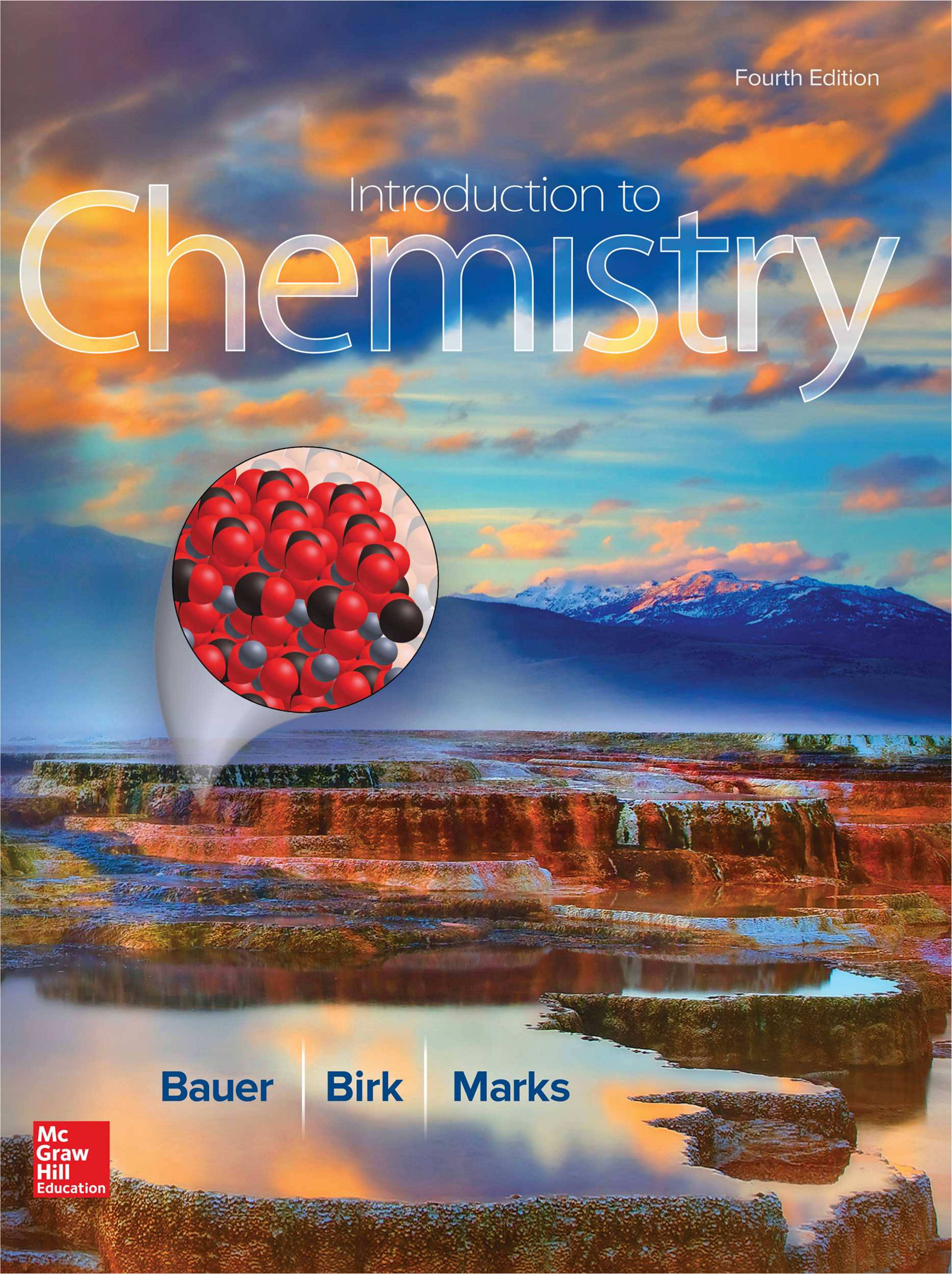


Fourth Edition

Introduction to Chemistry

Bauer | **Birk** | **Marks**

McGraw Hill Education

PERIODIC TABLE OF THE ELEMENTS

- 🟦 Metals (main-group)
- 🟦 Metals (transition)
- 🟦 Metals (inner-transition)
- 🟩 Metals (Metalloids)
- 🟨 Nonmetals

	MAIN-GROUP ELEMENTS		TRANSITION ELEMENTS										MAIN-GROUP ELEMENTS					
								VIIIB										
Period	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	(8)	(9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (18)
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.94	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
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	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** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.97	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.95	43 **Tc** (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
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)
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 (284)	114 **Fl** (289)	115 (288)	116 **Lv** (293)	117 (294)	118 (294)

INNER-TRANSITION ELEMENTS

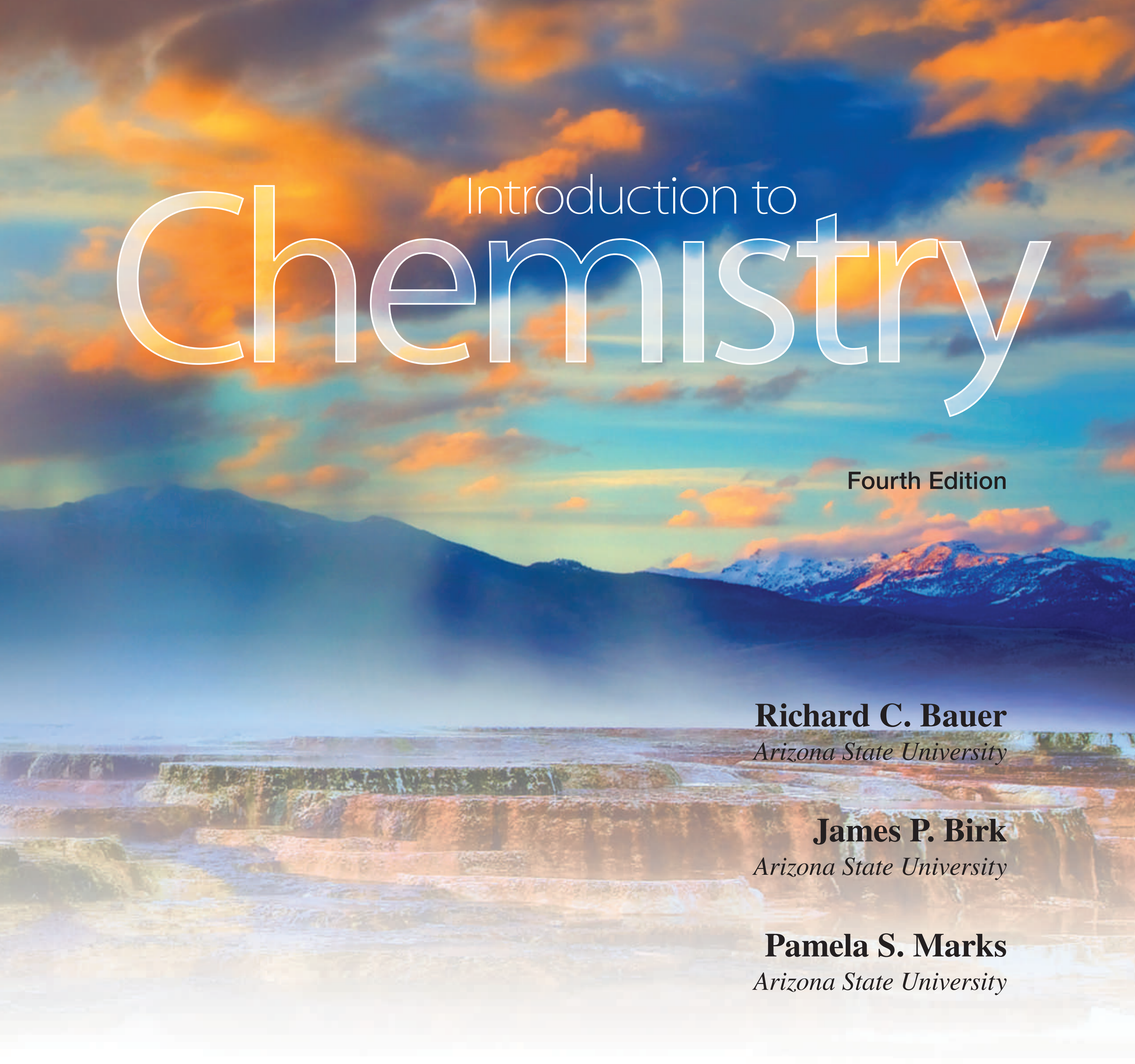
6	Lanthanides	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.1	71 **Lu** 175.0
7	Actinides	90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** (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)

THE ELEMENTS

Element	Symbol	Atomic Number	Relative Atomic Mass*	Element	Symbol	Atomic Number	Relative Atomic Mass*
Actinium	Ac	89	(227)	Mendelevium	Md	101	(258)
Aluminum	Al	13	26.98	Mercury	Hg	80	200.6
Americium	Am	95	(243)	Molybdenum	Mo	42	95.95
Antimony	Sb	51	121.8	Neodymium	Nd	60	144.2
Argon	Ar	18	39.95	Neon	Ne	10	20.18
Arsenic	As	33	74.92	Neptunium	Np	93	(237)
Astatine	At	85	(210)	Nickel	Ni	28	58.69
Barium	Ba	56	137.3	Niobium	Nb	41	92.91
Berkelium	Bk	97	(247)	Nitrogen	N	7	14.01
Beryllium	Be	4	9.012	Nobelium	No	102	(259)
Bismuth	Bi	83	209.0	Osmium	Os	76	190.2
Bohrium	Bh	107	(270)	Oxygen	O	8	16.00
Boron	B	5	10.81	Palladium	Pd	46	106.4
Bromine	Br	35	79.90	Phosphorus	P	15	30.97
Cadmium	Cd	48	112.4	Platinum	Pt	78	195.1
Calcium	Ca	20	40.08	Plutonium	Pu	94	(244)
Californium	Cf	98	(251)	Polonium	Po	84	(209)
Carbon	C	6	12.01	Potassium	K	19	39.10
Cerium	Ce	58	140.1	Praseodymium	Pr	59	140.9
Cesium	Cs	55	132.9	Promethium	Pm	61	(145)
Chlorine	Cl	17	35.45	Protactinium	Pa	91	(231.0)
Chromium	Cr	24	52.00	Radium	Ra	88	(226)
Cobalt	Co	27	58.93	Radon	Rn	86	(222)
Copernicium	Cn	112	(285)	Rhenium	Re	75	186.2
Copper	Cu	29	63.55	Rhodium	Rh	45	102.9
Curium	Cm	96	(247)	Roentgenium	Rg	111	(280)
Darmstadtium	Ds	110	(281)	Rubidium	Rb	37	85.47
Dubnium	Db	105	(268)	Ruthenium	Ru	44	101.1
Dysprosium	Dy	66	162.5	Rutherfordium	Rf	104	(265)
Einsteinium	Es	99	(252)	Samarium	Sm	62	150.4
Erbium	Er	68	167.3	Scandium	Sc	21	44.96
Europium	Eu	63	152.0	Seaborgium	Sg	106	(271)
Fermium	Fm	100	(257)	Selenium	Se	34	78.97
Flerovium	Fl	114	(289)	Silicon	Si	14	28.09
Fluorine	F	9	19.00	Silver	Ag	47	107.9
Francium	Fr	87	(223)	Sodium	Na	11	22.99
Gadolinium	Gd	64	157.3	Strontium	Sr	38	87.62
Gallium	Ga	31	69.72	Sulfur	S	16	32.06
Germanium	Ge	32	72.63	Tantalum	Ta	73	180.9
Gold	Au	79	197.0	Technetium	Tc	43	(98)
Hafnium	Hf	72	178.5	Tellurium	Te	52	127.6
Hassium	Hs	108	(277)	Terbium	Tb	65	158.9
Helium	He	2	4.003	Thallium	Tl	81	204.4
Holmium	Ho	67	164.9	Thorium	Th	90	232.0
Hydrogen	H	1	1.008	Thulium	Tm	69	168.9
Indium	In	49	114.8	Tin	Sn	50	118.7
Iodine	I	53	126.9	Titanium	Ti	22	47.87
Iridium	Ir	77	192.2	Tungsten	W	74	183.8
Iron	Fe	26	55.85	Uranium	U	92	238.0
Krypton	Kr	36	83.80	Vanadium	V	23	50.94
Lanthanum	La	57	138.9	Xenon	Xe	54	131.3
Lawrencium	Lr	103	(262)	Ytterbium	Yb	70	173.1
Lead	Pb	82	207.2	Yttrium	Y	39	88.91
Lithium	Li	3	6.94	Zinc	Zn	30	65.38
Livermorium	Lv	116	(293)	Zirconium	Zr	40	91.22
Lutetium	Lu	71	175.0			113[**]	(284)
Magnesium	Mg	12	24.31			115	(288)
Manganese	Mn	25	54.94			117	(294)
Meitnerium	Mt	109	(276)			118	(294)

*All relative atomic masses are given to four significant figures. 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.



Introduction to Chemistry

Fourth Edition

Richard C. Bauer
Arizona State University

James P. Birk
Arizona State University

Pamela S. Marks
Arizona State University

Mc
Graw
Hill
Education

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To my brother Lucas who, in spite of life's challenges, has a generous spirit; and to Trey for making me laugh, especially at myself.

—Rich Bauer

To my wife my wife, Kay Gunter, who encouraged me through battles with blank pages and shared the joys of completed chapters; and in memory of my parents, Albert and Christine Birk, who taught me to love books enough to see blank pages as a worthwhile challenge.

—Jim Birk

To my husband, Steve, for his love and support, and to my children, Lauren, Kelsey, and Michael, for helping me see aspects of the world differently; also to my mother, Jewel Nicholls, who inspired my love of chemistry at a young age.

—Pam Marks

About the Authors

Richard Bauer was born and raised in Saginaw, Michigan, and completed his B.S. degree in chemistry at Saginaw Valley State University. While pursuing his undergraduate degree he worked at Dow Chemical as a student technologist. He pursued master's and PhD degrees in chemistry education at Purdue University under the direction of Dr. George Bodner. After Purdue, he spent 2 years at Clemson University as a visiting assistant professor.

Dr. Bauer is currently the Faculty Head for Science, Mathematics, and Social Science at the Downtown Phoenix Campus of Arizona State University. He was the General Chemistry Coordinator on the Tempe Campus where he implemented an inquiry-based laboratory program. Dr. Bauer has taught introductory and general chemistry courses for 18 years, and has taught a methods of chemistry teaching course. He is especially fond of teaching introductory chemistry because of the diversity of students enrolled. In addition to general chemistry lab development, Dr. Bauer has interests in student visualization of abstract, molecular-level concepts; TA training; and methods of secondary school chemistry teaching. In addition to his scholarly interests, he plays the piano, sings, and directs choirs.

James Birk is Professor Emeritus of Chemistry and Biochemistry at Arizona State University. Born in Cold Spring, Minnesota, he received a B.A. degree in chemistry from St. John's University (Minnesota) and a PhD in physical chemistry from Iowa State University. After a postdoctorate at the University of Chicago, he started his academic career at the University of Pennsylvania, where he was appointed to the Rhodes-Thompson Chair of Chemistry. Initially doing research on mechanisms of inorganic reactions, he switched to research on various areas of chemical education after moving to Arizona State University as Coordinator of General Chemistry. Dr. Birk's teaching responsibilities have been in general chemistry, introductory chemistry, chemistry for engineers, inorganic chemistry, methods of teaching chemistry, and graduate courses on inorganic reaction mechanisms, chemical education, and science education. He has received several teaching awards, including awards for Distinction in Undergraduate Teaching, Teaching Innovation awards, the National Catalyst Award, and the President's Medal for Team Excellence. He has been a feature editor for the *Journal of Chemical Education*, editing the columns: Filtrates and Residues, The Computer Series, and Teaching with Technology. Recent research has focused on visualization (such as dynamic visualization in chemistry and the hidden earth), on inquiry-based instruction, and on misconceptions (Chemistry Concept Inventory).

Pamela Marks is a principal lecturer at Arizona State University, where her main focus has been teaching introductory and general chemistry for the past 19 years. She has been heavily involved in the general chemistry program at ASU, with a focus on curriculum development. Recently, she implemented changes involving collaborative and online learning. She has also taught in the general chemistry program at the College of St. Benedict and St. John's University in Minnesota. Previous publication materials include multimedia-based general chemistry education materials on CD. She received her B.A. in chemistry from St. Olaf College in 1984 and her M.A. in inorganic chemistry at the University of Arizona in 1988. She spends her free time with her husband, Steve, and their three children, Lauren, Kelsey, and Michael.

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Preface

As we head into our fourth edition of *Introduction to Chemistry*, we still believe in the premise under which we started this book. Our students learn best when we focus our class activities on a conceptual approach to chemistry. Our class meetings are significantly different from traditional lecture presentations in many ways. Beginning with the first week of classes and continuing through the rest of the semester, we follow a sequence of topics that allows us to explain macroscopic phenomena from a molecular perspective. This approach places emphasis on conceptual understanding over algorithmic problem solving. To help students develop conceptual understanding, we use numerous images, animations, video clips, and live demonstrations. Roughly a third of each class period is devoted to explaining chemical phenomena from a conceptual perspective. During the remaining time, students work in groups to discuss and answer conceptual and numerical questions.

Our desire to create a conceptually based text stems from our own classroom experience, as well as from educational research about how students learn. This book is grounded in educational research findings that address topic sequence, context, conceptual emphasis, and concept-embedded numerical problem solving. Throughout the text, we have made an effort to relate the content to students' daily lives and show them how chemistry allows us to understand the phenomena—both simple and complex—that we encounter on a regular basis. Students' initial exposure to chemical concepts should be in the realm of their personal experience, to give context to the abstract concepts we want them to understand later. This text presents macroscopic chemical phenomena early and uses familiar contexts to develop microscopic explanations.

This textbook is designed for the freshman-level Introductory Chemistry course that does not have a chemistry prerequisite and is suitable for either a one-semester course or a two-semester sequence. The book targets introductory courses taken by non-physical science majors who may be in allied health, agriculture, or other disciplines that do not require the rigor of a science major's General Chemistry course, or for students fulfilling university liberal arts requirements for science credits. In addition, students who lack a strong high school science background often take the course as a preparation for the regular general chemistry sequence.

New Features

Consider This questions have been added at the end of worked examples. These questions have been written to be conceptual in nature and often ask students to extend their understanding beyond the focus of the worked example.

Key Concepts have replaced the end-of-chapter summaries. These are presented in outline form and can be used by students as a guide for the most important ideas discussed in the chapter.

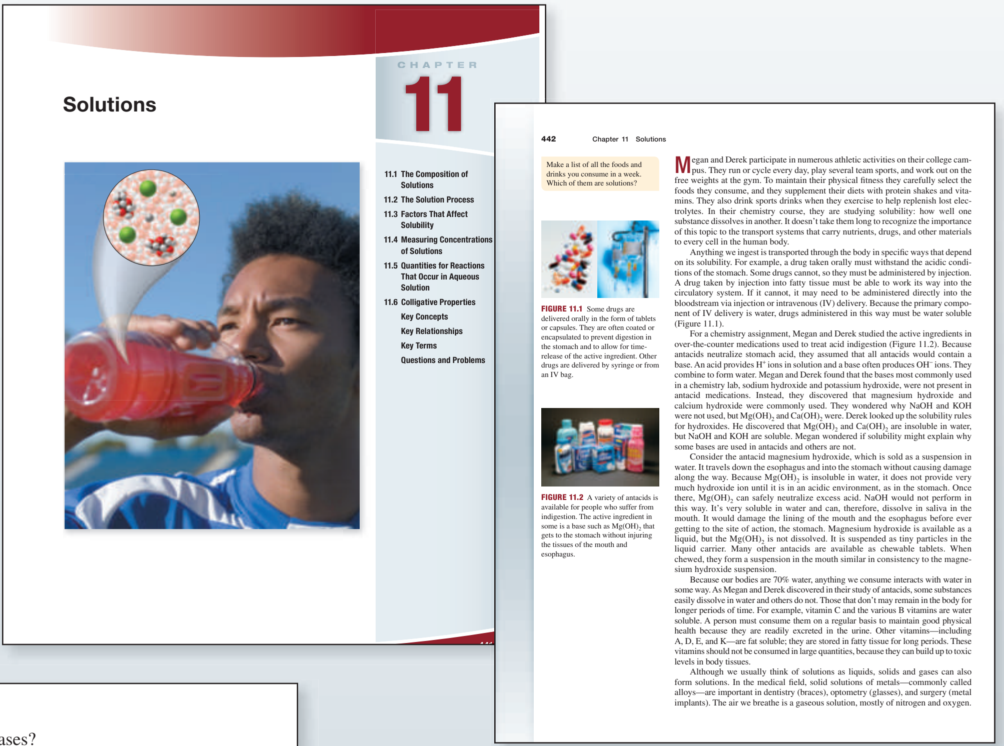
Concept Review multiple-choice questions have been added to end-of-chapter questions and problems. Most students enrolled in an introductory chemistry course will take multiple-choice exams, so these questions provide them with an opportunity to get some practice. The conceptual nature of these questions also helps develop deeper understanding and critical thinking skills. After each concept review question, there is a follow-up question that provides additional practice at analysis of multiple-choice responses.

Revised Periodic Table Since the third edition of this book, two new elements have been named: flerovium and livermorium. A few atomic weights have also been revised by IUPAC.

Features of this Text

Learning theory indicates that we should start with the concrete, macroscopic world of experience as the basis for developing student understanding of abstract, microscopic concepts. This textbook follows a topic sequence typically found in many general chemistry texts. That is, macroscopic ideas about chemical behavior are discussed before descriptions of abstract, molecular-level concepts associated with electron structure. The macroscopic ideas that begin chapters or sections are grounded in real-life experiences. Where appropriate, the macroscopic to molecular-level progression of ideas is carried over to topic sequence within individual chapters or sections in addition to the general sequence of chapters.

Each chapter begins with a chapter-opening outline and an **opening vignette** that personalizes the content by telling a story about chemical phenomena encountered by students. These applications help students see how chemistry relates to their daily lives.



Solutions

CHAPTER

11

11.1 The Composition of Solutions
11.2 The Solution Process
11.3 Factors That Affect Solubility
11.4 Measuring Concentrations of Solutions
11.5 Quantities for Reactions That Occur in Aqueous Solution
11.6 Colligative Properties
Key Concepts
Key Relationships
Key Terms
Questions and Problems

442 Chapter 11 Solutions

Make a list of all the foods and drinks you consume in a week. Which of them are solutions?

FIGURE 11.1 Some drugs are delivered orally in the form of tablets or capsules. They are often coated or encapsulated to prevent digestion in the stomach and to allow for time-release of the active ingredient. Other drugs are delivered by syringe or from an IV bag.

FIGURE 11.2 A variety of antacids is available for people who suffer from indigestion. The active ingredient in some is a base such as $Mg(OH)_2$ that gets to the stomach without injuring the tissues of the mouth and esophagus.

Megan and Derek participate in numerous athletic activities on their college campus. They run or cycle every day, play several team sports, and work out on the free weights at the gym. To maintain their physical fitness they carefully select the foods they consume, and they supplement their diets with protein shakes and vitamins. They also drink sports drinks when they exercise to help replenish lost electrolytes. In their chemistry course, they are studying solubility: how well one substance dissolves in another. It doesn't take them long to recognize the importance of this topic to the transport systems that carry nutrients, drugs, and other materials to every cell in the human body.

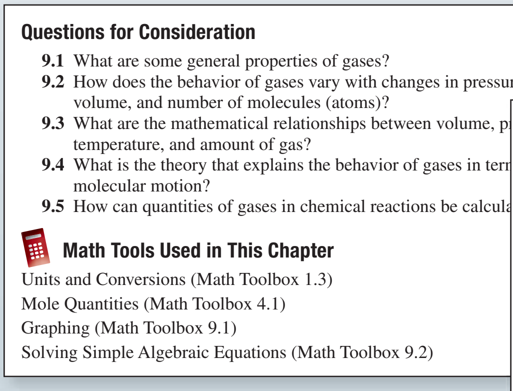
Anything we ingest is transported through the body in specific ways that depend on its solubility. For example, a drug taken orally must withstand the acidic conditions of the stomach. Some drugs cannot, so they must be administered by injection. A drug taken by injection into fatty tissue must be able to work its way into the circulatory system. If it cannot, it may need to be administered directly into the bloodstream via injection or intravenous (IV) delivery. Because the primary component of IV delivery is water, drugs administered in this way must be water soluble (Figure 11.1).

For a chemistry assignment, Megan and Derek studied the active ingredients in over-the-counter medications used to treat acid indigestion (Figure 11.2). Because antacids neutralize stomach acid, they assumed that all antacids would contain a base. An acid provides H^+ ions in solution and a base often produces OH^- ions. They combine to form water. Megan and Derek found that the bases most commonly used in a chemistry lab, sodium hydroxide and potassium hydroxide, were not present in antacid medications. Instead, they discovered that magnesium hydroxide and calcium hydroxide were commonly used. They wondered why NaOH and KOH were not used, but $Mg(OH)_2$ and $Ca(OH)_2$ were. Derek looked up the solubility rules for hydroxides. He discovered that $Mg(OH)_2$ and $Ca(OH)_2$ are insoluble in water, but NaOH and KOH are soluble. Megan wondered if solubility might explain why some bases are used in antacids and others are not.

Consider the antacid magnesium hydroxide, which is sold as a suspension in water. It travels down the esophagus and into the stomach without causing damage along the way. Because $Mg(OH)_2$ is insoluble in water, it does not provide very much hydroxide ion until it is in an acidic environment, as in the stomach. Once there, $Mg(OH)_2$ can safely neutralize excess acid. NaOH would not perform in this way. It's very soluble in water and can, therefore, dissolve in saliva in the mouth. It would damage the lining of the mouth and the esophagus before ever getting to the site of action, the stomach. Magnesium hydroxide is available as a liquid, but the $Mg(OH)_2$ is not dissolved. It is suspended as tiny particles in the liquid carrier. Many other antacids are available as chewable tablets. When chewed, they form a suspension in the mouth similar in consistency to the magnesium hydroxide suspension.

Because our bodies are 70% water, anything we consume interacts with water in some way. As Megan and Derek discovered in their study of antacids, some substances easily dissolve in water and others do not. Those that don't may remain in the body for longer periods of time. For example, vitamin C and the various B vitamins are water soluble. A person must consume them on a regular basis to maintain good physical health because they are readily excreted in the urine. Other vitamins—including A, D, E, and K—are fat soluble; they are stored in fatty tissue for long periods. These vitamins should not be consumed in large quantities, because they can build up to toxic levels in body tissues.

Although we usually think of solutions as liquids, solids and gases can also form solutions. In the medical field, solid solutions of metals—commonly called alloys—are important in dentistry (braces), optometry (glasses), and surgery (metal implants). The air we breathe is a gaseous solution, mostly of nitrogen and oxygen.



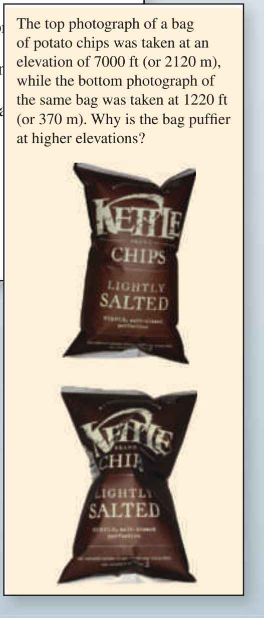
Questions for Consideration

9.1 What are some general properties of gases?
9.2 How does the behavior of gases vary with changes in pressure, temperature, volume, and number of molecules (atoms)?
9.3 What are the mathematical relationships between volume, p
temperature, and amount of gas?
9.4 What is the theory that explains the behavior of gases in ter
molecular motion?
9.5 How can quantities of gases in chemical reactions be calcula

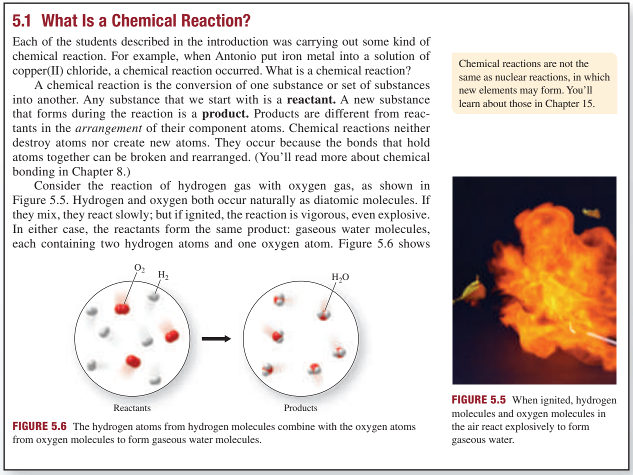
Math Tools Used in This Chapter

Units and Conversions (Math Toolbox 1.3)
Mole Quantities (Math Toolbox 4.1)
Graphing (Math Toolbox 9.1)
Solving Simple Algebraic Equations (Math Toolbox 9.2)

The top photograph of a bag of potato chips was taken at an elevation of 7000 ft (or 2120 m), while the bottom photograph of the same bag was taken at 1220 ft (or 370 m). Why is the bag puffier at higher elevations?



The chapter then offers some guiding questions typical of inquiry instruction. These **Questions for Consideration** serve as a guide in topic development through the chapter. **Margin notes** contain further explanations and chemical applications, combined with visuals, to help students conceptualize lessons.



5.1 What Is a Chemical Reaction?

Each of the students described in the introduction was carrying out some kind of chemical reaction. For example, when Antonio put iron metal into a solution of copper(II) chloride, a chemical reaction occurred. What is a chemical reaction?

A chemical reaction is the conversion of one substance or set of substances into another. Any substance that we start with is a **reactant.** A new substance that forms during the reaction is a **product.** Products are different from reactants in the *arrangement* of their component atoms. Chemical reactions neither destroy atoms nor create new atoms. They occur because the bonds that hold atoms together can be broken and rearranged. (You'll read more about chemical bonding in Chapter 8.)

Consider the reaction of hydrogen gas with oxygen gas, as shown in Figure 5.5. Hydrogen and oxygen both occur naturally as diatomic molecules. If they mix, they react slowly; but if ignited, the reaction is vigorous, even explosive. In either case, the reactants form the same product: gaseous water molecules, each containing two hydrogen atoms and one oxygen atom. Figure 5.6 shows

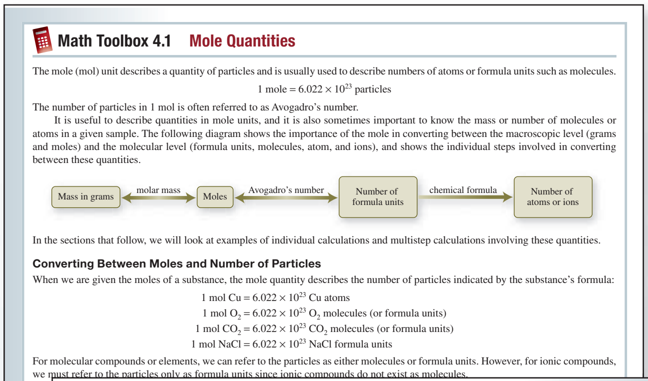
Chemical reactions are not the same as nuclear reactions, in which new elements may form. You'll learn about those in Chapter 15.

FIGURE 5.6 The hydrogen atoms from hydrogen molecules combine with the oxygen atoms from oxygen molecules to form gaseous water molecules.

FIGURE 5.5 When ignited, hydrogen molecules and oxygen molecules in the air react explosively to form gaseous water.

We believe that an introductory chemistry textbook should maintain a focus on *chemistry*, rather than on math. Students' interest must be captured early if they're going to persevere in the class. Early in this text, we introduce chemical reactions from macroscopic perspectives. A general fundamental knowledge of chemical behavior on a macroscopic level facilitates further development of molecular-level ideas, such as atomic structure.

We believe that the best approach to incorporating math involves development of associated math on an as-needed basis with an emphasis on concepts that problems are trying to illustrate. This text integrates need-to-know mathematical ideas that are important to chemists into conceptual discussions. **Math toolboxes** include a thorough explanation of the math, examples, worked-out solutions, and practice problems.



Math Toolbox 4.1 Mole Quantities

The mole (mol) unit describes a quantity of particles and is usually used to describe numbers of atoms or formula units such as molecules.

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ particles}$$

The number of particles in 1 mol is often referred to as Avogadro's number.

It is useful to describe quantities in mole units, and it is also sometimes important to know the mass or number of molecules or atoms in a given sample. The following diagram shows the importance of the mole in converting between the macroscopic level (grams and moles) and the molecular level (formula units, molecules, atom, and ions), and shows the individual steps involved in converting between these quantities.

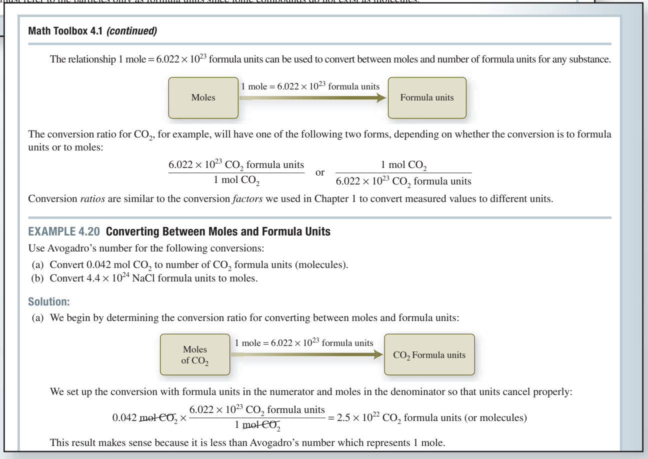
In the sections that follow, we will look at examples of individual calculations and multistep calculations involving these quantities.

Converting Between Moles and Number of Particles

When we are given the moles of a substance, the mole quantity describes the number of particles indicated by the substance's formula:

$$1 \text{ mol Cu} = 6.022 \times 10^{23} \text{ Cu atoms}$$
$$1 \text{ mol } O_2 = 6.022 \times 10^{23}\ O_2 \text{ molecules (or formula units)}$$
$$1 \text{ mol } CO_2 = 6.022 \times 10^{23}\ CO_2 \text{ molecules (or formula units)}$$
$$1 \text{ mol NaCl} = 6.022 \times 10^{23} \text{ NaCl formula units}$$

For molecular compounds or elements, we can refer to the particles as either molecules or formula units. However, for ionic compounds, we must refer to the particles only as formula units since ionic compounds do not exist as molecules.



Math Toolbox 4.1 ***(continued)***

The relationship 1 mole = 6.022×10^{23} formula units can be used to convert between moles and number of formula units for any substance.

The conversion ratio for CO_2, for example, will have one of the following two forms, depending on whether the conversion is to formula units or to moles:

$$\frac{6.022 \times 10^{23}\ CO_2 \text{ formula units}}{1 \text{ mol } CO_2} \quad \text{or} \quad \frac{1 \text{ mol } CO_2}{6.022 \times 10^{23}\ CO_2 \text{ formula units}}$$

Conversion *ratios* are similar to the conversion *factors* we used in Chapter 1 to convert measured values to different units.

EXAMPLE 4.20 Converting Between Moles and Formula Units

Use Avogadro's number for the following conversions:

(a) Convert 0.042 mol CO_2 to number of CO_2 formula units (molecules).
(b) Convert 4.4×10^{24} NaCl formula units to moles.

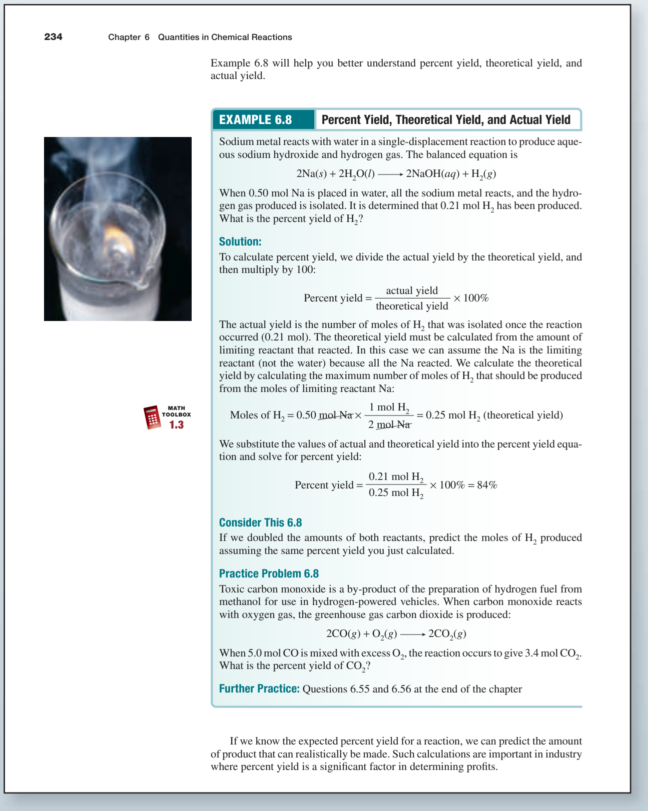
Solution:

(a) We begin by determining the conversion ratio for converting between moles and formula units:

We set up the conversion with formula units in the numerator and moles in the denominator so that units cancel properly:

$$0.042\ \cancel{\text{mol } CO_2} \times \frac{6.022 \times 10^{23}\ CO_2 \text{ formula units}}{1\ \cancel{\text{mol } CO_2}} = 2.5 \times 10^{22}\ CO_2 \text{ formula units (or molecules)}$$

This result makes sense because it is less than Avogadro's number which represents 1 mole.



234 Chapter 6 Quantities in Chemical Reactions

Example 6.8 will help you better understand percent yield, theoretical yield, and actual yield.

EXAMPLE 6.8 Percent Yield, Theoretical Yield, and Actual Yield

Sodium metal reacts with water in a single-displacement reaction to produce aqueous sodium hydroxide and hydrogen gas. The balanced equation is

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

When 0.50 mol Na is placed in water, all the sodium metal reacts, and the hydrogen gas produced is isolated. It is determined that 0.21 mol H_2 has been produced. What is the percent yield of H_2?

Solution:

To calculate percent yield, we divide the actual yield by the theoretical yield, and then multiply by 100:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

The actual yield is the number of moles of H_2 that was isolated once the reaction occurred (0.21 mol). The theoretical yield must be calculated from the amount of limiting reactant that reacted. In this case we can assume the Na is the limiting reactant (not the water) because all the Na reacted. We calculate the theoretical yield by calculating the maximum number of moles of H_2 that should be produced from the moles of limiting reactant Na:

MATH TOOLBOX 1.3

$$\text{Moles of } H_2 = 0.50\ \cancel{\text{mol Na}} \times \frac{1 \text{ mol } H_2}{2\ \cancel{\text{mol Na}}} = 0.25 \text{ mol } H_2 \text{ (theoretical yield)}$$

We substitute the values of actual and theoretical yield into the percent yield equation and solve for percent yield:

$$\text{Percent yield} = \frac{0.21 \text{ mol } H_2}{0.25 \text{ mol } H_2} \times 100\% = 84\%$$

Consider This 6.8

If we doubled the amounts of both reactants, predict the moles of H_2 produced assuming the same percent yield you just calculated.

Practice Problem 6.8

Toxic carbon monoxide is a by-product of the preparation of hydrogen fuel from methanol for use in hydrogen-powered vehicles. When carbon monoxide reacts with oxygen gas, the greenhouse gas carbon dioxide is produced:

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

When 5.0 mol CO is mixed with excess O_2, the reaction occurs to give 3.4 mol CO_2. What is the percent yield of CO_2?

Further Practice: Questions 6.55 and 6.56 at the end of the chapter

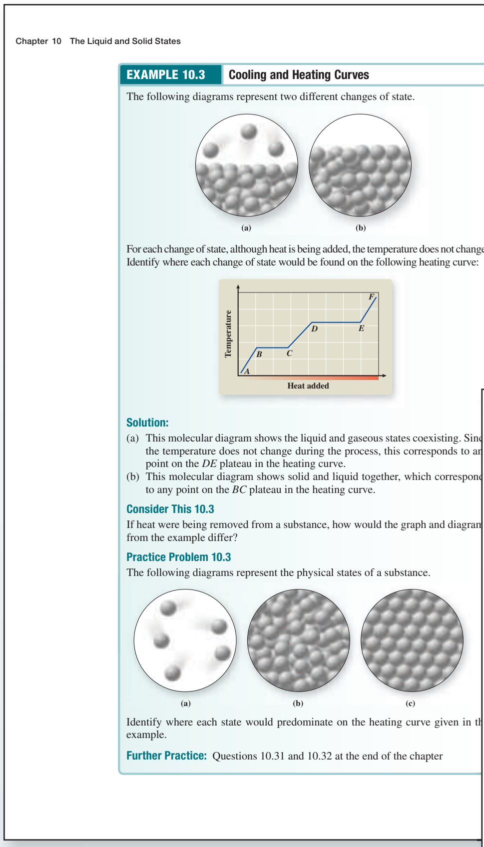
If we know the expected percent yield for a reaction, we can predict the amount of product that can realistically be made. Such calculations are important in industry where percent yield is a significant factor in determining profits.

Toolboxes are referenced with toolbox icons, where appropriate. As problem solving is developed within the text, emphasis is placed on the underlying concepts, letting the numerical solutions emerge from conceptual understanding. Numerical-type problems often ask students to estimate answers and to consider the physical meaning of calculated quantities.

The problem-solving approach used in this text is supported by worked example boxes that contain the following steps:

- question(s)
- solution
- consider this question
- practice problems
- further practice

Problem solving in chemistry is much more than algorithmic number crunching. It involves applying principles to solve conceptual as well as numerical problems. Conceptual problems are those that require students to apply their understanding of concepts instead of applying an algorithm. This text emphasizes the underlying concepts when discussing numerical problems within in-chapter worked examples. Many end-of-chapter problems also emphasize conceptual problem solving.



Chapter 10 The Liquid and Solid States

EXAMPLE 10.3 **Cooling and Heating Curves**

The following diagrams represent two different changes of state.

For each change of state, although heat is being added, the temperature does not change. Identify where each change of state would be found on the following heating curve:

Solution:

(a) This molecular diagram shows the liquid and gaseous states coexisting. Sin
the temperature does not change during the process, this corresponds to a
point on the *DE* plateau in the heating curve.

(b) This molecular diagram shows solid and liquid together, which correspon
to any point on the *BC* plateau in the heating curve.

Consider This 10.3

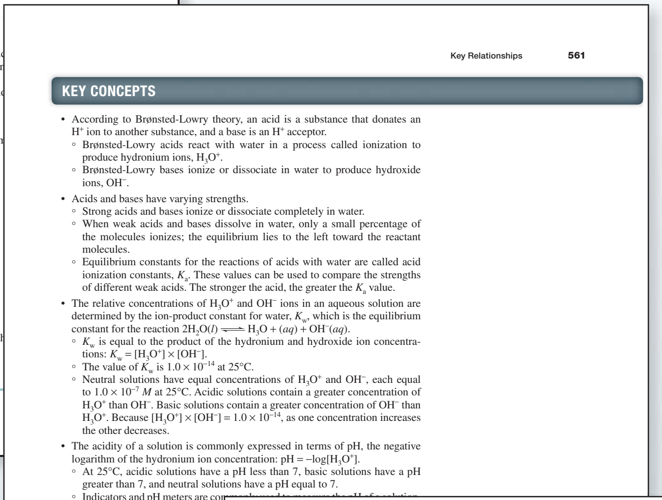
If heat were being removed from a substance, how would the graph and diagra
from the example differ?

Practice Problem 10.3

The following diagrams represent the physical states of a substance.

Identify where each state would predominate on the heating curve given in t
example.

Further Practice: Questions 10.31 and 10.32 at the end of the chapter



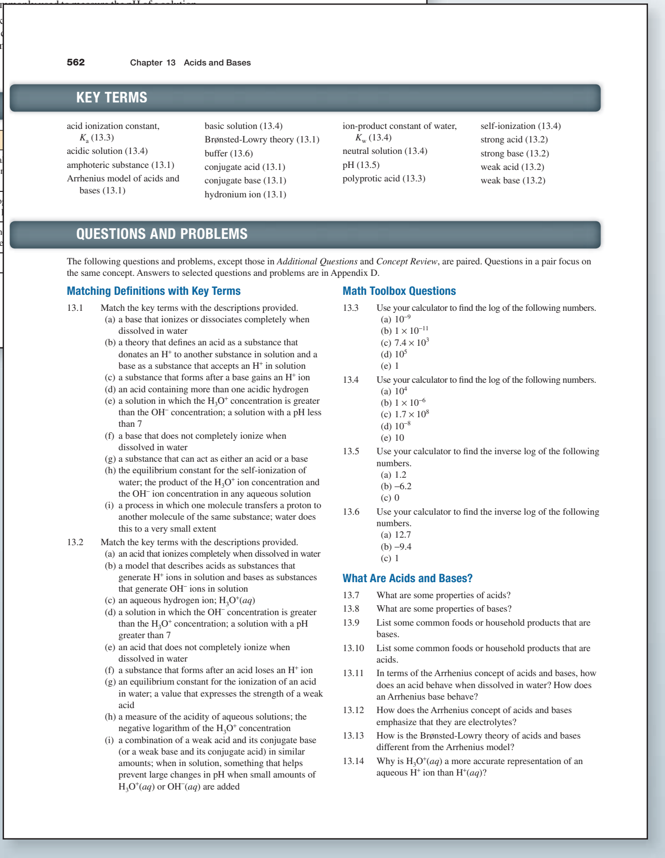
Key Relationships 561

KEY CONCEPTS

- According to Brønsted-Lowry theory, an acid is a substance that donates an H^+ ion to another substance, and a base is an H^+ acceptor.
  - Brønsted-Lowry acids react with water in a process called ionization to produce hydronium ions, H_3O^+.
  - Brønsted-Lowry bases ionize or dissociate in water to produce hydroxide ions, OH^-.
- Acids and bases have varying strengths.
  - Strong acids and bases ionize or dissociate completely in water.
  - When weak acids and bases dissolve in water, only a small percentage of the molecules ionizes; the equilibrium lies to the left toward the reactant molecules.
  - Equilibrium constants for the reactions of acids with water are called acid ionization constants, K_a. These values can be used to compare the strengths of different weak acids. The stronger the acid, the greater the K_a value.
- The relative concentrations of H_3O^+ and OH^- ions in an aqueous solution are determined by the ion-product constant for water, K_w, which is the equilibrium constant for the reaction $2H_2O(l) \rightleftharpoons H_3O + (aq) + OH^-(aq)$.
  - K_w is equal to the product of the hydronium and hydroxide ion concentrations: $K_w = [H_3O^+] \times [OH^-]$.
  - The value of K_w is 1.0×10^{-14} at 25°C.
  - Neutral solutions have equal concentrations of H_3O^+ and OH^-, each equal to 1.0×10^{-7} *M* at 25°C. Acidic solutions contain a greater concentration of H_3O^+ than OH^-. Basic solutions contain a greater concentration of OH^- than H_3O^+. Because $[H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}$, as one concentration increases the other decreases.
- The acidity of a solution is commonly expressed in terms of pH, the negative logarithm of the hydronium ion concentration: $pH = -\log[H_3O^+]$.
  - At 25°C, acidic solutions have a pH less than 7, basic solutions have a pH greater than 7, and neutral solutions have a pH equal to 7.
  - Indicators and pH meters are co
- Buffered solutions contain a weak
and its conjugate acid) in similar
changes in pH by reacting with sm

KEY RELATIONSHIPS

Relationship
The self-ionization constant for water equa concentration and the hydroxide ion conce 1.0×10^{-14}.
The pH of a solution equals the negative lo The pOH of a solution equals the negative
The hydronium ion concentration equals th The hydroxide ion concentration equals the
The pH and pOH of a solution total 14.00.



562 Chapter 13 Acids and Bases

KEY TERMS

acid ionization constant, K_a (13.3)
acidic solution (13.4)
amphoteric substance (13.1)
Arrhenius model of acids and bases (13.1)
basic solution (13.4)
Brønsted-Lowry theory (13.1)
buffer (13.6)
conjugate acid (13.1)
conjugate base (13.1)
hydronium ion (13.1)
ion-product constant of water, K_w (13.4)
neutral solution (13.4)
pH (13.5)
polyprotic acid (13.3)
self-ionization (13.4)
strong acid (13.2)
strong base (13.2)
weak acid (13.2)
weak base (13.2)

QUESTIONS AND PROBLEMS

The following questions and problems, except those in *Additional Questions* and *Concept Review*, are paired. Questions in a pair focus on the same concept. Answers to selected questions and problems are in Appendix D.

Matching Definitions with Key Terms

13.1 Match the key terms with the descriptions provided.
  (a) a base that ionizes or dissociates completely when dissolved in water
  (b) a theory that defines an acid as a substance that donates an H^+ to another substance in solution and a base as a substance that accepts an H^+ in solution
  (c) a substance that forms after a base gains an H^+ ion
  (d) an acid containing more than one acidic hydrogen
  (e) a solution in which the H_3O^+ concentration is greater than the OH^- concentration; a solution with a pH less than 7
  (f) a base that does not completely ionize when dissolved in water
  (g) a substance that can act as either an acid or a base
  (h) the equilibrium constant for the self-ionization of water; the product of the H_3O^+ ion concentration and the OH^- ion concentration in any aqueous solution
  (i) a process in which one molecule transfers a proton to another molecule of the same substance; water does this to a very small extent

13.2 Match the key terms with the descriptions provided.
  (a) an acid that ionizes completely when dissolved in water
  (b) a model that describes acids as substances that generate H^+ ions in solution and bases as substances that generate OH^- ions in solution
  (c) an aqueous hydrogen ion; $H_3O^+(aq)$
  (d) a solution in which the OH^- concentration is greater than the H_3O^+ concentration; a solution with a pH greater than 7
  (e) an acid that does not completely ionize when dissolved in water
  (f) a substance that forms after an acid loses an H^+ ion
  (g) an equilibrium constant for the ionization of an acid in water; a value that expresses the strength of a weak acid
  (h) a measure of the acidity of aqueous solutions; the negative logarithm of the H_3O^+ concentration
  (i) a combination of a weak acid and its conjugate base (or a weak base and its conjugate acid) in similar amounts; when in solution, something that helps prevent large changes in pH when small amounts of $H_3O^+(aq)$ or $OH^-(aq)$ are added

Math Toolbox Questions

13.3 Use your calculator to find the log of the following numbers.
  (a) 10^{-9}
  (b) 1×10^{-11}
  (c) 7.4×10^{3}
  (d) 10^{5}
  (e) 1

13.4 Use your calculator to find the log of the following numbers.
  (a) 10^{4}
  (b) 1×10^{-6}
  (c) 1.7×10^{8}
  (d) 10^{-8}
  (e) 10

13.5 Use your calculator to find the inverse log of the following numbers.
  (a) 1.2
  (b) −6.2
  (c) 0

13.6 Use your calculator to find the inverse log of the following numbers.
  (a) 12.7
  (b) −9.4
  (c) 1

What Are Acids and Bases?

13.7 What are some properties of acids?

13.8 What are some properties of bases?

13.9 List some common foods or household products that are bases.

13.10 List some common foods or household products that are acids.

13.11 In terms of the Arrhenius concept of acids and bases, how does an acid behave when dissolved in water? How does an Arrhenius base behave?

13.12 How does the Arrhenius concept of acids and bases emphasize that they are electrolytes?

13.13 How is the Brønsted-Lowry theory of acids and bases different from the Arrhenius model?

13.14 Why is $H_3O^+(aq)$ a more accurate representation of an aqueous H^+ ion than $H^+(aq)$?

There are several other features of this textbook that support student learning. End-of-chapter materials include **math toolboxes** (when appropriate), **key concepts summary, key terms list,** and **key relationships.** Each chapter has extensive **end-of-chapter questions and problems** that range in difficulty and conceptual/quantitative emphasis. Most of the questions and problems are sorted by section and are paired, with selected answers appearing in Appendix D. There are also **vocabulary identification questions** at the beginning of the end-of-chapter problems, as well as many questions involving interpretation of molecular-level images. The **concept review questions** provide students with practice at reasoning through multiple-choice questions. To support the text's problem-solving approach, video tutorials in the online homework system (ConnectPlus) provide assistance to students struggling with a particular question. These tutorials guide students through an approach to a similar type of question and emphasize the conceptual foundation and the process that leads to a reasonable answer. (See page xviii for a description of additional digital resources.)

The Art Program

A conceptual understanding of chemistry requires students to visualize molecular-level representations of macroscopic phenomena, as well as to connect macroscopic and molecular-level understandings to symbolic representations. To help students connect verbal descriptions to molecular-level representations, this book has an extensive **art program.** You'll notice many examples of zoomed art, where pictures or other macroscopic images have close-ups that show the particular phenomena at a molecular level.

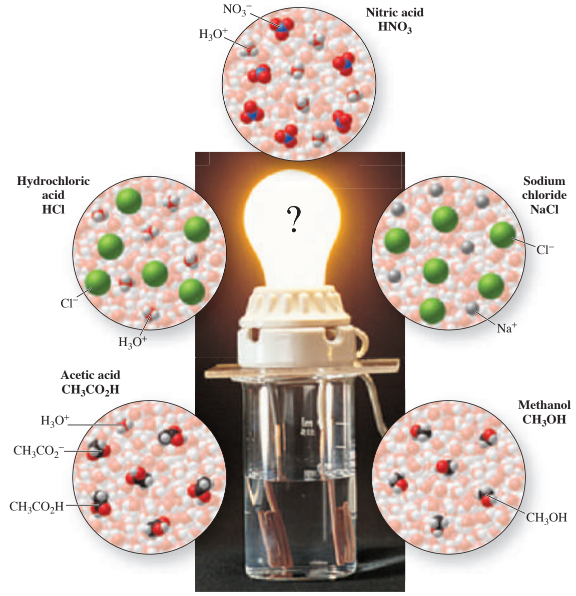


FIGURE 3.6 For each of the solutions, inspect the molecular-level images in which each type of particle is labeled. If any of these substances dissociate into ions, separate particles of single atoms (or atom groups) are visible. Which of the compounds dissociate completely in solution? Which only partially dissociate? Which do not dissociate at all? The acids do not simply dissociate, but form H_3O^+ ions in solution.

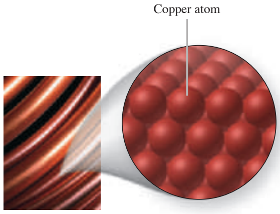


FIGURE 1.8 A copper pipe consists of a regular array of copper atoms.

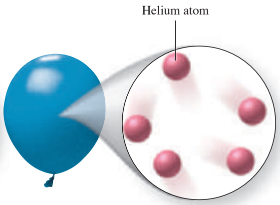


FIGURE 1.9 Helium atoms are present inside the balloon.

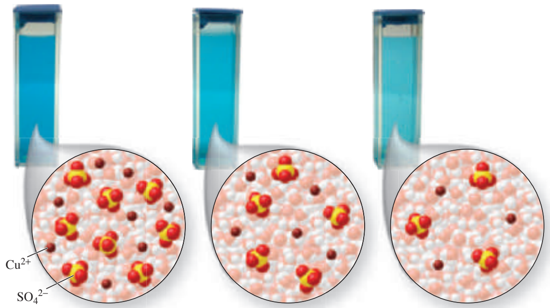


FIGURE 4.18 The color intensity of a colored solution decreases as the concentration of the solute decreases. Compare the concentrations of ions in the solution for these copper(II) sulfate solutions. Which has more copper(II) ions? More sulfate ions? Which solution has the greatest concentration?

Students who enroll in an introductory chemistry course often take an associated lab. Most of the experiments these students conduct involve working with solutions. To enhance this lab experience, a brief **introduction to solution behavior** appears early in the textbook (Chapter 4). This early introduction will allow students to better understand what they experience in the lab, as well as understand the multitude of solutions we encounter on a daily basis.

Detailed List of Changes

Chapter 1 Matter and Energy

- The captions to Figures 1.10 and 1.14 were modified to clarify molecular-level images.
- A margin note was added to the discussion of density to highlight different ways of representing the mass/volume relationship.
- A photo of a modern blimp was added to Figure 1.26.
- A margin note was added to the discussion of metal reactions with acids.
- Equations showing root relationships in Math Toolbox 1.1 were modified for clarity.
- Text and a new worked example were added to discuss and provide additional practice at solving problems involving multistep calculations in Math Toolbox 1.2.
- Several new end-of-chapter problems were added to provide additional practice at solving multistep problems involving significant figures and problems involving complex units
- A new worked example involving complex units was added to Math Toolbox 1.3.

Chapter 2 Atoms, Ions, and the Periodic Table

- New margin notes were added to cover the following: example of iron as an ion in iron(II) sulfate, a nutritional supplement, and the use of neutrons in fission reactions.
- The description of the relative sizes and masses of the atom and nucleus was rewritten to improve clarity.
- Figure 2.4 STM image was changed from gold to copper.
- Figure 2.20 was enlarged to improve visibility of element photos.

Chapter 3 Chemical Compounds

- Figures 3.11 and 3.18 were modified to improve clarity.
- Photos in Figure 3.16 were replaced for better agreement with the caption.
- Text and the net ionic equation were added to clarify discussion of acid neutralization by a base.
- Figure 3.36 was modified to include a compound containing ammonium ion.

Chapter 4 Chemical Composition

- The solution to Example 4.4 was expanded to help students connect atomic-level structure to mole-level quantities.
- Slight changes were made to the solutions to Examples 4.6 and 4.23 to improve clarity.
- Images of ions were added to Figure 4.21A to improve agreement with the text.
- The solution to Example 4.21 was modified for significant figures.
- Several new end-of-chapter problems were added to provide students with additional practice at solving complex empirical formula problems.

Chapter 5 Chemical Reactions and Equations

- Figures 5.9 and 5.10 were updated.
- The solution to Example 5.4 was modified to clarify the atom- and ion-counting procedure.
- A margin photo of cobalt glass was added to Example 5.6 to help students connect the problem to a real-life object.
- The caption to Figure 5.27 was modified to clarify discussion of double-displacement reactions.
- Several end-of-chapter questions were changed to improve relevance.
- A new end-of-chapter, multipart problem was added to provide students with additional practice at balancing equations.

Chapter 6 Quantities in Chemical Reactions

- Discussion of alternative fuels in the introduction was updated.
- Passages in the caption to Figure 6.2, Practice Problem 6.2, and Section 6.4 were rewritten to describe the space shuttles in a historical context instead of a current one.
- Example 6.3 was changed to describe a single-displacement reaction between copper metal and silver nitrate.
- The description of the reaction in Example 6.6 was rewritten to improve relevancy.
- Pedagogical changes were made in the solutions to Examples 6.6 and 6.7.
- A margin note about the efficiency of lightbulbs was rewritten to take into account the increasing use of fluorescent and LED lights.
- End-of-chapter question 6.107 was changed to provide students with practice at solving stoichiometry problems involving dissolving ionic compounds.

Chapter 7 Electron Structure of the Atom

- Discussion of lightbulbs and lasers in the introduction was updated.
- Example and Practice Problem 7.2 were divided into separate parts to improve pedagogy.
- Margin notes were added to weave inquiry into the discussions of the hydrogen line spectrum and the relative sizes of ions in an ionic solid.

Chapter 8 Chemical Bonding

- Figures 8.4 and 8.14 were modified for better agreement of images with caption descriptions.
- A new margin note was added and Figure 8.6 was modified to highlight that there are no electronegativity values known for most of the noble gases.
- A margin note was added to the discussion of Lewis symbols for ions to direct students to previously discussed ions in Chapter 3.
- A margin note about writing Lewis formulas was changed to clarify that valence electrons are shown in these structures.
- The explanation for drawing Lewis formulas was expanded.
- The subsection discussing bonding in carbon compounds was moved to its own section.
- The margin note describing isomers of large alkanes was moved to the body of the text.
- The discussion of multiple functional groups in the molecule glycine was expanded.

Chapter 9 The Gaseous State

- Temperatures expressed in both Celsius and Kelvin scales were provided throughout the text of the chapter.
- The solution to Example 9.12 was modified for clarity.
- An equation was inserted in Math Toolbox 9.1 to clarify proportional relationships.

Chapter 10 The Liquid and Solid States

- The caption to Figure 10.2 was expanded to clarify the composition of medical implants.
- Figure 10.14 was enlarged to clarify molecular-level images.
- A margin note with a graph was added and the text was modified to highlight energy associated with phase changes.
- Example 10.7 was modified to address a common misconception about hydrogen bonding.

Chapter 11 Solutions

- The chapter opening photo and Figures 11.3, 11.4, 11.5, and 11.22 were updated.
- The description and definition of entropy were rewritten to be consistent with current understanding of this concept.
- Figure 11.29 was changed to help students better visualize what happens in a reverse osmosis coil.
- A margin note on supersaturated solutions was inserted to complement the description of how these solutions are prepared.

Chapter 12 Reaction Rates and Chemical Equilibrium

- The margin note that describes reaction rate was expanded and moved from the introduction to Section 12.1.
- Margin notes were added to emphasize the constant value of the activation energy for an uncatalyzed reaction, even when temperature changes, and to emphasize the constant value of the equilibrium constant except when temperature changes.
- Margin notes were added to weave inquiry into the discussion of equilibrium and to show how catalysts are useful in industry.
- The photo in Figure 12.11 was changed to a schematic diagram of a catalytic converter to improve clarity.
- The explanation of how a chemical reaction approaches equilibrium was expanded.

Chapter 13 Acids and Bases

- The margin note about the basic properties of lime (CaO) was replaced by a margin note with photo about another basic substance, a drain cleaner. Information about the use of CaO to adjust soil acidity was added to the text.
- Two margin notes were added to weave inquiry into the discussion about the importance of an unshared electron pair on Brønsted-Lowry bases.
- Margin notes were added to emphasize the importance of acid-base properties in the effectiveness of cleaners, and to relate the sour taste of lemons to citric acid.
- The margin notes about variation of K_w with temperature were deleted as this topic is beyond the scope of this book.
- A margin note was added to clarify that calculating the concentrations of H_3O^+ and OH^- in weak acid and weak base solutions is beyond the scope of this book.
- The description of acidic hydrogen atoms was rewritten for clarity.
- Occurrences of the term *Lewis structure* were replaced by *Lewis formula* for consistency.

Chapter 14 Oxidation-Reduction Reactions

- The chapter opening photo and Figure 14.4 were updated.
- Figures in Example 14.1 were replaced with photos that clearly show formation of elemental silver.
- Discussion of mercury batteries was updated throughout the chapter.
- The discussion of the breathalyzer reaction was updated.

Chapter 15 Nuclear Chemistry

- The chapter opening photo and Figures 15.1, 15.3, 15.11, and 15.20 were updated.
- The synthesis of livermorium was added to the description of nuclear bombardment reactions.
- A margin note was added to emphasize the relationship between half-lives and radioactive nuclide quantities.
- Equations showing neutron bombardment of uranium-235 were modified for clarity.
- Discussion of the Fukushima Daiichi Nuclear Power Plant was added to the description of fission reactors.

Chapter 16 Organic Chemistry

- The explanation of use of R– to abbreviate a hydrocarbon part of a molecule was moved from Section 16.1 to the subsection on ethers in Section 16.4 where it is more relevant.
- The description of organic molecules was modified slightly to specify that the elements oxygen and nitrogen are often present.
- The description of isomers of hydrocarbons was expanded for clarity.
- The use of line structures to represent organic molecules was increased.
- Margin notes were added to show the common line structure for benzene and the general formula for cycloalkanes.
- Inquiry was added to the caption for Figure 16.3.
- The explanation for naming branched hydrocarbons in Section 16.3 was expanded.
- Example 16.6 was changed to a question about the reaction of an alkane with bromine in sunlight.
- The explanation of the differences between cis and trans isomers was rewritten for clarity.
- The explanation for naming benzene derivatives was modified to emphasize the regular use of the names toluene, aniline, and phenol accepted in the IUPAC nomenclature system. Similar modifications were also made in Table 16.6 and Example and Practice Problem 16.10.
- The description of carbohydrates in the section on alcohols was modified to include a statement about water solubility and structure.
- The description of acetic acid as a type of carboxylic acid was moved earlier in the subsection to improve content flow.

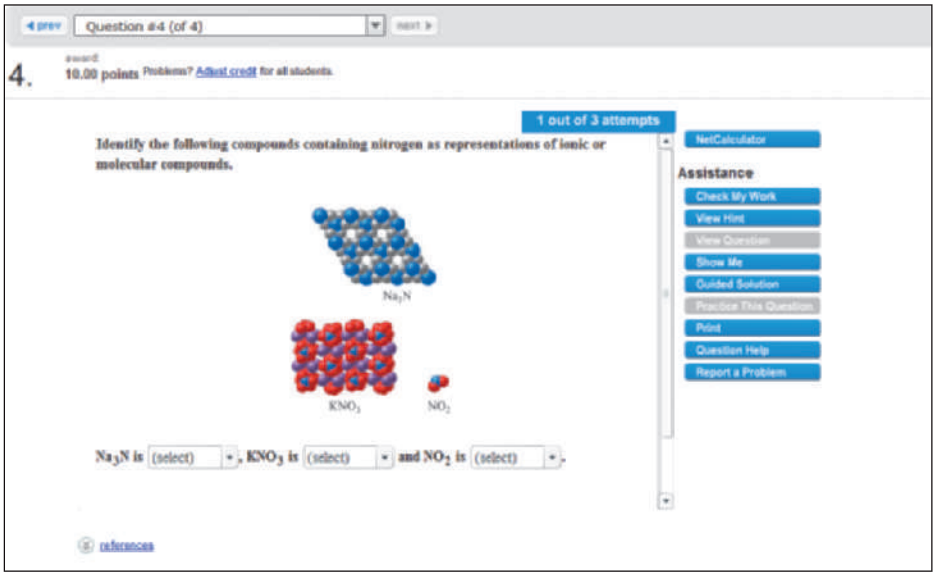
Chapter 17 Biochemistry

- An inquiry question was inserted in the nucleic acids section to assist students in distinguishing a ribonucleotide from a deoxyribonucleotide.
- Figure in 17.9 was modified to simplify the complex structure.
- Figure 17.35 was changed to be consistent with the representation shown in Figure 17.33.
- The caption to Figure 17.36 was modified to address multiple cis double bonds shown in the structure.

Digital Resources

McGraw-Hill ConnectPlus® Chemistry

McGraw-Hill ConnectPlus® Chemistry provides online presentation, assignment, and assessment solutions. It connects your students with the tools and resources they'll need to achieve success. With ConnectPlus Chemistry, you can deliver assignments, quizzes, and tests online. A robust set of questions, problems, and interactives is presented and aligned with the textbook's learning goals. The integration of **ChemDraw by PerkinElmer,** the industry standard in chemical drawing software, allows students to create accurate chemical structures in their online homework assignments. As an instructor, you can edit existing questions and author entirely new problems. Track individual student performance—by question, assignment, or in relation to the class overall—with detailed grade reports. Integrate grade reports easily with Learning Management Systems (LMS), such as WebCT and Blackboard—and much more. **ConnectPlus Chemistry** also provides 24/7 online access to an eBook. This media-rich version of the book is available through the McGraw-Hill Connect platform and allows seamless integration of text, media, and assessments. To learn more, visit **connect.mheducation.com**



Identify the following compounds containing nitrogen as representations of ionic or molecular compounds.

Na_3N is (select), KNO_3 is (select) and NO_2 is (select).

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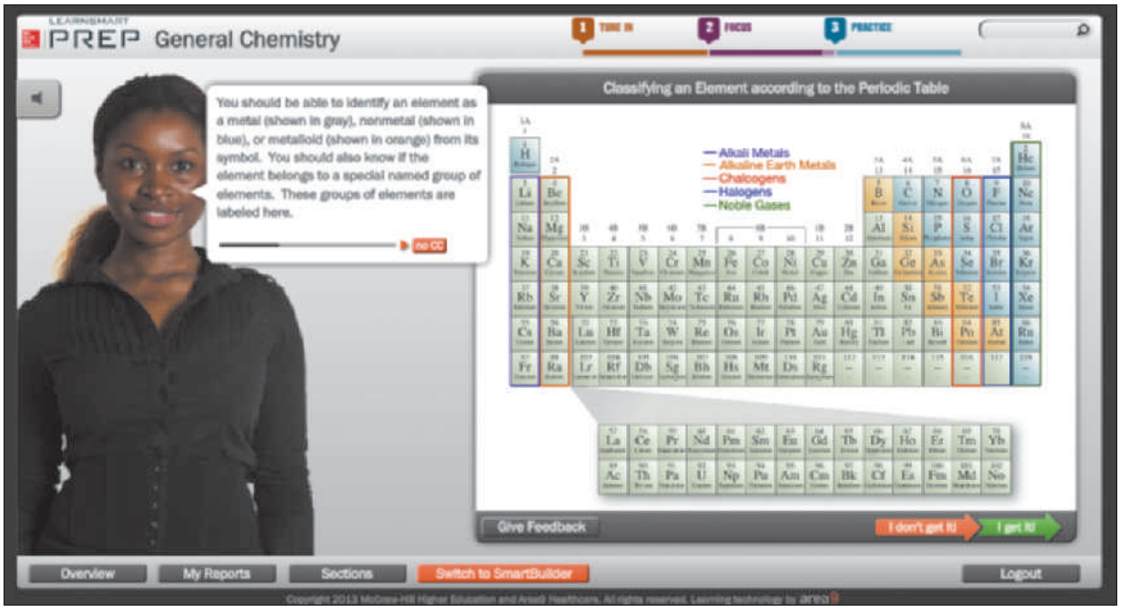

You should be able to identify an element as a metal (shown in gray), nonmetal (shown in blue), or metalloid (shown in orange) from its symbol. You should also know if the element belongs to a special named group of elements. These groups of elements are labeled here.

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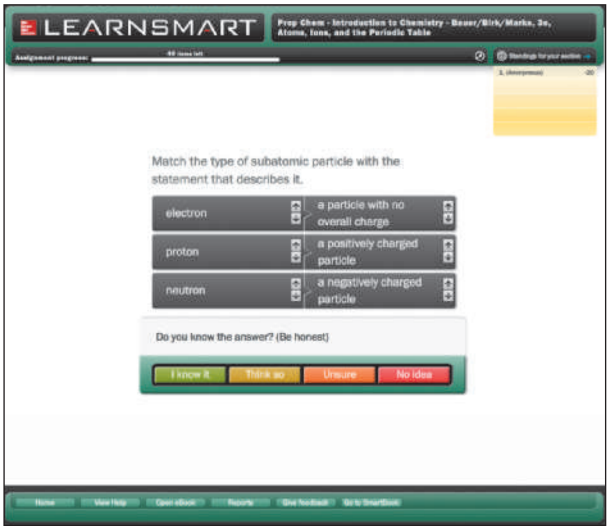

Match the type of subatomic particle with the statement that describes it.

electron	a particle with no overall charge
proton	a positively charged particle
neutron	a negatively charged particle

Do you know the answer? (Be honest)

McGraw-Hill SmartBook™

Powered by the intelligent and adaptive LearnSmart engine, SmartBook is the first and only continuously adaptive reading experience available today. Distinguishing what students know from what they don't, and honing in on concepts they are most likely to forget, SmartBook personalizes content for each student. Reading is no longer a passive and linear experience but an engaging and dynamic one, where students are more likely to master and retain important concepts, coming to class better prepared.

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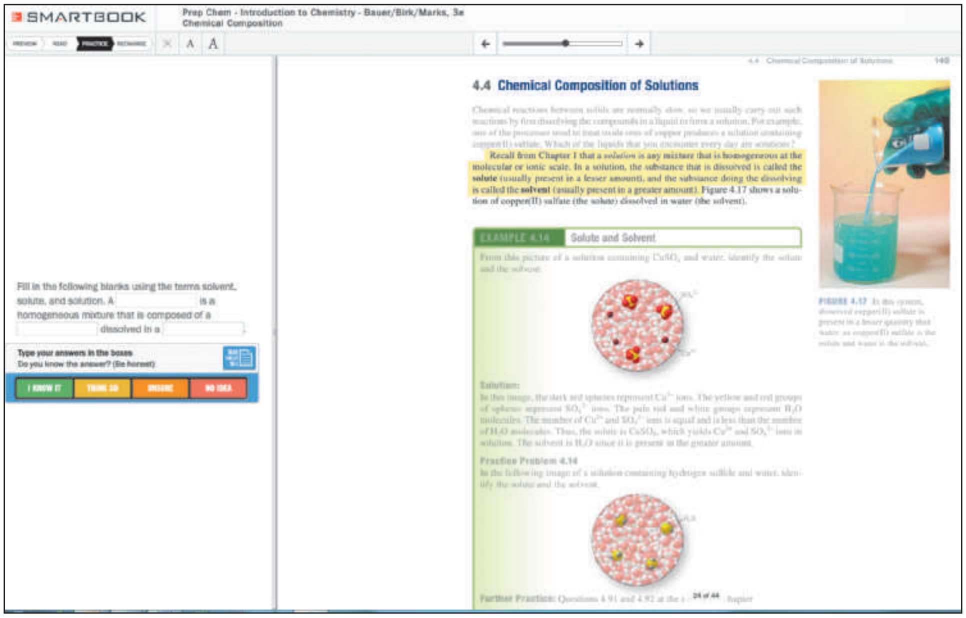

Prep Chem - Introduction to Chemistry - Bauer/Birk/Marks, 3e
Chemical Composition

4.4 Chemical Composition of Solutions

Recall from Chapter 1 that a *solution* is any mixture that is homogeneous at the molecular or ionic scale. In a solution, the substance that is dissolved is called the **solute** (usually present in a lesser amount), and the substance doing the dissolving is called the **solvent** (usually present in a greater amount). Figure 4.17 shows a solution of copper(II) sulfate (the solute) dissolved in water (the solvent).

EXAMPLE 4.14 Solute and Solvent

Practice Problem 4.14

Fill in the following blanks using the terms solvent, solute, and solution. A _____ is a homogeneous mixture that is composed of a _____ dissolved in a _____.

Type your answers in the boxes
Do you know the answer? (Be honest)

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Supplements for the Student

Student Solutions Manual. This separate manual contains detailed solutions and explanations for all odd-numbered problems and all concept review questions in the text.

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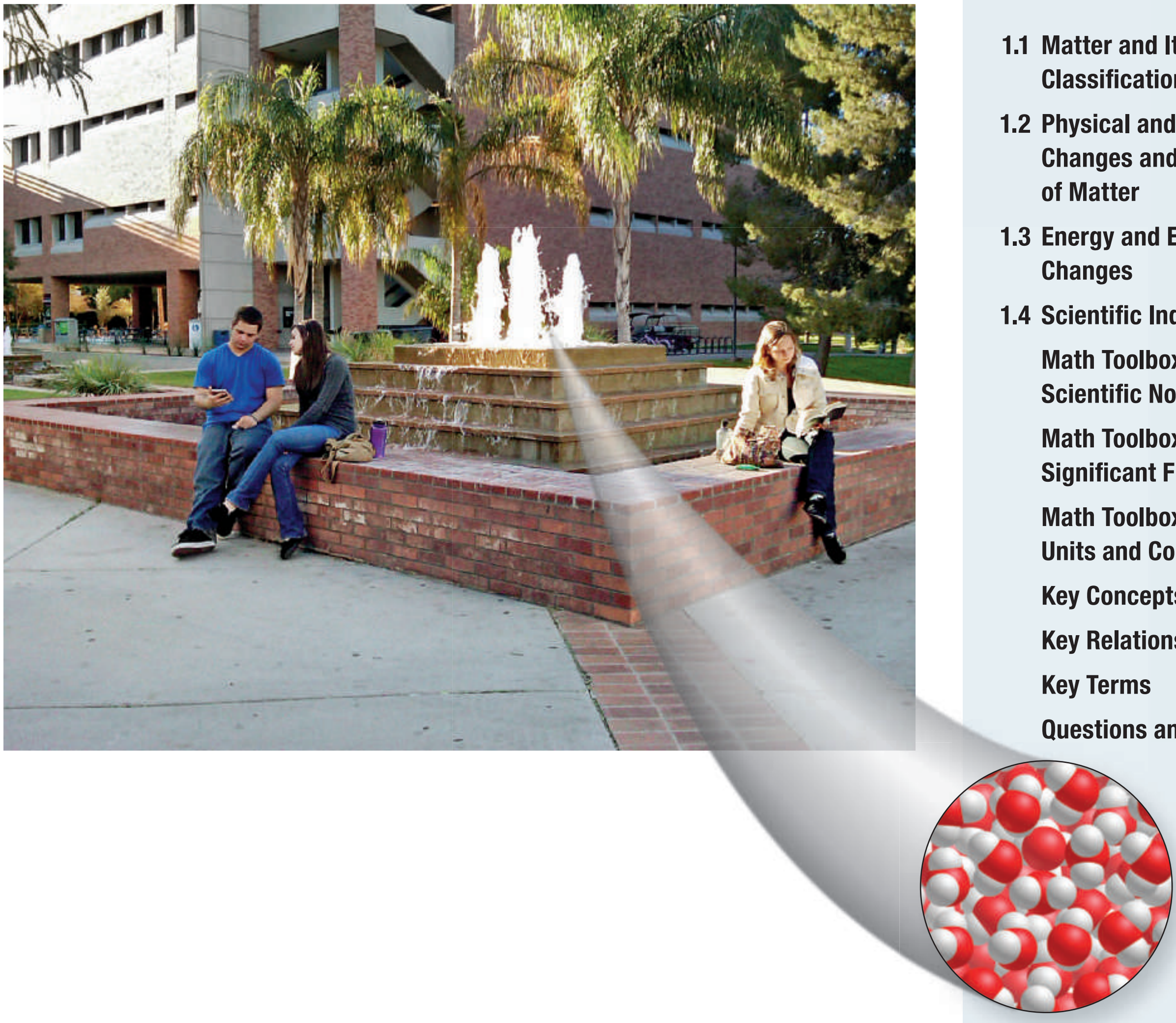
CHAPTER

1

Matter and Energy



Anna and Bill are college students enrolled in an introductory chemistry course. For their first assignment, the professor has asked them to walk around campus, locate objects that have something to do with chemistry, and classify the things they find according to characteristics of structure and form.

Anna and Bill begin their trek at the bookstore. They spot a fountain, a large metallic sculpture, a building construction site, and festive balloons decorating the front of the store. They notice water splashing in the fountain and coins that have collected at the bottom. The metallic sculpture has a unique color and texture. At the building construction site they notice murals painted on the wooden safety barricade. Through a hole in the fence, they see a construction worker doing some welding.

Bill and Anna make a list of the things that attracted their attention and start trying to classify them. Inspecting the fountain, they notice that it appears to be composed of pebbles embedded in cement. As water circulates in the fountain, it travels in waves on the water's surface. The coins in the fountain, mostly pennies, vary in their shininess. Some look new, with their copper color gleaming in the bright sunshine. Others look dingy, brown, and old. The metal sculpture has a unique, modern design, but it's showing signs of age. A layer of rust covers its entire surface. Anna and Bill decide to classify the sculpture as a metal, like the coins in the fountain. They also conclude that the water, pebbles, and concrete in the fountain are not metals.

As they approach the construction site, Anna and Bill examine the painted mural on a safety fence that surrounds the site. Through the peephole in the mural, they see gravel, cinder blocks, metallic tubes for ductwork, steel beams, and copper pipe. They add more nonmetals and metals to their list. A welder is joining two pieces of metal. Sparks are flying everywhere. Anna and Bill wonder what is in the sparks. Since the sparks are so small and vanish so rapidly, they don't know how to classify them.

Electric carts are an ideal way to get across a large campus. They run using an electric battery instead of fuel.

As they continue their walk, they pass the intramural fields and the gym where they see students using tennis rackets, baseball bats, bicycles, and weight belts. They wonder how they will classify these items. For lunch, Bill and Anna buy pizza. They sip soft drinks from aluminum cans. They settle on a bench to enjoy their lunch in the sunshine and watch students playing volleyball in a sandpit. As they put on their sunscreen, they wonder how they might classify sunlight. After lunch, they hurry off to an afternoon class. On the way, they notice a variety of vehicles on campus. Some are gasoline-powered cars and buses, but others have signs on them saying they operate on alternative fuels. Trucks lumber by, exhaust fumes spewing from their tailpipes. Bill and Anna feel the hoods of parked cars. Some are still warm from their engine's heat.

How are Bill's and Anna's observations related to chemistry? What characteristics have they identified that they can use for classification purposes? They have started their classification with metals and nonmetals. What other categories should they devise?

Now it's your turn. Make a list of things relevant to chemistry in the location where you are reading this. How will you classify the things on your list? What characteristics will you use to organize the items into categories? Most important, why bother to classify things at all?

In this chapter we will explore some answers to these questions. As you learn what chemistry is, you'll begin to develop explanations for how substances look, change, and behave.

Questions for Consideration

1.1 What characteristics distinguish different types of matter?
1.2 What are some properties of matter?
1.3 What is energy and how does it differ from matter?
1.4 What approaches do scientists use to answer these and other questions?

Math Tools Used in This Chapter

Scientific Notation (Math Toolbox 1.1)
Significant Figures (Math Toolbox 1.2)
Units and Conversions (Math Toolbox 1.3)

This icon refers to a Math Toolbox that provides more detail and practice.

1.1 Matter and Its Classification

All the *things* that Anna and Bill observed on campus are examples of matter. The fountain, the metal sculpture, the construction site, the balloons outside the bookstore, the exhaust fumes from buses, the pizza they had for lunch, even Bill and Anna themselves—all are matter. **Matter** is anything that occupies space and has mass. **Mass** is a measure of the quantity of matter. The interaction of mass with gravity creates weight, which can be measured on a scale or balance.

Some of Bill's and Anna's observations, however, were not of matter. Sunlight, the light from welding, and the heat of automobile engines are not matter. They do not occupy space, and they have no mass. They are forms of energy. *Energy* is the capacity to move an object or to transfer heat. We'll discuss energy in Section 1.3, but for now, let's focus on matter.

All of Anna's and Bill's observations are relevant to chemistry, because chemistry is the study of matter and energy. Since the entire physical world is matter and energy, chemistry would be an overwhelming subject of study if we did not classify phenomena in manageable ways. Anna and Bill used characteristics like shininess and hardness when they decided some materials were metals and others were not. Let's explore some other characteristics that can be used to classify matter.

Composition of Matter

One way to classify matter is by its chemical composition. Some types of matter always have the same chemical composition, no matter what their origin. Such matter is called a **pure substance** or more briefly, a *substance.* A pure substance has the same composition throughout and from sample to sample. It cannot be separated into components by physical means.

Are there any things where you are now that might be pure substances? Actually, pure substances are rare in our world. Most things are mixtures of some kind. Pure substances are found most often in laboratories where they are used to determine the properties and behavior of matter under controlled conditions.

Some pure substances can be observed. For example, the aluminum in Anna's soda can is pure. It is not combined with any other substances, although it is coated with plastic and paint. Consider also the sandpit where Bill and Anna watched the volleyball game. The sand is not a pure substance, but if we removed all the dirt, minerals, and other contaminants, it would be the pure substance, silica, which is one kind of sand (Figure 1.1). Grains of silica differ in size, but they all have the same chemical composition, which can be determined in the laboratory.

In contrast to pure substances, other materials are mixtures. A **mixture** consists of two or more pure substances and may vary in composition. The fountain, for example, is made from a mixture of gravel, concrete, and pebbles. Even the water in the fountain is not a pure substance since small amounts of gases and minerals are dissolved in it. Like sand, however, it could be made pure if all the other substances were removed.

There are two types of pure substances: elements and compounds. We will discuss these first, and then we'll describe types of mixtures.

FIGURE 1.1 Sand is composed of a mineral, silica. It contains the elements silicon and oxygen in specific proportions.

Elements All matter consists of pure substances or mixtures of substances. Pure substances, in turn, are of two types: elements and compounds. An **element** is a substance that cannot be broken down into simpler substances *even by a chemical reaction.* For example, suppose we first purified the water in a fountain to remove contaminants. Then we used a chemical process called *electrolysis* to separate it into its component elements. Water can be broken down by chemical means into

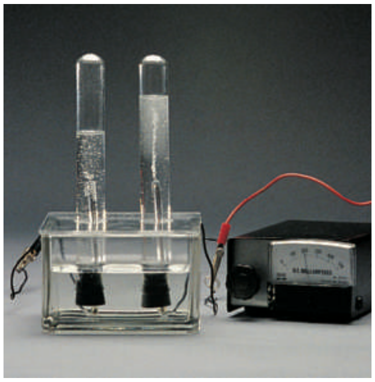
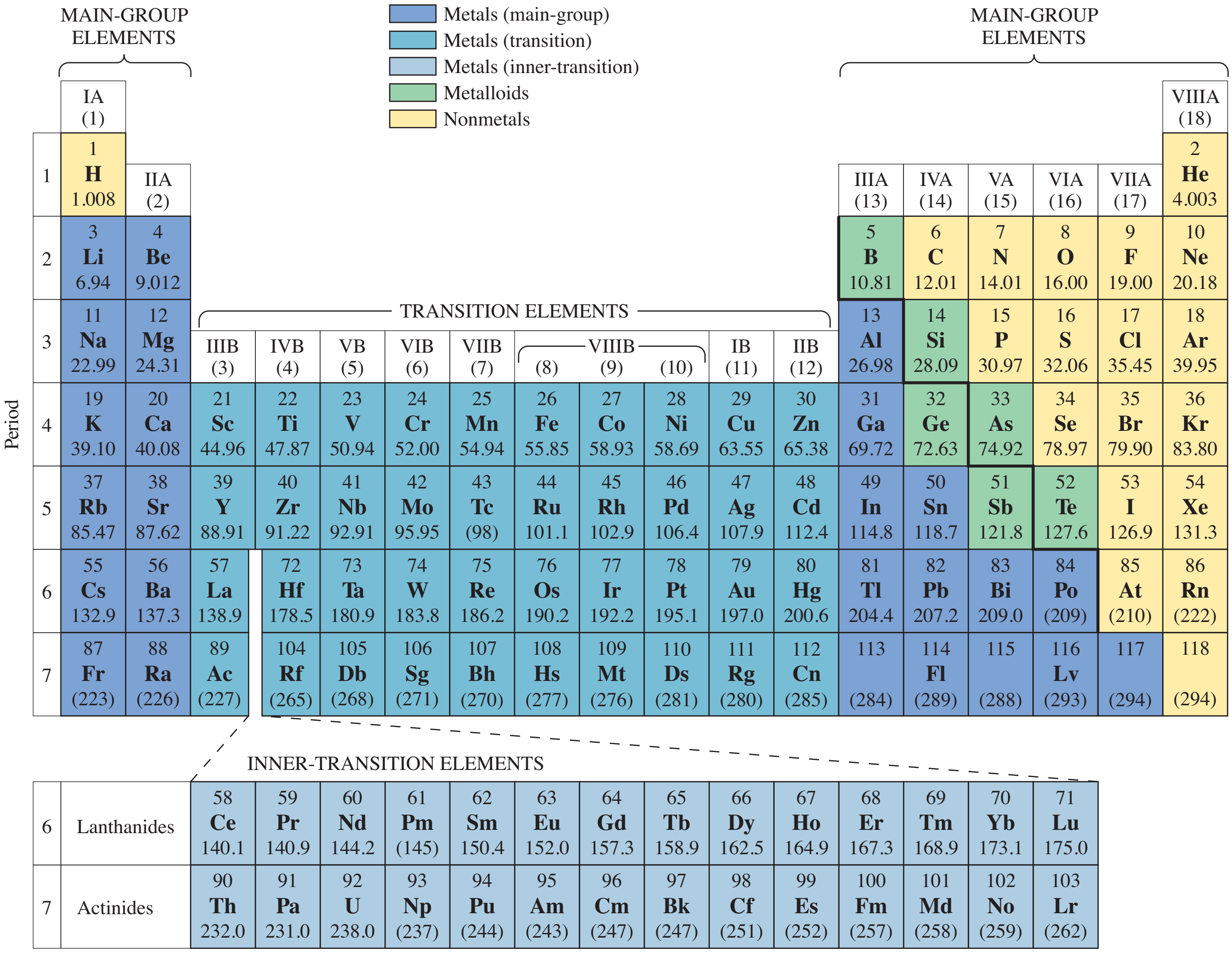


FIGURE 1.2 When electric current is passed through water, the water decomposes into the elements hydrogen and oxygen. The hydrogen (left) and oxygen (right) can be seen bubbling to the tops of the tubes.

hydrogen and oxygen, as shown in Figure 1.2, so water is not an element. The hydrogen and oxygen, however, are elements. We cannot break them down into any simpler substances using heat, light, electricity, or any chemical process. We can convert them into more complex substances, but not into simpler ones.

Elements are the building blocks of all matter. Of the 114 elements that have been given names, 83 can be found in natural substances and in sufficient quantity to isolate. The many examples of matter that we use, see, and read about are all built up of different elements in different combinations. The elements that are not isolated from natural sources on Earth have been synthesized by scientists. Some are so unstable that they have only a fleeting existence, including those that have not yet been formally named. To classify elements, chemists use a *periodic table,* like that shown in Figure 1.3. The elements in each column, called *groups* or *families* of elements in the periodic table, share similar characteristics, or *properties.*

Elements are generally classified into two main categories: metals and nonmetals. Generally, a **metal** can be distinguished from a **nonmetal** by its luster (shininess) and ability to conduct electricity (electrical conductivity). Copper, aluminum, iron, and other metals are good conductors of electricity. Nonmetal elements, such as carbon (in the form of diamond), chlorine, and sulfur, normally are not. Note the difference in appearance of the metals and nonmetals shown in Figure 1.4. Not all elements fit neatly into such categories. In Chapter 2 we'll discuss elements that have properties somewhere between metals and nonmetals.



MAIN-GROUP ELEMENTS

Metals (main-group)
Metals (transition)
Metals (inner-transition)
Metalloids
Nonmetals

MAIN-GROUP ELEMENTS

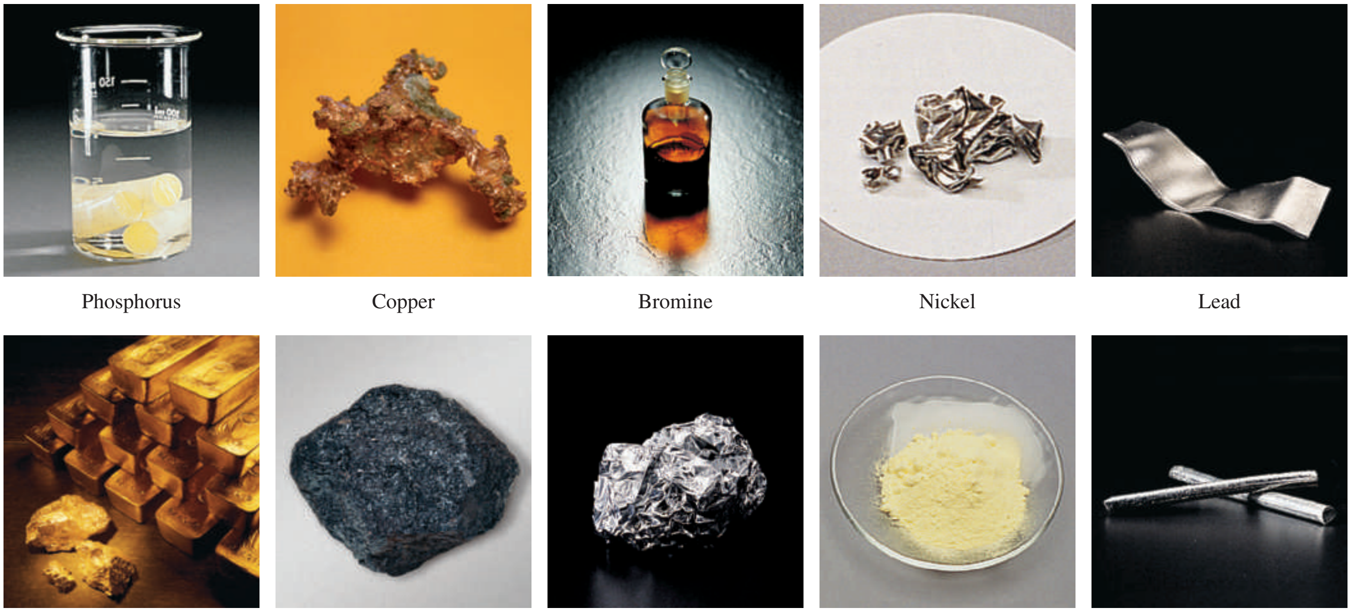
TRANSITION ELEMENTS

Period	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIIIB (8)	(9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (18)
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.94	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
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	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** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.97	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.95	43 **Tc** (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
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)
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 (284)	114 **Fl** (289)	115 (288)	116 **Lv** (293)	117 (294)	118 (294)

INNER-TRANSITION ELEMENTS

6	Lanthanides	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** 173.1	71 **Lu** 175.0
7	Actinides	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)

FIGURE 1.3 The periodic table organizes the known elements according to their properties. The letters are symbols for the names of the elements.



Phosphorus Copper Bromine Nickel Lead

Gold Carbon Aluminum Sulfur Tin

FIGURE 1.4 Some elements. Which of these are metals?

EXAMPLE 1.1 Metals and Nonmetals

Which of the elements pictured are metals? Why do you think so?

Iron

Carbon

Magnesium

Sulfur

Aluminum

Solution:

Notice that three of the elements—iron, aluminum, and magnesium—have a luster; that is, they shine. They are metals. If you could handle and test the substances, you could use other properties, such as electrical conductivity, to distinguish between metals and nonmetals.

Consider This 1.1

What if you were given element properties and had to determine the identity of the element from a list of possibilities? Suppose an element is rather dull in appearance, a poor conductor of electricity, and a gas at room temperature. Would this element be zinc, platinum, or chlorine?

Practice Problem 1.1

Identify the nonmetals in Figure 1.4. Explain the characteristics you considered in making your decision.

Further Practice: Questions 1.31 and 1.32 at the end of the chapter

TABLE 1.1 Symbols of Selected Elements

English Name	Original Name	Symbol	English Name	Original Name	Symbol
copper	cuprum	Cu	potassium	kalium	K
gold	aurum	Au	silver	argentum	Ag
iron	ferrum	Fe	sodium	natrium	Na
lead	plumbum	Pb	tin	stannum	Sn
mercury	hydrargyrum	Hg	tungsten	wolfram	W

To become familiar with the periodic table, you should learn the names and symbols for the first 36 elements, as well as the symbols for silver, tin, gold, mercury, and lead. Your instructor may ask you to learn others.

To avoid having to write out the name of an element every time we refer to it, we use a system of symbols. An **element symbol** is a shorthand version of an element's longer name. Often, the symbol is one or two letters of the element's name (C for carbon, He for helium, Li for lithium). The first letter is uppercase, and the second letter, if present, is lowercase. When the names of two elements start with the same two first letters (magnesium and manganese, for example), the symbol uses the first letter and a later letter to distinguish them (Mg for magnesium, Mn for manganese).

For a few elements, the symbols are based on their Latin names or on names from other languages. These are listed in Table 1.1. Some recently synthesized elements have been named for famous scientists or locations. Others have not been given permanent names. You'll find a list of the modern names and symbols on the inside front cover of this book.

EXAMPLE 1.2 Element Symbols

Potassium is a soft, silver-colored metal that reacts vigorously with water. Write the symbol for the element potassium.

Solution:

The symbol for potassium is K. In the periodic table, potassium is element 19 in group (column) IA (1) of the periodic table.

Consider This 1.2

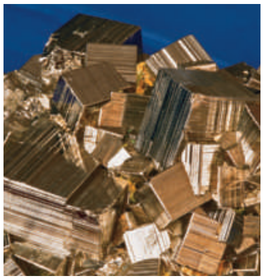
What if you instinctively identified the element symbol as P or Po? Why are these symbols incorrect for potassium?

Practice Problem 1.2

(a) Lead is a soft, dull, silver-colored metal. Write the symbol for the element lead.

(b) The symbol for a common element used to make jewelry is Ag. What is the name of this element?

Further Practice: Questions 1.39 and 1.40 at the end of the chapter



Iron pyrite

FIGURE 1.5 Iron pyrite is composed of the elements iron and sulfur. Iron is magnetic and can be separated from sulfur when the two exist as elements mixed together. Iron pyrite, a compound of iron and sulfur, is not magnetic.

Compounds A **compound,** sometimes called a *chemical compound,* is a substance composed of two or more elements combined in definite proportions. A compound has properties different from those of its component elements. For example, iron pyrite can be broken down into its component elements, iron and sulfur, but its characteristics are different from both (Figure 1.5). Anna and Bill saw many compounds that can be chemically separated into their component elements. Sand is a compound of silicon and oxygen. Water, as discussed earlier, is composed of hydrogen and oxygen. The cheese on their pizza contains many complex compounds, but each of the compounds contains carbon, hydrogen, oxygen, nitrogen, and a few other elements.

Chemists represent compounds with formulas based on the symbols for the elements that are combined in the compound. (Chemical formulas are not the same as the mathematical formulas that may be familiar to you, such as $A = \pi r^2$ for the area of a circle.) A **chemical formula** describes the composition of a compound, using the symbols for the elements that make up the compound. Subscript numbers show the relative proportions of the elements in the compound. If no subscript number is given for an element in a formula, then you may assume that the element has a relative proportion of one. For example, water is known to consist of one unit of oxygen and two units of hydrogen. This compound is represented by the formula H_2O. Sodium chloride, the chemical compound commonly called table salt, contains equal portions of the elements sodium and chlorine. Its formula is therefore NaCl. We will discuss formulas in detail in Chapter 3.

Graphite leaves a mark similar to that made by dragging a rod of lead along a surface, so it was called lead. A hardness number indicates the relative amounts of graphite and clay in a pencil lead. A number 2 pencil is fairly soft, while a number 6 pencil is quite hard. Which has more graphite?

Mixtures Some forms of matter, such as pencil lead, do not have the same composition in every sample. (Pencil lead isn't the element lead. It is a *mixture* of graphite and clay.) A mixture consists of two or more elements or compounds. It is possible to separate mixtures into their component pure substances. The separation can be done physically, using procedures such as grinding, dissolving, or filtering. Chemical processes are not needed to separate mixtures.

We can illustrate the difference between pure substances and mixtures by looking at salt water. Water that has been purified is a pure substance that is composed of hydrogen and oxygen, always in the same proportions. Salt water, on the other hand, is water mixed with salt and many other substances in varying proportions. For example, the Great Salt Lake in Utah is approximately 10% salt, while the Dead Sea is about 30% salt. In either case, we can readily separate salt from water by evaporating the water (Figure 1.6).

FIGURE 1.6 To collect salt, water is diverted into large ponds. The water evaporates, leaving solid salt behind.

Mixtures differ in uniformity of composition. A **homogeneous mixture** has a uniform composition throughout and is often called a **solution.** Most solutions that we commonly encounter are composed of compounds dissolved in water. They are often clear. For example, a well-mixed sample of salt water prepared in a kitchen is uniform in appearance. The salt dissolved in it is invisible. Furthermore, any microscopically small portion of the sample would have the same composition as any other. The particles in the mixture might not be arranged in exactly the same pattern, but each sample, regardless of size, would have the same components in the same proportions.

A mixture that is not uniform throughout—a mixture of salt and pepper, for instance—is a **heterogeneous mixture.** Different samples have their components present in different proportions. Which of the things that Bill and Anna had for lunch is a homogeneous mixture? Which is heterogeneous? How about your own lunch? How can you tell?

Not all solutions are liquids. For example, consider air that has been filtered to remove suspended solid particles. Filtered air is a gaseous solution containing a mixture of primarily oxygen and nitrogen. Solid solutions also exist and are called alloys. For example, brass is a solution of zinc and copper.

We have considered a number of classes and subclasses of matter: mixtures, homogeneous mixtures, heterogeneous mixtures, pure substances, compounds, elements, metals, and nonmetals. A method for classifying matter into these categories is outlined in Figure 1.7. Note in the figure that yes or no answers to several questions distinguish one type of matter from another. First, we ask if the material can be separated physically. If so, then it is a mixture. If not, it must be a pure substance. If this substance can be decomposed (broken down into simpler substances) by chemical reactions, it is a compound. If it cannot, it is an element.

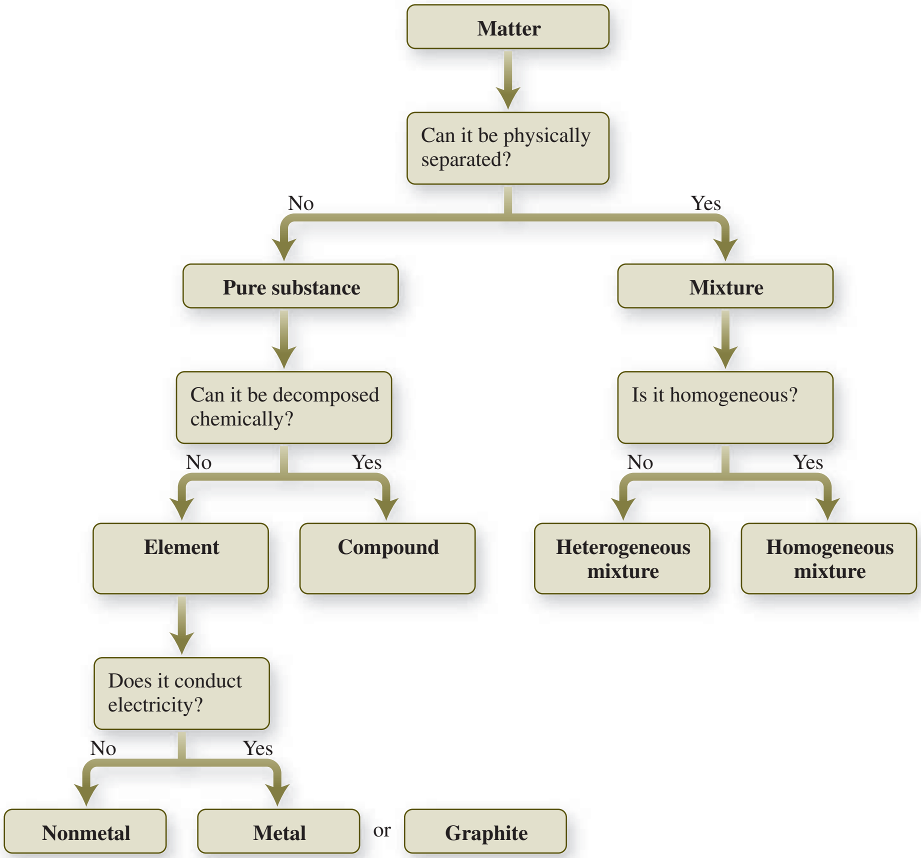


FIGURE 1.7 We can classify matter by answering the short series of questions in this flowchart.

EXAMPLE 1.3 Elements, Compounds, and Mixtures

Which of the following pictures represent pure substances?

Solution:

The copper on the outside of the coin and the helium inside the balloons are pure substances. (However, the helium and balloons considered together provide an example of a mixture.)

Consider This 1.3

Why isn't the water in the fountain considered a pure substance?

Practice Problem 1.3

Which of the pictures represent mixtures? Which are heterogeneous? Which are homogeneous?

Further Practice: Questions 1.45 and 1.46 at the end of the chapter

Representations of Matter

Chemists and other scientists view the world on several different levels. So far we have considered matter on a macroscopic scale. That is, we've discussed matter and phenomena we can see with our eyes. But simple observation is limited. Sometimes we cannot classify things merely by looking at them as Anna and Bill did. What do we do then? Chemists try to make sense of the structure of matter and its behavior on a scale that is much, much smaller than what we can see with our eyes.

Consider the copper pipe at the construction site, for example. If we could enlarge the tiniest unit that makes up the pipe, what would we see? Experimental evidence tells us copper is made up of discrete, spherical entities that all appear to be identical (Figure 1.8). Chemists identify these entities as atoms. An **atom** is the smallest unit of an element that has the chemical properties of that element. For example, we can imagine the helium inside a balloon as many, many atoms of helium, which we represent symbolically as He. In Figure 1.9, each sphere represents a single helium atom. Similarly, if we could magnify the structure of water, we would find two small hydrogen atoms bound separately to a single larger oxygen atom. Such a combination of elemental units is a **molecule.** Molecules are made up of two or more atoms bound together in a discrete arrangement. Several molecules of water, H_2O, are shown in Figure 1.10, where the central red sphere represents an oxygen atom and the two smaller, white spheres stand for hydrogen atoms. (Some compounds do not exist as molecules. We will discuss them in Chapter 3.)

Although chemists generally use color coding to distinguish between atoms of different elements in representations, the atoms themselves do not have colors. Macroscopic samples of matter may have color, but these colors do not usually match those used to represent atoms. In accurate representations, the sizes of the spheres change to reflect the relative differences in the sizes of atoms of different elements.

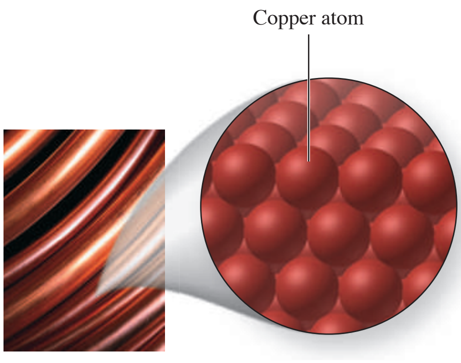


FIGURE 1.8 A copper pipe consists of a regular array of copper atoms.

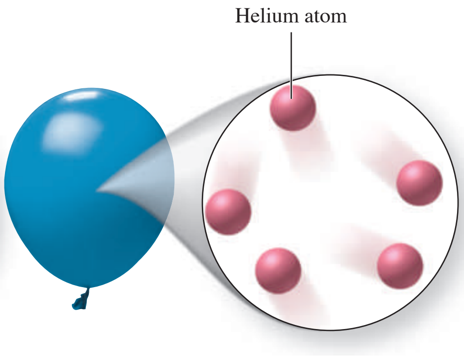


FIGURE 1.9 Helium atoms are present inside the balloon.