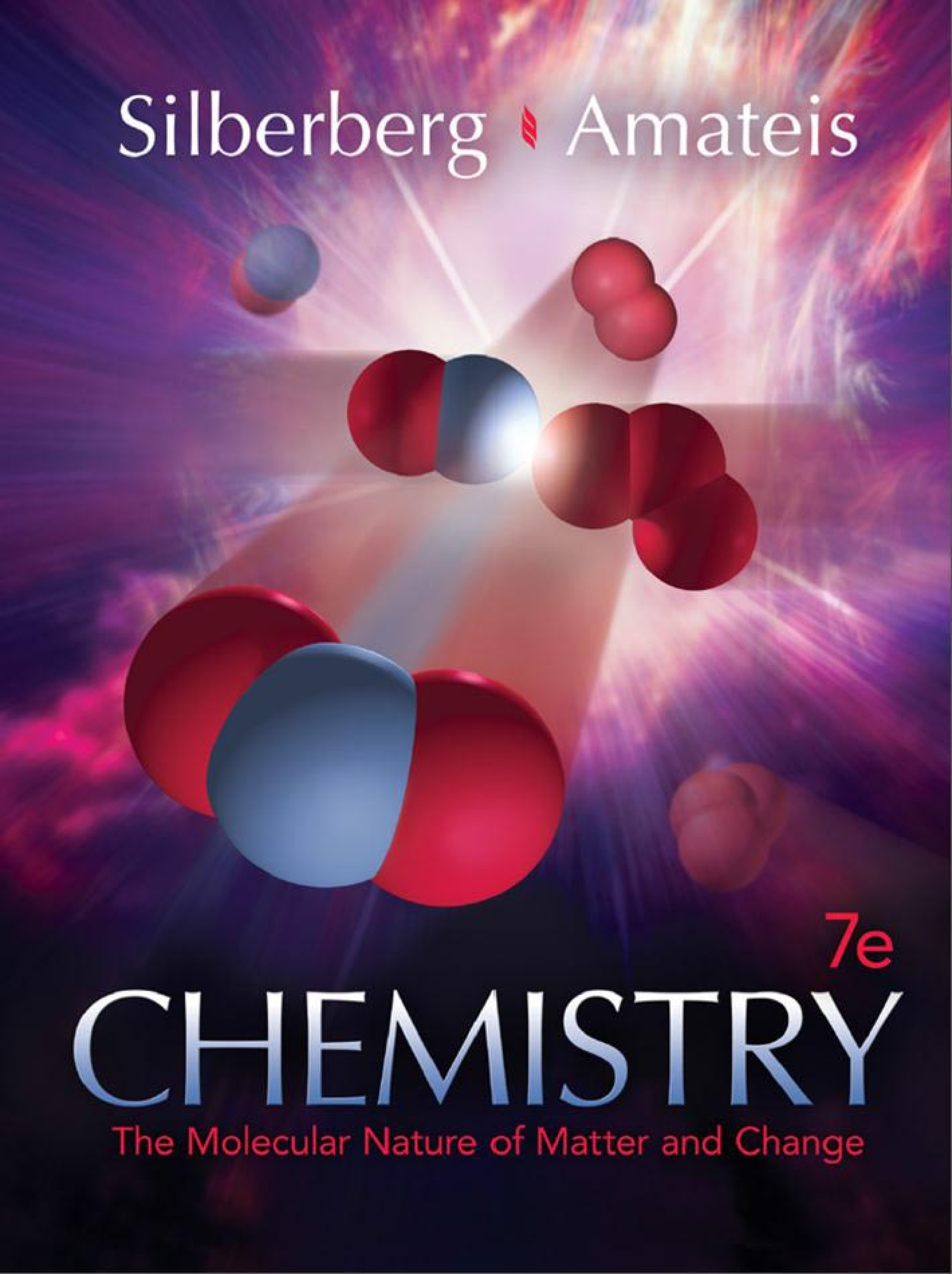
Silberberg Amateis



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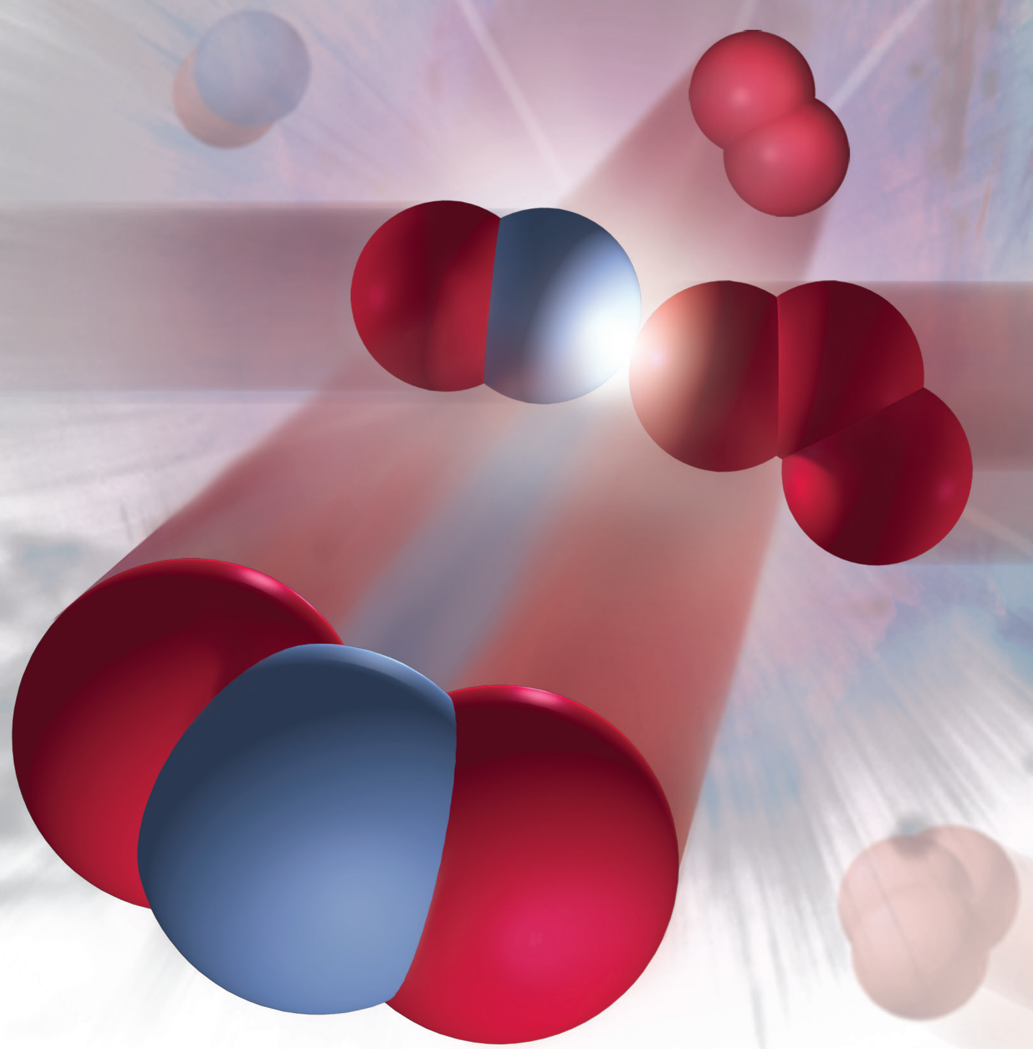
CHEMISTRY

The Molecular Nature of Matter and Change

Silberberg ◆ Amateis



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CHEMISTRY

The Molecular Nature of Matter and Change

Mc
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Hill
Education

CHEMISTRY: THE MOLECULAR NATURE OF MATTER AND CHANGE, SEVENTH EDITION

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To Ruth and Daniel, with all my love and gratitude.
MSS

To Ralph, Eric, Samantha, and Lindsay:
you bring me much joy.
PGA

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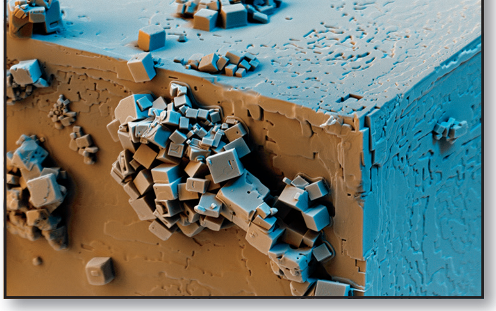

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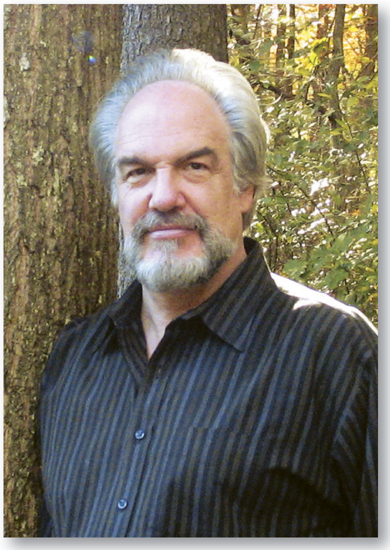


Martin S. Silberberg received a B.S. in Chemistry from the City University of New York and a Ph.D. in Chemistry from the University of Oklahoma. He then accepted a position as research associate in analytical biochemistry at the Albert Einstein College of Medicine in New York City, where he developed methods to study neurotransmitter metabolism in Parkinson's disease and other neurological disorders. Following six years in neurochemical research, Dr. Silberberg joined the faculty of Bard College at Simon's Rock, a liberal arts college known for its excellence in teaching small classes of highly motivated students. As head of the Natural Sciences Major and Director of Premedical Studies, he taught courses in general chemistry, organic chemistry, biochemistry, and liberal-arts chemistry. The small class size and close student contact afforded him insights into how students learn chemistry, where they have difficulties, and what strategies can help them succeed. Dr. Silberberg decided to apply these insights in a broader context and established a textbook writing, editing, and consulting company. Before writing his own texts, he worked as a consulting and development editor on chemistry, biochemistry, and physics texts for several major college publishers. He resides with his wife Ruth in the Pioneer Valley near Amherst, Massachusetts, where he enjoys the rich cultural and academic life of the area and relaxes by cooking, gardening, and singing.

Patricia G. Amateis graduated with a B.S. in Chemistry Education from Concord University in West Virginia and a Ph.D. in Analytical Chemistry from Virginia Tech. She has been on the faculty of the Chemistry Department at Virginia Tech for 28 years, teaching General Chemistry and Analytical Chemistry. For the past 13 years, she has served as Director of General Chemistry, responsible for the oversight of both the lecture and lab portions of the large General Chemistry program. She has taught thousands of students during her career and has been awarded the University Sporn Award for Introductory Teaching, the Alumni Teaching Award, and the William E. Wine Award for a history of university teaching excellence. She and her husband live in Blacksburg, Virginia and are the parents of three adult children. In her free time, she enjoys biking, hiking, competing in the occasional sprint triathlon, and playing the double second in Panjammers, Blacksburg's steel drum band.

PREFACE

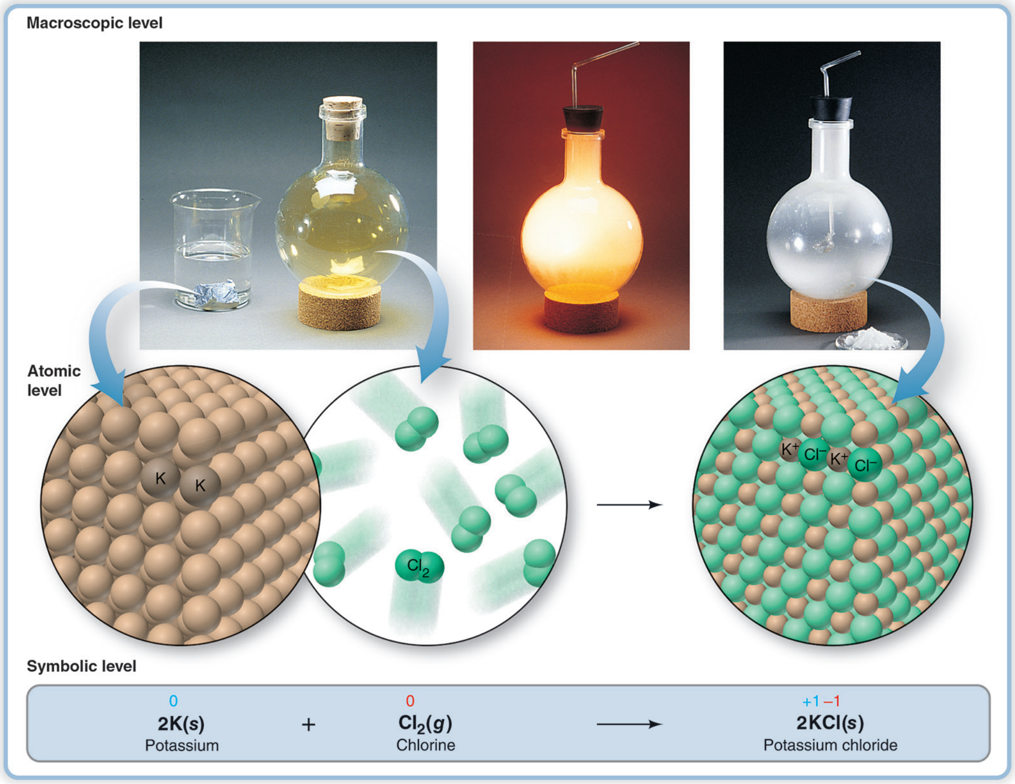
Chemistry is so crucial to an understanding of medicine and biology, environmental science, and many areas of engineering and industrial processing that it has become a requirement for an increasing number of academic majors. Furthermore, chemical principles lie at the core of some of the key societal issues we face in the 21st century—dealing with climate change, finding new energy options, and supplying nutrition and curing disease on an ever more populated planet.

SETTING THE STANDARD FOR A CHEMISTRY TEXT

The seventh edition of *Chemistry: The Molecular Nature of Matter and Change* maintains its standard-setting position among general chemistry textbooks by evolving further to meet the needs of professor and student. The text still contains the most accurate molecular illustrations, consistent step-by-step worked problems, and an extensive collection of end-of-chapter problems. And changes throughout this edition make the text more readable and succinct, the artwork more teachable and modern, and the design more focused and inviting. The three hallmarks that have made this text a market leader are now demonstrated in its pages more clearly than ever.

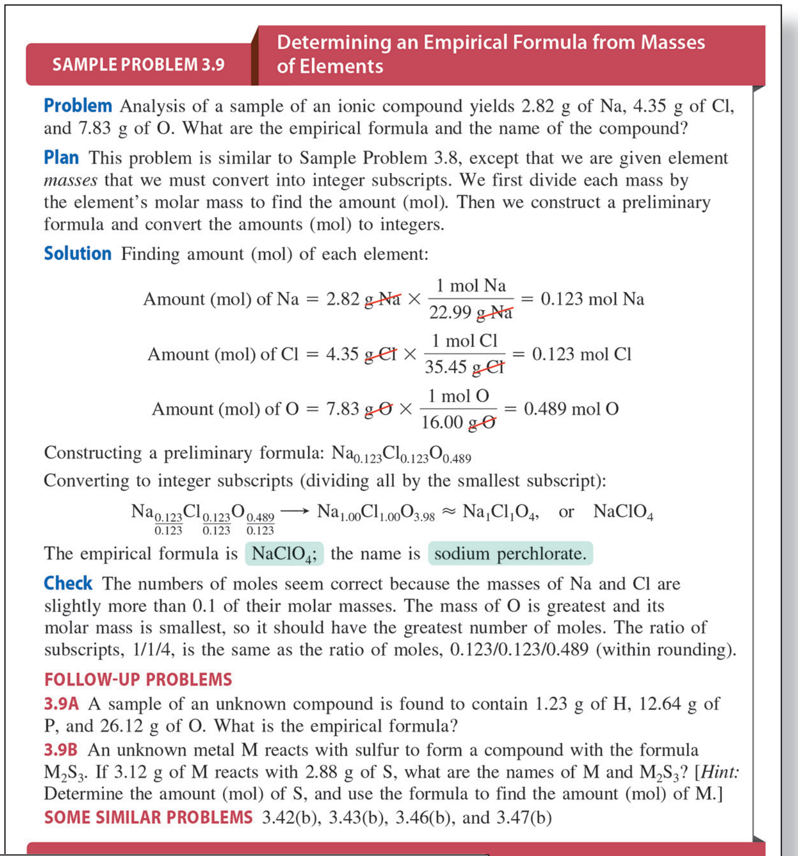
Visualizing Chemical Models—Macroscopic to Molecular

Chemistry deals with observable changes caused by unobservable atomic-scale events, requiring an appreciation of a size gap of mind-boggling proportions. One of the text's goals coincides with that of so many instructors: to help students visualize chemical events on the molecular scale. Thus, concepts are explained first at the macroscopic level and then from a molecular point of view, with pedagogic illustrations always placed next to the discussions to bring the point home for today's visually oriented students.



Thinking Logically to Solve Problems

The problem-solving approach, based on the four-step method widely accepted by experts in chemical education, is introduced in Chapter 1 and employed *consistently* throughout the text. It encourages students to *plan* a logical approach to a problem, and only then proceed to *solve* it. Each sample problem includes a *check,* which fosters the habit of "thinking through" both the chemical and the quantitative reasonableness of the answer. Finally, for *practice* and reinforcement, each sample problem is now followed immediately by *two* similar follow-up problems. And, *Chemistry* marries problem solving to visualizing models with molecular-scene problems, which appear not only in homework sets, as in other texts, but also in the running text, where they are worked out stepwise.



SAMPLE PROBLEM 3.9 Determining an Empirical Formula from Masses of Elements

Problem Analysis of a sample of an ionic compound yields 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What are the empirical formula and the name of the compound?

Plan This problem is similar to Sample Problem 3.8, except that we are given element *masses* that we must convert into integer subscripts. We first divide each mass by the element's molar mass to find the amount (mol). Then we construct a preliminary formula and convert the amounts (mol) to integers.

Solution Finding amount (mol) of each element:

$$\text{Amount (mol) of Na} = 2.82\ \cancel{\text{g Na}} \times \frac{1\ \text{mol Na}}{22.99\ \cancel{\text{g Na}}} = 0.123\ \text{mol Na}$$

$$\text{Amount (mol) of Cl} = 4.35\ \cancel{\text{g Cl}} \times \frac{1\ \text{mol Cl}}{35.45\ \cancel{\text{g Cl}}} = 0.123\ \text{mol Cl}$$

$$\text{Amount (mol) of O} = 7.83\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.489\ \text{mol O}$$

Constructing a preliminary formula: $Na_{0.123}Cl_{0.123}O_{0.489}$

Converting to integer subscripts (dividing all by the smallest subscript):

$$Na_{\frac{0.123}{0.123}}Cl_{\frac{0.123}{0.123}}O_{\frac{0.489}{0.123}} \longrightarrow Na_{1.00}Cl_{1.00}O_{3.98} \approx Na_1Cl_1O_4, \quad \text{or} \quad NaClO_4$$

The empirical formula is $NaClO_4$; the name is sodium perchlorate.

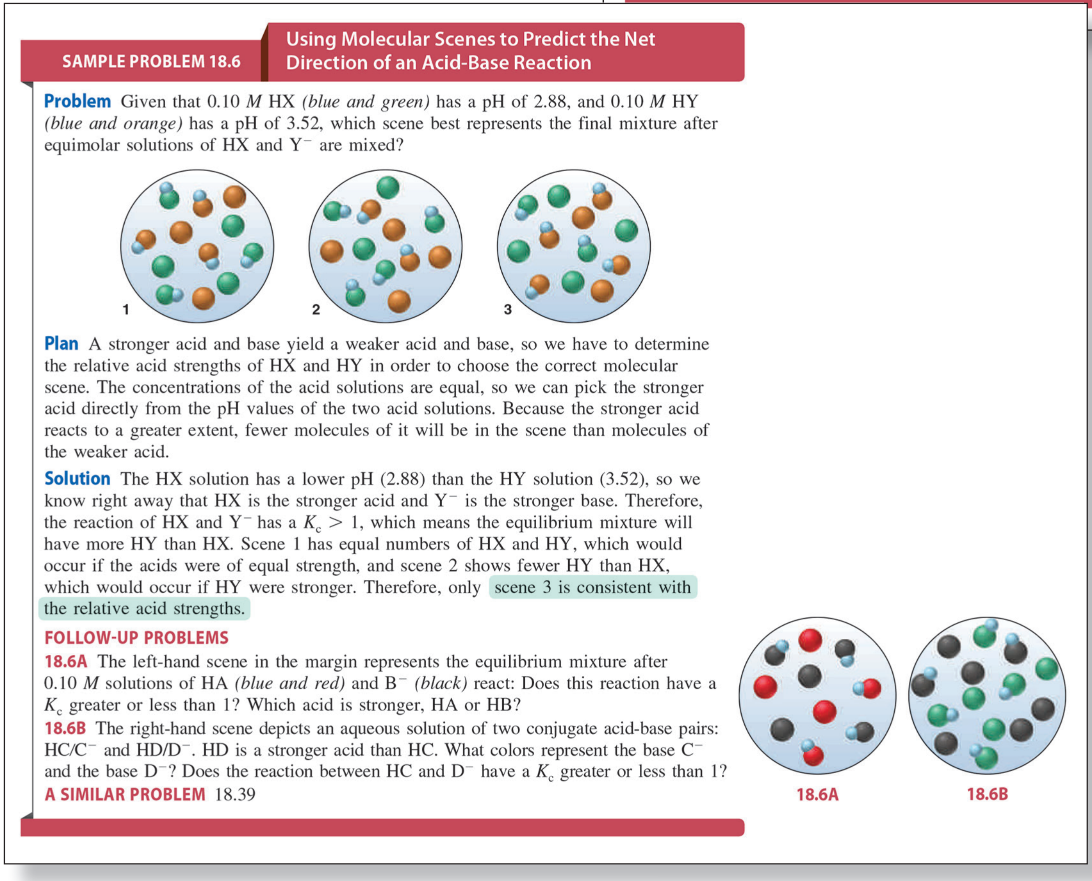
Check The numbers of moles seem correct because the masses of Na and Cl are slightly more than 0.1 of their molar masses. The mass of O is greatest and its molar mass is smallest, so it should have the greatest number of moles. The ratio of subscripts, 1/1/4, is the same as the ratio of moles, 0.123/0.123/0.489 (within rounding).

FOLLOW-UP PROBLEMS

3.9A A sample of an unknown compound is found to contain 1.23 g of H, 12.64 g of P, and 26.12 g of O. What is the empirical formula?

3.9B An unknown metal M reacts with sulfur to form a compound with the formula M_2S_3. If 3.12 g of M reacts with 2.88 g of S, what are the names of M and M_2S_3? [*Hint:* Determine the amount (mol) of S, and use the formula to find the amount (mol) of M.]

SOME SIMILAR PROBLEMS 3.42(b), 3.43(b), 3.46(b), and 3.47(b)



SAMPLE PROBLEM 18.6 Using Molecular Scenes to Predict the Net Direction of an Acid-Base Reaction

Problem Given that 0.10 *M* HX *(blue and green)* has a pH of 2.88, and 0.10 *M* HY *(blue and orange)* has a pH of 3.52, which scene best represents the final mixture after equimolar solutions of HX and Y^- are mixed?

Plan A stronger acid and base yield a weaker acid and base, so we have to determine the relative acid strengths of HX and HY in order to choose the correct molecular scene. The concentrations of the acid solutions are equal, so we can pick the stronger acid directly from the pH values of the two acid solutions. Because the stronger acid reacts to a greater extent, fewer molecules of it will be in the scene than molecules of the weaker acid.

Solution The HX solution has a lower pH (2.88) than the HY solution (3.52), so we know right away that HX is the stronger acid and Y^- is the stronger base. Therefore, the reaction of HX and Y^- has a $K_c > 1$, which means the equilibrium mixture will have more HY than HX. Scene 1 has equal numbers of HX and HY, which would occur if the acids were of equal strength, and scene 2 shows fewer HY than HX, which would occur if HY were stronger. Therefore, only scene 3 is consistent with the relative acid strengths.

FOLLOW-UP PROBLEMS

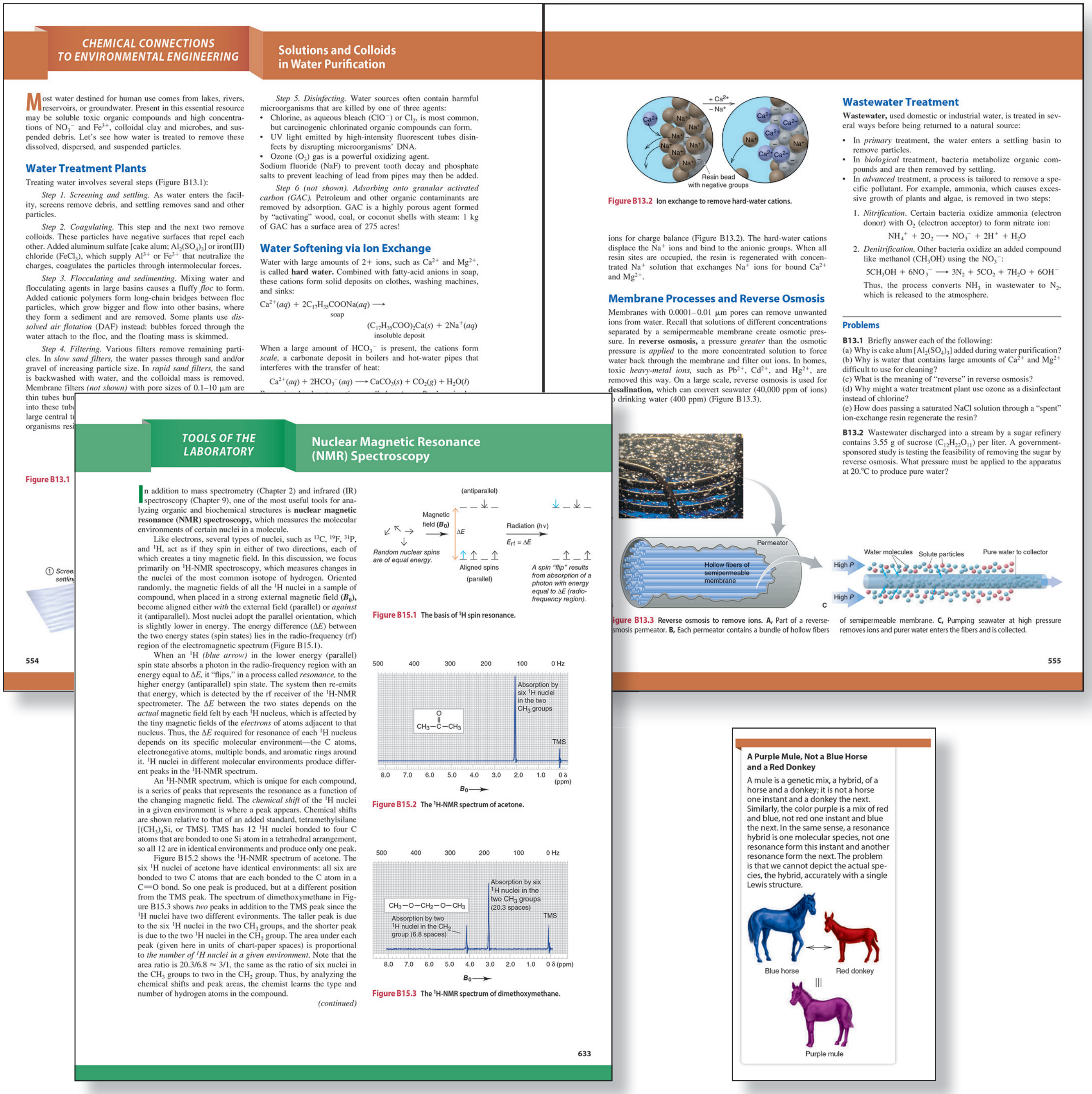
18.6A The left-hand scene in the margin represents the equilibrium mixture after 0.10 *M* solutions of HA *(blue and red)* and B^- *(black)* react: Does this reaction have a K_c greater or less than 1? Which acid is stronger, HA or HB?

18.6B The right-hand scene depicts an aqueous solution of two conjugate acid-base pairs: HC/C^- and HD/D^-. HD is a stronger acid than HC. What colors represent the base C^- and the base D^-? Does the reaction between HC and D^- have a K_c greater or less than 1?

A SIMILAR PROBLEM 18.39

Applying Ideas to the Real World

As the most practical science, chemistry should have a textbook that highlights its countless applications. Moreover, today's students may enter emerging chemistry-related hybrid fields, like biomaterials science or planetary geochemistry, and the text they use should point out the relevance of chemical concepts to such related sciences. The *Chemical Connections* and *Tools of the Laboratory* boxed essays (which include problems for added relevance), the more pedagogic margin notes, and the many applications woven into the chapter content are up-to-date, student-friendly features that are directly related to the neighboring content.



CHEMICAL CONNECTIONS TO ENVIRONMENTAL ENGINEERING

Solutions and Colloids in Water Purification

Most water destined for human use comes from lakes, rivers, reservoirs, or groundwater. Present in this essential resource may be soluble toxic organic compounds and high concentrations of NO_3^- and Fe^{3+}, colloidal clay and microbes, and suspended debris. Let's see how water is treated to remove these dissolved, dispersed, and suspended particles.

Water Treatment Plants

Treating water involves several steps (Figure B13.1):

Step 1. Screening and settling. As water enters the facility, screens remove debris, and settling removes sand and other particles.

Step 2. Coagulating. This step and the next two remove colloids. These particles have negative surfaces that repel each other. Added aluminum sulfate [cake alum; $Al_2(SO_4)_3$] or iron(III) chloride ($FeCl_3$), which supply Al^{3+} or Fe^{3+} that neutralize the charges, coagulates the particles through intermolecular forces.

Step 3. Flocculating and sedimenting. Mixing water and flocculating agents in large basins causes a fluffy *floc* to form. Added cationic polymers form long-chain bridges between floc particles, which grow bigger and flow into other basins, where they form a sediment and are removed. Some plants use *dissolved air flotation* (DAF) instead: bubbles forced through the water attach to the floc, and the floating mass is skimmed.

Step 4. Filtering. Various filters remove remaining particles. In *slow sand filters,* the water passes through sand and/or gravel of increasing particle size. In *rapid sand filters,* the sand is backwashed with water, and the colloidal mass is removed. Membrane filters *(not shown)* with pore sizes of 0.1–10 µm are thin tubes bun

into these tub

large central tu

organisms resi

Figure B13.1

① Scree settling

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Step 5. Disinfecting. Water sources often contain harmful microorganisms that are killed by one of three agents:

- Chlorine, as aqueous bleach (ClO^-) or Cl_2, is most common, but carcinogenic chlorinated organic compounds can form.
- UV light emitted by high-intensity fluorescent tubes disinfects by disrupting microorganisms' DNA.
- Ozone (O_3) gas is a powerful oxidizing agent.

Sodium fluoride (NaF) to prevent tooth decay and phosphate salts to prevent leaching of lead from pipes may then be added.

Step 6 (not shown). Adsorbing onto granular activated carbon (GAC). Petroleum and other organic contaminants are removed by adsorption. GAC is a highly porous agent formed by "activating" wood, coal, or coconut shells with steam: 1 kg of GAC has a surface area of 275 acres!

Water Softening via Ion Exchange

Water with large amounts of 2+ ions, such as Ca^{2+} and Mg^{2+}, is called **hard water.** Combined with fatty-acid anions in soap, these cations form solid deposits on clothes, washing machines, and sinks:

$$Ca^{2+}(aq) + \underset{\text{soap}}{2C_{17}H_{35}COONa(aq)} \longrightarrow \underset{\text{insoluble deposit}}{(C_{17}H_{35}COO)_2Ca(s)} + 2Na^+(aq)$$

When a large amount of HCO_3^- is present, the cations form *scale,* a carbonate deposit in boilers and hot-water pipes that interferes with the transfer of heat:

$$Ca^{2+}(aq) + 2HCO_3^-(aq) \longrightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$$

Figure B13.2 Ion exchange to remove hard-water cations.

ions for charge balance (Figure B13.2). The hard-water cations displace the Na^+ ions and bind to the anionic groups. When all resin sites are occupied, the resin is regenerated with concentrated Na^+ solution that exchanges Na^+ ions for bound Ca^{2+} and Mg^{2+}.

Membrane Processes and Reverse Osmosis

Membranes with 0.0001–0.01 µm pores can remove unwanted ions from water. Recall that solutions of different concentrations separated by a semipermeable membrane create osmotic pressure. In **reverse osmosis,** a pressure *greater* than the osmotic pressure is *applied* to the more concentrated solution to force water back through the membrane and filter out ions. In homes, toxic *heavy-metal ions,* such as Pb^{2+}, Cd^{2+}, and Hg^{2+}, are removed this way. On a large scale, reverse osmosis is used for **desalination,** which can convert seawater (40,000 ppm of ions) drinking water (400 ppm) (Figure B13.3).

Wastewater Treatment

Wastewater, used domestic or industrial water, is treated in several ways before being returned to a natural source:

- In *primary* treatment, the water enters a settling basin to remove particles.
- In *biological* treatment, bacteria metabolize organic compounds and are then removed by settling.
- In *advanced* treatment, a process is tailored to remove a specific pollutant. For example, ammonia, which causes excessive growth of plants and algae, is removed in two steps:
  1. *Nitrification.* Certain bacteria oxidize ammonia (electron donor) with O_2 (electron acceptor) to form nitrate ion:
  $$NH_4^+ + 2O_2 \longrightarrow NO_3^- + 2H^+ + H_2O$$
  2. *Denitrification.* Other bacteria oxidize an added compound like methanol (CH_3OH) using the NO_3^-:
  $$5CH_3OH + 6NO_3^- \longrightarrow 3N_2 + 5CO_2 + 7H_2O + 6OH^-$$
  Thus, the process converts NH_3 in wastewater to N_2, which is released to the atmosphere.

Problems

B13.1 Briefly answer each of the following:
(a) Why is cake alum [$Al_2(SO_4)_3$] added during water purification?
(b) Why is water that contains large amounts of Ca^{2+} and Mg^{2+} difficult to use for cleaning?
(c) What is the meaning of "reverse" in reverse osmosis?
(d) Why might a water treatment plant use ozone as a disinfectant instead of chlorine?
(e) How does passing a saturated NaCl solution through a "spent" ion-exchange resin regenerate the resin?

B13.2 Wastewater discharged into a stream by a sugar refinery contains 3.55 g of sucrose ($C_{12}H_{22}O_{11}$) per liter. A government-sponsored study is testing the feasibility of removing the sugar by reverse osmosis. What pressure must be applied to the apparatus at 20.°C to produce pure water?

igure B13.3 Reverse osmosis to remove ions. A, Part of a reverse-smosis permeator. B, Each permeator contains a bundle of hollow fibers of semipermeable membrane. C, Pumping seawater at high pressure removes ions and purer water enters the fibers and is collected.

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TOOLS OF THE LABORATORY

Nuclear Magnetic Resonance (NMR) Spectroscopy

In addition to mass spectrometry (Chapter 2) and infrared (IR) spectroscopy (Chapter 9), one of the most useful tools for analyzing organic and biochemical structures is **nuclear magnetic resonance (NMR) spectroscopy,** which measures the molecular environments of certain nuclei in a molecule.

Like electrons, several types of nuclei, such as ^{13}C, ^{19}F, ^{31}P, and ^1H, act as if they spin in either of two directions, each of which creates a tiny magnetic field. In this discussion, we focus primarily on ^1H-NMR spectroscopy, which measures changes in the nuclei of the most common isotope of hydrogen. Oriented randomly, the magnetic fields of all the ^1H nuclei in a sample of compound, when placed in a strong external magnetic field ($\boldsymbol{B_0}$), become aligned either *with* the external field (parallel) or *against* it (antiparallel). Most nuclei adopt the parallel orientation, which is slightly lower in energy. The energy difference (ΔE) between the two energy states (spin states) lies in the radio-frequency (rf) region of the electromagnetic spectrum (Figure B15.1).

When an ^1H *(blue arrow)* in the lower energy (parallel) spin state absorbs a photon in the radio-frequency region with an energy equal to ΔE, it "flips," in a process called *resonance,* to the higher energy (antiparallel) spin state. The system then re-emits that energy, which is detected by the rf receiver of the ^1H-NMR spectrometer. The ΔE between the two states depends on the *actual* magnetic field felt by each ^1H nucleus, which is affected by the tiny magnetic fields of the *electrons* of atoms adjacent to that nucleus. Thus, the ΔE required for resonance of each ^1H nucleus depends on its specific molecular environment—the C atoms, electronegative atoms, multiple bonds, and aromatic rings around it. ^1H nuclei in different molecular environments produce different peaks in the ^1H-NMR spectrum.

An ^1H-NMR spectrum, which is unique for each compound, is a series of peaks that represents the resonance as a function of the changing magnetic field. The *chemical shift* of the ^1H nuclei in a given environment is where a peak appears. Chemical shifts are shown relative to that of an added standard, tetramethylsilane [$(CH_3)_4Si$, or TMS]. TMS has 12 ^1H nuclei bonded to four C atoms that are bonded to one Si atom in a tetrahedral arrangement, so all 12 are in identical environments and produce only one peak.

Figure B15.2 shows the ^1H-NMR spectrum of acetone. The six ^1H nuclei of acetone have identical environments: all six are bonded to two C atoms that are each bonded to the C atom in a C=O bond. So one peak is produced, but at a different position from the TMS peak. The spectrum of dimethoxymethane in Figure B15.3 shows *two* peaks in addition to the TMS peak since the ^1H nuclei have two different evironments. The taller peak is due to the six ^1H nuclei in the two CH_3 groups, and the shorter peak is due to the two ^1H nuclei in the CH_2 group. The area under each peak (given here in units of chart-paper spaces) is proportional to *the number of ^1H nuclei in a given environment*. Note that the area ratio is $20.3/6.8 \approx 3/1$, the same as the ratio of six nuclei in the CH_3 groups to two in the CH_2 group. Thus, by analyzing the chemical shifts and peak areas, the chemist learns the type and number of hydrogen atoms in the compound.

(continued)

Figure B15.1 The basis of ^1H spin resonance.

Figure B15.2 The ^1H-NMR spectrum of acetone.

Figure B15.3 The ^1H-NMR spectrum of dimethoxymethane.

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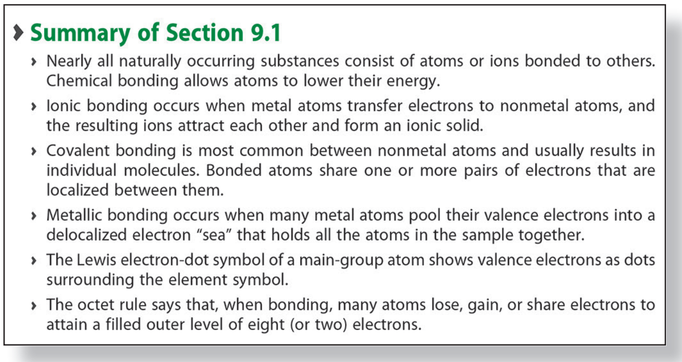
A Purple Mule, Not a Blue Horse and a Red Donkey

A mule is a genetic mix, a hybrid, of a horse and a donkey; it is not a horse one instant and a donkey the next. Similarly, the color purple is a mix of red and blue, not red one instant and blue the next. In the same sense, a resonance hybrid is one molecular species, not one resonance form this instant and another resonance form the next. The problem is that we cannot depict the actual species, the hybrid, accurately with a single Lewis structure.



Summary of Section 9.1

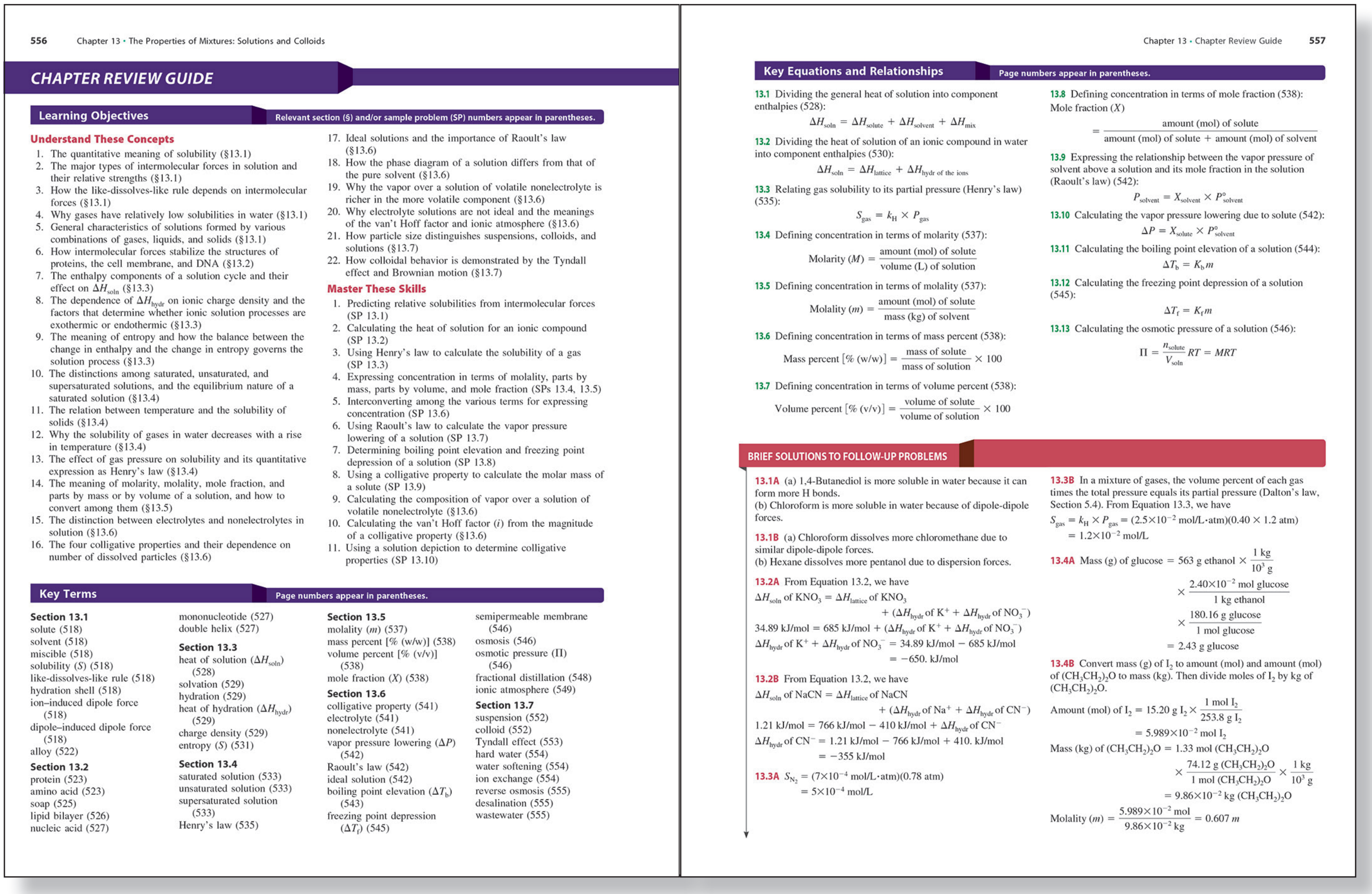
- Nearly all naturally occurring substances consist of atoms or ions bonded to others. Chemical bonding allows atoms to lower their energy.
- Ionic bonding occurs when metal atoms transfer electrons to nonmetal atoms, and the resulting ions attract each other and form an ionic solid.
- Covalent bonding is most common between nonmetal atoms and usually results in individual molecules. Bonded atoms share one or more pairs of electrons that are localized between them.
- Metallic bonding occurs when many metal atoms pool their valence electrons into a delocalized electron "sea" that holds all the atoms in the sample together.
- The Lewis electron-dot symbol of a main-group atom shows valence electrons as dots surrounding the element symbol.
- The octet rule says that, when bonding, many atoms lose, gain, or share electrons to attain a filled outer level of eight (or two) electrons.

Reinforcing through Review and Practice

A favorite feature, the section summaries that conclude every section restate the major ideas concisely and immediately (rather than postponing such review until the end of the chapter).

A rich catalog of study aids ends each chapter to help students review the content:

- **Learning Objectives,** with section and/or sample problem numbers, focus on the concepts to understand and the skills to master.
- **Key Terms,** boldfaced and defined within the chapter, are listed here by section (with page numbers), as well as being defined in the *Glossary.*
- **Key Equations and Relationships** are highlighted and numbered within the chapter and listed here with page numbers.
- **Brief Solutions to Follow-up Problems** triple the number of worked problems by providing multistep calculations at the end of the chapter, rather than just numerical answers at the back of the book.



556 Chapter 13 • The Properties of Mixtures: Solutions and Colloids

CHAPTER REVIEW GUIDE

Learning Objectives Relevant section (§) and/or sample problem (SP) numbers appear in parentheses.

Understand These Concepts

1. The quantitative meaning of solubility (§13.1)
2. The major types of intermolecular forces in solution and their relative strengths (§13.1)
3. How the like-dissolves-like rule depends on intermolecular forces (§13.1)
4. Why gases have relatively low solubilities in water (§13.1)
5. General characteristics of solutions formed by various combinations of gases, liquids, and solids (§13.1)
6. How intermolecular forces stabilize the structures of proteins, the cell membrane, and DNA (§13.2)
7. The enthalpy components of a solution cycle and their effect on ΔH_{soln} (§13.3)
8. The dependence of ΔH_{hydr} on ionic charge density and the factors that determine whether ionic solution processes are exothermic or endothermic (§13.3)
9. The meaning of entropy and how the balance between the change in enthalpy and the change in entropy governs the solution process (§13.3)
10. The distinctions among saturated, unsaturated, and supersaturated solutions, and the equilibrium nature of a saturated solution (§13.4)
11. The relation between temperature and the solubility of solids (§13.4)
12. Why the solubility of gases in water decreases with a rise in temperature (§13.4)
13. The effect of gas pressure on solubility and its quantitative expression as Henry's law (§13.4)
14. The meaning of molarity, molality, mole fraction, and parts by mass or by volume of a solution, and how to convert among them (§13.5)
15. The distinction between electrolytes and nonelectrolytes in solution (§13.6)
16. The four colligative properties and their dependence on number of dissolved particles (§13.6)
17. Ideal solutions and the importance of Raoult's law (§13.6)
18. How the phase diagram of a solution differs from that of the pure solvent (§13.6)
19. Why the vapor over a solution of volatile nonelectrolyte is richer in the more volatile component (§13.6)
20. Why electrolyte solutions are not ideal and the meanings of the van't Hoff factor and ionic atmosphere (§13.6)
21. How particle size distinguishes suspensions, colloids, and solutions (§13.7)
22. How colloidal behavior is demonstrated by the Tyndall effect and Brownian motion (§13.7)

Master These Skills

1. Predicting relative solubilities from intermolecular forces (SP 13.1)
2. Calculating the heat of solution for an ionic compound (SP 13.2)
3. Using Henry's law to calculate the solubility of a gas (SP 13.3)
4. Expressing concentration in terms of molality, parts by mass, parts by volume, and mole fraction (SPs 13.4, 13.5)
5. Interconverting among the various terms for expressing concentration (SP 13.6)
6. Using Raoult's law to calculate the vapor pressure lowering of a solution (SP 13.7)
7. Determining boiling point elevation and freezing point depression of a solution (SP 13.8)
8. Using a colligative property to calculate the molar mass of a solute (SP 13.9)
9. Calculating the composition of vapor over a solution of volatile nonelectrolyte (§13.6)
10. Calculating the van't Hoff factor (i) from the magnitude of a colligative property (§13.6)
11. Using a solution depiction to determine colligative properties (SP 13.10)

Key Terms Page numbers appear in parentheses.

Section 13.1
solute (518)
solvent (518)
miscible (518)
solubility (S) (518)
like-dissolves-like rule (518)
hydration shell (518)
ion–induced dipole force (518)
dipole–induced dipole force (518)
alloy (522)

Section 13.2
protein (523)
amino acid (523)
soap (525)
lipid bilayer (526)
nucleic acid (527)
mononucleotide (527)
double helix (527)

Section 13.3
heat of solution (ΔH_{soln}) (528)
solvation (529)
hydration (529)
heat of hydration (ΔH_{hydr}) (529)
charge density (529)
entropy (S) (531)

Section 13.4
saturated solution (533)
unsaturated solution (533)
supersaturated solution (533)
Henry's law (535)

Section 13.5
molality (m) (537)
mass percent [% (w/w)] (538)
volume percent [% (v/v)] (538)
mole fraction (X) (538)

Section 13.6
colligative property (541)
electrolyte (541)
nonelectrolyte (541)
vapor pressure lowering (ΔP) (542)
Raoult's law (542)
ideal solution (542)
boiling point elevation (ΔT_b) (543)
freezing point depression (ΔT_f) (545)
semipermeable membrane (546)
osmosis (546)
osmotic pressure (Π) (546)
fractional distillation (548)
ionic atmosphere (549)

Section 13.7
suspension (552)
colloid (552)
Tyndall effect (553)
hard water (554)
water softening (554)
ion exchange (554)
reverse osmosis (555)
desalination (555)
wastewater (555)

Chapter 13 • Chapter Review Guide 557

Key Equations and Relationships Page numbers appear in parentheses.

13.1 Dividing the general heat of solution into component enthalpies (528):

$$\Delta H_{soln} = \Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mix}$$

13.2 Dividing the heat of solution of an ionic compound in water into component enthalpies (530):

$$\Delta H_{soln} = \Delta H_{lattice} + \Delta H_{hydr\ of\ the\ ions}$$

13.3 Relating gas solubility to its partial pressure (Henry's law) (535):

$$S_{gas} = k_H \times P_{gas}$$

13.4 Defining concentration in terms of molarity (537):

$$\text{Molarity } (M) = \frac{\text{amount (mol) of solute}}{\text{volume (L) of solution}}$$

13.5 Defining concentration in terms of molality (537):

$$\text{Molality } (m) = \frac{\text{amount (mol) of solute}}{\text{mass (kg) of solvent}}$$

13.6 Defining concentration in terms of mass percent (538):

$$\text{Mass percent } [\% \text{ (w/w)}] = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

13.7 Defining concentration in terms of volume percent (538):

$$\text{Volume percent } [\% \text{ (v/v)}] = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

13.8 Defining concentration in terms of mole fraction (538):
Mole fraction (X)

$$= \frac{\text{amount (mol) of solute}}{\text{amount (mol) of solute} + \text{amount (mol) of solvent}}$$

13.9 Expressing the relationship between the vapor pressure of solvent above a solution and its mole fraction in the solution (Raoult's law) (542):

$$P_{solvent} = X_{solvent} \times P^{\circ}_{solvent}$$

13.10 Calculating the vapor pressure lowering due to solute (542):

$$\Delta P = X_{solute} \times P^{\circ}_{solvent}$$

13.11 Calculating the boiling point elevation of a solution (544):

$$\Delta T_b = K_b m$$

13.12 Calculating the freezing point depression of a solution (545):

$$\Delta T_f = K_f m$$

13.13 Calculating the osmotic pressure of a solution (546):

$$\Pi = \frac{n_{solute}}{V_{soln}} RT = MRT$$

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS

13.1A (a) 1,4-Butanediol is more soluble in water because it can form more H bonds.
(b) Chloroform is more soluble in water because of dipole-dipole forces.

13.1B (a) Chloroform dissolves more chloromethane due to similar dipole-dipole forces.
(b) Hexane dissolves more pentanol due to dispersion forces.

13.2A From Equation 13.2, we have

$$\begin{aligned}\Delta H_{soln} \text{ of } KNO_3 &= \Delta H_{lattice} \text{ of } KNO_3 \\ &\quad + (\Delta H_{hydr} \text{ of } K^+ + \Delta H_{hydr} \text{ of } NO_3^-) \\ 34.89 \text{ kJ/mol} &= 685 \text{ kJ/mol} + (\Delta H_{hydr} \text{ of } K^+ + \Delta H_{hydr} \text{ of } NO_3^-) \\ \Delta H_{hydr} \text{ of } K^+ + \Delta H_{hydr} \text{ of } NO_3^- &= 34.89 \text{ kJ/mol} - 685 \text{ kJ/mol} \\ &= -650. \text{ kJ/mol}\end{aligned}$$

13.2B From Equation 13.2, we have

$$\begin{aligned}\Delta H_{soln} \text{ of NaCN} &= \Delta H_{lattice} \text{ of NaCN} \\ &\quad + (\Delta H_{hydr} \text{ of } Na^+ + \Delta H_{hydr} \text{ of } CN^-) \\ 1.21 \text{ kJ/mol} &= 766 \text{ kJ/mol} - 410 \text{ kJ/mol} + \Delta H_{hydr} \text{ of } CN^- \\ \Delta H_{hydr} \text{ of } CN^- &= 1.21 \text{ kJ/mol} - 766 \text{ kJ/mol} + 410. \text{ kJ/mol} \\ &= -355 \text{ kJ/mol}\end{aligned}$$

13.3A $S_{N_2} = (7\times10^{-4} \text{ mol/L·atm})(0.78 \text{ atm}) = 5\times10^{-4} \text{ mol/L}$

13.3B In a mixture of gases, the volume percent of each gas times the total pressure equals its partial pressure (Dalton's law, Section 5.4). From Equation 13.3, we have

$$S_{gas} = k_H \times P_{gas} = (2.5\times10^{-2} \text{ mol/L·atm})(0.40 \times 1.2 \text{ atm}) = 1.2\times10^{-2} \text{ mol/L}$$

13.4A

$$\begin{aligned}\text{Mass (g) of glucose} &= 563 \text{ g ethanol} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \\ &\quad \times \frac{2.40\times10^{-2} \text{ mol glucose}}{1 \text{ kg ethanol}} \\ &\quad \times \frac{180.16 \text{ g glucose}}{1 \text{ mol glucose}} \\ &= 2.43 \text{ g glucose}\end{aligned}$$

13.4B Convert mass (g) of I_2 to amount (mol) and amount (mol) of $(CH_3CH_2)_2O$ to mass (kg). Then divide moles of I_2 by kg of $(CH_3CH_2)_2O$.

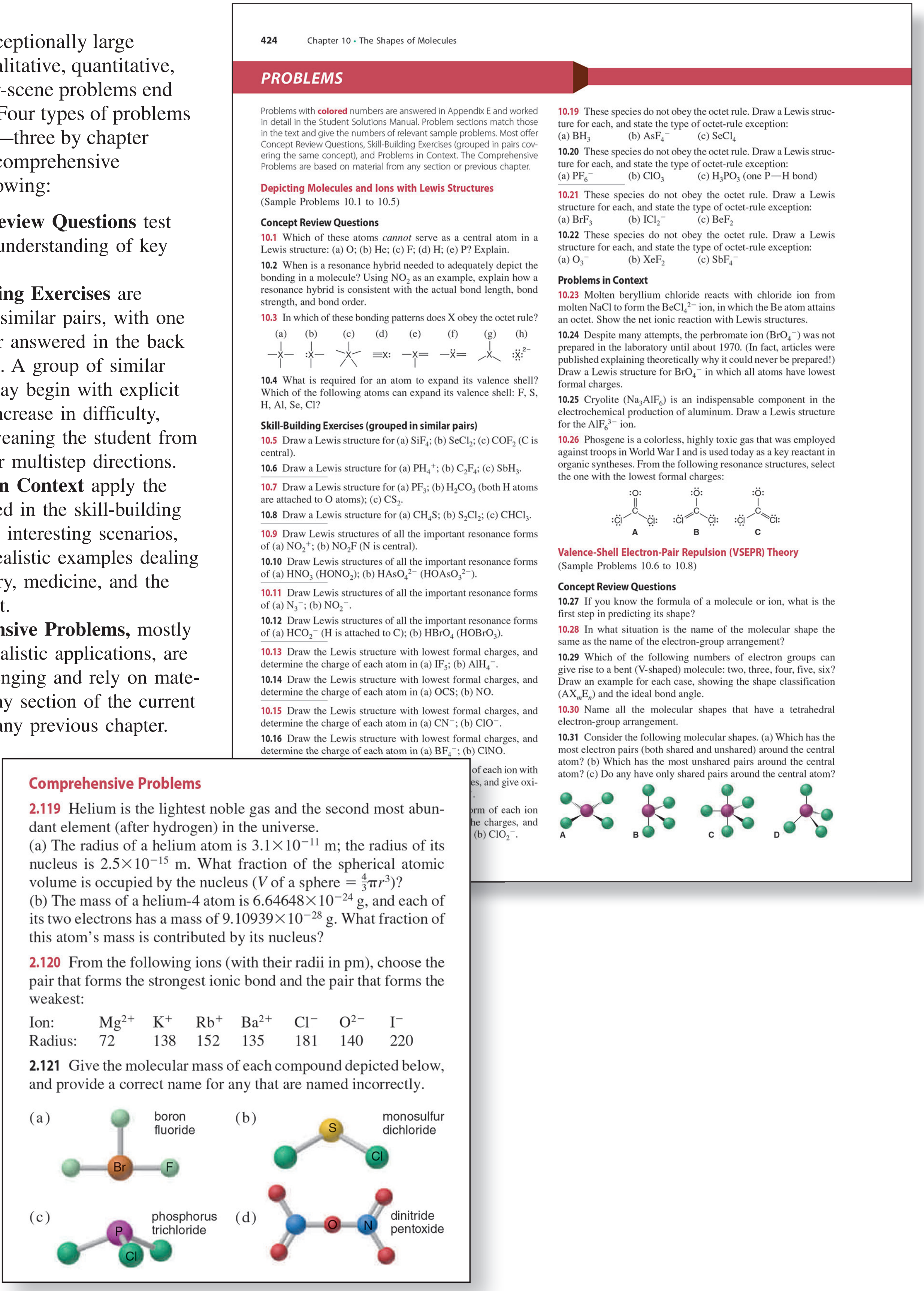
$$\text{Amount (mol) of } I_2 = 15.20 \text{ g } I_2 \times \frac{1 \text{ mol } I_2}{253.8 \text{ g } I_2} = 5.989\times10^{-2} \text{ mol } I_2$$

$$\begin{aligned}\text{Mass (kg) of } (CH_3CH_2)_2O &= 1.33 \text{ mol } (CH_3CH_2)_2O \\ &\quad \times \frac{74.12 \text{ g } (CH_3CH_2)_2O}{1 \text{ mol } (CH_3CH_2)_2O} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \\ &= 9.86\times10^{-2} \text{ kg } (CH_3CH_2)_2O\end{aligned}$$

$$\text{Molality } (m) = \frac{5.989\times10^{-2} \text{ mol}}{9.86\times10^{-2} \text{ kg}} = 0.607\ m$$

Finally, an exceptionally large number of qualitative, quantitative, and molecular-scene problems end each chapter. Four types of problems are presented—three by chapter section, with comprehensive problems following:

- **Concept Review Questions** test qualitative understanding of key ideas.
- **Skill-Building Exercises** are grouped in similar pairs, with one of each pair answered in the back of the book. A group of similar exercises may begin with explicit steps and increase in difficulty, gradually weaning the student from the need for multistep directions.
- **Problems in Context** apply the skills learned in the skill-building exercises to interesting scenarios, including realistic examples dealing with industry, medicine, and the environment.
- **Comprehensive Problems,** mostly based on realistic applications, are more challenging and rely on material from any section of the current chapter or any previous chapter.



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PROBLEMS

Problems with **colored** numbers are answered in Appendix E and worked in detail in the Student Solutions Manual. Problem sections match those in the text and give the numbers of relevant sample problems. Most offer Concept Review Questions, Skill-Building Exercises (grouped in pairs covering the same concept), and Problems in Context. The Comprehensive Problems are based on material from any section or previous chapter.

Depicting Molecules and Ions with Lewis Structures
(Sample Problems 10.1 to 10.5)

Concept Review Questions

10.1 Which of these atoms *cannot* serve as a central atom in a Lewis structure: (a) O; (b) He; (c) F; (d) H; (e) P? Explain.

10.2 When is a resonance hybrid needed to adequately depict the bonding in a molecule? Using NO_2 as an example, explain how a resonance hybrid is consistent with the actual bond length, bond strength, and bond order.

10.3 In which of these bonding patterns does X obey the octet rule?

10.4 What is required for an atom to expand its valence shell? Which of the following atoms can expand its valence shell: F, S, H, Al, Se, Cl?

Skill-Building Exercises (grouped in similar pairs)

10.5 Draw a Lewis structure for (a) SiF_4; (b) $SeCl_2$; (c) COF_2 (C is central).

10.6 Draw a Lewis structure for (a) PH_4^+; (b) C_2F_4; (c) SbH_3.

10.7 Draw a Lewis structure for (a) PF_3; (b) H_2CO_3 (both H atoms are attached to O atoms); (c) CS_2.

10.8 Draw a Lewis structure for (a) CH_4S; (b) S_2Cl_2; (c) $CHCl_3$.

10.9 Draw Lewis structures of all the important resonance forms of (a) NO_2^+; (b) NO_2F (N is central).

10.10 Draw Lewis structures of all the important resonance forms of (a) HNO_3 ($HONO_2$); (b) $HAsO_4^{2-}$ ($HOAsO_3^{2-}$).

10.11 Draw Lewis structures of all the important resonance forms of (a) N_3^-; (b) NO_2^-.

10.12 Draw Lewis structures of all the important resonance forms of (a) HCO_2^- (H is attached to C); (b) $HBrO_4$ ($HOBrO_3$).

10.13 Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) IF_5; (b) AlH_4^-.

10.14 Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) OCS; (b) NO.

10.15 Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) CN^-; (b) ClO^-.

10.16 Draw the Lewis structure with lowest formal charges, and determine the charge of each atom in (a) BF_4^-; (b) ClNO.

of each ion with
es, and give oxi-
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rm of each ion
he charges, and
(b) ClO_2^-.

10.19 These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:
(a) BH_3 (b) AsF_4^- (c) $SeCl_4$

10.20 These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:
(a) PF_6^- (b) ClO_3 (c) H_3PO_3 (one P—H bond)

10.21 These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:
(a) BrF_3 (b) ICl_2^- (c) BeF_2

10.22 These species do not obey the octet rule. Draw a Lewis structure for each, and state the type of octet-rule exception:
(a) O_3^- (b) XeF_2 (c) SbF_4^-

Problems in Context

10.23 Molten beryllium chloride reacts with chloride ion from molten NaCl to form the $BeCl_4^{2-}$ ion, in which the Be atom attains an octet. Show the net ionic reaction with Lewis structures.

10.24 Despite many attempts, the perbromate ion (BrO_4^-) was not prepared in the laboratory until about 1970. (In fact, articles were published explaining theoretically why it could never be prepared!) Draw a Lewis structure for BrO_4^- in which all atoms have lowest formal charges.

10.25 Cryolite (Na_3AlF_6) is an indispensable component in the electrochemical production of aluminum. Draw a Lewis structure for the AlF_6^{3-} ion.

10.26 Phosgene is a colorless, highly toxic gas that was employed against troops in World War I and is used today as a key reactant in organic syntheses. From the following resonance structures, select the one with the lowest formal charges:

Valence-Shell Electron-Pair Repulsion (VSEPR) Theory
(Sample Problems 10.6 to 10.8)

Concept Review Questions

10.27 If you know the formula of a molecule or ion, what is the first step in predicting its shape?

10.28 In what situation is the name of the molecular shape the same as the name of the electron-group arrangement?

10.29 Which of the following numbers of electron groups can give rise to a bent (V-shaped) molecule: two, three, four, five, six? Draw an example for each case, showing the shape classification (AX_mE_n) and the ideal bond angle.

10.30 Name all the molecular shapes that have a tetrahedral electron-group arrangement.

10.31 Consider the following molecular shapes. (a) Which has the most electron pairs (both shared and unshared) around the central atom? (b) Which has the most unshared pairs around the central atom? (c) Do any have only shared pairs around the central atom?

Comprehensive Problems

2.119 Helium is the lightest noble gas and the second most abundant element (after hydrogen) in the universe.
(a) The radius of a helium atom is 3.1×10^{-11} m; the radius of its nucleus is 2.5×10^{-15} m. What fraction of the spherical atomic volume is occupied by the nucleus (V of a sphere $= \frac{4}{3}\pi r^3$)?
(b) The mass of a helium-4 atom is 6.64648×10^{-24} g, and each of its two electrons has a mass of 9.10939×10^{-28} g. What fraction of this atom's mass is contributed by its nucleus?

2.120 From the following ions (with their radii in pm), choose the pair that forms the strongest ionic bond and the pair that forms the weakest:

Ion:	Mg^{2+}	K^+	Rb^+	Ba^{2+}	Cl^-	O^{2-}	I^-
Radius:	72	138	152	135	181	140	220

2.121 Give the molecular mass of each compound depicted below, and provide a correct name for any that are named incorrectly.

OPTIMIZING THE TEXT

The modern chemistry student's learning experience is changing dramatically. To address the changes that students face, a modern text partnered with a suite of robust electronic tools must continue to evolve. With each edition, students and instructors alike have been involved in refining this text. From one-on-one interviews, focus groups, and symposia, as well as extensive chapter reviews and class tests of the previous edition, we learned that everyone praises the pioneering molecular art, the stepwise problem-solving approach, the abundant mix of qualitative, quantitative, and applied end-of-chapter problems, and the rigorous *and* student-friendly coverage of mainstream topics.

Global Changes to Every Chapter

Our revision for this edition has gone further than ever to optimize the text. We completed the distillation of the text and the annotation of the illustrations so appreciated in the previous edition, and created an inviting, easy-to-follow page design. But, when professors *and* students were asked what they wanted most in a new edition, the answer inevitably cited more good problems, so that became a major focus. We are delighted to introduce the seventh edition of *Chemistry: The Molecular Nature of Matter and Change,* which takes several major steps closer to perfecting its content and approach.

Learning ideas through focused writing and content presentation. Once again, every discussion has been revised to optimize clarity, readability, and the conciseness and directness of the presentation. The use of additional subheads, numbered (and titled) paragraphs, and bulleted (and titled) lists that was introduced in the sixth edition has been completed in the seventh.

Applying ideas with an enhanced problem-solving approach. The much admired—and copied—four-part problem-solving format (plan, solution, check, follow-up) is still used throughout, in both data-based *and* molecular-scene *Sample Problems*. But, many changes make the seventh edition a problem-solving "powerhouse." Each sample problem now includes ***two*** *Follow-up Problems:* as professors requested, the first closely matches the worked-out problem, so the student gains confidence, while the second varies a bit to test comprehension. As always, a thoroughly worked-out *Brief Solution* for each follow-up problem appears at the end of the chapter (rather than providing just a numerical answer in a distant end-of-book appendix, as is typical). As students requested, for more practice, each sample problem now lists *Some Similar Problems* within the end-of-chapter problem set. Moreover, some remaining pedagogic gaps have been filled with 22 *new* sample problems: thus, the seventh edition has over 250 sample problems and over 500 follow-up problems.

Re-learning ideas with annotated illustrations. The innovative three-level figures and other art that raised the bar for molecular visualization in chemistry textbooks is still present. Many existing figures have been revised and several new ones added to create an even better teaching tool. And, continuing the innovation of last edition, wherever appropriate, figure legends have been turned into simple captions and their content into clarifying annotations within the figures themselves.

Easier studying with attractive and functional page design and layout. A more open layout with three-dimensional accents engages students by focusing attention on section heads, sample problems, tables, and other important features, while maintaining clean margins. The orderly presentation fosters easier viewing and studying and allows the intimate placement of figures and tables in relation to the discussion.

Mastering the content with abundant end-of-chapter problem sets. With the more open design for improved readability, traditional and molecular-scene problems updated and revised, and many new problems added, these problem sets are more extensive than in most other texts. They provide students and teachers with abundant choices in a wide range of difficulty and real-life scenarios.

Content Changes to Individual Chapters

The major revision that created this seventh edition of *Chemistry* included key improvements to nearly every chapter:

- **Chapter 1** has been rearranged so that units are now discussed *before* the problem-solving approach and unit conversions are introduced. The chapter includes a *new sample problem* on converting units raised to a power.
- **Chapter 2** now clarifies the meaning of mass fraction and introduces IUPAC's new recommendations on atomic-mass ranges. It presents the periodic table updated with the latest atomic masses and includes a *new sample problem* on identifying an element from its *Z* value.
- **Chapter 3** includes more information in its road maps, and road maps in later chapters also have these enhancements. Sample problems on calculating mass percent of an element and the mass of an element have been carefully revised. The section on calculations for reactions in solution has been moved to Chapter 4.
- **Chapter 4** introduces many types of reactions in solution, so it was thoroughly revised to, more logically, include the

material on molarity and solution stoichiometry previously in Chapter 3. The table on solubility rules was revised, and a new table focuses on the reactions of strong versus weak acids in strong base. *Two new sample problems* show how to determine the amount of ions in solution and the amounts of reactants and products in a precipitation reaction. And two sample problems from the last edition were combined to show the logical relation of their material. Balancing redox reactions by the oxidation-number method was deleted, but balancing them by the half-reaction method is covered in the electrochemistry chapter (Chapter 21).

- **Chapter 5** includes *two new sample problems* that apply various combinations of the individual gas laws, and it provides a derivation of Graham's law.
- **Chapter 6** includes *a new sample problem* on calculating *PV* work.
- **Chapter 7** incorporates carefully revised atomic and continuous spectra throughout.
- **Chapter 8** contains a new table on changes in Z_{eff} within the sublevels of an atom.
- **Chapter 9** includes *a new sample problem* on predicting relative lattice energy and a revised figure on the properties of the covalent bond.
- **Chapter 10** has pedagogic improvements to several figures and expanded coverage in two sample problems.
- **Chapter 12** includes *two new sample problems,* one on the use of phase diagrams to predict phase changes and the other on determining the number of particles in a unit cell and the coordination number.
- **Chapter 13** presents *a new sample problem* on calculating an aqueous ionic heat of solution.
- **Chapter 14** incorporates in its unique Family Portraits all the updated atomic masses and newly synthesized elements that fill out the periodic table.
- **Chapter 16** includes revisions to two important figures and *two new sample problems,* one on rate laws and the molecularity of elementary steps and the other on intermediates and the correlation of rate laws and mechanism.
- **Chapter 17** contains *a new sample problem* on writing a reaction quotient from the balanced equation, and two important figures have been revised.
- **Chapter 18** has undergone major improvements. It has *two new sample problems,* one on calculating hydronium and hydroxide ion concentrations in strong acids and bases and the other on finding the percent dissociation of a weak acid. Several summarizing tables were revised to display reactions, and a key figure was revised. Also, the discussion was rearranged to present the material on acids in a clearer sequence.
- **Chapter 19** incorporates improvements to several key figures, as well as a new figure that summarizes the effects of added acid or base to buffer-component concentration and pH.
- **Chapter 20** has clearer discussions of measuring the change in entropy and of free energy and work, as well as *a new sample problem* on the relationship between ΔG° and K.
- **Chapter 21** includes *a new sample problem* on using $E^\circ_{\text{half-cell}}$ to find E°_{cell}, as well as much clearer discussions of the activity series of the metals, the electrolysis of aqueous salts, and the stoichiometry of electrolysis.
- **Chapter 23** has improvements to many figures as well as *a new sample problem* on finding the coordination number and ion charge in a coordination compound.
- **Chapter 24** provides *a new sample problem* on calculating the specific activity and decay constant of a radionuclide.

Innovative Topic and Chapter Presentation

While the topic sequence coincides with that used in most mainstream courses, built-in flexibility allows a wide range of differing course structures:

For courses that follow their own topic sequence, the general presentation, with its many section and subsection breaks and bulleted lists, allows topics to be rearranged, or even deleted, with minimal loss of continuity.

For courses that present several chapters, or topics within chapters, in different orders:

- Redox balancing by the oxidation-number method (formerly covered in Chapter 4) has been removed from the text, and the half-reaction method is covered with electrochemistry in Chapter 21, but it can easily be taught with Chapter 4.
- Gases (Chapter 5) can be covered in sequence to explore the mathematical modeling of physical behavior or, with no loss of continuity, just before liquids and solids (Chapter 12) to show the effects of intermolecular forces on the three states of matter.

For courses that want an atoms-first approach for some of the material, Chapters 7 through 13 move smoothly from quantum theory (7) through electron configuration (8), bonding models (9), molecular shape (10), VB and MO bonding theories (11), intermolecular forces in liquids and solids (12), and solutions (13). Immediate applications of these concepts appear in the discussions of periodic patterns in main-group chemistry (Chapter 14) and in the survey of organic chemistry (Chapter 15). Some instructors have also brought forward the coverage of transition elements and coordination compounds (23) as further applications of bonding concepts. (Of course, Chapters 14, 15, and 23 can just as easily remain in their more traditional placement later in the course.)

For courses that want biological/medical applications, many chapters highlight these topics, including the role of intermolecular forces in biomolecular structure (12), the chemistry of polysaccharides, proteins, and nucleic acids (including protein synthesis, DNA replication, and DNA sequencing) (15), as well as introductions to enzyme catalysis (16), biochemical pathways (17), and trace elements in protein function (23).

For courses that want engineering applications of physical chemistry topics, Chapters 16 through 21 cover kinetics (16), equilibrium in gases (17), acids and bases (18), and aqueous ionic systems (19) and entropy and free energy (20) as they apply to electrochemical systems (21), all in preparation for coverage of the elements in geochemical cycles, metallurgy, and industry in Chapter 22.

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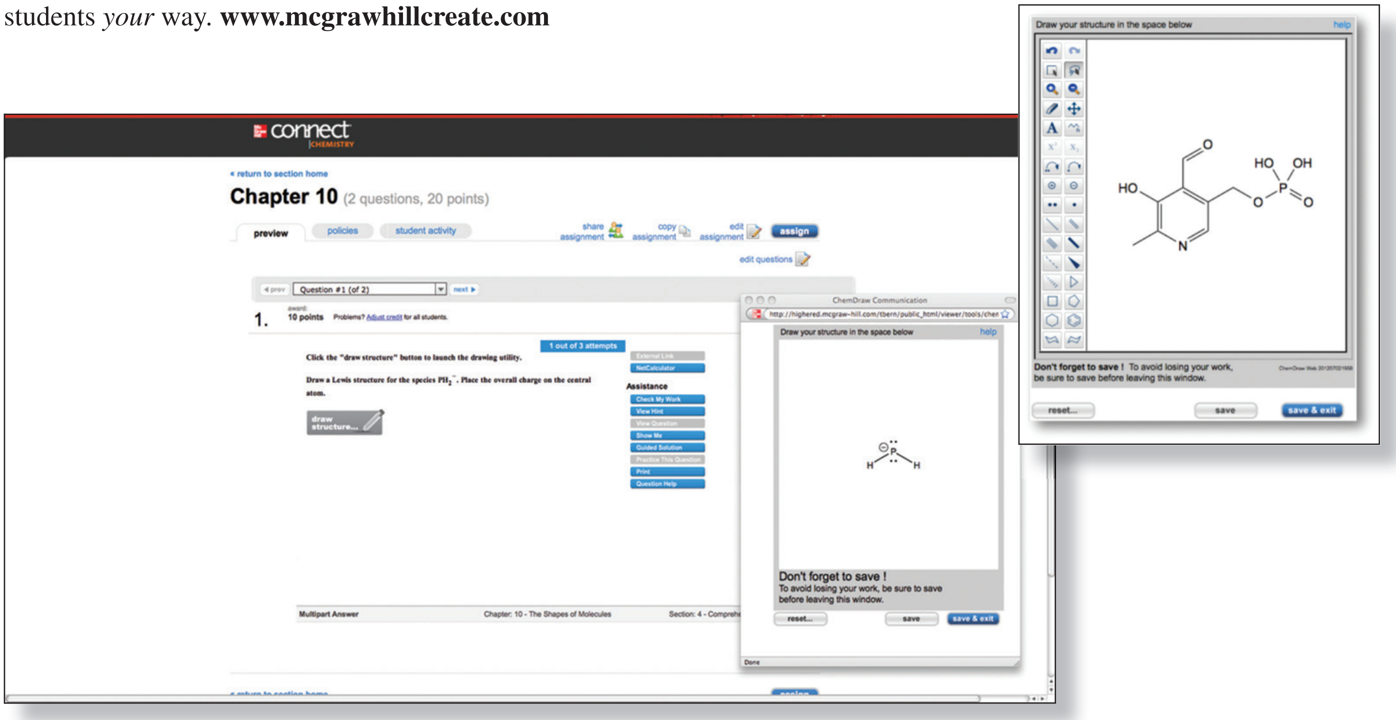

Chapter 10 (2 questions, 20 points)

Question #1 (of 2)

1. award: 10 points Problems? Adjust credit for all students.

Click the "draw structure" button to launch the drawing utility.

Draw a Lewis structure for the species PH_2^-. Place the overall charge on the central atom.

Multipart Answer Chapter: 10 - The Shapes of Molecules Section: 4 - Compreh

ChemDraw Communication

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Draw your structure in the space below

Don't forget to save ! To avoid losing your work, be sure to save before leaving this window.

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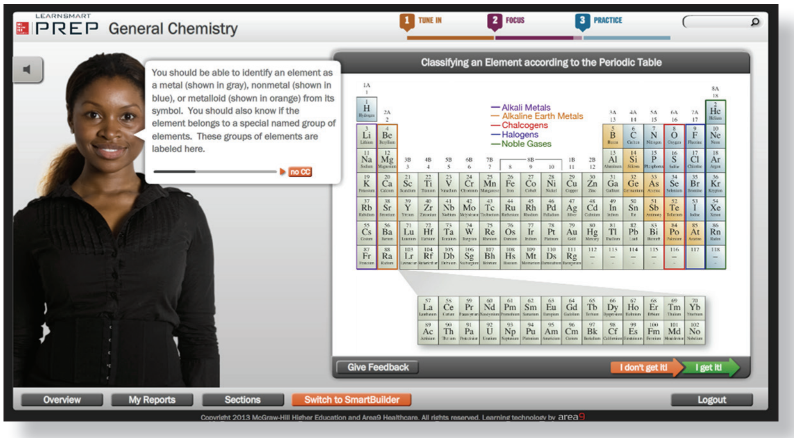

You should be able to identify an element as a metal (shown in gray), nonmetal (shown in blue), or metalloid (shown in orange) from its symbol. You should also know if the element belongs to a special named group of elements. These groups of elements are labeled here.

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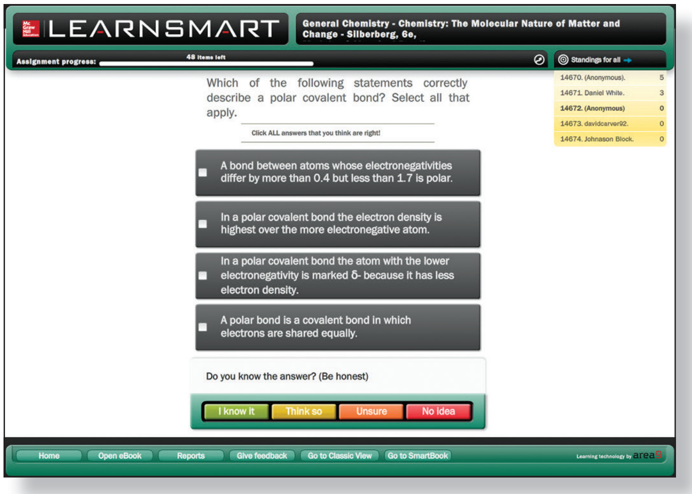


General Chemistry - Chemistry: The Molecular Nature of Matter and Change - Silberberg, 6e,

14670. (Anonymous).	5
14671. Daniel White.	3
14672. (Anonymous)	0
14673. davidcarver92.	0
14674. Johnason Block.	0

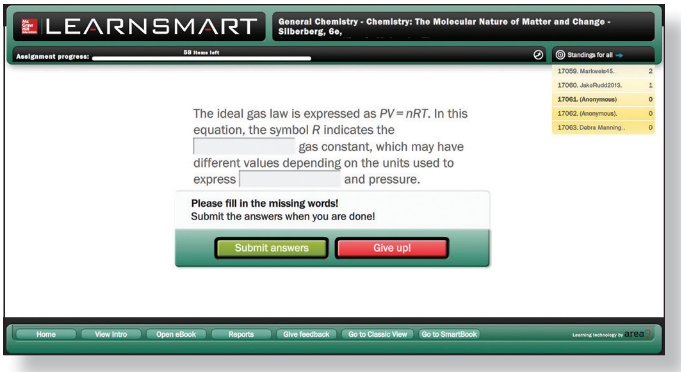
Which of the following statements correctly describe a polar covalent bond? Select all that apply.

Click ALL answers that you think are right!

- ☐ A bond between atoms whose electronegativities differ by more than 0.4 but less than 1.7 is polar.
- ☐ In a polar covalent bond the electron density is highest over the more electronegative atom.
- ☐ In a polar covalent bond the atom with the lower electronegativity is marked δ- because it has less electron density.
- ☐ A polar bond is a covalent bond in which electrons are shared equally.

Do you know the answer? (Be honest)

I know it | Think so | Unsure | No idea



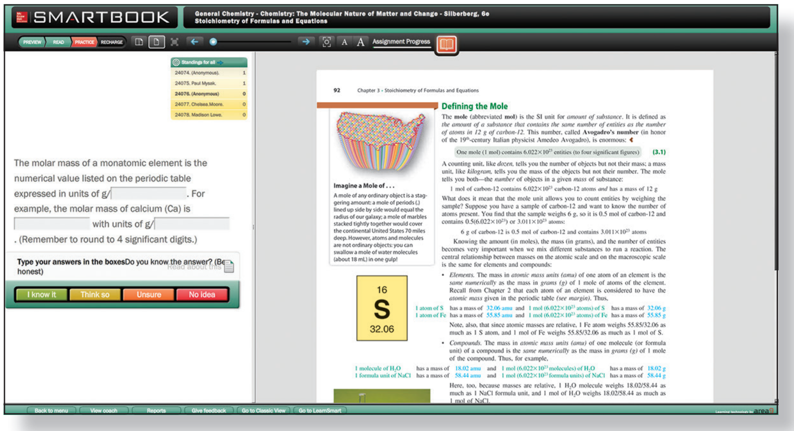
General Chemistry - Chemistry: The Molecular Nature of Matter and Change - Silberberg, 6e,

17059. Markweis45.	2
17060. JakeRudd2013.	1
17061. (Anonymous)	0
17062. (Anonymous).	0
17063. Debra Manning..	0

The ideal gas law is expressed as $PV = nRT$. In this equation, the symbol R indicates the ________ gas constant, which may have different values depending on the units used to express ________ and pressure.

Please fill in the missing words!
Submit the answers when you are done!

SMARTBOOK™



The molar mass of a monatomic element is the numerical value listed on the periodic table expressed in units of g/☐. For example, the molar mass of calcium (Ca) is ☐ with units of g/☐. (Remember to round to 4 significant digits.)

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

92 Chapter 3 • Stoichiometry of Formulas and Equations

Defining the Mole

The **mole** (abbreviated **mol**) is the SI unit for *amount of substance*. It is defined as *the amount of a substance that contains the same number of entities as the number of atoms in 12 g of carbon-12*. This number, called **Avogadro's number** (in honor of the 19th-century Italian physicist Amedeo Avogadro), is enormous:

One mole (1 mol) contains 6.022×10^{23} entities (to four significant figures) (3.1)

A counting unit, like *dozen*, tells you the number of objects but not their mass; a mass unit, like *kilogram*, tells you the mass of the objects but not their number. The mole tells you both—the *number* of objects in a given *mass* of substance:

1 mol of carbon-12 contains 6.022×10^{23} carbon-12 atoms *and* has a mass of 12 g

What does it mean that the mole unit allows you to count entities by weighing the sample? Suppose you have a sample of carbon-12 and want to know the number of atoms present. You find that the sample weighs 6 g, so it is 0.5 mol of carbon-12 and contains $0.5(6.022\times10^{23})$ or 3.011×10^{23} atoms:

6 g of carbon-12 is 0.5 mol of carbon-12 and contains 3.011×10^{23} atoms

Knowing the amount (in moles), the mass (in grams), and the number of entities becomes very important when we mix different substances to run a reaction. The central relationship between masses on the atomic scale and on the macroscopic scale is the same for elements and compounds:

- *Elements*. The mass in *atomic mass units (amu)* of one atom of an element is the *same numerically* as the mass in *grams (g)* of 1 mole of atoms of the element. Recall from Chapter 2 that each atom of an element is considered to have the *atomic mass* given in the periodic table *(see margin)*. Thus,

1 atom of S has a mass of 32.06 amu and 1 mol (6.022×10^{23} atoms) of S has a mass of 32.06 g
1 atom of Fe has a mass of 55.85 amu and 1 mol (6.022×10^{23} atoms) of Fe has a mass of 55.85 g

Note, also, that since atomic masses are relative, 1 Fe atom weighs 55.85/32.06 as much as 1 S atom, and 1 mol of Fe weighs 55.85/32.06 as much as 1 mol of S.

- *Compounds*. The mass in *atomic mass units (amu)* of one molecule (or formula unit) of a compound is the *same numerically* as the mass in *grams (g)* of 1 mole of the compound. Thus, for example,

1 molecule of H_2O has a mass of 18.02 amu and 1 mol (6.022×10^{23} molecules) of H_2O has a mass of 18.02 g
1 formula unit of NaCl has a mass of 58.44 amu and 1 mol (6.022×10^{23} formula units) of NaCl has a mass of 58.44 g

Here, too, because masses are relative, 1 H_2O molecule weighs 18.02/58.44 as much as 1 NaCl formula unit, and 1 mol of H_2O weighs 18.02/58.44 as much as 1 mol of NaCl.

Imagine a Mole of . . .

A mole of any ordinary object is a staggering amount: a mole of periods (.) lined up side by side would equal the radius of our galaxy; a mole of marbles stacked tightly together would cover the continental United States 70 miles deep. However, atoms and molecules are not ordinary objects: you can swallow a mole of water molecules (about 18 mL) in one gulp!

16
S
32.06

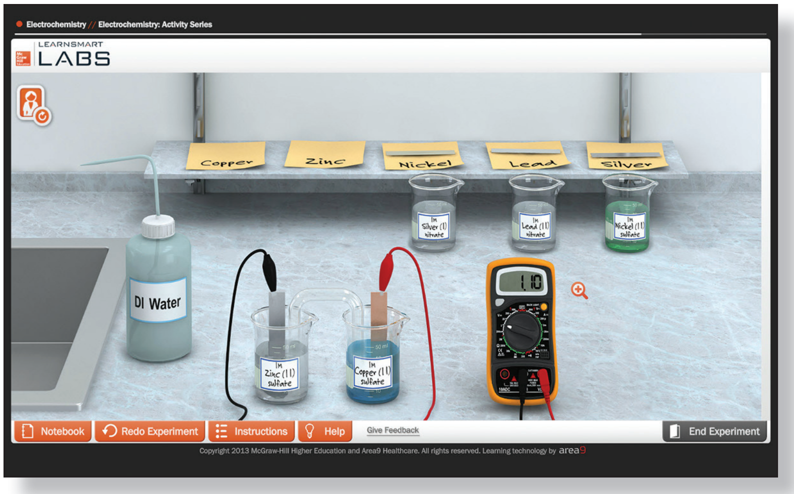
Powered by the intelligent and adaptive LearnSmart engine, **SmartBook** is the first and only continuously adaptive reading experience available. Distinguishing what students know from what they don't, and honing in on concepts they are most likely to forget, SmartBook personalizes content for each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

SmartBook includes powerful reports that identify specific topics and learning objectives students need to study. These valuable reports also provide instructors with insight into how students are progressing through textbook content and are useful for identifying class trends, focusing precious class time, providing personalized feedback to students, and tailoring assessment.

How does SmartBook work? Each SmartBook contains four components: Preview, Read, Practice, and Recharge. Starting with a preview of each chapter and key learning objectives, students read the material and are guided to topics on which they need the most practice based on their responses to a continuously adapting diagnostic. Reading and practice continue until SmartBook directs students to review—or recharge—important material they are most likely to forget to ensure concept mastery and retention.

LEARNSMART LABS™

THE Virtual Lab Experience



Based on the same world-class, superbly adaptive technology as LearnSmart, **McGraw-Hill LearnSmart Labs** is a must-see, outcomes-based lab simulation. It assesses a student's knowledge and adaptively corrects deficiencies, allowing the student to learn faster and retain more knowledge with greater success.

First, a student's knowledge is adaptively leveled on core learning outcomes: Questioning reveals knowledge deficiencies that are corrected by the delivery of content that is conditional on a student's response. Then, a simulated lab experience requires the student to think and act like a scientist: Recording, interpreting, and analyzing data using simulated equipment found in labs and clinics. The student is allowed to make mistakes—a powerful part of the learning experience! A virtual coach provides subtle hints when needed, asks questions about the student's choices, and allows the student to reflect on and correct those mistakes. Whether your need is to overcome the logistical challenges of a traditional lab, provide better lab prep, improve student performance, or make your online experience one that rivals the real world, LearnSmart Labs accomplishes it all. Learn more at **www.mhlearnsmart.com.**

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Presentation Tools

Accessed from your textbook's Connect website, the presentation tools include McGraw-Hill–owned photos, artwork, animations, and other types of media that can be used to create customized lectures, visually enhanced tests and quizzes, compelling course websites, or attractive printed support materials for classroom purposes. The visual resources in this collection include the following:

- **Art** Full-color digital files of all illustrations in the book can be readily incorporated into lecture presentations, exams, or custom-made classroom materials. In addition, all files have been inserted into PowerPoint slides for ease of lecture preparation.
- **Photos** The photo collection contains digital files of photographs from the text, which can be reproduced for multiple classroom uses.
- **Tables** Every table that appears in the text has been saved in electronic form for use in classroom presentations and/or materials.
- **Animations** Numerous full-color animations illustrating important processes are also provided. Harness the visual impact of concepts in motion by importing these files into classroom presentations or online course materials.

PowerPoint materials can also be accessed through your textbook's Connect website:

- **PowerPoint Lecture Outlines** Ready-made presentations that combine art and lecture notes are provided for each chapter of the text.
- **PowerPoint Slides** For instructors who prefer to create their lectures from scratch, all illustrations, photos, sample problems, and tables have been inserted into blank PowerPoint slides, arranged by chapter.

Computerized Test Bank

Prepared by Walter Orchard, Professor Emeritus of Tacoma Community College, over 2300 test questions that accompany *Chemistry: The Molecular Nature of Matter and Change* are available utilizing Brownstone's *Diploma* testing software. *Diploma's* software allows you to quickly create a customized test using McGraw-Hill's supplied questions or by writing your own. *Diploma* allows you to create your tests without an Internet connection—just download the software and question files directly to your computer.

Instructor's Solutions Manual

This supplement, prepared by MaryKay Orgill of the University of Nevada, Las Vegas, contains complete, worked-out solutions for *all* the end-of-chapter problems in the text. It can be found within the Instructors' Resources, on the Connect website.

Cooperative Chemistry Laboratory Manual

Prepared by Melanie Cooper of Clemson University, this innovative manual features open-ended problems designed to simulate experience in a research lab. Working in groups, students investigate one problem over a period of several weeks, so they might complete three or four projects during the semester, rather than one preprogrammed experiment per class. The emphasis is on experimental design, analytic problem solving, and communication.

Student Study Guide

This valuable study guide, prepared by Libby Bent Weberg, is designed to help students recognize learning style; understand how to read, classify, and create a plan for solving a problem; and practice problem-solving skills. For each section of each chapter, the guide provides study objectives and a summary of the corresponding text. Following the summary are sample problems with detailed solutions. Each chapter has true-false questions and a self-test, with all answers provided at the end of the chapter.

Student Solutions Manual

This supplement, prepared by MaryKay Orgill of the University of Nevada, Las Vegas, contains detailed solutions and explanations for all problems in the main text that have colored numbers.

Animations for MP3/iPod

A number of animations are available for downloading to an MP3 player or iPod through the textbook's Connect website.

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Beginning with the 7th edition, I am delighted to welcome Professor Patricia Amateis of Virginia Tech, who shared equally all the responsibilities for creating the text. Patricia's experience teaching large classes in a research university is the perfect complement to my teaching small classes in a liberal-arts college. She is thorough in her understanding of chemistry, meticulous in her execution of the text, and, as her numerous teaching awards attest, insightful in her appreciation of student learning and devoted to their comprehension of chemistry. I am very fortunate to have her as my co-author.

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The seventh edition and its robust learning system are the culmination of numerous chapter reviews, insightful suggestions, brainstorming sessions, symposia, focus groups, interviews, and class tests. Therefore, our deepest gratitude is extended to the following individuals, who participated in preparation of the seventh edition and its associated learning system:

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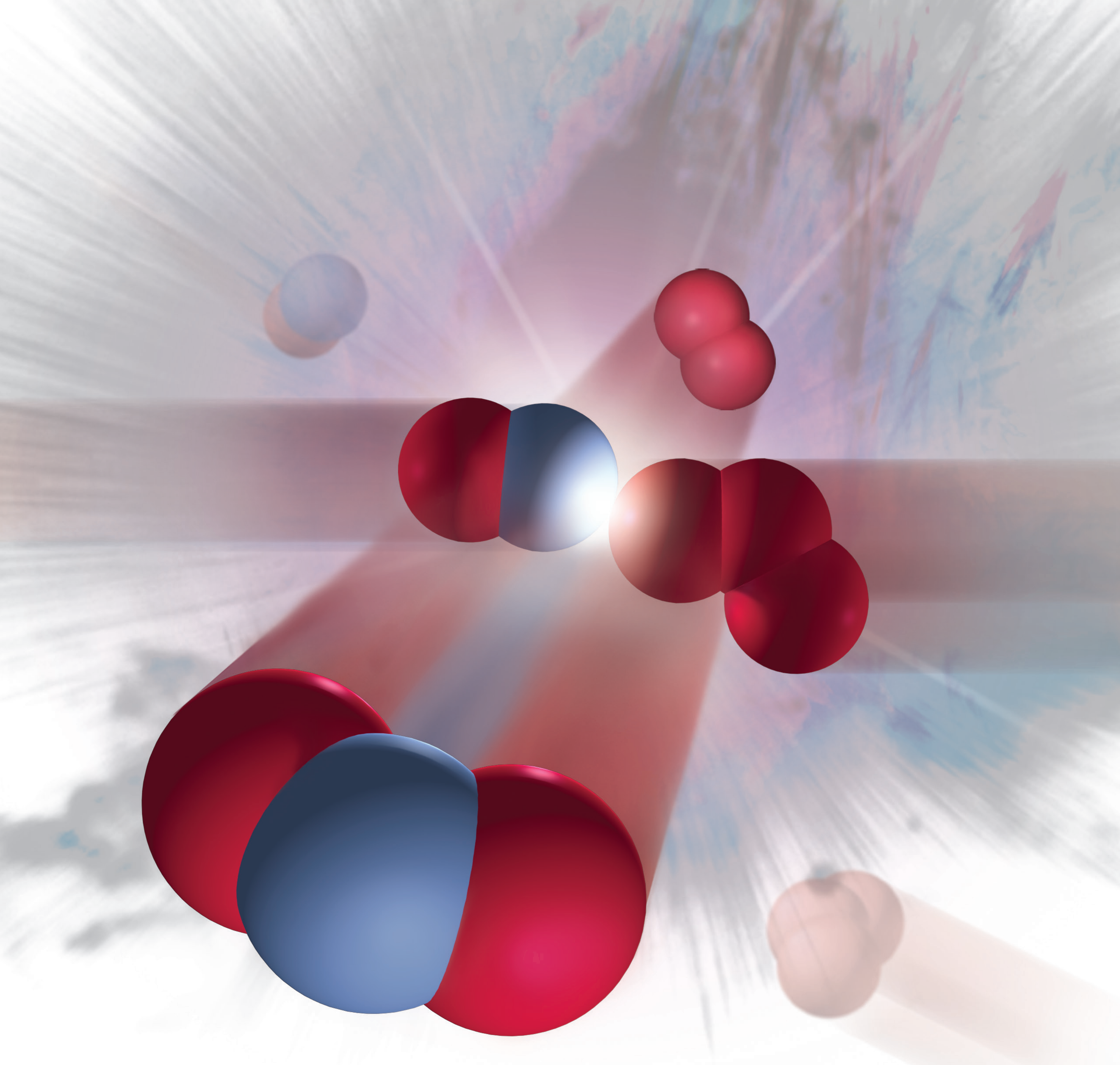
In this edition, perhaps more than in any other after the first, my wife Ruth's dedication and expertise shaped the quality of the text. She checked the style, revised the layout, and monitored changes in the pages, among countless other details—all while she kept smiling and supporting me.



CHEMISTRY

1

Keys to the Study of Chemistry

1.1 Some Fundamental Definitions
States of Matter
Properties of Matter and Its Changes
Central Theme in Chemistry
Importance of Energy

1.2 Chemical Arts and the Origins of Modern Chemistry
Prechemical Traditions
Impact of Lavoisier

1.3 The Scientific Approach: Developing a Model

1.4 Measurement and Chemical Problem Solving
Features of SI Units
SI Units in Chemistry
Units and Conversion Factors
Systematic Problem-Solving Approach
Temperature Scales
Extensive and Intensive Properties

1.5 Uncertainty in Measurement: Significant Figures
Determining Significant Digits
Calculations and Rounding Off
Precision, Accuracy, and Instrument Calibration

Concepts and Skills to Review Before You Study This Chapter

- exponential (scientific) notation (Appendix A)

Maybe you're taking this course because chemistry is fundamental to understanding other natural sciences. Maybe it's required for your medical or engineering major. Or maybe you just want to learn more about the impact of chemistry on society or even on a personal, everyday level. For example, were you awakened today by your alarm clock? In chemical terms, the sound of the alarm was created when molecules aligned in the liquid-crystal display of your clock and electrons flowed to create a noise. You might have thrown off a thermal insulator of manufactured polymer (blanket) and jumped in the shower to emulsify fatty substances on your skin and hair with purified water and formulated detergents. Perhaps you next adorned yourself in an array of pleasant-smelling pigmented gels, dyed polymeric fibers, synthetic footwear, and metal-alloy jewelry. After a breakfast of nutrient-enriched, spoilage-retarded carbohydrates (cereal) in a white emulsion of fats, proteins, and monosaccharides (milk) and a cup of hot aqueous extract containing a stimulating alkaloid (coffee), you probably abraded your teeth with a colloidal dispersion of artificially flavored, dental-hardening agents (toothpaste), grabbed your portable electronic device containing ultrathin, microetched semiconductor layers powered by a series of voltaic cells (laptop), collected some objects made from processed cellulose and plastic, electronically printed with light- and oxygen-resistant inks (books), hopped in your hydrocarbon-fueled, metal-vinyl-ceramic vehicle, electrically ignited a synchronized series of controlled gaseous explosions (started your car), and took off for class!

But the true impact of chemistry extends much farther than the commercial products of daily life. The truth is that the most profound biological and environmental questions ultimately have chemical answers: How does an organism reproduce, grow, and age? What are the underlying explanations for health and disease? How can we sustain a planetary ecosystem in which plant, animal, and human populations thrive? Is there life on other worlds?

So, no matter what your reason for studying chemistry, you're going to learn some amazing ideas. And, this course comes with a bonus for developing two mental skills. The first, common to all science courses, is the ability to solve problems systematically. The second is specific to chemistry, for as you comprehend its ideas, you begin to view a hidden reality, one filled with incredibly minute particles moving at fantastic speeds, colliding billions of times a second, and interacting in ways that allow your brain to translate fluxes of electric charge into thoughts and that determine how all the matter inside and outside of you behaves. This chapter holds the keys to unlock and enter this new world.

IN THIS CHAPTER . . . *We discuss some central ideas about matter and energy, the process of science, units of measurement, and how scientists handle data.*

- We begin with fundamental concepts about matter and energy and their changes.
- A brief discussion of chemistry's origins, including some major missteps, leads to an overview of how scientists build models to study nature.
- We examine modern units for mass, length, volume, density, and temperature and apply systematic chemical problem solving to unit conversions.
- We see that data collection always includes some uncertainty and examine the distinction between accuracy and precision.

1.1 SOME FUNDAMENTAL DEFINITIONS

A good place to begin our exploration of chemistry is to define it and a few central concepts. **Chemistry** is *the study of matter and its properties, the changes that matter undergoes, and the energy associated with those changes.* **Matter** is the "stuff" of the universe: air, glass, planets, students—*anything that has mass and volume.* (In Section 1.4, we discuss the meanings of mass and volume in terms of how they are measured.) Chemists want to know the **composition** of matter, *the types and amounts of simpler substances that make it up.* A *substance* is a type of matter that has a defined, fixed composition.

The States of Matter

Matter occurs commonly in *three physical forms* called **states:** solid, liquid, and gas. On the macroscopic scale, each state of matter is defined by the way the sample fills a container (Figure 1.1, *flasks at top*):

- A **solid** has a fixed shape that does not conform to the container shape. Solids are *not* defined by rigidity or hardness: solid iron is rigid and hard, but solid lead is flexible, and solid wax is soft.
- A **liquid** has a varying shape that conforms to the container shape, but only to the extent of the liquid's volume; that is, a liquid has *an upper surface.*
- A **gas** also has a varying shape that conforms to the container shape, but it fills the entire container and, thus, does *not* have a surface.

On the atomic scale, each state is defined by the relative positions of its particles (Figure 1.1, *circles at bottom*):

- In a *solid,* the particles lie next to each other in a regular, three-dimensional *array.*
- In a *liquid,* the particles also lie close together but move randomly around each other.
- In a *gas,* the particles have large distances between them and move randomly throughout the container.

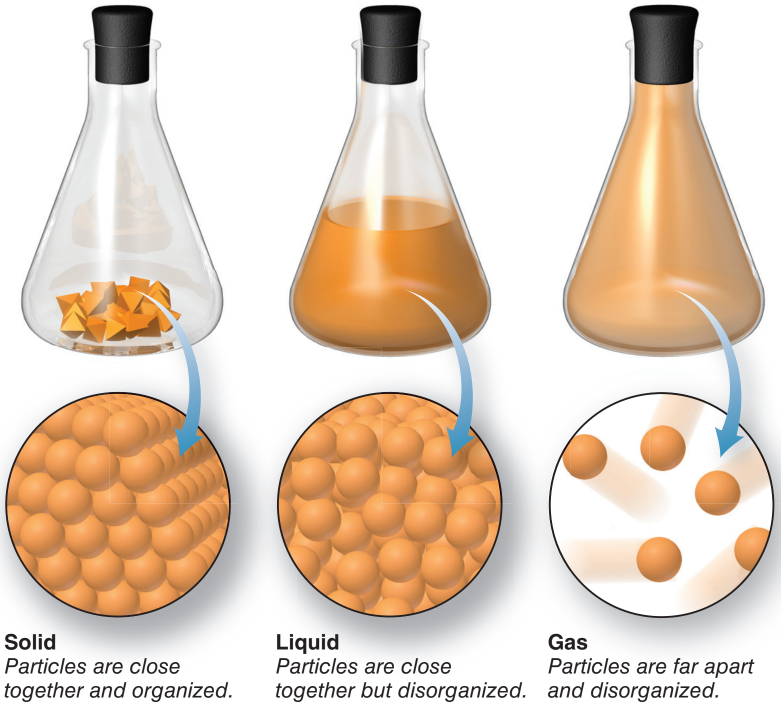


Figure 1.1 The physical states of matter.

The Properties of Matter and Its Changes

We learn about matter by observing its **properties,** *the characteristics that give each substance its unique identity.* To identify a person, we might observe height, weight, hair and eye color, fingerprints, and, now, even DNA pattern, until we arrive at a unique identification. To identify a substance, we observe two types of properties, physical and chemical, which are closely related to two types of change that matter undergoes.

Physical Change: No Change in Composition **Physical properties** are characteristics a substance shows *by itself, without changing into or interacting with another substance*. These properties include melting point, electrical conductivity, and density. A **physical change** occurs when a substance *alters its physical properties,* ***not*** *its composition*. For example, when ice melts, several physical properties change, such as hardness, density, and ability to flow. But the composition of the sample does *not* change: it is still water. The photograph in Figure 1.2A shows what this change looks like in everyday life. The "blow-up" circles depict a magnified view of the particles making up the sample. In the icicle, the particles lie in the repeating pattern characteristic of a solid, whereas they are jumbled in the liquid droplet, but *the particles are the same* in both states of water.

Physical change (same substance before and after):

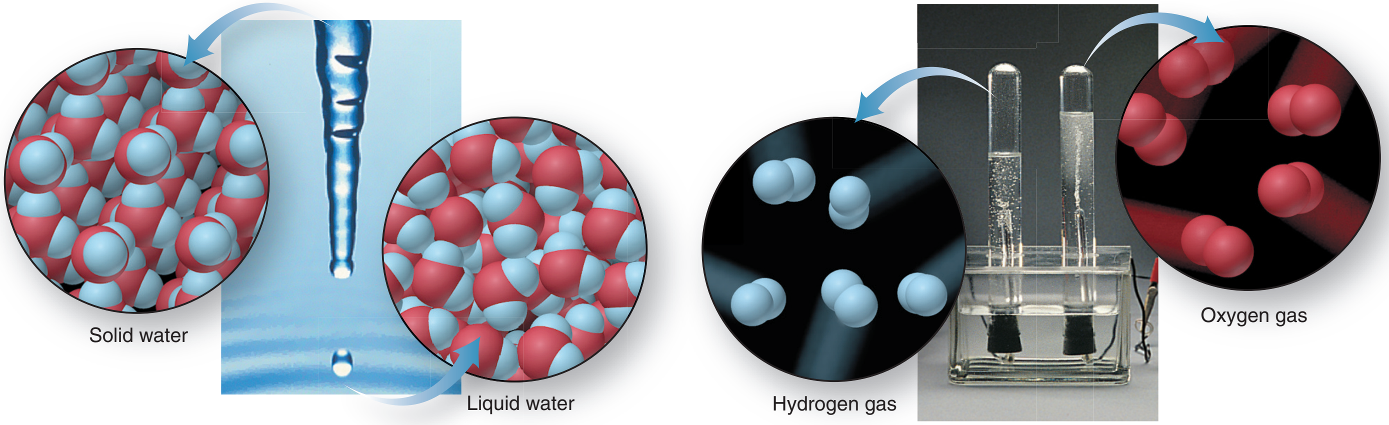
$$\text{Water (solid state)} \longrightarrow \text{water (liquid state)}$$

Chemical Change: A Change in Composition **Chemical properties** are characteristics a substance shows *as it changes into or interacts with another substance (or substances)*. Chemical properties include flammability, corrosiveness, and reactivity with acids. A **chemical change,** also called a **chemical reaction,** occurs when *a substance (or substances) is converted into a different substance (or substances)*. Figure 1.2B shows the chemical change (reaction) that occurs when you pass an electric current through water: the water decomposes (breaks down) into two other substances, hydrogen and oxygen, that bubble into the tubes. The composition *has* changed: the final sample is no longer water.

Chemical change (different substances before and after):

$$\text{Water} \xrightarrow{\text{electric current}} \text{hydrogen} + \text{oxygen}$$

Let's work through a sample problem that uses atomic-scale scenes to distinguish between physical and chemical change.



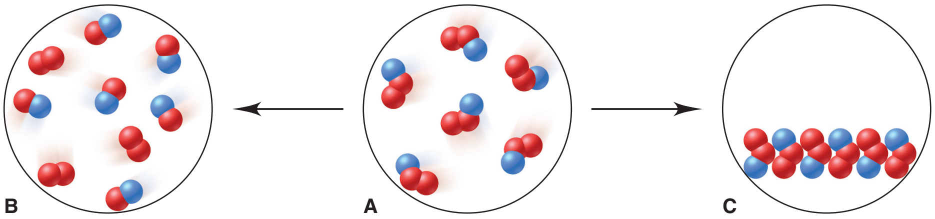
A Physical change:
Solid state of water becomes liquid state. Particles before and after remain the same, which means composition did ***not*** *change.*

B Chemical change:
Electric current decomposes water into different substances (hydrogen and oxygen). Particles before and after are different, which means composition ***did*** *change.*

Figure 1.2 The distinction between physical and chemical change.

SAMPLE PROBLEM 1.1 Visualizing Change on the Atomic Scale

Problem The scenes below represent an atomic-scale view of a sample of matter, A, undergoing two different changes, left to B and right to C:



Decide whether each depiction shows a physical or a chemical change.

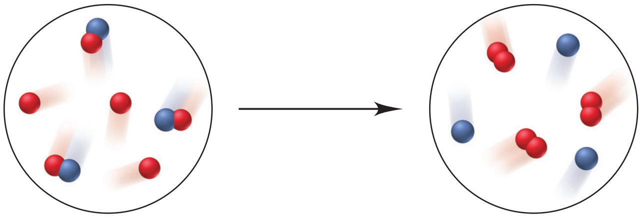
Plan Given depictions of the changes, we have to determine whether each represents a physical or a chemical change. The number and colors of the little spheres that make up each particle tell its "composition." Samples with particles of the *same* composition but in a different arrangement depict a *physical* change, whereas samples with particles of a *different* composition depict a *chemical* change.

Solution In A, each particle consists of one blue and two red spheres. The particles in A change into two types in B, one made of red and blue spheres and the other made of two red spheres; therefore, they have undergone a chemical change to form different particles. The particles in C are the same as those in A, but they are closer together and arranged in a different state of matter; therefore, they have undergone a physical change.

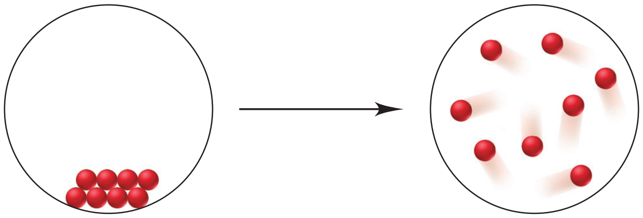
FOLLOW-UP PROBLEMS

*Brief Solutions for **all** Follow-up Problems appear at the end of the chapter.*

1.1A Is the following change chemical or physical?



1.1B Is the following change chemical or physical?

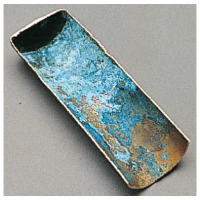


A SIMILAR PROBLEM 1.1

Like water, hydrogen, oxygen, or any other real substance, copper is also identified by *its own set* of physical and chemical properties (Table 1.1).

Temperature and Changes in Matter Depending on the temperature and pressure of the surroundings, many substances can exist in each of the three physical states and undergo changes in state as well. For example, as the temperature increases, solid water melts to liquid water, which boils to gaseous water (also called *water vapor*). Similarly, as the temperature drops, water vapor condenses to liquid water, and with further cooling, the liquid freezes to ice. In a steel plant, solid iron melts to liquid

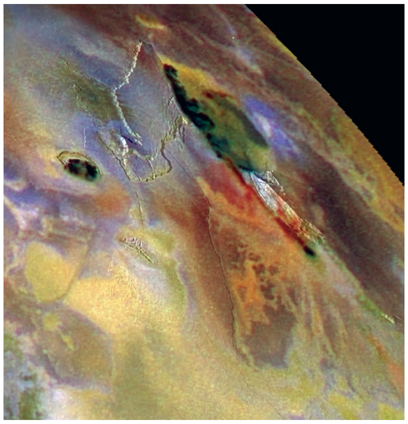
Table 1.1 Some Characteristic Properties of Copper

Physical Properties		Chemical Properties	
Easily shaped into sheets (malleable) and wires (ductile)		Slowly forms a blue-green carbonate in moist air	
Can be melted and mixed with zinc to form brass		Reacts with nitric or sulfuric acid	
Density = 8.95 g/cm^3 Melting point = 1083°C Boiling point = 2570°C		Slowly forms deep-blue solution in aqueous ammonia	

(molten) iron and then cools to the solid again. And, far beyond the confines of a laboratory or steel plant, lakes of molten sulfur lie on Jupiter's moon Io *(see photo),* which is capped by poles of frozen hydrogen sulfide, a gas on Earth.

The main point is that *a physical change caused by heating can generally be reversed by cooling*. This is *not* generally true for a chemical change. For example, heating iron in moist air causes a chemical reaction that yields the brown, crumbly substance known as rust. Cooling does not reverse this change; rather, another chemical change (or series of them) is required.

The following sample problem provides practice in distinguishing some familiar examples of physical and chemical change.



Many common substances occur in unusual states on other worlds.

SAMPLE PROBLEM 1.2 Distinguishing Between Physical and Chemical Change

Problem Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:
(a) Frost forms as the temperature drops on a humid winter night.
(b) A cornstalk grows from a seed that is watered and fertilized.
(c) A match ignites to form ash and a mixture of gases.
(d) Perspiration evaporates when you relax after jogging.
(e) A silver fork tarnishes slowly in air.

Plan The basic question we ask to decide whether a change is chemical or physical is, "Does the substance change composition or just change form?"

Solution **(a)** Frost forming is a physical change: the drop in temperature changes water vapor (gaseous water) in humid air to ice crystals (solid water).
(b) A seed growing involves chemical change: the seed uses water, substances from air, fertilizer, and soil, and energy from sunlight to make complex changes in composition.
(c) The match burning is a chemical change: the combustible substances in the match head are converted into other substances.
(d) Perspiration evaporating is a physical change: the water in sweat changes its state, from liquid to gas, but not its composition.
(e) Tarnishing is a chemical change: silver changes to silver sulfide by reacting with sulfur-containing substances in the air.

FOLLOW-UP PROBLEMS

1.2A Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

(a) Purple iodine vapor appears when solid iodine is warmed.

(b) Gasoline fumes are ignited by a spark in an automobile engine's cylinder.

(c) A scab forms over an open cut.

1.2B Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

(a) Clouds form in the sky.

(b) Old milk turns sour.

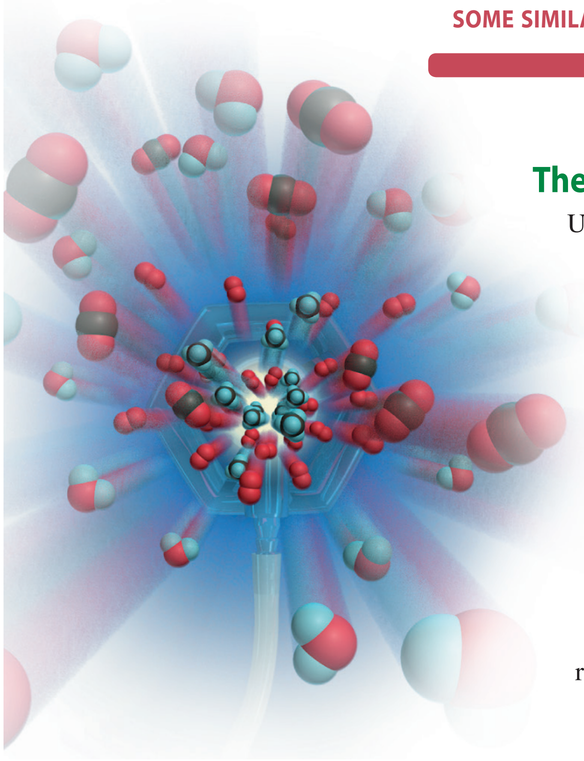
(c) Butter is melted to use on popcorn.

SOME SIMILAR PROBLEMS 1.6 and 1.7

The Central Theme in Chemistry

Understanding the properties of a substance and the changes it undergoes leads to the central theme in chemistry: *macroscopic-scale* properties and behavior, those we can see, are the results of *atomic-scale* properties and behavior that we cannot see. The distinction between chemical and physical change is defined by composition, which we study macroscopically. But composition ultimately depends on the makeup of substances at the atomic scale. Similarly, macroscopic properties of substances in any of the three states arise from atomic-scale behavior of their particles. Picturing a chemical event on the molecular scale, even one as common as the flame of a laboratory burner *(see margin),* helps clarify what is taking place. What is happening when water boils or copper melts? What events occur in the invisible world of minute particles that cause a seed to grow, a neon light to glow, or a nail to rust? Throughout the text, we return to this central idea:

> *We study **observable** changes in matter to understand their **unobservable** causes.*



Methane and oxygen form carbon dioxide and water in the flame of a lab burner. (Carbon is black, oxygen red, and hydrogen blue.)

The Importance of Energy in the Study of Matter

Physical and chemical changes are accompanied by energy changes. **Energy** is often defined as *the ability to do work.* Essentially, all work involves moving something. Work is done when your arm lifts a book, when a car's engine moves the wheels, or when a falling rock moves the ground as it lands. The object doing the work (arm, engine, rock) transfers some of the energy it possesses to the object on which the work is done (book, wheels, ground).

The total energy an object possesses is the sum of its potential energy and its kinetic energy.

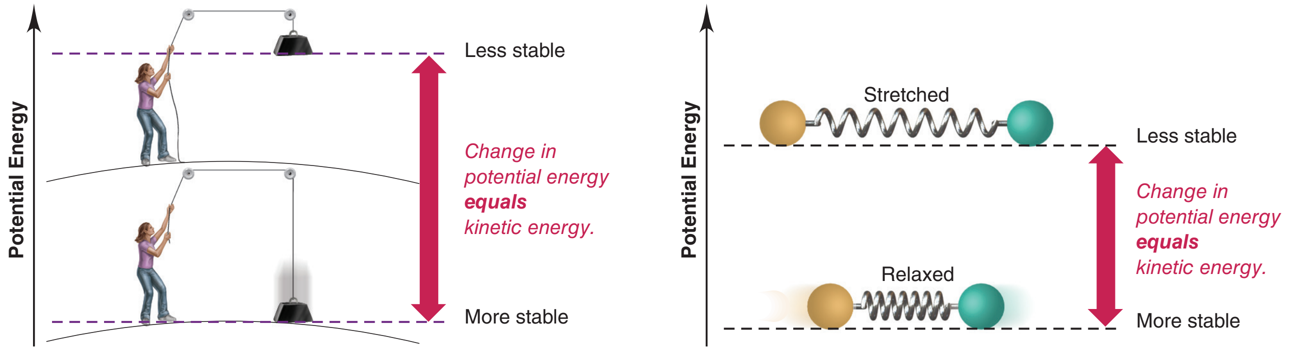
- **Potential energy** is *the energy due to the* ***position*** *of the object relative to other objects.*
- **Kinetic energy** is *the energy due to the* ***motion*** *of the object.*

Let's examine four systems that illustrate the relationship between these two forms of energy: a weight raised above the ground, two balls attached by a spring, two electrically charged particles, and a fuel and its waste products. Two concepts central to all these cases are

1. *When energy is converted from one form to the other, it is conserved, not destroyed.*
2. *Situations of lower energy (more stable) are favored over situations of higher energy (less stable).*

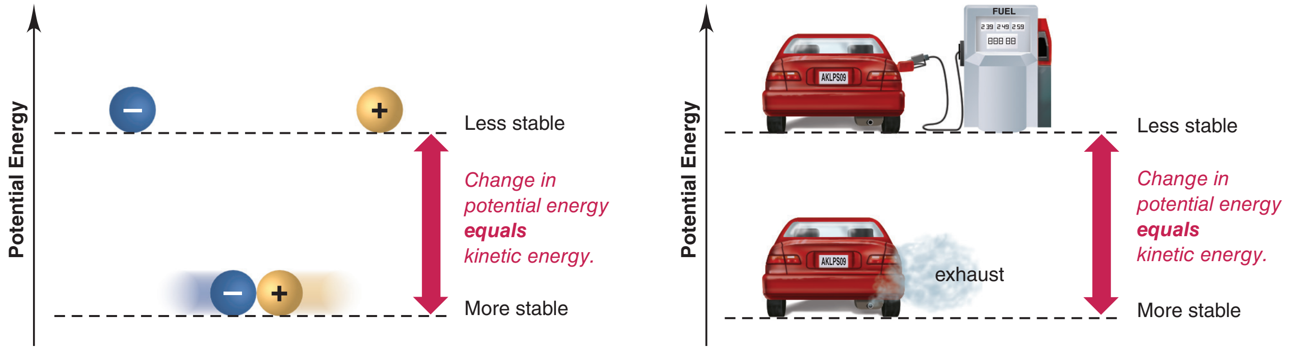
The four cases are

- *A weight raised above the ground* (Figure 1.3A). The energy you exert to lift a weight against gravity increases the weight's potential energy (energy due to its position). When you drop the weight, that additional potential energy is converted to kinetic energy (energy due to motion). The situation with the weight elevated and higher in potential energy is *less stable,* so the weight will fall when released, resulting in a situation that is lower in potential energy and *more stable.*
- *Two balls attached by a spring* (Figure 1.3B). When you pull the balls apart, the energy you exert to stretch the relaxed spring increases the system's potential energy. This change in potential energy is converted to kinetic energy when you release the balls. The system of balls and spring is less stable (has more potential energy) when the spring is stretched than when it is relaxed.
- *Two electrically charged particles* (Figure 1.3C). Due to interactions known as *electrostatic forces, opposite charges attract each other, and like charges repel each other.* When energy is exerted to move a positive particle away from a negative one, the potential energy of the system increases, and that increase is converted to kinetic energy when the particles are pulled together by the electrostatic attraction. Similarly, when energy is used to move two positive (or two negative) particles together, their potential energy increases and changes to kinetic energy when they are pushed apart by the electrostatic repulsion. Charged particles move naturally to a more stable situation (lower energy).
- *A fuel and its waste products* (Figure 1.3D). Matter is composed of positively and negatively charged particles. *The chemical potential energy of a substance results*



A A gravitational system. *Potential energy is gained when a weight is lifted. It is converted to kinetic energy as the weight falls.*

B A system of two balls attached by a spring. *Potential energy is gained when the spring is stretched. It is converted to the kinetic energy of the moving balls as the spring relaxes.*



C A system of oppositely charged particles. *Potential energy is gained when the charges are separated. It is converted to kinetic energy as the attraction pulls the charges together.*

D A system of fuel and exhaust. *A fuel is higher in chemical potential energy than the exhaust. As the fuel burns, some of its potential energy is converted to the kinetic energy of the moving car.*

Figure 1.3 Potential energy is converted to kinetic energy. The dashed horizontal lines indicate the potential energy of each system before and after the change.