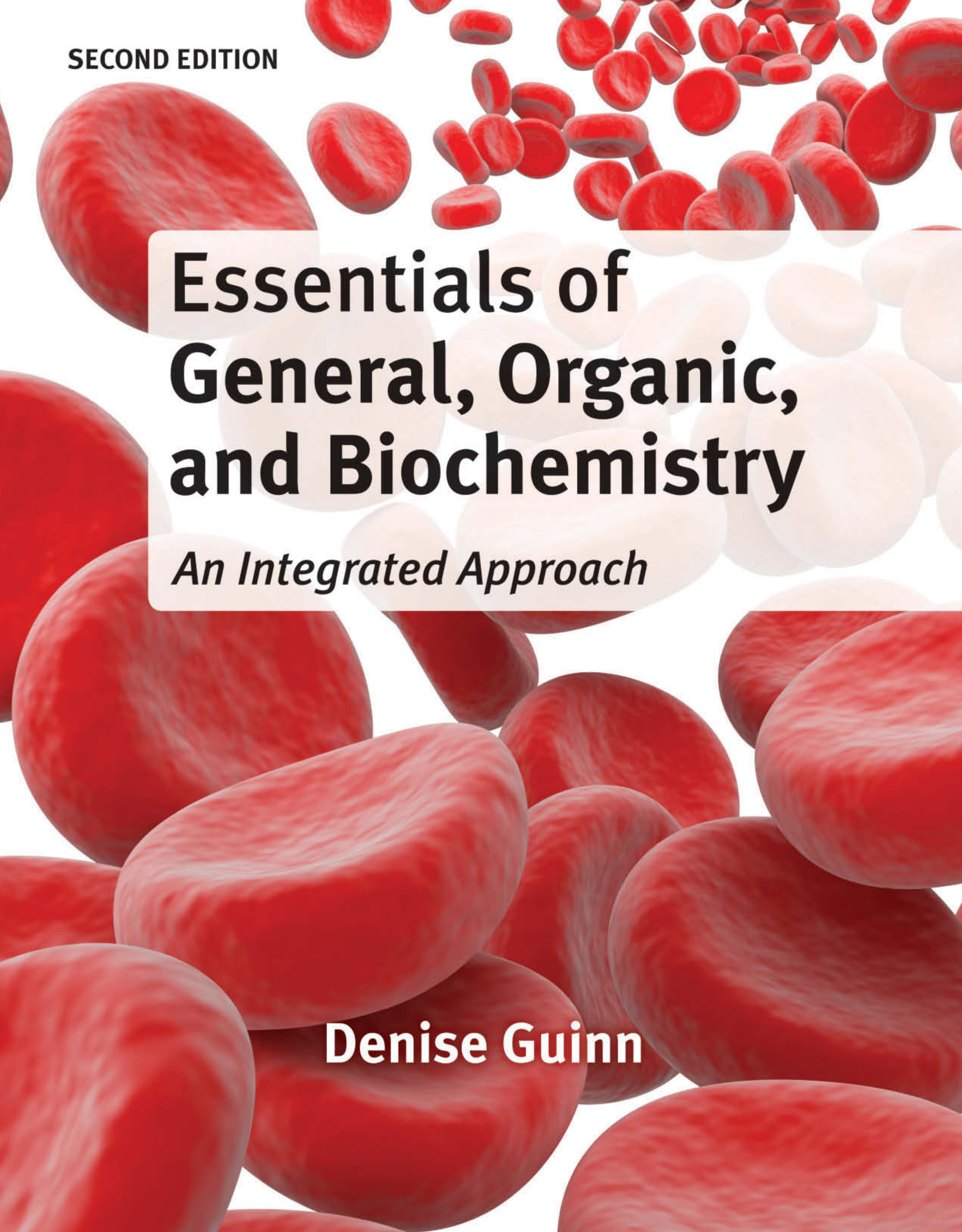

SECOND EDITION

Essentials of General, Organic, and Biochemistry

An Integrated Approach

Denise Guinn

Essentials of **General, Organic, and Biochemistry**

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An Integrated Approach

SECOND EDITION

Denise Guinn

The College of New Rochelle

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To my sons, Charles and Scott,
and to all the Vogels for their continued inspiration and support
–Denise

About the Author

Jorge Madrigal/Madrigal Studios

Denise Guinn received her B.A. in chemistry from the University of California at San Diego and her Ph.D. in organic chemistry from the University of Texas at Austin. She was a National Institutes of Health postdoctoral fellow at Harvard University before joining Abbott Laboratories as a Research Scientist in the Pharmaceutical Products Discovery Group. In 1992, Dr. Guinn joined the faculty at Regis University, in Denver, Colorado, as Clare Boothe Luce Professor of chemistry, where she taught courses in general chemistry, organic chemistry, and the general, organic, and biochemistry course for nursing and allied health majors. In 2008, she joined the chemistry department at The College of New Rochelle, New York where she teaches organic chemistry, biochemistry, and the one-semester GOB course for nursing students. She has published in the Journal of Organic Chemistry, the Journal of the American Chemical Society, and the Journal of Medicinal Chemistry. She currently resides in Nyack, New York.

Brief Contents

Contents

A Letter from The Author

In teaching the general, organic, and biochemistry course for the past 22 years, it has been a great pleasure and opportunity to be at the forefront of the integrated approach to teaching general, organic, and biochemistry. In writing the first edition of *Essentials of General, Organic, and Biochemistry,* our goal was to make it obvious why chemistry is a cornerstone in the education of today's health care professionals by using health and medicine as the framework for learning the fundamentals of chemistry. The second edition of this text has been further shaped by the hundreds of instructors and students who shared with us their experience using the first edition. There is a consensus that the integrated approach effectively engages students in the course early-on, while at the same time making it feasible to learn the fundamental concepts of organic chemistry and biochemistry in a one-semester course. I hope that you feel, as I do, that the second edition has retained the elements of the first edition that worked so well, while incorporating some organizational changes and new material that better support student learning in chemistry.

Denise Guinn

Preface

This text takes an integrated approach to general, organic, and biochemistry as applications of chemistry in health and medicine are used to illustrate the key concepts. To achieve this goal, we approach the course differently in several important ways.

Integration of Organic and Biochemistry in Every Chapter

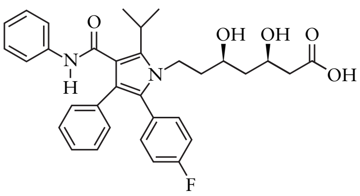
The GOB course has traditionally been taught sequentially covering the three major areas of chemistry—general chemistry, organic chemistry, and biochemistry—in that order. By relegating much of the interesting, relevant content to the end, some students may lose interest in the course.

In this textbook, organic and biochemistry concepts are included in every chapter, so that during every week of the course, students are engaged in topics directly related to their field of study. To make this integration even more effective, organic chemistry is introduced relatively early in the text (Chapters 6 and 7).

Historically, when **organic compounds** were first studied, it was believed that they could *not* be prepared in the laboratory and that only a living plant or animal could produce an organic compound. This is the origin of the term *organic*, which means "from living things." Compounds that do not contain carbon were known as **inorganic compounds**. Today, almost any organic compound can be prepared in the laboratory, even extremely complex organic molecules. Nevertheless, the terms *organic* and *inorganic* remain with us today to distinguish these two basic classes of compounds.

Organic compound:	Contains carbon
Inorganic compound:	Does not contain carbon

Many compounds produced in nature are synthesized in the laboratory. Some are used as **pharmaceuticals**—drugs used for therapeutic purposes. Plants and animals are rich sources of medicinally valuable organic compounds, such as Taxol, the lifesaving anti-cancer drug first isolated from the yew tree (Figure 6-4). Most pharmaceuticals, however, are synthesized in the laboratory. For example, Lipitor (atorvastatin), the best-selling pharmaceutical in the history of medicine, used for the treatment of high cholesterol, is entirely synthetic; it is not produced naturally by any plant or animal (Figure 6-5). You will see examples of pharmaceuticals throughout this chapter and the next as you are introduced to the fundamental principles of organic chemistry.



Lipitor

Figure 6-5 Lipitor (atorvastatin), a synthetic drug used for the treatment of high cholesterol.

NEW: Improved Organization

While retaining the integration of general, organic, and biochemistry, some topics have been rearranged to make the book an even better fit for most courses. Here are some of the organizational changes in the second edition:

- Chapter 1 from the first edition has been divided into two chapters. The benefits of this are that Chapter 1, Measuring Matter and Energy, is now shorter and focused on concepts relating to the measurement of matter and energy, thus allowing the introduction of energy—a central theme throughout the text—earlier in the course.

- Chapter 2, Atomic Structure and Nuclear Radiation, now covers atomic structure, including a streamlined section on nuclear radiation (previously in Chapter 16).
- Chemical reactions are now covered much earlier in the text. Chapter 4, Chemical Quantities and Chemical Reactions, introduces the concept of the mole, along with balancing equations, enthalpy, kinetics, and a new section on chemical equilibrium.
- The chapter on functional groups, Chapter 7, Organic Chemistry and Biomolecules, has a new section on stereochemistry and integrates the structure of the important biomolecules when introducing the functional groups.
- The chapter on solutions (Chapter 8, Mixtures, Solution Concentrations, and Diffusion) and the chapter on acids and bases (Chapter 9, Acids and Bases) are now next to each other while still following the chapters on organic chemistry (Chapters 6 and 7), thereby retaining the examples of organic compounds.
- The last five chapters (11–15) are presented in a new order: Carbohydrates, Lipids, Proteins, Nucleic Acids, and a capstone chapter on Energy and Metabolism. The structure and function of the biomolecules are described in Chapters 11–14, while the catabolic biochemical pathways and energy implications now appear in Chapter 15.

Content Tailored to Prepare Students for Their Careers in Health Care

As scientists, we already know that chemistry is the central science, and as such is an important foundation for understanding health and medicine. For students to be motivated to learn chemistry, they need to see how the concepts they are learning are relevant to their chosen field of study. Studies show that consistently motivated students are much more likely to succeed than those searching for relevance throughout the course, especially early in the course. While engineering examples are best for engineering students, medical examples, and other health- and consumer-based examples are most effective for teaching the fundamental concepts of chemistry to nursing students and allied health majors.

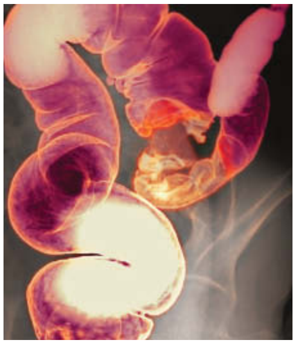


Figure 8-7 X-ray image of a patient with colon cancer, after receiving a barium enema—a suspension of radioactive barium sulfate, $BaSO_4$, in water. [ZEPHYR/Getty Images]

Saturated and Unsaturated Solutions Soluble ionic compounds dissolve in water because the ion-dipole forces of attraction are stronger than the ionic bonds of the intact solid lattice in water. Some ionic compounds, however, do not dissolve in water. For example, magnesium hydroxide, $Mg(OH)_2$, the active ingredient in milk of magnesia, used to relieve the symptoms of heart-burn, is insoluble in water. Radioactive barium sulfate, $BaSO_4$, a suspension taken orally to image the gastrointestinal tract or as a "barium enema," used to screen for colon cancer (Figure 8-7), is another example of an insoluble ionic compound. Insoluble ionic compounds have stronger electrostatic inter-

We make the connections between chemical concepts and health care in each chapter beginning with an intriguing opening story; throughout the main text as chemical concepts are covered; in all of the exercises; and at the end of each chapter in the Chemistry in Medicine feature, which utilizes the concepts of the chapter to explain a particular medical condition or disease.

- For example, when discussing units of concentration in Chapter 8, we include units commonly used in medicine. As a consequence, we are able to write exercises that ask students to calculate concentrations, dosages, and flow rates, all while providing practice with dimensional analysis and the metric system.
- When naming ionic compounds in Chapter 3, we include some that are commonly found in health and consumer products. When discussing the conversion of units in Chapter 4, we use actual blood test results so that students are performing calculations using real data.
- When describing organic functional groups, many prescription and over the counter medicines are used as examples.
- When discussing the gas laws in Chapter 5, we discuss the volume of air in the lungs during an asthmatic attack.
- Rather than introducing the traditional reactions such as displacement reactions between salts and redox reactions of metals, in Chapter 10 we use the five basic organic reactions catalyzed by enzymes in the cell as examples.

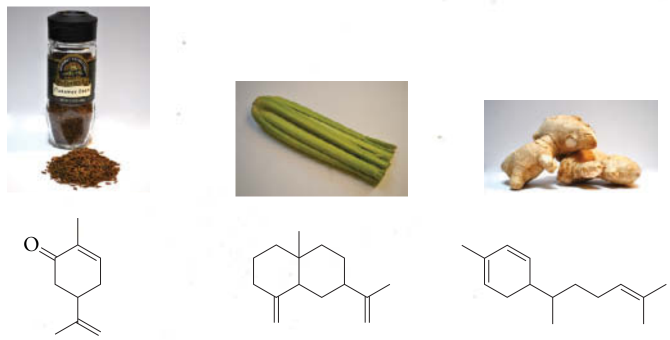
WORKED EXERCISE Applying Gas Laws to the Human Body

5-13 When an asthmatic has an asthma attack, a smaller amount of air enters the lungs, and therefore a lower oxygen partial pressure exists in the lungs and consequently the concentration of oxygen in the blood is lower. How is the concentration of oxygen in the blood affected during an asthma attack? Explain using Henry's law.

Solution

5-13 When an asthmatic has an asthma attack, the partial pressure of oxygen in the lungs is less because of the decreased amount of air inhaled, and therefore, according to Henry's law, the concentration of dissolved oxygen in the blood is less.

NEW: Many More Authentic Images. A liberal use of photos related to actual clinical practice, consumer health care products, and other natural products reinforce the chemical concepts, while showing an application.



Carvone β-Selinene (celery) Zingiberene (ginger)

Figure 6-16 Some dienes and polyenes found in various fruits, vegetables, and spices. [Catherine Bausinger]

Chapter Walkthrough

Chapter Opening Vignettes. Because context and connection are crucial to motivation and learning chemistry, the first encounter with new concepts arrives in the form of a short, practical, and real-life example of a health-related topic connected to the concepts in the chapter. These stories immediately immerse the student in a high-interest topic related to health and medicine.

The hip bone is often used for a bone mineral density (BMD) measurement because it is a good indicator of whether or not a person has osteoporosis. Shown is a healthy hip bone. [James Stevenson/Science Source]

1 Measuring Matter and Energy

Osteoporosis and Bone Density

One in five American women over the age of 50 is estimated to have osteoporosis, and over half of these women will break a bone at some point in their lives as a result of the condition. Osteoporosis is a condition characterized by the progressive loss of bone density, resulting in an increased risk of bone fractures. For example, even a cough or stumble is often enough to trigger a fracture. Most bone fractures occur in the hip, wrist, and spine.

Bone density is the measurement most commonly used to assess bone strength. Density is a physical property of a substance, which can be calculated from the mass of a sample of the substance divided by its volume: $d = m/V$. Mass is similar to weight and is a measure of the amount of material, in this case, bone mineral. Volume is a measure of the three-dimensional space occupied by a material, in this case, a segment of bone. The greater the bone density is, the greater the bone mineral per volume and the stronger the bone. Stronger bone is better able to withstand stress and less likely to fracture. The average bone density for a person is 1.5 g/cm^3, read as "one point five grams per centimeter cubed."

To screen for osteoporosis, several techniques have been developed for measuring or estimating the bone density of the hip, wrist, and lumbar spine. Currently, the most common technique is the DEXA scan (*d*ual *e*nergy *x*-ray *a*bsorptiometry), which

Detailed Worked Exercises, paired with Practice Exercises. Throughout each chapter, Worked Exercises reinforce the text's explanations. They give students a helpful roadmap for solving problems as well as the opportunity to practice the concepts they are learning in the text. Practice Exercises follow each set of Worked Exercises, offering students an immediate check on their understanding of the concepts. Answers to the Practice Exercises appear at the end of each chapter.

Figure 4-5 A can of Red Bull contains 80 mg of caffeine. An eight-ounce cup of coffee contains between 100 and 200 mg of caffeine. [Catherine Bausinger]

WORKED EXERCISES Interconverting between Mass and Moles

4-7 Sulfur reacts with oxygen, producing a blue flame as it forms sulfur dioxide, a compound with a noxious odor. How many grams of sulfur are there in 5.600 mol of sulfur?

4-8 In a 250 mL can of Red Bull there are 80. mg of caffeine (Figure 4-5). How many moles of caffeine, $C_8H_{10}N_4O_2$, are in 80. mg of caffeine?

Solutions

4-7 First determine the molar mass of sulfur because this will be the conversion between moles and mass. Remember, molar mass has the same numerical value as atomic mass, molecular mass, and formula mass but has units of g/mol instead of amu. According to the periodic table, the molar mass of sulfur is 32.07 g/mol.

Next, express the molar mass of S as two possible conversion factors:

$$\frac{32.07 \text{ g S}}{1 \text{ mol S}} \quad \text{or} \quad \frac{1 \text{ mol S}}{32.07 \text{ g S}}$$

Using dimensional analysis, multiply the supplied unit by the appropriate form of the conversion factor that allows the supplied units to cancel, leaving the requested unit. Note that the supplied unit here is moles and the requested unit is grams, so we use the conversion factor shown in red above:

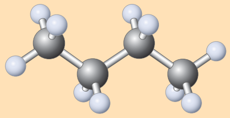
$$5.600 \cancel{\text{mol S}} \times \frac{32.07 \text{ g S}}{1 \cancel{\text{mol S}}} = 179.6 \text{ g S} \quad \text{(four significant figures)}$$

Using Molecular Models. These exercises walk students through the process of building and examining a ball-and-stick model to illustrate an important concept in the chapter. We have found that this type of tactile exercise is effective, and students begin to use the model kit on their own when solving problems. A small inexpensive model kit, specifically designed to be used for the modeling exercises, is available with the book. Building models is a superb way for students to experience and readily understand the three-dimensional aspects of chemistry. The exercises are accompanied by Inquiry Questions that encourage students to build a deeper understanding of molecular structure and relate it to the more abstract structures seen in print.

Using Molecular Models 6-4 Geometric Isomers

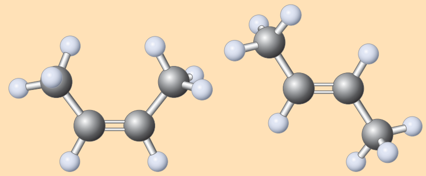
Construct a Model of Butane

1. Obtain eight black carbon atoms, 16 light blue hydrogen atoms, 23 straight bonds, and four bent bonds.
2. Construct a model of butane, C_4H_{10} and answer Inquiry Question 5.



Construct *cis*- and *trans*-but-2-ene

3. Remove one hydrogen atom and its C—H bond from both C(2) and C(3) of the model you made above. Next, replace the C(2)–C(3) single bond with two bent bonds to represent a double bond.
4. Build a model of the other geometric isomer. Your models should look like the ball-and-stick models shown below. Answer Inquiry Questions 6 through 12.



cis trans

Guidelines. Step-by-step instructions, such as naming organic compounds, can be used as both explanations and a quick reference when doing homework or other exercises. They are clearly set off in the text to emphasize their importance and to help students find them easily.

Guidelines for Solving a Dilution Calculation

Step 1. Begin by determining the three supplied variables and the one variable that needs to be calculated. Remember from algebra, that if you have one equation, you can solve for at most one unknown. Applied to a dilution calculation, any one of these variables can be determined if the other three are known. In this example:

$C_1 = 10\%$	concentration of stock solution
$C_2 = 1\%$	desired concentration of dilute solution
$V_1 = ?$	the variable being solved: how much of the stock solution we need to dilute
$V_2 = 100$ mL	the requested final volume of the dilute solution

Step 2. Rearrange the dilution equation so that the variable you are solving for—the unknown—is isolated on one side of the equality, by dividing both sides of the equation by the appropriate variable(s). In this example, we want to solve for V_1, so we isolate it on one side of the equality by dividing both sides by C_1:

$$V_1 = \frac{C_2 \times V_2}{C_1}$$

Step 3. Substitute the supplied variables listed in step 1 into the algebraically rearranged equation from step 2 and solve for the unknown variable. In this example:

$$V_1 = \frac{1\cancel{\%} \times 100 \text{ mL}}{10\cancel{\%}} = 10 \text{ mL}$$

Step 4. Prepare the dilute solution by transferring the calculated volume of stock solution, from step 3, to a volumetric flask with volume V_2. Then add water to the mark, with mixing. In this example, you would transfer 10 mL of the 10% stock saline solution to a 100. mL volumetric flask and then add water, the solvent, to the mark on the 100. mL volumetric flask, along with mixing. You now have a 1% saline solution.

Note that the amount of solute removed from the concentrated solution is the same amount of solute that is in present in the dilute solution since only solvent was added to the solution.

Chemistry in Medicine. A Chemistry in Medicine feature concludes each chapter by providing an in-depth look at how the chemical principles described in the chapter can be directly applied to a problem or issue in health care.

Chemistry in Medicine The Chemistry of Vision

In memory of Ingeborg Vogel (1935–2011), my mother, who suffered from macular degeneration in the last decade of her life.

The leading cause of blindness in people over the age of 60 is age-related macular degeneration (ARMD). It is estimated that 11 million people in the United States have some form of ARMD, and that this number is expected to double by the year 2050. Macular degeneration is a condition in which a person loses his or her central vision, preventing them from being able to read, recognize faces, drive, and see detail. Figure 6-25 shows what a person with ARMD sees compared to someone with normal vision.

We can better understand this disease if we consider the chemistry of vision, a process that involves an important chemical reaction initiated by light.

Figure 6-25 Image of what a person with age-related macular degeneration sees compared to someone with normal vision, showing loss of central vision. [National Eye Institute/National Institutes of Health]

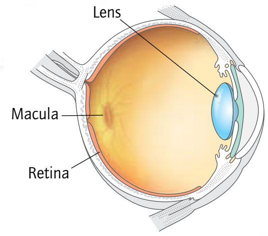


Figure 6-26 Parts of the eye: lens, retina, and macula. The macula is located at the center of the retina right behind the lens.

The retina, located at the back of the eye, contains millions of special photoreceptor cells known as rods and cones. In the center of the retina is a small area known as the macula, which contains a high concentration of cones. Light focused on the macula enables us to see small detail and color. ARMD is caused by a deterioration of the tissue that supports the macula (Figure 6-26).

from one geometric *isomer*—the cis—to the other geometric isomer—the trans. As a result of this cis-to-trans isomerization reaction, the overall shape of the protein-retinal complex changes significantly. The change in shape of the protein-retinal complex initiates a nerve impulse that travels along the optic nerve to the brain. Nerve impulses from the rods and cones are then interpreted in

Chapter Summaries. The chapter's key concepts are presented as a bulleted list at the end of each chapter, offering students a quick reference guide.

Key Words. All boldfaced terms in every chapter are defined in the Key Words section at the end of every chapter, as well as the Glossary at the end of the text.

Summary

Alcohols and Ethers

- An alcohol has the general structure R—O—H, where the —OH group is referred to as a hydroxyl group. The carbon atom bearing the hydroxyl group is an alkyl carbon.
- An ether has the general structure: R—O—R′. The carbon atoms bonded to the oxygen atom can be either alkyl or aromatic carbons.
- An R group is a carbon atom or chain of carbons of undefined length and composition.

have a bent molecular shape around the
.5° bond angles.
1°, 2°, and 3° alcohols, depending on the
the carbon atom bonded to the —OH
as two R groups, and 3° has three R

groups are known as diols; three hydroxyl
yl groups, polyols.
des are carbohydrates, a class of
yl groups—polyols.
ins an —OH group bonded to an

Key Words

Achiral An object or a molecule that is not chiral. It is superimposable on its mirror image and so identical to its mirror image.
Alcohol A functional group derived from water where one of the H atoms has been replaced with an —R group: R—O—H. The carbon atom bearing the hydroxyl group is an alkyl carbon.
Aldehyde A carbonyl containing functional group characterized by a hydrogen atom bonded to the carbonyl carbon. Also includes formaldehyde which contains two —H atoms bonded to the carbonyl carbon.
Alkaloid A compound containing an amine that is found in nature.
Amide A carbonyl compound with a nitrogen atom bonded to the carbonyl carbon. The nitrogen atom has two additional bonds to either —H or —R.
Amine A functional group derived from ammonia in which one, two, or all three of the hydrogen atoms have been replaced by R groups: RNH_2, R_2NH, or R_3N. There is no carbonyl group bonded to the nitrogen.
Amino acid Biological compounds used to build proteins, characterized by an amine and a carboxylic acid functional group bonded to the same carbon atom.
Analgesic A substance that reduces pain.
Carbohydrate A type of biomolecule that includes monosaccharides—simple sugars—and disaccharides. Carbohydrates are a source of energy for cells.
Carbonyl group A carbon-oxygen double bond, C═O.
Carboxylic acid A carbonyl containing functional group that contains an —OH group bonded directly to the carbonyl carbon.

Additional Exercises. At the end of each chapter, additional exercises reinforce the concepts and skills presented in the text. Clearly labeled by section to help both students and instructors, there are on average 100 exercises in each chapter, and they are also available in the textbook's accompanying online homework system. Answers to the odd-numbered exercises are available at the back of the book in Appendix B, and detailed solutions for all the exercises are available in the student solutions manual.

Student Ancillary Support

Supplemental learning materials allow students to interact with concepts in a variety of scenarios. By analyzing figures, reinforcing problem-solving methods, reviewing chapter objectives, and viewing podcasts of an instructor work out solutions, students obtain a practical understanding of the core concepts. With that in mind, W. H. Freeman has developed the most comprehensive student learning package available.

Printed Student Study Guide and Solutions Manual

by Rachel C. Lum,
ISBN: 1-4641-2506-6

The combined Student Study Guide and Solutions Manual provides students with a manual designed to help them avoid common mistakes and understand key concepts. After a brief review of each section's critical ideas, students are taken through step-by-step worked examples, try-it-yourself examples, and chapter quizzes, all structured to reinforce chapter objectives and build problem-solving techniques. The Solutions Manual includes detailed solutions to all odd-numbered exercises in the text.

Molecular Model Kit

ISBN: 1-4292-2687-0

Molecular models help students understand the physical and chemical properties of molecules by providing a way to visualize the three-dimensional arrangement of atoms. This model set, created specifically for *Essentials of General, Organic, and Biochemistry*, uses different color polyhedra to represent atoms and plastic connectors to represent bonds.

Electronic Student Resources

The *Essentials of General, Organic, and Biochemistry, Second Edition* Book Companion Website, accessed at www.whfreeman.com/guinn2e, provides a range of tools for problem solving and understanding concepts, from self-quizzes and pre-drawn molecules for use in homework problems to an interactive periodic table of the elements and an English/Spanish Glossary. In addition, the Book Companion Website includes these interactive activities:

Problem-Solving Tutorials reinforce the concepts presented in the classroom, providing students with self-paced explanations and practice. Each tutorial relates to a worked example from the text, and consists of the following four components:

- Conceptual Explanation: An easy-to-follow, thorough explanation of the topic.
- Worked Example: A step-by-step walkthrough of the problem-solving technique.
- Try It Yourself: An interactive version of the Worked Example that prompts students to complete the problem and supplies answer-specific feedback.
- Practice Problems: A series of problems designed to check understanding.

ChemCasts are videos that replicate the face-to-face experience of watching an instructor solve a problem. Using a virtual whiteboard, the ChemCast tutors show students the steps involved in solving key worked examples, while explaining the concepts along the way. The worked examples were chosen with the input of general, organic, and biochemistry students. ChemCasts can be viewed online or downloaded to a portable media device, such as an iPod.

The Virtual Model Kit allows students to build molecular models on the computer screen. Students select elements and bonds according to instructions in the text's Using Molecular Models features, as well as from additional exercises available within the associated Website and homework system. After constructing the models, students are presented with automatically graded questions to evaluate their understanding.

Electronic Textbooks

The CourseSmart eTextbook provides the full digital text, along with tools to take notes, search, and highlight passages. A free app allows access to CourseSmart eTextbooks on Android and Apple devices, such as the iPad. CourseSmart eTextbooks can be downloaded to your computer and accessed without an Internet connection, removing any limitations for students when it comes to reading a digital text. The CourseSmart eTexbook can be purchased at www.coursesmart.com

Online Tutoring

W. H. Freeman knows that learning chemistry can be difficult, and that sometimes students need some extra support. That's why we have partnered with NetTutor to bring students tutoring they can access anytime, anywhere.

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NetTutor® is an online tutoring service that specializes in a customized tutoring experience. Using the Socratic

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For the Laboratory Lab Manual

by Julie Klare, Fortis College
ISBN: 1-4641-2507-4

The lab manual provides a wide variety of classic and innovative experiments covering the basic topics of general, organic, and biochemistry. These experiments emphasize biological applications of chemical concepts, often within the context of the health sciences. Each experiment can easily be completed within a 3-hour time frame and is accompanied by data sheets and questions that guide students through the analysis of their data.

LabPartner® Chemistry

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W. H. Freeman's latest offering in custom lab manuals provides instructors with a diverse and extensive database of experiments published by W. H. Freeman and Hayden-McNeil Publishing—all in an easy-to-use, searchable online system. With the click of a button, instructors can choose from a variety of traditional and inquiry-based labs. LabPartner Chemistry sorts labs in a number of ways, from topic, title, and author, to page count, estimated completion time, and prerequisite knowledge level. You can add content on lab techniques and safety, reorder the labs to fit your syllabus, and include your original experiments with ease. Wrap it all up in an array of bindings, formats, and designs. It's the next step in lab publishing—the perfect partner for your course.

Instructor Resources

For instructors using *Essentials of General, Organic, and Biochemistry, Second Edition*, W. H. Freeman provides a complete suite of assessment tools and course materials.

Computerized Test Bank

by Rachel Jameton, Lewis-Clark State College
ISBN: 1-4641-2511-2

The Computerized Test Bank offers over 1600 multiple-choice questions. It is designed to assess student knowledge at all levels of learning, from basic definitions to application and synthesis. Utilizing diagrams, figures, and structures, the test bank emphasizes visual understanding of an array of concepts, including applications to the health sciences. While the test bank is also available as a printed manual, the easy-to-use CD-ROM includes Windows and Macintosh versions of the widely used test generation software, which allows instructors to add, edit, and re-sequence questions to suit their testing needs. The Test Bank is also available in Word Format.

Electronic Instructor Resources

Instructors can access valuable teaching tools through www.whfreeman.com/guinn2e. These password-protected resources are designed to enhance lecture presentations, and include textbook images (available in .jpeg and PowerPoint format), Clicker Questions, Lab Information, and more. Enhanced Lecture PowerPoints, which contain complete class-lecture content, highlight key chapter ideas and include worked examples, pre-class questions, and textbook figures and diagrams. There are approximately 60 slides per chapter.

Course Management System Cartridges

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The hip bone is often used for a bone mineral density (BMD) measurement because it is a good indicator of whether or not a person has osteoporosis. Shown is a healthy hip bone.
[James Stevenson/Science Source]

1 Measuring Matter and Energy

Osteoporosis and Bone Density

One in five American women over the age of 50 is estimated to have osteoporosis, and over half of these women will break a bone at some point in their lives as a result of the condition. Osteoporosis is a condition characterized by the progressive loss of bone density, resulting in an increased risk of bone fractures. For example, even a cough or stumble is often enough to trigger a fracture. Most bone fractures occur in the hip, wrist, and spine.

Bone density is the measurement most commonly used to assess bone strength. Density is a physical property of a substance, which can be calculated from the mass of a sample of the substance divided by its volume: $d = m/V$. Mass is similar to weight and is a measure of the amount of material, in this case, bone mineral. Volume is a measure of the three-dimensional space occupied by a material, in this case, a segment of bone. The greater the bone density is, the greater the bone mineral per volume and the stronger the bone. Stronger bone is better able to withstand stress and less likely to fracture. The average bone density for a person is 1.5 g/cm^3, read as "one point five grams per centimeter cubed."

To screen for osteoporosis, several techniques have been developed for measuring or estimating the bone density of the hip, wrist, and lumbar spine. Currently, the most common technique is the DEXA scan (*d*ual *e*nergy *x*-ray *a*bsorptiometry), which is a type of low dose x-ray that measures the *mass* and *area* of a section of bone and then estimates the third dimension to obtain a *volume* measurement.

From a DEXA scan, a bone density or BMD (*b*one *m*ineral *d*ensity) measurement is obtained. The result is then compared with the average BMD for healthy young adults of the same gender and ethnicity. A person is given a T-score that corresponds to how far they are from the average BMD of a similar population (**Figure 1-1**). A T-score of +1 to –1 means they fall within the "normal" range. A person with a T-score of –1 to –2.5 is classified as having osteopenia, a condition characterized by lower-than-average bone density, which may eventually develop into osteoporosis. People with a T-score less than or equal to –2.5 are classified as having osteoporosis. Basically, a T-score compares a patient's bone density to a similar segment of the population and classifies that patient as having normal bone density, osteopenia, or osteoporosis.

Outline

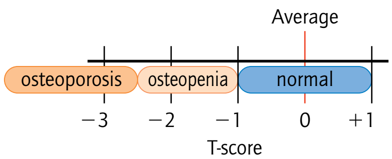

Figure 1-1 When a patient's BMD (bone mineral density) measurement is obtained, they are given a T-score that corresponds to how far they are from the average BMD of a similar population.

In this chapter, you will see how important measurements are in medicine and science. You will learn that there are limitations to any measurement, how measurements are reported, common units of measurement, and how to calculate density and other physical properties. While working in the health care field, you will often be called on to take measurements and to perform critical calculations, such as the dosage of medicine to administer to a patient. These skills are a foundational and critical part of your training, so it is with measurement that we start your venture into the fascinating field of chemistry.

1.1 Matter and Energy

Chemistry is the study of matter and changes in matter. ***Matter** is defined as anything that has mass and occupies volume.* Hence, matter is all the "stuff" around you and in you. It includes matter that you can see—the macroscopic—as well as matter that is too small for you to see—the microscopic and atomic scales.

Matter is found in three **states** or **phases**: solid, liquid, and gas. Examples of all three states of matter are found in the body. The oxygen you breathe is in the gas state, making it possible to quickly fill the lungs with each breath you take. Blood is in the liquid state, so it can be pumped throughout the circulatory system transporting important nutrients to cells. Skin and bone are in the solid state, providing structural integrity to the body.

The macroscopic differences between solids, liquids, and gases can be described by their shape and volume relative to their container, as illustrated in the top part of Figure 1-2:

- A **solid** (s) has a definite *volume* and *shape*, which are independent of its container.
- A **liquid** (l) has a definite *volume* but does not have a definite *shape*, as it conforms to the shape of its container.
- A **gas** (g) has neither a definite *volume* nor a definite *shape*. A gas conforms to the volume and shape of its container.

What makes a gas take on the shape of its container while a solid has a shape independent of its container? To understand these macroscopic properties, we must examine matter on the atomic scale—the incredibly small particles of matter that cannot be seen even with a light microscope, as illustrated in the bottom part of Figure 1-2. Experiments have shown that matter is made up of particles known as atoms, which is the subject of chapter 2. In the solid phase, atomic particles are very close to one another in an ordered lattice. In

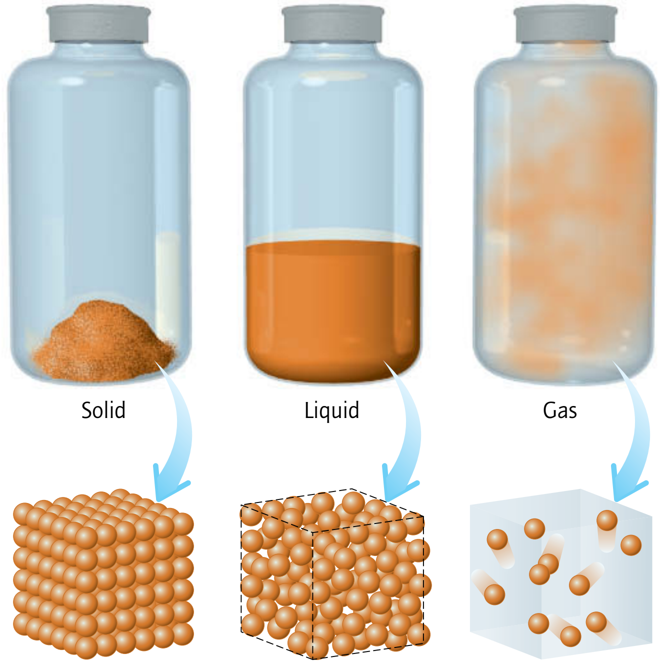

Figure 1-2 The differences in macroscopic properties of solids, liquids, and gases are a result of differences at the atomic scale.

the liquid phase, atomic particles are farther apart but still interacting. In the gas phase, atomic particles are far apart from one another and have little if any interaction.

As a science, chemistry endeavors to explain the macroscopic world—that which we can see—by understanding events on the atomic scale—that which we cannot see. Our understanding of the macroscopic world through knowledge of the atomic world has been achieved through centuries of experiments, measurements, and careful observation. Medicine has seen advances as a result, since medicine is a science with a foundation in chemistry and biology. As a person entering the health care professions, it is important for you to understand the fundamental principles of chemistry at both the macroscopic and the atomic level.

Kinetic and Potential Energy

Particles of matter are not stationary but instead are vibrating and moving about, to a degree that depends on the amount of energy they possess. To understand the physical properties of matter described above, we must first examine the relationship between matter and energy. Central to the physical properties of all matter is energy, one of the most important concepts in science. ***Energy**, broadly defined, is the capacity to do work, where **work** is defined as the act of moving an object against an opposing force.* Energy determines the state of matter. For a given substance, the gas phase has more energy than the liquid phase, which has more energy than the solid phase.

Kinetic Energy There are two basic forms of energy: *kinetic energy* and *potential energy*. ***Kinetic energy** is the energy of motion, including the energy a substance possesses as a result of the motion of its particles.* A skier has kinetic energy as she skies down a mountain (Figure 1-3). Similarly, the particles of matter in any state of matter have kinetic energy. The kinetic energy (KE) of an object or particle depends on its mass, m, and its velocity, v (speed), which can be expressed by the mathematical equation:

$$KE = \tfrac{1}{2}mv^2$$

Thus, faster-moving objects have greater kinetic energy than slower-moving objects with the same mass, and heavier objects have greater kinetic energy than lighter objects moving at the same velocity. For example, a high-speed car collision results in more damage than a low-speed car collision (velocity) if the mass of the cars is the same, and colliding with an SUV causes more damage than colliding with a compact car (mass) if they are each traveling at the same velocity.

Heat is a form of kinetic energy because it involves the *motion* of the particles of matter. *Heat energy always flows from the hotter object to the colder object.* Our bodies are able to perceive heat transfer. Thus, if you touch ice, heat transfers from your body to the ice, causing the particles of ice to increase in kinetic energy and the particles on the surface of your hand to lose kinetic energy. Your nerves detect this transfer of heat energy out of your hand, and your brain interprets it as interacting with something cold. If you touch a hot stove, a similar thing happens except in the opposite direction. Your brain knows it has touched something with more kinetic energy. You interpret it as something hot.

It is important to note that temperature is *not* the same as heat. Heat is kinetic energy—motion of particles of matter—while **temperature** is a *measure* of the particles' kinetic energy. For example, the temperature of the air is

Figure 1-3 This skier has kinetic energy by virtue of her motion. [Jupiterimages/Getty Images]

a measure of the average kinetic energy (heat energy) of the oxygen and nitrogen particles that constitute air. States of matter in which particles have a greater average kinetic energy will have a higher temperature than states of matter whose particles have a lower average kinetic energy, for a given substance. As particles move faster, their kinetic energy increases, and the temperature rises. Rub your hands together quickly and vigorously. Do you feel them getting warmer? Your hands become warmer because the skin particles on the top layer of skin are moving faster, resulting in greater kinetic energy.

Potential Energy ***Potential energy*** *is stored energy, the energy a substance possesses as a result of its position or composition* (Figure 1-4). A rock poised at the top of a precipice, for example, possesses potential energy as a result of its position. When the rock falls, its potential energy is converted into kinetic energy. A gallon of gasoline has potential energy as a result of its composition—the gasoline particles. When the gasoline is burned in a car engine, the potential energy of these gasoline particles is turned into kinetic energy, which is used to perform the work of moving your car. Food also contains potential energy, which is transferred to your cells after you eat and digest the food so that your cells can do the "work" they need for living.

Kinetic Molecular View of the States of Matter

Consider the physical states of matter from a **kinetic molecular view** (Figure 1-2). In the *gas* phase, particles of matter have the greatest kinetic energy. Particles in the gas phase are moving faster than when they are in the liquid or solid phase. Particles in the gas phase are far apart from one another, interacting only when they collide. In the *liquid* phase, these same particles are moving slower than in the gas phase but faster than in the solid phase. In the liquid state, the particles are much closer together, moving randomly and tumbling over one another. This is why liquids flow when poured. In the *solid* phase, the particles exist in a regular ordered pattern with much less kinetic energy than when in the liquid or gas phases, so they are closer together with mainly vibrational motion.

Figure 1-4 Left: The balanced rock has potential energy by virtue of its position. [iStockphoto/Thinkstock] Right: The gasoline in this container has potential energy by virtue of its composition. [© Ron Chapple/Corbis]

WORKED EXERCISES Matter and Energy

1-1 In which of the following states of matter is the velocity of the particles of a given substance the slowest?

a. solid **b.** liquid **c.** gas

1-2 Indicate whether each of the following is an example of kinetic energy or potential energy.

a. a ball rolling down a hill
b. standing on the edge of a diving board
c. a piece of bread

1-3 Is heat energy kinetic energy or potential energy?

1-4 Which direction does heat flow when you place your hand over a pot of steam: from the pot to your hand or from your hand to the pot?

1-5 Using the kinetic molecular view of matter, offer an explanation for why cooking odors travel quickly from the kitchen to neighboring rooms.

Solutions

1-1 **a.** Particles in the solid phase are the slowest because they have the least amount of kinetic energy.

1-2 **a.** Kinetic energy because the ball is moving
b. Potential energy because of the diver's position at the edge of the diving board. When he jumps, this potential energy will be converted to kinetic energy.
c. Potential energy because of the composition of bread, a food that can be turned into energy.

1-3 Heat is a form of kinetic energy because it represents the motion of particles of matter.

1-4 Heat flows from the pot to your hand because the pot is hotter than your hand and heat transfer always occurs in the direction from hot to cold.

1-5 Cooking odors are gas particles detected by your nose, and as a gas, they fill the volume of their container—the rooms, in this example.

PRACTICE EXERCISES*

1 Describe the macroscopic differences between the *solid*, *liquid*, and *gas states* by comparing their shape and volume relative to the container they occupy.

2 Indicate whether each of the following examples illustrates *potential energy* or *kinetic energy*:

a. a compressed spring
b. a windmill turning
c. particles in the gas phase colliding with the walls of their container
d. a skier skiing down a mountain
e. the breakfast you eat for sustenance during the day

*You can find the answers to the Practice Exercises at the end of each chapter.

Physical and Chemical Changes

Matter can undergo both physical changes and chemical changes. A **physical change** is a process that does *not* affect the composition of the substance. A **change of state** is an example of a physical change. For example, when water in the solid phase (ice) changes to a liquid (s → l, melting), the composition of the water molecules has not changed. The particles simply have greater kinetic energy, so they are moving more rapidly and are farther apart. One way to tell that the composition has not changed is that liquid water can be turned back into ice by removing heat. Thus, the observable properties of the substance have not changed physically.

A **chemical change**, also known as a **chemical reaction**, involves a change in the composition of the substance. For example, cooking food is a chemical change, because the food's properties have changed as a result of cooking. Cooling the food back to room temperature does not bring back the uncooked food, because a chemical change has occurred. Since all matter possesses energy, any changes in matter will have an associated change in energy. The energy associated with chemical changes is described in Chapter 4 and for physical changes in Chapter 5.

WORKED EXERCISE Physical and Chemical Changes

1-6 Indicate whether each of the following changes illustrates a *physical* or a *chemical change*.

a. water vapor condensing on the outside of a cold glass on a humid summer day
b. a shiny piece of iron rusting to form brown-colored iron(II) oxide
c. a sample of liquid water freezing
d. cooking a steak over hot coals

Solution

1-6 **a.** a physical change because water is undergoing a phase change from gas to liquid, but its composition has not changed—it is still water
b. a chemical change because the composition of matter changes, as observed in a color change
c. a physical change because water is undergoing a phase change from liquid to solid, but it is still water
d. a chemical change because the appearance of the steak changes, and it cannot return to uncooked steak on cooling

PRACTICE EXERCISES

3 Indicate whether each of the following changes is a *physical* or a *chemical change*:

a. silver tarnishing to form silver sulfide
b. sugar granules dissolving in water
c. combustion of gasoline in a car engine producing carbon dioxide and water
d. boiling water to make water vapor

1.2 Measurement in Science and Medicine

To understand matter and energy, we must be able to measure them. Thus, we turn our attention to measurement, a critical skill for anyone working in a field of science. Consider a drop of blood on the head of a pin, which you can *see* with the naked eye (Figure 1-5). If you look at this droplet of blood through a microscope, you will see that it is composed of millions of red blood cells, each with a diameter 1,000 times smaller than the droplet of blood. Now imagine—because you can't see it with a microscope—peering inside one of these red blood cells. Among other things, you will "see" millions of hemoglobin molecules, each composed of many atoms. Hemoglobin is the substance that carries oxygen from the lungs to the tissues throughout the body. A hemoglobin *molecule* has a diameter 1,000 times smaller than a red blood cell and a million times smaller than the droplet of blood on the head of a pin. If you imagine zooming in further, you will "see" that a single hemoglobin molecule is composed of about 10,000 atoms, including four iron atoms, about 100 times smaller than a hemoglobin molecule.

From the above exercise, we see that matter can be described on different scales, illustrated in Figure 1-5. The **macroscopic scale** includes all matter that

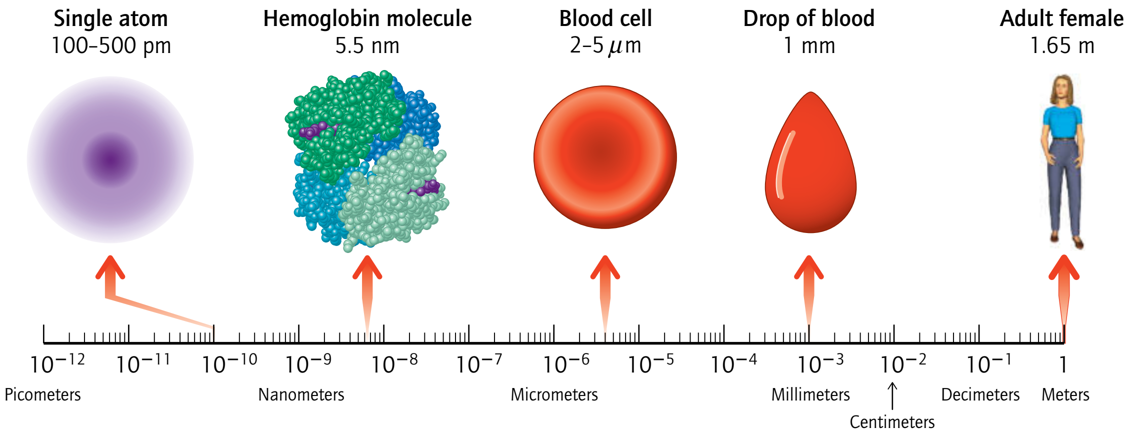

Figure 1-5 Metric lengths ranging from the incredibly small 100-pm diameter of a single iron atom—the atomic scale—to the 1.65-m height of the average adult female—the macroscopic scale.

you can see, such as the drop of blood and the human body, containing on average 5 liters of blood. The **microscopic scale** includes matter such as red blood cells that cannot be seen with the naked eye but can be seen with magnification through a microscope. The **atomic scale** describes matter such as a hemoglobin molecule or an atom that is far too small to be seen except with a very specialized type of microscope, the scanning electron microscope (SEM).

Throughout this text you will see that an understanding of the *atomic scale* brings with it a better understanding of the *macroscopic*—the things you can see—especially as the understanding applies to health and disease. You will learn that many disease processes are caused by some malfunction at the atomic level.

English and Metric Units

Medical professionals make measurements every day, whether to obtain a patient's weight, take a patient's temperature or blood pressure, or administer a particular dose of medication. *A measurement consists of two parts: a number and a unit.* For example, a baby might weigh 10 *lb*, not simply 10. Every measurement also has a margin of error associated with it, which is conveyed by the number of digits (figures) reported in the measurement. Thus, the baby's weight may be recorded as 10. lb, 10.0 lb, or 10.00 lb, depending on the precision of the balance used to weigh the baby, and each of these measurements conveys a different degree of uncertainty in the measurement.

The two most common *systems* of measurement that you will encounter in medicine are the metric system and the English system. The **English system** is used only in the United States and a few other countries. The **metric system** is the most widely used system of measurement in the world and the system of units used in the sciences. The metric system is convenient because it involves units that are multiples of 10 of the base unit.

The metric system employs various **base units** that measure a particular quantity, such as

- the *meter* (m) for measuring length
- the *gram* (g) for measuring mass
- the *second* (s) for measuring time
- the *calorie* (cal) for measuring energy

The *i*nternational *s*ystem of units (SI) was established by an international group of scientists for the purpose of setting a uniform set of units in the sciences. SI units are the preferred unit for science and commerce.

Prefixes in the Metric System If you have ever looked at the label on a bottle of multivitamins, you will see the mass of each vitamin and mineral reported with units such as mg and mcg. Large bags of flour and sugar have units such as kg, beverages have volume labels such as mL, and the thickness of a rope or string is given in units of cm or mm. These are all examples of metric units containing a *prefix* in front of a base unit (m, L, g). As you can see, metric prefixes are used not only in science but are part of our everyday world.

Metric *prefixes* are used when a measurement is much larger or smaller than the base unit so as to avoid the need to use many zeros in the numerical value, which makes the measurement cumbersome and difficult to interpret. For example, the head of a pin is 0.001 meter in diameter, a blood cell has a diameter of 0.000 001 meter, and there are 500,000,000,000 bytes in my computer hard drive. It is much simpler to report and interpret that the head of a pin is 1 millimeter in diameter, a blood cell has a diameter of 1 micrometer, and my hard drive has 500 gigabytes. To avoid writing numbers with many zeros, the metric system employs **prefixes**, which when placed in front of the base unit represent a multiplier or divider that makes the prefixed unit larger or smaller by some multiple of 10 or 1/10 (Table 1-1). For example, the prefix *milli* always represents the multiple 1,000 (10^3); thus, 10^3 millimeters (mm) is equal to 1 m. Similarly, 10^3 mg is equal to 1 gram (g), and 10^3 mL is equal to 1 liter (L). *Each prefix represents a specific multiplier independent of the base unit*. Other prefixes represent different multipliers. For example, the prefix *micro* represents the multiple 1,000,000 (10^6); thus, 10^6 micrometers (μm) is equal to 1 meter (m).

Some prefixes represent negative exponents. For example, the prefix *kilo* represents 1/1,000 or (10^{-3}), so 10^{-3} kilometers (km) is equal to 1 meter (m) and 10^{-3} kilogram (kg) is equal to 1 gram (g). The prefix *mega* represents 10^{-6} and the prefix *giga* represents 10^{-9}. Thus, my hard drive with 500,000,000,000 bytes can be written as 500 gigabytes, a number much easier to interpret. The common prefixes in the metric system, the multiplier/divider they represent, and their abbreviations are shown in Table 1-1.

The metric system with its prefixes is easy to learn because you need only memorize the prefixes (Table 1-1, column 2) and the associated multiple of 10 (columns 3 and 4) represented by each prefix. *Moreover, use Table 1-1 to create a mathematical equality, known as a* ***conversion****, between the base unit and the prefixed unit of interest when performing a metric conversion*. For example, 10^9 nm = 1 m. It is critical when writing a metric conversion that the exponent, 10^9 in this case, is always placed in front of the *prefixed* unit (nm) and 1 is

> Students often do a great job memorizing the metric prefixes and their associated exponents but then forget whether the exponent should be placed in front of the prefixed unit or the base unit. Remember that the exponent goes in front of the prefixed unit and a 1 goes in front of the base unit. For example, 10^3 mm = 1 m, not 1 mm = 10^3 m. One way to help you remember this is to note which is larger, the base unit or the prefixed unit. Clearly, more smaller units are required to equal a larger unit.

TABLE 1-1 Common Metric Prefixes and the Multipliers They Represent*

Prefix	Prefix Symbol	Multiplier	Multiplier in Scientific Notation
giga	G	0.000 000 001	10^{-9}
mega	M	0.000 001	10^{-6}
kilo	k	0.001	10^{-3}
deci	d	10	10
centi	c	100	10^2
milli	m	1000	10^3
micro	μ	1 000 000	10^6
nano	n	1 000 000 000	10^9
pico	p	1 000 000 000 000	10^{12}
femto	f	1 000 000 000 000 000	10^{15}

*Examples of how to use the table to create a metric conversion:
1 g = 10^3 mg; 1 g = 10^6 μg; 1 g = 10^{-3} kg

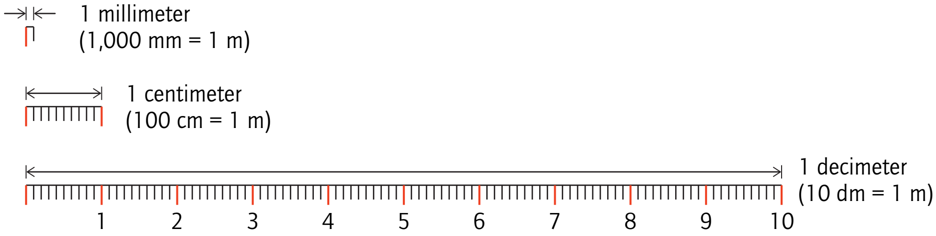

Figure 1-6 Actual size of 1 millimeter, 1 centimeter, and 1 decimeter. There are 10 millimeters in 1 centimeter, 10 centimeters in 1 decimeter, and 10 decimeters in 1 meter. With respect to the base unit, the meter, there are 10^3 millimeters in 1 meter, 10^2 centimeters in 1 meter, and 10 decimeters in 1 meter.

always placed in front of the *base* unit (m), not the other way around, when using this table.

You will notice that most of the multipliers in Table 1-1 are written in scientific notation. Scientific notation allows us to express numbers containing many zeros without writing the zeros. If you need a review of scientific notation, refer to Appendix A: Mathematics Review with Tips on How to Use a Calculator, where you will also find instructions for how to input a number that is given in scientific notation into a calculator. Let us now consider some common measurements using the metric system and the English system.

Length The meter is the base unit of **length** and distance in the metric system. There are 10 decimeters (dm) in 1 meter (m), 10^2 centimeters (cm) in 1 meter (m), and 10^3 millimeter (mm) in 1 meter (m) as shown in actual size in Figure 1-6. As you can see from Table 1-1, there are 10^6 micrometers (μm) in 1 meter, 10^9 nanometer (nm) in 1 meter, 10^{12} picometers (pm) in 1 meter, and so forth. Notice these units represent lengths smaller than the base unit, the meter. Table 1-1 also shows lengths and distances larger than the base unit, represented by the prefixes *kilo*, *mega*, and *giga*; however, the kilometer is the only commonly used prefix for distance measurements.

If you consider 1 mm the smallest length that you can reasonably see with the naked eye, then every 1,000-fold decrease takes you into the range of another scale: macroscale (greater than mm) → microscale (μm range) → atomic scale (smaller than a nm). Figure 1-5 shows a range of metric lengths from an iron atom with a diameter of 126 pm (the atomic scale) to a woman measuring 1.65 m tall (the macroscale).

The common English units of length are the inch (in), the foot (ft), and the mile (mi). The conversions between some common English and metric units of length are as follows:

$$
\begin{aligned}
1.00 \text{ in} &= 2.54 \text{ cm (exact)} \\
39.37 \text{ in} &= 1.00 \text{ m} \\
1.00 \text{ mi} &= 1.61 \text{ km}
\end{aligned}
$$

Measurements of length are routinely used to assess a developing fetus. An ultrasound scan of a developing fetus (Figure 1-7) is a noninvasive technique used to measure the size of various parts of the fetus, including the crown-rump length, the biparietal diameter (distance between the sides of the head), the femur length (thigh bone), and the abdominal circumference. These measurements of length assess gestational age, size, and growth of the fetus. For example, the biparietal diameter in a healthy fetus increases from about 2.4 cm at 13 weeks to about 9.5 cm at term. Structural abnormalities of the fetus such as spina bifida and cleft palate can be reliably diagnosed using ultrasound measurements taken before 20 weeks.

Figure 1-7 An ultrasound scan of a human fetus at 20 weeks. This fetus has a crown–rump length of 11.8 cm. Ultrasound scans are a noninvasive way to monitor the growth of the fetus. [© Craig Holmes Premium/Alamy]