If we compare other combinations of metals and aqueous solutions, we can arrange a series of elements based on their ability to gain electrons. The strongest oxidizing agent is the substance that is most easily reduced, that is, the substance that has the greater tendency to gain electrons. Conversely, the strongest reducing agent is the substance that is most easily oxidized and has the greater tendency to lose electrons.

The tendency for a substance to gain electrons and undergo reduction is called its **reduction potential**. The reduction potential for a substance is its relative strength as an oxidizing agent, that is, its relative strength to pull electrons to itself. Figure 17.4 shows the reduction potentials for several substances in order of their ability to undergo reduction. A substance that is high in the table has a strong reduction potential and is



STRONGEST Oxidizing Agent **Weakest Reducing Agent**

$\mathrm{F_2}(g)$	$+ 2\,\mathrm{e}^- \rightarrow$	$2\,\mathrm{F}^-(aq)$
$\mathrm{Cl_2}(g)$	$+ 2\,\mathrm{e}^- \rightarrow$	$2\,\mathrm{Cl}^-(aq)$
$\mathrm{Br_2}(l)$	$+ 2\,\mathrm{e}^- \rightarrow$	$2\,\mathrm{Br}^-(aq)$
$\mathrm{Ag}^+(aq)$	$+ \mathrm{e}^- \rightarrow$	$\mathrm{Ag}(s)$
$\mathrm{Fe}^{3+}(aq)$	$+ \mathrm{e}^- \rightarrow$	$\mathrm{Fe}^{2+}(aq)$
$\mathrm{I_2}(s)$	$+ 2\,\mathrm{e}^- \rightarrow$	$2\,\mathrm{I}^-(aq)$
$\mathrm{Cu}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Cu}(s)$
$2\,\mathrm{H}^+(aq)$	**$+ 2\,\mathrm{e}^- \rightarrow$**	**$\mathrm{H_2}(g)$**
$\mathrm{Pb}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Pb}(s)$
$\mathrm{Sn}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Sn}(s)$
$\mathrm{Ni}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Ni}(s)$
$\mathrm{Fe}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Fe}(s)$
$\mathrm{Cr}^{3+}(aq)$	$+ 3\,\mathrm{e}^- \rightarrow$	$\mathrm{Cr}(s)$
$\mathrm{Zn}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Zn}(s)$
$\mathrm{Mn}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Mn}(s)$
$\mathrm{Al}^{3+}(aq)$	$+ 3\,\mathrm{e}^- \rightarrow$	$\mathrm{Al}(s)$
$\mathrm{Mg}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Mg}(s)$
$\mathrm{Na}^+(aq)$	$+ \mathrm{e}^- \rightarrow$	$\mathrm{Na}(s)$
$\mathrm{Ca}^{2+}(aq)$	$+ 2\,\mathrm{e}^- \rightarrow$	$\mathrm{Ca}(s)$
$\mathrm{K}^+(aq)$	$+ \mathrm{e}^- \rightarrow$	$\mathrm{K}(s)$
$\mathrm{Li}^+(aq)$	$+ \mathrm{e}^- \rightarrow$	$\mathrm{Li}(s)$

Weakest Oxidizing Agent **STRONGEST Reducing Agent**

◀ **Figure 17.4 Oxidizing Agents and Reducing Agents** The strongest oxidizing agent is $\mathrm{F_2}$ (fluorine gas has the greatest tendency to gain electrons). The strongest reducing agent is Li (lithium metal has the greatest tendency to lose electrons).

easily reduced because it powerfully attracts electrons. A substance low in the table has a weak reduction potential and is not easily reduced because it does not powerfully attract electrons.

Spontaneous Reactions

The ability of a substance to be reduced, that is to attract electrons, decreases from top to bottom in the table. The strongest oxidizing agents are in the upper-left portion of the table. Thus, F_2 has the highest potential for reduction and is the strongest oxidizing agent. The strongest reducing agents are in the lower-right portion of the table. Thus, Li has the highest potential for oxidation and is the strongest reducing agent.

A metal ion can be reduced, that is gain electrons, by any reducing agent listed lower in the table. For example, Sn^{2+} can be reduced by Ni or Fe, but not by Pb. Any metal below H_2 in the table can react with an acid and be oxidized. A metal above H^+ in the table cannot react with acid. For example, Pb, Sn, and Ni react with acid and undergo oxidation, whereas Cu and Ag do not react with acid.

A nonmetal can oxidize any reducing agent listed lower in the table. For example, Cl_2 can oxidize Br^- or Ag, but not F^-. Figure 17.4 shows that F_2 is the strongest oxidizing agent and that Li is the strongest reducing agent. However, H_2 is the strongest *nonmetal* reducing agent listed in the table.

In Figure 17.4, we note that $Cu^{2+}(aq)$ is a stronger oxidizing agent than $Zn^{2+}(aq)$. We also note that Zn metal is a stronger reducing agent than Cu metal. Therefore, we can explain the reaction as follows:

$Cu^{2+}(aq)$	$+$	$Zn(s)$	$\longrightarrow$	$Cu(s)$	$+$	$Zn^{2+}(aq)$
stronger		stronger		weaker		weaker
oxidizing agent		reducing agent		reducing agent		oxidizing agent

Moreover, we see that the reverse process has very little tendency to occur and is nonspontaneous. We can therefore conclude that there is no reaction between aqueous zinc ion and copper metal.

$Zn^{2+}(aq)$	$+$	$Cu(s)$	$\longrightarrow$	$Zn(s)$	$+$	$Cu^{2+}(aq)$
weaker		weaker		stronger		stronger
oxidizing agent		reducing agent		reducing agent		oxidizing agent

Now, let's consider a nonmetal reaction in an aqueous solution. For example, consider what happens if we bubble chlorine gas through an aqueous sodium bromide solution. The net ionic equation for the possible reaction is

$$Cl_2(g) + 2\,Br^-(aq) \longrightarrow 2\,Cl^-(aq) + Br_2(l)$$

In Figure 17.4, we see that $Cl_2(g)$ is a stronger oxidizing agent than $Br_2(l)$. We also see that $Br^-(aq)$ is a stronger reducing agent than $Cl^-(aq)$. Therefore, we can explain the reaction as follows:

$Cl_2(g)$	$+$	$2\,Br^-(aq)$	$\longrightarrow$	$2\,Cl^-(aq)$	$+$	$Br_2(l)$
stronger		stronger		weaker		weaker
oxidizing agent		reducing agent		reducing agent		oxidizing agent

Because the reactants contain the stronger pair of oxidizing and reducing agents, the reaction takes place spontaneously. Is the reverse reaction spontaneous? Because the chloride ion and bromine liquid are the weaker pair of oxidizing and reducing agents, the reverse reaction is nonspontaneous. If a redox reaction is not spontaneous, we usually say that there is no reaction.

$2\,Cl^-(aq)$	$+$	$Br_2(l)$	$\longrightarrow$	$Cl_2(g)$	$+$	$2\,Br^-(aq)$
weaker		weaker		stronger		stronger
reducing agent		oxidizing agent		oxidizing agent		reducing agent

The following example exercise provides additional practice in predicting whether or not a redox reaction occurs spontaneously.

EXAMPLE EXERCISE 17.8 Predicting Spontaneous Redox Reactions

Predict whether the following reaction is spontaneous or nonspontaneous.

$$Ni^{2+}(aq) + Sn(s) \longrightarrow Ni(s) + Sn^{2+}(aq)$$

Solution

Let's refer to the table of reduction potentials to predict whether or not the reaction is spontaneous. Figure 17.4 lists $Ni^{2+}(aq)$ as a weaker oxidizing agent than $Sn^{2+}(aq)$. Moreover, $Sn(s)$ is a weaker reducing agent than $Ni(s)$.

$$\underset{\text{oxidizing agent}}{\underset{\text{weaker}}{Ni^{2+}(aq)}} + \underset{\text{reducing agent}}{\underset{\text{weaker}}{Sn(s)}} \longrightarrow \underset{\text{reducing agent}}{\underset{\text{stronger}}{Ni(s)}} + \underset{\text{oxidizing agent}}{\underset{\text{stronger}}{Sn^{2+}(aq)}}$$

Because the reactants are the weaker pair of oxidizing and reducing agents, the reaction is *nonspontaneous.* Conversely, the reverse reaction is spontaneous because the products are the stronger oxidizing and reducing agents.

Practice Exercise

Predict whether the following reaction is spontaneous or nonspontaneous.

$$Ni^{2+}(aq) + Al(s) \longrightarrow Ni(s) + Al^{3+}(aq)$$

Answer: spontaneous

Concept Exercise

Refer to Figure 17.4 and predict which of the following metals reacts spontaneously in an aqueous $FeSO_4$ solution.

(a) Ag
(b) Al
(c) Cr
(d) Ni

Answer: See Appendix G, 17.8.

17.6 Voltaic Cells

Batteries play an important role in our daily lives. They allow computer laptops to be portable, they start our automobiles, and they provide cordless electrical appliances. They extend the lives of heart patients who receive rhythmic pulses from pacemakers. All these devices use electrical energy supplied by batteries. Although different devices may use different types of batteries, all batteries involve a spontaneous redox reaction that produces electrical energy, and a flow of electrons.

LEARNING OBJECTIVES

- Identify the anode and cathode in a given voltaic cell.
- Identify the oxidation and reduction half-reactions in a given spontaneous electrochemical cell.

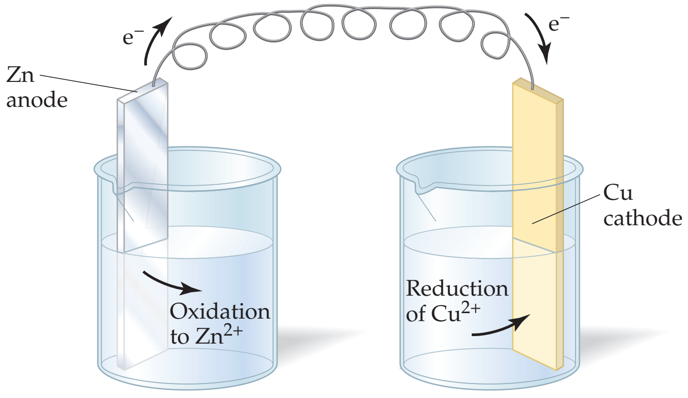
Redox Reactions in Voltaic Cells

The conversion of chemical energy to electrical energy, and vice versa, in a redox reaction is called **electrochemistry**. As an example, consider the reaction of zinc metal in a copper(II) sulfate solution:

$$Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$$

Experimentally, we can physically separate the oxidation half-reaction from the reduction half-reaction. We do this by placing aqueous solutions of zinc sulfate and copper(II) sulfate in separate containers or compartments. Then, we place a Zn metal electrode in the first compartment containing an aqueous solution of $ZnSO_4$, and a Cu metal electrode in the second compartment containing an aqueous solution of $CuSO_4$.

The two electrodes are connected by a wire, so electrons are free to travel between the two compartments. Each compartment is called a **half-cell**, and together the apparatus is called an **electrochemical cell** (Figure 17.5).

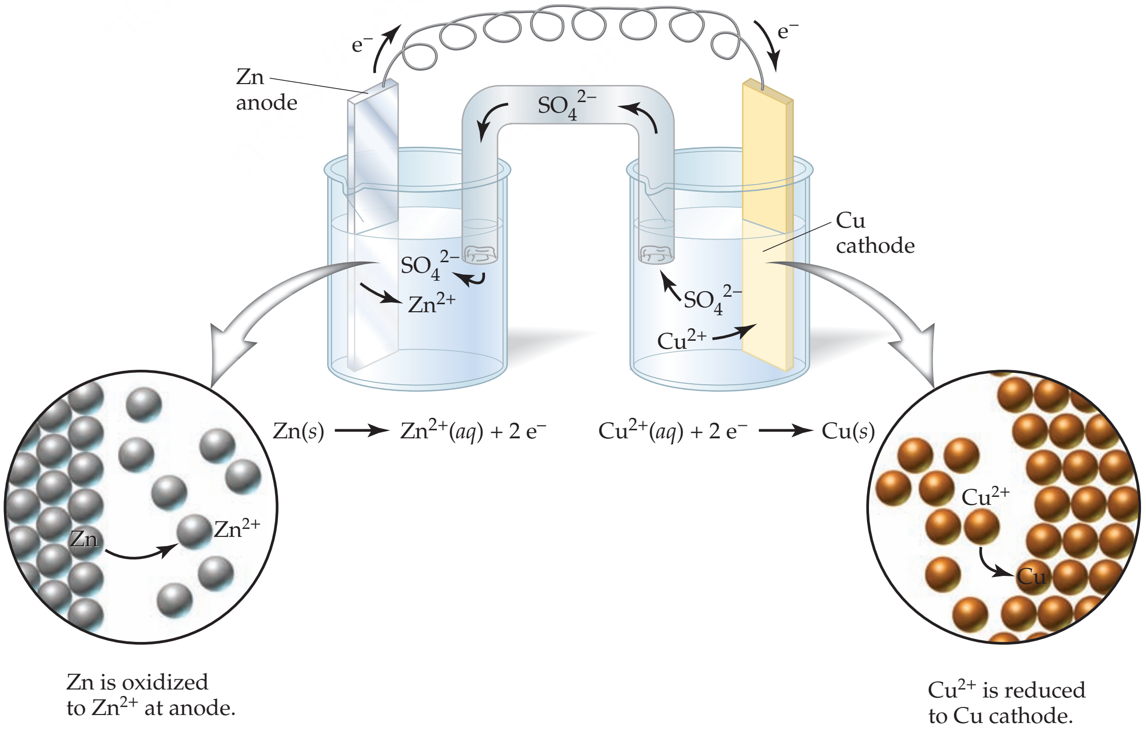


$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$

▶ **Figure 17.5 Redox Reaction in Half-Cells** In the compartment on the left, zinc metal is being oxidized. In the compartment on the right, copper(II) ion is being reduced. This redox process allows electrons to flow from the zinc electrode to the copper electrode. As the reaction proceeds, the number of zinc ions in the left half-cell increases, and the number of copper(II) ions in the right half-cell decreases.

As the reaction proceeds, the concentration of Zn^{2+} in the left compartment increases. The concentration of Cu^{2+} in the right compartment decreases. The concentration of SO_4^{2-} remains constant in both compartments. As a result, the left half-cell develops a net positive charge, as the right half-cell develops a net negative charge.

As the reaction continues, a charge develops in each compartment, which could halt the redox process. However, we can eliminate the charge buildup by introducing a **salt bridge**. A salt bridge placed between the compartments allows ions to travel between the two half-cells. The excess negative sulfate ions in the copper solution can move to the zinc solution, which has excess Zn^{2+}. Because a salt bridge eliminates charge buildup, the reaction can continue spontaneously. For spontaneous reactions that produce electrical energy, the cell is termed a **voltaic cell** (or galvanic cell). Figure 17.6 shows a voltaic cell based on the reaction of Zn in aqueous $CuSO_4$.



▲ **Figure 17.6 Voltaic Cell** The two half-cells are connected by a salt bridge, and the sulfate ions travel from the right half-cell to the left half-cell. The salt bridge reduces the positive charge buildup in the left half-cell and the negative charge buildup in the right half-cell. The cell continues to operate spontaneously as electrons flow from the zinc anode to the copper cathode.

In an electrochemical cell, the electrode at which oxidation occurs is called the **anode**. The electrode at which reduction occurs is called the **cathode**. For the voltaic cell in Figure 17.6, the zinc electrode is the anode, and Zn is being oxidized. The copper electrode is the cathode, and Cu^{2+} is being reduced. The following example exercise further illustrates the processes of oxidation and reduction that take place in a spontaneous electrochemical cell.

EXAMPLE EXERCISE 17.9 Voltaic Cells—Spontaneous Processes

Nickel can react with aqueous silver nitrate solution according to the following ionic equation:

$$\mathrm{Ni}(s) + 2\,\mathrm{Ag}^+(aq) \longrightarrow 2\,\mathrm{Ag}(s) + \mathrm{Ni}^{2+}(aq)$$

Assume the half-reactions are separated into two compartments. A Ni electrode is placed in a compartment with 1.00 *M* $Ni(NO_3)_2$, and a Ag electrode is placed in a compartment with 1.00 *M* $AgNO_3$. Indicate each of the following:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of NO_3^- in the salt bridge

Solution

Referring to Figure 17.4, we see that Ag^+ has a higher reduction potential than Ni^{2+}. Therefore, the process is spontaneous and Ni is oxidized as Ag^+ is reduced. The two half-cell processes are

(a) Oxidation: $\mathrm{Ni} \longrightarrow \mathrm{Ni}^{2+} + 2\,e^-$
(b) Reduction: $\mathrm{Ag}^+ + e^- \longrightarrow \mathrm{Ag}$
(c) The anode is where oxidation occurs; thus, Ni is the anode.
The cathode is where reduction occurs; thus, Ag is the cathode.
(d) Ni loses electrons while Ag^+ gains electrons, so the direction of electron flow is from Ni to Ag. (*Electrons flow from the anode to the cathode*.)
(e) As the Ni anode loses electrons, the Ni compartment acquires a net positive charge due to Ni^{2+}. As the Ag cathode gains electrons, the Ag compartment acquires a net negative charge due to excess NO_3^-. A salt bridge allows the cell to operate continuously as NO_3^- ions travel from the Ag compartment to the Ni compartment. (*Anions flow from the cathode to the anode*.)

Practice Exercise

Iron can react with an aqueous tin(II) nitrate solution, according to the following ionic equation:

$$\mathrm{Fe}(s) + \mathrm{Sn}^{2+}(aq) \longrightarrow \mathrm{Sn}(s) + \mathrm{Fe}^{2+}(aq)$$

An Fe electrode is placed in a compartment with 1.00 *M* $Fe(NO_3)_2$, and a Sn electrode is placed in another compartment with 1.00 *M* $Sn(NO_3)_2$. Indicate each of the following:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of NO_3^- in the salt bridge

Answers:

(a) Oxidation: $\mathrm{Fe} \longrightarrow \mathrm{Fe}^{2+} + 2\,e^-$
(b) Reduction: $\mathrm{Sn}^{2+} + 2\,e^- \longrightarrow \mathrm{Sn}$
(c) Fe is the anode; Sn is the cathode.
(d) Electrons flow from the Fe anode to the Sn cathode.
(e) A salt bridge allows NO_3^- ions to travel from the Sn cathode compartment to the Fe anode compartment.

Concept Exercise

Draw and illustrate the voltaic cell described in the practice exercise. Refer to Figure 17.4 and label the anode and cathode; show the direction of electron and anion flow.

Answer: See Appendix G, 17.9.

By convention, the anode in an electrochemical cell is usually shown as the left half-cell. If the anode is found on the left, electrons flow from the left half-cell to the right half-cell. To avoid a buildup of negative charge in the right half-cell, negative anions exit the right half-cell, travel through the salt bridge, and enter the left half-cell.

We should note that cations are also present in the salt bridge. Conversely, cations exit the left half-cell, travel through the salt bridge, and enter the right half-cell. In summary, anions travel toward the half-cell compartment where electrons are exiting (anode), and cations travel toward the half-cell compartment where electrons are entering (cathode).

Batteries

▲ **Figure 17.7 Batteries**
Batteries supply electrical energy from a redox reaction.

A **battery** is a general term for any electrochemical cell that spontaneously produces electrical energy. A battery has stored chemical energy, which can produce electrical energy measured in volts. A battery is made up of one or more voltaic cells. For example, the ordinary lead storage battery in an automobile is a series of six cells, each of which produces 2 volts. Operating together, the six cells generate 12 volts.

Unlike an ordinary voltaic cell, the batteries we use in flashlights and electronic toys do not have electrodes immersed in a conducting solution. Rather, their electrodes are connected by a solid paste that conducts electricity. A battery that does not use an electrolyte solution is called a **dry cell**. The common alkaline dry cell battery was invented more than 100 years ago by the French chemist Georges Leclanché.

A more compact version of the dry cell is the mercury dry cell. In this battery, a small zinc cup is filled with a conducting paste of HgO and KOH. Oxidation takes place at the zinc anode and reduction at the steel cathode. Mercury dry cells are used to power small devices such as electronic wristwatches and hearing aids.

An ordinary dry cell battery is not rechargeable, and the reactants stored in the battery are eventually depleted. However, a NiCad battery, which is a special type of dry cell, can be recharged almost indefinitely. That is, electricity can be used to reverse the spontaneous reaction, and the original reactants and products are reformed. A NiCad battery is so named because the redox reactions involve nickel and cadmium. NiCad batteries are used in rechargeable battery packs for calculators and portable electronic appliances. A collection of common batteries is shown in Figure 17.7.

17.7 Electrolytic Cells

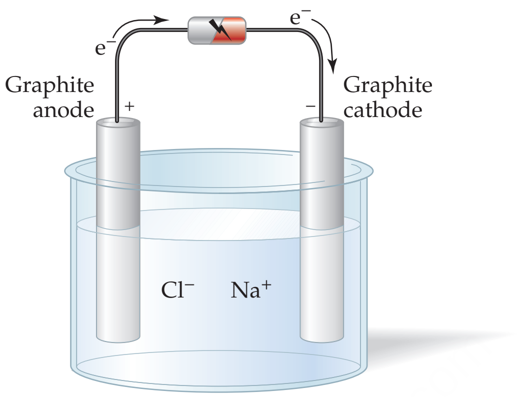
LEARNING OBJECTIVES

- Identify the anode and cathode in a given electrolytic cell.
- Identify the oxidation and reduction half-reactions in a given nonspontaneous electrochemical cell.

So far we have discussed electrochemical cells that operate spontaneously. Equally as important are electrochemical cells that operate nonspontaneously. Although a redox reaction may be nonspontaneous, we can supply electricity and force the reaction to occur. This process is referred to as **electrolysis**. When a cell phone battery runs down, it is recharged by using electricity to force nonspontaneous reactions. Thus, the application of electricity creates reactants that will react spontaneously to provide the energy to operate your cell phone.

Chromium metal is commonly plated onto steel (iron) by an electrolytic method. Although the reaction of iron in an aqueous chromium solution is not spontaneous, we can force the reaction to occur. In this process, an iron or steel object is placed in an aqueous chromium solution, and a direct electric current is applied to overcome the nonspontaneous reaction.

Many important elements, including sodium and chlorine, are produced from their compounds by electrolysis. In the preparation of sodium metal and chlorine gas, first salt is obtained by the evaporation of seawater. Next, the crystalline NaCl is placed in a crucible and heated until it melts. Then, two inert electrodes are dipped into the molten NaCl and electricity is applied. The electrodes, molten sodium chloride, and a source of electricity form an **electrolytic cell**, as shown in Figure 17.8.



◀ **Figure 17.8 Electrolytic Cell** An electrolytic cell can produce sodium metal and chlorine gas from molten NaCl. A source of direct current forces electrons to travel toward the cathode for the reduction of sodium ions, as electrons are released from the anode from the oxidation of chloride ions.

An electrolytic cell is a system in which a nonspontaneous reaction occurs as a result of the input of direct electric current. The electric current forces the nonspontaneous reaction, and in this example, NaCl decomposes into its elements.

Cathode reaction (reduction):

$$2\ \text{Na}^+(l) + 2\ e^- \longrightarrow 2\ \text{Na}(l)$$

Anode reaction (oxidation):

$$2\ \text{Cl}^-(l) \longrightarrow \text{Cl}_2(g) + 2\ e^-$$

We can write the overall equation by adding the two half-reactions together. After canceling 2 electrons from each side of the equation, we have the equation for the nonspontaneous redox reaction:

$$2\ \text{Na}^+(l) + 2\ \text{Cl}^-(l) \longrightarrow 2\ \text{Na}(l) + \text{Cl}_2(g)$$

Example Exercise 17.10 illustrates additional nonspontaneous redox reactions that occur in an electrolytic cell.

EXAMPLE EXERCISE 17.10 Electrolytic Cells—Nonspontaneous Processes

Aluminum metal is produced by passing an electric current through bauxite, which contains Al_2O_3 dissolved in the molten mineral cryolite. Graphite rods serve as electrodes, and we can write the redox equation as follows:

$$3\ \text{C}(s) + 2\ \text{Al}_2\text{O}_3(l) \xrightarrow{\text{electricity}} 4\ \text{Al}(l) + 3\ \text{CO}_2(g)$$

Indicate each of the following for this nonspontaneous process:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow

Solution

In the equation, we see that C is oxidized to CO_2 and that Al_2O_3 is reduced to Al metal. The two half-cell processes are

(a) Oxidation: $\text{C} + 2\ \text{O}^{2-} \longrightarrow \text{CO}_2 + 4\ e^-$
(b) Reduction: $\text{Al}^{3+} + 3\ e^- \longrightarrow \text{Al}$

(c) Oxidation occurs at the graphite anode, where CO_2 is released. Reduction occurs at the graphite cathode, where Al metal is produced.
(d) The electrons flow from the anode, where CO_2 gas is released, to the cathode, where Al metal is produced.

Practice Exercise

Magnesium metal is produced by passing an electric current through molten $MgCl_2$ obtained from evaporated seawater. Carbon and platinum rods serve as electrodes, and we can write the redox equation as follows:

$$C(s) + 2\,MgCl_2(l) \xrightarrow{\text{electricity}} 2\,Mg(l) + CCl_4(g)$$

Indicate each of the following for this nonspontaneous process:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow

Answers:

(a) Oxidation: $C + 4\,Cl^- \longrightarrow CCl_4 + 4\,e^-$
(b) Reduction: $Mg^{2+} + 2\,e^- \longrightarrow Mg$
(c) The C electrode is the anode; the Pt electrode is the cathode.
(d) The electrons flow from the C anode to the Pt cathode.

Concept Exercise

Draw and illustrate the electrolytic cell described in the practice exercise. Refer to Figure 17.6 and label the anode and cathode; show the direction of electron flow.

Answer: See Appendix G, 17.10.

A CLOSER LOOK Hybrid Vehicles

Q: What is the estimated gas mileage of the Chevy Volt, a hybrid vehicle that uses lithium ion batteries and a gasoline-powered generator for electricity?

Several automobile manufacturers are exploring various technologies to obtain greater gas mileage. Japanese manufacturers currently offer hybrid vehicles that utilize a combination of traditional gasoline power and lightweight voltaic cells that produce electrical energy.

▲ The Chevrolet Hybrid Gas–Electric Volt

The Chevrolet Volt, is a hybrid vehicle that uses lithium ion batteries and a gasoline-powered generator that produces electricity for the battery pack. The Volt can travel 53 miles gas-free, and over 1000 miles in hybrid mode. The batteries require 13 hours to recharge when plugged into a standard 120-V outlet.

Hybrid vehicles are designed to generate energy when braking or coasting, and store the energy as electricity in a battery pack. When accelerating, the vehicle uses the energy to power an electric motor, which reduces fuel consumption. While other hybrids use a battery-powered electric motor to supplement a gasoline engine, the Volt runs on electric power until the batteries run down. Then the gasoline-powered generator activates to produce electricity for the batteries.

The Chevrolet Volt has a front-mounted electric motor that generates 149 horsepower and has a top speed over 100 mph. The batteries are placed beneath the floor of the automobile, and a 8.9-gallon fuel tank is in the rear. The city/highway gas mileage is over 100 miles per gallon of gasoline. Similar to other hybrids, the Volt is environmentally friendly as it emits considerably less carbon dioxide greenhouse gas than conventional automobile engines.

A: Under normal driving conditions, the Chevrolet Volt achieves a combined city/highway gas mileage over 100 miles per gallon of gasoline.

Chapter Summary

Key Concepts | **Learning Objectives and Related Exercises**

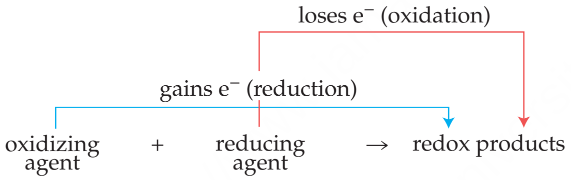
17.1 Oxidation Numbers

An oxidation–reduction reaction is characterized by the transfer of electrons. When a substance undergoes oxidation, it loses electrons, and its **oxidation number** increases. Conversely, when a substance undergoes reduction, it gains electrons, and its oxidation number decreases. When Cu is oxidized to Cu^{2+}, its oxidation number changes from 0 to +2; when Cu^{2+} is reduced to Cu, its oxidation number changes from +2 to 0. Transition metals, as well as nonmetals, demonstrate a variety of oxidation numbers. For example, Mn has oxidation numbers of +7, +6, +4, +3, +2, and 0. The oxidation numbers of Cl include +7, +5, +3, +1, 0, and −1.

- Explain the concept of an oxidation number.
  Related Exercises: 1–4
- Solve for the oxidation number of an element in each of the following:
  (a) metals and nonmetals
  (b) monoatomic and polyatomic ions
  (c) ionic and molecular compounds
  Related Exercises: 5–8

17.2 Oxidation–Reduction Reactions

In an oxidation–reduction reaction, or **redox reaction**, oxidation and reduction occur at the same time. In the process of **oxidation**, a substance loses electrons. In the process of **reduction**, a substance gains electrons. The substance that loses electrons causes reduction and is referred to as a **reducing agent**. On the other hand, the substance that gains electrons causes oxidation and is referred to as an **oxidizing agent**. We can illustrate the overall redox process as



- Identify the oxidized and reduced substances in a given redox reaction.
  Related Exercises: 9–14
- Identify the oxidizing and reducing agents in a given redox reaction.
- *Related Exercises: 15–16*

17.3 Balancing Redox Equations: Oxidation Number Method

To balance a redox equation, we keep track of the changes in oxidation number. This method allows us to balance a redox equation by determining changes in oxidation number for the oxidizing agent and reducing agent. If the oxidizing agent gains two electrons and the reducing agent loses one electron, the reducing agent must be multiplied by two to balance the electrons. The remaining elements in the equation are then balanced by inspection.

- Write a balanced chemical equation for a redox reaction using the oxidation number method.
  Related Exercises: 17–22

17.4 Balancing Redox Equations: Half-Reaction Method

Redox reactions can also be balanced by the **half-reaction** method. First, a balanced half-reaction is written for the oxidation process. Second, a balanced half-reaction is written for the reduction process. Each half-reaction is then multiplied by the number of electrons lost or gained in the opposite half-reaction. Thus, the number of electrons lost in oxidation equals the number of electrons gained in reduction. Finally, the two half-reactions are added together to give a balanced equation for the overall redox reaction.

- Explain the concept of a half-reaction.
  Related Exercises: 23–24
- Write a balanced chemical equation for a redox reaction using the half-reaction method:
  (a) in an acidic solution
  (b) in a basic solution.
  Related Exercises: 25–28

17.5 Predicting Spontaneous Redox Reactions

Half-reactions can be arranged according to their ability to undergo reduction. This tendency, called the **reduction potential**, is listed in Figure 17.4 for selected oxidizing and reducing agents. A substance that is easily reduced is a strong oxidizing agent, and a substance that is easily oxidized is a strong reducing agent. For a redox reaction to be spontaneous, the oxidizing agent must have a stronger attraction for electrons than the substance being oxidized.

- Predict the stronger oxidizing agent and reducing agent given a list of reduction potentials.
  Related Exercises: 29–32
- Predict whether a redox reaction is spontaneous or nonspontaneous given a list of reduction potentials.
  Related Exercises: 33–36

Key Concepts	Learning Objectives and Related Exercises
17.6 Voltaic Cells **Electrochemistry** is the study of the interconversion of chemical and electrical energy. In an **electrochemical cell**, a chemical reaction produces electrical energy. An electrochemical cell is composed of two **half-cell** compartments. Oxidation occurs in one half-cell and reduction in the other. If the redox reaction occurs spontaneously, the cell is referred to as a **voltaic cell**. The **anode** is the electrode where oxidation occurs, and the **cathode** is the electrode where reduction occurs. In the operation of an electrochemical cell, the electrodes are connected by a wire and the two half-cells are joined by a **salt bridge** that allows ions to travel from one compartment to the other. A **battery** is any device that supplies electricity and is composed of one or more voltaic cells. A **dry cell** is a special type of battery in which there is no conducting electrolyte solution.	• Identify the anode and cathode in a given voltaic cell. ***Related Exercises: 37–38*** • Identify the oxidation and reduction half-reactions in a given spontaneous electrochemical cell. ***Related Exercises: 39–40***
17.7 Electrolytic Cells Whereas the redox reaction is spontaneous in a voltaic cell, the redox reaction is not spontaneous in an **electrolytic cell**. In an electrolytic cell, it is necessary to supply electrical energy to produce a chemical reaction. For example, the **electrolysis** of water produces hydrogen and oxygen gases. Active metals, such as Na, K, and Ca, are obtained from electrolytic cells. In addition, chlorine gas and bromine liquid are produced commercially by passing a direct electric current through an electrolytic cell.	• Identify the anode and cathode in a given electrolytic cell. ***Related Exercises: 41–44*** • Identify the oxidation and reduction half-reactions in a given nonspontaneous electrochemical cell. ***Related Exercises: 41–44***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** a value assigned to an atom in a substance that indicates whether the atom is electron-poor or electron-rich compared to the free atom

_____ **2.** a chemical reaction in which an electron transfer takes place

_____ **3.** a chemical process in which a substance loses electrons

_____ **4.** a chemical process in which a substance gains electrons

_____ **5.** the substance undergoing reduction in a redox reaction

_____ **6.** the substance undergoing oxidation in a redox reaction

_____ **7.** a reaction that represents either an oxidation or a reduction process

_____ **8.** the relative ability of a substance to undergo reduction

_____ **9.** the study of the interconversion of chemical and electrical energy

_____ **10.** a general term for an apparatus containing two electrodes in separate compartments connected by a wire

_____ **11.** a portion of an electrochemical cell having a single electrode where either oxidation or reduction occurs

_____ **12.** a porous device that allows ions to travel between two half-cells to maintain an ionic charge balance in each compartment

_____ **13.** an electrochemical cell in which a spontaneous redox reaction occurs and generates electrical energy

_____ **14.** the providing of electricity to a nonspontaneous redox process to cause a reaction

_____ **15.** an electrochemical cell in which a nonspontaneous redox reaction occurs as a result of the input of direct electric current

_____ **16.** the electrode in an electrochemical cell at which oxidation occurs

_____ **17.** the electrode in an electrochemical cell at which reduction occurs

_____ **18.** a general term for any electrochemical cell that produces electrical energy

_____ **19.** an electrochemical cell in which the anode and cathode reactions do not take place in an aqueous solution

(a) anode (*Sec. 17.6*)
(b) battery (*Sec. 17.6*)
(c) cathode (*Sec. 17.6*)
(d) dry cell (*Sec. 17.6*)
(e) electrochemical cell (*Sec. 17.6*)
(f) electrochemistry (*Sec. 17.6*)
(g) electrolysis (*Sec. 17.7*)
(h) electrolytic cell (*Sec. 17.7*)
(i) half-cell (*Sec. 17.6*)
(j) half-reaction (*Sec. 17.4*)
(k) oxidation (*Sec. 17.2*)
(l) oxidation number (*Sec. 17.1*)
(m) oxidizing agent (*Sec. 17.2*)
(n) redox reaction (*Sec. 17.2*)
(o) reducing agent (*Sec. 17.2*)
(p) reduction (*Sec. 17.2*)
(q) reduction potential (*Sec. 17.5*)
(r) salt bridge (*Sec. 17.6*)
(s) voltaic cell (*Sec. 17.6*)

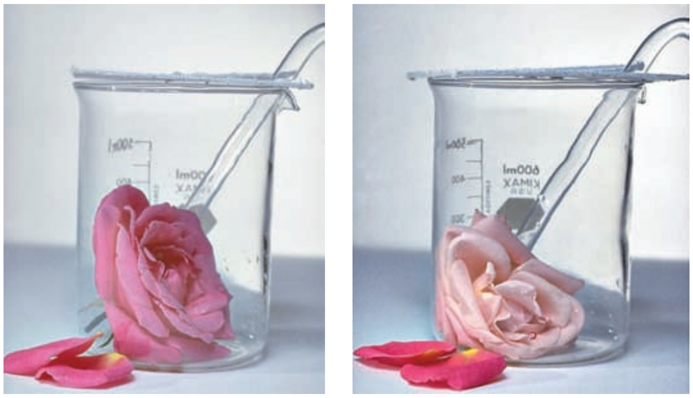
Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Oxidation Numbers (Sec. 17.1)

1. State the oxidation number for each of the following elements in the free state.
   (a) Ag
   (b) Mg
   (c) S
   (d) F_2
2. State the oxidation number for each of the following elements in the free state.
   (a) Au
   (b) Mn
   (c) P
   (d) O_2
3. State the oxidation number for each of the following monoatomic ions.
   (a) Ag^+
   (b) Mg^{2+}
   (c) S^{2-}
   (d) F^-
4. State the oxidation number for each of the following monoatomic ions.
   (a) Au^{3+}
   (b) Mn^{2+}
   (c) P^{3-}
   (d) O^{2-}
5. Calculate the oxidation number for nitrogen in the following compounds.
   (a) NH_3
   (b) N_2O_4
   (c) Li_3N
   (d) KNO_3
6. Calculate the oxidation number for silicon in the following compounds.
   (a) Si_2H_6
   (b) SiO_2
   (c) Si_3N_4
   (d) $CaSiO_3$
7. Calculate the oxidation number for sulfur in the following polyatomic ions.
   (a) $SO_3{}^{2-}$
   (b) $HSO_4{}^-$
   (c) HS^-
   (d) $S_2O_8{}^{2-}$
8. Calculate the oxidation number for carbon in the following polyatomic ions.
   (a) $CO_3{}^{2-}$
   (b) $HCO_3{}^-$
   (c) CN^-
   (d) CNO^-

Oxidation–Reduction Reactions (Sec. 17.2)

9. Supply the term that corresponds to each of the following.
   (a) a redox process characterized by electron gain
   (b) a substance undergoing reduction in a redox reaction
10. Supply the term that corresponds to each of the following.
    (a) a redox process characterized by electron loss
    (b) a substance undergoing oxidation in a redox reaction
11. State the substances undergoing oxidation and reduction in the following redox reactions:
    (a) $Co(s) + S(s) \longrightarrow CoS(s)$
    (b) $Cd(s) + Cl_2(g) \longrightarrow CdCl_2(s)$
12. State the substances undergoing oxidation and reduction in the following redox reactions.
    (a) $Mg(s) + S(s) \longrightarrow MgS(s)$
    (b) $Mn(s) + Br_2(l) \longrightarrow MnBr_2(s)$
13. State the substances undergoing oxidation and reduction in the following ionic redox reactions.
    (a) $Al(s) + Cr^{3+}(aq) \longrightarrow Al^{3+}(aq) + Cr(s)$
    (b) $F_2(g) + 2\,Cl^-(aq) \longrightarrow 2\,F^-(aq) + Cl_2(g)$
14. State the substances undergoing oxidation and reduction in the following ionic redox reactions.
    (a) $Fe^{3+}(aq) + SO_3{}^{2-}(aq) \longrightarrow Fe^{2+}(aq) + SO_4{}^{2-}(aq)$
    (b) $Cl_2(g) + 2\,I^-(aq) \longrightarrow 2\,Cl^-(aq) + I_2(s)$
15. Indicate the oxidizing and reducing agents in the following redox reactions.
    (a) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$
    (b) $PbO(s) + CO(g) \longrightarrow Pb(s) + CO_2(g)$
16. Indicate the oxidizing and reducing agents in the following redox reactions.
    (a) $Ca(s) + 2\,H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$
    (b) $Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$



▲ **Redox and a Rose** A red rose (left) undergoes a reduction in the presence of colorless SO_2 gas. After a few minutes, the color of the rose changes from red to pink (right).

Balancing Redox Equations: Oxidation Number Method (Sec. 17.3)

17. Does the electron loss by oxidation always equal the electron gain by reduction in a balanced redox equation?
18. Does the total ionic charge on the reactants always equal the total ionic charge on the products in a balanced redox equation?
19. Write a balanced equation for each of the following redox reactions using the oxidation number method.
    (a) $Br_2(l) + NaI(aq) \longrightarrow NaBr(aq) + I_2(s)$
    (b) $PbS(s) + O_2(g) \longrightarrow PbO(aq) + SO_2(g)$
20. Write a balanced equation for each of the following redox reactions using the oxidation number method.
    (a) $Cl_2(g) + KI(aq) \longrightarrow I_2(s) + KCl(aq)$
    (b) $Fe_2O_3(s) + CO(g) \longrightarrow Fe(s) + CO_2(g)$
21. Write a balanced equation for each of the following redox reactions using the oxidation number method.
    (a) $MnO_4^-(aq) + I^-(aq) + H^+(aq) \longrightarrow Mn^{2+}(aq) + I_2(s) + H_2O(l)$
    (b) $Cu(s) + H^+(aq) + SO_4^{2-}(aq) \longrightarrow Cu^{2+}(aq) + SO_2(g) + H_2O(l)$
22. Write a balanced equation for each of the following redox reactions using the oxidation number method.
    (a) $Fe^{2+}(aq) + H_2O_2(aq) + H^+(aq) \longrightarrow Fe^{3+}(aq) + H_2O(l)$
    (b) $Cr_2O_7^{2-}(aq) + Br^-(aq) + H^+(aq) \longrightarrow Cr^{3+}(aq) + Br_2(l) + H_2O(l)$

Balancing Redox Equations: Half-Reaction Method (Sec. 17.4)

23. Write a balanced half-reaction for each of the following in an *acidic* solution.
    (a) $H_2O_2(aq) \longrightarrow H_2O(l)$
    (b) $AsO_3{}^{3-}(aq) \longrightarrow AsO_3{}^-(aq)$

24. Write a balanced half-reaction for each of the following in a *basic* solution.
(a) $Ni(OH)_2(s) \longrightarrow NiO_2(s)$
(b) $NO_2^-(aq) \longrightarrow N_2O(g)$

25. Write a balanced equation for each of the following redox reactions in an *acidic* solution using the half-reaction method.
(a) $Zn(s) + NO_3^-(aq) \longrightarrow Zn^{2+}(aq) + NO(g)$
(b) $Mn^{2+}(aq) + BiO_3^-(aq) \longrightarrow MnO_4^-(aq) + Bi^{3+}(aq)$

26. Write a balanced equation for each of the following redox reactions in a *basic* solution using the half-reaction method.
(a) $MnO_4^-(aq) + S^{2-}(aq) \longrightarrow MnO_2(s) + S(s)$
(b) $Cu(s) + ClO^-(aq) \longrightarrow Cu^{2+}(aq) + Cl^-(aq)$

27. Chlorine can undergo a redox reaction in which it is simultaneously oxidized and reduced. Write a balanced equation for the following in an *acidic* solution.

$$Cl_2(aq) \longrightarrow Cl^-(aq) + HOCl(aq)$$

28. Chlorine can undergo a redox reaction in which it is simultaneously oxidized and reduced. Write a balanced equation for the following in a *basic* solution.

$$Cl_2(aq) \longrightarrow ClO_2^-(aq) + Cl^-(aq)$$

Predicting Spontaneous Redox Reactions (Sec. 17.5)

29. Refer to Figure 17.4 and indicate which substance in each of the following pairs has the greater tendency to be oxidized.
(a) $Li(s)$ or $K(s)$ **(b)** $Al(s)$ or $Mg(s)$
(c) $Fe^{2+}(aq)$ or $I^-(aq)$ **(d)** $Br^-(aq)$ or $Cl^-(aq)$

30. Refer to Figure 17.4 and indicate which substance in each of the following pairs has the greater tendency to be reduced.
(a) $Pb^{2+}(aq)$ or $Zn^{2+}(aq)$ **(b)** $Fe^{3+}(aq)$ or $Al^{3+}(aq)$
(c) $Ag^+(aq)$ or $I_2(s)$ **(d)** $Cu^{2+}(aq)$ or $Br_2(l)$

31. Refer to Figure 17.4 and indicate which substance in each of the following pairs is the stronger reducing agent.
(a) $Cu(s)$ or $Cr(s)$ **(b)** $H_2(g)$ or $Cu(s)$
(c) $Cu(s)$ or $I^-(aq)$ **(d)** $Cl^-(aq)$ or $H_2(g)$

32. Refer to Figure 17.4 and indicate which substance in each of the following pairs is the stronger oxidizing agent.
(a) $F_2(g)$ or $Cl_2(g)$ **(b)** $Ag^+(aq)$ or $Br_2(l)$
(c) $Cu^{2+}(aq)$ or $H^+(aq)$ **(d)** $Mg^{2+}(aq)$ or $Mn^{2+}(aq)$

33. Refer to Figure 17.4 and state whether each of the following reactions is spontaneous or nonspontaneous.
(a) $Br_2(l) + LiF(aq) \longrightarrow F_2(s) + LiBr(aq)$
(b) $Al(NO_3)_3(aq) + Mn(s) \longrightarrow Mn(NO_3)_2(aq) + Al(s)$

34. Refer to Figure 17.4 and state whether each of the following reactions is spontaneous or nonspontaneous.
(a) $Cr(s) + HCl(aq) \longrightarrow CrCl_3(aq) + H_2(g)$
(b) $FeCl_3(aq) + NaI(aq) \longrightarrow FeCl_2(aq) + NaCl(aq) + I_2(s)$

35. Refer to Figure 17.4 and state whether each of the following ionic redox reactions is spontaneous or nonspontaneous.
(a) $Mg(s) + Sn^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Sn(s)$
(b) $H^+(aq) + Ni(s) \longrightarrow H_2(g) + Ni^{2+}(aq)$

36. Refer to Figure 17.4 and state whether each of the following ionic redox reactions is spontaneous or nonspontaneous.
(a) $H^+(aq) + I^-(aq) \longrightarrow H_2(g) + I_2(s)$
(b) $Fe^{3+}(aq) + I^-(aq) \longrightarrow Fe^{2+}(aq) + I_2(s)$

Voltaic Cells (Sec. 17.6)

37. Sketch the following voltaic cell showing the two compartments, the anode and cathode, the wire connecting the two electrodes, and a salt bridge.

$$Mn(s) + Cd(NO_3)_2(aq) \longrightarrow Cd(s) + Mn(NO_3)_2(aq)$$

38. Sketch the following voltaic cell showing the two compartments, the anode and cathode, the wire connecting the two electrodes, and a salt bridge.

$$Ni(s) + 2\,AgNO_3(aq) \longrightarrow 2\,Ag(s) + Ni(NO_3)_2(aq)$$

39. Diagram the direction of electron flow and the movement of nitrate ions in the salt bridge that connects the two half-cells in Exercise 37.

40. Diagram the direction of electron flow and the movement of nitrate ions in the salt bridge that connects the two half-cells in Exercise 38.

Electrolytic Cells (Sec. 17.7)

41. The nonspontaneous redox reaction of nickel and aqueous iron(II) sulfate is an electrolytic process according to the following equation.

$$Ni(s) + FeSO_4(aq) \longrightarrow Fe(s) + NiSO_4(aq)$$

The half-reactions are separated into two compartments. A Ni electrode is placed in 1.00 *M* $NiSO_4$, and an Fe electrode in 1.00 *M* $FeSO_4$. Indicate each of the following for the electrolytic cell.
(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of SO_4^{2-} in the salt bridge

42. The nonspontaneous redox reaction of tin and aqueous cobalt(II) sulfate is an electrolytic process according to the following equation.

$$Sn(s) + CoSO_4(aq) \longrightarrow Co(s) + SnSO_4(aq)$$

The half-reactions are separated into two compartments. A Sn electrode is placed in 1.00 *M* $SnSO_4$, and a Co electrode in 1.00 *M* $CoSO_4$. Indicate each of the following for the electrolytic cell.
(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of SO_4^{2-} in the salt bridge

43. The nonspontaneous redox reaction of chlorine gas and aqueous sodium fluoride is an electrolytic process according to the following equation.

$$Cl_2(g) + 2\,NaF(aq) \longrightarrow 2\,NaCl(aq) + F_2(g)$$

The half-reactions are separated into two compartments. A Pt electrode is placed in 1.00 *M* NaCl, and a second Pt electrode in 1.00 *M* NaF. Indicate each of the following for the electrolytic cell.
(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of Na^+ in the salt bridge

44. The nonspontaneous redox reaction of bromine and aqueous sodium chloride is an electrolytic process according to the following equation.

$$Br_2(l) + 2\,NaCl(aq) \longrightarrow 2\,NaBr(aq) + Cl_2(g)$$

The half-reactions are separated into two compartments. A Pt electrode is placed in 1.00 *M* NaBr, and a second Pt electrode in 1.00 *M* NaCl. Indicate each of the following for the electrolytic cell.

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of Na^+ in the salt bridge

General Exercises

45. Calculate the oxidation number for sulfur in sodium tetrathionate, $Na_2S_4O_6$.

46. Calculate the oxidation number for sulfur in calcium thiosulfate, CaS_2O_3.

47. The redox reaction between cobalt metal and aqueous mercury(II) nitrate produces cobalt(II) nitrate and droplets of liquid mercury metal. Write a net ionic equation for the reaction.

48. The redox reaction between liquid mercury and aqueous gold(III) acetate produces mercury(I) acetate and gold metal. Write a net ionic equation for the reaction.

49. The redox reaction between zinc metal and sulfuric acid produces zinc sulfate and hydrogen gas. Write a net ionic equation for the reaction.

50. The redox reaction between aluminum metal and hydrochloric acid produces aluminum chloride and hydrogen gas. Write a net ionic equation for the reaction.

Challenge Exercises

51. Write a balanced equation for the reaction of a copper penny in a dilute nitric acid solution.

$$Cu(s) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO(g)$$

52. Write a balanced equation for the reaction of ethyl alcohol, C_2H_5OH, and an acidic potassium dichromate solution, which occurs in a "breathalyzer."

$$C_2H_5OH(aq) + Cr_2O_7^{2-}(aq) \longrightarrow HC_2H_3O_2(aq) + Cr^{3+}(aq) + H_2O(l)$$

Chapter 17 Self-Test

Answers to Self-Test are in Appendix J.

1. What is the oxidation number of Zn in zinc metal? (Sec. 17.1)

(a) 0
(b) −1
(c) −2
(d) −3
(e) none of the above

2. What substance is oxidized in the following redox reaction? (Sec. 17.2)

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

(a) Zn
(b) Cu^{2+}
(c) Zn^{2+}
(d) Cu
(e) none of the above

3. What is the coefficient of Co after balancing the following redox reaction using the oxidation number method? (Sec. 17.3)

$$Co_2O_3(s) + CO(g) \longrightarrow Co(s) + CO_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

4. What is the coefficient of H_2O after balancing the following redox reaction in *acidic* solution using the half-reaction method? (Sec. 17.4)

$$Cu(s) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO(aq)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

5. Which substance listed is the *strongest* oxidizing agent, given the following *spontaneous* redox reaction? (Sec. 17.5)

$$Mg(s) + Pb^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Pb(s)$$

(a) Mg
(b) Pb^{2+}
(c) Mg^{2+}
(d) Pb
(e) none of the above

6. Which of the statements listed is true regarding the following redox reaction occurring in a *spontaneous* electrochemical cell? (Sec. 17.6)

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Cu(s) + Mg^{2+}(aq)$$

(a) Mg is oxidized at the anode.
(b) Cu^{2+} is reduced at the cathode.
(c) Electrons flow from the Mg electrode to the Cu electrode.
(d) Anions in the salt bridge flow from the Cu half-cell to the Mg half-cell.
(e) all of the above

7. Which of the statements listed is true regarding the following redox reaction occurring in a *nonspontaneous* electrochemical cell? (Sec. 17.7)

$$3\,C(s) + 2\,Co_2O_3(l) \xrightarrow{\text{electricity}} 4\,Co(l) + 3\,CO_2(g)$$

(a) oxidation half-reaction: $C + 2\,O^{2-} \longrightarrow CO_2 + 4\,e^-$
(b) reduction half-reaction: $Co^{3+} + 3\,e^- \longrightarrow Co$
(c) CO_2 gas is produced at the anode.
(d) Co metal is produced at the cathode.
(e) all of the above

Key Concepts

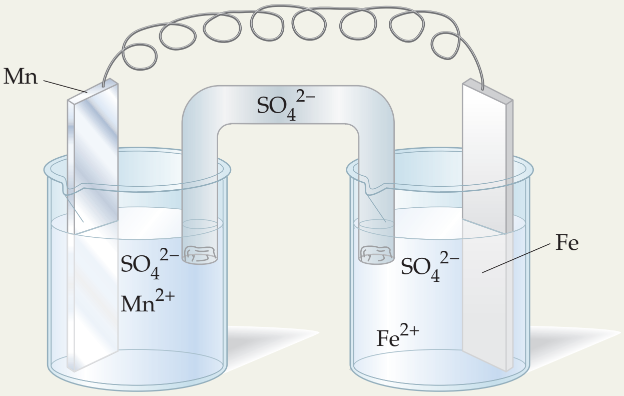
8. Refer to the redox reaction and identify the following items:

$$H_2O_2(aq) + NO_2^-(aq) \longrightarrow H_2O(l) + NO_3^-(aq)$$

(a) oxidation half-reaction
(b) reduction half-reaction
(c) substance oxidized
(d) substance reduced
(e) oxidizing agent
(f) reducing agent

9. Refer to the voltaic cell illustration and identify the following items:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of SO_4^{2-} ion flow



Critical Thinking

10. Nickel-cadmium batteries are used in rechargeable electronic calculators. Given the reaction for a *discharging* NiCad battery, what substance is being *oxidized*?

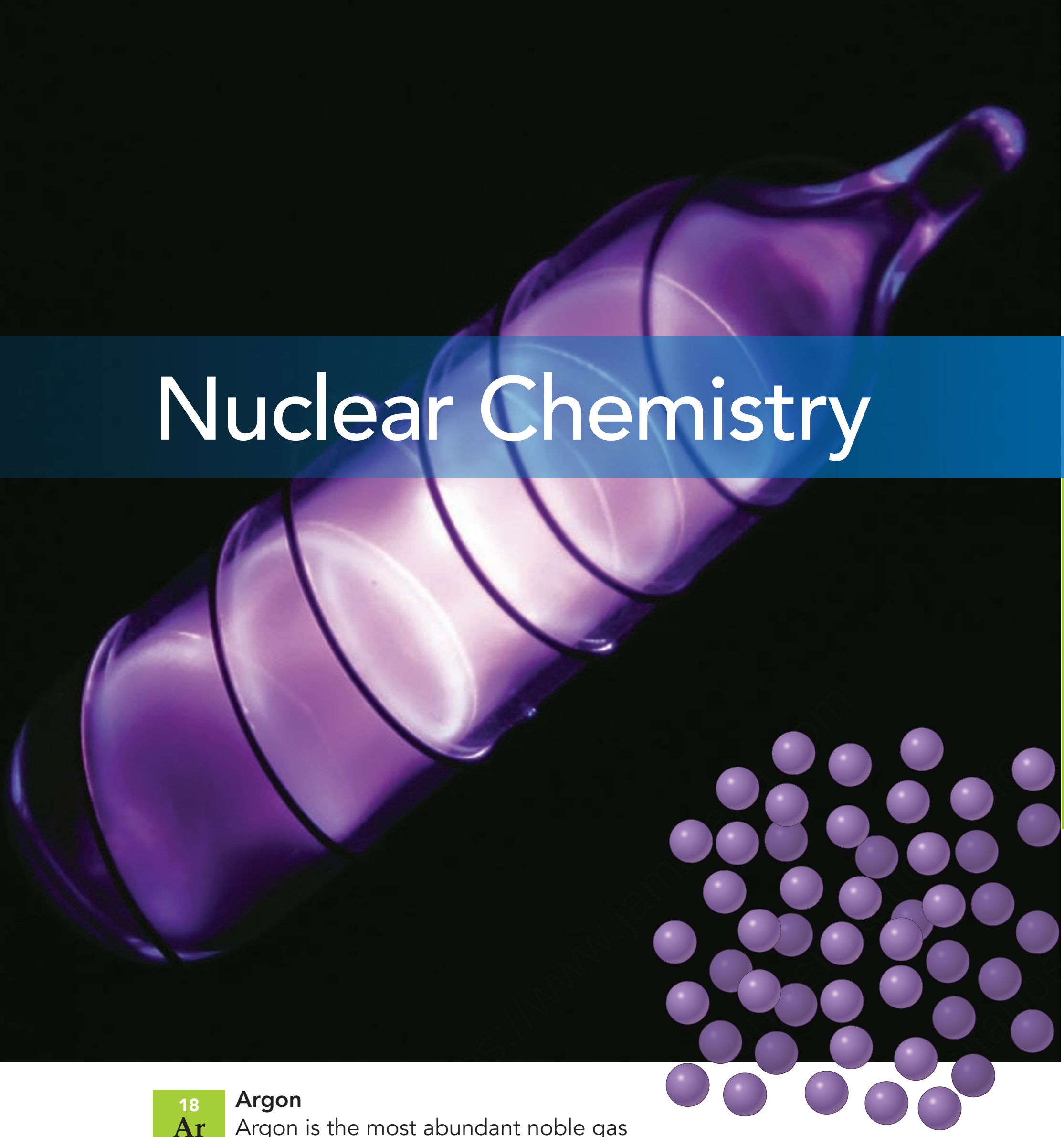
$$Cd(s) + NiO_2(s) + 2\,H_2O(l) \longrightarrow Cd(OH)_2(s) + Ni(OH)_2(s)$$

11. Nickel-cadmium batteries are used in rechargeable electronic calculators. Given the reaction for a *recharging* NiCad battery, what substance is being *oxidized*?

$$Cd(OH)_2(s) + Ni(OH)_2(s) \longrightarrow Cd(s) + NiO_2(s) + 2\,H_2O(l)$$

CHAPTER 18

Nuclear Chemistry



"I am among those who think that science has great beauty."

Marie Curie, Polish/French Physicist (1867–1934)

18
Ar
argon

Argon
Argon is the most abundant noble gas in the atmosphere and is inert to chemical reaction even at high temperatures. When excited electrically, argon gas emits a violet glow.

The energy requirements of industrial nations and the rapid consumption of nonrenewable fossil fuels have created a need to develop additional energy sources. To compensate for declining oil and natural gas resources, there is great interest in developing alternative energy sources. Nuclear energy is one possibility. The energy released from the splitting of atomic nuclei, most notably uranium nuclei, is called *nuclear fission energy*, or simply, *nuclear energy*. Although the United States has enough uranium to meet its nuclear energy needs through the middle of this century, the threat of nuclear accident and the disposal of radioactive waste are problems of great public concern.

Approximately 15% of all the world's electrical power comes from nuclear power plants. Interestingly, the United States generates the most nuclear energy, and about 20% of its electricity comes from nuclear power plants. In England, Germany, and Japan, the percentage is somewhat less than the United States. In China the percentage of electricity from nuclear power is only 1%, but in France the percentage is almost 80%, and in regions of Eastern Europe it is even higher.

18.1 Natural Radioactivity

LEARNING OBJECTIVE

- Describe the properties of alpha, beta, and gamma radiation.

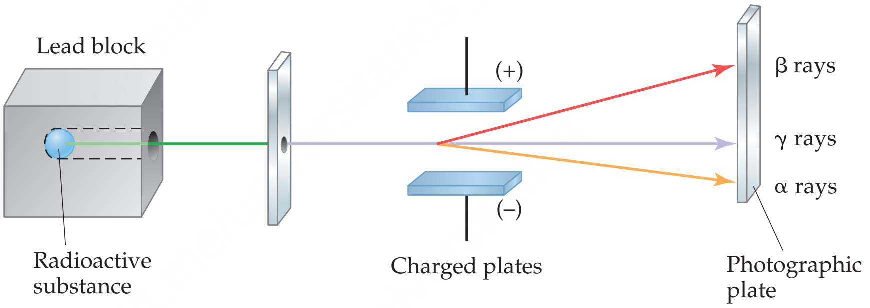
In 1896, the French physicist Henri Becquerel (1852–1908) noticed a fogged image on an unexposed photographic plate he had just removed from its heavy black protective paper. Moreover, the image was that of a uranium-containing crystal that had been resting on the paper wrapping around the photographic plate. Becquerel concluded that the uranium crystal had given off high-energy radiation to produce the fogged image. After further experiments, he discovered that the uranium crystal emitted small charged particles in addition to high-energy radiation.

▲ **Marie Curie** In 1903, Marie Curie shared the Nobel Prize in physics with Henri Becquerel and her husband, Pierre. In 1911, she won an unshared Nobel Prize in chemistry.

Radioactive Decay of Unstable Atoms

The rays Becquerel discovered were of immediate interest to the French physicist Pierre Curie (1859–1906) and his soon-to-be wife, Marie Sklodowska (1867–1934). In fact, it was Marie who coined the name **radioactivity** for Becquerel's rays. A radioactive element decays by emitting particles and/or energy from an unstable atomic nucleus. In 1898, the Curies discovered two other radioactive elements, polonium and radium. However, it took 4 years of meticulous labor to isolate a single gram of radium from 8 tons of uranium ore. In 1903, Pierre and Marie Curie shared the Nobel Prize in Physics with Becquerel for their work with radioactivity.

By 1903, the British physicist Ernest Rutherford had identified and named three different types of radioactivity. One type, an alpha ray, consists of a stream of positive particles. A second type, a beta ray, consists of negative particles. The third type, a gamma ray, does not consist of particles, but rather is high-energy radiation similar to X-rays. Rutherford was able to separate alpha, beta, and gamma rays experimentally by passing the radiation from a radioactive sample through an electric field. The apparatus he used is illustrated in Figure 18.1.



▶ **Figure 18.1 Separation of Alpha, Beta, and Gamma Rays** Notice that the alpha ray is deflected toward the negative electrode, and the beta ray toward the positive electrode. This indicates that the alpha ray is positively charged and the beta ray is negatively charged. The gamma ray is not affected because it has no charge.

After studying the radiation from unstable atoms, Rutherford eventually identified alpha particles as helium nuclei. That is, an **alpha particle (α)** is identical to a helium nucleus containing two protons and two neutrons. He found that a **beta particle (β)** is identical to an electron. A **gamma ray (γ)** is a powerful type of nuclear radiation that has neither mass nor charge. Table 18.1 lists the properties of alpha, beta, and gamma radiation.

TABLE 18.1 Properties of Nuclear Radiation

Radiation	Identity	Approximate Velocity*	Shielding Required	Penetrating Power
alpha, α	helium nucleus	$\leq$10% c	paper, clothing	low, stopped by the skin
beta, β	electron	$\leq$90% c	30 cm wood, aluminum foil	medium, ~1 cm of flesh
gamma, γ	high-energy radiation	100% c	10 cm lead, 30 cm concrete	high, passes through body

*The letter **c** is the symbol for the velocity of light, 3×10^8 m/s.

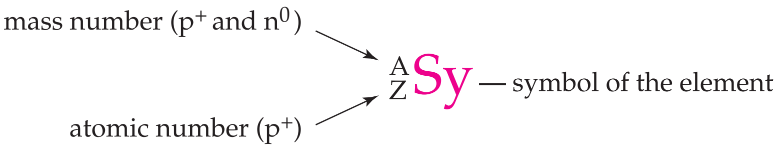
18.2 Nuclear Equations

In Section 4.4, we introduced the atomic notation used to indicate the composition of an atomic nucleus. In atomic notation, the symbol for the element is preceded by a subscript and superscript. The subscript indicates the number of protons (atomic number).

LEARNING OBJECTIVES

- Illustrate the following types of radiation using atomic notation: alpha, beta, gamma, positron, neutron, and proton.
- Write balanced nuclear equations involving natural radioactivity.

The superscript indicates the number of protons and neutrons (mass number). In the following example, the **atomic number** (symbol **Z**) appears as a subscript, and the **mass number** (symbol **A**) appears as a superscript.



In this chapter, we will use the term **nuclide** when referring to the nucleus of a specific atom. For example, the nucleus of strontium-90 is a radioactive nuclide that contains a total of 90 protons and neutrons. From the periodic table, we find that strontium has an atomic number of 38. Thus, the atomic notation for strontium-90 is

$$^{90}_{38}\mathrm{Sr}$$

If we subtract the atomic number from the mass number ($90 - 38 = 52$), we find that strontium-90 has 52 neutrons in addition to 38 protons. Although strontium-90 is radioactive, other strontium nuclides are not. For example, Sr-86, Sr-87, and Sr-88 are stable. Thus, we conclude that the number of neutrons can affect the stability of a Sr nucleus and perhaps the nuclides of other elements. Nuclear equations represent changes in nuclides due to radioactivity, and just like chemical equations they need to be balanced. Unlike chemical changes that involve valence electrons, nuclear reactions involve atomic nuclei.

Balancing Nuclear Equations

A **nuclear reaction** involves a high-energy change in an atomic nucleus. For example, a uranium-238 nucleus changes into a thorium-231 nucleus by releasing a helium-4 particle, and a great deal of energy.

$$^{235}_{92}\mathrm{U} \longrightarrow {}^{231}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}$$

To write a balanced **nuclear equation,** we must account for all atomic numbers and mass numbers. That is, the sums of the atomic numbers for reactants and products must be equal. Similarly, the sums of the mass numbers for reactants and products must be equal. The guidelines for balancing a nuclear equation are as follows.

General Guidelines for Balancing a Nuclear Equation

1. The sum of the atomic numbers (subscripts) on the left side of the equation must equal the sum of the atomic numbers on the right side.

   sum of atomic numbers for reactants = sum of atomic numbers for products

2. The sum of the mass numbers (superscripts) on the left side of the equation must equal the sum of the mass numbers on the right side.

   sum of mass numbers for reactants = sum of mass numbers for products

3. After completing the equation by writing all the nuclear particles in atomic notation, a coefficient may be necessary to balance the following:
   (a) atomic numbers of reactants and products
   (b) mass numbers of reactants and products

To balance the atomic numbers and mass numbers in a nuclear equation, we must be able to write the atomic notation for each particle in the reaction. Table 18.2 lists some common particles involved in nuclear reactions.

TABLE 18.2 Particles in Nuclear Reactions

Particle	Notation	Mass	Charge
alpha, α	${}^{4}_{2}\text{He}$	~4 amu	2+
beta, β^-	${}^{0}_{-1}\text{e}$	~0 amu	1−
gamma*, γ	${}^{0}_{0}\gamma$	0 amu	0
positron, β^+	${}^{0}_{+1}\text{e}$	~0 amu	1+
neutron, n^0	${}^{1}_{0}\text{n}$	~1 amu	0
proton, p^+	${}^{1}_{1}\text{H}$	~1 amu	1+

*Unlike particle radiation, a gamma ray is radiant energy and has no mass.

Alpha Particle Emission

Let's write a balanced nuclear equation for the radioactive decay of radium-226 by alpha emission. We can write the decay process using symbols as follows:

$$ {}^{226}_{88}\text{Ra} \longrightarrow {}^{A}_{Z}\text{X} + {}^{4}_{2}\text{He}\ (\alpha\ \text{particle}) $$

To balance the number of protons, the atomic number (represented by Z) of X the unknown element must be 86 ($88 = 86 + 2$). Therefore, we can write

$$ {}^{226}_{88}\text{Ra} \longrightarrow {}^{A}_{86}\text{X} + {}^{4}_{2}\text{He} $$

When we refer to the periodic table, we find that element 86 is radon, Rn. Thus,

$$ {}^{226}_{88}\text{Ra} \longrightarrow {}^{A}_{86}\text{Rn} + {}^{4}_{2}\text{He} $$

To balance the number of protons and neutrons, the mass number of X (represented by A) must be 222 ($226 = 222 + 4$). The balanced equation is

$$ {}^{226}_{88}\text{Ra} \longrightarrow {}^{222}_{86}\text{Rn} + {}^{4}_{2}\text{He} $$

▲ **Natural Radioactivity** Before the hazards of radiation exposure were fully understood, radioactive radium was used on the face of a watch. Radiation from the radium causes the watch to glow in the dark.

Beta Particle Emission

Unlike radium-226 that decays by alpha emission, radium-228 decays by beta emission. In beta emission, a neutron essentially decays into a proton and releases an electron. We can write this decay process in symbols as follows:

$$ {}^{228}_{88}\text{Ra} \longrightarrow {}^{A}_{Z}\text{X} + {}^{0}_{-1}\text{e}\ (\beta^-\ \text{particle}) $$

To balance the number of protons, the atomic number of X must be 89 ($88 = 89 - 1$). Since nuclide X has an atomic number of 89, the periodic table shows us that the element is actinium, Ac. The equation is

$$ {}^{228}_{88}\text{Ra} \longrightarrow {}^{A}_{89}\text{Ac} + {}^{0}_{-1}\text{e} $$

To balance the mass numbers for reactants and products, the A value for the Ac nuclide must be 228 since a beta particle has a mass number of zero ($228 = 228 + 0$). The balanced equation for the reaction is

$$ {}^{228}_{88}\text{Ra} \longrightarrow {}^{228}_{89}\text{Ac} + {}^{0}_{-1}\text{e} $$

Note that beta emission increases the atomic number by 1, thus producing the next higher atomic number element in the periodic table. The net result is that one neutron in the radium-228 nucleus decays into a proton as it releases an electron.

Gamma Ray Emission

Now let's try a more complex example. Uranium-233 decays by releasing both alpha particles and gamma rays. We can write the decay process as follows:

$$^{233}_{92}\text{U} \longrightarrow {}^{A}_{Z}\text{X} + {}^{4}_{2}\text{He} + {}^{0}_{0}\gamma \ (\gamma \text{ ray})$$

To balance subscripts, the atomic number of X must be 90 ($92 = 90 + 2 + 0$). Since nuclide X has an atomic number of 90, we see from the periodic table that the element is thorium, Th. The equation is

$$^{233}_{92}\text{U} \longrightarrow {}^{A}_{90}\text{Th} + {}^{4}_{2}\text{He} + {}^{0}_{0}\gamma$$

To balance the mass numbers for reactants and products, the *A* value for the Th nuclide must be 229 ($233 = 229 + 4 + 0$). The balanced equation for the reaction is

$$^{233}_{92}\text{U} \longrightarrow {}^{229}_{90}\text{Th} + {}^{4}_{2}\text{He} + {}^{0}_{0}\gamma$$

Note that the atomic number and mass number of a gamma ray are zero and have no effect on the equation. The gamma ray is included, however, to indicate that high-energy radiation is emitted during the decay process.

Positron Particle Emission

In 1932, a new particle was discovered. Its mass is identical to that of an electron, but its charge is 1+ rather than 1−. The newly discovered particle was given the name **positron (β^+),** for positive electron. When positrons strike electrons, they immediately annihilate each other. A collision between a positron and an electron causes mutual destruction and the release of two gamma rays.

Scientists have studied and cataloged the types of emission from radioactive nuclides. Sodium-22 emits positrons and is used in medicine to diagnose small lesions in the brain. The nuclear equation for the decay of Na-22 is as follows:

$$^{22}_{11}\text{Na} \longrightarrow {}^{A}_{Z}\text{X} + {}^{0}_{+1}\text{e} \ (\text{positron})$$

To balance the atomic numbers of reactants and products in this reaction, the atomic number of X must be 10 ($11 = 10 + 1$). Referring to the periodic table, we see that the atomic number 10 corresponds to neon, Ne. The equation is

$$^{22}_{11}\text{Na} \longrightarrow {}^{A}_{10}\text{Ne} + {}^{0}_{+1}\text{e}$$

To balance the mass numbers of reactants and products, the mass number of Ne must be 22 since a positron has a mass number of zero ($22 = 22 + 0$). The balanced equation for the reaction is

$$^{22}_{11}\text{Na} \longrightarrow {}^{22}_{10}\text{Ne} + {}^{0}_{+1}\text{e}$$

Note that positron emission decreases the atomic number by 1, thus producing the next lower element in the periodic table. The net result is that one proton in the Na-22 nucleus decays into a neutron while releasing a positron.

Electron Capture

A few large, unstable nuclides decay by a process called **electron capture (EC).** In this process, a heavy, positively charged nucleus strongly attracts an electron that lies close to the nucleus. If the attraction is sufficient, the negatively charged electron is captured by the nucleus and converts a positive proton to a neutral neutron. Since the number of protons is less, the atomic number decreases by one and a different element is formed. Consider a lead-205 nuclide that captures an electron as follows:

$$^{205}_{82}\text{Pb} + {}^{0}_{-1}\text{e} \longrightarrow {}^{A}_{Z}\text{X}$$

To balance the atomic numbers, the atomic number of X must be 81 ($82 - 1 = 81$). From the periodic table we see that element 81 is thallium, Tl. Since the mass number of an electron is 0, the mass number does not change (205 + 0 = 205). The balanced equation for the reaction is

$$^{205}_{82}\text{Pb} + {}^{0}_{-1}\text{e} \longrightarrow {}^{205}_{81}\text{Tl}$$

The following example exercise further illustrates how to write balanced nuclear equations for radioactively decaying nuclides.

EXAMPLE EXERCISE 18.1 Writing Nuclear Equations for Natural Radiation

Write a balanced nuclear equation for each of the following radioactive decay reactions.

(a) Radon-222 decays by alpha and gamma emission.
(b) Barium-133 decays by electron capture.

Solution

To write a balanced nuclear equation we must have

1. equal sums of atomic numbers for reactants and products
2. equal sums of mass numbers for reactants and products

(a) We can write the equation for the decay of radon-222 by alpha and gamma emission as follows:

$$^{222}_{86}\text{Rn} \longrightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He} + {}^{0}_{0}\gamma$$

Since the sums of the atomic numbers (86) and mass numbers (222) are the same for reactants and products, the equation is balanced.

(b) We can write the balanced nuclear equation for the decay of barium-133 by electron capture as follows:

$$^{133}_{56}\text{Ba} + {}^{0}_{-1}\text{e} \longrightarrow {}^{133}_{55}\text{Cs}$$

Since the sums of the atomic numbers (55) and mass numbers (133) are the same for reactants and products, the equation is balanced.

Practice Exercise

Identify nuclide X in the following nuclear reactions that represents the radioactive decay product from (a) beta emission and (b) positron emission.

(a) $^{99}_{43}\text{Tc} \longrightarrow {}^{A}_{Z}\text{X} + {}^{0}_{-1}\text{e}$
(b) $^{30}_{15}\text{P} \longrightarrow {}^{A}_{Z}\text{X} + {}^{0}_{+1}\text{e}$

Answers:

(a) $^{99}_{44}\text{Ru}$ (b) $^{30}_{14}\text{Si}$

Concept Exercise

Which type of natural radiation corresponds to each of the following descriptions?

(a) stopped by human skin and requires heavy cloth shielding
(b) penetrates human skin and requires aluminum shielding
(c) passes through the body and requires lead shielding

Answers: See Appendix G, 18.1.

LEARNING OBJECTIVES

- Identify products in the decay series for uranium-238 given the radiation emitted.

18.3 Radioactive Decay Series

In Section 18.2, we discussed the radioactive decay of unstable nuclides. During this process, unstable nuclides stabilize by emitting radiation. Many nuclides become stable by emitting radiation in a single step. Heavy nuclides, such as uranium-238, must go through a series of decay steps to reach a nuclide that is stable. This stepwise disintegration of a radioactive nuclide until a stable nucleus is reached is called a **radioactive decay series**.

- Identify products in the decay series for uranium-235 given the radiation emitted.
- Identify products in the decay series for thorium-232 given the radiation emitted.

Natural Decay Series

The decay series for three naturally occurring radioactive nuclides have been studied extensively by physicists and chemists. They are uranium-238, uranium-235, and thorium-232. Each of these radioactive nuclides decays into a stable isotope of lead in a series of steps.

$$\text{U-238} \longrightarrow \;\longrightarrow\; \longrightarrow \text{Pb-206}$$
$$\text{U-235} \longrightarrow \;\longrightarrow\; \longrightarrow \text{Pb-207}$$
$$\text{Th-232} \longrightarrow \;\longrightarrow\; \longrightarrow \text{Pb-208}$$

We use the term **parent–daughter nuclides** to describe a decaying parent nucleus and the resulting daughter nucleus. Although we have not shown the number of steps necessary to reach a stable nuclide, the actual sequence has been determined for each nuclide. In the case of uranium-238, 14 decay steps are required to become stable. Uranium-235 requires 11 steps, and thorium-232 requires 10 steps.

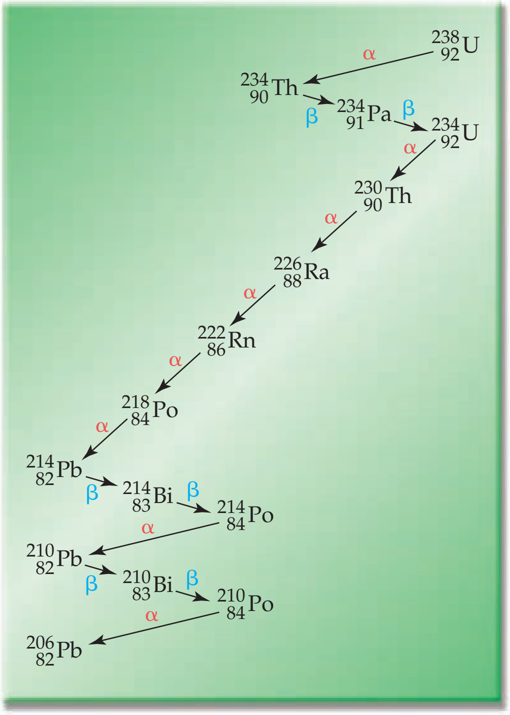
The first step in the radioactive decay of uranium-238 is the emission of an alpha particle to give thorium-234.

$$\underset{\textit{parent nuclide}}{{}^{238}_{92}\text{U}} \longrightarrow \underset{\textit{daughter product}}{{}^{234}_{90}\text{Th}} + \underset{\textit{alpha particle}}{{}^{4}_{2}\text{He}}$$

The second step is the radioactive decay of Th-234 into Pa-234 by the emission of a beta particle.

$${}^{234}_{90}\text{Th} \longrightarrow {}^{234}_{91}\text{Pa} + {}^{0}_{-1}\text{e}$$

The third step in the series is the decay of Pa-234 by beta emission. Several more steps in the decay series are required before a stable nuclide, Pb-206, is reached. Figure 18.2 illustrates the entire decay series for uranium-238. Notice that the atomic number *decreases* by 2 after each α emission (${}^{4}_{2}\text{He}$), whereas the atomic number *increases* by 1 for each β emission (${}^{0}_{-1}\text{e}$).



◀ **Figure 18.2 Radioactive Decay Series for Uranium-238** Uranium-238 requires 14 decay steps before reaching a stable nuclide. Notice that several radioactive nuclides of Po, Pb, and Bi are unstable and decay. Only Pb-206, which ends the series, is stable.

Note When we examine the periodic table, we see that a mass number in parentheses, rather than an atomic mass, is listed for all the elements following bismuth. This indicates that there are no stable nuclides of any elements higher in atomic number than bismuth. All the elements beyond atomic number 83 are unstable and therefore undergo spontaneous radioactive decay.

18.4 Radioactive Half-Life

LEARNING OBJECTIVE

- Calculate the amount of radioactive sample, given the initial amount, elapsed time, and half-life of the radionuclide.

In 1908, Hans Geiger, while working for Rutherford, developed an instrument for measuring radiation. Geiger filled a tube with a gas that ionizes when struck by radiation. The ionized gas particles allow passage of a pulse of electricity that can be converted to an audio signal heard as a click. Each click of a *Geiger counter* indicates an emission from the radioactive sample (Figure 18.3). A digital counter registers the number of clicks, providing a measure of the number of atomic nuclei undergoing radioactive decay. The number of nuclei that disintegrate in a given period of time is called the **activity** of the sample.

▲ **Figure 18.3 Geiger Counter** A Geiger counter gives a "click" each time a radioactive nucleus decays and releases a particle.

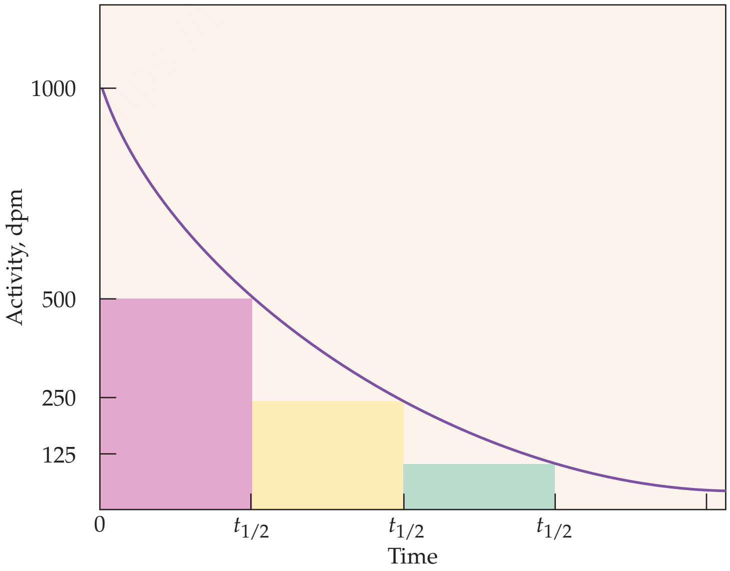
Half-Life of a Radioactive Sample

Scientists have observed that the level of radioactivity for all radioactive samples decreases with time. If we begin, for example, with a sample that has an activity of 1000 disintegrations per minute (dpm), after a given period of time the radioactivity will drop to 500 dpm, which is half the original level. Moreover, after the same period of time, the activity will drop from 500 dpm to 250 dpm.

The amount of time required for the activity of a sample to decrease by half is called the **half-life** (symbol $t_{1/2}$). If the half-life of a sample is 4 hours, it requires 4 hours for the activity to drop by half. That is, it takes 4 hours for each of the following half-life changes: 1000 to 500 dpm, 500 to 250 dpm, and 250 to 125 dpm. Since three half-lives are required for the activity to drop from 1000 to 125 dpm, the total time required is

$$3\, t_{1/2} \times \frac{4 \text{ hours}}{1\, t_{1/2}} = 12 \text{ hours}$$

The decrease in activity can be illustrated by the decay curve in Figure 18.4. The curve shows that each time the activity drops by 50%, one half-life has elapsed.



▶ **Figure 18.4 Decay Curve and Half-Life** The time required for the activity of a radioactive sample to decrease from 1000 to 500 dpm is one half-life. Notice that it requires the same amount of time, one half-life, for the activity to drop from 500 to 250 dpm.

Radioactive Waste

Disposing of radioactive waste material is a major concern because dangerous levels of radioactivity can persist for thousands of years. For example, the radioactive fuel from a plutonium-239 nuclear reactor emits high levels of radioactivity long after it is too weak to run a nuclear reactor. Let's find out how long it takes for the radioactivity of plutonium-239 waste to decrease from 20,000 to 625 dpm.

The half-life of Pu-239 is 24,000 years. By definition, we know that the activity of a radioactive sample drops by one-half after each half-life. We can list the activity of the sample corresponding to the elapsed time.

ACTIVITY	HALF-LIVES	ELAPSED TIME
20,000 dpm	—	0
10,000 dpm	$1\ t_{1/2}$	24,000 y
5000 dpm	$2\ t_{1/2}$	48,000 y
2500 dpm	$3\ t_{1/2}$	72,000 y
1250 dpm	$4\ t_{1/2}$	96,000 y
625 dpm	$5\ t_{1/2}$	120,000 y

We see that it takes five half-lives for the radioactivity of plutonium to drop from 20,000 to 625 dpm. Given that the half-life is 24,000 years, we can calculate the amount of time required for the radiation to drop to 625 dpm. The equation is

$$5\ t_{1/2} \times \frac{24{,}000\ \text{y}}{1\ t_{1/2}} = 120{,}000\ \text{y}$$

The activity of a sample is proportional to the number of radioactive nuclei, and thus to the mass of the sample. The time required for radioactive substances to decay to a safe level varies according to the nuclide. When debating the use of nuclear energy, we must consider the problem of radioactive waste disposal and its potential effect on the environment for thousands of years.

EXAMPLE EXERCISE 18.2 Half-Life of a Radioactive Nuclide

Iodine-131, in a sodium iodide tablet, is used to measure the activity of the thyroid gland. If a tablet initially contains 88 mg of I-131, what is the mass of the nuclide that remains after 24 days? The half-life of I-131 is 8 days.

Solution

This problem asks us to find the mass of I-131 remaining after 24 days. Let's begin by determining the number of half-lives that elapse. One half-life equals 8 days. Thus,

$$24\ \cancel{\text{days}} \times \frac{1\ t_{1/2}}{8\ \cancel{\text{days}}} = 3\ t_{1/2}$$

To calculate the mass of I-131 remaining after three half-lives, we decrease the mass of I-131 by one-half for each half-life. We can proceed as follows.

$$\begin{aligned} 88\ \text{mg I-131} \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} &= \text{mg I-131} \\ &= 11\ \text{mg I-131} \end{aligned}$$

Practice Exercise

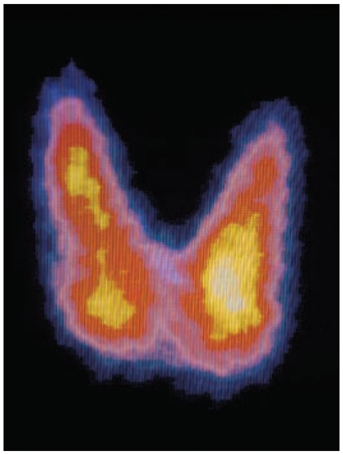
Iodine-128, in a potassium iodide solution, is used to treat a thyroid condition. If the initial activity is 10,000 dpm and the activity drops to 1250 dpm after 75 minutes, what is the half-life of I-128?

Answer: $t_{1/2} = 25$ minutes

Concept Exercise

When 500 mg of technetium-99 is administered for medical diagnosis, how much of the radionuclide remains after 24 hours? ($t_{1/2} = 6$ hours)

Answer: See Appendix G, 18.2.



▲ **Thyroid Scan** Physicians use radioactive iodine-131 as a diagnostic tool to assess activity of the thyroid gland, which regulates metabolism.

LEARNING OBJECTIVE

- Describe several applications of selected radionuclides.

18.5 Applications of Radionuclides

A nuclide that is unstable is called a **radionuclide**. Different radionuclides decay by emitting different types of radiation. These different types of radioactive emission have many applications in agriculture, industry, medicine, and research.

Radiocarbon Dating

Carbon occurs naturally as two stable nuclides, carbon-12 and carbon-13. It also occurs as an unstable nuclide, carbon-14, which decays by beta emission. The fact that carbon-14 is unstable allows its half-life to be used to estimate the age of any substance containing carbon. Scientists date objects by measuring the level of carbon-14 radioactivity in a sample. This process is called *carbon-14 dating*, or *radiocarbon dating*.

Among other things, radiocarbon dating has been used to establish the age of human remains that may be thousands of years old, the age of artifacts such as ancient wooden bowls, and the age of carbon dioxide gas from the atmosphere that has been trapped for thousands of years in glacial ice. This technique is considered reliable up to 50,000 years. Older objects usually emit too little radiation for accurate measurement.

Carbon-14 is continuously produced in Earth's upper atmosphere. The process begins when cosmic rays strike molecules and scatter high-energy neutrons. These neutrons in turn collide with nitrogen in the atmosphere. These collisions produce carbon-14 and protons. The nuclear equation is as follows:

$$^{14}_{7}\mathrm{N} + ^{1}_{0}\mathrm{n} \longrightarrow ^{14}_{6}\mathrm{C} + ^{1}_{1}\mathrm{H}$$

Carbon-14 atoms in the atmosphere combine with oxygen molecules to give radioactive CO_2. This carbon dioxide is then incorporated into plant life through the process of photosynthesis. The food chain begins with plants and continues through animals and eventually humans. A substance that contains carbon-14 emits beta radiation according to the following nuclear equation:

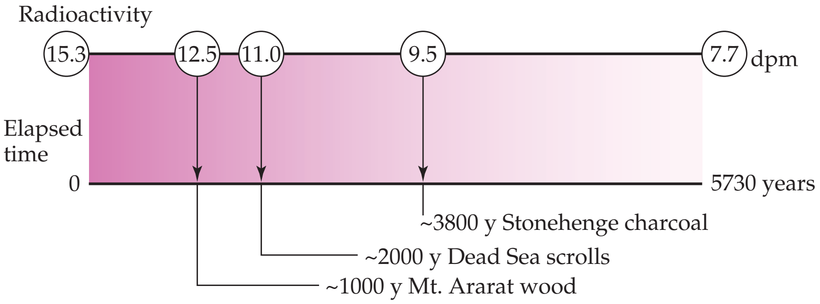
$$^{14}_{6}\mathrm{C} \longrightarrow ^{14}_{7}\mathrm{N} + ^{0}_{-1}\mathrm{e}$$

In a living plant or animal, the radiation from radiocarbon is 15.3 dpm per gram of carbon. When a plant or animal dies, it ceases to intake carbon-14, and the radiation level begins to decrease. After one half-life has elapsed, the activity declines to about 7.7 dpm. The half-life of carbon-14 is 5730 years. Therefore, if a radioactive sample has an activity of 7.7 dpm, its estimated age is 5730 years.

When a piece of wood was found on Mt. Ararat in eastern Turkey it was proposed that the wood might be from Noah's Ark. The wood sample underwent carbon-14 dating and gave an activity of 12.5 dpm per gram of carbon. This activity corresponds to that of a sample about 1000 years old. Thus, using radiocarbon dating, scientists concluded that the wood was not part of Noah's Ark.

When the Dead Sea Scrolls were found, radiocarbon dating was used to verify their authenticity. Actually, it was the linen wrapping around the Old Testament writings that gave a dating value of 2000 years. This value corresponds to a time that is consistent with biblical texts. It was this evidence from radiocarbon dating that led scholars to believe that the Dead Sea Scrolls are authentic writings.

In a similar fashion, Stonehenge in England was dated using charcoal from various campsites. The charcoal samples gave an activity of 9.5 dpm per gram of carbon, which corresponds to a sample about 3800 years old. Thus, the Stonehenge site is believed to have been active in about 1800 B.C. Figure 18.5 illustrates radiocarbon dating.



◀ **Figure 18.5 Radiocarbon Dating** The activity of carbon-14 in living tissue is about 15.3 dpm. When tissue expires, it continues to decay radioactively, but the tissue no longer incorporates carbon-14. After 5730 years, half the carbon-14 nuclides in the sample have decayed and the activity is about 7.7 dpm.

The accuracy of radiocarbon dating depends on a fundamental assumption: *the amount of carbon-14 in the atmosphere has remained constant for the last 50,000 years.* If this assumption is not true, radiocarbon dating is not valid.

Uranium-Lead Dating

Naturally occurring uranium-238 radioactively decays to lead-206 in a series of 14 steps. The half-life for the overall process is 4.5 billion years. We can indicate the net reaction according to the following equation:

$$^{238}_{92}\text{U} \longrightarrow {}^{206}_{82}\text{Pb} + 8\,^{4}_{2}\text{He} + 6\,^{0}_{-1}\text{e}$$

The ratio of uranium-238 to lead-206 in uranium samples provides an estimate of the age of very old geological events. For example, a meteorite that fell in Mexico is found to have a uranium-238/lead-206 ratio of 1:1. A ratio of 1:1 corresponds to a single half-life. Thus, the meteorite is about 4.5 billion years old.

The uranium-lead dating technique has also been used to estimate the age of Earth and of lunar rock samples collected during the *Apollo* missions to the Moon. Uranium-lead dating indicates that Earth and the Moon are about the same age. Thus, scientists have evidence for the theory that the Moon may have originally been part of Earth and broke away while Earth was in an early molten stage.

Agricultural Applications of Radionuclides

Pesticides like DDT can effectively control insects, but they often persist in the environment for a long time. Some of these pesticides are toxic to animals and humans. As an alternative, insect populations can be controlled using radiation. Gamma rays from a radionuclide such as cobalt-60 have been employed to sterilize male insects. When large numbers of sterilized males are released in an infested area, they mate with the females, but fertilization does not occur. The insect population is thus reduced and effectively controlled. A decrease in the insect population results in an increase in agricultural harvests.

Gamma irradiation of processed food destroys microorganisms. For example, cobalt-60 irradiation of pork destroys the parasite that causes trichinosis; irradiation of chicken destroys the bacteria that cause salmonella. Irradiation of other foods can extend shelf life without the use of preservatives.

▲ **Irradiation of Food** Gamma irradiation destroys microorganisms in food and extends freshness.

Medical Applications

Today, more than 100 radionuclides are available for medical diagnoses and treatments. An interesting example is plutonium-238, which decays by gamma emission. The gamma rays are used to power pacemakers in patients with irregular heartbeats. The radionuclide is sealed in a stainless steel case and implanted in the chest of the patient. The radiation from a plutonium-238 source can power a pacemaker for up to 10 years before it needs to be replaced.

A CLOSER LOOK Nuclear Medicine

Q: Which radionuclide emits gamma rays that are used for radiation therapy on tumors in cancer patients?

The term *nuclear medicine* refers to the use of radionuclides for medical purposes. For example, iodine-131 is used to measure the activity of the thyroid gland, which requires iodine to regulate metabolism. Patients diagnosed as having a thyroid condition are given an iodine-131 tablet, and after 24 hours a radiation scan is performed. The amount of radiation detected at the base of the throat, where the thyroid is located, is an indication of thyroid activity. Iodine-131 can also be used in veterinary medicine for the diagnosis of hyperthyroidism in cats.

Xenon-133 is used to diagnose respiratory problems. A patient inhales air containing the radionuclide. Both oxygen and xenon-133 are taken up by healthy lung tissue. As the chest is scanned for radiation, healthy areas of the lungs are revealed. An area where respiration is impaired is indicated by less radiation.

Iron-59 is used to diagnose anemia, a condition marked by a low red blood cell count. Red blood cells contain hemoglobin, which in turn contains iron atoms. A patient is given iron-59, and the radionuclide bonds to the hemoglobin. When the patient is checked for radiation, a scan reveals radioactive hemoglobin, providing an indirect count of red blood cells.

Brain tumors can be diagnosed using technetium-99, which concentrates in brain tissue. When a patient is given technetium-99, rapidly dividing cancer cells incorporate more of the radionuclide than normal cells. A scan shows radiation hot spots that can help locate the tumor. After locating the tumor, a cobalt-60 source of gamma radiation is focused on the tumor, which destroys the cancer cells and a small portion of the surrounding healthy tissue.

Breast cancer can be treated with iridium-192. A hypodermic needle is first injected into the tumor, followed by insertion of an iridium-192 bead through the needle. The needle is then withdrawn, and gamma rays from the radionuclide destroy the tumor. However, similar to radiation therapy, the gamma rays destroy some healthy tissue at the periphery of the tumor.

Although there is public concern about the use of radioactive substances, nuclear medicine is considered safe. In fact, diagnoses and treatments involving medical radionuclides save or prolong thousands of lives each year.

◀ **Radiation Therapy** A cancer patient receiving radiation therapy for a brain tumor.

A: Cobalt-60 is radioactive and emits gamma rays for radiation therapy treatment that destroys tumors in cancer patients.

LEARNING OBJECTIVE

- Write balanced nuclear equations involving induced radioactivity.

18.6 Induced Radioactivity

When a nuclide is bombarded with an atomic particle, the nuclide is converted to a different element. That is, a nuclear reaction occurs. Causing a nuclear reaction by particle bombardment is termed *induced radioactivity.* The conversion of one element to another by an induced nuclear reaction is called **transmutation**.

In 1919, Ernest Rutherford discovered the first transmutation reaction. When he bombarded nitrogen-14 with alpha particles, he detected the release of a proton, ${}^{1}_{1}H$. After careful analysis, he concluded that oxygen had also been produced. The equation for the transmutation is

$${}^{14}_{7}N + {}^{4}_{2}He \longrightarrow {}^{17}_{8}O + {}^{1}_{1}H$$

In 1932, James Chadwick, a student of Rutherford's, fired alpha particles at a sheet of beryllium metal. He found that some of the beryllium was changed to carbon and that a neutral particle was released. Chadwick called the neutral particle a neutron. Interestingly, Rutherford had predicted a neutral particle two decades earlier; however, at that time the neutron was too elusive to be confirmed experimentally. The equation for the transmutation is

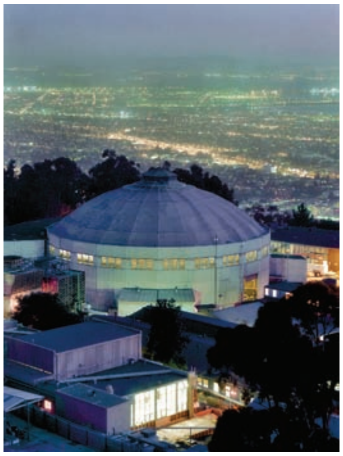
$$^{9}_{4}\text{Be} + ^{4}_{2}\text{He} \longrightarrow ^{12}_{6}\text{C} + ^{1}_{0}\text{n}$$

Synthesis of New Elements

After the discovery of the neutron, it was proposed that a neutral particle could easily invade another nucleus. In 1940, this proposal proved correct, as uranium was transmuted into a new element. The new element was named neptunium for the planet Neptune. (Neptune lies just beyond Uranus, the planet for which uranium is named.) The equation for the synthesis of neptunium is

$$^{238}_{92}\text{U} + ^{1}_{0}\text{n} \longrightarrow ^{239}_{93}\text{Np} + ^{0}_{-1}\text{e}$$

The elements beyond uranium in the periodic table do not occur naturally. Rather, they have been created in the laboratory. These transuranium elements are synthesized by accelerating a projectile particle to a high velocity and firing it at a heavy nuclide that serves as a target nucleus. One instrument used to accelerate particles to sufficient energy to overcome the repulsion of the target nucleus is a *cyclotron.* Most of the elements beyond uranium have been synthesized using the cyclotron at the University of California at Berkeley. In fact, element 103 is named lawrencium (Lr) for the Berkeley physicist who invented the cyclotron, Ernest O. Lawrence.



▲ **UC Berkeley Cyclotron** The cyclotron was invented at the University of California, Berkeley, and was used to create neptunium, plutonium, and other heavy elements.

Another instrument for accelerating particles is the *linear accelerator.* Linear accelerators and cyclotrons are sometimes referred to as "atom smashers" because they send high-velocity particles smashing into a target nucleus. The linear accelerator at Stanford University was completed in 1966, and is the longest in the world. It is capable of accelerating projectile particles to velocities approaching the speed of light (Figure 18.6).

▲ **Figure 18.6 Stanford Linear Accelerator** The Stanford linear accelerator is operated by Stanford University and is located near the campus. It is about two miles long and is the longest linear accelerator in the world.

In 1964 Soviet physicists claimed a successful synthesis of element 104. They named the element kurchatovium (Ku) in honor of their team leader, Igor Kurchatov. They used a linear accelerator to send a neon-22 projectile smashing into a plutonium target nucleus. Although the half-life of the new element was only 0.3 s, the Russians proposed the following reaction:

$$^{242}_{94}\text{Pu} + ^{22}_{10}\text{Ne} \longrightarrow ^{260}_{104}\text{Ku} + ^{1}_{0}\text{n}$$

Notice that the mass numbers in the equation are not balanced. The total of the reactant superscripts is 264 (242 + 22), but the total of the product superscripts is only 261 (260 + 1). To balance the equation, we must use 4 neutrons. The balanced nuclear equation is

$$^{242}_{94}\text{Pu} + ^{22}_{10}\text{Ne} \longrightarrow ^{260}_{104}\text{Ku} + 4^{1}_{0}\text{n}$$

When the Americans tried to confirm the Russian experiment, they were unsuccessful, which cast doubt on the credibility of element 104. In 1969, however, scientists at the University of California at Berkeley reported a different nuclide of element 104 that has a half-life of 5 s. The Berkeley group named the element rutherfordium (Rf) in honor of Ernest Rutherford. They induced the reaction by firing a carbon-12 projectile at a californium target nucleus.

$$^{249}_{98}\text{Cf} + ^{12}_{6}\text{C} \longrightarrow ^{257}_{104}\text{Rf} + ^{1}_{0}\text{n}$$

When we check the equation, we find that the atomic numbers are balanced but the mass numbers are not. The superscript total is 261 on the left side, but only 258 on the right side. To balance the equation, we place the coefficient 4 before the neutrons. The balanced nuclear equation is

$$^{249}_{98}\text{Cf} + ^{12}_{6}\text{C} \longrightarrow ^{257}_{104}\text{Rf} + 4^{1}_{0}\text{n}$$

The following example exercise provides additional practice in writing balanced nuclear equations for the synthesis of new elements.

EXAMPLE EXERCISE 18.3 Writing Nuclear Equations for induced Radiation

Write a balanced nuclear equation for the following nuclear reaction induced by an iron nuclide.

$$\text{Bi-209} + \text{Fe}-? \longrightarrow \text{Mt-266} + 1\ \text{neutron}$$

Solution

To write a nuclear equation, we must balance atomic numbers and mass numbers. Let's begin by writing the equation using atomic notation. We obtain the atomic number of each nuclide from the periodic table.

$$^{209}_{83}\text{Bi} + ^{?}_{26}\text{Fe} \longrightarrow ^{266}_{109}\text{Mt} + ^{1}_{0}\text{n}$$

We see that the sum of the atomic numbers is 109 on both sides of the equation. Since the sum of the mass numbers on the right side is 267, the sum on the left side must be 267. Thus, the mass number for Fe must be 58 ($209 + 58 = 267$). The balanced nuclear equation is

$$^{209}_{83}\text{Bi} + ^{58}_{26}\text{Fe} \longrightarrow ^{266}_{109}\text{Mt} + ^{1}_{0}\text{n}$$

Practice Exercise

Write a balanced nuclear equation for the following nuclear synthesis reaction.

$$\text{Cf-249} + \text{O-18} \longrightarrow \text{Sg-?} + 4\ \text{neutrons}$$

Answer:

$$^{249}_{98}\text{Cf} + ^{18}_{8}\text{O} \longrightarrow ^{263}_{106}\text{Sg} + 4^{1}_{0}\text{n}$$

Concept Exercise

In 1994, physicists in Darmstadt, Germany, synthesized element 110. If Ds-269 and a neutron were produced by firing a projectile into Pb-208, what was the projectile nuclide?

Answer: See Appendix G, 18.3.

18.7 Nuclear Fission

LEARNING OBJECTIVES

- Describe the process in a nuclear fission reaction.
- Write a nuclear equation for a given fission reaction.

Spontaneous **nuclear fission** is the process whereby a heavy nucleus, such as uranium-235, splits into lighter nuclei while releasing energy. Many heavy nuclei are unstable and decay by spontaneous nuclear fission. For example, californium-252 can split apart spontaneously into lighter nuclei as follows:

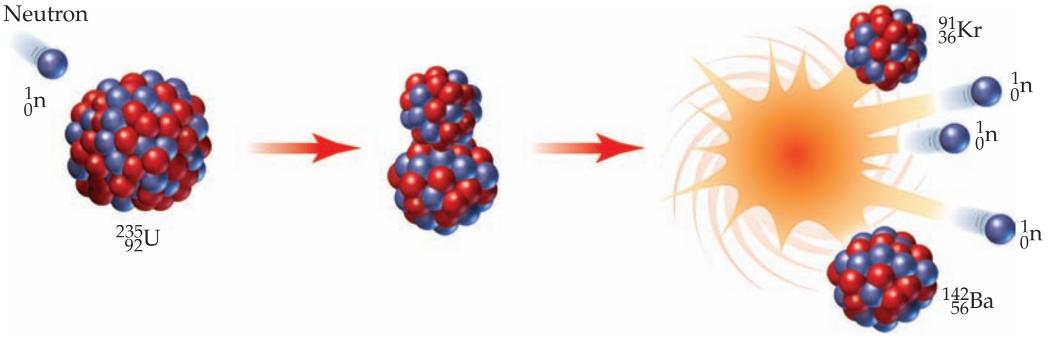
$$^{252}_{98}\text{Cf} \longrightarrow ^{142}_{56}\text{Ba} + ^{106}_{42}\text{Mo} + 4^{1}_{0}\text{n} + \text{energy}$$

There are only a few nuclides that can be made to undergo fission, for example, uranium-235, uranium-233, and plutonium-239. The fission process is induced by a slow-moving neutron. When the neutron strikes a heavy nucleus and is captured, the resulting nucleus becomes unstable. The unstable nucleus then decays by splitting into lighter nuclei and releasing a few neutrons. The fission of uranium-235 can occur as follows:

$$^{1}_{0}\text{n} + ^{235}_{92}\text{U} \longrightarrow ^{137}_{52}\text{Te} + ^{97}_{40}\text{Zr} + 2^{1}_{0}\text{n} + \text{energy}$$

After uranium-235 is struck by a neutron, the resulting unstable nucleus can undergo fission in a number of ways, one of which is depicted in Figure 18.7. As an analogy, a glass jar can split apart in various ways after being shot by a pellet. For example, uranium-235 can give the following nuclei as products:

$$ {}^{1}_{0}\text{n} + {}^{235}_{92}\text{U} \longrightarrow {}^{142}_{56}\text{Ba} + {}^{91}_{36}\text{Kr} + 3\,{}^{1}_{0}\text{n} + \text{energy} $$



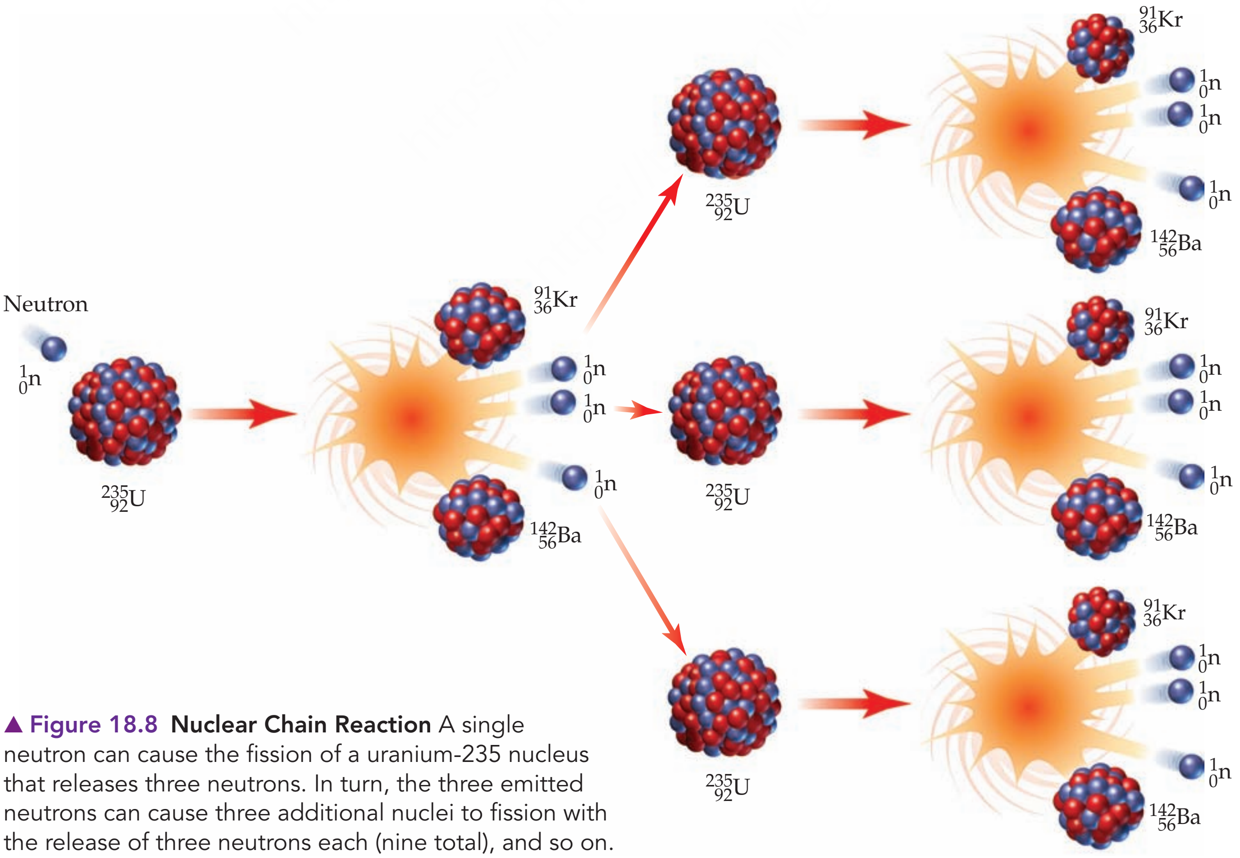
◀ **Figure 18.7 Nuclear Fission** A slow-moving neutron collides with a uranium-235 nucleus to form uranium-236. The resulting nucleus is unstable and splits into two lighter nuclei, releasing a few neutrons. In addition to particles, the fission reaction releases a huge burst of energy.

Nuclear Chain Reaction

The process of nuclear fission is analogous to a single domino knocking over three dominos, which in turn can ultimately topple an entire row of dominos. For each nucleus that undergoes fission, two or three neutrons are usually released. The emitted neutrons can cause additional fission reactions. In turn, the neutrons released from these fission reactions can initiate more fission reactions. If the fission process continues to repeat itself and is self-sustaining, it is called a **chain reaction**. We can indicate the growth in the number of neutrons in a chain reaction of uranium-235 as follows:

$$ 1\,{}^{1}_{0}\text{n} \longrightarrow 3\,{}^{1}_{0}\text{n} \longrightarrow 9\,{}^{1}_{0}\text{n} \longrightarrow 27\,{}^{1}_{0}\text{n} $$

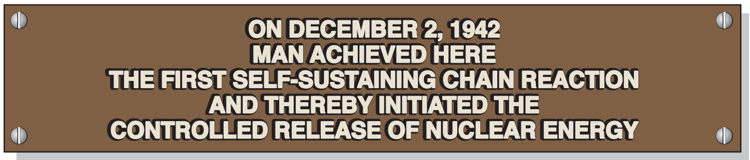
We can illustrate a chain reaction as shown in Figure 18.8. Starting with a single uranium-235 nucleus and 1 neutron, the chain reaction quickly escalates to give fission products and 9 neutrons.



▲ **Figure 18.8 Nuclear Chain Reaction** A single neutron can cause the fission of a uranium-235 nucleus that releases three neutrons. In turn, the three emitted neutrons can cause three additional nuclei to fission with the release of three neutrons each (nine total), and so on.

For a chain reaction to occur, there must be a fairly large mass of fissionable nuclide. Recall that atoms are mostly empty space. If the mass is too small, fission neutrons simply pass through the fissionable substance without striking a nucleus and causing it to fission. The minimum mass of fissionable nuclide required to sustain a chain reaction is called the **critical mass**. The critical mass depends on the purity of the sample and other factors, but for uranium-235 the critical mass is about 50 kg.

In 1939, nuclear fission was discovered in Germany. The possibility for a sustained chain reaction was experimentally confirmed three years later. In 1942, the Italian physicist Enrico Fermi directed an international group of scientists who carried out the first self-sustaining nuclear chain reaction. This landmark event took place on a converted squash court beneath the University of Chicago and is commemorated by a plaque at the university (Figure 18.9).



▶ **Figure 18.9 Commemoration Plaque** This large plaque at the University of Chicago signifies the ushering in of the nuclear age.

Uranium-235 Enrichment

Naturally occurring uranium is only 0.7% fissionable uranium-235. Most of the uranium, 99.3%, is nonfissionable uranium-238. To obtain a chain reaction, it is necessary to enrich the natural mixture with the fissionable nuclide. That is, the mixture must be enriched from 0.7% to about 3% uranium-235. The enrichment process is started by reacting uranium with fluorine gas to give gaseous uranium hexafluoride.

$$^{235}\text{U}(s) + 3\ \text{F}_2(g) \longrightarrow {}^{235}\text{UF}_6(g)$$
$$^{238}\text{U}(s) + 3\ \text{F}_2(g) \longrightarrow {}^{238}\text{UF}_6(g)$$

The nuclides $^{235}UF_6$ and $^{238}UF_6$ are too similar to separate chemically. However, the two gaseous nuclides can be separated physically. The lighter nuclide, $^{235}UF_6$, diffuses slightly faster than the heavier nuclide $^{238}UF_6$. When the first fraction that diffuses through the tube is collected, it is richer in the faster-diffusing nuclide, $^{235}UF_6$. The collected uranium fraction is then reduced to the element, which is enriched in the nuclide uranium-235.

A CLOSER LOOK Heavy Water

Q: Do all water molecules have the same mass?

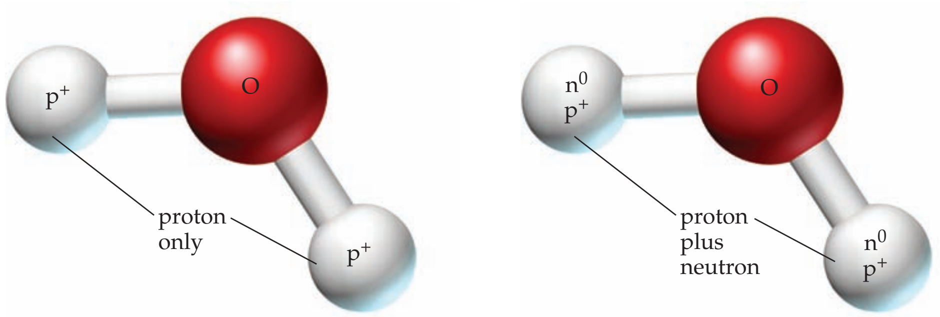
Hydrogen atoms usually have one proton as their nucleus. However, about 1 in 6000 hydrogen atoms has both a proton and a neutron as its nucleus. This isotope of hydrogen is called *deuterium* (D). The mass of deuterium is about twice that of hydrogen. When deuterium atoms replace the hydrogen atoms in water, the resulting compound is referred to as *heavy water.* The systematic name of heavy water is deuterium oxide, and its formula is D_2O.

Heavy water is colorless, odorless, and tasteless, but it is toxic to humans and animals. The physical and chemical properties of heavy water are similar to those of ordinary light water. The following table lists a few properties of light water versus heavy water.

Properties of Water versus Heavy Water

Property	Light Water, H_2O	Heavy Water, D_2O
appearance, smell, taste	colorless, odorless, tasteless liquid	colorless, odorless, tasteless liquid
molecular mass	18.02 amu	20.03 amu
density	1.000 g/mL at 4 °C	1.105 g/mL at 4 °C
freezing point	0.00 °C	3.82 °C
boiling point	100.00 °C	101.42 °C

Heavy water has two important uses in nuclear reactors. When an atomic nucleus undergoes fission, it releases neutrons. However, these neutrons are moving at a high velocity and must be slowed in order to be captured and cause further nuclear fissions of a radioactive substance such as uranium-235. Heavy water is ideal for slowing down fast neutrons released during the fission process. In addition, heavy water is used to cool the nuclear reactor and prevent meltdown.



A: Light water, H_2O, has a molecular mass of 18.02 amu, while heavy water, D_2O, is 20.03 amu. Heavy water molecules have a greater mass because deuterium atoms have a neutron and regular hydrogen atoms do not.

18.8 Nuclear Fusion

LEARNING OBJECTIVES

- Describe the process in a nuclear fusion reaction.
- Write a nuclear equation for a given fusion reaction.

Nuclear fusion is the process of combining two light nuclei into a heavier nucleus. It is more difficult to initiate a fusion reaction than a fission reaction, but a fusion reaction releases more energy than a comparable fission reaction. In addition, nuclear fusion is a cleaner process than fission because it produces very little radioactive waste.

The Sun is a huge nuclear fusion reactor operating at millions of degrees. The Sun is made up of 73% hydrogen, 26% helium, and only 1% of all other elements. Three of the fusion reactions thought to occur on the Sun are

$$ {}^{1}_{1}\text{H} + {}^{1}_{1}\text{H} \longrightarrow {}^{2}_{1}\text{H} + {}^{0}_{+1}\text{e} + \text{energy} $$

$$ {}^{2}_{1}\text{H} + {}^{1}_{1}\text{H} \longrightarrow {}^{3}_{2}\text{He} + \text{energy} $$

$$ {}^{3}_{2}\text{He} + {}^{1}_{1}\text{H} \longrightarrow {}^{4}_{2}\text{He} + {}^{0}_{+1}\text{e} + \text{energy} $$

Extreme temperatures, greater than 10,000,000°C, are needed to overcome the strong repulsive forces of the positively charged nuclei. Experimentally, a powerful laser beam is used to generate the necessary heat to initiate nuclear fusion. A major practical problem is how to confine a reaction that takes place at millions of degrees. Since ordinary materials vaporize at this temperature, a strong magnetic field is used to confine the fusion reaction. Thus, nuclear fusion is said to occur in a "magnetic bottle."

Nuclear fusion has been achieved in research fusion reactors on a limited scale. To date, continuous nuclear fusion has not been achieved, and no commercial fusion reactors are in operation. If the technology obstacles can be overcome, the fusion of two deuterium nuclei could be used to supply the world with electrical energy.

Helpful Hint

Nuclear Fission vs. Nuclear Fusion

Nuclear fission is the process whereby a heavy nucleus, such as uranium-235, splits into two lighter nuclei while releasing energy. The fission process begins when a large atomic nucleus captures a slow-moving neutron and subsequently splits into two heavy fragments and a few neutrons, which in turn, can initiate a second fission reaction.

Nuclear fusion is the process whereby two lighter nuclei, such as hydrogen or deuterium, combine to form a heavier nucleus. The fusion process requires temperatures above 10 million degrees, and must be contained in a "magnetic bottle."

Deuterium, ${}^{2}_{1}\text{H}$, is the nuclide of hydrogen that contains one proton and one neutron. Deuterium is available from the sea in the form of heavy water molecules, D_2O (Section 4.5). It is estimated that more energy is potentially available from the deuterium in one cubic mile of seawater than from all the current petroleum reserves in the world. A typical fusion reaction of deuterium is as follows:

$$ {}^{2}_{1}\text{H} + {}^{2}_{1}\text{H} \longrightarrow {}^{4}_{2}\text{He} + \text{energy} $$

Another promising nuclear fusion reaction involves deuterium and tritium. **Tritium**, ${}^{3}_{1}\text{H}$, is the nuclide of hydrogen that contains one proton and two neutrons. Here again, the fusion of nuclei requires temperatures above 10,000,000°C and is initiated by a powerful laser. A typical fusion reaction of deuterium and tritium produces an alpha particle and a neutron, while releasing a huge amount of energy.

$$ {}^{3}_{1}\text{H} + {}^{2}_{1}\text{H} \longrightarrow {}^{4}_{2}\text{He} + {}^{1}_{0}\text{n} + \text{energy} $$

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
18.1 Natural Radioactivity An unstable atomic nucleus disintegrates naturally by emitting **radioactivity** in the form of alpha, beta, or gamma radiation. An **alpha particle (α)** has a positive charge and is identical to a helium-4 nucleus. A **beta particle (β)** has a negative charge and is identical to an electron. A **gamma ray (γ)** is a powerful type of nuclear radiation that has neither mass nor charge.	• Describe the properties of alpha, beta, and gamma radiation. *Related Exercises 1–6*
18.2 Nuclear Equations The composition of a nucleus is indicated using atomic notation, where the **atomic number** of the element is indicated by a subscript and the **mass number** by a superscript. In this chapter, we have used the term **nuclide** to refer to the nucleus of an atom with a specific composition of protons and neutrons. A **nuclear reaction** involves a change in a nucleus, for example, **electron capture** or **positron** emission. A **nuclear equation** represents the nuclear change using atomic notation for each particle. To balance a nuclear equation, two criteria are necessary: (1) the sums of the atomic numbers must be the same on each side of the equation, and (2) the sums of the mass numbers must be equal for reactants and products.	• Illustrate the following types of radiation using atomic notation: alpha, beta, gamma, positron, neutron, and proton. *Related Exercises 7–8* • Write balanced nuclear equations involving natural radioactivity. *Related Exercises 9–12*
18.3 Radioactive Decay Series If a nuclide disintegrates through the emission of radiation in more than one step, the overall process is called a **radioactive decay series.** The decaying nucleus is referred to as the parent nuclide, and the resulting nucleus is the daughter nuclide. Together, they are referred to as **parent–daughter nuclides.** In nature, uranium-235, uranium-238, and thorium-232 decay naturally in a series of steps. Each of these nuclides continues to decay in a series of steps until a stable nuclide of lead is produced.	• Identify products in the decay series for uranium-238 given the radiation emitted. *Related Exercises 13–14* • Identify products in the decay series for uranium-235 given the radiation emitted. *Related Exercises 15–16* • Identify products in the decay series for thorium-232 and neptunium-237 given the radiation emitted. *Related Exercises 17–18*
18.4 Radioactive Half-Life We can keep track of the number of decaying nuclei in a sample using a Geiger counter. The **activity** of the sample refers to the number of unstable nuclei that decay in a given period of time. The activity of a sample decreases with time because there are fewer nuclei to decay. However, the time required for 50% of the radioactive nuclei to decay is constant and is called the **half-life.** After each half-life, only 50% of the radioactive nuclei remain.	• Calculate the amount of radioactive sample, given the initial amount, elapsed time, and half-life of the radionuclide. *Related Exercises 19–24*
18.5 Applications of Radionuclides Scientists have found important applications for a broad selection of radioactive nuclides. **Radionuclides** are available for dating objects, for improving agricultural harvests, and for providing energy. In addition, there is a long and growing list of radionuclides that are used for medical diagnosis and treatment.	• Describe several applications of selected radionuclides. *Related Exercises 25–30*
18.6 Induced Radioactivity Scientists can induce radioactivity by firing light, accelerated nuclides at heavy nuclei. In the event that the atomic number changes and the new nucleus corresponds to a different element, the change is called a **transmutation** reaction. With particle accelerators, it is possible to cause hundreds of nuclear reactions. A particle accelerator, such as a *cyclotron* or a *linear accelerator,* accelerates a projectile nuclide and sends it crashing into a target nucleus. This technique is called "atom smashing" and has created hundreds of synthetic nuclides.	• Write balanced nuclear equations involving induced radioactivity. *Related Exercises 31–38*

Key Concepts	Learning Objectives and Related Exercises
18.7 Nuclear Fission A neutron striking an unstable nucleus can induce **nuclear fission;** that is, the nucleus can split into two or more particles while releasing energy. The three most important nuclides that undergo fission are uranium-235, plutonium-239, and uranium-233. When a nucleus fissions, two or three neutrons are usually released. These neutrons can in turn cause a second and a third nuclear fission. If a single neutron initiates continuous fission, the process is called a **chain reaction.** For a chain reaction to occur, there must be a sufficient amount of fissionable nuclide. The minimum size of sample that can support a chain reaction is called the **critical mass.**	• Describe the process in a nuclear fission reaction. *Related Exercises 39–46* • Write a nuclear equation for a given fission reaction. *Related Exercises 39–46*
18.8 Nuclear Fusion The process of combining two light nuclei into a single nucleus is called **nuclear fusion.** In the interior of the Sun at temperatures above 10,000,000°C, there is continuous nuclear fusion. In the laboratory, scientists have had limited success and to date have not been able to create continuous nuclear fusion. One promising fusion reaction is that involving two **deuterium,** $^{2}_{1}H$, nuclei. Another possibility is the fusion of deuterium and **tritium,** $^{3}_{1}H$, nuclei. The technological problems are considerable because the fusion process requires temperatures of millions of degrees. However, nuclear fusion does not produce radioactive waste and is therefore an attractive possibility as a future energy source.	• Describe the process in a nuclear fusion reaction. *Related Exercises 47–52* • Write a nuclear equation for a given fusion reaction. *Related Exercises 47–52*

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

______ **1.** the emission of particles or energy from an unstable nucleus
______ **2.** nuclear radiation identical to a helium-4 nucleus
______ **3.** nuclear radiation identical to an electron
______ **4.** nuclear radiation that has neither mass nor charge
______ **5.** the value that indicates the number of protons in a nucleus
______ **6.** the value that indicates the number of protons and neutrons in a nucleus
______ **7.** an atom with a specific number of protons and neutrons
______ **8.** a high-energy change involving the atomic nucleus
______ **9.** a shorthand representation using atomic notation to describe a nuclear reaction
______ **10.** nuclear radiation identical in mass but opposite in charge to that of an electron
______ **11.** a nuclear reaction in which a heavy nuclide attracts one of its inner core electrons into the nucleus
______ **12.** the stepwise disintegration of a radioactive nucleus until a stable nuclide is reached
______ **13.** the relationship of a decaying nuclide and the resulting nuclide that is produced
______ **14.** the number of radioactive nuclei in a sample that disintegrate in a given period of time, for example, 500 dpm
______ **15.** the amount of time required for 50% of the radioactive nuclei in a given sample to decay
______ **16.** an atom whose nucleus is unstable and decays radioactively
______ **17.** the conversion of one element to another by a nuclear reaction
______ **18.** a nuclear reaction in which a nucleus splits into two or more lighter nuclei
______ **19.** a fission reaction in which the neutrons released initiate a second reaction, which in turn initiate a third reaction, and so on
______ **20.** the minimum mass of a fissionable nuclide necessary to sustain a chain reaction

(a) activity *(Sec. 18.4)*
(b) alpha particle (α) *(Sec. 18.1)*
(c) atomic number (Z) *(Sec. 18.2)*
(d) beta particle (β^-) *(Sec. 18.1)*
(e) chain reaction *(Sec. 18.7)*
(f) critical mass *(Sec. 18.7)*
(g) deuterium *(Sec. 18.8)*
(h) electron capture (EC) *(Sec. 18.2)*
(i) gamma ray (γ) *(Sec. 18.1)*
(j) half-life ($t_{1/2}$) *(Sec. 18.4)*
(k) mass number (A) *(Sec. 18.2)*
(l) nuclear equation *(Sec. 18.2)*
(m) nuclear fission *(Sec. 18.7)*
(n) nuclear fusion *(Sec. 18.8)*
(o) nuclear reaction *(Sec. 18.2)*
(p) nuclide *(Sec. 18.2)*
(q) parent–daughter nuclides *(Sec. 18.3)*
(r) positron (β^+) *(Sec. 18.2)*
(s) radioactive decay series *(Sec. 18.3)*

______**21.** a nuclear reaction in which two light nuclei combine into a single nucleus
______**22.** the nuclide of hydrogen having one neutron
______**23.** the nuclide of hydrogen having two neutrons

(t) radioactivity *(Sec. 18.1)*
(u) radionuclide *(Sec. 18.5)*
(v) transmutation *(Sec. 18.6)*
(w) tritium *(Sec. 18.8)*

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

Natural Radioactivity (Sec. 18.1)

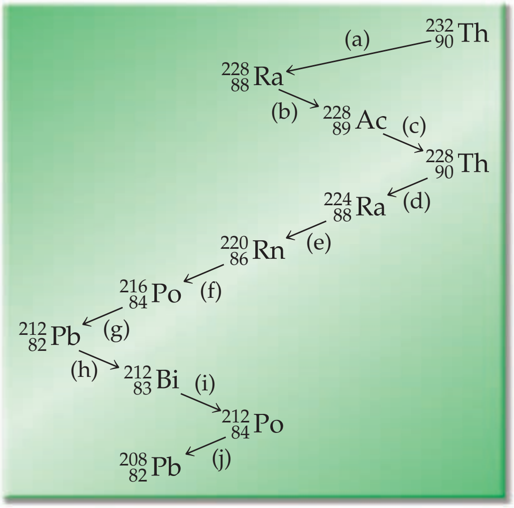
1. Which type of radiation is identical to a helium nucleus and is deflected toward the negative electrode as it passes between electrically charged plates?
2. Which type of radiation cannot penetrate human skin and requires heavy clothing as minimum protective shielding?
3. Which type of radiation is identical to an electron and is deflected toward the positive electrode as it passes between electrically charged plates?
4. Which type of radiation can penetrate about 1 cm of human flesh and requires wood or aluminum as minimum protective shielding?
5. Which type of radiation is powerful radiant energy that is *not* deflected as it passes between electrically charged plates?
6. Which type of radiation can pass through the human body and requires thick lead or concrete as minimum protective shielding?

Nuclear Equations (Sec. 18.2)

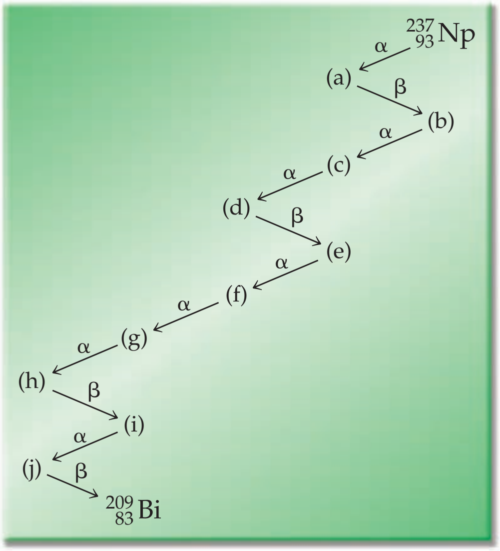
7. What is the approximate mass for each of the following?
   (a) alpha particle
   (b) beta particle
   (c) gamma ray
   (d) positron
   (e) neutron
   (f) proton
8. What is the relative charge for each of the following?
   (a) α
   (b) β^-
   (c) γ
   (d) β^+
   (e) n^0
   (f) p^+
9. Write an equation for each of the following natural radioactive decay processes.
   (a) U-238 decays by alpha emission
   (b) Al-28 decays by beta emission
   (c) O-15 decays by positron emission
   (d) Fe-55 decays by electron capture
10. Write an equation for each of the following natural radioactive decay processes.
   (a) W-160 decays by alpha emission
   (b) P-32 decays by beta emission
   (c) Co-55 decays by positron emission
   (d) Ti-44 decays by electron capture
11. Identify X, an unknown radioactive nuclide, given the following information.
   (a) Nuclide X decays by alpha emission to give ^{217}Rn.
   (b) Nuclide X decays by beta emission to give ^{43}Ca.
   (c) Nuclide X decays by positron emission to give ^{73}Br.
   (d) Nuclide X decays by electron capture to give ^{133}Cs.
12. Identify X, an unknown radioactive nuclide, given the following information.
   (a) Nuclide X decays by alpha emission to give ^{218}Ra.
   (b) Nuclide X decays by beta emission to give ^{56}Fe.
   (c) Nuclide X decays by positron emission to give ^{19}F.
   (d) Nuclide X decays by electron capture to give ^{37}Cl.

Radioactive Decay Series (Sec. 18.3)

13. The uranium-238 decay series begins with the emission of an alpha particle. The daughter product emits a beta particle to give which nuclide?
14. In the final step of the uranium-238 disintegration series, the parent nuclide decays into lead-206 and an alpha particle. Identify the parent nuclide.
15. The uranium-235 decay series begins with the emission of an alpha particle. The daughter product emits a beta particle to give which nuclide?
16. In the final step of the uranium-235 disintegration series, the parent nuclide decays into lead-207 and a beta particle. Identify the parent nuclide.
17. Supply each of the following emission particles in the 10-step decay series for radioactive thorium-232.



18. Supply each of the following decaying nuclides in the 11-step decay series for radioactive neptunium-237.



Radioactive Half-Life (Sec. 18.4)

19. What percentage of U-238 radionuclides in a sample remain after two half-lives?
20. What percentage of U-238 radionuclides in a sample remain after three half-lives?
21. If the carbon-14 reading of a fossil bone is 60 dpm, and a recent bone is 240 dpm, what is the estimated age of the fossil? ($t_{1/2} = 5730$ years)
22. If the carbon-14 radioactivity of an ancient wooden artifact is 6.25% that of a reference sample, what is the estimated age of the artifact? ($t_{1/2} = 5730$ years)
23. Sodium-24 in the form of NaCl is given as an injection to measure the sodium electrolyte balance. If 80 mg of the medical radionuclide is injected, how much Na-24 remains after 60 hours? ($t_{1/2} = 15$ hours)
24. If 160 mg of technetium-99 is administered for a medical diagnosis, how much of the nuclide remains after 24 hours? ($t_{1/2} = 6$ hours)

Applications of Radionuclides (Sec. 18.5)

25. What radionuclide technique can be used to estimate the age of fossils up to 50,000 years old?
26. What radionuclide technique can be used to estimate the age of geological events up to a few billion years?
27. What radionuclide can be used to irradiate food and destroy microorganisms?
28. What radionuclide can be used to sterilize male insects and control pests?
29. What γ-emitting radionuclide can be used to diagnose and locate brain tumors?
30. What γ-emitting radionuclide can be used to treat breast cancer?

Induced Radioactivity (Sec. 18.6)

31. Bombarding Na-23 with a proton gives a radioactive nuclide and a neutron. Identify the radionuclide.
32. Bombarding Li-6 with a neutron produces a radioactive nuclide and an alpha particle. Identify the radionuclide.
33. Bombarding Br-81 with gamma rays gives a radioactive nuclide and a neutron. Identify the radionuclide.
34. Bombarding U-238 with a hydrogen-2 nucleus produces a radioactive nuclide, a beta particle, and two neutrons. Identify the radionuclide.
35. In 1967, a Russian team smashed a neon nucleus into an americium target and claimed it had synthesized a new element, X. Given the equation for the reaction, what nuclide did the team create?

$$^{243}_{95}\text{Am} \quad + \quad ^{22}_{10}\text{Ne} \quad \rightarrow \quad ^{A}_{Z}\text{X} + \quad 5\,^{1}_{0}\text{n}$$

36. In 1970, a group at the University of California at Berkeley fired a nitrogen nucleus into a californium target and claimed it had synthesized a new element, X. Given the equation for the reaction, what nuclide did the group create?

$$^{249}_{98}\text{Cf} \quad + \quad ^{15}_{7}\text{N} \quad \rightarrow \quad ^{A}_{Z}\text{X} + 4\,^{1}_{0}\text{n}$$

37. In 1974, a team of Russian physicists smashed a chromium-54 nucleus into a lead-207 target. If a new element, X, and one neutron were produced from the collision, what nuclide did the team create?
38. In 1976, a team of German physicists smashed a chromium-54 nucleus into a bismuth-209 target. If a new element, X, and one neutron were produced from the collision, what nuclide did the team create?

Nuclear Fission (Sec. 18.7)

39. If a neutron causes a fission reaction that releases two neutrons, how many neutrons have been produced from the third fission step? (Assume each step in the fission process releases two neutrons.)
40. If a neutron causes a fission reaction that releases three neutrons, how many neutrons have been produced from the third fission step? (Assume each step in the fission process releases three neutrons.)
41. Why must uranium ore be enriched in U-235 before it can be used in the fuel rods of a nuclear reactor?
42. The fission of uranium-235 produces 2.4 neutrons per nucleus. Why is there a fractional value for the number of neutrons released?
43. How many neutrons are produced from the following fission reaction?

$$^{235}_{92}\text{U} \quad + \quad ^{1}_{0}\text{n} \quad \rightarrow \quad ^{144}_{54}\text{Xe} \quad + \quad ^{90}_{38}\text{Sr} \quad + \quad ^{1}_{0}\text{n}$$

44. How many neutrons are produced from the following fission reaction?

$$^{235}_{92}\text{U} \quad + \quad ^{1}_{0}\text{n} \quad \rightarrow \quad ^{142}_{56}\text{Ba} \quad + \quad ^{91}_{36}\text{Kr} \quad + \quad ^{1}_{0}\text{n}$$

45. What nuclide undergoes nuclear fission to give barium-143, krypton-88, and three neutrons?

46. What nuclide undergoes nuclear fission to give xenon-142, strontium-90, and two neutrons?

Nuclear Fission In a uranium-235 atomic bomb explosion, the large uranium nucleus fissions into two smaller fragments, and releases neutrons and energy.

Nuclear Fusion (Sec. 18.8)

47. The nuclear fusion of two deuterium nuclei gives a tritium nucleus, a positron, and particle X. Identify particle X.

48. The nuclear fusion of two helium-3 nuclei gives two protons and particle X. Identify particle X.

49. The nuclear fusion of a helium-3 nucleus and particle X releases an alpha particle and a positron. Identify particle X.

50. The nuclear fusion of a lithium-7 nucleus and particle X releases two alpha particles and a neutron. Identify particle X.

51. The nuclear fusion of two identical nuclides produces an alpha particle and a gamma ray. Identify the fusion nuclide.

52. The nuclear fusion of two identical nuclides produces a deuterium nucleus and a positron. Identify the fusion nuclide.

General Exercises

53. Crater Lake in Oregon was formed by volcanic explosion. If a tree charred by the explosion gave an activity of ~7.7 dpm and the half-life of C-14 is 5730 years, what is the approximate age of Crater Lake?

54. Uranium-lead dating of lunar rock samples gave a U-238 to Pb-206 ratio of ~1.0. If the half-life of U-238 is 4.5 billion years, what is the approximate age of the lunar rocks?

55. A radionuclide ionizes the air in a smoke detector, which in turn allows the flow of electricity. When smoke particles enter the detector, they disrupt the electric current and set off an alarm. What α-emitting radionuclide is used to ionize the air in a smoke detector? (*Hint:* The radionuclide decays into Np-237.)

56. A PET scan is a medical imaging technique for the circulatory system. First, a radionuclide is injected that decays by positron emission. When the positron is emitted, it strikes an electron, causing mutual annihilation and the release of two gamma rays. Write an equation for the annihilation reaction.

Challenge Exercises

57. The nuclear fusion of deuterium and tritium releases an alpha particle, energy, and nuclide X. Identify nuclide X.

58. The nuclear fission of lithium-6 by a neutron releases an alpha particle, energy, and nuclide X. Identify nuclide X.

Chapter 18 Self-Test

Answers to Self-Test are in Appendix J.

1. Which type of nuclear radiation is powerful light energy that is *not* deflected as it passes between electrically charged plates? (Sec. 18.1)

(a) alpha
(b) beta
(c) gamma
(d) all of the above
(e) none of the above

2. What is the approximate mass and relative charge of gamma radiation? (Sec. 18.2)

(a) 4 amu and 2+
(b) 1 amu and 1+
(c) 1 amu and 0
(d) 0 amu and 0
(e) 0 amu and 1−

3. What particle is emitted when an Ar-39 nucleus decays to K-39? (Sec. 18.2)

(a) alpha
(b) beta
(c) neutron
(d) positron
(e) proton

4. What nuclide is produced when Ar-39 decays by beta and gamma emission? (Sec. 18.2)

(a) ${}^{39}_{17}\mathrm{Cl}$
(b) ${}^{40}_{17}\mathrm{Cl}$
(c) ${}^{40}_{18}\mathrm{Ar}$
(d) ${}^{39}_{19}\mathrm{K}$
(e) ${}^{40}_{19}\mathrm{K}$

5. In the final step of the uranium-235 decay series, the parent nuclide decays to a lead-207 nuclide and a beta particle. What is the parent nuclide? (Sec. 18.3)

(a) ${}^{206}_{82}\mathrm{Pb}$
(b) ${}^{208}_{82}\mathrm{Pb}$
(c) ${}^{207}_{83}\mathrm{Bi}$
(d) ${}^{207}_{81}\mathrm{Tl}$
(e) none of the above

6. If an iron-59 sample has an initial activity of 160 dpm, how much time is required for the activity to drop to 20 dpm? ($t_{1/2} = 45$ days) (Sec. 18.4)

(a) 45 days
(b) 90 days
(c) 135 days
(d) 180 days
(e) 210 days

7. What is the principle assumption for radiocarbon dating to be reliable? (Sec. 18.5)

(a) The amount of nitrogen-14 in the atmosphere has remained constant.
(b) The amount of carbon-14 in the atmosphere has remained constant.
(c) The number of neutrons in the atmosphere has remained constant.
(d) The number of protons in the atmosphere has remained constant.
(e) The temperature of the atmosphere has remained constant.

8. In 1919 Rutherford performed the first transmutation of an element. He bombarded N-14 with an alpha particle and obtained a proton and what nuclide? (Sec. 18.6)
(a) fluorine-17 **(b)** fluorine-18
(c) oxygen-17 **(d)** oxygen-18
(e) none of the above

9. How many neutrons are produced from the following fission reaction? (Sec. 18.7)

$$^{235}_{92}\text{U} + ^{1}_{0}\text{n} \rightarrow ^{141}_{56}\text{Ba} + ^{91}_{36}\text{Kr} + ^{1}_{0}\text{n}$$

(a) 1 **(b)** 2
(c) 3 **(d)** 4
(e) none of the above

10. The nuclear fusion of a helium-3 nucleus and nuclide X releases an alpha particle and a positron. What is nuclide X? (Sec. 18.8)
(a) deuterium **(b)** tritium
(c) neutron **(d)** positron
(e) proton

Key Concepts

11. Write a nuclear equation for each of the following natural decay reactions.
(a) U-238 decays by alpha emission
(b) U-239 decays by beta emission
(c) K-40 decays by positron emission
(d) K-40 decays by electron capture

12. A single neutron causes uranium-235 to undergo nuclear fission and release three neutrons. Assuming each of the neutrons released causes a fission that releases three more neutrons, how many neutrons are released in the second step?

13. The nuclear fusion of two protons releases a deuterium nucleus and nuclide X. Identify nuclide X.

Critical Thinking

14. Write a nuclear equation corresponding to the following illustration of radioactive decay.

15. In 2010 a team of German physicists created an atom of Element 117. If the atom had a mass number of 294, and three neutrons were released from the collision of a Ca-48 projectile and a target nucleus, what was the target nucleus?

CHAPTER

19

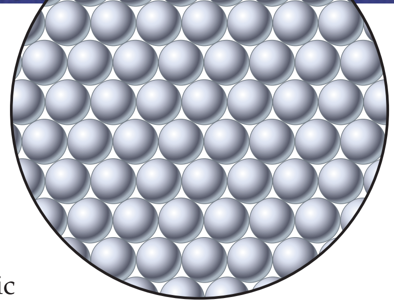
hemistry

"*Life exists in the universe only because the carbon atom possesses certain exceptional properties.*"

Sir James Jeans, English Physicist (1877–1946)

19
K
potassium

Element 19: Potassium
Potassium is a very reactive alkali metal. When dropped into water, it reacts violently to produce potassium hydroxide and hydrogen gas that ignites to give a flame.



In the early 1800s, chemists believed that inorganic compounds were different from organic compounds. Inorganic compounds originate from rocks and minerals. Organic compounds, on the other hand, originate from plants and animals. Because organic compounds are derived from living organisms, they were thought to contain a life force. This popular idea was referred to as the *vital force theory*.

In 1828, the German chemist Friedrich Wöhler (1800–1882) heated ammonium cyanate and obtained urea, a compound present in urine. The result was surprising because ammonium cyanate was considered an *inorganic* compound, whereas urea from animals is an *organic* compound. Wöhler's experiment helped disprove the vital force theory, although reluctant skeptics insisted that the vital force from Wöhler's hands had contaminated the result.

From simple and somewhat controversial beginnings, organic chemistry has grown to the point that many of our modern day materials and medicines are the result of research in organic chemistry. For example, plastics, fibers for fabrics and carpets, medical products, and parts for appliances, automobiles, and airplanes all derive from discoveries in organic chemistry. In addition, the molecules associated with animal and plant life are organic molecules.

LEARNING OBJECTIVES

- Classify a hydrocarbon having single, double, or triple bonds.

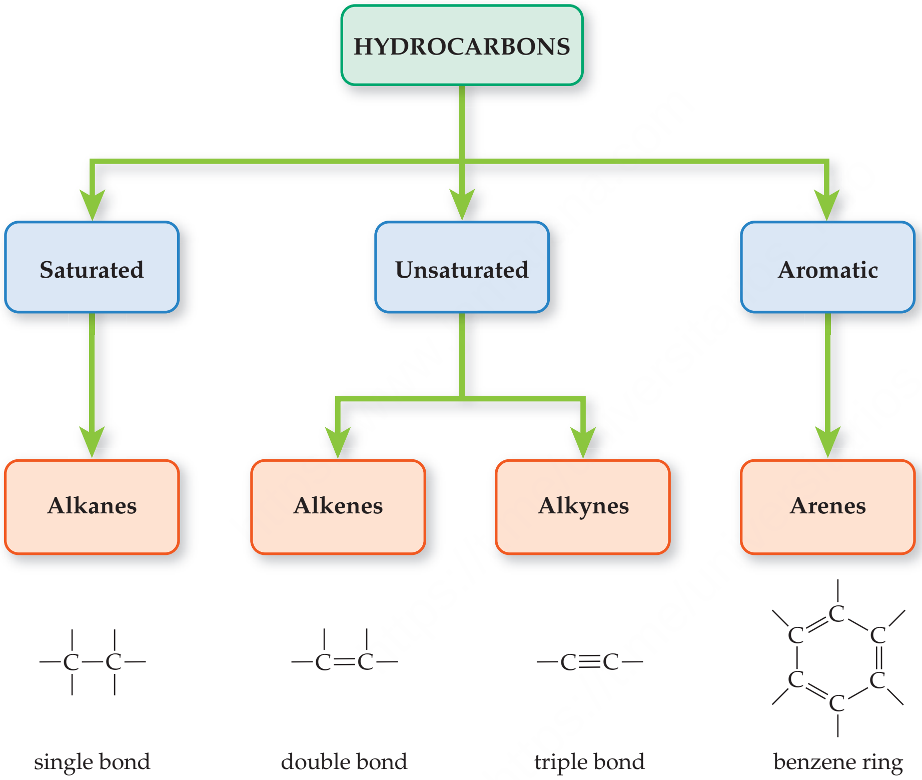
19.1 Hydrocarbons

The definition of **organic chemistry** is the study of compounds that contain carbon. Currently, about 50 million organic compounds have been identified and account for about 90% of all known substances. In addition, over 100,000 new organic compounds are identified each year. One reason that there are so many organic compounds is that carbon atoms have the ability to link together and form long chains. The chief sources of carbon are petroleum, natural gas, and coal, all of which are referred to as *fossil fuels.*

An organic compound composed of hydrogen and carbon is called a **hydrocarbon**. Hydrocarbons can be classified as saturated or unsaturated. A **saturated hydrocarbon** has a single bond between each of its carbon atoms. Since a carbon atom has four valence electrons, it can bond to four other atoms. That is, a carbon atom can form four single covalent bonds. If a hydrocarbon compound has all single bonds, it belongs to the *alkane* family.

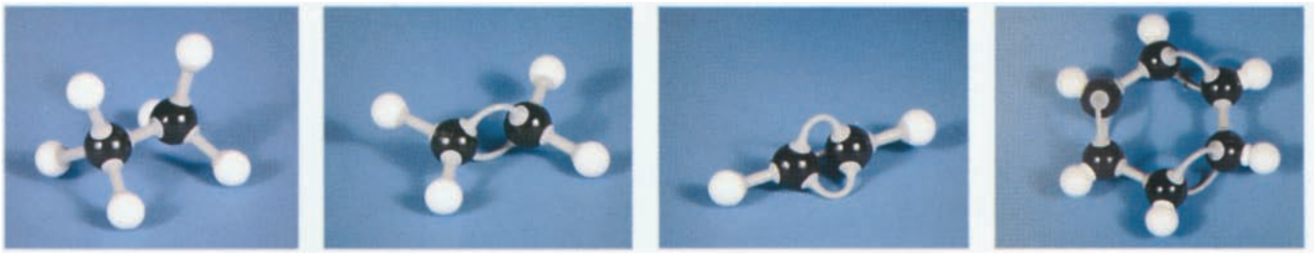
An **unsaturated hydrocarbon** has either a double bond or a triple bond between two carbon atoms. If a compound has a double bond, it is a member of the *alkene* family. If it has a triple bond, it belongs to the *alkyne* family. An **aromatic hydrocarbon** is usually characterized by having a benzene ring. If a compound contains a benzene ring, it is a member of the *arene* family. Figure 19.1 shows the classification of hydrocarbons.

- Classify a hydrocarbon as saturated, unsaturated, or aromatic.
- Classify a hydrocarbon as an alkane, alkene, alkyne, or arene.



◀ **Figure 19.1 Classification of Hydrocarbons** A hydrocarbon can be classified as an alkane, an alkene, an alkyne, or an arene.

We can construct molecular models of hydrocarbons to gain an appreciation of their three-dimensional structure. Ball-and-stick models are commonly used to show the arrangement of atoms in a molecule. Figure 19.2 illustrates the four classes of compounds with selected molecular models.


(a) (b) (c) (d)

▲ **Figure 19.2 Molecular Models of Hydrocarbons** The photographs show molecular models of (a) an alkane (saturated), (b) an alkene (unsaturated), (c) an alkyne (unsaturated), and (d) an arene (aromatic hydrocarbon). (C-black, H-white)

LEARNING OBJECTIVES

- Write names and formulas for simple alkanes.
- Write combustion reactions for alkanes.

19.2 Alkanes

Alkanes are a family of saturated hydrocarbons whose names end in the suffix −*ane*. The general molecular formula for the alkanes is C_nH_{2n+2}, where n is the number of carbon atoms. The alkane in Figure 19.2(a) has two carbon atoms ($n = 2$), six hydrogen atoms ($2n + 2 = 6$), and the molecular formula is C_2H_6. The sixth member of the alkane family has 6 carbon atoms ($n = 6$) and 14 hydrogen atoms ($2n + 2 = 14$) and the molecular formula is C_6H_{14}.

Alkane Family of Compounds

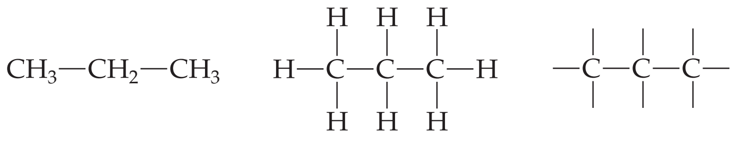
The smaller members of the alkane family, ranging from 1–10 carbons, are used as gaseous and liquid fuels. The larger members, ranging from 20–40 carbons, are waxy solids. These solid alkanes are referred to as *paraffin* and are used to make waxes and candles. Table 19.1 lists the first 10 members of the alkane family.

TABLE 19.1 Alkanes*

IUPAC Name*	Molecular Formula	Structural Formula	Boiling Point
methane	CH_4	CH_4	−161 °C
ethane	C_2H_6	$CH_3—CH_3$	−89 °C
propane	C_3H_8	$CH_3—CH_2—CH_3$	−44 °C
butane	C_4H_{10}	$CH_3—CH_2—CH_2—CH_3$	−1 °C
pentane	C_5H_{12}	$CH_3—CH_2—CH_2—CH_2—CH_3$	36 °C
hexane	C_6H_{14}	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_3$	68 °C
heptane	C_7H_{16}	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$	98 °C
octane	C_8H_{18}	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$	125 °C
nonane	C_9H_{20}	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$	151 °C
decane	$C_{10}H_{22}$	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$	174 °C

*The first four members of the alkanes are gases at normal temperature (25°C); pentane through decane are liquids.

The molecular and structural formulas of alkanes are shown in Table 19.1. For example, the molecular formula of propane is C_3H_8. The condensed **structural formula** is $CH_3—CH_2—CH_3$. To see the structure even more clearly, we can write an expanded structural formula. For simplicity, we can omit the hydrogen atoms and write a skeletal structural formula. Each of the following formulas can be used to represent a molecule of propane, C_3H_8.



condensed structural formula

expanded structural formula

skeletal structural formula

Helpful Hint

IUPAC vs. "Common Names"

In this textbook, a common name such as "isobutane" is placed in quotes to indicate it is not an official IUPAC name. The systematic IUPAC name for "isobutane" is 2–methylpropane.

Branched Hydrocarbon Chains

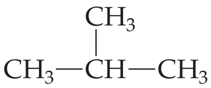
Each structural formula in Table 19.1 represents a molecule that is a straight chain and does not have branches. However, most carbon chains do have branches. When branching occurs, a different compound results. Two compounds that have the same molecular formula, but a different structural formula are called structural **isomers**.

Butane has the molecular formula C_4H_{10}. However, we can draw the structural formula in two ways: $CH_3CH_2CH_2CH_3$ and $CH_3CH(CH_3)CH_3$. These formulas represent the two isomers—butane and "isobutane." In addition to different structural formulas, isomers have different physical properties such as melting points and boiling points.

$CH_3—CH_2—CH_2—CH_3$

butane

Mp = –138 °C, Bp = –0.5 °C



$$CH_3—\overset{\overset{\displaystyle CH_3}{|}}{C}H—CH_3$$

“isobutane”

Mp = –159 °C, Bp = –12 °C

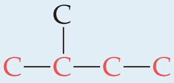
As the number of carbons in the molecule increases, the number of possible isomers increases. For example, C_5H_{12} has 3 isomers, C_6H_{14} has 5 isomers, and $C_{10}H_{22}$ has 75 isomers! To draw the structural isomers corresponding to a molecular formula, we use the following guidelines.

Guidelines for Drawing Hydrocarbon Isomers

1. Draw a continuous chain of carbon atoms. For C_5H_{12}, the chain is five carbon atoms long. To simplify, we omit hydrogen atoms and show the skeletal formula.

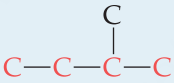
$$C—C—C—C—C$$

2. To construct a different isomer, draw a shorter continuous chain; in this case, the chain is now four carbon atoms long. Attach the fifth carbon to either carbon atom in the middle of the chain.



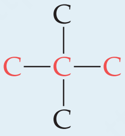
$$C—\overset{\overset{\displaystyle C}{|}}{C}—C—C$$

It makes no difference which middle carbon atom we attach the fifth carbon to since we can flip the molecule over without changing the structure. The branch is still on the second carbon from the end. The following structure is identical to the previous one.



$$C—C—\overset{\overset{\displaystyle C}{|}}{C}—C$$

3. To construct another isomer, shorten the chain to three carbon atoms. Attach the fourth and fifth carbon atoms to the central carbon atom.



$$C—\underset{\underset{\displaystyle C}{|}}{\overset{\overset{\displaystyle C}{|}}{C}}—C$$

4. For higher members of the alkane series that contain more than five carbons, apply a similar procedure. Although the process can become complex, this systematic procedure is very effective.

Alkyl and Aryl Groups

When a hydrogen atom is removed from an alkane, an **alkyl group** (symbol **R**—) results. Because an alkyl group is missing a hydrogen atom, it can bond to another atom or molecule. The name of the alkyl group is formed by changing the suffix —*ane* of the alkane to the suffix —*yl*. For example, methane becomes a methyl group.

$$\underset{\text{meth}\mathbf{ane}}{CH_4} \quad - \quad H \quad = \quad \underset{\text{meth}\mathbf{yl}}{CH_3—}$$

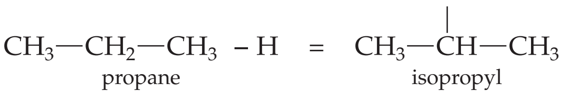
Similarly, removing a hydrogen atom from ethane gives an ethyl group.

$$\underset{\text{eth}\mathbf{ane}}{CH_3—CH_3} \quad - \quad H \quad = \quad \underset{\text{eth}\mathbf{yl}}{CH_3—CH_2—}$$

Removal of a hydrogen atom from propane can occur in two ways—from the end of the molecule or from the middle carbon atom. The structural formula of propane, $CH_3-CH_2-CH_3$, shows that there are six hydrogen atoms on the end carbons and two hydrogen atoms on the central carbon. Since the end carbon atoms are identical, we can remove any one of the six hydrogen atoms to produce a propyl group.

$$\underset{\text{prop}\textbf{ane}}{CH_3-CH_2-CH_3} \quad - \quad H \quad = \quad \underset{\text{prop}\textbf{yl}}{CH_3-CH_2-CH_2-}$$

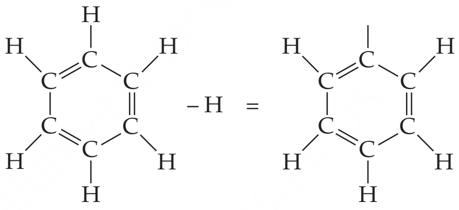
Since the middle two hydrogen atoms are equivalent, we can remove either of these atoms to form an isopropyl group.



$$\underset{\text{propane}}{CH_3-CH_2-CH_3} - H = \underset{\text{isopropyl}}{CH_3-\overset{|}{CH}-CH_3}$$

It is easy to see that we can substitute a chlorine atom onto a propane molecule at two different positions. If the chlorine is attached to an end carbon, the result is "propyl chloride," $CH_3-CH_2-CH_2-Cl$. If the chlorine atom is attached to the central carbon, the result is "isopropyl chloride," $CH_3-CH(Cl)-CH_3$. "Propyl chloride" and "isopropyl chloride" are isomers, and thus have different properties.

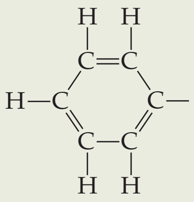
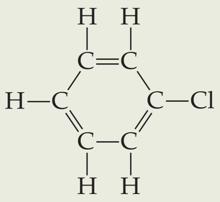
When a hydrogen atom is removed from an aromatic hydrocarbon, the result is an **aryl group** (symbol **Ar**—). The most common aryl group results from removal of a hydrogen atom from a benzene ring. The resulting aryl group is called a **phenyl group**.



benzene, C_6H_6 phenyl, C_6H_5-

Now, let's summarize the structures of alkyl groups. The names and formulas of selected alkyl and aryl groups are shown in Table 19.2.

TABLE 19.2 Alkyl and Aryl Groups

Name	Molecular Formula	Structural Formula	Example
methyl	CH_3-	CH_3-	CH_3-Cl "methyl chloride"
ethyl	C_2H_5-	CH_3-CH_2-	CH_3-CH_2-Cl "ethyl chloride"
propyl	C_3H_7-	$CH_3-CH_2-CH_2-$	$CH_3-CH_2-CH_2-Cl$ "propyl chloride"
isopropyl	C_3H_7-	$CH_3-\underset{\underset{CH_3}{\mid}}{CH}-$	$CH_3-\underset{\underset{CH_3}{\mid}}{CH}-Cl$ "isopropyl chloride"
phenyl	C_6H_5-		 "phenyl chloride"

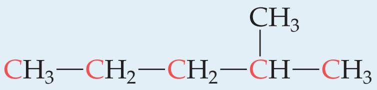
Nomenclature of Alkanes

In 1892, the International Union of Chemistry met in Geneva, Switzerland, and recommended a systematic set of rules for naming organic compounds. Although these rules are revised periodically by the International Union of Pure and Applied Chemistry (IUPAC), the basic guidelines for naming organic compounds are still referred to as the *Geneva convention*. In naming an alkane, we will apply the following nomenclature guidelines.

▲ **Organic Nomenclature** This Swiss stamp commemorates the 100th anniversary of systematic nomenclature rules for organic compounds first proposed by IUPAC in 1892 at a conference in Geneva, Switzerland.

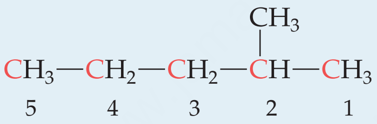
Guidelines for Naming Alkanes

1. Name an alkane for its longest continuous carbon chain, which is referred to as the "parent" chain. If the longest chain is five carbon atoms, the compound is a *pentane*. Regardless of the branches on the carbon chain, the alkane is named a pentane.



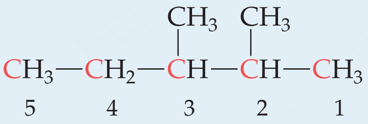
$$\mathrm{CH_3{-}CH_2{-}CH_2{-}\overset{\overset{CH_3}{|}}{C}H{-}CH_3}$$

2. Number the longest continuous chain of carbon atoms starting from the end closest to the first branch on the chain. In the preceding example, we number from right to left.



$$\mathrm{\underset{5}{CH_3}{-}\underset{4}{CH_2}{-}\underset{3}{CH_2}{-}\underset{2}{\overset{\overset{CH_3}{|}}{C}H}{-}\underset{1}{CH_3}}$$

3. Indicate the position of the alkyl group(s) by name and number. In the preceding example, a methyl group is on the second carbon atom. The IUPAC name is thus *2-methylpentane*.
4. If there are two or more of the same group attached to the chain, use the prefix *di–* for two, *tri–* for three, or *tetra–* for four.

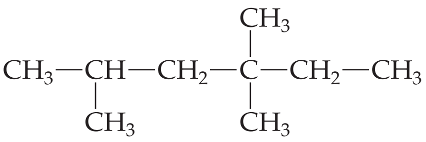


$$\mathrm{\underset{5}{CH_3}{-}\underset{4}{CH_2}{-}\underset{3}{\overset{\overset{CH_3}{|}}{C}H}{-}\underset{2}{\overset{\overset{CH_3}{|}}{C}H}{-}\underset{1}{CH_3}}$$

Since there are two methyl groups on the second and third carbon atoms, the name of the hydrocarbon is *2,3-dimethylpentane*.

EXAMPLE EXERCISE 19.1 Nomenclature of Alkanes

Give the IUPAC name for the following structure:



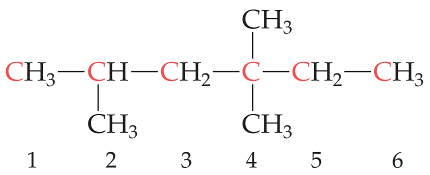
$$\mathrm{CH_3{-}\underset{\underset{CH_3}{|}}{C}H{-}CH_2{-}\underset{\underset{CH_3}{|}}{\overset{\overset{CH_3}{|}}{C}}{-}CH_2{-}CH_3}$$

Solution

Let's apply the systematic rules for naming alkanes.

1. The longest continuous carbon chain has six carbon atoms, and so the compound is a *hexane*.

2. Since the first branch is closest to the left end of the chain, we number from left to right.



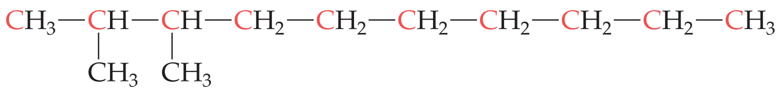
$$CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2-\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_3}{\underset{|}{C}}}}-CH_2-CH_3$$

3. Each of the three attached groups is a methyl. Thus, the compound is a *methylhexane*.
4. Since there are three methyl groups, we use the prefix *tri–* and indicate the position of each methyl. The IUPAC name for this hydrocarbon is *2,4,4-trimethylhexane*.

Practice Exercise

Draw the condensed structural formula for 2,3-dimethyldecane.

Answer:



$$CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$$

Concept Exercise

Which of the following compounds is an example of an alkane: decane ($C_{10}H_{22}$), decene ($C_{10}H_{20}$), decyne ($C_{10}H_{18}$), or naphthalene ($C_{10}H_8$)?

Answer: See Appendix G, 19.1.

Reactions of Alkanes

Since alkanes have strong C—C and C—H bonds, they are unreactive at room temperature. At high temperatures, however, alkanes burn rapidly with oxygen in the air. The reaction is called a **combustion reaction**, and burning alkanes supplies much of the world's energy.

The products from the complete combustion of an alkane are carbon dioxide and water. If the amount of oxygen is limited, carbon monoxide may be produced. For example, the equation for the complete combustion of methane gas is

$$CH_4(g) + 2\,O_2(g) \xrightarrow{\text{spark}} CO_2(g) + 2H_2O(g)$$

19.3 Alkenes and Alkynes

LEARNING OBJECTIVES

- Write names and formulas for simple alkenes.
- Write names and formulas for simple alkynes.
- Write addition reactions for alkenes and alkynes.

Alkenes are unsaturated hydrocarbons that contain a carbon–carbon double bond. The general molecular formula for alkenes with one double bond is C_nH_{2n}. The alkene in Figure 19.2(b) has two carbon atoms ($n = 2$) and four hydrogen atoms ($2n = 4$) and the molecular formula is C_2H_4. The sixth member of the alkene family has 6 carbon atoms ($n = 6$) and 12 hydrogen atoms ($2n = 12$) and the molecular formula is C_6H_{12}.

Alkynes are unsaturated hydrocarbons that contain a carbon–carbon triple bond. The general molecular formula for alkynes with one triple bond is C_nH_{2n-2}. The alkyne in Figure 19.2(c) has two carbon atoms ($n = 2$) and two hydrogen atoms ($2n - 2 = 2$) and the molecular formula is C_2H_2. The sixth member of the alkyne family has 6 carbon atoms ($n = 6$) and 10 hydrogen atoms ($2n - 2 = 10$) and the molecular formula is C_6H_{10}.

Alkene Family of Compounds

The first, smallest member of the alkene family is $H_2C{=}CH_2$, as shown in Figure 19.2(b). Its IUPAC name is ethene, but it is commonly called "ethylene." The second member of the alkene family is $CH_2{=}CH{-}CH_3$. Its IUPAC name is propene, but it is commonly referred to as "propylene." "Ethylene" and "propylene" are used to manufacture the plastics polyethylene and polypropylene.

The next member of the alkene family, C_4H_8, is butene. Butene has two structural isomers because there are two possible positions for the double bond. The double bond can be between the first and second carbon atoms, $CH_2{=}CH{-}CH_2{-}CH_3$, or between the second and third, $CH_3{-}CH{=}CH{-}CH_3$. The name of the first isomer is 1-butene and the name of the second is 2-butene. Table 19.3 lists the first few members of the alkene family.

TABLE 19.3 Alkenes

IUPAC Name	Common Name	Molecular Formula	Structural Formula	Boiling Point
ethene	"ethylene"	C_2H_4	$CH_2{=}CH_2$	−104 °C
propene	"propylene"	C_3H_6	$CH_2{=}CH{-}CH_3$	−47 °C
1-butene	"1-butylene"	C_4H_8	$CH_2{=}CH{-}CH_2{-}CH_3$	−6 °C
2-butene	"2-butylene"	C_4H_8	$CH_3{-}CH{=}CH{-}CH_3$	*

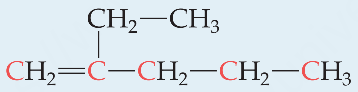
*There are two isomers of 2-butene, which have boiling points of 3 °C and 1 °C.

Nomenclature of Alkenes

An alkene is named after the corresponding alkane. That is, the name of an alkene is formed by changing the suffix −*ane* to the suffix −*ene*. For example, ethane is changed to ethene. The rules for naming alkenes are similar to the rules for naming alkanes. We will use the following nomenclature guidelines.

Guidelines for Naming Alkenes

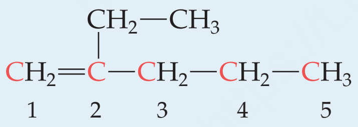
1. Name an alkene for the longest continuous carbon chain that contains the double bond. Consider the following alkene.



$$CH_2{=}\overset{\displaystyle CH_2{-}CH_3}{\overset{|}{C}}{-}CH_2{-}CH_2{-}CH_3$$

   Although the longest chain is six carbon atoms, only five carbon atoms are in the chain with the double bond. Thus, the compound is a *pentene*.

2. Number the longest continuous chain starting from the end closest to the double bond. Thus, we number from left to right to give *1-pentene*.



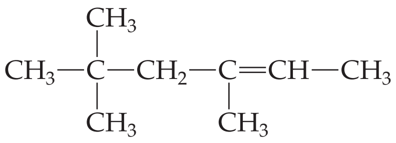
$$\underset{1}{CH_2}{=}\underset{2}{\overset{\displaystyle CH_2{-}CH_3}{\overset{|}{C}}}{-}\underset{3}{CH_2}{-}\underset{4}{CH_2}{-}\underset{5}{CH_3}$$

3. Indicate the position of each attached alkyl group(s) by name and number. Since there is an ethyl group, $CH_3{-}CH_2{-}$, on the second carbon, the compound is named *2-ethyl-1-pentene*.

4. If there are two or more of the same group on the chain, use the prefix *di*– for two, *tri*– for three, or *tetra*– for four. In this case there is one group, so the name remains *2-ethyl-1-pentene*.

EXAMPLE EXERCISE 19.2 Nomenclature of Alkenes

Give the IUPAC name for the following structure:

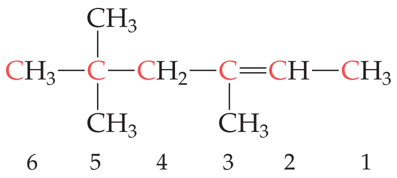


$$CH_3{-}\underset{\underset{\displaystyle CH_3}{|}}{\overset{\overset{\displaystyle CH_3}{|}}{C}}{-}CH_2{-}\underset{\underset{\displaystyle CH_3}{|}}{C}{=}CH{-}CH_3$$

Solution

Let's apply the systematic rules for naming alkenes.

1. The longest continuous carbon chain containing the double bond has six carbons, and so the compound is a *hexene.*
2. To assign the double bond the lowest value, we start numbering from the right. This compound is a *2-hexene*.

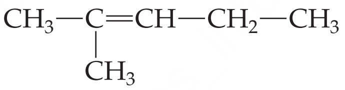

$$CH_3-\underset{CH_3}{\overset{CH_3}{C}}-CH_2-\underset{CH_3}{C}=CH-CH_3$$

3. Each of the three branches is a methyl group. Thus, the compound is a *methyl-2-hexene*.
4. Since there are three methyl groups, we use the prefix *tri–* and indicate the position of each group. The IUPAC name for this alkene is *3,5,5-trimethyl-2-hexene*.

Practice Exercise

Draw the condensed structural formula for 2-methyl-2-pentene.

Answer:


$$CH_3-\underset{CH_3}{C}=CH-CH_2-CH_3$$

Concept Exercise

Which of the following compounds is an example of an alkene: decane ($C_{10}H_{22}$), decene ($C_{10}H_{20}$), decyne ($C_{10}H_{18}$), or naphthalene ($C_{10}H_8$)?

Answer: See Appendix G, 19.2.

Alkyne Family of Compounds

The first, smallest member of the alkyne family is $CH{\equiv}CH$, see Figure 19.2(c). Its IUPAC name is ethyne, but it is commonly called "acetylene." Oxygen and acetylene gases are used for oxyacetylene welding. When oxygen and acetylene gases are ignited, the temperature reaches 3000 °C. The second member of the alkyne family is $CH{\equiv}C-CH_3$. Its IUPAC name is propyne, but it is commonly referred to as "methyl acetylene."

The next member of the alkyne family, C_4H_6, is butyne. Butyne has two structural isomers because there are two possible positions for the triple bond. The triple bond can be between the first and second carbon atoms, $CH{\equiv}C-CH_2-CH_3$, or between the second and third, $CH_3-C{\equiv}C-CH_3$. The name of the first isomer is 1-butyne and the name of the second is 2-butyne. Table 19.4 lists the first few members of the alkyne family.

TABLE 19.4 Alkynes

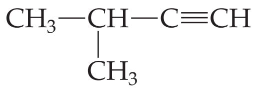
IUPAC Name	Common Name	Molecular Formula	Structural Formula	Boiling Point
ethyne	"acetylene"	C_2H_2	$CH{\equiv}CH$	−84 °C
propyne	"methyl acetylene"	C_3H_4	$CH_3-C{\equiv}CH$	−23 °C
1-butyne	"ethyl acetylene"	C_4H_6	$CH_3-CH_2-C{\equiv}CH$	8 °C
2-butyne	"dimethyl acetylene"	C_4H_6	$CH_3-C{\equiv}C-CH_3$	27 °C

Nomenclature of Alkynes

Like alkenes, alkynes are named after the corresponding alkane. That is, the name of an alkyne is formed by changing the suffix *−ane* to the suffix *−yne*. For example, ethane is changed to ethyne. The rules for naming complex alkynes are similar to the general rules for naming alkenes. The following example exercise illustrates these nomenclature rules.

EXAMPLE EXERCISE 19.3 Nomenclature of Alkynes

Write the IUPAC name for the following structure:

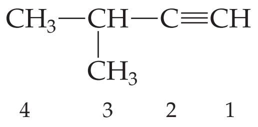


$$CH_3{-}\underset{\displaystyle CH_3}{\underset{|}{CH}}{-}C{\equiv}CH$$

Solution

We can follow the systematic rules for naming alkenes, but we substitute the suffix *−yne* for alkynes.

1. The longest continuous carbon chain containing the triple bond has four carbons, and so the compound is a *butyne*.
2. To assign the triple bond the lowest value, we start numbering from the right. This compound is *1-butyne*.



$$CH_3{-}\underset{\displaystyle CH_3}{\underset{|}{CH}}{-}C{\equiv}CH$$

3. Since there is a methyl group on the third carbon, the complete name of the compound is *3-methyl-1-butyne*.

Practice Exercise

Draw the condensed structural formula for 1-pentyne.

Answer: $CH{\equiv}C{-}CH_2{-}CH_2{-}CH_3$

Concept Exercise

Which of the following compounds is an example of an alkyne: decane ($C_{10}H_{22}$), decene ($C_{10}H_{20}$), decyne ($C_{10}H_{18}$), or naphthalene ($C_{10}H_8$)?

Answer: See Appendix G, 19.3.

Reactions of Alkenes and Alkynes

Like alkanes, alkenes and alkynes also burn with oxygen to provide energy. For example, acetylene, C_2H_2, and oxygen gases give the following combustion reaction in oxyacetylene welding.

$$2\,C_2H_2(g) + 5\,O_2(g) \xrightarrow{\text{spark}} 4\,CO_2(g) + 2\,H_2O(g)$$

Alkenes and alkynes are more reactive than alkanes because of their double and triple bonds. Unlike an alkane, an alkene or alkyne has unsaturated bonds that can react with hydrogen gas. When an alkene or alkyne adds hydrogen gas, it is referred to as a **hydrogenation reaction**. For example, when an unsaturated vegetable oil undergoes hydrogenation, it is converted to a semisolid margarine. Hydrogenation reactions occur readily at room temperature and atmospheric pressure when a metal catalyst, such as Ni or Pt, is introduced.

Consider the following stepwise addition of hydrogen gas to acetylene. First, ethyne is converted to ethene; then, ethene is converted to ethane.

$$\underset{\text{ethyne}}{CH{\equiv}CH(g)} + H_2(g) \xrightarrow{\text{Ni}} \underset{\text{ethene}}{CH_2{=}CH_2(g)}$$

$$\underset{\text{ethene}}{CH_2{=}CH_2(g)} + H_2(g) \xrightarrow{\text{Ni}} \underset{\text{ethane}}{CH_3{-}CH_3(g)}$$

Polymerization Reactions of Alkenes

A **polymer** is a giant molecule made up of many small molecules, each of which is called a **monomer**. A polymer can be formed when an alkene molecule reacts with a second molecule, which in turn reacts with a third molecule, and so on. For instance, “ethylene,” $CH_2{=}CH_2$, can react to give a giant molecule called *polyethylene.* We can represent the polymer as $[-CH_2-CH_2-]_n$, where *n* indicates the number of repeating “ethylene” monomers. During the polymerization reaction, the double bond in an “ethylene” molecule opens and links to another “ethylene” molecule as follows:

$$\begin{array}{ccccccc} \textit{monomer} & + & \textit{monomer} & + & \ldots\ldots\ldots \rightarrow & \textit{polymer} \\ CH_2{=}CH_2 & + & CH_2{=}CH_2 & + & \ldots\ldots\ldots \rightarrow & [-CH_2-CH_2-]_n \\ \textit{ethylene} & + & \textit{ethylene} & + & \ldots\ldots\ldots \rightarrow & \textit{polyethylene} \end{array}$$

LEARNING OBJECTIVE

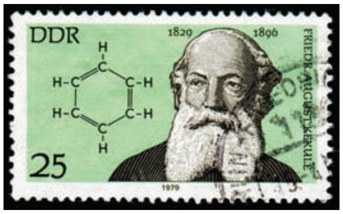
- Write names and formulas for simple arenes.

19.4 Arenes

In 1825, the English scientist Michael Faraday (1791–1867) isolated a substance from the fuel oil used in lamps. After analysis, the formula of the compound was determined to be C_6H_6. A few years later, the same compound was isolated from the natural product benzoin, a topical antiseptic. This compound, C_6H_6, was originally called benzin, but the name evolved into benzine, and eventually benzene.

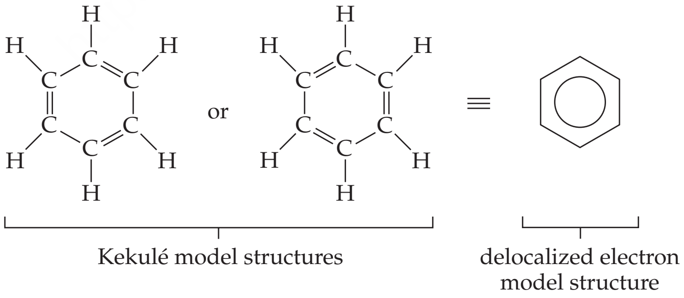
Arene Family of Compounds

As other organic compounds were isolated and identified, chemists observed that those containing a benzene structure often had a fragrant odor. Thus, these compounds were termed *aromatic* because of their pleasant aroma. Further analysis revealed that the benzene structure and the fragrant odor were not related. The term persisted, however, and today an aromatic hydrocarbon is one that usually contains the benzene structure. **Arenes** are a family of aromatic hydrocarbon compounds whose names end in the suffix −*ene*.



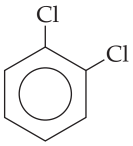
▲ **August Kekulé** The stamp commemorates Kekulé’s proposal for the cyclic structure of benzene in 1865.

In 1865, German chemist August Kekulé (1829–1896) proposed that benzene, C_6H_6, had a cyclic structure with three double bonds. He suggested that six carbon atoms formed a ring with alternating single and double bonds. The structure is called the *Kekulé structure of benzene*. Since the double bonds are not in fixed positions and shift between carbon atoms, the electrons are said to be delocalized. Because of this phenomenon, and for simplicity, the benzene ring is usually shown by a hexagon enclosing a circle to represent C_6H_6.



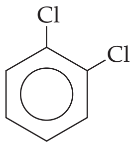
Isomers of Benzene

Using the delocalized electron model of benzene, we can represent the structure of dichlorobenzene as follows:

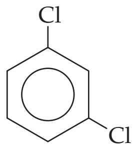


Notice that the two chlorine atoms are next to each other on the benzene ring. When two groups are adjacent on a benzene ring, their positions are indicated by the prefix *ortho*–. If the groups are separated by a carbon, their positions are indicated by the prefix *meta*-. If the groups are opposite each other on the benzene ring, their positions are indicated by the prefix *para*-.

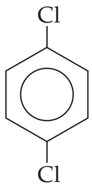
Since *ortho*-, *meta*-, and *para*- compounds have the same molecular formula but are different, they are structural isomers. Recall that isomers have different physical properties. We can illustrate these isomers as follows:



ortho-dichlorobenzene
Mp = –17 °C, Bp = 181 °C



meta-dichlorobenzene
Mp = 25 °C, Bp = 174 °C



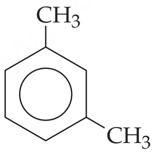
para-dichlorobenzene
Mp = 53 °C, Bp = 175 °C

EXAMPLE EXERCISE 19.4 Benzene Isomers

Draw the structural formula for *meta*-dimethylbenzene.

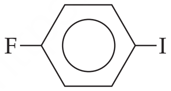
Solution

The compound contains a benzene ring. It also has two $—CH_3$ groups on the benzene ring in the *meta* position.



Practice Exercise

Give the systematic IUPAC name for the following:



Answer: *para*-fluoroiodobenzene

Concept Exercise

Which of the following compounds is an example of an arene: decane ($C_{10}H_{22}$), decene ($C_{10}H_{20}$), decyne ($C_{10}H_{18}$), or naphthalene ($C_{10}H_8$)?

Answer: See Appendix G, 19.4.

EXAMPLE EXERCISE 19.5 Hydrocarbons

Identify each of the following compounds as an alkane, alkene, or alkyne based on the name of the hydrocarbon.

(a) cyclopropane
(b) methylpropyne

Solution

(a) Cyclopropane has an –*ane* suffix; thus, the compound is an *alkane*.
(b) Methylpropyne has a –*yne* suffix; thus, the compound is an *alkyne*.

Practice Exercise

Identify each of the following compounds as an alkane, alkene, or alkyne based on the name of the hydrocarbon.

(a) cyclohexene
(b) ethylpentyne

Answers:

(a) alkene (b) alkyne

Concept Exercise

Is it possible to distinguish an alkene from an arene based on the hydrocarbon suffix?

Answer:

See Appendix G, 19.5.

19.5 Hydrocarbon Derivatives

LEARNING OBJECTIVES

- Describe the concept of an organic functional group.
- Identify the functional group in each of the following: organic halide, alcohol, phenol, ether, and amine.
- Identify the functional group in each of the following: aldehyde, ketone, carboxylic acid, ester, and amide.

In the 1800s, chemists began to realize that there was a huge number of organic compounds. As more compounds were identified and studied, the subject became overwhelming. According to the German chemist Friedrich Wöhler, organic chemistry seemed to be a "dreadful endless jungle." Gradually, chemists began to realize that organic compounds could be classified into only a few categories. Each of these categories is referred to as a **class of compounds** and all the compounds in a given class have similar names and properties.

Hydrocarbon Derivatives

To comprehend the vast subject of organic chemistry, chemists have divided a massive amount of information into divisions of organic compounds. The first division separates hydrocarbons and their derivatives. A hydrocarbon contains only hydrogen and carbon. A **hydrocarbon derivative** is derived from a hydrocarbon, but contains other elements, such as oxygen, nitrogen, or halogen (F, Cl, Br, I).

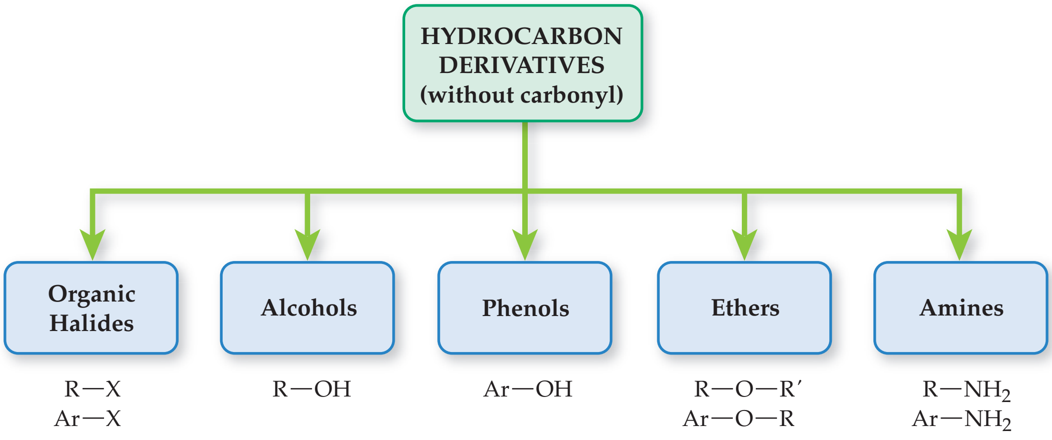
We will study 10 classes of hydrocarbon derivatives. Our task is simplified by the fact that each member in a class has similar names and properties. The classes are organic halides (R—X), alcohols (R—OH), phenols (Ar—OH), ethers (R—O—R′), amines ($R—NH_2$), aldehydes (R—CHO), ketones (R—COR′), carboxylic acids (R—COOH), esters (R—COOR′), and amides ($R—CONH_2$). The use of R and R′ indicates that the groups may or may not be the same.

The structural feature in a molecule that characterizes a class of compounds is called the **functional group**. For example, in an alcohol (R—OH), the functional group is —OH attached to a carbon atom. For an ether (R—O—R′), the functional group is the oxygen atom between two carbon atoms.

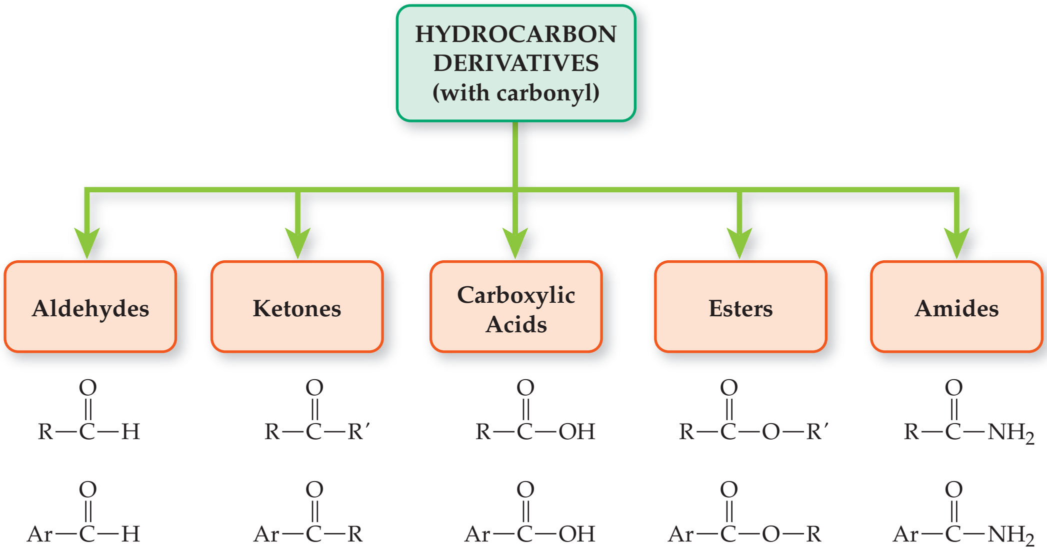
We can further divide these 10 hydrocarbon derivatives into five classes that have a carbon atom double bonded to an oxygen atom (C═O), and five classes that do not (Figures 19.3 and 19.4). A carbon atom joined to an oxygen atom by a double bond is called a **carbonyl group**. The presence of a carbonyl group gives distinct properties to a given class of compounds, such as amides that have a carbonyl group and amines that do not.

Helpful Hint
Alkyl (R) vs. Aryl (Ar) Groups
The symbol **R** represents an alkyl group and **Ar** represents an aryl group. The designations R and R′ indicate different alkyl groups.

Helpful Hint
Multiple Functional Groups
An organic compound may contain more than one functional group. Thus, an organic compound may represent two or more hydrocarbon derivatives.



▶ **Figure 19.3 Hydrocarbon Derivatives (without Carbonyl)** For organic halides, X can be F, Cl, Br, or I. For alcohols, R can be any alkyl group that is not an aryl (Ar); for phenols, the —OH group must be attached directly to an aryl group.



▲ **Figure 19.4** **Hydrocarbon Derivatives (with Carbonyl)** Aldehydes, ketones, carboxylic acids, esters, and amides all contain a carbonyl group with different attached groups. The symbol R can be any alkyl group, and Ar can be any aryl group. (R can also be a hydrogen atom except in ketones.)

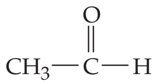
EXAMPLE EXERCISE 19.6 Classifying Hydrocarbon Derivatives

Indicate the class of hydrocarbon derivative for each of the following compounds:

(a) $CH_3—Br$

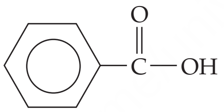
(b) $CH_3—CH_2—NH_2$

(c)



$\mathrm{CH_3-\overset{\displaystyle O}{\overset{\|}{C}}-H}$

(d)



Solution

Let's first identify the functional group in each compound. We can then refer to the general formulas to determine the class of compound.

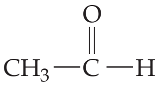
(a) This hydrocarbon derivative does not have a carbonyl group. Since the functional group is a halogen (—Br),

$CH_3—Br$ is an *organic halide.*

(b) This hydrocarbon derivative does not have a carbonyl group. Since the functional group is an amine ($—NH_2$),

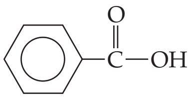
$CH_3—CH_2—NH_2$ is an *amine.*

(c) This hydrocarbon derivative has a carbonyl group. Since the functional group has a carbonyl attached to a H atom,



$\mathrm{CH_3-\overset{\displaystyle O}{\overset{\|}{C}}-H}$ is an *aldehyde.*

(d) This hydrocarbon derivative has a carbonyl group attached to a benzene ring. Since the functional group has a carbonyl attached to an OH,



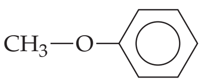
is a *carboxylic acid.*

Practice Exercise

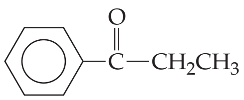
Indicate the class of hydrocarbon derivative for each of the following compounds.

(a) $(CH_3)_2CH—OH$

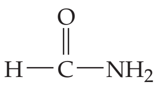
(b)



(c)



(d)



$$H—\overset{\overset{\large O}{\|}}{C}—NH_2$$

Answers: (a) alcohol (b) ether (c) ketone (d) amide

Concept Exercise

Which classes of hydrocarbon derivatives contain a carbonyl group? Which classes of hydrocarbon derivatives do not contain a carbonyl group?

Answers: See Appendix G, 19.6.

EXAMPLE EXERCISE 19.7 Hydrocarbon Derivatives

Identify each of the following compounds as an organic halide, alcohol, phenol, ether, or amine based on the name of the hydrocarbon derivative.

(a) caffeine
(b) chloroform
(c) cholesterol
(d) diethylethe

Solution

(a) Caffeine has an −*ine* suffix; thus, the compound is an *amine*.
(b) Chloroform has −*chloro* in the name; thus, the compound is an *organic halide*.
(c) Cholesterol has an −*ol* suffix; thus, the compound is an *alcohol*, or *phenol*.
(d) Diethylether has −*ether* in the name; thus, the compound is an *ether*.

Practice Exercise

Identify each of the following compounds as an aldehyde, ketone, carboxylic acid, ester, or amide based on the name of the hydrocarbon derivative.

(a) cinnamal
(b) cortisone
(c) methylacetate
(d) sulfanilamide

Answers: (a) aldehyde (b) ketone (c) ester (d) amide

Concept Exercise

Is it possible to distinguish an alcohol from a phenol based on the hydrocarbon suffix?

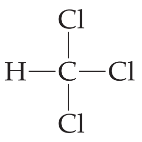
Answer: See Appendix G, 19.7.

LEARNING OBJECTIVE

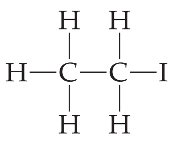
- Write names and formulas for simple organic halides.

19.6 Organic Halides

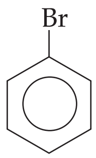
If a halogen (F, Cl, Br, I) atom replaces a hydrogen atom in a hydrocarbon, it gives rise to a class of compounds called *organic halides* (Figure 19.5). In this class of compounds, the hydrocarbon group can be alkyl, R—X, or aryl, Ar—X. For example,



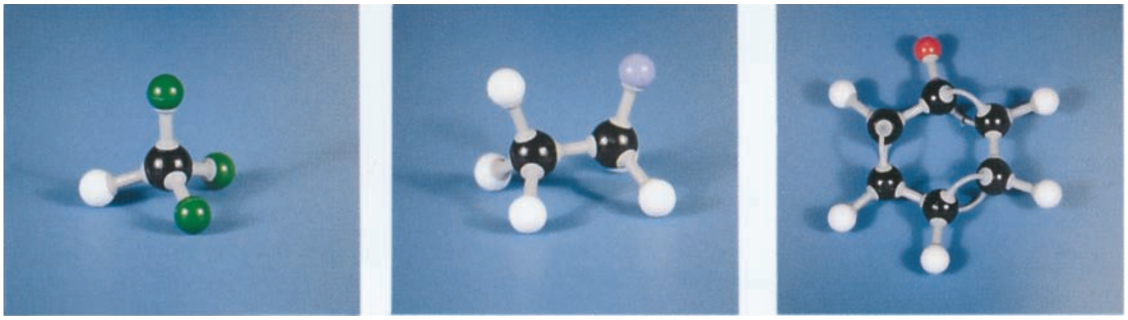
(a) "chloroform"



(b) "ethyl iodide"



(c) "phenyl bromide"


(a) (b) (c)

▲ **Figure 19.5** **Molecular Models of Organic Halides** Model structures of (a) "chloroform," $CHCl_3$; (b) "ethyl iodide," C_2H_5I; and (c) "phenyl bromide," C_6H_5Br. (Cl-green, I-purple, Br-orange)

Organic halides are used primarily as industrial and household solvents. Although "carbon tetrachloride," CCl_4, was once widely used for dry cleaning and spot removing, it has been replaced because of its toxicity. Now less toxic solvents, such as "trichloroethane," $C_2H_3Cl_3$, are used in dry cleaning.

"Chloroform," $CHCl_3$, is a common solvent originally used for general anesthesia. When its vapor was found to be harmful to the respiratory system, it was replaced by safer anesthetics such as halothane, $CF_3CHClBr$. "Ethyl chloride," CH_3CH_2Cl, is used as a mild topical anesthetic. After it is sprayed on the skin, ethyl chloride rapidly evaporates and produces a cold sensation that numbs the skin.

Chlorinated fluorocarbons (CFCs) are organic halides composed of chlorine, fluorine, and carbon. The best known example is Freon–12, CF_2Cl_2, which has been used as an aerosol propellant and a refrigerant gas. Carbon dioxide has replaced Freon-12 in most aerosol cans. Currently, CFCs are being replaced by hydrofluorocarbons (HFCs) in refrigeration and air-conditioning units. This change was mandated after CFCs in the upper atmosphere were found to be responsible for depletion of the ozone layer.

Organic halides are found in many pesticides, including Aldrin, Chlordane, and DDT. In 1972, the Environmental Protection Agency banned DDT because of its threat to wildlife. Before the ban, DDT was responsible for boosting agricultural harvests and reducing insect populations that carried malaria and yellow fever.

According to IUPAC nomenclature, organic halides are named for the halogen attached to the parent alkane. For example, CH_3CH_2Br is named bromoethane. Alternatively, we can designate an acceptable common name for the compound by stating the alkyl group attached to the bromine atom; that is, we can also name CH_3CH_2Br "ethyl bromide."

Most organic halides are nonpolar, although some are slightly polar. In general, organic halides have physical properties similar to those of alkanes. That is, they have low boiling points, are soluble in hydrocarbon solvents, and are insoluble in water. Table 19.5 briefly describes a few simple organic halides.

TABLE 19.5 Organic Halides

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
chloromethane	"methyl chloride"	$CH_3—Cl$	−24 °C	insoluble
chloroethane	"ethyl chloride"	$CH_3—CH_2—Cl$	12 °C	insoluble
1-chloropropane	"propyl chloride"	$CH_3—CH_2—CH_2—Cl$	47 °C	insoluble
2-chloropropane	"isopropyl chloride"	$CH_3—CH(Cl)—CH_3$	36 °C	insoluble
chlorobenzene	"phenyl chloride"	$C_6H_5—Cl$	132 °C	insoluble

*Most organic halides with low molecular mass are liquids at room temperature.

LEARNING OBJECTIVES

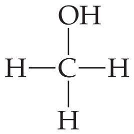
- Write names and formulas for simple alcohols.
- Write names and formulas for simple phenols.
- Write names and formulas for simple ethers.

19.7 Alcohols, Phenols, and Ethers

Alcohols, phenols, and ethers all contain oxygen in addition to hydrogen and carbon. Alcohols and phenols contain an —OH group and, in general, are polar. Ethers, however, contain a C—O—C functional group and are mostly nonpolar.

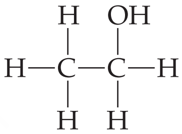
Alcohol and Phenol Families of Compounds

When an —OH group replaces a hydrogen atom on an alkane, the result is an *alcohol*, R—OH. When an —OH group replaces a hydrogen atom on a benzene ring, the result is a *phenol*, Ar—OH. The —OH group on an alcohol or a phenol is called a **hydroxyl group** (Figure 19.6). For example,



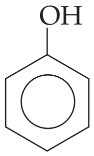
$$\begin{array}{ccccc} & & \mathrm{OH} & & \\ & & | & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{H} \\ & & | & & \\ & & \mathrm{H} & & \end{array}$$

(a) “methyl alcohol”

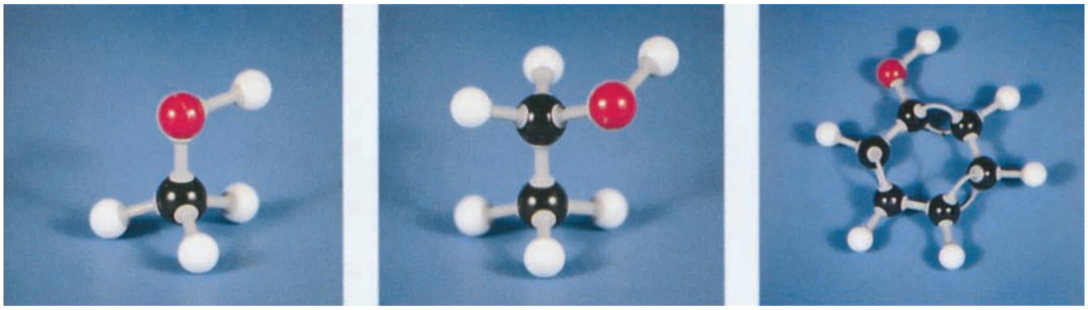


$$\begin{array}{ccccccc} & & \mathrm{H} & & \mathrm{OH} & & \\ & & | & & | & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{C} & - & \mathrm{H} \\ & & | & & | & & \\ & & \mathrm{H} & & \mathrm{H} & & \end{array}$$

(b) “ethyl alcohol”



(c) phenol



(a) (b) (c)

▲ **Figure 19.6** **Molecular Models of Alcohols and Phenol** Model structures of (a) “methyl alcohol,” CH_3OH; (b) “ethyl alcohol,” C_2H_5OH; and (c) phenol, C_6H_5OH. (C-black, H-white, O-red)

Alcohols are everyday substances with many practical uses. The simplest family member, “methyl alcohol,” CH_3OH, is often used as an antifreeze, solvent, and in the production of biodiesel fuel. It is referred to as “wood alcohol” because it can be produced from the distillation of wood. “Wood alcohol” has a infamous history because it is found in “moonshine” whiskey. Unlike grain alcohol, small amounts of wood alcohol can cause blindness, and even death, if taken internally.

“Ethyl alcohol,” CH_3CH_2OH, is blended with gasoline to give gasohol. In fact, the popular E85 fuel at gas stations is 85% ethanol. It is sometimes called *grain alcohol* because it is produced by the fermentation of sugars found in various grains such as corn, rye, and barley. “Isopropyl alcohol,” $CH_3CH(OH)CH_3$, is common in most homes and is referred to as rubbing alcohol.

Alcohols may have two or more hydroxyl groups in the same molecule. Ethylene glycol, $CH_2(OH)CH_2(OH)$, and glycerol, $CH_2(OH)CH(OH)CH_2(OH)$, are important examples. Ethylene glycol is the main ingredient in permanent automobile antifreeze. Glycerol, also called *glycerin*, is a syrupy liquid used as a moisturizer in cosmetics.

According to IUPAC nomenclature, the name of an alcohol is derived from the name of the parent alkane. The name is formed by dropping the final −*e* and adding the suffix −*ol*. For example, CH_3OH has an —OH group on a methane molecule; the IUPAC name is methanol, and the common name is “methyl alcohol.” In CH_3CH_2OH, the —OH is on an ethane molecule; the systematic name is ethanol, and the common name is “ethyl alcohol.”

If a hydroxyl group is attached to a benzene ring, the resulting alcohol belongs to a special class of compounds called *phenols*, Ar—OH. The simplest member of this class, phenol, C_6H_5OH, is an important industrial chemical. Phenol is used in the manufacture of plastics and for the preparation of dyes. It is somewhat acidic and is referred to as "carbolic acid." Phenol derivatives have antimicrobial properties and are used in medical and dental offices, and in cosmetics as a preservative.

Most alcohols are quite polar because of the —OH group. Alcohols, like water, can form hydrogen bonds between an oxygen atom and a hydrogen atom in separate molecules. Figure 19.7 depicts hydrogen bonding between alcohol molecules.

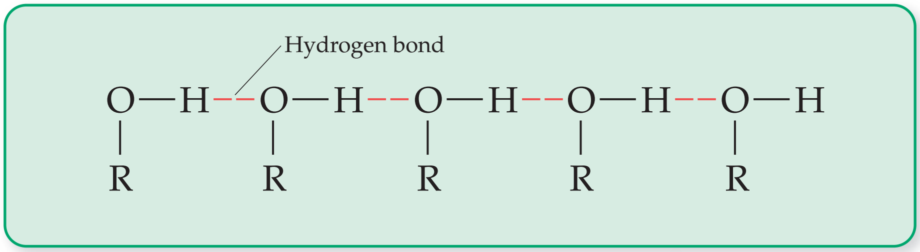


Figure 19.7 Hydrogen Bonding in Alcohols Notice that hydrogen bonds connect alcohol molecules in a long chain.

As a result of hydrogen bonding, the physical properties of alcohols are quite different from those of the parent alkane. Alcohols have higher boiling points than alkanes, and are soluble, or partially soluble, in polar solvents such as water. Table 19.6 lists a few simple alcohols as well as phenol.

TABLE 19.6 Alcohols and Phenol

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
methanol	"methyl alcohol"	$CH_3—OH$	65 °C	soluble
ethanol	"ethyl alcohol"	$CH_3—CH_2—OH$	78 °C	soluble
1-propanol	"propyl alcohol"	$CH_3—CH_2—CH_2—OH$	97 °C	soluble
2-propanol	"isopropyl alcohol"	$CH_3—CH(OH)—CH_3$	82 °C	soluble
phenol	"carbolic acid"	$C_6H_5—OH$	182 °C	~7%

*Most alcohols with low molecular mass are liquids at room temperature; phenol is a solid.

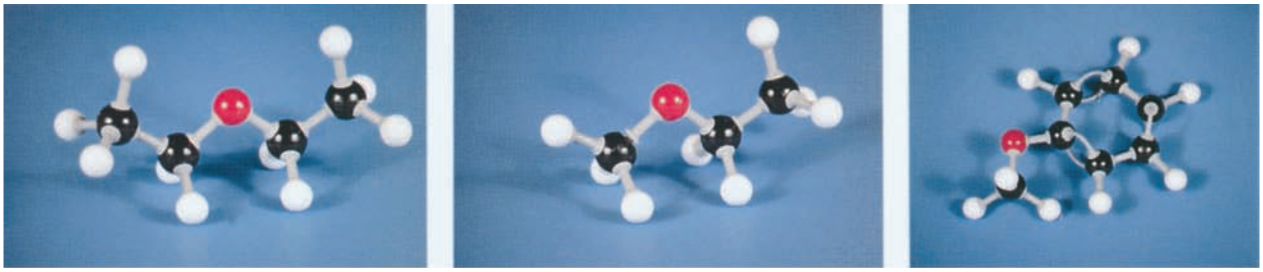
Ether Family of Compounds

If two alkyl groups are attached to an oxygen atom, we have a class of compounds called *ethers*, R—O—R′. Commercially, the most significant member of this class is diethyl ether, $CH_3CH_2—O—CH_2CH_3$. It is usually referred to as "ethyl ether," or simply "ether."

"Ethyl ether" is a common laboratory solvent for organic reactions. It was once used as a general surgical anesthetic, but has been replaced because it is highly flammable and causes patients to become nauseous. "Ethyl ether" is synthesized by dehydrating "ethyl alcohol." In the reaction sulfuric acid is used to split out a molecule of water from two molecules of alcohol as follows:

$$\underset{\text{"ethyl alcohol"}}{CH_3CH_2—OH} + \underset{\text{"ethyl alcohol"}}{H—OCH_2CH_3} \xrightarrow{H_2SO_4} \underset{\text{"diethyl ether"}}{CH_3CH_2—O—CH_2CH_3} + H_2O$$

In the ether class of compounds, the attached groups may be alkyl or aryl. That is, an ether may be R—O—Ar, or Ar—O—Ar′, as well as R—O—R′ (Figure 19.8). The IUPAC names of ethers are not often used, and we will refer to an ether by indicating the two attached groups. For example, the name of $CH_3—O—CH_2CH_3$ is "methyl ethyl ether." Similarly, the name of $CH_3—O—C_6H_5$ is "methyl phenyl ether." Figure 19.8 illustrates the model structure of selected ethers.



(a) (b) (c)

▲ **Figure 19.8** **Molecular Models of Ethers** Model structures of (a) "diethyl ether," $C_2H_5—O—C_2H_5$; (b) "methyl ethyl ether," $CH_3—O—C_2H_5$; and (c) "methyl phenyl ether," $CH_3—O—C_6H_5$. (C-black, H-white, O-red)

In general, the physical properties of ethers lie between those of nonpolar hydrocarbons and those of polar alcohols of comparable molecular mass. Ethers are relatively nonpolar and do not form hydrogen bonds between molecules. Table 19.7 briefly describes a few simple ethers.

TABLE 19.7 Ethers

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
—	"dimethyl ether"	$CH_3—O—CH_3$	−25 °C	soluble
—	"methyl ethyl ether"	$CH_3—O—CH_2CH_3$	8 °C	soluble
—	"diethyl ether"	$CH_3CH_2—O—CH_2CH_3$	35 °C	~7%
—	"diphenyl ether"	$C_6H_5—O—C_6H_5$	259 °C	insoluble

*Most ethers with low molecular mass are liquids at room temperature; "diphenyl ether" is a solid.

LEARNING OBJECTIVE

- Write names and formulas for simple amines.

19.8 Amines

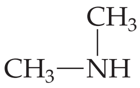
The amines often have an unpleasant odor. The simplest amine, CH_3NH_2, smells like ammonia. As the carbon chain increases, the odor becomes progressively unpleasant, and $CH_3CH_2CH_2CH_2NH_2$ has an offensive smell. The amines putrescine and cadaverine have a revolting smell, and are found in decaying flesh. Other examples of amines include the dental anesthetic novocaine, the hormone epinephrine, as well as nicotine, amphetamine, and cocaine.

Amine Family of Compounds

If an alkyl group (R—) replaces a hydrogen in ammonia, NH_3, the resulting $R—NH_2$ is called an *amine* (Figure 19.9). Further substitution may occur to give amines with two or three attached alkyl groups, that is, $R_2—NH$ or $R_3—N$. In this class of compounds, the hydrocarbon group can be alkyl or aryl. For example,

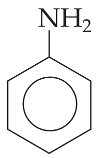
$CH_3—CH_2—NH_2$

(a) "ethylamine"



$$CH_3—\overset{\displaystyle CH_3}{\overset{|}{N}}H$$

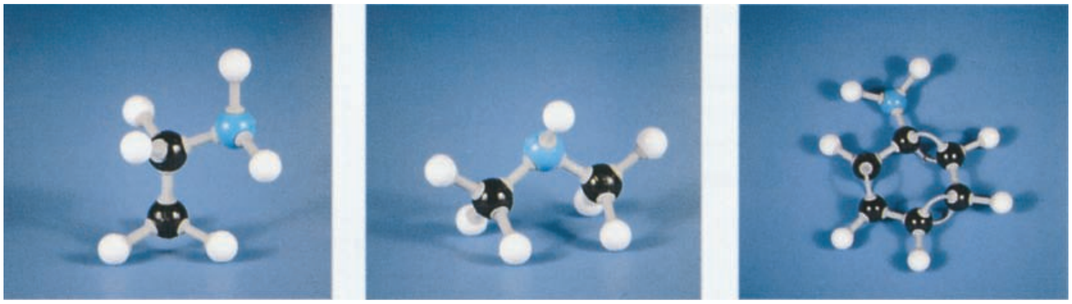
(b) "dimethylamine"



(c) "phenylamine"

The IUPAC names of amines are often not used; thus, we will refer to the amines by their common name. For example, the common name of CH_3NH_2 is "methylamine" and the name of $CH_3CH_2NH_2$ is "ethylamine." Similarly, we can refer to $C_6H_5NH_2$ as "phenylamine."

Most amines are polar because of the $—NH_2$ group. In fact, amines are similar to alcohols in that they can form intermolecular hydrogen bonds. As we might expect, the



(a) (b) (c)

▲ **Figure 19.9 Molecular Models of Amines** Model structures of (a) “ethylamine,” $C_2H_5NH_2$; (b) “dimethylamine,” $(CH_3)_2NH$; and (c) “phenylamine,” $C_6H_5NH_2$. (C-black, H-white, N-blue)

physical properties of amines and alcohols are comparable. That is, they both have high boiling points and are soluble in water. Table 19.8 briefly describes a few simple amines.

TABLE 19.8 Amines

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
—	“methylamine”	$CH_3{-}NH_2$	−7 °C	soluble
—	“ethylamine”	$CH_3{-}CH_2{-}NH_2$	17 °C	soluble
—	“propylamine”	$CH_3{-}CH_2{-}CH_2{-}NH_2$	48 °C	soluble
—	“isopropylamine”	$CH_3{-}CH(NH_2){-}CH_3$	33 °C	soluble
—	“phenylamine”	$C_6H_5{-}NH_2$	185 °C	~5%

*Most amines with low molecular mass are liquids at room temperature.

19.9 Aldehydes and Ketones

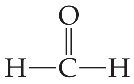
LEARNING OBJECTIVES

- Write names and formulas for simple aldehydes.
- Write names and formulas for simple ketones.

Many aldehydes and ketones have an appealing taste and a fragrant odor. For this reason they are frequently used as flavorings in food and candy and as fragrance in inhalants and perfumes. They are also used as solvents and for the manufacture of other organic compounds.

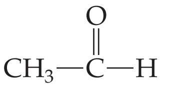
Aldehyde Family of Compounds

Aldehydes contain a carbonyl group (C═O) and, in general, are polar. In *aldehydes,* RCHO, the carbonyl group is attached to a hydrogen atom and either an alkyl or an aryl group (Figure 19.10). For example,



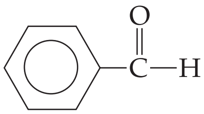
$$\mathrm{H{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}H}$$

(a) “formaldehyde”

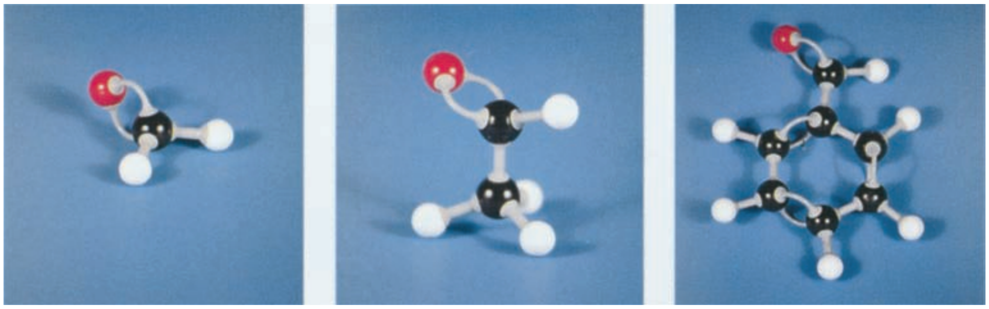


$$\mathrm{CH_3{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}H}$$

(b) “acetaldehyde”



(c) benzaldehyde



(a) (b) (c)

◄ **Figure 19.10 Molecular Models of Aldehydes** Model structures of (a) “formaldehyde,” HCHO; (b) “acetaldehyde,” CH_3CHO; and (c) benzaldehyde, C_6H_5CHO. (C-black, H-white, O-red)

The simplest aldehyde, "formaldehyde," HCHO, is one of the top 50 industrial chemicals. Formaldehyde is perhaps best known as the solution used to preserve specimens in biology classes. The preservative is actually formalin, an aqueous solution of formaldehyde. Formaldehyde is manufactured by the oxidation of methanol using a variety of metal catalysts.

According to IUPAC nomenclature, the name of an aldehyde is derived from the name of the parent alkane. The name is formed by dropping the final −*e* and adding the suffix −*al*. For example, the IUPAC name for HCHO is methanal, and the name of CH_3CHO is ethanal. We can also refer to CH_3CHO by its common name "acetaldehyde."

Aldehydes with low molecular mass are slightly polar because of the carbonyl group. They are not as polar as alcohols and do not form hydrogen bonds. Accordingly, the physical properties of aldehydes fall between those of polar alcohols and nonpolar hydrocarbons. In general, aldehydes are less soluble in water and have lower boiling points than similar alcohols. Table 19.9 lists a few simple aldehydes.

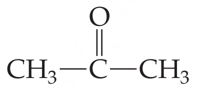
TABLE 19.9 Aldehydes

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
methanal	"formaldehyde"	$H{-}CHO$	−21 °C	soluble
ethanal	"acetaldehyde"	$CH_3{-}CHO$	21 °C	soluble
propanal	"propionaldehyde"	$CH_3{-}CH_2{-}CHO$	49 °C	soluble
benzaldehyde	"benzoic aldehyde"	$C_6H_5{-}CHO$	179 °C	~1%

*Most aldehydes with low molecular mass are liquids at room temperature.

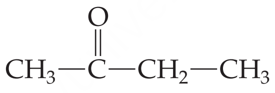
Ketone Family of Compounds

Ketones contain a carbonyl group (C═O) and, in general, are polar. In *ketones*, RCOR′, either an alkyl group or an aryl group is attached on each side of the carbonyl group (Figure 19.11). For example,



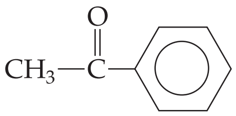
$$CH_3{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}CH_3$$

(a) "dimethyl ketone"
or
"acetone"

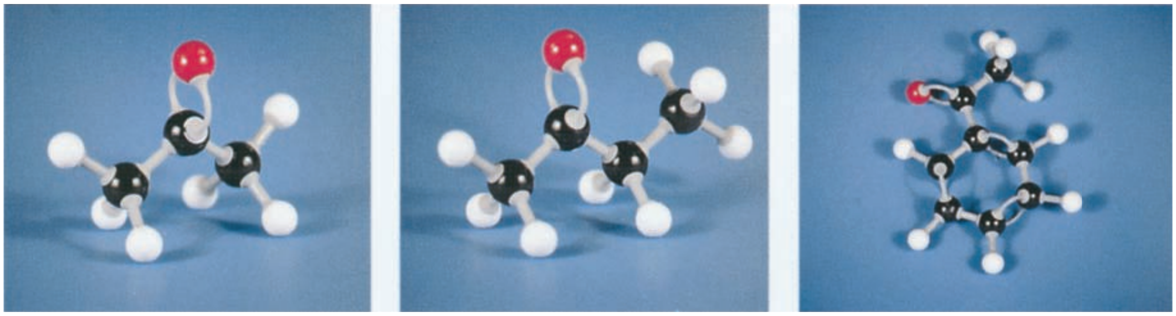


$$CH_3{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}CH_2{-}CH_3$$

(b) "methyl ethyl ketone"
or
butanone



(c) "methyl phenyl ketone"
or
acetophenone



(a) (b) (c)

▲ **Figure 19.11 Molecular Models of Ketones** Model structures of (a) "acetone," CH_3COCH_3; (b) butanone, $CH_3COC_2H_5$; and acetophenone, $CH_3COC_6H_5$. (C-black, H-white, O-red)

"Acetone," CH_3COCH_3, is the simplest ketone and is one of the top 50 industrial chemicals. It evaporates rapidly, and its vapor is quite flammable. Although acetone is soluble in water, it can dissolve a wide range of polar and nonpolar compounds. "Methyl ethyl ketone," $CH_3COCH_2CH_3$, is a solvent in the paint industry. Ketones are produced by the oxidation of alcohols. For example, large quantities of acetone are obtained from the oxidation of isopropyl alcohol.

According to IUPAC nomenclature, the name of a ketone is derived from the name of the parent alkane. The name is formed by dropping the final −*e* and adding the suffix −*one*. For example, the IUPAC name for CH_3COCH_3 is propanone, but propanone is also referred to by its common names, "dimethyl ketone" or "acetone." The IUPAC name for $CH_3COCH_2CH_3$, is butanone, and it is referred to by its common name, "methyl ethyl ketone." The next member of the series is pentanone. However, the carbonyl group may be on the second carbon, $CH_3COCH_2CH_2CH_3$, or the third carbon, $CH_3CH_2COCH_2CH_3$, of the parent chain. Thus, there are two isomers that are named 2-pentanone and 3-pentanone, respectively.

Ketones with low molecular mass are slightly polar because of the carbonyl group. They are not as polar as alcohols and do not form hydrogen bonds. In general, ketones are less soluble in water and have lower boiling points than similar alcohols. Table 19.10 lists three of the most common ketones.

TABLE 19.10 Ketones

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
propanone	"acetone" or "dimethyl ketone"	$CH_3—CO—CH_3$	56 °C	soluble
butanone	"methyl ethyl ketone"	$CH_3—CO—CH_2CH_3$	80 °C	soluble
acetophenone	"methyl phenyl ketone"	$CH_3—CO—C_6H_5$	202 °C	~1

*Most ketones with low molecular mass are liquids at room temperature; acetophenone melts at 20 °C.

19.10 Carboxylic Acids, Esters, and Amides

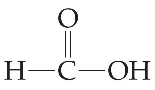
Carboxylic acids, esters, and amides are similar in that members of each class of compound contain a carbonyl group (C═O) and, in general, are polar. Carboxylic acids undergo many reactions, some of which produce esters and amides. Thus, esters and amides are derived from carboxylic acids.

LEARNING OBJECTIVES

- Write names and formulas for simple carboxylic acids.
- Write names and formulas for simple esters.
- Identify the parent alcohol and acid in an ester.
- Write names and formulas for simple amides.

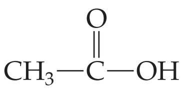
Carboxylic Acid Family of Compounds

In a *carboxylic acid*, RCOOH, we find a carbonyl group and a hydroxyl group bonded together to form a **carboxyl group**, —COOH (Figure 19.12). In this class of compounds, the hydrocarbon group can be an alkyl or an aryl. For example,



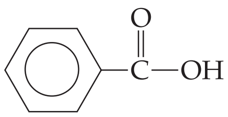
$$H—\overset{\overset{\displaystyle O}{\|}}{C}—OH$$

(a) "formic acid"

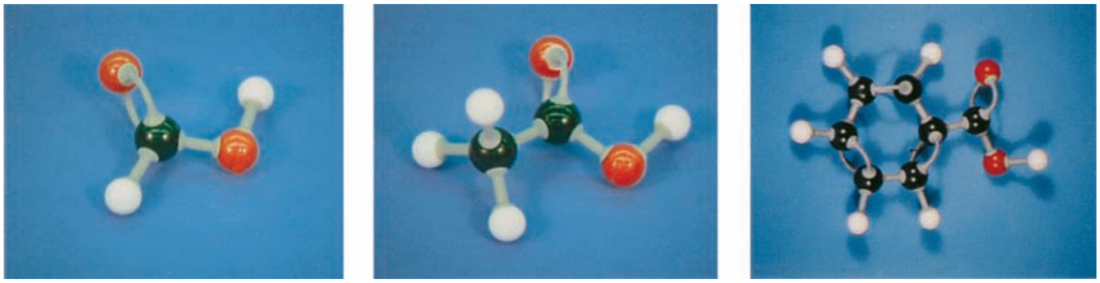


$$CH_3—\overset{\overset{\displaystyle O}{\|}}{C}—OH$$

(b) "acetic acid"



(c) benzoic acid



(a) (b) (c)

▲ **Figure 19.12 Molecular Models of Carboxylic Acids** Model structures of (a) "formic acid," HCOOH; (b) "acetic acid," CH_3COOH; and (c) benzoic acid, C_6H_5COOH. (C-black, H-white, O-red)

The simplest carboxylic acid is “formic acid,” HCOOH. Formic acid was first extracted from ants. Its name, in fact, is derived from the Latin *formica*, meaning “ant.” Formic acid is responsible for the stinging sensation of a red ant sting. It is used industrially to tan leather and to coagulate rubber latex.

“Acetic acid” was first isolated from vinegar, and its name is derived from the Latin *acetum*, meaning “vinegar.” Acetic acid gives vinegar its sour taste and is the most common carboxylic acid. It is used for manufacturing other chemicals and for preparing foods such as pickles and salad dressing.

According to IUPAC nomenclature, the name of a carboxylic acid is derived from the name of the parent alkane. The name is formed by dropping the final −*e* and adding the suffix −*oic acid*. For example, the IUPAC name for HCOOH is methanoic acid, and for CH_3COOH is ethanoic acid. We can also refer to CH_3COOH by its common name, “acetic acid.”

Inasmuch as carboxylic acids have a polar —COOH group, they are capable of forming intermolecular hydrogen bonds. The properties of carboxylic acids are similar to those of alcohols, and they have high boiling points. Since the carboxyl group is quite polar, we would predict that carboxylic acids are usually soluble in water. When we refer to Table 19.11, we see that our predictions are correct for low molecular mass carboxylic acids.

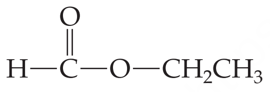
TABLE 19.11 Carboxylic Acids

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
methanoic acid	“formic acid”	$H-COOH$	101 °C	soluble
ethanoic acid	“acetic acid”	CH_3-COOH	118 °C	soluble
propanoic acid	“propionic acid”	CH_3-CH_2-COOH	141 °C	soluble
benzoic acid	—	C_6H_5-COOH	249 °C	~1%

*Most carboxylic acids with low molecular mass are liquids at room temperature; benzoic acid is a solid.

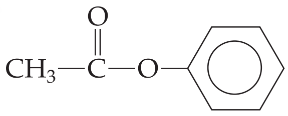
Ester Family of Compounds

In *esters*, RCOOR′, the carbonyl is attached to an —R and to an —OR′ group (Figure 19.13). In this class of compounds, the hydrocarbon group can be an alkyl or an aryl. For example,

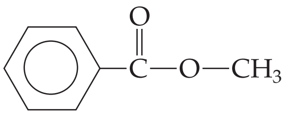


$$H-\overset{\overset{\displaystyle O}{\|}}{C}-O-CH_2CH_3$$

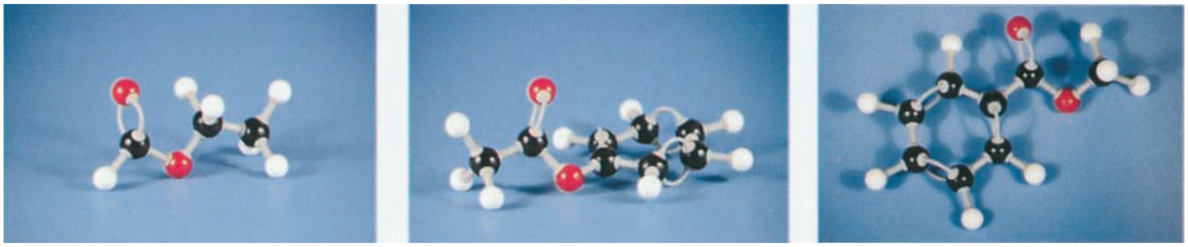
(a) “ethyl formate”



(b) “phenyl acetate”



(c) methyl benzoate

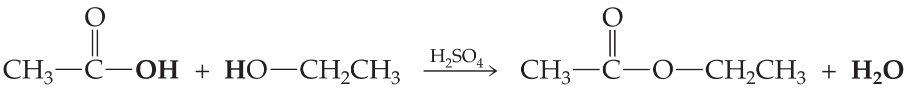


(a) (b) (c)

▲ **Figure 19.13 Molecular Models of Esters** Model structures of (a) “ethyl formate,” $HCOOC_2H_5$; (b) “phenyl acetate,” $CH_3COOC_6H_5$; and (c) methyl benzoate, $C_6H_5COOCH_3$. (C-black, H-white, O-red)

Esters typically have a pleasant fruity odor. For example, “butyl acetate” has the aroma of banana, “ethyl butyrate” has the smell of pineapple, and “ethyl formate” has the smell of rum.

An ester is produced by the reaction of a carboxylic acid and an alcohol. In the presence of sulfuric acid catalyst, the reaction takes place by splitting out a molecule of water. For example, "acetic acid" and "ethyl alcohol" react to give "ethyl acetate" and water.



$$CH_3-\overset{\overset{O}{\|}}{C}-\mathbf{OH} + \mathbf{HO}-CH_2CH_3 \xrightarrow{H_2SO_4} CH_3-\overset{\overset{O}{\|}}{C}-O-CH_2CH_3 + \mathbf{H_2O}$$

ethanoic acid "acetic acid" | ethanol "ethyl alcohol" | ethyl ethanoate "ethyl acetate"

According to IUPAC nomenclature, the name of an ester is formed from the names of the parent alcohol and carboxylic acid and the suffix –*oate*. For example, the IUPAC name for $HCOOCH_2CH_3$ is formed from the parent alcohol (ethanol) and the parent acid (methanoic acid). After adding the suffix –*oate*, the resulting name is ethyl methanoate. The common name for the ester $HCOOCH_2CH_3$ is "ethyl formate."

Esters are slightly polar because of the carbonyl group. They are not as polar as alcohols and do not form hydrogen bonds. Consequently, their physical properties fall between those of polar alcohols and those of nonpolar hydrocarbons. In general, esters are not soluble in water and have lower boiling points than similar alcohols. Table 19.12 describes a few common esters.

▲ **Ester Fragrance** The aroma of a banana is "butyl acetate," the aroma of an apple is "methyl butyrate," and the aroma of an orange is "octyl acetate."

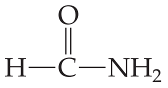
TABLE 19.12 Esters

IUPAC Name	Common Name	Structural Formula	Boiling Point*	Solubility in Water
methyl methanoate	"methyl formate"	$H-CO-OCH_3$	32 °C	soluble
methyl ethanoate	"methyl acetate"	$CH_3-CO-OCH_3$	57 °C	soluble
ethyl ethanoate	"ethyl acetate"	$CH_3-CO-OCH_2CH_3$	77 °C	soluble
methyl benzoate	—	$C_6H_5-CO-OCH_3$	200 °C	~1%

*Most esters with low molecular mass are liquids at room temperature.

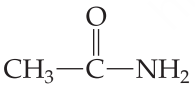
Amide Family of Compounds

In *amides*, $RCONH_2$, the carbonyl is attached to an —R group and to an $-NH_2$ group (Figure 19.14). In this class of compounds, the hydrocarbon group can be an alkyl or an aryl. For example,



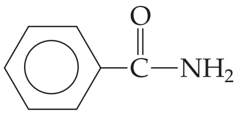
$$H-\overset{\overset{O}{\|}}{C}-NH_2$$

(a) "formamide"

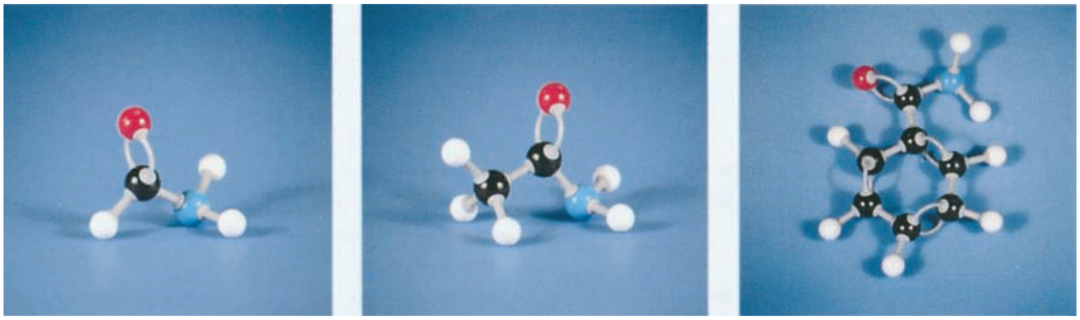


$$CH_3-\overset{\overset{O}{\|}}{C}-NH_2$$

(b) "acetamide"

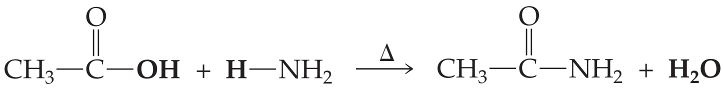


(c) benzamide



(a) (b) (c)

▲ **Figure 19.14 Molecular Models of Amides** Model structures of (a) "formamide," $HCONH_2$; (b) "acetamide," CH_3CONH_2; and (c) benzamide, $C_6H_5CONH_2$. (C-black, H-white, O-red, N-blue)



$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-\mathbf{OH} + \mathbf{H}-NH_2 \xrightarrow{\Delta} CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-NH_2 + \mathbf{H_2O}$$

ethanoic acid
"acetic acid"

ammonia

ethanamide
"acetamide"

According to IUPAC nomenclature, the name of an amide compound ends in the suffix −*amide*. The name of an amide is formed by dropping −*oic acid* from the name of the parent acid and adding the suffix −*amide*. For example, the amide derivative of ethanoic acid is CH_3CONH_2; it is named ethanamide. The amide derivative of propanoic acid is $CH_3CH_2CONH_2$; it is named propanamide.

Inasmuch as amides have a polar $-CONH_2$ group, they are capable of forming intermolecular hydrogen bonds. The properties of amides are similar to those of carboxylic acids, and amides are usually soluble in water. If we refer to Table 19.13, we see that amides are solids at room temperature.

TABLE 19.13 Amides

IUPAC Name	Common Name	Structural Formula	Melting Point*	Solubility in Water
methanamide	"formamide"	$H-CONH_2$	3 °C	soluble
ethanamide	"acetamide"	CH_3-CONH_2	82 °C	soluble
propanamide	"propionamide"	$CH_3CH_2-CONH_2$	79 °C	soluble
benzamide	"benzoic acid amide"	$C_6H_5-CONH_2$	132 °C	~1%

*Most amides are solids at room temperature; methanamide is a liquid.

A CLOSER LOOK Toxic Nail Polish

Q: Is nail polish toxic?

Medical researchers have discovered a chemical found in nail polish that is potentially dangerous. Although the research findings are preliminary, the chemical triphenyl phosphate (TPHP) may initiate hormonal changes and cause reproductive issues in women. Approximately half of all nail polishes have been found to contain the chemical.

Interestingly, research has shown that TPHP in nail polish is absorbed into the human body. In a research study, it was shown that the chemical TPHP was absorbed after a couple of hours, and after twelve hours the TPHP levels increased almost sevenfold. Moreover, clear nail polish contains more of the chemical than colored polish. Moreover, it made no difference whether the nails were natural or acrylic.

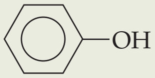
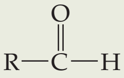
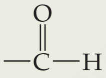
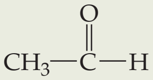
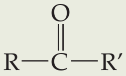
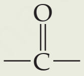
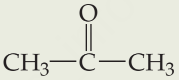
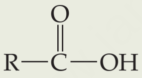
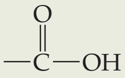
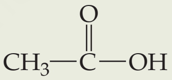
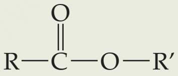
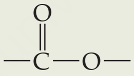
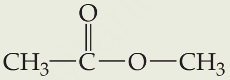
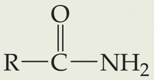
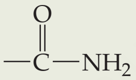
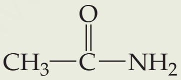
TPHP is used in plastics and nail polish to make the substance more flexible. In a study at Duke University, researchers found TPHP linked to several health issues including obesity, reproduction, and hormonal irregularities. Because teenagers are especially vulnerable to hormonal changes, it has been suggested that they avoid nail polish, and other associated beauty products. When using nail polish, do so in a well-ventilated room.

A: Nail polish contains triphenyl phosphate, TPHP, which can interfere with hormone metabolism, especially in teenagers and young children.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
19.1 Hydrocarbons **Organic chemistry** is the study of carbon compounds. A **hydrocarbon** is a carbon and hydrogen compound. **Saturated hydrocarbons** have single bonds and compose a class of compounds called *alkanes*. **Unsaturated hydrocarbons** have either a double bond and compose the family alkenes, or a triple bond and compose the family alkynes. **Aromatic hydrocarbons** typically contain a benzene ring and compose the family arenes.	• Classify a hydrocarbon having single, double, or triple bonds. ***Related Exercises: 1–6*** • Classify a hydrocarbon as saturated, unsaturated, or aromatic. ***Related Exercises: 1–6*** • Classify a hydrocarbon as an alkane, alkene, alkyne, or arene. ***Related Exercises: 1–6***
19.2 Alkanes **Alkanes** have single bonds and the general molecular formula C_nH_{2n+2}. The **structural formula** shows the arrangement of atoms in a molecule. Molecules with the same molecular formula, but a different structural formula, are called **isomers**. The name of each member of the alkane family ends in the suffix −*ane*. If a H atom is removed from an alkane, an **alkyl group** results. If a H atom is removed from an aromatic hydrocarbon, an **aryl group** results. If a H atom is removed from a benzene molecule, a **phenyl group** results. When an alkane burns with oxygen in air, it is said to undergo a **combustion reaction**.	• Write names and formulas for simple alkanes. ***Related Exercises: 7–18*** • Write combustion reactions for alkanes. ***Related Exercises: 19–20***
19.3 Alkenes and Alkynes **Alkenes** have at least one carbon–carbon double bond and the general molecular formula C_nH_{2n}. **Alkynes** have at least one carbon–carbon triple bond and the general molecular formula C_nH_{2n-2}. The name of each member of the alkene family ends in the suffix −*ene*. The names of alkynes end in the suffix −*yne*. Alkenes and alkynes can each undergo a **hydrogenation reaction**. That is, a hydrogen molecule can add to the unsaturated bond in an alkene or alkyne. An alkene can react with itself to form a long chain molecule, which is called a **polymer**, and each repeating molecular unit is called a **monomer**.	• Write names and formulas for simple alkenes. ***Related Exercises: 21–24*** • Write names and formulas for simple alkynes. ***Related Exercises: 25–34*** • Write hydrogenation reactions for alkenes and alkynes. ***Related Exercises: 35–36***
19.4 Arenes **Arenes** are hydrocarbons that contain a benzene ring. If there are two groups attached to the ring, there are three possible isomers: *ortho-*, *meta-*, and *para-*. For example, if two bromine atoms are in adjacent positions on the ring, the isomer is called *ortho*-dibromobenzene. If the bromine atoms are separated by a carbon on the ring, the isomer is called *meta*-dibromobenzene. If the bromine atoms are on opposite sides of the ring, the isomer is called *para*-dibromobenzene.	• Write names and formulas for simple arenes. ***Related Exercises: 37–40***
19.5 Hydrocarbon Derivatives Derivatives A **hydrocarbon derivative** contains carbon, hydrogen, and an additional element such as O, N, or a halogen. The hydrocarbon derivatives can be divided into families, each of which is called a **class of compounds**. There are 10 classes of hydrocarbon derivatives, five of which have a **carbonyl group** $(C{=}O)$ and five of which do not. The atom or group of atoms that characterizes a class of compounds is referred to as a **functional group**.	• Describe the concept of an organic functional group. ***Related Exercises: 41–44*** • Identify the functional group in each of the following: organic halide, alcohol, phenol, ether, and amine. ***Related Exercises: 41–44*** • Identify the functional group in each of the following: aldehyde, ketone, carboxylic acid, ester, and amide. ***Related Exercises: 41–44***
19.6 Organic Halides *Organic halides* are compounds containing hydrogen, carbon, and a halogen. They are generally nonpolar and insoluble in water. The IUPAC systematic name for an organic halide is derived from the names of the halogen and the parent alkane. For example, CH_3Br is named bromomethane, and its common name is "methyl bromide."	• Write names and formulas for simple organic halides. ***Related Exercises: 45–48***

TABLE 19.14 Summary of Hydrocarbon Derivatives

Family Name	General Formula	Functional Group	Example
Organic halide	$R-X$	$-X$	CH_3-CH_2-Cl "ethyl chloride"
Alcohol	$R-OH$	$-OH$	CH_3-CH_2-OH "ethyl alcohol"
Phenol	$Ar-OH$	$-OH$	 phenol
Ether	$R-O-R'$	$-O-$	CH_3-O-CH_3 "dimethyl ether"
Amine	$R-NH_2$	$-NH_2$	$CH_3-CH_2-NH_2$ "ethyl amine"
Aldehyde	 $R-\overset{\overset{O}{\|}}{C}-H$	 $-\overset{\overset{O}{\|}}{C}-H$	 $CH_3-\overset{\overset{O}{\|}}{C}-H$ "acetaldehyde"
Ketone	 $R-\overset{\overset{O}{\|}}{C}-R'$	 $-\overset{\overset{O}{\|}}{C}-$	 $CH_3-\overset{\overset{O}{\|}}{C}-CH_3$ "acetone"
Carboxylic acid	 $R-\overset{\overset{O}{\|}}{C}-OH$	 $-\overset{\overset{O}{\|}}{C}-OH$	 $CH_3-\overset{\overset{O}{\|}}{C}-OH$ "acetic acid"
Ester	 $R-\overset{\overset{O}{\|}}{C}-O-R'$	 $-\overset{\overset{O}{\|}}{C}-O-$	 $CH_3-\overset{\overset{O}{\|}}{C}-O-CH_3$ "methyl acetate"
Amide	 $R-\overset{\overset{O}{\|}}{C}-NH_2$	 $-\overset{\overset{O}{\|}}{C}-NH_2$	 $CH_3-\overset{\overset{O}{\|}}{C}-NH_2$ "acetamide"

Key Concepts

Learning Objectives and Related Exercises

19.7 Alcohols, Phenols, and Ethers

Alcohols and *phenols* are generally polar and are soluble in water. Alcohols have high boiling points because the **hydroxyl group** (—OH) can attract other alcohol molecules to form a hydrogen bond. The systematic name for an alcohol is formed by applying the suffix *–ol* to the name of the parent alkane. For example, CH_3OH is named methanol, and CH_3CH_2OH is named ethanol. The most important member of the phenol class is C_6H_5OH; its systematic name is phenol. *Ethers* are generally nonpolar and only slightly soluble in water. Ethers are usually referred to by their common names; for example, $CH_3CH_2-O-CH_2CH_3$ is referred to as "diethyl ether."

- Write names and formulas for simple alcohols.
  Related Exercises: 49–58
- Write names and formulas for simple phenols.
  Related Exercises: 51–58
- Write names and formulas for simple ethers.
  Related Exercises: 51–58

19.8 Amines

Amines are generally polar and are soluble in water. The amines have unusually high melting and boiling points because they can hydrogen bond. The common name for an amine is formed by adding "amine" to the name of the parent alkyl group. For example, the name of CH_3NH_2 is "methylamine"; $CH_3CH_2NH_2$ is named "ethylamine." The name of the aromatic amine $C_6H_5NH_2$ is "phenylamine."

- Write names and formulas for simple amines.
  Related Exercises: 59–62

Key Concepts	Learning Objectives and Related Exercises
19.9 Aldehydes and Ketones *Aldehydes* and *ketones* are slightly polar, and those with low molecular mass are soluble in water. The systematic name for an aldehyde is formed by applying the suffix −*al* to the name of the parent alkane. For example, HCHO is named methanal, but is usually referred to by its common name "formaldehyde." The systematic name for a ketone is formed by applying the suffix −*one* to the name of the parent alkane. For example, CH_3COCH_3 is named propanone, but it is referred to as "acetone" or "dimethyl ketone."	• Write names and formulas for simple aldehydes. ***Related Exercises: 63–64*** • Write names and formulas for simple ketones. ***Related Exercises: 65–68***
19.10 Carboxylic Acids, Esters, and Amides *Carboxylic acids* are soluble in water and have high boiling points because the **carboxyl group** (—COOH) can hydrogen bond. The systematic name for an acid is formed by applying the suffix −*oic acid* to the name of the parent alkane. For example, CH_3COOH is named ethanoic acid, but it is usually referred to as "acetic acid." *Esters* are made from the reaction of a carboxylic acid and an alcohol. The systematic name for an ester is formed by applying the suffix −*oate* to the name of the parent carboxylic acid. For example, CH_3COOCH_3 is named methyl ethanoate. *Amides* are made from the reaction of a carboxylic acid and ammonia. The systematic name for an amide is formed by applying the suffix −*amide* to the name of the parent carboxylic acid. For example, CH_3CONH_2 is synthesized from ethanoic acid and is named ethanamide.	• Write names and formulas for simple carboxylic acids. ***Related Exercises: 69–70*** • Write names and formulas for simple esters. ***Related Exercises: 71–72*** • Identify the parent alcohol and acid in an ester. ***Related Exercises: 73–74*** • Write names and formulas for simple amides. ***Related Exercises: 75–76***

Key Terms *Answers to Key Terms are in Appendix H.*

Select the key term that corresponds to each of the following definitions.

_____ **1.** the study of carbon compounds
_____ **2.** a compound containing only hydrogen and carbon
_____ **3.** a hydrocarbon containing only single bonds
_____ **4.** a hydrocarbon containing a carbon–carbon double or triple bond
_____ **5.** a hydrocarbon compound typically containing a benzene ring
_____ **6.** a family of saturated hydrocarbon compounds
_____ **7.** a family of unsaturated hydrocarbon compounds having a double bond
_____ **8.** a family of unsaturated hydrocarbon compounds having a triple bond
_____ **9.** a giant molecule composed of small molecules bonded together repeatedly in a long, continuous chain
_____ **10.** a small molecule that bonds to itself repeatedly to form a long, continuous chain
_____ **11.** a family of aromatic hydrocarbon compounds
_____ **12.** a chemical formula that shows the arrangement of atoms in a molecule
_____ **13.** compounds with the same molecular formula, but different structures
_____ **14.** the part that remains after a hydrogen atom is removed from an alkane
_____ **15.** the part that remains after a hydrogen atom is removed from an arene
_____ **16.** the part that remains after a hydrogen atom is removed from benzene
_____ **17.** a chemical reaction in which a hydrocarbon reacts rapidly with oxygen
_____ **18.** a chemical reaction in which an unsaturated hydrocarbon adds hydrogen
_____ **19.** a family of compounds in which all members have similar names and properties
_____ **20.** a compound containing carbon, hydrogen, and another element
_____ **21.** an atom or group of atoms that characterizes a class of compounds
_____ **22.** the C═O group
_____ **23.** the —OH group
_____ **24.** the —COOH group

(a) alkanes (*Sec. 19.2*)
(b) alkenes (*Sec. 19.3*)
(c) alkyl group (R—) (*Sec. 19.2*)
(d) alkynes (*Sec. 19.3*)
(e) arenes (*Sec. 19.4*)
(f) aromatic hydrocarbon (*Sec. 19.1*)
(g) aryl group (Ar—) (*Sec. 19.2*)
(h) carbonyl group (*Sec. 19.5*)
(i) carboxyl group (*Sec. 19.10*)
(j) class of compounds (*Sec. 19.5*)
(k) combustion reaction (*Sec. 19.2*)
(l) functional group (*Sec. 19.5*)
(m) hydrocarbon (*Sec. 19.1*)
(n) hydrocarbon derivative (*Sec. 19.5*)
(o) hydrogenation reaction (*Sec. 19.3*)
(p) hydroxyl group (*Sec. 19.7*)
(q) isomers (*Sec. 19.2*)
(r) monomer (*Sec. 19.3*)
(s) organic chemistry (*Sec. 19.1*)
(t) phenyl group (*Sec. 19.2*)
(u) polymer (*Sec. 19.3*)
(v) saturated hydrocarbon (*Sec. 19.1*)
(w) structural formula (*Sec. 19.2*)
(x) unsaturated hydrocarbon (*Sec. 19.1*)

Exercises

Answers to odd-numbered Exercises are in Appendix I.

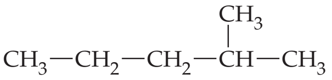
Hydrocarbons (Sec. 19.1)

1. Which class of hydrocarbon is saturated?
2. Which classes of hydrocarbon are unsaturated?
3. Which class of hydrocarbon has single bonds?
4. Which class of hydrocarbon has double bonds?
5. Which class of hydrocarbon has triple bonds?
6. Which class of hydrocarbon contains a benzene ring?

Alkanes (Sec. 19.2)

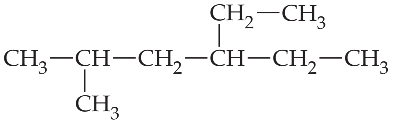
7. Which of the following molecular formulas represents an alkane?
   (a) C_6H_{14} **(b)** C_6H_{12}
8. What is the molecular formula for each of the following alkanes?
   (a) $C_8H_?$ **(b)** $C_{10}H_?$
9. State the name for each of the following alkanes.
   (a) $CH_3-CH_2-CH_2-CH_3$
   (b) $CH_3-CH_2-CH_2-CH_2-CH_2-CH_3$
   (c) $CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$
   (d) $CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$
10. Draw the condensed structural formula for each of the following alkanes.
    (a) propane
    (b) pentane
    (c) heptane
    (d) nonane
11. Draw the condensed structural formula for the three isomers of pentane, C_5H_{12}.
12. Draw the condensed structural formula for the five isomers of hexane, C_6H_{14}.
13. Draw the two isomers of bromopropane, C_3H_7Br.
14. Draw the two isomers of dibromoethane, $C_2H_4Br_2$.
15. Draw each of the following alkyl groups.
    (a) methyl **(b)** ethyl
16. Draw the structure for each of the following alkyl groups.
    (a) propyl **(b)** isopropyl
17. Give the IUPAC names for each of the following structural formulas.

(a)



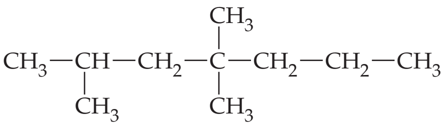
$$CH_3-CH_2-CH_2-\overset{\substack{CH_3\\|}}{CH}-CH_3$$

(b)



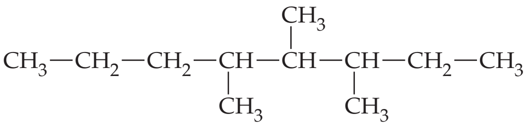
$$CH_3-\underset{\substack{|\\CH_3}}{CH}-CH_2-\overset{\substack{CH_2-CH_3\\|}}{CH}-CH_2-CH_3$$

(c)



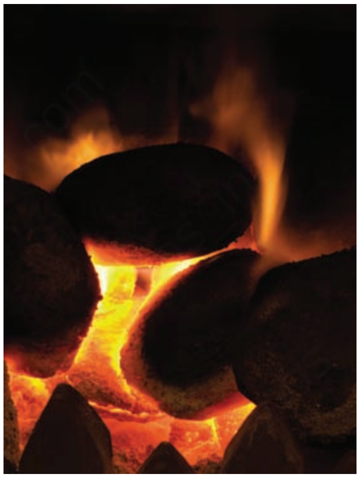
$$CH_3-\underset{\substack{|\\CH_3}}{CH}-CH_2-\overset{\substack{CH_3\\|}}{\underset{\substack{|\\CH_3}}{C}}-CH_2-CH_2-CH_3$$

(d)



$$CH_3-CH_2-CH_2-\underset{\substack{|\\CH_3}}{CH}-\overset{\substack{CH_3\\|}}{CH}-\underset{\substack{|\\CH_3}}{CH}-CH_2-CH_3$$

18. Draw the condensed structural formula for each of the following compounds.
    (a) 2-methylpropane **(b)** 2,2-dimethylbutane
    (c) 3-ethylheptane **(d)** 3,3-diethylpentane
19. Write a balanced equation for the complete combustion for each of the following.
    (a) methane **(b)** propane
    (c) hexane **(d)** octane
20. Write a balanced equation for the complete combustion for each of the following.
    (a) ethane **(b)** butane
    (c) pentane **(d)** heptane



◀ **Combustion Reaction** Charcoal burns with oxygen in air to give carbon dioxide, water, and heat.

Alkenes and Alkynes (Sec. 19.3)

21. Which of the following molecular formulas represents an alkene?
    (a) C_6H_{14} **(b)** C_6H_{12}
22. What is the molecular formula for each of the following alkenes?
    (a) $C_8H_?$ **(b)** $C_{10}H_?$
23. State the name for each of the following alkenes.
    (a) $CH_2=CH-CH_2-CH_3$
    (b) $CH_2=CH-CH_2-CH_2-CH_3$
    (c) $CH_3-CH=CH-CH_2-CH_2-CH_3$
    (d) $CH_3-CH_2-CH_2-CH=CH-CH_2-CH_2-CH_3$
24. Draw the condensed structural formula for each of the following alkenes.
    (a) 2-pentene **(b)** 3-heptene
    (c) 1-propene **(d)** 4-nonene
25. Which of the following molecular formulas represents an alkyne?
    (a) C_6H_{12} **(b)** C_6H_{10}
26. What is the molecular formula for each of the following alkynes?
    (a) $C_8H_?$ **(b)** $C_{10}H_?$
27. State the name for each of the following alkynes.
    (a) $CH_3-C\equiv C-CH_2-CH_3$
    (b) $CH\equiv C-CH_2-CH_2-CH_3$
    (c) $CH_3-CH_2-C\equiv C-CH_2-CH_3$
    (d) $CH_3-CH_2-CH_2-CH_2-C\equiv C-CH_3$

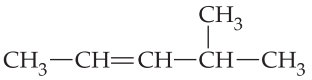
28. Draw the condensed structural formula for each of the following alkynes.
(a) 1-propyne
(b) 4-octyne
(c) 2-butyne
(d) 3-decyne

29. Draw the two structural isomers of straight-chain pentene, C_5H_{10}.

30. Draw the two structural isomers of straight-chain pentyne, C_5H_8.

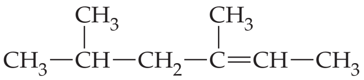
31. Give the IUPAC name for each of the following structural formulas.

(a)



$$\mathrm{CH_3{-}CH{=}CH{-}\overset{CH_3}{\overset{|}{C}H}{-}CH_3}$$

(b)

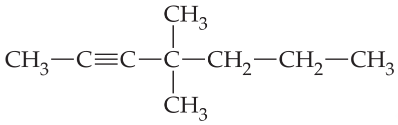


$$\mathrm{CH_3{-}\overset{CH_3}{\overset{|}{C}H}{-}CH_2{-}\overset{CH_3}{\overset{|}{C}}{=}CH{-}CH_3}$$

32. Draw the condensed structural formula for each of the following compounds.
(a) 2-methyl-1-propene
(b) 3,3-dimethyl-1-butyne

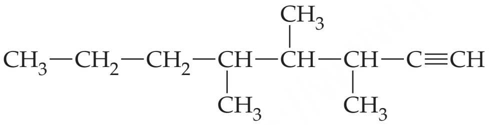
33. Give the IUPAC names for each of the following structural formulas.

(a)



$$\mathrm{CH_3{-}C{\equiv}C{-}\underset{\underset{CH_3}{|}}{\overset{\overset{CH_3}{|}}{C}}{-}CH_2{-}CH_2{-}CH_3}$$

(b)



$$\mathrm{CH_3{-}CH_2{-}CH_2{-}\underset{\underset{CH_3}{|}}{CH}{-}\overset{\overset{CH_3}{|}}{CH}{-}\underset{\underset{CH_3}{|}}{CH}{-}C{\equiv}CH}$$

34. Draw the condensed structural formula for each of the following compounds.
(a) 3-ethyl-2-heptene
(b) 4,4-dimethyl-2-pentyne

35. Write a balanced equation for each of the following reactions.
(a) $CH_2{=}CH_2 + O_2 \rightarrow$
(b) $CH_3{-}CH{=}CH_2 + H_2 \rightarrow$

36. Write a balanced equation for each of the following reactions.
(a) $CH{\equiv}C{-}CH_3 + O_2 \rightarrow$
(b) $CH_3 - C{\equiv}CH + 2\,H_2 \rightarrow$

Arenes (Sec. 19.4)

37. Draw and name the three isomers of the solvent xylene, $C_6H_4(CH_3)_2$.

38. Draw and name the three isomers of dinitrobenzene, $C_6H_4(NO_2)_2$.

39. What are the two products from the complete combustion of benzene?

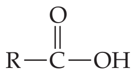
40. Write a balanced chemical equation for the complete combustion of benzene.

Hydrocarbon Derivatives (Sec. 19.5)

41. Identify the class of compound for each of the following hydrocarbon derivatives.
(a) $R{-}O{-}R'$
(b) $R{-}X$
(c) $Ar{-}OH$
(d) $R{-}NH_2$

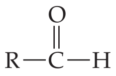
42. Identify the class of compound for each of the following hydrocarbon derivatives.

(a)



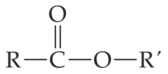
$$\mathrm{R{-}\overset{O}{\overset{\|}{C}}{-}OH}$$

(b)



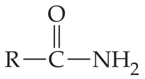
$$\mathrm{R{-}\overset{O}{\overset{\|}{C}}{-}H}$$

(c)



$$\mathrm{R{-}\overset{O}{\overset{\|}{C}}{-}O{-}R'}$$

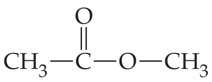
(d)



$$\mathrm{R{-}\overset{O}{\overset{\|}{C}}{-}NH_2}$$

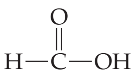
43. Identify the class of compound for each of the following hydrocarbon derivatives.
(a) $CH_3{-}NH_2$
(b) $CH_3{-}CH_2{-}F$

(c)



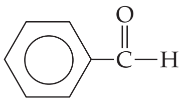
$$\mathrm{CH_3{-}\overset{O}{\overset{\|}{C}}{-}O{-}CH_3}$$

(d)

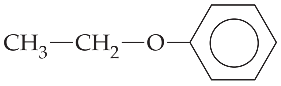


$$\mathrm{H{-}\overset{O}{\overset{\|}{C}}{-}OH}$$

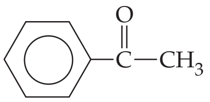
(e)



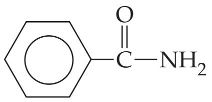
(f)



(g)

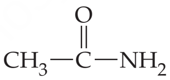


(h)



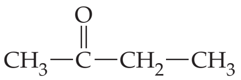
44. Identify the class of compound for each of the following hydrocarbon derivatives.
(a) $CH_3{-}OH$
(b) $CH_3{-}CH_2{-}O{-}CH_2{-}CH_3$

(c)



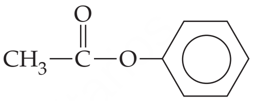
$$\mathrm{CH_3{-}\overset{O}{\overset{\|}{C}}{-}NH_2}$$

(d)

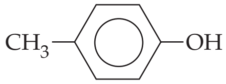


$$\mathrm{CH_3{-}\overset{O}{\overset{\|}{C}}{-}CH_2{-}CH_3}$$

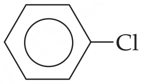
(e)



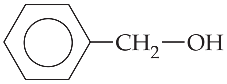
(f)



(g)



(h)



Organic Halides (Sec. 19.6)

45. Give the common name for each of the following organic halides.
(a) $CH_3{-}I$
(b) $CH_3{-}CH_2{-}Br$
(c) $CH_3{-}CH_2{-}CH_2{-}F$
(d) $(CH_3)_2{-}CH{-}Cl$

46. Draw the structure for each of the following organic halides.
(a) dichloromethane
(b) iodoethane
(c) 2-fluoropropane
(d) 1,2-dibromopropane

47. Draw the structure for "trichloroethylene," a common solvent used for degreasing electronic and automobile parts.

48. Draw the structure for 1,1,1-trichloroethane, a solvent formerly used for dry cleaning clothes.

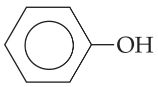
Alcohols, Phenols, and Ethers (Sec. 19.7)

49. Give the systematic IUPAC name for each of the following alcohols.
(a) $CH_3{-}CH_2{-}CH_2{-}CH_2{-}OH$
(b) $CH_3{-}CH_2{-}CH(OH){-}CH_3$
(c) $CH_2(OH){-}CH_2{-}CH_2{-}CH_3$
(d) $CH_3{-}CH(OH){-}CH_2{-}CH_3$

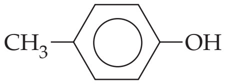
50. Draw the structure for each of the following alcohols.
(a) "methyl alcohol"
(b) "ethyl alcohol"
(c) "propyl alcohol"
(d) "isopropyl alcohol"

51. Give an acceptable name for each of the following phenols.

(a)

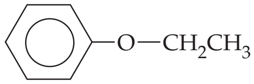


(b)



52. Draw the structure for each of the following phenols.
(a) *ortho*-chlorophenol **(b)** *meta*-ethylphenol

53. Give the common name for each of the following ethers.
(a) CH_3-O-CH_3
(b) $CH_3-O-CH_2CH_3$
(c) $CH_3CH_2CH_2-O-CH_2CH_2CH_3$
(d)



54. Draw the structure for each of the following ethers.
(a) "diisopropyl ether" **(b)** "diphenyl ether"
(c) "ethyl propyl ether" **(d)** "methyl phenyl ether"

55. "Ethyl alcohol" and "dimethyl ether" have the same molecular mass. Explain why the boiling point of "ethyl alcohol" (78 °C) is much higher than the boiling point of "dimethyl ether" (−25 °C).

56. Methanol, ethanol, and propanol are soluble in water. Explain why butanol is only slightly soluble.

57. There are two compounds that have the molecular formula C_2H_6O. Draw the structure for each isomer and classify the hydrocarbon derivative.

58. There are three compounds that have the molecular formula C_3H_8O. Draw the structure for each isomer and classify the hydrocarbon derivative.

Amines (Sec. 19.8)

59. Give the common name for each of the following amines.
(a) $CH_3CH_2NH_2$ **(b)** $CH_3CH_2CH_2NH_2$

60. Draw the structure for each of the following amines.
(a) "methylamine" **(b)** "isopropylamine"

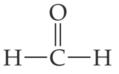
61. "Ethylamine" and propane have about the same molecular mass. Explain why the boiling point of "ethylamine" (17 °C) is much higher than that of propane (−44 °C).

62. "Ethylamine" and "diethylamine" are soluble in water. Explain why "triethylamine" is only slightly soluble in water.

Aldehydes and Ketones (Sec. 19.9)

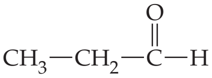
63. Give the systematic IUPAC name for each of the following aldehydes.

(a)



$H-\overset{O}{\overset{\|}{C}}-H$

(b)

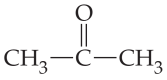


$CH_3-CH_2-\overset{O}{\overset{\|}{C}}-H$

64. Draw the structure for each of the following aldehydes.
(a) "formaldehyde" **(b)** "acetaldehyde"
(c) "propionaldehyde" **(d)** benzaldehyde

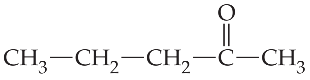
65. Give the systematic IUPAC name for each of the following ketones.

(a)



$CH_3-\overset{O}{\overset{\|}{C}}-CH_3$

(b)



$CH_3-CH_2-CH_2-\overset{O}{\overset{\|}{C}}-CH_3$

66. Draw the structure for each of the following ketones.
(a) "methyl ethyl ketone" **(b)** "diethyl ketone"
(c) "methyl phenyl ketone" **(d)** "diphenyl ketone"

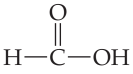
67. An aldehyde and a ketone each have the molecular formula C_3H_6O. Draw the structure for each isomer.

68. An aromatic aldehyde and an aromatic ketone each have the molecular formula C_8H_8O. Draw the structure for each isomer.

Carboxylic Acids, Esters, and Amides (Sec. 19.10)

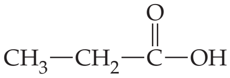
69. Give the systematic IUPAC name for each of the following carboxylic acids.

(a)



$H-\overset{O}{\overset{\|}{C}}-OH$

(b)

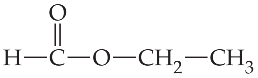


$CH_3-CH_2-\overset{O}{\overset{\|}{C}}-OH$

70. Draw the structure for each of the following carboxylic acids.
(a) "formic acid" **(b)** "acetic acid"

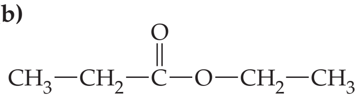
71. Give the systematic IUPAC name for each of the following esters.

(a)



$H-\overset{O}{\overset{\|}{C}}-O-CH_2-CH_3$

(b)

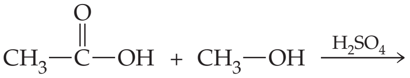


$CH_3-CH_2-\overset{O}{\overset{\|}{C}}-O-CH_2-CH_3$

72. Draw the structure for each of the following esters.
(a) "propyl formate" **(b)** "ethyl acetate"

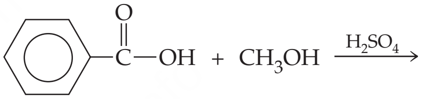
73. Write a balanced equation for each of the following reactions.

(a)



$CH_3-\overset{O}{\overset{\|}{C}}-OH + CH_3-OH \xrightarrow{H_2SO_4}$

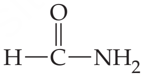
(b)



74. Identify the carboxylic acid and alcohol that react to produce the following.
(a) ethyl methanoate **(b)** phenyl ethanoate

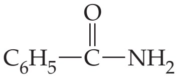
75. Give an acceptable name for each of the following amides.

(a)



$H-\overset{O}{\overset{\|}{C}}-NH_2$

(b)



$C_6H_5-\overset{O}{\overset{\|}{C}}-NH_2$

76. Draw the structure for each of the following amides.
(a) "acetamide" **(b)** "propionamide"

General Exercises

77. Predict whether each of the following compounds is saturated or unsaturated as indicated by the name.
(a) eicosane **(b)** dodecene

78. Predict whether each of the following compounds is saturated or unsaturated as indicated by the name.
(a) cyclohexene **(b)** cyclopropane

79. Acetylsalicylic acid is the active ingredient in aspirin. Based on its name and suffix, which functional group is present?

80. Sulfanilamide is an antibacterial. Based on its name and suffix, which functional group is present?

81. Androsterone is a male hormone. Based on its name and suffix, which functional group is present?

82. Progesterone is a female hormone. Based on its name and suffix, which functional group is present?

83. Chloral hydrate is a veterinary anesthetic. Based on its name, which two functional groups are present?

84. Ethynyl estradiol is an oral contraceptive. Based on its name, which two functional groups are present?

Challenge Exercises

85. The following pairs of compounds have about the same molar mass. Predict which compound in each pair has the higher boiling point.
(a) CH_3OCH_3 or CH_3CH_2OH
(b) $CH_3CH_2CH_2NH_2$ or $CH_3CH_2CH_2F$

86. The following pairs of compounds have about the same molar mass. Predict which compound in each pair has the higher boiling point.
(a) CH_3COOH or $HCOOCH_3$
(b) $CH_3CH_2CONH_2$ or CH_3COOCH_3

Chapter 19 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following hydrocarbons is saturated? (Sec. 19.1)
(a) alkanes **(b)** alkenes
(c) alkynes **(d)** arenes
(e) none of the above

2. What is the IUPAC name for $CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_3$? (Sec. 19.2)
(a) decane **(b)** heptane
(c) hexane **(d)** octane
(e) none of the above

3. What is the product from the reaction of one mole of acetylene and two moles of hydrogen gas using platinum catalyst? (Sec. 19.3)
(a) ethane **(b)** ethene
(c) propane **(d)** propene
(e) none of the above

4. How many structures of benzene, C_6H_6, are represented by the Kekulé model? (Sec. 19.4)
(a) 1 **(b)** 2
(c) 3 **(d)** 4
(e) none of the above

5. Which of the following classes of compounds has a carbonyl group? (Sec. 19.5)
(a) aldehyde **(b)** amide
(c) ester **(d)** ketone
(e) all of the above

6. What is the condensed structural formula for "propyl bromide?" (Sec. 19.6)
(a) $CH_3—Br$ **(b)** $CH_3—CH_2—Br$
(c) $CH_3—CH_2—CH_2—Br$ **(d)** $CH_3—CHBr—CH_3$
(e) none of the above

7. What is the condensed structural formula for 2–propanol? (Sec. 19.7)
(a) $CH_3—CH_2—CH_2—OH$ **(b)** $CH_3—CH(OH)—CH_3$
(c) $CH_3—CH_2—O—CH_3$ **(d)** $CH_3—O—CH_2—CH_3$
(e) none of the above

8. Which of the following compounds can form hydrogen bonds? (Sec. 19.8)
(a) $CH_3—CH_2—CH_2—CH_2—Cl$
(b) $CH_3—CH_2—CH_2—CH_2—CH_3$
(c) $CH_3—CH_2—CH_2—CH{=}CH_2$
(d) $CH_3—CH_2—CH_2—CH_2—OH$
(e) none of the above

9. What is the condensed structural formula for 2–pentanone? (Sec. 19.9)
(a) $CH_3—CH_2—CH_2—CH_2—CH_3$
(b) $CH_3—CH{=}CH—CH_2—CH_3$
(c) $CH_3—CH_2—CO—CH_2—CH_3$
(d) $CH_3—CH_2—CH_2—CH_2—COOH$
(e) none of the above

10. What are the products from the reaction of ethanoic acid and ethanol using sulfuric acid catalyst? (Sec. 19.10)
(a) $CH_3COOH + CH_3CH_2OH \xrightarrow{H_2SO_4} CH_3COCH_2CH_3 + H_2O$
(b) $CH_3COOH + CH_3CH_2OH \xrightarrow{H_2SO_4} CH_3CH_2CH_2CHO + H_2O$
(c) $CH_3COOH + CH_3CH_2OH \xrightarrow{H_2SO_4} CH_3CH_2CH_2COOH + H_2O$
(d) $CH_3COOH + CH_3CH_2OH \xrightarrow{H_2SO_4} CH_3COOCH_2CH_3 + H_2O$
(e) $CH_3COOH + CH_3CH_2OH \xrightarrow{H_2SO_4}$ no reaction

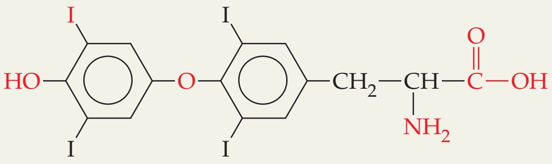
Key Concepts

11. Draw the alcohol and ether isomers with the molecular formula C_2H_6O.

12. Draw the aldehyde and ketone isomers with the molecular formula C_3H_6O.

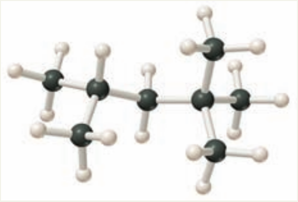
13. Draw the carboxylic acid and ester isomers with the molecular formula $C_2H_4O_2$.

14. Identify five functional groups in thyroxine (thyroid hormone).

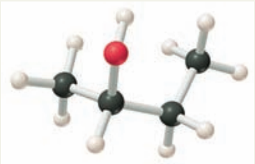


Critical Thinking

15. Draw the condensed structural formula for each of the following models.
(a)



(b)



16. The organic compound caffeine is a stimulant found in coffee. Based on its name, to which class of compound does caffeine belong?

17. The organic compound cholesterol has been linked to arterial plaque. Based on its name, to which class of compound does cholesterol belong?

18. The organic compound cortisone is an anti-inflammatory drug. Based on its name, to which class of compound does cortisone belong?

CHAPTER 20

try

"Art is I; science is we."

Claude Bernard, French Physiologist (1813–1878)

- **20.1** Biological Compounds
- **20.2** Proteins
- **20.3** Enzymes
- **20.4** Carbohydrates
- **20.5** Lipids
- **20.6** Nucleic Acids

20
Ca
calcium

Element 20: Calcium
Calcium is an alkaline earth metal that reacts with water. When the metal is dropped into water, it reacts to produce somewhat insoluble calcium hydroxide and releases hydrogen gas bubbles. Calcium is an important dietary mineral, which is responsible for bone formation.

There are about 30 chemical elements essential for life. The pictorial periodic table of elements in Figure 3.7 indicates the *essential elements* that are necessary to support life. The elements found in significant amounts in humans are oxygen, carbon, hydrogen, nitrogen, phosphorus, chlorine, sulfur, calcium, magnesium, sodium, and potassium. In addition, there are several *trace elements* that are present in only small amounts, but nonetheless are necessary to sustain life. These trace elements include iron, zinc, cobalt, iodine, selenium, and a few others.

The study of the chemistry of living things is referred to as *biochemistry*. Many biochemists are employed in the pharmaceutical industry, which manufactures drugs that affect chemical reactions in plants and animals. Some biochemists are referred to as *molecular biologists* because they study the relationships among different types of biological molecules and their role in life processes. As we proceed into the twenty-first century, applications of biotechnology and genetic engineering offer the potential for transforming our lives in ways that previously were unimaginable.

20.1 Biological Compounds

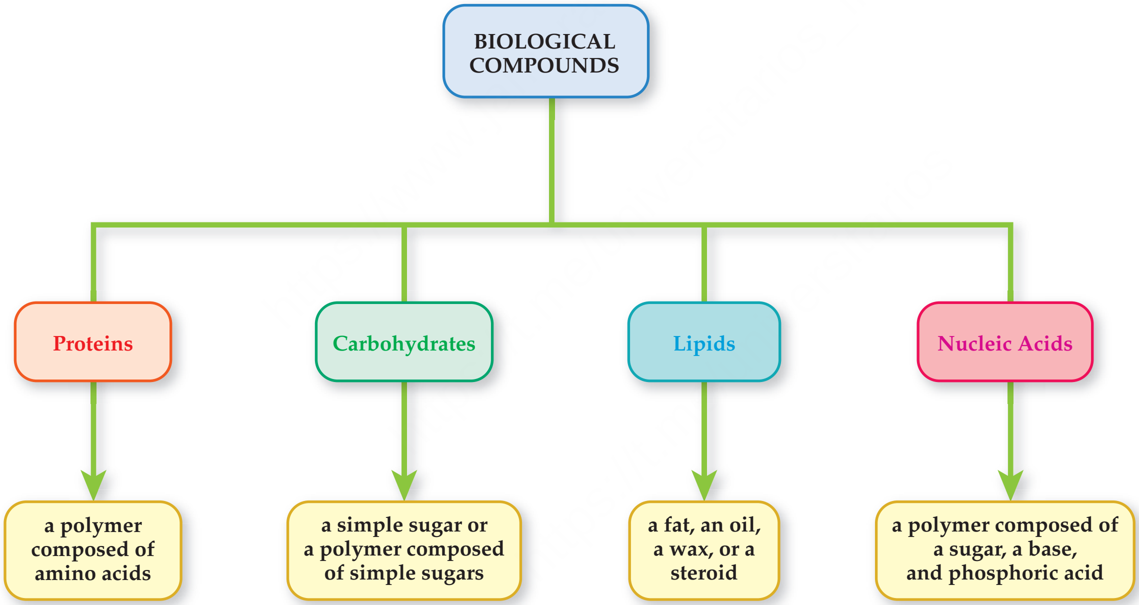
LEARNING OBJECTIVES

- Recognize a protein given its structural formula.
- Recognize a carbohydrate given its structural formula.
- Recognize a lipid given its structural formula.
- Recognize a nucleic acid given its structural formula.

Biochemistry can be defined as the study of biological compounds and their chemical reactions. Although biological compounds are often large and complex, they have the same types of functional groups we encountered in simple organic compounds. For instance, biochemical compounds frequently contain an alcohol, an amine, an aldehyde, a ketone, a carboxylic acid, an ester, or an amide functional group.

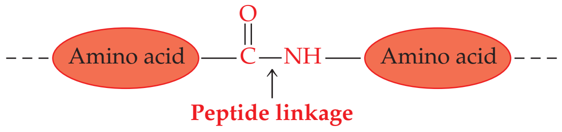
Biological compounds are often large molecules with molar masses greater than 1,000,000 g/mol. A large molecule is composed of small molecules linked together in a long continuous chain. A giant molecule containing many repeating smaller molecules is called a *polymer.* Proteins, starch, and nucleic acids are examples of naturally occurring polymers found in plants and animals.

In this chapter, we will study the following types of biochemical compounds: *proteins, carbohydrates, lipids,* and *nucleic acids.* These four types of compounds are classified in Figure 20.1.



▲ **Figure 20.1 Classification of Biological Compounds** Biochemical compounds can be classified as one of the following: protein, carbohydrate, lipid, or nucleic acid. In this chapter, protein structures will be shown in red, carbohydrates in green, lipids in blue, and nucleic acids in pink.

Before discussing biological compounds in detail, let's briefly describe the general features of proteins, carbohydrates, lipids, and nucleic acids. A **protein** is a naturally occurring polymer composed of many amino acids. An amino acid has both an amine and a carboxylic acid functional group. In a protein, amino acid molecules are linked together by amide bonds. Each of these amide bonds is called a *peptide linkage,* which we can show as follows:



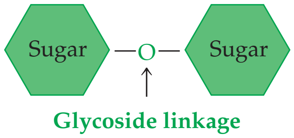
Helpful Hint
Biomolecules

Although proteins, carbohydrates, lipids, and nucleic acids appear to be complex, each of these structures has the same functional groups we studied for simple organic molecules in Chapter 19.

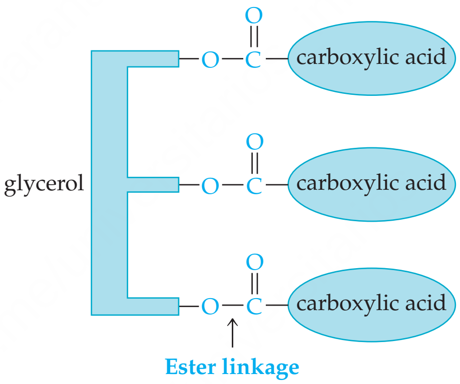
Notice that proteins contain amide groups; sugars contain alcohol groups as well as an aldehyde or a ketone group; lipids contain ester linkages; and nucleic acids have alcohol, amine, and amide functional groups.

A **carbohydrate** is a compound that can be either a simple sugar or a polymer composed of many simple sugar molecules. A carbohydrate usually contains either an aldehyde or a ketone functional group. Furthermore, a typical carbohydrate is characterized by having several alcohol groups: glucose is a simple sugar that contains an aldehyde and five alcohol groups; fructose is a simple sugar that contains a ketone and five alcohol groups.

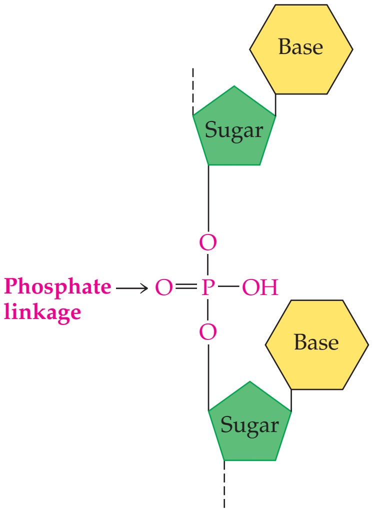
Starch is a carbohydrate polymer composed of glucose units linked together in a large molecule. In carbohydrates containing more than one sugar unit, the sugar molecules are joined together by an —O— linkage. The —O— bond between sugar molecules in a carbohydrate is called a *glycoside linkage*, which we can show as follows:



A **lipid** is a water-insoluble compound usually composed of an alcohol and one or more carboxylic acid molecules. The lipids include animal fats and vegetable oils. Fats and oils are esters of glycerol, an alcohol that has three —OH groups. As a result, each molecule of a fat or an oil contains three ester groups resulting from three carboxylic acid molecules joined to one glycerol molecule. We can show the *ester linkages* in a fat or oil molecule as follows:



A **nucleic acid** is a biochemical polymer composed of a very large number of individual units. Each unit in the nucleic acid contains a sugar molecule attached to a nitrogen base and a phosphate group. The nitrogen base has an amine functional group and is simply referred to as a *base.* The backbone of the nucleic acid is formed by sugar molecules joined together in a long polymer chain. In a nucleic acid molecule, there may be millions of sugar molecules, each with an attached base, linked by phosphate ester bonds. We can show the *phosphate linkage* in a typical nucleic acid as follows:



EXAMPLE EXERCISE 20.1 Biological Compounds

Identify each of the following biological compounds as a protein, carbohydrate, lipid, or nucleic acid based on the following descriptions.

(a) a chain of amino acids joined by an amide linkage
(b) sugar molecules joined by a glycoside linkage

Solution

(a) A chain of amino acids joined by an amide linkage describes a *protein*.
(b) Sugar molecules joined by a glycoside linkage describes a *carbohydrate*.

Practice Exercise

Identify each of the following biological compounds as a protein, carbohydrate, lipid, or nucleic acid based on the following descriptions.

(a) three fatty acids joined to a glycerol molecule by ester linkages
(b) a chain of ribose molecules and base joined by phosphate linkages

Answers:

(a) Three fatty acids joined to a glycerol molecule by ester linkages describes a fat or oil, which is a *lipid*.
(b) A chain of ribose molecules and base joined by phosphate linkages describes a *nucleic acid*.

Concept Exercise

What physical property distinguishes a lipid from a protein, carbohydrate, and nucleic acid?

Answers: See Appendix G, 20.1.

20.2 Proteins

LEARNING OBJECTIVES

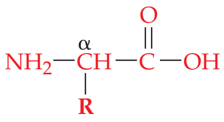
- Describe the primary structure of a protein.
- Describe the secondary structure of a protein.
- Describe the tertiary structure of a protein.
- Identify amino acids and a peptide linkage.

Proteins play a crucial role in virtually all biological processes. In fact, the word *protein* is derived from the Greek word *proteios,* meaning "of first importance." The shapes of protein molecules have been observed to be either long fibers or compact globules. Moreover, biochemists have found that the molecular shape of a protein is related to its biological function.

◀ **Protein** Meat, fish, poultry, cheese, and milk products are all food sources rich in protein.

Long protein fibers are responsible for the structure, support, and motion of an organism. Accordingly, these protein fibers are found in muscle tissue and cartilage. In addition, protein fibers compose the hair, skin, and nails. On the other hand, compact protein globules are responsible for metabolic processes. For example, many hormones are globular proteins such as those secreted by the pituitary gland and the pancreas. Globular proteins are involved in controlling growth, transporting small molecules and ions, and providing immune protection from foreign substances.

Human protein is composed of 20 different amino acids. Of these 20 amino acids, about half are essential and must be included in the diet. The other half are nonessential, and the human body can synthesize them from one of the essential amino acids. An **amino acid** has an amine group, $—NH_2$, attached to a carbon that is also bonded to a carboxyl group, —COOH. The carbon atom in the molecule that is attached to the carboxyl group is called the *alpha carbon* (α-carbon). Since the amine group is attached to the α-carbon, an amino acid is referred to as an α-amino. acid. The α-amino acids have the following general structural formula:

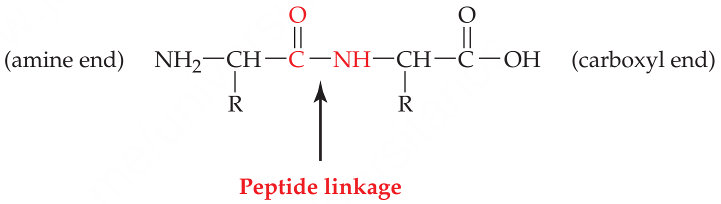


$$NH_2—\overset{\alpha}{C}H(—R)—\overset{O}{\overset{\|}{C}}—OH$$

Alpha (α) amino acid

Amino acids differ from one another according to the R— group attached to the α-carbon. This R— group is referred to as the *side chain.* In some amino acids, the R— group is neutral, and in other amino acids the side chain has either acidic or basic properties. The structures of the 20 α-amino acids found in human protein are shown in Figure 20.2. For convenience, chemists abbreviate the names of amino acids by a three-letter symbol; for example, the name of the amino acid glycine is abbreviated Gly.

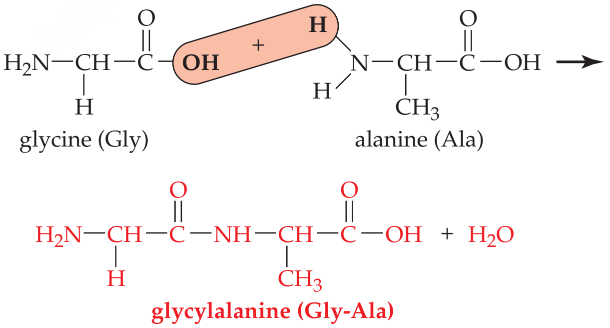
We mentioned previously that proteins consist of amino acids linked together by amide bonds. In a protein, this amide bond is called a **peptide linkage** or a peptide bond. A peptide linkage has the following general structural features:



$$\text{(amine end)}\quad NH_2—\underset{R}{CH}—\overset{O}{\overset{\|}{C}}—NH—\underset{R}{CH}—\overset{O}{\overset{\|}{C}}—OH\quad\text{(carboxyl end)}$$

Peptide linkage

In a **dipeptide**, 2 amino acid units are joined by an amide bond. In a tripeptide, there are 3 amino acids, and in a **polypeptide**, there are up to 50 amino acids. Human insulin, for example, is a polypeptide composed of 48 amino acid units. We can illustrate a specific dipeptide by joining the 2 amino acids glycine and alanine.

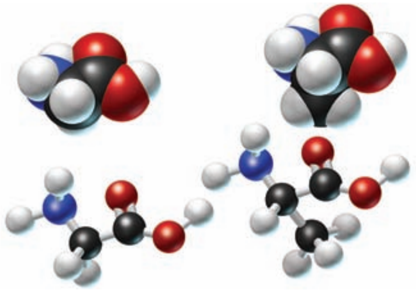


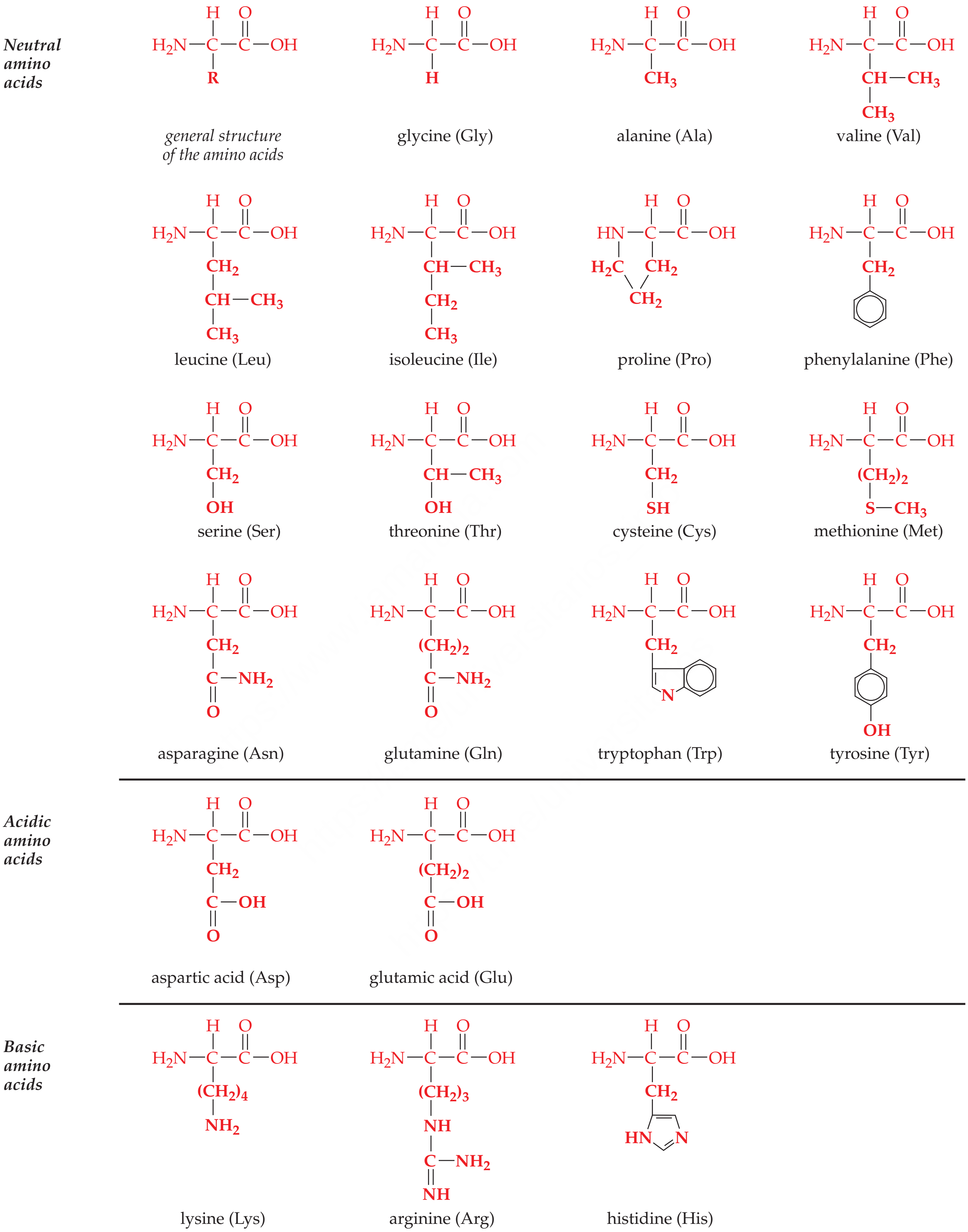
$$H_2N—\underset{H}{CH}—\overset{O}{\overset{\|}{C}}—OH + H—N(H)—\underset{CH_3}{CH}—\overset{O}{\overset{\|}{C}}—OH \longrightarrow$$

glycine (Gly) alanine (Ala)

$$H_2N—\underset{H}{CH}—\overset{O}{\overset{\|}{C}}—NH—\underset{CH_3}{CH}—\overset{O}{\overset{\|}{C}}—OH + H_2O$$

glycylalanine (Gly-Ala)





▲ **Figure 20.2** **Common Amino Acids** Notice that the amino acids are similar except for the R— group. The R— side chain is responsible for the amino acid having neutral, acidic, or basic properties.

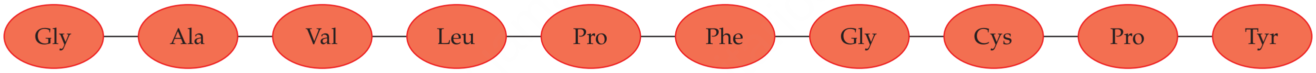
The dipeptide of glycine and alanine is named glycylalanine and is abbreviated Gly-Ala. The sequence of amino acids in a peptide or protein is always listed beginning with the amine end of the chain on the left. The other amino acids in the chain are listed in order, and the amino acid with the free carboxyl group is mentioned last. In the above mentioned dipeptide example, the amine end of the peptide chain is glycine (Gly) and the carboxyl end is alanine (Ala).

Primary Structure of Proteins

Proteins are polypeptides that may contain hundreds or thousands of amino acid units linked together in a long-chain molecule. Egg whites, for example, contain a protein chain with nearly 400 amino acid molecules joined by peptide linkages.

One of the most amazing scientific achievements of the twentieth century was the unraveling of the mystery of protein synthesis in living organisms. In the 1960s, biochemists discovered how living organisms synthesize huge protein molecules in which thousands of amino acids are always in exactly the same sequence. When an organism synthesizes a protein, the sequence of amino acids is determined by a genetic code embodied in the structure of a nucleic acid.

Using currently available instruments, biochemists can determine the exact sequence of amino acids in small proteins. The sequence of amino acids in a peptide chain is referred to as the *primary structure* of the peptide. Figure 20.3 shows the primary structure of a polypeptide with 10 amino acids.

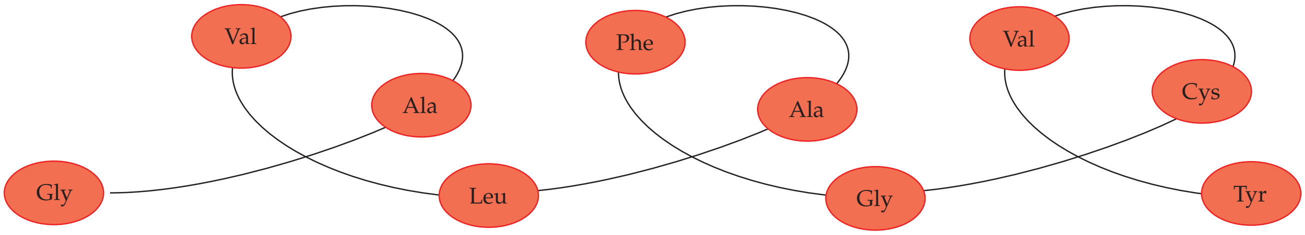


▲ **Figure 20.3 Primary Structure of a Protein** The primary structure of a protein refers to the sequence of amino acids present in the polymer chain. (Each ellipse represents an amino acid.)

The replacement of a single amino acid by another amino acid in the primary structure of a protein can alter its biological activity. Even a small change in the primary sequence of amino acids can drastically alter the biochemistry of a protein. For example, the disease sickle-cell anemia results from an amino acid substitution in hemoglobin. Hemoglobin is the protein responsible for oxygen transport through the circulatory system. Sickle-cell anemia, an extremely serious, often fatal condition, results when a single amino acid, glutamic acid, is replaced by the amino acid valine.

Secondary Structure of Proteins

When a peptide chain twists and bends, the protein acquires a *secondary structure*. One type of secondary structure exhibited by proteins is referred to as an $\alpha-$helix. This is the structure of protein molecules found in hair and wool fibers. An $\alpha-$helix structure is analogous to that of a coiled telephone cord. That is, an $\alpha-$helix is a coiled chain of amino acids. Figure 20.4 illustrates the $\alpha-$helix secondary structure of a hypothetical polypeptide composed of ten amino acids.



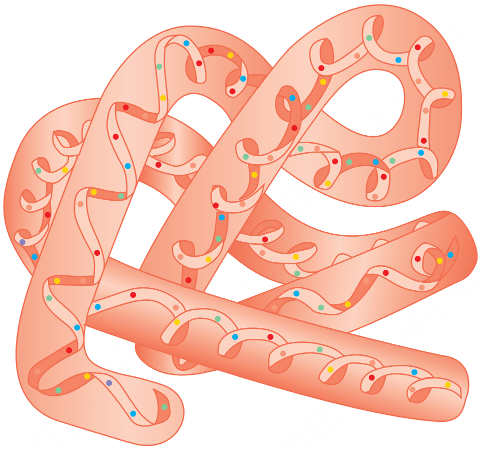
▲ **Figure 20.4 Secondary Structure of a Protein: $\alpha-$Helix** The secondary structure of a protein may be in the shape of a coiled chain of amino acids referred to as an $\alpha-$helix. (Each ellipse represents an amino acid.)

What maintains the secondary structure of a protein? The answer is hydrogen bonding (Section 12.9). Hydrogen bonds between amino acids are responsible for the secondary structure of a protein. In an α−helix, the hydrogen bonds are between nearby amino acid units in the same coiled protein chain.

Tertiary Structure of Proteins

The overall three-dimensional structure of a protein is referred to as its *tertiary structure.* The tertiary structure of a protein may be long and extended or compact and folded. As an analogy, the tertiary structure of a protein may be compared to a telephone cord that is straight and extended or twisted about itself. The twisting of an α−helix illustrates a tertiary structure. The formation of a twisted helix results in a protein molecule with the appearance of a globule.

The tertiary structure of a protein is held together by intermolecular forces, primarily hydrogen bonds. Figure 20.5 illustrates the tertiary structure of myoglobin. Myoglobin is a globular protein that stores oxygen in animal tissue until it is required for a metabolic activity.

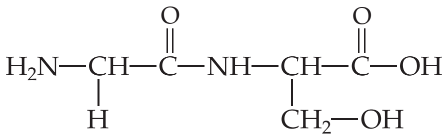


◀ **Figure 20.5 Tertiary Structure of a Protein** A twisted α−helix can create the tertiary structure of a protein. This model structure shows the coiled secondary structure of myoglobin twisted into a tertiary structure.

Hydrogen bonds are weaker than covalent bonds, and the tertiary structure of a protein can collapse with a change in pH or temperature. The process of breaking hydrogen bonds in a protein molecule is called *denaturation*. A familiar example of denaturing a protein is boiling an egg in hot water.

EXAMPLE EXERCISE 20.2 Proteins

Identify the amino acids in the following dipeptide.



$$\mathrm{H_2N{-}\underset{\displaystyle H}{\underset{|}{CH}}{-}\overset{\displaystyle O}{\overset{\|}{C}}{-}NH{-}\underset{\displaystyle CH_2{-}OH}{\underset{|}{CH}}{-}\overset{\displaystyle O}{\overset{\|}{C}}{-}OH}$$

Solution

A dipeptide has two amino acids joined by a peptide linkage. If we refer to Figure 20.2, we can identify the two amino acids in the dipeptide as glycine (left) and serine (right).

Practice Exercise

Classify glycine and serine as neutral, acidic, or basic amino acids.

Answers:

If we refer to Figure 20.2, we find that glycine and serine are both examples of neutral amino acids. Note that neither amino acid has a carboxyl group or an amine group on the R— portion of the amino acid structure.

Concept Exercise

The structure of a protein is described as a helix of amino acids in a spiral protein chain. Does this description correspond to the primary, secondary, or tertiary structure of a protein molecule?

Answers:

See Appendix G, 20.2.

20.3 Enzymes

LEARNING OBJECTIVES

- Analyze the lock-and-key model for an enzyme.
- Analyze the action of an enzyme as a catalyst.

A single human cell may contain many enzymes that are crucial for life processes. An **enzyme** is a protein that acts as a biological catalyst. Enzymes are remarkably selective for specific molecules. An enzyme can speed up a biochemical reaction so that the rate is a million times faster than it would be in the absence of the enzyme. In fact, without enzymes, these biochemical reactions would take place so slowly that life would be impossible. Furthermore, only a very small amount of enzyme catalyst is necessary because a biochemical reaction can occur in microseconds, and the enzyme is then free to repeat its role as catalyst.

Biochemical reactions require sensitive conditions that cannot withstand the conditions of extreme pH or high temperature often associated with nonbiochemical reactions. If the pH or temperature is too high or too low, an enzyme becomes inactive and a biochemical reaction cannot take place. A chemical that renders an enzyme inactive is called an *inhibitor*. Sarin and phosgene are inhibitors that have been used as chemical weapons. They act as nerve agents and block the enzyme responsible for breaking down lethal toxins that accumulate naturally in the body.

Lock-and-Key Model

The details of an enzyme's function are complex, but we can use a **lock-and-key model** to describe the general mechanism. The lock-and-key theory proposes that an enzyme molecule has a given structural feature complementary in shape to the shape of the reacting molecule. This location on the enzyme where a reaction can occur is called the *active site* and the reacting molecule is referred to as the *substrate.* Figure 20.6 illustrates the lock-and-key model using a lock and key as an analogy.

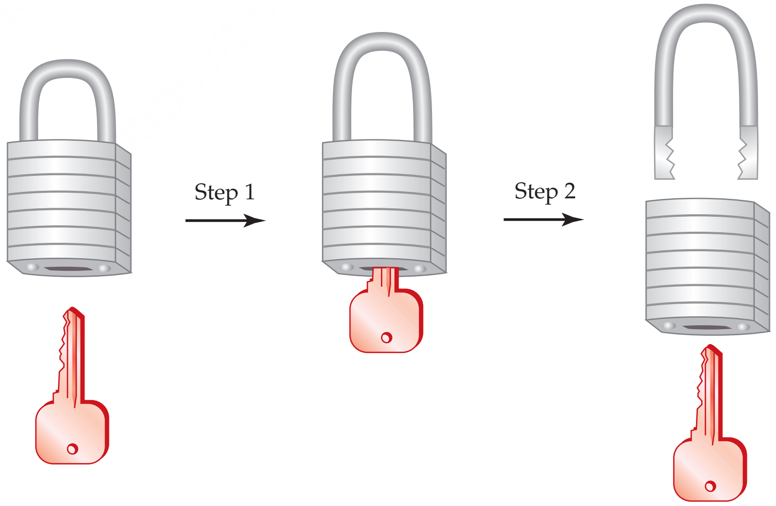
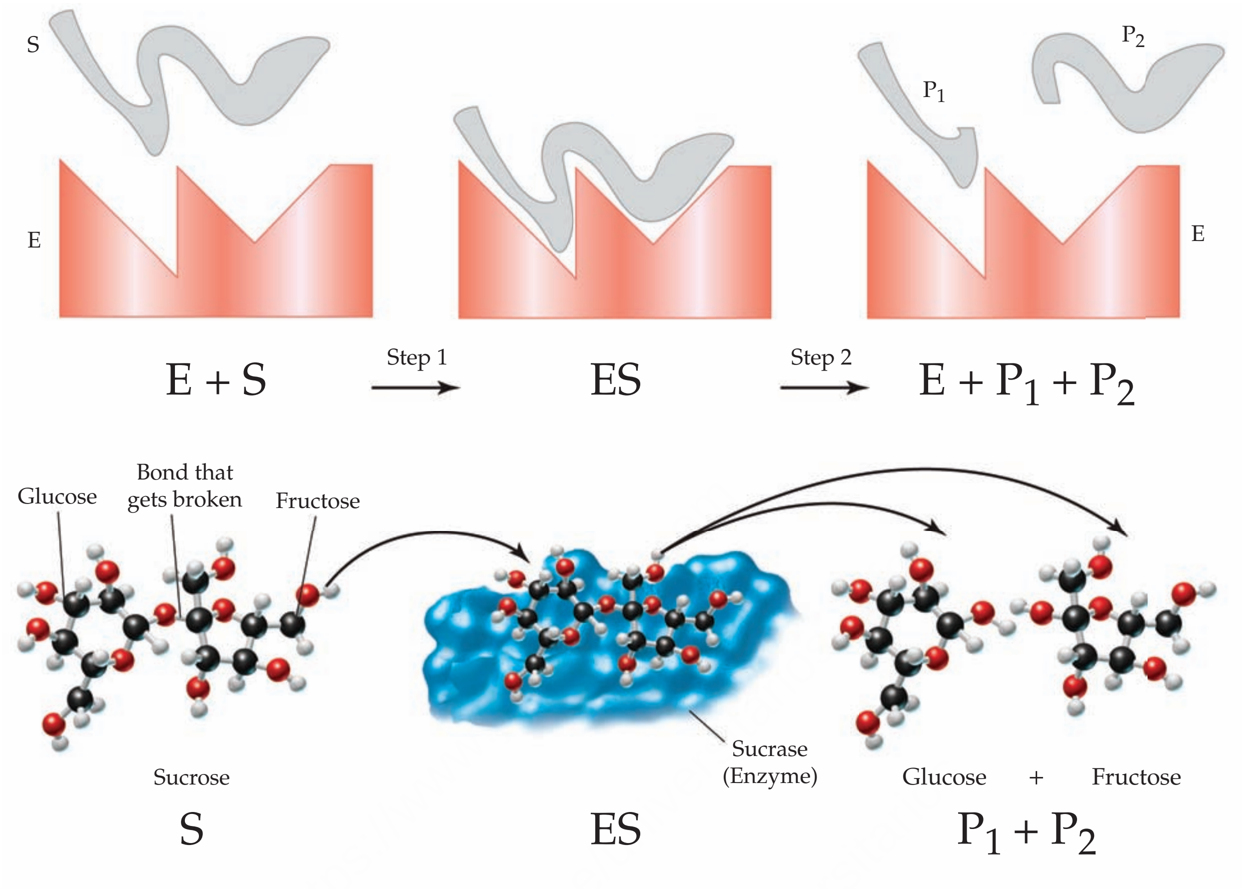


Figure 20.6 Analogy for Enzyme Catalysis In Step 1, the key (enzyme) is inserted into the lock (substrate). In Step 2, the key opens the lock and releases the two parts (products).

An enzyme catalyzed reaction takes place in two steps. First, the substrate (S) binds to the active site on the enzyme (E). Second, the substrate molecule undergoes a reaction and the enzyme releases two or more products (P_1 and P_2).

$$\text{Step 1:} \quad E + S \longrightarrow ES$$
$$\text{Step 2:} \quad ES \longrightarrow E + P_1 + P_2$$

Figure 20.7 illustrates this two-step process for an enzyme catalyzed reaction of a substrate molecule to give two products.



▲ **Figure 20.7 Enzyme Catalysis** In Step 1, the substrate molecule (S) binds to the active site on the enzyme (E). In Step 2, the enzyme cleaves the substrate and releases the two products, P_1 and P_2.

After the enzyme releases the products, the enzyme is then free to bind to another substrate molecule and repeat the process all over again. Enzyme catalysis is quite rapid, and a trace amount of enzyme can yield a large amount of product in a short period of time.

You can demonstrate enzyme catalysis with a saltine cracker. Although the cracker is not sweet, it contains starch, which is a polymer of glucose sugar. If you place the cracker in your mouth, in a short time you will notice a sweet taste. The explanation is that saliva contains the enzyme ptyalin, which begins the digestion process by breaking down starch molecules into sugar units. If the sweet taste takes a minute to develop with the enzyme, it would take about a million minutes for the taste to develop without the enzyme—that is almost two years!

20.4 Carbohydrates

Carbohydrates are an important food source for animals and provide the structure of plants. The word *carbohydrate* literally means "hydrates of carbon" ($C \cdot H_2O$), and indeed many carbohydrates have the empirical formula CH_2O. The names of carbohydrates usually end in the suffix *−ose*, for example, glucose, fructose, and sucrose. A carbohydrate is characterized by having an aldehyde or ketone functional group and, in addition, several hydroxyl (—OH) groups.

A simple sugar molecule typically has three to six carbons, an aldehyde or ketone group, and a few hydroxyl groups. A simple sugar molecule is referred to as

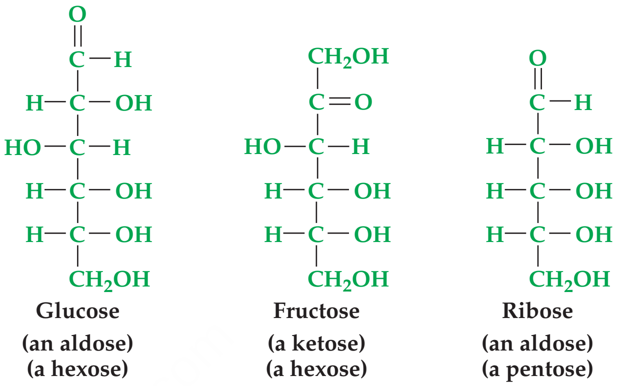
LEARNING OBJECTIVES

- Recognize the structure of mono-, di-, and polysaccharides.
- Identify a glycoside linkage between monosaccharides.

▲ **Carbohydrates** Bread, fruits, and vegetables are all food sources rich in carbohydrate.

a **monosaccharide**. If a monosaccharide has an aldehyde group, it is referred to as an *aldose;* if it has a ketone group, it is called a *ketose*. If the sugar molecule contains five carbons, it is referred to as a *pentose;* if it contains six carbons, it is called a *hexose.*

Glucose, $C_6H_{12}O_6$, and fructose, $C_6H_{12}O_6$, are both hexoses. However, glucose has an aldehyde group and is an aldose; fructose has a ketone group and is a ketose. Ribose, $C_5H_{10}O_5$, is an important pentose found in nucleic acids. Figure 20.8 illustrates the structures of these three simple sugars.



▶ **Figure 20.8 Common Monosaccharides** Glucose and fructose are both hexoses, but notice that glucose is an aldose and fructose is a ketose. Glucose and ribose are both aldoses, but notice that glucose is a hexose and ribose is a pentose.

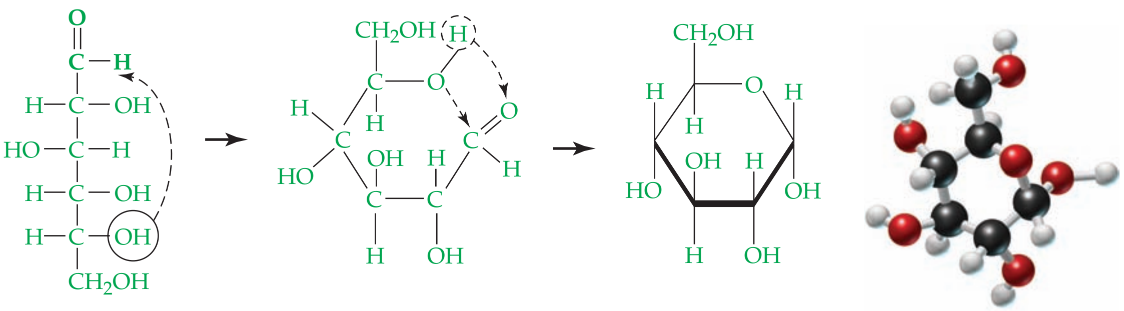
Glucose occurs in grapes and blood and, accordingly, is referred to as *grape sugar* or *blood sugar*. Fructose occurs naturally in fruit and honey and is commonly referred to as fruit sugar. Although they have the same molecular formula, fructose is twice as sweet as glucose.

A **disaccharide** is composed of two monosaccharides bonded together. For instance, maltose is malt sugar, which is composed of two glucose molecules. Lactose is milk sugar, which is comprised of glucose and galactose. Sucrose is ordinary table sugar derived from sugarcane and sugar beets. Sucrose is a disaccharide comprised of glucose and fructose. In the presence of an enzyme or aqueous acid, a disaccharide breaks down into its two component monosaccharides. The equation for the acid *hydrolysis* of maltose, lactose, and sucrose is

$$\begin{array}{ccccccc}
C_{12}H_{22}O_{11} & + & H_2O & \xrightarrow{H^+} & C_6H_{12}O_6 & + & C_6H_{12}O_6 \\
\text{maltose} & + & \text{water} & \longrightarrow & \text{glucose} & + & \text{glucose} \\
\text{lactose} & + & \text{water} & \longrightarrow & \text{galactose} & + & \text{glucose} \\
\text{sucrose} & + & \text{water} & \longrightarrow & \text{glucose} & + & \text{fructose}
\end{array}$$

Structures of Sugars in an Aqueous Solution

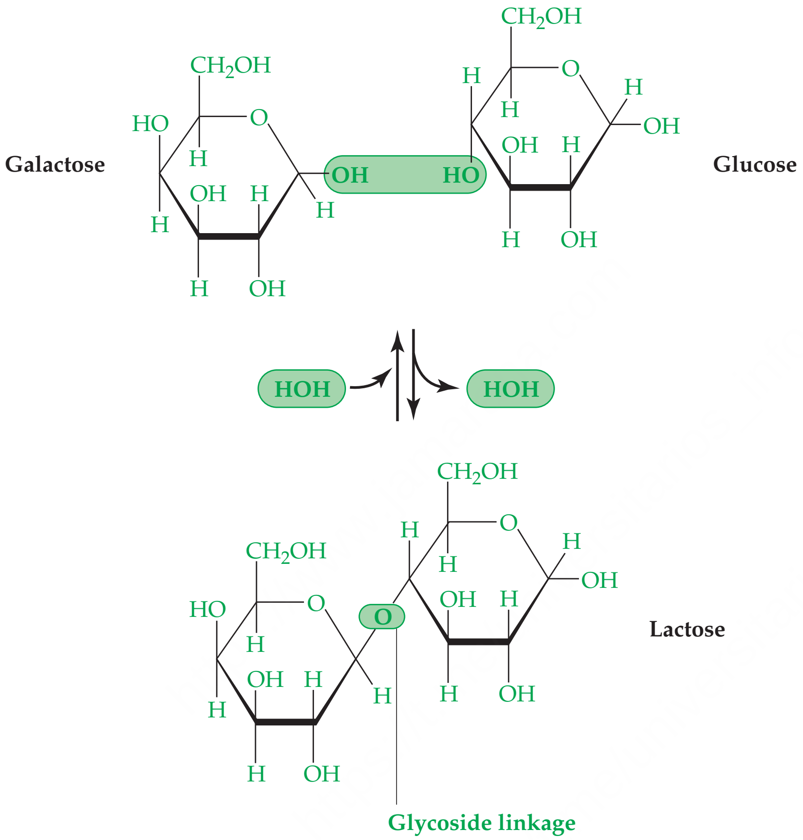
In an aqueous solution, sugar molecules usually exist as ringlike structures. In water, glucose forms a cyclic structure by joining a hydroxyl group to the aldehyde group. The resulting ring structure contains five carbon atoms and one oxygen atom. Figure 20.9 illustrates the formation of a glucose ring in the solution.



▲ **Figure 20.9 Glucose in an Aqueous Solution** In an aqueous solution, a glucose molecule forms a ring structure containing five carbon atoms and one oxygen atom.

In the formation of a disaccharide, two simple monosaccharides eliminate HOH and are joined together by a special —O— bond called a **glycoside linkage**. For instance, a molecule of galactose and glucose can eliminate water and form a molecule of lactose.

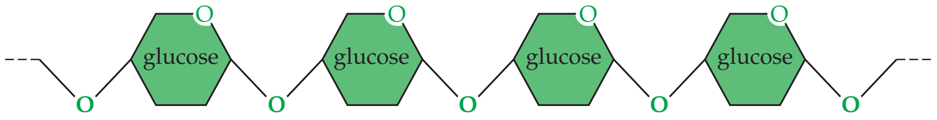
In the process of digestion, the enzyme lactase breaks down lactose to yield galactose and glucose. It is interesting to note that adults have a less active form of this enzyme than infants. As a result, some adults cannot break down lactose and consequently may suffer indigestion. This condition, known as *lactose intolerance*, is treated by avoiding milk products or by taking a lactase supplement in the form of a liquid or tablet. Figure 20.10 illustrates the formation and breakage of a glycoside linkage in lactose.



◀ **Figure 20.10 Formation and Breakage of a Glycoside Linkage** Lactose is composed of two simple sugars, glucose and galactose. In the formation of a glycoside linkage, a molecule of galactose and glucose split out a water molecule. In the reverse reaction, the glycoside linkage is cleaved in the presence of an enzyme or acid.

Polysaccharides

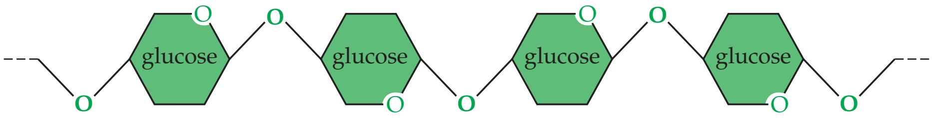
A **polysaccharide** is a polymer composed of many, often thousands, of monosaccharide units joined by glycoside linkages. Starch and cellulose are important examples of polysaccharides. Starch is a polysaccharide composed of the same repeating unit. When starch undergoes hydrolysis and is broken down with dilute acid, or an enzyme, the only product obtained is glucose. Figure 20.11 illustrates a segment of a starch polymer showing a simplified representation of glucose molecules.



▲ **Figure 20.11 A Segment of Starch** Starch is a polymer of repeating glucose units. In addition to the straight chain molecule shown, a starch molecule has numerous branches that are also composed of glucose units.

In the human body, glucose monomers can bond to one another to form *glycogen*, a small polysaccharide stored in the muscles and liver as an energy resource. Excess glucose, however, is converted to lipids and stored in the body as fat. When energy is expended for physical activity, glycogen is released. The glycogen in turn is broken down to release glucose molecules. Glucose molecules are further broken down to produce carbon dioxide, water, and energy for the body.

Cellulose is a polysaccharide found in plants and trees. When cellulose is broken down with dilute acid, the only product obtained is glucose. Thus, starch and cellulose are both polysaccharides composed exclusively of glucose units. Although the distinction is subtle, the glucose units in starch and cellulose are not linked together in exactly the same way. In cellulose every other glucose unit is flipped upside down. Figure 20.12 illustrates a segment of a cellulose polymer showing a simplified version of glucose molecules.



◀ **Figure 20.12 A Segment of Cellulose** Cellulose is a polymer of repeating glucose units. Note the position of the "O" in each glucose ring, which indicates that every other glucose unit is flipped upside down.

The subtle structural difference between starch and cellulose reveals that nature is highly selective; that is, most animals possess an enzyme that digests starch, but not cellulose. However, bacteria in the digestive tract of grazing animals and termites have enzymes that can cleave the inverted glycoside linkages in cellulose. Thus, cows can digest the cellulose in grass, and termites can digest wood. Humans, however, cannot digest either grass or wood because we lack the enzyme to cleave the —O— bonds in cellulose.

20.5 Lipids

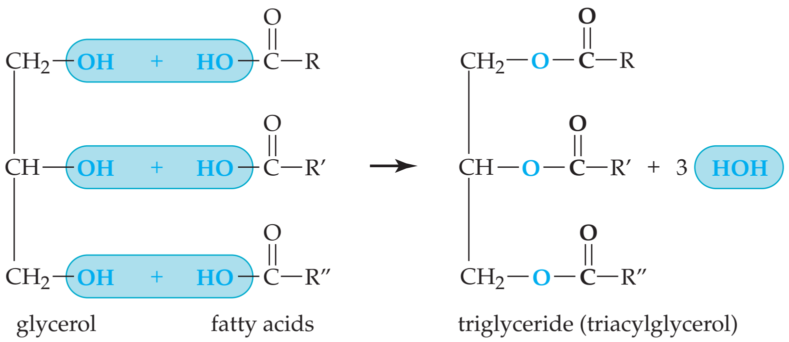
LEARNING OBJECTIVES

- Recognize the structure of a triglyceride.
- Recognize the structure of a phospholipid.
- Recognize the structure of a lipid wax.
- Recognize the structure of a steroid.

Lipids are biochemical compounds that are not soluble in water. This property sets them apart from proteins, carbohydrates, and nucleic acids that are soluble in water. Lipids include the familiar compounds we refer to as fats, oils, and waxes. Lipids also include steroids, such as cholesterol, testosterone, and progesterone.

Triglycerides (Triacylglycerols)

A **triglyceride** (or triacylglycerol) is a lipid formed from glycerol, a trihydroxy alcohol, and three long-chain carboxylic acids. A carboxylic acid with a long hydrocarbon chain is referred to as a **fatty acid**. When three fatty acids (RCOOH, R′COOH, R″COOH) react with glycerol, $C_3H_5(OH)_3$, a triglyceride is produced that has three ester linkages. We can represent the general reaction for the formation of a triglyceride as follows:



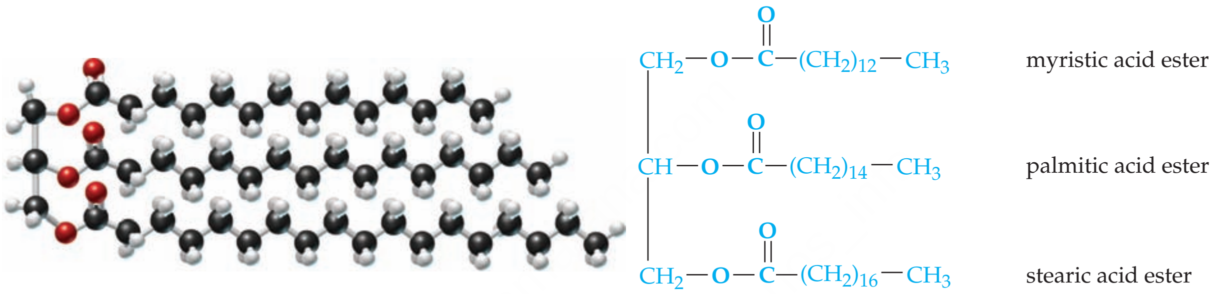
The notations R, R′, and R″ represent long hydrocarbon chains on the three fatty acids. If the triglyceride is obtained from an animal source, the fatty acid chains are mostly saturated and have few double bonds. A semisolid lipid obtained from an animal source is called a **fat**. Table 20.1 lists some common saturated fatty acids present in animal fat.

TABLE 20.1 Common Saturated Fatty Acids

Name	Structure
lauric acid	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—COOH$ $CH_3—(CH_2)_{10}—COOH$
myristic acid	$CH_3—(CH_2)_{12}—COOH$
palmitic acid	$CH_3—(CH_2)_{14}—COOH$
stearic acid	$CH_3—(CH_2)_{16}—COOH$

▲ **Lipids** Vegetable oil, margarine, and creamy foods are rich in lipids.

The following structure illustrates an example of a saturated triglyceride found in animal fat.

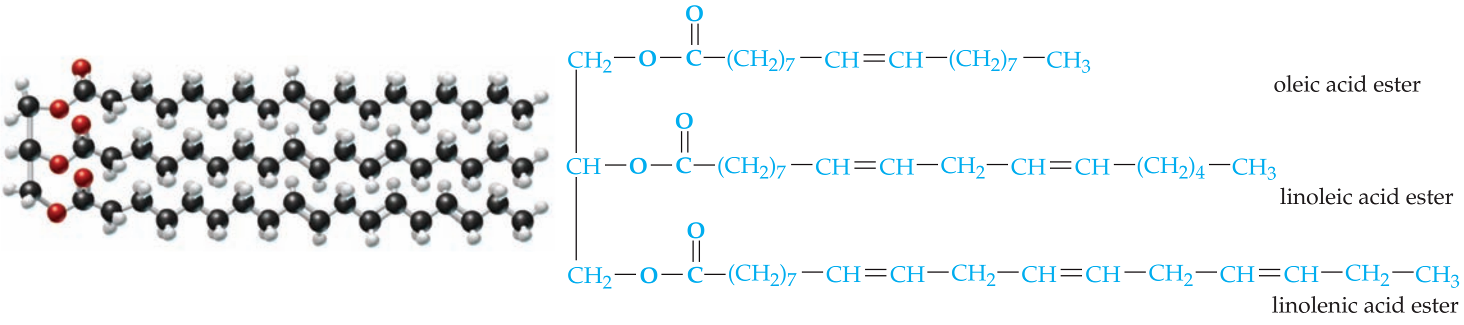


A liquid lipid obtained from a plant source is called an **oil**. If a triglyceride is obtained from a plant source, the fatty acid chains are mostly unsaturated and have one or more double bonds. Table 20.2 lists some common unsaturated fatty acids that occur in vegetable oils.

TABLE 20.2 Common Unsaturated Fatty Acids

Name	Structure
oleic acid	$CH_3—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH_2—CH=CH—(CH_2)_7—COOH$
linoleic acid	$CH_3—CH_2—CH_2—CH_2—CH_2—CH=CH—CH_2—CH=CH—(CH_2)_7—COOH$
linolenic acid	$CH_3—CH_2—CH=CH—CH_2—CH=CH—CH_2—CH=CH—(CH_2)_7—COOH$

The following structure illustrates an example of an unsaturated triglyceride found in vegetable oils. Note that the fatty acids have one, two, and three double bonds, respectively.

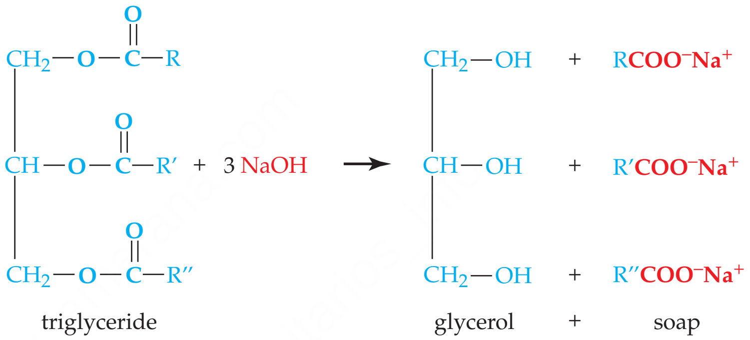


The distinction between a fat and an oil, based on the saturation of the fatty acids, is a generalization with numerous exceptions. That is, even fats have a small percentage of unsaturated fatty acids, and oils contain a few saturated fatty acids. Moreover, a natural source of a fat or oil is a complex mixture of triglycerides.

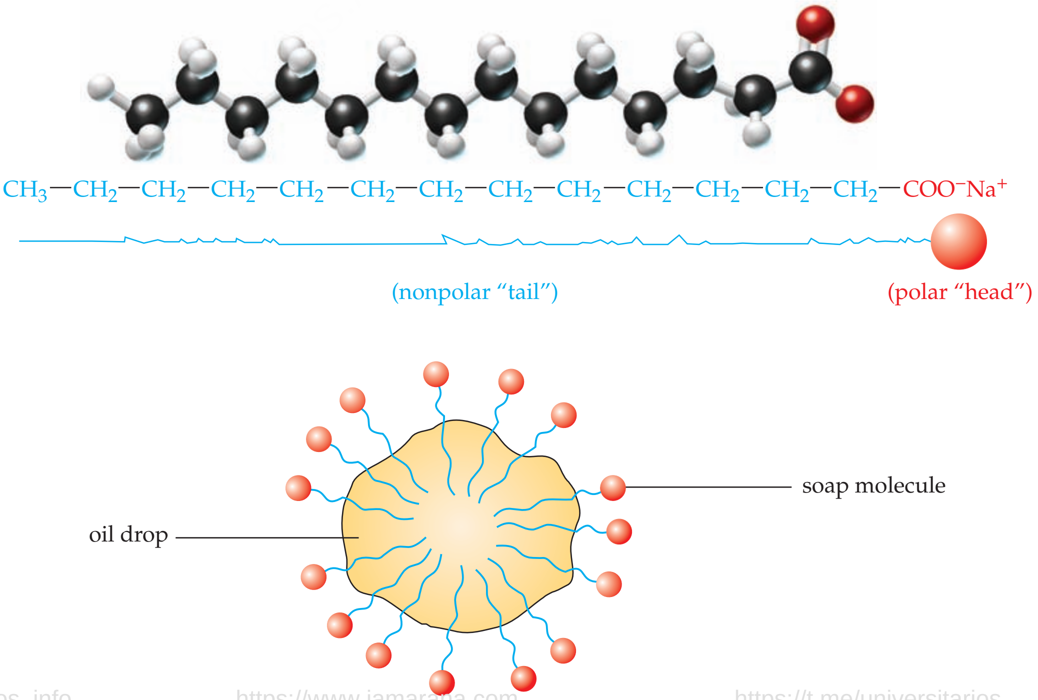
A detailed analysis of a fat or oil indicates the average percentage of each fatty acid. Corn oil, for example, averages about 2% myristic acid, 8% palmitic acid, 5% stearic acid, 35% oleic acid, and 50% linoleic acid. Thus, corn oil averages about 15% saturated fatty acids and 85% unsaturated fatty acids.

Saponification of Fats and Oils

The ester linkage in a triglyceride can be broken by treatment with aqueous sodium hydroxide. The products of the reaction are glycerol and three sodium salts of fatty acids. Sodium salts of fatty acids constitute what we refer to as *soap*. This reaction is called **saponification**, which literally means "soapmaking." We can represent the general reaction for the saponification of a triglyceride as follows:



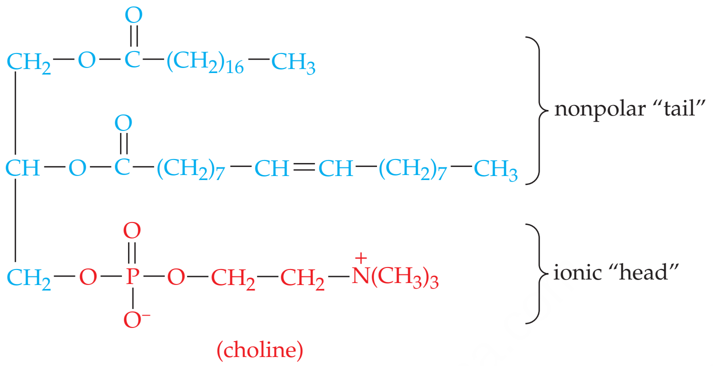
Soap has a polar ionic "head" (COO^- Na^+) and a long nonpolar "tail" (R, R′, or R″). Although oil and water do not mix, we can use soap to remove an oily residue from our hands. Soap orients its fatty acid nonpolar "tail" toward the oily layer on our hands, and its polar "head" toward the water. Soap thereby provides a link between nonpolar oil molecules and polar water molecules (Figure 20.13). Thus, washing our hands with soap and water can remove an oily residue even though oil and water are not miscible.



▶ **Figure 20.13 The Action of Soap** A soap molecule ($R-COO^-Na^+$) has a long nonpolar "tail" and a polar "head" ($-COO^-Na^+$). Soap is effective in removing oil because soap has a "tail" that is soluble in oil, and a "head" that is soluble in water.

Phospholipids

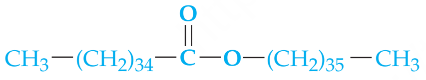
A **phospholipid** is a glyceride found in living cells and cell membranes. Unlike fats and oils, which have *three* fatty acids attached to glycerol, phospholipids have only *two* fatty acids. The third —OH in glycerol is linked to phosphoric acid. Thus, the properties of phospholipids are different from those of fats and oils. With the addition of a phosphate group, a phospholipid has an ionic "head" that is soluble in water, as well as a nonpolar "tail" that is insoluble. In fact, the phospholipid lecithin is used to emulsify margarine and chocolate, which are otherwise insoluble in water. The following example structure illustrates a typical composition of lecithin:



The phosphate ester in lecithin is bonded to choline, $HO—CH_2CH_2N^+(CH_3)_3$, as well as glycerol. Choline is a B-complex vitamin found in egg yolks and is essential to liver function.

Waxes

A lipid **wax** is a mixture of naturally occurring esters of fatty acids and long-chain alcohols. An example is beeswax, which is found in honeycombs and is used to make candles and shoe polish. Another example is carnauba wax, an ingredient in car polish and furniture polish. A lipid wax is water insoluble and occurs naturally as the slippery coating on fruits, vegetables, and plant leaves. Since waxes are insoluble in water, they can be used to waterproof fabrics and wood. The following example structure illustrates one of the typical esters found in beeswax:



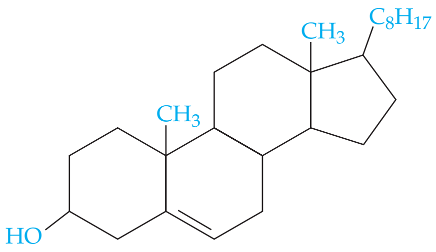
$$CH_3—(CH_2)_{34}—\overset{\overset{\displaystyle O}{\|}}{C}—O—(CH_2)_{35}—CH_3$$

Beeswax

Waxes are esters found in nature and, like triglycerides, demonstrate variety in their fatty acid composition. A typical analysis of carnauba wax reveals a 12-carbon unsaturated fatty acid and a broad mixture of long-chain alcohols.

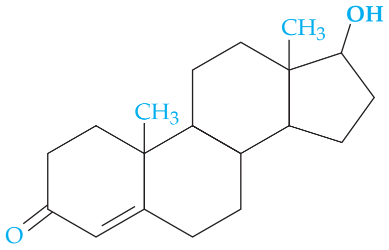
Steroids

A **steroid** belongs to a special class of lipids that has a structure composed of four rings of carbon atoms fused together. Cholesterol is an example of an important steroid and is the precursor of many other steroid molecules essential to human life. On the other hand, high levels of cholesterol have been linked to the formation of plaque in arteries, which can contribute to heart disease. We can diagram the ring structure of cholesterol as follows:

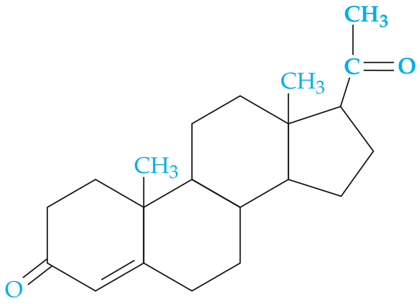


Cholesterol

As mentioned previously, sex hormones are steroids. This includes the hormones testosterone and progesterone secreted by the testes and ovaries, respectively. Testosterone is responsible for the sex characteristics of males, whereas progesterone helps regulate ovulation in females. It is interesting to note that the two hormones are quite similar and differ only slightly in molecular structure.



Testosterone



Progesterone

CHEMISTRY CONNECTION **Vitamins**

Q: Which vitamins are fat-soluble and should not be taken in excessive amounts?

Vitamins are biochemical compounds required for normal metabolism. Vitamins are not synthesized in the human body, and therefore must be included in the diet. If a vitamin is not present in the diet, its absence can lead to a vitamin-deficiency disease.

The U.S. Food and Nutrition Board has established minimal vitamin requirements. The board periodically publishes a new listing of the Recommended Dietary Allowance (RDA) for vitamins based on current research. These RDA values are intended as a guideline for estimating the needs of adults and children and may vary depending on individual circumstances. Ingesting polluted air or water can increase vitamin requirements. Some infectious diseases, accompanied by a high fever, may also result in increased vitamin requirements.

In most instances, a diet containing fresh fruits, vegetables, dairy products, and meat ensures an adequate intake of necessary vitamins. Vitamin supplements, however, may be recommended by a physician during periods of stress, growth, or pregnancy. There is a growing body of evidence that vitamins A, C, and E may lessen the risk of cancer. These vitamins are antioxidants, and the so-called "free-radical theory" proposes that free radicals can alter the structure of DNA. Vitamins C and E, in conjunction with the mineral selenium, help to clear tissues of toxic free radicals.

Excess amounts of water-soluble vitamins (B_1, B_2, B_3, B_6, B_{12}, and C) are rapidly excreted in the urine and are rarely toxic. Although vitamin C is relatively nontoxic, daily doses in excess of a gram may lead to skin rashes and diarrhea. A deficiency of vitamin C, ascorbic acid, can lead to the disease scurvy, which is characterized by loose teeth and bleeding gums. Citrus fruits prevent and cure this disease. In the early 1800s, the British navy issued rations of lemons and limes to prevent scurvy and British sailors acquired the nickname "limeys."

Excess amounts of fat-soluble vitamins (A, D, E, and K) can accumulate in body fat tissue. Excess vitamin A, retinol, can lead to symptoms of toxin poisoning. The symptoms include irritability, vomiting, loss of appetite, headache, and dry skin. On the other hand, a deficiency of vitamin A has long been known to cause night blindness.

Vitamin D is synthesized by the body when skin is exposed to sunlight. An excess of vitamin D can cause fatigue, headache, and nausea. Vitamin D deficiencies are rare in tropical climates and are more common in northern climates. A vitamin D deficiency results in failure to absorb calcium and phosphorus, causing malformation of the bones. In children, the disease is known as rickets and is characterized by a deformed rib cage and bowed legs.

A supplement of vitamin E, alpha tocopherol, has been advocated for many disorders, especially coronary artery disease. In moderate excess, it is considered relatively nontoxic. Vitamin K is available in milk products and is essential for the clotting of blood. In addition, a wide variety of foods such as vegetables, eggs, and fish contain vitamin K, so a deficiency of this vitamin is rare.

▲ **Vitamins** A sampling of various vitamins that are available as tablets and gelatin capsules.

A: Vitamins A, D, E, and K are fat-soluble and should not be ingested in excessive amounts.

20.6 Nucleic Acids

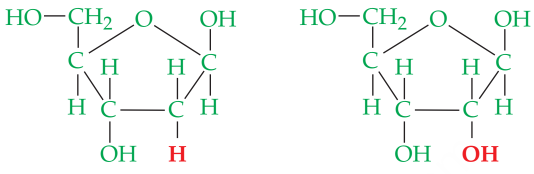
LEARNING OBJECTIVES

- Describe the three components of a nucleotide.
- Distinguish between DNA and RNA nucleotides.
- Explain the process of replication.
- Explain the process of transcription.

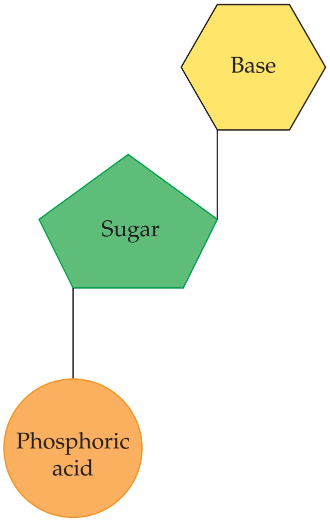
Nucleic acids are biochemical compounds found in every living cell. Nucleic acids contain the genetic information responsible for the reproduction of a species. There are two types of nucleic acids: deoxyribonucleic acid (DNA) and ribonucleic acid (RNA).

A nucleic acid is a polymer composed of many repeating units, each of which is called a **nucleotide**. All nucleotides are similar and consist of a sugar, a base, and phosphoric acid. These three components form the general structure of a nucleotide shown in Figure 20.14.

The sugars found in DNA and RNA differ slightly. That is, a DNA nucleotide contains the sugar deoxyribose, and RNA contains the sugar ribose. These sugars are identical, except for a hydroxyl group (—OH).

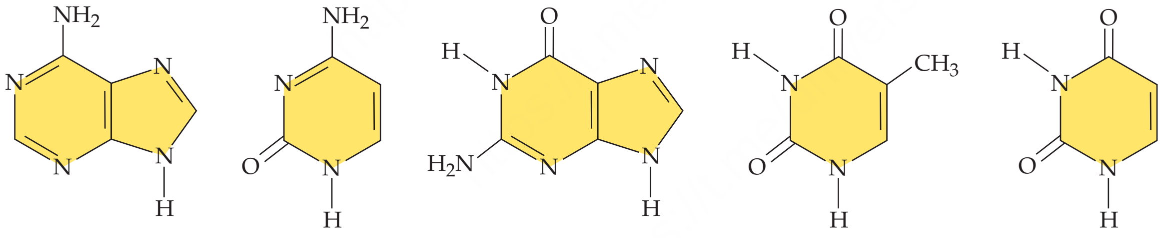

Deoxyribose (DNA) Ribose (RNA)

▲ **Deoxyribose and Ribose** Notice that deoxyribose, the sugar in DNA, and ribose, the sugar in RNA, differ by a single hydroxyl group (—OH).



▲ **Figure 20.14 Nucleic Acid Nucleotide** A nucleic acid is a polymer of nucleotides. A nucleotide is composed of ribose (or deoxyribose) and an attached nitrogen base. Nucleotides are joined together in the nucleic acid by phosphoric acid linkages.

With one exception, the bases found in DNA and RNA are the same. DNA and RNA each contain adenine (A), cytosine (C), and guanine (G). The difference is that DNA contains the base thymine (T), whereas RNA contains the base uracil (U). Thymine and uracil are identical, except for a methyl group. Figure 20.15 shows the structures for the bases in DNA and/or RNA.


Adenine (A) [DNA] [RNA]

Cytosine (C) [DNA] [RNA]

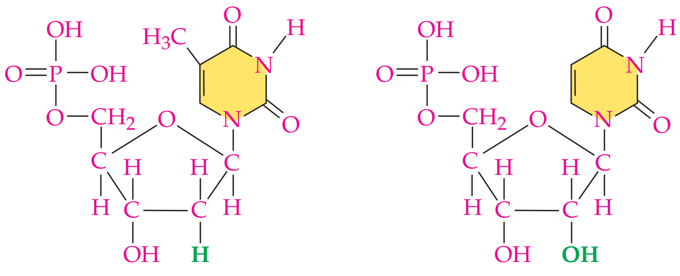
Guanine (G) [DNA] [RNA]

Thymine (T) [DNA]

Uracil (U) [RNA]

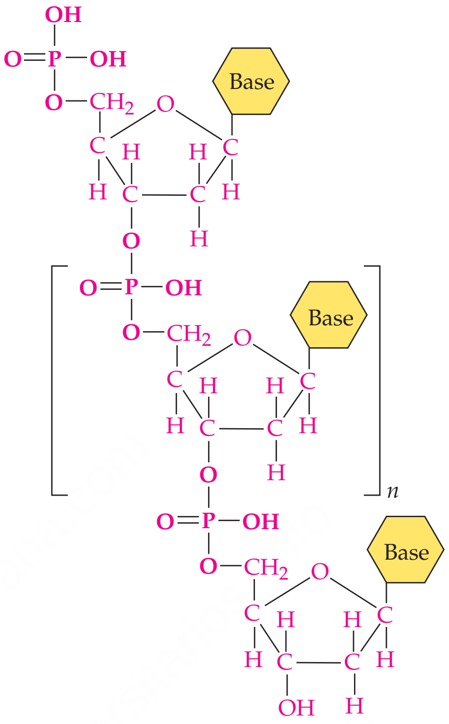
▲ **Figure 20.15 Nucleic Acid Bases** Note that DNA contains the four bases adenine (A), cytosine (C), guanine (G), and thymine (T). RNA contains the four bases A, C, G, and uracil (U).

Now that we are familiar with the three components of a nucleotide, let's see how the sugar, base, and phosphoric acid group are arranged. Figure 20.16 illustrates a DNA nucleotide and an RNA nucleotide.


DNA nucleotide RNA nucleotide

◀ **Figure 20.16 DNA and RNA Nucleotides** Note that the DNA nucleotide contains deoxyribose sugar and that the RNA nucleotide contains ribose sugar. In this example, deoxyribose in the DNA nucleotide has an attached thymine base, and ribose has an attached uracil base.

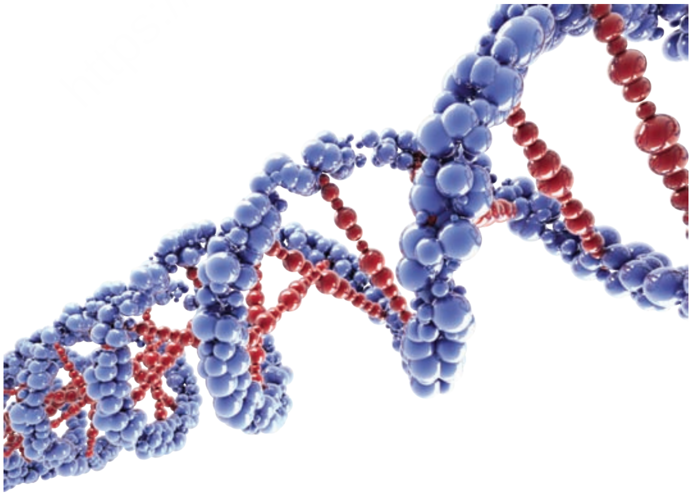
We can draw a short segment of a nucleic acid. That is, we can diagram a few nucleotides joined together in a chain. For simplicity, we will not specify the bases. Figure 20.17 illustrates a small chain of nucleotides found in a segment of a nucleic acid.



▶ **Figure 20.17 A Chain of Nucleotides** A nucleic acid is composed of many nucleotide molecules joined by phosphate linkages. Here the central nucleotide can vary, but is repeated a number of times (n) in the long chain.

The DNA Double Helix

A landmark scientific achievement was the determination of the structure of DNA by the English chemist Francis Crick (1916–2004) and the American biochemist James Watson (1928–). They determined that a DNA molecule is composed of two long chains of DNA nucleotides. These two strands of DNA are wound together in a spiral structure that is referred to as a *double helix*.



▶ **DNA Spiral Helix** The molecular model shows a portion of DNA whose two strands of nucleotides are twisted into an α − helix.

The two strands of the DNA helix are held together by hydrogen bonds between bases on the two separate chains. Specifically, adenine (A) on one strand *always* hydrogen bonds to thymine (T) on the other strand. In Figure 20.18, note that there are two hydrogen bonds between each pair of these bases (A═T or T═A). In contrast, cytosine (C) on one strand *always* hydrogen bonds to guanine (G) on the other strand. Note that there are three hydrogen bonds between each pair of these bases (C≡G or G≡C).

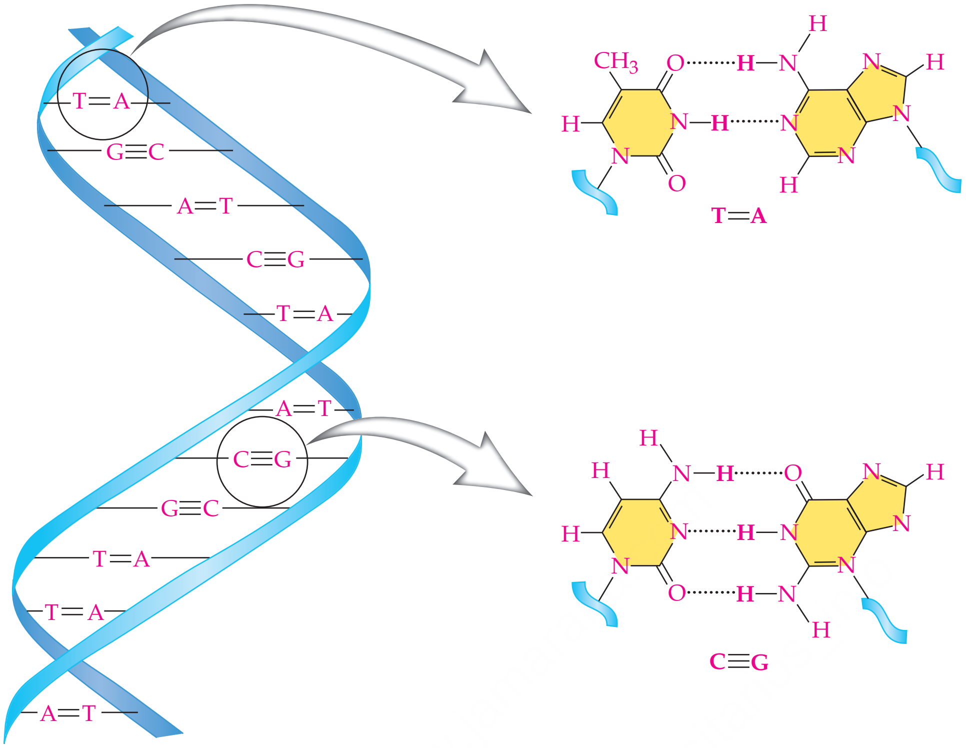


Figure 20.18 **DNA Double Helix** The two DNA strands of nucleotides are held together by hydrogen bonds. The hydrogen bonds are always between the same two bases. Notice there are two hydrogen bonds between thymine (T) and adenine (A) and three hydrogen bonds between cytosine (C) and guanine (G).

The term **replication** describes the process by which a single molecule of DNA synthesizes two molecules of DNA. First, the DNA molecule unwinds by breaking hydrogen bonds between bases on the two strands of the double helix. Then, each single strand of DNA acts as a template to synthesize a new strand of DNA, one nucleotide at a time.

Since the bases always pair in the same way, A=T, T=A, C≡G, or G≡C, each base in the DNA template strand always codes for its complementary base on the growing strand of DNA. That is, adenine (A) on the template strand codes for thymine (T) on the growing strand, thymine (T) codes for adenine (A), cytosine (C) codes for guanine (G), and guanine (G) codes for cytosine (C).

When replication is complete, each strand in the original molecule of DNA has produced a complementary strand of DNA. Each template strand is hydrogen bonded to its complementary strand in a new DNA double helix. The final result is two molecules of DNA, each of which is identical to the original molecule of DNA.

Protein Synthesis

DNA is the parent molecule responsible for transmitting genetic information. DNA can transfer the genetic code to RNA, which contains the instructions for protein synthesis. First, a molecule of DNA in the cell nucleus synthesizes a molecule of RNA; that is, a DNA molecule transcribes its genetic code into an RNA molecule. Next, the RNA molecule moves out of the cell nucleus into the surrounding cytoplasm, where protein synthesis takes place.

Unlike DNA, RNA is a single stranded polymer of nucleotides. The term **transcription** describes the process by which a DNA double helix unwinds and synthesizes two complementary strands of RNA. That is, a string of nucleotides in each DNA strand codes for a complementary strand of nucleotides in a growing RNA chain. Specifically, thymine (T) on the DNA template strand codes for adenine (A) on the growing RNA chain, cytosine (C) codes for guanine (G), and guanine (G) codes for cytosine (C). But unlike replication, adenine (A) codes for uracil (U), not thymine (T), on a growing RNA chain.

The genetic code involves three consecutive nucleotides in an RNA molecule, and each of these trinucleotides is referred to as a **codon**. Each codon specifies a particular amino acid to be used in building a protein. For example, the GCU codon represents three consecutive nucleotides having the bases guanine, cytosine, and uracil, respectively.

The GCU codon is specific for the amino acid alanine, and is responsible for adding alanine to a growing protein chain. Other examples include the CGU codon, which specifies arginine; and the UUU codon, which specifies phenylalanine. Thus, each RNA codon designates a specific amino acid, in a given sequence, in a growing protein chain. This process is termed **translation**.

There are 64 codons corresponding to all the possible combinations of the four bases (4^3). Since there are only 20 amino acids, more than one codon may specify a given amino acid; for example, GCU, GCC, GCA, and GCG all code for the amino acid alanine. There are three so-called "stop codons" (UAA, UAG, UGA) and their function is to terminate a growing protein chain.

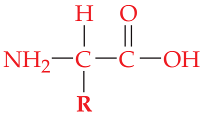
A CLOSER LOOK Left- and Right-Handed Molecules

Q: **Which of the cups shown below has a nonidentical mirror image?**

One of the reasons that there are so many different organic compounds is explained by isomers. In Chapter 19, we learned about structural isomers such as butane and isobutane, and functional group isomers such as alcohols and ethers having the same molecular formula.

In addition to structural isomers, organic compounds can exist as "stereoisomers" that differ in their spatial arrangement of atoms about a central carbon atom. That is, if a carbon atom has four different attached groups (for example, F, Cl, Br, and I), there are two possible spatial orientations and two molecules exist that are nonidentical mirror images. Human hands are an example of nonidentical mirror images. Although our left and right hands are similar, they are not superimposable. If you hold a mirror in front of your left hand, the reflected image in the mirror is not identical to your left hand, but is similar to your right hand. Carefully observe the figure (right) and you will see that the two organic molecules are nonidentical mirror images.

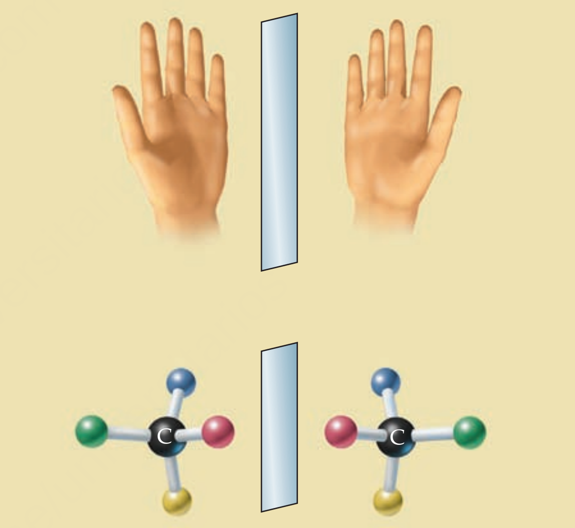
In this chapter we learned that amino acids have an amine group on the alpha (α) carbon. Examine the figure (below) and note that there are four nonidentical groups attached to the alpha carbon atom.



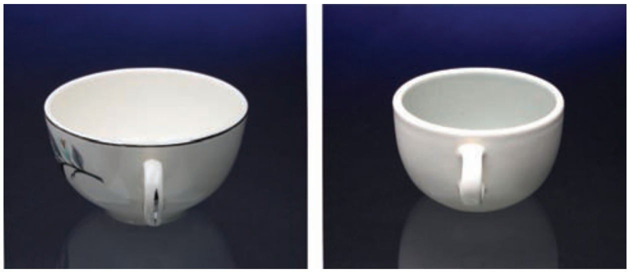
$$\mathrm{NH_2}-\underset{\mathbf{R}}{\overset{\mathrm{H}}{\mathrm{C}}}-\overset{\mathrm{O}}{\overset{\|}{\mathrm{C}}}-\mathrm{OH}$$

Alpha (α) amino acid

Since amino acids have four nonidentical groups on the alpha carbon, we know that stereoisomers are possible. When chemists synthesize amino acids in the laboratory, they obtain both the left- and right-handed stereoisomers, unless a special left- or right-handed reagent is used in the synthesis. However, only one of the stereoisomers is found in nature. Nature produces only the left-handed amino acid and not the right-handed molecule. Similarly, sugar molecules like glucose have four nonidentical groups attached to carbon atoms, which gives rise to stereoisomers. Interestingly, nature produces only the right-handed sugar and not the left-handed molecule.



When we studied isomers in Chapter 19, we learned that isomers had different physical and chemical properties. Stereoisomers have identical properties with one exception. A beam of polarized light passed through a solution containing a stereoisomer rotates the light either to the left or to the right. The nonidentical mirror image rotates the polarized light the same number of degrees in the opposite direction. A solution containing equal concentrations of each stereoisomer exhibits zero rotation of polarized light.

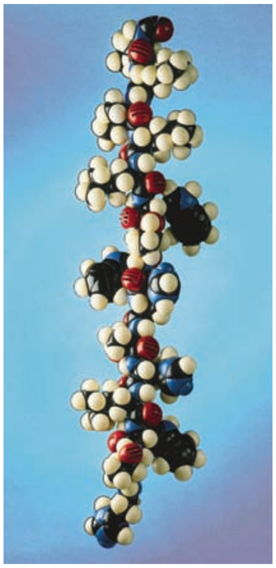


▲ **Unsymmetrical Cup and Symmetrical Cup**

A: The cup on the left has a design on one side; the cup on the right does not. Thus, the cup on the left has a nonidentical mirror image; the cup on the right has an identical mirror image.

Chapter Summary

Key Concepts	Learning Objectives and Related Exercises
20.1 Biological Compounds **Biochemistry** is the study of biological compounds and their chemical reactions. A **protein** is a polymer composed of amino acids. A **carbohydrate** is a compound consisting of one or more simple sugar molecules. A **lipid** is a water-insoluble biological compound such as a fat, oil, wax, or steroid. A **nucleic acid** is a polymer composed of a sugar molecule, a nitrogen base, and phosphoric acid. Although biological compounds can be large and complex, they have familiar functional groups such as amines, alcohols, and esters.	• Recognize a protein given its structural formula. *Related Exercises: 1, 3* • Recognize a carbohydrate given its structural formula. *Related Exercises: 1, 3* • Recognize a lipid given its structural formula. *Related Exercises: 2, 4* • Recognize a nucleic acid given its structural formula. *Related Exercises: 2, 4*
20.2 Proteins A carboxylic acid having an amine group is called an **amino acid**. A **dipeptide** is two amino acids joined by an amide bond called a **peptide linkage**. A **polypeptide** has several amino acids joined by peptide linkages. A protein is a large polypeptide that can contain thousands of amino acids. The structure of a protein is related to its biological function. Extended protein molecules are responsible for shape and structure, whereas globular proteins often regulate hormonal activity.	• Describe the primary structure of a protein. *Related Exercises: 5* • Describe the secondary structure of a protein. *Related Exercises: 6* • Describe the tertiary structure of a protein. *Related Exercises: 7 – 8* • Identify amino acids and a peptide linkage. *Related Exercises: 9 – 18*
20.3 Enzymes An **enzyme** is a protein that acts as a catalyst for biochemical reactions. An enzyme is highly specific for a given biological molecule and is quite sensitive to changes in pH and temperature. Although the mechanism of enzyme activity is complex, the **lock-and-key model** explains the selectivity of enzymes. To speed up a biochemical reaction, the shape of the substrate molecule must fit the contour of the active site on an enzyme.	• Analyze the lock-and-key model for an enzyme. *Related Exercises: 19 – 22* • Analyze the action of an enzyme as a catalyst. *Related Exercises: 23 – 24*
20.4 Carbohydrates A carbohydrate is a biological compound that contains one or more sugar molecules. Glucose is a **monosaccharide** corresponding to one simple sugar molecule. Sucrose is a **disaccharide** and contains molecules of glucose and fructose joined by a **glycoside linkage**. Starch is a **polysaccharide** and contains many molecules of glucose joined by glycoside linkages.	• Recognize the structure of mono-, di-, and polysaccharides. *Related Exercises: 25 – 30* • Identify a glycoside linkage between monosaccharides. *Related Exercises: 31 – 32*
20.5 Lipids Many lipids contain glycerol, an alcohol with three hydroxyl groups. A **fatty acid** is a carboxylic acid bearing a long hydrocarbon chain. A **triglyceride** is an ester of glycerol and three fatty acids. An animal **fat** is a triglyceride with mostly saturated fatty acids. A vegetable **oil** is a triglyceride with mostly unsaturated fatty acids. Triglycerides react with sodium hydroxide to produce soap and glycerol in a chemical reaction called **saponification**. A **phospholipid** contains two fatty acids and phosphoric acid joined to glycerol by ester linkages. A lipid **wax** is a simple ester of a fatty acid and a long-chain alcohol. A **steroid** is a lipid with a characteristic structure of four rings of carbon atoms. Steroids include many hormones and fat-soluble vitamins.	• Recognize the structure of a triglyceride. *Related Exercises: 33 – 36* • Recognize the structure of a phospholipid. *Related Exercises: 37 – 38* • Recognize the structure of a lipid wax. *Related Exercises: 39 – 40* • Recognize the structure of a steroid. *Related Exercises: 41 – 42*



▲ **α-Helix Model** The molecular model shows a portion of a protein chain whose secondary structure is twisted into an alpha helix.

Key Concepts

20.6 Nucleic Acids

A nucleic acid is a biological compound that carries genetic information. Nucleic acids are polymers, which have repeating units called nucleotides. A **nucleotide** is composed of a sugar, a base, and phosphoric acid. Deoxyribonucleic acid (DNA) and ribonucleic acid (RNA) are two types of nucleic acids. DNA and RNA have slightly different nucleotides whose composition is summarized in Table 20.3.

When a single molecule of DNA synthesizes an identical molecule of DNA, the process is termed **replication**. First, the DNA molecule unwinds by breaking hydrogen bonds between bases in the double helix. Then, each strand of DNA synthesizes a complementary strand of DNA, one nucleotide at a time. Similarly, when a molecule of DNA unwinds and acts as a template for the synthesis of a complementary strand of RNA, the process is termed **transcription**. The addition of an amino acid to a growing protein chain during protein synthesis is specified by an RNA trinucleotide, which is called a **codon,** in a process termed **translation.**

Learning Objectives and Related Exercises

- Describe the three components of a nucleotide.
  Related Exercises: 43 – 48
- Distinguish between DNA and RNA nucleotides.
  Related Exercises: 49 – 50
- Explain the process of replication.
  Related Exercises: 51 – 52
- Explain the process of transcription.
  Related Exercises: 53 – 54

TABLE 20.3 DNA versus RNA Nucleotides

Component	DNA	RNA
Sugar	deoxyribose	ribose
Base	adenine (A)	adenine (A)
	cytosine (C)	cytosine (C)
	guanine (G)	guanine (G)
	thymine (T)	uracil (U)
Inorganic Acid	phosphoric acid	phosphoric acid

Key Terms

Answers to Key Terms are in Appendix H.

Select the key term that corresponds to each of the following definitions.

_____ **1.** the study of biological compounds
_____ **2.** a biological compound that is a polymer of amino acids
_____ **3.** a biological compound composed of one or more sugar molecules
_____ **4.** a biological compound that is insoluble in water, such as a fat, oil, or wax
_____ **5.** a biological polymer compound that can transmit genetic information
_____ **6.** a carboxylic acid with an amine group on the alpha carbon
_____ **7.** an amide bond that joins two amino acids
_____ **8.** two amino acids joined by a peptide linkage
_____ **9.** 10–50 amino acids joined by peptide linkages
_____ **10.** a protein molecule that catalyzes a biochemical reaction
_____ **11.** a theory that explains enzyme catalysis
_____ **12.** a carbohydrate composed of a simple sugar molecule
_____ **13.** a carbohydrate composed of two simple sugar molecules
_____ **14.** a carbohydrate composed of many simple sugar molecules
_____ **15.** an —O— bond that joins two simple sugars
_____ **16.** a carboxylic acid with a long hydrocarbon chain
_____ **17.** a lipid composed of glycerol and three fatty acids
_____ **18.** a triglyceride from an animal source that has mostly saturated fatty acids
_____ **19.** a triglyceride from a plant source that has mostly unsaturated fatty acids
_____ **20.** a chemical reaction of a fat or oil with sodium hydroxide to produce soap
_____ **21.** a lipid composed of glycerol, two fatty acids, and phosphoric acid
_____ **22.** a lipid composed of a fatty acid and a long-chain alcohol
_____ **23.** a lipid hormone composed of four rings of carbon atoms
_____ **24.** a repeating unit in a nucleic acid composed of a sugar, a base, and phosphoric acid
_____ **25.** the process by which a DNA molecule synthesizes an identical molecule of DNA
_____ **26.** the process by which a DNA molecule synthesizes a complementary single strand of RNA
_____ **27.** three consecutive nucleotides in RNA that designates a specific amino acid to be added to a growing protein chain
_____ **28.** the process by which an RNA codon specifies which amino acid is added to a growing protein chain

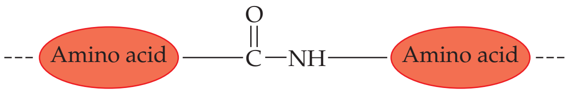
(a) amino acid (*Sec. 20.2*)
(b) biochemistry (*Sec. 20.1*)
(c) carbohydrate (*Sec. 20.1*)
(d) codon (*Sec. 20.6*)
(e) dipeptide (*Sec. 20.2*)
(f) disaccharide (*Sec. 20.4*)
(g) enzyme (*Sec. 20.3*)
(h) fat (*Sec. 20.5*)
(i) fatty acid (*Sec. 20.5*)
(j) glycoside linkage (*Sec. 20.4*)
(k) lipid (*Sec. 20.1*)
(l) lock-and-key model (*Sec. 20.3*)
(m) monosaccharide (*Sec. 20.4*)
(n) nucleic acid (*Sec. 20.1*)
(o) nucleotide (*Sec. 20.6*)
(p) oil (*Sec. 20.5*)
(q) peptide linkage (*Sec. 20.2*)
(r) phospholipid (*Sec. 20.5*)
(s) polypeptide (*Sec. 20.2*)
(t) polysaccharide (*Sec. 20.4*)
(u) protein (*Sec. 20.1*)
(v) replication (*Sec. 20.6*)
(w) saponification (*Sec. 20.5*)
(x) steroid (*Sec. 20.5*)
(y) transcription (*Sec. 20.6*)
(z) translation (*Sec. 20.6*)
(aa) triglyceride (*Sec. 20.5*)
(bb) wax (*Sec. 20.5*)

Exercises *Answers to odd-numbered Exercises are in Appendix I.*

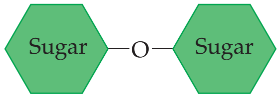
Biological Compounds (Sec. 20.1)

1. Identify the type of biological compound having each of the following structures.

(a)

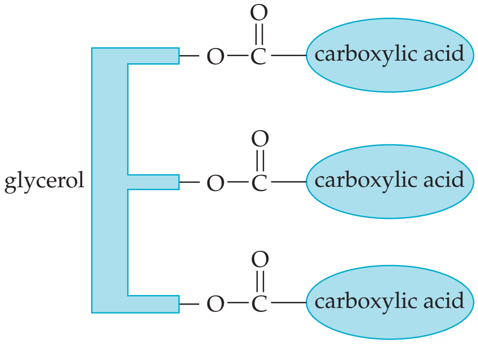


(b)

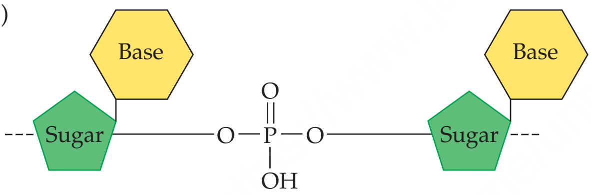


2. Identify the type of biological compound having each of the following structures.

(a)



(b)

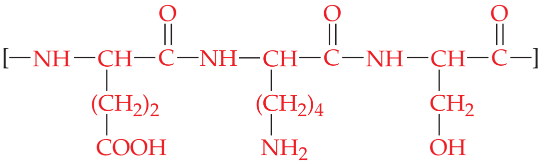


3. Identify the type of linkage that joins repeating units in each of the following.
(a) protein
(b) two sugars in a carbohydrate

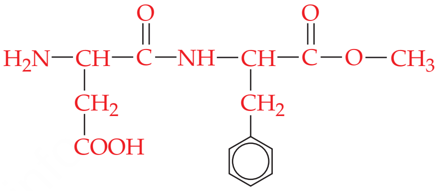
4. Identify the type of linkage that joins each of the following.
(a) glycerol and a carboxylic acid in a lipid
(b) nucleic acid

Proteins (Sec. 20.2)

5. What type of bonds are responsible for the primary structure of a protein?
6. What type of bonds are responsible for the secondary structure of a protein?
7. Explain the difference between the primary and secondary structure of a protein.
8. Explain the difference between the secondary and tertiary structure of a protein.
9. Draw the structure of the dipeptide formed by two molecules of the amino acid alanine. Circle the peptide linkage.
10. Draw the structure of the dipeptide formed by two molecules of the amino acid cysteine. Circle the peptide linkage.
11. Draw the structure of the dipeptide seryltyrosine (Ser—Tyr).
12. Draw the structure of the dipeptide tyrosylserine (Tyr—Ser).
13. Indicate all the possible sequences for a tripeptide containing alanine, glutamine, and methionine; use the standard abbreviation for each amino acid (Ala, Gln, Met).
14. Indicate all the possible sequences for a tripeptide containing glycine, histidine, and tryptophan; use the standard abbreviation for each amino acid (Gly, His, Trp).
15. A segment of a protein chain contains the following tripeptide. Identify the three amino acids given the structure for the tripeptide.


$$[-NH-\underset{\underset{COOH}{|}{(CH_2)_2}}{\underset{|}{CH}}-\overset{\overset{O}{\|}}{C}-NH-\underset{\underset{NH_2}{|}{(CH_2)_4}}{\underset{|}{CH}}-\overset{\overset{O}{\|}}{C}-NH-\underset{\underset{OH}{|}{CH_2}}{\underset{|}{CH}}-\overset{\overset{O}{\|}}{C}-]$$

16. The artificial sweetener NutraSweet is the methyl ester of a dipeptide. Identify the two amino acids given the structure for NutraSweet.



17. The hormones vasopressin and oxytocin are polypeptides with nine amino acids. Given the primary sequence for each hormone, identify the two amino acids that differ in each polypeptide.

vasopressin: Cys—Tyr—Phe—Gln—Asn—Cys—Pro—Arg—Gly
oxytocin: Cys—Tyr—Ile—Gln—Asn—Cys—Pro—Leu—Gly

18. Vasopressin raises blood pressure in humans, and oxytocin induces lactation in mammals. Suggest a reason why the biological activity of the polypeptides differs.

Enzymes (Sec. 20.3)

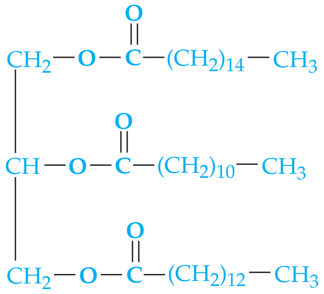
19. In the lock-and-key model of enzyme catalysis, what does the key represent?
20. In the lock-and-key model of enzyme catalysis, what does the lock represent?
21. What do the “teeth” on the key represent in the lock-and-key model of enzyme catalysis?
22. What is the term for a location on an enzyme that conforms to the shape of the substrate?
23. What are the experimental conditions that make an enzyme inactive?
24. What is the term for a molecule that blocks an enzyme and makes it inactive?

Carbohydrates (Sec. 20.4)

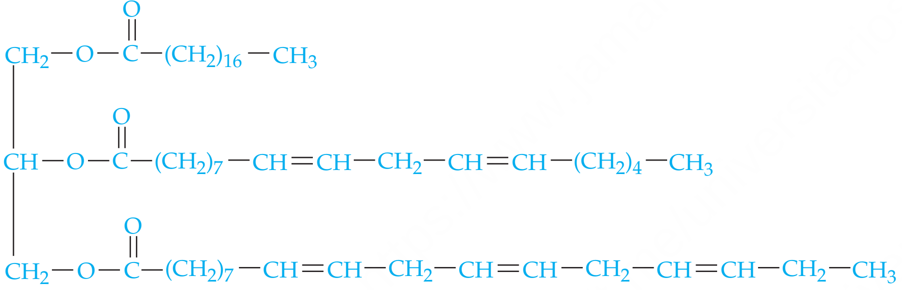
25. How many carbon atoms are in a molecule of a pentose sugar?
26. How many carbon atoms are in a molecule of a hexose sugar?
27. What are the two types of functional groups in an aldose sugar?
28. What are the two types of functional groups in a ketose sugar?
29. What is the distinction between a monosaccharide and a disaccharide?
30. What is the distinction between a disaccharide and a polysaccharide?
31. Which two monosaccharides are joined by a glycoside linkage in sucrose?
32. Which two monosaccharides are joined by a glycoside linkage in lactose?

Lipids (Sec. 20.5)

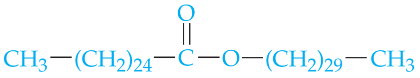
33. State whether each of the following characteristics indicates a fat, an oil, or both.
    (a) animal source
    (b) plant source
    (c) semisolid
    (d) insoluble in water
34. State whether each of the following constituents typifies a fat, an oil, or both.
    (a) saturated fatty acids
    (b) unsaturated fatty acids
    (c) glycerol
    (d) ester linkages
35. Identify the fatty acids in the following triglyceride fat.



36. Identify the fatty acids in the following triglyceride oil.



37. How many fatty acids are in a phospholipid?
38. What phospholipid is used to emulsify chocolate?
39. Carnauba wax is a lipid obtained from the leaves of palm trees and used in polishes for metal and wood surfaces. If the given structure is treated with aqueous NaOH, what are the formulas of the two products?



$CH_3-(CH_2)_{24}-\overset{\overset{\displaystyle O}{\|}}{C}-O-(CH_2)_{29}-CH_3$

Carnauba wax

40. Lanolin is a lipid wax obtained from sheep's wool and is used in ointments and cosmetic lotions. If lanolin is treated with aqueous NaOH, what two types of organic structures are obtained?
41. Draw the unique ring structure that is typical of steroid compounds.
42. What functional group differs in testosterone and progesterone?

Nucleic Acids (Sec. 20.6)

43. What are the three general components of a DNA nucleotide?
44. What are the three general components of an RNA nucleotide?
45. What are the four bases in DNA?
46. What are the four bases in RNA?
47. How many strands of nucleotides are in a DNA molecule?
48. How many strands of nucleotides are in an RNA molecule?
49. How many hydrogen bonds are there between adenine and thymine nucleotides in the DNA double helix?
50. How many hydrogen bonds are there between cytosine and guanine nucleotides in the DNA double helix?
51. During DNA replication, an adenine base (A) in the template strand codes for which base in the complementary strand?
52. During DNA replication, a cytosine base (C) in the template strand codes for which base in the complementary strand?
53. During RNA transcription, an adenine base (A) in the template strand codes for which base on the growing strand?
54. During RNA transcription, a cytosine base (C) in the template strand codes for which base on the growing strand?

General Exercises

55. What is the overall shape of a protein that provides strength, such as a protein in muscle tissue?
56. What is the overall shape of a protein that serves a metabolic role, such as a protein for oxygen transport?
57. What happens to a protein when it undergoes denaturation in an aqueous acid?
58. What happens to a polysaccharide when it undergoes hydrolysis in an aqueous acid?
59. Which vitamins belong to the lipid class of biological compounds?
60. Which vitamins are antioxidants and help eliminate free radicals?
61. How many codons are possible for an RNA trinucleotide having an adenine base, a cytosine base, and guanine base?
62. How many codons are possible for an RNA trinucleotide having an adenine base, and two cytosine bases?

Challenge Exercises

63. Explain the difference between the terms *replication* and *transcription*.
64. Explain the difference between the terms *transcription* and *translation*.
65. Explain the process of replication and how a DNA molecule synthesizes two identical molecules of DNA.
66. Explain the process of transcription and how a DNA molecule synthesizes a molecule of RNA.

Chapter 20 Self-Test

Answers to Self-Test are in Appendix J.

1. Which of the following types of linkage is found in a protein? (Sec. 20.1)
(a) peptide linkage **(b)** glycoside linkage
(c) ester linkage **(d)** phosphate linkage
(e) none of the above

2. What type of bonds shape the secondary structure of a protein? (Sec. 20.2)
(a) amide bonds **(b)** ester bonds
(c) hydrogen bonds **(d)** ionic bonds
(e) none of the above

3. How many possible tripeptides contain glycine (Gly), valine (Val), and proline (Pro) amino acids? (Sec. 20.3)
(a) 1 **(b)** 3
(c) 6 **(d)** 9
(e) none of the above

4. An enzyme is an example of which type of biological compound? (Sec. 20.3)
(a) protein **(b)** carbohydrate
(c) lipid **(d)** nucleic acid
(e) none of the above

5. Which of the following functional groups are found in an aldose sugar? (Sec. 20.4)
(a) alcohol and aldehyde **(b)** alcohol and ketone
(c) aldehyde and ketone **(d)** aldehyde and phenol
(e) none of the above

6. Which monosaccharides are produced from the hydrolysis of lactose? (Sec. 20.4)
(a) glucose and fructose **(b)** glucose and galactose
(c) glucose and ribose **(d)** glucose and threose
(e) none of the above

7. Which of the following is characteristic of a lipid oil? (Sec. 20.5)
(a) animal source **(b)** saturated fatty acids
(c) semisolid **(c)** water insoluble
(e) all of the above

8. Which of the following is found in a lipid wax? (Sec. 20.5)
(a) ester linkage **(b)** glycerol
(c) long-chain alcohol **(d)** saturated fatty acid
(e) unsaturated fatty acid

9. Which of the following is found in a DNA nucleotide? (Sec. 20.6)
(a) deoxyribose sugar **(b)** nitrogen base
(c) phosphoric acid **(d)** all of the above
(e) none of the above

10. Which of the following bases is *not* found in a DNA nucleotide? (Sec. 20.6)
(a) adenine **(b)** cytosine
(c) guanine **(d)** thymine
(e) uracil

Key Concepts

11. Draw the structure for the dipeptide glycylvaline (Gly–Val).

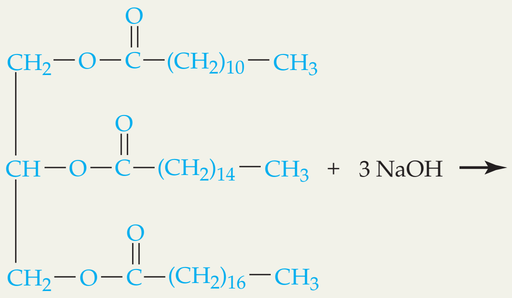
12. Complete the following reaction for the acid hydrolysis of the disaccharide sucrose and label the two products.

$$\underset{\text{sucrose}}{C_{12}H_{22}O_{11}} + \underset{\text{water}}{H_2O} \xrightarrow{H^+}$$

13. What type of biological compound is a water-insoluble fat, oil, or wax?

14. What type of biological compound is a polymer of a sugar attached to a base, joined together by phosphoric acid?

15. Complete the following reaction for the saponification of an animal fat. Label the products and identify the fatty acids in the animal fat.



16. One of the components of beeswax is an ester composed of the following fatty acid and alcohol. Draw the structure of the beeswax ester.

Fatty acid: $CH_3-(CH_2)_{34}-COOH$
Alcohol: $CH_3-(CH_2)_{35}-OH$

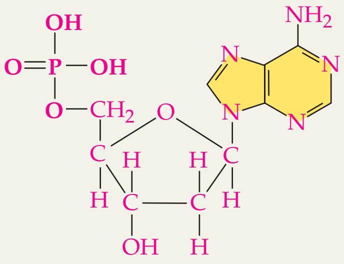
17. During DNA replication, an adenine base on the template strand will code for which base on the complementary strand?

18. During DNA transcription, an adenine base on the template strand will code for which base on the growing RNA strand?

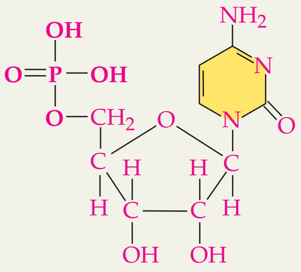
Critical Thinking

19. Identify the sugar and base in each of the following.

(a) nucleotide A



(b) nucleotide B



APPENDIX

A

Weights and Measures

METRIC SYSTEM EXACT EQUIVALENTS

Length:	1 meter (m)	=	100 centimeters (cm)
	1 meter (m)	=	1000 millimeters (mm)
	1 kilometer (km)	=	1000 meters (m)
Mass:	1 gram (g)	=	1000 milligrams (mg)
	1 kilogram (kg)	=	1000 grams (g)
	1 metric ton	=	1000 kilograms (kg)
Volume:	1 liter (L)	=	1000 milliliters (mL)
	1 liter (L)	=	10 deciliters (dL)
	1 milliliter (mL)	=	1 cubic centimeter (cm^3)

ENGLISH SYSTEM EXACT EQUIVALENTS

Length:	1 foot (ft)	=	12 inches (in.)
	1 mile (mi)	=	1760 yards (yd)
	1 mile (mi)	=	5280 feet (ft)
Mass:	1 pound (lb)	=	16 ounces (oz)
	1 troy pound (t lb)	=	12 troy ounces (t oz)
	1 ton	=	2000 pounds (lb)
Volume:	1 quart (qt)	=	32 fluid ounces (fl oz)
	1 quart (qt)	=	2 pints (pt)
	1 gallon (gal)	=	4 quarts (qt)

ENGLISH–METRIC APPROXIMATE EQUIVALENTS

Length:	1 inch (in.)	=	2.54 centimeters (cm)
Mass:	1 pound (lb)	=	454 grams (g)
Volume:	1 quart (qt)	=	946 milliliters (mL)
Time:	1 second (sec)	=	1.00 second (s)

TEMPERATURE EQUIVALENTS

$$\frac{9}{5}{}^{\circ}\mathrm{C} + 32 = {}^{\circ}\mathrm{F}$$

$$\frac{5}{9}({}^{\circ}\mathrm{F} - 32) = {}^{\circ}\mathrm{C}$$

$${}^{\circ}\mathrm{C} + 273 = \mathrm{K}$$

ENERGY EQUIVALENTS

1 calorie (cal)	=	4.184 joules (J)
1 kilocalorie (kcal)	=	1000 calories (cal)
1 kilocalorie (kcal)	=	4.184 kilojoules (kJ)

APPENDIX B

Physical Constants

Avogadro's number = 6.02×10^{23}

Absolute zero
Kelvin scale = 0 K
Celsius scale = −273.15 °C

Standard temperature and pressure (STP)
standard temperature = 0° C (273 K)
standard atmospheric pressure = 1 atm
= 760 mm Hg (760 torr)
= 76 cm Hg
= 29.9 in. Hg
= 14.7 psi
= 101 kPa

Molar volume of a gas (STP) = 22.4 L/mol

Ideal gas constant, R = 0.0821 L · atm/mol · K

Ionization constant of water, K_w = 1.00×10^{-14}

Mass of proton = 1.0073 amu

Mass of neutron = 1.0087 amu

Mass of electron = 0.00055 amu

Velocity of light = 3.00×10^{8} m/s

APPENDIX C

Activity Series for Metals

Most Active Metal: $\mathrm{Li} \rightarrow \mathrm{Li}^{+} + \mathrm{e}^{-}$
$\mathrm{K} \rightarrow \mathrm{K}^{+} + \mathrm{e}^{-}$
$\mathrm{Ba} \rightarrow \mathrm{Ba}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Sr} \rightarrow \mathrm{Sr}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Ca} \rightarrow \mathrm{Ca}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Na} \rightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$
(*metals above react with water at* 25 °C)

$\mathrm{Mg} \rightarrow \mathrm{Mg}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Al} \rightarrow \mathrm{Al}^{3+} + 3\,\mathrm{e}^{-}$
$\mathrm{Mn} \rightarrow \mathrm{Mn}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Zn} \rightarrow \mathrm{Zn}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Cr} \rightarrow \mathrm{Cr}^{3+} + 3\,\mathrm{e}^{-}$
$\mathrm{Fe} \rightarrow \mathrm{Fe}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Cd} \rightarrow \mathrm{Cd}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Co} \rightarrow \mathrm{Co}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Ni} \rightarrow \mathrm{Ni}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Sn} \rightarrow \mathrm{Sn}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Pb} \rightarrow \mathrm{Pb}^{2+} + 2\,\mathrm{e}^{-}$
(*metals above react with acid*)

$\mathbf{H_2 \rightarrow 2H^{+} + 2e^{-}}$

(*metals below do not react with acid*)
$\mathrm{Cu} \rightarrow \mathrm{Cu}^{2+} + 2\,\mathrm{e}^{-}$
$\mathrm{Ag} \rightarrow \mathrm{Ag}^{+} + \mathrm{e}^{-}$
$\mathrm{Hg} \rightarrow \mathrm{Hg}^{2+} + 2\,\mathrm{e}^{-}$
Least Active Metal: $\mathrm{Au} \rightarrow \mathrm{Au}^{3+} + 3\,\mathrm{e}^{-}$

APPENDIX

D

Solubility Rules for Ionic Compounds

Ionic compounds containing the following ions are generally ***soluble*** in water.

1. alkali metal ions and the ammonium ion, Li^+, Na^+, K^+, NH_4^+
2. acetate ion, $C_2H_3O_2^-$
3. nitrate ion, NO_3^-
4. halide ions ($X = Cl^-, Br^-, I^-$)
(AgX, Hg_2X_2, and PbX_2 are exceptions and are *insoluble*)
5. sulfate ion, SO_4^{2-}
($SrSO_4$, $BaSO_4$, and $PbSO_4$ are exceptions and are *insoluble*)

Ionic compounds containing the following ions are generally ***insoluble**** in water.

6. carbonate ion, CO_3^{2-}
(see Rule 1 exceptions, which are *soluble*)
7. chromate ion, CrO_4^{2-}
(see Rule 1 exceptions, which are *soluble*)
8. phosphate ion, Po_4^{3-}
(see Rule 1 exceptions, which are *soluble*)
9. sulfide ion, S^{2-}
(CaS, SrS, BaS, and Rule 1 exceptions are *soluble*)
10. hydroxide ion, OH^-
[$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, and Rule 1 exceptions are *soluble*]

---

*These compounds are actually slightly soluble, or very slightly soluble, in water.

APPENDIX

E

Vapor Pressure of Water

Temperature (°C)	Vapor Pressure (mm Hg)	Temperature (°C)	Vapor Pressure (mm Hg)	Temperature (°C)	Vapor Pressure (mm Hg)
0	4.6	21	18.7	35	41.2
5	6.5	22	19.8	40	55.3
10	9.2	23	21.1	45	71.9
12	10.5	24	22.4	50	92.5
14	12.0	25	23.8	55	118.0
16	13.6	26	25.2	60	149.4
17	14.5	27	26.7	70	233.7
18	15.5	28	28.4	80	355.1
19	16.5	29	30.0	90	525.8
20	17.5	30	31.8	100	760.0

APPENDIX

F

Properties of Water

Density of H_2O: 0.99987 g/mL at 0 °C
1.00000 g/mL at 4 °C
0.99707 g/mL at 25 °C

Heat of fusion at 0 °C: 80.0 cal/g (335 J/g)
Heat of solidification at 0 °C: 80.0 cal/g (335 J/g)

Heat of vaporization at 100 °C: 540 cal/g (2260 J/g)
Heat of condensation at 100 °C: 540 cal/g (2260 J/g)

Specific heat of H_2O
ice: 0.50 cal/(g × °C) [2.1 J/(g × °C)]
water: 1.00 cal/(g × °C) [4.18 J/(g × °C)]
steam: 0.48 cal/(g × °C) [2.0 J/(g × °C)]

Ionization constant of H_2O
K_w : 1.00×10^{-14} at 25 °C

APPENDIX

G Answers to Concept Exercises

Chapter 1

1.1 Alchemists believed in a magic potion that had the power to heal and to transmute lead into gold.

1.2 By studying the "vase/profiles" example in Section 1.3, and the "stacked blocks" example in the Exercises, a student begins to expand their perceptions.

Prerequisite Science Skills

PSS.1 No instrument is capable of an exact measurement.

PSS.2 No measurement is exact due to uncertainty.

PSS.3 No measurement has infinite significant digits because no measurement is exact.

PSS.4 The rules for significant digits do not apply to exact numbers; the rules apply only to measurements.

PSS.5 When *adding* or *subtracting* measurements, the answer is limited by the decimal place with the most uncertainty. That is, the measurement that is least certain in a set of data limits the answer.

PSS.6 When *multiplying* or *dividing* measurements, the answer is limited by the number of significant digits in the data. That is, the measurement with the least number of significant digits limits the number of digits in the answer.

PSS.7 The length 1×10^{-3} cm is less; note the negative power of 10.

PSS.8 The mass 0.000 001 g is less; note the extra placeholder zero.

PSS.9 The mass 0.000 01 g is greater; that is, 1×10^{-5} g is greater than 1×10^{-6} g.

PSS.10 A length of 1.55×10^{0} cm equals 1.55 cm expressed as an ordinary number.

Chapter 2

2.1 The basic units of the metric system are meter, gram, and liter.

2.2 Because 1 m equals 1000 mm exactly, there is an infinite number of significant digits in the unit equation.

2.3 Because 1 L = 1000 mL exactly, there is an infinite number of significant digits in the unit factor.

2.4 The volume of a cube 1 cm on a side (1 cm^3) equals 1 mL.

2.5 There is an infinite number of significant digits in the unit factor 1 g/1000 mg because 1 g = 1000 mg is an exact relationship.

2.6 There are an infinite number of significant digits in the unit factor 1 gal/4 qt because 1 gal = 4 qt is an exact relationship.

2.7 There are three significant digits in the unit factor 1 kg/2.20 lb because 1 kg = 2.20 lb is not an exact relationship. The kilogram and pound come from different systems of measurement.

2.8 The percent of gold in the alloy is 75%; that is, pure gold is 100% minus 20% silver and 5% copper.

2.9 The unit factors are: $\dfrac{11.2\text{ g hydrogen}}{100\text{ g water}}$ and $\dfrac{100\text{ g water}}{11.2\text{ g hydrogen}}$.

2.10 The volume of a cube 10 cm on a side (1000 cm^3) equals 1 L.

2.11 The thicknesses are all equal; that is, 1 mm = 0.1 cm = 0.001 m.

2.12 A volume of 500 in.^3 is greater than 500 cm^3. Since 1 in. is greater than 1 cm, 1 in^3 is greater than 1 cm^3.

2.13 The volumes are equal because 1 mL = 1 cm^3.

2.14 Because ice floats in water, water has the greater density.

2.15 Because 1 kg = 1000 g and 1 L = 1000 mL, the two densities are equal.

2.16 The temperatures are the same, but the term *degrees centigrade* is discouraged.

2.17 The Kelvin temperature scale is assigned a value of zero for the coldest possible temperature; thus, −100 K cannot exist.

2.18 If the can feels cold, heat is flowing away from your hand into the can.

Chapter 3

3.1 (a) solid, (b) liquid, (c) gas

3.2 (a) element, (b) element, (c) compound, (d) mixture

3.3 Calcium is a solid metal under normal conditions; phosphorus is a solid nonmetal; mercury is a liquid metal; and silicon is a solid semimetal.

3.4 Bromine is a liquid nonmetal under normal conditions.

3.5 The molecular formula for vitamin C is $C_6H_8O_6$.

3.6 Nitroglycerin has a total of 20 atoms in one molecule.

3.7 Copper turning black when heated is an example of a chemical change.

3.8 An Alka–Seltzer® tablet dissolving in water and producing gas bubbles is an example of a chemical change.

3.9 Applying the conservation of mass law, the mass of sulfur is found by difference to be 0.252 g (1.252 g − 1.000 g).

3.10 As the temperature increases, the kinetic energy of air molecules increases. As kinetic energy increases, the velocity of air molecules increases.

3.11 The energy required to heat 10.0 g of alcohol from 20 °C to 50 °C is identical to the heat released when 10.0 g of alcohol cools from 50 °C to 20 °C.

3.12 Chemical, electrical, heat, light, mechanical, and nuclear are all basic forms of energy; solar energy is an example of light energy.

Chapter 4

4.1 Atoms of different elements have a different number of protons, and *cannot have* the same atomic number.

4.2 Atoms of different elements *can have* isotopes with the same mass number.

4.3 (a) Simple Average Mass: $\dfrac{2.00\text{ g} + 1.00\text{ g}}{2} = 1.50\text{ g}$

(b) Weighted Average Mass:

$$2.00\text{ g}\left(\frac{75\text{ large marbles}}{100\text{ total marbles}}\right) + 1.00\text{ g}\left(\frac{25\text{ small marbles}}{100\text{ total marbles}}\right)$$
$$= 1.50\text{ g} + 0.25\text{ g} = 1.75\text{ g}$$

4.4 Because the periodic table lists the atomic mass of sodium as 22.99, and Na-23 is the only stable isotope, the mass of a Na-23 isotope is 22.99 amu.

4.5 Atomic number and mass number are always whole numbers; the atomic mass is never a whole number because it is obtained from a measurement of mass.

4.6 The energy *decreases* as the wavelength increases; but the energy *increases* as frequency increases.

4.7 A water wave is to a drop of water, as a light wave is to a *photon*.

4.8 All three statements are true.

4.9 The theoretical number of sublevels corresponds to the energy level. Thus, the 10th energy level has 10 sublevels.

4.10 The sublevels are closer together and the energy difference between sublevels *decreases* moving away from the nucleus. (Refer to Figure 4.13.)

4.11 The atomic number of Mn is one more than Cr, so Mn has one more *d* electron in the outermost sublevel.

4.12 All three statements are true.

Chapter 5

5.1 The modern periodic law states that elements are arranged according to *increasing atomic number*.

5.2 The group number for each family of elements is as follows:

(a) IA/1
(b) IIA/2
(c) VIIA/17
(d) VIIIA/18

5.3 Period 1 at the top of the periodic table represents the smallest atomic radius. Group IA/1 at the far left of the periodic table represents the most metallic character.

5.4 Given the density of Ag (10.5 g/cm^3) and Au (19.3 g/cm^3), the difference is 8.8 g/cm^3. The estimated density for Rg is 28.1 g/cm^3.

5.5 Because Zn and Cd are in the same group, they have very similar chemical properties.

5.6 Because Ag is in Period 5 and Group 1B/11, the highest energy sublevel is 4*d*.

5.7 Because Pt is the 8th element in the 5*d* sublevel, we predict the number of 5*d* electrons is 8.

5.8 Representative elements have valence electrons only in the outermost *s* and *p* sublevels. Transition elements have electrons in the outermost *d* and *f* sublevels, which complicates making reliable predictions.

5.9 Because Mg has two valence electrons in a 3*s* sublevel, we can draw the two electrons next to each other, or separate.

Mg: which may be preferable to ·Mg·

Either formula is acceptable. However, valence electrons are not stationary and the electron dot formula is only a tool for keeping track of valence electrons.

5.10 The noble gases have the highest ionization energies since they have filled valence shells. The alkali metals have the lowest ionization energies since they assume a noble gas electron configuration by losing one electron.

5.11 Group IA/1 elements have an ionic charge of 1+, Group IIA/2 elements have a charge of 2+, and Group IIIA/13 elements have a charge of 3+.

5.12 Group VIIA/17 elements have an ionic charge of 1−, Group VIA/16 elements have a charge of 2−, and Group VA/15 elements have a charge of 3−.

5.13 The next ion in the isoelectronic series Se^{2-}, Br^-, Kr, Rb^+, Sr^{2+} is Y^{3+}.

5.14 The electron configuration is simply [Rn].

Chapter 6

6.1 A binary acid is in an aqueous solution and is indicated by (*aq*), for example, HCl(*aq*).

6.2 Hg^{2+} is a monoatomic cation; whereas, Hg_2^{2+} is a polyatomic cation.

6.3 The Stock system specifies the charge on a cation using a Roman numeral in parentheses.

6.4 Iodide is a monoatomic anion because it has an −*ide* suffix.

6.5 A monoatomic anion has an −*ide* suffix; a polyatomic anion has an −*ate* or −*ite* suffix.

6.6 The formula for the metal oxide is M_2O_3.

6.7 The formula for the metal carbonate is $M_2(CO_3)_3$.

6.8 The charge on an oxide ion is 2−, and there are two oxide ions. The total negative charge is 4−. Thus, the positive charge on metal M is 4+.

6.9 Sodium chloride is a binary ionic compound as indicated by the −*ide* suffix.

6.10 LiCl is a binary ionic compound because it has two elements.

6.11 Because potassium and sodium are in the same family, and fluorine and iodine are in the same family, the formula for potassium iodide is KI.

6.12 The charge on a carbonate ion is 2−. Thus, the positive charge on metal M is 2+.

6.13 Potassium nitride is a binary ionic compound. Potassium nitrate and potassium nitrite are ternary ionic compounds as indicated by the −*ate* and −*ite* suffixes.

6.14 LiCl is a binary ionic compound. LiClO and $LiClO_2$ are ternary ionic compounds because they have three elements.

6.15 Since calcium and strontium are in the same family, the formula for strontium sulfate is $SrSO_4$.

6.16 P_2O_3 is a binary molecular compound and requires Greek prefixes to specify the nonmetal ratio; that is, diphosphorus trioxide. Fe_2O_3 is an ionic compound and does not use Greek prefixes; that is, iron(III) oxide.

6.17 Nitrogen oxide is a binary molecular compound because it contains a nonmetal and oxygen.

6.18 $HBr(aq)$ is a binary acid named hydrobromic acid.

6.19 $HBrO_3(aq)$ is a ternary oxyacid named bromic acid, similar to chloric acid, $HClO_3(aq)$. $HBrO_2(aq)$ is a ternary oxyacid named bromous acid, similar to chorus acid, $HClO_2(aq)$.

Chapter 7

7.1 If the beaker becomes warm, the reaction is exothermic. If the beaker becomes cool, the reaction is endothermic.

7.2 The nonmetals that occur naturally as diatomic molecules are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.

7.3 No, changing a subscript gives an incorrect chemical formula.

7.4 Writing water as HOH often makes a neutralization equation easier to balance, as it is a special type of a double replacement reaction.

7.5 Decomposition reactions are easily recognized as they only have one reactant, unlike other types of reactions.

7.6 Heating a metal and oxygen produces a metal oxide.

7.7 Heating a nonmetal and oxygen produces a nonmetal oxide.

7.8 Heating a metal and nonmetal produces an ionic compound.

7.9 Heating a metal hydrogen carbonate produces a metal carbonate, water, and carbon dioxide gas.

7.10 Heating a metal carbonate produces a metal oxide and carbon dioxide gas.

7.11 Heating an oxygen-containing compound often releases oxygen gas.

7.12 Zn and Cu metals react with aqueous $AgNO_3$ (above Ag in the activity series), but not Au.

7.13 The metals that react with aqueous acid (above H in the activity series) include Zn, but not Cu and Au.

7.14 $Pb(NO_3)_2$ is soluble in water; $PbCl_2$ and $PbCO_3$ are insoluble.

7.15 $AgCl(s)$ is the insoluble product.

7.16 $NaCl(aq)$ is the salt product.

Chapter 8

8.1 (a) A carbon atom has a mass of 12.01 amu. (b) Avogadro's number of carbon atoms has a mass of 12.01 g.

8.2 (a) One platinum atom has a mass of 195.08 amu. (b) Avogadro's number of platinum atoms has a mass of 195.08 g.

8.3 The number of atoms in 1.00 mol of a metallic element is 6.02×10^{23}.

8.4 The number of molecules in 1.00 mol of a diatomic nonmetal is 6.02×10^{23}.

8.5 The mass of Avogadro's number of water molecules is 18.02 g.

8.6 The mass of Avogadro's number of sulfur molecules is 256.56 g (8×32.07 g).

8.7 The mass of Avogadro's number of ozone molecules is 48.00 g (3×16.00 g).

8.8 The mass of a carbon monoxide molecule is 28.01 amu.

8.9 The He balloon rises because the density of helium gas is much less than air. The N_2 balloon does not rise because the density of nitrogen gas is about the same as air, which is composed of 78% N_2 and 21% O_2.

8.10 The volume of 1.00 mol of any gas at STP is 22.4 L.

8.11 The number of molecules in 1.00 mol of ozone gas is 6.02×10^{23}.

8.12 The mass of 1.00 mol of hydrogen gas is 2.02 g.

8.13 The percent oxygen is 53.3% (100.0% − 40.0% − 6.7%).

8.14 The empirical formula is SO_2.

8.15 The empirical formula is SO_3.

8.16 The empirical formula is HO.

Chapter 9

9.1 The coefficients of a balanced equation indicate the ratio of moles and volumes of gases, but *not* the masses of gas, which vary with the molar mass of each gas.

9.2 One unit factor is required for a mole-mole conversion problem.

9.3 The three types of stoichiometry problem are mass–mass, mass–volume, and volume–volume.

9.4 The three steps in the unit analysis method of problem solving are as follows: (1) Write down the unit asked for in the answer. (2) Write down the given value that is related to the answer. (3) Apply unit factor(s) to convert the given value to the units in the answer.

9.5 Three unit factors are usually required to convert the mass given to the mass in the answer.

9.6 The three steps in unit analysis are as follows: (1) Write down the unit in the answer. (2) Write down the given value. (3) Apply unit factor(s) to convert the given value to the answer.

9.7 Three unit factors are usually required to convert the mass given to the volume in the answer.

9.8 One unit factor is usually required to convert the volume given to the volume in the answer.

9.9 Because equal volumes of gas have the same number of molecules, the coefficients in the balanced chemical equation indicate the volume ratio of gases, as well as the mole ratio. Thus, the mole ratio can be used to solve volume–volume stoichiometry problems in one step.

9.10 Each bicycle frame requires two tires for a total of six tires. Thus, only three bicycles can be assembled.

9.11 Zn metal is the limiting reactant.

9.12 S powder is the limiting reactant.

9.13 A percent yield greater than 100% is possible as some experimental errors lead to high results, and others to low results.

Chapter 10

10.1 An atmosphere is the largest unit, and the greatest gas pressure is 1 atm.

10.2 Increasing the volume produces a *decrease* in gas pressure because molecules are further apart and collide less frequently.

10.3 The gas pressure *increases* because the volume decreases, and there are more molecular collisions.

10.4 The volume *decreases* when the temperature cools; the air molecules move slower, and the balloon shrinks to maintain constant pressure.

10.5 The pressure *decreases* when the temperature cools; the air molecules move slower, and the collisions are fewer and less energetic.

10.6 The pressure doubles when the temperature doubles, but the pressure halves when the volume doubles. Thus, the volume does not change.

10.7 The partial pressure of NO gas is found by subtracting the partial pressure of N_2 and O_2 from 2.50 atm. Thus, the partial pressure of NO is 0.50 atm.

10.8 The pressure doubles when the volume halves, but the pressure halves when the temperature is halved. Thus, the pressure does not change.

10.9 No real gases are ideal; however, He behaves most like an ideal gas.

10.10 The proportionality constant (R) in the ideal gas law equals PV/nT.

Chapter 11

11.1 Because $CH_3{-}CH_2{-}OH$ can hydrogen bond, it has a stronger intermolecular attraction than $CH_3{-}O{-}CH_3$.

11.2 Because $CH_3{-}CH_2{-}OH$ can hydrogen bond, it has a higher boiling point than $CH_3{-}O{-}CH_3$.

11.3 A paper clip can float on water due to the surface tension between water molecules. If you disturb the paper clip, it will sink because it is more dense.

11.4 An *ionic crystalline solid* has a high melting point, and is a conductor of electricity only when melted. A *molecular crystalline solid* has a low melting point, and is a nonconductor of electricity.

11.5 Because the heat of vaporization is greater than the heat of fusion, it requires more heat to vaporize 1 g of water (540 cal) than to melt 1 g of water (80 cal).

11.6 The systematic names for the hydrates are calcium chloride *monohydrate*, calcium chloride *dihydrate*, and calcium chloride *hexahydrate*, respectively.

11.7 The hexahydrate, $CaCl_2 \cdot 6H_2O$, has the highest percentage of water.

11.8 The hexahydrate, $CaCl_2 \cdot 6H_2O$, has the most water of hydration (6).

Chapter 12

12.1 The attraction between a metal cation and a nonmetal anion gives an *ionic bond*; the attraction between two nonmetal atoms produces a *covalent bond.*

12.2 Xenon is isoelectronic with both Ba^{2+} and I^-.

12.3 All three statements are true.

12.4 All three statements are true.

12.5 There are 2 pairs of nonbonding electrons in $H_2\mathbf{S}$.

12.6 There are 6 pairs of nonbonding electrons in $\mathbf{S}O_2$.

12.7 There are 2 pairs of nonbonding electrons in $\mathbf{C}O$.

12.8 There are 10 pairs of nonbonding electrons in $\mathbf{S}O_3{}^{2-}$.

12.9 The general trends for increasing electronegativity in the periodic table are from left to right in a period, and from bottom to top in a group.

12.10 HF has the greatest difference in electronegativity ($4.0 - 2.1 = 1.9$); thus, the bond in $H{-}\mathbf{F}$ is the most polar.

12.11 The bond in $C{-}\mathbf{I}$ is the least polar because there is no difference in the electronegativity of carbon and iodine ($2.5 - 2.5 = 0$).

12.12 The structural formula for the molecule is $H{-}\mathbf{Cl}{-}O$. Since the oxygen atom can be removed with its six valence electrons, the $\mathbf{Cl}{-}O$ bond is coordinate covalent.

Chapter 13

13.1 Because atmospheric pressure is less at higher elevations, Lake Tahoe has less dissolved oxygen gas than Lake Havasu.

13.2 According to the *like dissolves like* rule, two liquids are miscible when both liquids are polar, or both liquids are nonpolar.

13.3 According to the *like dissolves like* rule, a solid is soluble in a liquid when the solid and liquid are both polar, or both nonpolar. An ionic compound is more soluble in a polar liquid than a nonpolar liquid.

13.4 Because the oxygen atom in a water molecule is partially negative, it is attracted to the positive cation in the solid solute.

13.5 Grinding the solute gives smaller particles with more surface area. Because solvent molecules attack the particle surface, the solute dissolves faster.

13.6 A hot cup of coffee can dissolve more sugar than coffee at a lower temperature.

13.7 By saturating a solution at a higher temperature, and allowing the solution to cool without disturbing it, the solute remains in the solution at a lower temperature and is said to be supersaturated.

13.8 Although dextrose and fructose have the same molecular formula, they have a different structural formula (see Figure 20.9).

13.9 There are six mass percent concentration unit factors; that is, three unit factors and their reciprocals.

13.10 There are four molar concentration unit factors; that is, two unit factors and their reciprocals.

13.11 Because equal volumes of 6 *M* sulfuric acid and distilled water are added together, the concentration is halved, and the diluted acid is 3 *M*.

13.12 Before solving a stoichiometry problem, first write a balanced chemical equation for the reaction.

Chapter 14

14.1 The solution with the lowest pH (1) has the highest acidity.

14.2 Aqueous HNO_3 is a strong Arrhenius acid (refer to Table 14.2).

14.3 Aqueous $HC_2H_3O_2$ and NaOH react to give sodium acetate, $NaC_2H_3O_2$.

14.4 Phenolphthalein is colorless in acidic and neutral solutions (refer to Table 14.4).

14.5 The titration of H_3PO_4 with NaOH is the titration of a weak acid with a strong base.

14.6 The titration of HCl with $Ba(OH)_2$ is the titration of a strong acid with a strong base.

14.7 Phenolphthalein is pink at the endpoint of the titration of oxalic acid.

14.8 Phenolphthalein is pink at the endpoint of the titration of benzoic acid.

14.9 Distilled water is neutral because the concentration of hydrogen ions and hydroxide ions is equal.

14.10 Apple juice at pH 4 is weakly acidic.

14.11 The value increases at higher temperatures; for example, at 50 °C, $K_w = 5.5 \times 10^{-14}$. The pH is slightly less than 7.0; for example, at 50 °C, pH = 6.6.

14.12 Milk of magnesia at pH 10 is weakly basic.

14.13 Grape juice at pH 4.05 is weakly acidic.

14.14 Wine at pH 3.25 is weakly acidic.

14.15 The reaction of HF*aq* and NH_4OH(*aq*) produces NH_4F and H_2O; NH_4F is a soluble salt, so the conductivity apparatus will glow brightly.

14.16 The reaction of HNO_3(*aq*) and $Ba(OH)_2$(*aq*) produces $Ba(NO_3)_2$ and H_2O; $Ba(NO_3)_2$ is a soluble salt, so the apparatus will glow brightly.

14.17 The reaction of $AgNO_3$(*aq*) and $AlCl_3$(*aq*) gives AgCl and $Al(NO_3)_3$. Although AgCl is insoluble, $Al(NO_3)_3$ is soluble, so the apparatus will glow brightly.

Chapter 15

15.1 A mole of gas is related to: (1) mass of gas; (2) number of gaseous particles; (3) volume of a gas at STP; and (4) molar concentration of a solution.

15.2 A balanced chemical equation is required to solve a stoichiometry problem.

15.3 According to the balanced equation, the mole ratio is 2 mol H_2S to 3 mol O_2. Since the starting volumes of gas are equal, the limiting reactant is O_2. That is, 2.00 L H_2S produces 2.00 L SO_2, but 2.00 L O_2 produces only 1.33 LSO_2.

15.4 Because ice floats in water, 1.00 cm^3 of water weighs more than a 1.00-cm cube of ice.

15.5 Similar to a plant greenhouse that retains heat energy from sunlight, carbon dioxide gas in the atmosphere traps heat energy from the Sun, and contributes to global warming (see the Chemistry Connection in Section 10.2).

15.6 Because $pV = nRT$, $n = pV/RT$. The number of moles of gas (n) is equal to its mass divided by its molar mass (MM); that is, $n = g/MM$. Substituting g/MM for n in the ideal gas law, and rearranging, we have $MM = gRT/PV$.

15.7 The volume of wet hydrogen gas is greater than the same number of hydrogen molecules in a dry gas assuming the temperature and pressure are equal.

15.8 The beaker of water containing AgCl contains the least number of chloride ions because silver chloride is insoluble.

15.9 Because the molar mass of $NaHCO_3$ is less than $KHCO_3$, a 1.00-g $NaHCO_3$ tablet has more moles of bicarbonate and neutralizes more stomach acid.

Chapter 16

16.1 (a) At higher temperatures molecules have a greater velocity, and the collision frequency increases. As the number of collisions increases, there are more effective collisions.

(b) At higher temperatures molecules have more kinetic energy, and collision energy increases. If the collisions are more energetic, more molecules have the necessary energy to react.

16.2 Notice that the progress of reaction (a) goes from lower to higher energy, which characterizes an endothermic reaction. The progress of reaction (b) is higher to lower, which characterizes an *exothermic reaction*.

16.3 *After* a reaction reaches chemical equilibrium, the concentrations of reactants and products are constant.

16.4 *Before* a reaction reaches chemical equilibrium, the concentration of reactants is decreasing, and the concentration of products is increasing.

16.5 The equilibrium constant expression is: $K_{eq} = \dfrac{[X_2Y_3]^2}{[X_2]^2[Y_2]^3}$

16.6 (a) Increasing temperature for an endothermic reaction shifts the equilibrium to the right, toward the products.

(b) Increasing pressure for this reaction has no effect on the equilibrium because there are the same number of reactant and product molecules.

16.7 The ionization constant expression is: $K_i = \dfrac{[H^+][H_2PO_4^-]}{[H_3PO_4]}$

16.8 (a) Adding NH_4Cl increases the NH_4^+ concentration; thus, shifting the equilibrium to the left toward the reactants.

(b) Adding NaOH increases the OH^- concentration; thus, shifting the equilibrium to the left toward the reactants.

16.9 The solubility product constant expression is: $K_{sp} = [Al^{3+}][OH^-]^3$.

16.10 (a) Adding $Ca(NO_3)_2$ increases the Ca^{2+} concentration; thus, shifting the equilibrium to the left toward the reactants.

(b) Adding Na_2CO_3 increases the CO_3^{2-} concentration; thus, shifting the equilibrium to the left toward the reactants.

Chapter 17

17.1 (a) 0 (b) +1; (c) −1; (d) +7

17.2 (a) −1; (b) +1; (c) +3; (d) +5

17.3 oxidizing agent: HX; reducing agent: M

17.4 oxidizing agent: Y_2; reducing agent: X^-

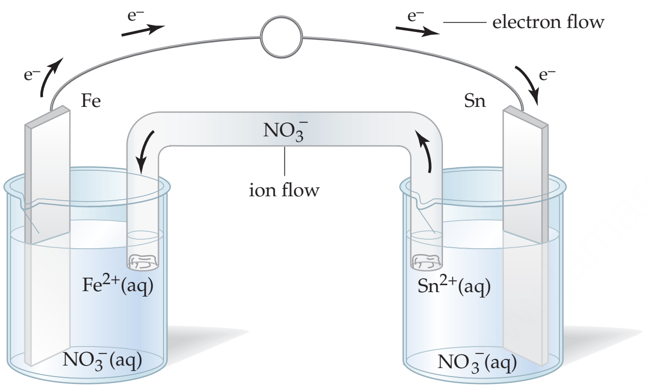
17.5 $2\,M_2O_3(s) + 6\,X_2(g) \rightarrow 4\,MX_3(aq) + 3\,O_2(g)$

17.6 $2\,MnO_4^-(aq) + 5\,SO_3^{2-}(aq) + 6\,H^+(aq) \rightarrow$
$2\,Mn^{2+}(aq) + 5\,SO_4^{2-}(aq) + 3\,H_2O(l)$

17.7 $2\,MnO_4^-(aq) + 3\,SO_3^{2-}(aq) + H_2O(l) \rightarrow$
$2\,MnO_2(s) + 3\,SO_4^{2-}(aq) + 2\,OH^-(aq)$

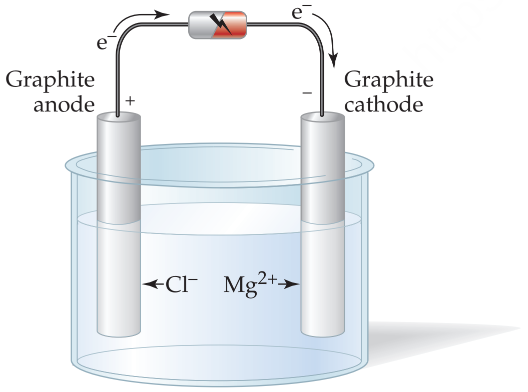
17.8 Figure 17.4 lists Cr and Al as stronger reducing agents than $Fe^{2+}(aq)$. Thus, Cr and Al react spontaneously with $FeSO_4(aq)$; Ag and Ni are weaker and do not react.

17.9 $Fe(s) + Sn^{2+}(aq) \rightarrow Sn(s) + Fe^{2+}(aq)$



Voltaic Cell

17.10 $C(s) + 2\,MgCl_2(l) \xrightarrow{\text{electricity}} 2\,Mg(l) + CCl_4(g)$



Electrolytic Cell

Chapter 18

18.1 (a) alpha ray, (b) beta ray, (c) gamma ray

18.2 31 mg

18.3 ${}^{208}_{82}Pb + {}^{62}_{28}Ni \rightarrow {}^{269}_{110}Ds + {}^{1}_{0}n$

Chapter 19

19.1 Note that decane ($C_{10}H_{22}$) has an *−ane* suffix, and fits the general molecular formula for an alkane (C_nH_{2n+2}).

19.2 Note that decene ($C_{10}H_{20}$) has an *−ene* suffix, and fits the general molecular formula for an alkene (C_nH_{2n}).

19.3 Note that decyne ($C_{10}H_{18}$) has a *−yne* suffix, and fits the general molecular formula for an alkyne (C_nH_{2n-2}).

19.4 Note that decene and naphthalene each have an *−ene* suffix; however, decene fits the general molecular formula for an alkene (C_nH_{2n}). Naphthalene has a different molecular formula (C_nH_{n-2}), so it is the arene by process of elimination.

19.5 Because an alkene and arene each end with an *−ene* suffix, it is not possible to distinguish an alkene from an arene based on the suffix.

19.6 Aldehydes, ketones, carboxylic acids, esters, and amides contain a carbonyl group. Organic halides, alcohols, phenols, ethers, and amines do not.

19.7 Because an alcohol and phenol each end with an *−ol* suffix, it is not possible to distinguish an alcohol from an phenol based on the suffix.

Chapter 20

20.1 A lipid is insoluble in water; whereas, a protein, carbohydrate, and nucleic acid are soluble.

20.2 A spiral helix describes the *secondary* structure of a protein. The sequence of amino acids in the protein corresponds to the *primary* structure, and a twisted helix describes the *tertiary* structure.

APPENDIX H

Answers to Key Term Exercises

Chapter 1

1. j, **2.** d, **3.** k, **4.** f, **5.** l, **6.** h, **7.** a, **8.** c, **9.** i, **10.** g, **11.** b, **12.** e

Prerequisite Science Skills

1. f, **2.** a, **3.** c, **4.** g, **5.** d, **6.** m, **7.** e, **8.** l, **9.** h, **10.** j, **11.** b, **12.** i, **13.** k

Chapter 2

1. e, **2.** o, **3.** n, **4.** h, **5.** m, **6.** r, **7.** j, **8.** w, **9.** f, **10.** x, **11.** q, **12.** v, **13.** p, **14.** c, **15.** y, **16.** d, **17.** s, **18.** i, **19.** u, **20.** g, **21.** b, **22.** l, **23.** t, **24.** a, **25.** k

Chapter 3

1. z, **2.** cc, **3.** i, **4.** l, **5.** m, **6.** dd, **7.** a, **8.** h, **9.** k, **10.** b, **11.** g, **12.** t, **13.** s, **14.** j, **15.** v, **16.** bb, **17.** c, **18.** w, **19.** r, **20.** u, **21.** e, **22.** y, **23.** f, **24.** d, **25.** x, **26.** aa, **27.** n, **28.** p, **29.** o, **30.** q

Chapter 4

1. h, **2.** t, **3.** q, **4.** d, **5.** e, **6.** p, **7.** c, **8.** n, **9.** b, **10.** a, **11.** y, **12.** m, **13.** o, **14.** x, **15.** v, **16.** g, **17.** s, **18.** f, **19.** k, **20.** j, **21.** l, **22.** i, **23.** w, **24.** u, **25.** r

Chapter 5

1. r, **2.** h, **3.** q, **4.** b, **5.** c, **6.** i, **7.** p, **8.** t, **9.** u, **10.** o, **11.** a, **12.** j, **13.** s, **14.** v, **15.** f, **16.** d, **17.** w, **18.** g, **19.** k, **20.** m, **21.** l, **22.** n, **23.** e

Chapter 6

1. i, **2.** h, **3.** d, **4.** p, **5.** e, **6.** b, **7.** c, **8.** q, **9.** f, **10.** a, **11.** l, **12.** n, **13.** o, **14.** j, **15.** m, **16.** g, **17.** k

Chapter 7

1. h, **2.** o, **3.** n, **4.** d, **5.** q, **6.** g, **7.** s, **8.** r, **9.** f, **10.** l, **11.** i, **12.** v, **13.** j, **14.** k, **15.** u, **16.** m, **17.** p, **18.** c, **19.** b, **20.** a, **21.** e, **22.** t

Chapter 8

1. a, **2.** g, **3.** e, **4.** f, **5.** j, **6.** b, **7.** d, **8.** i, **9.** c, **10.** h

Chapter 9

1. j, **2.** b, **3.** d, **4.** h, **5.** m, **6.** f, **7.** g, **8.** o, **9.** i, **10.** l, **11.** c, **12.** e, **13.** a, **14.** n, **15.** k

Chapter 10

1. k, **2.** v, **3.** b, **4.** d, **5.** u, **6.** c, **7.** p, **8.** i, **9.** e, **10.** f, **11.** l, **12.** g, **13.** t, **14.** w, **15.** h, **16.** r, **17.** x, **18.** q, **19.** m, **20.** s, **21.** j, **22.** a, **23.** o, **24.** n

Chapter 11

1. g, **2.** f, **3.** n, **4.** x, **5.** b, **6.** y, **7.** w, **8.** d, **9.** o, **10.** r, **11.** q, **12.** v, **13.** k, **14.** l, **15.** c, **16.** s, **17.** i, **18.** p, **19.** t, **20.** m, **21.** a, **22.** z, **23.** j, **24.** u, **25.** e, **26.** h

Chapter 12

1. e, **2.** bb, **3.** v, **4.** p, **5.** g, **6.** n, **7.** r, **8.** c, **9.** b, **10.** d, **11.** t, **12.** k, **13.** z, **14.** y, **15.** j, **16.** aa, **17.** s, **18.** x, **19.** m, **20.** h, **21.** w, **22.** u, **23.** i, **24.** f, **25.** o, **26.** cc, **27.** l, **28.** q, **29.** a

Chapter 13

1. c, **2.** n, **3.** p, **4.** o, **5.** b, **6.** i, **7.** k, **8.** j, **9.** e, **10.** g, **11.** d, **12.** q, **13.** a, **14.** s, **15.** m, **16.** l, **17.** t, **18.** r, **19.** f, **20.** h

Chapter 14

1. g, **2.** c, **3.** d, **4.** k, **5.** h, **6.** j, **7.** q, **8.** e, **9.** f, **10.** p, **11.** o, **12.** b, **13.** a, **14.** u, **15.** i, **16.** s, **17.** l, **18.** n, **19.** t, **20.** w, **21.** v, **22.** r, **23.** m

Chapter 15

1. e, **2.** l, **3.** a, **4.** b, **5.** r, **6.** p, **7.** q, **8.** c, **9.** g, **10.** t, **11.** s, **12.** k, **13.** v, **14.** m, **15.** f, **16.** j, **17.** d, **18.** u, **19.** i, **20.** h, **21.** n, **22.** o

Chapter 16

1. d, **2.** e, **3.** f, **4.** o, **5.** r, **6.** a, **7.** h, **8.** b, **9.** p, **10.** n, **11.** c, **12.** l, **13.** g, **14.** j, **15.** i, **16.** m, **17.** k, **18.** q

Chapter 17

1. l, **2.** n, **3.** k, **4.** p, **5.** m, **6.** o, **7.** j, **8.** q, **9.** f, **10.** e, **11.** i, **12.** r, **13.** s, **14.** g, **15.** h, **16.** a, **17.** c, **18.** b, **19.** d

Chapter 18

1. t, **2.** b, **3.** d, **4.** i, **5.** c, **6.** k, **7.** p, **8.** o, **9.** l, **10.** r, **11.** h, **12.** s, **13.** q, **14.** a, **15.** j, **16.** u, **17.** v, **18.** m, **19.** e, **20.** f, **21.** n, **22.** g, **23.** w

Chapter 19

1. s, **2.** m, **3.** v, **4.** x, **5.** f, **6.** a, **7.** b, **8.** d, **9.** u, **10.** r, **11.** e, **12.** w, **13.** q, **14.** c, **15.** g, **16.** t, **17.** k, **18.** o, **19.** j, **20.** n, **21.** l, **22.** h, **23.** p, **24.** i

Chapter 20

1. b, **2.** u, **3.** c, **4.** k, **5.** n, **6.** a, **7.** q, **8.** e, **9.** s, **10.** g, **11.** l, **12.** m, **13.** f, **14.** t, **15.** j, **16.** i, **17.** aa, **18.** h, **19.** p, **20.** w, **21.** r, **22.** bb, **23.** x, **24.** o, **25.** v, **26.** y, **27.** d, **28.** z

APPENDIX

I Answers to Odd-Numbered Exercises

Chapter 1

1. *yin* and *yang*
3. air, earth, fire, and water
5. Plan and *perform an experiment*, *make observations*, and *record data*.
7. *Conduct additional experiments* in order to test the hypothesis.
9. A *scientific theory* explains the behavior of nature using a model or concept, whereas a *natural law* states a relationship that is measurable under various experimental conditions.
11. (a) and (c)
13. *Chemistry* is referred to as the central science because it supports other physical sciences, life sciences, and earth sciences. An understanding of chemistry is also prerequisite to the health sciences.
15. Scientists and engineers must have formal training in chemistry, as well as doctors, nurses, dentists, physical therapists, chiropractors, dietitians, and veterinarians.
17. A solution to the nine–dot problem with only one straight line is to use a *very wide* line; the unconscious assumption regards the thickness of the line.
19. Although the perception is that line segment AB is shorter than BC, the two lines are of equal length. Measure each line segment to confirm this statement.

Prerequisite Science Skills

1. approximately 5 g
3. (a) length; (b) mass; (c) volume; (d) time
5. (a) 5.0 cm; (d) 0.5 cm
7. (d) 75.518 g
9. (b) 32.00 mL; (c) 32.05 mL
11. (a) 1; (b) 1; (c) 1; (d) 1
13. (a) 1; (b) 4; (c) 2; (d) 3
15. (a) 3; (b) 4; (c) 1; (d) 2
17. (a) 10.3; (b) 10.2; (c) 0.0103; (d) 10,200
19. (a) 9.12×10^5; (b) 9.46×10^{10}; (c) 9.00×10^{-7}; (d) 9.08×10^{-12}
21. (a) 55.0 cm; (b) 22.3 cm
23. (a) 11.6 cm; (b) 9.9 cm
25. (a) 0.6 cm^2; (b) 2.8 cm^2; (c) 12.0 cm^2; (d) 17 cm^2
27. (a) 8.8 g/mL; (b) 3.0 g/mL; (c) 4.26 g/mL; (d) 9.124 g/mL
29. (a) 10^3; (b) $(1/10)^3$ or 10^{-3}
31. (a) 2^3; (b) $(1/2)^3$ or 2^{-3}
33. (a) 1000; (b) 0.000 000 1
35. (a) 1×10^9; (b) 1×10^{-8}
37. (a) 10; (b) 0.1
39. (a) 8.0916×10^7; (b) 1.5×10^{-8}; (c) 3.356×10^{14}; (d) 9.27×10^{-13}
41. 2.69×10^{22} oxygen molecules
43. 5.32×10^{-23} g/oxygen molecule
45. 10.0 mL (± 0.1 mL)
47. 1.86×10^5 miles per second
49. 205 cm
51. (a) 3.52×10^6; (b) 1.91×10^{-6}
53. 3.3475×10^{-24} g
55. 4.200×10^3 lb

Chapter 2

1. (a) meter; (b) gram; (c) liter; (d) second
3. (a) Em; (b) Tg; (c) aL; (d) ps
5. (a) kilometer; (b) megagram; (c) nanoliter; (d) decisecond
7. (a) length; (b) mass; (c) volume; (d) time
9. (a) gigameter, Gm; (b) kilogram, kg; (c) milliliter, mL; (d) nanosecond, ns
11. (a) 1 km $= 1 \times 10^3$ m; (b) 1 Mg $= 1 \times 10^6$ g; (c) 1 L $= 1 \times 10^3$ mL; (d) 1 s $= 1 \times 10^6$ μs
13. (a) $\dfrac{1\ \text{km}}{1 \times 10^3\ \text{m}}$ and $\dfrac{1 \times 10^3\ \text{m}}{1\ \text{km}}$

    (b) $\dfrac{1\ \text{Mg}}{1 \times 10^6\ \text{g}}$ and $\dfrac{1 \times 10^6\ \text{g}}{1\ \text{Mg}}$

    (c) $\dfrac{1\ \text{L}}{1 \times 10^3\ \text{mL}}$ and $\dfrac{1 \times 10^3\ \text{mL}}{1\ \text{L}}$

    (d) $\dfrac{1\ \text{s}}{1 \times 10^6\ \mu\text{s}}$ and $\dfrac{1 \times 10^6\ \mu\text{s}}{1\ \text{s}}$
15. (a) 5.00×10^{-3} km; (b) 5.00×10^2 cg; (c) 50.0 dL; (d) 5.00×10^9 ns
17. (a) 6.50×10^6 Mm; (b) 6.5×10^8 kg; (c) 6.50×10^{-2} dL; (d) 0.650 ps
19. (a) 1 in. = 2.54 cm; (b) 1 lb = 454 g; (c) 1 qt = 946 mL; (d) 1 sec = 1.00 s
21. (a) 9.8 in.; (b) 0.055 lb; (c) 2,370 mL; (d) 2.50 s
23. (a) 1.8 m; (b) 79.5 kg; (c) 0.473 L; (d) 0.273 ds
25. 541 m
27. 45 mi/gal
29. 4.91 g copper
31. 40.5% of solution is ethanol
33. 2.5% copper
35. (a) 1 cm^3 (exactly); (b) 16.4 cm^3

37. 1.09 cm^3
39. 1.90×10^{-3} cm
41. 5.82 L
43. 2.5 mL
45. 62.5 mL
47. (a) float; (b) sink
49. (a) rise; (b) drop
51. 1.00 g/mL
53. (a) 0.788 g/mL; (b) 0.714 g/mL
55. (a) 170 g; (b) 120 g
57. (a) 35.0 mL; (b) 18.9 mL
59. (a) 32 °F; (b) 0 °C; (c) 273 K
61. (a) 38 °C; (b) −137 °C
63. (a) 768 K; (b) 88 K
65. −128 °F
67. gold
69. 1900 kJ
71. 67.5 cal
73. 0.0309 cal/g × °C
75. 52.5 g titanium
77. 1.5×10^6 MB
79. 320 DVDs
81. 35,000 μs; 3.5×10^4 μs
83. 5.87×10^{12} mi
85. 435 m^2
87. A 400-meter race (0.248 mi) is less than a quarter mile (0.250 mi).
89. 7690 tablets
91. 0.801
93. 8.33 lb/gal
95. 3.90 cm

Chapter 3

1. solid state
3. gaseous state
5. solid state
7. (a) condensing; (b) freezing
9. (a) cooling; (b) cooling
11. cooling
13. (a) element; (b) element; (c) compound; (d) mixture
15. (a) element; (b) heterogeneous mixture; (c) compound; (d) homogeneous mixture
17. (a) Li; (b) Ar; (c) Mg; (d) Mn; (e) Be; (f) Si; (g) Hg; (h) Ti
19. (a) phosphorus; (b) sulfur; (c) krypton; (d) iron; (e) antimony; (f) nitrogen; (g) platinum; (h) silver
21. (a) 1; (b) 5; (c) 13; (d) 22; (e) 33; (f) 38; (g) 50; (h) 83
23. (a) metal; (b) nonmetal; (c) metal; (d) nonmetal
25. (a) nonmetal; (b) nonmetal; (c) metal; (d) nonmetal
27. (a) nonmetal; (b) nonmetal; (c) metal; (d) metal
29. (a) semimetal; (b) metal; (c) metal; (d) metal
31. (a) gas; (b) solid; (c) gas; (d) solid
33. (a) gas; (b) solid; (c) gas; (d) liquid
35. (a) 27 carbon atoms; 46 hydrogen atoms; 1 oxygen atom; (b) 20 carbon atoms; 24 hydrogen atoms; 2 nitrogen atoms; 2 oxygen atoms
37. (a) $C_6H_{14}N_2O_2$; (b) $C_5H_{11}NO_2S$
39. (a) 21 atoms; (b) 12 atoms
41. The mass ratio of the elements in water is constant: 1:8.
43. (a) physical; (b) physical; (c) physical; (d) chemical
45. (a) chemical; (b) physical; (c) chemical; (d) physical
47. (a) chemical; (b) physical; (c) chemical; (d) physical
49. (a) chemical; (b) physical; (c) chemical; (d) chemical
51. (a) chemical; (b) physical; (c) physical; (d) chemical
53. (a) physical; (b) chemical; (c) physical; (d) chemical
55. (a) chemical; (b) physical; (c) physical; (d) chemical
57. (a) physical; (b) chemical; (c) chemical; (d) physical
59. 1.49 g zinc sulfide
61. 1.60 g carbon dioxide
63. solid state
65. increase
67. increase
69. 10.5 kcal
71. 807 cal
73. 68.2 kcal
75. (a) chemical → heat; (b) heat → mechanical; (c) mechanical → mechanical; (d) mechanical → electrical
77. (a) heat → mechanical; (b) mechanical → mechanical; (c) mechanical → mechanical; (d) mechanical → electrical
79. A pre-1982 penny is an alloy, thus a homogeneous mixture.
81. homogeneous mixture
83. germanium
85. oxygen, silicon, and aluminum
87. (a) 26; (b) 29; (c) 50; (d) 82
89. 3250 cal
91. In theory, the mass should be slightly less.
93. E = energy, m = mass, c = speed of light
95. physical change

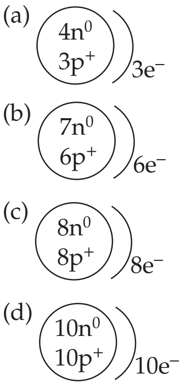
Chapter 4

1. Empedocles
3. Robert Boyle; Antoine Lavoisier; and Joseph Proust
5. (c) Atoms are indivisible
7. electrons
9. electron (e^-)
11. electron (−1)
13. Protons and neutrons are located in the nucleus.
15. 1×10^{-8} cm
17. 1 and 1836
19. (a) 0 n^0; (b) 2 n^0; (c) 8 n^0; (d) 10 n^0
21. (a) 0 n^0; (b) 7 n^0; (c) 32 n^0; (d) 74 n^0

23.

Atomic Notation	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
$^{11}_{5}\text{B}$	5	11	5	6	5
$^{15}_{7}\text{N}$	7	15	7	8	7
$^{40}_{20}\text{Ca}$	20	40	20	20	20
$^{200}_{80}\text{Hg}$	80	200	80	120	80

25.



27. The atomic number refers to the number of protons in the nucleus; whereas, the mass number refers to the number of protons *and* neutrons.

29. The carbon-12 reference isotope is assigned a mass of 12 amu.

31. 22.99 amu

33. 19.00 amu

35. 6.941 amu

37. 55.85 amu

39. red light

41. yellow light

43. yellow light

45. 400 nm

47. photon

49. (a) continuous; (b) quantized

51. (a) continuous; (b) quantized

53. 5 → 2

55. 4 → 1

57. red line

59. ultraviolet

61. red line

63. (a) 1; (b) 1

65. (a) ultraviolet; (b) red; (c) infrared

67. lines in the emission spectrum of hydrogen

69. (a) 1 sublevel; (b) 2 sublevels; (c) 3 sublevels; (d) 4 sublevels

71. (a) 2 e^-; (b) 6 e^-; (c) 10 e^-; (d) 14 e^-

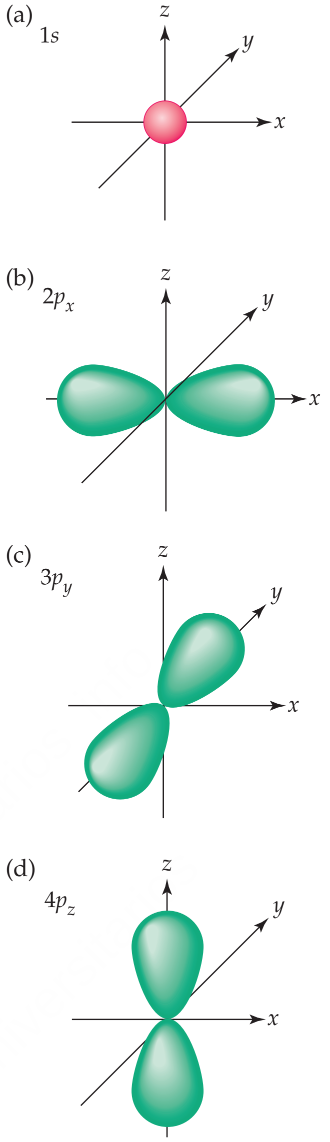
73. 8 e^-

75. See Figure 4.16; 3*d* follows 4*s*

77. (a) $1s^2$; (b) $1s^2\ 2s^2$; (c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$;
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}$

79. (a) Li; (b) Si; (c) Ti; (d) Sr

81.



83. (a) 3*s*; (b) $3p_x$; (c) equal; (d) equal

85. (a) 5*s*; (b) 4*p*

87. (a) 2 e^-; (b) 2 e^-; (c) 2 e^-; (d) 2 e^-

89. 106.9 amu

91. no stable isotopes

93. (a) ultraviolet; (b) infrared

95. Ag-107

97. An atom is more stable with a completely filled *d* sublevel than a partially filled sublevel. If one of the silver 5*s* electrons drops into the 4*d* sublevel, the 4*d* sublevel is completely filled, and thus the silver atom is more stable.

Chapter 5

1. Neon was not discovered until 1898 (see Figure 3.7).

3. Mendelevium was not synthesized until 1955 (see Figure 3.7).

5. increasing atomic mass

7. Co and Ni

9. groups or families

11. representative elements

13. the lanthanides

15. the lanthanides

17. metals

19. Sc and Y

21. (a) IA/1; (b) IIA/2; (c) VIIA/17; (d) VIIIA/18

23. (a) 1; (b) 13; (c) 15; (d) 17; (e) 11; (f) 3; (g) 5; (h) 7

25. (a) B; (b) Ge; (c) Cl; (d) At

27. (a) Na; (b) Mg; (c) Lu; (d) Pm

29. decreases

31. decreases

33. (a) K; (b) Mg; (c) Fe; (d) S

35. (a) Na; (b) P; (c) Ca; (d) Kr

37. 284 pm; 2.33 g/mL; 17.5 °C

39. 151 pm; 6.76 g/mL; 309 °C

41. (a) KF; (b) CaF_2; (c) BBr_3; (d) GeI_4

43. (a) CaF_2; (b) $CaCl_2$; (c) $CaBr_2$; (d) CaI_2

45. (a) As_2O_3 and As_2O_5; (b) Sb_2O_3 and Sb_2O_5

47. *s* sublevels

49. *d* sublevels

51. 4*f* sublevel

53. (a) 1*s*; (b) 3*s*; (c) 4*f*; (d) 4*p*; (e) 5*s*; (f) 2*p*; (g) 5*p*; (h) 6*s*

55. (a) $1s^2\ 2s^1$
(b) $1s^2\ 2s^2\ 2p^5$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^2$
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
(e) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
(f) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$
(g) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$
(h) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$

57. (a) 1; (b) 3; (c) 5; (d) 7

59. (a) 1; (b) 3; (c) 5; (d) 7; (e) 2; (f) 4; (g) 6; (h) 8

61. (a)  (b) 
(c)  (d) 
(e)  (f) 
(g)  (h) 

63. increases

65. Group IA/1

67. (a) Mg; (b) S; (c) Sn; (d) N

69. (a) Cs; (b) Ar; (c) Al; (d) I

71. (a) 1+; (b) 2+; (c) 3+; (d) 4+

73. (a) 2+; (b) 4+; (c) 3−; (d) 2−

75. (a) K^+; (b) Sr^{2+}; (c) Cl^-; (d) Se^{2-}

77. (a) [Ne]; (b) [Ar]; (c) [Ar] $3d^6$; (d) [Kr] $4d^2$

79. (a) [He] $2s^2\ 2p^6$, or [Ne]
(b) [Ne] $3s^2\ 3p^6$, or [Ar]
(c) [He] $2s^2\ 2p^6$, or [Ne]
(d) [Kr] $5s^2\ 4d^{10}\ 5p^6$, or [Xe]

81. (a) Group IA; (b) Group IB; (c) Group IIIB; (d) Group IIIA

83. 131 pm; 13.42 g/mL; 1824 °C

85. (a) [Kr] $5s^2$; (b) [Kr] $5s^2\ 4d^6$; (c) [Kr] $5s^2\ 4d^{10}\ 5p^3$; (d) [Xe] $6s^1$

87. Although H and the Group IA/1 metals each have one valence electron, in H the electron is closer to its nucleus. Because the negatively charged electron is closer to the positively charged nucleus, it requires more energy to remove the electron from hydrogen than a Group IA/1 metal.

89. scandium (Sc)

91. When an alkali metal atom loses one electron, it gains a noble gas configuration. An alkaline earth metal atom must lose two electrons to gain a noble gas electron configuration.

Chapter 6

1. (a) binary molecular; (b) binary ionic; (c) ternary ionic; (d) binary acid

3. (a) polyatomic cation; (b) monoatomic cation; (c) monoatomic anion; (d) polyatomic anion

5. (a) calcium ion; (b) cadmium ion; (c) cobalt(II) ion; (d) copper(II) ion

7. (a) fluoride ion; (b) chloride ion; (c) bromide ion; (d) iodide ion

9. (a) K^+; (b) Ba^{2+}; (c) Fe^{2+}; (d) Sn^{2+}

11. (a) nitrate ion; (b) nitrite ion; (c) chlorate ion; (d) chlorite ion

13. (a) $CrO_4{}^{2-}$; (b) $Cr_2O_7{}^{2-}$; (c) OH^-; (d) CN^-

15. (a) $AlBr_3$; (b) CdS; (c) Mn_3P_2; (d) PbF_4

17. (a) KNO_3; (b) $(NH_4)_2Cr_2O_7$; (c) Cu_2CO_3; (d) $MnSO_3$

19. (a) $Sr(NO_2)_2$; (b) $Zn(MnO_4)_2$; (c) $CaCrO_4$; (d) $Ni(ClO_4)_2$

21. (a) magnesium oxide; (b) zinc oxide; (c) cadmium oxide; (d) barium oxide

23. (a) cobalt(II) oxide; (b) iron(II) oxide; (c) mercury(II) oxide; (d) tin(II) oxide

25. (a) Cu_2O; (b) Fe_3N_2; (c) $HgCl_2$; (d) PbS_2

27. (a) RbCl; (b) NaBr

29. (a) GaN; (b) AlAs

31. (a) lithium permanganate; (b) strontium chlorite; (c) barium carbonate; (d) ammonium dichromate

33. (a) copper(II) sulfate; (b) iron(II) chromate; (c) mercury(II) nitrite; (d) lead(II) acetate

35. (a) $Mn(C_2H_3O_2)_2$; (b) $Pb(ClO_2)_4$; (c) $Sn_3(PO_4)_2$; (d) $Fe(ClO)_3$

37. (a) Fr_2SO_4; (b) Na_2SO_3

39. (a) $Ra(ClO_3)_2$; (b) $Ba(BrO_3)_2$

41. (a) cobaltous ion; cobaltic ion; stannous ion; stannic ion

43. (a) Cu^+; Cu^{2+}; Pb^{2+}; Pb^{4+}

45. (a) cupric sulfide; ferrous oxide; mercuric nitrite; plumbous acetate

47. (a) $CuCl_2$; FeS; $Co(C_2H_3O_2)_3$; $SnCrO_4$

49. (a) N_2O_5; (b) CCl_4; (c) IBr; (d) H_2S

51. (a) bromine monofluoride; (b) carbon tetrafluoride; (c) diiodine tetraoxide; (d) dichlorine trioxide

53. (a) hydrosulfuric acid; (b) hydrobromic acid

55. (a) $HC_2H_3O_2(aq)$; (b) $HNO_3(aq)$

57. (a) chlorous acid; (b) phosphoric acid

59. (a) 0; (b) 2+; (c) 3+; (d) 0

61. (a) IO_4^-; (b) SiO_3^{2-}

63.

Ions	chloride ion	sulfide ion	phosphide ion
copper(I) ion	$CuCl$ copper(I) chloride	Cu_2S copper(I) sulfide	Cu_3P copper(I) phosphide
cobalt(III) ion	$CoCl_3$ cobalt(III) chloride	Co_2S_3 cobalt(III) sulfide	CoP cobalt(III) phosphide
lead(IV) ion	$PbCl_4$ lead(IV) chloride	PbS_2 lead(IV) sulfide	Pb_3P_4 lead(IV) phosphide

65.

Ions	hydroxide ion	sulfate ion	phosphate ion
mercurous ion	$Hg_2(OH)_2$ mercurous hydroxide	Hg_2SO_4 mercurous sulfate	$(Hg_2)_3(PO_4)_2$ mercurous phosphate
ferrous ion	$Fe(OH)_2$ ferrous hydroxide	$FeSO_4$ ferrous sulfate	$Fe_3(PO_4)_2$ ferrous phosphate
stannous ion	$Sn(OH)_2$ stannous hydroxide	$SnSO_4$ stannous sulfate	$Sn_3(PO_4)_2$ stannous phosphate

67. (a) −*ide*; (b) −*ic acid*

69. (a) −*ite*; (b) −*ous acid*

71. (a) −*ate*; (b) −*ic acid*

73. (a) H_2O; (b) $NaClO$; (c) $NaOH$; (d) $NaHCO_3$

75. (a) boron trifluoride; (b) silicon tetrachloride; (c) diarsenic pentaoxide; (d) diantimony trioxide

77. TiO_2

79. $AcCl_3$

Chapter 7

1. (a) chemical reaction; (b) chemical reaction

3. (a) chemical reaction; (b) physical change

5. (a) chemical reaction; (b) chemical reaction

7. $Mg(s) + Br_2(l) \rightarrow MgBr_2(s)$

9. $Zn(HCO_3)_2(s) \rightarrow ZnCO_3(s) + H_2O(g) + CO_2(g)$

11. $Cd(s) + Co(NO_3)_2(aq) \rightarrow Cd(NO_3)_2(aq) + Co(s)$

13. $LiI(aq) + AgNO_3(aq) \rightarrow AgI(s) + LiNO_3(aq)$

15. $HC_2H_3O_2(aq) + NH_4OH(aq) \rightarrow NH_4C_2H_3O_2(aq) + HOH(l)$

17. (a) $3\,H_2(g) + N_2(g) \rightarrow 2\,NH_3(g)$
(b) $Al_2(CO_3)_3(s) \rightarrow Al_2O_3(s) + 3\,CO_2(g)$
(c) $Sr(s) + 2\,H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$
(d) $K_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2\,KOH(aq)$
(e) $2\,H_3PO_4(aq) + 3\,Mn(OH)_2(s)$
$\rightarrow Mn_3(PO_4)_2(s) + 6\,HOH(l)$

19. (a) $2\,Pb(s) + O_2(g) \rightarrow 2\,PbO(s)$
(b) $2\,LiNO_3(s) \rightarrow 2\,LiNO_2(s) + O_2(g)$
(c) $Mg(s) + 2\,HC_2H_3O_2(aq) \rightarrow Mg(C_2H_3O_2)_2(aq) + H_2(g)$
(d) $Hg_2(NO_3)_2(aq) + 2\,NaBr(aq)$
$\rightarrow Hg_2Br_2(s) + 2\,NaNO_3(aq)$
(e) $H_2CO_3(aq) + 2\,NH_4OH(aq)$
$\rightarrow (NH_4)_2CO_3(aq) + 2\,HOH(l)$

21. (a) combination; (b) decomposition; (c) single replacement; (d) double replacement; (e) neutralization

23. (a) combination; (b) decomposition; (c) single replacement; (d) double replacement; (e) neutralization

25. metal + oxygen gas → metal oxide
(a) $2\,Ni(s) + O_2(g) \rightarrow 2\,NiO(s)$
(b) $4\,Fe(s) + 3\,O_2(g) \rightarrow 2\,Fe_2O_3(s)$

27. (a) $2\,C(s) + O_2(g) \rightarrow 2\,CO(g)$
(b) $4\,P(s) + 5\,O_2(g) \rightarrow 2\,P_2O_5(s)$

29. (a) $2\,Cu(s) + Cl_2(g) \rightarrow 2\,CuCl(s)$
(b) $Co(s) + S(s) \rightarrow CoS(s)$

31. (a) $4\,Cr(s) + 3\,O_2(g) \rightarrow 2\,Cr_2O_3(s)$
(b) $2\,Cr(s) + N_2(g) \rightarrow 2\,CrN(s)$

33. (a) $4\,Li + O_2 \rightarrow 2\,Li_2O$
(b) $2\,Ca + O_2 \rightarrow 2\,CaO$

35. (a) $2\,Na + I_2 \rightarrow 2\,NaI$
(b) $3\,Ba + N_2 \rightarrow Ba_3N_2$

37. (a)$2\,KHCO_3(s) \rightarrow K_2CO_3(s) + H_2O(g) + CO_2(g)$
(b) $Sr(HCO_3)_2(s) \rightarrow SrCO_3(s) + H_2O(g) + CO_2(g)$

39. (a)$Co_2(CO_3)_3(s) \rightarrow Co_2O_3(s) + 3\,CO_2(g)$
(b) $Sn(CO_3)_2(s) \rightarrow SnO_2(s) + 2\,CO_2(g)$

41. (a) $2\,AgHCO_3(s) \rightarrow Ag_2CO_3(s) + H_2O(g) + CO_2(g)$
(b) $Zn(HCO_3)_2(s) \rightarrow ZnCO_3(s) + H_2O(g) + CO_2(g)$

43. (a) $Ca(NO_3)_2(s) \rightarrow Ca(NO_2)_2(s) + O_2(g)$
(b) $2\,Ag_2SO_4(s) \rightarrow 2\,Ag_2SO_3(s) + O_2(g)$

45. (a) $2\,NaClO_3(s) \rightarrow 2\,NaCl(s) + 3\,O_2(g)$
(b) $Ca(NO_3)_2(s) \rightarrow Ca(NO_2)_2(s) + O_2(g)$

47. (b) Sn; (c) Co; (d) Mn

49. (b) Sn; (c) Co; (d) Mn

51. (a) Li; (c) Ca

53. (a) $2\,Al(s) + 3\,Cu(NO_3)_2(aq) \rightarrow 3\,Cu(s) + 2\,Al(NO_3)_3(aq)$
(b) $Cu(s) + Al(NO_3)_3(aq) \rightarrow NR$

55. (a) $Ni(s) + Pb(C_2H_3O_2)_2(aq) \rightarrow Pb(s) + Ni(C_2H_3O_2)_2(aq)$
(b) $Pb(s) + Ni(C_2H_3O_2)_2(aq) \rightarrow NR$

57. (a) $Mg(s) + 2\,HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
(b) $Mn(s) + 2\,HCl(aq) \rightarrow MnCl_2(s) + H_2(g)$

59. (a) $2\,Li(s) + 2\,H_2O(l) \rightarrow 2\,LiOH(aq) + H_2(g)$
(b) $Ba(s) + 2\,H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$

61. (a) $Zn(s) + Pb(NO_3)_2(aq) \rightarrow Pb(s) + Zn(NO_3)_2(aq)$
(b) $Cd(s) + Fe(NO_3)_2(aq) \rightarrow NR$

63. (a) $Zn(s) + 2\,HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
(b) $Cd(s) + 2\,HNO_3(aq) \rightarrow Cd(NO_3)_2(aq) + H_2(g)$

65. (a) $Zn(s) + H_2O(l) \rightarrow NR$
(b) $Ba(s) + 2\,H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$

67. (a) $(NH_4)_2S$; (d) $Pb(C_2H_3O_2)_2$

69. (a) insoluble; (c) insoluble; (d) insoluble

71. (a) $SnCl_2(aq) + K_2S(aq) \rightarrow SnS(s) + 2\ KCl(aq)$
(b) $Co(NO_3)_2(aq) + Na_2CrO_4(aq) \rightarrow CoCrO_4(s) + 2\ NaNO_3(aq)$

73. (a) $MgSO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + MgCl_2(aq)$
(b) $FeSO_4(aq) + Ca(OH)_2(aq) \rightarrow Fe(OH)_2(s) + CaSO_4(aq)$

75. (a) $H_2CO_3(aq) + 2\ LiOH(aq) \rightarrow Li_2CO_3(aq) + 2\ HOH(l)$
(b) $2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ HOH(l)$

77. (a) $2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ HOH(l)$
(b) $2\ HNO_2(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_2)_2(aq) + 2\ HOH(l)$

79. Not necessarily. The sum of the coefficients may or may not be the same.

81. (a) $3\ Fe(s) + 4\ H_2O(g) \rightarrow Fe_3O_4(s) + 4\ H_2(g)$
(b) $4\ FeS(s) + 7\ O_2(g) \rightarrow 2\ Fe_2O_3(s) + 4\ SO_2(g)$

83. (a) $F_2(g) + 2\ NaBr(aq) \rightarrow Br_2(l) + 2\ NaF(aq)$
(b) $Sb_2S_3(s) + 6\ HCl(aq) \rightarrow 2\ SbCl_3(aq) + 3\ H_2S(g)$

85. (a) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$
(b) $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

87. $Fe_2O_3(l) + 3\ CO(g) \rightarrow 2\ Fe(l) + 3\ CO_2(g)$

Chapter 8

1. (a) 22.99 amu; (b) 87.62 amu; (c) 28.09 amu; (d) 78.97 amu

3. (a) 22.99 g; (b) 87.62 g; (c) 28.09 g; (d) 78.97 g

5. (a) 6.02×10^{23} atoms C; (b) 6.02×10^{23} molecules CH_4

7. (a) 1 mol S; (b) 1 mol SO_2

9. (a) 6.02×10^{22} atoms Mg; (b) 1.20×10^{23} molecules Cl_2; (c) 1.81×10^{23} formula units $MgCl_2$

11. (a) 0.0415 mol Fe; (b) 0.831 mol SO_3; (c) 12.5 mol $FeSO_4$

13. (a) 26.98 g/mol; (b) 28.09 g/mol; (c) 74.92 g/mol; (d) 32.07 g/mol

15. (a) 72.15 g/mol; (b) 136.15 g/mol; (c) 150.90 g/mol; (d) 92.11 g/mol

17. (a) 98.3 g Hg; (b) 0.540 g N_2; (c) 1.75 g $BaCl_2$

19. (a) 2.31×10^{22} atoms K; (b) 8.84×10^{21} molecules O_2; (c) 1.75×10^{21} formula units $AgClO_3$

21. (a) 1.15×10^{-23} g/atom; (b) 4.66×10^{-23} g/atom; (c) 1.06×10^{-22} g/molecule; (d) 7.31×10^{-23} g/molecule

23. 0 °C and 1 atm

25. (a) 1.78 g/L; (b) 1.70 g/L; (c) 0.717 g/L; (d) 1.97 g/L

27. (a) 32.0 g/mol; (b) 34.0 g/mol; (c) 44.1 g/mol; (d) 121 g/mol

29.

Gas	Molecules	Mass	Volume at STP
nitrogen, N_2	6.02×10^{23}	28.02 g	22.4 L
oxygen, O_2	6.02×10^{23}	32.00 g	22.4 L
nitrogen dioxide, NO_2	6.02×10^{23}	46.01 g	22.4 L

31. (a) 3.22×10^{21} molecules CO; (b) 2.69×10^{21} molecules H_2

33. (a) 1.60 g H_2S; (b) 3.97 g C_3H_8

35. (a) 1.40 L He; (b) 0.826 L CH_4

37.

Gas	Molecules	Mass	Volume at STP
methane, CH_4	1.50×10^{23}	4.00 g	5.58 L
ethane, C_2H_6	1.50×10^{23}	7.50 g	5.58 L
propane, C_3H_8	1.50×10^{23}	11.0 g	5.58 L

39. The law of definite composition states that a compound always contains the same elements in the same proportion by mass. Thus, the percent sodium in a kilogram of salt 39.34%.

41. 42.10% C, 6.49% H, 51.42% O

43. 40.24% C, 7.444% H, 9.388% N, 21.49% S, 21.44% O

45. 26.722% Hg, 6.125% Na, 31.999% C, 1.08% H, 21.288% Br, 12.79% O

47. CoS

49. CuO

51. SnO_2

53. CuCl

55. $SnBr_2$

57. C_2HCl_3

59. $C_9H_8O_4$

61. $C_6H_{10}O_4$

63. CH_3O, $C_2H_6O_2$

65. CHCl, $C_6H_6Cl_6$

67. C_5H_7N, $C_{10}H_{14}N_2$

69. $GeCl_4$

71. 5-g nickel coin

73. Earth

75. 2.11×10^{22} atoms C

77. 11.3 g Cu_2O

79. 2.99×10^{-23} cm^3 H_2O

81. 6.04×10^{23} atoms Cu/mol

Chapter 9

1. (a) 1 mol C; (b) 2 L B

3. (a) 2.50 g C; (b) 3.00 g B

5. (a) 152.04 g → 152.04 g; (b) 216.04 g → 216.04 g

7. 2.50 mol O_2

9. 2.50 mol O_2

11. 0.150 mol Cl_2 react; 0.100 mol $FeCl_3$

13. 0.500 mol C_3H_8 react; 1.50 mol CO_2

15. mass–mass problem

17. mass–volume problem

19. volume–volume problem

21. 1.24 g ZnO

23. 5.21 g $BiCl_3$

25. 3.39 g Ag

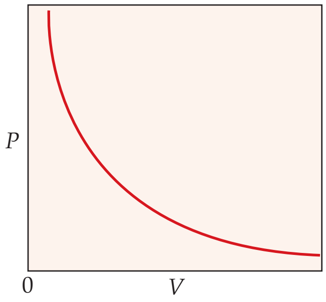
27. 6.38 g Hg

29. 1.68 g $Ca_3(PO_4)_2$
31. 233 mL CO_2
33. 1150 mL O_2
35. 129 mL O_2
37. 510 mL O_2
39. 16.7 L N_2
41. 5.00 mL O_2
43. 1750 mL Cl_2
45. 50.0 L SO_3
47. 500.0 cm^3 N_2
49. N_2 or O_2 are each a limiting reactant because each produces 10.0 mol NO.
51. The limiting reactant is NO, which produces 1.50 mol of NO_2.
53. The limiting reactant is H_2, which produces 5.00 mol of H_2O.
55. The limiting reactant is C_2H_6, which produces 3.00 mol of H_2O.
57. (a) Neither Co nor S is a limiting reactant; the reaction produces 1.00 mol CoS.
After the reaction:
Co = 0.00 mol; S = 0.00 mol; CoS = 1.00 mol.

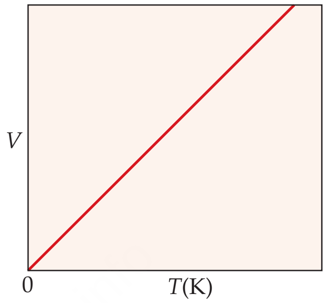
(b) The limiting reactant is Co, which produces 2.00 mol of CoS.
After the reaction:
Co = 0.00 mol; S = 1.00 mol; CoS = 2.00 mol.
59. Mg is the limiting reactant and produces 23.0 g Fe.
61. Al is the limiting reactant and produces 31.1 g Fe.
63. 0.742 g $MgSO_4$
65. 0.693 g H_2O
67. 90.0 mL NO_2
69. 30.0 mL N_2O_3
71. 2.50 L SO_3
73. 20.0 L Cl_2
75. 104%
77. 106%
79. g/mol
81. 0.929 g Cr_2O_3
83. 655 g Al
85. 0.593 L H_2S
87. 16.3 g H_2O
89. 124 L H_2

Chapter 10

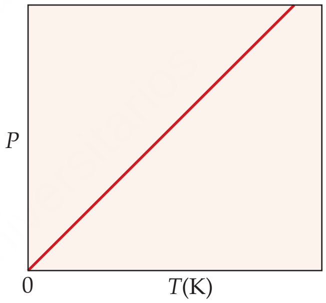
1. (a), (b), and (d)
3. (a) 760 mm Hg; (b) 14.7 psi
5. (a) 1.90×10^3 mm Hg; (b) 36.8 psi
7. (a) 104 kPa; (b) 783 torr
9. (a) pressure decreases; (b) pressure increases; (c) pressure increases
11. (a) pressure decreases; (b) pressure increases; (c) pressure increases
13. pressure decreases
15.



17. 3.50 atm
19. 1.19 L
21.



23. 602 mL
25. −31 °C
27.



29. 1.31 atm
31. 215 °C
33. 8.83 L
35. 1.33 L
37. 8.19 atm
39. 5187 °C (rounds to 5190 °C)
41. 0.704 atm
43. vapor pressure of water is greater
45. (a) 23.8 mm Hg; (b) 92.5 mm Hg
47. 760 mm Hg
49. 135 mm Hg
51. A gas may be collected over water in a graduated cylinder by displacing an equal volume of water (see Figure 10.11).
53. 749 mm Hg
55. zero pressure (0 mm Hg)
57. zero kinetic energy
59. Refer to Section 10.10 in the textbook.
61. (a) all gases are equal; (b) all gases are equal; (c) He atoms; (d) Ar atoms
63. 0.286 mol He

65. 0.411 atm

67. 1030 cm H_2O

69. 2350 psi

71. 38.5 mL

73. 1.23 atm

75. CH_4 molecules

77. 3.92 g Cl_2

79. 3750 lb

81. O_3 (48.0 g/mol)

83. 12.5 g/mol

Chapter 11

1. (a), (b), and (d) are true, and generally observed properties of liquids.

3. (a) solid; (b) gas; (c) solid; (d) liquid

5. A *temporary dipole* is caused by electrons moving about in a molecule. The resulting positive and negative charges in a molecule give rise to a temporary attraction between molecules. A *permanent dipole* is caused by electrons being drawn closer to one atom than another within a molecule. The resulting positive charge in one molecule is attracted to the negative charge in another molecule.

7. The vapor pressure of water is the pressure exerted by water vapor molecules when the *rate of evaporation* is equal to the *rate of condensation*.

9. The intermolecular attraction in water is strong; thus the viscosity of water is higher than other liquids with molecules the same size.

11. (a) low; (b) high; (c) high; (d) high

13. (a) C_2H_5Cl; (b) CH_3OCH_3

15. (a) CH_3COOH; (b) C_2H_5OH

17. As the temperature of a liquid increases, the vapor pressure increases.

19. (a) ~180 mm Hg; (b) ~330 mm Hg

21. 56 °C

23. (a) true; (b) true; (c) true; (d) false

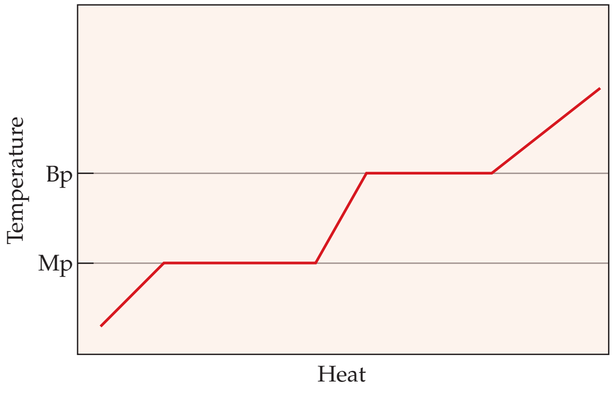
25. (a) liquid; (b) liquid

27. NaCl, Cl_2, and Na

29. (a) ions; (b) molecules; (c) atoms

31. (a) molecular; (b) molecular; (c) metallic; (d) ionic

33.



35. 2.00×10^3 cal (2.00 kcal)

37. 15,400 cal (15.4 kcal)

39. 82,800 cal (82.8 kcal)

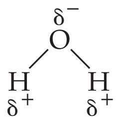
41. 28,100 cal (28.1 kcal)

43. Water has an unusually *low* vapor pressure.

45. Water has an unusually *high* viscosity.

47. 104.5°

49.



51. Water has the highest specific heat of any liquid.

53. Water is more dense than ice.

55. (a) H_2O; (b) H_2Se

57. (a) H_2O; (b) H_2Se

59. (a) increases; (b) increases; (c) increases; (d) increases

61. $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(l)$

63. (a) $2\,Li(s) + 2\,H_2O(l) \rightarrow 2\,LiOH(aq) + H_2(g)$

(b) $Na_2O(s) + H_2O(l) \rightarrow 2\,NaOH(aq)$

(c) $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

65. (a) $Mg(s) + H_2O(l) \rightarrow NR$

(b) $SrO(s) + H_2O(l) \rightarrow Sr(OH)_2(aq)$

(c) $N_2O_5(s) + H_2O(l) \rightarrow 2\,HNO_3(aq)$

67. (a) $2\,C_3H_6(g) + 9\,O_2(g) \rightarrow 6\,CO_2(g) + 6\,H_2O(g)$

(b) $Na_2Cr_2O_7 \cdot 2\,H_2O(s) \rightarrow Na_2Cr_2O_7(s) + 2\,H_2O(g)$

(c) $2\,HF(aq) + Ca(OH)_2(aq) \rightarrow CaF_2(aq) + 2\,HOH(l)$

69. (a) $2\ C_3H_7OH(l) + 9\ O_2(g) \xrightarrow{spark} 6\ CO_2(g) + 8\ H_2O(g)$

(b) $Ca(NO_3)_2 \cdot 4H_2O(s) \xrightarrow{\Delta} Ca(NO_3)_2(s) + 4\,H_2O(g)$

(c) $H_2CO_3(aq) + 2\,KOH(aq) \rightarrow K_2CO_3(aq) + 2\,HOH(l)$

71. (a) magnesium chloride hexahydrate; (b) manganese(II) sulfate heptahydrate; (c) cobalt(III) cyanide trihydrate;

(d) chromium(III) nitrate nonahydrate

73. (a) $NaC_2H_3O_2 \cdot 3\,H_2O$; (b) $CaSO_4 \cdot 2\,H_2O$;

(c) $K_2CrO_4 \cdot 4H_2O$; (d) $ZnSO_4 \cdot 7H_2O$

75. (a) 10.66% H_2O; (b) 33.81% H_2O; (c) 28.94% H_2O;

(d) 40.52% H_2O

77. (a) $SrCl_2 \cdot 2\,H_2O$; (b) $Sr(NO_3)_2 \cdot 6\,H_2O$; (c) $Ca(NO_3)_2 \cdot 4\,H_2O$;

(d) $Na_2B_4O_7 \cdot 5\,H_2O$

79. ~65%

81. (a) dispersion force; (b) dipole force; (c) hydrogen bond;

(d) dipole force

83. ~65 °C

85. Sulfuric acid, H_2SO_4, has O—H bonds and molecules show strong intermolecular attraction. Thus, sulfuric acid has high surface tension, minimal surface area, and sulfuric acid "raindrops" on Venus will be spherical.

87. 96,600 cal + 1,040,000 cal = 1,140,000 cal (1,140 kcal)

Chapter 12

1. Na (1 valence e^-); Cl (7 valence e^-)

3. Na^+ (0 valence e^-); Cl^- (8 valence e^-)

5. (a) ionic; (b) covalent; (c) covalent; (d) ionic

7. (a) molecule; (b) formula unit; (c) formula unit;

(d) molecule

9. (a) molecule; (b) atom; (c) formula unit; (d) molecule

11. (a) 1+; (b) 2+; (c) 3+; (d) 4+

13. (a) 1−; (b) 1−; (c) 2−; (d) 3−

15. (a) $1s^2\,2s^2\,2p^6$
(b) $1s^2\,2s^2\,2p^6$
(c) $1s^2\,2s^2\,2p^6$
(d) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^1$

17. (a) $1s^2\,2s^2\,2p^6$
(b) $1s^2\,2s^2\,2p^6$
(c) $1s^2\,2s^2\,2p^6$
(d) $1s^2\,2s^2\,2p^6$

19. (a) Ar; (b) Ar; (c) Ar; (d) Ar

21. (a) He; (b) Ne; (c) Ar; (d) Ne

23. (a) Ar; (b) Xe; (c) Ar; (d) Ar

25. (a) Li atom; (b) Mg atom; (c) F ion; (d) O ion

27. The true statements are (a) and (c).

29. (a) The bond length in H—Cl is less.
(b) The bond length in O═O is less.

31. The true statements are (a) and (d).

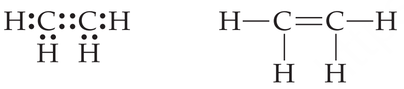
33. (a) $H{:}H$ $\quad H-H$
(b) $:F{:}F:$ $\quad :F-F:$
(c) $H{:}Br:$ $\quad H-Br:$
(d)

$H{:}\overset{H}{N}{:}H \qquad H-\overset{\overset{H}{|}}{N}-H$

35. (a) $:N{:::}N:$ $\quad :N{\equiv}N:$
(b)

$:I{:}\overset{:I:}{P}{:}I: \qquad :I-\overset{\overset{:I:}{|}}{P}-I:$

(c) $H{:}O{:}N{::}O$ $\quad H-O-N{=}O$
(d)

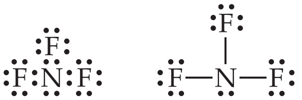


$H{:}\underset{H}{C}{::}\underset{H}{C}{:}H \qquad H-\underset{\underset{H}{|}}{C}=\underset{\underset{H}{|}}{C}-H$

37. (a)

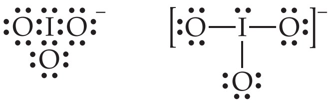
$H{:}\overset{H}{\underset{H}{C}}{:}H \qquad H-\overset{\overset{H}{|}}{\underset{\underset{H}{|}}{C}}-H$

(b) $:F{:}O{:}F:$ $\quad :F-O-F:$
(c) $H{:}O{:}O{:}H$ $\quad H-O-O-H$
(d)



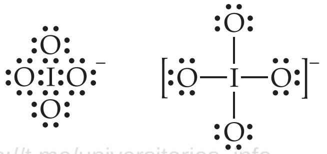
$:F{:}\overset{:F:}{N}{:}F: \qquad :F-\overset{\overset{:F:}{|}}{N}-F:$

39. (a) $:I{:}O:^-$ $\quad [:I-O:]^-$
(b) $:O{:}I{:}O:^-$ $\quad [:O-I-O:]^-$
(c)



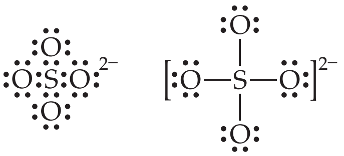
$:O{:}\underset{:O:}{I}{:}O:^- \qquad [:O-\underset{\underset{:O:}{|}}{I}-O:]^-$

(d)



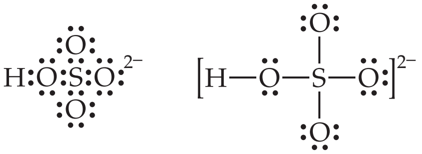
$:O{:}\overset{:O:}{\underset{:O:}{I}}{:}O:^- \qquad [:O-\overset{\overset{:O:}{|}}{\underset{\underset{:O:}{|}}{I}}-O:]^-$

41. (a)



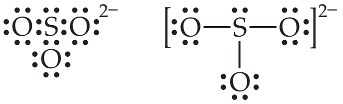
$:O{:}\overset{:O:}{\underset{:O:}{S}}{:}O:^{2-} \qquad [:O-\overset{\overset{:O:}{|}}{\underset{\underset{:O:}{|}}{S}}-O:]^{2-}$

(b)



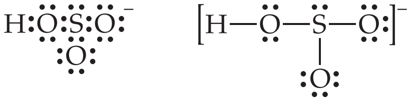
$H{:}O{:}\overset{:O:}{\underset{:O:}{S}}{:}O:^{2-} \qquad [H-O-\overset{\overset{:O:}{|}}{\underset{\underset{:O:}{|}}{S}}-O:]^{2-}$

(c)



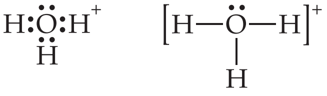
$:O{:}\underset{:O:}{S}{:}O:^{2-} \qquad [:O-\underset{\underset{:O:}{|}}{S}-O:]^{2-}$

(d)



$H{:}O{:}\underset{:O:}{S}{:}O:^{-} \qquad [H-O-\underset{\underset{:O:}{|}}{S}-O:]^{-}$

43. (a)



$H{:}\underset{H}{O}{:}H^+ \qquad [H-\underset{\underset{H}{|}}{O}-H]^+$

(b) $:O{:}H^-$ $\quad [:O-H]^-$
(c) $H{:}S:^-$ $\quad [H-S:]^-$
(d) $:C{:::}N:$ $\quad [:C{\equiv}N:]^-$

45. Electronegativity increases from left to right in the periodic table.

47. nonmetals

49. (a) Cl; (b) O; (c) Se; (d) F

51. (a) 0.2; (b) 1.2; (c) 0.5; (d) 0.3

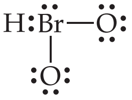
53. (a) δ^- C—H δ^+ (b) δ^+ Se—O δ^- (c) δ^+ P—I δ^-
(d) δ^+ H—Br δ^-

55. (a), (b), and (d) are nonpolar

57. H_2, N_2, Cl_2, and Br_2 occur naturally as diatomic molecules.

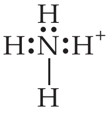
59. $H{:}Br-O:$

61.



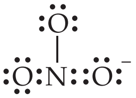
$H{:}\underset{\underset{:O:}{|}}{Br}-O:$

63.



$H{:}\overset{H}{\underset{\underset{H}{|}}{N}}{:}H^+$

65.



$:O{:}\overset{\overset{:O:}{|}}{N}{::}O:^-$

67. H—F: — — — H—F:

69. The *bond energy is much less* in a hydrogen bond than in a polar covalent bond.

71. tetrahedral, tetrahedral, 109.5°

73. tetrahedral, trigonal pyramidal, 107°

75. tetrahedral, bent, 104.5°

77. In CF_4, each of the four **C**—F bonds is polar, but each fluorine atom pulls equally in opposite directions to give a *nonpolar molecule*. All four F atoms are at the corners of a tetrahedron and the molecular shape is described as *tetrahedral*.

79. (a) atom; (b) molecule; (c) formula unit; (d) molecule

81. (a) Sr_3As_2; (b) RaO; (c) $Al_2(CO_3)_3$; (d) $Cd(OH)_2$

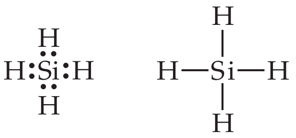
83. The radius of a sodium ion is about half that of a sodium atom because the sodium ion has one less energy level (3*s*). In addition, the sodium ion has one more proton than electron, which draws the electrons closer to the nucleus.

85. 1.0 (polar)

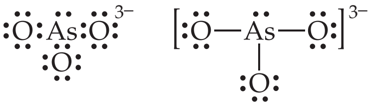
87. 0 (nonpolar)

89. δ^+ Ge—Cl δ^-

91.



93.



95. linear, linear, 180°

97. trigonal planar, trigonal planar, 120°

99.

Chapter 13

1. (a) solubility of CO_2 gas increases;
(b) solubility of CO_2 gas decreases

3. 8.55 mL N_2/1 dL blood

5. 2.1 atm

7. (a) miscible; (b) immiscible

9. (a) polar; (b) nonpolar; (c) polar; (d) nonpolar

11. (a) miscible; (b) immiscible; (c) immiscible; (d) miscible

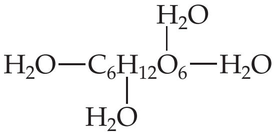
13. Add a few drops of water to the unknown liquid.
If the unknown liquid is miscible with water, it is polar.
If the unknown liquid is immiscible with water, it is nonpolar.

15. (a) insoluble; (b) soluble; (c) insoluble

17. (a) insoluble; (b) soluble; (c) soluble; (d) insoluble

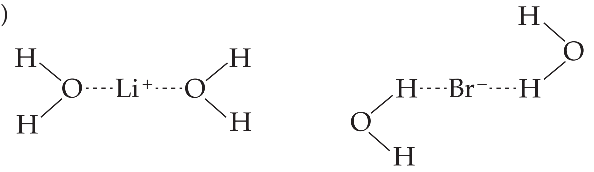
19. (a) fat soluble; (b) water soluble; (c) water soluble;
(d) fat soluble

21.

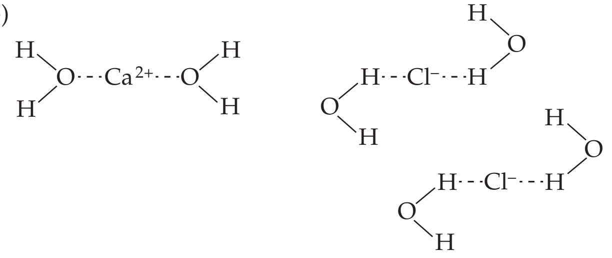


$$\mathrm{H_2O{-}C_6\underset{\underset{H_2O}{|}}{H_{12}}\overset{\overset{H_2O}{|}}{O_6}{-}H_2O}$$

23. (a)



(b)



25. (1) heating the solution; (2) stirring the solution;
(3) grinding the solute

27. (a) ~35 g NaCl; (b) ~35 g KCl

29. (a) ~35 g NaCl/100 g H_2O; (b) ~40 g KCl/100 g H_2O

31. (a) ~0 °C; (b) ~70 °C

33. (a) ~100 °C; (b) ~35 °C

35. (a) supersaturated; (b) saturated; (c) unsaturated

37. (a) ~80 g LiCl in solution; (b) ~20 g LiCl crystallizes

39. (a) supersaturated; (b) saturated; (c) unsaturated

41. saturated

43. (a) 1.25%; (b) 2.63%; (c) 4.00%; (d) 52.0%

45. (a) $\dfrac{1.50\text{ g KBr}}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{1.50\text{ g KBr}}$

$\dfrac{98.50\text{ g }H_2O}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{98.50\text{ g }H_2O}$

$\dfrac{1.50\text{ g KBr}}{98.50\text{ g }H_2O}$ and $\dfrac{98.50\text{ g }H_2O}{1.50\text{ g KBr}}$

(b) $\dfrac{2.50\text{ g }AlCl_3}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{2.50\text{ g }AlCl_3}$

$\dfrac{97.50\text{ g }H_2O}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{97.50\text{ g }H_2O}$

$\dfrac{2.50\text{ g }AlCl_3}{97.50\text{ g }H_2O}$ and $\dfrac{97.50\text{ g }H_2O}{2.50\text{ g }AlCl_3}$

(c) $\dfrac{3.75\text{ g }AgNO_3}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{3.75\text{ g }AgNO_3}$

$\dfrac{96.25\text{ g }H_2O}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{96.25\text{ g }H_2O}$

$\dfrac{3.75\text{ g }AgNO_3}{96.25\text{ g }H_2O}$ and $\dfrac{96.25\text{ g }H_2O}{3.75\text{ g }AgNO_3}$

(d) $\dfrac{4.25\text{ g }Li_2SO_4}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{4.25\text{ g }Li_2SO_4}$

$\dfrac{95.75\text{ g }H_2O}{100\text{ g solution}}$ and $\dfrac{100\text{ g solution}}{95.75\text{ g }H_2O}$

$\dfrac{4.25\text{ g }Li_2SO_4}{95.75\text{ g }H_2O}$ and $\dfrac{95.75\text{ g }H_2O}{4.25\text{ g }Li_2SO_4}$

47. (a) 53.6 g solution; (b) 2.00×10^2 g solution

49. (a) 1.70 g $FeBr_2$; (b) 5.25 g Na_2CO_3

51. (a) 95.0 g H_2O; (b) 225 g H_2O

53. (a) 0.257 *M* NaCl; (b) 0.0510 *M* $K_2Cr_2O_7$;
(c) 0.400 *M* $CaCl_2$; (d) 0.313 *M* Na_2SO_4

55. (a) $\dfrac{0.100\text{ mol LiI}}{1\text{ L solution}}$ and $\dfrac{1\text{ L solution}}{0.100\text{ mol LiI}}$

$\dfrac{0.100\text{ mol LiI}}{1000\text{ mL solution}}$ and $\dfrac{1000\text{ mL solution}}{0.100\text{ mol LiI}}$

(b) $\dfrac{0.100\text{ mol }NaNO_3}{1\text{ L solution}}$ and $\dfrac{1\text{ L solution}}{0.100\text{ mol }NaNO_3}$

$$\frac{0.100 \text{ mol NaNO}_3}{1000 \text{ mL solution}} \text{ and } \frac{1000 \text{ mL solution}}{0.100 \text{ mol NaNO}_3}$$

(c) $$\frac{0.500 \text{ mol K}_2\text{CrO}_4}{1 \text{ L solution}} \text{ and } \frac{1 \text{ L solution}}{0.500 \text{ mol K}_2\text{CrO}_4}$$

$$\frac{0.500 \text{ mol K}_2\text{CrO}_4}{1000 \text{ mL solution}} \text{ and } \frac{1000 \text{ mL solution}}{0.500 \text{ mol K}_2\text{CrO}_4}$$

(d) $$\frac{0.500 \text{ mol ZnSO}_4}{1 \text{ L solution}} \text{ and } \frac{1 \text{ L solution}}{0.500 \text{ mol ZnSO}_4}$$

$$\frac{0.500 \text{ mol ZnSO}_4}{1000 \text{ mL solution}} \text{ and } \frac{1000 \text{ mL solution}}{0.500 \text{ mol ZnSO}_4}$$

57. (a) 0.0761 L solution; (b) 0.0288 L solution; (c) 0.0141 L solution; (d) 0.0168 L solution

59. (a) 73.0 g $FeCl_3$; (b) 104 g KIO_4; (c) 2.38 g $ZnSO_4$; (d) 2.69 g $Ni(NO_3)_2$

61. 0.0250 M $Ca(OH)_2$

63. (a) 5.00%; (b) 0.291 M $C_6H_{12}O_6$

65. 0.48 M HCl

67. 0.014 L (14 mL)

69. (a) 0.208 M $Mg(NO_3)_2$; (b) 1.47 g AgI

71. Fog demonstrates the Tyndall effect when it contains small water droplets ranging in size from 1–100 nm.

73. Helium gas is less soluble in blood and allows for quicker vdecompression.

75. 228 g SO_2

77. Although ethyl ether, $C_4H_{10}O$, contains two C—O polar bonds, it is only partially miscible in water because it has mostly nonpolar C—H bonds.

79. Water can hydrogen bond with the polar —OH in an alcohol, and draw it into solution. As the nonpolar C_xH_y– – portion of the alcohol molecule increases, the alcohol becomes less soluble, and eventually immiscible with water.

81. (a) In a 40% solution, ethanol is the solute and water is the solvent.

(b) In a 95% solution, water is the solute and ethanol is the solvent.

83. 0.841 M CH_3COOH

Chapter 14

1. (a) and (b) are general properties of an acidic solution.

3. (a) strongly acidic; (b) weakly acidic; (c) neutral; (d) strongly basic

5. (a) weak acid; (b) weak acid; (c) strong acid; (d) strong acid

7. (a) Arrhenius acid; (b) salt; (c) Arrhenius base; (d) salt

9. (a) $HBr(aq)$, $NaOH(aq)$; (b) $H_2SO_4(aq)$, $Mg(OH)_2(aq)$

11. (a) $HF(aq)$, $NaOH(aq)$; (b) $HI(aq)$, $Mg(OH)_2\ (aq)$; (c) $HNO_3(aq)$, $Ca(OH)_2(aq)$; (d) $H_2CO_3(aq)$, $LiOH(aq)$

13. (a) $2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$

(b) $H_2CO_3(aq) + Ba(OH)_2(aq) \rightarrow BaCO_3(aq) + 2\ H_2O(l)$

15. (a) $HC_2H_3O_2(aq)$, $LiOH(aq)$; (b) $HCl(aq)$, $NH_3(aq)$

17. (a) $HI(aq)$, $H_2O(l)$; (b) $HC_2H_3O_2(aq)$, $HS^-(aq)$

19. (a) $HF(aq) + NaHS(aq) \rightarrow H_2S(aq) + NaF(aq)$

(b) $HNO_2(aq) + NaC_2H_3O_2(aq) \rightarrow NaNO_2(aq) + HC_2H_3O_2(aq)$

21. (a) colorless; (b) pink

23. (a) red; (b) yellow

25. (a) yellow; (b) blue

27. 0.198 M HNO_3

29. 0.141 M H_2SO_4

31. 0.142 M H_3PO_4

33. (a) 19.9% HCl; (b) 5.95% $HC_2H_3O_2$; (c) 3.12% HNO_3; (d) 24.9% H_2SO_4

35. 0.3992 M HCl

37. 0.390 M LiOH

39. 28.0 mL solution

41. 176 g/mol

43. (a) $H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$; (b) $K_w = [H^+][OH^-]$; (c) $K_w = 1.0 \times 10^{-14}$

45. (a) $[OH^-] = 4.0 \times 10^{-13}\ M$; (b) $[OH^-] = 5.9 \times 10^{-10}\ M$

47. (a) $[H^+] = 6.3 \times 10^{-12}\ M$; (b) $[H^+] = 3.4 \times 10^{-11}\ M$

49. (a) pH = 3; (b) pH = 5

51. (a) $[H^+] = 0.000001\ M$; (b) $[H^+] = 0.00000001\ M$

53. (a) pH = 6.52; (b) pH = 7.80

55. (a) $[H^+] = 2.8 \times 10^{-7}\ M$; (b) $[H^+] = 3.2 \times 10^{-8}\ M$

57. (a) pH = 13.04; (b) pH = 10.74

59. (a) $[OH^-] = 3.6 \times 10^{-10}\ M$; (b) $[OH^-] = 1.6 \times 10^{-9}\ M$

61. (a) highly ionized; (b) highly ionized; (c) highly ionized

63. (a) weak; (b) strong; (c) strong; (d) weak

65. (a) weak; (b) strong; (c) strong; (d) weak

67. (a) $HF(aq)$; (b) $H^+(aq)$ and $Br^-(aq)$; (c) $H^+(aq)$ and $NO_3^-(aq)$; (d) $HNO_2(aq)$

69. (a) $Ag^+(aq)$ and $F^-(aq)$; (b) $AgI(s)$; (c) $Hg_2Cl_2(s)$; (d) $Ni^{2+}(aq)$ and $2\ Cl^-(aq)$

71. (1) Complete and balance the nonionized equation.

(2) Convert the nonionized equation to the total ionic equation.

(3) Cancel spectator ions to obtain the net ionic equation.

(4) Check ($\surd$) each ion or atom on both sides of the equation.

73. (a) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(b) $HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

75. (a) $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

(b) $Ba^{2+}(aq) + CrO_4^{2-}(aq) \rightarrow BaCrO_4(s)$

77. yellow

79. orange (red + yellow)

81. NaH_2PO_4

83. pH = 0.30

85. $6 \times 10^{13}\ H^+$

87. $NH_4OH(aq) + HC_2H_3O_2(aq) \rightarrow NH_4^+(aq) + C_2H_3O_2^-(aq) + H_2O(l)$

89. 89.2 g/mol

Chapter 15

1. (a) unit analysis; (b) algebraic analysis

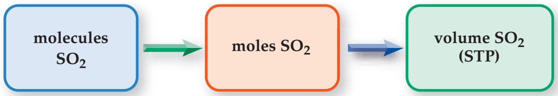
3. (a) meter (m); (b) gram (g); (c) liter (L);
(d) second (s); (e) degree Celsius (°C); (f) calorie (cal)

5. (a) length; (b) mass; (c) volume; (d) time;
(e) temperature; (f) heat energy

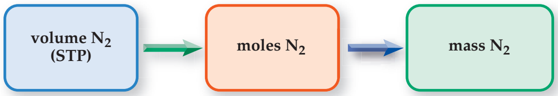
7. (a) velocity; (b) density; (c) specific heat;
(d) m/m% concentration

9. (a) estimated answer: ~2 g MgO; calculated answer: 2.036 g MgO;
(b) estimated answer: ~1 g Ag_2O; calculated answer: 1.235 g Ag_2O

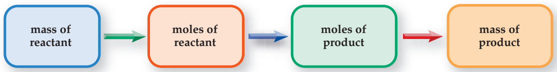
11. (a)



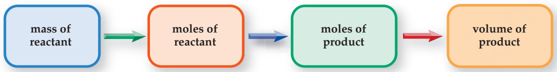
(b)



13. (a)



(b)



15. Refer to Figure 15.2.
(a) 0.518 g HCl (b) 0.318 L HCl
(c) 8.55 × 10^{21} molecules HCl (d) 0.299 *M* HCl

17. Refer to Figure 15.2.
(a) 1.12 L HF; (b) 3.01 × 10^{22} molecules HF; (c) 0.500 *M* HF

19. Refer to Figure 15.2.
(a) 1.96 g CO_2; (b) 2.69 × 10^{22} molecules CO_2;
(c) 0.0893 *M* H_2CO_3

21. Refer to Figure 15.2.
(a) 0.826 L H_2S; (b) 1.26 g H_2S; (c) 0.0819 *M* H_2S

23. Refer to Figure 15.3.
(a) 2.23 g CH_4; (b) 3.11 L CO_2

25. Refer to Figure 15.3.
(a) 0.580 L H_2; (b) 518 mL HCl solution

27. Refer to Figure 15.3.
(a) 0.719 *M* $AgNO_3$; (b) 1.88 g AgBr

29. (a) 91.6 g CO_2; (b) 46.6 L CO_2;
(c) 1st reaction + 2nd reaction:
33.3 g + 33.3 g = 66.6 g O_2

31. (a) 2500 g SO_3; (b) 698 L SO_3; (c) 1680 g H_2SO_4

33. (a) 6.30 g N_2; (b) 5.04 L N_2; (c) 5.41 g H_2O

35. 44.9 g/mol

37. 0.622 *M* Ca^{2+}; 1.24 *M* Cl^-

39. 44.6 mL solution

41. 3.64 × 10^{21} molecules N_2O; 0.266 g N_2O

43. $C_6H_{14}O_6$

45. 0.683 g $NaClO_3$

47. 2,000 gross

49. (a) SO_2 is the limiting reactant; (b) 65.0 mL NO_2

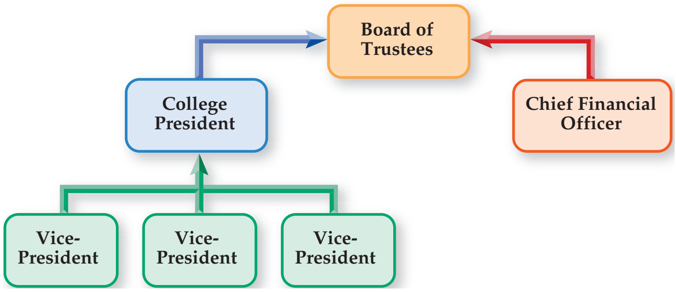
51. (a) H_2SO_4 is the limiting reactant; (b) 0.451 g H_2O

53. 3.16 × 10^9 kg seawater

55. 0.250 *M* Na^+

57. 0.784 g Na_2CO_3

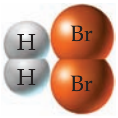
59. A concept map for the college administration is:



Chapter 16

1. (a), (b), and (c) decrease the rate of a reaction

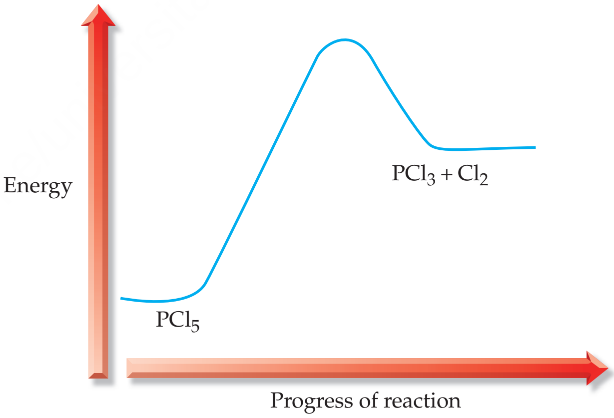
3.



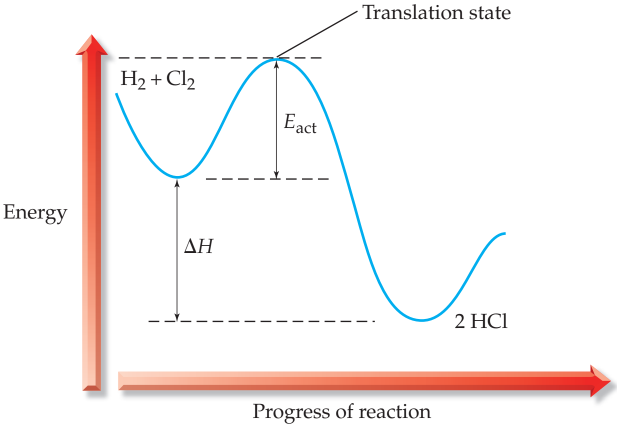
5. (a), (b), and (c) increase the rate of a reaction

7. No, a catalyst speeds up the rate of reaction, but does not affect the amount of product.

9. $PCl_5(g) + \textit{heat} \rightleftarrows PCl_3(g) + Cl_2(g)$



11. $H_2(g) + Cl_2(g) \rightleftarrows 2\,HCl(g) + \textit{heat}$



13. lowers E_{act}

15. (a) false; (b) true

17. (a) $K_{eq} = \dfrac{[HF]^2}{[H_2][F_2]}$

(b) $K_{eq} = \dfrac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7}$

(c) $K_{eq} = [CO_2]$

19. $K_{eq} = 1.10$

21. (a) NH_3 decreases; (b) NH_3 increases

23. (a) right; (b) left; (c) left; (d) right; (e) right; (f) left; (g) right; (h) left; (i) no shift; (j) no shift

25. (a) $K_i = \dfrac{[H^+][CHO_2^-]}{[HCHO_2]}$

(b) $K_i = \dfrac{[H^+][HC_2O_4^-]}{[H_2C_2O_4]}$

(c) $K_i = \dfrac{[H^+][H_2C_6H_5O_7^-]}{[H_3C_6H_5O_7]}$

27. $K_i = 4.5 \times 10^{-4}$

29. $K_i = 7.2 \times 10^{-4}$

31. (a) right; (b) left; (c) left; (d) right; (e) left; (f) left; (g) right; (h) right

33. (a) right; (b) left; (c) left; (d) right; (e) left; (f) no shift; (g) right; (h) right

35. (a) $K_{sp} = [Ag^+][I^-]$

(b) $K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

(c) $K_{sp} = [Ag^+]^3[PO_4^{3-}]$

37. $K_{sp} = 5.9 \times 10^{-21}$

39. (a) left; (b) left; (c) right; (d) right; (e) no shift; (f) left; (g) no shift; (h) right

41. (a) left; (b) left; (c) right; (d) right; (e) no shift; (f) left; (g) no shift; (h) right

43. $K_{eq} = 0.20$

45. (a) left; (b) right; (c) right; (d) left

47. $K_{sp} = 5.3 \times 10^{-6}$

Chapter 17

1. (a) 0; (b) 0; (c) 0; (d) 0

3. (a) +1; (b) +2; (c) −2; (d) −1

5. (a) −3; (b) +4; (c) −3; (d) +5

7. (a) +4; (b) +6; (c) −2; (d) +7

9. (a) reduction; (b) oxidizing agent

11. (a) oxidized—Co, reduced—S;

(b) oxidized—Cd, reduced—Cl_2

13. (a) oxidized—Al, reduced—Cr^{3+};

(b) oxidized—Cl^-, reduced—F_2

15. (a) oxidizing agent—CuO; reducing agent—H_2;

(b) oxidizing agent—PbO; reducing agent—CO

17. Yes, the total electron loss by oxidation must always equal the total electron gain by reduction.

19. (a) $Br_2(l) + 2\,NaI(aq) \rightarrow 2\,NaBr(aq) + I_2(s)$

(b) $2\,PbS(s) + 3\,O_2(g) \rightarrow 2\,PbO(s) + 2\,SO_2(g)$

21. (a) $2\,MnO_4^-(aq) + 10\,I^-(aq) + 16\,H^+(aq) \rightarrow 2\,Mn^{2+}(aq) + 5\,I_2(s) + 8\,H_2O(l)$

(b) $Cu(s) + SO_4^{2-}(aq) + 4\,H^+(aq) \rightarrow Cu^{2+}(aq) + SO_2(g) + 2\,H_2O(l)$

23. (a) $H_2O_2(aq) + 2\,H^+(aq) + 2\,e^- \rightarrow 2\,H_2O(l)$

(b) $AsO_3^{3-}(aq) \rightarrow AsO_3^-(aq) + 2\,e^-$

25. (a) $3\,Zn(s) + 2\,NO_3^-(aq) + 8\,H^+(aq) \rightarrow 3\,Zn^{2+}(aq) + 2\,NO(g) + 4\,H_2O(l)$

(b) $2\,Mn^{2+}(aq) + 5\,BiO_3^-(aq) + 14\,H^+(aq) \rightarrow 2\,MnO_4^-(aq) + 5\,Bi^{3+}(aq) + 7\,H_2O(l)$

27. $Cl_2(g) + H_2O(l) \rightarrow Cl^-(aq) + HOCl(aq) + H^+(aq)$

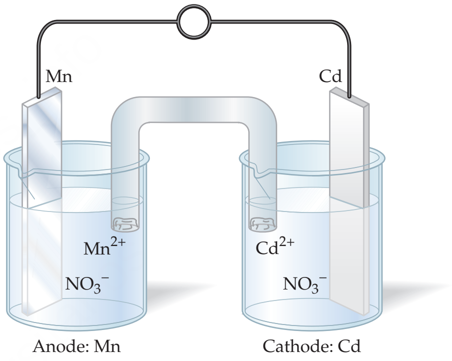
29. (a) Li(*s*); (b) Mg(*s*); (c) $I^-(aq)$; (d) Br^-(aq)

31. (a) Cr(*s*); (b) $H_2(g)$; (c) Cu(*s*); (d) $H_2(g)$

33. (a) nonspontaneous; (b) spontaneous

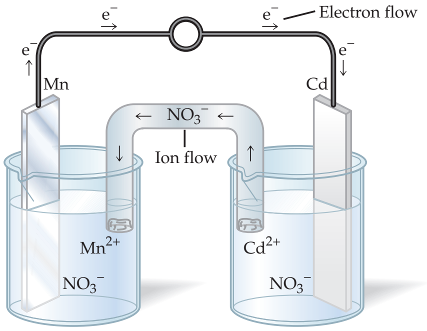
35. (a) spontaneous; (b) spontaneous

37.



$Mn(s) + Cd(NO_3)_2(aq) \rightarrow Cd(s) + Mn(NO_3)_2(aq)$

39.



$Mn(s) + Cd(NO_3)_2(aq) \rightarrow Cd(s) + Mn(NO_3)_2(aq)$

41. (a) $Ni(s) \rightarrow Ni^{2+}(aq) + 2\,e^-$;

(b) $Fe^{2+}(aq) + 2\,e^- \rightarrow Fe(s)$;

(c) Ni anode, Fe cathode;

(d) Ni anode → Fe cathode;

(e) Fe half-cell → Ni half-cell

43. (a) $2\,F^-(aq) \rightarrow F_2(g) + 2\,e^-$;

(b) $Cl_2(g) + 2\,e^- \rightarrow 2\,Cl^-(aq)$;

(c) Pt(F_2) anode, Pt(Cl_2) cathode;

(d) Pt(F_2) anode → Pt(Cl_2) cathode; cathode;

(e) Pt(F_2) half-cell → Pt(Cl_2) half-cell half-cell

45. +2.5

47. $Co(s) + Hg^{2+}(aq) \rightarrow Co^{2+}(aq) + Hg(l)$

49. $Zn(s) + 2\,H^{+}(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

51. Net ionic equation:
$3\,Cu(s) + 8\,H^{+}(aq) + 2\,NO_3^{-}(aq)$
$\rightarrow 3\,Cu^{2+}(aq) + 2NO(g) + 4\,H_2O(l)$

Chapter 18

1. alpha particle (α)

3. beta particle (β)

5. gamma ray (γ)

7. (a) ~4 amu; (b) ~0 amu; (c) 0 amu; (d) ~0 amu; (e) ~1 amu; (f) ~1 amu

9. (a) ${}^{238}_{92}\text{U} \rightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He}$;
(b) ${}^{28}_{13}\text{AI} \rightarrow {}^{28}_{14}\text{SI} + {}^{0}_{-1}\text{e}$;
(c) ${}^{15}_{8}\text{O} \rightarrow {}^{15}_{7}\text{N} + {}^{0}_{+1}\text{e}$;
(d) ${}^{55}_{26}\text{Fe} + {}^{0}_{-1}\text{e} + {}^{55}_{25}\text{Mn}$

11. (a) ${}^{221}_{88}\text{Ra}$; (b) ${}^{43}_{19}\text{K}$; (c) ${}^{73}_{36}\text{Kr}$; (d) ${}^{133}_{56}\text{Ba}$

13. ${}^{234}_{91}\text{Pa}$

15. ${}^{231}_{91}\text{Pa}$

17. (a) α; (b) β; (c) β; (d) α; (e) α; (f) α; (g) α; (h) β; (i) β; (j) α;

19. 25%

21. 11,500 years

23. 5 mg

25. radiocarbon dating

27. cobalt-60

29. technetium-99

31. ${}^{23}_{12}\text{Mg}$

33. ${}^{80}_{35}\text{Br}$

35. ${}^{260}_{105}\text{Db}$

37. ${}^{260}_{106}\text{Sg}$

39. 8 neutrons

41. The number of neutrons varies, but the *average* number is 2.4.

43. 2 neutrons

45. ${}^{233}_{92}\text{U}$

47. ${}^{1}_{0}\text{n}$

49. ${}^{1}_{1}\text{H}$

51. ${}^{2}_{1}\text{H}$

53. ~6000 years

55. ${}^{241}_{95}\text{Am}$

57. ${}^{1}_{0}\text{n}$

Chapter 19

1. alkanes

3. alkanes

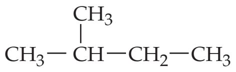
5. alkynes

7. (a) C_6H_{14} is an alkane; (b) C_6H_{12} is not an alkane

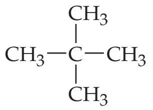
9. (a) butane; (b) hexane; (c) octane; (d) decane

11. Three isomers of pentane (C_5H_{12})

$CH_3{-}CH_2{-}CH_2{-}CH_2{-}CH_3$



$CH_3{-}\overset{\overset{\large CH_3}{|}}{CH}{-}CH_2{-}CH_3$



$CH_3{-}\underset{\underset{\large CH_3}{|}}{\overset{\overset{\large CH_3}{|}}{C}}{-}CH_3$

13. Two isomers of bromopropane (C_3H_7Br)
$CH_3{-}CH_2{-}CH_2{-}Br$ $CH_3{-}CH(Br){-}CH_3$

15. (a) CH_3^{-}; (b) $CH_3CH_2^{-}$

17. (a) 2-methylpentane; (b) 4-ethyl-2-methylhexane; (c) 2,4,4-trimethylheptane; (d) 3,4,5-trimethyloctane

19. (a) $CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$
(b) $C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$
(c) $2\,C_6H_{14} + 19\,O_2 \rightarrow 12\,CO_2 + 14\,H_2O$
(d) $2\,C_8H_{18} + 25\,O_2 \rightarrow 16\,CO_2 + 18\,H_2O$

21. (a) C_6H_{14} is not an alkene; (b) C_6H_{12} is an alkene

23. (a) 1-butene; (b) 1-pentene; (c) 2-hexene; (d) 4-octene

25. (a) C_6H_{12} is not an alkyne; (b) C_6H_{10} is an alkyne

27. (a) 2-pentyne; (b) 1-pentyne; (c) 3-hexyne; (d) 2-heptyne

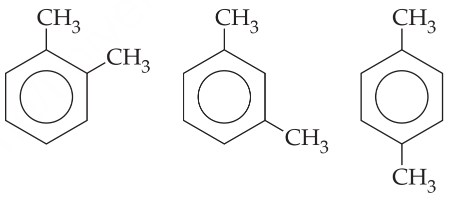
29. $CH_2{=}CH{-}CH_2{-}CH_2{-}CH_3$ and
$CH_3{-}CH{=}CH{-}CH_2{-}CH_3$

31. (a) 4-methyl-2-pentene; (b) 3,5-dimethyl-2-hexene

33. (a) 4,4-dimethyl-2-heptyne; (b) 3,4,5-trimethyl-1-octyne

35. (a) $CH_2{=}CH_2 + 3\,O_2 \rightarrow 2\,CO_2 + 2\,H_2O$
(b) $CH_3{-}CH{=}CH_2 + H_2 \rightarrow CH_3{-}CH_2{-}CH_3$

37.



ortho *meta* *para*

39. carbon dioxide and water

41. (a) ether; (b) organic halide; (c) phenol; (d) amine

43. (a) amine; (b) organic halide; (c) ester; (d) carboxylic acid; (e) aldehyde; (f) ether; (g) ketone; (h) amide

45. (a) "methyl iodide"; (b) "ethyl bromide"; (c) "propyl fluoride"; (d) "isopropyl chloride"

47. $CCl_2{\equiv}CHCl$

49. (a) 1-butanol; (b) 2-butanol; (c) 1-butanol; (d) 2-butanol

51. (a) phenol; (b) para-methylphenol

53. (a) "dimethyl ether"; (b) "methyl ethyl ether"; (c) "dipropyl ether"; (d) "phenyl ethyl ether"

55. "Ethyl alcohol" has a higher boiling point because it can hydrogen bond.

57. $CH_3{-}CH_2{-}OH$ and $CH_3{-}O{-}CH_3$

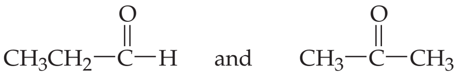
59. (a) "ethylamine"; (b) "propylamine"

61. "Ethylamine" has a higher boiling point because it can hydrogen bond.

63. (a) methanal; (b) ethanal

65. (a) 2-propanone; (b) 2-butanone

67.

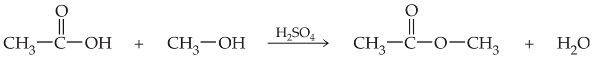


$CH_3CH_2-\overset{O}{\overset{\|}{C}}-H$ and $CH_3-\overset{O}{\overset{\|}{C}}-CH_3$

69. (a) methanoic acid; (b) ethanoic acid

71. (a) ethyl methanoate; (b) methyl ethanoate

73. (a)



$CH_3-\overset{O}{\overset{\|}{C}}-OH + CH_3-OH \xrightarrow{H_2SO_4} CH_3-\overset{O}{\overset{\|}{C}}-O-CH_3 + H_2O$

(b)

75. (a) methanamide ("formamide"); (b) benzamide

77. (a) saturated; (b) unsaturated

79. carboxylic acid

81. ketone

83. organic chloride, aldehyde

85. (a) CH_3CH_2-OH; (b) $CH_3CH_2CH_2-NH_2$

Chapter 20

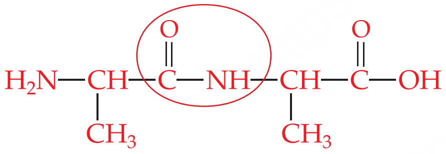
1. (a) protein; (b) carbohydrate

3. (a) peptide linkage; (b) glycoside linkage

5. covalent amide bond

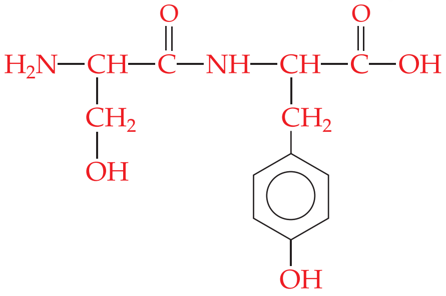
7. The primary structure of a protein refers to the amino acid sequence. The secondary structure of a protein refers to the shape of the polymer chain.

9.



$H_2N-\underset{CH_3}{CH}-\overset{O}{\overset{\|}{C}}-NH-\underset{CH_3}{CH}-\overset{O}{\overset{\|}{C}}-OH$

11.



13. (1) Ala–Gln–Met; (2) Ala–Met–Gln; (3) Gln–Ala–Me; (4) Gln–Met–Ala; (5) Met–Ala–Gln; (6) Met–Gln–Ala

15. glutamic acid, lysine, and serine

17. In oxytocin, *isoleucine* replaces phenyl alanine and *leucine* replaces arginine.

19. The key represents an *enzyme (E)*.

21. The "teeth" on the key represent an active site on the enzyme molecule.

23. An enzyme is likely to be inactive under experimental conditions where the pH or temperature is too high or too low.

25. A pentose sugar molecule has five carbon atoms.

27. An aldose sugar has an *aldehyde* group and one or more *alcohol* groups.

29. A *monosaccharide* is a simple sugar molecule such as glucose, whereas a *disaccharide* has two simple sugar molecules joined by a glycoside linkage.

31. *Glucose* and *fructose* are joined by a glycoside linkage in sucrose.

33. (a) fat; (b) oil; (c) fat; (d) fat and oil

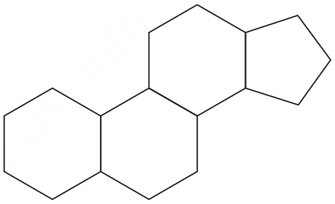
35.

Name of Fatty Acid	Structure of Fatty Acid
(a) palmitic acid	$HOOC-(CH_2)_{14}-CH_3$
(b) lauric acid	$HOOC-(CH_2)_{10}-CH_3$
(c) myristic acid	$HOOC-(CH_2)_{12}-CH_3$

37. A phospholipid has *two* fatty acids and phosphoric acid esterified to glycerol.

39. $CH_3-(CH_2)_{24}-COO^-\ Na^+$ and $CH_3-(CH_2)_{29}-OH$

41.



43. deoxyribose sugar, a nitrogen base, and phosphoric acid

45. adenine (A), cytosine (C), guanine (G), and thymine (T)

47. 2 strands of nucleotides

49. 2 hydrogen bonds

51. thymine (T)

53. uracil (U)

55. long and extended

57. When a protein is denatured in aqueous acid, hydrogen bonds are broken and the molecule assumes a random shape.

59. A, D, E, and K

61. ACG, AGC, CAG, CGA, GAC, and GCA

63. *Replication* is the process by which a DNA molecule synthesizes an identical molecule of DNA; whereas, *transcription* is the process by which a DNA molecule synthesizes a complementary strand of RNA.

65. The first step in *replication* is for a DNA molecule to unwind by breaking the hydrogen bonds (A═T and C≡G) that hold the double helix together.

In the second step, each strand of DNA acts as a template to synthesize a complementary strand of DNA. That is, a nucleotide in the template strand that contains the base adenine (A) codes for thymine (T) on the growing strand; cytosine (C) codes for guanine (G); guanine (G) codes for cytosine (C); and thymine (T) codes for adenine (A).

At the conclusion, each pair of template strands and complementary strands are joined by hydrogen bonds (A═T and C≡G) and the two resulting structures are exact duplicates of the original DNA molecule. Thus, two identical molecules of DNA are produced from one molecule of DNA.

APPENDIX

J

Answers to Chapter Self-Tests

Chapter 1 — Self-Test

1. (c) **2.** (b) **3.** (e)

Key Concept

4. The image of a cube creates a dual perception. By mentally pulling the red dot, the front face of the cube appears to be down and to the right. By mentally pushing the red dot, the front face of the cube appears to be up and to the left.

Critical Thinking

5. Chemical evolution is a *scientific theory* because it is not measurable, and the relationship between elements is not expressed as a mathematical equation.

6. Biological evolution is a *scientific theory* because it is not measurable, and the relationship between species is not expressed as a mathematical equation.

Prerequisite Science Skills — Self-Test

1. (d) **2.** (b) **3.** (e) **4.** (b) **5.** (c) **6.** (a) **7.** (c)

Key Concepts

8. Round off the answer in the display after the final calculation. Rounding off the final answer is more accurate, and convenient.

Critical Thinking

9. Note the exponent for the mass of an alpha and beta particle. Since 10^{-24} g is four powers of 10 greater than 10^{-28} g, the alpha particle is much heavier.

10. Precious metals such as silver and gold are weighed in units of troy ounces, whereas base metals such as aluminum and iron are weighed in avoirdupois ounces. However, a gram is a constant unit of mass in the metric system. Thus, an ounce of silver (31.1 g) weighs more than an ounce of aluminum (28.4 g).

Chapter 2 — Self-Test

1. (e) **2.** (a) **3.** (b) **4.** (c) **5.** (d) **6.** (e) **7.** (a) **8.** (c) **9.** (b) **10.** (d)

Key Concepts

11. Fool's gold ($d = 5.00$ g/mL) sinks in water, but rests on mercury. The gold nugget ($d = 18.3$ g/mL) sinks in both water and liquid mercury ($d = 13.6$ g/mL), and comes to rest at the bottom of the glass cylinder.

12. Since the absolute scale begins at zero Kelvin, the coldest possible temperature is 0 K; for comparison, 0 °F is 523 K, and 0 °C is 273 K.

Critical Thinking

13. The mass of a silver dime is greater than a zinc penny because silver metal is more dense than zinc metal.

14. Gasoline and chloroform are insoluble in water. Add a few drops of each liquid to water and observe if they sink or float. Gasoline ($d = 0.70$ g/mL) is *less dense* than water ($d = 1.00$ g/mL), so it floats. Chloroform ($d = 1.5$ g/mL) is *more dense* than water, so it sinks.

Chapter 3 — Self-Test

1. (d) **2.** (a) **3.** (e) **4.** (c) **5.** (d) **6.** (b) **7.** (c) **8.** (a) **9.** (b) **10.** (d)

Key Concepts

11. (a) liquid; (b) gas; (c) solid

12. (a) heterogeneous mixture; (b) compound; (c) element; (d) homogeneous mixture

13. The photo shows a mineral and illustrates a heterogeneous mixture. The model illustrates a molecule of S_8, which is an element.

14. The molecular formula for aspirin is $C_9H_8O_4$.

Critical Thinking

15. Ice floating in water is an example of a heterogeneous mixture because ice and water have different densities, and can be separated by physical methods.

16. The properties of an element (Co) and its compounds (CoO, Co_2O_3, $Co(NO_3)_3$) are not related, and their physical and chemical properties are different.

17. When a playground swing is at its lowest point, the potential energy is minimum, and kinetic energy is maximum.

18. On a roller coaster ride the potential energy is greatest when the coaster is at its highest point. As the coaster travels down, it loses potential energy and gains kinetic energy. Thus, the potential energy is maximum, and the kinetic energy is minimum, when the roller coaster is at its highest point.

19. The conservation of mass law applies to ordinary chemical reactions. For nuclear reactions, the energy change is so large that there is a measurable loss of mass. That is, the Sun is a huge nuclear reactor that converts hydrogen into helium while losing some mass, which is converted to energy.

20. The conservation of mass and energy law explains the mass difference for the conversion of two hydrogen atoms to a helium atom.

Chapter 4 — Self-Test

1. (d) **2.** (a) **3.** (d) **4.** (b) **5.** (a) **6.** (d) **7.** (c) **8.** (a) **9.** (e) **10.** (b) **11.** (d)

Key Concepts

12. In the analogy ”like a marble in Dallas Cowboys Stadium” the marble represents the atomic nucleus, and the stadium represents the entire atom.

13. In the analogy, the missile represents an alpha particle and the planets in our solar system represent atomic nuclei.

14. An *electron* is to electricity, as a *photon* is to light.

15. The evidence for electrons occupying fixed energy states is a line spectrum; that is, the ”atomic fingerprint” of an element.

Critical Thinking

16. If the two isotopes, Cu-63 and Cu-65, were each 50% abundant, the atomic mass would be approximately 64 amu. Since the atomic mass of copper is 63.55, the lighter isotope, Cu-63, is more abundant.

17. No, atoms of different elements have different atomic numbers, and cannot have the same number of protons.

18. Yes, atoms of different elements can have the same mass number, and the same number of neutrons.

Chapter 5 — Self-Test

1. (e) **2.** (a) **3.** (c) **4.** (c) **5.** (c) **6.** (a) **7.** (d) **8.** (e) **9.** (b) **10.** (c)

Key Concepts

11. The metallic character *decreases* up a group of elements, and generally *decreases* from left to right in a given period.

12. The atomic radius *decreases* up a group of elements, and generally *decreases* from left to right in a given period.

13. The ionization energy *increases* up a group of elements, and generally *increases* from left to right in a given period.

14. Elements in the same group react to form compounds with similar chemical formulas; for example, P and As are in Group VA/15 and react with oxygen to form P_2O_5 and As_2O_5, respectively.

Critical Thinking

15. *Yes*, ions of different elements can have the same number of electrons. For example, Na^+ and Mg^{2+} each have 10 electrons.

16. *No*, ions of different elements have a different number of protons.

17. *Yes*, ions of different elements can have the same number of neutrons. For example, O^{2-} and F^- each have isotopes with 10 neutrons (O-18 and F-19).

18. The yellow substance appears to be sulfur powder. The gray substance appears metallic, so it must be the semimetal antimony. By elimination, the orange substance must be antimony sulfide, Sb_2S_3.

Chapter 6 — Self-Test

1. (c) **2.** (b) **3.** (a) **4.** (c) **5.** (d) **6.** (d) **7.** (b) **8.** (a) **9.** (b) **10.** (d)

Key Concepts

11. NaCl is a binary ionic compound; $NaClO_3$ is a ternary ionic compound; HCl is a binary molecular compound; HCl(*aq*) is a binary acid; $HClO_3$(*aq*) is a ternary oxyacid

12. NaCl is named using an −*ide* suffix; that is, sodium chloride.

13. $NaClO_3$ is named using an −*ate* suffix; that is, sodium chlorate.

14. $NaClO_2$ is named using an −*ite* suffix; that is, sodium chlorite.

15. HCl(*aq*) is named as a *hydro* + *nonmetal stem* + *ic* acid; that is, hydrochloric acid.

16. $HClO_3$(*aq*) is named *nonmetal stem* + *ic* acid; that is, chloric acid.

17. $HClO_2$(*aq*) is named *nonmetal stem* + *ous* acid; that is, chlorous acid.

Critical Thinking

18. Copper predictably loses two 4*s* electrons giving Cu^{2+}. Copper can also lose one 4*s* electron giving Cu^+. This ion is especially stable as the remaining 4*s* electron can move into the 3*d* sublevel, thus filling the third energy level.

19. Predictably, iron loses two 4*s* electrons giving Fe^{2+}. Iron can lose an additional 3*d* electron giving Fe^{3+}. This ion is especially stable as the 3*d* sublevel is half-filled. A half-filled 3*d* sublevel has one electron in each of the five 3*d* orbitals.

20. White chalk is made by grinding limestone into a fine powder and pressing the powder into chalk. Marble is a crystalline form of the mineral found naturally.

Chapter 7 — Self-Test

1. (d) **2.** (a) **3.** (a) **4.** (a) **5.** (a) **6.** (d) **7.** (c) **8.** (c) **9.** (b) **10.** (e) **11.** (d)

Key Concepts

12. $4\,Fe(s) \quad + \quad 3\,O_2(g) \quad \rightarrow \quad 2\,Fe_2O_3(s)$

13. $Ca(HCO_3)_2(s) \quad \xrightarrow{\Delta} \quad CaCO_3(s) \quad + \quad H_2O(g) \quad + \quad CO_2(g)$

14. $Zn(s) \quad + \quad Ni(NO_3)2(aq) \quad \rightarrow \quad Zn(NO_3)_2(aq) \quad + \quad Ni(s)$

15. $HNO_3(aq) \quad + \quad NH_4OH(aq) \quad \rightarrow \quad NH_4NO_3(aq) \quad + \quad HOH(l)$

16. $C_2H_5OH(l) \quad + \quad 3\,O_2(g) \quad \xrightarrow{\text{spark}} \quad 2\,CO_2(g) \quad + \quad 3\,H_2O(g)$

Critical Thinking

17. $HC_2H_3O_2(aq) \quad + \quad NaHCO_3(aq) \quad \rightarrow \quad NaC_2H_3O_2\,(aq) \quad + \quad H_2O(l) \quad + \quad CO_2(g)$

18. $Pb(C_2H_3O_2)_2(aq) \quad + \quad 2\,NaI(aq) \quad \rightarrow \quad PbI_2(s) \quad + \quad 2\,NaC_2H_3O_2\,(aq)$

Chapter 8 — Self-Test

1. (a) **2.** (c) **3.** (b) **4.** (c) **5.** (e) **6.** (b) **7.** (c) **8.** (b) **9.** (d)

Key Concepts

10. The mass is found in the periodic table. One atom of aluminum has a mass of 26.98 amu, and Avogadro's number of atoms has a mass of 26.98 g.

11. The mass is found in the periodic table. One atom of iodine has a mass of 126.90 amu, and Avogadro's number of atoms has a mass of 126.90 g.

12. Avogadro's number of steel balls has an estimated mass equal to Earth.

13. Avogadro's number of small marbles has a volume equal to the Moon.

Critical Thinking

14. No. Ten trillion (1×10^{13}) copper atoms weigh 1×10^{-9} g, which is much less than a mass of 1 1 μg(1×10^{-6} g).

15. A 1-carat diamond has about 1×10^{22} carbon atoms. A googol (1×10^{100}) of atoms is many more than the number of atoms in a 1-carat diamond.

Chapter 9 — Self-Test

1. (b) **2.** (c) **3.** (c) **4.** (c) **5.** (d) **6.** (a) **7.** (b) **8.** (b) **9.** (c)

Key Concepts

10. A balanced chemical equation is necessary before solving a stoichiometry problem.

11. Typically, *three* unit factors are required to solve a mass-mass problem.

12. Typically, *three* unit factors are required to solve a mass-volume problem.

13. Typically, only *one* unit factor is required to solve a volume–volume problem because the coefficients of the balanced equation are in the same ratio as the volumes of gaseous reactants and products.

Critical Thinking

14. The limiting reactant is gasoline; the excess reactant is oxygen in air.

15. (a) 2 Ag_2S, 2 excess S atoms (Ag is limiting reactant)
(b) 2 Ag_2S, 2 excess Ag atoms (S is limiting reactant)

Chapter 10 — Self-Test

1. (c) **2.** (e) **3.** (c) **4.** (d) **5.** (b) **6.** (d) **7.** (c) **8.** (a) **9.** (b) **10.** (e) **11.** (a)

Key Concepts

12. As the steam cools inside the capped can, it condenses to a liquid, thus reducing gaseous molecules and the pressure inside the can. The pressure outside the can is greater than inside, and atmospheric pressure crushes the can.

13. You would be repulsed first because H_2S (34 amu) has less molecular mass than $C_3H_6O_2$ (74 amu), so the molecules of H_2S move faster.

14. If you check tire pressure after driving a few miles, the pressure increases because road friction has heated the tire and the air molecules inside, which are moving faster and colliding more frequently.

15. At the top of the mountain the atmospheric pressure is less than the pressure inside the bottle, so the plastic bottle bulges.

Critical Thinking

16. The balloon rises because the density of air inside the balloon is less than outside. Heating air in a hot-air balloon causes air molecules to move faster and escape so as to equalize pressure inside and outside the balloon.

17. If you release a helium balloon into the air, the volume of the balloon increases as it ascends because atmospheric pressure outside the balloon is less at higher elevations.

18. The water bottle bulges on a hot day because the vapor pressure of water in the bottle increases and the pressure inside the bottle is greater than outside.

19. The water bottle at a colder temperature in a refrigerator collapses because the vapor pressure of water inside the bottle is less than outside.

Chapter 11 — Self-Test

1. (a) **2.** (c) **3.** (d) **4.** (c) **5.** (b) **6.** (e) **7.** (c) **8.** (e) **9.** (b) **10.** (b)

Key Concepts

11. A beehive surrounded by flying honey bees is analogous to the gaseous physical state. If honey bees move about faster near a fire, their movement is analogous to gas molecules when a gas is heated.

12. As the water freezes to a solid, water molecules will form hydrogen bonds and the volume will expand. Thus, the cap will pop off, or the bottle will break.

Critical Thinking

13. In Los Angeles, water boils at 100 °C. In Denver at a mile-high elevation, water boils at 95 °C because the atmospheric pressure is lower. Eggs cook slower at lower temperatures, and it takes longer to hard boil an egg.

14. Water has a concave lens because there is *attraction* between water and glass. Mercury has a convex lens because there is *repulsion* between mercury and glass.

Chapter 12 — Self-Test

1. (a) **2.** (d) **3.** (e) **4.** (d) **5.** (b) **6.** (e) **7.** (c) **8.** (b) **9.** (c) **10.** (a)

Key Concepts

11. The three diagrams represent (a) atoms, (b) molecules, and (c) formula units; for example, (a) Na atoms, (b) H_2O molecules, and (c) NaCl formula units.

12. The HF molecule is most polar because F is the most electronegative element in Group VIIA/17.

13. A water molecule has two pairs of nonbonding electrons on the oxygen atom, and two hydrogen atoms. Thus, *four* H_2O molecules can attach directly to a single H_2O molecule by hydrogen bonds (see Figure 11.12).

14. The electron pair geometry of $\mathbf{C}HCl_3$ is *tetrahedral,* and the molecular shape is *tetrahedral.*

Critical Thinking

15. Sulfur dioxide has two resonance structures: $O{=}\mathbf{S}{-}O$ and $O{-}\mathbf{S}{=}O$. A double bond is stronger than a single bond, but due to resonance, the two bonds are equivalent, and each has the same bond energy and bond length.

16. Sulfur trioxide has three resonance structures; one double bond and two single bonds. A double bond is stronger than a single bond, but due to resonance, the three bonds are equivalent, and each has the same bond energy and bond length.

17. The electron pair geometry in both H_2O and H_3O^+ is tetrahedral. However, in H_2O the molecular shape is *bent,* and in H_3O^+ the molecular shape is *trigonal pyramidal.*

18. The electron pair geometry in both NH_3 and NH_4^+ is tetrahedral. However, in NH_3 the molecular shape is *trigonal pyramidal,* and in NH_4^+ the molecular shape is *tetrahedral.*

Chapter 13 — Self-Test

1. (a) **2.** (c) **3.** (b) **4.** (d) **5.** (d) **6.** (c) **7.** (b) **8.** (b) **9.** (c) **10.** (d) **11.** (c)

Key Concepts

12. Bubbles form on the inside surface of a pan when the water is heated because water contains air and the solubility of a gas in water is less as the temperature increases.

13. Peanut butter can dissolve grease in clothing because peanut butter and grease are both nonpolar.

14. Powdered sugar dissolves faster because the solute particles are smaller than granulated sugar, and thus greater surface area to be attacked by solvent water molecules.

Critical Thinking

15. Lake Green has more dissolved oxygen gas. Lake Nuclear cools a nuclear power plant and the heat raises the temperature of the water. At a higher temperature (Lake Nuclear), the solubility of oxygen gas is less.

16. At the conceptual level for NaCl, there is less attraction between Na^+ and Cl^- ions than the ion attraction for water molecules. Conversely, with AgCl there is more attraction between Ag^+ and Cl^- ions than the ion attraction for solvent water molecules.

Chapter 14 — Self-Test

1. (e) **2.** (d) **3.** (b) **4.** (d) **5.** (a) **6.** (d) **7.** (e) **8.** (c) **9.** (b) **10.** (c) **11.** (a)

Key Concepts

12. We must always start with a balanced equation before applying unit analysis to solve an acid-base titration problem.

13. The bright light indicates Beaker X contains a strong electrolyte such as H_2SO_4, which is highly ionized. The dim light indicates Beaker Y has a weak electrolyte such as H_2SO_3, which is slightly ionized.

Critical Thinking

14. A "fishy smell" due to a basic organic compound can be neutralized using an acid. Household vinegar contains acetic acid, which can neutralize the fish odor.

15. Since phenolphthalein is colorless in acid, and bromthymol blue is yellow in acid, the color of the solution in the test tube will be yellow.

16. At the molecular level, it requires less energy to break H—Cl bonds than to break H—F bonds. In addition, more energy is released by the solvent water molecules surrounding H^+ and Cl^- than H^+ and F^-.

Chapter 15 — Self-Test

1. (b) **2.** (d) **3.** (e) **4.** (c) **5.** (a) **6.** (b) **7.** (c) **8.** (d) **9.** (e) **10.** (b)

Key Concepts

11. The *unit analysis method* converts a given value to an unknown value by applying one or more unit factors.

12. The *algebraic method* requires rearranging variables in an equation to calculate an unknown quantity.

13. *Concept mapping* traces the relationship of chemical concepts by connected boxes.

14. *Visualization* helps to understand concepts that cannot be observed directly.

Critical Thinking

15. The given value must have compound units, for example, g/mL.

16. a. The molar mass of a substance can be found by summing the masses of each element in the periodic table; or, by using an online reference.

b. The physical properties of a substance can be found online, or, by using a reference book, such as the *Handbook of Chemistry and Physics.*

Chapter 16 - Self-Test

1. (d) **2.** (a) **3.** (b) **4.** (a) **5.** (e) **6.** (b) **7.** (d) **8.** (e) **9.** (d)

Key Concepts

10. The reaction profile shows products at a higher energy level than reactants; thus, the profile represents an endothermic reaction.

11. $K_{eq} = [O_3]^2/[O_2]^3$

12. After heating an endothermic reaction, the equilibrium shifts to the right.

13. After reducing the volume of a reaction in which products have fewer molecules than reactants, the equilibrium shifts to products on the right.

14. After adding a catalyst, the rate increases, but there is *no shift* in equilibrium.

Critical Thinking

15. In a coal mine, small particles of coal dust in the air provide a large surface area for reaction. In a charcoal barbecue, the briquettes have less surface area to contact oxygen in air.

16. At equilibrium, the rate of evaporation of the liquid equals the rate of condensation of the vapor.

Chapter 17 — Self-Test

1. (a) **2.** (a) **3.** (b) **4.** (d) **5.** (b) **6.** (e) **7.** (e)

Key Concepts

8. Given the redox reaction:

$H_2O_2(aq) \quad + \quad NO_2^-(aq) \quad \rightarrow \quad H_2O(l) \quad + \quad NO_3^-(aq)$

a. $H_2O(l) + NO_2^-(aq) \rightarrow NO_3^-(aq) + 2\,H^+(aq) + 2\,e^-$
b. $2\,e^- \quad + \quad 2\,H^+(aq) \quad + \quad H_2\,O_2(aq) \quad \rightarrow \quad 2H_2O(l)$
c. $NO_2^-(aq)$
d. $H_2O_2(aq)$
e. $H_2O_2(aq)$
f. $NO_2^-(aq)$

9. Given the voltaic cell shown:

a. $Mn(s) \quad \rightarrow \quad Mn^{2+}(aq) \quad + \quad 2\,e^-$
b. $Fe^{2+}(aq) \quad + \quad 2\,e^- \quad \rightarrow \quad Fe(s)$
c. anode: Mn; cathode: Fe
d. Electrons flow from Mn anode to Fe cathode.
e. SO_4^{2-} ions flow from Fe compartment to Mn compartment.

Critical Thinking

10. As a NiCad battery is *discharging,* Cd is being *oxidized.*

11. As a NiCad battery is *recharging,* $Ni(OH)_2$ is being *oxidized.*

12. Zinc metal is undergoing *oxidation,* while silver oxide is undergoing *reduction.*

Chapter 18 — Self-Test

1. (c) **2.** (d) **3.** (b) **4.** (d) **5.** (d) **6.** (c) **7.** (b) **8.** (c) **9.** (d) **10.** (e)

Key Concepts

11. (a) ${}^{238}_{92}\text{U} \quad \rightarrow \quad {}^{234}_{92}\text{U} \quad + \quad {}^{4}_{2}\text{He}$

(b) ${}^{239}_{92}\text{U} \quad \rightarrow \quad {}^{239}_{93}\text{Np} \quad + \quad {}^{0}_{-1}\text{e}$

(c) ${}^{40}_{19}\text{K} \quad \rightarrow \quad {}^{40}_{18}\text{Ar} \quad + \quad {}^{0}_{+1}\text{e}$

(d) ${}^{40}_{19}\text{K} \quad + {}^{0}_{-1}\text{e} \quad \rightarrow \quad {}^{40}_{18}\text{Ar}$

12. $1\,n^0 \quad \rightarrow \quad 3\,n^0 \quad \rightarrow \quad 9\,n^0$

13. positron, ${}^{0}_{+1}\text{e}$

Critical Thinking

14. ${}^{210}_{82}\text{Pb} \quad \rightarrow \quad {}^{210}_{83}\text{Bi} \quad + \quad {}^{0}_{-1}\text{e}$

15. ${}^{249}_{97}\text{Bk} \quad + \quad {}^{48}_{20}\text{Ca} \quad \rightarrow \quad {}^{294}_{117}\text{Ts} \quad + \quad 3\,{}^{1}_{0}\text{n}$

Chapter 19 — Self-Test

1. (a) **2.** (b) **3.** (a) **4.** (b) **5.** (e) **6.** (c) **7.** (b) **8.** (d) **9.** (e) **10.** (d)

Key Concepts

11. alcohol: $CH_3—CH_2—OH$; ether: $CH_3—O—CH_3$

12. aldehyde: $CH_3—CH_2—CHO$; ketone: $CH_3—CO—CH_3$

13. carboxylic acid: $CH_3—CO—OH$; ester: $H—CO—OCH_3$

14. Thyroxine contains the following functional groups (left to right): phenol, organic halide, ether, amine, carboxylic acid.

Critical Thinking

15. a. $CH_3-CH(CH_3)-CH_2-C(CH_3)_2-CH_3$
b. $CH_3-CH(OH)-CH_2-CH_3$

16. Caffeine is an *amine;* the *−ine* suffix suggests an amine.

17. Cholesterol is an *alcohol;* the *−ol* suffix suggests an alcohol.

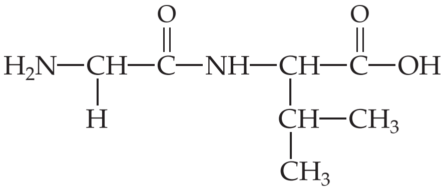
18. Cortisone is a *ketone;* the *−one* suffix suggests a ketone.

Chapter 20 — Self-Test

1. (a) **2.** (c) **3.** (c) **4.** (a) **5.** (a) **6.** (b) **7.** (d) **8.** (e) **9.** (d) **10.** (e)

Key Concepts

11.



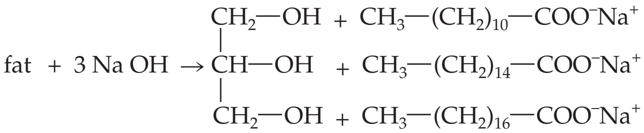
$$\mathrm{H_2N{-}\underset{\substack{|\\ H}}{CH}{-}\overset{\substack{O\\ \|}}{C}{-}NH{-}\underset{\substack{|\\ CH{-}CH_3\\ |\\ CH_3}}{CH}{-}\overset{\substack{O\\ \|}}{C}{-}OH}$$

12. $$\underset{\text{sucrose}}{C_{12}H_{22}O_{11}} + \underset{\text{water}}{H_2O} \xrightarrow{H^+} \underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}$$

13. A *lipid* is a water-insoluble fat, oil, or wax.

14. A *nucleic acid* is a polymer of a sugar, a base, and phosphoric acid.

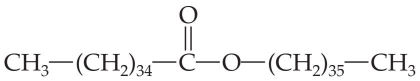
15.



$$\text{fat} + 3\,\mathrm{Na\,OH} \rightarrow \begin{array}{l} CH_2-OH \\ | \\ CH-OH \\ | \\ CH_2-OH \end{array} \begin{array}{l} + CH_3-(CH_2)_{10}-COO^-Na^+ \\ \\ + CH_3-(CH_2)_{14}-COO^-Na^+ \\ \\ + CH_3-(CH_2)_{16}-COO^-Na^+ \end{array}$$

The products are glycerol and soap. The soap is composed of the sodium salts of lauric acid, palmitic acid, and stearic acid.

16.



$$CH_3-(CH_2)_{34}-\overset{\substack{O\\ \|}}{C}-O-(CH_2)_{35}-CH_3$$

17. During DNA replication, an adenine base will code for *thymine.*

18. During DNA transcription, an adenine base will code for *uracil.*

Critical Thinking

19. a. Nucleotide A shows the sugar deoxyribose and the base adenine; thus, nucleotide A is found in DNA.
b. Nucleotide B shows the sugar ribose and the base cytosine; thus, nucleotide B is found in RNA.

Glossary

A

absolute zero The theoretical temperature at which the kinetic energy of a gas is zero. *(Sec. 10.10)*

acid A substance that releases hydrogen ions (H^+) when dissolved in water. *(Sec. 7.11)*

acid–base indicator A chemical substance that changes color according to the pH of the solution. *(Sec. 14.4)*

acidic oxide See *nonmetal oxide.*

actinide series The elements with atomic numbers 90–103. *(Sec. 5.3)*

activation energy **(E_{act})** The energy necessary for reactants to achieve the transition state and form products. *(Sec. 16.2)*

active metal A metal that is sufficiently active to react with water at room temperature; the active metals include Li, Na, K, Ca, Sr, and Ba. *(Sec. 7.7)*

activity The number of radioactive nuclei in a given sample that decay per unit time; for example, 500 disintegrations per minute (dpm). *(Sec. 18.4)*

activity series A relative order of metals arranged according to their ability to undergo a reaction; also called an *electromotive series*. *(Sec. 7.7)*

actual yield The amount of product that is experimentally obtained from given amounts of reactants. *(Sec. 9.9)*

alchemy A pseudoscience that attempted to convert a base metal, such as lead, into gold; a medieval science that sought to discover a universal cure for disease and a magic potion for immortality. *(Sec. 1.2)*

alkali metals The Group IA/1 elements, excluding hydrogen. *(Sec. 5.3)*

alkaline earth metals The Group IIA/2 elements. *(Sec. 5.3)*

alkanes A family of saturated hydrocarbon compounds. *(Sec. 19.2)*

alkenes A family of unsaturated hydrocarbon compounds having a double bond. *(Sec. 19.3)*

alkyl group **(R—)** The part that remains after a hydrogen atom is removed from an alkane. *(Sec. 19.2)*

alkynes A family of unsaturated hydrocarbon compounds having a triple bond. *(Sec. 19.3)*

alloy A homogeneous mixture of two or more metals. *(Sec. 3.2)*

alpha particle (α)A particle in an alpha ray that is identical to a helium-4 nucleus. *(Sec. 18.1)*

amino acid A carboxylic acid with an amine group on the alpha carbon; also referred to as an α-amino acid. *(Sec. 20.2)*

amphiprotic A substance capable of either donating or accepting a proton in an acid–base reaction. *(Sec. 14.3)*

anhydrous compound A compound that does not contain water. *(Sec. 11.10)*

anion A negatively charged ion. *(Sec. 6.1)*

anode The electrode in an electrochemical cell at which oxidation occurs. *(Sec. 17.6)*

aqueous solution (*aq*) A homogeneous mixture of a substance dissolved in water. *(Secs. 6.1 and 7.1)*

arenes A family of aromatic hydrocarbon compounds. *(Sec. 19.4)*

aromatic hydrocarbon A hydrocarbon compound that typically contains a benzene ring. *(Sec. 19.1)*

Arrhenius acid A substance that releases hydrogen ions when dissolved in water. *(Sec. 14.2)*

Arrhenius base A substance that releases hydroxide ions when dissolved in water. *(Sec. 14.2)*

aryl group **(Ar—)** The part that remains after a hydrogen atom is removed from an arene. *(Sec. 19.2)*

atmospheric pressure The pressure exerted by the air molecules in Earth's atmosphere. The atmospheric pressure at sea level is 1 atm or 760 mm of Hg. *(Sec. 10.2)*

atom The smallest particle that represents an element. *(Sec. 3.3)*

atomic mass The weighted average mass of all the naturally occurring isotopes of an element. *(Sec. 4.5)*

atomic mass unit (amu) A unit of mass exactly equal to 1/12 the mass of a carbon-12 atom. *(Sec. 4.5)*

atomic notation A symbolic method for expressing the composition of an atomic nucleus; the mass number and atomic number are indicated to the left of the chemical symbol for the element. *(Sec. 4.4)*

atomic nucleus A region of very high density in the center of the atom that contains protons and neutrons. *(Sec. 4.3)*

atomic number (*Z*) A number that identifies a specific element. *(Sec. 3.4)* A value that indicates the number of protons in the nucleus of an atom. *(Secs. 4.4 and 18.2)*

Avogadro's number (*N*) The value that corresponds to the number of carbon atoms in 12.01 g of carbon; 6.02×10^{23} particles. *(Sec. 8.1)*

Avogadro's theory The principle that equal volumes of gases, at the same temperature and pressure, contain equal numbers of molecules; also referred to as *Avogadro's law* or *Avogadro's hypothesis*. *(Secs. 8.5, 9.1, and 10.3)*

B

barometer An instrument for measuring atmospheric pressure. *(Sec. 10.2)*

base A substance that releases hydroxide ions (OH^-) when dissolved in water. *(Sec. 7.11)*

basic oxide See *metal oxide.*

battery A general term for any electrochemical cell that produces electrical energy spontaneously. *(Sec. 17.6)*

beta particle **(β)** A particle in a beta ray that is identical to an electron. *(Sec. 18.1)*

binary acid A compound that contains hydrogen and a nonmetal dissolved in water. *(Sec. 6.1)*

binary ionic compound A compound that contains one metal and one nonmetal. *(Sec. 6.1)*

binary molecular compound A compound that contains two nonmetals. *(Sec. 6.1)*

biochemistry The study of chemical substances derived from plants and animals. *(Sec. 1.2)* The study of biological compounds. *(Sec. 20.1)*

Bohr atom A model of the atom that describes an electron circling the atomic nucleus in an orbit of specific energy. *(Sec. 4.8)*

boiling point (Bp) The temperature at which the vapor pressure of a liquid is equal to the atmospheric pressure. *(Sec. 11.3)*

bond angle The angle formed by two atoms bonded to the central atom in a molecule. *(Secs. 11.7 and 12.10)*

bond energy The amount of energy required to break a covalent bond between two atoms. *(Sec. 12.3)*

bond length The distance between the nuclei of two atoms that are joined by a covalent bond. *(Sec. 12.3)*

bonding electrons The valence electrons in a molecule that are shared between two atoms. *(Sec. 12.4)*

Boyle's law The statement that the pressure and the volume of a gas are inversely proportional at constant temperature; that is, $P_1V_1 = P_2V_2$. *(Sec. 10.4)*

Brønsted-Lowry acid A substance that donates a proton in an acid–base reaction. *(Sec. 14.3)*

Brønsted-Lowry base A substance that accepts a proton in an acid–base reaction. *(Sec. 14.3)*

buffer A solution that resists changes in pH when an acid or a base is added. *(Sec. 14.1)*

C

calorie (cal) The amount of heat required to raise the temperature of 1 g of water 1 °C. *(Sec. 2.10)*

carbohydrate A biological compound that contains one or more sugar molecules; a sugar is a polyhydroxy aldehyde or ketone. *(Sec. 20.1)*

carbonyl group **($-$C$=$O)** The C$=$O group, which is present in aldehydes, ketones, carboxylic acids, esters, and amides. *(Sec. 19.5)*

carboxyl group ($-$COOH) The $-$COOH group, which is present in carboxylic acids. *(Sec. 19.10)*

catalyst A substance that speeds up a reaction without being permanently changed. *(Sec. 7.2)* A substance that allows a reaction to proceed faster by lowering the energy of activation. *(Sec. 16.2)*

cathode The electrode in an electrochemical cell at which reduction occurs. *(Sec. 17.6)*

cation A positively charged ion. *(Sec. 6.1)*

Celsius degree (°C) The basic unit of temperature in the metric system. *(Sec. 2.9)*

chain reaction A fission reaction in which the neutrons released initiate a second reaction, which in turn initiates a third reaction, and so on. *(Sec. 18.7)*

Charles's law The statement that the volume and the Kelvin temperature of a gas are directly proportional at constant pressure; that is, $V_1/T_1 = V_2/T_2$. *(Sec. 10.5)*

centimeter (cm) A common unit of length in the metric system that is equal to one-hundredth of a meter. *(PSS.1)*

chemical bond The attraction between positive and negative ions, or two nonmetal atoms. *(Sec. 12.1)*

chemical change A modification of a substance that alters its chemical composition. *(Sec. 3.7)*

chemical equation A shorthand representation using formulas and symbols to describe a chemical reaction. *(Sec. 7.2)*

chemical equilibrium A dynamic state for a reversible reaction in which the rates of the forward and reverse reactions are equal and the amounts of reactants and products are constant. *(Sec. 16.3)*

chemical formula An abbreviation for the name of a chemical compound that indicates the number of atoms of each element; for example, H_2O is the formula for water. *(Sec. 3.5)*

chemical property A characteristic of a substance that cannot be observed without changing its chemical formula. *(Sec. 3.6)*

chemical reaction The process of undergoing a chemical change. *(Sec. 7.1)*

chemical symbol An abbreviation for the name of a chemical element; for example, Cu is the symbol for copper. *(Sec. 3.3)*

chemistry The branch of science that studies the composition and properties of matter. *(Sec. 1.2)*

class of compounds A family of compounds in which all the members have the same structural feature (that is, an atom or group of atoms) and similar chemical properties. *(Sec. 19.5)*

codon A term for three consecutive nucleotides in RNA that designates a specific amino acid to be added to a growing protein chain. *(Sec. 20.6)*

coefficient A digit in front of a chemical formula that helps to balance an equation; a digit that helps to balance the number of atoms of each element on the left side of an equation with those on the right side of the equation. *(Sec. 7.3)*

collision theory The principle that the rate of a chemical reaction is regulated by the collision frequency, collision energy, and the orientation of molecules striking each other. *(Sec. 16.1)*

colloid A homogeneous mixture in which the diameters of the dispersed particles range from 1 to 100 nm. *(Sec. 13.3)*

combination reaction A type of reaction in which two substances produce a single compound. *(Sec. 7.4)*

combined gas law The statement that the pressure exerted by a gas is inversely proportional to its volume and directly proportional to its Kelvin temperature; that is, $P_1V_1/T_1 = P_2V_2/T_2$. *(Sec. 10.7)*

combustion reaction A chemical reaction in which a hydrocarbon reacts rapidly with oxygen (burns) to produce carbon dioxide, water, and heat. *(Sec. 19.2)*

compound A pure substance that can be broken down into two or more simpler substances by a chemical reaction. *(Sec. 3.2)*

continuous spectrum A single, broad band of radiant energy. *(Sec. 4.6)*

coordinate covalent bond A bond in which an electron pair is shared, but both electrons have been donated by a single atom. *(Sec. 12.8)*

core electrons The inner electrons in an atom, which are not available for bonding. *(Sec. 5.8)*

core notation A method of writing the electron configuration in which the inner core electrons are represented by a noble gas symbol in brackets followed by valence electrons; for example, [Ne] $3s^2$. *(Sec. 5.10)*

covalent bond A bond resulting from the sharing of one or more pairs of valence electrons between two nonmetal atoms. *(Sec. 12.1)*

critical mass The minimum mass of a fissionable nuclide necessary to sustain a continuous chain reaction. *(Sec. 18.7)*

crystalline solid A solid substance composed of ions or molecules that repeat in a regular geometric pattern. *(Sec. 11.4)*

cubic centimeter **(cm^3)** A unit of volume occupied by a cube 1 cm on a side; 1 cm^3 is exactly equal to 1 mL. *(Sec. 2.6)*

D

Dalton's law of partial pressures The statement that the pressure exerted by a mixture of gases is equal to the sum of the individual pressures exerted by each gas; that is, $P_1 + P_2 + P_3 + \ldots = P_{total}$. *(Sec. 10.9)*

decay series See *radioactive decay series.*

decomposition reaction A type of reaction in which a single compound produces two or more substances. *(Sec. 7.4)*

degree Celsius See *Celsius degree.*

degree Fahrenheit See *Fahrenheit degree.*

deionized water Water purified by removing ions using an ion exchange method; also termed *demineralized water*. *(Sec. 11.7)*

delta (δ) **notation** A method of indicating the partial positive charge (δ^+) and the partial negative charge (δ^-) in a polar covalent bond. *(Sec. 12.6)*

density (*d*) The amount of mass in one unit volume of matter. *(Sec. 2.8)*

deposition A direct change of state from a gas to a solid without forming a liquid. *(Sec. 3.1)*

deuterium ($^2_1$**H**) The nuclide of hydrogen with one neutron in the nucleus. *(Sec. 18.8)*

diatomic molecule A particle composed of two nonmetal atoms. *(Sec. 7.2)* A molecule composed of two nonmetal atoms held together by a covalent bond. *(Sec. 12.7)*

dipeptide A molecule that contains two amino acids joined by an amide bond. *(Sec. 20.2)*

dipole A region in a molecule having partial negative charge and partial positive charge resulting from a polar bond. *(Sec. 13.2)*

dipole force A type of attraction between two molecules that have permanent dipoles. *(Sec. 11.2)*

directly proportional A relationship between two variables that have a direct relationship; for example, if one variable doubles the other doubles, or if one variable is halved the other is halved. *(Sec. 10.5)*

disaccharide A carbohydrate composed of two simple sugar molecules joined by a glycoside linkage. *(Sec. 20.4)*

dispersion force A type of attraction between two molecules that have temporary dipoles. A temporary dipole arises from the unequal distribution of charge as electrons move around a molecule; also called a *London dispersion force. (Sec. 11.2)*

dissociation The process of an ionic compound dissolving in water and separating into positive and negative ions; for example, NaOH dissolves in water to give sodium ions and hydroxide ions. *(Sec. 14.2)*

distilled water Water purified by boiling and collecting the condensed vapor. *(Sec. 11.7)*

double bond A bond composed of two electron pairs shared between two atoms. A double bond is represented by two dashes between the symbols of the two atoms. *(Sec. 12.4)*

double-replacement reaction A type of reaction in which two cations in different compounds exchange anions. *(Sec. 7.4)*

dry cell An electrochemical cell in which the anode and cathode reactions do not take place in an aqueous solution. *(Sec. 17.6)*

ductile The property of a metal that allows it to be drawn into a wire. *(Sec. 3.4)*

E

elastic collision An impact between gas molecules in which the total energy remains constant. *(Sec. 10.10)*

electrochemical cell A general term for an apparatus that contains two solutions with electrodes in separate compartments that are connected by a conducting wire and salt bridge. *(Sec. 17.6)*

electrochemistry The study of the interconversion of chemical and electrical energy from redox reactions. *(Sec. 17.6)*

electrolysis The chemical reaction produced by passing electricity through an aqueous solution. *(Sec. 11.9)* The providing of electricity to a nonspontaneous redox process to cause a reaction. *(Sec. 17.7)*

electrolytic cell An electrochemical cell in which a nonspontaneous redox reaction occurs as a result of the input of direct electric current. *(Sec. 17.7)*

electromotive series See *activity series.*

electron (e^-) A negatively charged subatomic particle having a tiny mass. *(Sec. 4.2)*

electron capture (EC) A nuclear reaction in which a heavy nuclide attracts one of its core electrons into the nucleus. *(Sec. 18.2)*

electron configuration A shorthand description of the arrangement of electrons by sublevels according to increasing energy. *(Secs. 4.10 and 5.6)*

electron dot formula The chemical symbol of an element surrounded by a dot representing each valence electron. (*Sec. 5.8*) A diagram of a molecule or polyatomic ion that shows the chemical symbol of each atom and a dot representing each bonding and nonbonding electron. An electron dot formula is also called a *Lewis structure*. (*Sec. 12.4*)

electron pair geometry The geometric shape formed by bonding and nonbonding electron pairs surrounding the central atom in a molecule. *(Sec. 12.9)*

electronegativity The ability of an atom to attract a pair of electrons in a covalent chemical bond. *(Sec. 12.6)*

element A pure substance that cannot be broken down by an ordinary chemical reaction. *(Sec. 3.2)*

emission line spectrum Several narrow bands of radiant energy that result from excited atoms releasing energy. *(Sec. 4.8)*

empirical formula A chemical formula that expresses the simplest whole-number ratio of atoms of each element in a molecule, or ions in an ionic compound. *(Sec. 8.8)*

endothermic reaction A chemical reaction that absorbs heat energy. *(Secs. 7.1 and 16.2)*

endpoint The stage in a titration when the indicator changes color permanently. *(Sec. 14.5)*

energy level An orbit of specific energy that electrons occupy at a fixed distance from the nucleus; also referred to as a *main energy level*; designated as 1, 2, 3, 4 . . . .*(Sec. 4.8)*

energy sublevel An electron energy level that results from splitting a main energy level; designated as *s*, *p*, *d*, *f* . . . . *(Sec. 4.9)*

English system A nondecimal system of measurement without any basic unit for length, mass, or volume. *(Sec. 2.1)*

enzyme A protein molecule that catalyzes a specific biochemical reaction. *(Sec. 20.3)*

equilibrium constant See *general equilibrium constant* (K_{eq}), *ionization equilibrium constant* (K_i), or *solubility product equilibrium constant* (K_{sp}).

evaporation See *vaporization.*

exact equivalent A statement that relates two values that are exactly equal; for example, 1 yd = 36 in. and 1 m = 100 cm. *(Sec. 2.2)*

exothermic reaction A chemical reaction that releases heat energy. *(Secs. 7.1 and 16.2)*

experiment A scientific procedure for collecting data and recording observations under controlled conditions. *(Sec. 1.1)*

exponent A number written as a superscript indicating that a value is multiplied by itself; for example, $10^4 = 10 \times 10 \times 10 \times 10$ and $cm^3 = cm \times cm \times cm$. *(PSS.6)*

F

Fahrenheit degree (°F) A basic unit of temperature in the English system. *(Sec. 2.9)*

fat A triglyceride from an animal source that contains mostly saturated fatty acids. *(Sec. 20.5)*

fatty acid A carboxylic acid with a long hydrocarbon chain. *(Sec. 20.5)*

formula unit The simplest representative particle in a compound composed of ions. *(Sec. 6.4)* The simplest representative particle in a substance held together by ionic bonds. *(Sec. 12.1)*

frequency The number of times a light wave travels a complete cycle in one second. *(Sec. 4.6)*

functional group An atom or group of atoms that characterizes a class of compounds and contributes to their similar chemical properties. *(Sec. 19.5)*

G

galvanic cell See *voltaic cell.*

gamma ray (γ) A powerful type of nuclear radiation that has neither mass nor charge. *(Sec. 18.1)*

gas density The ratio of mass per unit volume for a gas, usually expressed in grams per liter. *(Sec. 8.5)*

gas pressure A measure of the frequency and energy of gas molecules colliding against the walls of their container. *(Sec. 10.2)*

Gay-Lussac's law The statement that the pressure and the Kelvin temperature of a gas are directly proportional at constant volume; that is, $P_1/T_1 = P_2/T_2$. *(Sec. 10.6)*

Gay-Lussac's law of combining volumes See *law of combining volumes.*

general equilibrium constant (K_{eq}) A constant that expresses the molar equilibrium concentration of each substance participating in a reversible reaction at a given temperature. *(Sec. 16.3)*

glycoside linkage An —O— bond that joins two simple sugar molecules. *(Sec. 20.4)*

gram (g) A common metric unit of mass. *(PSS.1)* The basic unit of mass in the metric system. *(Sec. 2.1)*

green chemistry A term that refers to the design of products and processes that reduce or eliminate hazardous chemical substances; also referred to as the *sustainable chemistry*. (*Sec. 1.2*)

group A vertical column in the periodic table; a family of elements with similar properties. *(Sec. 5.3)*

H

half-cell A portion of an electrochemical cell having a single electrode where either oxidation or reduction occurs. *(Sec. 17.6)*

half-life ($t_{1/2}$) The amount of time required for 50% of the radioactive nuclei in a given sample to decay. *(Sec. 18.4)*

half-reaction A reaction that represents either an oxidation or reduction process separately and indicates the number of electrons lost or gained. *(Sec. 17.4)*

halogens The Group VIIA/17 elements. *(Sec. 5.3)*

hard water Water that contains a variety of cations and anions such as Ca^{2+}, Mg^{2+}, Fe^{3+}, CO_3^{2-}, SO_4^{2-}, and PO_4^{3-}. *(Sec. 11.7)*

heat The flow of energy from an object at a higher temperature to an object at a lower temperature. *(Sec. 2.10)*

heat of fusion (H_{fusion}) The heat required to convert a solid to a liquid at its melting point. Conversely, the heat released when a liquid changes to a solid is called the *heat of solidification* (H_{solid}). *(Sec. 11.6)*

heat of reaction (ΔH) The difference in heat energy between the reactants and the products for a given chemical reaction. *(Sec. 16.2)*

heat of vaporization (H_{vapor}) The heat required to convert a liquid to a gas at its boiling point. Conversely, the heat released when a gas changes to a liquid is called the *heat of condensation* (H_{cond}). *(Sec. 11.6)*

Henry's law The principle that the solubility of a gas in a liquid is proportional to the partial pressure of the gas above the liquid. *(Sec. 13.1)*

heterogeneous equilibrium A type of equilibrium in which all of the participating species are not in the same physical state; for example, an equilibrium between an insoluble precipitate and its ions in an aqueous solution. *(Sec. 16.4)*

heterogeneous mixture Matter having an indefinite composition and variable properties. *(Sec. 3.2)*

homogeneous equilibrium A type of equilibrium in which all of the participating species are in the same physical state; for example, an equilibrium between reactants and products in the gaseous state. *(Sec. 16.4)*

homogeneous mixture Matter having a definite composition, but variable properties from sample to sample; examples include alloys and solutions. *(Sec. 3.2)*

hydrate A substance that contains a specific number of water molecules in a crystalline compound. *(Sec. 11.9)*

hydrocarbon A compound that contains only hydrogen and carbon. *(Sec. 19.1)*

hydrocarbon derivative A compound that contains carbon, hydrogen, and another element such as oxygen, nitrogen, or a halogen. *(Sec. 19.5)*

hydrogenation reaction A chemical reaction in which a hydrogen molecule adds to an unsaturated bond in an alkene or alkyne. *(Sec. 19.3)*

hydrogen bond An intermolecular attraction based on permanent dipoles involving a H atom bonded to an O or N atom. *(Sec. 11.9)* The attraction between a H atom bonded to an O or N atom, and a nonbonding electron pair on an electronegative atom in another molecule. *(Sec. 12.2)*

hydronium ion (H_3O^+) The ion that results when a hydrogen ion attaches to a water molecule by a coordinate covalent bond. The hydronium ion is the predominant form of the hydrogen ion in an aqueous acid solution. *(Sec. 14.2)*

hydroxyl group (—OH) The —OH group, which is present in alcohols and phenols. *(Sec. 19.7)*

hypothesis An initial, tentative proposal of a scientific principle that attempts to explain the meaning of a set of data collected in an experiment. *(Sec. 1.1)*

I

ideal gas A theoretical gas that obeys the kinetic theory under all conditions of temperature and pressure. *(Sec. 10.10)*

ideal gas constant (*R*) The proportionality constant R in the equation $PV = nRT$. *(Sec. 10.11)*

ideal gas law The principle stated by the relationship $PV = nRT$; also called the *ideal gas equation. (Sec. 10.11)*

immiscible A term that refers to liquids that do not dissolve in one another and separate into two layers. *(Sec. 13.2)*

inner transition elements The elements in the lanthanide and actinide series. *(Sec. 5.3)*

inorganic chemistry The study of chemical substances that do not contain the element carbon. *(Sec. 1.2)*

inorganic compound A compound that does not contain carbon. *(Sec. 6.1)*

instrument A device for recording a measurement such as length, mass, volume, time, or temperature. *(PSS.1)*

International System (SI) A sophisticated system of measurement that is more comprehensive than the metric system and has seven base units. *(Sec. 2.1)*

inversely proportional A relationship between two variables that have an inverse relationship; for example, if one variable doubles, the other variable is halved. *(Sec. 10.4)*

ion An atom (or group of atoms) that has a negative or a positive charge as the result of gaining or losing valence electrons. *(Sec. 5.9)*

ionic bond A bond resulting from the attraction between a positively charged cation and a negatively charged anion. *(Sec. 12.1)*

ionic charge A term for the positive charge on a metal atom that has lost electrons, or the negative charge on a nonmetal atom that has gained electrons. *(Sec. 5.10)*

ionic solid A crystalline solid composed of ions that repeat in a regular pattern. *(Sec. 11.5)*

ionization The process of a polar compound dissolving in water and forming positive and negative ions; for example, HCl dissolves in water to give hydrogen ions and chloride ions. *(Sec. 14.2)*

ionization constant of water (K_w)A constant that equals the product of the molar hydrogen ion concentration and the molar hydroxide ion concentration in water; $K_w = 1.0 \times 10^{-14}$ at 25°C. *(Sec. 14.7)*

ionization energy The amount of energy necessary to remove an electron from a neutral atom in the gaseous state. *(Sec. 5.9)*

ionization equilibrium constant (K_i) A constant that expresses the molar equilibrium concentrations of ions in an aqueous solution for a slightly ionized acid or base. *(Sec. 16.6)*

isoelectronic A term for two or more ions (or ions and an atom) with the same electron configuration; for example, Na^+ and F^- each have 10 electrons, and their electron configurations are identical to the noble gas neon. *(Sec. 5.10)*

isomers Compounds with the same molecular formula, but with different structural formulas. Isomers have different physical and chemical properties. *(Sec. 19.2)*

isotopes Atoms with the same atomic number, but a different mass number; atoms of the same element that differ by the number of neutrons in the nucleus. *(Sec. 4.4)*

IUPAC nomenclature The system of rules for naming chemical compounds set forth by the International Union of Pure and Applied Chemistry. *(Sec. 6.1)*

J

joule (J) A unit of energy in the SI system; 1 cal = 4.184 J. *(Sec. 2.10)*

K

Kelvin unit (K) The basic unit of temperature in the SI system. *(Sec. 2.9)*

kinetic energy (KE) The energy associated with the motion of particles; the energy associated with the mass and velocity of a particle. *(Sec. 3.9)*

kinetic theory A theoretical description of gas molecules demonstrating ideal behavior; also referred to as the *kinetic molecular theory. (Sec. 10.10)*

L

lanthanide series The elements with atomic numbers 58–71. *(Sec. 5.3)*

Latin system A system for naming ionic compounds that designates the variable charge on a metal cation with an −*ic* or −*ous* suffix attached to the stem of the Latin name. *(Sec. 6.2)*

law See *natural law.*

law of chemical equilibrium The principle that the molar concentrations of the products in a reversible reaction divided by the molar concentrations of the reactants (each raised to a power corresponding to a coefficient in the balanced equation) is equal to a constant. The law of chemical equilibrium can be written mathematically as $K_{eq} = [C]^a [D]^d/[A]^a [B]^b$. *(Sec. 16.3)*

law of combining volumes The principle that volumes of gases that combine in a chemical reaction, at the same temperature and pressure, are in the ratio of small whole numbers; also called *Gay-Lussac's law of combining volumes. (Sec. 9.6)*

law of conservation of energy The principle that states energy can neither be created nor destroyed, but can be changed from one form to another; for example, nuclear energy to heat energy. *(Sec. 3.10)*

law of conservation of mass The principle that states matter can neither be created nor destroyed. *(Sec. 3.8)* The statement that mass is neither gained nor lost during a chemical reaction. *(Sec. 9.1)*

law of conservation of mass and energy The principle that states the total mass and energy in the universe is constant. The statement that the total mass and energy, before and after a chemical change, is constant. *(Sec. 3.10)*

law of definite composition The principle that states a compound always contains the same elements in the same proportion by mass; also called the *law of definite proportion. (Sec. 3.5)*

law of multiple proportions The principle that two elements that can form more than one compound combine in the ratio of small whole numbers; for example, CO and CO_2. *(Sec. 4.1)*

Le Chatelier's principle The statement that any reversible reaction at equilibrium when stressed by a change in concentration, temperature, or pressure, shifts to relieve the stress. *(Sec. 16.5)*

Lewis structure See *electron dot formula.*

light A specific term that refers to visible radiant energy; that is, violet to red. A general term that refers to both visible and invisible radiant energy. *(Sec. 4.6)*

***like dissolves like* rule** The general principle that solubility is greatest when the polarity of the solute and solvent are similar. *(Sec. 13.2)*

limiting reactant The substance in a chemical reaction that controls or limits the maximum amount of product formed. *(Sec. 9.7)*

line spectrum See *emission line spectrum.*

lipid A biological compound such as a fat, oil, wax, or steroid that is insoluble in water. *(Sec. 20.1)*

liter (L) The basic unit of volume in the metric system equal to the volume of a cube 10 cm on a side. *(Sec. 2.1)*

lock-and-key model A theory that explains enzyme catalysis; the theory states that the shape of the substrate molecule must fit the contour of the active site on an enzyme. *(Sec. 20.3)*

M

main-group elements See *representative elements.*

malleable The property of a metal that allows it to be hammered or machined into a foil. *(Sec. 3.4)*

mass The quantity of matter in an object that is measured by a balance. *(PSS.1)*

mass/mass percent (m/m %) A solution concentration expression that relates the mass of solute in grams dissolved in each 100 g of solution. *(Sec. 13.8)*

mass–mass problem A type of stoichiometry calculation that relates the masses of two substances according to a balanced equation. *(Sec. 9.3)*

mass number (*A*) A value that indicates the number of protons and neutrons in the nucleus of a given atom. *(Secs. 4.4 and 18.2)*

mass–volume problem A type of stoichiometry calculation that relates the mass of a substance to the volume of a gas according to a balanced equation. *(Sec. 9.3)*

measurement A numerical value with an attached unit that expresses a physical quantity such as length, mass, volume, time, or temperature. *(PSS.1)*

metal An element that is generally shiny in appearance, has a high density, high melting point, and is a good conductor of heat and electricity. *(Sec. 3.4)*

metal oxide A compound that reacts with water to form an alkaline solution; also termed a *basic oxide. (Sec. 11.9)*

metallic solid A crystalline solid composed of metal atoms that repeat in a regular pattern. *(Sec. 11.5)*

metalloid See *semimetal.*

meter (m) The basic unit of length in the metric system of measurement. *(Sec. 2.1)*

metric system A decimal system of measurement using prefixes and a basic unit to express physical quantities such as length, mass, and volume. *(Sec. 2.1)*

milliliter (mL) A common unit of volume in the metric system equal to one-thousandth of a liter. *(PSS.1)*

miscible A term that refers to liquids that dissolve completely in one another. *(Sec. 13.2)*

molar mass (MM) The mass of 1 mol of substance expressed in grams. *(Secs. 8.3 and 9.1)*

molar volume The volume occupied by 1 mol of gas at STP; at 0 °C and 1.00 atm the volume of 1 mol of gas is 22.4 L. *(Secs. 8.5 and 9.5)*

molarity (*M*) A solution concentration expression that relates the moles of solute dissolved in each liter of solution. *(Sec. 13.9)*

mole (mol) The amount of substance that contains Avogadro's number of particles; that is, an amount of substance that contains 6.02×10^{23} particles. *(Sec. 8.2)*

mole ratio The ratio of moles of reactants and products according to the coefficients in the balanced chemical equation. *(Sec. 9.1)*

molecular formula A chemical formula that expresses the actual number of atoms of each element in a molecule. *(Sec. 8.9)*

molecular shape The geometric shape formed by the atoms bonded to the central atom in a molecule; also called *molecular geometry. (Sec. 12.10)*

molecular solid A crystalline solid composed of molecules that repeat in a regular pattern. *(Sec. 11.5)*

molecule A single particle composed of nonmetal atoms. *(Sec. 3.5)* The simplest representative particle in a compound composed of nonmetals. *(Sec. 6.7)* The simplest particle in a substance held together by covalent bonds. *(Sec. 12.1)*

monoatomic ion A single atom that has a negative or positive a charge as the result of gaining or losing valence electrons. *(Secs. 6.1 and 12.5)*

monomer A small molecule that bonds to itself repeatedly to form a long, continuous chain. *(Sec. 19.3)*

monosaccharide A carbohydrate composed of a simple sugar molecule. A simple sugar is characterized by an aldehyde or ketone functional group and usually several hydroxyl groups. *(Sec. 20.4)*

N

natural law An extensively tested proposal of a scientific principle that states a measurable relationship under different experimental conditions. A natural law is often expressed as an equation; for example, $P_1V_1 = P_2V_2$. *(Sec. 1.1)*

net dipole The overall direction of partial negative charge in a molecule with two or more dipoles. *(Secs. 11.7 and 13.2)*

net ionic equation A chemical equation that results from canceling spectator ions in a total ionic equation. *(Sec. 14.11)*

neutralization reaction A type of reaction in which an acid and a base produce a salt and water. *(Sec. 7.4)*

neutron (n^0) A neutral subatomic particle with a mass of 1 amu, which is approximately equal to the mass of a proton. *(Sec. 4.3)*

noble gases The relatively unreactive Group VIIIA/18 elements. *(Sec. 5.3)*

nonbonding electrons The valence electrons in a molecule that are not shared; also referred to as *lone pairs*. *(Sec. 12.4)*

nonmetal An element that is generally dull in appearance, has a low density, low melting point, and is not a good conductor of heat and electricity. *(Sec. 3.4)*

nonmetal oxide A compound that reacts with water to form an acidic solution; also termed an *acidic oxide. (Sec. 11.9)*

nonpolar bond A covalent bond in which one or more pairs of electrons are shared equally. *(Sec. 12.7)*

nonpolar solvent A dissolving liquid composed of nonpolar molecules. *(Sec. 13.2)*

nonsignificant digits The digits in a measurement that exceed the certainty of the instrument. *(PSS.3)*

nuclear equation A shorthand representation using atomic notation to describe a nuclear reaction. *(Sec. 18.2)*

nuclear fission A nuclear reaction in which a heavy nucleus splits into two or more lighter nuclei and releases energy. *(Sec. 18.7)*

nuclear fusion A nuclear reaction in which two light nuclei combine into a single heavy nucleus and releases energy. *(Sec. 18.8)*

nuclear reaction A high-energy change involving the atomic nucleus. *(Sec. 18.2)*

nucleic acid A biological polymer compound that carries genetic information. *(Sec. 20.1)*

nucleotide A repeating unit in a nucleic acid composed of a sugar, a base, and phosphoric acid. *(Sec. 20.6)*

nucleus See *atomic nucleus.*

nuclide A specific atom with a designated number of protons and neutrons in its nucleus. *(Sec. 18.2)*

O

octet rule The statement that an atom tends to bond in such a way so as to acquire eight valence electrons. A hydrogen atom is an exception to the rule and acquires only two valence electrons. *(Sec. 12.1)*

oil A triglyceride from a plant source that contains mostly unsaturated fatty acids. *(Sec. 20.5)*

orbital A region in space surrounding the nucleus of an atom in which there is a high probability (~95%) of finding an electron with a given energy. *(Sec. 4.11)*

organic chemistry The study of chemical substances that contain the element carbon. *(Sec. 1.2)* The study of compounds that contain one or more carbon atoms. *(Sec. 19.1)*

oxidation A chemical process characterized by the loss of electrons. A process in which a substance undergoes an increase in oxidation number. *(Sec. 17.2)*

oxidation number A positive or negative value assigned to an atom in a substance according to a set of rules; a value that indicates whether an atom is electron-poor or electron-rich compared to the free atom. *(Sec. 17.1)*

oxidation–reduction reaction See *redox reaction.*

oxidizing agent A substance that causes the oxidation of another substance in a redox reaction; the substance that is reduced in a redox reaction. *(Sec. 17.2)*

oxyanion A polyatomic anion that contains one or more elements combined with oxygen; for example, NO_3^- and SO_4^{2-}. *(Sec. 6.3)*

P

parent–daughter nuclides A term for a decaying nuclide and the resulting nuclide that is produced. *(Sec. 18.3)*

partial pressure The pressure exerted by an individual gas in a mixture of two or more gases. *(Sec. 10.9)*

peptide linkage An amide bond that joins two amino acid molecules in a small peptide, polypeptide, or protein. *(Sec. 20.2)*

percent (%) An expression for the amount of a single quantity compared to an entire sample; an expression of parts per hundred parts. *(Sec. 2.5)*

percent composition A list of the mass percent of each element in a compound. *(Sec. 8.7)*

percent yield The actual yield compared to the theoretical yield expressed as a percent. *(Sec. 9.9)*

period A horizontal row in the periodic table; a series of elements with properties that vary from metallic to nonmetallic. *(Sec. 5.3)*

periodic law The properties of the elements recur in a repeating pattern when arranged according to increasing atomic number. *(Sec. 5.2)*

periodic table A chart that arranges elements according to their properties and includes metals, nonmetals, and semimetals. *(Sec. 3.4)*

pH The molar concentration of hydrogen ions expressed on an exponential scale; the negative logarithm of the molar hydrogen ion concentration. *(Sec. 14.8)*

phenyl group The part that remains after a hydrogen atom is removed from a benzene molecule. *(Sec. 19.2)*

phospholipid A lipid that has two fatty acids and phosphoric acid joined to glycerol by ester bonds. *(Sec. 20.5)*

photon A particle of radiant energy. *(Sec. 4.7)*

physical change A modification of a substance that does not alter its chemical composition; for example, a change in physical state. *(Sec. 3.7)*

physical property A characteristic of a substance that can be observed without changing its chemical formula. *(Sec. 3.6)*

physical state The condition of matter existing as a solid, liquid, or gas. *(Sec. 3.1)*

polar bond A covalent bond in which one or more pairs of electrons are shared unequally. *(Sec. 12.6)*

polar solvent A dissolving liquid composed of polar molecules. *(Sec. 13.2)*

polyatomic ion A group of atoms bound together that bears an overall negative or positive charge. *(Sec. 6.1)* A group of atoms held together by covalent bonds that bears an overall negative or positive charge. *(Sec. 12.5)*

polymer A giant molecule composed of small molecules bonded together repeatedly in a long, continuous chain. *(Sec. 19.3)*

polypeptide A large molecule composed of 10–50 amino acids joined by amide bonds. *(Sec. 20.2)*

polysaccharide A carbohydrate polymer composed of simple sugar molecules joined by glycoside linkages. *(Sec. 20.4)*

positron (β^+ or ${}^{0}_{+1}e$) A nuclear radiation identical in mass, but opposite in charge, to that of an electron. The term *positron* is derived from "positive electron." *(Sec. 18.2)*

potential energy (PE) The stored energy that matter possesses owing to its position or chemical composition. *(Sec. 3.9)*

power of 10 A positive or negative exponent of 10. *(PSS.6)*

precipitate An insoluble solid substance produced by a chemical reaction in an aqueous solution. *(Sec. 7.1)*

product A substance resulting from a chemical reaction. *(Sec. 7.2)*

protein A biological compound that is a polymer composed of amino acids joined by peptide linkages. *(Sec. 20.1)*

proton (p^+) A positively charged subatomic particle having a mass of 1 amu, which is approximately 1836 times the mass of an electron. *(Sec. 4.2)*

proton acceptor A term for a Brønsted–Lowry base, which is a synonym for hydrogen ion acceptor. *(Sec. 14.3)*

proton donor A term for a Brønsted–Lowry acid, which is a synonym for hydrogen ion donor. *(Sec. 14.3)*

pure substance See *substance.*

Q

quantum level See *energy level.*

quantum mechanical atom A model of the atom that describes an electron in terms of its probability of being found in a particular location around the nucleus. *(Sec. 4.11)*

R

radiant energy spectrum A range of light energy extending from short-wavelength gamma rays to long-wavelength microwaves. *(Sec. 4.6)*

radioactive decay series The stepwise disintegration of an unstable nuclide until a stable nucleus is reached; also called *decay series*. *(Sec. 18.3)*

radioactivity The emission of particles or energy from an unstable atomic nucleus. *(Sec. 18.1)*

radionuclide An atom whose nucleus is unstable and decays radioactively. *(Sec. 18.5)*

rare earth elements The elements with atomic numbers 21, 39, 57, and 58–71. *(Sec. 5.3)*

rate of reaction The rate at which the concentrations of reactants decrease, or the concentrations of products increase, per unit time. *(Sec. 16.3)*

reactant A substance undergoing a chemical reaction. *(Sec. 7.2)*

reaction profile A graph of the energy of reactants and products as a reaction progresses. *(Sec. 16.2)*

real gas A gas that deviates from ideal behavior at a low temperature and a high pressure. *(Sec. 10.10)*

reciprocal The relationship between a fraction and its inverse; for example, 1 yard/3 feet and 3 feet/1 yard. *(Sec. 2.2)*

redox reaction A chemical reaction that involves electron transfer between two reacting substances. *(Sec. 17.2)*

reducing agent A substance that causes the reduction of another substance in a redox reaction; the substance that is oxidized in a redox reaction. *(Sec. 17.2)*

reduction A chemical process characterized by the gain of electrons. A process in which a substance undergoes a decrease in oxidation number. *(Sec. 17.2)*

reduction potential The relative ability of a substance to undergo reduction; the relative strength of an oxidizing agent. *(Sec. 17.5)*

replication The process by which a DNA molecule synthesizes an identical molecule of DNA. *(Sec. 20.6)*

representative elements The Group A (1, 2, and 13–18) elements in the periodic table; also termed *main-group elements. (Sec. 5.3)*

reversible reaction A reaction that proceeds simultaneously in both the forward direction toward products and the opposite direction toward reactants. *(Sec. 16.3)*

rounding off The process of eliminating digits that are not significant. *(PSS.3)*

S

salt An ionic compound produced by an acid–base reaction; the product of a neutralization reaction in addition to water. *(Secs. 7.11 and 14.2)*

salt bridge A porous device that allows ions to travel between two half-cells to maintain an ionic charge balance in each compartment. *(Sec. 17.6)*

saponification A chemical reaction of a fat or oil with sodium hydroxide to produce soap and glycerol. *(Sec. 20.5)*

saturated hydrocarbon A hydrocarbon that contains only single bonds. *(Sec. 19.1)*

saturated solution A solution that contains the maximum solute that can dissolve at a given temperature. *(Sec. 13.7)*

science The methodical exploration of nature and the logical explanation of the observations. *(Sec. 1.1)*

scientific method A systematic investigation that involves performing an experiment, proposing a hypothesis, testing the hypothesis, and finally, stating a theory or law that explains a scientific principle. *(Sec. 1.1)*

scientific notation A method for expressing numbers by moving the decimal place after the first significant digit and indicating the number of decimal moves by a power of 10. *(PSS.7)*

scientific theory An extensively tested proposal of a scientific principle that explains the behavior of nature. A theory offers a model to describe nature; for example, the atomic theory describes the particle nature of matter. *(Sec. 1.1)*

second (s) The basic unit of time in the metric system. *(Sec. 2.1)*

semimetal An element that is generally metallic in appearance and has properties midway between those of a metal and a nonmetal; also called a *metalloid. (Sec. 3.4)*

SI See *International System.*

significant digits The digits in a measurement known with certainty plus one digit that is estimated; also referred to as *significant figures,* or simply, *"sig figs." (PSS.2)*

single bond A bond composed of one electron pair shared between two atoms. A single bond is represented by a dash between the symbols of the two atoms. *(Sec. 12.4)*

single-replacement reaction A type of reaction in which a more active element displaces a less active element from a solution or compound. *(Sec. 7.4)*

soft water Water that contains sodium cations, Na^+, and a variety of anions such as CO_3^{2-}, SO_4^{2-}, and PO_4^{3-}. *(Sec. 11.7)*

solubility The maximum amount of solute that can dissolve in a solvent at a given temperature; usually expressed in grams of solute per 100 g of solvent. *(Sec. 13.6)*

solubility product equilibrium constant (K_{sp}) A constant that expresses the molar equilibrium concentrations of ions in an aqueous solution for a slightly dissociated ionic compound. *(Sec. 16.8)*

solute The component of a solution that is the lesser quantity. *(Sec. 13.2)*

solution The general term for a solute dissolved in a solvent. A solution is an example of a homogeneous mixture. *(Sec. 13.2)*

solvent The component of a solution that is the greater quantity. *(Sec. 13.2)*

solvent cage A cluster of solvent molecules surrounding a solute particle in a solution. *(Sec. 13.4)*

specific gravity (sp gr) The ratio of the density of a liquid compared to the density of water at 4 °C; a unitless expression. *(Sec. 2.8)*

specific heat The amount of heat required to raise the temperature of 1 g of any substance 1 °C; the specific heat of water is 1.00 cal/(g × °C). *(Secs. 2.10, 11.6)*

spectator ions Those ions in aqueous solution that do not participate in a reaction, and do not appear in the net ionic equation. *(Sec. 14.11)*

standard conditions See *standard temperature and pressure (STP).*

standard solution A solution whose concentration has been established accurately (usually by titration to 3 or 4 significant digits). *(Sec. 14.6)*

standard temperature and pressure (STP) A temperature of 0 °C and a pressure of 1 atm. *(Secs. 8.5 and 9.5)* A temperature of 273 K and a pressure of 760 mm Hg for a gas. *(Sec. 10.7)*

steroid A lipid hormone composed of four rings of carbon atoms fused into a single molecular structure. *(Sec. 20.5)*

Stock system A system for naming ionic compounds that designates the variable charge on a metal cation with Roman numerals in parentheses. *(Sec. 6.2)*

stoichiometry The relationship of quantities (i.e., mass of substance or volume of gas) in a chemical reaction according to the balanced chemical equation. *(Sec. 9.3)*

strong electrolyte An aqueous solution that is a good conductor of electricity; for example, strong acids, strong bases, and soluble salts. *(Sec. 14.10)*

structural formula A diagram of a molecule or polyatomic ion that shows the chemical symbol of each atom and a dash representing each pair of bonding electrons. *(Sec. 12.4)* A chemical formula that shows the arrangement of atoms in a molecule. *(Sec. 19.2)*

sublimation A direct change of state from a solid to a gas without forming a liquid. *(Sec. 3.1)*

subscript A digit in a chemical formula that represents the number of atoms or ions in the substance. *(Sec. 7.3)*

substance Matter having a definite composition and constant properties; examples include elements and compounds. It is also referred to as *pure substance. (Sec. 3.2)*

supersaturated solution A solution that contains more than the maximum solute that can dissolve at a given temperature. *(Sec. 13.7)*

surface tension The resistance of a liquid to spread out, and its tendency to form spherical drops with minimum surface area. *(Sec. 11.3)*

T

temperature The average energy of molecules in motion. *(Sec. 2.9)*

ternary ionic compound A compound that contains three elements, including at least one metal. *(Sec. 6.1)*

ternary oxyacid A compound that contains hydrogen, a nonmetal, and oxygen dissolved in water. *(Sec. 6.1)*

theoretical yield The amount of product that is calculated to be obtained from given amounts of reactants. *(Sec. 9.9)*

titration A laboratory procedure for delivering a measured volume of solution using a buret. *(Sec. 14.5)*

torr A unit of gas pressure equal to 1 mm Hg. *(Sec. 10.2)*

total ionic equation A chemical equation that shows ionized substances in solution as ions and slightly ionized substances as nonionized. *(Sec. 14.11)*

transcription The process by which a DNA molecule synthesizes a complementary single strand of RNA. *(Sec. 20.6)*

transition elements The Group B (3–12) elements in the periodic table. *(Sec. 5.3)*

transition state The highest point on the reaction profile at which there is the greatest potential energy. *(Sec. 16.2)*

translation The process by which an RNA codon specifies which amino acid is added to a growing protein chain. *(Sec. 20.6)*

transmutation The conversion of one element to another by inducing a nuclear reaction using particle bombardment. *(Sec. 18.6)*

transuranium elements The elements with atomic numbers beyond 92. All of the elements following uranium are synthetic. *(Sec. 5.3)*

triglyceride A lipid that has three fatty acids joined to glycerol by ester bonds; also called a *triacylglycerol*. *(Sec. 20.5)*

triple bond A bond composed of three electron pairs shared between two atoms. A triple bond is represented by three dashes between the symbols of the two atoms. *(Sec. 12.4)*

tritium ($_1^3$H) The nuclide of hydrogen with two neutrons in the nucleus. *(Sec. 18.8)*

Tyndall effect The phenomenon of scattering a beam of light by colloid particles. *(Sec. 13.3)*

U

uncertainty The degree of inexactness in a measurement obtained from an instrument. *(PSS.1)*

uncertainty principle The statement that it is impossible to precisely measure both the location and energy of a particle at the same time. *(Sec. 4.11)*

unit analysis method A systematic procedure for solving problems that converts the units in a given value to the units in the answer; also referred to as *dimensional analysis* or *factor label method. (Sec. 2.3)*

unit equation A statement that relates two values that are equivalent; for example, 1 ft = 12 in. and 1 in. = 2.54 cm. *(Sec. 2.2)*

unit factor A ratio of two quantities that are equivalent and can be applied to convert from one unit to another; for example, 1 m/100 cm. *(Sec. 2.2)*

unsaturated hydrocarbon A hydrocarbon that contains a carbon-carbon double bond or triple bond. *(Sec. 19.1)*

unsaturated solution A solution that contains less than the maximum solute that can dissolve at a given temperature. *(Sec. 13.7)*

V

vacuum A volume of space that does not contain gas molecules. *(Sec. 10.2)*

valence electrons The electrons that occupy the outermost *s* and *p* sublevels of an atom. *(Sec. 5.7)* The electrons in the outermost *s* and *p* sublevels of an atom that undergo reaction and form chemical bonds. *(Sec. 12.1)*

valence shell electron pair repulsion See *VSEPR theory.*

vapor pressure The pressure exerted by gaseous vapor above a liquid in a sealed container when the rates of evaporation and condensation are equal. *(Secs. 10.8 and 11.3)*

viscosity The resistance of a liquid to flow. *(Sec. 11.3)*

visible spectrum A range of light energy observed as violet, blue, green, yellow, orange, or red; the wavelengths of light from 400–700 nm. *(Sec. 4.6)*

voltaic cell An electrochemical cell in which a spontaneous redox reaction occurs and generates electrical energy; also called a *galvanic cell*. *(Sec. 17.6)*

volume by displacement A technique for determining the volume of a solid or a gas by measuring the volume of water it displaces. *(Sec. 2.7)* A technique for determining the volume of a gas by measuring the volume of water it displaces. *(Sec. 10.9)*

volume–volume problem A type of stoichiometry calculation that relates the volumes of two gases (at the same temperature and pressure) according to a balanced equation. *(Sec. 9.3)*

VSEPR theory A model that explains the shapes of molecules as a result of bonding and nonbonding electron pairs around the central atom repelling each other. *(Sec. 12.10)*

W

water of hydration The number of water molecules bound to a formula unit in a hydrate; also called the *water of crystallization. (Sec. 11.10)*

wavelength The distance a light wave travels to complete one cycle. *(Sec. 4.6)*

wax A lipid that contains a fatty acid joined to an alcohol by an ester linkage. *(Sec. 20.5)*

weak electrolyte An aqueous solution that is a poor conductor of electricity; for example, weak acids, weak bases, and slightly soluble salts. *(Sec. 14.10)*

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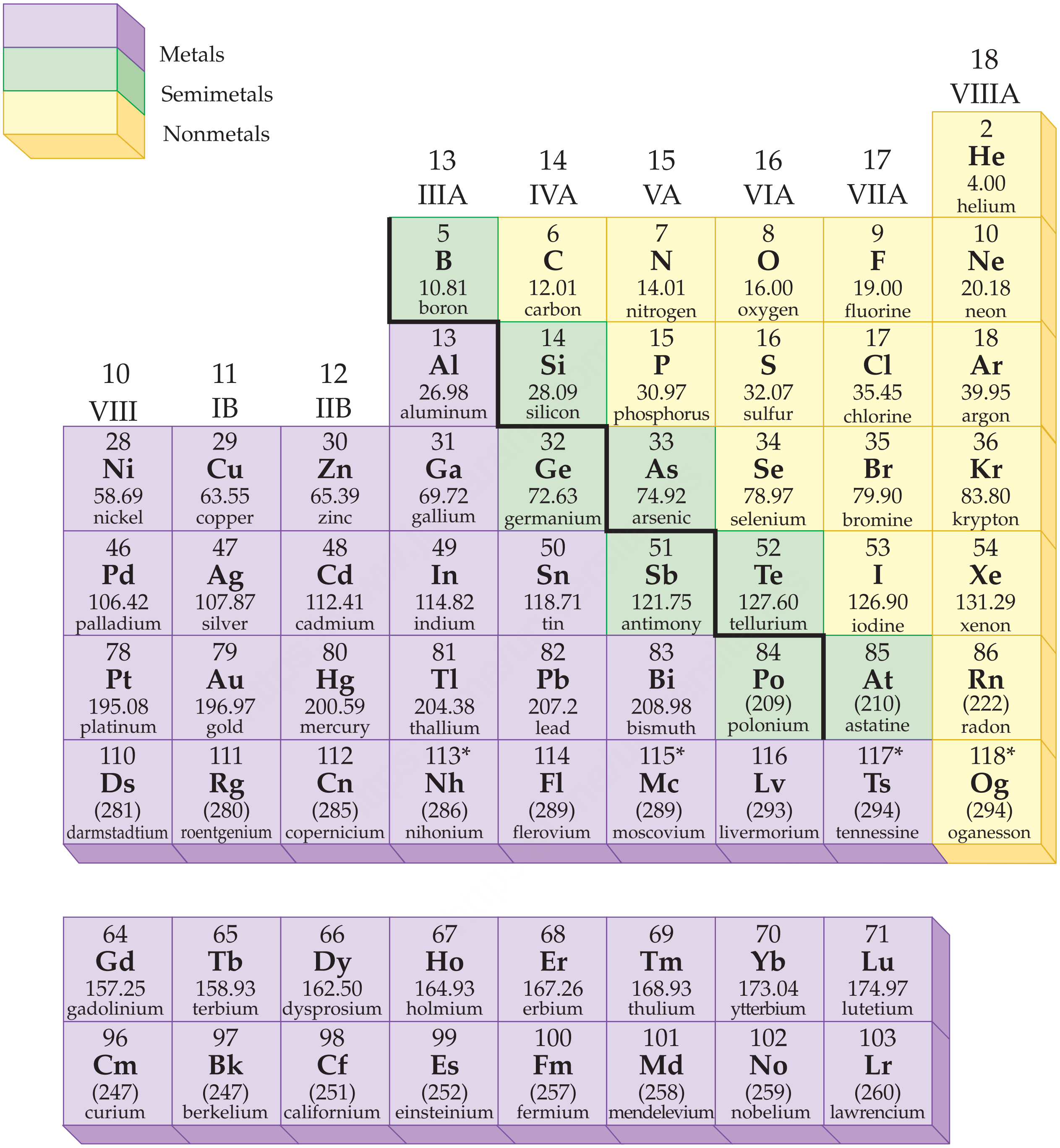
Periodic Table of the Elements

1	Atomic number
H	Element symbol
1.01	Atomic mass†
hydrogen	Element name

PERIOD	GROUP 1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII
1	1 **H** 1.01 hydrogen								
2	3 **Li** 6.94 lithium	4 **Be** 9.01 beryllium							
3	11 **Na** 22.99 sodium	12 **Mg** 24.31 magnesium							
4	19 **K** 39.10 potassium	20 **Ca** 40.08 calcium	21 **Sc** 44.96 scandium	22 **Ti** 47.88 titanium	23 **V** 50.94 vanadium	24 **Cr** 52.00 chromium	25 **Mn** 54.94 manganese	26 **Fe** 55.85 iron	27 **Co** 58.93 cobalt
5	37 **Rb** 85.47 rubidium	38 **Sr** 87.62 strontium	39 **Y** 88.91 yttrium	40 **Zr** 91.22 zirconium	41 **Nb** 92.91 niobium	42 **Mo** 95.95 molybdenum	43 **Tc** (99) technetium	44 **Ru** 101.07 ruthenium	45 **Rh** 102.91 rhodium
6	55 **Cs** 132.91 cesium	56 **Ba** 137.33 barium	57 **La** 138.91 lanthanum	72 **Hf** 178.49 hafnium	73 **Ta** 180.95 tantalum	74 **W** 183.85 tungsten	75 **Re** 186.21 rhenium	76 **Os** 190.2 osmium	77 **Ir** 192.22 iridium
7	87 **Fr** (223) francium	88 **Ra** (226) radium	89 **Ac** (227) actinium	104 **Rf** (261) rutherfordium	105 **Db** (262) dubnium	106 **Sg** (263) seaborgium	107 **Bh** (262) bohrium	108 **Hs** (265) hassium	109 **Mt** (266) meitnerium

Lanthanide series	58 **Ce** 140.12 cerium	59 **Pr** 140.91 praseodymium	60 **Nd** 144.24 neodymium	61 **Pm** (147) promethium	62 **Sm** 150.36 samarium	63 **Eu** 151.97 europium
Actinide series	90 **Th** (232) thorium	91 **Pa** (231) protactinium	92 **U** (238) uranium	93 **Np** (237) neptunium	94 **Pu** (244) plutonium	95 **Am** (243) americium

†The mass number of an important radioactive isotope—not the atomic mass—is shown in parentheses for an element with no stable isotopes.



Metals
Semimetals
Nonmetals

10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
								2 **He** 4.00 helium
			5 **B** 10.81 boron	6 **C** 12.01 carbon	7 **N** 14.01 nitrogen	8 **O** 16.00 oxygen	9 **F** 19.00 fluorine	10 **Ne** 20.18 neon
			13 **Al** 26.98 aluminum	14 **Si** 28.09 silicon	15 **P** 30.97 phosphorus	16 **S** 32.07 sulfur	17 **Cl** 35.45 chlorine	18 **Ar** 39.95 argon
28 **Ni** 58.69 nickel	29 **Cu** 63.55 copper	30 **Zn** 65.39 zinc	31 **Ga** 69.72 gallium	32 **Ge** 72.63 germanium	33 **As** 74.92 arsenic	34 **Se** 78.97 selenium	35 **Br** 79.90 bromine	36 **Kr** 83.80 krypton
46 **Pd** 106.42 palladium	47 **Ag** 107.87 silver	48 **Cd** 112.41 cadmium	49 **In** 114.82 indium	50 **Sn** 118.71 tin	51 **Sb** 121.75 antimony	52 **Te** 127.60 tellurium	53 **I** 126.90 iodine	54 **Xe** 131.29 xenon
78 **Pt** 195.08 platinum	79 **Au** 196.97 gold	80 **Hg** 200.59 mercury	81 **Tl** 204.38 thallium	82 **Pb** 207.2 lead	83 **Bi** 208.98 bismuth	84 **Po** (209) polonium	85 **At** (210) astatine	86 **Rn** (222) radon
110 **Ds** (281) darmstadtium	111 **Rg** (280) roentgenium	112 **Cn** (285) copernicium	113* **Nh** (286) nihonium	114 **Fl** (289) flerovium	115* **Mc** (289) moscovium	116 **Lv** (293) livermorium	117* **Ts** (294) tennessine	118* **Og** (294) oganesson

64 **Gd** 157.25 gadolinium	65 **Tb** 158.93 terbium	66 **Dy** 162.50 dysprosium	67 **Ho** 164.93 holmium	68 **Er** 167.26 erbium	69 **Tm** 168.93 thulium	70 **Yb** 173.04 ytterbium	71 **Lu** 174.97 lutetium
96 **Cm** (247) curium	97 **Bk** (247) berkelium	98 **Cf** (251) californium	99 **Es** (252) einsteinium	100 **Fm** (257) fermium	101 **Md** (258) mendelevium	102 **No** (259) nobelium	103 **Lr** (260) lawrencium

*The names and symbols for elements 113, 115, 117, and 118 are currently under review by IUPAC.

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Metric Prefixes

Prefix	Symbol	Multiple or Fraction
exa	E	1,000,000,000,000,000,000 = 1×10^{18}
peta	P	1,000,000,000,000,000 = 1×10^{15}
tera	T	1,000,000,000,000 = 1×10^{12}
giga	G	1,000,000,000 = 1×10^{9}
mega	M	1,000,000 = 1×10^{6}
kilo	k	1000 = 1×10^{3}
hecto	h	100 = 1×10^{2}
deca	da	10 = 1×10^{1}
deci	d	0.1 = 1×10^{-1}
centi	c	0.01 = 1×10^{-2}
milli	m	0.001 = 1×10^{-3}
micro	μ	0.000 001 = 1×10^{-6}
nano	n	0.000 000 001 = 1×10^{-9}
pico	p	0.000 000 000 001 = 1×10^{-12}
femto	f	0.000 000 000 000 001 = 1×10^{-15}
atto	a	0.000 000 000 000 000 001 = 1×10^{-18}

English–Metric Equivalents

Quantity	English Unit	Metric Equivalent
length	1 inch (in.)	1 in. = 2.54 cm
mass	1 pound (lb)	1 lb = 454 g
volume	1 quart (qt)	1 qt = 946 mL
time	1 second (sec)	1 sec = 1.00 s

Temperature Equivalents

$\frac{9}{5}\,^{\circ}\text{C} + 32 = {}^{\circ}\text{F}$

$^{\circ}\text{C} + 273 = \text{K}$

Energy Equivalents

1 calorie (cal) = 4.184 joules (J)

1 kilocalorie (kcal) = 1000 calories (cal)

Physical Constants

Avogadro's number = 6.02×10^{23}

Molar volume of a gas = 22.4 L @ STP

Common Monoatomic Cations

—*Stock System*

Cation	Name	Cation	Name
Al^{3+}	aluminum ion	Li^{+}	lithium ion
Ba^{2+}	barium ion	Mg^{2+}	magnesium ion
Cd^{2+}	cadmium ion	Mn^{2+}	manganese(II) ion
Ca^{2+}	calcium ion	Hg_2^{2+}	mercury(I) ion
Co^{2+}	cobalt(II) ion	Hg^{2+}	mercury(II) ion
Co^{3+}	cobalt(III) ion	Ni^{2+}	nickel(II) ion
Cu^{+}	copper(I) ion	K^{+}	potassium ion
Cu^{2+}	copper(II) ion	Ag^{+}	silver ion
Cr^{3+}	chromium(III) ion	Na^{+}	sodium ion
H^{+}	hydrogen ion	Sr^{2+}	strontium ion
Fe^{2+}	iron(II) ion	Sn^{2+}	tin(II) ion
Fe^{3+}	iron(III) ion	Sn^{4+}	tin(IV) ion
Pb^{2+}	lead(II) ion	Zn^{2+}	zinc ion
Pb^{4+}	lead(IV) ion		

—*Latin System*

Cation	Name	Cation	Name
Co^{2+}	cobaltous ion	Pb^{2+}	plumbous ion
Co^{3+}	cobaltic ion	Pb^{4+}	plumbic ion
Cu^{+}	cuprous ion	Hg_2^{2+}	mercurous ion
Cu^{2+}	cupric ion	Hg^{2+}	mercuric ion
Fe^{2+}	ferrous ion	Sn^{2+}	stannous ion
Fe^{3+}	ferric ion	Sn^{4+}	stannic ion

Common Monoatomic Anions

Anion	Name	Anion	Name
Br^-	bromide ion	N^{3-}	nitride ion
Cl^-	chloride ion	O^{2-}	oxide ion
F^-	fluoride ion	P^{3-}	phosphide ion
I^-	iodide ion	S^{2-}	sulfide ion

Common Polyatomic Cations

Cation	Name	Cation	Name
NH_4^+	ammonium ion	H_3O^+	hydronium ion

Common Polyatomic Anions

Anion	Name	Anion	Name
$C_2H_3O_2^-$	acetate ion	OH^-	hydroxide ion
CO_3^{2-}	carbonate ion	ClO^-	hypochlorite ion
ClO_3^-	chlorate ion	NO_3^-	nitrate ion
ClO_2^-	chlorite ion	NO_2^-	nitrite ion
CrO_4^{2-}	chromate ion	ClO_4^-	perchlorate ion
CN^-	cyanide ion	MnO_4^-	permanganate ion
$Cr_2O_7^{2-}$	dichromate ion	PO_4^{3-}	phosphate ion
HCO_3^-	hydrogen carbonate ion	SO_4^{2-}	sulfate ion
HSO_4^-	hydrogen sulfate ion	SO_3^{2-}	sulfite ion