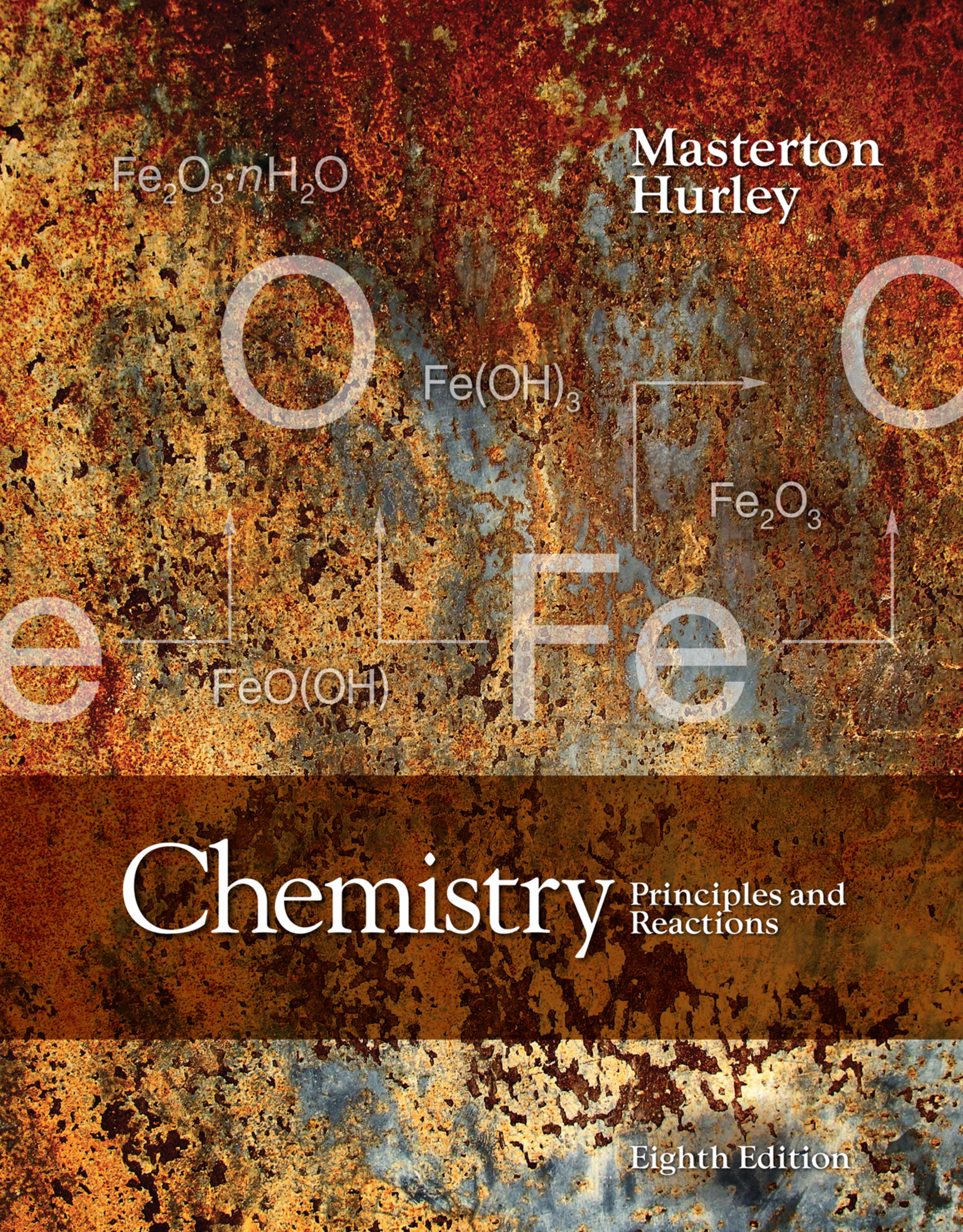
Masterton
Hurley



Chemistry
Principles and Reactions

Eighth Edition

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Principles and Reactions

William L. Masterton
University of Connecticut

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Eighth Edition
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To Jim, Joe, and Regina

They also serve who only stand and wait.
—John Milton
On His Blindness

Brief Contents

Contents

Preface

It is always difficult for an author to praise the virtues of one's own book. I could tell the instructors that the book is so inspiring that students will be turned on to chemistry with little or no effort on the instructor's part. I doubt you would believe that. I could also tell you that the text is so clearly written, so attuned to the students in the twenty-first century that your students will learn chemistry with little or no effort on their part. You certainly would not believe that. I can tell you that the two goals in writing this edition have been to make it as clear and as interesting as possible. I hope you believe that, because it is true.

Today's freshmen are quite different from those of a few years ago. Text messaging and Twitter™ have strongly influenced sentence length and structure. In current writing and conversation, short sentences or sentence fragments convey straight-to-the-point information. Multimedia presentations are a way of life. This edition, like the seventh, is written to be fully in tune with today's technology and speech.

Why Write a Short Book?

Rising tuition costs, depleted forests, and students' aching backs have kept me steadfast in my belief that it should be possible to cover a text completely (or at least *almost* completely) in a two-semester course. The students (and their parents) justifiably do not want to pay for 1000-page books with material that is never discussed in the courses taught with those texts.

The common perception is that a short book is a low-level book. I believe, however, that treating general concepts in a concise way can be done without sacrificing depth, rigor, or clarity. The criterion for including material continues to be its importance and relevance to the student, not its difficulty. To achieve this, the following guidelines are used.

1. Eliminate repetition and duplication wherever possible. Like its earlier editions, this text uses
   - Only one method for balancing redox reactions, the half-equation method introduced in Chapter 17.
   - Only one way of working gas-law problems, using the ideal gas law in all cases (Chapter 5).
   - Only one way of calculating ΔH (Chapter 8), using enthalpies of formation.
   - Only one equilibrium constant for gas-phase reactions (Chapter 12), the thermodynamic constant K, often referred to as K_p. This simplifies not only the treatment of gaseous equilibrium but also the discussion of reaction spontaneity (Chapter 16) and electrochemistry (Chapter 17).
2. Relegate to the Appendices or Beyond the Classroom essays topics ordinarily covered in longer texts. Items in this category include
   - MO (molecular orbital) theory (Appendix 4). Experience has shown (and continues to show) that although this approach is important to chemical bonding, most general chemistry students do not understand it but only memorize the principles discussed in the classroom.
   - Nomenclature of organic compounds. This material is of little value in a beginning course and is better left to a course in organic chemistry.
   - Qualitative analysis. This is summarized in a few pages in an essay in Chapter 15 in the Beyond the Classroom section. An extended discussion of the qualitative scheme and the chemistry behind it belongs in a laboratory manual, not a textbook.
   - Biochemistry. This material is traditionally covered in the last chapter of general chemistry texts. Although there are several biochemical topics included in the text (among them a discussion of heme in Chapter 19 and

carotenoids in Chapter 6), an entire chapter is not devoted to biochemistry. Interesting as this material is, it requires a background in organic chemistry that first-year students lack.

3. Avoid superfluous asides, applications to the real world, or stories about scientists in the exposition of principles. There are many applications incorporated in the context of problems and some of the exposition of general principles. In general, however, a bare-bones approach is used. Students can easily be distracted by interesting but peripheral tidbits while they are striving hard to understand the core concepts. Favorite real-world applications and personal stories about scientists are in separate sections, Beyond the Classroom and Chemistry: The Human Side. Students say that they read these two sections first and that these are the parts of the book that "we really enjoy the most." (Talk about faint praise!) They do admit to enjoying the marginal notes too.

What Changes Have Been Made?

The eighth edition has not been as radically changed as the seventh. I talked to students, instructors, and TAs and listened to suggestions and complaints.

While all the changes made to the seventh edition were enthusiastically received, there were areas where making small changes would make them better. For the eighth edition, the following changes were made:

- The Example format has been revised, so that the strategy, analysis, and solution follow each part of the example. The most common comment was: "Show me first how to do part (a) before asking me about part (b)."
- More flowcharts have been added. There was unanimous support and requests for more of them. We revised some of the existing ones and added a few more.
- The discussion of balancing redox equations has been moved from Chapter 4 to Chapter 17. Instructors comment that they have had to reintroduce redox equations in Chapter 17 and treat it like new material. Students and TAs both agree that Chapter 4 is a dense and heavy chapter. Thus, redox reactions are treated in Chapter 4 only as far as stoichiometric calculations are involved. Balanced equations are provided for these reactions.

Detailed List of Changes by Chapter

Global Changes
- Changes in about 25% of the topical end-of-chapter problems
- Almost all of the summary problems have either been revised or are completely new to this edition
- Revised artwork with enhanced labeling and several new photos
- Almost all of the chapter opening art is new

Chapter 1
- Redrawn flowchart for matter classification
- New Examples 1.3 and 1.6

Chapter 2
- Redrawn Figure 2.5 (Rutherford Experiment)
- New Examples 2.1 and 2.5
- New Beyond the Classroom essay on the origin of some elements' names
- Redrawn flowchart for naming molecular compounds

Chapter 3
- New Examples 3.1 (b), 3.5, 3.7, and 3.11

Chapter 4
- Redrawn Figure 4.1
- New Example 4.5
- Revised discussion of redox reactions excluding balancing of half and complete reactions
- New flowchart for determining oxidation number
- New figure summarizing differences between oxidation and reduction
- New Beyond the Classroom essay on antacids

Chapter 5
- New Figure 5.6
- New Examples 5.2 and 5.4 (a) and (b)
- Discussion of the relation of time and molar mass to the rate of effusion

continued

Chapter 6
- New Example 6.1
- New discussion on electron capacities for principal levels and subshells
- New table summarizing electron capacities in principal levels and subshells

Chapter 7
- New Figure 7.5

Chapter 8
- Replaced Figure 8.10 with a new table

Chapter 9
- Figure 9.3 redrawn

Chapter 10
- New Examples 10.9 and 10.10
- Figure 10.14 redrawn

Chapter 11
- New Figure 11.4
- Example 11.3 redone

Chapter 12
- Examples 12.1 and 12.6 rewritten

Chapter 13
- New Example 13.3
- Examples 13.2 and 13.9 rewritten
- Revised Figure 13.14

Chapter 14
- Example 14.3 rewritten

Chapter 15
- Example 15.3 rewritten
- Example 15.1 reclassified as nongraded
- Examples 15.8 and 15.9 reclassified as graded

Chapter 16
- Example 16.2 rewritten

Chapter 17
- New Section 17.1 on balancing redox half and complete reactions
- Review of oxidation, reduction, oxidizing agents, and reducing agents
- New schematic on balancing half-reactions
- New Examples 17.1 and 17.2

Chapter 18
- New Table 18.1—summary of different modes of radioactive decay
- New paragraph about SPECT (single proton emission computer tomography) scans

Chapters 19–23
No changes

Alternate Editions

***Chemistry: Principles and Reactions*, Eighth Edition Hybrid Version with Access (24 months) to OWLv2 with MindTap Reader**

ISBN: 978-1-305-08215-1

This briefer, paperbound version of *Chemistry: Principles and Reactions*, Eighth Edition does not contain the end-of-chapter problems, which can be assigned in OWLv2, the online homework and learning system for this book. Access to OWLv2 and the MindTap Reader eBook is included with the Hybrid version. The MindTap Reader is the full version of the text, with all end-of-chapter questions and problem sets.

Supporting Materials

Please visit **http://www.cengage.com/chemistry/masterton/CPAR8e** for information about student and instructor resources for this text, including custom versions and laboratory manuals.

Acknowledgments

Many people who have used this book—instructors, teaching assistants, students, and former students now teaching general chemistry—have e-mailed, written, and called with suggestions on how to improve the exposition. I am grateful to them all.

Reviewers who have helped in the preparation of this edition include the following:

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This edition would not have been possible without the superb guidance of my content developer, Ed Dodd. He was a real gift. He smoothed rough patches and demanded perfection from everyone on the team. It was a real pleasure working with him.

Many people worked on the editorial and production team for this text. They took pages of manuscript, rough ideas, crude sketches, and long wish lists and put them together to create this edition. They prodded, cajoled, and set impossible deadlines. They are:

Mary Finch, Product Director
Maureen Rosener, Product Manager
Lisa Lockwood, Product Manager
Peter McGahey, Managing Developer
Elizabeth Woods, Associate Content Developer
Karolina Kiwak, Product Assistant
Lisa Weber, Media Developer
Brendan Killion, Media Developer
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Jill Traut, Project Manager at MPS Limited
Dhanalakshmi Singaravelu, and Padmapriya Soundararajan, Image Researchers at Lumina Datamatics
Pinky Subi, Text Researcher at Lumina Datamatics

One person who does not belong to any team deserves special recognition. Jim Hurley picked up the slack when time was short, deadlines were imminent, and the list of tasks was long. He listened to endless complaints and commiserated. Thank you once again for continuing on this journey with me.

Cecile N. Hurley
University of Connecticut
Storrs, CT
November 2014

To the Student

You've probably already heard a lot about your general chemistry course. Many think it is more difficult than other courses. There may be some justification for that opinion. Besides having its very own specialized vocabulary, chemistry is a quantitative science—which means that you need mathematics as a tool to help you understand the concepts. As a result, you will probably receive a lot of advice from your instructor, teaching assistant, and fellow students about how to study chemistry. We would, however, like to acquaint you with some of the learning tools in this text. They are described in the pages that follow.

Learning Tools in *Chemistry: Principles and Reactions,* Eighth Edition

Examples

In a typical chapter, you will find ten or more examples each designed to illustrate a particular principle. These examples are either general (green bars), graded (purple bars), or conceptual (blue bars). These have answers, screened in color. They are presented in a two-column format. Most of them contain three parts:

- **Analysis**, which lists
  1. The information given.
  2. The information implied—information not directly stated in the problem but data that you can find elsewhere.
  3. What is asked for.
- **Strategy**

  This part gives you a plan to follow in solving the problem. It may lead you through a schematic pathway or remind you of conversion factors you have to consider or suggest equations that are useful.
- **Solution**

  This portion shows in a stepwise manner how the strategy given is implemented.
- Many of the examples end with a section called **End Points.** These are either checks on the reasonableness of your answer or relevant information obtained from the problem.

You should find it helpful to get into the habit of working all problems this way.

EXAMPLE

Calculate the wavelength in nanometers of the line in the Balmer series that results from the transition $\mathbf{n} = 4$ to $\mathbf{n} = 2$.

ANALYSIS	
Information given:	$\mathbf{n} = 2$; $\mathbf{n} = 4$
Information implied:	speed of light (2.998×10^{8} m/s) Rydberg constant (2.180×10^{-18} J) Planck constant (6.626×10^{-34} J · s)
Asked for:	wavelength in nm

continued

STRATEGY	
1. Substitute into Equation 6.4 to find the frequency due to the transition. $$\nu = \frac{R_H}{h}\left(\frac{1}{\mathbf{n}_{lo}^2} - \frac{1}{\mathbf{n}_{hi}^2}\right)$$ Use the lower value for $\mathbf{n}$ as $\mathbf{n}_{lo}$ and the higher value for $\mathbf{n}_{hi}$.	
2. Use Equation 6.1 to find the wavelength in meters and then convert to nanometers.	
SOLUTION	
1. Frequency	$\nu = \frac{2.180 \times 10^{-18}\ \text{J}}{6.626 \times 10^{-34}\ \text{J}\cdot\text{s}}\left(\frac{1}{(2)^2} - \frac{1}{(4)^2}\right) = 6.169 \times 10^{14}\ \text{s}^{-1}$
2. Wavelength	$\lambda = \frac{2.998 \times 10^{8}\ \text{m/s}}{6.169 \times 10^{14}\ \text{s}^{-1}} \times \frac{1\ \text{nm}}{1 \times 10^{-9}\ \text{m}} = 486.0\ \text{nm}$
END POINT	
Compare this value with that listed in Table 6.2 for the second line of the Balmer series.	

Graded Examples

Throughout the text, you will encounter special *graded* examples. Note that they are the problems with the purple bars. A typical graded example looks like the following:

EXAMPLE — GRADED

For the reaction

$$A + 2B \longrightarrow C$$

determine

(a) the number of moles of A required to react with 5.0 mol of B.

(b) the number of grams of A required to react with 5.0 g of B.

(c) the volume of a 0.50 *M* solution of A required to react with 5.0 g of B.

(d) the volume of a 0.50 *M* solution of A required to react with 25 mL of a solution that has a density of 1.2 g/mL and contains 32% by mass of B.

There are two advantages to working a graded example:

1. By working parts (a) through (d) in succession, you can see how many different ways there are to ask a question about mass relations in a reaction. That should cushion the shock should you see only part (d) in an exam.
2. The parts of the graded example do not just progress from an easy mass relations question to a more difficult one. The value of the graded example is that the last question *assumes the ability to answer the earlier ones*. You may be able to answer parts (a) and (b) with a limited understanding of the material, but to answer part (d) you need to have mastered the material.

Use the graded example as you review for exams. Try to skip the earlier parts [in this case (a), (b) and (c)] and go directly to the last part (d). If you can solve (d), you do not need to try (a), (b), and (c); you know how to do them. If you can't, then try (c) to see where you may have a problem. If you can't do (c), then try (b). As a last resort, start at (a) and work your way back through (d).

Marginal Notes

Sprinkled throughout the text are a number of short notes in the margin. Many of these are of the "now, hear this" variety, others are mnemonics, and still others make points that we forgot to put in the text. (These were contributed by your fellow students.) Some—probably fewer than we think—are supposed to be humorous.

Chemistry: The Human Side

Throughout the text, short biographies of some of the pioneers of chemistry appear in sections with this heading. They emphasize not only the accomplishments of these individuals but also their personalities.

Chemistry: Beyond the Classroom

Each chapter contains a Beyond the Classroom feature. It is a self-contained essay that illustrates a current example either of chemistry in use in the world or an area of chemical research. It does not intrude into the explanation of the concepts, so it won't distract you. But we promise that those essays—if you read them—will make you more scientifically literate.

Chapter Highlights

At the end of each chapter, you will find a brief review of its concepts. A review is always helpful not only to refresh yourself about past material but also to organize your time and notes when preparing for an examination. The chapter highlights include

- The ***Key Terms*** in the chapter. If a particular term is unfamiliar, refer to the index at the back of the book. You will find the term in the glossary that is incorporated in the index and also the pages in the text where it appears (if you need more explanation).
- The ***Key Concepts*** and ***Key Equations*** introduced in the chapter. These are indexed to the corresponding examples and end-of-chapter problems. End-of-chapter problems available on OWLv2 are also cross-referenced. If you have trouble working a particular problem here, it may help to go back and reread the example that covers the same concept.

Summary Problem

Each chapter is summarized by a multistep problem that covers all or nearly all of the key concepts in the chapter. You can test your understanding of the chapter by working this problem. A major advantage of the summary problems is that they tie together many different ideas, showing how they correlate with one another. An experienced general chemistry professor always tells his class, "If you can answer the summary problem without help, you are ready for a test on its chapter."

Questions and Answers

At the end of each chapter is a set of questions and problems that your instructor may assign for homework. They are also helpful in testing the depth of your knowledge about the chapter. These sets include

- Conceptual problems that test your understanding of principles. A calculator is not (or should not be) necessary to answer these questions.
- Questions that test your knowledge of the specialized vocabulary that chemists use (e.g., write the names of formulas, write the chemical equation for a reaction that is described).
- Quantitative problems that require a calculator and some algebraic manipulations.

Classified problems start the set and are grouped by type under a particular heading that indicates the section and/or topic from the chapter that they address. The classified problems occur in matched pairs, so the second member illustrates the same principle as the first. This allows you more than one opportunity to test yourself. The second problem (whose number is even) is answered in Appendix 5. If your instructor assigns the odd problems without answers for homework, wait until the problem solution is discussed and solve the even problem to satisfy yourself that you understand how to solve the problem of that type.

Each chapter also contains a smaller number of **Unclassified** problems, which may involve more than one concept, including, perhaps, topics from a preceding chapter.

The section of **Challenge** problems presents problems that may require extra skill and/or insight and effort. They are all answered in Appendix 5.

Even-numbered questions and Challenge Problems answered in Appendix 5 have fully worked solutions available in the *Student Solutions Manual*. Please visit **http://www.cengage.com/chemistry/masterton/CPAR8e** for information about the *Student Solutions Manual*.

Appendices

The appendices at the end of the book provide not only the answers to the even-numbered problems but also additional materials you may find useful. Among them are

- Appendix 1, which includes a review of SI base units as well as tables of thermodynamic data and equilibrium constants.
- Appendix 3, which contains a mathematical review touching on just about all the mathematics you need for general chemistry. Exponential notation and logarithms (natural and base 10) are emphasized.

Other Resources to Help You Pass Your General Chemistry Course

Besides the textbook, several other resources are available to help you study and master general chemistry concepts. Please visit **http://www.cengage.com/chemistry/masterton/CPAR8e** for information about student resources for this text, including custom versions and laboratory manuals.

Matter and Measurements

1

Museum of the History of Science in Oxford, England

The painting shows measuring instruments used in the Middle Ages. We still use many of them today.

There is measure in everything.

—HORACE

Chapter Outline

Almost certainly, this is your first college course in chemistry; perhaps it is your first exposure to chemistry at any level. Unless you are a chemistry major, you may wonder why you are taking this course and what you can expect to gain from it. To address that question, it is helpful to look at some of the ways in which chemistry contributes to other disciplines.

If you're planning to be an engineer, you can be sure that many of the materials you will work with have been synthesized by chemists. Some of these materials are organic (carbon-containing). They could be familiar plastics like polyethylene (Chapter 23) or the more esoteric plastics used in unbreakable windows and nonflammable clothing. Other materials, including metals (Chapter 20) and semiconductors, are inorganic in nature.

Perhaps you are a health science major, looking forward to a career in medicine or pharmacy. If so, you will want to become familiar with the properties of aqueous solutions (Chapters 4, 10, 14, and 16), which include blood and other body fluids. Chemists today are involved in the synthesis of a variety of life-saving products. These range from drugs used in chemotherapy (Chapter 19) to new antibiotics used against resistant microorganisms.

Beyond career preparation, an objective of a college education is to make you a better-informed citizen. In this text, we'll look at some of the chemistry-related topics that make the news:

- depletion of the ozone layer (Chapter 11).
- alternative sources of fuel (Chapter 17).
- the pros and cons of nuclear power (Chapter 18).

Another goal of this text is to pique your intellectual curiosity by trying to explain the chemical principles behind such recent advances as

- "self-cleaning" windows (Chapter 1).
- "the ice that burns" (Chapter 3).
- "maintenance-free" storage batteries (Chapter 17).
- "chiral" drugs (Chapter 22).

We hope that when you complete this course you too will be convinced of the importance of chemistry in today's world. We should, however, caution you on one point.

Chemistry deals with the properties and reactions of substances.

Although we will talk about many of the applications of chemistry, *our main concern will be with the principles that govern chemical reactions.* Only by mastering those principles will you understand the basis of the applications mentioned above.

This chapter begins the study of chemistry by

- considering the different types of matter: pure substances versus mixtures, elements versus compounds (Section 1-1).
- looking at the kinds of measurements fundamental to chemistry, the uncertainties associated with those measurements, and a method to convert measured quantities from one unit to another (Section 1-2).
- focusing on certain physical properties, including density and water solubility, which can be used to identify substances (Section 1-3).

1-1 Matter and Its Classifications

Matter is anything that has mass and occupies space. It can be classified either with respect to its physical phases or with respect to its composition (Figure 1.1).

The three phases of matter are solid, liquid, and gas. A **solid** has a fixed shape and volume. A **liquid** has a fixed volume but is not rigid in shape; it takes the shape of its container. A **gas** has neither a fixed volume nor a shape. It takes on both the shape and the volume of its container.

Matter can also be classified with respect to its composition:

- pure substances, each of which has a fixed composition and a unique set of properties.
- mixtures, composed of two or more substances.

Pure substances are either elements or compounds (Figure 1.1), whereas mixtures can be either homogeneous or heterogeneous.

Elements

An **element** is a type of matter that cannot be broken down into two or more pure substances. There are 118 known elements, of which 91 occur naturally.

Many elements are familiar to all of us. The charcoal used in outdoor grills is nearly pure carbon. Electrical wiring, jewelry, and water pipes are often made from copper, a metallic element. Another such element, aluminum, is used in many household utensils.

Some elements come in and out of fashion, so to speak. Sixty years ago, elemental silicon was a chemical curiosity. Today, ultrapure silicon has become the basis for the multibillion-dollar semiconductor industry. Lead, on the other hand, is an element moving in the other direction. A generation ago it was widely used to make paint pigments, plumbing connections, and gasoline additives. Today, because of the toxicity of lead compounds, all of these applications have been banned in the United States.

In chemistry, an element is identified by its **symbol**. This consists of one or two letters, usually derived from the name of the element. Thus the symbol for carbon is C; that for aluminum is Al. Sometimes the symbol comes from the Latin name of the element or one of its compounds. The two elements copper and mercury, which were known in ancient times, have the symbols Cu *(cuprum)* and Hg *(hydrargyrum)*.

Table 1.1 (p. 4) lists the names and symbols of several elements that are probably familiar to you. In either free or combined form, they are commonly found in the laboratory or in commercial products. The abundances listed measure the relative amount of each element in the earth's crust, the atmosphere, and the oceans.

Curiously, several of the most familiar elements are really quite rare. An example is mercury, which has been known since at least 500 B.C., even though its abundance is only 0.00005%. It can easily be prepared by heating the red mineral cinnabar (Figure 1.2, p. 4).

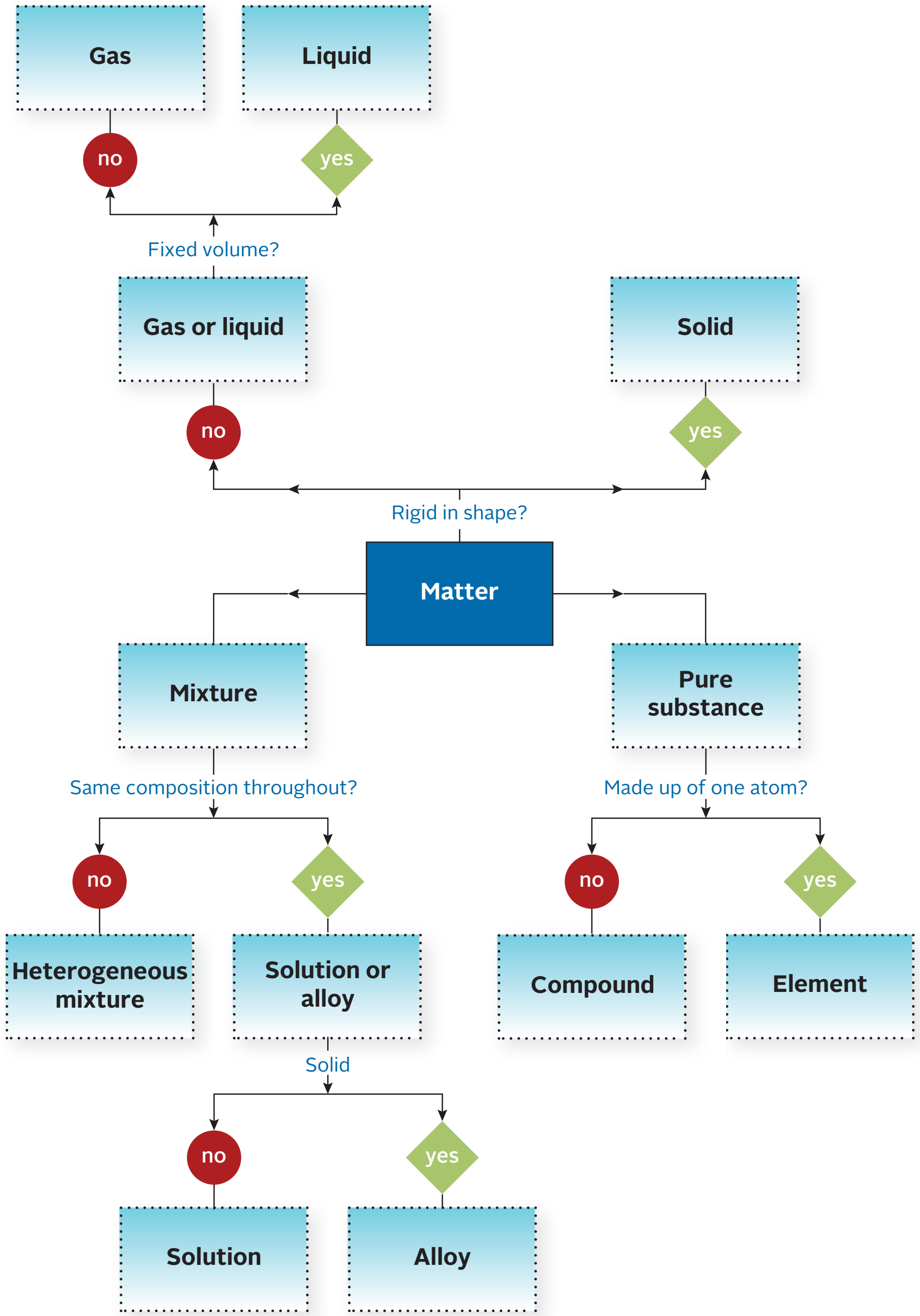


Figure 1.1 **Classification of matter into solid, liquid, and gas.**

Mercury is the only metal that is a liquid at room temperature. It is also one of the densest elements. Because of its high density, mercury was the liquid extensively used in thermometers and barometers. In the 1990s all instruments using mercury were banned because of environmental concerns. Another useful quality of mercury is its ability to dissolve many metals, forming solutions (amalgams). A silver-mercury-tin amalgam is still used to fill tooth cavities, but many dentists now use tooth-colored composites because they adhere better and are aesthetically more pleasing.

Mercury thermometers, both for laboratory and clinical use, have been replaced by digital ones.

In contrast, aluminum (abundance = 7.5%), despite its usefulness, was little more than a chemical curiosity until about a century ago. It occurs in combined form in clays and rocks, from which it cannot be extracted. In 1886 two young chemists, Charles Hall in the United States and Paul Héroult in France, independently worked out a process for extracting aluminum from a relatively rare ore,

Table 1.1 Some Familiar Elements with Their Percentage Abundances

Element	Symbol	Percentage Abundance	Element	Symbol	Percentage Abundance
Aluminum	Al	7.5	Manganese	Mn	0.09
Bromine	Br	0.00025	Mercury	Hg	0.00005
Calcium	Ca	3.4	Nickel	Ni	0.010
Carbon	C	0.08	Nitrogen	N	0.03
Chlorine	Cl	0.2	Oxygen	O	49.4
Chromium	Cr	0.018	Phosphorus	P	0.12
Copper	Cu	0.007	Potassium	K	2.4
Gold	Au	0.0000005	Silicon	Si	25.8
Hydrogen	H	0.9	Silver	Ag	0.00001
Iodine	I	0.00003	Sodium	Na	2.6
Iron	Fe	4.7	Sulfur	S	0.06
Lead	Pb	0.0016	Titanium	Ti	0.56
Magnesium	Mg	1.9	Zinc	Zn	0.008

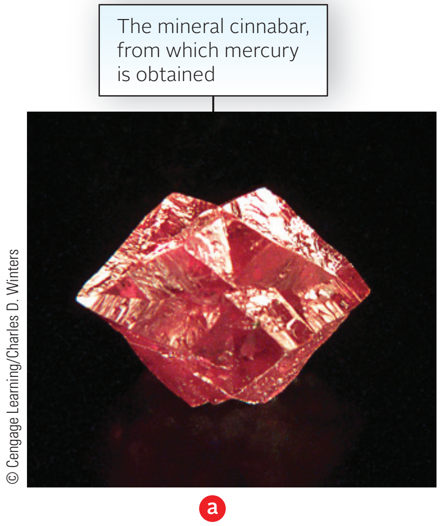


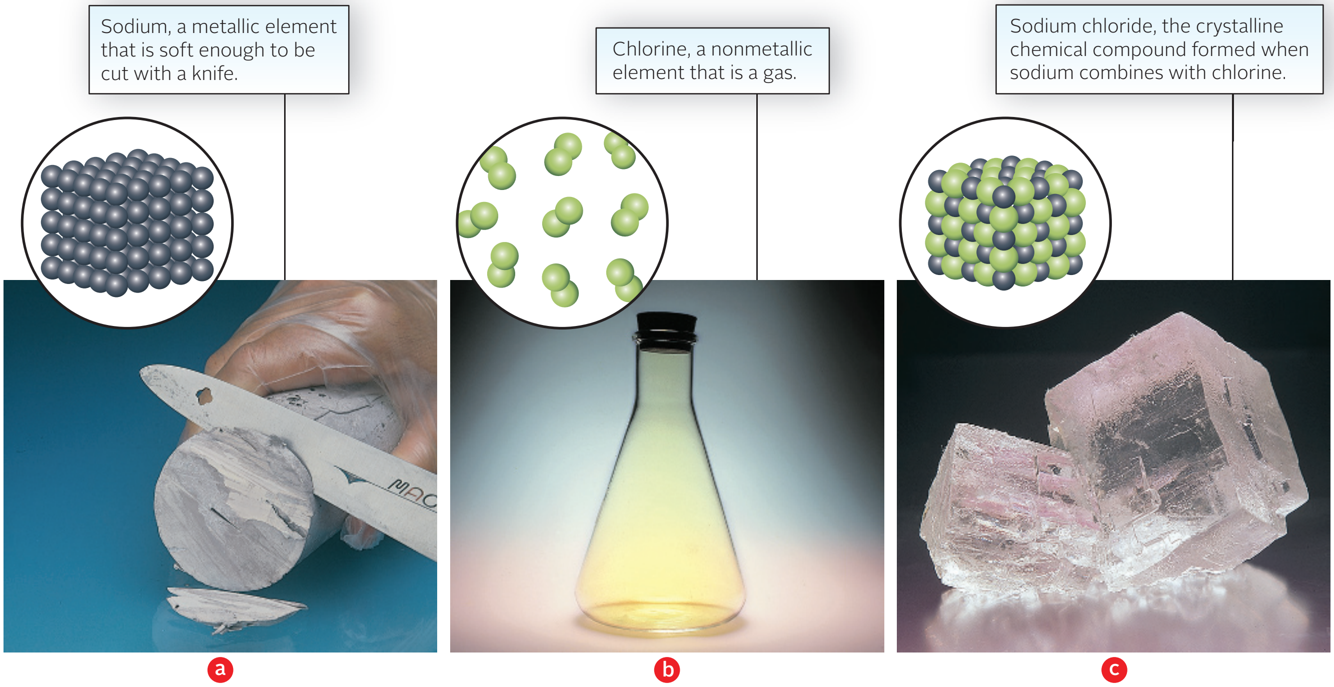
Figure 1.2 Cinnabar and mercury.

bauxite. That process is still used today to produce the element. By an odd coincidence, Hall and Héroult were born in the same year (1863) and died in the same year (1914).

Compounds

A **compound** is a pure substance that contains more than one element. Water is a compound of hydrogen and oxygen. The compounds methane, acetylene, and naphthalene all contain the elements carbon and hydrogen, in different proportions.

Compounds have fixed compositions. That is, a given compound always contains the same elements in the same percentages by mass. A sample of pure water contains precisely 11.19% hydrogen and 88.81% oxygen. In contrast, mixtures can vary in composition. For example, a mixture of hydrogen and oxygen might contain 5, 10, 25, or 60% hydrogen, along with 95, 90, 75, or 40% oxygen.



© Cengage Learning/Charles D. Winters (Photos a–c)

Figure 1.3 **Sodium, chlorine, and sodium chloride.**

The properties of compounds are usually very different from those of the elements they contain. Ordinary table salt, sodium chloride, is a white, unreactive solid. As you can guess from its name, it contains the two elements sodium and chlorine. Sodium (Na) is a shiny, extremely reactive metal. Chlorine (Cl) is a poisonous, greenish-yellow gas. Clearly, when these two elements combine to form sodium chloride, a profound change takes place (Figure 1.3).

Many different methods can be used to resolve compounds into their elements. Sometimes, but not often, heat alone is sufficient. Mercury(II) oxide, a compound of mercury and oxygen, decomposes to its elements when heated to 600°C. Joseph Priestley, an English chemist, discovered oxygen more than 200 years ago when he carried out this reaction by exposing a sample of mercury(II) oxide to an intense beam of sunlight focused through a powerful lens. The mercury vapor formed is a deadly poison. Sir Isaac Newton, who distilled large quantities of mercury in his laboratory, suffered the effects in his later years.

Another method of resolving compounds into elements is *electrolysis*, which involves passing an electric current through a compound, usually in the liquid state. By electrolysis it is possible to separate water into the gaseous elements hydrogen and oxygen. Several decades ago it was proposed to use the hydrogen produced by electrolysis to raise the *Titanic* from its watery grave off the coast of Newfoundland. It didn't work.

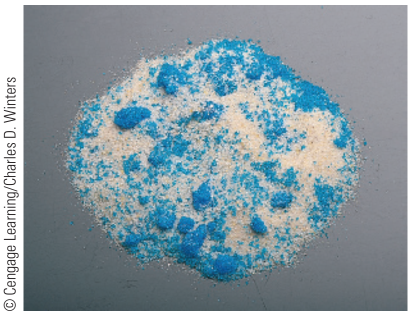


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Figure 1.4 **A heterogeneous mixture of copper sulfate crystals (blue) and sand.**

Mixtures

A **mixture** ◀ *i* contains two or more substances combined in such a way that each substance retains its chemical identity. When you shake copper sulfate with sand (Figure 1.4), the two substances do not react with one another. In contrast, when sodium is exposed to chlorine gas, a new compound, sodium chloride, is formed.

i Most materials you encounter are mixtures.

There are two types of mixtures:

1. Homogeneous or uniform mixtures are ones in which the composition is the same throughout. Another name for a homogeneous mixture is a **solution,** which is made up of a solvent, usually taken to be the substance present in largest amount, and one or more solutes. Most commonly, the **solvent** is a liquid, whereas

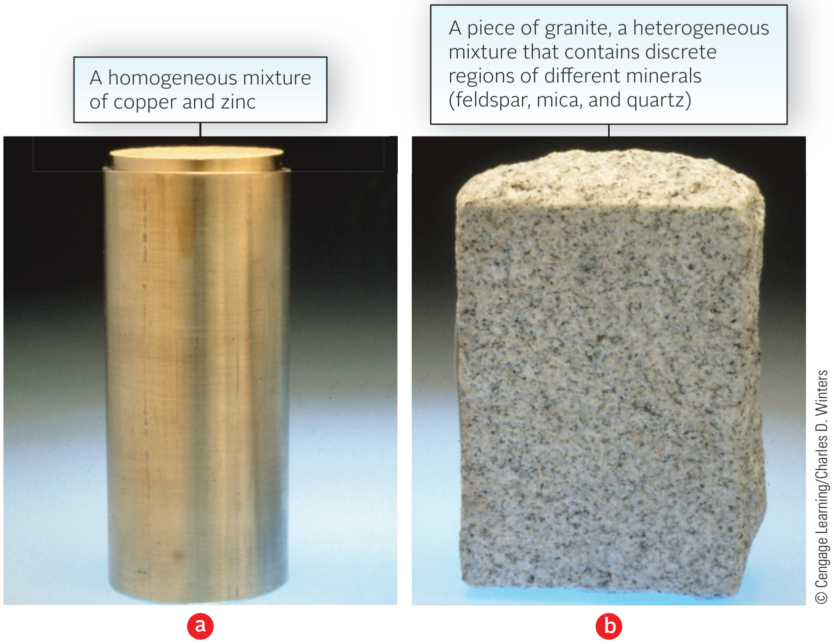

Figure 1.5 **Two mixtures.**

a **solute** may be a solid, liquid, or gas. Soda water is a solution of carbon dioxide (solute) in water (solvent). Seawater is a more complex solution in which there are several solid solutes, including sodium chloride; the solvent is water. It is also possible to have solutions in the solid state. Brass (Figure 1.5a) is a solid solution containing the two metals copper (67%–90%) and zinc (10%–33%).

All gaseous mixtures, including air, are solutions.

2. Heterogeneous or nonuniform mixtures are those in which the composition varies throughout. Most rocks fall into this category. In a piece of granite (Figure 1.5b), several components can be distinguished, differing from one another in color.

Many different methods can be used to separate the components of a mixture from one another. A couple of methods that you may have carried out in the laboratory are

- **filtration**, used to separate a heterogeneous solid-liquid mixture. The mixture is passed through a barrier with fine pores, such as filter paper. Copper sulfate, which is water-soluble, can be separated from sand by shaking with water. On filtration the sand remains on the paper and the copper sulfate solution passes through it.
- *distillation,* used to resolve a homogeneous solid-liquid mixture. The liquid vaporizes, leaving a residue of the solid in the distilling flask. The liquid is obtained by condensing the vapor. Distillation can be used to separate the components of a water solution of copper sulfate (Figure 1.6).

A more complex but more versatile separation method is *chromatography,* a technique widely used in teaching, research, and industrial laboratories to separate all kinds of mixtures. This method takes advantage of differences in solubility and/or extent of adsorption on a solid surface. In *gas-liquid chromatography,* a mixture of volatile liquids and gases is introduced into one end of a heated glass tube. As little as one microliter (10^{-6} L) of sample may be used. The tube is packed with an inert solid whose surface is coated with a viscous liquid. An unreactive "carrier gas," often helium, is passed through the tube. The components of the sample gradually separate as they vaporize into the helium or condense into the viscous liquid. Usually the more volatile fractions move faster and emerge first; successive fractions activate a detector and recorder.

GLC is a favorite technique in the forensics labs of many TV shows.

Gas-liquid chromatography (GLC) (Figure 1.7) finds many applications outside the chemistry laboratory. If you've ever had an emissions test on the exhaust system of your car, GLC was almost certainly the analytical method used. Pollutants such as carbon monoxide and unburned hydrocarbons appear as peaks on a graph. A computer determines the areas under these peaks, which are proportional to the concentrations of pollutants, and prints out a series of numbers that tells the inspector whether your car passed or failed the test. Many of the techniques used to test people for drugs (marijuana, cocaine, and others) or alcohol also make use of gas-liquid chromatography.

Ultra-high-speed gas chromatography (GC) fitted with an odor sensor is a powerful tool for analyzing the chemical vapors produced by explosives or other chemical or biological weapons.

In this section we will look at four familiar properties that you will almost certainly measure in the laboratory: length, volume, mass, and temperature. Other physical and chemical properties will be introduced in later chapters as they are needed.

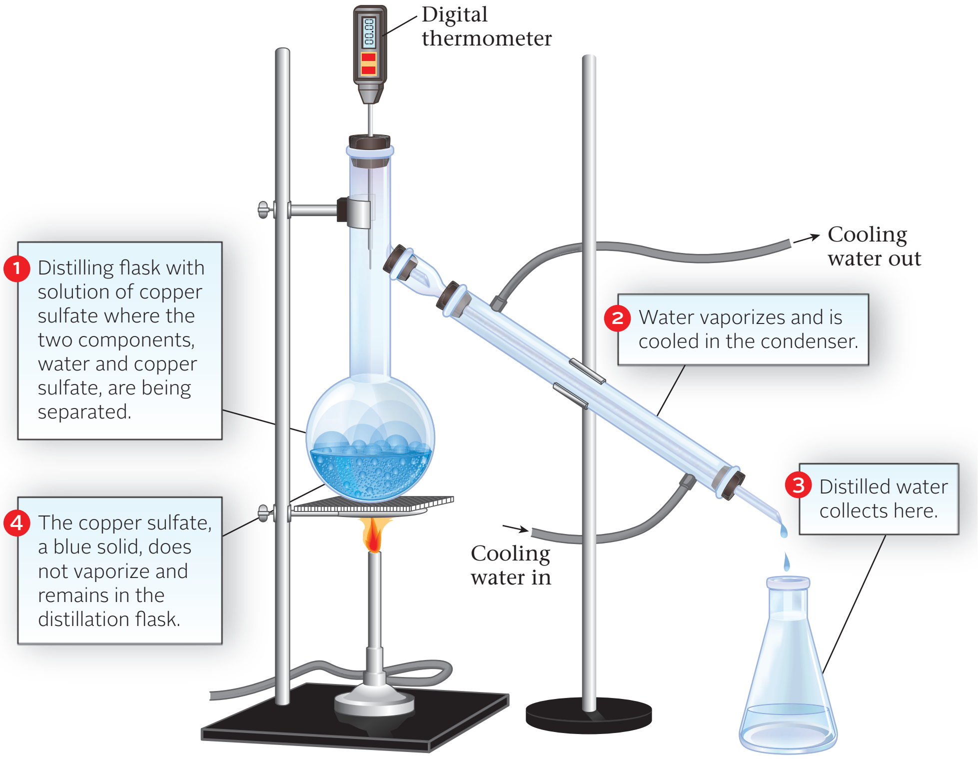

Figure 1.6 Apparatus for a simple distillation.

1-2 Measurements

Chemistry is a quantitative science. The experiments that you carry out in the laboratory and the calculations that you perform almost always involve measured quantities with specified numerical values. Consider, for example, the following set of directions for the preparation of aspirin (measured quantities are shown in italics).

> Add *2.0 g* of salicylic acid, *5.0 mL* of acetic anhydride, and *5 drops* of *85%* H_3PO_4 to a 50-mL Erlenmeyer flask. Heat in a water bath at *75°C* for *15 minutes*. Add cautiously *20 mL* of water and transfer to an ice bath at *0°C*. Scratch the inside of the flask with a stirring rod to initiate crystallization. Separate aspirin from the solid-liquid mixture by filtering through a Buchner funnel *10 cm* in diameter.

Scientific measurements are expressed in the **metric system**. As you know, this is a decimal-based system in which all of the units of a particular quantity are related to one another by factors of 10. The more common prefixes used to express these factors are listed in Table 1.2.

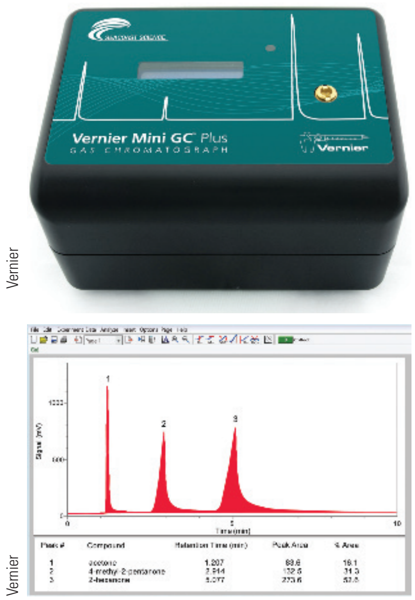

Figure 1.7 Mass spectrometer and gas chromatograph. Airport security uses these instruments to separate mixtures (chromatograph) and to detect the presence of nitrogen containing explosives (mass spectrograph). Larger versions of these instruments are used to scan checked baggage.

Table 1.2 Metric Prefixes

Factor	Prefix	Abbreviation	Factor	Prefix	Abbreviation
10^6	mega	M	10^{-3}	**milli**	m
10^3	**kilo**	k	10^{-6}	micro	μ
10^{-1}	deci	d	10^{-9}	**nano**	n
10^{-2}	**centi**	c	10^{-12}	pico	p

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Figure 1.8 Measuring volume. A buret (*left*) delivers an accurately measured variable volume of liquid. A pipet (*right*) delivers a fixed volume (e.g., 25.00 mL) of liquid.

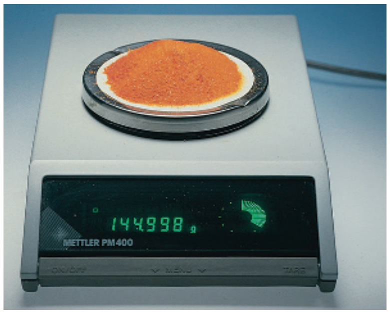

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Figure 1.9 Weighing a solid. The solid sample plus the paper on which it rests weighs 144.998 g. The pictured balance is a single-pan analytical balance.

Writing "m" in upper case or lower case makes a big difference.

Many countries still use degrees centigrade.

Instruments and Units

The standard unit of *length* in the metric system is the meter, which is a little larger than a yard. The meter was originally intended to be 1/40,000,000 of the earth's meridian that passes through Paris. It is now defined as the distance light travels in a vacuum in 1/299,792,458 of a second.

Other units of length are expressed in terms of the meter, using the prefixes listed in Table 1.2. You are familiar with the centimeter, the millimeter, and the kilometer:

$$1\ \text{cm} = 10^{-2}\ \text{m} \qquad 1\ \text{mm} = 10^{-3}\ \text{m} \qquad 1\ \text{km} = 10^{3}\ \text{m}$$

The dimensions of very tiny particles are often expressed in nanometers:

$$1\ \text{nm} = 10^{-9}\ \text{m}$$

Volume is most commonly expressed in one of three units

- cubic centimeters $1\ \text{cm}^3 = (10^{-2}\ \text{m})^3 = 10^{-6}\ \text{m}^3$
- liters (L) $1\ \text{L} = 10^{-3}\ \text{m}^3 = 10^{3}\ \text{cm}^3$
- milliliters (mL) $1\ \text{mL} = 10^{-3}\ \text{L} = 10^{-6}\ \text{m}^3$

Notice that a milliliter is equal to one cubic centimeter:

$$1\ \text{mL} = 1\ \text{cm}^3$$

The device most commonly used to measure volume in general chemistry is the graduated cylinder. A pipet or buret (Figure 1.8) is used when greater accuracy is required. A pipet is calibrated to deliver a fixed volume of liquid—for example, 25.00 mL—when filled to the mark and allowed to drain. Different volumes can be delivered accurately by a buret, perhaps to ±0.01 mL.

In the metric system, *mass* is most commonly expressed in grams, kilograms, or milligrams:

$$1\ \text{g} = 10^{-3}\ \text{kg} \qquad 1\ \text{mg} = 10^{-3}\ \text{g}$$

This book weighs about 1.5 kg. The megagram, more frequently called the *metric ton,* is

$$1\ \text{Mg} = 10^{6}\ \text{g} = 10^{3}\ \text{kg}$$

Properly speaking, there is a distinction between mass and weight. *Mass* is a measure of the amount of matter in an object; *weight* is a measure of the gravitational force acting on the object. Chemists often use these terms interchangeably; we determine the mass of an object by "weighing" it on a balance (Figure 1.9).

Temperature is the factor that determines the direction of heat flow. When two objects at different temperatures are placed in contact with one another, heat flows from the one at the higher temperature to the one at the lower temperature.

Thermometers used in chemistry are marked in degrees **Celsius** (referred to as degrees centigrade until 1948). On this scale, named after the Swedish astronomer Anders Celsius (1701–1744), the freezing point of water is taken to be 0°C. The normal boiling point of water is 100°C. Household thermometers in the United States are commonly marked in *Fahrenheit* degrees. Daniel Fahrenheit (1686–1736) was a German instrument maker who was the first to use the mercury-in-glass thermometer. On this scale, the normal freezing and boiling points of water are taken to be 32° and 212°, respectively (Figure 1.10). It follows that (212°F − 32°F) = 180°F covers the same temperature interval as (100°C − 0°C) = 100°C. This leads to the general relation between the two scales:

$$t_{°F} = 1.8\, t_{°C} + 32° \qquad \textbf{(1.1)}$$

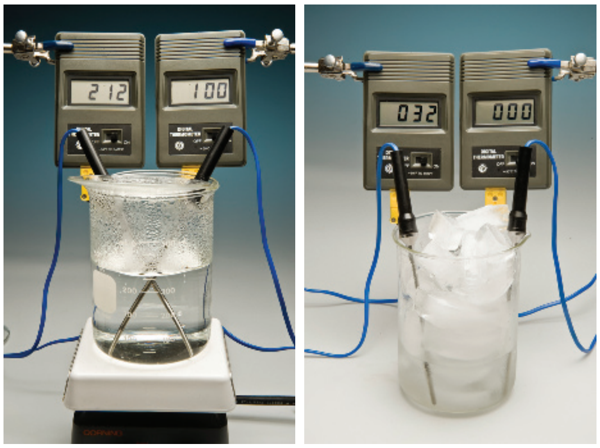

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Figure 1.10 Relationship between Fahrenheit and Celsius scales. This figure shows the relationship between the Fahrenheit and Celsius temperature scales. Note that there are 180 degrees F for 100 degrees C (1.8 F/C) and 0°C = 32°F.

The two scales coincide at $-40°$; as you can readily see from equation 1.1:

$$\text{At } -40°\text{C:} \qquad t_{°F} = 1.8(-40°) + 32° = -72° + 32° = -40°$$

For many purposes in chemistry, the most convenient unit of temperature is the **kelvin** (K); note the absence of the degree sign. The kelvin is defined to be 1/273.16 of the difference between the lowest attainable temperature (0 K) and the triple point of water* (0.01°C). The relationship between temperature in K and in °C is

$$T_K = t_{°C} + 273.15 \qquad \textbf{(1.2)}$$

This scale is named after Lord Kelvin (1824–1907), a British scientist who showed in 1848, at the age of 24, that it is impossible to reach a temperature lower than 0 K.

EXAMPLE 1.1

Mercury thermometers have been phased out because of the toxicity of mercury vapor. A common replacement for mercury in glass thermometers is the organic liquid isoamyl benzoate, which boils at 262°C. What is its boiling point in (a) °F? (b) K?

ANALYSIS	
Information given:	Boiling point (262°C)
Asked for:	boiling point in °F and K

STRATEGY

1. Substitute into Equation 1.1 for (a).
2. Substitute into Equation 1.2 for (b).

SOLUTION	
(a) °F	°F = 1.8(°C) + 32 = 1.8(262°C) + 32 = 504°F
(b) K	K = 273.15 + 262°C = 535 K

As you can see from this discussion, a wide number of different units can be used to express measured quantities in the metric system. This proliferation of units has long been of concern to scientists. In 1960 a self-consistent set of metric units was proposed. This so-called International System of Units (SI) is discussed in Appendix 1. The SI units for the four properties we have discussed so far are

Length:	meter (m)	*Mass:*	kilogram (kg)
Volume:	cubic meter (m^3)	*Temperature:*	kelvin (K)

Uncertainties in Measurements: Significant Figures

Every measurement carries with it a degree of uncertainty. Its magnitude depends on the nature of the measuring device and the skill of its operator. Suppose, for example, you measure out 8 mL of liquid using the 100-mL graduated cylinder shown in Figure 1.11. Here the volume is uncertain to perhaps ±1 mL. With such a crude measuring device, you would be lucky to obtain a volume between 7 and 9 mL. To obtain greater precision, you could use a narrow 10-mL cylinder, which has divisions in small increments. You might now measure a volume within 0.1 mL of the desired value, in the range of 7.9 to 8.1 mL. By using a buret, you could reduce the uncertainty to ±0.01 mL.

Anyone making a measurement has a responsibility to indicate the uncertainty associated with it. Such information is vital to someone who wants to repeat the

*The triple point of water (Chapter 9) is the one unique temperature and pressure pair at which ice, liquid water, and water vapor can coexist in contact with one another.

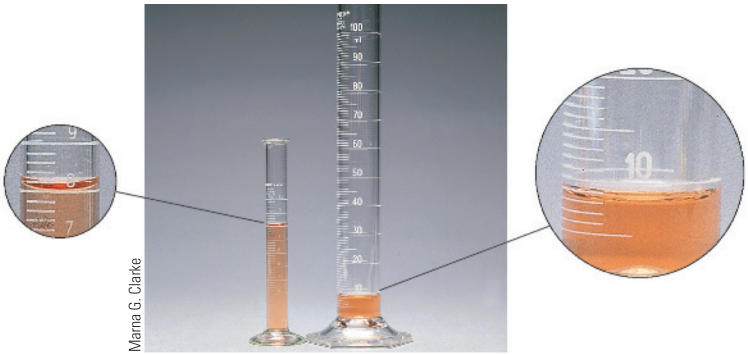

Figure 1.11 Uncertainty in measuring volume. The uncertainty depends on the nature of the measuring device. Eight mL of liquid can be measured with less uncertainty in the 10-mL graduated cylinder than in the 100-mL graduated cylinder.

experiment or judge its precision. The three volume measurements referred to earlier could be reported as

8 ± 1 mL	(large graduated cylinder)
8.0 ± 0.1 mL	(small graduated cylinder)
8.00 ± 0.01 mL	(buret)

In this text, we will drop the ± notation and simply write

8 mL   8.0 mL   8.00 mL

There's a big difference between 8 mL and 8.00 mL, perhaps as much as half a milliliter.

When we do this, it is understood that there is an ***uncertainty of at least one unit in the last digit***—that is, 1 mL, 0.1 mL, 0.01 mL, respectively. This method of citing the degree of confidence in a measurement is often described in terms of **significant figures,** the meaningful digits obtained in a measurement. In 8.00 mL there are three significant figures; each of the three digits has experimental meaning. Similarly, there are two significant figures in 8.0 mL and one significant figure in 8 mL.

Frequently we need to know the number of significant figures in a measurement reported by someone else (Example 1.2).

EXAMPLE 1.2

Using different balances, three different students weigh the same object. They report the following masses:

(a) 1.611 g   (b) 1.60 g   (c) 0.001611 kg

How many significant figures does each value have?

STRATEGY

Assume each student reported the mass in such a way that the last number indicates the uncertainty associated with the measurement.

SOLUTION

(a) 1.611 g	4	
(b) 1.60 g	3	The zero after the decimal point is significant. It indicates that the object was weighed to the nearest 0.01 g.
(c) 0.001611 kg	4	The zeros at the left are not significant. They are only there because the mass was expressed in kilograms rather than grams. Note that 1.611 g and 0.001611 kg represent the same mass.

END POINT

If you express these masses in exponential notation as 1.611×10^0 g, 1.60×10^0 g, and 1.611×10^{-3} kg, the number of significant figures becomes obvious.

Sometimes the number of significant figures in a reported measurement is ambiguous. Suppose that a piece of metal is reported to weigh 500 g. You cannot be sure how many of these digits are meaningful. Perhaps the metal was

Unfortunately, the uncertainty here is uncertain.

weighed to the nearest gram (500 ± 1 g). If so, the 5 and the two zeros are significant; there are three significant figures. Then again, the metal might have been weighed only to the nearest 10 g (500 ± 10 g). In this case, only the 5 and one zero are known accurately; there are two significant figures. About all you can do in such cases is to wish the person who carried out the weighing had used exponential notation. The mass should have been reported as

$$5.00 \times 10^2 \text{ g} \qquad \text{(3 significant figures)}$$

or

$$5.0 \times 10^2 \text{ g} \qquad \text{(2 significant figures)}$$

or

$$5 \times 10^2 \text{ g} \qquad \text{(1 significant figure)}$$

In general, *any ambiguity concerning the number of significant figures in a measurement can be resolved by using exponential notation* (often referred to as "scientific notation"), discussed in Appendix 3.

The number of significant figures is the number of digits shown when a quantity is expressed in exponential notation.

Most measured quantities are not end results in themselves. Instead, they are used to calculate other quantities, often by multiplication or division. The precision of any such derived result is limited by that of the measurements on which it is based. ***When measured quantities are multiplied or divided, the number of significant figures in the result is the same as that in the quantity with the smallest number of significant figures.***

The rule is approximate, but sufficient for our purposes.

EXAMPLE 1.3

A patient is given an antibiotic intravenously. The rate of infusion is set so that the patient receives 1.15 mg of antibiotic per minute. How many milligrams of antibiotic are received after 35 minutes of infusion?

ANALYSIS

Information given:	rate of infusion (1.15 mg/min) time elapsed (35 minutes)
Asked for:	amount of antibiotic infused

STRATEGY

1. Substitute into the formula.

$$\text{rate} = \frac{\text{amount infused}}{\text{time}}$$

2. Recall the rules for significant figures.

SOLUTION

amount infused	amount infused = rate × time = (1.15 mg/min)(35 min) = 40.25 mg
significant figures	rate: 3; time: 2 The answer should have 2 significant figures.
amount infused	4.0×10^1 mg

The rules for "rounding off" a measurement, which were applied in Example 1.3, are as follows:

1. *If the digits to be discarded are less than – – 500 . . ., leave the last digit unchanged.* Masses of 23.315 g and 23.487 g both round off to 23 g if only two significant digits are required.

2. *If the digits to be discarded are greater than – – 500 . . ., add one to the last digit.* Masses of 23.692 g and 23.514 g round off to 24 g.
3. *If, perchance, the digits to be discarded are – – 500 . . . (or simply – – 5 by itself), round off so that the last digit is an even number.* Masses of 23.500 g and 24.5 g both round off to 24 g (two significant figures).

This way, you round up as often as you round down.

When measured quantities are added or subtracted, the uncertainty in the result is found in a quite different way than when they are multiplied and divided. It is determined by counting the number of decimal places, that is, the number of digits to the right of the decimal point for each measured quantity. ***When measured quantities are added or subtracted, the number of decimal places in the result is the same as that in the quantity with the greatest uncertainty and hence the smallest number of decimal places.***

In a multi-step calculation, round off only in the final step.

To illustrate this rule, suppose you want to find the total volume of a vanilla latté made up of 2 shots of espresso (1 shot = 46.1 mL), 301 mL of milk, and 2 tablespoons of vanilla syrup (1 tablespoon = 14.787 mL).

	Volume	Uncertainty	
Espresso coffee	92.2 mL	±0.1 mL	1 decimal place
Milk	301 mL	±1 mL	0 decimal place
Vanilla syrup	29.574 mL	±0.001 mL	3 decimal places
Total volume	423 mL		

Because there are no digits after the decimal point in the volume of milk, there are none in the total volume. Looking at it another way, we can say that the total volume, 423 mL, has an uncertainty of ±1 mL, as does the volume of milk, the quantity with the greatest uncertainty.

In applying the rules governing the use of significant figures, you should keep in mind that certain numbers involved in calculations are exact rather than approximate. To illustrate this situation, consider the equation relating Fahrenheit and Celsius temperatures:

$$t_{°F} = 1.8t_{°C} + 32°$$

The numbers 1.8 and 32 are exact. Hence they do not limit the number of significant figures in a temperature conversion; that limit is determined only by the precision of the thermometer used to measure temperature.

A different type of exact number arises in certain calculations. Suppose you are asked to determine the amount of heat evolved when *one kilogram* of coal burns. The implication is that because "one" is spelled out, *exactly* one kilogram of coal burns. The uncertainty in the answer should be independent of the amount of coal.

A number that is spelled out (one, two, . . .) does not affect the number of significant figures.

Conversion of Units

It is often necessary to convert a measurement expressed in one unit to another unit in the same system or to convert a unit in the English system to one in the metric system. To do this we follow what is known as a **conversion factor** approach or dimensional analysis. For example, to convert a volume of 536 cm^3 to liters, the relation

$$1\text{ L} = 1000\text{ cm}^3$$

There are exactly 1000 cm^3 in exactly 1 L.

is used. Dividing both sides of this equation by 1000 cm^3 gives a quotient equal to 1:

$$\frac{1\text{ L}}{1000\text{ cm}^3} = \frac{1000\text{ cm}^3}{1000\text{ cm}^3} = 1$$