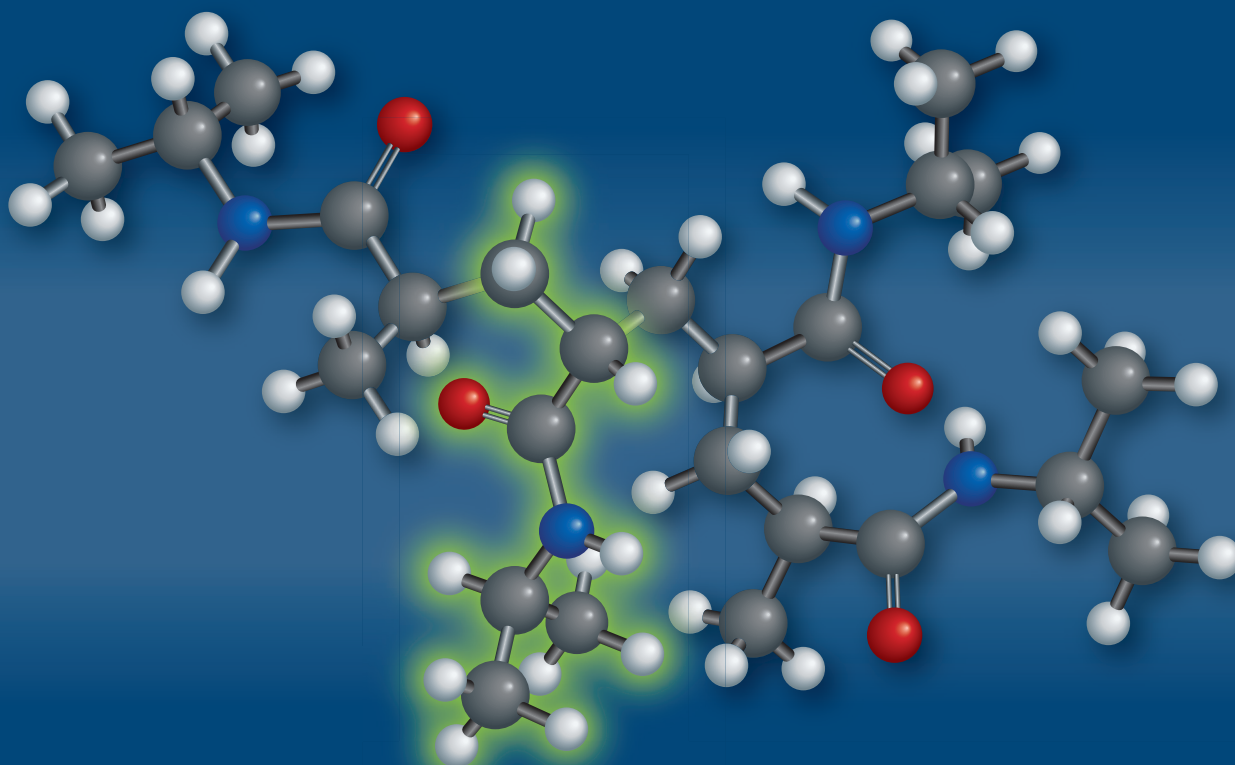


GENERAL CHEMISTRY

PRINCIPLES AND MODERN APPLICATIONS

ELEVENTH EDITION



PETRUCCI

HERRING

MADURA

BISSONNETTE

GENERAL CHEMISTRY

PRINCIPLES AND MODERN APPLICATIONS

ELEVENTH EDITION

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WARNING: Many of the compounds and chemical reactions described or pictured in this book are hazardous. Do not attempt any experiment pictured or implied in the text except with permission in an authorized laboratory setting and under adequate supervision.

PEARSON

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We, the authors, dedicate this edition to Ralph H. Petrucci who passed away as the final edits of this edition were being completed. The first edition of *General Chemistry: Principles and Modern Applications* was published in 1972 with Ralph as the sole author. Although the book is now in its eleventh edition, with more authors, it is still shaped by Ralph's original vision and his belief that students are very much interested in the practical applications, social significance, and historical roots of the subject areas they study, as well as their conceptual frameworks, facts, and theories. Ralph was an inspiring mentor who warmly welcomed each of us to the authoring team. We envied his clear and precise writing style and impeccable eye for detail. He was an excellent advisor to us during the preparation of the most recent editions, all of which benefited greatly from his valuable input. We will miss him dearly.

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modeling dynamical processes of polyatomic molecules in the gas phase. He has won awards for excellence in teaching, including the University of Waterloo's Distinguished Teacher Award in 2005. Dr. Bissonnette has made extensive use of technology in both the classroom and the laboratory to create an interactive environment for his students to learn and explore. For the past several years, he has been actively engaged in undergraduate curriculum development, high-school liaison activities, and the coordination of the university's high-school chemistry contests, which are written each year by students around the world.

Preface

“Know your audience.” For this new edition, we have tried to follow this important advice by attending even more to the needs of students who are taking a serious journey through this material. We also know that most general chemistry students have career interests not in chemistry but in other areas such as biology, medicine, engineering, environmental science, and agricultural sciences. And we understand that general chemistry will be the only university or college chemistry course for some students, and thus their only opportunity to learn some practical applications of chemistry. We have designed this book for all these students.

Students of this text should have already studied some chemistry. But those with no prior background and those who could use a refresher 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 may need help identifying and applying principles and 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 students’ skills in problem solving and critical thinking. Thus, we have tried to strike the proper balances between principles and applications, qualitative and quantitative discussions, and rigor and simplification.

Throughout the text and on the MasteringChemistry® site (www.masteringchemistry.com) we provide real-world examples to enhance the discussion. Examples relevant to the biological sciences, engineering, and the environmental sciences are found in numerous places. This should help to bring chemistry alive for these students and help them understand its relevance to their career interests. It also, in most cases, should help them master core concepts.

ORGANIZATION

In this edition we retain the core organization of the previous edition with two notable exceptions. First, we have moved the chapter entitled Spontaneous Change: Entropy and Gibbs Energy forward in the text. It is now Chapter 13. By moving the introduction of entropy and Gibbs energy forward in the text, we are able to use these concepts in subsequent chapters. Second, we have moved the chapter on chemical kinetics to Chapter 20. Consequently, the discussion of chemical kinetics now appears after the chapters that rely on equilibrium and thermodynamic concepts.

Like the previous edition, this edition begins with a brief overview of core concepts in Chapter 1. Then, we introduce atomic theory, including the periodic table, in Chapter 2. The periodic table is an extraordinarily useful tool, and presenting it early allows us to use the periodic table in different ways throughout the early chapters of the text. In Chapter 3, we introduce chemical compounds and their stoichiometry. Organic compounds are included in this presentation. The early introduction of organic compounds allows us to use organic examples throughout the book. Chapters 4 and 5 introduce chemical reactions. We discuss gases in Chapter 6, partly because they are familiar to students (which helps them build confidence), but also because some instructors prefer to cover this material early to better integrate their lecture and lab programs. (Chapter 6 can easily be deferred for coverage with the other states of matter, in Chapter 12.)

In Chapter 7, we introduce thermochemistry and discuss the energy changes that accompany physical and chemical transformations. Chapter 8 introduces quantum mechanical concepts that are needed to understand the energy changes we encounter at the atomic level. This chapter includes a discussion of wave

mechanics, although this topic may be omitted at the instructor's discretion. Collectively, Chapters 8 through 11 provide the conceptual basis for describing the electronic structure of atoms and molecules, and the physical and chemical properties of these entities. The properties of atoms and molecules are then used in Chapter 12 to rationalize the properties of liquids and solids.

Chapter 13 is a significant revision of Chapter 19 from the tenth edition. It introduces the concept of entropy, the criteria for predicting the direction of spontaneous change, and the thermodynamic equilibrium condition. In Chapters 14–19, we apply and extend concepts introduced in Chapter 13. However, Chapters 14–19 can be taught without explicitly covering, or referring back to, Chapter 13.

As with previous editions, we have emphasized real-world chemistry in the final chapters that cover descriptive chemistry (Chapters 21–24), 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. The text ends with comprehensive chapters on nuclear chemistry (Chapter 25) and organic chemistry (Chapters 26 and 27). Please note that an additional chapter on biochemistry (Chapter 28) is available online.

CHANGES TO THIS EDITION

We have made the following important changes in specific chapters and appendices:

- In Chapter 2 (Atoms and the Atomic Theory), new material is included to describe the use of atomic mass intervals and conventional atomic masses for elements such as H, Li, B, C, N, O, Mg, Si, S, Cl, Br, and Tl. Atomic mass intervals are recommended by the IUPAC because the isotopic abundances of these elements vary from one source to another, and therefore, their atomic masses cannot be considered constants of nature.
- Chapter 4 (Chemical Reactions) includes a new section that discusses the extent of reaction, and introduces a tabular approach for representing the changes in amount in terms of a single variable, representing the extent of reaction.
- In Chapter 5 (Introduction to Reactions in Aqueous Solutions), we revised Section 5-1 to differentiate between dissociation and ionization, and introduced a new figure to illustrate the dissociation of an ionic compound in water.
- Chapter 6 (Gases) makes increased use of the recommended units of pressure (e.g., Pa, kPa, and bar). Section 6-7 on the kinetic-molecular theory has been significantly revised. For example, the subsection on Derivation of Boyle's Law has been simplified and now comes after the subsections on Distribution of Molecular Speeds and The Meaning of Temperature. Section 6-8 has also been revised so that Graham's law is presented first, as an empirical law, which is then justified by using the kinetic-molecular theory.
- In Chapter 7 (Thermochemistry), we have updated the notation to ensure that we are using, for the most part, symbols that are recommended by the IUPAC. For example, standard enthalpies of reaction are represented by the symbol $\Delta_r H^\circ$ (not ΔH°) and are expressed in kJ mol^{-1} (not kJ). We have added a molecular interpretation of specific heat capacities (in Section 7-2) and an introduction to entropy (in Section 7-10).
- Chapter 8 (Electrons in Atoms) has been substantially rewritten to provide a logical introduction to the ideas leading to wave mechanics. Sections 8-2 and 8-3 of the previous edition have been combined and the

material reorganized. This chapter includes a new section that focuses on the energy level diagram and spectrum of the hydrogen atom. The section entitled Interpreting and Representing the Orbitals of the Hydrogen Atom has been rewritten to include a discussion of the radial functions. A new subsection describing the conceptual model for multielectron atoms has been added to the section entitled Multielectron Atoms. The sections on multielectron atoms and electron configurations have been rewritten to emphasize more explicitly that the observed ground-state electron configuration for an atom is the one that minimizes E_{atom} and that the energies of the orbitals is only one consideration. There are two new Are You Wondering? boxes in this chapter: Is the Born interpretation an idea we use to determine the final form of a wave function? and Are all orbital transitions allowed in atomic absorption and emission spectra?

- In Chapter 9 (The Periodic Table and Some Atomic Properties), a number of sections have been rewritten to emphasize the importance of effective nuclear charge in determining atomic properties. A new section on polarizability has been introduced. Several new figures have been created to illustrate the variation of effective nuclear charge and atomic properties across a period or down a group (e.g., effective nuclear charges for the first 36 elements; the variation of effective nuclear charge and percent screening with atomic number; the variation of average distance from the nucleus with atomic number; first ionization energies of the third row *p*-block elements; electron affinities of some of the main group elements; polarization of an atom; the variation of polarizability and atomic volume with atomic number). The sections on ionization energies and electron affinities have been significantly revised. Of particular note, we have revised the discussion of the decrease in ionization energy that occurs as we move from group 2 to 13 and from group 15 to 16. Our discussion points out that various explanations have been used. The section from the tenth edition entitled Periodic Properties of the Elements has been deleted.
- Chapter 11 (Chemical Bonding II: Valence Bond and Molecular Orbital Theories) has been revised to include an expanded discussion of the redistribution of electron density that occurs during bond formation, an improved introduction to Section 11-5 Molecular Orbital Theory, and an improved discussion of molecular orbital theory of the CO molecule. We have moved the section entitled Bonding in Metals online.
- Chapter 13 (Spontaneous Change: Entropy and Gibbs Energy) is a totally revised version of Chapter 19 from the previous edition. The chapter focuses first on Boltzmann's view of entropy, which is based on microstates, and then on Clausius's view, which relates entropy change to reversible heat transfer. The connection between microstates and particle-in-a-box model is developed to reinforce Boltzmann's view of entropy. Clausius's view of entropy change is used to develop expressions for important and commonly encountered physical changes (e.g., phase transitions; heating or cooling at constant pressure; isothermal expansion or compression of an ideal gas). These expressions are subsequently used to develop the criterion for predicting the direction of spontaneous change. The chapter includes a proper description of the difference between the Gibbs energy change of a system, ΔG , and the reaction Gibbs energy, $\Delta_r G$. The reaction Gibbs energy ($\Delta_r G$) is used as the basis for describing how the Gibbs energy of a system changes with composition (i.e., with respect to the extent of reaction). The derivation of the equation is done in a separate section that may be used or skipped at the instructor's discretion. The concepts of chemical potential and activity are also introduced.

- In Chapter 14 (Solutions and Their Physical Properties), we have added a section to describe the standard thermodynamic properties of aqueous ions. We use the concepts of entropy and chemical potential in Chapter 13 to explain vapor pressure lowering and why gasoline and water don't mix.
- Chapter 15 (Principles of Chemical Equilibrium) has been significantly revised to emphasize the thermodynamic basis of equilibrium and to de-emphasize aspects of kinetics. There is an increased emphasis on the thermodynamic equilibrium constant, which is expressed in terms of activities, along with an updated discussion of Le Châtelier's principle to emphasize certain limitations associated with its use (e.g., for certain reactions and initial conditions, the addition of a reactant may actually cause net change to the left). Several new worked examples are included to show how equilibrium constant expressions may be simplified and solved when the equilibrium constant is either very small or very large.
- In Chapter 16 (Acids and Bases), significant changes have been made. Sections 16-1 through 16-3 have been significantly revised to provide a more logical flow and to emphasize and demonstrate that the distinction between strong and weak acids is based on the degree of ionization, which in turn depends on the magnitude of the acid ionization constant. There are two new sections, namely Sections 16-7 (Simultaneous or Consecutive Acid–Base Reactions: A General Approach) and 16-9 (Qualitative Aspects of Acid–Base Reactions). Section 16-7 focuses on writing and using material balance and charge equations. Section 16-9 focuses on predicting the equilibrium position of a general acid–base reaction. A new subsection entitled Rationalization of Acid Strengths: An Alternative Approach has been added to Section 16-10, Molecular Structure and Acid–Base Behavior. This new subsection focuses on factors that stabilize the anion formed by an acid.
- In Chapter 19 (Electrochemistry), we have modified the Nernst equation to have the form $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0257 \text{ V}}{z} \ln Q$. We have changed the text so that the standard hydrogen electrode is defined with respect to a pressure of 1 bar instead of 1 atm, and added a problem to the Integrative and Advanced Exercises to illustrate that this change in pressure causes only a small change in the standard reduction potentials (see Exercise 108). We have also added a section on reserve batteries.
- In Appendix D, we have modified the table of Standard Electrode (Reduction) Potentials at 25 °C so that it now includes a column with the cell notation for the half-reactions.

In addition to the specific changes noted above, we have also changed much of the artwork throughout the textbook. In particular, all of the atomic and molecular orbital representations have been modified to be consistent across all chapters. We have redone all of the electrostatic potential maps (EPMs) to have the same potential energy color scale unless noted in the textbook.

OVERALL APPROACH

The pedagogical apparatus and overall approach in this edition continue to reflect contemporary thoughts on how best to teach general chemistry. We have retained the following key features of the text:

- *Logical approach to solving problems.* All worked examples are presented consistently throughout the text by using a tripartite structure of Analyze–Solve–Assess. This presentation not only encourages students to use a logical approach in solving problems but also provides them

with a way to start when they are trying to solve a problem that may seem, at first, impossibly difficult. The approach is used implicitly by those who have had plenty of practice solving problems, but for those who are just starting out, the Analyze–Solve–Assess structure will serve to remind students to (1) analyze the information and plan a strategy, (2) implement the strategy, and (3) check or assess their answer to ensure that it is a reasonable one.

- *Integrative Practice Examples and End-of-Chapter Exercises.* Users of previous editions have given us very positive feedback about the quality of the integrative examples at the end of each chapter and the variety of the end-of-chapter exercises. We have added two practice examples (Practice Example A and Practice Example B) to every Integrative Example in the text. Rather than replace end-of-chapter exercises with new exercises, we have opted to increase the number of exercises. In most chapters, at least 10 new exercises have been added; and in many chapters, 20 or more exercises have been added.
- *Use of IUPAC recommendations.* We are pleased that our book serves the needs of instructors and students around the globe. Because communication among scientists in general, and chemists in particular, is made easier when we agree to use the same terms and notations, we have decided to follow—with relatively few exceptions—recommendations made by the International Union of Pure and Applied Chemistry (IUPAC). In particular, the version of the periodic table that now appears throughout the text is based on the one currently endorsed by IUPAC. The IUPAC-endorsed version places the elements lanthanum (La) and actinium (Ac) in the lanthanides and actinides series, respectively, rather than in group 3. Interestingly, almost every other chemistry book still uses the old version of the periodic table, even though the proper placement of La and Ac has been known for more than 20 years! An important change is the use of IUPAC-recommended symbols and units for thermodynamic quantities. For example, in this edition, standard enthalpies of reaction are represented by the symbol $\Delta_r H^\circ$ (not ΔH_r°) and are expressed in kJ mol^{-1} (not kJ).

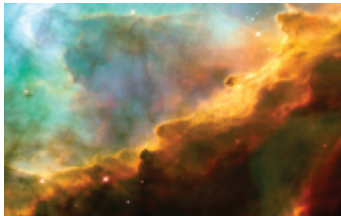
FEATURES OF THIS EDITION

We have made a careful effort with this edition to incorporate features that will facilitate the teaching and learning of chemistry.

Matter: Its Properties and Measurement

1

CONTENTS	
1.1 The Scientific Method	1.5 Density and Percent Composition: Their Use in Problem Solving
1.2 Properties of Matter	1.6 Uncertainties in Scientific Measurements
1.3 Classification of Matter	1.7 Significant Figures
1.4 Measurement of Matter: SI (Metric) Units	



LEARNING OBJECTIVES

1.1 Describe the purpose and process of the scientific method.

1.2 Discuss the meaning of matter and the changes it can undergo physically and chemically.

1.3 Classify matter based on its basic building blocks (atoms), and identify the three states of matter.

1.4 Identify the SI unit for length, mass, time, temperature, amount of substance, electric current, and luminous intensity.

1.5 Use percent composition and the relationship among density, volume, and mass, as conversion factors in problem solving.

1.6 Differentiate between precision and accuracy.

1.7 Use the standard rules for significant figures to determine the number of significant figures needed at the end of a calculation.

The result of multiplication or division may contain only as many significant figures as the least precisely known quantity in the calculation.

Chapter Opener

Each chapter opens with listing of the main headings to provide a convenient overview of the chapter's **Contents**. The opener also contains a list of numbered **Learning Objectives** that correspond with the main sections of the chapter.

Key Terms

Key terms are boldfaced where they are defined in the text. A **Glossary** of key terms with their definitions is presented in Appendix F.

Highlighted Boxes

Significant equations, concepts, and rules are highlighted against a color background for easy reference.

Concept Assessment

Concept Assessment questions (many of which are qualitative) are distributed throughout the body of the chapters. They enable students to test their understanding of basic concepts before proceeding further. Full solutions are provided in Appendix H.

Examples with Practice Examples A and B Worked-out Examples throughout the text illustrate how to apply the concepts. In many instances, a drawing or photograph is included to help students visualize what is going on in the problem. More importantly, all worked-out Examples now follow a tripartite structure of **Analyze–Solve–Assess** to encourage students to adopt a logical approach to problem solving.

Two **Practice Examples** are provided for each worked-out Example. The first, **Practice Example A**, provides immediate practice in a problem very similar to the given Example. The second, **Practice Example B**, often takes the student one step further than the given Example and is similar to the end-of-chapter problems in terms of level of difficulty. Answers to all the Practice Examples are given in Appendix G.

Marginal Notes

Marginal notes help clarify important points.

Keep In Mind Notes

Keep In Mind margin notes remind students about ideas introduced earlier in the text that are important to an understanding of the topic under discussion. In some instances they also warn students about common pitfalls.

Are You Wondering?

Are You Wondering? boxes pose and answer good questions that students often ask. Some are designed to help students avoid common misconceptions; others provide analogies or alternate explanations of a concept; and still others address apparent inconsistencies in the material that the students are learning. These topics can be assigned or omitted at the instructor's discretion.

Focus On Discussions

References are given near the end of each chapter to a **Focus On** essay that is found on the MasteringChemistry® site (www.masteringchemistry.com). These essays describe interesting and significant applications of the chemistry discussed in the chapter. They help show the importance of chemistry in all aspects of daily life.

2-4 CONCEPT ASSESSMENT

What is the single exception to the statement that all atoms comprise protons, neutrons, and electrons?

EXAMPLE 14-5 Using Henry's Law

At 0 °C and an O₂ pressure of 1.00 atm, the aqueous solubility of O₂(g) is 48.9 mL O₂ per liter. What is the molarity of O₂ in a saturated water solution when the O₂ is under its normal partial pressure in air, 0.2095 atm?

Analyze

Think of this as a two-part problem. (1) Determine the molarity of the saturated O₂ solution at 0 °C and 1 atm. (2) Use Henry's law in the manner just outlined.

Solve

Determine the molarity of O₂ at 0 °C when P_{O₂} = 1 atm. We are given the information that, at an O₂ pressure of 1.00 atm, a saturated solution of O₂ in water contains 48.9 mL (0.0489 L) of O₂. We also know that, at 0 °C and 1.00 atm, 1 mol O₂ occupies a volume of 22.4 L. Therefore,

$$\text{molarity} = \frac{0.0489 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2}}{1 \text{ L soln}} = 2.18 \times 10^{-3} \frac{\text{mol O}_2}{\text{L soln}} = 2.18 \times 10^{-3} \text{ M}$$

Evaluate the Henry's law constant.

$$k = \frac{C}{P_{\text{gas}}} = \frac{2.18 \times 10^{-3} \text{ M}}{1.00 \text{ atm}}$$

Apply Henry's law.

$$C = k \times P_{\text{gas}} = \frac{2.18 \times 10^{-3} \text{ M}}{1.00 \text{ atm}} \times 0.2095 \text{ atm} = 4.57 \times 10^{-4} \text{ M}$$

Assess

When working problems involving gaseous solutes in a solution in which the solute is at very low concentration, use Henry's law.

PRACTICE EXAMPLE A: Use data from Example 14-5 to determine the partial pressure of O₂ above an aqueous solution at 0 °C known to contain 5.00 mg O₂ per 100.0 mL of solution.

PRACTICE EXAMPLE B: A handbook lists the solubility of carbon monoxide in water at 0 °C and 1 atm pressure as 0.0354 mL CO per milliliter of H₂O. What pressure of CO(g) must be maintained above the solution to obtain 0.0100 M CO?

► Other atomic symbols not based on English names include Cu, Ag, Sn, Sb, Au, and Hg.

KEEP IN MIND

that all we know is that the second oxide is twice as rich in oxygen as the first. If the first is CO, the possibilities for the second are CO₂, C₂O₄, C₃O₆, and so on. (See also Exercise 18.)



1-1 ARE YOU WONDERING?

Why is it so important to attach units to a number?

In 1993, NASA started the Mars Surveyor program to conduct an ongoing series of missions to explore Mars. In 1995, two missions were scheduled that would be launched in late 1998 and early 1999. The missions were the Mars Climate Orbiter (MCO) and the Mars Polar Lander (MPL). The MCO was launched December 11, 1998, and the MPL, January 3, 1999.

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What is the most abundant element? This seemingly simple question does not have a simple answer. To learn more about the abundances of elements in the universe and in the Earth's crust, go to the Focus On feature for Chapter 2, entitled Occurrence and Abundances of the Elements, on the MasteringChemistry site.

Summary

2-1 Early Chemical Discoveries and the Atomic Theory—Modern chemistry began with eighteenth-century discoveries leading to the formulation of two basic laws of chemical combination, the **law of conservation of mass** and the **law of constant composition (definite proportions)**. These discoveries led to Dalton's atomic theory—that matter is composed of indestructible particles called atoms, that the atoms of an element are identical to one another but different from atoms of all other elements, and that chemical compounds are combinations of atoms of different elements. Based on this theory, Dalton proposed still another law of chemical combination, the **law of multiple proportions**.

2-2 Electrons and Other Discoveries in Atomic Physics—The first clues to the structures of atoms came through the discovery and characterization of **cathode rays (electrons)**. Key experiments were those that established

the mass-to-charge ratio (Fig. 2-7) and then the charge on an electron (Fig. 2-8). Two important accidental discoveries made in the course of cathode-ray research were of X-rays and **radioactivity**. The principal types of radiation emitted by radioactive substances are **alpha (α) particles**, **beta (β) particles**, and **gamma (γ) rays** (Fig. 2-10).

2-3 The Nuclear Atom—Studies on the scattering of α particles by thin metal foils (Fig. 2-11) led to the concept of the nuclear atom—a tiny, but massive, positively charged nucleus surrounded by lightweight, negatively charged electrons (Fig. 2-12). A more complete description of the nucleus was made possible by the discovery of **protons** and **neutrons**. An individual atom is characterized in terms of its **atomic number (proton number) Z** and **mass number, A** . The difference, $A - Z$, is the **neutron number**. The masses of individual atoms and their component parts are expressed in **atomic mass units (u)**.

Integrative Example

For use in analytical chemistry, sodium thiosulfate solutions must be carefully prepared. In particular, the solutions must be kept from becoming acidic. In strongly acidic solutions, thiosulfate ion disproportionates into $\text{SO}_2(\text{g})$ and $\text{S}_8(\text{s})$.



▲ Decomposition of thiosulfate ion

When an aqueous solution of $\text{Na}_2\text{S}_2\text{O}_3$ is acidified, the sulfur is in the colloidal state when first formed (right).

Show that the disproportionation of $\text{S}_2\text{O}_3^{2-}(\text{aq})$ is spontaneous for standard-state conditions in acidic solution, but not in basic solution.

Analyze

Begin by writing the half-equations and an overall equation for the disproportionation reaction. Determine E_{cell}° for the reaction and thus whether the reaction is spontaneous for standard-state conditions in acidic solution. Then make a qualitative assessment of whether the reaction is likely to be more spontaneous or less spontaneous in basic solution.

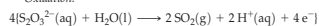
Solve

Balance the overall equation on the verbal description of the reaction.

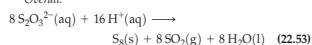
Reduction:



Oxidation:



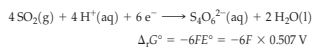
Overall:



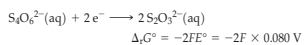
PRACTICE EXAMPLE A: Use information from Figure 22-17 to decide whether the nitrite anion, NO_2^- , disproportionates spontaneously in basic solution to NO_3^- and NO . Assume standard-state conditions.

PRACTICE EXAMPLE B: Does HNO_2 spontaneously disproportionate to NO_3^- and NO in acidic solution? Assume standard-state conditions. [Hint: Use data from Figure 22-17.]

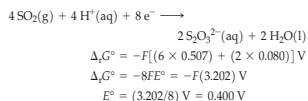
To determine E_{cell}° for the reaction (22.53), use data from Figure 22-13. That figure gives an E° value for the reduction half-reaction (0.465 V) but no value for the oxidation. To obtain this missing E° , use additional data from Figure 22-13 together with the method of Example 22-1. That is, the sum of the half-equation



and the half-equation



yields the desired new half-equation and its E° value.



Now we can calculate E_{cell}° for reaction (22.53).

$$E_{\text{cell}}^{\circ} = E^{\circ}(\text{reduction}) - E^{\circ}(\text{oxidation}) \\ = 0.465 \text{ V} - 0.400 \text{ V} = 0.065 \text{ V}$$

The disproportionation is spontaneous for standard-state conditions in acidic solution.

Increasing $[\text{OH}^-]$, as would be the case in making the solution basic, means decreasing $[\text{H}^+]$. In fact, $[\text{OH}^-] = 1 \text{ M}$ corresponds to $[\text{H}^+] = 1 \times 10^{-14} \text{ M}$. Because equation (22.53) has $\text{H}^+(\text{aq})$ on the left side of the equation, a decrease in $[\text{H}^+]$ favors the reverse reaction (by Le Châtelier's principle). At some point before the solution becomes basic, the forward reaction is no longer spontaneous.

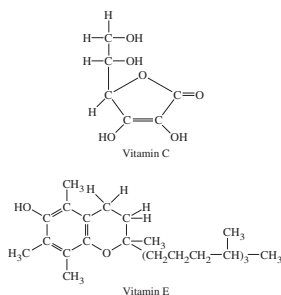
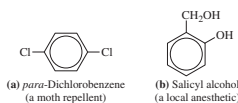
Assess

This calculation demonstrated in a qualitative way that $\text{S}_2\text{O}_3^{2-}(\text{aq})$ is stable in basic solutions and spontaneously disproportionates in acidic solutions. To determine the pH at which the disproportionation becomes spontaneous, one can use the Nernst equation, as seen in Exercise 100.

Exercises

Homogeneous and Heterogeneous Mixtures

- Which of the following do you expect to be most water soluble, and why? $\text{C}_{10}\text{H}_8(\text{s})$, $\text{NH}_2\text{OH}(\text{s})$, $\text{C}_6\text{H}_6(\text{l})$, $\text{CaCO}_3(\text{s})$.
- Which of the following is moderately soluble both in water and in benzene [$\text{C}_6\text{H}_6(\text{l})$], and why? (a) 1-butanol, $\text{CH}_3(\text{CH}_2)_2\text{CH}_2\text{OH}$; (b) naphthalene, C_{10}H_8 ; (c) hexane, C_6H_{14} ; (d) $\text{NaCl}(\text{s})$.
- Substances that dissolve in water generally do not dissolve in benzene. Some substances are moderately soluble in both solvents, however. One of the following is such a substance. Which do you think it is and why?



Summary

A prose **Summary** is provided for each chapter. The Summary is organized by the main headings in the chapter and incorporates the key terms in bold-faced type.

Integrative Example

An **Integrative Example** is provided near the end of each chapter. These challenging examples show students how to link various concepts from the chapter and earlier chapters to solve complex problems. Each Integrative Example is now accompanied by a **Practice Example A** and **Practice Example B**. Answers to these Practice Examples are given in Appendix G.

End-of-Chapter Questions and Exercises

Each chapter ends with four categories of questions:

Exercises are organized by topic subheads and are presented in pairs. Answers to selected questions (i.e., those numbered in red) are given in Appendix G.

Integrative and Advanced Exercises are more advanced than the preceding Exercises. They are not grouped by topic or type. They integrate material from sections of the chapter and sometimes from multiple chapters. In some instances, they introduce new ideas or pursue specific ideas further than is done in the chapter. Answers to selected questions (i.e., those numbered in red) are given in Appendix G.

Feature Problems require the highest level of skill to solve. Some deal with classic experiments; some require students to interpret data or graphs; some suggest alternative techniques for problem solving; some are comprehensive in their scope; and some introduce new material. These problems are a resource that can be used in several ways: for discussion in class, for individually assigned homework, or for collaborative group work. Answers to selected questions (i.e., those numbered in red) are given in Appendix G.

Self-Assessment Exercises are designed to help students review and prepare for some of the types of questions that often appear on quizzes and exams. Students can use these questions to decide whether they are ready to move on to the next chapter or first spend more time working with the concepts in the current chapter. Answers with explanations to selected questions (i.e., those numbered in red) are given in Appendix G.

Appendices

The Appendices at the back of the book provide important information:

Appendix A succinctly reviews of some basic **Mathematical Operations**.

Appendix B concisely describes **Some Basic Physical Concepts**.

Appendix C summarizes the conventions of **SI Units**.

Appendix D provides five useful **Data Tables**.

Appendix E provides guidelines, along with an example, for constructing **Concept Maps**.

Appendix F consists of a **Glossary** of all the key terms in the book.

Appendix G provides **Answers to Practice Examples and Selected Exercises**.

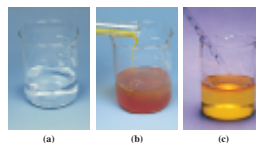
Appendix H provides **Answers to Concept Assessment Questions**.

For easy reference, the **Periodic Table of Elements** and a **Tabular Listing of Elements** are presented on the inside of the front cover.

For convenience, listings of **Selected Physical Constants**, **Some Common Conversion Factors**, **Some Useful Geometric Formulas**, and **Location of Important Data and Other Useful Information** are presented on the inside of the back cover.

Integrative and Advanced Exercises

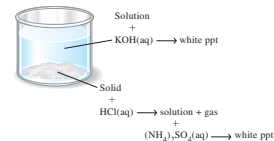
69. Write net ionic equations for the reactions depicted in photo (a) sodium metal reacts with water to produce hydrogen; photo (b) an excess of aqueous iron(III) chloride is added to the solution in (a); and photo (c) the precipitate from (b) is collected and treated with an excess of HCl(aq).



70. Following are some laboratory methods occasionally used for the preparation of small quantities of chemicals. Write a balanced equation for each.
- (a) preparation of $\text{H}_2\text{S}(\text{g})$: $\text{HCl}(\text{aq})$ is heated with $\text{FeS}(\text{s})$
- (b) preparation of $\text{Cl}_2(\text{g})$: $\text{HCl}(\text{aq})$ is heated with $\text{MnO}_2(\text{s})$; $\text{MnCl}_2(\text{aq})$ and $\text{H}_2\text{O}(\text{l})$ are other products
- (c) preparation of $\text{N}_2(\text{g})$: Br_2 and NH_3 react in aqueous solution; NH_4Br is another product

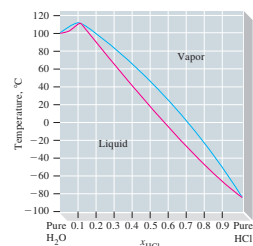
concentration of 0.250 M Cl⁻? Assume that the solution volumes are additive.

76. An unknown white solid consists of two compounds, each containing a different cation. As suggested in the illustration, the unknown is partially soluble in water. The solution is treated with NaOH(aq) and yields a white precipitate. The part of the original solid that is insoluble in water dissolves in HCl(aq) with the evolution of a gas. The resulting solution is then treated with $(\text{NH}_4)_2\text{SO}_4(\text{aq})$ and yields a white precipitate. (a) Is it possible that any of the cations Mg^{2+} , Cu^{2+} , Ba^{2+} , Na^+ , or NH_4^+ were present in the original unknown? Explain your reasoning. (b) What compounds could be in the unknown mixture (that is, what anions might be present)?



Feature Problems

113. Cinnamaldehyde is the chief constituent of cinnamon oil, which is obtained from the twigs and leaves of cinnamon trees grown in tropical regions. Cinnamon oil is used in the manufacture of food flavorings, perfumes, and cosmetics. The normal boiling point of cinnamaldehyde, $\text{C}_9\text{H}_8\text{CH}=\text{CHCHO}$, is 246.0 °C, but at this temperature it begins to decompose. As a result, cinnamaldehyde cannot be easily purified by ordinary distillation. A method that can be used instead is *steam distillation*. A heterogeneous mixture of cinnamaldehyde and water is heated until the sum of the vapor pressures of the two liquids is equal to barometric pressure. At this point, the temperature remains constant as the liquids vaporize. The mixed vapor condenses to produce two immiscible liquids; one liquid is essentially pure water and the other, pure cinnamaldehyde. The following vapor pressures of cinnamaldehyde are given: 1 mmHg at 76.1 °C; 5 mmHg at 105.8 °C; and 10 mmHg at 120.0 °C. Vapor pressures of water are given in Table 14.3.



Self-Assessment Exercises

77. In your own words, define the following terms or symbols: (a) sp^2 ; (b) σ_{2p}^* ; (c) bond order; (d) π bond.
78. Briefly describe each of the following ideas: (a) hybridization of atomic orbitals; (b) σ -bond framework; (c) Kekulé structures of benzene, C_6H_6 .
79. Explain the important distinctions between the terms in each of the following pairs: (a) σ and π bonds; (b) localized and delocalized electrons; (c) bonding and antibonding molecular orbitals.
80. A molecule in which sp^2 hybrid orbitals are used by the central atom in forming covalent bonds is (a) PCl_3 ; (b) N_2 ; (c) SO_2 ; (d) He_2 .
81. The bond angle in H_2Se is best described as (a) between 109° and 120°; (b) less than in H_2S ; (c) less than in H_2S , but not less than 90°; (d) less than 90°.
82. The hybridization scheme for the central atom includes a d orbital contribution in (a) I_3^- ; (b) PCl_3 ; (c) NO_3^- ; (d) H_2Se .
83. Of the following, the species with a bond order of 1 is (a) H_2^+ ; (b) Li_2 ; (c) He_2 ; (d) H_2^- .
84. The hybridization scheme for Xe in XeF_2 is (a) sp ; (b) sp^2 ; (c) sp^3 ; (d) sp^3d ; (e) sp^3d^2 .
85. Delocalized molecular orbitals are found in (a) H_2 ; (b) HS^- ; (c) CH_2 ; (d) CO_3^{2-} .
86. Explain why the molecular structure of BF_3 cannot be adequately described through overlaps involving pure s and p orbitals.
87. Why does the hybridization sp^3d not account for bonding in the molecule BF_3 ? What hybridization scheme does work? Explain.
88. What is the total number of (a) σ bonds and (b) π bonds in the molecule CH_3NCO ?
89. Which of the following species are paramagnetic? (a) B_2 ; (b) B_2^+ ; (c) B_2^- . Which species has the strongest bond?
90. Use the valence molecular orbital configuration to determine which of the following species is expected to have the lowest ionization energy: (a) C_2^+ ; (b) C_2^- ; (c) C_2 .
91. Use the valence molecular orbital configuration to determine which of the following species is expected to have the greatest electron affinity: (a) C_2^+ ; (b) Be_2 ; (c) F_2 ; (d) B_2^- .
92. Which of these diatomic molecules do you think has the greater bond energy, Li_2 or C_2^+ ? Explain.
93. For each of the following ions or molecules, decide whether the structure is best described by a single Lewis structure or by resonance structures. (a) CO_3^{2-} ; (b) HCO_2^- ; (c) NO_3^- .
94. Draw Lewis structures for the NO_2^+ and NO_2^- ions, and determine the likely geometry for each by using VSEPR theory. How does the hybridization of N differ in these two species?

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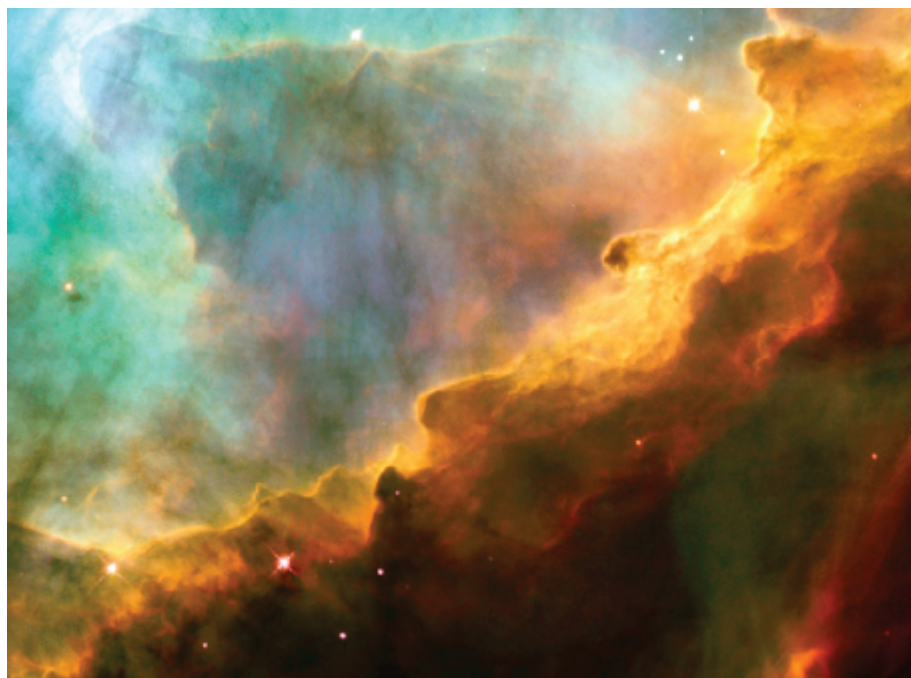
WARNING: Many of the compounds and chemical reactions described or pictured in this book are hazardous. Do not attempt any experiment pictured or implied in the text except with permission in an authorized laboratory setting and under adequate supervision.

Matter: Its Properties and Measurement

1

CONTENTS

1-1	The Scientific Method	1-5	Density and Percent Composition: Their Use in Problem Solving
1-2	Properties of Matter	1-6	Uncertainties in Scientific Measurements
1-3	Classification of Matter	1-7	Significant Figures
1-4	Measurement of Matter: SI (Metric) Units		



ESA, J. Hester (ASU) / NASA

A Hubble Space Telescope image of a cloud of hydrogen gas and dust (lower right half of the image) that is part of the Swan Nebula (M17). The colors correspond to light emitted by hydrogen (green), sulfur (red), and oxygen (blue). The chemical elements discussed in this text are those found on Earth and, presumably, throughout the universe.

From the clinic that treats chemical dependency to a theatrical performance with good chemistry to the food label stating “no chemicals added,” chemistry and chemicals seem an integral part of life, even if everyday references to them are often misleading. A label implying the absence of chemicals in a food makes no sense. All foods consist entirely of chemicals, even if organically grown. In fact, all material objects—whether living or inanimate—are made up only of chemicals, and we should begin our study with that thought clearly in mind.

By manipulating materials in their environment, people have always practiced chemistry. Among the earliest applications were glazing pottery, smelting ores to produce metals, tanning hides, dyeing fabrics, and making cheese, wine, beer, and soap. With modern knowledge, though, chemists

LEARNING OBJECTIVES

- 1.1 Describe the purpose and process of the scientific method.
- 1.2 Discuss the meaning of *matter* and the changes it can undergo physically and chemically.
- 1.3 Classify matter based on its basic building blocks (atoms), and identify the three states of matter.
- 1.4 Identify the SI unit for length, mass, time, temperature, amount of substance, electric current, and luminous intensity.
- 1.5 Use percent composition and the relationship among density, volume, and mass as conversion factors in problem solving.
- 1.6 Differentiate between precision and accuracy.
- 1.7 Use the standard rules for significant figures to determine the number of significant figures needed at the end of a calculation.

can decompose matter into its smallest components (atoms) and reassemble those components into materials that do not exist naturally and that often exhibit unusual properties. Thus, motor fuels and thousands of chemicals used in the manufacture of plastics, synthetic fabrics, pharmaceuticals, and pesticides can all be made from petroleum. Modern chemical knowledge is also needed to understand the processes that sustain life and to understand and control processes that are detrimental to the environment, such as the formation of smog and the destruction of stratospheric ozone. Because it relates to so many areas of human endeavor, chemistry is sometimes called the *central science*.

Early chemical knowledge consisted of the “how to” of chemistry, discovered through trial and error. Modern chemical knowledge answers the “why” as well as the “how to” of chemical change. It is grounded in principles and theory, and mastering the principles of chemistry requires a systematic approach to the subject. Scientific progress depends on the way scientists do their work—asking the right questions, designing the right experiments to supply the answers, and formulating plausible explanations of their findings. We begin with a closer look into the scientific method.

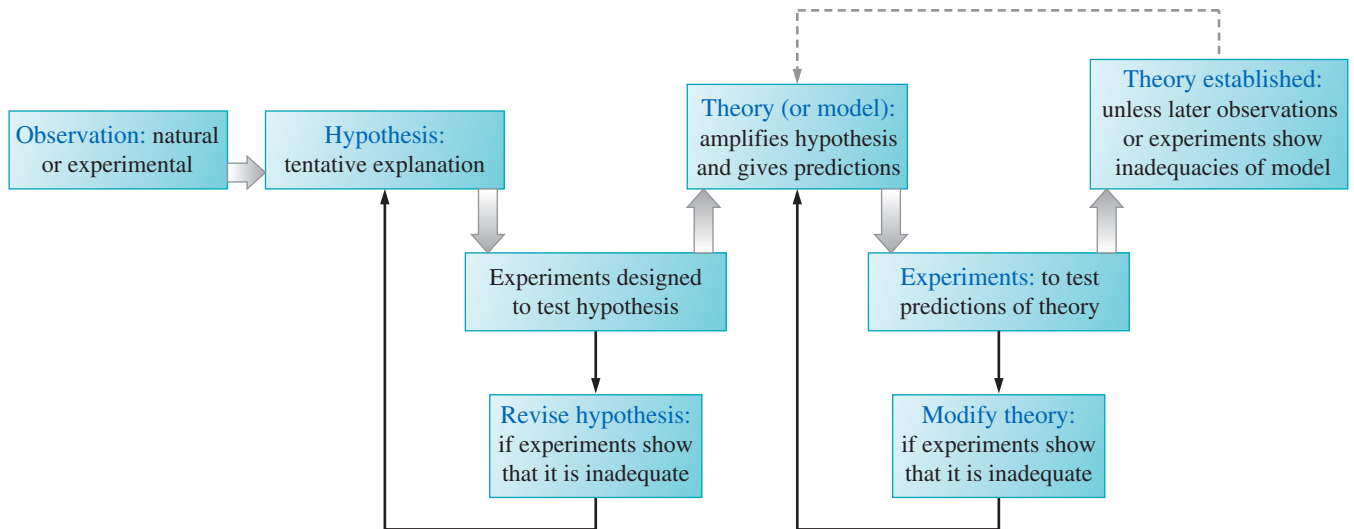
1-1 The Scientific Method

Science differs from other fields of study in the *method* that scientists use to acquire knowledge and the special significance of this knowledge. Scientific knowledge can be used to explain natural phenomena and, at times, to *predict* future events.

The ancient Greeks developed some powerful methods of acquiring knowledge, particularly in mathematics. The Greek approach was to start with certain basic assumptions or premises. Then, by the method known as *deduction*, certain conclusions must logically follow. For example, if $a = b$ and $b = c$, then $a = c$. Deduction alone is not enough for obtaining scientific knowledge, however. The Greek philosopher Aristotle *assumed* four fundamental substances: air, earth, water, and fire. All other materials, he believed, were formed by combinations of these four elements. Chemists of several centuries ago (more commonly referred to as alchemists) tried, in vain, to apply the four-element idea to turn lead into gold. They failed for many reasons, one being that the four-element assumption is false.

The scientific method originated in the seventeenth century with such people as Galileo, Francis Bacon, Robert Boyle, and Isaac Newton. The key to the method is to make no initial assumptions, but rather to make careful observations of natural phenomena. When enough observations have been made so that a pattern begins to emerge, a generalization or natural law can be formulated describing the phenomenon. **Natural laws** are concise statements, often in mathematical form, about natural phenomena. The form of reasoning in which a general statement or natural law is inferred from a set of observations is called *induction*. For example, early in the sixteenth century, Polish astronomer Nicolaus Copernicus (1473–1543), through careful study of astronomical observations, concluded that Earth revolves around the sun in a circular orbit, although the general teaching of the time, not based on scientific study, was that the sun and other heavenly bodies revolved around Earth. We can think of Copernicus’s statement as a natural law. Another example of a natural law is the radioactive decay law, which dictates how long it takes for a radioactive substance to lose its radioactivity.

The success of a natural law depends on its ability to explain, or account for, observations and to predict new phenomena. Copernicus’s work was a great success because he was able to predict future positions of the planets more accurately than his contemporaries. We should not think of a natural law as an *absolute* truth, however. Future experiments may require us to modify the law. For example, Copernicus’s ideas were refined a half-century later by Johannes Kepler, who showed that planets travel in elliptical, not circular, orbits. To verify



▲ FIGURE 1-1
The scientific method illustrated

a natural law, a scientist designs *experiments* that show whether the conclusions deduced from the natural law are supported by experimental results.

A **hypothesis** is a tentative explanation of a natural law. If a hypothesis survives testing by experiments, it is often referred to as a theory. In a broader sense, a **theory** is a model or way of looking at nature that can be used to explain natural laws and make further predictions about natural phenomena. When differing or conflicting theories are proposed, the one that is most successful in its predictions is generally chosen. Also, the theory that involves the smallest number of assumptions—the simplest theory—is preferred. Over time, as new evidence accumulates, most scientific theories undergo modification, and some are discarded.

The **scientific method** is the combination of observation, experimentation, and the formulation of laws, hypotheses, and theories. The method is illustrated by the flow diagram in Figure 1-1. Scientists may develop a pattern of thinking about their field, known as a *paradigm*. Some paradigms may be successful at first but then become less so. When that happens, a new paradigm may be needed or, as is sometimes said, a *paradigm shift* occurs. In a way, the method of inquiry that we call the scientific method is itself a paradigm, and some people feel that it, too, is in need of change. That is, the varied activities of modern scientists are more complex than the simplified description of the scientific method presented here.* In any case, merely following a set of procedures, rather like using a cookbook, will not guarantee scientific success.

Another factor in scientific discovery is chance, or serendipity. Many discoveries have been made by accident. For example, in 1839, American inventor Charles Goodyear was searching for a treatment for natural rubber that would make it less brittle when cold and less tacky when warm. During this work, he accidentally spilled a rubber–sulfur mixture on a hot stove and found that the resulting product had exactly the properties he was seeking. Other chance discoveries include X-rays, radioactivity, and penicillin. So scientists and inventors always need to be alert to unexpected observations. Perhaps no one was more aware of this than Louis Pasteur, who wrote, “Chance favors the prepared mind.”



Prof. Marvin Lang and Gary J. Shuller/University of Wisconsin–Stevens Point

▲ Louis Pasteur (1822–1895). This great practitioner of the scientific method was the developer of the germ theory of disease, the sterilization of milk by pasteurization, and vaccination against rabies. He has been called the greatest physician of all time by some. He was, in fact, not a physician at all, but a chemist—by training and by profession.



1-1 CONCEPT ASSESSMENT

Is the common saying “The exception proves the rule” a good statement of the scientific method? Explain.

◀ Answers to Concept Assessment questions are given in Appendix H.

*W. Harwood, *JCST*, 33, 29 (2004). *JCST* is an abbreviation for *Journal of College Science Teaching*.

1-2 Properties of Matter

Dictionary definitions of chemistry usually include the terms *matter*, *composition*, and *properties*, as in the statement that “chemistry is the science that deals with the composition and properties of matter.” In this and the next section, we will consider some basic ideas relating to these three terms in hopes of gaining a better understanding of what chemistry is all about.

Matter is anything that occupies space and displays the properties of *mass* and *inertia*. Every human being is a collection of matter. We all occupy space, and we describe our mass in terms of weight, a related property. (Mass and weight are described in more detail in Section 1-4. Inertia is described in Appendix B.) All the objects that we see around us consist of matter. The gases of the atmosphere, even though they are invisible, are matter—they occupy space and have mass. Sunlight is *not* matter; rather, it is a form of energy. Energy is discussed in later chapters.

Composition refers to the parts or components of a sample of matter and their relative proportions. Ordinary water is made up of two simpler substances—hydrogen and oxygen—present in certain fixed proportions. A chemist would say that the composition of water is 11.19% hydrogen and 88.81% oxygen by mass. Hydrogen peroxide, a substance used in bleaches and antiseptics, is also made up of hydrogen and oxygen, but it has a different composition. Hydrogen peroxide is 5.93% hydrogen and 94.07% oxygen by mass.

Properties are those qualities or attributes that we can use to distinguish one sample of matter from others; and, as we consider next, the properties of matter are generally grouped into two broad categories: physical and chemical.

Physical Properties and Physical Changes

A **physical property** is one that a sample of matter displays without changing its composition. Thus, we can distinguish between the reddish brown solid, copper, and the yellow solid, sulfur, by the physical property of *color* (Fig. 1-2).

Another physical property of copper is that it can be hammered into a thin sheet of foil (see Figure 1-2). Solids having this ability are said to be *malleable*. Sulfur is not malleable. If we strike a chunk of sulfur with a hammer, it crumbles into a powder. Sulfur is *brittle*. Another physical property of copper that sulfur does not share is the ability to be drawn into a fine wire (ductility). Also, sulfur is a far poorer conductor of heat and electricity than is copper.

Sometimes a sample of matter undergoes a change in its physical appearance. In such a **physical change**, some of the physical properties of the sample may change, but its composition remains unchanged. When liquid water freezes into solid water (ice), it certainly looks different and, in many ways, it is different. Yet the water remains 11.19% hydrogen and 88.81% oxygen by mass.

Chemical Properties and Chemical Changes

In a **chemical change**, or **chemical reaction**, one or more kinds of matter are converted to new kinds of matter with different compositions. The key to

► FIGURE 1-2
Physical properties of sulfur and copper

A lump of sulfur (left) crumbles into a yellow powder when hammered. Copper (right) can be obtained as large lumps of native copper, formed into pellets, hammered into a thin foil, or drawn into a wire.



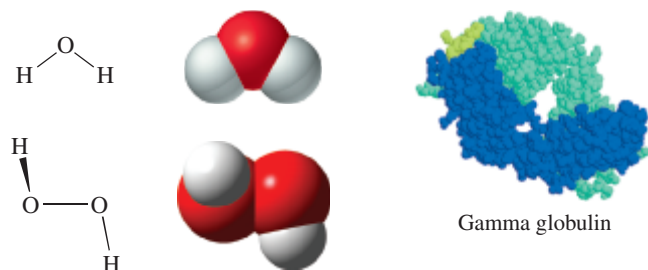
identifying chemical change, then, comes in observing a *change in composition*. The burning of paper involves a chemical change. Paper is a complex material, but its principal constituents are carbon, hydrogen, and oxygen. The chief products of the combustion are two gases, one consisting of carbon and oxygen (carbon dioxide) and the other consisting of hydrogen and oxygen (water, as steam). The ability of paper to burn is an example of a chemical property. A **chemical property** is the ability (or inability) of a sample of matter to undergo a change in composition under stated conditions.

Zinc reacts with hydrochloric acid solution to produce hydrogen gas and a solution of zinc chloride in water (Fig. 1-3). This reaction is one of zinc's distinctive chemical properties, just as the inability of gold to react with hydrochloric acid is one of gold's chemical properties. Sodium reacts not only with hydrochloric acid but also with water. In some of their physical properties, zinc, gold, and sodium are similar. For example, each is malleable and a good conductor of heat and electricity. In most of their chemical properties, though, zinc, gold, and sodium are quite different. Knowing these differences helps us to understand why zinc, which does not react with water, is used in roofing nails, roof flashings, and rain gutters, and sodium is not. Also, we can appreciate why gold, because of its chemical inertness, is prized for jewelry and coins: It does not tarnish or rust. In our study of chemistry, we will see why substances differ in properties and how these differences determine the ways in which we use them.

1-3 Classification of Matter

Matter is made up of very tiny units called **atoms**. Each different type of atom is the building block of a different chemical **element**. Presently, the International Union of Pure and Applied Chemistry (IUPAC) recognizes 118 elements, but four do not yet have names or symbols. The known elements range from common substances, such as carbon, iron, and silver, to uncommon ones, such as lutetium and thulium. About 90 of the elements can be obtained from natural sources. The remainder do not occur naturally and have been created only in laboratories. On the inside front cover you will find a complete listing of the elements and also a special tabular arrangement of the elements known as the *periodic table*. The periodic table is the chemist's directory of the elements. We will describe it in Chapter 2 and use it throughout most of the text.

Chemical **compounds** are substances comprising atoms of two or more elements joined together. Scientists have identified millions of different chemical compounds. In some cases, we can isolate a molecule of a compound. A **molecule** is the smallest entity having the same proportions of the constituent atoms as does the compound as a whole. A molecule of water consists of three atoms: two hydrogen atoms joined to a single oxygen atom. A molecule of hydrogen peroxide has two hydrogen atoms and two oxygen atoms; the two oxygen atoms are joined together and one hydrogen atom is attached to each oxygen atom. By contrast, a molecule of the blood protein gamma globulin is made up of 19,996 atoms, but they are of just four types: carbon, hydrogen, oxygen, and nitrogen.



▲ Structures of water, hydrogen peroxide, and gamma globulin. Gamma globulin consists of three subunits (shown in different colors). Each subunit consists of carbon, hydrogen, oxygen, and nitrogen.



Diane Hirschi/Fundamental Photographs

▲ FIGURE 1-3
A chemical property of zinc and gold: reaction with hydrochloric acid
The zinc-plated (galvanized) nail reacts with hydrochloric acid, producing the bubbles of hydrogen gas seen on its surface. The gold bracelet is unaffected by hydrochloric acid. In this photograph, the zinc plating has been consumed, exposing the underlying iron nail. The reaction of iron with hydrochloric acid imparts some color to the acid solution.

◀ The International Union of Pure and Applied Chemistry (IUPAC) is recognized as the world authority on chemical nomenclature, terminology, standardized methods for measurement, atomic mass, and more. Along with many other activities, IUPAC publishes journals, technical reports, and chemical databases, most of which are available at www.iupac.org.

◀ The identity of an atom is established by a feature called its atomic number (see Section 2-3). Characterizing "superheavy" elements is a daunting challenge; they are produced only a few atoms at a time and the atoms disintegrate almost instantaneously.



▲ Is it homogeneous or heterogeneous? When viewed through a microscope, homogenized milk is seen to consist of globules of fat dispersed in a watery medium. Homogenized milk is a *heterogeneous* mixture.

► It is composition, particularly its variability, that helps us distinguish the several classifications of matter.

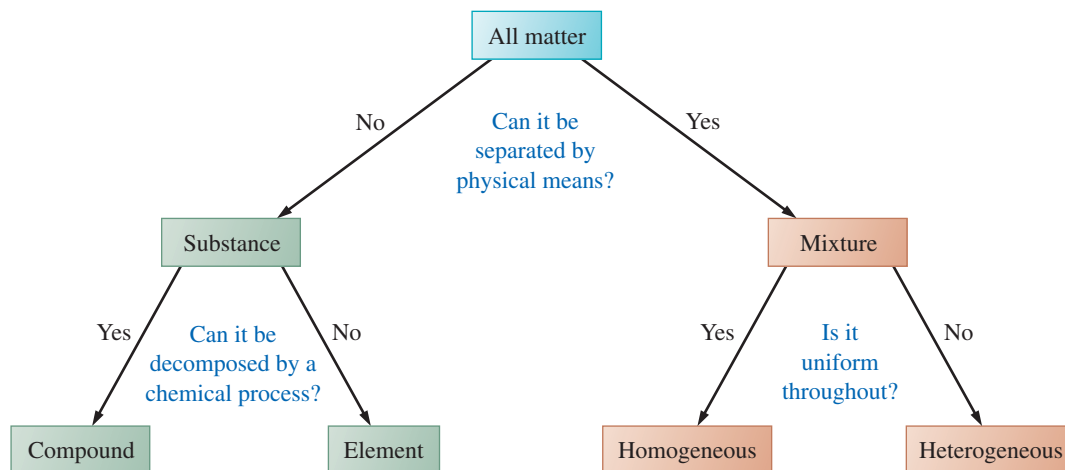
► Solutions can be gaseous and liquids as described here, but they can also be solids. Some alloys are examples of solid solutions.

The composition and properties of an element or a compound are uniform throughout a given sample and from one sample to another. Elements and compounds are called **substances**. (In the chemical sense, the term *substance* should be used only for elements and compounds.) A *mixture* of substances can vary in composition and properties from one sample to another. One that is uniform in composition and properties throughout is said to be a **homogeneous mixture** or a *solution*. A given solution of sucrose (cane sugar) in water is uniformly sweet throughout the solution, but the sweetness of another sucrose solution may be rather different if the sugar and water are present in different proportions. Ordinary air is a homogeneous mixture of several gases, principally the *elements* nitrogen and oxygen. Seawater is a solution of the *compounds* water, sodium chloride (salt), and a host of others. Gasoline is a homogeneous mixture or solution of dozens of compounds.

In **heterogeneous mixtures**—sand and water, for example—the components separate into distinct regions. Thus, the composition and physical properties vary from one part of the mixture to another. Salad dressing, a slab of concrete, and the leaf of a plant are all heterogeneous. It is usually easy to distinguish heterogeneous from homogeneous mixtures. A scheme for classifying matter into elements and compounds and homogeneous and heterogeneous mixtures is summarized in Figure 1-4.

Separating Mixtures

A mixture can be separated into its components by appropriate physical means. Consider again the heterogeneous mixture of sand in water. When we pour this mixture into a funnel lined with porous filter paper, the water passes through and sand is retained on the paper. This process of separating a solid from the liquid in which it is suspended is called *filtration* (Fig. 1-5a). You will probably use this procedure in the laboratory. Conversely, we cannot separate a homogeneous mixture (solution) of copper(II) sulfate in water by filtration because all components pass through the paper. We can, however, boil the solution of copper(II) sulfate and water. In the process of *distillation*, a pure liquid is condensed from the vapor given off by a boiling solution. When all



▲ FIGURE 1-4

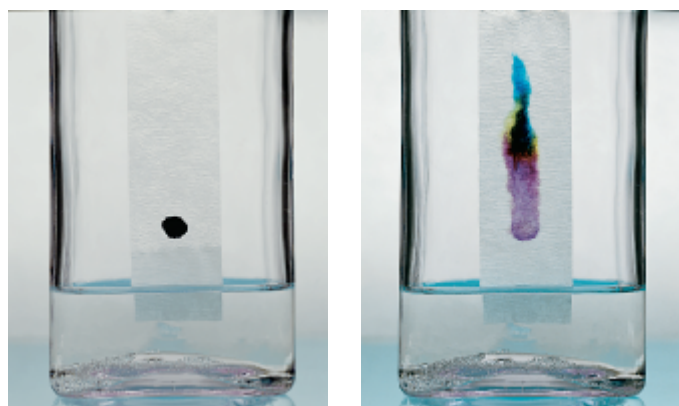
A classification scheme for matter

Every sample of matter is either a single substance (an element or compound) or a mixture of substances. At the molecular level, an element consists of atoms of a single type and a compound consists of two or more different types of atoms, usually joined into molecules. In a homogeneous mixture, atoms or molecules are randomly mixed at the molecular level. In heterogeneous mixtures, the components are physically separated, as in a layer of octane molecules (a constituent of gasoline) floating on a layer of water molecules.



(a)

(b)



(c)

(d)

◀ FIGURE 1-5

Separating mixtures: a physical process

(a) Separation of a heterogeneous mixture by filtration: Solid copper(II) sulfate is retained on the filter paper, while liquid hexane passes through. (b) Separation of a homogeneous mixture by distillation: Copper(II) sulfate remains in the flask on the left as water passes to the flask on the right, by first evaporating and then condensing back to a liquid. (c) Separation of the components of ink by using chromatography: A dark spot of black ink can be seen just above the water line as water moves up the paper. (d) Water has dissolved the colored components of the ink, and these components are retained in different regions on the paper according to their differing tendencies to adhere to the paper.

(a) Carey B. Van Loon; (b) Carey B. Van Loon; (c) Richard Megna/Fundamental Photographs; (d) Richard Megna/Fundamental Photographs

the water has been removed by boiling a solution of copper(II) sulfate in water, solid copper(II) sulfate remains behind (Fig. 1-5b).

Another method of separation available to modern chemists depends on the differing abilities of compounds to adhere to the surfaces of various solid substances, such as paper and starch. The technique of *chromatography* relies on this principle. The dramatic results that can be obtained with chromatography are illustrated by the separation of ink on a filter paper (Fig. 1-5c, d).

Decomposing Compounds

A chemical compound retains its identity during physical changes, but it can be decomposed into its constituent elements by *chemical changes*. The decomposition of compounds into their constituent elements is a more difficult matter than the mere physical separation of mixtures. The extraction of iron from iron oxide ores requires a blast furnace. The industrial production of pure magnesium from magnesium chloride requires electricity. It is generally easier to convert a compound into other compounds by a chemical reaction than it is to separate a compound into its constituent elements. For example, when heated, ammonium dichromate decomposes into the substances chromium(III) oxide, nitrogen, and water. This reaction, once used in movies to simulate a volcano, is illustrated in Figure 1-6.

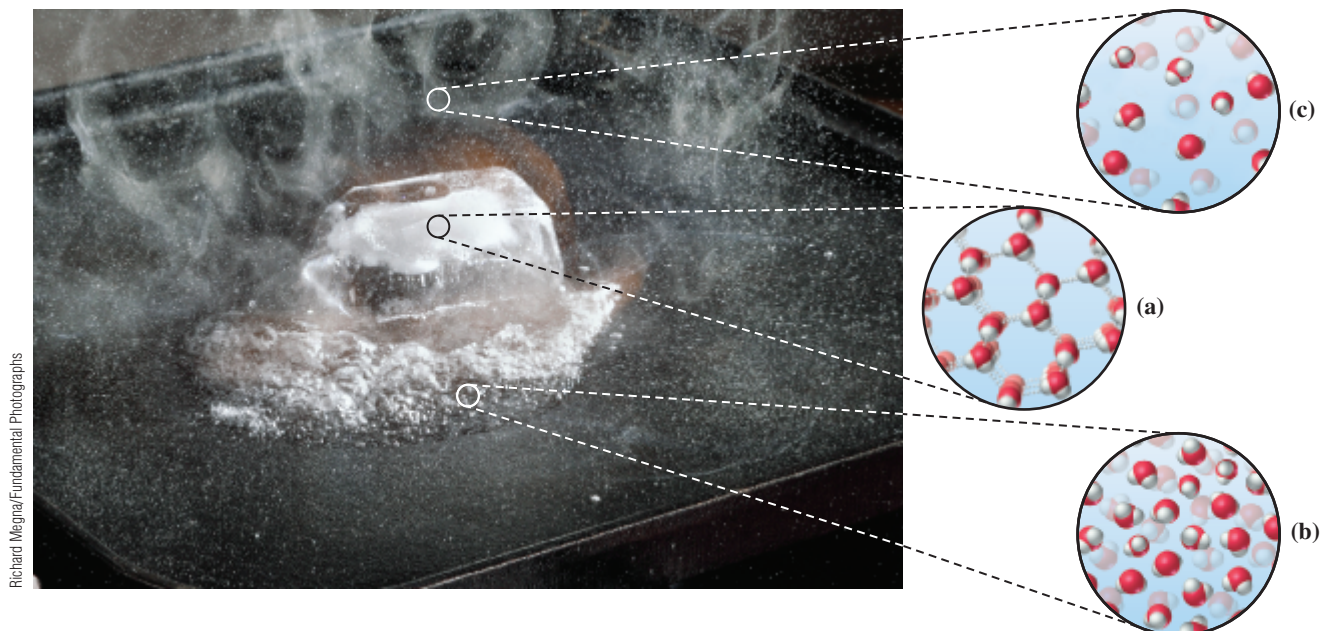
States of Matter

Matter is generally found in one of three *states*: solid, liquid, or gas. In a **solid**, atoms or molecules are in close contact, sometimes in a highly organized arrangement called a *crystal*. A solid has a definite shape. In a **liquid**, the atoms or molecules are usually separated by somewhat greater distances than in a solid. Movement of these atoms or molecules gives a liquid its most distinctive property—the ability to flow, covering the bottom and assuming the shape of its container. In a **gas**, distances between atoms or molecules are much greater



Carey B. Van Loon

▲ FIGURE 1-6
**A chemical change:
decomposition of
ammonium dichromate**



▲ FIGURE 1-7
Macroscopic and microscopic views of matter

The picture shows a block of ice on a heated surface and the three states of water. The circular insets show how chemists conceive of these states microscopically, in terms of molecules with two hydrogen atoms joined to one of oxygen. In ice (a), the molecules are arranged in a regular pattern in a rigid framework. In liquid water (b), the molecules are rather closely packed but move freely. In gaseous water (c), the molecules are widely separated.

than in a liquid. A gas always expands to fill its container. Depending on conditions, a substance may exist in only one state of matter, or it may be present in two or three states. Thus, as the ice in a small pond begins to melt in the spring, water is in two states: solid and liquid (actually, three states if we also consider water vapor in the air above the pond). The three states of water are illustrated at two levels in Figure 1-7.

The *macroscopic level* refers to how we perceive matter with our eyes, through the outward appearance of objects. The *microscopic level* describes matter as chemists conceive of it—in terms of atoms and molecules and their behavior. In this text, we will describe many macroscopic, observable properties of matter, but to explain these properties, we will often shift our view to the atomic or molecular level—the microscopic level.

1-4 Measurement of Matter: SI (Metric) Units

► Nonnumerical information is *qualitative*, such as the color blue.

► The *definition of the meter*, formerly based on the atomic spectrum of ^{86}Kr , was changed to the speed of light in 1983. Effectively, the speed of light is now defined as 2.99792458×10^8 m/s.

Chemistry is a *quantitative* science, which means that in many cases we can measure a property of a substance and compare it with a standard having a known value of the property. We express the measurement as the product of a *number* and a *unit*. The unit indicates the standard against which the measured quantity is being compared. When we say that the length of the playing field in football is 100 yd, we mean that the field is 100 times as long as a standard of length called the yard (yd). In this section, we will introduce some basic units of measurement that are important to chemists.

The scientific system of measurement is called the *Système Internationale d'Unités* (International System of Units) and is abbreviated **SI**. It is a modern version of the metric system, a system based on the unit of length called a *meter* (m). The meter was originally defined as 1/10,000,000 of the distance from the equator to the North Pole and translated into the length of a metal bar kept in Paris. Unfortunately, the length of the bar is subject to change with temperature, and it cannot be exactly reproduced. The SI system substitutes for the

TABLE 1.1 SI Base Quantities

Physical Quantity	Unit	Symbol
Length	meter ^a	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Amount of substance	mole ^b	mol
Electric current ^c	ampere	A
Luminous intensity ^d	candela	cd

^aThe official spelling of this unit is “metre,” but we will use the American spelling.

^bThe mole is introduced in Section 2-7.

^cElectric current is described in Appendix B and in Chapter 19.

^dLuminous intensity is not discussed in this text.

standard meter bar an unchanging, reproducible quantity: 1 meter is the distance traveled by light in a vacuum in $1/299,792,458$ of a second. Length is one of the seven fundamental quantities in the SI system (see Table 1.1). All other physical quantities have units that can be derived from these seven. SI is a *decimal* system. Quantities differing from the base unit by powers of ten are noted by the use of prefixes. For example, the prefix *kilo* means “one thousand” (10^3) times the base unit; it is abbreviated as k. Thus 1 *kilometer* = 1000 meters, or 1 km = 1000 m. The SI prefixes are listed in Table 1.2.

Most measurements in chemistry are made in SI units. Sometimes we must convert between SI units, as when converting kilometers to meters. At other times we must convert measurements expressed in non-SI units into SI units, or from SI units into non-SI units. In all these cases, we can use a *conversion factor* or a series of conversion factors in a scheme called a conversion pathway. Later in this chapter, we will apply conversion pathways in a method of problem solving known as *dimensional analysis*. The method itself is described in some detail in Appendix A.

Mass

Mass describes the quantity of matter in an object. In SI the standard of mass is 1 *kilogram* (kg), which is a fairly large unit for most applications in chemistry. More commonly we use the unit *gram* (g).

Weight is the force of gravity on an object. It is directly proportional to mass, as shown in the following mathematical expressions.

$$W \propto m \text{ and } W = g \times m \quad (1.1)$$

An object has a fixed mass (m), which is independent of where or how the mass is measured. Its weight (W), however, may vary because the acceleration caused by gravity (g) varies slightly from one point on Earth to another. Thus, an object that weighs 100.0 kg in St. Petersburg, Russia, weighs only 99.6 kg in Panama (about 0.4% less). The same object would weigh only about 17 kg on the moon. Although the weight of an object varies from place to place, its mass is the same in all locations. The terms *weight* and *mass* are often used interchangeably, but only mass is a measure of the quantity of matter. A common laboratory device for measuring mass is called a balance. A balance is often called, incorrectly, a scale.

The principle used in a balance is that of counteracting the force of gravity on an unknown mass with a force of equal magnitude that can be precisely measured. In older two-pan beam balances, the object whose mass is being determined is placed on one pan and counterbalancing is achieved through the force of gravity acting on *weights*, objects of precisely known mass, placed on the other pan. In the type of balance most commonly seen in laboratories today—the electronic balance—the counterbalancing force is a magnetic force produced by passing an electric current through an electromagnet. First, an initial balance condition is achieved when no object is present on the balance pan.

TABLE 1.2 SI Prefixes

Multiple	Prefix
10^{18}	exa (E)
10^{15}	peta (P)
10^{12}	tera (T)
10^9	giga (G)
10^6	mega (M)
10^3	kilo (k)
10^2	hecto (h)
10^1	deka (da)
10^{-1}	deci (d)
10^{-2}	centi (c)
10^{-3}	milli (m)
10^{-6}	micro (μ) ^a
10^{-9}	nano (n)
10^{-12}	pico (p)
10^{-15}	femto (f)
10^{-18}	atto (a)
10^{-21}	zepto (z)
10^{-24}	yocto (y)

^aThe Greek letter μ (pronounced “mew”).

◀ It is a good idea to *memorize the most common SI prefixes* (such as G, M, k, d, c, m, μ , n, and p) because you can’t survive in a world of science without knowing these SI prefixes.

◀ The symbol \propto means “proportional to.” It can be replaced by an equality sign and a proportionality constant. In expression (1.1), the constant is the acceleration caused by gravity, g . (See Appendix B.)