



# CHEMISTRY

an atoms-focused approach

GILBERT  
KIRSS  
FOSTER

# CHEMISTRY

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AN ATOMS-FOCUSED APPROACH

**Thomas R. Gilbert**

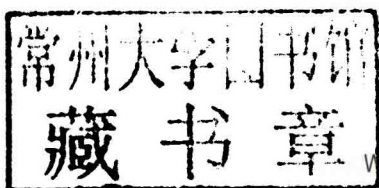
NORTHEASTERN UNIVERSITY

**Rein V. Kirss**

NORTHEASTERN UNIVERSITY

**Natalie Foster**

LEHIGH UNIVERSITY



With Contributions by

**Geoffrey Davies**

NORTHEASTERN UNIVERSITY



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## 18 8A

18 8A

7 Actinides

We have used the U.S. system as well as the system recommended by the International Union of Pure and Applied Chemistry (IUPAC) to label the groups in this periodic table. The system used in the United States includes a letter and a number (1A, 2A, 3B, 4B, etc.), which is close to the system developed by Mendeleev. The IUPAC system uses numbers 1–18 and has been recommended by the American Chemical Society (ACS). While we show both numbering systems here, we use the IUPAC system exclusively in the book.

Element	Symbol	Atomic Number	Average Atomic Mass <sup>a</sup>	Element	Symbol	Atomic Number	Average Atomic Mass <sup>a</sup>	Element	Symbol	Atomic Number	Average Atomic Mass <sup>a</sup>
Actinium	Ac	89	[227]	Gold	Au	79	196.97	Promethium	Pm	61	[145]
Aluminum	Al	13	26.982	Hafnium	Hf	72	178.49	Protactinium	Pa	91	231.04
Americium	Am	95	[243]	Hassium	Hs	108	[277]	Radium	Ra	88	[226]
Antimony	Sb	51	121.76	Helium	He	2	4.0026	Radon	Rn	86	[222]
Argon	Ar	18	39.948	Holmium	Ho	67	164.93	Rhenium	Re	75	186.21
Arsenic	As	33	74.922	Hydrogen	H	1	1.0079	Rhodium	Rh	45	102.91
Astatine	At	85	[210]	Indium	In	49	114.82	Roentgenium	Rg	111	[280]
Barium	Ba	56	137.33	Iodine	I	53	126.90	Rubidium	Rb	37	85.468
Berkelium	Bk	97	[247]	Iridium	Ir	77	192.22	Ruthenium	Ru	44	101.07
Beryllium	Be	4	9.0122	Iron	Fe	26	55.845	Rutherfordium	Rf	104	[265]
Bismuth	Bi	83	208.98	Krypton	Kr	36	83.798	Samarium	Sm	62	150.36
Bohrium	Bh	107	[270]	Lanthanum	La	57	138.91	Scandium	Sc	21	44.956
Boron	B	5	10.811	Lawrencium	Lr	103	[262]	Seaborgium	Sg	106	[271]
Bromine	Br	35	79.904	Lead	Pb	82	207.2	Selenium	Se	34	78.96
Cadmium	Cd	48	112.41	Lithium	Li	3	6.941	Silicon	Si	14	28.086
Calcium	Ca	20	40.078	Livermorium	Lv	116	[293]	Silver	Ag	47	107.87
Californium	Cf	98	[251]	Lutetium	Lu	71	174.97	Sodium	Na	11	22.990
Carbon	C	6	12.011	Magnesium	Mg	12	24.305	Strontium	Sr	38	87.62
Cerium	Ce	58	140.12	Manganese	Mn	25	54.938	Sulfur	S	16	32.065
Cesium	Cs	55	132.91	Meitnerium	Mt	109	[276]	Tantalum	Ta	73	180.95
Chlorine	Cl	17	35.453	Mendelevium	Md	101	[258]	Technetium	Tc	43	[98]
Chromium	Cr	24	51.996	Mercury	Hg	80	200.59	Tellurium	Te	52	127.60
Cobalt	Co	27	58.933	Molybdenum	Mo	42	95.96	Terbium	Tb	65	158.93
Copernicium	Cn	112	[285]	Neodymium	Nd	60	144.24	Thallium	Tl	81	204.38
Copper	Cu	29	63.546	Neon	Ne	10	20.180	Thorium	Th	90	232.04
Curium	Cm	96	[247]	Neptunium	Np	93	[237]	Thulium	Tm	69	168.93
Darmstadtium	Ds	110	[281]	Nickel	Ni	28	58.693	Tin	Sn	50	118.71
Dubnium	Db	105	[268]	Niobium	Nb	41	92.906	Titanium	Ti	22	47.867
Dysprosium	Dy	66	162.50	Nitrogen	N	7	14.007	Tungsten	W	74	183.84
Einsteinium	Es	99	[252]	Nobelium	No	102	[259]	Ununocium	Uuo	118	[294]
Erbium	Er	68	167.26	Osmium	Os	76	190.23	Ununpentium	Uup	115	[288]
Europium	Eu	63	151.96	Oxygen	O	8	15.999	Ununseptium	Uus	117	[294]
Fermium	Fm	100	[257]	Palladium	Pd	46	106.42	Ununtrium	Uut	113	[284]
Flerovium	Fl	114	[289]	Phosphorus	P	15	30.974	Uranium	U	92	238.03
Fluorine	F	9	18.998	Platinum	Pt	78	195.08	Vanadium	V	23	50.942
Francium	Fr	87	[223]	Plutonium	Pu	94	[244]	Xenon	Xe	54	131.29
Gadolinium	Gd	64	157.25	Polonium	Po	84	[209]	Ytterbium	Yb	70	173.05
Gallium	Ga	31	69.723	Potassium	K	19	39.098	Yttrium	Y	39	88.906
Germanium	Ge	32	72.63	Praseodymium	Pr	59	140.91	Zinc	Zn	30	65.38
								Zirconium	Zr	40	91.224

<sup>a</sup> Average atomic mass values for most elements are from *Pure Appl. Chem.* (2011) **83**, 359. Those for B, C, Cl, H, Li, N, O, Si, S, and Tl are from *Pure Appl. Chem.* (2009) **81**, 2131 and are within the ranges cited in the first reference. Atomic masses in brackets are the mass numbers of the longest-lived isotopes of elements with no stable isotopes.

# CHEMISTRY

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AN ATOMS-FOCUSED APPROACH





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## About the Authors



**Thomas R. Gilbert** has a BS in chemistry from Clarkson and a PhD in analytical chemistry from MIT. After 10 years with the Research Department of the New England Aquarium in Boston, he joined the faculty of Northeastern University, where he is currently associate professor of chemistry and chemical biology. His research interests are in chemical and science education. He teaches general chemistry and science education courses and conducts professional development workshops for K–12 teachers. He has won Northeastern's Excellence in Teaching Award and Outstanding Teacher of First-Year Engineering Students Award. He is a fellow of the American Chemical Society and in 2012 was elected to the ACS Board of Directors.



**Rein V. Kirss** received both a BS in chemistry and a BA in history as well as an MA in chemistry from SUNY Buffalo. He received his PhD in inorganic chemistry from the University of Wisconsin, Madison, where the seeds for this textbook were undoubtedly planted. After two years of postdoctoral study at the University of Rochester, he spent a year at Advanced Technology Materials, Inc., before returning to academics at Northeastern University in 1989. He is an associate professor of chemistry with an active research interest in organometallic chemistry.



**Natalie Foster** is emeritus professor of chemistry at Lehigh University in Bethlehem, Pennsylvania. She received a BS in chemistry from Muhlenberg College and MS, DA, and PhD degrees from Lehigh University. Her research interests included studying poly(vinyl alcohol) gels by NMR as part of a larger interest in porphyrins and phthalocyanines as candidate contrast enhancement agents for MRI. She taught both semesters of the introductory chemistry class to engineering, biology, and other nonchemistry majors and a spectral analysis course at the graduate level. She is the recipient of the Christian R. and Mary F. Lindback Foundation Award for distinguished teaching.



# Preface

Dear Student,

An old adage says that you can't (or shouldn't) judge a book by its cover. However, the cover of this book—and its title—say a great deal about what's inside. The cover photo is an image of graphite, the most common form of a common element: carbon. The illustration shows what might happen if we could pull back an ultrathin surface layer of graphite and view it using an extremely powerful microscope. What we would see is a layer of carbon atoms, each chemically bonded to three others, forming a seemingly endless array of hexagons. Here's the cool part: peeling back and viewing a single layer of graphite is more than a theoretical possibility—it's actually happened! In addition, scientists have been able to study the properties of these monolayers of carbon atoms, and what they've discovered is pretty amazing. These layers represent the strongest material known—much stronger than the strongest steel—and they conduct electricity better than the most conductive metal. This remarkable material even has its own name: graphene.

Why does graphene have such remarkable properties? The answer to that question is contained in this book's title: *Chemistry: An Atoms-Focused Approach*. The properties of graphene are tied directly to the presence of those hexagonal arrays of carbon atoms and the nature of the chemical bonds that join them together. The geometry and strength of those bonds contribute to the overall strength of graphene, and the unusually high mobilities of the electrons that make up those bonds contribute to the material's outstanding conductivity.

Our cover illustrates a central message of this book: the properties of substances are directly linked to their atomic and molecular structures. In our book we start with the smallest particles of matter and assemble them into more elaborate structures: from subatomic particles to single atoms to monatomic ions and polyatomic ions, and from atoms to small molecules to bigger ones to truly gigantic polymers. By constructing this layered particulate view of matter, we hope our book helps you visualize the underlying chemistry of a wide range of substances and the changes they undergo. With this ability to visualize atoms and molecules, you won't have to resort to memorizing formulas and reactions as a strategy for *surviving* general chemistry. Instead, you will be able to understand *why* elements combine to form compounds with particular formulas and *why* substances react with each other the way they do. For example, you won't have to memorize the charges of the common ions that make seawater (and your blood plasma and tears) salty; instead, you will understand why the many billions of tons of sodium dissolved in the sea (and the 100 grams of it inside your body) exist entirely as  $\text{Na}^+$  ions.



## Context

While our primary learning goal is for you to be able to interpret and even predict the physical and chemical properties of substances based on their atomic and molecular structures, we would also like you to understand how chemistry is linked to other scientific disciplines. We illustrate these connections using contexts drawn from fields such as biology, environmental science, materials science, astronomy, geology, and medicine. We hope that this approach helps you better understand how scientists apply the principles of chemistry to treat and cure diseases, to make more-efficient use of natural resources, and to minimize the impact of human activity on our planet and its climate.

## Problem-Solving Strategies

Another major goal of our book is to help you improve your problem-solving skills. To solve problems in chemistry, you first need to recognize the connections between the information provided in a problem and the answer you are asked to find. Sometimes the hardest part of solving a problem is distinguishing between information that is relevant and information that is not. Once you are clear on where you are starting and where you are going, planning for and carrying out a solution become much easier.

To help you hone your problem-solving skills, we have developed a framework that we introduce in Chapter 1. It is a four-step approach we call **COAST**, which is our acronym for (1) **Collect** and **Organize**, (2) **Analyze**, (3) **Solve**, and (4) **Think**

about it. We use these four steps in *every* Sample Exercise and in the solutions to *odd* problems in the Student's Solutions Manual. They are also used in the hints and feedback embedded in the SmartWork online homework program. To summarize the four steps:

**COLLECT AND ORGANIZE** helps you understand where to begin to solve the problem. In this step we often rephrase the problem and the answer that is sought, and we identify the relevant information that is provided in the problem statement or available elsewhere in the book.

**ANALYZE** is where we map out a strategy for solving the problem. As part of that strategy we often estimate what a reasonable answer might be.

**SOLVE** applies our analysis of the problem from the second step to the information and relations from the first step to actually solve the problem. We walk you through each step in the solution so that you can follow the logic and the math.

**THINK ABOUT IT** reminds us that an answer is not the last step in solving a problem. We should check the accuracy of the solution and think about the value of a quantitative answer. Is it realistic? Are the units correct? Is the number of significant figures appropriate? Does it agree with our estimate from the Analyze step?

### SAMPLE EXERCISE 10.5 Applying Amontons's Law

L04

Labels on aerosol cans caution against their incineration because the cans may explode when the pressure inside them exceeds 3.00 atm. At what temperature in degrees Celsius might an aerosol can burst if its internal pressure is 2.00 atm at 25°C?

**COLLECT AND ORGANIZE** We are given the temperature ( $T_1 = 25^\circ\text{C}$ ) and pressure ( $P_1 = 2.00$  atm) of a gas and asked to determine the temperature ( $T_2$ ) at which the pressure ( $P_2$ ) reaches 3.00 atm.

**ANALYZE** Because the gas is isolated in a rigid aerosol can, we know that the quantity of gas and its volume are constant. Amontons's law (Equation 10.18) relates the pressures of a confined quantity of gas at two different temperatures. To estimate our answer, we note that the pressure in the can must increase by 50% to reach 3.00 atm. Pressure is directly proportional to absolute temperature, so it, too, must increase by 50%. The initial temperature of  $25^\circ\text{C}$  is nearly 300 K, so a 50% increase in absolute temperature corresponds to a final temperature near 450 K, or about  $175^\circ\text{C}$ .

**SOLVE** Rearranging Equation 10.18 to solve for  $T_2$ :

$$T_2 = \frac{T_1 P_2}{P_1}$$

and inserting the given  $T$  and  $P$  values:

$$T_2 = \frac{[(25 + 273) \text{ K}](3.00 \text{ atm})}{2.00 \text{ atm}} = 447 \text{ K}$$

Converting  $T_2$  to degrees Celsius:

$$T_2 = 447 \text{ K} - 273 = 174^\circ\text{C}$$

**THINK ABOUT IT** This temperature is close to our estimated value. It is also well below the temperatures that solid waste experiences in the fires of an operating incinerator, which makes the warning label on the can all the more important.

**Practice Exercise** Air pressure in each of the tires of an automobile is adjusted to 34 psi at a gas station in San Diego, California, where the air temperature is  $68^\circ\text{F}$ . After a 3-hour drive along Interstate Highway 8, the car and driver are in Yuma, Arizona, where the temperature is  $110^\circ\text{F}$ . What is the pressure in the tires?



Many students use the **Sample Exercises** more than any other part of the book. Sample Exercises take the concept being discussed and illustrate how to apply it to solve a problem. We think that repeated application of the COAST framework will help you refine your problem-solving skills and hope that the approach becomes habit-forming for you. When you finish a Sample Exercise, you'll find a Practice Exercise to try on your own. If you have the ebook, the Practice Exercises are “live,” meaning that you can solve them and receive hints and answer-specific feedback when you need help. The next few pages describe how to use the tools built into each chapter to gain a conceptual understanding of chemistry.

## Chapter Structure

Each chapter begins with an **opening story**, which provides glimpses of how the chemistry in the chapter that follows connects to the world. We have used topics that should be familiar to you, but we place them in chemical contexts that may surprise you.

If you are trying to decide what is most important in a chapter, check the **Learning Outcomes** listed on the first page. Whether you are reading the chapter from first page to last or reviewing it for an exam, the Learning Outcomes should help you focus on the key information you need and the skills you should acquire. You will also see which Learning Outcomes are linked to which Sample Exercises in the chapter.


As you study each chapter, you will find **key terms** in boldface in the text and in a running glossary in the margin. We have deliberately duplicated these definitions so that you can continue reading without interruption but quickly find them when doing homework or reviewing for a test. All key terms are also defined in the Glossary in the back of the book.

Many concepts are related to others described earlier in the book. We point out these relationships with **Connection** icons in the margins. We hope they help you draw your own connections between major themes covered in the book.

To help you develop your own microscale view of matter, we use **molecular art** to enhance photos and figures, and to illustrate what is happening at the atomic and molecular levels.

If you're looking for additional help visualizing a concept, we have about 100 **ChemTours**, denoted by the ChemTour icon, available online at [wwnpag.es/chemtours](http://wwnpag.es/chemtours). ChemTours demonstrate dynamic processes and help you visualize events at the molecular level. Many of the ChemTours allow you to manipulate variables and observe the resulting changes. Questions at the end of the ChemTour tutorials offer step-by-step assistance in solving problems and provide useful feedback.

**Concept Tests** are short, conceptual questions that serve as a self-check by asking you to stop and answer a question relating to what you just read. We designed them to help you see for yourself whether you have grasped a key concept and can apply it. We have an average of one Concept Test per section and many have a visual component. You may find some Concept Tests challenging. We provide the answers to all Concept Tests in the back of the book.

 **CONNECTION** In Chapter 9 we defined standard conditions of temperature and pressure as they apply to thermochemistry. Note that *STP* and *standard conditions* are not the same. STP applies strictly to calculations involving the gas laws, while standard conditions apply to thermochemical data.

 **CHEMTOUR**

### CONCEPT TEST

Which of the graphs in Figure 10.17 correctly describes the relationship between the product of pressure and volume ( $PV$ ) as a function of pressure ( $P$ ) for a given quantity of gas at constant temperature?

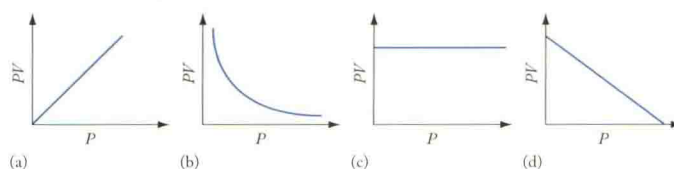


FIGURE 10.17



At the end of each chapter is a special Sample Exercise that draws on several key concepts from the chapter and occasionally others from preceding chapters to solve a problem that is framed in the context of a real-world scenario or incident. We call these **Integrated Sample Exercises**. You may find them more challenging than most of those that precede them in each chapter, but please invest your time in working through them because they represent authentic exercises that will enhance your problem-solving skills.

### SAMPLE EXERCISE 10.16 Integrating Concepts: Air for a Jet Engine

The Boeing 767 (Figure 10.42) is one of the most popular wide-body commercial airliners (over 1000 have been built), and it's also one of the most fuel efficient. While cruising at 851 km/hr (530 mph) at an altitude of 11,000 m (36,000 ft), a 767-200ER (the extended-range version of the plane) consumes about 1720 U.S. gallons of jet fuel per hour.

We are given the following facts: (1) the density of jet fuel is 0.80 g/mL; (2) at an altitude of 11,000 m,  $P_{\text{atm}} = 210 \text{ mmHg}$  and  $T = -56^\circ\text{C}$ ; (3) 1 U.S. gallon = 3.785 L; and (4) dodecane,  $\text{C}_{12}\text{H}_{26}$ , is considered an appropriate model hydrocarbon for jet fuel.

- What volume of air, in liters, does a cruising 767 need so that it can completely burn an hour's worth of jet fuel?
- Fuel efficiency is often based on number of passengers times distance traveled per volume of fuel consumed. In the United States, this value is typically expressed in units of passenger-miles per gallon. However, in much of the rest of the world, efficiency units are inverted and are typically expressed in liters per 100 km per passenger. Express the fuel efficiency of a full 767-200ER (which holds 224 passengers) in both sets of units.

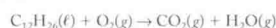
**COLLECT AND ORGANIZE** We know the quantity of dodecane to be combusted, and we are asked to calculate the volume of air needed for complete combustion—that is, to convert its C and H content into  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . We know the pressure and temperature of the air. According to Table 10.1, dry air is 20.95%  $\text{O}_2$ .

**ANALYZE** This exercise involves a chemical reaction, so writing a balanced chemical equation describing it is a good place to start. Next we need to convert the volume of fuel into an equivalent number of moles of fuel, and then to use the stoichiometry of the reaction to convert that value to moles of  $\text{O}_2$ . We will then calculate the equivalent volume of  $\text{O}_2$  using the ideal gas equation. (The pressure is below 1 atm, so there should be no need to correct for nonideal behavior.) Finally we will convert the volume of  $\text{O}_2$  to the corresponding volume of air. The volume of air needed each hour by the engines of a 767 should be enormous.

#### SOLVE

- Calculating volume of air needed in 1 hour:

- Write the balanced chemical equation describing the combustion reaction. The reactants and products are



We first balance the numbers of C and H atoms:



This leaves us with an odd number of O atoms on the right, requiring that we multiply all the terms by 2 and then balance the number of O atoms:

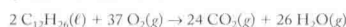


FIGURE 10.42 A Boeing 767-200ER.

- Converting the volume of fuel consumed in 1 hour into an equivalent number of moles of  $\text{C}_{12}\text{H}_{26}$ :

$$1720 \text{ gal} \left( \frac{3.785 \text{ L}}{1 \text{ gal}} \right) \left( \frac{1000 \text{ mL}}{1 \text{ L}} \right) \left( \frac{0.80 \text{ g}}{\text{mL}} \right) \left( \frac{1 \text{ mol}}{170.33 \text{ g}} \right) = 3.06 \times 10^4 \text{ mol C}_{12}\text{H}_{26}$$

- Converting moles of  $\text{C}_{12}\text{H}_{26}$  into moles of  $\text{O}_2$ :

$$3.06 \times 10^4 \text{ mol C}_{12}\text{H}_{26} \left( \frac{37 \text{ mol O}_2}{2 \text{ mol C}_{12}\text{H}_{26}} \right) = 5.66 \times 10^5 \text{ mol O}_2$$

- In converting moles of  $\text{O}_2$  into a volume of air using the ideal gas law, we need to convert pressure units from mmHg to atm and temperature to the Kelvin scale:

$$V = \frac{nRT}{P} = \frac{(5.66 \times 10^5 \text{ mol}) \left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273 - 56) \text{ K}}{210 \text{ mmHg} \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right)} = 3.65 \times 10^7 \text{ L O}_2$$

- Air is 20.95%  $\text{O}_2$  by volume, so the volume of air the engines must take in each hour is

$$3.65 \times 10^7 \text{ L O}_2 \left( \frac{100 \text{ L air}}{20.95 \text{ L O}_2} \right) = 1.74 \times 10^8 \text{ L air}$$

- Calculating fuel efficiency in U.S. Customary units based on the distance traveled and fuel consumed in 1 hour:

$$\frac{224 \text{ passengers} \times 530 \text{ miles}}{1720 \text{ gal}} = 69.0 \text{ passenger-miles/gallon}$$

The corresponding efficiency value in liters per 100 km per passenger is

$$\frac{1720 \text{ gal} \times \frac{3.785 \text{ L}}{1 \text{ gal}}}{224 \text{ passengers} \times 851 \text{ km}} \times 100 = 3.41 \text{ L/100 km-passenger}$$

**THINK ABOUT IT** The calculated volume of air is 174 million liters. That is a lot of air, but then 1720 U.S. gallons of jet fuel is a lot of fuel. Perhaps a more interesting value is the 69.0 passenger-miles per gallon: better than that of most automobiles with two occupants, and you reach your destination much faster.

Also at the end of each chapter are a thematic **Summary** and a **Problem-Solving Summary**. The first is a brief synopsis of the chapter, organized by section. Key figures provide visual cues as you review. The Problem-Solving Summary is unique to this general chemistry book—it outlines the different types of problems you should be able to solve, where to find examples of them in the Sample Exercises, and reiterates relevant concepts and equations.

## PROBLEM-SOLVING SUMMARY

TYPE OF PROBLEM	CONCEPTS AND EQUATIONS	SAMPLE EXERCISES
Calculating relative rates of effusion	$\frac{\text{effusion rate}_A}{\text{effusion rate}_B} = \sqrt{\frac{M_B}{M_A}}$	(10.7) 10.1
Measuring gas pressure with a manometer; converting pressure units	See the conversion factors inside the back cover and Table 10.2.	10.2
Applying Boyle's law	$P_1V_1 = P_2V_2$	(10.10) 10.3
Applying Charles's law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	(10.14) 10.4
Applying Amontons's law	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	(10.18) 10.5
Applying the combined gas law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	(10.19) 10.6
Applying the ideal gas law	$PV = nRT$	(10.20) 10.7–10.10
Calculating mole fractions and partial pressures	$x_i = \frac{n_i}{n_{\text{total}}}$ $P_i = x_i P_{\text{total}}$	(10.23) 10.11 (10.24)
Calculating the quantity of a gas collected by water displacement	Calculate the partial pressure ( $P_i$ ) of the collected gas using the equation $P_i = P_{\text{atm}} - P_{\text{H}_2\text{O}}$	10.12
Calculating gas solubility using Henry's law	$C_{\text{gas}} = k_H P_{\text{gas}}$	(10.26) 10.13
Calculating root-mean-square speeds	$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$	(10.27) 10.14
Using the van der Waals equation	$\left[ P + a \left( \frac{n}{V} \right)^2 \right] (V - nb) = nRT$	(10.28) 10.15

Following the summaries are groups of questions and problems. The first group consists of **Visual Problems**. In many of them, you are asked to interpret a molecular view of a sample or a graph of experimental data.

**Concept Review Questions and Problems** come next, arranged by topic in the same order as they appear in the chapter. Concept Reviews are qualitative and often ask you to explain why or how something happens. Problems are paired and can be quantitative, conceptual, or a combination of both. **Contextual problems** have a title that describes the context in which the problem is placed. Finally, **Additional Problems** can come from any section or combination of sections in the chapter. Some of them incorporate concepts from previous chapters. Problems marked with an asterisk (\*) are more challenging and often take multiple steps to solve.

We want you to have confidence in using the answers in the back of the book as well as the Student's Solutions Manual, so we used a rigorous triple-check accuracy program for this book. Each end-of-chapter question and problem was solved independently by the Solutions Manual author, Karen Brewer, and by two additional chemical educators. Karen compared her solutions to those from the two reviewers and resolved any discrepancies. This process is designed to ensure clearly written problems and accurate answers in the appendices and Solutions Manual.

**10.99.** A sample of oxygen is collected over water at 25°C and 1.00 atm. If the total sample volume is 0.480 L, how many moles of O<sub>2</sub> are collected?

**10.100.** Water vapor is removed from the O<sub>2</sub> sample in Problem 10.99. What is the volume of the dry O<sub>2</sub> at 25°C and 1.00 atm?

**10.101.** The following reactions are carried out in sealed containers. Will the total pressure after each reaction be complete be greater than, less than, or equal to the total pressure before the reaction? Assume all reactants and products are gases at the same temperature.

- $\text{N}_2\text{O}_4(\text{g}) + \text{NO}_2(\text{g}) \rightarrow 3 \text{NO}(\text{g}) + 2 \text{O}_2(\text{g})$
- $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})$
- $\text{C}_2\text{H}_6(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{g})$

**10.102.** In each of the following gas-phase reactions, determine whether the total pressure at the end of the reaction (carried out in a sealed, rigid vessel) will be greater than, less than, or equal to the total pressure at the beginning. Assume all reactants and products are gases at the same temperature.

- $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{HCl}(\text{g})$
- $4 \text{NH}_3(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 4 \text{NO}(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$
- $2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$

**\*10.103. High-Altitude Mountaineering** Most alpine climbers breathe pure oxygen near the summits of the world's highest mountains. How much more O<sub>2</sub> is there in a lungful of pure O<sub>2</sub> at an elevation where atmospheric pressure is 266 mmHg than in a lungful of air at sea level? Express your answer as a percentage.

**10.104. Scuba Diving** A scuba diver is at a depth of 50 m, where the pressure is 5.0 atm. What should be the mole fraction of O<sub>2</sub> in the gas mixture the diver breathes to achieve the same  $P_{\text{O}_2}$  as at sea level?

## Solubilities of Gases and Henry's Law

## CONCEPT REVIEW

**10.109.** Why is the Henry's law constant for CO<sub>2</sub> so much larger than those for N<sub>2</sub> and O<sub>2</sub> at the same temperature?

*Hint:* Does CO<sub>2</sub> react with water?

**10.110.** As water in a beaker is heated, bubbles form inside the beaker at temperatures well below the boiling point of water. What gas is in the bubbles?

**10.111.** What type of intermolecular interaction accounts for the limited solubility of methane in water?

**\*10.112.** Air is primarily a mixture of nitrogen and oxygen. Is the Henry's law constant for the solubility of air in water the sum of  $k_{\text{H}}$  for N<sub>2</sub> and  $k_{\text{H}}$  for O<sub>2</sub>? Explain why or why not.

## PROBLEMS

**\*10.113. Arterial Blood** Arterial blood contains about 0.25 g of oxygen per liter at 37°C and standard atmospheric pressure. What is the Henry's law constant, in mol/(L · atm), for O<sub>2</sub> dissolution in blood?

**10.114.** The solubility of O<sub>2</sub> in water is 6.5 mg/L at an atmospheric pressure of 1 atm and a temperature of 40°C. Calculate the Henry's law constant of O<sub>2</sub> at 40°C.

**\*10.115. Oxygen for Climbers and Divers** Use the Henry's law constant for O<sub>2</sub> dissolved in arterial blood from Problem 10.113 to calculate the solubility of O<sub>2</sub> in the blood of (a) a climber on Mt. Everest ( $P_{\text{atm}} = 0.35 \text{ atm}$ ) and (b) a scuba diver breathing air at a depth of 20 meters ( $P = 3.0 \text{ atm}$ ).

**\*10.116.** The solubility of air in water is approximately  $7.9 \times 10^{-4} \text{ M}$  at 20°C and 1.0 atm. Calculate the Henry's law constant for air.

## Gas Diffusion: Molecules Moving Rapidly

## PROBLEMS



## Acknowledgments

As we launch this first edition of *Chemistry: An Atoms-Focused Approach*, our thanks go out to our publisher, W. W. Norton, for supporting us in writing a book that is written the way we much prefer to teach general chemistry. We especially wish to acknowledge the hard work and dedication of our editor/motivator/taskmaster, Erik Fahlgren. Erik has been an indefatigable source of guidance, perspective, persuasion, and inspiration to all of us. He has kept this project on track and on time over many months of conceptualization, development, and production. Erik is the single greatest reason the first copies of this book will roll off the presses in the summer of 2013, and, we hope, help change the way general chemistry is taught in many colleges and universities in the years that follow.

We are pleased to acknowledge the contributions of an outstanding developmental editor, Andrew Sobel. Andrew provided invaluable advice on how the chapters and topics of this book should be organized to achieve our learning goals, and he was also tenacious in reading and weighing every word of text and every illustration to ensure that we achieved those goals. Thanks to him, the following pages contain presentations and explanations that are more clear, more concise, and more likely to engage readers just setting out on their college careers.

Carla Talmadge is our project editor and doyenne of the many features in the book designed to help students visualize matter and the chemical changes it undergoes on an atomic and molecular scale. Assistant editor Renee Cotton kept information flowing in a timely fashion among all of the players involved in writing, reviewing, and writing some more. Thanks as well to Stephanie Romeo and Rona Tuccillo for finding great photos; production manager Eric Pier-Hocking for his work behind the scenes; Jennifer Barnhardt for managing the print ancillaries; Rob Bellinger for his diligence with the book's many electronic enhancements; and Stacy Loyal for her encouragement, fellowship, and marketing prowess. The entire Norton team was staffed by skilled, dedicated professionals who, as a bonus, were delightful people to work with and, on occasion, relax with.

Many reviewers, listed here, contributed to the development and production of this book. First among them are the Editorial Consultants: Lee Friedman, Jeremy Kua, Jeff Macedone, and Wayne Wesolowski. These insightful colleagues provided a continuous stream of useful suggestions on how to make a general chemistry textbook truly *atoms focused*. Essentially every section of every chapter in this book benefited from their input. We also owe an extra special thanks to Karen Brewer for her dedicated work on the Solutions Manuals and for her invaluable suggestions on how to improve the inventory and organization of problems and concept questions at the end of each chapter. She, along with Jordan Fantini and Amy Irwin, comprised the triple-check accuracy team who helped ensure the quality of the back-of-book answers and Solutions Manuals. Finally, we wish to acknowledge the care and thoroughness of Petia Bobadova-Parvanova, Tara Carpenter, Garry Crosson, Greg Domski, Doug English, Daniel Groh, Megan Grunert, Tracy Hamilton, Maria Kolber, Willem Leenstra, Douglas Magde, Gellert Mezei, Nancy Mullins, Sherine Obare, Edith Osborne, Robert Parson, James Patterson, Garry Pennycuff, and John Pollard for checking the accuracy of the myriad facts that frame the contexts and the science in the pages that follow.

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Natalie Foster