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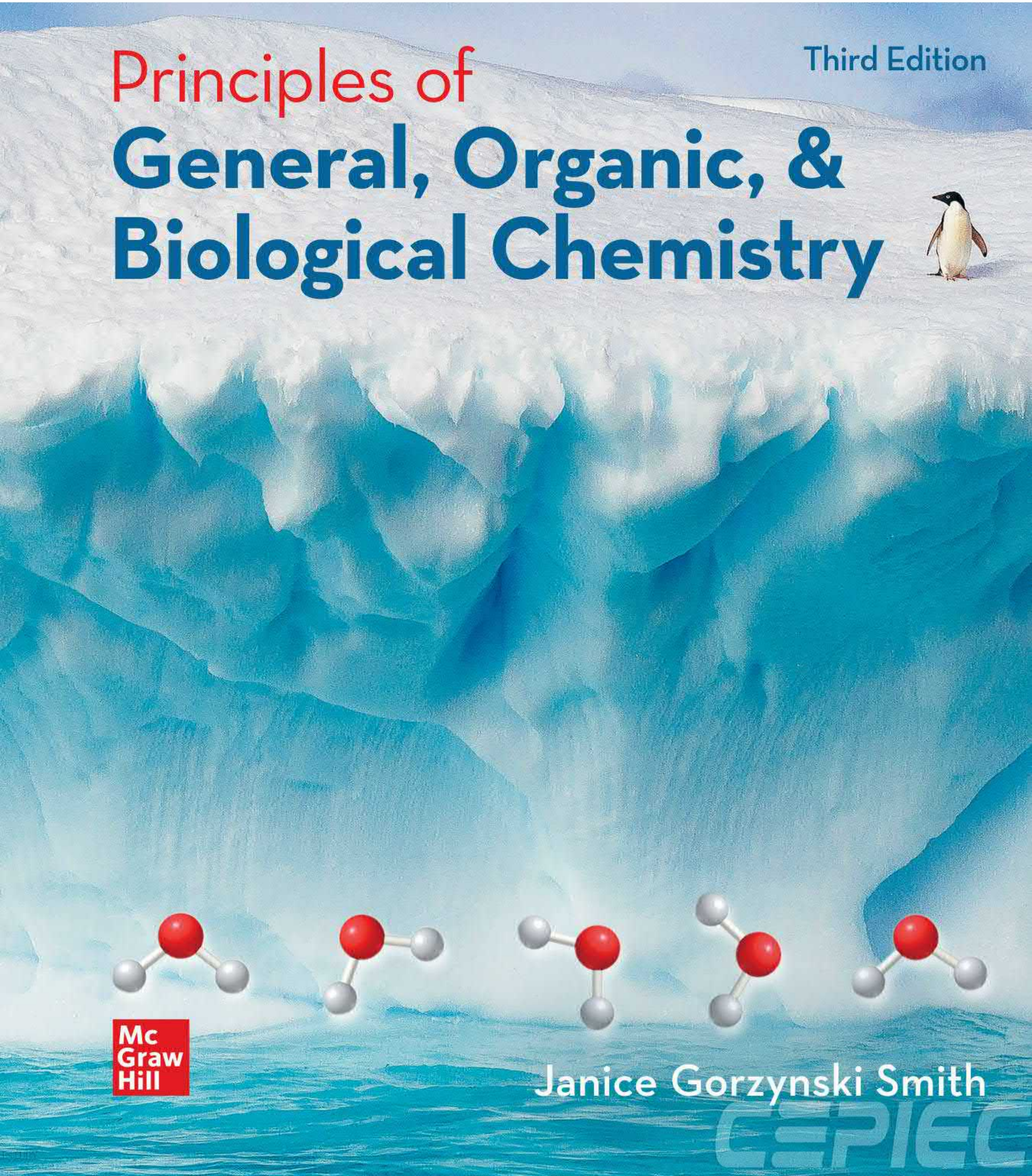

Third Edition

Principles of General, Organic, & Biological Chemistry

Mc Graw Hill

Janice Gorzynski Smith

PRINCIPLES of
General, Organic, & Biological Chemistry

Third Edition

Janice Gorzynski Smith
University of Hawai'i at Mānoa

Mc Graw Hill

PRINCIPLES OF GENERAL, ORGANIC, & BIOLOGICAL CHEMISTRY

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About the Author


Daniel C. Smith

Janice Gorzynski Smith was born in Schenectady, New York. She received an A.B. degree *summa cum laude* in chemistry at Cornell University and a Ph.D. in Organic Chemistry from Harvard University under the direction of Nobel Laureate E. J. Corey. During her tenure with the Corey group she completed the total synthesis of the plant growth hormone gibberellic acid.

Following her postdoctoral work, Jan joined the faculty of Mount Holyoke College, where she was employed for 21 years. During this time, she was active in teaching organic chemistry lecture and lab courses, conducting a research program in organic synthesis, and serving as department chair. Her organic chemistry class was named one of Mount Holyoke's "Don't-miss courses" in a survey by *Boston* magazine. After spending two sabbaticals amidst the natural beauty and diversity of Hawai'i in the 1990s, Jan and her family moved there permanently in 2000. She has been a faculty member at the University of Hawai'i at Mānoa, where she has taught a one-semester organic and biological chemistry course for nursing students, as well as the two-semester organic chemistry lecture and lab courses. In 2003, she received the Chancellor's Citation for Meritorious Teaching.

Jan resides in Hawai'i with her husband Dan, an emergency medicine physician, pictured with her at an event on Oahu. She has four children and nine grandchildren. When not teaching, writing, or enjoying her family, Jan bikes, hikes, snorkels, and scuba dives in sunny Hawai'i, and time permitting, enjoys travel and Hawaiian quilting.

To my family

Contents in Brief

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Appendices

Preface

This textbook is written for students who have an interest in nursing, nutrition, environmental science, food science, and a wide variety of other health-related professions. The content of this book is designed for an introductory chemistry course with no chemistry prerequisite and is suitable for either a one- or two-semester course. This text relates the principal concepts of general, organic, and biological chemistry to the world around us, and in this way illustrates how chemistry explains many aspects of daily life.

The learning style of today's students relies heavily on visual imagery. In this text, new concepts are introduced one at a time, keeping the basic themes in focus, and breaking down complex problems into manageable chunks of information. Relevant applications are provided for all basic chemical concepts, and molecular art illustrates and explains common everyday phenomena. Students learn step-by-step problem solving throughout the chapter within sample problems and *How To* boxes. Students are given enough detail to understand basic concepts so that they may acquire a new appreciation of the human body and the larger world around them.

New to This Edition

General

Problem Solving Sample Problems are now paired with Practice Problems to allow students to apply what they have just learned. The Practice Problems are followed by More Practice lists that point students to end-of-chapter problems that are similar in concept. Concept Check problems have replaced other in-chapter problems to give students an immediate check on the topic that has just been presented. The answers to all Practice Problems, all Concept Check problems, and odd-numbered end-of-chapter problems are given at the end of each chapter.

Chapter Review Chapter Review, which replaces Chapter Highlights at the end of each chapter, consists of Key Terms that are defined in the Glossary, Key Concepts, Key Equations, Key Reactions, and Key Skills. The Key Concepts and Key Skills sections use art and chemical structures to more clearly explain the key features detailed within the chapter. Key Skills, which presents the steps needed to solve important topics within the chapter, should be especially valuable for students learning stepwise processes.

Self-Test Each chapter contains a Self-Test, which consists of short-answer questions that test an understanding of definitions, equations, and other material encountered within the chapter. Answers to each question are provided at the end of the chapter.

Study Tips Brief Study Tips have been added to the margins in Chapters 1, 3, 6, 10, and 14 to help students develop general methods for solving recurrent types of problems, such as those that require a specific equation.

Photos Three-fourths of the chapter-opening photos have been replaced with photos emphasizing relevant material within the chapter. More marginal photos of applications have also been added on topics including non-contact thermometers (Chapter 1), radioactive seeds for cancer treatment (Chapter 9), leghemoglobin in plant-based burgers (Chapter 16), and many others.

Problems Over 150 new problems have been added.

Other New Coverage

Some of the new material added within specific chapters is listed below.

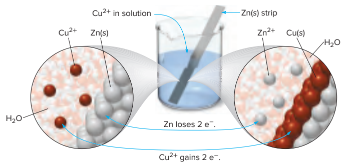
- Coverage on using a scientific calculator with scientific notation and logarithms has been expanded in Sections 1.6B and 8.5B. Tables with art that indicates what buttons should be pressed and what calculator displays will be shown are given.
- Chapter 3 opens by presenting a new current topic, the effect of sunscreens like oxybenzone on the bleaching of coral reefs. Recent research on oxybenzone is also discussed in Section 11.10.
- A new section on determining types of reactions—combination, decomposition, single replacement, and double replacement—has been added to Chapter 5. The chapter has been reorganized to place oxidation and reduction reactions immediately following this section, so that all of these different reaction types are in proximity.
- The discussion of dialysis and osmosis in Section 7.8 has been edited to emphasize the distinction between these related concepts. Three new problems on this subject have been added within the chapter.
- New material on using PET scans to visualize the brain in Alzheimer's patients has been added to Section 9.5.
- Material on methane, a greenhouse gas, has been expanded in Section 10.8.
- New material on the human milk oligosaccharides in breast milk has been added in Section 14.6D.
- Figure 15.1 now presents data on saturated fats, unsaturated oils, and trans fats in bar graph form for easier visualization of lipid content.
- The material on enzymes has been expanded into two sections (Sections 16.9 and 16.10), which include classes of enzymes, naming enzymes, and factors that affect enzyme activity.
- A section on the Human Genome Project (Section 17.10B) has been added.
- Section 17.11 on viruses has been expanded with material on coronaviruses and mRNA vaccines.

The Construction of a Learning System

Writing a textbook and its supporting learning tools is a multifaceted endeavor. McGraw Hill's development process is an ongoing, market-oriented approach to building accurate and innovative learning systems. It is dedicated to continual large scale and incremental improvement, driven by multiple customer feedback loops and checkpoints. This is initiated during the early planning stages of new products and intensifies during the development and production stages, and then begins again upon publication, in anticipation of the next version of each print and digital product. This process is designed to provide a broad, comprehensive spectrum of feedback for refinement and innovation of learning tools for both student and instructor. The development process includes market research, content reviews, faculty and student focus groups, course- and product-specific symposia, accuracy checks, and art reviews.

The Learning System Used in *Principles of General, Organic, & Biological Chemistry,* Third Edition

Figure 5.2 A Redox Reaction—The Transfer of Electrons from Zn to Cu^{2+}



A redox reaction occurs when a strip of Zn metal is placed in a solution of Cu^{2+} ions. In this reaction, Zn loses two electrons to form Zn^{2+}, which goes into solution. Cu^{2+} gains two electrons to form Cu metal, which precipitates out of solution, forming a coating on the zinc strip.

CONSUMER NOTE

Benzoyl peroxide ($C_{14}H_{10}O_4$) is the active ingredient in several acne medications. Benzoyl peroxide kills bacteria by oxidation reactions. *Jill Braaten*

Each of these processes can be written as individual reactions, called **half reactions,** to emphasize which electrons are gained and lost.

Oxidation half reaction: $Zn \longrightarrow Zn^{2+} + 2\,e^-$ (Loss of electrons = **oxidation**)

Reduction half reaction: $Cu^{2+} + 2\,e^- \longrightarrow Cu$ (Gain of electrons = **reduction**)

- **A compound that gains electrons (is reduced) while causing another compound to be oxidized is called an *oxidizing agent.***
- **A compound that loses electrons (is oxidized) while causing another compound to be reduced is called a *reducing agent.***

In this example, Zn loses electrons to Cu^{2+}. We can think of Zn as a **reducing agent** because it causes Cu^{2+} to gain electrons and become reduced. We can think of Cu^{2+} as an **oxidizing agent** because it causes Zn to lose electrons and become oxidized.

To draw the products of an oxidation–reduction reaction, we must decide which element or ion gains electrons and which element or ion loses electrons. Use the following guidelines.

- **When considering neutral atoms, metals *lose* electrons and nonmetals *gain* electrons.**
- **When considering ions, cations tend to *gain* electrons and anions tend to *lose* electrons.**

Thus, the metals sodium (Na) and magnesium (Mg) readily lose electrons to form the cations Na^+ and Mg^{2+}, respectively; that is, they are oxidized. The nonmetals O_2 and Cl_2 readily gain electrons to form $2\,O^{2-}$ and $2\,Cl^-$, respectively; that is, they are reduced. A positively charged ion like Cu^{2+} is reduced to Cu by gaining two electrons, whereas two negatively charged Cl^- anions are oxidized to Cl_2 by losing two electrons. These reactions and additional examples are shown in Figure 5.3.

Writing Style

A succinct writing style weaves together key points of general, organic, and biological chemistry, along with attention-grabbing applications to consumer, environmental, and health-related fields. Concepts and topics are broken into small chunks of information that are more easily learned.

Chapter Goals, Tied to End-of-Chapter Review

Chapter Goals at the beginning of each chapter identify what students will learn, and are tied to the end-of-chapter Key Concepts and Key Skills, which serve as bulleted summaries of the most important concepts for study.

CHAPTER OUTLINE

CHAPTER GOALS

In this chapter you will learn how to:

1. Describe the different types of radiation emitted by a radioactive nucleus
2. Write equations for nuclear reactions
3. Define half-life
4. Recognize the units used for measuring radioactivity
5. Give examples of common radioisotopes used in medical diagnosis and treatment
6. Describe the general features of nuclear fission and nuclear fusion
7. Describe the features of medical imaging techniques that do not use radioactivity

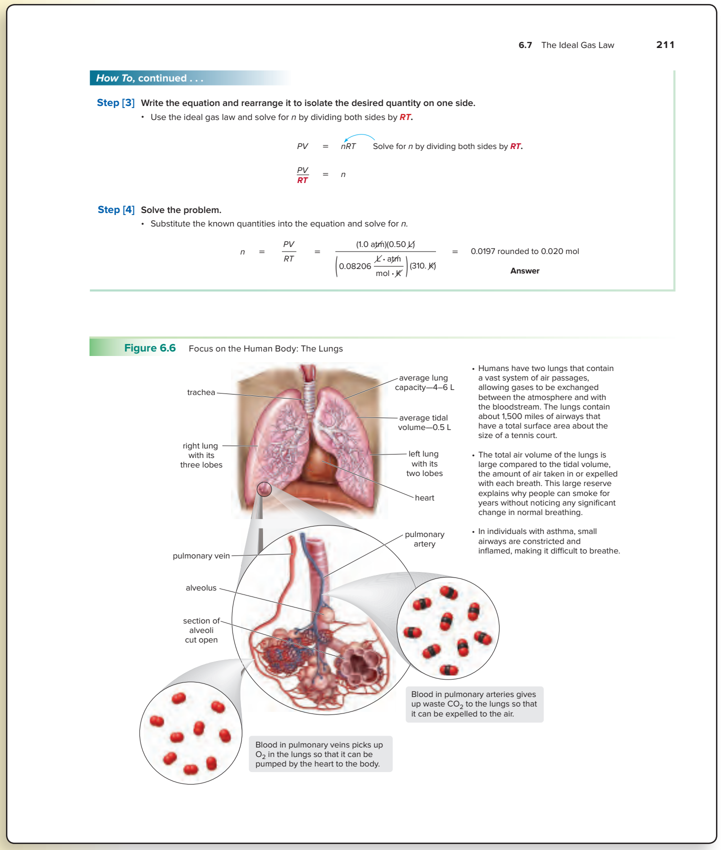
KEY CONCEPTS

1 Types of nuclear radiation (9.1)

1 Alpha particle	2 Beta particle	3 Positron	4 Gamma ray
α or ${}^{4}_{2}He$	β or ${}^{\ 0}_{-1}e$	β^+ or ${}^{\ 0}_{+1}e$	γ
• A high-energy nucleus that contains two protons and two neutrons	• A high-energy electron that has a −1 charge and negligible mass	• An antiparticle of a β particle that has a +1 charge and negligible mass	• High-energy radiation with no mass or charge

2 Nuclear fission and nuclear fusion (9.6)

1 Nuclear fission	2 Nuclear fusion
${}^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{91}_{36}Kr + {}^{142}_{56}Ba + 3\,{}^{1}_{0}n$	${}^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n$
• Nuclear fission is the splitting apart of a heavy nucleus into lighter nuclei and neutrons. • Nuclear fission releases a great deal of energy. • Fission is used in nuclear power plants to generate electricity.	• Nuclear fusion is the joining together of two light nuclei to form a larger nucleus. • Nuclear fusion releases a great deal of energy. • Nuclear fusion occurs in stars.



How To, continued . . .

Step [3] Write the equation and rearrange it to isolate the desired quantity on one side.

- Use the ideal gas law and solve for *n* by dividing both sides by ***RT***.

$$PV = nRT \quad \text{Solve for } n \text{ by dividing both sides by } RT.$$

$$\frac{PV}{RT} = n$$

Step [4] Solve the problem.

- Substitute the known quantities into the equation and solve for *n*.

$$n = \frac{PV}{RT} = \frac{(1.0\ \text{atm})(0.50\ \text{L})}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(310.\ \text{K})} = 0.0197 \text{ rounded to } 0.020 \text{ mol}$$

Answer

Figure 6.6 Focus on the Human Body: The Lungs

- Humans have two lungs that contain a vast system of air passages, allowing gases to be exchanged between the atmosphere and with the bloodstream. The lungs contain about 1,500 miles of airways that have a total surface area about the size of a tennis court.
- The total air volume of the lungs is large compared to the tidal volume, the amount of air taken in or expelled with each breath. This large reserve explains why people can smoke for years without noticing any significant change in normal breathing.
- In individuals with asthma, small airways are constricted and inflamed, making it difficult to breathe.

Macro-to-Micro Illustrations

Visualizing molecular-level representations of macroscopic phenomena is critical to the understanding of any chemistry course. Many illustrations in this text include photos or drawings of everyday objects, paired with their molecular representation, to help students visualize and understand the chemistry behind ordinary things. Many illustrations of the human body include magnifications for specific anatomic regions, as well as representations at the microscopic level, for today's visual learners.

Applications

Relevant, interesting applications of chemistry to everyday life are included for all basic chemical concepts. These are interspersed in margin-placed Health Notes, Consumer Notes, and Environmental Notes, as well as sections entitled "Focus on Health & Medicine," "Focus on the Environment," and "Focus on the Human Body."

HEALTH NOTE

Individuals with cystic fibrosis, the most common genetic disease in Caucasians, produce thick mucus in the lungs, resulting in a higher-than-normal level of CO_2 and respiratory acidosis.

HEALTH NOTE

During strenuous exercise, the lungs expel more CO_2 than usual and the pH of the blood increases.
CS Productions/Jupiterimages/Brand X Pictures/Alamy Stock Photo

8.9 FOCUS ON THE HUMAN BODY Buffers in the Blood

The normal blood pH of a healthy individual is in the range of 7.35 to 7.45. A pH above or below this range is generally indicative of an imbalance in respiratory or metabolic processes. The body is able to maintain a very stable pH because the blood and other tissues are buffered. The principal buffer in the blood is carbonic acid/bicarbonate (H_2CO_3/HCO_3^-).

In examining the carbonic acid/bicarbonate buffer system in the blood, two reactions are important. First of all, carbonic acid (H_2CO_3) is formed from CO_2 dissolved in the bloodstream (Section 8.6). Second, because carbonic acid is a weak acid, it is also dissociated in water to form its conjugate base, bicarbonate (HCO_3^-). Bicarbonate is also generated in the kidneys.

$$CO_2(g) + H_2O(l) \rightleftharpoons \underset{\text{carbonic acid}}{H_2CO_3(aq)} \overset{H_2O}{\rightleftharpoons} H_3O^+(aq) + \underset{\text{bicarbonate}}{HCO_3^-(aq)}$$

principal buffer in the blood

CO_2 is constantly produced by metabolic processes in the body and then transported to the lungs to be eliminated. Le Châtelier's principle explains the effect of increasing or decreasing the level of dissolved CO_2 on the pH of the blood. A higher-than-normal CO_2 concentration shifts the equilibrium to the right, increasing the H_3O^+ concentration and lowering the pH. **Respiratory acidosis** results when the body fails to eliminate adequate amounts of CO_2 through the lungs. This may occur in patients with advanced lung disease or respiratory failure.

A lower respiratory rate **increases [CO_2]**.

$$CO_2(g) + 2\,H_2O(g) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

This drives the reaction to the right, **increasing [H_3O^+]**.
Blood has a higher [H_3O^+] ⟶ **lower pH**

A lower-than-normal CO_2 concentration shifts the equilibrium to the left, decreasing the H_3O^+ concentration and raising the pH. **Respiratory alkalosis** is caused by hyperventilation, very rapid breathing that occurs when an individual experiences excitement or panic.

$$CO_2(g) + 2\,H_2O(g) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

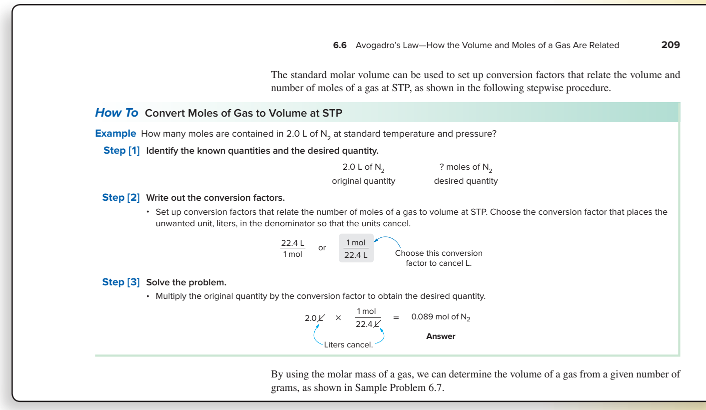
A faster respiratory rate **decreases [CO_2]**.

This drives the reaction to the left, **decreasing [H_3O^+]**.
Blood has a lower [H_3O^+] ⟶ **higher pH**

The pH of the blood may also be altered when the metabolic processes of the body are not in balance. **Metabolic acidosis** results when excessive amounts of acid are produced and the blood pH falls. This may be observed in patients with severe infections (sepsis). It may also occur in poorly controlled diabetes. **Metabolic alkalosis** may occur when recurrent vomiting decreases the amount of acid in the stomach, thus causing a rise in pH.

How To's

Key processes are taught to students in a straightforward and easy-to-understand manner by using examples and multiple, detailed steps to solving problems.



6.6 Avogadro's Law—How the Volume and Moles of a Gas Are Related 209

The standard molar volume can be used to set up conversion factors that relate the volume and number of moles of a gas at STP, as shown in the following stepwise procedure.

How To **Convert Moles of Gas to Volume at STP**

Example How many moles are contained in 2.0 L of N_2 at standard temperature and pressure?

Step [1] **Identify the known quantities and the desired quantity.**

2.0 L of N_2 (original quantity) — ? moles of N_2 (desired quantity)

Step [2] **Write out the conversion factors.**

- Set up conversion factors that relate the number of moles of a gas to volume at STP. Choose the conversion factor that places the unwanted unit, liters, in the denominator so that the units cancel.

$$\frac{22.4\ L}{1\ mol} \quad or \quad \frac{1\ mol}{22.4\ L}$$

Choose this conversion factor to cancel L.

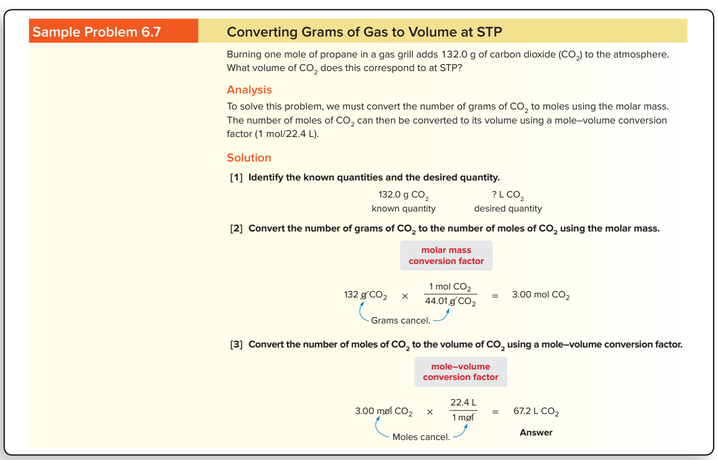
Step [3] **Solve the problem.**

- Multiply the original quantity by the conversion factor to obtain the desired quantity.

$$2.0\ L \times \frac{1\ mol}{22.4\ L} = 0.089\ mol\ of\ N_2$$

Liters cancel. **Answer**

By using the molar mass of a gas, we can determine the volume of a gas from a given number of grams, as shown in Sample Problem 6.7.



Sample Problem 6.7 **Converting Grams of Gas to Volume at STP**

Burning one mole of propane in a gas grill adds 132.0 g of carbon dioxide (CO_2) to the atmosphere. What volume of CO_2 does this correspond to at STP?

Analysis

To solve this problem, we must convert the number of grams of CO_2 to moles using the molar mass. The number of moles of CO_2 can then be converted to its volume using a mole–volume conversion factor (1 mol/22.4 L).

Solution

[1] Identify the known quantities and the desired quantity.

132.0 g CO_2 (known quantity) — ? L CO_2 (desired quantity)

[2] Convert the number of grams of CO_2 to the number of moles of CO_2 using the molar mass.

molar mass conversion factor

$$132\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.01\ g\ CO_2} = 3.00\ mol\ CO_2$$

Grams cancel.

[3] Convert the number of moles of CO_2 to the volume of CO_2 using a mole–volume conversion factor.

mole–volume conversion factor

$$3.00\ mol\ CO_2 \times \frac{22.4\ L}{1\ mol} = 67.2\ L\ CO_2$$

Moles cancel. **Answer**

Problem Solving

Stepwise sample problems lead students through the thought process tied to successful problem solving by employing *Analysis* and *Solution* steps. Sample Problems are categorized sequentially by topic to match chapter organization, and are paired with practice problems to allow students to apply what they have just learned. Students can immediately verify their answers to the follow-up problems in the answers at the end of each chapter.

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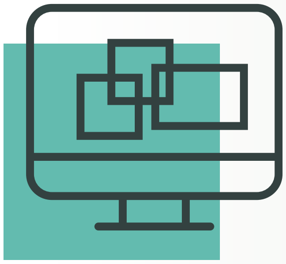

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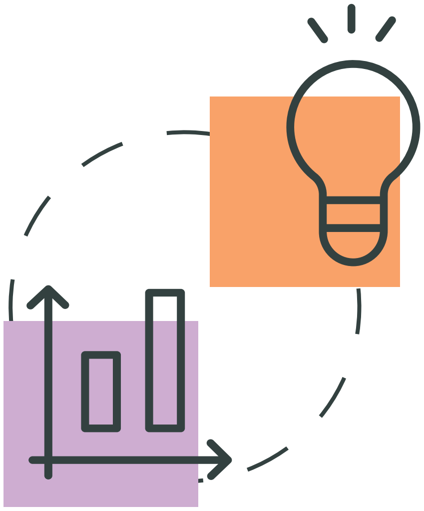

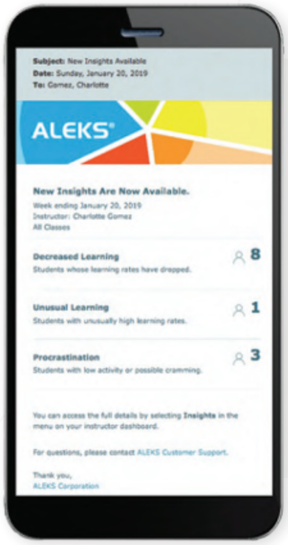

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- **Presentation Tools Table:** Instructors have access to fully editable **accessible PowerPoint lecture outlines,** which appear as ready-made presentations that combine art and lecture notes for each chapter of the text. For instructors who prefer to create their lectures from scratch, all illustrations, photos, and tables are pre-inserted by chapter into blank **PowerPoint slides** and are also available as **downloadable jpeg files.**
- **Instructor's Solutions Manual:** This supplement contains complete, worked out solutions for all the end-of-chapter problems in the text.
- **Computerized Test Bank:** Over 1,800 test questions that accompany *Principles of General, Organic, & Biological Chemistry* are available for creating exams or quizzes.
- **Videos and Animations:** More than 90 videos and animations available through the eBook and the instructor resource center, supplement the textbook material in much the same way as instructor demonstrations. However, they are only a few mouse-clicks away, any time, day or night. Because many students are visual learners, the animations add another dimension of learning; they bring a greater degree of reality to the written word.

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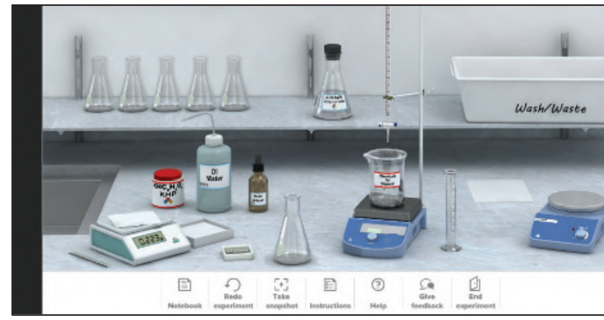

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Student Solutions Manual

The *Student Solutions Manual* contains the solutions to all in-chapter problems as well as the solutions to all odd-numbered end-of-chapter problems.

Acknowledgments

When I first began textbook writing 20 years ago, I had no idea how many people I would have to rely upon to see a project from manuscript preparation to published text. Special thanks for this edition go to Senior Product Developer Mary Hurley, Senior Core Content Project Manager Laura Bies, and freelance Developmental Editor John Murdzek, who handled all the day-to-day steps needed to publish an accurate and engaging text. Thanks are also due to Executive Portfolio Manager Michelle Hentz and Associate Portfolio Manager Hannah Downing for spearheading the revision of *Principles,* and the art, production, marketing, and sales teams for their support and contributions.

I would also like to acknowledge the following individuals for their masterful authoring of the ancillaries to accompany the third edition: Lauren McMills of Ohio University for the Solutions Manuals; Andrea Leonard of the University of Louisiana, Lafayette for the Accessible PowerPoint Lecture Outlines; and Cari Gigliotti of Sinclair Community College for the Test Bank.

Finally, I thank my family for their support and patience during the long process of publishing a textbook. My husband Dan, an emergency medicine physician, took several photos that appear in the text, and served as a consultant for many medical applications.

Matter and Measurement

1

Determining the weight and length of a newborn are common measurements performed by healthcare professionals.
Daniel C. Smith

CHAPTER OUTLINE

1.1 Chemistry—The Science of Everyday Experience
1.2 States of Matter
1.3 Classification of Matter
1.4 Measurement
1.5 Significant Figures
1.6 Scientific Notation
1.7 Problem Solving Using Conversion Factors
1.8 FOCUS ON HEALTH & MEDICINE: Problem Solving Using Clinical Conversion Factors
1.9 Temperature
1.10 Density and Specific Gravity

CHAPTER GOALS

In this chapter you will learn how to:

1. Describe the three states of matter
2. Classify matter as a pure substance, mixture, element, or compound
3. Report measurements using the metric units of length, mass, and volume
4. Use significant figures
5. Use scientific notation for very large and very small numbers
6. Use conversion factors to convert one unit to another
7. Convert temperature from one scale to another
8. Define density and specific gravity and use density to calculate the mass or volume of a substance

Why Study ...

Matter and Measurement?

Everything you touch, feel, or taste is composed of chemicals—that is, **matter**—so an understanding of its composition and properties is crucial to our appreciation of the world around us. Some matter—lakes, trees, sand, and soil—is naturally occurring, whereas other examples of matter—aspirin, nylon fabric, plastic syringes, and vaccines—are made by humans. To understand the properties of matter, as well as how one form of matter is converted to another, we must also learn about measurements. Following a recipe, pumping gasoline, and figuring out drug dosages involve manipulating numbers. Thus, Chapter 1 begins our study of chemistry by examining the key concepts of matter and measurement.

1.1 Chemistry—The Science of Everyday Experience

What activities might occupy the day of a typical student? You may have done some or all of the following tasks: eaten some meals, drunk coffee or cola, showered with soap, checked email on a computer, ridden a bike or car to a part-time job, taken an aspirin to relieve a headache, and spent some of the evening having snacks and refreshments with friends. Perhaps, without your awareness, your life was touched by chemistry in each of these activities. What, then, is this discipline we call **chemistry?**

- ***Chemistry*** **is the study of matter—its composition, properties, and transformations.**

What is **matter?**

- ***Matter*** **is anything that has mass and takes up volume.**

In other words, **chemistry studies anything that we touch, feel, see, smell, or taste,** from simple substances like water or salt, to complex substances like proteins and carbohydrates that combine to form the human body. Some matter—cotton, sand, an apple, and the cardiac drug digoxin—is **naturally occurring,** meaning it is isolated from natural sources. Other substances—nylon, Styrofoam, the plastic used in soft drink bottles, and the pain reliever ibuprofen—are **synthetic,** meaning they are produced by chemists in the laboratory (Figure 1.1).

Figure 1.1 Naturally Occurring and Synthetic Materials

a. Naturally occurring materials

b. Synthetic materials

Matter occurs in nature or is synthesized in the lab. (a) Sand and apples are two examples of natural materials. Cotton fabric is woven from cotton fiber, obtained from the cotton plant. The drug digoxin (trade name Lanoxin), widely prescribed for decades for patients with congestive heart failure, is extracted from the leaves of the woolly foxglove plant. (b) Nylon was the first synthetic fiber made in the laboratory. It quickly replaced the natural fiber silk in parachutes and ladies' stockings. Styrofoam and PET (polyethylene terephthalate), the plastic used for soft drink bottles, are strong yet lightweight synthetic materials used for food storage. Over-the-counter pain relievers like ibuprofen are synthetic. The starting materials for all of these useful products are obtained from petroleum.
(a)–(b): Jill Braaten/McGraw Hill

Figure 1.2

Transforming a Natural Material into a Useful Synthetic Product

a. b.

(a) Latex, the sticky liquid that oozes from a rubber tree when it is cut, is too soft for most applications. (b) Vulcanization converts latex to the stronger, elastic rubber used in tires and other products.
(a): Suphatthra China/Shutterstock; (b): Roy McMahon/Fuse/Getty Images

Sometimes a chemist studies what a substance is made of, whereas at other times, the focus may be how to convert one material into a new material with unique and useful properties. As an example, naturally occurring rubber exists as the sticky liquid latex, which is too soft for most applications. The laboratory process of vulcanization converts it to the stronger, more elastic material used in tires and other products (Figure 1.2).

Chemistry is truly the science of everyday experience. Soaps and detergents, newspapers and DVDs, condoms and oral contraceptives, Tylenol and penicillin—all of these items are products of chemistry. Without a doubt, advances in chemistry have transformed life in modern times.

1.2 States of Matter

Matter exists in three common states—solid, liquid, and gas.

- A *solid* has a definite volume, and maintains its shape regardless of the container in which it is placed. The particles of a solid lie close together, and are arranged in a regular three-dimensional array.
- A *liquid* has a definite volume, but takes on the shape of the container it occupies. The particles of a liquid are close together, but they can randomly move around, sliding past one another.
- A *gas* has no definite shape or volume. The particles of a gas move randomly and are separated by a distance much larger than their size. The particles of a gas expand to fill the volume and assume the shape of whatever container they are put in.

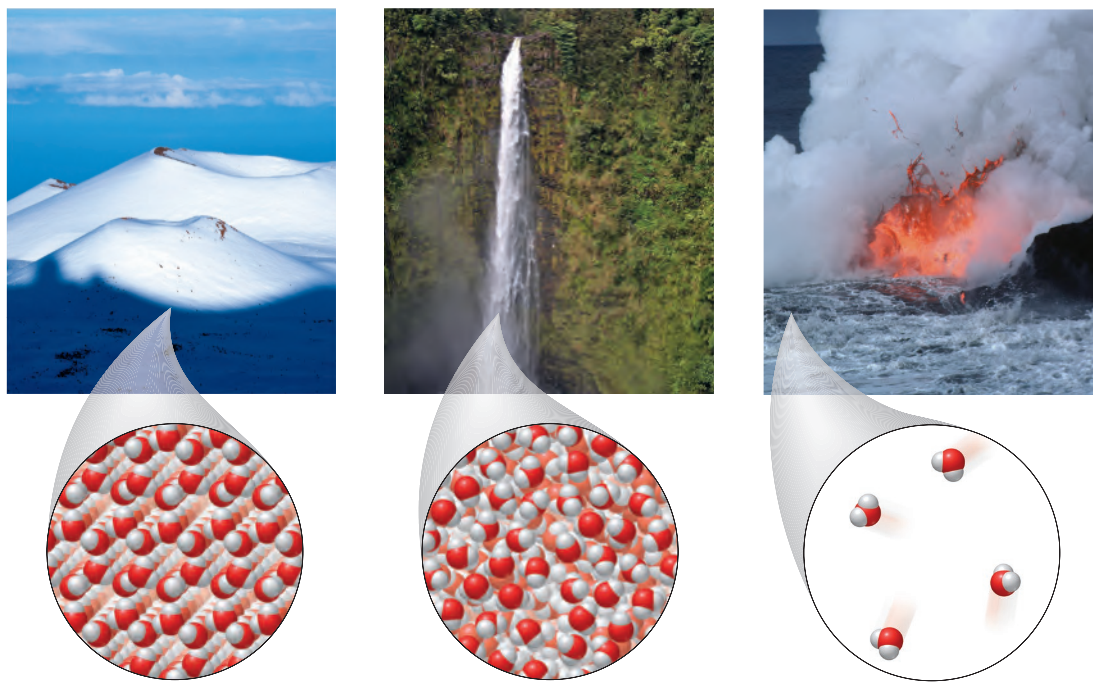
For example, water exists in its solid state as ice or snow, liquid state as liquid water, and gaseous state as steam or water vapor. Blow-up circles like those in Figure 1.3 will be used commonly in this text to indicate the composition and state of the particles that compose a substance. In this molecular art, different types of particles are shown in color-coded spheres, and the distance between the spheres signals its state—solid, liquid, or gas.

Matter is characterized by its **physical properties** and **chemical properties.**

- *Physical properties* are those that can be observed or measured without changing the composition of the material.

Figure 1.3 The Three States of Water—Solid, Liquid, and Gas

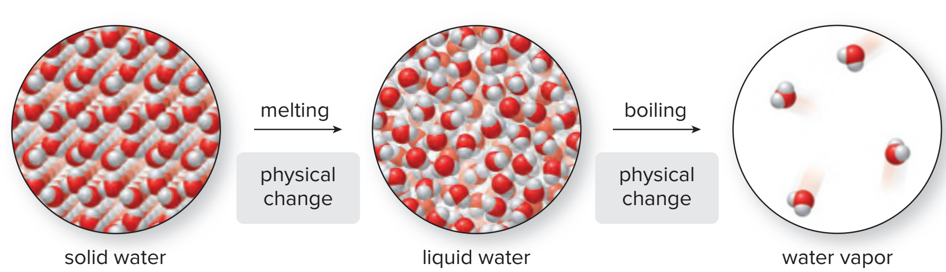
a. Solid water b. Liquid water c. Gaseous water



- The particles of a solid are close together and highly organized. (Photo: snow-capped Mauna Kea on the Big Island of Hawaii)
- The particles of a liquid are close together but more disorganized than the solid. (Photo: Akaka Falls on the Big Island of Hawaii)
- The particles of a gas are far apart and disorganized. (Photo: steam formed by a lava flow on the Big Island of Hawaii)

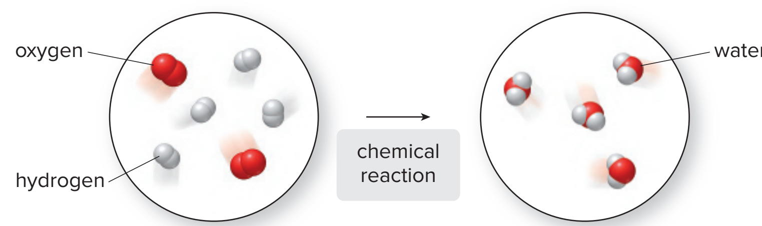
Each red sphere joined to two gray spheres represents a single water particle. In proceeding from left to right, from solid to liquid to gas, the molecular art shows that the level of organization of the water particles decreases. Color-coding and the identity of the spheres within the particles will be addressed in Chapter 2. *(a): Alvis Upitis/Getty Images; (b): Daniel C. Smith; (c): Source:T.J. Takahash/USGS*

Common physical properties include melting point (mp), boiling point (bp), solubility, color, and odor. **A *physical change* alters a substance without changing its composition.** The most common physical changes are **changes in state**—that is, the **conversion of matter from one state to another.** Melting an ice cube to form liquid water, and boiling liquid water to form steam are two examples of physical changes. Water is the substance at the beginning and end of both physical changes. More details about physical changes are discussed in Chapter 4.



- ***Chemical properties*** **are those that determine how a substance can be converted to another substance.**

A *chemical change*, or a *chemical reaction*, converts one material to another. The conversion of hydrogen and oxygen to water is a chemical reaction because the composition of the material is different at the beginning and end of the process. Chemical reactions are discussed in Chapter 5.

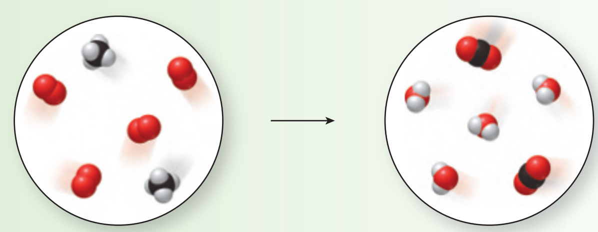


Concept Check 1.1

Characterize each process as a physical change or a chemical change: (a) making ice cubes; (b) burning natural gas; (c) silver jewelry tarnishing; (d) a pile of snow melting; (e) fermenting grapes to produce wine.

Concept Check 1.2

Does the molecular art represent a chemical change or a physical change? Explain your choice.



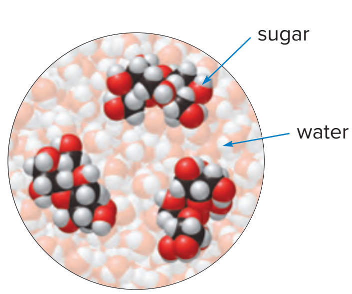
1.3 Classification of Matter

All matter can be classified as either a **pure substance** or a **mixture.**

- **A *pure substance* is composed of a single component and has a constant composition, regardless of the sample size and the origin of the sample.**

A pure substance, such as water or table sugar, can be characterized by its physical properties, because these properties do not change from sample to sample. **A pure substance cannot be broken down to other pure substances by any physical change.**

- **A *mixture* is composed of more than one substance. The composition of a mixture can vary depending on the sample.**



sugar dissolved in water

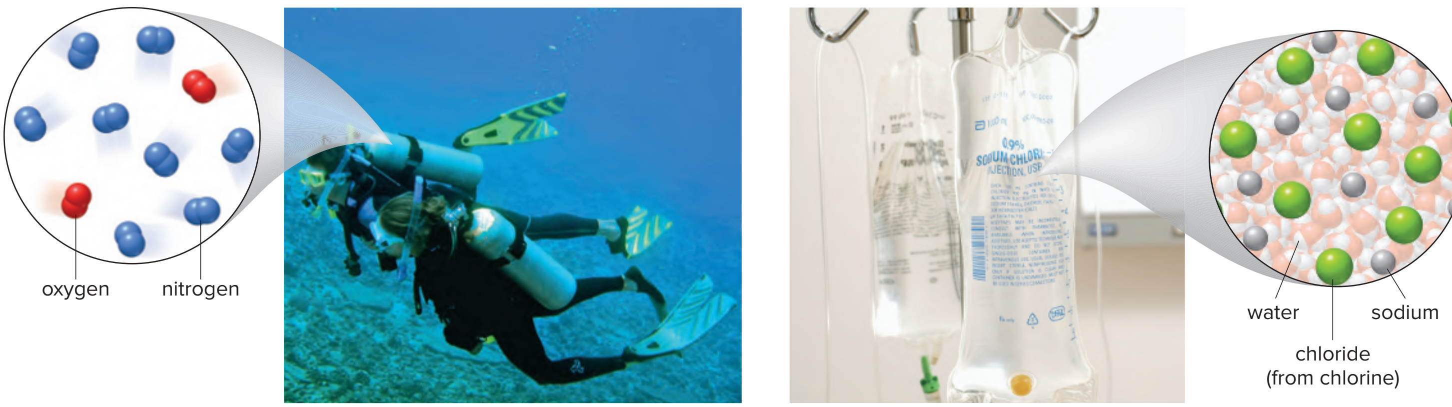
mixture

The physical properties of a mixture may also vary from one sample to another. **A mixture can be separated into its components by physical changes.** Dissolving table sugar in water forms a mixture, whose sweetness depends on the amount of sugar added. If the water is allowed to evaporate from the mixture, pure table sugar and pure water are obtained.

Mixtures can be formed from solids, liquids, and gases, as shown in Figure 1.4. The compressed air breathed by a scuba diver consists mainly of the gases oxygen and nitrogen. A saline solution used in an IV bag contains solid sodium chloride (table salt) dissolved in water.

Figure 1.4 Two Examples of Mixtures

a. Two gases b. A solid and a liquid



(a): Daniel C. Smith; (b): Janis Christie/Digital Vision/Alamy Stock Photo

A pure substance is classified as either an **element** or a **compound.**

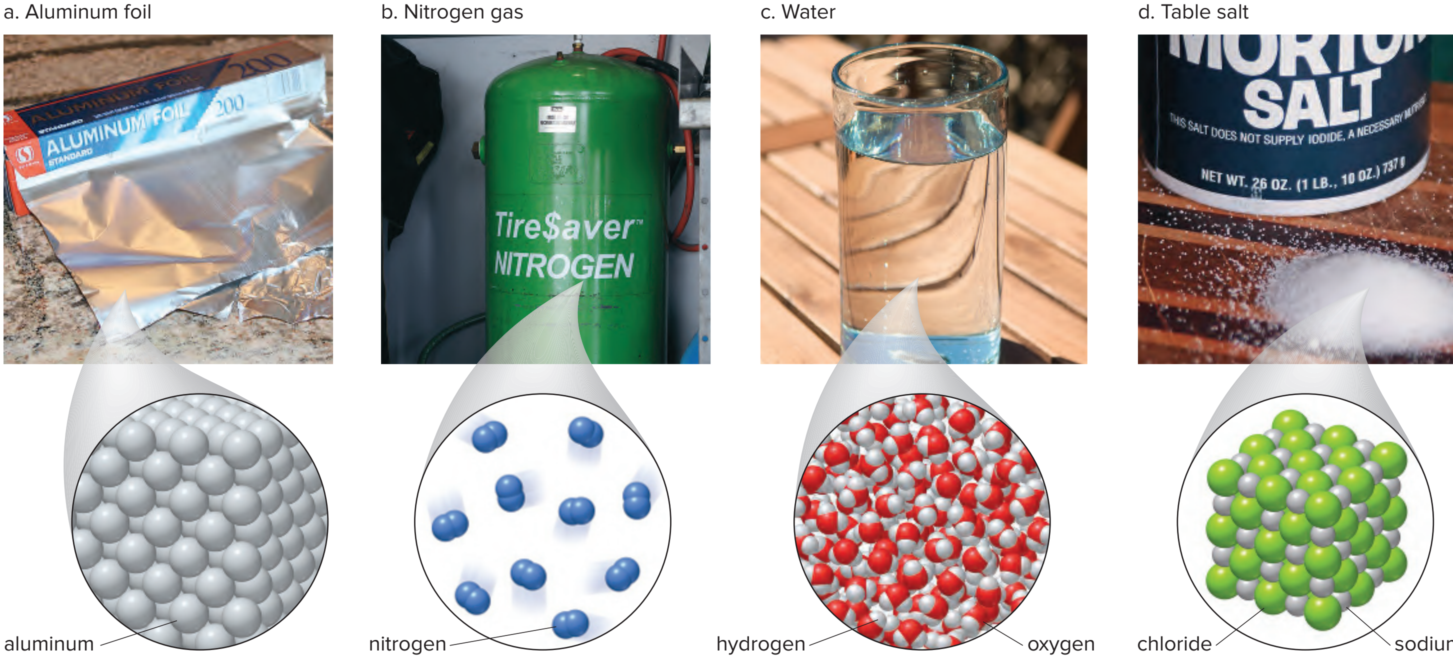
- **An *element* is a pure substance that cannot be broken down into simpler substances by a chemical reaction.**
- **A *compound* is a pure substance formed by chemically combining (joining together) two or more elements.**

An alphabetical list of elements is located in Appendix A. The elements are commonly organized into a periodic table, shown in Appendix B, and discussed in much greater detail in Chapter 2.

Nitrogen gas, aluminum foil, and copper wire are all elements. Water is a compound because it is composed of the elements hydrogen and oxygen. Table salt, sodium chloride, is also a compound because it is formed from the elements sodium and chlorine (Figure 1.5). Although

Figure 1.5 Elements and Compounds

a. Aluminum foil b. Nitrogen gas c. Water d. Table salt

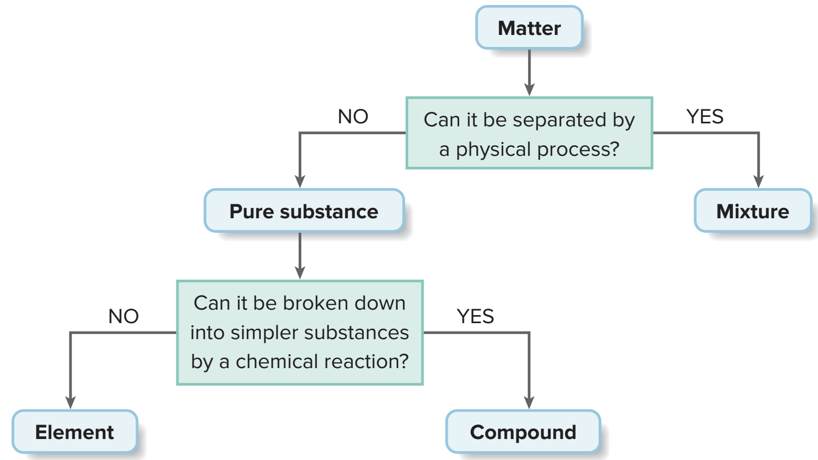


- Aluminum foil and nitrogen gas are elements. **The molecular art used for an element shows spheres of *one* color only.** Thus, aluminum is a solid shown with gray spheres, whereas nitrogen is a gas shown with blue spheres. Water and table salt are compounds. Color-coding of the spheres used in the molecular art indicates that water is composed of two elements—hydrogen shown as gray spheres and oxygen shown in red. Likewise, the gray (sodium) and green (chlorine) spheres illustrate that sodium chloride is formed from two elements as well.

(a): Daniel C. Smith; (b): Keith Eng, 2008; (c): Jill Braaten/McGraw Hill; (d): Daniel C. Smith

Figure 1.6

Classification of Matter



only 118 elements are currently known, over 50 million compounds occur naturally or have been synthesized in the laboratory. We will learn much more about elements and compounds in Chapters 2 and 3.

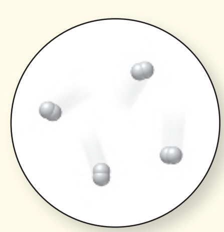
Concept Check 1.3

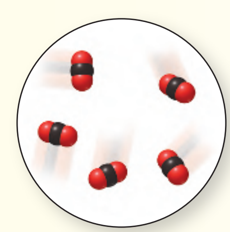
Use the list of elements in Appendix A to classify each item as an element or a compound: (a) the gas inside a helium balloon; (b) table sugar; (c) the rust on an iron nail; (d) aspirin.

Figure 1.6 summarizes the categories into which matter is classified.

Sample Problem 1.1 Using Molecular Art for an Element and a Compound

Classify each example of molecular art as an element or a compound:

a. 

b. 

Analysis

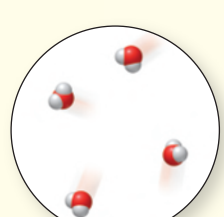
In molecular art, an element is composed of spheres of the same color, whereas a compound is composed of spheres of different colors.

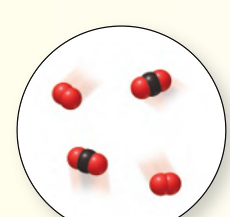
Solution

Representation (a) is an element because each particle contains only gray spheres. Representation (b) is a compound because each particle contains both red and black spheres.

Practice Problem 1.1

Classify each example of molecular art as a pure substance or a mixture:

a. 

b. 

Practice Problem 1.2

Classify each item as a pure substance or a mixture: (a) blood; (b) ocean water; (c) a piece of wood; (d) a chunk of ice.

More Practice: Try Problems 1.1, 1.2, 1.15, 1.16.

1.4 Measurement

Any time you check your weight on a scale, measure the ingredients of a recipe, or figure out how far it is from one location to another, you are measuring a quantity. Measurements are routine for healthcare professionals who use weight, blood pressure, pulse, and temperature to chart a patient's progress.

Roy Hsu/Photographer's Choice RF/Getty Images

In 1960, the **International System of Units** was formally adopted as the uniform system of units for the sciences. **SI units,** as they are called, are based on the metric system, but the system recommends the use of some metric units over others. SI stands for the French words, *Système Internationale.*

- Every measurement is composed of a *number* and a *unit.*

Reporting the value of a measurement is meaningless without its unit. For example, if you were told to give a patient an aspirin dosage of 325, does this mean 325 ounces, pounds, grams, milligrams, or tablets? Clearly there is a huge difference among these quantities.

CONSUMER NOTE

The metric system is slowly gaining acceptance in the United States, as seen in the gallon jug of milk and the two-liter bottle of soda. *Jill Braaten*

1.4A The Metric System

In the United States, most measurements are made with the **English system,** using units like miles (mi), gallons (gal), pounds (lb), and so forth. A disadvantage of this system is that the units are not systematically related to each other and require memorization. For example, 1 lb = 16 oz, 1 gal = 4 qt, and 1 mi = 5,280 ft.

Scientists, health professionals, and people in most other countries use the **metric system,** with units like meter (m) for length, gram (g) for mass, and liter (L) for volume. The metric system is slowly gaining popularity in the United States. The weight of packaged foods is often given in both ounces and grams. Distances on many road signs are shown in miles and kilometers. Most measurements in this text will be reported using the metric system, but learning to convert English units to metric units is also a necessary skill that will be illustrated in Section 1.7.

The important features of the metric system are the following:

- Each type of measurement has a base unit—the meter (m) for length; the gram (g) for mass; the liter (L) for volume; the second (s) for time.
- All other units are related to the base unit by powers of 10.
- The prefix of the unit name indicates if the unit is larger or smaller than the base unit.

Table 1.1 Metric Units

Quantity	Base Unit	Symbol
Length	Meter	m
Mass	Gram	g
Volume	Liter	L
Time	Second	s

The base units of the metric system are summarized in Table 1.1, and the most common prefixes used to convert the base units to smaller or larger units are summarized in Table 1.2. **The same prefixes are used for all types of measurement.** For example, the prefix *kilo-* means 1,000 times as large. Thus,

1 **kilo**meter = **1,000** meters or 1 km = 1,000 m
1 **kilo**gram = **1,000** grams or 1 kg = 1,000 g
1 **kilo**liter = **1,000** liters or 1 kL = 1,000 L

The prefix *milli-* means one thousandth as large (1/1,000 or 0.001). Thus,

1 **milli**meter = **0.001** meters or 1 mm = 0.001 m
1 **milli**gram = **0.001** grams or 1 mg = 0.001 g
1 **milli**liter = **0.001** liters or 1 mL = 0.001 L

Table 1.2 Common Prefixes Used for Metric Units

Prefix	Symbol	Meaning	Numerical Value[a]	Scientific Notation[b]
Giga-	G	Billion	1,000,000,000.	10^{9}
Mega-	M	Million	1,000,000.	10^{6}
Kilo-	k	Thousand	1,000.	10^{3}
Deci-	d	Tenth	0.1	10^{-1}
Centi-	c	Hundredth	0.01	10^{-2}
Milli-	m	Thousandth	0.001	10^{-3}
Micro-	μ[c]	Millionth	0.000 001	10^{-6}
Nano-	n	Billionth	0.000 000 001	10^{-9}

[a]Numbers that contain five or more digits to the right of the decimal point are written with a small space separating each group of three digits.
[b]How to express numbers in scientific notation is explained in Section 1.6.
[c]The symbol μ is the lower case Greek letter mu. The prefix *micro-* is sometimes abbreviated as **mc.**

The metric symbols are all lower case except for the unit **liter** (L) and the prefixes **mega-** (M) and **giga-** (G).

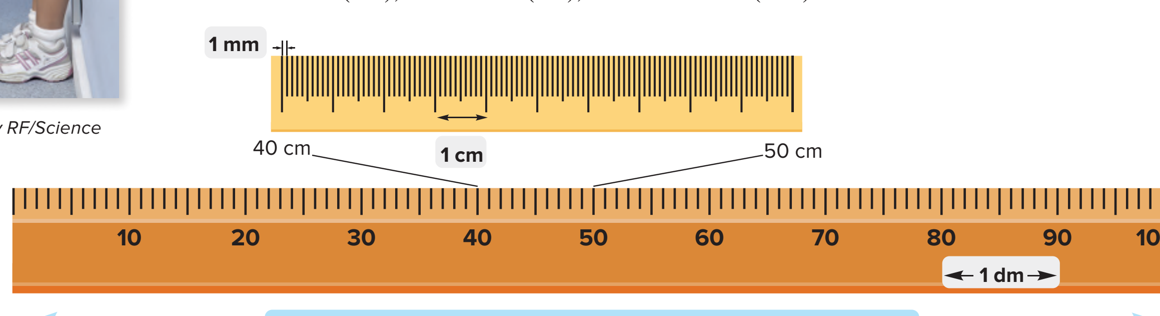
Adam Gault/Science Photo Library RF/Science Source

Concept Check 1.4

What term is used for each of the following units: (a) a million liters; (b) a thousandth of a second; (c) a hundredth of a gram; (d) a tenth of a liter?

1.4B Measuring Length

The base unit of length in the metric system is the *meter* (m). A meter, 39.37 inches in the English system, is slightly longer than a yard (36 inches). Common units derived from a meter are the decimeter (dm), centimeter (cm), and millimeter (mm).



Note how these values are related to those in Table 1.2. Because a centimeter is one *hundredth* of a meter (0.01 m), there are *100* centimeters in a meter.