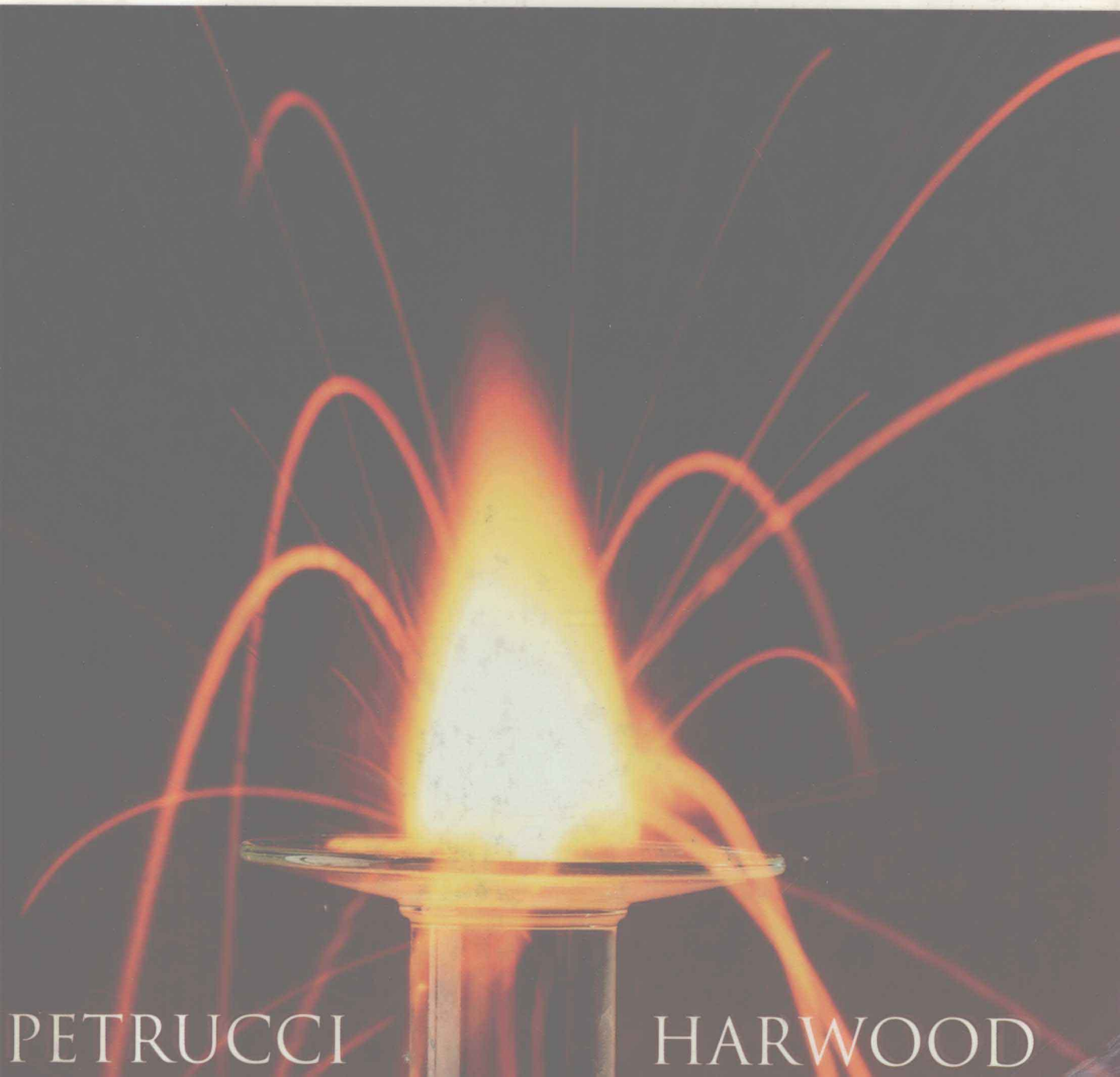# GENERAL CHEMISTRY

*Principles and Modern Applications*

## Seventh Edition

PETRUCCI                    HARWOOD
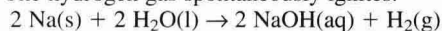
# General Chemistry

## Principles and Modern Applications

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Photographs, NYC. Sodium, an alkali metal, reacts violently with
water to form hydrogen gas and an aqueous solution of sodium
hydroxide. The hydrogen gas spontaneously ignites.

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

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# General Chemistry

## Principles and Modern Applications

### SEVENTH EDITION

**R**alph H. Petrucci
*California State University, San Bernardino*

**W**illiam S. Harwood
*University of Maryland, College Park*

# About
# the Authors

## Ralph H. Petrucci

Ralph Petrucci received his B.S. in Chemistry from Union College and his Ph.D. from the University of Wisconsin–Madison. Following several years of teaching, research, consulting, and directing NSF Institutes for Secondary School Science Teachers at Case Western Reserve University, Professor Petrucci joined the planning staff of the new California State University campus at San Bernardino in 1964. There, in addition to his faculty appointment, he has served as Chairman of the Natural Sciences Division and Dean of Academic Planning. Professor Petrucci is the author of several books, including *General Chemistry* with John W. Hill.

## William S. Harwood

Bill Harwood received his Ph.D. in Inorganic Chemistry from Purdue University in 1986. He is currently Assistant Dean for Undergraduate Studies at the University of Maryland at College Park. Previously, Dr. Harwood directed the undergraduate program in the Department of Chemistry and Biochemistry. In his current role, Dr. Harwood continues to teach General Chemistry and conduct research in chemical education. He has received several awards for teaching excellence and is locally popular for his chemistry demonstrations. Dr. Harwood is also active in the American Chemical Society and the Division of Chemical Education and is a consultant to AAAS Project 2061. In Maryland, he is involved in the science reform efforts at both the precollege and college levels. His research focuses on how best to use technology to improve learning in chemistry.

# Preface

We are aware that most general chemistry students have career interests not in chemistry but in biology, medicine, engineering, and environmental and agricultural sciences as well as in many other fields. We also know that general chemistry will be the only college chemistry course for many students and their only opportunity to learn some practical applications of chemistry. We have designed this text for these "typical" students.

Students of this text probably have already studied some chemistry, but those with no prior background will find that the early chapters develop fundamental concepts from the most elementary ideas. Students who do plan to become professional chemists will also find opportunities in the text to pursue their own special interests.

The typical student needs help in identifying and applying principles and in visualizing their physical significance. The pedagogical features of this text are designed to provide this help. At the same time, we hope the text serves to sharpen student skills in problem solving and critical thinking. Thoughout, we have tried to strike the proper balances between principles and applications, qualitative and quantitative discussions, and rigor and simplification.

## Organizational Changes

In the previous edition, we made some major organizational changes. We retain the overall organization in this new edition but have made a number of smaller organizational changes. Several subsections have been moved from their sixth-edition locations to new chapters in this edition. We think these changes will improve the flow of information to the student.

For example, we have moved the introduction to the periodic table from Chapters 3 and 10 to Chapter 2. The periodic table is an extraordinarily useful tool, and presenting it early allows us to use the periodic table in new ways throughout the early chapters of the text. We have also moved the discussion of electron configurations and the periodic table from Chapter 10 to Chapter 9. The section on energy changes in the formation of ionic crystals has been moved from Chapter 11 to Chapter 13, where we discuss crystal structure more fully. The discussion of formal charge in Chapter 11 is more focused than in the previous edition. The discussion of bonding in hypercoordinated species (Chaper 12) carefully reflects the current consensus among theorists, though in a limited way that is appropriate for a first-year chemistry course. The brief discussion of the acidity of aquated metal ions

has been moved from Chapter 17 to Chapter 25, where it becomes part of a larger discussion of the acid–base reactions of complex ions. As with previous editions, we have emphasized real-world chemistry in the final descriptive chapters, and we have tried to make this material easy to bring forward into earlier parts of the text. Moreover, many topics in these chapters can be covered selectively, without requiring the study of entire chapters.

Throughout the text we also provide real-world examples to enhance the discussion. For instance, we include examples relevant to the biological sciences, engineering, and the environmental sciences.

## Special Features

The text contains a number of pedagogical features:

*Important Expressions.* The most significant equations, concepts, and rules are highlighted with a yellow panel so that students can readily find them.

*Summary/Key Terms/Glossary.* Each chapter concludes with a comprehensive verbal *Summary* of important concepts and factual information. The Summary is followed by a list of *Key Terms*—terms that appear in boldface type in the text and are defined again in the *Glossary* (Appendix E). Students can use Key Terms lists and the Glossary to help them master the terminology of general chemistry.

*Are You Wondering* . . . . To help clarify matters that often puzzle students, questions are posed and answered under this heading. These questions are cast in the form in which students typically ask them. Some are designed to help students avoid common pitfalls; others provide analogies or alternative explanations of a concept. In response to reviewer and student comments and suggestions, several of these have been changed in this edition, and some new ones have been added. A few such questions have been incorporated directly into the text.

*In-Text Illustrative Examples.* In each chapter, most concepts—especially those that students will be expected to apply in homework assignments and examinations—are illustrated with worked-out examples. In many cases a line drawing or photograph accompanies an example to help students visualize what is going on in the problem.

*Practice Examples.* These were first introduced in the previous edition as part of the in-text illustrative examples. They were designed to give students immediate practice in applying the principle(s) illustrated in the example. In this new edition, we have expanded the *Practice Examples:* We now offer two for every illustrative example. The first Practice Example provides the immediate practice in a problem very similar to the illustrative example. The second Practice Example generally takes the student one step further than the illustrative example. This combination helps students to integrate and extend their knowledge and problem-solving skills. Answers to all Practice Examples are given in Appendix F. Complete solutions are given in the Student Study Guide.

*Summarizing Examples.* Each chapter concludes with a multi-part example, generally of a practical nature. *Summarizing Examples* link various important problem types introduced in the chapter—with each other and, in many cases, with problem types from earlier chapters. As in the previous edition, the solution is only outlined, and intermediate results and a final answer are given. Students are expected to work out the details of the solution. Complete solutions are given in the Student Study Guide.

***End-of-Chapter Exercises.*** Each chapter ends with four categories of exercises. *Review Questions* require straightforward application of principles introduced in the chapter, each generally involving a single concept. *Exercises* are grouped by subject matter and are of a broader nature than the Review Questions. In this edition, the Exercises have also been paired, so that there are now two problems covering the same idea. The *Advanced Exercises* are not grouped by type. As expected, some of these are more difficult than those in the previous sections, but some are advanced only in the sense that they introduce new ideas or pursue certain ideas further than is done in the text. The fourth category, *Feature Problems,* is discussed in the next paragraph. Answers to all Review Questions, odd-numbered Exercises, selected Advanced Exercises, and Feature Problems are given in Appendix F.

***Feature Problems.*** New to this edition are end-of-chapter *Feature Problems,* which are of special interest. Some of these problems retrace aspects of the history of chemistry; a few deal with classic experiments; others require students to interpret data or graphs; some present new material; some suggest alternative techniques for problem solving; and a few summarize main points of the chapter in a comprehensive manner. The Feature Problems are a resource that can be used in several ways: for example, as discussion points in class, as assigned homework for individuals, or for collaborative group work. The Feature Problems are cited in margin notes in the text itself.

***Focus On.*** We believe that relevant applications should be an integral part of the text, that asides should be limited to margin notes and occasional Are You Wondering . . . features, and that interesting but less vital issues should follow the main text of a chapter. With this view in mind, we have concluded the text of each chapter with a short essay on a practical topic appropriate to the chapter content. These essays, which may be considered optional reading, focus on ideas introduced in the chapter.

# Supplements

## For the Instructor

***MATTER: The Visual Presentation Resource CD-ROM*** **[IBM (0-13-596909-3); Macintosh (0-13-700006-5)]** includes animated concept sequences, video laboratories, current and informative ABC News Video segments, still illustrations, and version 3.0 of chemistry Presentation Manager.

The ***Instructor's Resource Manual*** **(0-13-271370-5) by Robert K. Wismer, Millersville University,** integrates ancillary material, offers Notes for the Instructor and key concepts, lists Chapter Objectives, and provides worked-out solutions to all Advanced Exercises.

The ***Laboratory Demonstration Videos*** **(0-13-129158-0) by David Humphreys, McMaster University,** is a two-cassette series with approximately 100 laboratory demonstrations of experiments that may be either too expensive or dangerous for the classroom.

The ***Chemical Concepts & Techniques Video*** **(0-13-338666-X)** is a 100-minute videotape divided into 10 segments; each highlights important concepts through a variety of visual media, including electron microscopy and computer animation.

The ***ABC NEWS/Prentice Hall Video Library*** **(0-13-578428-X)** contains 5-20 minute segments from award-winning shows such as *20/20, World News Tonight,*

and *The American Agenda.* This innovative resource shows chemical principles at work and teaches students to analyze media messages critically based on their scientific knowledge.

The *Chemical Animation Video Series* **(0-13-719022-0)** contains over 50 animated concept sequences authored by Theodore L. Brown. This series shows chemistry in action. They provide students with a view of matter and its transformation beyond what can be illustrated on the printed page.

The *Transparencies* **(0-13-578394-1)** include approximately 250 full-color images from this text.

The *Prentice Hall Custom Test* **[Macintosh (0-13-271412-4); Windows (0-13-271420-5); Test Item File (Hard Copy Manual) (0-13-271404-3)]** consists of more than 1900 multiple-choice questions. PH Custom Test allows you to create and tailor exams to your own needs.

The *Solutions Manual* **(0-13-271347-0)** and *Selected Solutions* **(0-13-271354-3) by Robert K. Wismer, Millersville University,** contain the solutions to the end-of-chapter exercises. With instructor permission, these manuals may be made available to students.

### For the Laboratory

*Experiments in General Chemistry* **(0-13-271362-4)** is a comprehensive laboratory manual containing 37 experiments that parallel the text, including a final group of six experiments on qualitative cation analysis. There is an accompanying instructor's manual **(0-13-271338-8).**

### For the Student

The *Prentice Hall Interactive General Chemistry CD-ROM* **(0-13-598012-7) by Steven D. Gammon, et al.,** is an interactive tutorial that focuses on a number of core chemical concepts covered in General Chemistry. The student can interact with and visualize chemical concepts in ways that are not possible through traditional learning programs.

The award-winning *LOGAL Chemistry Explorer Runtime Simulation Software Version 2.0* **[Windows (0-13-51986-0); Mac (0-13-350646-0)]** allows students to explore the molecular world in over 75 interactive experiments by altering parameters and exploring "what if" scenarios.

The *Student Study Guide* **(0-13-281990-2)** is organized around a set of learning objectives for each chapter and features brief discussions of these objectives, drill problems, self quizzes, and sample tests.

*The New York Times/Prentice Hall Themes of the Times Supplement,* in newspaper format, brings together a collection of recent chemistry articles from the pages of *The New York Times.* This free supplement, available in quantity, encourages students to make connections between the classroom and the world around them.

## Acknowledgments

Many people have given of their time, creativity, and support during the preparation of this textbook. Numerous colleagues from California State University, San Bernardino, have offered helpful suggestions through the several editions of this text. From the University of Maryland, D. Webster Keogh and Allessandra Quadrelli acted as sounding boards and gave excellent suggestions for a

number of the Feature Problems. Robert K. Wismer of Millersville University has again provided many of the supplementary materials and has done a careful and complete examination of the manuscript. We cannot thank him enough for his tireless efforts.

We are grateful to many people at Prentice Hall for their help and encouragement, starting with our editor, Ben Roberts. Al MacDonell and Karen Karlin, our development editors, have done a fabulous job of keeping us on track, handling the myriad details, maintaining a sense of humor, and asking dozens of questions after the manner of the best students. We would also like to thank Kirsten Kauffman, our production editor at York Production Services, for her careful and caring work in leading this book through production. It has been a real joy to work with all of them. Cindy-Lee Overton and her crew of photo researchers have sought out and found many exciting images in a brief amount of time. Her grace under pressure and good humor are much appreciated. We applaud Richard Megna, who shot several special photographs, for his willingness to try almost anything and for his artfulness in capturing the drama of chemistry.

We are especially grateful for the patience and encouragement of our wives, Ruth Petrucci and Diana Harwood. They have been willing to give up precious family time so that we could produce this textbook. Without their love and support, this book would not have been possible.

We also appreciate the significant contributions of those who commented on the sixth edition or reviewed manuscript chapters of this edition. Their comments and suggestions have been very helpful.

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W.S.H.

**WARNING:** Many of the compounds described or pictured in this text are hazardous, as are many of the chemical reactions. The reader should not attempt any experiment pictured or implied in the text. Experiments should be performed only in authorized laboratory settings and under adequate supervision.

# A Student's Guide to Using This Text

*T*o some students, chemistry is an almost magical art: There are both mystery and excitement to chemistry. No one knows all the answers, and many aspects of our science are not well understood. Chemists are continuously seeking new materials, new methods, and, most important, new ideas that enhance our understanding of basic scientific principles and may be applied to improve our lives.

In studying chemistry, you will discover some of the ideas and concepts chemists use to understand and direct chemical changes. We urge you not to settle for memorizing facts and mathematical equations. Demand of yourself, this text, and your instructors explanations of the *why* and *how* of chemistry. Seek the concepts behind the facts and equations. Skill in algebra and the assimilation of facts are important, but imagination is the key to mastering chemistry. At the heart of most chemistry problems is perceiving a connection between an observation in the "real," macroscopic world and an imagined change in the microscopic world—the world of atoms, ions, and molecules. Once you perceive this connection, finding the solution to a problem should become much simpler.

This book contains a number of special features designed to help you gain an understanding of the concepts and methods used by chemists. We illustrate several of them on these next few pages.

## CHAPTER OPENERS

Each chapter opens with a striking full-page photograph that illustrates a topic discussed in the chapter. On the facing page in the margin, a descriptive caption ties the photo in to the chapter concepts. In each chapter, an outline of chapter sections leads you into introductory text.
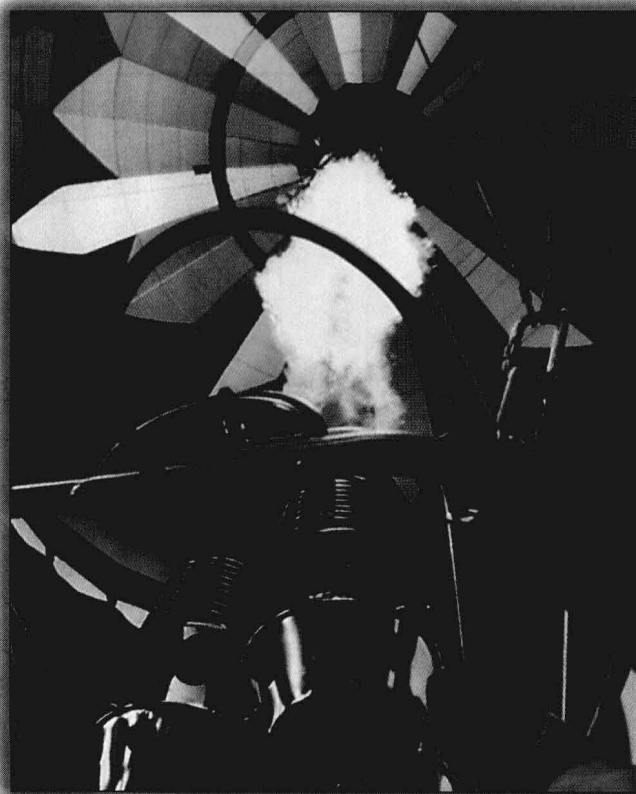
# Gases

## 6

6-1 Properties of Gases: Gas Pressure
6-2 The Simple Gas Laws
6-3 The Ideal Gas Equation
6-4 Applications of the Ideal Gas Equation
6-5 Gases in Chemical Reactions
6-6 Mixtures of Gases
6-7 Kinetic-Molecular Theory of Gases
6-8 Gas Properties Relating to the Kinetic-Molecular Theory
6-9 Nonideal (Real) Gases
*Focus On The Chemistry of Air-Bag Systems*

*A* bicycle tire must not be overinflated because it might rupture. An aerosol can should not be discarded in an incinerator. Carbon dioxide gas vaporizing from a block of dry ice sinks to the floor. A balloon filled with helium rises in air, and so does a balloon filled with hot air. One should never search for a gas leak with an open flame. These and many other observations concerning gases can be explained by concepts considered in this chapter.

For example, the behavior of the bicycle tire and the aerosol can are based on relationships among the variables pressure, temperature, volume, and amount of gas. An understanding of the lifting power of lighter-than-air balloons comes in large part from a knowledge of gas densities and their dependence on molar mass, temperature, and pressure. Predicting how far and how fast gas molecules migrate through air requires knowing something about the phenomenon of diffusion.

To do quantitative calculations about the behavior of gases, we will use some simple gas laws and a more general expression called the ideal gas equation. To explain these laws, we will use a theory known as the kinetic-molecular theory of gases. Ideas presented in this chapter reappear in new contexts in several later chapters.

Hot-air balloons have intrigued people from the time the simple gas laws fundamental to their operation came to be understood over 200 years ago.

### Standard Conditions of Temperature and Pressure

Because gas properties depend on temperature and pressure, at times we find it useful to choose standard conditions of temperature and pressure. This is especially true when we compare gases to one another. The standard temperature for gases is taken to be 0 °C = 273.15 K and standard pressure, as 1 atm = 760 mmHg, exactly. Standard conditions of temperature and pressure are usually abbreviated as **STP**.

## TWO PRACTICE EXAMPLES

Each worked Example is followed by *two* Practice Examples. Practice Example A allows you to work a problem similar to the worked Example that precedes it. Practice Example B takes you one step further or calls on a previously learned concept. These sets of Practice Examples should help you integrate and extend your knowledge and problem-solving skills and prepare you for the end-of-chapter Exercises.

▶ Feature Problem C on page 201 tests your qualitative understanding of the gas laws at both the microscopic and molecular levels.

▶ If the volume of $N_2(g)$ at STP had been somewhat *greater* than 1.20 L, we could not have answered the question with qualitative reasoning alone. Where calculations are required, we can do these most readily with the ideal gas equation (Section 6-3).

### EXAMPLE 6-5

*Testing Your Qualitative Understanding of the Gas Laws.* Which will have the greater volume when the gases are compared at STP: (a) 1.20 L $N_2(g)$ at 25 °C and 748 mmHg or (b) 1.25 L $O_2(g)$ at STP?

**Solution**

First, we can assume that the simple gas laws apply to both gases. The fact that one gas is $N_2$ and the other $O_2$ does not enter into our assessment. Next, we see that the gas in (b) is already at STP and its volume is 1.25 L. The crux of the matter is whether the volume of gas increases or decreases when the conditions in (a) are changed to STP. Lowering the temperature from 25 to 0 °C causes the volume to decrease (Charles's law). Raising the pressure from 748 to 760 mmHg also causes the volume to decrease (Boyle's law). At STP the gas in (a) will occupy a volume smaller than its initial 1.20 L. Thus the greater volume at STP is the 1.25 L $O_2(g)$.

**Practice Example A:** Suppose that the initial condition in case (a) in Example 6-5 is 0.60 L $N_2(g)$ at 150 K and 0.98 atm. Which gas, the $N_2$ or the $O_2$, will have the greater volume at STP?

**Practice Example B:** Suppose that the initial condition in case (a) in Example 6-5 is 1.20 L $N_2(g)$ at 4.2 °C and 758 mmHg, whereas that in (b) is 1.25 L $O_2(g)$ at −2.5 °C and 1.02 atm. Which gas will have the greater volume at STP?

## FEATURE PROBLEMS

Some key ideas in the chapter are called out in margin notes (as above) and featured in problems of special interest. These Feature Problems may involve the interpretation of data or graphs, retrace aspects of the history of chemistry, present new material or alternative techniques of problem solving, or address situations with a practical application. These problems vary in difficulty and may be discussed as in-class exercises, worked on collaboratively in groups, or done individually.

### Feature Problems

A. The figure shows how the emission of pollutants is related to the air/fuel ratio in an internal combustion engine.
1. Use information given in this chapter to provide a general interpretation of the figure.
2. Establish by calculation that the stoichiometric ratio of air to fuel (on a mass basis) is about 14.5 : 1. To do so, assume that (a) octane, $C_8H_{18}$, is a typical gasoline molecule; (b) the maximum quantity of heat is evolved when $C_8H_{18}(l)$ and $O_2(g)$ are in their stoichiometric proportions and $CO_2(g)$ and $H_2O(l)$ are the products; (c) air used in the combustion is the naturally occurring mixture of $O_2(g)$, $N_2(g)$, and $Ar(g)$.

B. In some research that required the careful measurement of gas densities, John Rayleigh, a physicist, found that the density of $O_2(g)$ had the same value whether the gas was obtained from air or derived from one of its compounds. The situation with $N_2(g)$ was different, however. The density of $N_2(g)$ had the same value when the $N_2(g)$ was derived from any of various compounds, but a *different* value if the $N_2(g)$ was extracted from air. In 1894, Rayleigh enlisted the aid of William Ramsay, a chemist, to solve this apparent mystery; in the course of their work they discovered the noble gases.
1. Why do you suppose that the $N_2(g)$ extracted from liquid air did *not* have the same density as $N_2(g)$ obtained from its compounds?
2. Which gas do you suppose had the greater density, $N_2(g)$ extracted from air or $N_2(g)$ prepared from nitrogen compounds? Explain.
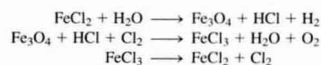3. The way in which Ramsay proved that nitrogen gas extracted from air was itself a mixture of gases involved allowing this nitrogen to react with magnesium metal to form magnesium nitride. Explain the significance of this experiment.
4. Use data on the composition of air from Table 8.1 to calculate the *percent difference* in the densities at STP of
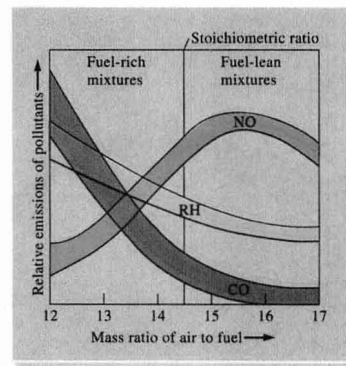


Rayleigh's $N_2(g)$ extracted from air and $N_2(g)$ derived from nitrogen compounds.

C. Various thermochemical cycles are being explored as possible sources of $H_2(g)$. The object is to find a series of reactions that can be conducted at moderate temperatures (about 500 °C) and has as its net result the decomposition of water into $H_2$ and $O_2$. Show that the following series of reactions meets this requirement. (*Hint:* Balance the equations, multiply by the appropriate coefficients, and combine them into a net equation.)

$$FeCl_2 + H_2O \longrightarrow Fe_3O_4 + HCl + H_2$$
$$Fe_3O_4 + HCl + Cl_2 \longrightarrow FeCl_3 + H_2O + O_2$$
$$FeCl_3 \longrightarrow FeCl_2 + Cl_2$$

## Exercises

### The Periodic Law

19. Use data from Figure 10-1 and Equation (10.1) to estimate the density expected for the undiscovered element 114. Assume a mass number of 298.

21. The following densities, in grams per cubic centimeter, are for the listed elements in their standard states at 298 K. Show that density is a periodic property of these elements: Al, 2.699; Ar, 0.0018; As, 5.778; Br, 3.100; Ca, 1.550; Cl, 0.0032; Ga, 5.904; Ge, 5.323; Kr, 0.0037; Mg, 1.738; P, 1.823; K, 0.856; Se, 4.285; Si, 2.336; Na, 0.968; S, 2.069.

20. Suppose that lanthanum ($Z = 57$) were a newly discovered element having a density of 6.145 g/cm³. Estimate its molar mass. (*Hint:* Use Figure 10-1.)

22. The following melting points are in degrees Celsius. Show that melting point is a periodic property of these elements: Al, 660; Ar, −189; Be, 1278; B, 2300; C, 3350; Cl, −101; F, −220; Li, 179; Mg, 651; Ne, −249; N, −210; O, −218; P, 590; Si, 1410; Na, 98; S, 119.

### The Periodic Table

23. Mendeleev's periodic table did not preclude the possibility of a new group of elements that would fit within the existing table, as was the case with the noble gases. Moseley's work did preclude this possibility. Explain this difference.

25. Assuming that the seventh period is 32 members long, what should be the atomic number of the noble gas following radon (Rn)? of the alkali metal following francium (Fr)? What would you expect their approximate atomic masses to be?

24. Explain why the several periods in the periodic table do not all have the same number of members.

26. Concerning the incomplete seventh period of the periodic table, what should be the atomic number of the element **(a)** for which the filling of the $6d$ subshell is completed; **(b)** that should most closely resemble bismuth; **(c)** that should be a noble gas?

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the masses of other atoms can be established by mass spectrometry. From the masses of the different isotopes of an element and their percent natural abundances, the weighted-average atomic mass (weight) of an element can be determined.

... atom, of atoms ... ned its molar mass. Molar ... and the Avogadro constant are used in a variety of calculations involving the mass, amount (in moles), or number of atoms in a sample of an element.

## Summarizing Example

A stainless steel ball bearing has a radius of 6.35 mm and a density of 7.75 g/cm³. Iron is the principal element in steel. Carbon is a key minor element. The ball bearing consists of 0.25% carbon, by mass.

Given that the percent natural abundance of $^{13}C$ is 1.108%, how many $^{13}C$ atoms are present in the ball bearing?

**1.** *Determine the volume of the ball bearing, in cubic centimeters.* The formula for the volume ($V$) of a sphere is $V = \frac{4}{3}\pi r^3$. The radius ($r$) must be expressed in centimeters. *Result:* 1.07 cm³.

**2.** *Determine the mass of carbon present.* (a) Use the density of steel and the volume of the ball bearing to calculate its

mass. (b) Use the percent carbon in the steel to convert from mass of steel to mass of carbon. *Result:* 0.021 g C.

**3.** *Determine the total number of carbon atoms present.* (a) Use the molar mass of carbon to convert the mass of carbon to the amount of carbon expressed in moles. (b) Use the Avogadro constant to establish the number of carbon atoms. *Result:* $1.1 \times 10^{21}$ C atoms.

**4.** *Determine the number of $^{13}C$ atoms.* The percent natural abundance leads to the conversion factor: 1.108 $^{13}C$ atoms/100 C atoms. Multiply the result of step **3** by this factor to obtain the number of $^{13}C$ atoms. *Answer:* $1.2 \times 10^{19}$ $^{13}C$ atoms.

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## Key Terms

| | | |
|---|---|---|
| actinides (2-6) | group (2-6) | neutron (2-3) |
| alpha ($\alpha$) particle (2-2) | ion (2-4) | neutron number (2-3) |
| atomic mass (weight) (2-5) | isotope (2-4) | nonmetals (2-6) |
| atomic mass unit (u) (2-3) | lanthanides (2-6) | nuclide (2-4) |
| atomic number (proton number), Z (2-3) | law of conservation of mass (2-1) | percent natural abundance (2-4) |
| | law of constant composition (definite proportions) (2-1) | period (2-6) |
| Avogadro constant, $N_A$ (2-7) | | periodic table (2-6) |
| beta ($\beta$) particle (2-2) | main group elements (2-6) | proton (2-3) |
| cathode ray (2-2) | mass number, A (2-3) | radioactivity (2-2) |
| chemical symbol (2-4) | mass spectrometer (2-4) | transition elements (transition metals) (2-6) |
| electron (2-2) | metals (2-6) | |
| family (2-6) | molar mass, $\mathcal{M}$ (2-7) | |
| gamma ($\gamma$) ray (2-2) | mole (2-7) | |

## ARE YOU WONDERING . . .

New ideas may prompt you to raise important questions as you read the text. These boxes highlight some questions that students frequently ask and provide responses at that point in the text where such questions usually arise.

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**48** CHAPTER 2 *Atoms and the Atomic Theory*

### Are You Wondering . . .

**Why some atomic masses (for example, F = 18.9984032 u) are stated so much more precisely than others?**

There is *one* naturally occurring type of fluorine atom: fluorine-19. Determining the atomic mass of fluorine means establishing the mass of this type of atom as precisely as possible. On the other hand, krypton, with an atomic mass 83.80 u, has *six* naturally occurring isotopes. Because the percent distribution of the isotopes of krypton may differ very slightly from one sample to another, we can't state the weighted-average atomic mass of krypton with high precision.

### 2-6 Introduction to the Periodic Table

Scientists spend a lot of time organizing information into useful patterns. Before they can organize information, however, they must possess it, and it must be correct. Botanists had enough information about plants to organize their field in the eighteenth century. Because of uncertainties in atomic masses and because several elements remained undiscovered, chemists were not able to organize the elements until a century later.

We can distinguish one element from all others through its particular set of observable physical properties. For example, sodium has a low density of 0.971 g/cm³ and a low melting point of 97.81 °C. No other element has this same combination of density and melting point. Potassium, though, also has a low density (0.862 g/cm³) and low melting point (63.65 °C), much like sodium. Sodium and potassium further resemble each other in that both are good conductors of heat and electricity. Moreover, both react vigorously with water to liberate hydrogen gas. Gold, on the other hand, has a density (19.32 g/cm³) and melting point ( °C) that are very ... those of sodium or pota...

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## Focus On   Fats, Carbohydrates, and Energy Storage

Aerobic exercise is an excellent way to burn fat and build muscle.

Aerobics, tennis, weight lifting, and jogging are popular forms of exercise, which is important to maintaining a healthy body. Where do we get the energy to do these things? Surprisingly, mostly from fat, the body's chief energy storage system.

During exercise, fat molecules react with water (hydrolyze) to form a group of compounds called fatty acids. Through a complex series of reactions, these fatty acids are converted to carbon dioxide and water. Energy released in these reactions is used to power muscles. A
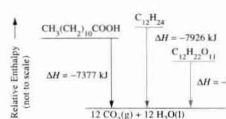
typical human fatty acid is palmitic acid, $CH_3(CH_2)_{14}COOH$.

The direct combustion of $CH_3(CH_2)_{14}COOH$ in a bomb calorimeter yields the same products as its metabolism in the body, together with a large quantity of heat.

$$CH_3(CH_2)_{14}COOH + 23 O_2(g) \longrightarrow 16 CO_2(g) + 16 H_2O(l)$$
$$\Delta H = -9977 \text{ kJ}$$
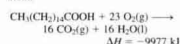
A hydrocarbon with a similar carbon and hydrogen content, $C_{16}H_{34}$, yields a similar quantity of heat on complete combustion, $-10,700$ kJ. The fat stored in our bodies (our fuel) is comparable to the jet fuel in an airplane. In both cases the fuel adds weight, but can be burned to produce the large amount of energy needed for movement and life support.

Energy from simple carbohydrates is released more rapidly than from fats, and this is why we consume sugars (fruit juice, candy bars, . . .) for a quick burst of energy. The combustion of 1 mol $C_{12}H_{22}O_{11}$ (sucrose or cane sugar), however, yields much less energy

**FIGURE 7-15**
**Energy values of a fatty acid, a carbohydrate, and a hydrocarbon.**
Each of these three comparable 12-carbon compounds produces 12 mol $CO_2(g)$ and some $H_2O(l)$ as combustion products. The hydrocarbon, $C_{12}H_{24}$, releases the most energy on combustion. The fatty acid lauric acid, $CH_3(CH_2)_{10}COOH$, yields nearly as much, but the carbohydrate sucrose, $C_{12}H_{22}O_{11}$, yields considerably less.

($-5640$ kJ/mol) than does a fatty acid of similar carbon and hydrogen content. In the enthalpy diagram of Figure 7-15, we can think of the sucrose molecule as in a condition somewhere between a hydrocarbon and the ultimate oxidation products: $CO_2$ and $H_2O$. A fatty acid, on the other hand, is in a condition closer to that of the hydrocarbon.

A common energy unit for expressing the energy values of food is the food Calorie (Cal), which is actually a kilocalorie.

$$1 \text{ Cal} = 1 \text{ kcal} = 4.184 \text{ kJ}$$

As energy sources, fats yield about 9 Cal/g (38 kJ/g), whereas carbohydrates and proteins both yield about 4 Cal/g (17 kJ/g).

Consider a 65-kg (143-lb) person of average height and build. Typically this person would have about 11 kg (24 lb ) of stored fats—a fuel reserve of about $4.2 \times 10^5$ kJ. On the other hand, if the body stored only carbohydrates, the stored reservoir would have to be about 25 kg. The "65-kg" person would have to weigh about 80 kg (176 lb). A more dramatic picture might be that of a bird carrying its energy reserves as carbohydrates instead of fat. The excess "baggage" might make it impossible for the bird to fly.

It is important to understand some other matters about fat. First, the quantity of stored fat that an individual typically carries is greater than what is needed to meet energy

requirements. Excess body fat leads to an increased risk of heart disease, diabetes, and other medical problems. Second, one does not have to consume fats in order to store body fats. All food types are broken down into small molecules during digestion, and these small molecules are reassembled into the more complex structures that the body needs. In the body, fat molecules can be synthesized from the small molecules produced from carbohydrates, for example. Current medical findings indicate that the most healthful diet provides the bulk of food calories through carbohydrates, with the recommended percent of caloric intake from fats being less than 30%.

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## FOCUS ON

Practical and real-world examples of chemistry are integrated throughout the text. And at the end of each chapter is a short essay on a practical topic that relates the chapter subject to real-world situations. This feature gives you an opportunity to read about a current topic in more depth.

# Brief Contents

# Contents

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