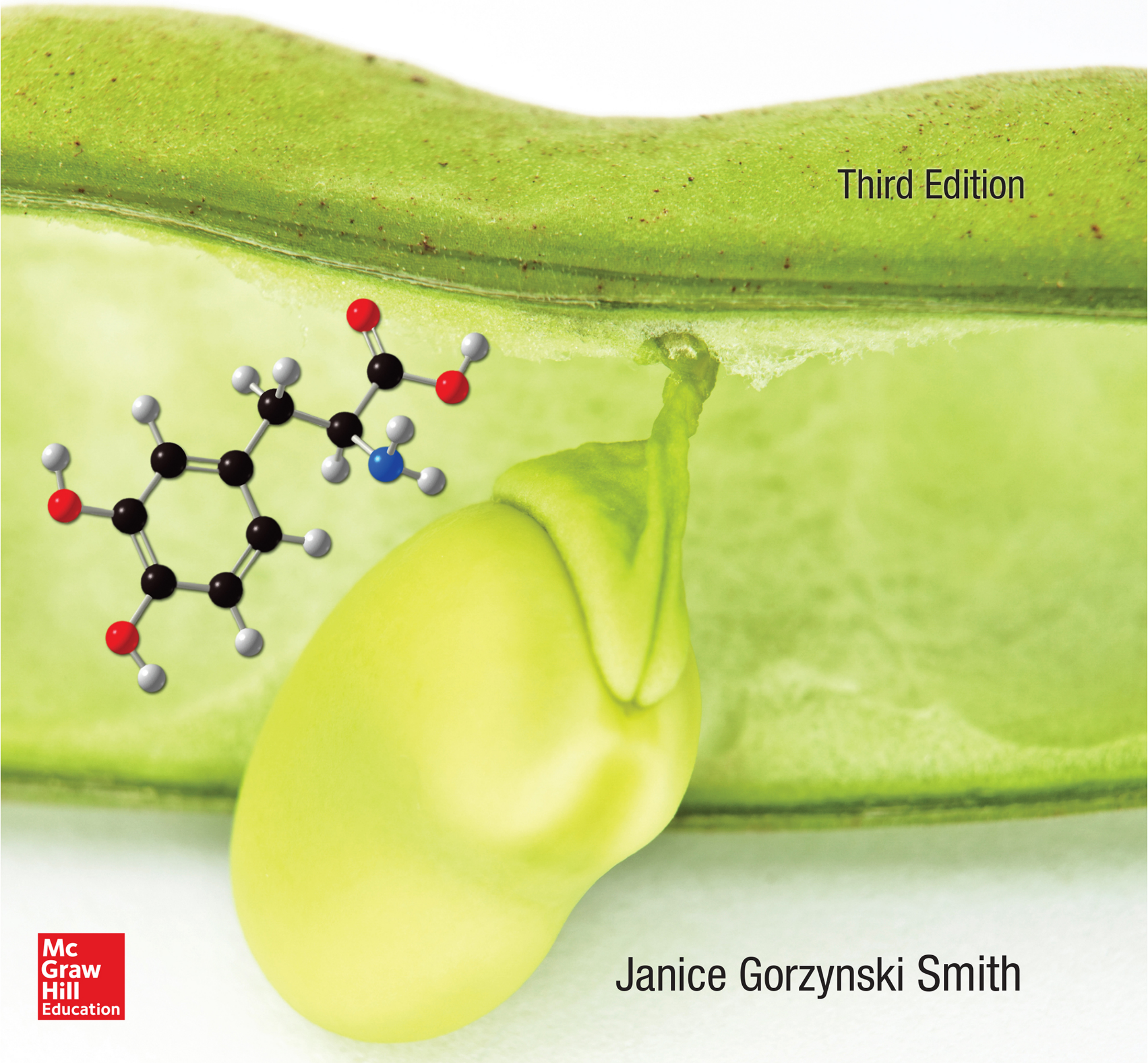
General, Organic, & Biological Chemistry



Third Edition

Janice Gorzynski Smith

McGraw Hill Education

Periodic Table of the Elements

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	← 8	8B 9	10 →	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	
1	1 **H** 1.008																	2 **He** 4.003	1
2	3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18	2
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95	3
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.41	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80	4
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3	5
6	55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.8	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)	6
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **–** (284)	114 **Fl** (289)	115 **–** (289)	116 **Lv** (293)	117 **–** (294)	118 **–** (294)	7

6	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0	6
7	90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)	7

The Elements

Element	Symbol	Atomic Number	Relative Atomic Mass*
Actinium	Ac	89	(227)
Aluminum	Al	13	26.98
Americium	Am	95	(243)
Antimony	Sb	51	121.8
Argon	Ar	18	39.95
Arsenic	As	33	74.92
Astatine	At	85	(210)
Barium	Ba	56	137.3
Berkelium	Bk	97	(247)
Beryllium	Be	4	9.012
Bismuth	Bi	83	209.0
Bohrium	Bh	107	(272)
Boron	B	5	10.81
Bromine	Br	35	79.90
Cadmium	Cd	48	112.4
Calcium	Ca	20	40.08
Californium	Cf	98	(251)
Carbon	C	6	12.01
Cerium	Ce	58	140.1
Cesium	Cs	55	132.9
Chlorine	Cl	17	35.45
Chromium	Cr	24	52.00
Cobalt	Co	27	58.93
Copernicium	Cn	112	(285)
Copper	Cu	29	63.55
Curium	Cm	96	(247)
Darmstadtium	Ds	110	(281)
Dubnium	Db	105	(268)
Dysprosium	Dy	66	162.5
Einsteinium	Es	99	(252)
Erbium	Er	68	167.3
Europium	Eu	63	152.0
Fermium	Fm	100	(257)
Flerovium	Fl	114	(289)
Fluorine	F	9	19.00
Francium	Fr	87	(223)
Gadolinium	Gd	64	157.3
Gallium	Ga	31	69.72
Germanium	Ge	32	72.64
Gold	Au	79	197.0
Hafnium	Hf	72	178.5
Hassium	Hs	108	(270)
Helium	He	2	4.003
Holmium	Ho	67	164.9
Hydrogen	H	1	1.008
Indium	In	49	114.8
Iodine	I	53	126.9
Iridium	Ir	77	192.2
Iron	Fe	26	55.85
Krypton	Kr	36	83.80
Lanthanum	La	57	138.9
Lawrencium	Lr	103	(262)
Lead	Pb	82	207.2
Lithium	Li	3	6.941
Livermorium	Lv	116	(293)
Lutetium	Lu	71	175.0
Magnesium	Mg	12	24.31
Manganese	Mn	25	54.94
Meitnerium	Mt	109	(276)
Mendelevium	Md	101	(258)
Mercury	Hg	80	200.6
Molybdenum	Mo	42	95.94
Neodymium	Nd	60	144.2
Neon	Ne	10	20.18
Neptunium	Np	93	(237)
Nickel	Ni	28	58.69
Niobium	Nb	41	92.91
Nitrogen	N	7	14.01
Nobelium	No	102	(259)
Osmium	Os	76	190.2
Oxygen	O	8	16.00
Palladium	Pd	46	106.4
Phosphorus	P	15	30.97
Platinum	Pt	78	195.1
Plutonium	Pu	94	(244)
Polonium	Po	84	(209)
Potassium	K	19	39.10
Praseodymium	Pr	59	140.9
Promethium	Pm	61	(145)
Protactinium	Pa	91	231.0
Radium	Ra	88	(226)
Radon	Rn	86	(222)
Rhenium	Re	75	186.2
Rhodium	Rh	45	102.9
Roentgenium	Rg	111	(280)
Rubidium	Rb	37	85.47
Ruthenium	Ru	44	101.1
Rutherfordium	Rf	104	(267)
Samarium	Sm	62	150.4
Scandium	Sc	21	44.96
Seaborgium	Sg	106	(271)
Selenium	Se	34	78.96
Silicon	Si	14	28.09
Silver	Ag	47	107.9
Sodium	Na	11	22.99
Strontium	Sr	38	87.62
Sulfur	S	16	32.07
Tantalum	Ta	73	180.9
Technetium	Tc	43	(98)
Tellurium	Te	52	127.6
Terbium	Tb	65	158.9
Thallium	Tl	81	204.4
Thorium	Th	90	232.0
Thulium	Tm	69	168.9
Tin	Sn	50	118.7
Titanium	Ti	22	47.88
Tungsten	W	74	183.8
Uranium	U	92	238.0
Vanadium	V	23	50.94
Xenon	Xe	54	131.3
Ytterbium	Yb	70	173.0
Yttrium	Y	39	88.91
Zinc	Zn	30	65.41
Zirconium	Zr	40	91.22
		113**	(284)
		115	(289)
		117	(294)
		118	(294)

*Values in parentheses represent the mass number of the most stable isotope.

**The names and symbols for elements 113, 115, 117, and 118 have not been chosen.

General, Organic, & Biological CHEMISTRY

Third Edition

Janice Gorzynski Smith
University of Hawai‘i at Mānoa

Mc Graw Hill Education

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About the Author

Janice Gorzynski Smith was born in Schenectady, New York, and grew up following the Yankees, listening to the Beatles, and water skiing on Sacandaga Reservoir. She became interested in chemistry in high school, and went on to major in chemistry at Cornell University where she received an A.B. degree *summa cum laude*. Jan earned a Ph.D. in Organic Chemistry from Harvard University under the direction of Nobel Laureate E. J. Corey, and she also spent a year as a National Science Foundation National Needs Postdoctoral Fellow at Harvard. 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 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 in Hawai'i in the 1990s, Jan and her family moved there permanently in 2000. Most recently, she has served as 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. She has also served as the faculty advisor to the student affiliate chapter of the American Chemical Society. In 2003, she received the Chancellor's Citation for Meritorious Teaching.

Jan resides in Hawai'i with her husband Dan, an emergency medicine physician. She has four children and three 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.

Dedicated to my family

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Brief Contents

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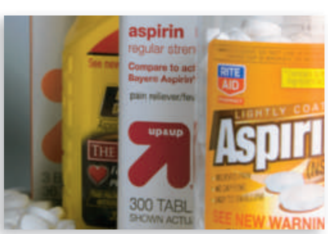

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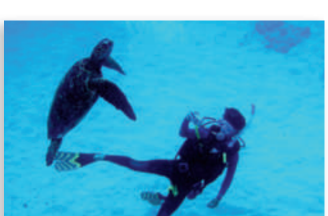

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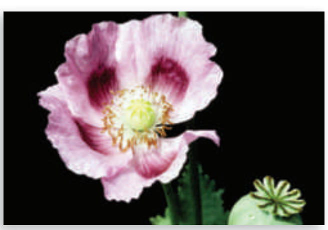

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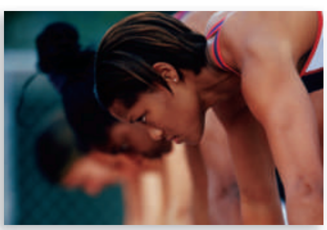

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Preface

My goal in writing this text was to relate the fundamental concepts of general, organic, and biological chemistry to the world around us, and in this way illustrate how chemistry explains many aspects of everyday life. I have followed two guiding principles:

- use relevant and interesting applications for all basic chemical concepts
- present the material in a student-friendly fashion using bulleted lists, extensive illustrations, and step-by-step problem solving

This text is different—by design. Since today's students rely more heavily on visual imagery to learn than ever before, this text uses less prose and more diagrams and figures to reinforce the major themes of chemistry. A key feature is the use of molecular art to illustrate and explain common phenomena we encounter every day. Each topic is broken down into small chunks of information that are more manageable and easily learned. Students are given enough detail to understand basic concepts, such as how soap cleans away dirt and why trans fats are undesirable in the diet, without being overwhelmed.

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 two-semester sequence or a one-semester course. I have found that by introducing one new concept at a time, keeping the basic themes in focus, and breaking down complex problems into small pieces, many students in these chemistry courses acquire a new appreciation of both the human body and the larger world around them.

Building the Text

Writing a textbook is a multifaceted process. McGraw-Hill's 360° Development Process is an ongoing, never ending market-oriented approach to building accurate and innovative print and digital products. 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, intensifies during the development and production stages, and then begins again upon publication, in anticipation of the next edition. 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 360° Development Process includes market research, content reviews, faculty and student focus groups, course- and product-specific symposia, accuracy checks, and art reviews.

"The writing style is accessible and gets the point across. I think that there are ample everyday examples to keep the students interested in the information. I like the bulleted-approach because it allows the students to glance over the text when they are reviewing for either the exam or doing a homework assignment. By looking at the bullets they can easily find the information."

—Daniel Eves,
Southern Utah University

The Learning System Used in *General, Organic, & Biological Chemistry*

- **Writing Style** A concise writing style allows students to focus on learning major concepts and themes of general, organic, and biological chemistry. Relevant materials from everyday life are used to illustrate concepts, and topics are broken into small chunks of information that are more easily learned.
- **Chapter Outline** The chapter outline lists the main headings of the chapter, to help students map out the organization of each chapter's content.

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

KEY CONCEPTS

1 What is pressure and what units are used to measure it? (7.2)
- Pressure is the force per unit area. The pressure of a gas is the force exerted when gas particles strike a surface. Pressure is measured by a barometer and recorded in atmospheres (atm), millimeters of mercury (mm Hg), or pounds per square inch (psi).
- 1 atm = 760 mm Hg = 14.7 psi.

2 What are gas laws and how are they used to describe the relationship between the pressure, volume, and temperature of a gas? (7.3)
- Because gas particles are far apart and behave independently, a set of gas laws describes the behavior of all gases regardless of their identity. Three gas laws—Boyle's law, Charles's law, and Gay–Lussac's law—describe the relationship between the pressure, volume, and temperature of a gas. These gas laws are summarized in "Key Equations—The Gas Laws" on page 254.
- For a constant amount of gas, the following relationships exist.
  - The pressure and volume of a gas are inversely related, so increasing the pressure decreases the volume at constant temperature.
  - The volume of a gas is proportional to its Kelvin temperature, so increasing the temperature increases the volume at constant pressure.
  - The pressure of a gas is proportional to its Kelvin temperature, so increasing the temperature increases the pressure at constant volume.

3 Describe the relationship between the volume and number of moles of a gas. (7.4)
- Avogadro's law states that when temperature and pressure are held constant, the volume of a gas is proportional to its number of moles.
- One mole of any gas has the same volume, the standard molar volume of 22.4 L, at 1 atm and 273 K (STP).

4 What is the ideal gas law? (7.5)
- The ideal gas law is an equation that relates the pressure (*P*), volume (*V*), temperature (*T*), and number of moles (*n*) of a gas; $PV = nRT$, where *R* is the universal gas constant. The ideal gas law can be used to calculate any one of the four variables, as long as the other three variables are known.

5 What is Dalton's law and how is it used to relate partial pressures and the total pressure of a gas mixture? (7.6)
- Dalton's law states that the total pressure of a gas mixture is the sum of the partial pressures of its component gases. The partial pressure is the pressure exerted by each component of a mixture.

6 What types of intermolecular forces exist and how do they determine a compound's boiling point and melting point? (7.7)
- Intermolecular forces are the forces of attraction between molecules. Three types of intermolecular forces exist in covalent compounds. London dispersion forces are due to momentary changes in electron density in a molecule. Dipole–dipole interactions are due to permanent dipoles. Hydrogen bonding, the strongest intermolecular force, results when a H atom bonded to an O, N, or F, is attracted to an O, N, or F atom in another molecule.
- The stronger the intermolecular forces, the higher the boiling point and melting point of a compound.

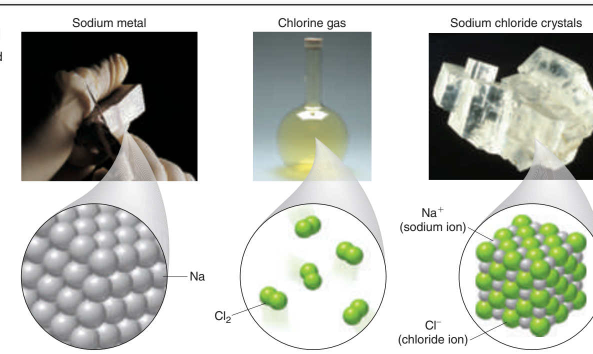
CHAPTER GOALS

In this chapter you will learn how to:

1. Measure pressure and convert one unit of pressure to another
2. Describe the relationship between the pressure, volume, and temperature of a gas using gas laws
3. Describe the relationship between the volume and number of moles of a gas
4. Write the equation for the ideal gas law and use it in calculations
5. Use Dalton's law to determine the partial pressure and total pressure of a gas mixture
6. Determine the types of intermolecular forces in a compound and how these forces determine a compound's boiling point and melting point
7. Describe the properties of a liquid, including vapor pressure, viscosity, and surface tension
8. Describe the features of different types of solids
9. Define specific heat and use specific heat to determine the amount of heat gained or lost by a substance
10. Describe the energy changes that accompany changes of state
11. Interpret the changes depicted in heating and cooling curves

- **Macro-to-Micro Illustrations** Because today's students are visual learners, and because 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 understand the chemistry behind ordinary occurrences.

Figure 3.1
Sodium Chloride, an Ionic Compound



Sodium metal and chlorine gas are both elements. Sodium chloride is an ionic compound composed of sodium ions and chloride ions.

- **Problem Solving** Sample Problems lead students through the thought process tied to successful problem solving by employing Analysis and Solution parts. Sample Problems are categorized sequentially by topic to match chapter organization, and are often 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 appendix at the end of the book.

How To Name an Ionic Compound That Contains a Metal with Variable Charge

Example: Give the name for $CuCl_2$.

Step [1] Determine the charge on the cation.
- Since there are two Cl^- anions, each of which has a −1 charge, the copper cation must have a +2 charge to make the overall charge zero.

$CuCl_2$ → 2 Cl^- anions ---→ The total negative charge is −2.
Cu must have a +2 charge to balance the −2 charge of the anions.
→ Cu^{2+}

Step [2] Name the cation and anion.
- Name the cation using its element name followed by a Roman numeral to indicate its charge. In the common system, use the suffix *-ous* or *-ic* to indicate charge.
- Name the anion by changing the ending of the element name to the suffix *-ide*.

Cu^{2+} ---→ copper(II) or cupric
Cl^- ---→ chloride

Step [3] Write the name of the cation first, then the anion.
- **Answer:** Copper(II) chloride or cupric chloride

- ***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.
- **Applications** Common applications of chemistry to everyday life are found 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."

New to This Edition

Chapter Specific

- A new section on determining the correct number of significant figures when using an electronic calculator has been added to Chapter 1. To help students understand density, a new sample problem and several problems with line art have been added as well.
- Three new sample problems on isotopes, atomic size, and ionization energy are added to Chapter 2 to further assist students in developing problem-solving skills. A new Figure 2.7 better illustrates the order of orbital filling.
- Chapter 3 includes new Sample Problem 3.2 on compounds, elements, and molecules and new Sample Problem 3.3 on ions.
- In Chapter 4, the *How To* on naming covalent compounds has been updated, and a new Sample Problem 4.11, beginning with a molecular model, has been added.
- In response to reviewer feedback, Chapter 5 includes a new Section 5.3 (Types of Reactions), with two new sample problems. The chapter has also been reorganized so that oxidation and reduction appear earlier in the chapter, bringing all the types of reactions together. This addition and reorganization will help students understand and categorize common reactions.
- Chapter 6 contains a new sample problem that illustrates how to use molar mass to convert the number of grams of a reactant to the number of kilocalories released. There is also a new Sample Problem 6.11 on the important topic of equilibrium.
- Coverage of gas laws and gas pressure related to the kinetic-molecular theory has been added to Section 7.2. New Sections 7.10 (Specific Heat) and 7.12C (Combining Energy Calculations) are new, too.
- New material on colloids and suspensions was added to Chapter 8, a topic viewed as particularly useful for nursing students who sometimes give medications that must be shaken before they are administered. Section 8.2 expands the discussion of electrolytes, also now covers equivalents, and includes two new sample problems. It is hoped that this addition will be helpful to many nursing students who deal with equivalents in blood plasma and IV solutions.
- The topic of naming acids was added to Section 9.1 to aid students in identifying common acids.
- Since nuclear chemistry is very different from the material students are exposed to in Chapters 1–9, a new Table 10.2 summarizes the types of nuclear reactions; new art appears in the *How To* in Section 10.3 to better visualize the concept of half-life; and a new Sample Problem 10.5 relates half-life to radioactivity.
- Thiols have been added to Table 11.3. Chapter 11 also includes a new Sample Problem 11.11 on solubility and a new Figure 11.4 on carboxylic acids in oil and vinegar.
- In response to reviewer feedback, the topic of halogenation has been added to Chapter 12, Section 12.9.
- Chapter 16 includes a new *How To* on forming acetals.
- The material on Key Reactions was expanded in Chapter 17 for clarification.
- Based on reviewer feedback, ball-and-stick models replace line structures for bufotenin and psilocin in Section 18.8 in order to make the material more intuitive for students.
- To give students a better understanding of the many aspects of enzyme chemistry, the material on enzymes in Chapter 21 was expanded to include enzyme classification and naming, the effect of temperature and pH, and allosteric control.
- Section 23.6 includes new material on rotenone disrupting the electron transport chain.

General

- **Problem sets.** More problems with molecular art and 3-D models have been added to the text and the ends of the chapters.
- **Design and layout.** An effort has been made with the revised third edition design and layout to move all photos, graphics, and tables closer to related material in the text.
- **Photos.** Roughly one-half 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.
- **Art.** The colors of subatomic particles in all nuclear art were revised for clarity and consistency (Chapters 2, 3, and 10).

Our Commitment to Serving Teachers and Learners

TO THE INSTRUCTOR Writing a chemistry textbook is a colossal task. Teaching chemistry for over 20 years at both a private, liberal arts college and a large state university has given me a unique perspective with which to write this text. I have found that students arrive with vastly different levels of preparation and widely different expectations for their college experience. As an instructor and now an author, I have tried to channel my love and knowledge of chemistry into a form that allows this spectrum of students to understand chemical science more clearly, and then see everyday phenomena in a new light.

TO THE STUDENT I hope that this text and its ancillary program will help you to better understand and appreciate the world of chemistry. My interactions with thousands of students in my long teaching career have profoundly affected the way I teach and write about chemistry, so please feel free to email me with any comments or questions at **jgsmith@hawaii.edu.**

P.A.V.E. the Way to Student Learning

- **To divide two numbers in scientific notation, divide the coefficients and subtract the exponents in the powers of 10.**

For a number written in scientific notation as $y \times 10^x$, y is the coefficient and x is the exponent in the power of 10 (Section 1.6).

$$\frac{6.0 \times 10^{2}}{2.0 \times 10^{20}} = 3.0 \times 10^{-18}$$

Divide coefficients. (6.0 ÷ 2.0) Subtract exponents. (2 − 20)

Sample Problems 5.9 and 5.10 illustrate how to interconvert moles and molecules. In both problems we follow the stepwise procedure for problem solving using conversion factors outlined in Section 1.7B.

SAMPLE PROBLEM 5.9

Converting moles to number of molecules: How many molecules are contained in 5.0 mol of carbon dioxide (CO_2)?

Analysis and Solution

[1] Identify the original quantity and the desired quantity.

5.0 mol of CO_2	? number of molecules of CO_2
original quantity	desired quantity

[2] Write out the conversion factors.

- Choose the conversion factor that places the unwanted unit, mol, in the denominator so the units cancel.

$$\frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{molecules}} \quad \text{or} \quad \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \text{mol}}$$

Choose this conversion factor to cancel mol.

[3] Set up and solve the problem.

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

$$5.0\ \text{mol} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \text{mol}} = 30. \times 10^{23}\ \text{molecules} = 3.0 \times 10^{24}\ \text{molecules of } CO_2$$

Moles cancel. Convert to a number between 1 and 10.

Answer

- Multiplication first gives an answer that is not written in scientific notation since the coefficient (30.) is greater than 10. Moving the decimal point one place to the *left* and *increasing* the exponent by one gives the answer written in the proper form.

PROBLEM 5.18

How many carbon atoms are contained in each of the following number of moles: (a) 2.00 mol; (b) 6.00 mol; (c) 0.500 mol; (d) 25.0 mol?

PROBLEM 5.19

How many molecules are contained in each of the following number of moles?

a. 2.5 mol of penicillin molecules
b. 0.25 mol of NH_3 molecules
c. 0.40 mol of sugar molecules
d. 55.3 mol of acetaminophen molecules

Practice chemistry through stepped-out practice problems and end-of-chapter problems categorized sequentially by topic to match chapter organization. ***How To* boxes** offer step-by-step strategies for difficult concepts.

How To **Use Boyle's Law to Calculate a New Gas Volume or Pressure**

Example: If a 4.0-L container of helium gas has a pressure of 10.0 atm, what pressure does the gas exert if the volume is increased to 6.0 L?

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

- To solve an equation using Boyle's law, we must know three quantities and solve for one quantity. In this case P_1, V_1, and V_2 are known and the final pressure, P_2, must be determined.

$P_1 = 10.0$ atm, $V_1 = 4.0$ L, $V_2 = 6.0$ L — known quantities
$P_2 = ?$ — desired quantity

Step [2] Write the equation and rearrange it to isolate the desired quantity on one side.

- Rearrange the equation for Boyle's law so that the unknown quantity, P_2, is present alone on one side.

$$P_1V_1 = P_2V_2$$

Solve for P_2 by dividing both sides by V_2.

$$\frac{P_1V_1}{V_2} = P_2$$

Step [3] Solve the problem.

- Substitute the known quantities into the equation and solve for P_2. Identical units must be used for two similar quantities (liters in this case) so that the units cancel.

$$P_2 = \frac{P_1V_1}{V_2} = \frac{(10.0\ \text{atm})(4.0\ \text{L})}{6.0\ \text{L}} = 6.7\ \text{atm}$$

Liters cancel. **Answer**

- In this example, the volume increased so the pressure decreased.

SAMPLE PROBLEM 7.2

A tank of compressed air for scuba diving contains 8.5 L of gas at 204 atm pressure. What volume of air does this gas occupy at 1.0 atm?

Analysis

Boyle's law can be used to solve this problem since an initial pressure and volume (P_1 and V_1) and a final pressure (P_2) are known, and a final volume (V_2) must be determined.

Solution

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

$P_1 = 204$ atm, $P_2 = 1.0$ atm, $V_1 = 8.5$ L — known quantities
$V_2 = ?$ — desired quantity

[2] Write the equation and rearrange it to isolate the desired quantity, V_2, on one side.

$$P_1V_1 = P_2V_2$$

Solve for V_2 by dividing both sides by P_2.

$$\frac{P_1V_1}{P_2} = V_2$$

[3] Solve the problem.

- Substitute the three known quantities into the equation and solve for V_2.

"The sample problems in this chapter are very well written and solved with the appropriate level of detail. They illustrate the concepts expected to be learned according to our course objectives."

—Edward Alexander,
San Diego Mesa College

Apply chemistry through **"Focus on Health & Medicine," "Focus on the Human Body,"** and **"Focus on the Environment"** sections woven throughout the text. Chemistry applications are also woven into marginal notes that cover topics on consumer, health, and environmental issues.

5.9C FOCUS ON HEALTH & MEDICINE
The Importance of Percent Yield in the Pharmaceutical Industry

Although some drugs, like the cardiac drug digoxin (used to treat congestive heart failure, Section 1.1), are isolated directly from a natural source, most widely used drugs are synthesized in the laboratory. All common pain relievers—aspirin, acetaminophen, and ibuprofen—are synthetic. The same is true for the bronchodilator albuterol (trade names Proventil or Ventolin), the antidepressant fluoxetine (trade name Prozac), and the cholesterol-lowering medication atorvastatin (trade name Lipitor), whose three-dimensional structures are shown in Figure 5.8.

Once it has been determined that a drug is safe and effective, a pharmaceutical company must be able to prepare large quantities of the material cost-efficiently. This means that cheap and readily available starting materials must be used. It also means that the reactions used to synthesize a drug must proceed in high yield. Rarely is a drug prepared in a single step, and typically, five or more steps may be required in a synthesis.

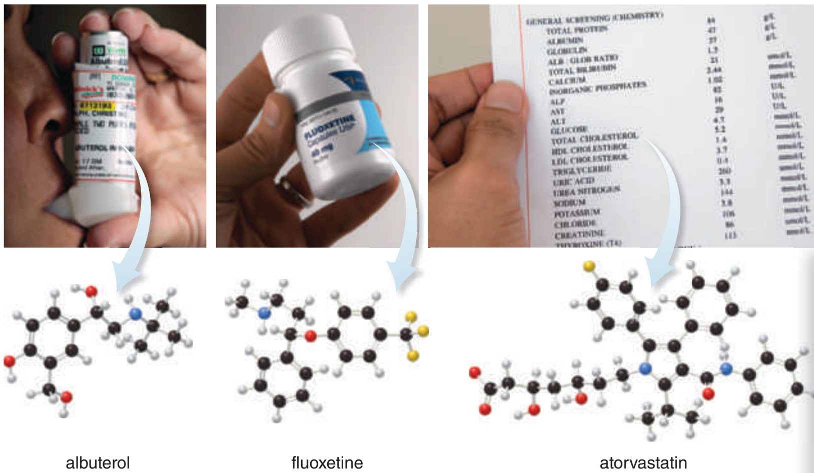
- **To determine the overall percent yield in a synthesis that has more than one step, multiply the percent yield for each step.**

For example, if a synthesis has five steps and each step has a 90.% yield (0.90 written as a decimal), the overall yield is

$$0.90 \times 0.90 \times 0.90 \times 0.90 \times 0.90 = 0.59 = 59\%$$

yield for each step, written as a decimal — **overall yield for five steps**

Figure 5.8 Three Widely Used Synthetic Drugs—Albuterol, Fluoxetine, and Atorvastatin

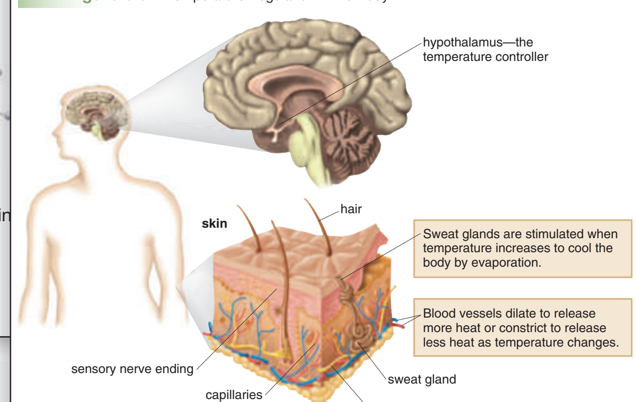


Most commonly prescribed drugs are synthesized in the laboratory. Albuterol (Proventil, Ventolin) is a bronchodilator—that is, it widens airways—and so it is used to treat asthma. Fluoxetine (Prozac) is one of the most common antidepressants currently on the market, used by over 40 million individuals since 1986. Atorvastatin (Lipitor) lowers cholesterol levels and in this way decreases the risk of heart attack and stroke.

"Love it! It is simple, straight forward, complete and logical."

—Susan T. Thomas,
The University of Texas at San Antonio

Figure 6.6 Temperature Regulation in the Body



When the temperature in the environment around the body changes, the body works to counteract the change, in a method similar to Le Châtelier's principle. The hypothalamus acts as a thermostat, which signals the body to respond to temperature changes. When the temperature increases, the body must dissipate excess heat by dilating blood vessels and sweating. When the temperature decreases, blood vessels constrict and the body shivers.

Visualize chemistry through a dynamic art program that brings together macroscopic and microscopic representations of images to help students comprehend on a molecular level. Many illustrations include photos or drawings of everyday objects, paired with their molecular representation, to help students understand the chemistry behind ordinary occurrences. 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.

7.8 The Liquid State

Since liquid molecules are much closer together than gas molecules, many properties of a liquid are determined by the strength of its intermolecular forces. The molecules in a liquid are still much more mobile than those of a solid, though, making liquids fluid and giving them no definite shape. Some liquid molecules move fast enough that they escape the liquid phase altogether and become gas molecules that are very far apart from each other.

7.8A Vapor Pressure

When a liquid is placed in an open container, liquid molecules near the surface that have enough kinetic energy to overcome the intermolecular forces escape to the gas phase. This process, **evaporation,** will continue until all of the liquid has become gas. A puddle of water formed after a rainstorm evaporates as all of the liquid water is converted to gas molecules called water **vapor. Evaporation is an *endothermic* process**—it absorbs heat from the surroundings. This explains why the skin is cooled as sweat evaporates.

HEALTH NOTE

Chloroethane (CH_3CH_2Cl), commonly called ethyl chloride, is a local anesthetic. When chloroethane is sprayed on a wound it quickly evaporates, causing a cooling sensation that numbs the site of an injury.

H_2O

evaporation

In a closed container, some liquid molecules evaporate from the surface and enter the gas phase. As more molecules accumulate in the gas phase, some molecules re-enter the liquid phase in the process of **condensation. Condensation is an *exothermic* process**—it gives off heat to the surroundings. At equilibrium, the rate of evaporation and the rate of condensation are equal.

In **evaporation,** molecules go from liquid to gas.

In **condensation,** molecules go from gas to liquid.

H_2O

liquid in a closed container

The gas laws we have already learned can describe the behavior of the gas molecules above a liquid. In particular, these gas molecules exert pressure, called **vapor pressure.**

- Vapor pressure is the pressure exerted by gas molecules in equilibrium with the liquid phase.

The vapor pressure exerted by a particular liquid depends on the identity of the liquid and the temperature. As the temperature is increased, the kinetic energy of the molecules increases and more molecules escape into the gas phase.

- Vapor pressure increases with increasing temperature.

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—Bobbie Grey,
Riverside City College

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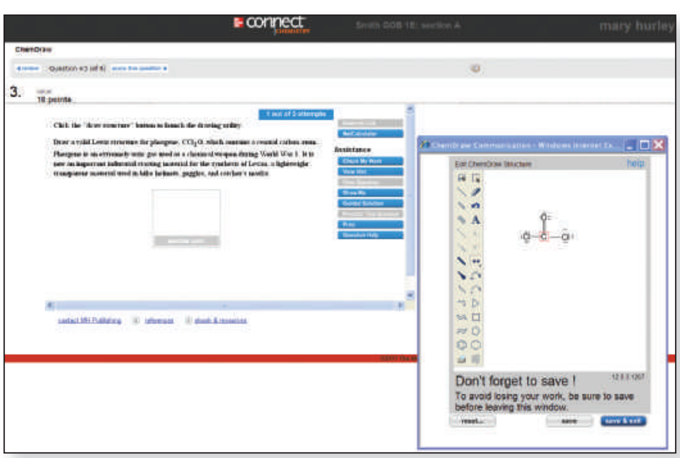

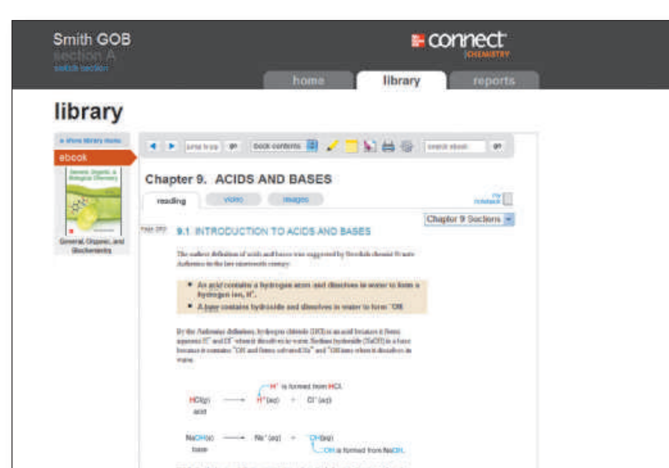

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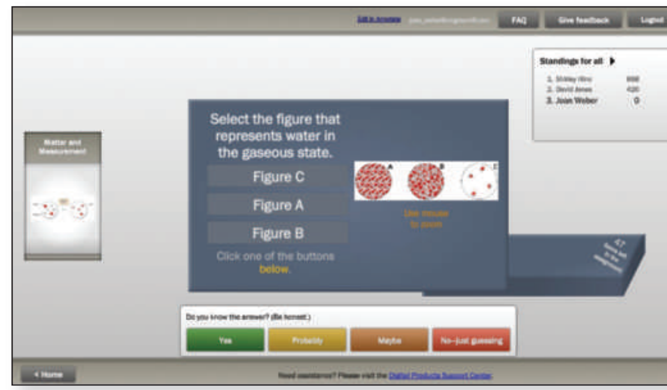

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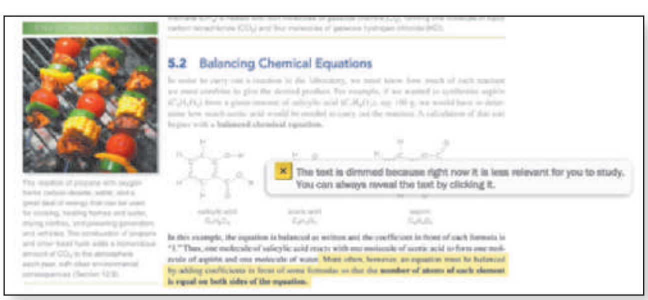

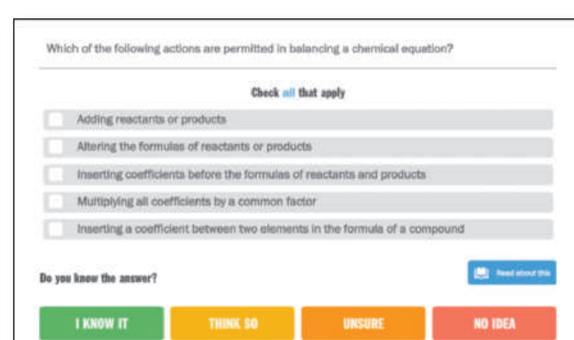

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- Structure and Bonding (31 topics)
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—Chemical Reactions (12 topics)
—Structure and Bonding (37 topics)
—Gases, Liquids, and Solids (24 topics)
—Solutions (9 topics)
—Kinetics and Equilibrium (40 topics)
—Acids and Bases (46 topics)
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REVIEWERS

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List of *How To*'s

List of Applications

1

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

Matter and Measurement

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

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, while other examples of matter—aspirin, CDs, 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, taken a shower with soap, taken notes in a class, checked email on a computer, watched some television, 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

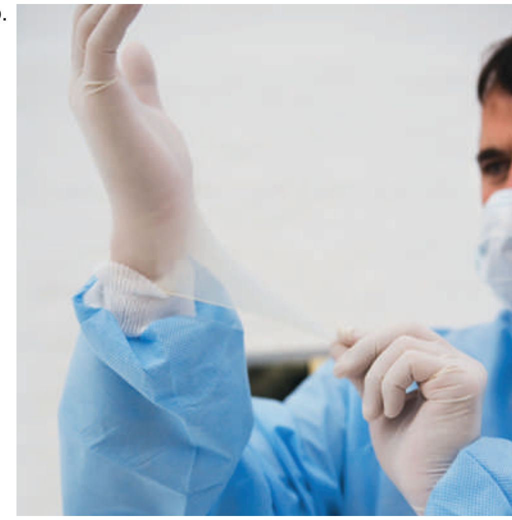
a. Naturally occurring materials b. Synthetic materials

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.

Figure 1.2

Transforming Natural Materials into Useful Synthetic Products

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.

Sometimes a chemist studies what a substance is made of, while at other times he or she might be interested in its properties. Alternatively, 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 CDs, 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.

PROBLEM 1.1

Imagine that your job as a healthcare professional is to take a blood sample from a patient and store it in a small container in a refrigerator until it is picked up for analysis in the hospital lab. You might have to put on gloves and a mask, use a plastic syringe with a metal needle, store the sample in a test tube or vial, and place it in a cold refrigerator. Pick five objects you might encounter during the process and decide if they are made of naturally occurring or synthetic materials.

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.**

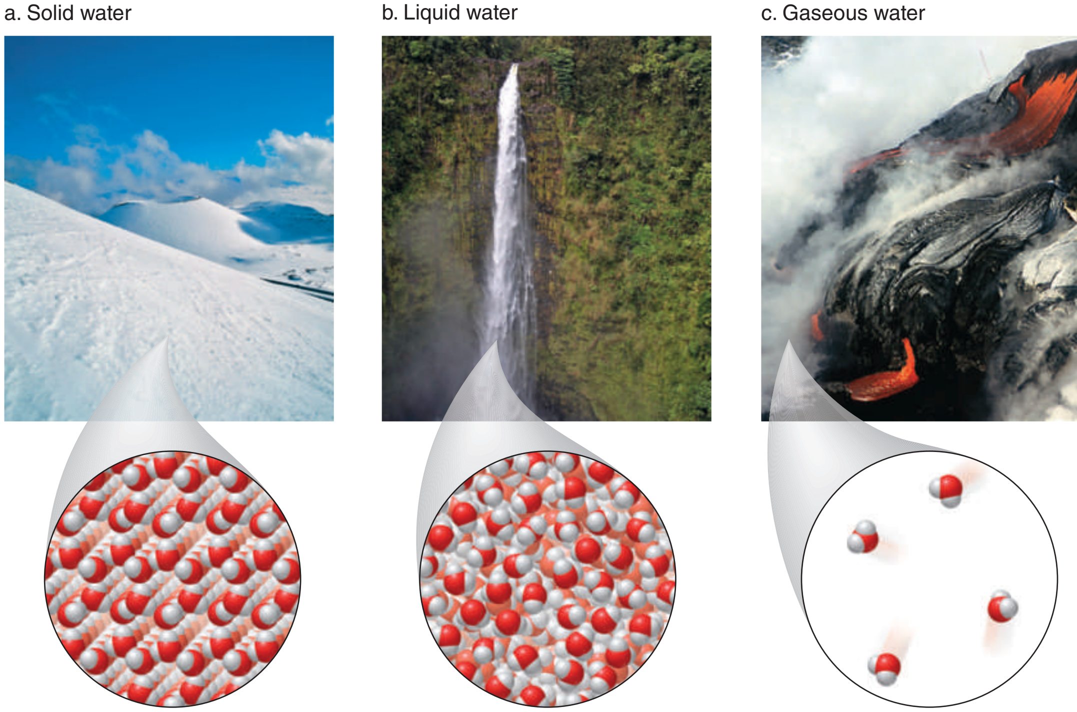
Common physical properties include melting point (mp), boiling point (bp), solubility, color, and odor. **A *physical change* alters a substance without changing its composition.** The most common physical changes are **changes in state.** Melting an ice cube to form liquid water, and boiling liquid 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 7.

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

a. Solid water

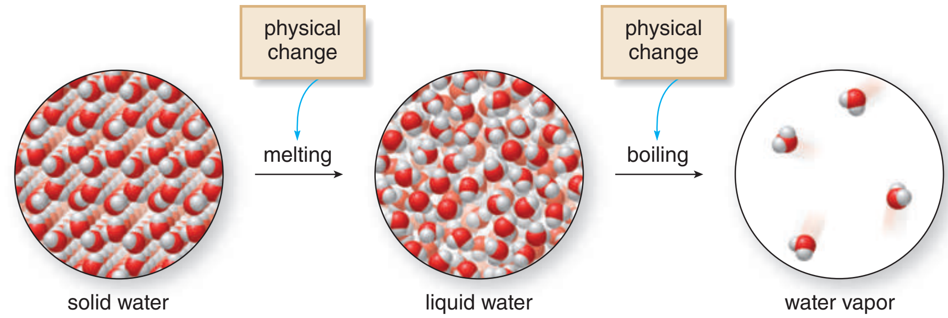
b. Liquid water

c. Gaseous water



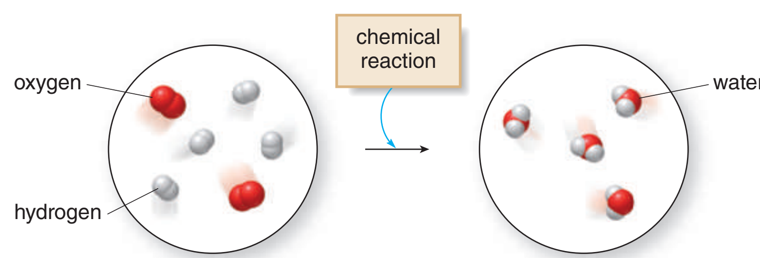
- The particles of a solid are close together and highly organized. (Photo: snow-capped Mauna Kea on the Big Island of Hawaii)
- The particles of a liquid are close together but more disorganized than the solid. (Photo: Akaka Falls on the Big Island of Hawaii)
- The particles of a gas are far apart and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



- *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 Chapters 5 and 6.

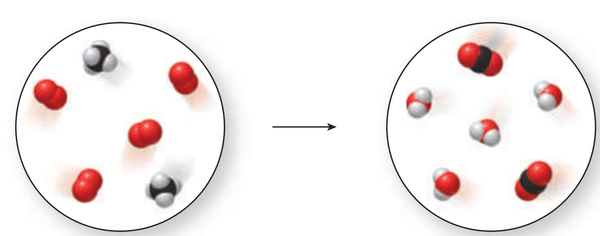


PROBLEM 1.2

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.

PROBLEM 1.3

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



1.3 Classification of Matter

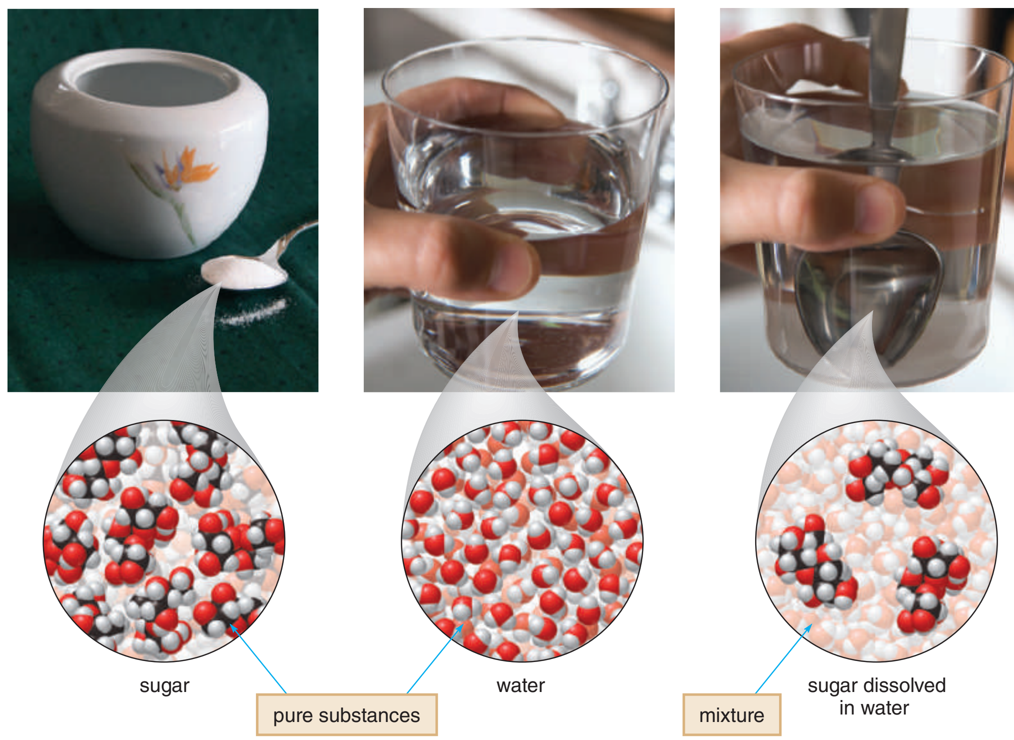
All matter can be classified as either a **pure substance** or a **mixture.**

- 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.

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 liquid water.

Figure 1.4 Two Examples of Mixtures

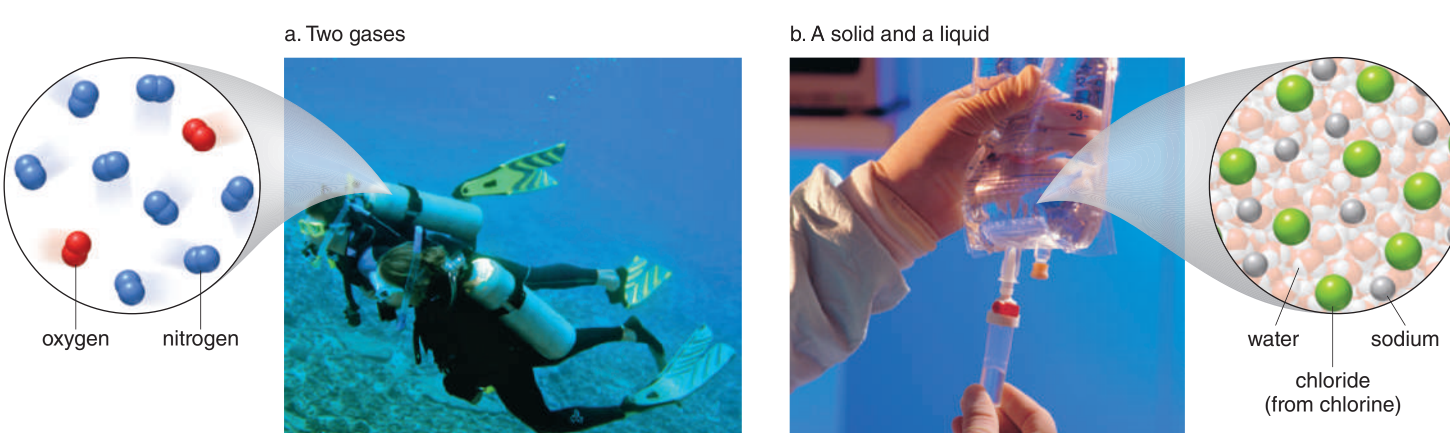
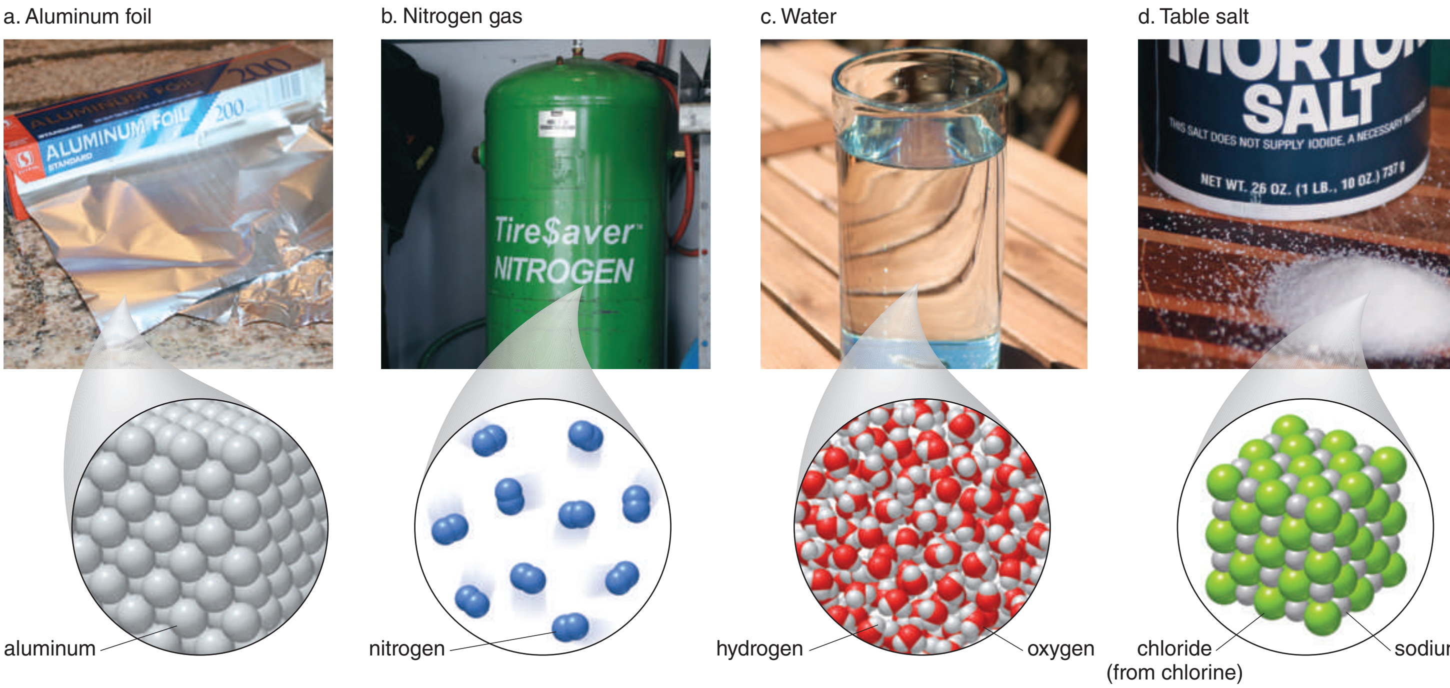


Figure 1.5 Elements and Compounds



- Aluminum foil and nitrogen gas are elements. The molecular art for an element shows spheres of one color only. Thus, aluminum is a solid shown with gray spheres, while 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 pure substance is classified as either an **element** or a **compound.**

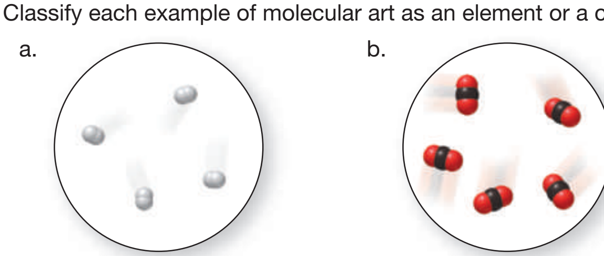
- An *element* is a pure substance that cannot be broken down into simpler substances by a chemical reaction.
- A *compound* is a pure substance formed by chemically combining (joining together) two or more elements.

An alphabetical list of elements is located on the inside front cover of this text. The elements are commonly organized into a periodic table, also shown on the inside front cover, and discussed in much greater detail in Section 2.4.

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 since it is formed from the elements sodium and chlorine (Figure 1.5). Although 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 Chapter 2. Figure 1.6 summarizes the categories into which matter is classified.

SAMPLE PROBLEM 1.1

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



Analysis

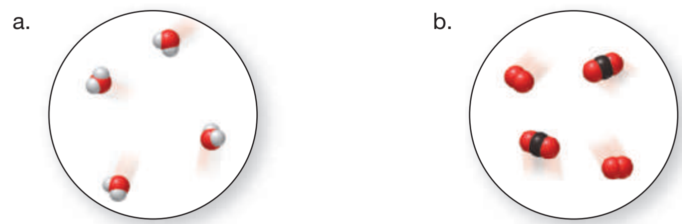
In molecular art, an element is composed of spheres of the same color, while a compound is composed of spheres of different colors.

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.

PROBLEM 1.4

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



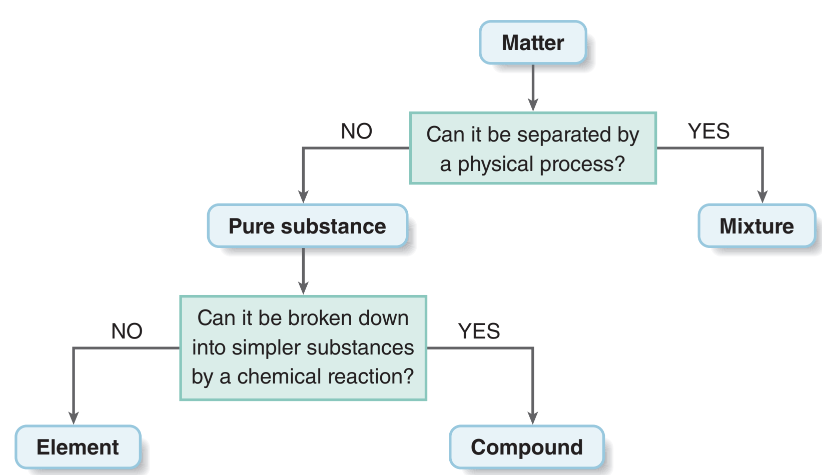
PROBLEM 1.5

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.

PROBLEM 1.6

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. All elements are listed alphabetically on the inside front cover.

Figure 1.6 Classification of Matter



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.

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.*

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.

Table 1.1 Metric Units

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

- **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.

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.**

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