GLOBAL EDITION

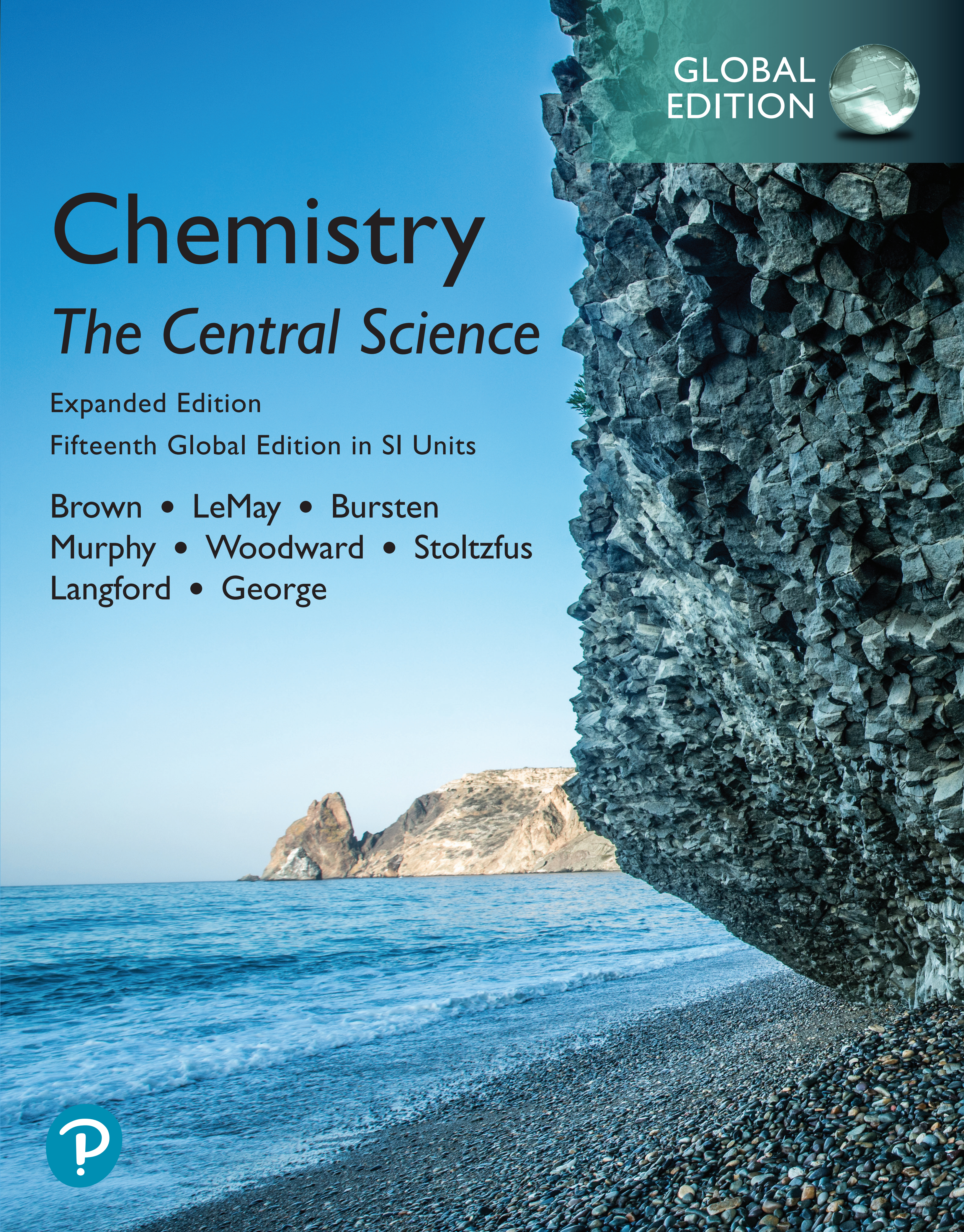
Chemistry

The Central Science

Expanded Edition

Fifteenth Global Edition in SI Units

Brown • LeMay • Bursten
Murphy • Woodward • Stoltzfus
Langford • George



Periodic Table of the Elements

Metals | Metalloids | Nonmetals

Period	1A[a] 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H** 1.00794																	2 **He** 4.002602
2	3 **Li** 6.941	4 **Be** 9.012182											5 **B** 10.811	6 **C** 12.0107	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.998403	10 **Ne** 20.1797
3	11 **Na** 22.989770	12 **Mg** 24.3050											13 **Al** 26.981538	14 **Si** 28.0855	15 **P** 30.973761	16 **S** 32.065	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.955910	22 **Ti** 47.867	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.938049	26 **Fe** 55.845	27 **Co** 58.933200	28 **Ni** 58.6934	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.64	33 **As** 74.92160	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.90585	40 **Zr** 91.224	41 **Nb** 92.90638	42 **Mo** 95.94	43 **Tc** [98]	44 **Ru** 101.07	45 **Rh** 102.90550	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.411	49 **In** 114.818	50 **Sn** 118.710	51 **Sb** 121.760	52 **Te** 127.60	53 **I** 126.90447	54 **Xe** 131.293
6	55 **Cs** 132.90545	56 **Ba** 137.327	71 **Lu** 174.967	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.23	77 **Ir** 192.217	78 **Pt** 195.078	79 **Au** 196.96655	80 **Hg** 200.59	81 **Tl** 204.3833	82 **Pb** 207.2	83 **Bi** 208.98038	84 **Po** [208.98]	85 **At** [209.99]	86 **Rn** [222.02]
7	87 **Fr** [223.02]	88 **Ra** [226.03]	103 **Lr** [262.11]	104 **Rf** [261.11]	105 **Db** [262.11]	106 **Sg** [266.12]	107 **Bh** [264.12]	108 **Hs** [269.13]	109 **Mt** [268.14]	110 **Ds** [281.15]	111 **Rg** [272.15]	112 **Cn** [285]	113 [284]	114 **Fl** [289.2]	115 [288]	116 **Lv** [293]	117 ** [294]	118 [294]

Main Group Representative Elements: groups 1–2 and 13–18. Transition metals: groups 3–12.

Lanthanide series	57 **La** 138.9055	58 **Ce** 140.116	59 **Pr** 140.90765	60 **Nd** 144.24	61 **Pm** [145]	62 **Sm** 150.36	63 **Eu** 151.964	64 **Gd** 157.25	65 **Tb** 158.92534	66 **Dy** 162.50	67 **Ho** 164.93032	68 **Er** 167.259	69 **Tm** 168.93421	70 **Yb** 173.04
Actinide series	89 **Ac** [227.03]	90 **Th** 232.0381	91 **Pa** 231.03588	92 **U** 238.02891	93 **Np** [237.05]	94 **Pu** [244.06]	95 **Am** [243.06]	96 **Cm** [247.07]	97 **Bk** [247.07]	98 **Cf** [251.08]	99 **Es** [252.08]	100 **Fm** [257.10]	101 **Md** [258.10]	102 **No** [259.10]

[a] The labels on top (1A, 2A, etc.) are common American usage. The labels below these (1, 2, etc.) are those recommended by the International Union of Pure and Applied Chemistry (IUPAC).

Except for elements 114 and 116, the names and symbols for elements above 113 have not yet been decided.

Atomic weights in brackets are the names of the longest-lived or most important isotope of radioactive elements.

Further information is available at http://www.webelements.com

** Discovered in 2010, element 117 is currently under review by IUPAC.

List of Elements with Their Symbols and Atomic Weights

Element	Symbol	Atomic Number	Atomic Weight	Element	Symbol	Atomic Number	Atomic Weight	Element	Symbol	Atomic Number	Atomic Weight
Actinium	Ac	89	227.03a	Hafnium	Hf	72	178.49	Praseodymium	Pr	59	140.90766
Aluminum	Al	13	26.981538	Hassium	Hs	108	269.1a	Promethium	Pm	61	145a
Americium	Am	95	243.06a	Helium	He	2	4.002602a	Protactinium	Pa	91	231.03588
Antimony	Sb	51	121.760	Holmium	Ho	67	164.93033	Radium	Ra	88	226.03a
Argon	Ar	18	39.948	Hydrogen	H	1	1.00794	Radon	Rn	86	222.02a
Arsenic	As	33	74.92160	Indium	In	49	114.818	Rhenium	Re	75	186.207a
Astatine	At	85	209.99a	Iodine	I	53	126.90447	Rhodium	Rh	45	102.90550
Barium	Ba	56	137.327	Iridium	Ir	77	192.217	Roentgenium	Rg	111	282.2a
Berkelium	Bk	97	247.07a	Iron	Fe	26	55.845	Rubidium	Rb	37	85.4678
Beryllium	Be	4	9.012183	Krypton	Kr	36	83.80	Ruthenium	Ru	44	101.07
Bismuth	Bi	83	208.98038	Lanthanum	La	57	138.9055	Rutherfordium	Rf	104	267.1a
Bohrium	Bh	107	270.1a	Lawrencium	Lr	103	262.11a	Samarium	Sm	62	150.36
Boron	B	5	10.81	Lead	Pb	82	207.2	Scandium	Sc	21	44.955908
Bromine	Br	35	79.904	Lithium	Li	3	6.941	Seaborgium	Sg	106	269.1a
Cadmium	Cd	48	112.414	Livermorium	Lv	116	293[a]	Selenium	Se	34	78.97
Calcium	Ca	20	40.078	Lutetium	Lu	71	174.967	Silicon	Si	14	28.0855
Californium	Cf	98	251.08a	Magnesium	Mg	12	24.3050	Silver	Ag	47	107.8682
Carbon	C	6	12.0107	Manganese	Mn	25	54.938044	Sodium	Na	11	22.989770
Cerium	Ce	58	140.116	Meitnerium	Mt	109	278.2a	Strontium	Sr	38	87.62
Cesium	Cs	55	132.905452	Mendelevium	Md	101	258.10a	Sulfur	S	16	32.065
Chlorine	Cl	17	35.453	Mercury	Hg	80	200.59	Tantalum	Ta	73	180.9479
Chromium	Cr	24	51.9961	Molybdenum	Mo	42	95.95	Technetium	Tc	43	98a
Cobalt	Co	27	58.933194	Moscovium	Mc	115	289.2a	Tellurium	Te	52	127.60
Copernicium	Cn	112	285.2[a]	Neodymium	Nd	60	144.24	Tennessine	Ts	117	293.2a
Copper	Cu	29	63.546	Neon	Ne	10	20.1797	Terbium	Tb	65	158.92534
Curium	Cm	96	247.07a	Neptunium	Np	93	237.05a	Thallium	Tl	81	204.3833
Darmstadtium	Ds	110	281.2a	Nickel	Ni	28	58.6934	Thorium	Th	90	232.0377
Dubnium	Db	105	268.1a	Nihonium	Nh	113	286.2[a]	Thulium	Tm	69	168.93422
Dysprosium	Dy	66	162.50	Niobium	Nb	41	92.90637	Tin	Sn	50	118.710
Einsteinium	Es	99	252.08a	Nitrogen	N	7	14.0067	Titanium	Ti	22	47.867
Erbium	Er	68	167.259	Nobelium	No	102	259.10a	Tungsten	W	74	183.84
Europium	Eu	63	151.964	Oganesson	Og	118	294.2a	Uranium	U	92	238.02891
Fermium	Fm	100	257.10a	Osmium	Os	76	190.23	Vanadium	V	23	50.9415
Flerovium	Fl	114	289.2a	Oxygen	O	8	15.9994	Xenon	Xe	54	131.293
Fluorine	F	9	18.9984016	Palladium	Pd	46	106.42	Ytterbium	Yb	70	173.04
Francium	Fr	87	223.02a	Phosphorus	P	15	30.973762	Yttrium	Y	39	88.90584
Gadolinium	Gd	64	157.25	Platinum	Pt	78	195.078	Zinc	Zn	30	65.39
Gallium	Ga	31	69.723	Plutonium	Pu	94	244.06a	Zirconium	Zr	40	91.224
Germanium	Ge	32	72.64	Polonium	Po	84	208.98a				
Gold	Au	79	196.966569	Potassium	K	19	39.0983				

[a]Mass of longest-lived or most important isotope.

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To our students,
whose enthusiasm and curiosity
have often inspired us,
and whose questions and suggestions
have sometimes taught us.

BRIEF CONTENTS

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PREFACE

To the Instructor

Philosophy

We the authors of *Chemistry: The Central Science* are delighted and honored that you have chosen us as your instructional partners for your chemistry class. Collectively we have taught chemistry to multiple generations of students. So we understand the challenges and opportunities of teaching a class that so many students take. We have also been active researchers who appreciate both the learning and the discovery aspects of the chemical sciences. Our varied, wide-ranging experiences have formed the basis of the close collaborations we have enjoyed as coauthors. In writing our book, our focus is on the students: we try to ensure that the text is not only accurate and up-to-date but also clear and readable. We strive to convey the breadth of chemistry and the excitement that scientists experience in making new discoveries that contribute to our understanding of the physical world. We want the student to appreciate that chemistry is not a body of specialized knowledge that is separate from most aspects of modern life, but central to any attempt to address a host of societal concerns, including renewable energy, environmental sustainability, and improved human health.

Publishing the fifteenth edition of this text bespeaks an exceptionally long record of successful textbook writing. We are appreciative of the loyalty and support the book has received over the years, and mindful of our obligation to justify each new edition. We begin our approach to each new edition with an intensive author retreat, in which we ask ourselves the deep questions that we must answer before we can move forward. What justifies yet another edition? What is changing in the world not only of chemistry, but with respect to science education and the qualities of the students we serve? How can we help your students not only learn the principles of chemistry, but also become critical thinkers who can think more like chemists? The answers lie only partly in the changing face of chemistry itself. The introduction of many new technologies has changed the landscape in the teaching of sciences at all levels. The use of the Internet in accessing information and presenting learning materials has markedly changed the role of the textbook as one element among many tools for student learning. Our challenge as authors is to maintain the text as the primary source of chemical knowledge and practice while at the same time integrating it with the new avenues for learning made possible by technology. This edition continues to incorporate a number of those new methodologies, including use of computer-based classroom tools, such as Learning Catalytics™, a cloud-based active learning analytics and assessment system, and web-based tools, particularly Pearson Mastering Chemistry, which is continually evolving to provide more effective means of testing and evaluating student performance, while giving the student immediate and helpful feedback. Pearson Mastering Chemistry not only provides feedback on a question by question basis but, using Knewton-enhanced adaptive follow-up assignments, it now continually adapts to each student, offering a personalized learning experience.

As authors, we want this text to be a central, indispensable learning tool for students. Whether as a physical book or in electronic form, it can be carried everywhere and used at any time. It is the best resource for students to obtain the information outside of the classroom needed for learning, skill development, reference, and test preparation. The text, more effectively than any other instrument, provides the depth of coverage and coherent background in modern chemistry that students need to serve their professional interests and, as appropriate, to prepare for more advanced chemistry courses.

If the text is to be effective in supporting your role as instructor, it must be addressed to the students. We have done our best to keep our writing clear and interesting and the book attractive and well illustrated. The book has numerous in-text study aids for students including carefully placed descriptions of problem-solving strategies. We hope that our cumulative experiences as teachers is evident in our pacing, choice of examples, and the kinds of study aids and motivational tools we have employed. We believe students are more enthusiastic about learning chemistry when they see its importance relative to their own goals and interests; therefore, we have highlighted many important applications of chemistry in everyday life. We hope you make use of this material.

It is our philosophy, as authors, that the text and all the supplementary materials provided to support its use must work in concert with you, the instructor. A textbook is only as useful to students as the instructor permits it to be. This book is replete with features that help students learn and that can guide them as they acquire both conceptual understanding and problem-solving skills. There is a great deal here for the students to use, too much for all of it to be absorbed by any student in a one-year course. You will be the guide to the best use of the book. Only with your active help will the students be able to utilize most effectively all that the text and its supplements offer. Students care about grades, of course, and with encouragement they will also become interested in the subject matter and care about learning. Please consider emphasizing features of the book that can enhance student appreciation of chemistry, such as the *Chemistry Put To Work* and *Chemistry and Life* boxes that show how chemistry impacts modern life and its relationship to health and life processes. Also consider emphasizing conceptual understanding (placing less emphasis on simple manipulative, algorithmic problem solving) and urging students to use the rich online resources available.

Organization and Contents

The first five chapters give a largely macroscopic, phenomenological view of chemistry. The basic concepts introduced—such as nomenclature, stoichiometry, and thermochemistry—provide necessary background for many of the laboratory experiments usually performed in chemistry. We believe that an early introduction to thermochemistry is desirable because so much of our understanding of chemical processes is based on considerations of energy changes. As before, we discuss bond enthalpies in the Thermochemistry chapter to emphasize the connection between the macroscopic properties of substances and the submicroscopic world of atoms and bonds. We believe this enables an effective, balanced approach to teaching thermodynamics in general chemistry, as well as provides students with an introduction to some of the global issues involving energy production and consumption. It is no easy matter to walk the narrow pathway between—on the one hand—trying to teach too much at too high a level and—on the other hand—resorting to oversimplifications. As with the book as a whole, the emphasis has been on imparting *conceptual* understanding, as opposed to presenting equations into which students are supposed to plug numbers.

The next four chapters (Chapters 6–9) deal with electronic structure and bonding. For more advanced students, *A Closer Look* boxes in Chapters 6 and 9 highlight radial probability functions and the phases of orbitals. Our approach of placing this latter discussion in *A Closer Look* box in Chapter 9 enables those who wish to cover this topic to do so, while others may wish to bypass it.

In Chapters 10–13, the focus of the text changes to the next level of the organization of matter: examining the states of matter. Chapters 10 and 11 deal with gases, liquids, and intermolecular forces, while Chapter 12 is devoted to solids, presenting a contemporary view of the solid state as well as of modern materials accessible to general chemistry students. The chapter provides an opportunity to show how abstract chemical bonding concepts impact real-world applications. The modular organization of the chapter allows instructors to tailor coverage to focus on the materials (semiconductors, polymers, nanomaterials, and so forth) that are most relevant to students and instructors alike. This section of the book concludes with Chapter 13, which covers the formation and properties of solutions.

The next several chapters examine the factors that determine the speed and extent of chemical reactions: kinetics (Chapter 14), equilibria (Chapters 15–17), thermodynamics (Chapter 19), and electrochemistry (Chapter 20). Also in this section is a chapter on environmental chemistry (Chapter 18), in which the concepts developed in preceding chapters are applied to a discussion of the atmosphere and hydrosphere. This chapter has increasingly come to be focused on green chemistry and the impacts of human activities on Earth's water and atmosphere.

After a discussion of nuclear chemistry (Chapter 21), the book has two survey chapters. Chapter 22 deals with nonmetals, and Chapter 23 with the chemistry of transition metals, including coordination compounds. These last three chapters are developed in an independent, modular fashion and can be covered in any order.

Organic chemistry is central to all living things and Chapters 24–32 lead us on a journey from elementary hydrocarbons to elaborate bio-organic molecules. Much of what we discuss is treated from a fundamental level so students' transition to tertiary studies in organic chemistry is smooth and rapid. We place emphasis on the core reactions observed in organic chemistry and treat many cases mechanistically. This fosters a deep understanding of why organic molecules react in the way they do, thereby giving students an opportunity to understand much more chemistry than is discussed.

Chapter 24 provides a foundation to our examination of organic chemistry by using hydrocarbons to illustrate how we represent and name organic molecules. It goes on to provide an overview of the functional groups—the reactive parts of the molecule—on which we build our understanding of organic chemistry. The shape of a molecule may be pivotal in determining its reactivity, particularly in a biological context, and Chapter 25 leads to an in-depth discussion of stereochemistry. The next six chapters cover the fundamental reactions encountered in organic chemistry, at each step building to the application of these reactions in a modern world (for example, polymerisation in Chapters 26 and 29) and their essential role in the chemistry of life (for example, carbohydrates in Chapter 28, fats in Chapter 29, proteins and nucleic acids in Chapter 31). Chapter 30 investigates aromatic compounds as a separate class. Here, it is important for the student to note the differences in reactivity to the alkenes studied in Chapter 26.

Finally, Chapter 32 stands alone as a reference guide to mass spectrometry, NMR spectroscopy, and IR spectroscopy. Whether these topics are taught with much emphasis on the technology is up to the instructor. What we believe is most important is students' development at complex problem-solving, bringing two or more concepts together to draw a logical conclusion. The approach to solving molecular structure also confirms their knowledge of the basic principles of organic chemistry, bonding, functional groups and drawing structural formulas. Our coverage of organic chemistry gives students a unique perspective and challenges the very 'standard format' often seen in a first-year text.

Our chapter sequence provides a fairly standard organization, but we recognize that not everyone teaches all the topics in the order we have chosen. We have, therefore, made sure that instructors can make common changes in teaching sequence with no loss in student comprehension. In particular, many instructors prefer to introduce gases (Chapter 10) after stoichiometry (Chapter 3) rather than with states of matter. The chapter on gases has been written to permit this change with *no* disruption in the flow of material. It is also possible to treat balancing redox equations (Sections 20.1 and 20.2) earlier, after the introduction of redox reactions in Section 4.4.

We have brought students into greater contact with descriptive organic and inorganic chemistry by integrating examples throughout the text. Students will find pertinent and relevant examples of "real" chemistry woven into all the chapters to illustrate principles and applications. Some chapters, of course, more directly address the "descriptive" properties of elements and their compounds, especially Chapters 4, 7, 11, 18, 22, and 23. We also incorporate descriptive organic and inorganic chemistry in the exercises found throughout each chapter.

New to This Edition

It is perhaps a natural tendency for chemistry textbooks to grow in length with succeeding editions, but it is one that we have resisted. There are, nonetheless, many updates to features to serve students and instructors better in the classroom. *Chemistry: The Central Science* has traditionally been valued for its clarity of writing, its scientific accuracy and currency, its strong end-of-chapter exercises, and its consistency in level of coverage. The book was updated in a way that did not compromise these characteristics, and we have also continued to employ an open, clean design in the layout of the book.

The art program for the fifteenth edition continues the trajectory set in the previous two editions: to make greater and more effective use of the figures as learning tools, by drawing the reader more directly into the figure. The style of the art enhances clarity with a clean and modern look. This includes white-background annotation boxes with crisp, thin leaders; rich and saturated colors in the art, and use of 3D renderings. Using statistics from Pearson Mastering Chemistry, we have shifted some Exercises to the ends of sections, where students are more likely to attempt them before moving on to more complex questions. Also in the ends of sections are new Self-Assessment Exercises that provide immediate assessment and feedback content in the form of multiple-choice questions meant to test the concepts learnt in the section. In the Pearson eText, these exercises provide specific wrong-answer feedback.

Updates to subject matter in chapter text, Sample Exercises, and assessment content reflect current trends in teaching chemistry.

Each section now opens with new section-opening text and images that enhance students' understanding of the concepts introduced in that section as well as explicate the historical contexts around key inventions and discoveries in chemistry.

This edition features eight detailed chapters on organic chemistry for instructors and students who have more in-depth course discussions on organic chemistry than those covered in the shorter, 24-chapter variant of this book. An additional chapter on spectrometry is also available. All these additional chapters come with the wealth of Sample Exercises, essay features, assessment content, and updated art that has made the title a favorite with students and instructors the world over.

- The essays titled *Strategies in Chemistry,* which provide advice to students on problem solving and "thinking like a chemist," have been renamed *Strategies for Success* to better convey their usefulness to the student.

Key Features in This Edition

Chemistry: The Central Science, continues to provide relevant, up-to-date content—be it art or assessment material—that enhances the clarity and effectiveness of the text. Key features for this edition include the following:

- The treatment of energy and thermochemistry draws on significant revisions to previous editions. The introduction of the concept of energy in Chapter 1 allows instructors greater freedom in the order in which they cover the material. For example, this arrangement facilitates coverage of Chapters 6 and 7 immediately following Chapter 2, a sequence that is in line with an atoms-first approach to teaching general chemistry. The discussion of bond enthalpies in Chapter 5 emphasizes the connection between macroscopic quantities, like reaction enthalpies, and the submicroscopic world of atoms and bonds. We feel this leads to a better integration of thermochemical concepts with the surrounding chapters. Bond enthalpies are revisited in Chapter 8 after students have developed a more sophisticated view of chemical bonding.
- The text continues to provide students with a clear discussion, superior problem sets, and better real-time feedback on students' understanding of the material. This is based on the authors' insight into student usage of the interactive e-book platform, such as the most frequently highlighted passages and the accompanying notes and questions.
- Extensive effort has gone into creating enhanced content for the Pearson eText for the book. These features make the eText so much more than just an electronic copy of the physical textbook. Self-Assessment Exercises at the end of each section are enhanced with specific wrong-answer feedback in the Pearson eText. New Smart Figures take key figures from the text and bring them to life through animation and narration. Smart Sample Exercises animate key sample exercises from the text, offering students a more in-depth and detailed discussion than can be provided in the printed text. These interactive features also include follow-up questions, which can be assigned in Pearson Mastering Chemistry.
- Finally, Subtle but important changes have been made to allow students to quickly reference important concepts and assess their knowledge of the material. Key points are set in italic with line spaces above and below for greater emphasis. The skills-based *How To . . .* features offer step-by-step guidance for solving specific types of problems such as Drawing Lewis Structures, Balancing Redox Equations, and Naming Acids. These features, with numbered steps encased by a thin rule, are integrated into the main discussion and are easy to find. Finally, each Learning Objective is now correlated to specific end-of-chapter exercises. This allows students to test their mastery of each learning objective when preparing for quizzes and exams.

We have continued to emphasize conceptual exercises in the end-of-chapter, problems. In each chapter, we begin the exercises with the well-received *Visualizing Concepts* category. These exercises are designed to facilitate conceptual understanding through use of models, graphs, photographs, and other visual materials. They precede the regular end-of-chapter exercises and are identified in each case with the relevant chapter section number. A generous selection of *Integrative Exercises,* which give students the opportunity to solve problems that integrate concepts from the present chapter with those of previous chapters, is included at the end of each chapter. The importance of integrative problem solving is highlighted by the *Sample Integrative Exercise,* which ends each chapter beginning with Chapter 4. In general, we have included more conceptual end-of-chapter exercises and have made sure that there is a good representation of somewhat more difficult exercises to provide a better mix in terms of topic and level of difficulty. Many of the exercises are structured in a way that makes it easy to use them in Pearson Mastering Chemistry. We have made extensive use of the metadata from student use of Pearson Mastering Chemistry to analyze end-of-chapter exercises and make appropriate changes, as well as to develop *Learning Outcomes* for each chapter.

The essays in our well-received *Chemistry Put To Work* and *Chemistry and Life* series emphasize world events, scientific discoveries, and medical breakthroughs relevant to topics developed in each chapter. We maintain our focus on the positive aspects of chemistry without neglecting the problems that can arise in an increasingly technological world. Our goal is to help students appreciate the real-world perspective of chemistry and the ways in which chemistry affects their lives.

To the Student

Chemistry: The Central Science, Fifteenth Edition, has been written to introduce you to modern chemistry. As authors, we have, in effect, been engaged by your instructor to help you learn chemistry. Based on the comments of students and instructors who have used this book in its previous editions, we believe that we have done that job well. Of course, we expect the text to continue to evolve through future editions. We invite you to write to tell us what you like about the book so that we will know where we have helped you most. Also, we would like to learn of any shortcomings so we may further improve the book in subsequent editions. Our addresses are given at the end of the Preface.

Advice for Learning and Studying Chemistry

Learning chemistry requires both the assimilation of many concepts and the development of analytical skills. In this text, we have provided you with numerous tools to help you succeed in both tasks. If you are going to succeed in your chemistry course, you will have to develop good study habits. Science courses, and chemistry in particular, make different demands on your learning skills than do other types of courses. We offer the following tips for success in your study of chemistry:

Don't fall behind! As the course moves along, new topics will build on material already presented. If you don't keep up in your reading and problem solving, you will find it much harder to follow the lectures and discussions on current topics. Experienced teachers know that students who read the relevant sections of the text *before* coming to a class learn more from the class and retain greater recall. "Cramming" just before an exam has been shown to be an ineffective way to study any subject, chemistry included. So now you know. How important to you, in this competitive world, is a good grade in chemistry?

Focus your study. The amount of information you will be expected to learn may seem overwhelming. It is essential to recognize those concepts and skills that are particularly important. Pay attention to what your instructor is emphasizing. As you work through the *Sample Exercises* and homework assignments, try to see what general principles and skills they employ. A single reading of a chapter will generally not be enough for successful learning of chapter concepts and problem-solving skills. You will often need to go over assigned materials more than once. Don't skip the *Go Figure* features, *Sample Exercises,* and *Practice Exercises.* These are your guides to whether you are learning the material. They are also good preparation for test-taking. The *Learning Outcomes* and *Key Equations* at the end of the chapter will also help you focus your study.

Keep good lecture notes. Your lecture notes will provide you with a clear and concise record of what your instructor regards as the most important material to learn. Using your lecture notes in conjunction with this text is the best way to determine which material to study.

Skim topics in the text before they are covered in lecture. Reviewing a topic before lecture will make it easier for you to take good notes. First read the end-of-chapter *Summary;* then quickly read through the chapter, skipping Sample Exercises and supplemental sections. Paying attention to the titles of sections and subsections gives you a feeling for the scope of topics. Try to avoid thinking that you must learn and understand everything right away.

You need to do a certain amount of preparation before lecture. More than ever, instructors are using the lecture period not simply as a one-way channel of communication from teacher to student. Rather, they expect students to come to class ready to work on problem solving and critical thinking. Coming to class unprepared is not a good idea for any lecture environment, but it certainly is not an option for an active learning classroom if you aim to do well in the course.

After lecture, carefully read the topics covered in class. As you read, pay attention to the concepts presented and to the application of these concepts in the *Sample Exercises.* Once you think you understand a *Sample Exercise,* test your understanding by working the accompanying *Practice Exercise.*

Learn the language of chemistry. As you study chemistry, you will encounter many new words. It is important to pay attention to these words and to know their meanings or the entities to which they refer. Knowing how to identify chemical substances from their names is an important skill; it can help you avoid painful mistakes on examinations. For example, "chlorine" and "chloride" refer to very different things.

Attempt the assigned end-of-chapter exercises. Working the exercises selected by your instructor provides necessary practice in recalling and using the essential ideas of the chapter. You cannot learn merely by observing; you must be a participant. If you get stuck on an exercise, however, get help from your instructor, your teaching assistant, or another student. Spending more than 20 minutes on a single exercise is rarely effective unless you know that it is particularly challenging.

Learn to think like a scientist. This book is written by scientists who love chemistry. We encourage you to develop your critical thinking skills by taking advantage of features in this new edition, such as exercises that focus on conceptual learning, and the *Design an Experiment* exercises.

Use online resources. Some things are more easily learned by discovery, and others are best shown in three dimensions. If your instructor has included Pearson Mastering Chemistry with your book, take advantage of the unique tools it provides to get the most out of your time in chemistry.

The bottom line is to work hard, study effectively, and use the tools available to you, including this textbook. We want to help you learn more about the world of chemistry and why chemistry is the central science. If you really learn chemistry, you can be the life of the party, impress your friends and parents, and . . . well, also pass the course with a good grade.

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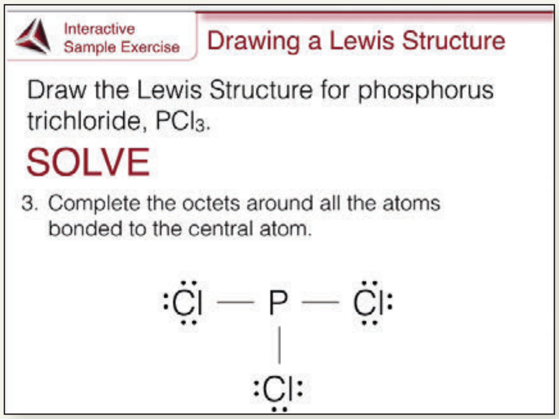
Adrian V. George received his BSc(Hons) and PhD degrees from The University of Reading in England and joined the staff there as a lecturer in 1984. After a short spell as a guest scientist at The University of California, Berkeley he moved to The University of Sydney in 1988. His research has ranged from organic synthesis at extremely high pressures and the development of new organometallic materials to the use of isotope ratio mass spectrometry in the detection of doping in competitive sports and chemistry education. He has conducted research in Japan and taught University level chemistry in Sweden. He has always had a passion for teaching and obtained a graduate certificate of education in 2000. He has been awarded a University of Sydney Excellence in Teaching award (1999), Vice Chancellor's award for Support of the Student Experience twice (2007, 2011), the inaugural Royal Australian Chemical Institute Centenary of Federation Teaching Award (2001), Australian College of Education Teaching Award (2001) and was part of a team that received the Carrick Institute Award for Programs that Enhance Learning (2007). He has been Director of First Year Studies in the School of Chemistry and the Associate Dean (Teaching and Learning) in the Faculty of Science at The University of Sydney. He currently divides his time between academic pursuits at The University of Sydney and rain forest regeneration in northern New South Wales.

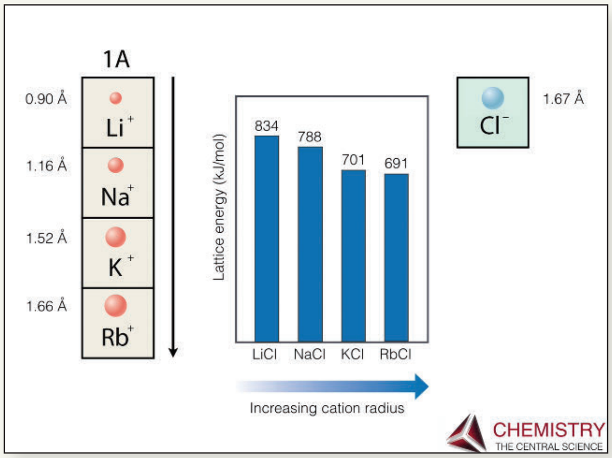
Michael W. Lufaso received his B.S. degree in Chemistry from Youngstown State University in 1998 and his Ph.D. in Chemistry from the Ohio State University in 2002. He was a National Research Council Postdoctoral Fellow at the National Institute for Standards and Technology and a postdoctoral fellow at the University of South Carolina. In 2006 he joined the University of North Florida where he currently holds the rank of Associate Professor in the Department of Chemistry. He was a Brian Andreen Cottrell College Science Award winner from Research Corporation. He was named a Munoz Presidential Professor in 2011 and received an Outstanding Faculty Scholarship award in 2014. He has authored laboratory manuals and taught ten different undergraduate courses primarily in the areas of general, inorganic, and solid state chemistry. His undergraduate research program focuses on structure prediction, synthesis, and characterization of the structure and properties of solid state materials.

New Levels of Student Interaction for Improved Conceptual Understanding

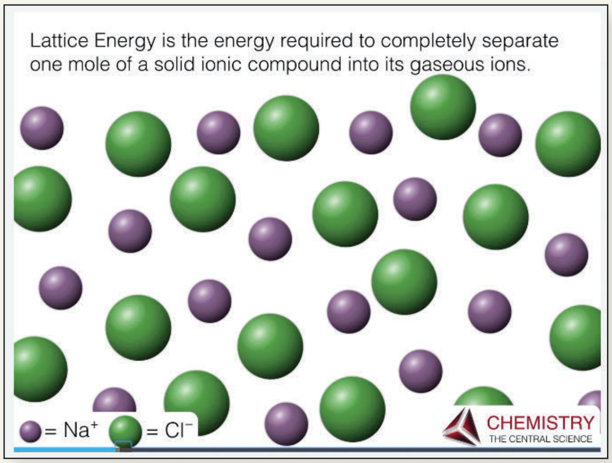
Assignable in Pearson Mastering Chemistry, unique features engage students through interactivity to enhance the reading experience and help them learn challenging chemistry concepts.



Interactive Sample Exercises bring key *Sample Exercises* in the text to life through animation and narration. Author Matt Stoltzfus uses the text's *Analyze–Plan–Solve–Check* technique to guide students through the problem-solving process. Play icons within the text identify each Interactive Sample Exercise. Clicking the icon in the eText launches a visual and conceptual presentation which goes beyond the static page. The *Practice Exercises* within each *Sample Exercise* can also be assigned in Pearson Mastering Chemistry where students will receive answer-specific feedback.

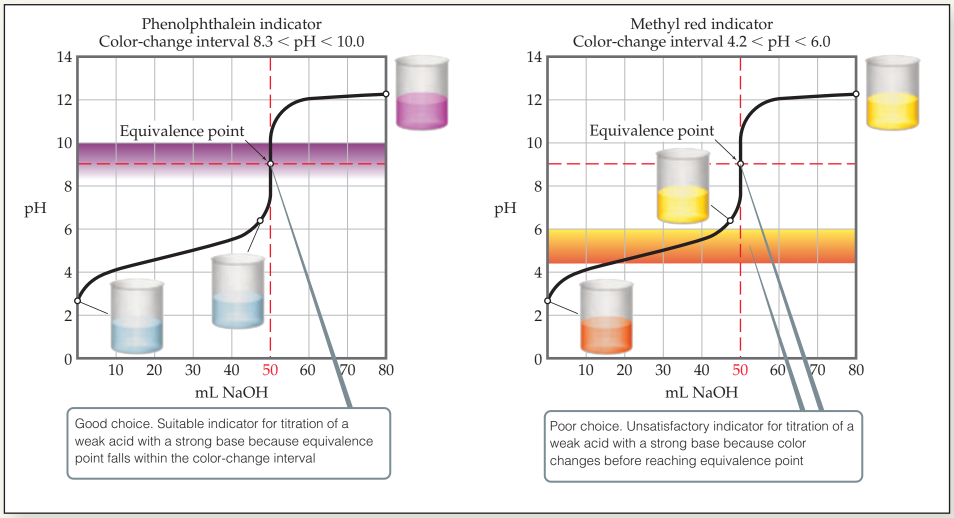


Smart Figures walk students through complex visual representations, dispelling common misconceptions before they take root. Each *Smart Figure* converts a static in-text figure into a dynamic process narrated by author Matt Stoltzfus. Play icons within the text identify each *Smart Figure*. Clicking the icon in the eText launches the animation. Smart Figures are assignable in Pearson Mastering Chemistry where they are accompanied by a multiple-choice question with answer-specific video feedback. Selecting the correct answer launches a brief wrap-up video that highlights the key concepts behind the answer.



Visually Revised to Better Help Students Build Chemistry

The visual program enhances clarity with a clean, modern look. Style changes include: expanded use of 3D renderings, new white annotation boxes with crisp leader lines, and a more saturated art palette.



Annotations offer expanded explanations; additional new leaders emphasize key relationships and key points in figures.

Before and after photos clearly show characteristics of endothermic and exothermic reactions. Added reaction equations connect the chemistry to what's happening in the photos.

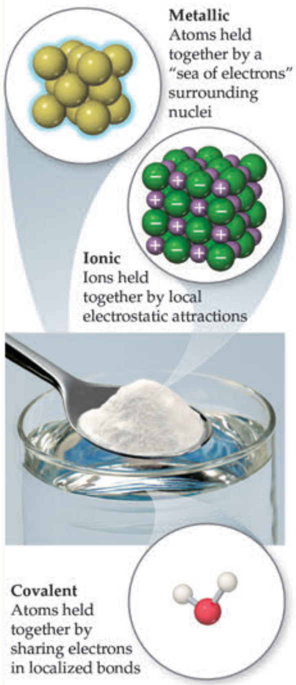
◀ **Figure 5.8 Endothermic and exothermic reactions.** In both instances, the system is defined as the reactants and products, while surroundings are the containers and everything else in the universe.

Knowledge and Understanding

The authors used the wealth of student data in Pearson Mastering Chemistry to identify the areas where students struggle most, revising discussions, figures, and exercises throughout the text to address misconceptions and encourage thinking about the real-world use of chemistry.

Go Figure

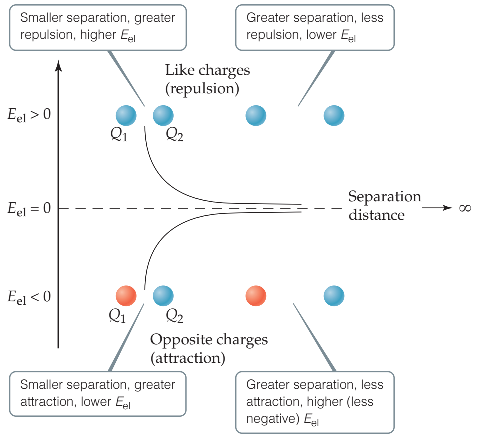
If the white powder were sugar, $C_{12}H_{22}O_{11}$, how would we have to change this picture?



▲ **Figure 8.1 Ionic, covalent, and metallic bonds.** The three different substances shown here are held together by different types of chemical bonds.

Go Figure

A positively charged particle and a negatively charged particle are initially far apart. What happens to their electrostatic potential energy as they are brought closer together?



▲ **Figure 5.2 Electrostatic potential energy.** At finite separation distances the electrostatic potential energy, E_{el}, is positive for objects with like charges, and negative for objects that are oppositely charged. As the particles move farther apart, their electrostatic potential energy approaches zero.

The author team utilized Mastering metadata to edit and clarify in-chapter *Go Figure* questions, as well as end-of-chapter problems. User data helped them to identify problematic questions and then modify, replace, or delete—resulting in a more diverse and polished set of problems.

A CLOSER LOOK Lead Contamination in Drinking Water

Access to clean drinking water is something most people in the industrialized world take for granted. Unfortunately, there are rare instances in which tap water is not safe to drink, as illustrated by the discovery in 2015 of elevated levels of lead in the municipal water supply of Flint, a city in Michigan State, USA.

Lead is detrimental to many organs in the human body, but the brain and central nervous system are particularly sensitive to its presence. In the brain, Pb^{2+} ions interfere with cell communication and growth by mimicking Ca^{2+} ions. One of the most serious side effects of lead poisoning occurs in young children, where it leads to cognitive impairment. Although lead compounds were once used in a variety of applications—as a gasoline additive, in pigments, shotgun pellets, glass, and water pipes—our daily exposure to lead dropped dramatically once governmental agencies started regulating its use in the 1970s. According to the National Health and Nutrition Examination Survey, the mean concentration of lead in the blood for an average U.S. resident dropped from 150 ppb in 1976 to 16 ppb by 2002, nearly an order-of-magnitude decrease.

The regulatory limit set by the U.S. Environmental Protection Agency (EPA) for lead in drinking water is 15 parts per billion (ppb). According to EPA regulations, utilities serving more than 50,000 people must monitor the level of lead in their water and take corrective action if more than 10% of the homes sampled exceed the 15 ppb limit. Tests performed on samples collected in September 2015 by researchers from Virginia Tech found that in 10% of the 252 Flint homes tested the lead concentration exceeded 25 ppb, and in several homes the concentration exceeded 100 ppb. At the same time, a local pediatrician analyzed the results of infant blood tests and found that the percentage of children with elevated concentrations of lead in their bloodstream (>50 ppb) had doubled from 2.4% in the 2013 to 4.9% in 2015.

The troubles began in April 2014 when the city began using the nearby Flint River as the natural source of its municipal water. Prior to that, Flint obtained water from Detroit, where water taken from Lake Huron was treated before piping it to Flint. The source of the lead was not the Flint River itself, but corrosion from lead pipes that are present in the underground water distribution network. Why did the change in water supply dramatically increase the leaching of lead from the old pipes? Many factors were at play, but basically they all came down to solubility considerations.

When water is properly treated, a passivation layer of insoluble lead salts builds up on the inner surface of the lead pipes (Figure 17.21). This layer prevents corrosion that would otherwise allow lead to be oxidized and dissolve into the water as Pb^{2+} ions. The water treatment facility in Detroit was adding phosphate ions, PO_4^{3-}, to their water to inhibit corrosion, whereas the people managing water treatment in Flint elected not to do so. The presence of PO_4^{3-} ions promotes the formation of highly insoluble phosphate salts on the inner surface of the pipes that helps prevent corrosion.

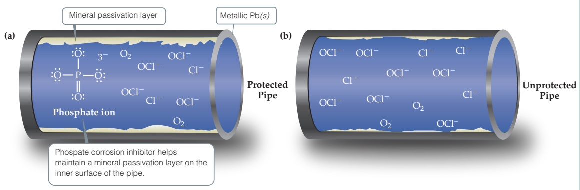
Another factor that appears to have contributed to the problem is a drop in the pH of the water, from 8.0 in December 2014 to 7.3 in August 2015. Because the insoluble lead salts that form the passivating layer, like $Pb_3(PO_4)_2$, $PbHSO_4$, and $PbCO_3$, contain anions that can act as weak bases, anything that makes the water more acidic increases their solubility.

Another contributing factor was the presence of high levels of chloride ions. The treated water from Detroit had chloride levels of about 11 ppm, whereas the treated Flint water had chloride levels of 85 ppm in August 2015. While $PbCl_2$ is fairly insoluble ($K_{sp} = 1.7 \times 10^{-5}$), high concentrations of chloride ions can lead to the formation of soluble complex ions such as $PbCl_3^-$ and $PbCl_4^{2-}$. The increased chloride levels were due in part to the addition of $FeCl_3$, which was used to help coagulate and filter out unwanted organic matter that was leading to problems with *E. coli* contamination. Chloride ions are also produced when unwanted organic matter is oxidized by hypochlorite ions, which are added to kill bacteria. Runoff containing chloride salts used to treat icy roads in the winter may have also contributed.

In the face of mounting evidence that the lead in Flint's drinking water was at unsafe levels, in October of 2015 the city switched back to water piped in from Detroit. Over a period of many months, the circulation of properly treated water should restore the passivating layers inside the pipes. However, cleanup costs are expected to be in excess of $120 million and possibly much more, a large sum compared to the estimated $50,000 per year that it would have cost to treat the water with phosphates. More importantly, the damage done to people who drank the contaminated water, especially children, cannot be undone.

Although the use of lead in plumbing has been banned in the United States since 1986, it is estimated that millions of kilometers of buried lead pipe are still in use in America's cities. Vigilance by water treatment facilities and environmental protection agencies is needed to avoid a repeat of the tragedy in Flint.

Related Exercises: 17.107, 17.111



▲ **Figure 17.21 Protected and unprotected lead pipes.** (**a**) A lead pipe that has a protective passivation layer, and (**b**) a lead pipe where the lack of a phosphate corrosion inhibitor causes the passivation layer to dissolve and fall off, exposing the lead to oxidizing agents such as O_2 and OCl^-.

A Closer Look features reflect news and discoveries in the field of chemistry, providing relevance and applications for students. End-of-chapter questions give students the chance to test whether they understand the concept or not.

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  * SmartFigures walk students through complex visual representations, dispelling common misconceptions before they take root.
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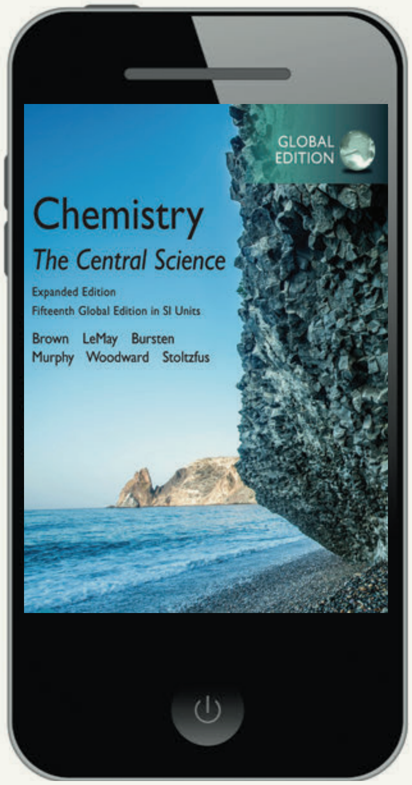

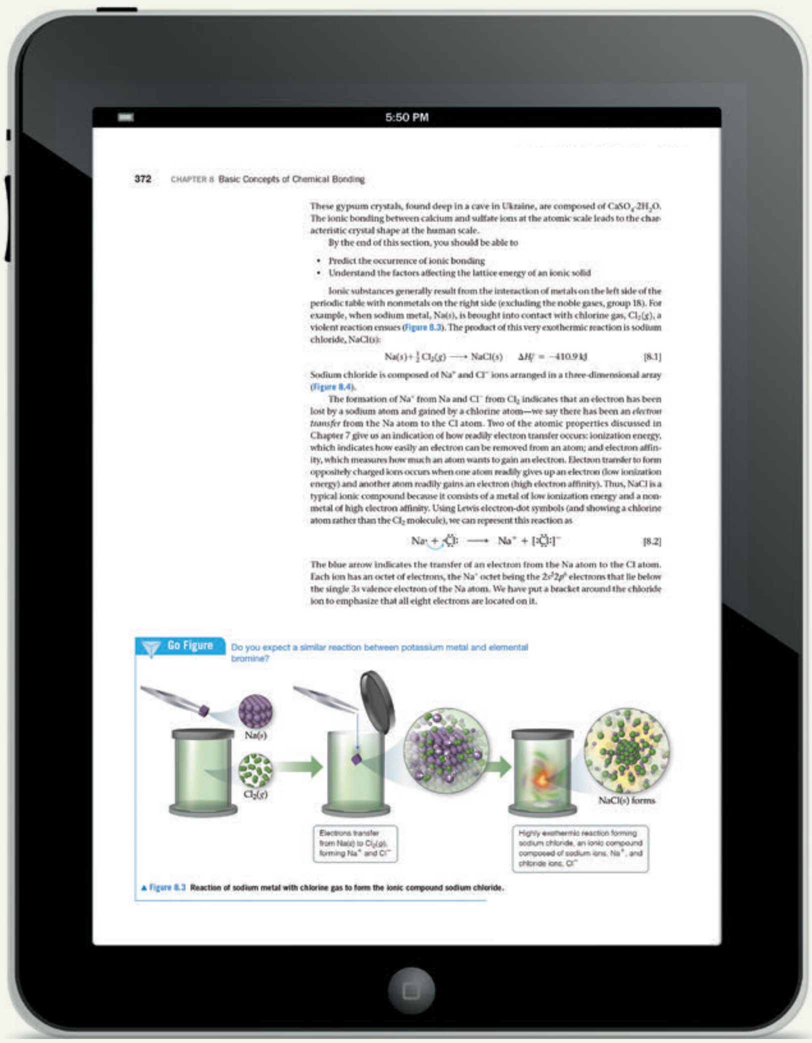

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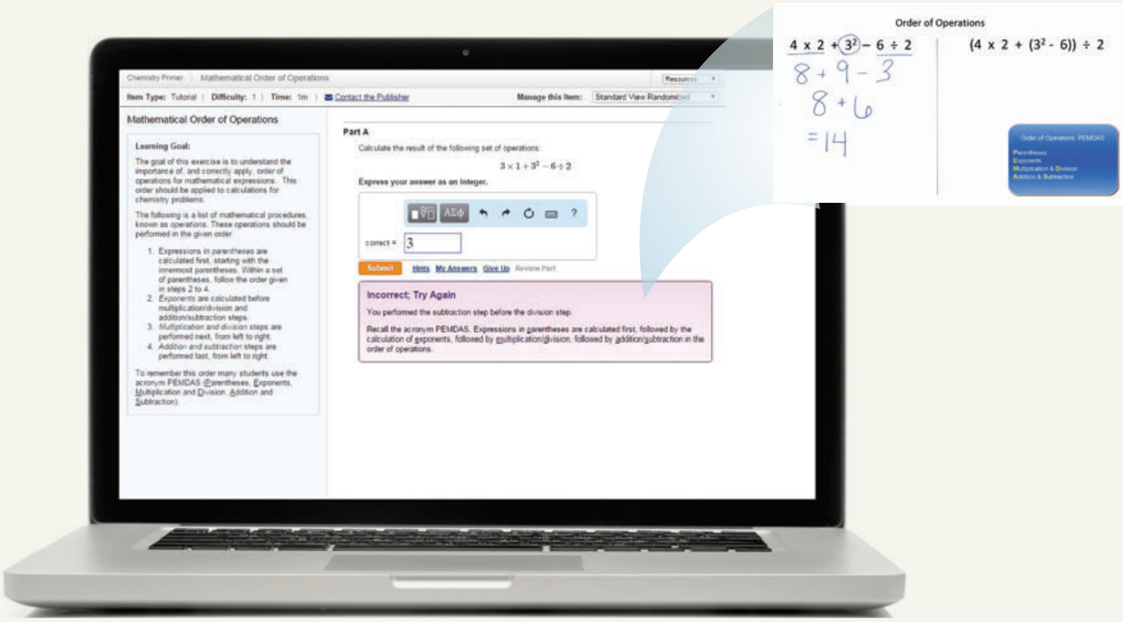

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- Upload a full PowerPoint® deck for easy creation of slide questions.
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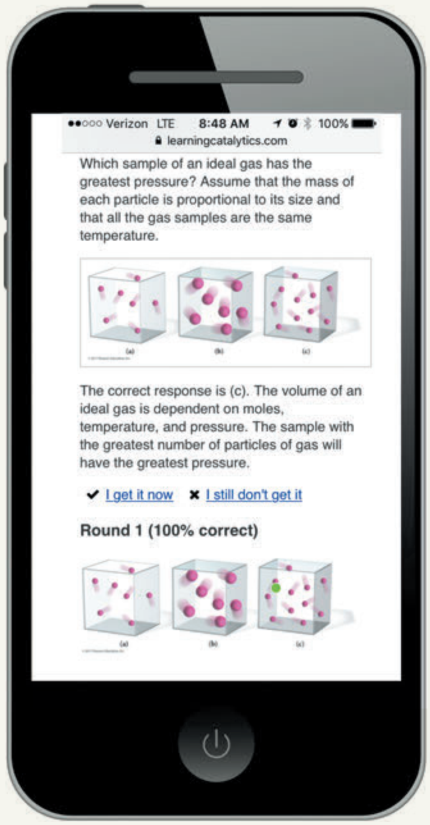

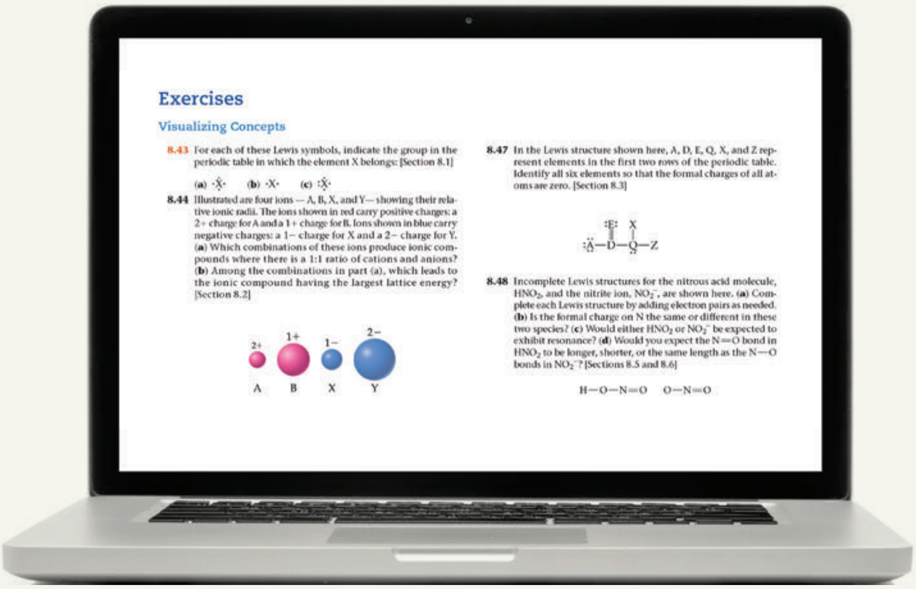

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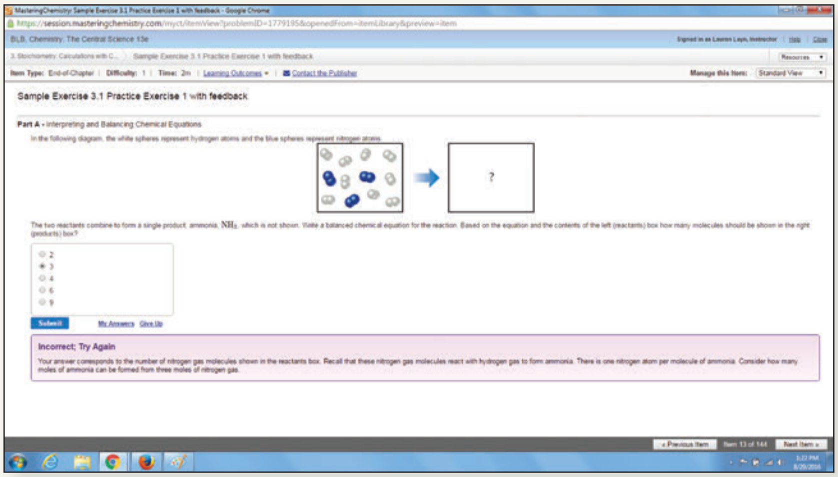
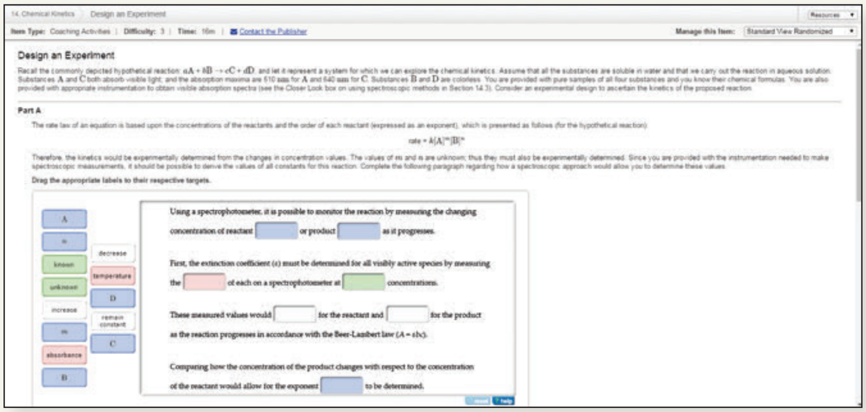
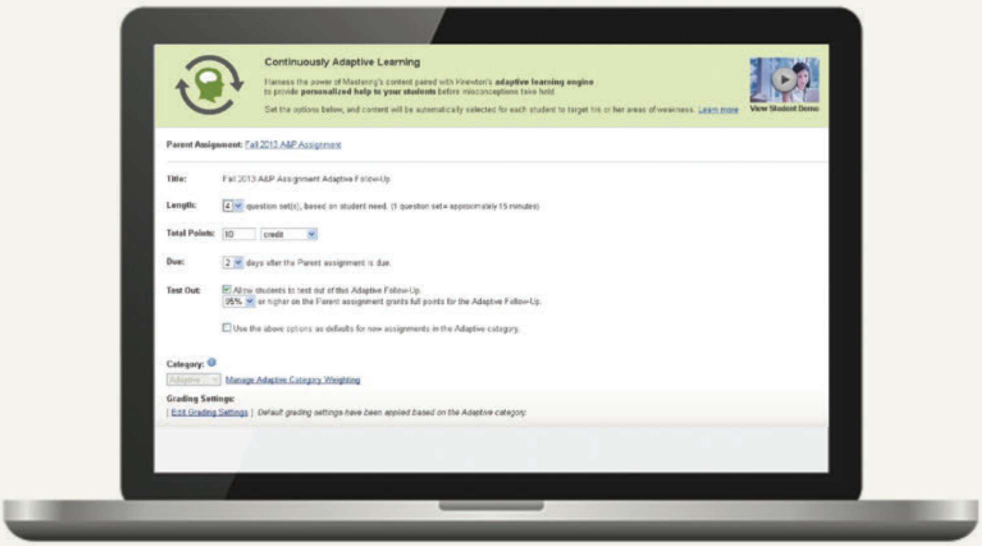


The **Design An Experiment** feature provides a departure from the usual kinds of end-of-chapter exercises with an inquiry-based, open-ended approach that tries to stimulate the student to "think like a scientist." Designed to foster critical thinking, each exercise presents the student with a scenario in which various unknowns require investigation. The student is called upon to ponder how experiments might be set up to provide answers to particular questions about observations.





Adaptive Follow-Up Assignments allow instructors to deliver content to students–automatically personalized for each individual based on the strengths and weaknesses identified by his or her performance on initial Mastering assignments.

Instructor and Student Resources

Resource	Available in Print	Available Online	Instructor or Student Resource	Description
TestGen Test Bank		√	Instructor	TestGen® is a computerized test generator that lets teachers view and edit Test Bank questions, transfer questions to tests, and print tests in a variety of customized formats. This Test Bank includes over 3000 multiple choice, true/false, and answer/essay questions. Questions are rated by difficulty and are correlated to the book's Learning Outcomes.
Instructor Manual		√	Instructor	Organized by chapter, this useful guide includes objectives, lecture outlines, references to figures and solved problems, as well as teaching tips.
Instructor Resource Materials		√	Instructor	The material available for download includes: • All illustrations, tables, and photos from the text in JPEG format • Pre-built PowerPoint™ Presentations (lecture, worked examples, images) • TestGen computerized software with the TestGen version of the Testbank • Word.doc files of the Test Item File
Solutions Manual (provided on request)		√	Instructor	Full solutions to all of the exercises in the text are provided.
Laboratory Experiments	√		Student	This manual contains 43 finely-tuned experiments chosen to introduce students to basic lab techniques and to illustrate core chemical principles.

CHEMISTRY

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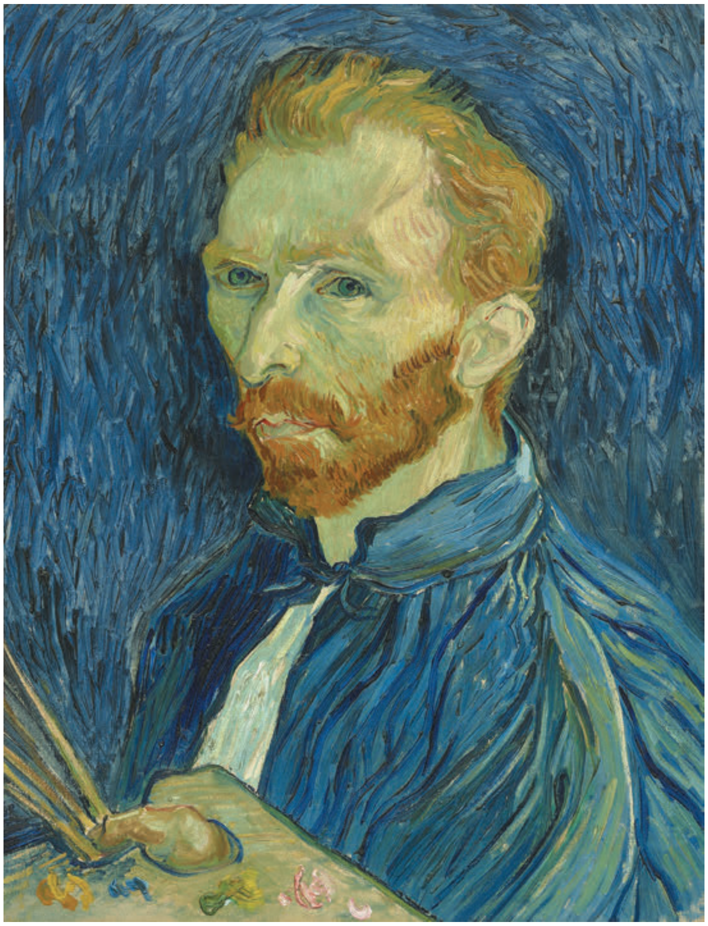
WHAT'S AHEAD

1

INTRODUCTION: MATTER, ENERGY, AND MEASUREMENT

1.1 | The Study of Chemistry



The title of this book—*Chemistry: The Central Science*—reflects the fact that much of what goes on in the world around us involves chemistry. Everyday chemical processes include the changes that produce the brilliant colors of flowers, the ways our bodies process the food we eat, and the electrical energy that powers our cell phones.

This first chapter provides an overview of what chemistry is about and what chemists do. The "What's Ahead" list gives an overview of the chapter organization and of some of the ideas we will consider.

Chemistry is the study of matter, its properties, and the changes that matter undergoes. As you progress in your study, you will come to see how chemical principles operate in all aspects of our lives, from everyday activities like food preparation to more complex processes such as those that operate in the environment. We will also learn how the properties of substances can be tailored for specific applications by controlling their composition and structure. For example, the synthetic pigments chemists developed in the nineteenth century were used extensively by impressionist artists such as van Gogh.

Chemistry is at the heart of many changes we see in the world around us, and it accounts for the myriad different properties we see in matter. To understand how these changes and properties arise, we need to look far beneath the surfaces of our everyday observations.

By the end of this section, you should be able to

- Appreciate the scope of chemistry

The Atomic and Molecular Perspective of Chemistry

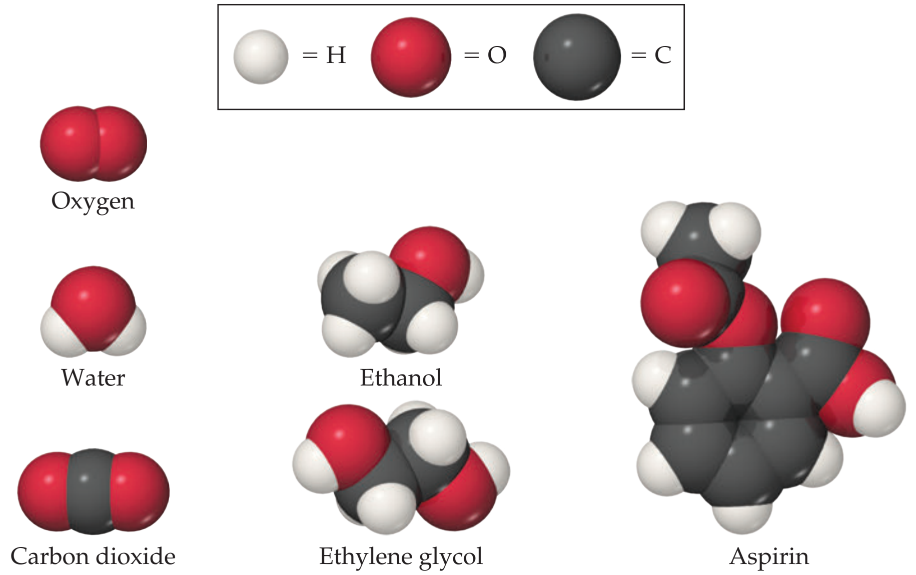
Chemistry is the study of the properties and behavior of matter. **Matter** is the physical material of the universe; it is anything that has mass and occupies space. A **property** is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types. This book, your body, the air you are breathing, and the clothes you are wearing are all samples of matter. We observe a tremendous variety of matter in our world, but countless experiments have shown that all matter is comprised of combinations of only about 100 substances called ***elements***. One of our major goals will be to relate the properties of matter to its composition, that is, to the particular elements it contains.

Chemistry also provides a background for understanding the properties of matter in terms of **atoms**, the almost infinitesimally small building blocks of matter. Each element is composed of a unique kind of atom. We will see that the properties of matter relate to both the kinds of atoms the matter contains (*composition*) and the arrangements of these atoms (*structure*).

In **molecules**, two or more atoms are joined in specific shapes. Throughout this text you will see molecules represented using colored spheres to show how the atoms are connected (**Figure 1.1**). The color provides a convenient way to distinguish between atoms of different elements. For example, notice that the molecules of ethanol and ethylene glycol in Figure 1.1 have different compositions and structures. Ethanol contains one oxygen atom, depicted by one red sphere. In contrast, ethylene glycol contains two oxygen atoms.

Even apparently minor differences in the composition or structure of molecules can cause profound differences in properties. For example, let's compare ethanol and ethylene glycol, which appear in Figure 1.1 to be quite similar. Ethanol is the alcohol in beverages such as beer and wine, whereas ethylene glycol is a viscous liquid used as automobile antifreeze. The properties of these two substances differ in many ways, as do their biological activities. Ethanol is consumed throughout the world, but you should *never* consume ethylene glycol because it is highly toxic. One of the challenges chemists undertake is to alter the composition or structure of molecules in a controlled way, creating new substances with different properties. For example, the common drug aspirin, shown in Figure 1.1, was first synthesized in 1897 in a successful attempt to

Go Figure Which molecule has the most carbon atoms? How many carbon atoms does it contain?



▲ **Figure 1.1 Molecular models.** The white, black, and red spheres represent atoms of hydrogen, carbon, and oxygen, respectively.

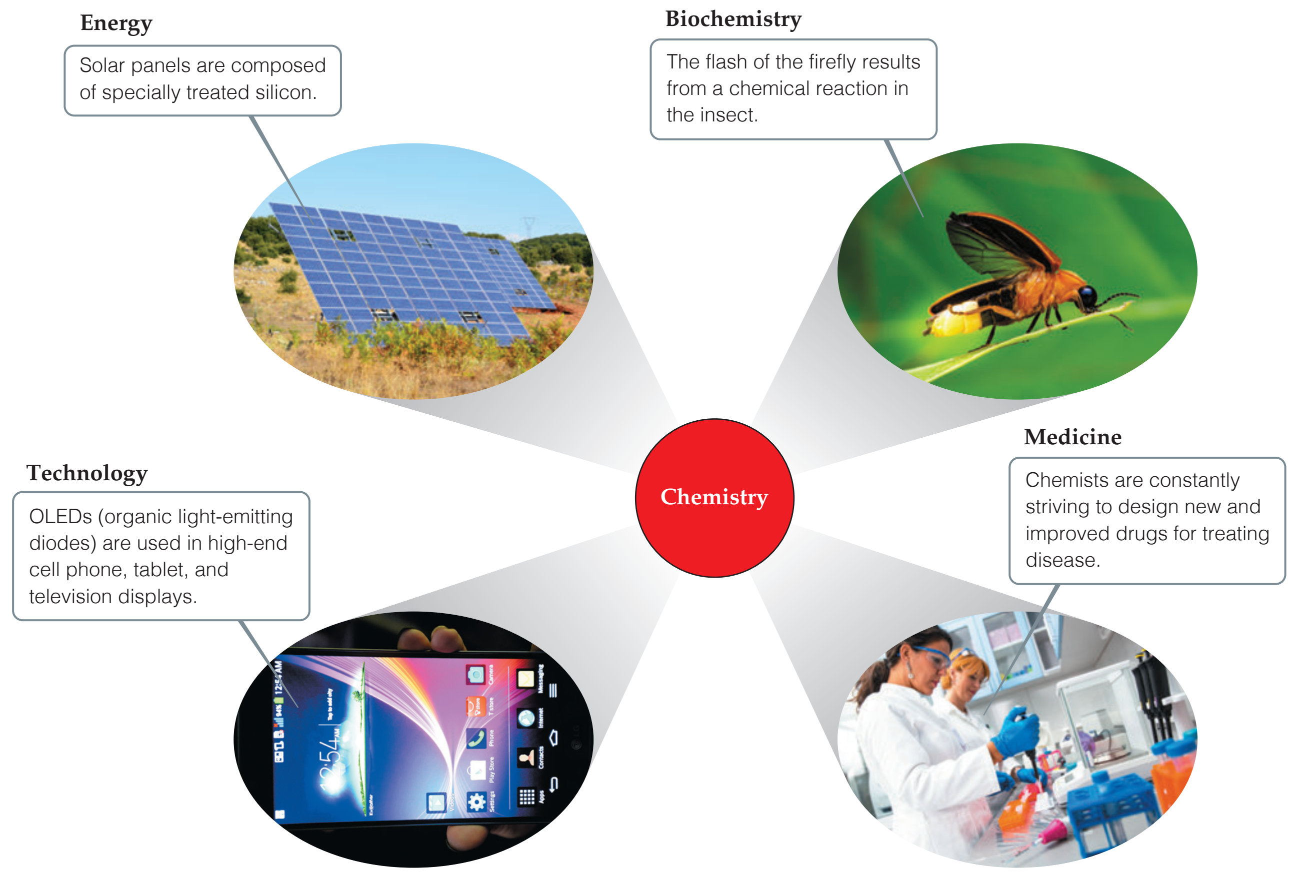
improve on a natural product extracted from willow bark that had long been used to alleviate pain.

Every change in the observable world—from boiling water to the changes that occur as our bodies combat invading viruses—has its basis in the world of atoms and molecules. Thus, as we proceed with our study of chemistry, we will find ourselves thinking in three realms: the *macroscopic* realm of ordinary-sized objects (*macro* = large), the *microscopic* realm of atoms and molecules, and the symbolic realm of how we represent these particles. We make our observations in the macroscopic world, but to understand that world, we must visualize how atoms and molecules behave at the microscopic level. Chemistry is the science that seeks to understand the properties and behavior of matter by studying the properties and behavior of atoms and molecules.

Why Study Chemistry?

Chemistry lies near the heart of many matters of public concern, such as improvement of health care, conservation of natural resources, protection of the environment, and the supply of energy needed to keep society running. Using chemistry, we have discovered and continually improved upon pharmaceuticals, fertilizers and pesticides, plastics, solar panels, light-emitting diodes (LEDs), and building materials. We have also discovered that some chemicals are harmful to our health or the environment. This means that we must be sure that the materials with which we come into contact are safe. As a citizen and consumer, it is in your best interest to understand the effects, both positive and negative, that chemicals can have, in order to arrive at a balanced outlook regarding their uses.

You may be studying chemistry because it is an essential part of your curriculum. Your major might be chemistry, or it could be biology, engineering, pharmacy, agriculture, geology, or some other field. Chemistry is central to a fundamental understanding of governing principles in many science-related fields. For example, our interactions with the material world raise basic questions about the materials around us. **Figure 1.2** illustrates how chemistry is central to several different realms of modern life.



▲ **Figure 1.2** **Chemistry is central to our understanding of the world around us.**

CHEMISTRY PUT TO WORK Chemistry and the Chemical Industry

Chemistry is all around us. For example, sugar is a chemical extracted from natural sources, refined by industry and sold across the world in a highly purified form. However, few realize the size and importance of the chemical industry. The chemical industry in the United States is estimated to be an $800 billion enterprise that employs over 800,000 people and accounts for 14% of all U.S. exports.

Who are chemists, and what do they do? People who have degrees in chemistry hold a variety of positions in industry, government, and academia. Those in industry work as laboratory chemists, developing new products (research and development); analyzing materials (quality control); or assisting customers in using products (sales and service). Those with more experience or training may work as managers or company directors. Chemists are important members of the scientific workforce in government (the National Institutes of Health, Department of Energy, and Environmental Protection Agency all employ chemists) and at universities. A chemistry degree is also good preparation for careers in teaching, medicine, biomedical research, information science, environmental work, technical sales, government regulatory agencies, and patent law.

Fundamentally, chemists do three things: They (1) make new types of matter: materials, substances, or combinations of substances with desired properties; (2) measure the properties of matter; and (3) develop models that explain and/or predict the properties of matter. One chemist, for example, may work in the laboratory to discover new drugs. Another may concentrate on the development of new instrumentation to measure properties of matter at the atomic level. Other chemists may use existing materials and methods to understand how pollutants are transported in the environment or how drugs are processed in the body. Yet another chemist will develop theory, write computer code, and run computer simulations to understand how molecules move and react. The collective chemical enterprise is a rich mix of all of these activities.

▲ **Figure 1.3** **Common chemicals used for household cleaning.**

Self-Assessment Exercise

1.1 Which of the following groups of substances involve the use of chemicals? Indicate all that apply.

(a) Paints, printer toner, food coloring

(b) Computer displays, LED lights, barcode readers

(c) Antiseptic cream, pain killers, energy drinks

(d) A light-weight bicycle frame, food packaging, a car exhaust catalytic converter

(e) Soap, shampoo, washing powder

Answers to Self-Assessment Exercises

1.1 All of them

1.2 Classifications of Matter

The study of any science starts with classification. This brings order to what we study much as the sorting of books in a library enables you to find the book you want.

Let's begin our study of chemistry by examining two fundamental ways in which matter is classified. Matter is typically characterized by (1) its physical state (gas, liquid, or solid) and (2) its composition (whether it is an element, a *compound*, or a *mixture*).

By the end of this section, you should be able to

- Understand the basic classification of matter

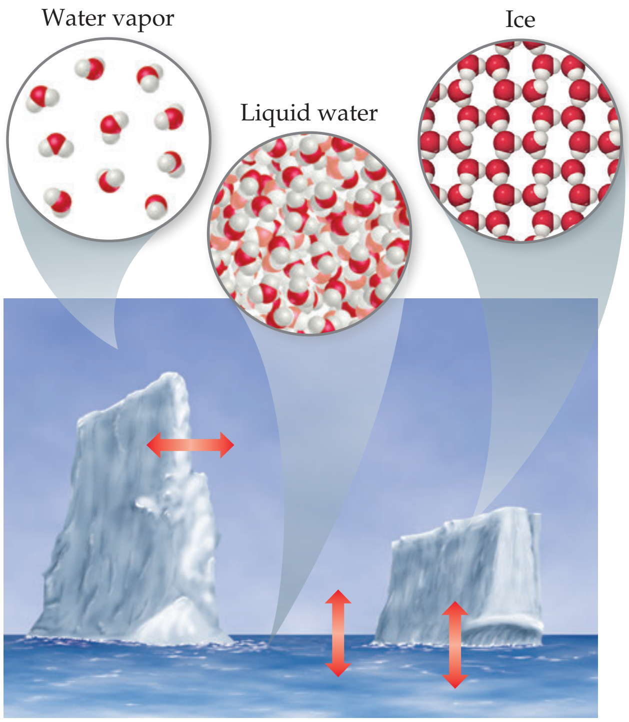
States of Matter

A sample of matter can be a gas, a liquid, or a solid. These three forms, called the **states of matter**, differ in some of their observable properties.

- A **gas** (also known as vapor) has no fixed volume or shape; rather, it uniformly fills its container. A gas can be compressed to occupy a smaller volume, or it can expand to occupy a larger one.
- A **liquid** has a distinct volume independent of its container, assumes the shape of the portion of the container it occupies, and is not compressible to any appreciable extent.
- A **solid** has both a definite shape and a definite volume and is not compressible to any appreciable extent.

The properties of the states of matter can be understood on the molecular level (**Figure 1.4**). In a gas the molecules are far apart and moving at high speeds, colliding repeatedly with one another and with the walls of the container. Compressing a gas decreases the amount of space between molecules and increases the frequency of collisions between molecules but does not alter the size or shape of the molecules. In a liquid, the molecules are packed closely together but still move rapidly. The rapid movement allows the molecules to slide over one another; thus, a liquid pours easily. In a solid the molecules are held tightly together, usually in definite arrangements in which the molecules can wiggle only slightly in their otherwise fixed positions. Thus, the distances between molecules are similar in the liquid and solid states, but while the molecules are for the most part locked in place in a solid, they retain considerable freedom of motion in a liquid. Changes in temperature and/or pressure can lead to conversion from one state of matter to another, illustrated by such familiar processes as ice melting or water vapor condensing.

Go Figure In which form of water are the water molecules farthest apart?



▲ **Figure 1.4 The three physical states of water—water vapor, liquid water, and ice.** We see the liquid and solid states but cannot see the gas (vapor) state. The red arrows show that the three states of matter interconvert.

Pure Substances

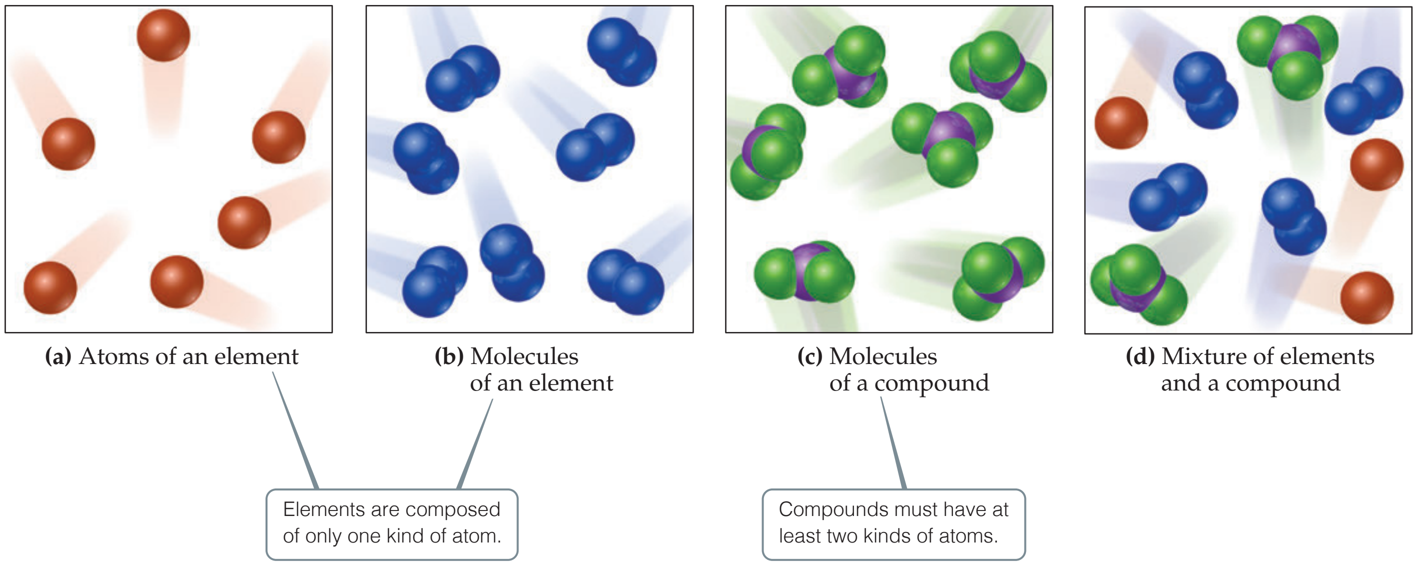
Most forms of matter we encounter—the air we breathe (a gas), the fuel we burn in our cars (a liquid), and the road they run on (a solid)—are not chemically pure. We can, however, separate these forms of matter into pure substances. A **pure substance** (usually referred to simply as a *substance*) is matter that has distinct properties and a composition that does not vary from sample to sample. Water and table salt (sodium chloride) are examples of pure substances.

All substances are either elements or compounds.

- **Elements** are substances that cannot be decomposed into simpler substances. On the molecular level, each element is composed of only one kind of atom [Figure 1.5(**a** and **b**)].
- **Compounds** are substances composed of two or more elements; they contain two or more kinds of atoms [Figure 1.5(**c**)]. Water, for example, is a compound composed of two elements: hydrogen and oxygen.

Figure 1.5(**d**) shows a mixture of substances. **Mixtures** are combinations of two or more substances in which each substance retains its chemical identity.

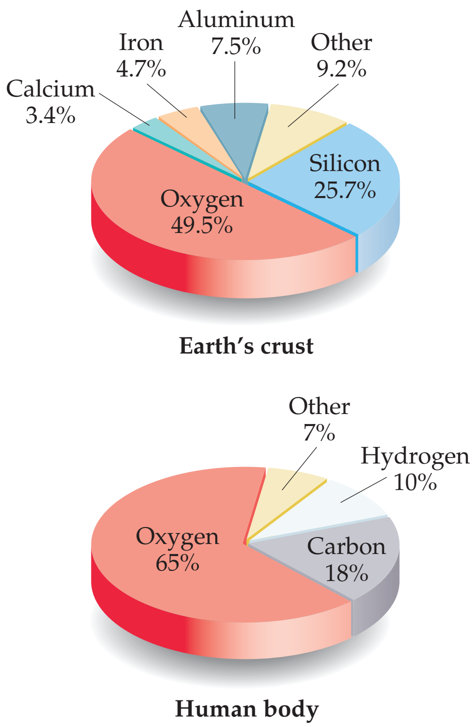
Go Figure How do the molecules of a compound differ from the molecules of an element?



▲ **Figure 1.5 Representation of elements, compounds, and mixtures.**

Go Figure

If the lower pie chart was drawn as the percentage in terms of number of atoms rather than the percentage in terms of mass, would the hydrogen slice of the pie get larger or smaller?



▲ **Figure 1.6 Relative abundances of elements.*** Elements in percent by mass in Earth's crust (including oceans and atmosphere) and the human body.

Elements

Currently, 118 elements are known, though they vary widely in abundance. Hydrogen constitutes about 74% of the mass in the Milky Way galaxy, and helium constitutes 24%. Closer to home, only five elements—oxygen, silicon, aluminum, iron, and calcium—account for over 90% of Earth's crust (including oceans and atmosphere), and only three—oxygen, carbon, and hydrogen—account for over 90% of the mass of the human body (**Figure 1.6**).

Table 1.1 lists some common elements, along with the chemical *symbols* used to denote them. The symbol for each element consists of one or two letters, with the first letter capitalized. These symbols are derived mostly from the English names of the elements, but sometimes they are derived from a foreign name instead (last column in Table 1.1). You will need to know these symbols and learn others as we encounter them in the text.

All of the known elements and their symbols are listed on the front inside cover of this text in a table known as the *periodic table*. In the periodic table, the elements are arranged in columns so that closely related elements are grouped together. We describe the periodic table in more detail in Section 2.5 and consider the periodically repeating properties of the elements in Chapter 7.

Compounds

Most elements can interact with other elements to form compounds. For example, when hydrogen gas burns in oxygen gas, the elements hydrogen and oxygen combine to form the compound water. Conversely, water can be decomposed into its elements by passing an electrical current through it (**Figure 1.7**).

Decomposing pure water into its constituent elements shows that it contains 11% hydrogen and 89% oxygen by mass, regardless of its source. This ratio is constant because every water molecule has the same number of hydrogen and oxygen atoms. While the mass percentages make it seem that water is mostly oxygen, there are actually two hydrogen atoms and only one oxygen atom per molecule. The explanation for this apparent discrepancy comes from the fact that hydrogen atoms are much lighter than oxygen atoms. The mass composition, a macroscopic determination, corresponds to

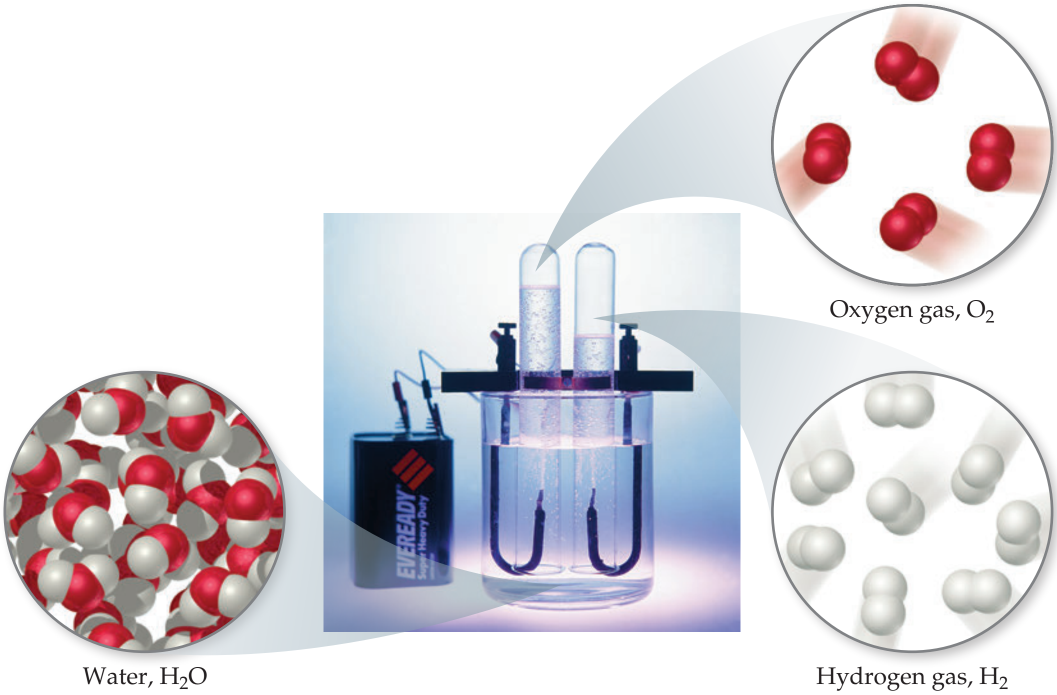
TABLE 1.1 Some Common Elements and Their Symbols

Carbon	C	Aluminum	Al	Copper	Cu (from *cuprum*)
Fluorine	F	Bromine	Br	Iron	Fe (from *ferrum*)
Hydrogen	H	Calcium	Ca	Lead	Pb (from *plumbum*)
Iodine	I	Chlorine	Cl	Mercury	Hg (from *hydrargyrum*)
Nitrogen	N	Helium	He	Potassium	K (from *kalium*)
Oxygen	O	Lithium	Li	Silver	Ag (from *argentum*)
Phosphorus	P	Magnesium	Mg	Sodium	Na (from *natrium*)
Sulfur	S	Silicon	Si	Tin	Sn (from *stannum*)

**U.S. Geological Survey Circular* 285, U.S Department of the Interior.

Go Figure

Is the volume of H_2 produced larger than the volume of O_2 produced because (a) hydrogen atoms are lighter than oxygen atoms, (**b**) hydrogen atoms are larger than oxygen atoms, or (**c**) each water molecule contains one oxygen atom and two hydrogen atoms?



▲ **Figure 1.7 Electrolysis of water.** Water decomposes into its component elements, hydrogen and oxygen, when an electrical current is passed through it. The volume of hydrogen, collected in the right test tube, is twice the volume of oxygen.

the molecular composition, which consists of two hydrogen atoms combined with one oxygen atom:

 Hydrogen atom (written H)

 Oxygen atom (written O)

 Water molecule (written H_2O)

The elements hydrogen and oxygen themselves exist naturally as diatomic (two-atom) molecules:

Oxygen molecule  (written O_2)

Hydrogen molecule  (written H_2)

As seen in **Table 1.2**, the properties of water bear no resemblance to the properties of its component elements. Hydrogen, oxygen, and water are each a unique substance, a consequence of the uniqueness of their respective molecules.

TABLE 1.2 Comparison of Water, Hydrogen, and Oxygen

	Water	Hydrogen	Oxygen
State[a]	Liquid	Gas	Gas
Normal boiling point	100 °C	−253 °C	−183 °C
Density[a]	1000 kg/m^3	0.084 kg/m^3	1.33 kg/m^3
Flammable	No	Yes	No

[a] At room temperature and atmospheric pressure.

The observation that the elemental composition of a compound is always the same is known as the **law of constant composition** (or the **law of definite proportions**). French chemist Joseph Louis Proust (1754–1826) first stated the law in about 1800. Although this law has been known for 200 years, the belief persists among some people that a fundamental difference exists between compounds prepared in the laboratory and the corresponding compounds found in nature. This simply is not true. Regardless of its source—nature or a laboratory—a pure compound has the same composition and properties under the same conditions. Both chemists and nature must use the same elements and operate under the same natural laws. When two materials differ in composition or properties, either they are composed of different compounds or they differ in purity.

Mixtures

Most of the matter we encounter consists of mixtures of different substances. Each substance in a mixture retains its chemical identity and properties. In contrast to a pure substance, which by definition has a fixed composition, the composition of a mixture can vary. A cup of sweetened coffee, for example, can contain either a little sugar or a lot. The substances making up a mixture are called *components* of the mixture.

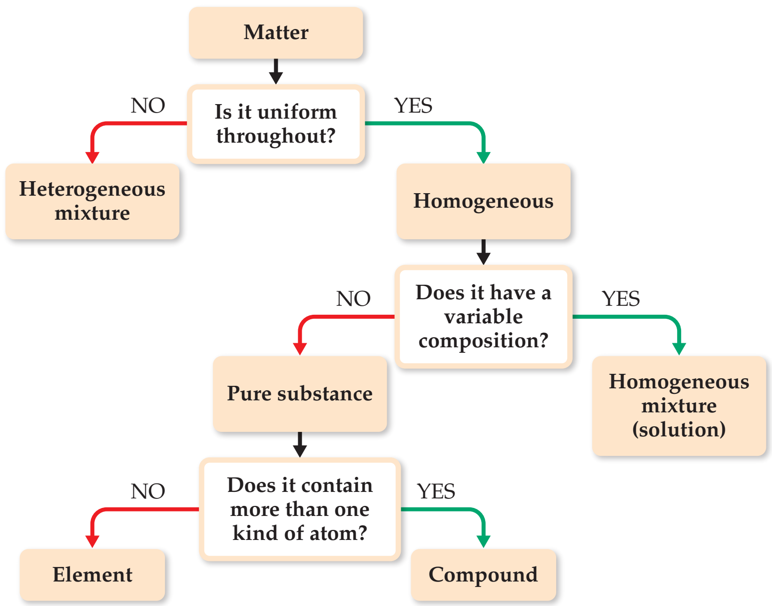
Some mixtures do not have the same composition, properties, and appearance throughout. Rocks and wood, for example, vary in texture and appearance in any typical sample. Such mixtures are *heterogeneous* [Figure 1.8(**a**)]. Mixtures that are uniform throughout are *homogeneous*. Air is a homogeneous mixture of nitrogen, oxygen, and smaller amounts of other gases. The nitrogen in air has all the properties of pure nitrogen because both the pure substance and the mixture contain the same nitrogen molecules. Salt, sugar, and many other substances dissolve in water to form **solutions**, which are homogeneous mixtures [Figure 1.8(**b**)]. Although the term *solution* conjures an image of a liquid, the term can also apply to solids, liquids, or gases.

Figure 1.9 summarizes the classification of matter into elements, compounds, and mixtures.

▶ **Figure 1.8 Mixtures.** (**a**) Many common materials, including rocks, are heterogeneous mixtures. This photograph of granite shows a heterogeneous mixture of silicon dioxide and other metal oxides. (**b**) Homogeneous mixtures are called solutions. Many substances, including the blue solid shown here [copper(II) sulfate pentahydrate], dissolve in water to form solutions.

(a)

(b)



◀ **Figure 1.9 Classification of matter.** All pure matter is classified ultimately as either an element or a compound.

Sample Exercise 1.1

Distinguishing among Elements, Compounds, and Mixtures

"White gold" contains gold and a "white" metal, such as palladium. Two samples of white gold differ in the relative amounts of gold and palladium they contain. Both samples are uniform in composition throughout. Use Figure 1.9 to classify white gold.

SOLUTION

Because the material is uniform throughout, it is homogeneous. Because its composition differs for the two samples, it cannot be a compound. Instead, it must be a homogeneous mixture.

▶ Practice Exercise

Aspirin is composed of 60.0% carbon, 4.5% hydrogen, and 35.5% oxygen by mass, regardless of its source. Use Figure 1.9 to classify aspirin.

Self-Assessment Exercise

1.2 Carbonated water, 'Soda Water', is water in which carbon dioxide has been dissolved under pressure. Classify soda water.

(a) A homogeneous solution of elements

(b) A homogeneous solution of compounds

(c) A heterogeneous mixture of elements

(d) A heterogeneous mixture of compounds

Exercises

1.3 Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: (**a**) air, (**b**) chocolate with almond, (**c**) aluminium, (**d**) iodine tincture.

1.4 Give the chemical symbol or name for the following elements, as appropriate: **(a)** helium, **(b)** platinum, **(c)** cobalt, **(d)** tin, **(e)** silver, **(f)** Sb, **(g)** Pb, **(h)** Br, **(i)** V, **(j)** Hg.

1.5 A solid white substance A is heated strongly in the absence of air. It decomposes to form a new white substance B and a gas C. The gas has exactly the same properties as the product obtained when carbon is burned in an excess of oxygen. Based on these observations, can we determine whether solids A and B and gas C are elements or compounds?

Answers to Self-Assessment Exercises

1.2 (b)

1.3 Properties of Matter

We are familiar with many of the properties of materials around us—the hardness of steel, the flexibility of rubber, or the glow of a wood fire. If we delve a little deeper, we can classify these properties as one of two types. By the end of this section, you should be able to

- Distinguish between physical and chemical changes

Every substance has unique properties. For example, the properties listed in Table 1.2 allow us to distinguish hydrogen, oxygen, and water from one another. The properties of matter can be categorized as physical or chemical. **Physical properties** can be observed without changing the identity and composition of the substance. These properties include color, odor, density, melting point, boiling point, and hardness. **Chemical properties** describe the way a substance may change, or *react*, to form other substances. A common chemical property is flammability, the ability of a substance to burn in the presence of oxygen.

Some properties, such as temperature and melting point, are *intensive properties*. **Intensive properties** do not depend on the amount of sample being examined and are particularly useful in chemistry because many intensive properties can be used to *identify* substances. **Extensive properties** depend on the amount of sample, with two examples being mass and volume. Extensive properties relate to the *amount* of substance present.

Physical and Chemical Changes

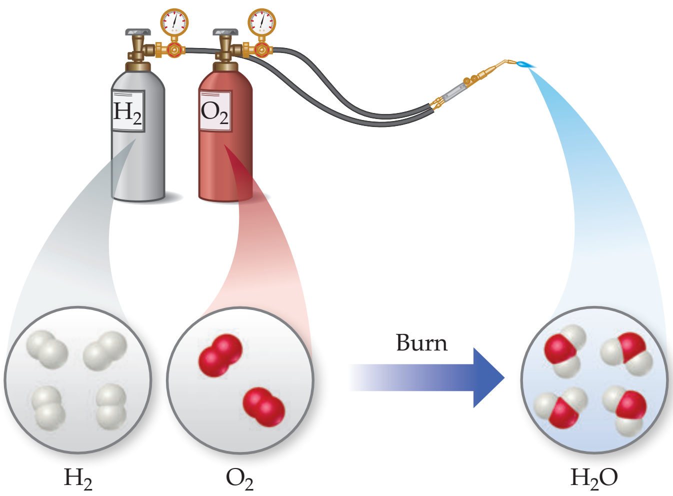
The changes substances undergo are either physical or chemical. During a **physical change**, a substance changes its physical appearance but not its composition. (That is, it is the same substance before and after the change.) The evaporation of water is a physical change. When water evaporates, it changes from the liquid state to the gas state, but it is still composed of water molecules, as depicted in Figure 1.4. All **changes of state** (for example, from liquid to gas or from liquid to solid) are physical changes.

In a **chemical change** (also called a **chemical reaction**), a substance is transformed into a chemically different substance. When hydrogen burns in air, for example, it undergoes a chemical change because it combines with oxygen to form water (**Figure 1.10**).

Chemical changes can be dramatic. In the account given in **Figure 1.11**, Ira Remsen, author of a popular chemistry text published in 1901, describes his first experiences with chemical reactions.

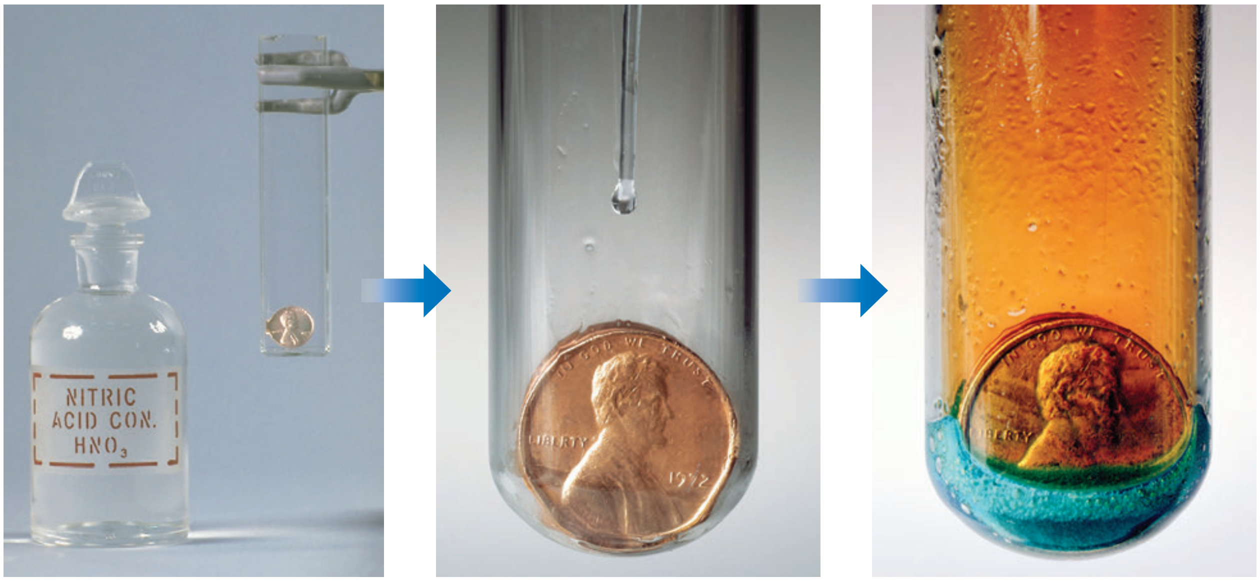
Separation of Mixtures

We can separate a mixture into its components by taking advantage of differences in their properties. For example, a heterogeneous mixture of iron filings and gold filings could be sorted by color into iron and gold. A less tedious approach would be to use a magnet to



▲ **Figure 1.10 A chemical reaction.**

While reading a textbook of chemistry, I came upon the statement "nitric acid acts upon copper," and I determined to see what this meant. Having located some nitric acid, I had only to learn what the words "act upon" meant. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table, opened a bottle labeled "nitric acid," poured some of the liquid on the copper, and prepared to make an observation. But what was this wonderful thing which I beheld? The cent was already changed, and it was no small change either. A greenish-blue liquid foamed and fumed over the cent and over the table. The air became colored dark red. How could I stop this? I tried by picking the cent up and throwing it out the window. I learned another fact: nitric acid acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and discovered nitric acid acts upon trousers. That was the most impressive experiment I have ever performed. I tell of it even now with interest. It was a revelation to me. Plainly the only way to learn about such remarkable kinds of action is to see the results, to experiment, to work in the laboratory.*



▲ **Figure 1.11 The chemical reaction between a copper penny and nitric acid.** The dissolved copper produces the blue-green solution; the reddish brown gas produced is nitrogen dioxide.

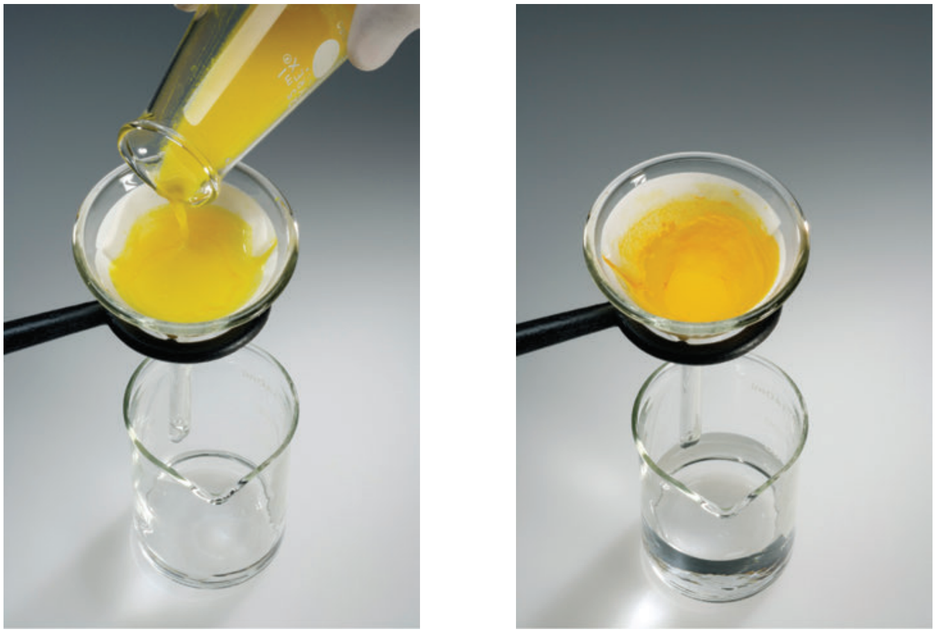
attract the iron filings, leaving the gold ones behind. We can also take advantage of an important chemical difference between these two metals: Many acids dissolve iron but not gold. Thus, if we put our mixture into an appropriate acid, the acid would dissolve the iron and the solid gold would be left behind. The two could then be separated by

*Remsen, *The Principles of Theoretical Chemistry,* 1887.

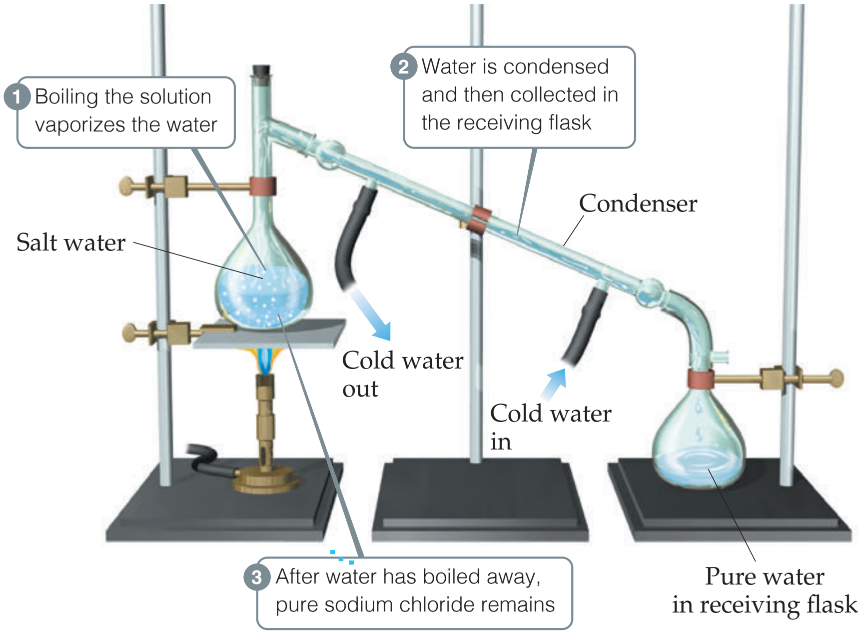
filtration (Figure 1.12). We would have to use other chemical reactions, which we will learn about later, to transform the dissolved iron back into metal.

An important method of separating the components of a homogeneous mixture is **distillation**, a process that depends on the different abilities of substances to form gases. For example, if we boil a solution of salt and water, the water evaporates, forming a gas, and the salt is left behind. The gaseous water can be converted back to a liquid on the walls of a condenser, as shown in Figure 1.13.

The differing abilities of substances to adhere to the surfaces of solids can also be used to separate mixtures. This ability is the basis of *chromatography*, a technique shown in Figure 1.14.

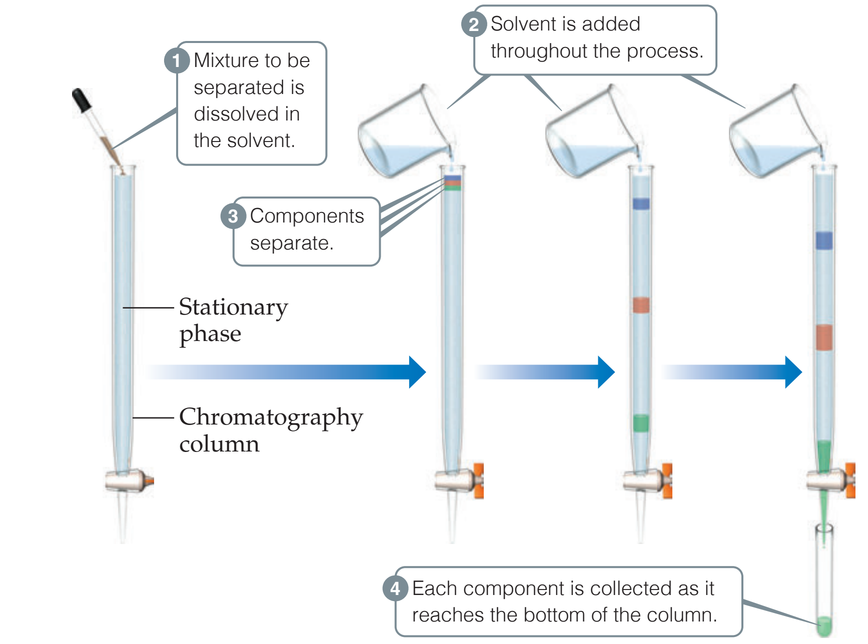


▶ **Figure 1.12 Separation by filtration.** A mixture of a solid and a liquid is poured through filter paper. The liquid passes through the paper while the solid remains on the paper.



▲ **Figure 1.13 Distillation.** Apparatus for separating a sodium chloride solution (salt water) into its components.

Go Figure Is the separation of the mixture shown here a physical or chemical process?



▲ **Figure 1.14 Separation of three substances using column chromatography.**

Self-Assessment Exercise

1.6 When water boils in a pot, bubbles of water vapor can be seen rising through the liquid. What type of change does this represent?

(a) a physical change
(b) a chemical change

Exercises

1.7 Zirconia, an oxide of zirconium, is often used as an affordable diamond substitute. Just like diamond, it is a colorless crystal which sparkles under sunlight. Which of the following physical properties do you think would help in differentiating between diamond and Zirconia—melting point, density, or physical state?

1.8 In the process of attempting to characterize a substance, a chemist makes the following observations: The substance is a silvery white, lustrous metal. It melts at 649 °C and boils at 1105 °C. Its density at 20 °C is 1.738 g/cm^3. The substance burns in air, producing an intense white light. It reacts with chlorine to give a brittle white solid. The substance can be pounded into thin sheets or drawn into wires. It is a good conductor of electricity. Which of these characteristics are physical properties, and which are chemical properties?

1.9 Label each of the following as either a physical process or a chemical process: **(a)** crushing a metal can, **(b)** production of urine in the kidneys, **(c)** melting a piece of chocolate, **(d)** burning fossil fuel, **(e)** discharging a battery.

1.10 Which separation method is better suited for obtaining sugar from cane juice—filtration or evaporation?

Answers to Self-Assessment Exercises

1.6 (a)

1.4 | The Nature of Energy

Since the earliest time, humans have recognized fire as a form of energy. In chemistry, it is convenient to distinguish between different types of energy, and, by the end of this section, you should be able to

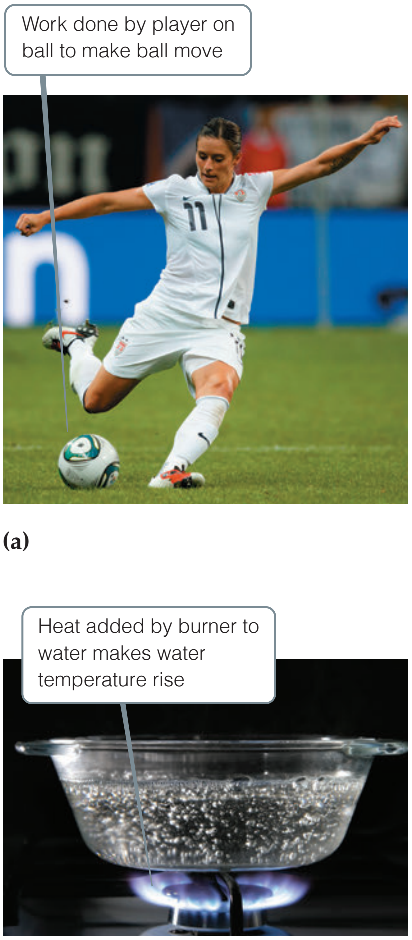
- Recognize the difference between kinetic and potential energy

All objects in the universe are made of matter, but matter alone is not enough to describe the behavior of the world around us. The water in an alpine lake and a pot of boiling water are both made from the same substance, but your body will experience a very different sensation if you put your hand in each. The difference between the two is their energy content; boiling water has more energy than chilled water. To understand chemistry, we must also understand energy and the changes in energy that accompany chemical processes.

Unlike matter, energy does not have mass and cannot be held in our hands, but its effects can be observed and measured. **Energy** *is defined as the capacity to do work or transfer heat*. **Work** is *the energy transferred when a force exerted on an object causes a displacement of that object*, and **heat** is *the energy used to cause the temperature of an object to increase* (Figure 1.15). Although the temperature of an object is intuitive to most people, the definition of work is less apparent. We define work, w, as the product of the force exerted on the object, F, and the distance, d, that it moves:

$$w = F \times d \qquad [1.1]$$

where **force** is defined as any push or pull exerted on the object.* Familiar examples include gravity and the attraction between opposite poles of a bar magnet. It takes work to lift an object off of the floor, or to pull apart two magnets that have come together at the opposite poles.



(a)

(b)

▲ **Figure 1.15 Work and heat, two forms of energy.** (**a**) *Work* is energy used to cause an object to move against an opposing force. (**b**) *Heat* is energy used to increase the temperature of an object.

Kinetic Energy and Potential Energy

To understand energy, we need to grasp its two fundamental forms, kinetic energy and potential energy. Objects, whether they are automobiles, soccer balls, or molecules, can possess **kinetic energy**, the energy of *motion*. The magnitude of kinetic energy, E_k, of an object depends on its mass, m, and velocity, v:

$$E_k = \frac{1}{2}mv^2 \qquad [1.2]$$

*In using this equation, only the component of the force that is acting parallel to the distance traveled is used. That will generally be the case for problems we will encounter in this chapter.

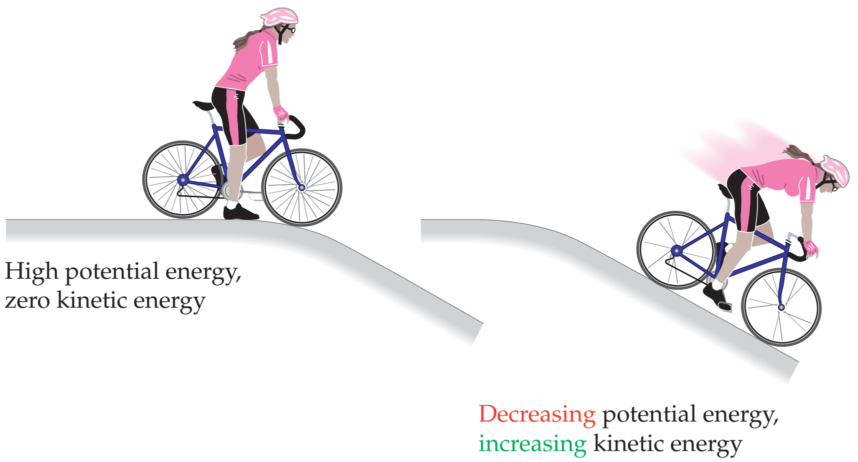
Thus, the kinetic energy of an object increases as its velocity or speed** increases. For example, a car has greater kinetic energy moving at 65 kilometers per hour (km/h) than it does at 25 km/h. For a given velocity, the kinetic energy increases with increasing mass. Thus, a large truck traveling at 65 km/h has greater kinetic energy than a motorcycle traveling at the same velocity because the truck has the greater mass.

In chemistry, we are interested in the kinetic energy of atoms and molecules. Although these particles are too small to be seen, they have mass and are in motion, and therefore, possess kinetic energy. When a substance is heated, be it a pot of water on the stove or an aluminum can sitting in the sun, the atoms and molecules in that substance gain kinetic energy and their average speed increases. Hence, we see that the transfer of heat is simply the transfer of kinetic energy at the molecular level.

All other forms of energy—the energy stored in a stretched spring, in a weight held above your head, or in a chemical bond—are classified as potential energy. An object has **potential energy** by virtue of its position relative to other objects. Potential energy is, in essence, the "stored" energy that arises from the attractions and repulsions an object experiences in relation to other objects.

We are familiar with many instances in which potential energy is converted into kinetic energy. For example, think of a cyclist poised at the top of a hill (**Figure 1.16**). Because of the attractive force of gravity, the potential energy of the bicycle is greater at the top of the hill than at the bottom. As a result, the bicycle easily rolls down the hill with increasing speed. As it does so, potential energy is converted into kinetic energy. The potential energy decreases as the bicycle rolls down the hill, while at the same time its kinetic energy increases as it picks up speed (Equation 1.2). This example illustrates that kinetic and potential energy are interconvertible.

Gravitational forces play a negligible role in the ways that atoms and molecules interact with one another. Forces that arise from electrical charges are more important when dealing with atoms and molecules. One of the most important forms of potential energy in chemistry is *electrostatic potential energy*, which arises from the interactions between charged particles. You are probably familiar with the fact that opposite charges attract each other and like charges repel. The strength of this interaction increases as the magnitude of the charges increase, and decreases as the



▲ **Figure 1.16 Potential energy and kinetic energy.** The potential energy initially stored in the motionless bicycle and rider at the top of the hill is converted to kinetic energy as the bicycle moves down the hill and loses potential energy.

**Strictly speaking, velocity is a vector quantity that has a direction; that is, it tells you how fast an object is moving and in what direction. Speed is a scalar quantity that tells you how fast an object is moving but not the direction of the motion. Unless otherwise stated, we will not be concerned with the direction of motion, and thus velocity and speed are interchangeable quantities for the treatment in this book.

distance between charges increases. We will return to electrostatic energy several times throughout the book.

One of our goals in chemistry is to relate the energy changes seen in the macroscopic world to the kinetic or potential energy of substances at the molecular level. Many substances, fuels for example, release energy when they react. The *chemical energy* of a fuel is due to the potential energy stored in the arrangements of its atoms. As we will learn in later chapters, *chemical energy is released when bonds between atoms are formed, and consumed when bonds between atoms are broken*. When a fuel burns, some bonds are broken and others are formed, but the net effect is to convert chemical potential energy to thermal energy, the energy associated with temperature. The increase in thermal energy arises from increased molecular motion and hence increased kinetic energy at the molecular level.

Self-Assessment Exercise

1.11 What happens to the chemical energy of a nickel-metal hydride battery as it is recharged after powering an electric razor?

(a) It increases
(b) It decreases
(c) It remains the same

Exercises

1.12 Two positively charged particles are first brought close together and then released. Once released, the repulsion between particles causes them to move away from each other. (**a**) This is an example of potential energy being converted into what form of energy? (**b**) Does the potential energy of the two particles prior to release increase or decrease as the distance between them is increased.

1.13 For each of the following processes, does the potential energy of the object(s) increase or decrease? (**a**) The charge of two oppositely charged particles is increased. (**b**) H_2O molecule is split into two oppositely charged ions, H^+ and OH^-. (**c**) A person skydives from a height of 600 meters.

Answers to Self-Assessment Exercises

1.11 (a)

1.5 Units of Measurement

Many properties of matter are *quantitative*, that is, associated with numbers. When a number represents a measured quantity, the units of that quantity must be specified. The volume of a soft drink expressed in different units demonstrates how important this is. To say that the length of a pencil is 17.5 is meaningless. Expressing the number with its units, 17.5 centimeters (cm), properly specifies the length. The units used for scientific measurements are those of the **metric system**.

The metric system, developed in France during the late eighteenth century, is used as the system of measurement in most countries. By the end of this section, you should be able to

- Convert between different multiples of SI units

SI Units

In 1960, an international agreement was reached specifying a particular choice of metric units for use in scientific measurements. These preferred units are called **SI units**, after the French *Système International d'Unités*. This system has seven *base units* from which all other units are derived (Table 1.3). In this section, we will consider the base units for length, mass, and temperature.

With SI units, prefixes are used to indicate decimal fractions or multiples of various units. For example, the prefix *milli-* represents a 10^{-3} fraction, one-thousandth, of a unit: A milligram (mg) is 10^{-3} gram (g), a millimeter (mm) is 10^{-3} meter (m), and so forth.

TABLE 1.3 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Length	Meter	m
Mass	Kilogram	kg
Temperature	Kelvin	K
Time	Second	s or sec
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

A CLOSER LOOK The Scientific Method

Where does scientific knowledge come from? How is it acquired? How do we know it is reliable? How do scientists add to it, or modify it?

There is nothing mysterious about how scientists work. The first idea to keep in mind is that scientific knowledge is gained through observations of the natural world. A principal aim of the scientist is to organize these observations by identifying patterns and regularity, making measurements, and associating one set of observations with another. The next step is to ask *why* nature behaves in the manner we observe. To answer this question, the scientist constructs a model, known as a **hypothesis**, to explain the observations. Initially, the hypothesis is likely to be pretty tentative. There could be more than one reasonable hypothesis. If a hypothesis is correct, then certain results and observations should follow from it. In this way, hypotheses can stimulate the design of experiments to learn more about the system being studied. Scientific creativity comes into play in thinking of hypotheses that are fruitful in suggesting good experiments to do, ones that will shed new light on the nature of the system.

As more information is gathered, the initial hypotheses get winnowed down. Eventually, just one may stand out as most consistent with a body of accumulated evidence. We then begin to call this hypothesis a **theory**, a model that has predictive powers and that accounts for all the available observations. A theory also generally is consistent with other, perhaps larger and more general theories. For example, a theory of what goes on inside a volcano has to be consistent with more general theories regarding heat transfer, chemistry at high temperature, and so forth.

We will be encountering many theories as we proceed through this book. Some of them have been found over and over again to be consistent with observations. However, no theory can be proven to be absolutely true. We can treat it as though it is, but there always remains a possibility that there is some respect in which a theory is wrong. A famous example is Isaac Newton's theory of mechanics, which yielded such precise results for the mechanical behavior of matter that no exceptions to it were found before the twentieth century. But Albert Einstein showed that Newton's theory of the nature of space and time is incorrect.

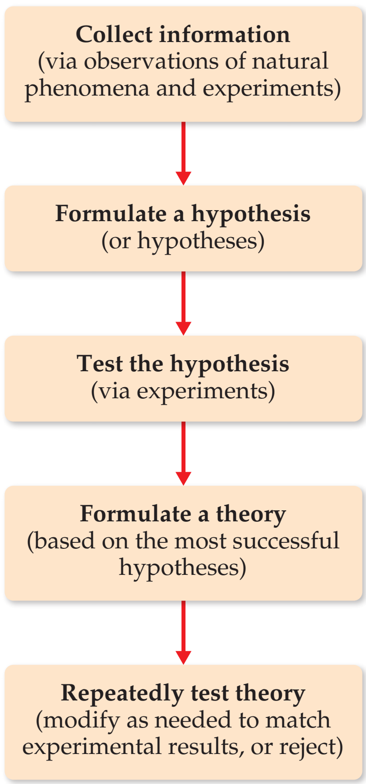
Continued

Einstein's theory of relativity represented a fundamental shift in how we think of space and time. He predicted where the exceptions to predictions based on Newton's theory might be found. Although only small departures from Newton's theory were predicted, they *were* observed. Einstein's theory of relativity became accepted as the correct model. However, for most uses, Newton's laws of motion are quite accurate enough.

The overall process we have just considered, illustrated in Figure 1.17, is often referred to as *the scientific method*. But there is no single scientific method. Many factors play a role in advancing scientific knowledge. The one unvarying requirement is that our explanations be consistent with observations and that they depend solely on natural phenomena.

When nature behaves in a certain way over and over again, under all sorts of different conditions, we can summarize that behavior in a **scientific law**. For example, it has been repeatedly observed that in a chemical reaction there is no change in the total mass of the materials reacting as compared with the materials that are formed; we call this observation the *law of conservation of mass*. It is important to make a distinction between a theory and a scientific law. On the one hand, a scientific law is a statement of what always happens, to the best of our knowledge. A theory, on the other hand, is an *explanation* for what happens. If we discover some law fails to hold true, then we must assume the theory underlying that law is wrong in some way.

Related Exercises: 1.74, 1.96



▲ Figure 1.17 The scientific method.

Table 1.4 presents the prefixes commonly encountered in chemistry. In using SI units and in working problems throughout this text, you must be comfortable using exponential notation. If you are unfamiliar with exponential notation or want to review it, refer to Appendix A.1.

TABLE 1.4 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\ J = 1\ kg \cdot m^2/s^2$ and $1\ W = 1\ J/s$.
[b]Greek letter mu, pronounced "mew."

Although non–SI units are being phased out, some are still commonly used by scientists. Whenever we first encounter a non–SI unit in the text, the SI unit will also be given. The relations between the non–SI and SI units we will use most frequently in this text appear on the back inside cover. We will discuss how to convert from one to the other in Section 1.7.

Length and Mass

The SI base unit of *length* is the meter, a distance slightly longer than a yard. **Mass*** is a measure of the amount of material in an object. The SI base unit of mass is the kilogram (kg). This base unit is unusual because it uses a prefix, *kilo-*, instead of the word *gram* alone. We obtain other units for mass by adding prefixes to the word *gram*.

Sample Exercise 1.2
Using SI Prefixes

What is the name of the unit that equals (**a**) 10^{-9} gram, (**b**) 10^{-6} second, (**c**) 10^{-3} meter?

SOLUTION

We can find the prefix related to each power of ten in Table 1.4: (**a**) nanogram, ng; (**b**) microsecond, μs; (**c**) millimeter, mm.

Practice Exercise

(**a**) How many picometers are there in 1 m? (**b**) Express 6.0×10^3 m using a prefix to replace the power of ten. (**c**) Use exponential notation to express 4.22 mg in grams. (**d**) Use decimal notation to express 4.22 mg in grams.

Temperature

Temperature, a measure of the hotness or coldness of an object, is a physical property that determines the direction of heat flow. Heat always flows spontaneously from a substance at higher temperature to one at lower temperature. Thus, the influx of heat we feel when we touch a hot object tells us that the object is at a higher temperature than our hand.

The temperature scales commonly employed in science are the Celsius and Kelvin scales. The **Celsius scale** was originally based on the assignment of 0 °C to the freezing point of water and 100 °C to its boiling point at sea level (**Figure 1.18**).

The **Kelvin scale** is the SI temperature scale, and the SI unit of temperature is the *kelvin* (K). Zero on the Kelvin scale is the temperature at which all thermal motion ceases, a temperature referred to as **absolute zero**. On the Celsius scale, absolute zero has the value −273.15 °C. The Celsius and Kelvin scales have equal-sized units—that is, a kelvin is the same size as a degree Celsius. Thus, the Kelvin and Celsius scales are related according to

$$K = {}^\circ C + 273.15 \qquad [1.3]$$

The freezing point of water, 0 °C, is 273.15 K (Figure 1.18). Notice that we do not use a degree sign (°) with temperatures on the Kelvin scale.

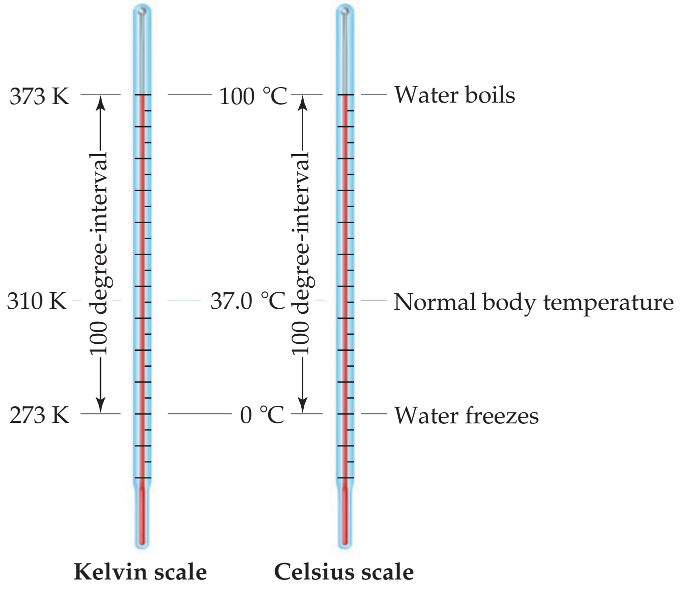
Derived SI Units

The SI base units are used to formulate *derived units*. A **derived unit** is obtained by multiplication or division of one or more of the base units. We begin with the defining

*Mass and weight are not the same. Mass is a measure of the amount of matter; weight is the force exerted on this mass by gravity. For example, an astronaut weighs less on the Moon than on Earth because the Moon's gravitational force is less than Earth's. The astronaut's mass on the Moon, however, is the same as it is on Earth.

Go Figure

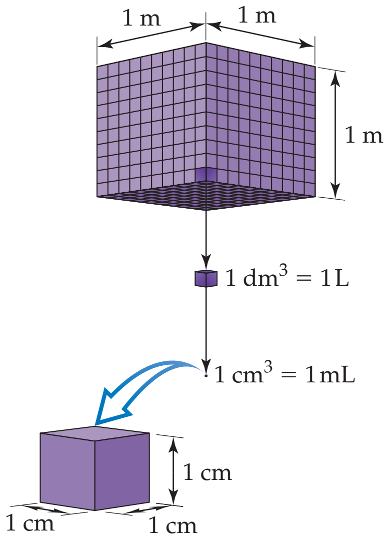
True or false: The "size" of a degree on the Celsius scale is the same as the "size" of a degree on the Kelvin scale.



▲ **Figure 1.18 Comparison of the Kelvin and Celsius temperature scales.**

Go Figure

How many 1-L bottles are required to contain 1 m^3 of liquid?



▲ **Figure 1.19 Volume relationships.** The volume occupied by a cube 1 m on each edge is one cubic meter, 1 m^3. Each cubic meter contains 1000 dm^3, 1 m^3 = 1000 dm^3. One liter is the same volume as one cubic decimeter, 1 L = 1 dm^3. Each cubic decimeter contains 1000 cubic centimeters, 1 dm^3 = 1000 cm^3. One cubic centimeter equals one milliliter, 1 cm^3 = 1 mL.

equation for a quantity and, then substitute the appropriate base units. For example, *speed* is defined as the ratio of distance traveled to elapsed time. Thus, the derived SI unit for speed is the SI unit for distance (length), m, divided by the SI unit for time, s, which gives m/s, read "meters per second." Two common derived units in chemistry are those for volume and density.

Volume

The *volume* of a cube is its length cubed, length^3. Thus, the derived SI unit of volume is the SI unit of length, m, raised to the third power. The cubic meter, m^3, is the volume of a cube that is 1 m on each edge (**Figure 1.19**). Smaller units, such as cubic centimeters, cm^3 (sometimes written cc), are frequently used in chemistry. Another volume unit used in chemistry is the *liter* (L), which equals a cubic decimeter, dm^3. (The liter is the first metric unit we have encountered that is *not* an SI unit.) There are 1000 milliliters (mL) in a liter, and 1 mL is the same volume as 1 cm^3: 1 mL = 1 cm^3.

Sample Exercise 1.3

Converting Units of Temperature

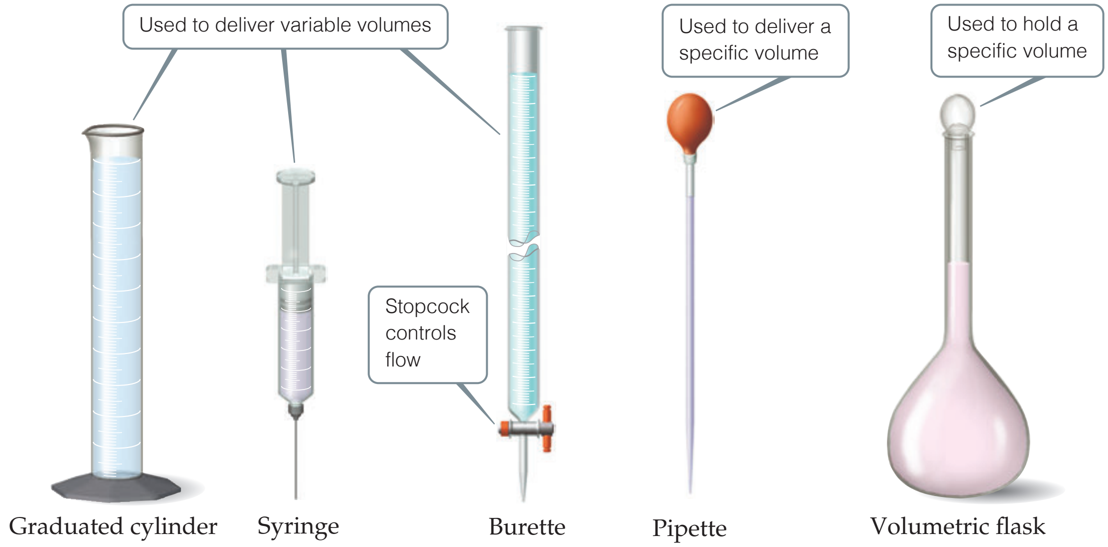
A weather forecaster predicts the temperature will reach 31 °C. What is this temperature in K?

SOLUTION

Equation 1.3, we have K = 31 + 273 = 304 K.

▶ **Practice Exercise**

Using Wolfram Alpha (http://www.wolframalpha.com/) or some other reference, determine which of these elements would be liquid at 525 K (assume samples are protected from air): (**a**) bismuth, Bi; (**b**) platinum, Pt; (**c**) selenium, Se; (**d**) calcium, Ca; (**e**) copper, Cu.



▲ **Figure 1.20** **Common volumetric glassware.**

In the lab, you will likely use the devices in **Figure 1.20** to measure and deliver volumes of liquids. Syringes, burettes, and pipettes deliver amounts of liquids with more precision than graduated cylinders. Volumetric flasks are used to contain specific volumes of liquid.

Density

Density is defined as the amount of mass in a unit volume of a substance:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad [1.4]$$

The densities of solids and liquids are commonly expressed in either grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL). The densities of some common substances are listed in **Table 1.5**. It is no coincidence that the density of water is 1.00 g/mL; the gram was originally defined as the mass of 1 mL of water at a specific temperature. Because most substances change volume when they are heated or cooled, densities are temperature dependent, and so temperature should be specified when reporting densities. If no temperature is reported, we assume 25 °C, close to normal room temperature.

The terms *density* and *weight* are sometimes confused. A person who says that iron weighs more than air generally means that iron has a higher density than air—1 kg of air has the same mass as 1 kg of iron, but the iron occupies a smaller volume, thereby giving it a higher density. If we combine two liquids that do not mix, the less dense liquid will float on the denser liquid.

TABLE 1.5 Densities of Selected Substances at 25 °C

Substance	Density (g/cm^3)
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

Units of Energy

The SI unit for energy is the **joule** (pronounced "jool"), J, in honor of James Joule (1818–1889), a British scientist who investigated work and heat. If we return to Equation 1.2 where kinetic energy was defined, we immediately see that joules are a derived unit, $1\ J = 1\ kg\text{-}m^2/s^2$. Numerically, a 2-kg mass moving at a velocity of 1 m/s possesses a kinetic energy of 1 J:

$$E_k = \frac{1}{2}mv^2 = \frac{1}{2}(2\ kg)(1\ m/s)^2 = 1\ kg\text{-}m^2/s^2 = 1\ J$$

Sample Exercise 1.4

Determining Density and Using Density to Determine Volume or Mass

(**a**) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.

(**b**) Calculate the volume of 65.0 g of liquid methanol if its density is 0.791 g/mL.

(**c**) What is the mass in grams of a cube of gold (density = 19.32 g/cm^3) if the length of the cube is 2.00 cm?

SOLUTION

(**a**) We are given mass and volume, so Equation 1.3 yields

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.00 \times 10^2\ \text{g}}{7.36\ \text{cm}^3} = 13.6\ \text{g/cm}^3$$

(**b**) Solving Equation 1.3 for volume and then using the given mass and density gives

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{65.0\ \cancel{\text{g}}}{0.791\ \cancel{\text{g}}/\text{mL}} = 82.2\ \text{mL}$$

(**c**) We can calculate the mass from the volume of the cube and its density. The volume of a cube is given by its length cubed:

$$\text{Volume} = (2.00\ \text{cm})^3 = (2.00)^3\ \text{cm}^3 = 8.00\ \text{cm}^3$$

Solving Equation 1.3 for mass and substituting the volume and density of the cube, we have

$$\text{Mass} = \text{volume} \times \text{density} = (8.00\ \cancel{\text{cm}^3})(19.32\ \text{g}/\cancel{\text{cm}^3}) = 155\ \text{g}$$

Practice Exercise

(**a**) Calculate the density of a 374.5-g sample of copper if it has a volume of 41.8 cm^3. (**b**) A student needs 15.0 g of ethanol for an experiment. If the density of ethanol is 0.789 g/mL, how many milliliters of ethanol are needed? (**c**) What is the mass, in grams, of 25.0 mL of mercury (density = 13.6 g/mL)?

Because a joule is not a very large amount of energy, we often use *kilojoules* (kJ) in discussing the energies associated with chemical reactions. For example, the amount of heat released when hydrogen and oxygen react to form 1 g of water is 16 kJ.

It is still quite common in chemistry, biology, and biochemistry to find energy changes associated with chemical reactions expressed in the non-SI unit of calories. A **calorie** (cal) was originally defined as the amount of energy required to raise the temperature of 1 g of water from 14.5 to 15.5 °C. It has since been defined in terms of a joule:

$$1\ \text{cal} = 4.184\ \text{J (exactly)}$$

A related energy unit that is familiar to anyone who has read a food label is the nutritional *Calorie* (note the capital C), which is 1000 times larger than calorie with a lowercase c: 1 Cal = 1000 cal = 1 kcal.

Sample Exercise 1.5

Identifying and Calculating Energy Changes

A standard propane (C_3H_8) tank used in an outdoor grill holds approximately 9.0 kg of propane. When the grill is operating, propane reacts with oxygen to form carbon dioxide and water. For every gram of propane that reacts with oxygen in this way, 46 kJ of energy is released as heat. (**a**) How much energy is released if the entire contents of the propane tank react with oxygen? (**b**) As the propane reacts, does the potential energy stored in chemical bonds increase or decrease? (**c**) If you were to store an equivalent amount of potential energy by pumping water to an elevation of 75 m above the ground, what mass of water would be needed? (Note: The force due to gravity acting on the water, which is the water's weight, is $F = m \times g$, where m Is the mass of the object and g is the gravitational constant, $g = 9.8\ \text{m/s}^2$.)

SOLUTION

(**a**) We can calculate the amount of energy released from the propane as heat by converting the mass of propane from kg to g and then using the fact that 46 kJ of heat are released per gram:

$$E = 9.0\ \cancel{\text{kg}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{46\ \text{kJ}}{1\ \cancel{\text{g}}} = 4.1 \times 10^5\ \text{kJ} = 4.1 \times 10^8\ \text{J}$$

(**b**) When propane reacts with oxygen, the potential energy stored in the chemical bonds is converted to an alternate form of energy, heat. Therefore, the potential energy stored as chemical energy must decrease.

(**c**) The amount of work done to pump the water to a height of 75 m can be calculated using Equation 1.1:

$$w = F \times d = (m \times g) \times d$$

rearranging to solve for the mass of water:

$$m = \frac{w}{g \times d} = \frac{4.1 \times 10^8\ \text{J}}{(9.8\ \text{m/s}^2)(75\ \text{m})} = \frac{4.1 \times 10^8\ \text{kg-}\cancel{\text{m}^2}/\cancel{\text{s}^2}}{(9.8\ \cancel{\text{m}}/\cancel{\text{s}^2})(75\ \cancel{\text{m}})}$$

$$= 5.6 \times 10^5\ \text{kg}$$

At 25 °C, this mass of water would have a volume of 560,000 L. Thus, we see that large amounts of potential energy can be stored as chemical energy.

Practice Exercise

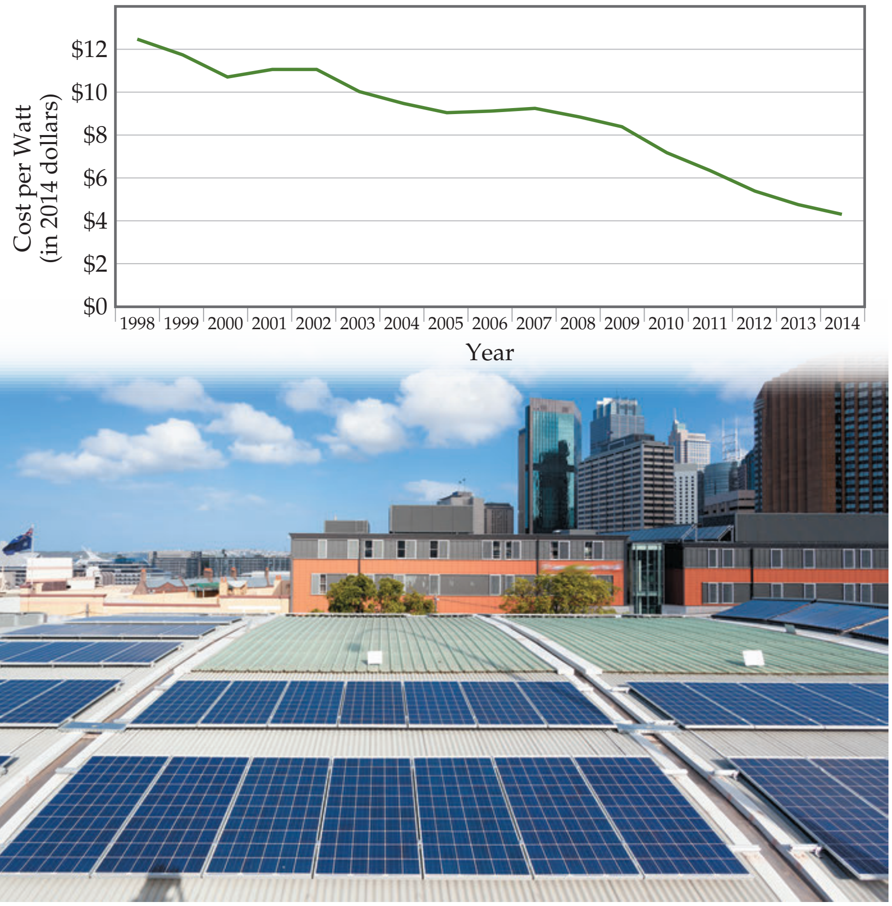
Which of the following objects has the greatest kinetic energy? (**a**) a 500-kg motorcycle moving at 100 km/h (**b**) a 1,000-kg car moving at 50 km/h (**c**) a 1500-kg car moving at 30 km/h (**d**) a 5000 kg truck moving at 10 km/h (**e**) a 10,000-kg truck moving at 5 km/h

CHEMISTRY PUT TO WORK Chemistry in the News

Solar Energy Steps Up. Given the challenges society faces in trying to mitigate the effects of climate change, the need for affordable clean energy has never been greater. Given the massive amount of energy our planet receives from the Sun, solar energy has long been touted as a technology of the future. The problem for decades has been the relatively high cost of solar power compared to energy generated from burning fossil fuels. However, in recent years the cost of solar energy has decreased more rapidly than most people thought possible, decreasing more than 50% in just the past five years (Figure 1.21).

Not surprisingly, over the last six years the number of solar panel installations worldwide has increased sixfold. Over the same period, the number of coal plants operating in the United States has decreased by 38%, from 523 to 323. China is on track to install plants capable of generating more than 18 gigawatts of solar energy in 2015, nearly equal to the entire solar energy capacity of the United States. Furthermore, with the price of solar energy dropping rapidly, there is reason to believe that developing nations may forgo building fossil fuel power plants and skip straight to green energy technologies like solar and wind.

Recently, chemists have discovered a new class of materials called halide perovskites that have the potential to bring the cost of solar energy down even further. Solar cells made with halide perovskites have been shown to be nearly as efficient as single crystal silicon solar cells, but can be prepared from inexpensive solution methods that differ from the more costly and energy-intensive methods used to produce silicon solar cells. There are still many challenges to be overcome before halide perovskite solar cells are produced commercially, but their future looks very promising.



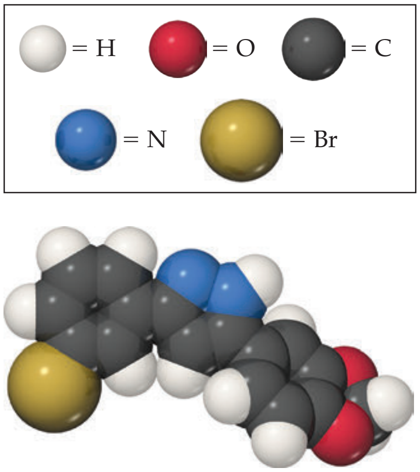
▲ **Figure 1.21 The cost of energy generated by photovoltaic modules.** The median price per watt of electricity generated has dropped by two-thirds since 1998. The cost figures here include the cost for fully installed residential solar panels.

Slowing a Progressive Disease. Proteins are very large molecules that play an essential role in biology and the functioning of living organisms. Out of the myriad different types of proteins, scientists have identified one class of proteins called *prions* that play an important role in certain neurological diseases. We all have prion proteins in our brains, and for nearly all of us, they cause no harmful effects. For a small fraction of people, though, something causes the prions to change form and to adopt an incorrectly folded molecular shape. This process, once begun, is cumulative, propagating throughout the brain; the misfolded proteins somehow trigger the same misfolding in other prion proteins. Eventually, the misfolded proteins aggregate into clusters that can destroy neurons, producing symptoms such as those seen in Alzheimer's and Parkinson's diseases.

Currently, no therapies have been developed to stop the progression of prion diseases. However, there are encouraging signs that certain small molecules could disrupt propagation of the disease. They might be able to do this by blocking the interaction between one prion protein and another that causes the misfolding to propagate. So far, the experimental work has involved studies of mice infected with prions. One such molecule, called Anle-138b, is shown in Figure 1.22. Administration of this compound more than doubled the life spans of treated mice.

Compounds such as this one are not suitable for use in treating prion disease in humans, but studies conducted thus far point the way toward discovery of molecules that might in the future provide

Continued



▲ **Figure 1.22** **Anle-138b, a promising molecule for the treatment of degenerative brain diseases.**

effective therapies. This work illustrates the complex and sometimes tortuous path that takes the scientist from promising beginnings to a much desired goal.

Chemical Approaches to Art Restoration. Works of art are made from materials that over time can undergo chemical reactions that alter their appearance or undermine their mechanical stability. Layers of dirt and pollutants can build up on the surface, discoloring statues, murals, and paintings. Plasters that serve as the base for frescoes react with gases in the atmosphere. One approach to art restoration is to use selected chemical reactions to undo or reverse the effects of detrimental reactions that have occurred over time.

One such example involves murals at the ancient Mayan ruins of Mayapan in Mexico's Yucatán peninsula. In the 1960s, polymer coatings were applied to these murals in an attempt to preserve them, but within a decade it became clear that the polymer coatings were doing more harm than good. Unfortunately, years of oxidation and cross-linking reactions had rendered the polymer coatings insoluble to nearly every type of organic solvent. In 2008, a team of Italian chemists found a way to remove the polymer coatings without damaging the underlying fresco. They treated the murals with a microemulsion, a special type of mixture where molecules called surfactants encapsulate and carry nanometer-sized droplets of organic solvents through an aqueous solution to the surface of the fresco. Once the droplets reach the surface, they can dissolve and carry away the unwanted polymer coating.

Self-Assessment Exercise

1.14 A standard Olympic-size swimming pool contains 2.5 ML of water. What volume is this in m^3?

(a) 2.5 m^3

(b) 250 m^3

(c) 2 500 m^3

(d) 2 500 000 m^3

Exercises

1.15 Convert the following expressions into exponential notation: (**a**) 3 terameters (Tm) (**b**) 2.5 femtoseconds (fs) (**c**) 57 micrometers (μm) (**d**) 8.3 megagrams (Mg).

1.16 Make the following conversions: **(a)** 25 °C to K (**b**) 110 °C to K (**c**) 300 K to °C (**d**) 1100 K to °C

1.17 **(a)** A sample of tetrachloroethene, a liquid used in dry cleaning that is being phased out because of its potential to cause cancer, has a mass of 40.55 g and a volume of 25.0 mL at 25 °C. What is its density at this temperature? Will tetrachloroethene float on water? (Materials that are less dense than water will float.) **(b)** Carbon dioxide (CO_2) is a gas at room temperature and pressure. However, carbon dioxide can be put under pressure to become a "supercritical fluid" that is a much safer dry-cleaning agent than tetrachloroethylene. At a certain pressure, the density of supercritical CO_2 is 0.469 g/cm^3. What is the mass of a 25.0-mL sample of supercritical CO_2 at this pressure?

1.18 **(a)** To identify a liquid substance, a student determined its density. Using a graduated cylinder, she measured out a 45-mL sample of the substance. She then measured the mass of the sample, finding that it weighed 38.5 g. She knew that the sub-stance had to be either isopropyl alcohol (density 0.785 g/mL) or toluene (density 0.866 g/mL). What are the calculated density and the probable identity of the substance? **(b)** An experiment requires 45.0 g of ethylene glycol, a liquid whose density is 1.114 g/mL. Rather than weigh the sample on a balance, a chemist chooses to dispense the liquid using a graduated cylinder. What volume of the liquid should he use? **(c)** Is a graduated cylinder such as that shown in Figure 1.20 likely to afford the accuracy of measurement needed? **(d)** A cubic piece of metal measures 5.00 cm on each edge. If the metal is nickel, whose density is 8.90 g/cm^3, what is the mass of the cube?

1.19 If on a certain year, an estimated amount of 4 million metric tons (1 metric ton = 1000 kg) of nitrous oxide (N_2O) was emitted worldwide due to agricultural activities, express this mass of N_2O in grams without exponential notation, using an appropriate metric prefix.

Answers to Self-Assessment Exercises

1.14 (c)

1.6 | Uncertainty in Measurement

Weighing a single postage stamp on a pair of old-fashioned scales is unlikely to give you a meaningful result. As scientists, when we make measurements, we must always consider the appropriateness of the numbers we report. There are specific ways in which to do this that can give us confidence in interpreting data collected by ourselves and others. By the end of this section, you should be able to

- Determine the correct number of significant figures to report at the conclusion of an experiment

Two kinds of numbers are encountered in scientific work: *exact numbers* (those whose values are known exactly) and *inexact numbers* (those whose values have some uncertainty). Most of the exact numbers we will encounter in this book have defined values. For example, there are exactly 12 eggs in a dozen and exactly 1000 g in a kilogram. The number 1 in any conversion factor, such as 1 m = 100 cm or 1 kg = 1000 g, is an exact number. Exact numbers can also result from counting objects. For example, we can count the exact number of marbles in a jar or the exact number of people in a classroom.

Numbers obtained by measurement are always *inexact*. The equipment used to measure quantities always has inherent limitations (equipment errors), and there are differences in how different people make the same measurement (human errors). Suppose ten students with ten balances are to determine the mass of the same coin. The ten measurements will probably vary slightly for various reasons. The balances might be calibrated slightly differently, and there might be differences in how each student reads the mass from the balance. Remember: *Uncertainties always exist in measured quantities.*

Precision and Accuracy

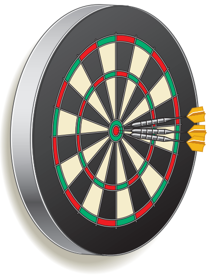
The terms *precision* and *accuracy* are often used in discussing the uncertainties of measured values. **Precision** is a measure of how closely individual measurements agree with one another. **Accuracy** refers to how closely individual measurements agree with the correct, or "true," value. The dart analogy in Figure 1.23 illustrates the difference between these two concepts.

In the laboratory, we often perform several "trials" of an experiment and average the results. The precision of the measurements is often expressed in terms of the *standard deviation* (Appendix A.5), which reflects how much the individual measurements differ

High precision can be achieved on a scale like this one, which has 0.1 mg accuracy.

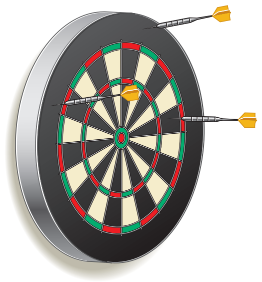
Go Figure

How would the darts be positioned on the target for the case of "good accuracy, poor precision"?



Good accuracy
Good precision

Poor accuracy
Good precision



Poor accuracy
Poor precision

▲ **Figure 1.23** Precision and accuracy.

from the average. We gain confidence in our measurements if we obtain nearly the same value each time—that is, when the standard deviation is small. Figure 1.23 reminds us, however, that precise measurements can be inaccurate. For example, if a very sensitive balance is poorly calibrated, the masses we measure will be consistently either high or low. They will be inaccurate even if they are precise.

Significant Figures

Suppose you determine the mass of a coin on a balance capable of measuring to the nearest 0.0001 g. You could report the mass as 2.2405 ± 0.0001 g. The ± notation (read "plus or minus") expresses the magnitude of the uncertainty of your measurement. In much scientific work, we drop the ± notation with the understanding that *there is always some uncertainty in the last digit reported for any measured quantity.*

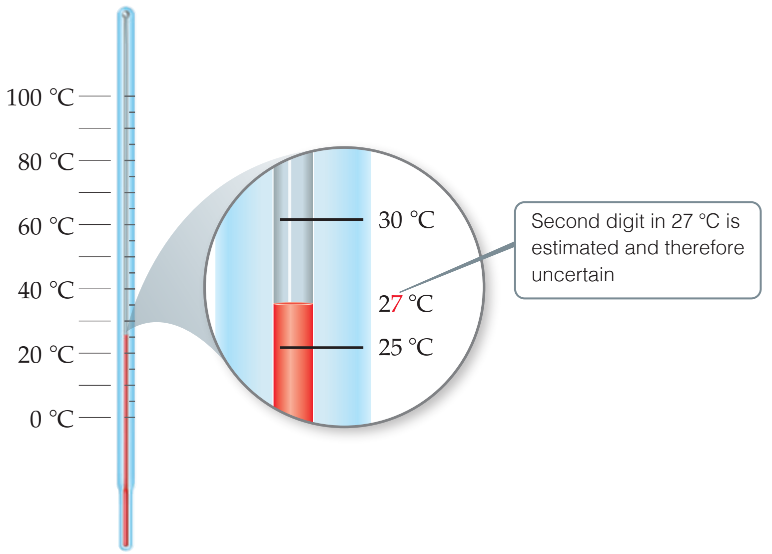
Figure 1.24 shows a thermometer with its liquid column between two scale marks. We can read the certain digits from the scale and estimate the uncertain one. Seeing that the liquid is between the 25 °C and 30 °C marks, we estimate the temperature to be 27 °C, being uncertain of the second digit of our measurement. By *uncertain* we mean that the temperature is reliably 27 °C and not 28 °C or 26 °C, but we can't say that it is *exactly* 27 °C.

All digits of a measured quantity, including the uncertain one, are called **significant figures**. A measured mass reported as 2.2 g has two significant figures, whereas one reported as 2.2405 g has five significant figures. The greater the number of significant figures, the greater the precision implied for the measurement.

To determine the number of significant figures in a reported measurement, read the number from left to right, counting the digits starting with the first digit that is not zero. *In any measurement that is properly reported, all nonzero digits are significant.* Because zeros can be used either as part of the measured value or merely to locate the decimal point, they may or may not be significant:

- Zeros *between* nonzero digits are always significant—1005 kg (four significant figures); 7.03 cm (three significant figures).
- Zeros *at the beginning* of a number are never significant; they merely indicate the position of the decimal point—0.02 g (one significant figure); 0.0026 cm (two significant figures).
- Zeros *at the end* of a number are significant if the number contains a decimal point—0.0200 g (three significant figures); 3.0 cm (two significant figures).

A problem arises when a number ends with zeros but contains no decimal point. In such cases, it is normally assumed that the zeros are not significant. Exponential notation (Appendix A.1) can be used to indicate whether end zeros are significant.



▲ **Figure 1.24** Uncertainty and significant figures in a measurement.

For example, a mass of 10 300 g can be written to show three, four, or five significant figures depending on how the measurement is obtained:

1.03×10^4 g	(three significant figures)
1.030×10^4 g	(four significant figures)
1.0300×10^4 g	(five significant figures)

In these numbers, all the zeros to the right of the decimal point are significant (rules 1 and 3). (The exponential term 10^4 does not add to the number of significant figures.)

Sample Exercise 1.6

Assigning Appropriate Significant Figures

The US state of Colorado is listed in a road atlas as having a population of 5,546,574 and an area of 269,837 square kilometers. Do the numbers of significant figures in these two quantities seem reasonable? If not, what seems to be wrong with them?

SOLUTION

The population of Colorado must vary from day to day as people move in or out, are born, or die. Thus, the reported number suggests a much higher degree of *accuracy* than is possible. Moreover, it would not be feasible to actually count every individual resident in the state at any given time. Thus, the reported number suggests far greater *precision* than is possible. A reported number of 5,500,000 would better reflect the actual state of knowledge.

The area of Colorado does not normally vary from time to time, so the question here is whether the accuracy of the measurements is good to six significant figures. It would be possible to achieve such accuracy using satellite technology, provided the legal boundaries are known with sufficient accuracy.

Practice Exercise

The back inside cover of the book tells us that there are 1000 cm^3 in 1 liter. Does this make the liter an exact volume?

Sample Exercise 1.7

Determining the Number of Significant Figures in a Measurement

How many significant figures are in each of the following numbers (assume that each number is a measured quantity): (**a**) 4.003, (**b**) 6.023×10^{23}, (**c**) 5000?

SOLUTION

(**a**) Four; the zeros are significant figures. (**b**) Four; the exponential term does not add to the number of significant figures. (**c**) One; we assume that the zeros are not significant when there is no decimal point shown. If the number has more significant figures, a decimal point should be employed or the number written in exponential notation. Thus, 5000. has four significant figures, whereas 5.00×10^3 has three.

Practice Exercise

An object is determined to have a mass of 0.01080 g. How many significant figures are there in this measurement?
(**a**) 2 (**b**) 3 (**c**) 4 (**d**) 5 (**e**) 6

Significant Figures in Calculations

Apply the following rule when carrying measured quantities through calculations.

The least certain measurement limits the certainty of the calculated quantity and thereby determines the number of significant figures in the final answer.

The final answer should be reported with only one uncertain digit. To keep track of significant figures in calculations, we will make frequent use of two rules: one for addition and subtraction, and another for multiplication and division.

1. *For addition and subtraction,* the result has the same number of decimal places as the measurement with the *fewest decimal places.* When the result contains more than the correct number of significant figures, it must be rounded off. Consider the following

example where the first three numbers are added to give the fourth and in which the uncertain digits appear in color:

This number limits the number of the significant figures in the result

20.42	⟵ two decimal places
1.322	⟵ three decimal places
83.1	⟵ one decimal places
104.842	⟵ round off to one decimal place (104.8)

We report the result as 104.8 because 83.1 has only one decimal place.

2. *For multiplication and division*, the result contains the same number of significant figures as the measurement with the *fewest significant figures*. When the result contains more than the correct number of significant figures, it must be rounded off. For example, the area of a rectangle whose measured edge lengths are 6.221 and 5.2 cm should be reported with two significant figures, 32 cm^2, even though a calculator shows the product to have more digits:

$$\text{Area} = (6.221\text{ cm})(5.2\text{ cm}) = 32.3492\text{ cm}^2 \Rightarrow \text{round off to } 32\text{ cm}^2$$

because 5.2 has two significant figures.

In determining the final answer for a calculated quantity, *exact numbers* are assumed to have an infinite number of significant figures. Thus, when we say, "There are 1000 meters in 1 kilometer," the number 1000 is exact, and we need not worry about the number of significant figures in it.

When *rounding off numbers* look at the leftmost digit to be removed:

- If the leftmost digit removed is less than 5, the preceding number is left unchanged. Thus, rounding off 7.248 to two significant figures gives 7.2.
- If the leftmost digit removed is 5 or greater, the preceding number is increased by 1. Rounding off 4.735 to three significant figures gives 4.74, and rounding 2.376 to two significant figures gives 2.4.*

Sample Exercise 1.8
Determining the Number of Significant Figures in a Calculated Quantity

The width, length, and height of a small box are 15.5, 27.3, and 5.4 cm, respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

SOLUTION

In reporting the volume, we can show only as many significant figures as given in the dimension with the fewest significant figures, which is that for the height (two significant figures):

$$\begin{aligned}\text{Volume} &= \text{width} \times \text{length} \times \text{height}\\ &= (15.5\text{ cm})(27.3\text{ cm})(5.4\text{ cm})\\ &= 2285.01\text{ cm}^3 \Rightarrow 2.3 \times 10^3\text{ cm}^3\end{aligned}$$

A calculator used for this calculation shows 2285.01, which we must round off to two significant figures. Because the resulting number is 2300, it is best reported in exponential notation, 2.3×10^3, to clearly indicate two significant figures.

▶ Practice Exercise

It takes 10.5 s for a sprinter to run 100.00 m. Calculate her average speed in meters per second and express the result to the correct number of significant figures.

*Your instructor may want you to use a slight variation on the rule when the leftmost digit to be removed is exactly 5, with no following digits or only zeros following. One common practice is to round up to the next higher number if that number will be even and otherwise leave the number unchanged. Thus, 4.7350 would be rounded to 4.74, and 4.7450 would also be rounded to 4.74.

Sample Exercise 1.9

Determining the Number of Significant Figures in a Calculated Quantity

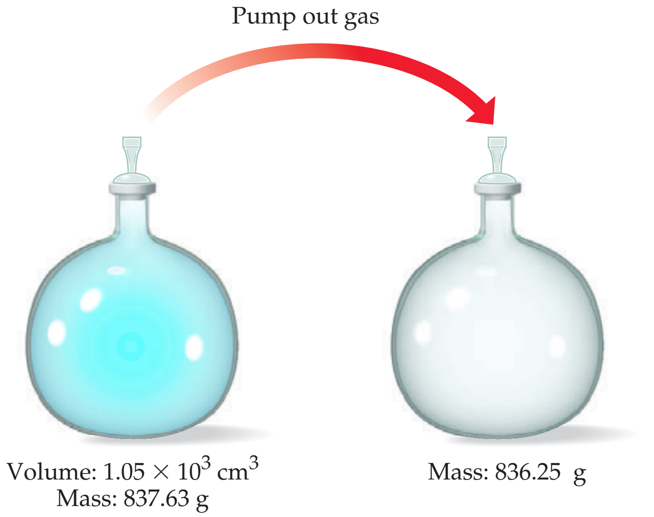
A vessel containing a gas at 25 °C is weighed, emptied, and then reweighed as depicted in **Figure 1.25**. From the data provided, calculate the density of the gas at 25 °C.

SOLUTION

To calculate the density, we must know both the mass and the volume of the gas. The mass of the gas is just the difference in the masses of the full and empty container:

$$\begin{array}{r} 837.63\ \text{g} \\ -836.25\ \text{g} \\ \hline 1.38\ \text{g} \end{array}$$

In subtracting numbers, we determine the number of significant figures in our result by counting decimal places in each quantity.



▲ **Figure 1.25 Uncertainty and significant figures in a measurement.**

In this case each quantity has two decimal places. Thus, the mass of the gas, 1.38 g, has two decimal places.

Using the volume given in the question, $1.05 \times 10^3\ \text{cm}^3$, and the definition of density, we have

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.38\ \text{g}}{1.05 \times 10^3\ \text{cm}^3}$$
$$= 1.31 \times 10^{-3}\ \text{g/cm}^3 = 0.00131\ \text{g/cm}^3$$

In dividing numbers, we determine the number of significant figures our result should contain by counting the number of significant figures in each quantity. There are three significant figures in our answer, corresponding to the number of significant figures in the two numbers that form the ratio. Notice that in this example, following the rules for determining significant figures gives an answer containing only three significant figures, even though the measured masses contain five significant figures.

▶ Practice Exercise

You are asked to determine the mass of a piece of copper using its reported density, 8.96 g/mL, and a 150-mL graduated cylinder. First, you add 105 mL of water to the graduated cylinder; then you place the piece of copper in the cylinder and record a volume of 137 mL. What is the mass of the copper reported with the correct number of significant figures? **(a)** 287 g **(b)** 3.5×10^{-3} g/mL **(c)** 286.72 g/mL **(d)** 3.48×10^{-3} g/mL **(e)** 2.9×10^2 g/mL

When a calculation involves two or more steps and you write answers for intermediate steps, retain at least one nonsignificant digit for the intermediate answers. This procedure ensures that small errors from rounding at each step do not combine to affect the final result. When using a calculator, you may enter the numbers one after another, rounding only the final answer. Accumulated rounding-off errors may account for small differences among results you obtain and answers given in the text for numerical problems.

Self-Assessment Exercise

1.20 Do the following numbers, all representing the thickness of a typical piece of paper, have the same number of significant figures: 0.10 mm, 1.0×10^{-4} m, 100 μm?

(a) Yes

(b) No

Exercises

1.21 Indicate which of the following are exact numbers: **(a)** the mass of a 7.5 by 12.5 cm index card, **(b)** the number of grams in a kilogram, **(c)** the volume of a cup of Seattle's Best coffee, **(d)** the number of centimeters in a kilometer, **(e)** the number of microseconds in a week, **(f)** the number of pages in this book.

1.22 What is the number of significant figures in each of the following measured quantities? **(a)** 902.5 kg, **(b)** 3×10^{-6} m, **(c)** 0.0096 L, **(d)** $2.94 \times 10^3\ \text{m}^2$, **(e)** 92.03 km **(f)** 782.234 g.

1.23 Round each of the following numbers to three significant figures and express the result in standard exponential notation: **(a)** 2048732.23 **(b)** 0.000292945 **(c)** −82454.09 **(d)** 942.057024 **(e)** −0.00000324683.

1.24 Carry out the following operations and express the answers with the appropriate number of significant numbers.

(a) $43.029 + 0.02348$ **(b)** $952.72 - 73.4201$

(c) $(2.93 \times 10^3)(0.732)$ **(d)** $0.06324/0.624$

1.25 You weigh an object on a balance and read the mass in grams according to the picture. How many significant figures are in this measurement?

Answers to Self-Assessment Exercises

1.20 (b)

1.7 Dimensional Analysis

Anyone who has travelled to another country and looked at the price of fuel there will experience the need to convert from one set of units to another. Most often, this conversion is between currencies but, in some cases, the standard volume may have different units—liters or gallons for example.

The SI system of units provides a common 'language' in which to communicate measurement. However, there are still some industries and some countries in which historic units are employed. By the end of this section, you should be able to

- Convert between different units for a particular measurement

Because measured quantities have units associated with them, it is important to keep track of units as well as numerical values when using the quantities in calculations. Throughout the text we use dimensional analysis in solving problems. In **dimensional analysis**, units are multiplied together or divided into each other along with the numerical values. Equivalent units cancel each other. Using dimensional analysis helps ensure that solutions to problems yield the proper units. Moreover, it provides a systematic way of solving many numerical problems and of checking solutions for possible errors.

Conversion Factors

The key to using dimensional analysis is the correct use of *conversion factors* to change one unit into another. A **conversion factor** is a fraction whose numerator and denominator are the same quantity expressed in different units. For example, 2.54 cm and 1 inch are the same length: 2.54 cm = 1 in. This relationship allows us to write two conversion factors:

$$\frac{2.54\text{ cm}}{1\text{ in.}} \quad \text{and} \quad \frac{1\text{ in.}}{2.54\text{ cm}}$$

We use the first factor to convert inches to centimeters. For example, the length in centimeters of an object that is 8.50 in. long is

$$\text{Number of centimeters} = (8.50\ \cancel{\text{in.}})\frac{2.54\text{ cm}}{1\ \cancel{\text{in.}}} = 21.6\text{ cm}$$

Desired unit

Given unit

The unit inches in the denominator of the conversion factor cancels the unit inches in the given data (8.50 *inches*), so that the centimeters unit in the numerator of the conversion factor becomes the unit of the final answer. Because the numerator and denominator of a conversion factor are equal, multiplying any quantity by a conversion factor is equivalent to multiplying by the number 1 and so does not change the intrinsic value of the quantity. The length 8.50 in. is the same as the length 21.6 cm.

In general, we begin any conversion by examining the units of the given data and the units we desire. We then ask ourselves what conversion factors we have available to take us from the units of the given quantity to those of the desired one. When we multiply a quantity by a conversion factor, the units multiply and divide as follows:

$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

If the desired units are not obtained in a calculation, an error must have been made somewhere. Careful inspection of units often reveals the source of the error.

Sample Exercise 1.10
Converting Units

If a woman has a mass of 115 lb, what is her mass in grams? (The measurement of a 'pound' has the units 'lb'. 1 lb = 453.6 g)

SOLUTION

Because we want to change from pounds to grams, we look for a relationship between these units of mass. The conversion factor table found on the back inside cover tells us that 1 lb = 453.6 g. To cancel pounds and leave grams, we write the conversion factor with grams in the numerator and pounds in the denominator:

$$\text{Mass in grams} = (115\ \cancel{\text{lb}})\left(\frac{453.6\text{ g}}{1\ \cancel{\text{lb}}}\right) = 5.22 \times 10^4\text{ g}$$

The answer can be given to only three significant figures, the number of significant figures in 115 lb. The process we have used is diagrammed on the top right column.

Given: lb

Use $\dfrac{453.6\text{ g}}{1\text{ lb}}$

Find: g

Practice Exercise

At a particular instant in time, the Earth is judged to be 92,955,000 miles from the Sun. What is the distance in kilometers to four significant figures? (1 mile = 1.6093 km.)

(**a**) 5763 $\times$ 10^4 km (**b**) 1.496 $\times$ 10^8 km (**c**) 1.49596 $\times$ 10^8 km (**d**) 1.483 $\times$ 10^4 km (**e**) 57,759,000 km

STRATEGIES FOR SUCCESS Estimating Answers

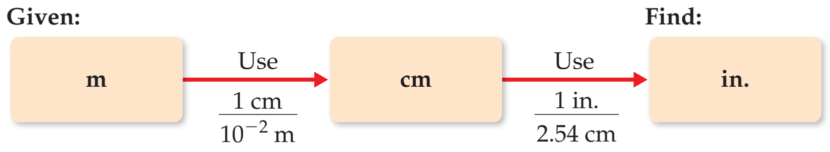
Calculators are wonderful devices; they enable you to get to the wrong answer very quickly. Of course, that's not the destination you want. You can take certain steps to avoid putting that wrong answer into your homework set or on an exam. One is to keep track of the units in a calculation and use the correct conversion factors. Second, you can do a quick mental check to be sure that your answer is reasonable: you can try to make a "ballpark" estimate.

A ballpark estimate involves making a rough calculation using numbers that are rounded off in such a way that the arithmetic can be done without a calculator. Even though this approach does not give an exact answer, it gives one that is roughly the correct size. By using dimensional analysis and by estimating answers, you can readily check the reasonableness of your calculations.

You can get better at making estimates by practicing in everyday life. How far is it from your house to the chemistry lecture hall? How many bikes are there on campus? If you respond "I have no idea" to these questions, you're giving up too easily. Try estimating familiar quantities and you'll get better at making estimates in science and in other aspects of your life where a misjudgment can be costly.

Using Two or More Conversion Factors

It is often necessary to use several conversion factors in solving a problem. As an example, let's convert the length of an 8.00 m rod to inches. We may be given the relationship between centimeters and inches (1 in. = 2.54 cm). We combine this with our knowledge of SI prefixes, ie 1 cm = 10^{-2} m. Thus, we can convert step by step, first from meters to centimeters and then from centimeters to inches:



Combining the given quantity (8.00 m) and the two conversion factors, we have

$$\text{Number of inches} = (8.00\ \cancel{\text{m}})\left(\frac{1\ \cancel{\text{cm}}}{10^{-2}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{in.}}{2.54\ \cancel{\text{cm}}}\right) = 315\ \text{in.}$$

The first conversion factor is used to cancel meters and convert the length to centimeters. Thus, meters are written in the denominator and centimeters in the numerator. The second conversion factor is used to cancel centimeters and convert the length to inches, so it has centimeters in the denominator and inches, the desired unit, in the numerator.

Note that you could have used 100 cm = 1 m as a conversion factor as well in the second parentheses. As long as you keep track of your given units and cancel them properly to obtain the desired units, you are likely to be successful in your calculations.

Sample Exercise 1.11
Converting Units Using Two or More Conversion Factors

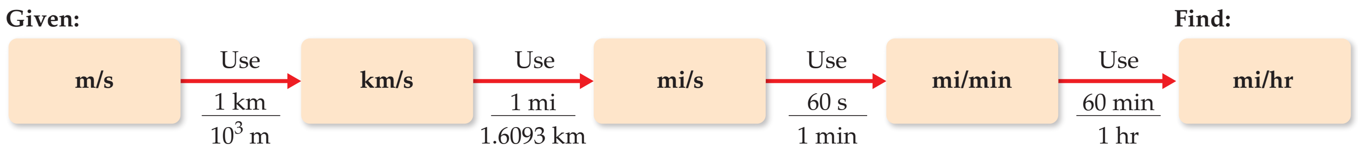
The average speed of a nitrogen molecule in air at 25 °C is 515 m/s. Convert this speed to miles per hour.

SOLUTION

To go from the given units, m/s, to the desired units, mi/hr, we must convert meters to miles and seconds to hours. From our knowledge of SI prefixes we know that 1 km = 10^3 m and we are given 1 mi = 1.6093 km. Thus, we can convert m to km and then convert km to mi. From our knowledge of time we know that 60 s = 1 min and 60 min = 1 hr. Thus, we can convert s to min and then convert min to hr. The overall process is

Applying first the conversions for distance and then those for time, we can set up one long equation in which unwanted units are canceled:

$$\text{Speed in mi/hr} = \left(515\frac{\cancel{\text{m}}}{\cancel{\text{s}}}\right)\left(\frac{1\ \cancel{\text{km}}}{10^3\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{mi}}{1.6093\ \cancel{\text{km}}}\right)\left(\frac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}}\right)\left(\frac{60\ \cancel{\text{min}}}{1\ \text{hr}}\right)$$
$$= 1.15 \times 10^3\ \text{mi/hr}$$



Our answer has the desired units. We can check our calculation, using the estimating procedure described in the "Strategies in Chemistry" box. The given speed is about 500 m/s. Dividing by 1000 converts m to km, giving 0.5 km/s. Because 1 mi is about 1.6 km, this speed corresponds to $0.5/1.6 = 0.3$ mi/s. Multiplying by 60 gives about $0.3 \times 60 = 20$ mi/min. Multiplying again by 60 gives $20 \times 60 = 1200$ mi/hr. The approximate solution (about 1200 mi/hr) and the detailed solution (1150 mi/hr) are reasonably close.

The answer to the detailed solution has three significant figures, corresponding to the number of significant figures in the given speed in m/s.

Practice Exercise

A car travels 28 mi per gallon of gasoline. What is the mileage in kilometers per liter?

Conversions Involving Volume

The conversion factors previously noted convert from one unit of a given measure to another unit of the same measure, such as from length to length. We also have conversion factors that convert from one measure to a different one. The density of a substance, for example, can be treated as a conversion factor between mass and volume. Suppose we want to know the mass in grams of 2 cubic inches (2.00 in.^{3}) of gold, which has a density of 19.3 g/cm^3. The density gives us the conversion factors:

$$\frac{19.3\ \text{g}}{1\ \text{cm}^3} \quad \text{and} \quad \frac{1\ \text{cm}^3}{19.3\ \text{g}}$$

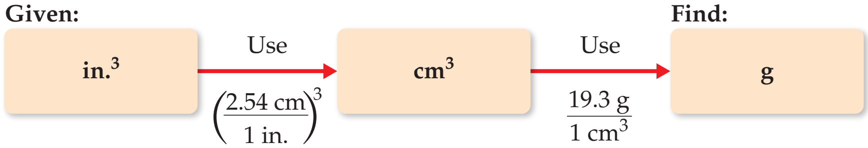
Because we want a mass in grams, we use the first factor, which has mass in grams in the numerator. To use this factor, however, we must first convert cubic inches to cubic centimeters. The relationship between in.^{3} and cm^3 may not be given directly, but the relationship between inches and centimeters is given: 1 in. = 2.54 cm (exactly). Cubing both sides of this equation gives $(1\ \text{in.})^3 = (2.54\ \text{cm})^3$, from which we write the desired conversion factor:

$$\frac{(2.54\ \text{cm})^3}{(1\ \text{in.})^3} = \frac{(2.54)^3\ \text{cm}^3}{(1)^3\ \text{in.}^3} = \frac{16.39\ \text{cm}^3}{1\ \text{in.}^3}$$

Notice that both the numbers and the units are cubed. Also, because 2.54 is an exact number, we can retain as many digits of $(2.54)^3$ as we need. We have used four, one more than the number of digits in the density (19.3 g/cm^3). Applying our conversion factors, we can now solve the problem:

$$\text{Mass in grams} = (2.00\ \cancel{\text{in.}^3})\left(\frac{16.39\ \cancel{\text{cm}^3}}{1\ \cancel{\text{in.}^3}}\right)\left(\frac{19.3\ \text{g}}{1\ \cancel{\text{cm}^3}}\right) = 633\ \text{g}$$

The procedure is diagrammed here. The final answer is reported to three significant figures, the same number of significant figures as in 2.00 in.^{3} and 19.3 g.



Sample Exercise 1.12
Converting Volume Units

Earth's oceans contain approximately 1.36×10^9 km^3 of water. Calculate the volume in liters.

SOLUTION

We know 1 L = 10^{-3} m^3 (Figure 1.19), but there is no relationship listed involving km^3. From our knowledge of SI prefixes, however, we know 1 km = 10^3 m and we can use this relationship between lengths to write the desired conversion factor between volumes:

$$\left(\frac{10^3 \text{ m}}{1 \text{ km}}\right)^3 = \frac{10^9 \text{ m}^3}{1 \text{ km}^3}$$

Thus, converting from km^3 to m^3 to L, we have

$$\text{Volume in liters} = (1.36 \times 10^9 \cancel{\text{km}^3})\left(\frac{10^9 \cancel{\text{m}^3}}{1 \cancel{\text{km}^3}}\right)\left(\frac{1 \text{ L}}{10^{-3} \cancel{\text{m}^3}}\right)$$

$$= 1.36 \times 10^{21} \text{ L}$$

How many liters of water do Earth's oceans contain?

Practice Exercise

A barrel of oil as measured on the oil market is equal to 1.333 U.S. barrels. A U.S. barrel is equal to 31.5 gal. If oil is on the market at \$94.0 per barrel, what is the price in dollars per gallon?

(a) \$2.24/gal **(b)** \$3.98/gal **(c)** \$2.98/gal **(d)** \$1.05/gal **(e)** \$8.42/gal

STRATEGIES FOR SUCCESS The Importance of Practice

If you have ever played a musical instrument or participated in athletics, you know that the keys to success are practice and discipline. You cannot learn to play a piano merely by listening to music, and you cannot learn how to play basketball merely by watching games on television. Likewise, you cannot learn chemistry by merely watching your instructor give lectures. Simply reading this book, listening to lectures, or reviewing notes will not usually be sufficient when exam time comes around. Your task is to master chemical concepts to a degree that you can put them to use in solving problems and answering questions. Solving problems correctly takes practice—actually, a fair amount of it. You will do well in your chemistry course if you embrace the idea that you need to master the materials presented and then learn how to apply them in solving problems. Even if you're a brilliant student, this will take time; it's what being a student is all about. Almost no one fully absorbs new material on a first reading, especially when new concepts are being presented. You are sure to master the content of the chapters more fully by reading them through at least twice, even more for passages that present you with difficulties in understanding.

Throughout the book, we have provided sample exercises in which the solutions are shown in detail. For practice exercises, we supply only the answer, at the back of the book. It is important that you use these exercises to test yourself.

The practice exercises in this text and the homework assignments given by your instructor provide the minimal practice that you will need to succeed in your chemistry course. Only by working all the assigned problems will you face the full range of difficulty and coverage that your instructor expects you to master for exams. There is no substitute for a determined and perhaps lengthy effort to work problems on your own. If you are stuck on a problem, however, ask for help from your instructor, a teaching assistant, a tutor, or a fellow student. Spending an inordinate amount of time on a single exercise is rarely effective unless you know that it is particularly challenging and is expected to require extensive thought and effort.

STRATEGIES FOR SUCCESS The Features of This Book

If, like most students, you haven't yet read the part of the Preface to this text entitled TO THE STUDENT, *you should do it now*. In less than two pages of reading you will encounter valuable advice on how to navigate your way through this book and through the course. We're serious! This is advice you can use.

The TO THE STUDENT section describes how text features such as "What's Ahead," Key Terms, Learning Outcomes, and Key Equations will help you remember what you have learned. If you have registered for MasteringChemistry®, you will have access to many helpful animations, tutorials, and additional problems correlated to specific topics and sections of each chapter. An eBook is also available online. In addition, the Pearson eText brings to life the content of each chapter with animations and videos. Interactive end-of-section Self-Assessment Exercises are featured with specific wrong-answer feedback.

As previously mentioned, working exercises is very important—in fact, essential. You will find a large variety of exercises at the end of each section and chapter that are designed to test your problem-solving skills in chemistry. Your instructor will very likely assign some of these exercises as homework.

- The first few exercises called "Visualizing Concepts" are meant to test how well you understand a concept without plugging a lot of numbers into a formula.
- Additional Exercises appear after the regular exercises; the chapter sections that they cover are not identified.
- Integrative Exercises, which start appearing in Chapter 3, are problems that require skills learned in previous chapters.

- Also first appearing in Chapter 3 are Design an Experiment exercises consisting of problem scenarios that challenge you to design experiments to test hypotheses.

Many chemical databases are available, usually on the Web.

- The *CRC Handbook of Chemistry and Physics* is the standard reference for many types of data and is available in libraries.
- The *Merck Index* is a standard reference for the properties of many organic compounds, especially those of biological interest.
- WebElements (http://www.webelements.com/) is a good website for looking up the properties of the elements.
- Wolfram Alpha (http://www.wolframalpha.com/) can also be a source of useful information on substances, numerical values, and other data.

Self-Assessment Exercise

1.26 A soft drink indicates that it contain 21 kJ of energy. Given 1 J = 0.2390 calories, what is the energy content in calories (hint: mind the significant figures)?

(a) 5.0 cal

(b) 88 cal

(c) 5.0 kcal

(d) 5019 cal

Exercises

1.27 Using your knowledge of metric units and the information given in the chapter, write down the conversion factors needed to convert **(a)** in. to cm **(b)** lb to g **(c)** μg to g **(d)** ft^2 to cm^2.

1.28 **(a)** A bumblebee flies with a ground speed of 15.2 m/s. Calculate its speed in km/hr. **(b)** The lung capacity of the blue whale is 5.0×10^3 L. Convert this volume into gallons. **(c)** The Statue of Liberty is 151 ft tall. Calculate its height in meters. **(d)** Bamboo can grow up to 60.0 cm/day. Convert this growth rate into inches per hour.

1.29 Perform the following conversions: **(a)** 5.00 days to s, **(b)** 0.0550 miles to m, **(c)** \$1.89/gal to dollars per liter, **(d)** 0.510 in/ms to km/h, **(e)** 22.50 gal/min to L/s, **(f)** 0.02500 ft^3 to cm^3.

1.30 **(a)** How many liters of wine can be held in a wine barrel whose capacity is 31 gal? **(b)** The recommended adult dose of Elixophyllin®, a drug used to treat asthma, is 6 mg/kg of body mass. Calculate the dose in milligrams for a 185 lb person. **(c)** If an automobile is able to travel 400 km on 47.3 L of fuel, what is the gas mileage in miles per gallon? **(d)** When the coffee is brewed according to directions, a pound of coffee beans yields 50 cups of coffee. How many kg of coffee are required to produce 200 cups of coffee?

1.31 The density of air at ordinary atmospheric pressure and 25 °C is 1.19 g/L. What is the mass, in kilograms, of the air in a room that measures 4.5 m $\times$ 5.0 m $\times$ 2.5 m?

1.32 Gold can be hammered into extremely thin sheets called gold leaf. An architect wants to cover a 30 m $\times$ 25 m ceiling with gold leaf that is twelve-millionths of a centimeter thick. The density of gold is 19.32 g/cm^3, and gold costs \$1654 per troy ounce (1 troy ounce = 31.1034768 g). How much will it cost the architect to buy the necessary gold?

Answers to Self-Assessment Exercises

1.26 (c)

Chapter Summary and Key Terms

THE STUDY OF CHEMISTRY (SECTION 1.1) **Chemistry** is the study of the composition, structure, properties, and changes of **matter**. The composition of matter relates to the kinds of **elements** it contains. The structure of matter relates to the ways the **atoms** of these elements are arranged. A **property** is any characteristic that gives a sample of matter its unique identity. A **molecule** is an entity composed of two or more atoms, with the atoms attached to one another in a specific way.

The **scientific method** is a dynamic process used to answer questions about the physical world. Observations and experiments lead to tentative explanations or **hypotheses**. As a hypothesis is tested and refined, a **theory** may be developed that can predict the results of future observations and experiments. When observations repeatedly lead to the same consistent results, we speak of a **scientific law**, a general rule that summarizes how nature behaves.

CLASSIFICATIONS OF MATTER (SECTION 1.2) Matter exists in three physical states, **gas**, **liquid**, and **solid**, which are known as the **states of matter**. There are two kinds of **pure substances: elements** and **compounds**. Each element has a single kind of atom and is represented by a chemical symbol consisting of one or two letters, with the first letter capitalized. Compounds are composed of two or more elements joined chemically. The **law of constant composition**, also called the **law of definite proportions**, states that the elemental composition of a pure compound is always the same. Most matter consists of a mixture of substances. **Mixtures** have variable compositions and can be either homogeneous or heterogeneous; homogeneous mixtures are called **solutions**.

PROPERTIES OF MATTER (SECTION 1.3) Each substance has a unique set of **physical properties** and **chemical properties** that can be used to identify it. During a **physical change**, matter does not change its composition. **Changes of state** are physical changes. In a **chemical change (chemical reaction)**, a substance is transformed into a chemically different substance. **Intensive properties** are independent of the amount of matter examined and are used to identify substances. **Extensive properties** relate to the amount of substance present. Differences in physical and chemical properties are used to separate substances.

THE NATURE OF ENERGY (SECTION 1.4) **Energy** is defined as the capacity to do work or transfer heat. **Work** is the energy transferred when a force exerted on an object causes a displacement of that object, and **heat** is the energy used to cause the temperature of an object to increase. An object can possess energy in two forms: **kinetic energy**, which is the energy associated with the motion of an object, and **potential energy**, which is the energy that an object possesses by virtue of its position relative to other objects. Important forms of potential energy include gravitational energy and electrostatic energy.

UNITS OF MEASUREMENT (SECTION 1.5) Measurements in chemistry are made using the **metric system**. Special emphasis is placed on **SI units**, which are based on the meter, the kilogram, and the second as the basic units of length, **mass**, and time, respectively. SI units use prefixes to indicate fractions or multiples of base units. The SI **temperature** scale is the **Kelvin scale**, although the **Celsius scale** is frequently used as well. **Absolute zero** is the lowest temperature attainable. It has the value 0 K. A **derived unit** is obtained by multiplication or division of SI base units. Derived units are needed for defined quantities such as speed or volume. **Density** is an important derived unit that equals mass divided by volume.

UNCERTAINTY IN MEASUREMENT (SECTION 1.6) All measured quantities are inexact to some extent. The **precision** of a measurement indicates how closely different measurements of a quantity agree with one another. The **accuracy** of a measurement indicates how well a measurement agrees with the accepted or "true" value. The **significant figures** in a measured quantity include one estimated digit, the last digit of the measurement. The significant figures indicate the extent of the uncertainty of the measurement. Certain rules must be followed so that a calculation involving measured quantities is reported with the appropriate number of significant figures.

DIMENSIONAL ANALYSIS (SECTION 1.7) In the **dimensional analysis** approach to problem solving, we keep track of units as we carry measurements through calculations. The units are multiplied together, divided into each other, or canceled like algebraic quantities. Obtaining the proper units for the final result is an important means of checking the method of calculation. When converting units and when carrying out several other types of problems, **conversion factors** can be used. These factors are ratios constructed from valid relations between equivalent quantities.

Learning Outcomes After studying this chapter, you should be able to:

- Distinguish among elements, compounds, and mixtures. (Section 1.2) *Related Exercises: 1.3, 1.5, 1.45*
- Identify symbols of common elements. (Section 1.2) *Related Exercises: 1.4, 1.46*
- Distinguish between chemical and physical changes. (Section 1.3) *Related Exercises: 1.8, 1.9, 1.47, 1.48*
- Distinguish between kinetic and potential energy. (Section 1.4) *Related Exercises: 1.12, 1.13*
- Calculate the kinetic energy of an object. (Section 1.4). *Related Exercises: 1.50, 1.51, 1.52, 1.53*
- Identify common metric prefixes. (Section 1.5) *Related Exercises: 1.15, 1.54*
- Demonstrate the use of significant figures, scientific notation, and SI units in calculations. (Section 1.6) *Related Exercises: 1.22, 1.24, 1.63, 1.65*
- Use appropriate SI units for defined quantities, and employ dimensional analysis in calculations. (Sections 1.5 and 1.7) *Related Exercises: 1.28, 1.30, 1.68, 1.70*

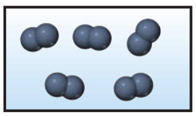
Key Equations

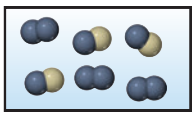
- $w = F \times d$ [1.1] Work done by a force in the direction of displacement
- $E_k = \frac{1}{2}mv^2$ [1.2] Kinetic energy
- $K = {}^\circ C + 273.15$ [1.3] Converting between Celsius (°C) and Kelvin (K) temperature scales
- $\text{Density} = \frac{\text{mass}}{\text{volume}}$ [1.4] Definition of density

Exercises

Visualizing Concepts

1.33 Which of the following figures represents (**a**) a pure element, (**b**) a mixture of two elements, (**c**) a pure compound, (**d**) a mixture of an element and a compound? (More than one picture might fit each description.) [Section 1.2]


(i)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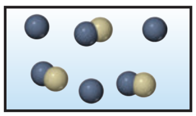
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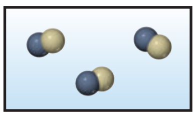
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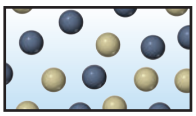
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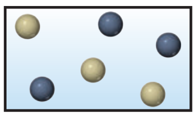
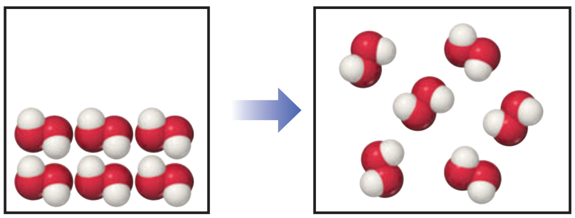
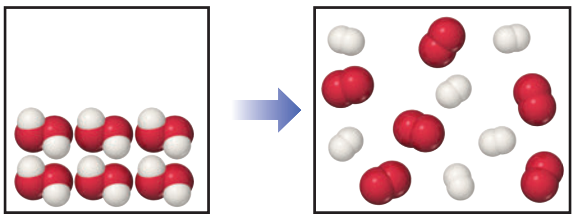

(vi)

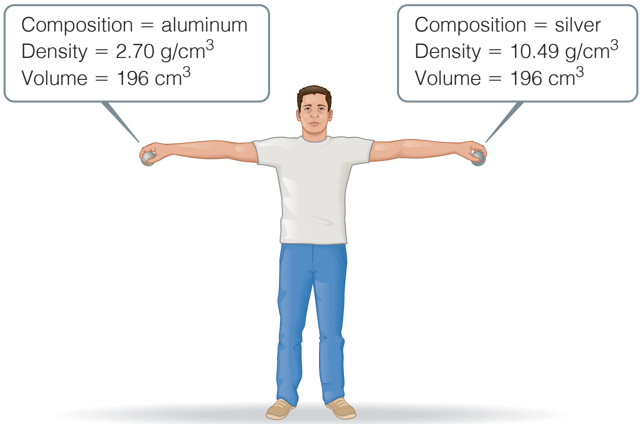
1.34 Which of the following diagrams represents a chemical change? [Section 1.3]

(a)


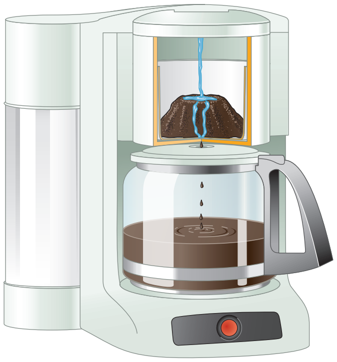
(b)


1.35 Musical instruments like trumpets and trombones are made from an alloy called brass. Brass is composed of copper and zinc atoms and appears homogeneous under an optical microscope. The approximate composition of most brass objects is a 2:1 ratio of copper to zinc atoms, but the exact ratio varies somewhat from one piece of brass to another. (**a**) Would you classify brass as an element, a compound, a homogeneous mixture, or a heterogeneous mixture? (**b**) Would it be correct to say that brass is a solution? [Section 1.2]

1.36 Consider the two spheres shown here, one made of silver and the other of aluminum. (**a**) What is the mass of each sphere in kg? (**b**) The force of gravity acting on an object is $F = mg$, where m is the mass of an object and g is the acceleration of gravity (9.8 m/s^2). How much work do you do on each sphere it you raise it from the floor to a height of 2.2 m? (**c**) Does the act of lifting the sphere off the ground increase the potential energy of the aluminum sphere by a larger, smaller, or same amount as the silver sphere? (**d**) If you release the spheres simultaneously, they will have the same velocity when they hit the ground. Will they have the same kinetic energy? If not, which sphere will have more kinetic energy? [Section 1.4]



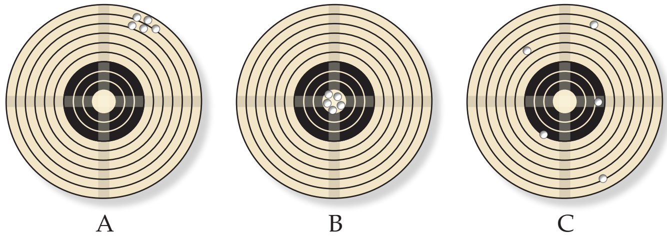
1.37 Is the separation method used in brewing a cup of coffee best described as distillation, filtration, or chromatography? [Section 1.3]



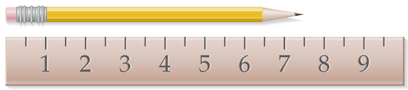
1.38 Identify each of the following as measurements of length, area, volume, mass, density, time, or temperature: (**a**) 25 ps, (**b**) 374.2 mg, (**c**) 77 K, (**d**) 100,000 km^2, (**e**) 1.06 μm, (**f**) 16 nm^2, (**g**) −78 °C, (**h**) 2.56 g/cm^3, (**i**) 28 cm^3. [Section 1.5]

1.39 (**a**) Three spheres of equal size are composed of aluminum (density = 2.70 g/cm^3), silver (density = 10.49 g/cm^3), and nickel (density = 8.90 g/cm^3). List the spheres from lightest to heaviest. (**b**) Three cubes of equal mass are composed of gold (density = 19.32 g/cm^3), platinum (density = 21.45 g/cm^3), and lead (density = 11.35 g/cm^3). List the cubes from smallest to largest. [Section 1.5]

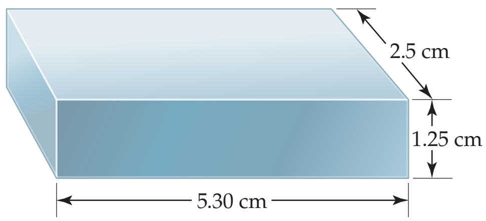
1.40 The three targets from a rifle range shown here were produced by: (**A**) the instructor firing a newly acquired target rifle; (**B**) the instructor firing his personal target rifle; and (**C**) a student who has fired his target rifle only a few times. (**a**) Comment on the accuracy and precision for each of these three sets of results. (**b**) For the A and C results in the future to look like those in B, what needs to happen? [Section 1.6]



1.41 (**a**) What is the length of the pencil in the following figure if the ruler reads in centimeters? How many significant figures are there in this measurement? (**b**) An automobile speedometer with circular scales reading both miles per hour and kilometers per hour is shown. What speed is indicated, in both units? How many significant figures are in the measurements? [Section 1.6]



1.42 (**a**) How many significant figures should be reported for the volume of the metal bar shown here? (**b**) If the mass of the bar is 104.72 g, how many significant figures should be reported when its density is determined using the calculated volume? [Section 1.6]



1.43 Consider the jar of jelly beans in the photo. To get an estimate of the number of beans in the jar you weigh six beans and obtain masses of 3.15, 3.12, 2.98, 3.14, 3.02, and 3.09 g. Then you weigh the jar with all the beans in it, and obtain a mass of 2082 g. The empty jar has a mass of 653 g. Based on these data, estimate the number of beans in the jar. Justify the number of significant figures you use in your estimate. [Section 1.6]

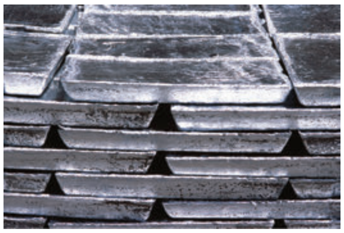
1.44 This photo shows a picture of an agate stone. Jack, who picked up the stone on the Lake Superior shoreline and polished it, insists that agate is a chemical compound. Ellen argues that it cannot be a compound. Discuss the relative merits of their positions. [Section 1.2]

Classification and Properties of Matter (Sections 1.2 and 1.3)

1.45 Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: (**a**) milk, (**b**) beer, (**c**) diamond, (**d**) mayonnaise.

1.46 Give the chemical symbol or name for each of the following elements, as appropriate: (**a**) rhenium, (**b**) tungsten, (**c**) caesium, (**d**) hydrogen, (**e**) indium, (**f**) As, (**g**) Xe, (**h**) Kr, (**i**) Te, (**j**) Ge.

1.47 (**a**) Read the following description of the element zinc and indicate which are physical properties and which are chemical properties.



Zinc melts at 420 °C. When zinc granules are added to dilute sulfuric acid, hydrogen is given off and the metal dissolves. Zinc has a hardness on the Mohs scale of 2.5 and a density of 7.13 g/cm^3 at 25 °C. It reacts slowly with oxygen gas at elevated temperatures to form zinc oxide, ZnO.

(**a**) Which properties of zinc can you describe from the photo? Are these physical or chemical properties?

1.48 A match is lit to light a candle. The following observations are made: (**a**) The candle burns. (**b**) Some wax melts. (**c**) Melted wax solidifies on the candleholder. (**d**) Soot (carbon) is produced by the burning of the match and the candle. Which of these occurrences are due to physical changes, and which are due to chemical changes?

1.49 A silvery metal is put inside a beaker of water. Bubbles form on the surface of the metal and it dissolves gradually. (**a**) Is this an example of a chemical or a physical change? (**b**) Do you expect the remaining solution to be a pure substance or a mixture?

The Nature of Energy (Section 1.4)

1.50 (**a**) Calculate the kinetic energy, in joules, of a 15-g bullet moving at 120 m/s. (**b**) When the bullet is stopped by a bulletproof vest, which form of energy does the kinetic energy of the bullet convert to?

1.51 (**a**) A baseball weighs 145.4 g. What is the kinetic energy, in joules, of this baseball when it is thrown by a major league pitcher at 150 km/h? (**b**) By what factor will the kinetic energy change if the speed of the baseball is decreased to 90 km/h? (**c**) What happens to the kinetic energy when the baseball is caught by the catcher? Is it converted mostly to heat or to some form of potential energy?

1.52 What is the kinetic energy and velocity of the aluminum sphere in Problem 1.36 at the moment it hits the ground? (Assume that energy is conserved during the fall and that 100% of the sphere's initial potential energy is converted to kinetic energy by the time impact occurs.)

1.53 What is the kinetic energy and velocity of the silver sphere in Problem 1.36 at the moment it hits the ground? (Assume that energy is conserved during the fall and that 100% of the sphere's initial potential energy is converted to kinetic energy by the time impact occurs.)

Units of Measurement (Section 1.5)

1.54 Use appropriate metric prefixes to write the following measurements without use of exponents: (**a**) 7.29×10^6 g (**b**) 6.1×10^{-10} m (**c**) 1.828×10^{-3} s (**d**) 3.523×10^9 m^3 (**e**) 9.62×10^2 m/s (**f**) 8.923×10^{-12} kg (**g**) 3.552×10^{12} L.

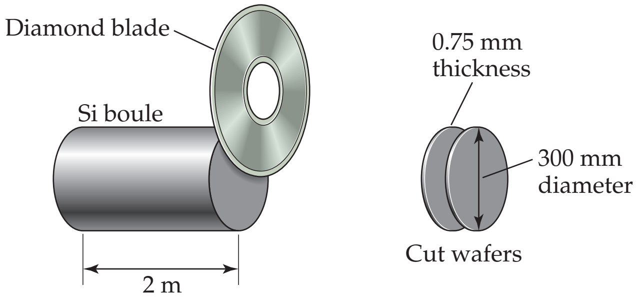
1.55 Make the following conversions: (**a**) 83 °F to °C (**b**) 29 °C to °F (**c**) 294 °C to K (**d**) 832 K to °C (**e**) 721 K to °F (**f**) 35 °F to K.

1.56 (**a**) A child has a fever of 101 °F. What is the temperature in °C? (**b**) In a desert, the temperature can be as high as 45 °C, what is the temperature in °F? (**c**) During winter, the temperature of the Arctic region can drop below −50 °C, what is the temperature in degree Fahrenheit and in Kelvin? (**d**) The sublimation temperature of dry ice is −78.5 °C. Convert this temperature to degree Fahrenheit and Kelvin. **(e)** Ethanol boils at 351 K. Convert this temperature to degree Fahrenheit and degree Celsius.

1.57 (**a**) What is the mass of a silver cube whose edges measure 2.00 cm each at 25 °C? The density of silver is 10.49 g/cm^3 at 25 °C. (**b**) The density of aluminum is 2.70 g/cm^3 at 25 °C. What is the weight of the aluminum foil with an area of 0.5 m^2 and a thickness of 0.5 mm? (**c**) The density of hexane is 0.655 g/mL at 25 °C. Calculate the mass of 1.5 L of hexane at this temperature.

1.58 (**a**) After the label fell off a bottle containing a clear liquid believed to be benzene, a chemist measured the density of the liquid to verify its identity. A 25.0-mL portion of the liquid had a mass of 21.95 g. A chemistry handbook lists the density of benzene at 15 °C as 0.8787 g/mL. Is the calculated density in agreement with the tabulated value? (**b**) An experiment requires 15.0 g of cyclohexane, whose density at 25 °C is 0.7781 g/mL. What volume of cyclohexane should be used? (**c**) A spherical ball of lead has a diameter of 5.0 cm. What is the mass of the sphere if lead has a density of 11.34 g/cm^3? (The volume of a sphere is $(4/3)\pi r^3$, where r is the radius.)

1.59 Silicon for computer chips is grown in large cylinders called "boules" that are 300 mm in diameter and 2 m in length, as shown. The density of silicon is 2.33 g/cm^3. Silicon wafers for making integrated circuits are sliced from a 2.0 m boule and are typically 0.75 mm thick and 300 mm in diameter. (**a**) How many wafers can be cut from a single boule? (**b**) What is the mass of a silicon wafer? (The volume of a cylinder is given by $\pi r^2 h$, where r is the radius and h is its height.)



1.60 Use of the British thermal unit (Btu) is common in some types of engineering work. A Btu is the amount of heat required to raise the temperature of 1 lb of water by 1°F. Calculate the number of joules in a Btu.

1.61 A watt is a measure of power (the rate of energy change) equal to 1 J/s. **(a)** Calculate the number of joules in a kilowatt-hour. (**b**) An adult person radiates heat to the surroundings at about the same rate as a 100-watt electric incandescent light bulb. What is the total amount of energy in kcal radiated to the surroundings by an adult over a 24 h period?

Uncertainty in Measurement (Section 1.6)

1.62 Indicate which of the following are exact numbers: (**a**) the mass of a 945-mL can of coffee, (**b**) the number of students in your chemistry class, (**c**) the temperature of the surface of the Sun, (**d**) the mass of a postage stamp, (**e**) the number of milliliters in a cubic meter of water, (**f**) the average height of NBA basketball players.

1.63 Indicate the number of significant figures in each of the following measured quantities: (**a**) 62.65 km/hr, (**b**) 78.00 K, (**c**) 36.9 mL, (**d**) 250 mm, (**e**) 89.2 metric tons, (**f**) 6.4224×10^2 m^3.

1.64 (**a**) The diameter of Earth at the equator is 12756.27 km. Round this number to three significant figures and express it in standard exponential notation. (**b**) The circumference of Earth through the poles is 40,008 km. Round this number to four significant figures and express it in standard exponential notation.

1.65 Carry out the following operations and express the answers with the appropriate number of significant numbers.

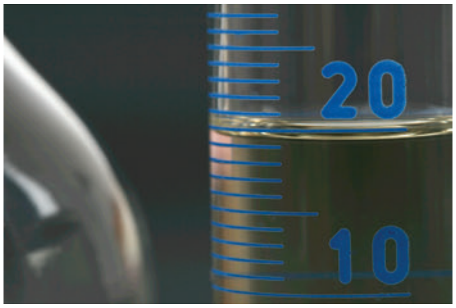
(**a**) $(6.234 + 8.72) \times 0.6746$

(**b**) $732.1 - (892.5/8.2)$

(**c**) $[(3.696 \times 10^5) - (6.234 \times 10^3)] \times 0.0742$

(**d**) $0.006438 \times 108 - (8.639 + 8.52)$

1.66 You have a graduated cylinder that contains a liquid (see photograph). Write the volume of the liquid, in milliliters, using the proper number of significant figures.



Dimensional Analysis (Section 1.7)

Note: 12 inches = 1 foot

1 m = 39.37 inches

1 mile = 1.609 km

1 pound = 454 g

1 gallon = 3.7854 L

1 m^3 = 264 gallon

1.67 Using your knowledge of metric units and the given information, write down the conversion factors needed to convert (**a**) km/h to m/s (**b**) mL to μL (**c**) ps to s (**d**) m^3 to gal.

1.68 (**a**) The speed of light in a vacuum is 2.998×10^8 m/s. Calculate its speed in miles per hour. (**b**) The Sears Tower in Chicago is 1454 ft tall. Calculate its height in meters. (**c**) The Vehicle Assembly Building at the Kennedy Space Center in Florida has a volume of 3,666,500 m^3. Convert this volume to liters and express the result in standard exponential notation. (**d**) An individual suffering from a high cholesterol level in her blood has 242 mg of cholesterol per 100 mL of blood. If the total blood volume of the individual is 5.2 L, how many grams of total blood cholesterol does the individual's body contain?

1.69 Carry out the following conversions: (**a**) 0.105 in. to mm, (**b**) 8.75 mm/s to km/h, (**c**) \$3.99/lb to dollars per kg, (**d**) 8.75 lb/ft^3 to g/mL.

1.70 (**a**) In March 1989, the *Exxon Valdez* ran aground and spilled 240,000 barrels of crude petroleum off the coast of Alaska. One barrel of petroleum is equal to 42 gal. How many liters of petroleum were spilled?

1.71 The indoor concentration of ozone above 300 μg/m^3 is considered to be unhealthy. What mass of ozone in grams is present in a room measuring 3.2 m $\times$ 2.8 m $\times$ 4.1 m?

1.72 A copper refinery produces a copper ingot weighing 70 kg. If the copper is drawn into wire whose diameter is 7.50 mm, how many meters of copper can be obtained from the ingot? The density of copper is 8.94 g/cm^3. (Assume that the wire is a cylinder whose volume $V = \pi r^2 h$, where r is its radius and h is its height or length.)

Additional Exercises

1.73 Classify each of the following as a pure substance, a solution, or a heterogeneous mixture: (**a**) a leaf, (**b**) a 999 gold bar, (**c**) stainless steel.

1.74 (**a**) Which is more likely to eventually be shown to be incorrect: an hypothesis or a theory? (**b**) A(n) _________ reliably predicts the behavior of matter, while a(n) _________ provides an explanation for that behavior.

1.75 A sample of ascorbic acid (vitamin C) is synthesized in the laboratory. It contains 1.50 g of carbon and 2.00 g of oxygen. Another sample of ascorbic acid isolated from citrus fruits contains 6.35 g of carbon. According to the law of constant composition, how many grams of oxygen does it contain?

1.76 Ethyl chloride is sold as a liquid (see photo) under pressure for use as a local skin anesthetic. Ethyl chloride boils at 12 °C at atmospheric pressure. When the liquid is sprayed onto the skin, it boils off, cooling and numbing the skin as it vaporizes. (**a**) What changes of state are involved in this use of ethyl chloride? (**b**) What is the boiling point of ethyl chloride in degrees Fahrenheit? (**c**) The bottle shown contains 103.5 mL of ethyl chloride. The density of ethyl chloride at 25 °C is 0.765 g/cm^3. What is the mass of ethyl chloride in the bottle?

1.77 Two students determine the percentage of lead in a sample as a laboratory exercise. The true percentage is 22.52%. The students' results for three determinations are as follows:

(**1**) 22.52, 22.48, 22.54

(**2**) 22.64, 22.58, 22.62

(**a**) Calculate the average percentage for each set of data and state which set is the more accurate based on the average. (**b**) Precision can be judged by examining the average of the deviations from the average value for that data set. (Calculate the average value for each data set; then calculate the average value of the absolute deviations of each measurement from the average.) Which set is more precise?

1.78 Is the use of significant figures in each of the following statements appropriate? (**a**) The 2005 circulation of *National Geographic* was 7,812,564. (**b**) On July 1, 2005, the population of Cook County, Illinois, was 5,303,683. (**c**) In the United States, 0.621% of the population has the surname Brown. (**d**) You calculate your grade point average to be 3.87562.

1.79 What type of quantity (for example, length, volume, density) do the following units indicate? (**a**) m^3, (**b**) ns, (**c**) mm, (**d**) g/dm^3, (**e**) °C, (**f**) ms^{-1}, (**g**) Pa.

1.80 Give the derived SI units for each of the following quantities in base SI units:

(**a**) acceleration = distance/time^2

(**b**) force = mass × acceleration

(**c**) work = force × distance

(**d**) pressure = force/area

(**e**) power = work/time

(**f**) velocity = distance/time

(**g**) energy = mass × (velocity)^2

1.81 The distance from Earth to the Moon is approximately 240,000 mi. (**a**) What is this distance in meters? (**b**) The peregrine falcon has been measured as traveling up to 350 km/hr in a dive. If this falcon could fly to the Moon at this speed, how many seconds would it take? (**c**) The speed of light is 3.00×10^8 m/s. How long does it take for light to travel from Earth to the Moon and back again? (**d**) Earth travels around the Sun at an average speed of 29.783 km/s. Convert this speed to miles per hour.

1.82 Which of the following would you characterize as pure or nearly pure substance? (**a**) stomach acid; (**b**) dry ice; (**c**) ice-cream; (**d**) stainless steel; (**e**) petroleum; (**f**) distilled water; (**g**) carbon monoxide gas; (**h**) compressed air in balloon.

1.83 The U.S. quarter has a mass of 5.67 g and is approximately 1.55 mm thick. (**a**) How many quarters would have to be stacked to reach 575 ft, the height of the Washington Monument? (**b**) How much would this stack weigh? (**c**) How much money would this stack contain? (**d**) The U.S. National Debt Clock showed the outstanding public debt to be $16,213,166,914,811 on October 28, 2012. How many stacks like the one described would be necessary to pay off this debt?

1.84 In the United States, water used for irrigation is measured in acre-feet. An acre-foot of water covers an acre to a depth of exactly 1 ft. An acre is 4840 yd^2. An acre-foot is enough water to supply two typical households for 1.00 yr. (**a**) If desalinated water costs $1950 per acre-foot, how much does desalinated water cost per liter? (**b**) How much would it cost one household per day if it were the only source of water?

1.85 By using estimation technique, determine which of the following is the heaviest and which is the lightest: a 10-lb bag of fertilizer, a 10-kg bag of rice, or 2 gal of olive oil (density = 0.918 g/cm^3).

1.86 Suppose you decide to define your own temperature scale with units of O, using the freezing point (13 °C) and boiling point (360 °C) of oleic acid, the main component of olive oil. If you set the freezing point of oleic acid as 0 °O and the boiling point as 100 °O, what is the freezing point of water on this new scale?

1.87 Hexane (density = 0.659 g/mL) and acetic acid (density = 1.0446 g/mL) do not form a solution when mixed but are separate in distinct layers. A piece of oak wood (density = 900 kg/m^3) is placed inside a test tube containing hexane and acetic acid solution; sketch how the three substances would position themselves.

1.88 Two spheres of equal volume are placed on the scales as shown. Which one is more dense?

1.89 Water has a density of 0.997 g/cm^3 at 25 °C; ice has a density of 0.917 g/cm^3 at −10 °C. (**a**) If a soft-drink bottle whose volume is 1.50 L is completely filled with water and then frozen to −10 °C, what volume does the ice occupy? (**b**) Can the ice be contained within the bottle?

1.90 A 32.65-g sample of a solid is placed in a flask. Toluene, in which the solid is insoluble, is added to the flask so that the total volume of solid and liquid together is 50.00 mL. The solid and toluene together weigh 58.58 g. The density of toluene at the temperature of the experiment is 0.864 g/mL. What is the density of the solid?

1.91 A thief plans to steal a cylindrical platinum medal with a radius of 2.3 cm and a thickness of 0.8 cm from a jewellery store. If the platinum has a density of 21.45 g/cm^3, what is the mass of the medal in kg? [The volume of a cylinder is $V = \pi r^2 h$.]

1.92 Saline solution used in hospital contains 0.9% sodium chloride by mass. Calculate the number of grams of sodium chloride in 0.5 gal of saline solution if the solution has a density of 1.01 g/mL.

1.93 A 40-lb container of peat moss measures $14 \times 20 \times 30$ in. A 40-lb container of topsoil has a volume of 1.9 gal. (**a**) Calculate the average densities of peat moss and topsoil in units of g/cm^3. Would it be correct to say that peat moss is "lighter" than topsoil? (**b**) How many bags of peat moss are needed to cover an area measuring 15.0 ft $\times$ 20.0 ft to a depth of 3.0 in.?

1.94 A 10.0 g block of gold is hammered into a thin gold sheet which has an area of 150 cm^2. Given the density of gold is 19.3 g/cm^3, what is the approximate thickness of the gold sheet in millimeters?

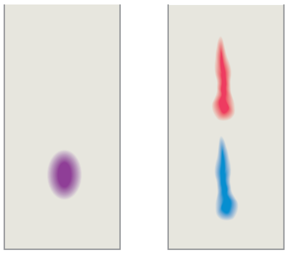
1.95 The total rate at which power is used by humans worldwide is approximately 15 TW (terawatts). The solar flux averaged over the sunlit half of Earth is 680 W/m^2 (assuming no clouds). The area of Earth's disc as seen from the Sun is 1.28×10^{14} m^2. The surface area of Earth is approximately 197,000,000 square miles. How much of Earth's surface would we need to cover with solar energy collectors to power the planet for use by all humans? Assume that the solar energy collectors can convert only 10% of the available sunlight into useful power.

1.96 In 2005, J. Robin Warren and Barry J. Marshall shared the Nobel Prize in Medicine for discovering the bacterium *Helicobacter pylori* and for establishing experimental proof that it plays a major role in gastritis and peptic ulcer disease. The story began when Warren, a pathologist, noticed that bacilli were associated with the tissues taken from patients suffering from ulcers. Look up the history of this case and describe Warren's first hypothesis. What sorts of evidence did it take to create a credible theory based on it?

1.97 A 30.0-cm-long cylindrical plastic tube, sealed at one end, is filled with acetic acid. The mass of acetic acid needed to fill the tube is found to be 89.24 g. The density of acetic acid is 1.05 g/mL. Calculate the inner diameter of the tube in centimeters.

1.98 Gold is alloyed (mixed) with other metals to increase its hardness in making jewelry. (**a**) Consider a piece of gold jewelry that weighs 9.85 g and has a volume of 0.675 cm^3. The jewelry contains only gold and silver, which have densities of 19.3 and 10.5 g/cm^3, respectively. If the total volume of the jewelry is the sum of the volumes of the gold and silver that it contains, calculate the percentage of gold (by mass) in the jewelry. (**b**) The relative amount of gold in an alloy is commonly expressed in units of carats. Pure gold is 24 carat, and the percentage of gold in an alloy is given as a percentage of this value. For example, an alloy that is 50% gold is 12 carat. State the purity of the gold jewelry in carats.

1.99 Paper chromatography is a simple but reliable method for separating a mixture into its constituent substances. You have a mixture of two vegetable dyes, one red and one blue, that you are trying to separate. You try two different chromatography procedures and achieve the separations shown in the figure. Which procedure worked better? Can you suggest a method to quantify how good or poor the separation was?



1.100 Judge the following statements as true or false. If you believe a statement to be false, provide a corrected version.

(**a**) Air and water are both elements.

(**b**) All mixtures contain at least one element and one compound.

(**c**) Compounds can be decomposed into two or more other substances; elements cannot.

(**d**) Elements can exist in any of the three states of matter.

(**e**) When yellow stains in a kitchen sink are treated with bleach water, the disappearance of the stains is due to a physical change.

(**f**) A hypothesis is more weakly supported by experimental evidence than a theory.

(**g**) The number 0.0033 has more significant figures than 0.033.

(**h**) Conversion factors used in converting units always have a numerical value of one.

(**i**) Compounds always contain at least two different elements.

1.101 You are assigned the task of separating a desired granular material with a density of 3.62 g/cm^3 from an undesired granular material that has a density of 2.04 g/cm^3. You want to do this by shaking the mixture in a liquid in which the heavier material will fall to the bottom and the lighter material will float. A solid will float on any liquid that is more dense. Using an Internet-based source or a handbook of chemistry, find the densities of the following substances: carbon tetrachloride, hexane, benzene, and diiodomethane. Which of these liquids will serve your purpose, assuming no chemical interaction takes place between the liquid and the solids?

1.102 In 2009, a team from Northwestern University and Western Washington University reported the preparation of a new "spongy" material composed of nickel, molybdenum, and sulfur that excels at removing mercury from water. The density of this new material is $0.20 \, g/cm^3$, and its surface area is $1242 \, m^2$ per gram of material. (**a**) Calculate the volume of a 10.0-mg sample of this material. (**b**) Calculate the surface area for a 10.0-mg sample of this material. (**c**) A 10.0-mL sample of contaminated water had 7.748 mg of mercury in it. After treatment with 10.0 mg of the new spongy material, 0.001 mg of mercury remained in the contaminated water. What percentage of the mercury was removed from the water? (**d**) What is the final mass of the spongy material after the exposure to mercury?

2

ATOMS, MOLECULES, AND IONS

WHAT'S AHEAD

2.1 | The Atomic Theory of Matter

Take a moment to appreciate the great variety of colors, textures, and other properties in the materials that surround you—the array of colors in a flower, the texture of the fabric in your clothes, the solubility of sugar in a cup of coffee or the transparency and beauty of a diamond. How can we explain the striking and seemingly infinite variety of properties of the materials that make up our world? What makes diamond transparent and hard whereas table salt is brittle and dissolves in water? Aluminum conducts electricity, but aluminum oxide does not. Paper burns in the presence of oxygen gas but not in the presence of nitrogen gas. What accounts for these differences? The answers to all such questions lie in the structures of atoms, which determine the physical and chemical properties of matter.

Although materials vary greatly in their properties, everything is formed from only about 100 different kinds of atoms. In a sense, these different atoms are like the 26 letters of the English alphabet that join in different combinations to form the immense number of words in our language. But what rules govern the ways in which atoms combine? How do the properties of a substance relate to the kinds of atoms it contains? Indeed, what is an atom like, and what makes the atoms of one element different from those of another?

A helicopter engine is composed of many smaller parts, just as any substance on Earth is composed of countless atoms and molecules to give it its unique characteristics.

In this chapter, we introduce the basic structure of atoms and discuss the formation of molecules and ions. This knowledge provides you with the foundation you need to understand the chapters that follow.

By the end of this section, you should be able to

- Understand Dalton's postulates.

Philosophers from the earliest times speculated about the nature of the fundamental "stuff" from which the world is made. Democritus (460–370 BCE) and other early Greek philosophers described the material world as made up of tiny indivisible particles that they called *atomos*, meaning "indivisible" or "uncuttable." Later, however, Plato and Aristotle formulated the notion that there can be no ultimately indivisible particles, and the "atomic" view of matter faded for many centuries during which Aristotelean philosophy dominated Western culture.

The notion of **atoms** reemerged in Europe during the seventeenth century. As chemists learned to measure the amounts of elements that reacted with one another to form new substances, the ground was laid for an atomic theory that linked the idea of elements with the idea of atoms. That theory came from the work of John Dalton during the period from 1803 to 1807. Dalton's atomic theory was based on four postulates (see **Figure 2.1**).

A good theory explains known facts; Dalton's theory explained several laws of chemical combination known at the time.

- The *law of constant composition*, based on postulate 4:

    In a given compound, the relative numbers and kinds of atoms are constant.

- The **law of conservation of mass,** based on postulate 3:

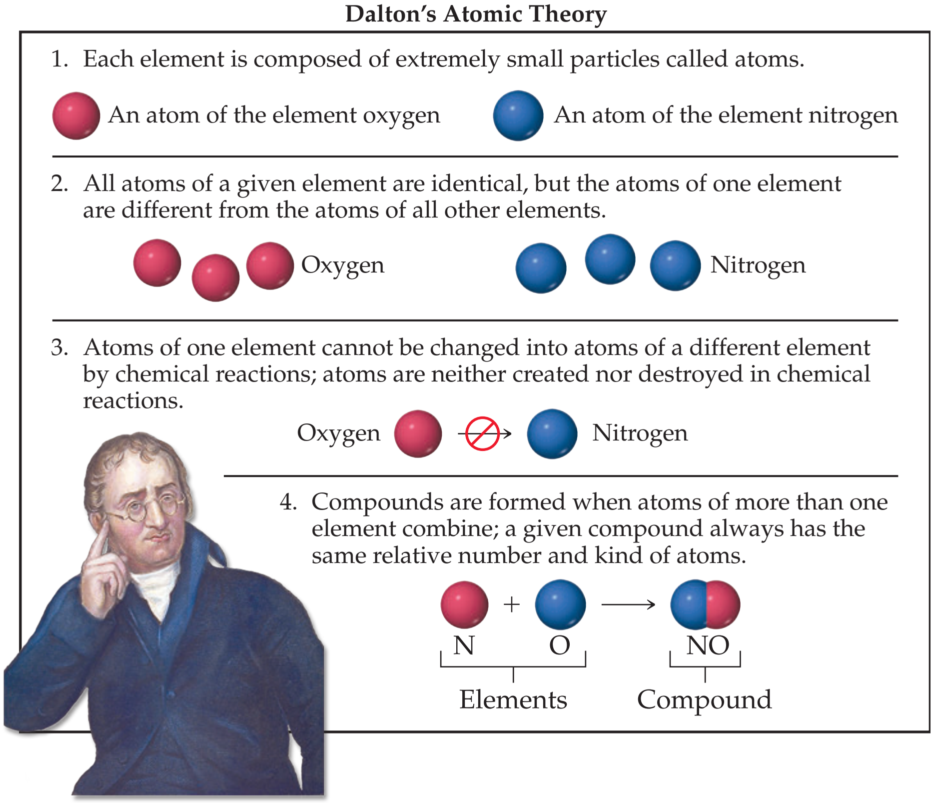
    The total mass of materials present after a chemical reaction is the same as the total mass present before the reaction.

A good theory also predicts new facts; Dalton used his theory to deduce

- The **law of multiple proportions:**

    If two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in the ratio of small whole numbers.

We can illustrate this law by considering water and hydrogen peroxide, both of which consist of the elements hydrogen and oxygen. In forming water, 8.0 g of oxygen combines with 1.0 g of hydrogen. In forming hydrogen peroxide, 16.0 g of oxygen combines with 1.0 g of hydrogen. Thus, the ratio of the masses of oxygen per gram of hydrogen in the two compounds is 2:1. Using Dalton's atomic theory, we conclude that hydrogen peroxide contains twice as many atoms of oxygen per hydrogen atom than does water.



◀ **Figure 2.1 Dalton's atomic theory.*** John Dalton (1766–1844), the son of a poor English weaver, began teaching at age 12. He spent most of his years in Manchester, where he taught both grammar school and college. His lifelong interest in meteorology led him to study gases, then chemistry, and eventually atomic theory. Despite his humble beginnings, Dalton gained a strong scientific reputation during his lifetime.

Self-Assessment Exercise

2.1 In an experiment, 7.0 g of nitrogen reacted with exactly 16.0 g of oxygen to form a single compound. What would be the total mass of the compound?

(a) 7.0 g

(b) 16 g

(c) 23 g

Exercises

2.2 A 1.0-g sample of carbon dioxide (CO_2) is fully decomposed into its elements, yielding 0.273 g of carbon and 0.727 g of oxygen. (**a**) What is the ratio of the mass of O to C? (**b**) If a sample of a different compound decomposes into 0.429 g of carbon and 0.571 g of oxygen, what is its ratio of the mass of O to C? (**c**) According to Dalton's atomic theory, what is the empirical formula of the second compound?

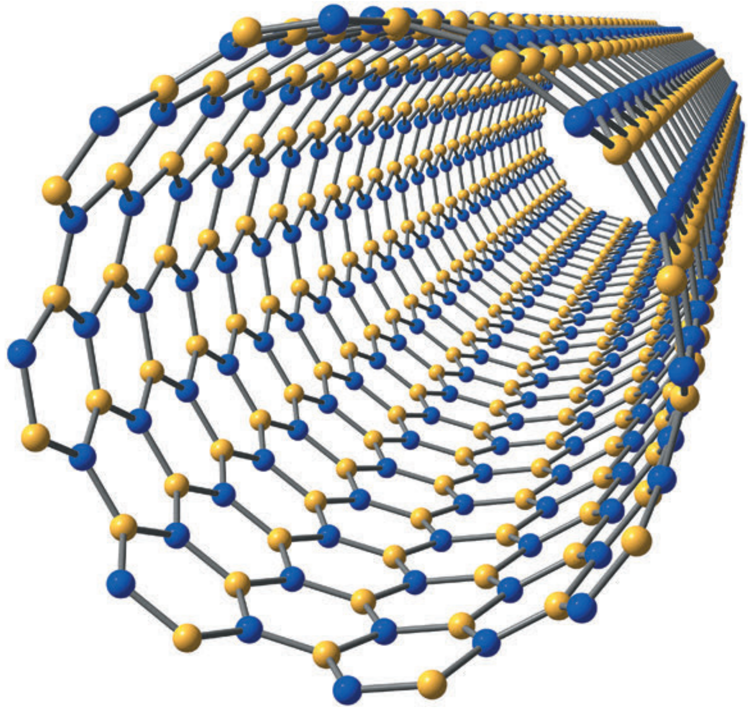
2.3 A chemist finds that 30.82 g of nitrogen will react with 17.60, 35.20, 70.40, or 88.00 g of oxygen to form four different compounds. (**a**) Calculate the mass of oxygen per gram of nitrogen in each compound. (**b**) How do the numbers in part (a) support Dalton's atomic theory?

Answers to Self-Assessment Exercises

2.1 (c)

*Dalton, "Atomic Theory" 1844.

2.2 | The Discovery of Atomic Structure



Dalton based his conclusions about atoms on chemical observations made in the laboratory. By assuming the existence of atoms, he was able to account for the laws of constant composition and of multiple proportions. But neither Dalton nor those who followed him during the century after his work was published had any direct evidence for the existence of atoms. Today, however, we can measure the properties of individual atoms and even provide images of them. The picture was obtained by a technique called scanning tunneling microscopy. The color was added to the image by computer to help distinguish its features. Each gold sphere is a silicon atom.

By the end of this section, you should be able to

- Describe the main experiments that led to the discovery of the electron and to the nuclear model of the atom

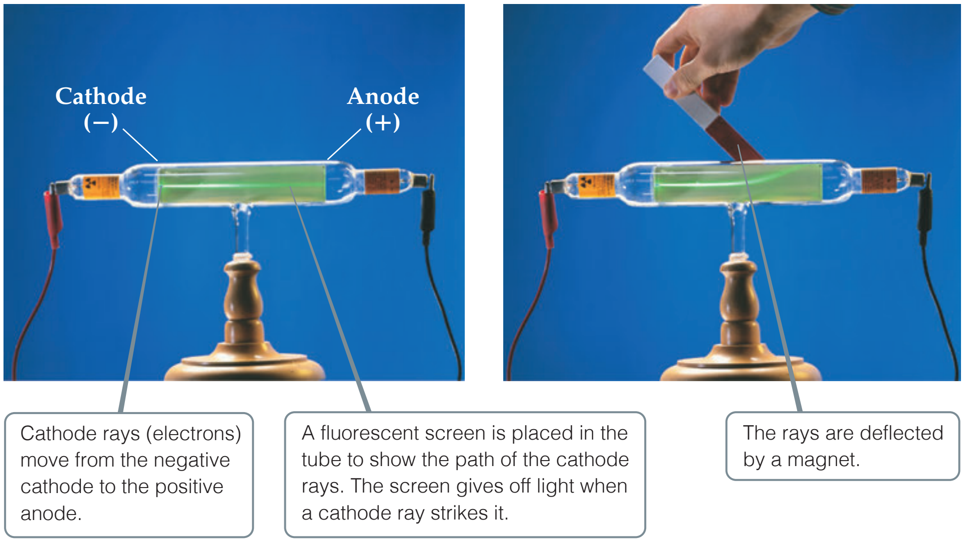
As scientists developed methods for probing the nature of matter, the supposedly indivisible atom began to show signs of a more complex structure, and today we know that the atom is composed of **subatomic particles**. Before we summarize the current model, we briefly consider a few of the landmark discoveries that led to that model. We will see that the atom is composed in part of electrically charged particles, some with a positive charge and some with a negative charge. As we discuss the development of our current model of the atom, keep in mind this fact: *Particles with the same charge repel one another, whereas particles with opposite charges attract one another*.

Cathode Rays and Electrons

During the mid-1800s, scientists began to study electrical discharge through a glass tube pumped almost empty of air (**Figure 2.2**). When a high voltage was applied to the electrodes in the tube, radiation was produced between the electrodes. This radiation, called **cathode rays**, originated at the negative electrode and traveled to the positive electrode. Although the rays could not be seen, their presence was detected because they cause certain materials to *fluoresce,* or to give off light.

Experiments showed that cathode rays are deflected by electric or magnetic fields in a way consistent with there being a stream of negative electrical charge. The British scientist J. J. Thomson (1856–1940) observed that cathode rays are the same regardless of the identity of the cathode material. In a paper published in 1897, Thomson described cathode rays as streams of negatively charged particles that we now call **electrons**.

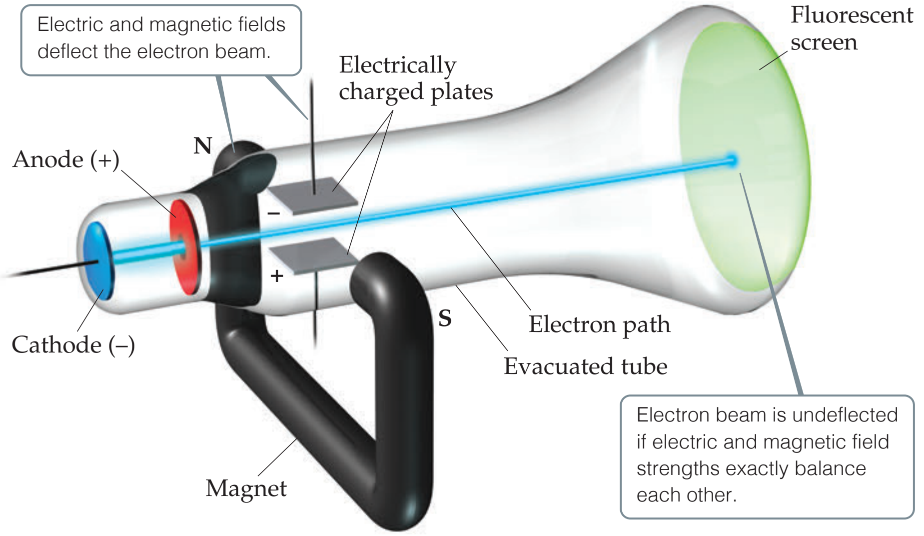
Go Figure If the fluorescent screen were removed from the tube, would cathode rays still be generated? Would you be able to see them?



▲ **Figure 2.2 Cathode-ray tube.**

Thomson constructed a cathode-ray tube having a hole in the anode through which the cathode rays could pass. Electrically charged plates and a magnet were positioned perpendicular to the beam, and a fluorescent screen that would give off light when struck with a cathode ray was located at one end (**Figure 2.3**). Because the electron is a negatively charged particle, the electric field deflected the rays in one direction, and the magnetic field deflected them in the opposite direction. Thomson adjusted the strengths of the fields so that the effects balanced each other, allowing the electrons to travel in a straight path to the screen. Knowing the strengths that resulted in the straight path made it possible to calculate a value of 1.76×10^8 coulombs* per gram for the ratio of the electron's electrical charge to its mass.

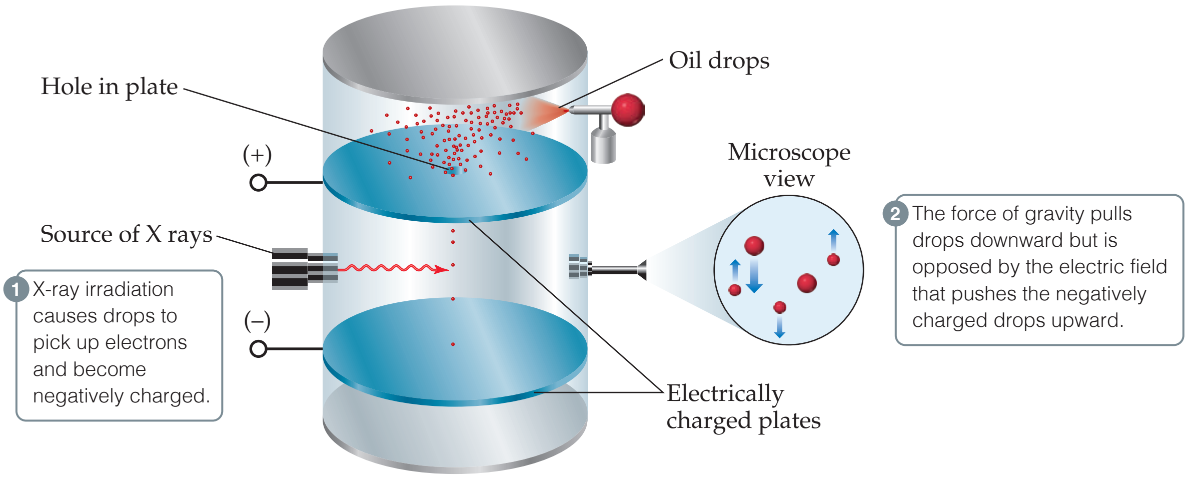
Go Figure If no magnetic field were applied, would you expect the electron beam to be deflected upward or downward by the electric field?



▲ **Figure 2.3 Cathode-ray tube with perpendicular magnetic and electric fields.** The cathode rays (electrons) originate at the cathode and are accelerated toward the anode, which has a hole in its center. A narrow beam of electrons passes through the hole and travels to the fluorescent screen that glows when struck by a cathode ray.

*The coulomb (C) is the SI unit for electrical charge.

Go Figure Are the masses of the oil drops changed significantly when electrons accumulate on them?



▲ **Figure 2.4 Millikan's oil-drop experiment to measure the charge of the electron.** Small drops of oil are allowed to fall between electrically charged plates. Millikan measured how varying the voltage between the plates affected the rate of fall. From these data he calculated the negative charge on the drops. Because the charge on any drop was always some integral multiple of 1.602×10^{-19} C, Millikan deduced this value to be the charge of a single electron.

Once the charge-to-mass ratio of the electron was known, measuring either quantity allowed scientists to calculate the other. In 1909, Robert Millikan (1868–1953) of the University of Chicago succeeded in measuring the charge of an electron by performing the experiment described in **Figure 2.4**. He then calculated the mass of the electron by using his experimental value for the charge, 1.602×10^{-19} C, and Thomson's charge-to-mass ratio, 1.76×10^{8} C/g:

$$\text{Electron mass} = \frac{1.602 \times 10^{-19}\,\cancel{C}}{1.76 \times 10^{8}\,\cancel{C}/g} = 9.10 \times 10^{-28}\,g$$

This result agrees well with the currently accepted value for the electron mass, 9.10938×10^{-28} g. This mass is about 2000 times smaller than that of hydrogen, the lightest atom.

▲ **Figure 2.5 Marie Sklodowska Curie (1867–1934).** In 1903, Henri Becquerel, Marie Curie, and her husband, Pierre, were jointly awarded the Nobel Prize in Physics for their pioneering work on radioactivity (a term she introduced). In 1911, Marie Curie won a second Nobel Prize, this time in chemistry for her discovery of the elements polonium and radium.

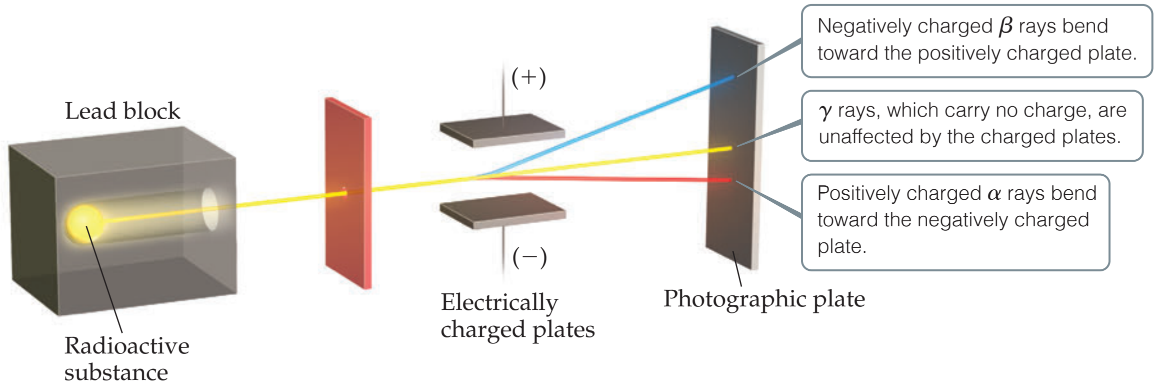
Radioactivity

In 1896 the French scientist Henri Becquerel (1852–1908) discovered that a compound of uranium spontaneously emits high-energy radiation. This spontaneous emission of radiation is called **radioactivity**. At Becquerel's suggestion, Marie Curie (**Figure 2.5**) and her husband, Pierre, began experiments to identify and isolate the source of radioactivity in the compound. They concluded that it was the uranium atoms.

Further study of radioactivity, principally by the British scientist Ernest Rutherford, revealed three types of radiation: alpha (α), beta (β), and gamma (γ). Rutherford (1871–1937) was a very important figure in this period of atomic science. After working at Cambridge University with J. J. Thomson, he moved to McGill University in Montreal, where he did research on radioactivity that led to his 1908 Nobel Prize in Chemistry. In 1907 he returned to England as a faculty member at Manchester University, where he did his famous α-particle scattering experiments, described further in this chapter.

Rutherford showed that the paths of α and β radiation are bent by an electric field, although in opposite directions; while γ radiation is unaffected by the field (**Figure 2.6**). From this finding he concluded that α and β rays consist of fast-moving electrically

Go Figure Which subatomic particle—proton, neutron, or electron—is equivalent to a β ray? β rays are deflected to a greater extent than α rays because (**a**) they are lighter, or (**b**) they are more highly charged.

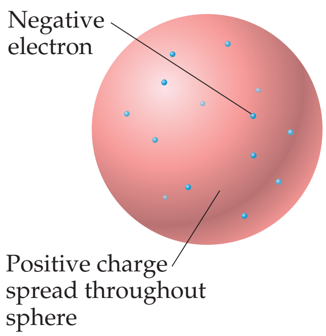


▲ **Figure 2.6 Behavior of alpha (α), beta (β), and gamma (γ) rays in an electric field.**

charged particles. In fact, β particles are nothing more than high-speed electrons that can be considered the radioactive equivalent of cathode rays. Because of their negative charge, they are attracted to a positively charged plate. The α particles have a positive charge and are attracted to a negative plate. In units of the charge of the electron, β particles have a charge of 1– and α particles a charge of 2+. Each α particle has a mass about 7400 times that of an electron. Gamma radiation is high-energy electromagnetic radiation similar to X rays; it does not consist of particles and it carries no charge.

The Nuclear Model of the Atom

With growing evidence that the atom is composed of smaller particles, scientists gave attention to how the particles fit together. During the early 1900s, Thomson reasoned that because electrons contribute only a very small fraction of an atom's mass, they probably are responsible for an equally small fraction of the atom's size. He proposed that the atom consists of a uniform positive sphere of matter in which the mass is evenly distributed and in which the electrons are embedded like raisins in a pudding or seeds in a watermelon (**Figure 2.7**). This *plum-pudding model,* named after a traditional English dessert, was very short-lived.



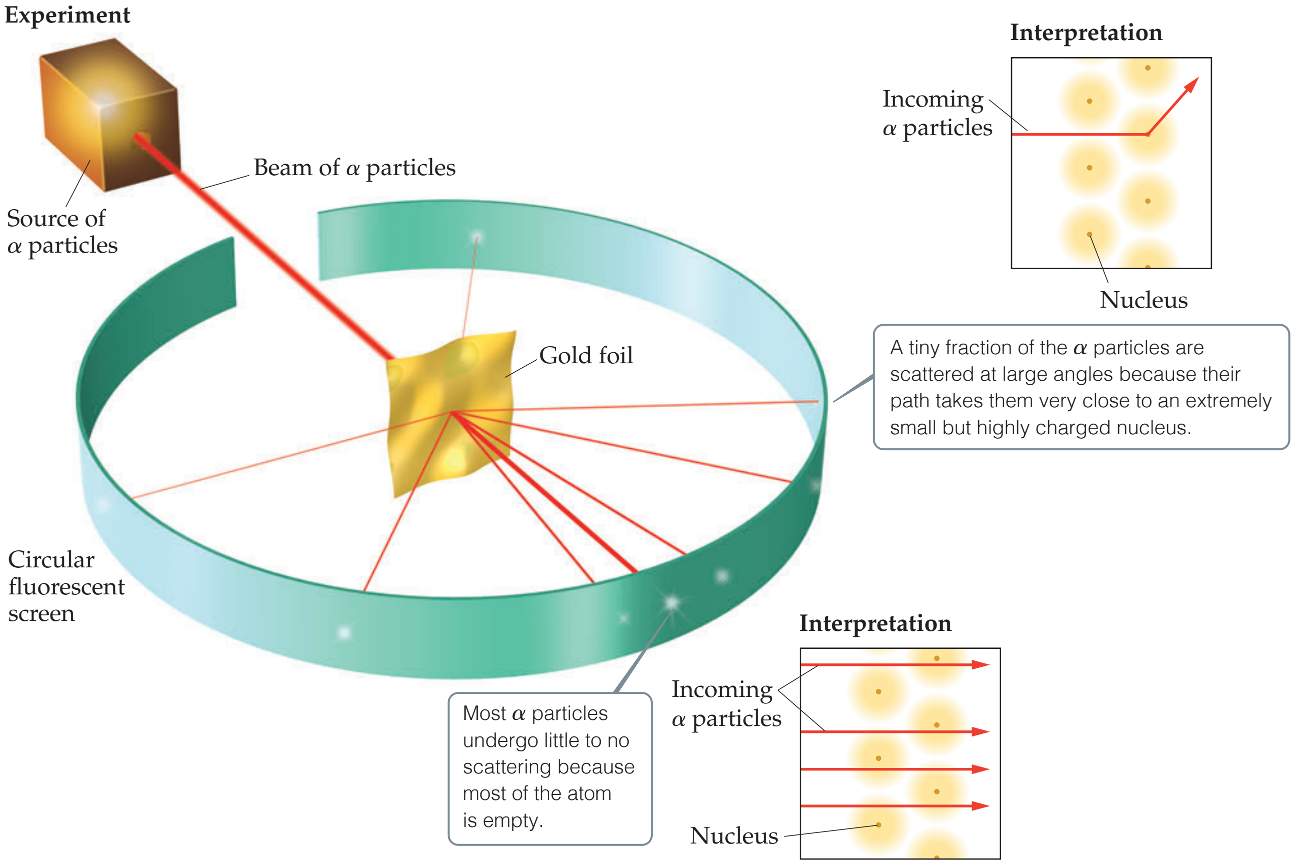
▲ **Figure 2.7 J. J. Thomson's plum-pudding model of the atom.** Ernest Rutherford and Ernest Marsden proved this model wrong.

In 1910, Rutherford was studying the angles at which α particles were deflected, or *scattered*, as they passed through a thin sheet of gold foil (**Figure 2.8**). He discovered that almost all the particles passed directly through the foil without deflection, with a few particles deflected about 1°, consistent with Thomson's plum-pudding model. For the sake of completeness, Rutherford suggested that Ernest Marsden (1889–1970), an undergraduate student working in the laboratory, look for scattering at large angles. To everyone's surprise, a small amount of scattering was observed at large angles, with some particles scattered back in the direction from which they had come. The explanation for these results was not immediately obvious, but they were clearly inconsistent with Thomson's plum-pudding model.

Rutherford explained the results by postulating the **nuclear model** of the atom, in which most of the mass of each gold atom and all of its positive charge reside in a very small, extremely dense region that he called the **nucleus**. He postulated further that most of the volume of an atom is empty space in which electrons move around the nucleus. In the α-scattering experiment, most of the particles passed through the foil unscattered because they did not encounter the minute nucleus of any gold atom. Occasionally, however, an α particle came close to a gold nucleus. In such encounters, the repulsion between the highly positive charge of the gold nucleus and the positive charge of the α particle was strong enough to deflect the particle, as shown in Figure 2.8.

Subsequent experiments led to the discovery of positive particles (**protons**) and neutral particles (**neutrons**) in the nucleus. Protons were discovered in 1919 by Rutherford and neutrons in 1932 by British scientist James Chadwick (1891–1972). Thus, the atom is composed of electrons, protons, and neutrons.

Go Figure What is the charge on the particles that form the beam? Will they be attracted to or repelled from the positively charged gold nuclei?



▲ **Figure 2.8 Rutherford's α-scattering experiment.** When α particles pass through a gold foil, most pass through undeflected but some are scattered, a few at very large angles. According to the plum-pudding model of the atom, the particles should experience only very minor deflections. The nuclear model of the atom explains why a few α particles are deflected at large angles. Although the nuclear atom has been depicted here as a yellow sphere, it is important to realize that most of the space around the nucleus contains only the low-mass electrons.

Self-Assessment Exercise

2.4 Which experiment enabled the mass of an electron to be calculated?

(a) J J Thompson's cathode ray tube experiment

(b) R Millikan's oil drop experiment

(c) E Rutherford's gold foil experiment

Exercises

2.5 An unknown particle is caused to move between two electrically charged plates, as illustrated in Figure 2.6. You hypothesize that the particle is a proton. (**a**) If your hypothesis is correct, would the particle be deflected in the same or opposite direction as the β rays? (**b**) Would it be deflected by a smaller or larger amount than the β rays?

2.6 What fraction of the α particles in Rutherford's gold foil experiment are scattered at large angles? Assume the gold foil is two layers thick, as shown in Figure 2.8, and that the approximate diameters of a gold atom and its nucleus are 270 pm and 1.0×10^{-2} pm, respectively. *Hint:* Calculate the cross sectional area occupied by the nucleus as a fraction of that occupied by the atom. Assume that the gold nuclei in each layer are offset from each other.

Answers to Self-Assessment Exercises

2.4 (b)

2.3 The Modern View of Atomic Structure

Ernest Rutherford (1871–1937), whom Einstein called 'the second Newton', was born and educated in New Zealand. In 1895, he was the first overseas student ever to be awarded a position at the Cavendish Laboratory at Cambridge University in England, where he worked with J J Thompson. In 1898, he joined the faculty of McGill University in Canada, where he did the research on radioactivity that led to his being awarded the 1908 Nobel Prize in Chemistry. In 1907, he moved back to England to join Manchester University where, in 1910, he performed his famous α-particle scattering experiments that led to the nuclear model of the atom. In 1992, his native New Zealand honored Rutherford by putting his likeness, along with his Nobel Prize medal, on its \$100 currency note. Element 104 is named in his honor.

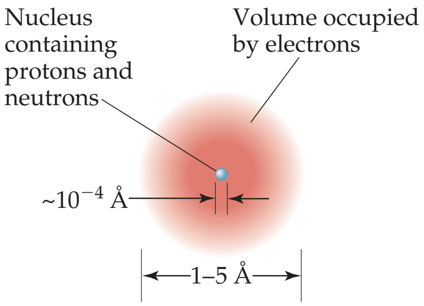
By the end of this section, you should be able to

- Describe the structure of the atom in terms of its subatomic particles
- Use symbols to indicate the composition of an isotope

Since Rutherford's time, as physicists have learned more and more about atomic nuclei, the list of particles that make up nuclei has grown and continues to increase. As chemists, however, we can take a simple view of the atom because only three subatomic particles—the proton, neutron, and electron—have a bearing on chemical behavior.

Go Figure

What is the approximate diameter of the nucleus in units of pm?



▲ **Figure 2.9 The structure of the atom.** A cloud of rapidly moving electrons occupies most of the volume of the atom. The nucleus occupies a tiny region at the center of the atom and is composed of the protons and neutrons. The nucleus contains virtually all the mass of the atom.

As noted earlier, the charge of an electron is -1.602×10^{-19} C. The charge of a proton is opposite in sign but equal in magnitude to that of an electron: $+1.602 \times 10^{-19}$ C. The quantity 1.602×10^{-19} C is called the **electronic charge**. For convenience, the charges of atomic and subatomic particles are usually expressed as multiples of this charge rather than as coulombs. Thus, the charge of an electron is 1− and that of a proton is 1+. Neutrons are electrically neutral (which is how they received their name). *Every atom has an equal number of electrons and protons, so atoms have no net electrical charge.*

Protons and neutrons reside in the tiny nucleus of the atom. The vast majority of an atom's volume is the space in which the electrons reside (Figure 2.9). Most atoms have diameters between 30 pm and 300 pm. The diameter of a chlorine atom, for example, is 175 pm.

Electrons are attracted to the protons in the nucleus by the electrostatic force that exists between particles of opposite electrical charge. In later chapters, we will see that the strength of the attractive forces between electrons and nuclei can be used to explain many of the differences among different elements.

Atoms have extremely small masses. The mass of the heaviest known atom, for example, is approximately 4×10^{-22} g. Because it would be cumbersome to express such small masses in grams, we use the **atomic mass unit** (u), where 1 u $= 1.66054 \times 10^{-24}$ g. A proton has a mass of 1.0073 u, a neutron 1.0087 u, and an electron 5.486×10^{-4} u (Table 2.1). Because it takes 1836 electrons to equal the mass of one proton, and 1839 electrons to equal the mass of a single neutron, the nucleus accounts for nearly the entire mass of an atom.

TABLE 2.1 Comparison of the Proton, Neutron, and Electron

Particle	Charge	Mass (u)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1−)	5.486×10^{-4}

The diameter of an atomic nucleus is approximately 1^{-10} fm, only a small fraction of the diameter of the atom as a whole. You can appreciate the relative sizes of the atom and its nucleus by imagining that if the hydrogen atom were as large as a football stadium, the nucleus would be the size of a small marble. Because the tiny nucleus carries most of the mass of the atom in such a small volume, it has an incredibly high density—on the order of 10^{13}–10^{14} g/cm^3. A matchbox full of material of such density would weigh over 2.5 billion tons!

Figure 2.9 incorporates the features we have just discussed. Electrons play the major role in chemical reactions. The significance of representing the region containing electrons as an indistinct cloud will become clear in later chapters when we consider the energies and spatial arrangements of the electrons. For now, however, we have all the information we need to discuss many topics that form the basis of everyday uses of chemistry.

Atomic Numbers, Mass Numbers, and Isotopes

What makes an atom of one element different from an atom of another element? The atoms of each element have a *characteristic number of protons*. The number of protons in an atom of any particular element is called that element's **atomic number**. Because an atom has no net electrical charge, the number of electrons it contains must equal the number of protons. All atoms of carbon, for example, have six protons and six electrons, whereas all atoms of oxygen have eight protons and eight electrons. Thus, carbon has atomic number 6, and oxygen has atomic number 8. The atomic number of each element is listed with the name and symbol of the element on the front inside cover of the text.

Atoms of a given element can differ in the number of neutrons they contain and, consequently, in mass. For example, while most atoms of carbon have six neutrons, some have more and some have less. The symbol $^{12}_{6}C$ (read "carbon twelve," carbon-12) represents the carbon atom containing six protons and six neutrons, whereas carbon atoms

Sample Exercise 2.1
Atomic Size

The diameter of a small coin is 17.9 mm, and the diameter of a silver atom is 288 pm. How many silver atoms could be arranged side by side across the diameter of the coin?

SOLUTION

The unknown is the number of silver (Ag) atoms. Using the relationship 1 Ag atom = 288 pm as a conversion factor relating number of atoms and distance, we start with the diameter of the coin, first converting this distance into picometers and then using the diameter of the Ag atom to convert distance to number of Ag atoms:

$$\text{Ag atoms} = (17.9\ \cancel{\text{mm}})\left(\frac{10^{-3}\ \cancel{\text{m}}}{1\ \cancel{\text{mm}}}\right)\left(\frac{1\ \text{pm}}{10^{-12}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{Ag atom}}{288\ \text{pm}}\right)$$

$$= 6.22 \times 10^{7}\ \text{Ag atoms}$$

That is, 62.2 million silver atoms could sit side by side across the coin!

Practice Exercise

The diameter of a carbon atom is 1.54 Å. (**a**) Express this diameter in picometers. (**b**) How many carbon atoms could be aligned side by side across the width of a pencil line that is 0.20 mm wide?

A CLOSER LOOK Basic Forces

Four basic forces are known in nature: (1) gravitational, (2) electromagnetic, (3) strong nuclear, and (4) weak nuclear. *Gravitational forces* are attractive forces that act between all objects in proportion to their masses. Gravitational forces between atoms or between subatomic particles are so small that they are of no chemical significance.

Electromagnetic forces are attractive or repulsive forces that act between either electrically charged or magnetic objects. The magnitude of the electric force between two charged particles is given by *Coulomb's law:* $F = kQ_1Q_2/d^2$, where Q_1 and Q_2 are the magnitudes of the charges on the two particles, d is the distance between their centers, and k is a constant determined by the units for Q and d. A negative value for the force indicates attraction, whereas a positive value indicates repulsion. Electric forces are of primary importance in determining the chemical properties of elements.

All nuclei except those of hydrogen atoms contain two or more protons. Because like charges repel, electrical repulsion would cause the protons to fly apart if the *strong nuclear force* did not keep them together. As the name implies, this force can be quite strong but only when particles are extremely close together, as are the protons and neutrons in a nucleus. At this distance, the attractive strong nuclear force is stronger than the positive–positive repulsive electric force and holds the nucleus together.

The *weak nuclear force* is weaker than the electric force and the strong nuclear force but stronger than the gravitational force. We are aware of its existence only because it shows itself in certain types of radioactivity.

Related Exercise: 2.124

that contain six protons and eight neutrons have mass number 14, are represented as ${}^{14}_{6}\text{C}$ and are referred to as carbon-14.

The atomic number is indicated by the subscript; the superscript, called the **mass number**, is the number of protons plus neutrons in the atom:

Mass number (number of protons plus neutrons) →

$${}^{12}_{6}\text{C} \longleftarrow \text{Symbol of element}$$

Atomic number (number of protons or electrons) →

Because all atoms of a given element have the same atomic number, the subscript is redundant and is often omitted. Thus, the symbol for carbon-12 can be represented simply as ${}^{12}\text{C}$.

Atoms with identical atomic numbers but different mass numbers (that is, the same number of protons but different numbers of neutrons) are called **isotopes** of one another. Several isotopes of carbon are listed in **Table 2.2**. We will generally use the notation with superscripts only when referring to a particular isotope of an element. It is important to keep in mind that the isotopes of any given element are all alike chemically. A carbon dioxide molecule that contains a ${}^{13}\text{C}$ atom behaves for all practical purposes identically to one that contains a ${}^{12}\text{C}$ atom.

TABLE 2.2 Some Isotopes of Carbon[a]

Symbol	Number of Protons	Number of Electrons	Number of Neutrons
^{11}C	6	6	5
^{12}C	6	6	6
^{13}C	6	6	7
^{14}C	6	6	8

[a] Almost 99% of the carbon found in nature is ^{12}C.

Sample Exercise 2.2
Determining the Number of Subatomic Particles in Atoms

How many protons, neutrons, and electrons are in an atom of (**a**) ^{197}Au, (**b**) strontium-90?

SOLUTION

(**a**) The superscript 197 is the mass number (protons + neutrons). According to the list of elements given on the front inside cover, gold has atomic number 79. Consequently, an atom of ^{197}Au has 79 protons, 79 electrons, and 197 − 79 = 118 neutrons. (**b**) The atomic number of strontium is 38. Thus, all atoms of this element have 38 protons and 38 electrons. The strontium-90 isotope has 90 − 38 =52 neutrons.

Practice Exercise

Which of these atoms has the largest number of neutrons?
(**a**) ^{148}Eu (**b**) ^{157}Dy (**c**) ^{149}Nd (**d**) ^{162}Ho (**e**) ^{159}Gd

Sample Exercise 2.3
Writing Symbols for Atoms

Magnesium has three isotopes with mass numbers 24, 25, and 26. (**a**) Write the complete chemical symbol (superscript and subscript) for each. (**b**) How many neutrons are in an atom of each isotope?

SOLUTION

(**a**) Magnesium has atomic number 12, so all atoms of magnesium contain 12 protons and 12 electrons. The three isotopes are therefore represented by $^{24}_{12}Mg$, $^{25}_{12}Mg$, and $^{26}_{12}Mg$. (**b**) The number of neutrons in each isotope is the mass number minus the number of protons. The numbers of neutrons in an atom of each isotope are therefore 12, 13, and 14, respectively.

Practice Exercise

Give the complete chemical symbol for the atom that contains 82 protons, 82 electrons, and 126 neutrons.

Self-Assessment Exercises

2.7 What subatomic particle(s) account for the vast majority of the mass of an atom of helium?

(a) Electrons

(b) Protons

(c) Neutrons

(d) Electrons and protons

(e) Protons and neutrons

2.8 What is the symbol for the isotope of oxygen-18?

(a) $^{10}_{8}O$

(b) $^{18}_{8}O$

(c) $^{10}_{18}O$

Exercises

2.9 The radius of an atom of tungsten (W) is about 2.10 units. (**a**) Express this distance in nanometers (nm) and in picometers (pm). (**b**) How many tungsten atoms would have to be lined up to create a wire of 2.0 mm? (**c**) If the atom is assumed to be a sphere, what is the volume in m^3 of a single W atom?

2.10 Answer the following questions without referring to Table 2.1: (**a**) What are the main subatomic particles that make up the atom? (**b**) What is the relative charge (in multiples of the electronic charge) of each of the particles? (**c**) Which of the particles is the most massive? (**d**) Which is the least massive?

2.11 Consider an atom of ^{32}P. (**a**) How many protons, neutrons, and electrons does this atom contain? (**b**) What is the symbol of the atom obtained by adding one proton to ^{32}P? (**c**) What is the symbol of the atom obtained by adding one neutron to ^{32}P? (**d**) Are either of the atoms obtained in parts (b) and (c) isotopes of ^{32}P? If so which one?

2.12 (**a**) Define atomic number and mass number. (**b**) Which of these can vary without changing the identity of the element?

2.13 How many protons, neutrons, and electrons are in the following atoms? (**a**) ^{84}Kr, (**b**) ^{200}Hg, (**c**) ^{59}Co, (**d**) ^{55}Mn, (**e**) ^{239}U, (**f**) ^{181}Ta.

2.14 Fill in the gaps in the following table, assuming each column represents a neutral atom.

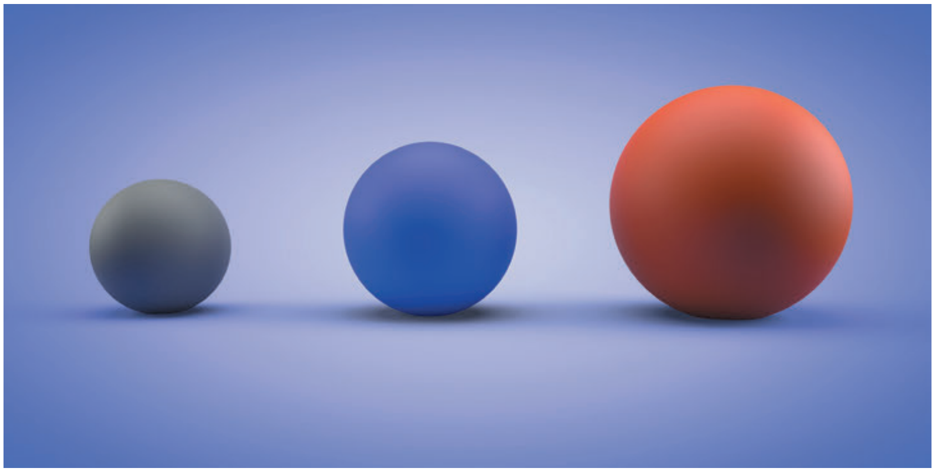
Symbol	^{159}Tb				
Protons		29			37
Neutrons		34	53		
Electrons			42	34	
Mass no.				79	85

2.15 Write the correct symbol, with both superscript and subscript, for each of the following. Use the list of elements in the front inside cover as needed: (**a**) the isotope of hafnium that contains 106 neutrons, (**b**) the isotope of mercury with mass number 201, (**c**) the isotope of rhenium with mass number 187, (**d**) the isotope of calcium that has an equal number of protons and neutrons.

Answers to Self-Assessment Exercises

2.7 (e) **2.8** (b)

2.4 | Atomic Weights



We are quite used to objects of the same shape but of different sizes—a golf ball, a tennis ball, and a soccer ball for example. Atoms of different elements also have a different size and weight. By the end of this section, you should be able to

- Calculate the atomic weight of an element from its isotopic composition.

Atoms are small pieces of matter, so they have mass. In this section, we discuss the mass scale used for atoms and introduce the concept of *atomic weights*.

The Atomic Mass Scale

Scientists of the nineteenth century were aware that atoms of different elements have different masses. They found, for example, that each 100.0 g of water contains 11.1 g of hydrogen and 88.9 g of oxygen. Thus, water contains $88.9/11.1 = 8$ times as much oxygen, by mass, as hydrogen. Once scientists understood that water contains two hydrogen atoms for each oxygen atom, they concluded that an oxygen atom must have $2 \times 8 = 16$ times as much mass as a hydrogen atom. Hydrogen, the lightest atom, was arbitrarily assigned a relative mass of 1 (no units). Atomic masses of other elements were at first determined relative to this value. Thus, oxygen was assigned an atomic mass of 16.

Today we can determine the masses of individual atoms with a high degree of accuracy. For example, we know that the ^{1}H atom has a mass of 1.6735×10^{-24} g and the ^{16}O atom has a mass of 2.6560×10^{-23} g. As we noted in Section 2.3, it is convenient to use the **atomic mass unit** when dealing with these extremely small masses:

$$1\ \text{u} = 1.66054 \times 10^{-24}\ \text{g} \quad \text{and} \quad 1\ \text{g} = 6.02214 \times 10^{23}\ \text{u}$$

The atomic mass unit is presently defined by assigning a mass of exactly 12 u to a chemically unbound atom of the ^{12}C isotope of carbon. In these units, an ^{1}H atom has a mass of 1.0078 u and an ^{16}O atom has a mass of 15.9949 u.

Atomic Weight

Most elements occur in nature as mixtures of isotopes. We can determine the *average atomic mass* of an element, usually called the element's **atomic weight**, by summing (indicated by the Greek sigma, Σ) over the masses of its isotopes multiplied by their relative abundances:

$$\text{Atomic weight} = \sum_{\substack{\text{over all}\\ \text{isotopes of}\\ \text{the element}}} [(\text{isotope mass}) \times (\text{fractional isotope abundance})] \qquad [2.1]$$

Naturally occurring carbon, for example, is composed of 98.93% ^{12}C and 1.07% ^{13}C. The masses of these isotopes are 12 u (exactly) and 13.00335 u, respectively, making the atomic weight of carbon

$$(0.9893)(12\ \text{u}) + (0.0107)(13.00335\ \text{u}) = 12.01\ \text{u}$$

The atomic weights of the elements are listed in both the periodic table and the table of elements on the front inside cover of this text.

Sample Exercise 2.4
Calculating the Atomic Weight of an Element from Isotopic Abundances

Naturally occurring chlorine is 75.78% ^{35}Cl (atomic mass 34.969 u) and 24.22% ^{37}Cl (atomic mass 36.966 u). Calculate the atomic weight of chlorine.

SOLUTION

We can calculate the atomic weight by multiplying the abundance of each isotope by its mass and summing these products. Because 75.78% = 0.7578 and 24.22% = 0.2422, we have

$$\begin{aligned}\text{Atomic weight} &= (0.7578)(34.969\ \text{u}) + (0.2422)(36.966\ \text{u})\\ &= 26.50\ \text{u} + 8.953\ \text{u}\\ &= 35.45\ \text{u}\end{aligned}$$

This answer makes sense: The atomic weight, which is actually the average atomic mass, is between the masses of the two isotopes and is closer to the value of ^{35}Cl, the more abundant isotope.

Practice Exercise

Three isotopes of silicon occur in nature: ^{28}Si (92.23%), atomic mass 27.97693 u; ^{29}Si (4.68%), atomic mass 28.97649 u; and ^{30}Si (3.09%), atomic mass 29.97377 u. Calculate the atomic weight of silicon.

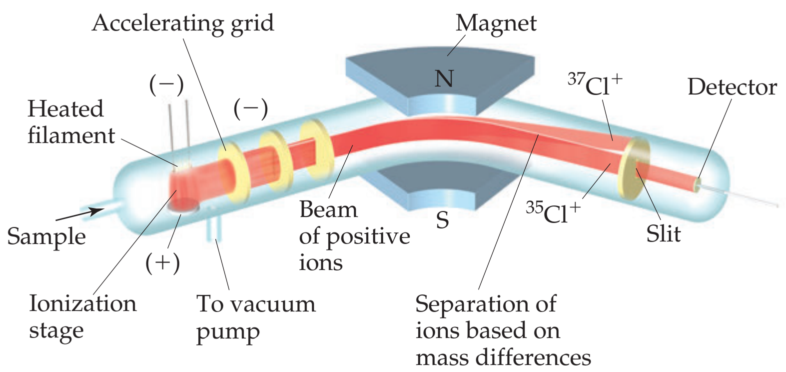
A CLOSER LOOK The Mass Spectrometer

The most accurate means for determining atomic weights is provided by the **mass spectrometer** (Figure 2.10). There are various designs of mass spectrometers, but they all operate on similar principles. The first step is to get atoms or molecules into the gas phase. Sometimes the sample to be analyzed is already a gas, whereas in other cases heating, application of an electric field, or a pulse of laser light may be needed to create gas-phase atoms or molecules. Next, the gas-phase species must be converted to positively charged particles called *ions*. There are many approaches to creating ions, including bombardment with beams of high-energy electrons or chemical reactions with other gas-phase molecules. Once gas-phase ions have been produced, they are accelerated toward a negatively charged grid. After the ions pass through the grid, they encounter two slits that allow only a narrow beam of ions to pass. This beam then passes between the poles of a magnet, which deflects the ions into a curved path. For ions with the same charge, the extent of deflection depends on mass—the more massive the ion, the less the deflection. The ions are thereby separated according to their masses. By changing the strength of the magnetic field or the accelerating voltage on the grid, ions of various masses can be selected to enter the detector.

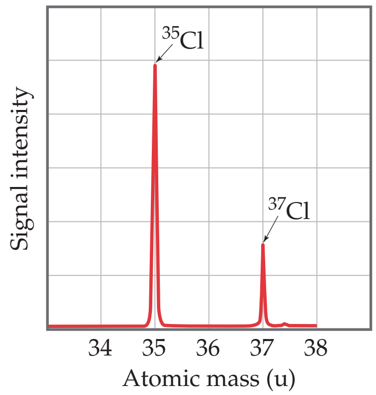
A graph of the intensity of the detector signal versus ion atomic mass is called a *mass spectrum* (Figure 2.11). Analysis of a mass spectrum gives both the masses of the ions reaching the detector and their relative abundances, which are obtained from the signal intensities. Knowing the atomic mass and the abundance of each isotope allows us to calculate the atomic weight of an element, as shown in Sample Exercise 2.4.

Mass spectrometers are used extensively today to identify chemical compounds and analyze mixtures of substances. Any molecule that loses electrons can fall apart, forming an array of positively charged fragments. The mass spectrometer measures the masses of these fragments, producing a chemical "fingerprint" of the molecule and providing clues about how the atoms were connected in the original molecule. Thus, a chemist might use this technique to determine the molecular structure of a newly synthesized compound, to analyze proteins in the human genome, or to identify a pollutant in the environment.

Related Exercises: 2.19, 2.73, 2.74, 2.99, 2.109, 2.110



▲ **Figure 2.10 A mass spectrometer.** Cl atoms are first ionized to form Cl^+ ions, accelerated with an electric field, and finally their path is directed by a magnetic field. The paths of the ions of the two Cl isotopes diverge as they pass through the field.



▲ **Figure 2.11 Mass spectrum of atomic chlorine.** The fractional abundances of the isotopes ^{35}Cl and ^{37}Cl are indicated by the relative signal intensities of the beams reaching the detector of the mass spectrometer.

Self-Assessment Exercise

2.16 Bromine has two main isotopes of similar abundance: ^{79}Br and ^{81}Br. Estimate (without using a calculator) the atomic weight of bromine.

(a) ~79 u
(b) ~80 u
(c) ~81 u

Exercises

2.17 (**a**) What isotope is used as the standard in establishing the atomic mass scale? (**b**) The atomic weight of boron is reported as 10.81, yet no atom of boron has the mass of 10.81 u. Explain.

2.18 Iron has three major isotopes: ^{54}Fe (atomic mass = 53.9396 u; abundance 5.85%), ^{56}Fe (atomic mass = 55.9349 u; abundance 91.75%), and ^{57}Fe (atomic mass = 56.9354 u; abundance 2.12 %). Calculate the atomic weight of iron.

2.19 (**a**) Thomson's cathode-ray tube (Figure 2.3) and the mass spectrometer (Figure 2.10) both involve the use of electric or magnetic fields to deflect charged particles. What are the charged particles involved in each of these experiments? (**b**) What are the labels on the axes of a mass spectrum? (**c**) To measure the mass spectrum of an atom, the atom must first lose one or more electrons. Which would you expect to be deflected more by the same setting of the electric and magnetic fields, a Cl^+ or a Cl^{2+} ion?

2.20 Naturally occurring lead has the following isotopic abundances:

Isotope	Abundance (%)	Atomic mass (u)
^{204}Pb	1.4	203.9730
^{206}Pb	24.1	205.9744
^{207}Pb	22.1	206.9759
^{208}Pb	52.4	207.9766

(**a**) What is the average atomic mass of Pb? (**b**) Sketch the mass spectrum of Pb.

Answers to Self-Assessment Exercises

2.16 (b)

2.5 The Periodic Table

The shininess of metals is an obvious similarity that they share, but many elements show pattens in their reactivity that are not so evident from their physical appearance. One triumph of classification was the ordering of the elements into the Periodic Table. It enables us to understand trends in the properties of the elements, make predictions about their chemical reactivity and even suggest the characteristics of elements that are yet to be discovered.

By the end of this section, you should be able to

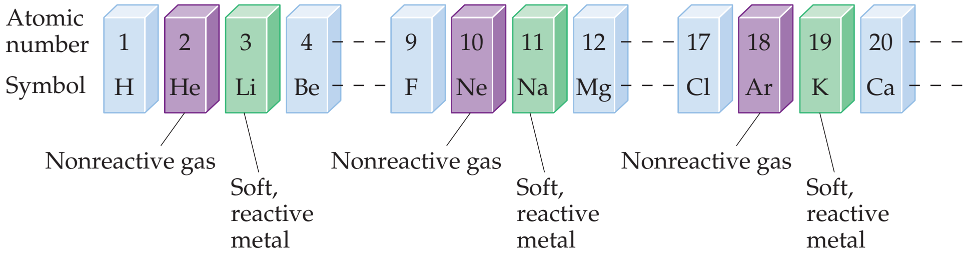
- Describe the general features of the periodic table

As the list of known elements expanded during the early 1800s, attempts were made to find patterns in chemical behavior. These efforts culminated in the development of the periodic table in 1869. We will have much to say about the periodic table in later chapters, but it is so important and useful that you should become acquainted with it now. You will quickly learn that *the periodic table is the most significant tool that chemists use for organizing and remembering chemical facts*.

Many elements show strong similarities to one another. The elements lithium (Li), sodium (Na), and potassium (K) are all soft, very reactive metals, for example. The elements helium (He), neon (Ne), and argon (Ar) are all nonreactive gases. If the elements are arranged in order of increasing atomic number, their chemical and physical properties show a repeating, or *periodic*, pattern. For example, each of the soft, reactive metals—lithium, sodium, and potassium—comes immediately after one of the nonreactive gases—helium, neon, and argon, respectively—as shown in **Figure 2.12**.

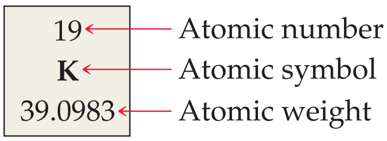
Go Figure

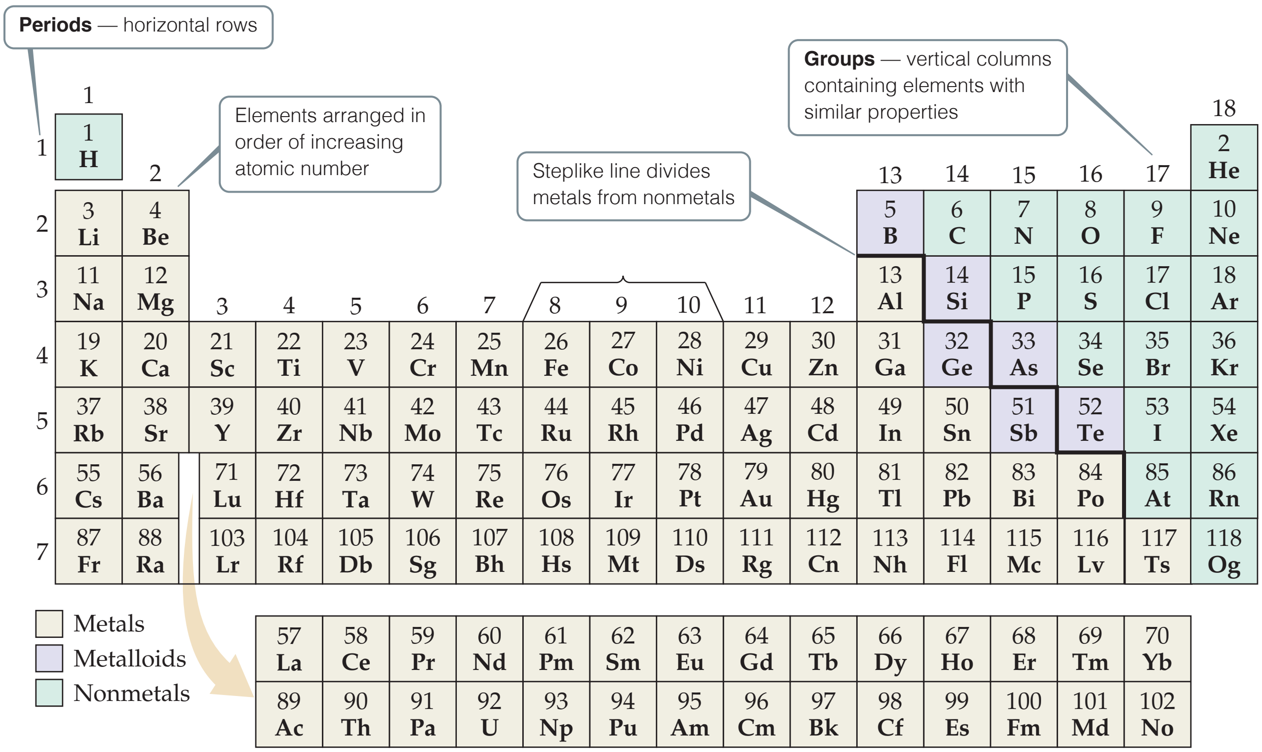
If F is a reactive nonmetal, which other element or elements shown here are likely to be reactive nonmetals?



▲ **Figure 2.12** **Arranging elements by atomic number reveals a periodic pattern of properties.** This pattern is the basis of the periodic table.

The arrangement of elements in order of increasing atomic number, with elements having similar properties placed in vertical columns, is known as the **periodic table** (Figure 2.13). The table shows the atomic number and atomic symbol for each element, and the atomic weight is often given as well, as in this typical entry for potassium:





▲ **Figure 2.13** **Periodic table of elements.**

You might notice slight variations in periodic tables from one book to another or between those in the lecture hall and in the text. These are simply matters of style, or they might concern the particular information included. There are no fundamental differences.

The horizontal rows of the periodic table are called **periods**. The first period consists of only two elements, hydrogen (H) and helium (He). The second and third periods consist of eight elements each. The fourth and fifth periods contain 18 elements. The sixth and seventh periods have 32 elements each, but, in order to fit on a page, 14 of the elements from each period (atomic numbers 57–70 and 89–102) appear at the bottom of the table.

The vertical columns are **groups**. These are numbered from 1 to 18, as shown in Figure 2.13.

We will use the IUPAC's proposed convention and number the groups from 1 through 18.

Elements in a group often exhibit similarities in physical and chemical properties. For example, the "coinage metals"—copper (Cu), silver (Ag), and gold (Au)—belong to Group 11. These elements are less reactive than most metals, which is why they have been traditionally used throughout the world to make coins. Many other groups in the periodic table also have names, listed in Table 2.3.

We will learn in Chapters 6 and 7 that elements in a group have similar properties because they have the same arrangement of electrons at the periphery of their atoms. However, we need not wait until then to make good use of the periodic table; after all, the chemists who developed the table knew nothing about electrons! We can use the table, as they intended, to correlate behaviors of elements and to help us remember many facts.

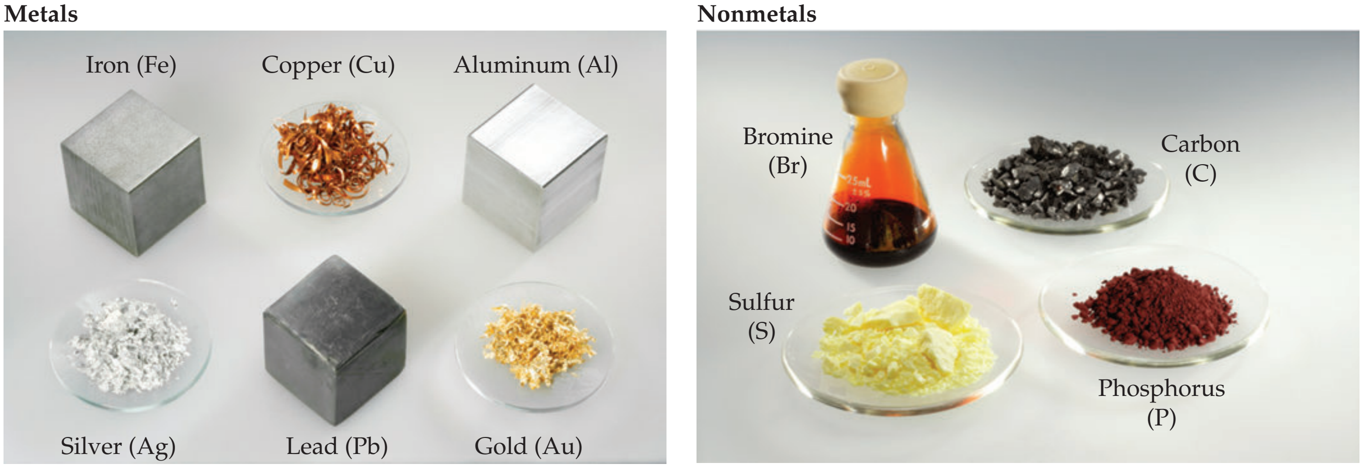
The color code of Figure 2.13 shows that, except for hydrogen, all the elements on the left and in the middle of the table are **metallic elements**, or **metals**. All the metallic elements share characteristic properties, such as luster and high electrical and heat conductivity, and all of them except mercury (Hg) are solid at room temperature.* The metals are separated from the **nonmetallic elements**, or **nonmetals**, by a stepped line that runs from boron (B) to astatine (At). (Note that hydrogen, although on the left side of the table, is a nonmetal.) At room temperature and pressure, some of the nonmetals are gaseous, some are solid, and one is liquid. Nonmetals generally differ from metals in appearance (Figure 2.14) and in other physical properties. Many of the elements that lie along the line that separates metals from nonmetals have properties that fall between those of metals and nonmetals. These elements are often referred to as **metalloids**.

TABLE 2.3 Names of Some Groups in the Periodic Table

Group	Name	Elements
1	Alkali metals	Li, Na, K, Rb, Cs, Fr
2	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
16	Chalcogens	O, S, Se, Te, Po
17	Halogens	F, Cl, Br, I, At
18	Noble gases	He, Ne, Ar, Kr, Xe, Rn

*All metals become liquids if heated sufficiently. Hg simply has the lowest melting point of any metallic element. Although sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and gallium (Ga) are solids at room temperature, they all melt at temperatures below 100 °C.

Go Figure Name two ways in which the metals shown here differ in general appearance from the nonmetals.



▲ **Figure 2.14** Examples of metals and nonmetals.

Sample Exercise 2.5
Using the Periodic Table

Which two of these elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?

SOLUTION

Elements in the same group of the periodic table are most likely to exhibit similar properties. We therefore expect Ca and Mg to be most alike because they are in the same group (2, the alkaline earth metals).

Practice Exercise

Locate Na (sodium) and Br (bromine) in the periodic table. Give the atomic number of each and classify each as metal, metalloid, or nonmetal.

Self-Assessment Exercise

2.21 The halogens are a group of elements, two of which are gases, one of which is a liquid, and two solid at room temperature. In spite of these obvious physical differences, there are many chemical similarities between the elements. What are the elements that are classed as halogens?

(a) Li, Be, B, C, N, O, F, Ne
(b) Be, Mg, Ca, Sr, Ba, Ra
(c) F, Cl, Br, I, At

Exercises

2.22 For each of the following elements, write its chemical symbol, locate it in the periodic table, give its atomic number, and indicate whether it is a metal, metalloid, or nonmetal: (**a**) radon, (**b**) tellurium, (**c**) cadmium, (**d**) chromium, (**e**) barium, (**f**) selenium, (**g**) sulphur.

2.23 The elements of Group 14 show an interesting change in properties moving down the group. Give the name and chemical symbol of each element in the group and label it as a nonmetal, metalloid, or metal.

Answers to Self-Assessment Exercises

2.21 (c)

2.6 | Molecules and Molecular Compounds

DNA, responsible for carrying the genetic instructions of all organisms, is one of the most complex molecules, yet it is composed of just five elements: C, H, O, N, and P. Knowing how the elements join together is fundamental to our understanding of chemistry. By the end of this section, you should be able to

- Recognize the different types of formula used to describe molecules

Even though the atom is the smallest representative sample of an element, only the noble-gas elements are normally found in nature as isolated atoms. Most matter is composed of molecules or ions. We examine molecules here and ions in Section 2.7.

Molecules and Chemical Formulas

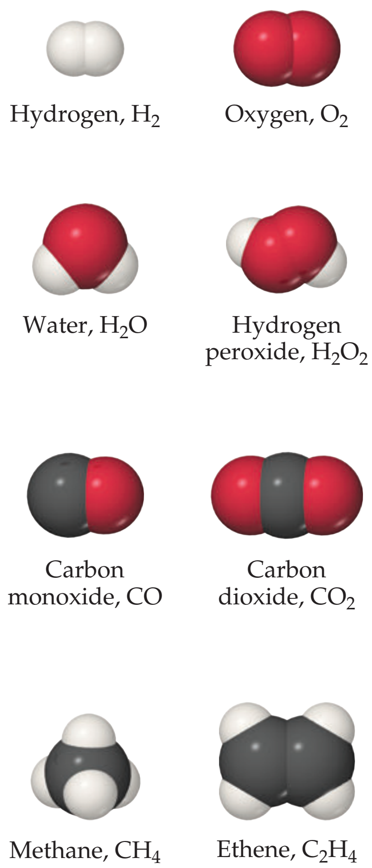
Several elements are found in nature in molecular form—two or more of the same type of atom bound together. For example, most of the oxygen in air consists of molecules that contain two oxygen atoms. As we saw in Section 1.2, we represent this molecular oxygen by the **chemical formula** O_2 (read "oh two"). The subscript tells us that two oxygen atoms are present in each molecule. A molecule made up of two atoms is called a **diatomic molecule**.

Oxygen also exists in another molecular form known as *ozone*. Molecules of ozone consist of three oxygen atoms, making the chemical formula O_3. Even though "normal" oxygen (O_2) and ozone (O_3) are both composed only of oxygen atoms, they exhibit very different chemical and physical properties. For example, O_2 is essential for life, but O_3 is toxic; O_2 is odorless, whereas O_3 has a sharp, pungent smell.

The elements that normally occur as diatomic molecules are hydrogen, oxygen, nitrogen, and the halogens (H_2, O_2, N_2, F_2, Cl_2, Br_2, and I_2). Except for hydrogen, these diatomic elements are clustered on the right side of the periodic table.

Compounds composed of molecules contain more than one type of atom and are called **molecular compounds**. A molecule of the compound methane, for example, consists of one carbon atom and four hydrogen atoms and is therefore represented by the chemical formula CH_4. Lack of a subscript on the C indicates one atom of C per methane molecule. Several common molecules of both elements and compounds are shown in **Figure 2.15**. Notice how the composition of each substance is given by its chemical formula. Notice also that these substances are composed only of nonmetallic elements.

Most of the molecular substances we will encounter contain only nonmetals.



▲ **Figure 2.15 Molecular models.** Notice how the chemical formulas of these simple molecules correspond to their compositions.

Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers of atoms in a molecule are called **molecular formulas.** (The formulas in Figure 2.17 are molecular formulas.) Chemical formulas that give only the relative number of atoms of each type in a molecule are called **empirical formulas**. The subscripts in an empirical formula are always the smallest possible whole-number ratios. The molecular formula for hydrogen peroxide is H_2O_2, for example, whereas its empirical formula is HO. The molecular formula for ethene is C_2H_4, and its empirical formula is CH_2. For many substances, the molecular formula and the empirical formula are identical, as in the case of water, H_2O.

Whenever we know the molecular formula of a compound, we can determine its empirical formula. The converse is not true, however. If we know the empirical formula of a substance, we cannot determine its molecular formula unless we have more information. So why do chemists bother with empirical formulas? As we will see in Chapter 3, certain common methods of analyzing substances lead to the empirical formula only. For example, if you decomposed hydrogen peroxide H_2O_2 into its elements and weighed them, you could determine that there were equal numbers of hydrogen and oxygen atoms, but you would not know if the molecular formula was HO, H_2O_2, H_3O_3, or the like. Once the empirical formula is known, additional experiments can give the information needed to convert the empirical formula to the molecular one. In addition, there are many substances that do not exist as isolated molecules, one example being ionic compounds that are discussed later in this chapter. For these substances, we must rely on empirical formulas.

Sample Exercise 2.6
Relating Empirical and Molecular Formulas

Write the empirical formulas for (**a**) glucose, a substance also known as either blood sugar or dextrose—molecular formula $C_6H_{12}O_6$; (**b**) nitrous oxide, a substance used as an anesthetic and commonly called laughing gas—molecular formula N_2O.

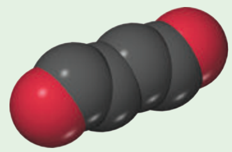
SOLUTION

(**a**) The subscripts of an empirical formula are the smallest whole-number ratios. The smallest ratios are obtained by dividing each subscript by the largest common factor, in this case 6. The resultant empirical formula for glucose is CH_2O.

(**b**) Because the subscripts in N_2O are already the lowest integral numbers, the empirical formula for nitrous oxide is the same as its molecular formula, N_2O.

Practice Exercise

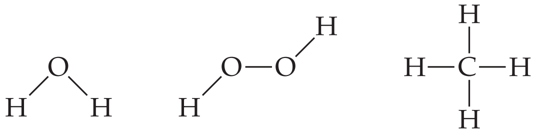
Tetracarbon dioxide is an unstable oxide of carbon with the following molecular structure:



What are the molecular and empirical formulas of this substance? (**a**) C_2O_2, CO_2 (**b**) C_4O, CO (**c**) CO_2, CO_2 (**d**) C_4O_2, C_2O (**e**) C_2O, CO_2

Picturing Molecules

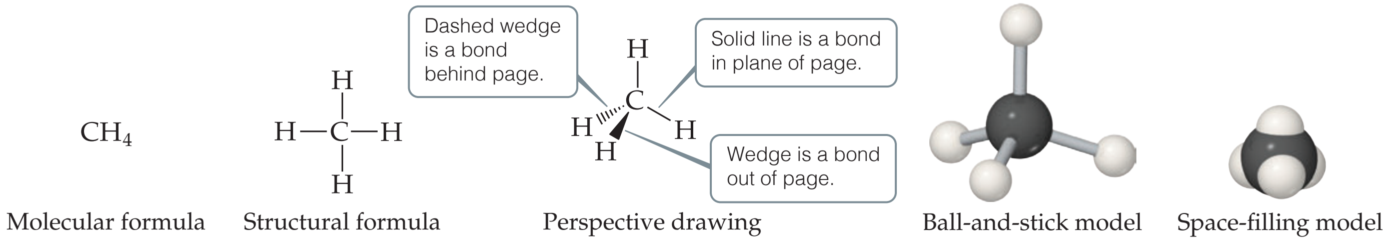
The molecular formula of a substance does not show how its atoms are joined together. A **structural formula** is needed to convey that information, as in the following examples:



Water Hydrogen peroxide Methane

The atoms are represented by their chemical symbols, and lines are used to represent the bonds that hold the atoms together.

Go Figure Which model, the ball-and-stick or the space-filling, more effectively shows the angles between bonds around a central atom?



▲ **Figure 2.16 Different representations of the methane (CH_4) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models.

A structural formula does not typically depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined; for that, more sophisticated representations are needed (**Figure 2.16**).

- **Perspective drawings** use wedges and dashed lines to depict bonds that are not in the plane of the paper. This gives a crude sense of the three-dimensional shape of a molecule.
- **Ball-and-stick models** show atoms as spheres and bonds as sticks. This type of model has the advantage of accurately representing the angles at which the atoms are attached to one another in a molecule (Figure 2.16). Sometimes the chemical symbols of the elements are superimposed on the balls, but often the atoms are identified simply by color.
- **Space-filling models** depict what a molecule would look like if the atoms were scaled up in size (Figure 2.16). These models show the relative sizes of the atoms, but the angles between atoms, which help define their molecular geometry, are often more difficult to see than in ball-and-stick models. Because space-filling models give a good representation of the true size of a molecule, they are useful for picturing how two molecules might fit together or pack in the solid state. As with ball-and-stick models, the identities of the atoms are typically indicated by color.

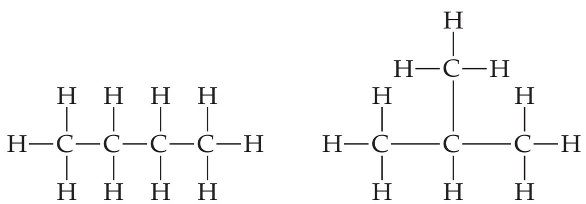
Self-Assessment Exercise

2.24 Lactic acid can build up in muscle tissue during strenuous exercise, sometimes leading to painful 'cramps'. The molecular formula of lactic acid is $C_3H_6O_3$. What is its empirical formula?

(a) CHO
(b) CH_2O
(c) $C_3H_6O_3$

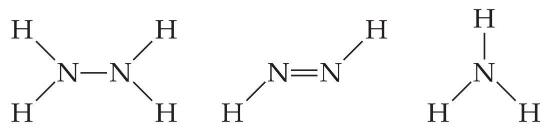
Exercises

2.25 The structural formulas of the compounds *n*-butane and isobutane are shown here. (**a**) Determine the molecular formula of each. (**b**) Determine the empirical formula of each. (**c**) Which formulas—empirical, molecular, or structural—allow you determine these are different compounds?



n-butane isobutane

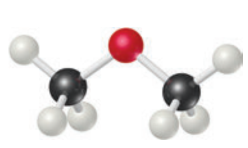
2.26 What are the molecular and empirical formulas for each of the following compounds?



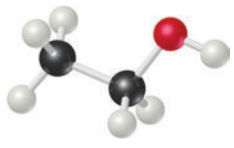
2.27 Write the empirical formula corresponding to each of the following molecular formulas: (**a**) Al_2Br_6, (**b**) C_8H_{10}, (**c**) $C_4H_8O_2$, (**d**) P_4O_{10}, (**e**) $C_6H_4Cl_2$, (**f**) $B_3N_3H_6$.

2.28 How many hydrogen atoms are in each of the following: (**a**) C_2H_5OH, (**b**) $Ca(C_2H_5COO)_2$, (**c**) $(NH_4)_3PO_4$?

2.29 Write the molecular and structural formulas for the compounds represented by the following molecular models:

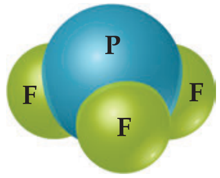


(a)



(b)

(c)



(d)

Answers to Self-Assessment Exercises

2.24 (b)

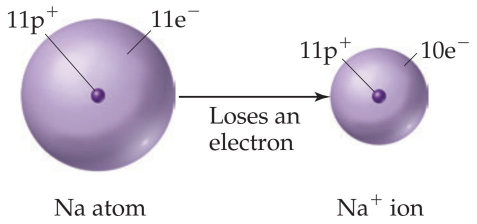
2.7 Ions and Ionic Compounds

Salt has always been an essential part of our diet. In regions where it is scarce, it commands a high price. During the Roman Empire, soldiers were paid, in part, with an allocation of salt (*sal* being the Latin word for salt) and this is the origin of our word 'salary'. By the end of this section, you should be able to

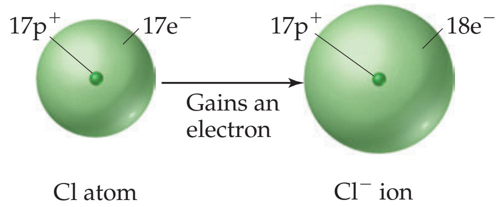
- Predict the charge on the ion formed by elements of Group 1, 2, 15, 16, and 17.
- Write the formula of an ionic compound given the charges of the ions

If electrons are removed from or added to an atom, a charged particle called an **ion** is formed. An ion with a positive charge is a **cation** (pronounced CAT-ion); a negatively charged ion is an **anion** (AN-ion).

To see how ions form, consider the sodium atom, which has 11 protons and 11 electrons. If this atom loses one electron, the resulting cation has 11 protons and 10 electrons, which means it has a net charge of 1+.



The net charge on an ion is represented by a superscript. The superscripts +, 2+, and 3+, for instance, mean a net charge resulting from the *loss* of one, two, and three electrons, respectively. The superscripts −, 2−, and 3− represent net charges resulting from the *gain* of one, two, and three electrons, respectively. Chlorine, with 17 protons and 17 electrons, for example, can gain an electron in chemical reactions, producing the Cl^- ion:



In general, metal atoms tend to lose electrons to form cations and nonmetal atoms tend to gain electrons to form anions. Thus, ionic compounds tend to be composed of both metal cations and nonmetal anions, as in NaCl.

Sample Exercise 2.7
Writing Chemical Symbols for Ions

Give the chemical symbol, including superscript indicating mass number, for (**a**) the ion with 22 protons, 26 neutrons, and 19 electrons; and (**b**) the ion of sulfur that has 16 neutrons and 18 electrons.

SOLUTION

(**a**) The number of protons is the atomic number of the element. A periodic table or list of elements tells us that the element with atomic number 22 is titanium (Ti). The mass number (protons plus neutrons) of this isotope of titanium is 22 + 26 = 48. Because the ion has three more protons than electrons, it has a net charge of 3+ and is designated $^{48}Ti^{3+}$.

(**b**) The periodic table tells us that sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur contains 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number is 16 + 16 = 32. Because the ion has 16 protons and 18 electrons, its net charge is 2− and the ion symbol is $^{32}S^{2-}$.

In general, we will focus on the net charges of ions and ignore their mass numbers unless the circumstances dictate that we specify a certain isotope.

Practice Exercise

In which of the following species is the difference between the number of protons and the number of electrons largest?
(**a**) Ti^{2+} (**b**) P^{3-} (**c**) Mn (**d**) Se^{2-} (**e**) Ce^{4+}

In addition to simple ions such as Na^+ and Cl^-, there are **polyatomic ions**, such as NH_4^+ (ammonium ion) and SO_4^{2-} (sulfate ion), which consist of atoms joined as in a molecule, but carrying a net positive or negative charge. Polyatomic ions will be discussed in Section 2.8.

It is important to realize that the chemical properties of ions are very different from the chemical properties of the atoms from which the ions are derived. The addition or removal of one or more electrons produces a charged species with behavior very different from that of its associated atom or group of atoms. For example, sodium metal reacts violently with water, but an ionic compound containing sodium ions, such as NaCl, does not.

Predicting Ionic Charges

As noted in Table 2.3, the elements of Group 18 are called the noble-gas elements. The noble gases are chemically nonreactive elements that form very few compounds. Many atoms gain or lose electrons to end up with the same number of electrons as the noble gas closest to them in the periodic table. We might deduce that atoms tend to acquire the electron arrangements of the noble gases because these electron arrangements are very stable. Nearby elements can obtain these same stable arrangements by losing or gaining electrons. For example, the loss of one electron from an atom of sodium leaves it with the same number of electrons as in a neon atom (10). Similarly, when chlorine gains an electron, it ends up with 18, the same number of electrons as in argon. This simple observation will be helpful for now to account for the formation of ions. A deeper explanation awaits us in Chapter 8, where we discuss chemical bonding.

Sample Exercise 2.8

Predicting Ionic Charge

Predict the charge expected for the most stable ion of barium and the most stable ion of oxygen.

SOLUTION

We will assume that barium and oxygen form ions that have the same number of electrons as the nearest noble-gas atom. From the periodic table, we see that barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two electrons, forming the Ba^{2+} cation.

Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, forming the O^{2-} anion.

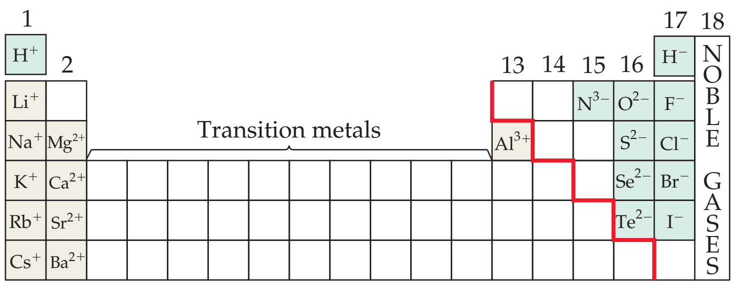
Practice Exercise

Although it is helpful to know that many ions have the electron arrangement of a noble gas, many elements, especially among the metals, form ions that do not have a noble-gas electron arrangement. Use the periodic table, Figure 2.13, to determine which of the following ions has a noble-gas electron arrangement, and which do not. For those that do, indicate the noble-gas arrangement they match: **(a)** Ti^{4+}, **(b)** Mn^{2+}, **(c)** Pb^{2+}, **(d)** Te^{2-}, **(e)** Zn^{2+}.

The periodic table is very useful for remembering ionic charges, especially those of elements on the left and right sides of the table. As **Figure 2.17** shows, the charges of these ions relate in a simple way to their positions in the table: The Group 1 elements (alkali metals) form 1+ ions, the Group 2 elements (alkaline earth metals) form 2+ ions, the Group 17 elements (halogens) form 1− ions, and the Group 16 elements form 2− ions. As noted in the Practice Exercise of Sample Exercise 2.8, many of the other groups do not lend themselves to such simple rules.

Go Figure

The most common ions for silver, zinc, and scandium are Ag^{+}, Zn^{2+}, and Sc^{3+}. Locate the boxes in which you would place these ions in this table. Which of these ions has the same number of electrons as a noble-gas element?

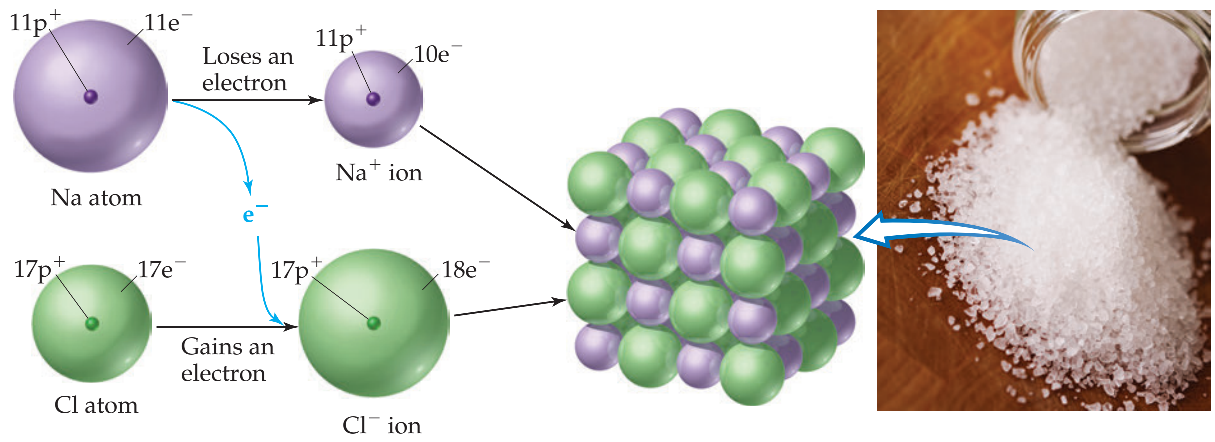


▲ **Figure 2.17 Predictable charges of some common ions.** Notice that the red stepped line that divides metals from nonmetals also separates cations from anions. Hydrogen forms both 1+ and 1− ions.

Ionic Compounds

A great deal of chemical activity involves the transfer of electrons from one substance to another. **Figure 2.18** shows that when elemental sodium is allowed to react with elemental chlorine, an electron transfers from a sodium atom to a chlorine atom, forming a Na^{+} ion and a Cl^{-} ion. Because objects of opposite charges attract, the Na^{+} and the Cl^{-} ions bind together to form the compound sodium chloride (NaCl). Sodium chloride, which we know better as common table salt, is an example of an **ionic compound**, a compound made up of cations and anions.

Go Figure Is it possible to pick out a unique cluster of atoms that can be thought of as a sodium chloride molecule?



▲ **Figure 2.18 Formation of an ionic compound.** The transfer of an electron from a sodium atom to a chlorine atom leads to the formation of a Na^+ ion and a Cl^- ion. These ions are arranged in a lattice in solid sodium chloride, NaCl.

We can often tell whether a compound is ionic (consisting of ions) or molecular (consisting of molecules) from its composition. As a general rule, cations are metal ions and anions are nonmetal ions. Consequently,

Ionic compounds are generally combinations of metals and nonmetals, as in NaCl.

Sample Exercise 2.9
Identifying Ionic and Molecular Compounds

Which of these compounds would you expect to be ionic: N_2O, Na_2O, $CaCl_2$, SF_4?

SOLUTION

We predict that Na_2O and $CaCl_2$ are ionic compounds because they are composed of a metal combined with a nonmetal. We predict (correctly) that N_2O and SF_4 are molecular compounds because they are composed entirely of nonmetals.

▶ Practice Exercise

Give a reason why each of the following statements is a safe prediction:

(**a**) Every compound of Rb with a nonmetal is ionic in character.
(**b**) Every compound of nitrogen with a halogen element is a molecular compound.
(**c**) The compound $MgKr_2$ does not exist.
(**d**) Na and K are very similar in the compounds they form with nonmetals.
(**e**) If contained in an ionic compound, calcium (Ca) will be in the form of the doubly charged ion, Ca^{2+}.

The ions in ionic compounds are arranged in three-dimensional structures, as Figure 2.18 shows for NaCl. Because there is no discrete “molecule” of NaCl, we are able to write only an empirical formula for this substance. This is true for most other ionic compounds.

We can write the empirical formula for an ionic compound if we know the charges of the ions. Because chemical compounds are always electrically neutral, the ions in an ionic compound always occur in such a ratio that the total positive charge equals the total negative charge. Thus, there is one Na^+ to one Cl^- in NaCl, one Ba^{2+} to two Cl^- in $BaCl_2$, and so forth.

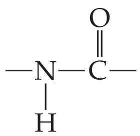
As you consider these and other examples, you will see that if the charges on the cation and anion are equal, the subscript on each ion is 1. If the charges are not equal, the charge on one ion (without its sign) will become the subscript on the other ion. For example, the ionic compound formed from Mg (which forms Mg^{2+} ions) and N (which forms N^{3-} ions) is Mg_3N_2:

$$Mg^{2+} \quad N^{3-} \longrightarrow Mg_3N_2$$

There is one caveat to using this approach. Remember that the empirical formula should be the smallest possible whole-number ratio of the two elements. So the empirical formula for the ionic compound formed between Ti^{4+} and O^{2-} is TiO_2 rather than Ti_2O_4.

CHEMISTRY AND LIFE Elements Required by Living Organisms

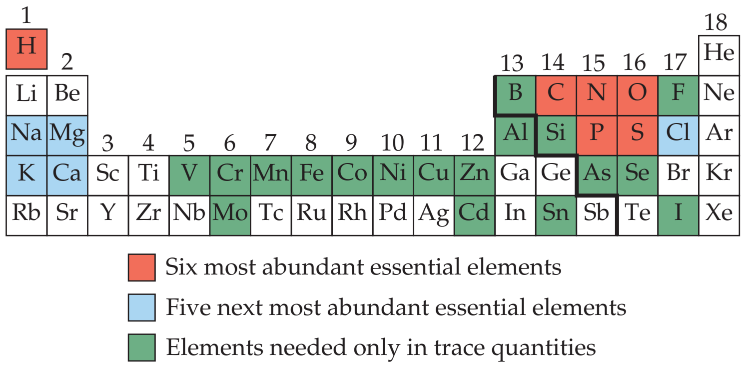
The elements essential to life are highlighted in color in **Figure 2.19.** More than 97% of the mass of most organisms is made up of just six of these elements—oxygen, carbon, hydrogen, nitrogen, phosphorus, and sulfur. Water is the most common compound in living organisms, accounting for at least 70% of the mass of most cells. In the solid components of cells, carbon is the most prevalent element by mass. Carbon atoms are found in a vast variety of organic molecules, bonded either to other carbon atoms or to atoms of other elements. Nearly all proteins, for example, contain the carbon-based group



which occurs repeatedly in the molecules.

In addition, 23 other elements have been found in various living organisms. Five are ions required by all organisms: Ca^{2+}, Cl^-, Mg^{2+}, K^+, and Na^+. Calcium ions, for example, are necessary for the formation of bone and transmission of nervous system signals. Many other elements are needed in only very small quantities and consequently are called *trace* elements. For example, trace quantities of copper are required in the diet of humans to aid in the synthesis of hemoglobin.

Related Exercise: 2.113



▲ **Figure 2.19 Elements essential to life.**

Sample Exercise 2.10
Using Ionic Charge to Write Empirical Formulas for Ionic Compounds

Write the empirical formula of the compound formed by **(a)** Al^{3+} and Cl^- ions, **(b)** Al^{3+} and O^{2-} ions, **(c)** Mg^{2+} and NO_3^- ions.

SOLUTION

(a) Three Cl^- ions are required to balance the charge of one Al^{3+} ion, making the empirical formula $AlCl_3$.

(b) Two Al^{3+} ions are required to balance the charge of three O^{2-} ions. A 2:3 ratio is needed to balance the total positive charge of 6+ and the total negative charge of 6–. The empirical formula is Al_2O_3.

(c) Two NO_3^- ions are needed to balance the charge of one Mg^{2+}, yielding $Mg(NO_3)_2$. Note that the formula for the polyatomic ion, NO_3^-, must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

▶ Practice Exercise

Which of the following nonmetals will form an ionic compound with Sc^{3+} that has a 1:1 ratio of cations to anions?
(a) Ne **(b)** F **(c)** O **(d)** N

Self-Assessment Exercises

2.30 Which group in the periodic table contains elements that form a 2- anion?

(a) Group 2

(b) Group 12

(c) Group 16

2.31 Which is the formula of the ionic compound formed from sodium cations and sulfur anion?

(a) Na_2S

(b) NaS

(c) NaS_2

Exercises

2.32 Predict whether each of the following compounds is molecular or ionic: (**a**) $HClO_4$ (**b**) CH_3OCH_3 (**c**) $Mg(NO_3)_2$ (**d**) H_2S (**e**) $TiCl_4$ (**f**) K_2O_2 (**g**) PCl_5 (**h**) $P(OH)_3$.

2.33 Fill in the gaps in the following table:

Symbol	$^{58}Fe^{2+}$			
Protons		50		40
Neutrons		68	78	50
Electrons			54	38
Net charge		4+	2−	

2.34 Each of the following elements is capable of forming an ion in chemical reactions. By referring to the periodic table, predict the charge of the most stable ion of each: (**a**) Be, (**b**) Rb, (**c**) As, (**d**) In, (**e**) At.

2.35 Using the periodic table to guide you, predict the chemical formula of the compound formed by the following elements: (**a**) Ga and F, (**b**) Li and H, (**c**) Al and I, (**d**) K and S.

2.36 Predict the chemical formulas of the ionic compound formed by (**a**) Fe^{3+} and OH^-, (**b**) Cs^+ and NO_3^-, (**c**) V^{2+} and CH_3COO^-, (**d**) Li^+ and PO_4^{3-}, (**e**) In^{3+} and O^{2-}.

2.37 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	K^+	NH_4^+	Mg^{2+}	Fe^{3+}
Cl^-	KCl			
OH^-				
CO_3^{2-}				
PO_4^{3-}				

Answers to Self-Assessment Exercises

2.30 (c) 2.31 (a)

2.8 Naming Inorganic Compounds

The names and chemical formulas of compounds are essential vocabulary in chemistry. The system used in naming substances is called **chemical nomenclature**, from the Latin words *nomen* (name) and *calare* (to call). By the end of this section, you should be able to

- Convert between the name and formula of an inorganic compound

There are more than 50 million known chemical substances. Naming them all would be a hopelessly complicated task if each had a name independent of all others. Many important substances that have been known for a long time, such as water (H_2O) and ammonia (NH_3), do have traditional names (called *common names*). For most substances, however, we rely on a set of rules that leads to an informative and unique name that conveys the composition of the substance.

The rules for chemical nomenclature are based on the division of substances into categories. The major division is between organic and inorganic compounds. *Organic compounds* contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements. All others are *inorganic compounds*. Early chemists associated organic compounds with plants and animals and inorganic compounds with the nonliving portion of our world, for example minerals and water. Although this distinction is no longer pertinent, the classification between organic and inorganic compounds continues to be useful. In this section, we consider the basic rules for naming three categories of inorganic compounds: ionic compounds, acids, and molecular compounds.

Names and Formulas of Ionic Compounds

Recall from Section 2.7 that ionic compounds usually consist of metal ions combined with nonmetal ions. The metals form the cations, and the nonmetals form the anions.

1. Cations

a. *Cations formed from metal atoms have the same name as the metal:*

Na^+	sodium ion	Zn^{2+}	zinc ion	Al^{3+}	aluminum ion

b. *If a metal can form cations with different charges, the positive charge is indicated by a Roman numeral in parentheses following the name of the metal:*

Fe^{2+}	iron(II) ion	Cu^+	copper(I) ion
Fe^{3+}	iron(III) ion	Cu^{2+}	copper(II) ion

Ions of the same element that have different charges have different chemical and physical properties, such as color (**Figure 2.20**).

Most metals that form cations with different charges are *transition metals*, elements that occur in the middle of the periodic table, from Group 3 to Group 12. The metals that form only one cation (only one possible charge) are those of Group 1 and Group 2, as well as Al^{3+} (Group 13) and two transition-metal ions: Ag^+ (Group 12) and Zn^{2+} (Group 11). Charges are not expressed when naming these ions. However, if there is any doubt in your mind whether a metal forms more than one cation, use a Roman numeral to indicate the charge. It is never wrong to do so, even though it may be unnecessary.

An older method still widely used for distinguishing between differently charged ions of a metal uses the endings *-ous* and *-ic* added to the root of the element's Latin name:

Fe^{2+}	ferrous ion	Cu^+	cuprous ion
Fe^{3+}	ferric ion	Cu^{2+}	cupric ion

Although we will only rarely use these older names in this text, you might encounter them elsewhere.

▲ **Figure 2.20 Different ions of the same element have different properties.** Both substances shown are compounds of iron. The grey substance on the left is Fe_3O_4, which contains Fe^{2+} and Fe^{3+} ions. The red substance on the right is Fe_2O_3, which contains Fe^{3+} ions.

c. *Cations formed from molecules composed of nonmetal atoms have names that end in* -ium:

NH_4^+	ammonium ion	H_3O^+	hydronium ion

These two ions are the only ions of this kind that we will encounter frequently in the text.

The names and formulas of some common cations are shown in Table 2.4 and on the back inside cover of the text. The ions on the left side in Table 2.4 are the monatomic ions that do not have more than one possible charge. Those on the right side are either polyatomic cations or cations with more than one possible charge. The Hg_2^{2+} ion is unusual because, even though it is a metal ion, it is not monatomic. It is called the mercury(I) ion because it can be thought of as two Hg^+ ions bound together. The cations that you will encounter most frequently in this text are shown in boldface. You should learn these cations first.

TABLE 2.4 Common Cations[a]

Charge	Formula	Name	Formula	Name
1+	$\mathbf{H^+}$	**hydrogen ion**	$\mathbf{NH_4^+}$	**ammonium ion**
	Li^+	lithium ion	Cu^+	copper(I) or cuprous ion
	$\mathbf{Na^+}$	**sodium ion**		
	$\mathbf{K^+}$	**potassium ion**		
	Cs^+	cesium ion		
	$\mathbf{Ag^+}$	**silver ion**		
2+	$\mathbf{Mg^{2+}}$	**magnesium ion**	Co^{2+}	cobalt(II) or cobaltous ion
	$\mathbf{Ca^{2+}}$	**calcium ion**	$\mathbf{Cu^{2+}}$	**copper(II)** or cupric ion
	Sr^{2+}	strontium ion	$\mathbf{Fe^{2+}}$	**iron(II)** or ferrous ion
	Ba^{2+}	barium ion	Mn^{2+}	manganese(II) or manganous ion
	$\mathbf{Zn^{2+}}$	**zinc ion**	Hg_2^{2+}	mercury(I) or mercurous ion
	Cd^{2+}	cadmium ion	Hg^{2+}	mercury(II) or mercuric ion
			Ni^{2+}	nickel(II) or nickelous ion
			$\mathbf{Pb^{2+}}$	**lead(II)** or plumbous ion
			Sn^{2+}	tin(II) or stannous ion
3+	$\mathbf{Al^{3+}}$	**aluminum ion**	Cr^{3+}	chromium(III) or chromic ion
			$\mathbf{Fe^{3+}}$	**iron(III)** or ferric ion

[a]The ions we use most often in this course are in boldface. Learn them first.

2. Anions

a. *The names of monatomic anions are formed by replacing the ending of the name of the element with* -ide:

H^-	hydride ion	O^{2-}	oxide ion	N^{3-}	nitride ion

A few polyatomic anions also have names ending in *-ide*:

OH^-	hydroxide ion	CN^-	cyanide ion	O_2^{2-}	peroxide ion

b. *Polyatomic anions containing oxygen have names ending in either* -ate *or* -ite *and are called* **oxyanions**. The *-ate* is used for the most common or representative oxyanion of an element, and *-ite* is used for an oxyanion that has the same charge but one O atom fewer:

NO_3^-	nitrate ion	SO_4^{2-}	sulfate ion
NO_2^-	nitrite ion	SO_3^{2-}	sulfite ion

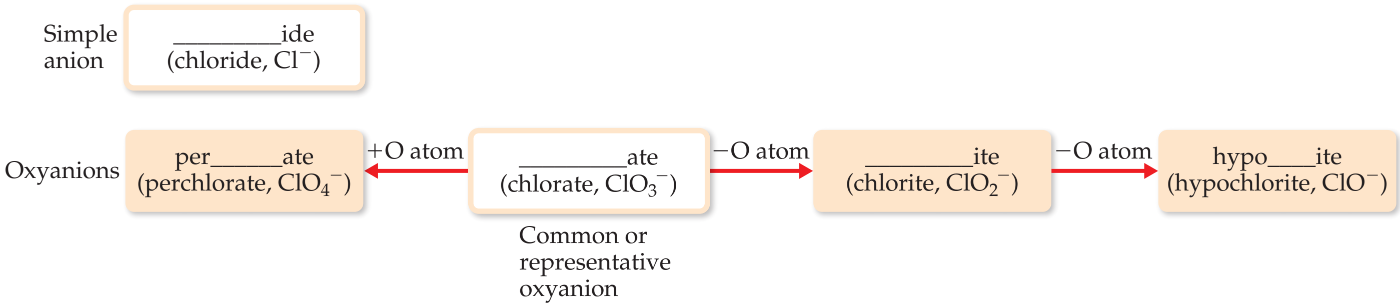
Prefixes are used when the series of oxyanions of an element extends to four members, as with the halogens. The prefix *per-* indicates one more O atom than the oxyanion ending in *-ate; hypo-* indicates one O atom fewer than the oxyanion ending in *-ite*:

ClO_4^-	perchlorate ion (one more O atom than chlorate)
ClO_3^-	chlorate ion
ClO_2^-	chlorite ion (one O atom fewer than chlorate)
ClO^-	hypochlorite ion (one O atom fewer than chlorite)

These rules are summarized in **Figure 2.21**.

Go Figure

Name the anion obtained by removing one oxygen atom from the perbromate ion, BrO_4^-.



▲ **Figure 2.21 Procedure for naming anions.** Here the simple anion and polyatomic oxyanions derived from chlorine are used as an example.

Figure 2.22 can help you remember the charge and number of oxygen atoms in the various oxyanions. Notice that C and N, both Period 2 elements, have only three O atoms each, whereas the Period 3 elements P, S, and Cl have four O atoms each. Beginning at the lower right in Figure 2.22, note that ionic charge increases from right to left, from 1− for ClO_4^- to 3− for PO_4^{3-}. In the second period the charges also increase from right to left, from 1− for NO_3^- to 2− for CO_3^{2-}. Notice also that although each of the anions in Figure 2.22 ends in *-ate*, the ClO_4^- ion also has a *per-* prefix.

Maximum of three O atoms in Period 2.

	Group 14	Group 15	Group 16	Group 17
Period 2	CO_3^{2-} **Carbonate ion**	NO_3^- **Nitrate ion**		
Period 3		PO_4^{3-} **Phosphate ion**	SO_4^{2-} **Sulfate ion**	ClO_4^- **Perchlorate ion**

Charges increase right to left.

Maximum of four O atoms in Period 3.

▲ **Figure 2.22 Common oxyanions.** The composition and charges of common oxyanions are related to their location in the periodic table.

Sample Exercise 2.11

Determining the Formula of an Oxyanion from Its Name

Based on the formula for the sulfate ion, predict the formula for (**a**) the selenate ion and (**b**) the selenite ion. (Sulfur and selenium are both in Group 16 and form analogous oxyanions.)

SOLUTION

(**a**) The sulfate ion is SO_4^{2-}. The analogous selenate ion is therefore SeO_4^{2-}.

(**b**) The ending *-ite* indicates an oxyanion with the same charge but one O atom fewer than the corresponding oxyanion that ends in *-ate*. Thus, the formula for the selenite ion is SeO_3^{2-}.

Practice Exercise

Which of the following oxyanions is incorrectly named? (**a**) ClO_2^-, chlorate (**b**) IO_4^-, periodate (**c**) SO_3^{2-}, sulfite (**d**) IO_3^-, iodate (**e**) NO_2^-, nitrite

c. *Anions derived by adding* H^+ *to an oxyanion are named by adding as a prefix the word* hydrogen *or* dihydrogen, *as appropriate*:

CO_3^{2-}	carbonate ion	PO_4^{3-}	phosphate ion
HCO_3^-	hydrogen carbonate ion	$H_2PO_4^-$	dihydrogen phosphate ion

Notice that each H^+ added reduces the negative charge of the parent anion by one. An older method for naming some of these ions uses the prefix *bi-*. Thus, the HCO_3^- ion is commonly called the bicarbonate ion, and HSO_4^- is sometimes called the bisulfate ion.

The names and formulas of the common anions are listed in Table 2.5 and on the back inside cover of the text. Those anions whose names end in *-ide* are listed on the left portion of Table 2.5, and those whose names end in *-ate* are listed on the right. The most common of these ions are shown in boldface. You should learn the names and formulas of these anions first. The formulas of the ions whose names end with *-ite* can be derived from those ending in *-ate* by removing an O atom. Notice the location of the monatomic ions in the periodic table. Those of Group 17 always have a 1− charge (F^-, Cl^-, Br^-, and I^-), and those of Group 16 have a 2− charge (O^{2-} and S^{2-}).

TABLE 2.5 Common Anions[a]

Charge	Formula	Name	Formula	Name
1−	H^-	hydride ion	**CH_3COO^-** (or $C_2H_3O_2^-$)	**acetate ion**
	F^-	**fluoride ion**	ClO_3^-	chlorate ion
	Cl^-	**chloride ion**	**ClO_4^-**	**perchlorate ion**
	Br^-	**bromide ion**	**NO_3^-**	**nitrate ion**
	I^-	**iodide ion**	MnO_4^-	permanganate ion
	CN^-	cyanide ion		
	OH^-	**hydroxide ion**		
2−	**O^{2-}**	**oxide ion**	**CO_3^{2-}**	**carbonate ion**
	O_2^{2-}	peroxide ion	CrO_4^{2-}	chromate ion
	S^{2-}	**sulfide ion**	$Cr_2O_7^{2-}$	dichromate ion
			SO_4^{2-}	**sulfate ion**
3−	N^{3-}	nitride ion	**PO_4^{3-}**	**phosphate ion**

[a]The ions we use most often are in boldface. Learn them first.

3. Ionic Compounds

Names of ionic compounds consist of the cation name followed by the anion name:

$CaCl_2$	calcium chloride
$Al(NO_3)_3$	aluminum nitrate
$Cu(ClO_4)_2$	copper(II) perchlorate (or cupric perchlorate)

In the chemical formulas for aluminum nitrate and copper(II) perchlorate, parentheses followed by the appropriate subscript are used because the compounds contain two or more polyatomic ions.

Sample Exercise 2.12

Determining the Names of Ionic Compounds from Their Formulas

Name the ionic compounds (**a**) K_2SO_4, (**b**) $Ba(OH)_2$, (**c**) $FeCl_3$.

SOLUTION

In naming ionic compounds, it is important to recognize polyatomic ions and to determine the charge of cations with variable charge.

(**a**) The cation is K^+, the potassium ion, and the anion is SO_4^{2-}, the sulfate ion, making the name potassium sulfate. (If you thought the compound contained S^{2-} and O^{2-} ions, you failed to recognize the polyatomic sulfate ion.)

(**b**) The cation is Ba^{2+}, the barium ion, and the anion is OH^-, the hydroxide ion: barium hydroxide.

(**c**) You must determine the charge of Fe in this compound because an iron atom can form more than one cation. Because the compound contains three chloride ions, Cl^-, the cation must be Fe^{3+}, the iron(III), or ferric, ion. Thus, the compound is iron(III) chloride or ferric chloride.

Practice Exercise

Which of the following ionic compounds is incorrectly named? (**a**) $Zn(NO_3)_2$, zinc nitrate (**b**) $TeCl_4$, tellurium(IV) chloride (**c**) Fe_2O_3, diiron oxide (**d**) BaO, barium oxide (**e**) $Mn_3(PO_4)_2$, manganese(II) phosphate

Names and Formulas of Acids

Acids are an important class of hydrogen-containing compounds, and they are named in a special way. For our present purposes, an *acid* is a substance whose molecules yield hydrogen ions (H^+) when dissolved in water. When we encounter the chemical formula for an acid at this stage of the course, it will be written with H as the first element, as in HCl and H_2SO_4.

An acid is composed of an anion connected to enough H^+ ions to neutralize, or balance, the anion's charge. Thus, the SO_4^{2-} ion requires two H^+ ions, forming H_2SO_4. The name of an acid is related to the name of its anion, as summarized in **Figure 2.23**.

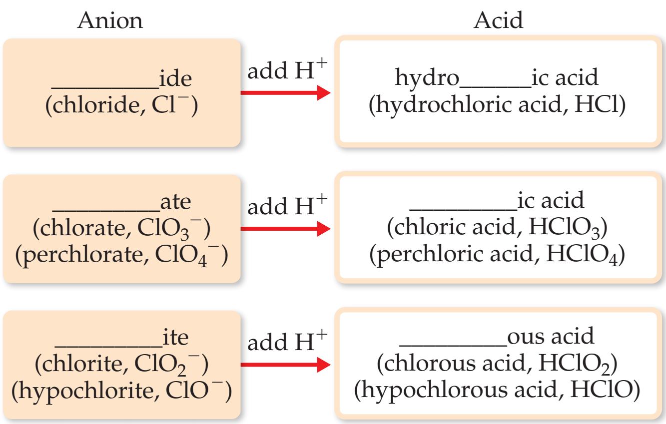
How to Name Acids

1. *Acids containing anions whose names end in* -ide *are named by changing the* -ide *ending to* -ic, *adding the prefix* hydro- *to this anion name, and then following with the word* acid:

Anion	Corresponding Acid
Cl^- (chloride)	HCl (hydrochloric acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

2. *Acids containing anions whose names end in* -ate *or* -ite *are named by changing* -ate *to* -ic *and* -ite *to* -ous *and then adding the word* acid. Prefixes in the anion name are retained in the name of the acid:

Anion		Corresponding Acid	
ClO_4^-	(perchlorate)	$HClO_4$	(perchloric acid)
ClO_3^-	(chlorate)	$HClO_3$	(chloric acid)
ClO_2^-	(chlorite)	$HClO_2$	(chlorous acid)
ClO^-	(hypochlorite)	$HClO$	(hypochlorous acid)



▲ **Figure 2.23 Procedure for naming acids.** Here acids containing chlorine are used as an example. Prefixes used for oxyanions, such as, *per-* and *hypo-*, are retained in the acids derived from those anions.

Sample Exercise 2.13
Relating the Names and Formulas of Acids

Name the acids (**a**) HCN, (**b**) HNO_3, (**c**) H_2SO_4, (**d**) H_2SO_3.

SOLUTION

(**a**) The anion from which this acid is derived is CN^-, the cyanide ion. Because this ion has an *-ide* ending, the acid is given a *hydro-* prefix and an *-ic* ending: hydrocyanic acid. Only water solutions of HCN are referred to as hydrocyanic acid. The pure compound, which is a gas under normal conditions, is called hydrogen cyanide. Both hydrocyanic acid and hydrogen cyanide are *extremely* toxic.

(**b**) Because NO_3^- is the nitrate ion, HNO_3 is called nitric acid (the *-ate* ending of the anion is replaced with an *-ic* ending in naming the acid).

(**c**) Because SO_4^{2-} is the sulfate ion, H_2SO_4 is called sulfuric acid.

(**d**) Because SO_3^{2-} is the sulfite ion, H_2SO_3 is sulfurous acid (the *-ite* ending of the anion is replaced with an *-ous* ending).

▶ Practice Exercise

Which of the following acids are incorrectly named? For those that are, provide a correct name or formula.
(**a**) hydrofluoric acid, HF (**b**) nitrous acid, HNO_3 (**c**) perbromic acid, $HBrO_4$ (**d**) iodic acid, HI (**e**) selenic acid, H_2SeO_4

TABLE 2.6 Prefixes Used in Naming Binary Compounds Formed between Nonmetals

Prefix	Meaning
mono-	*1*
di-	*2*
tri-	*3*
tetra-	*4*
penta-	*5*
hexa-	*6*
hepta-	*7*
octa-	*8*
nona-	*9*
deca-	*10*

Names and Formulas of Binary Molecular Compounds

The procedures used for naming *binary* (two-element) molecular compounds are similar to those used for naming ionic compounds:

How to Name Binary Molecular Compounds

1. *The name of the element farther to the left in the periodic table (closest to the metals) is usually written first.* An exception occurs when the compound contains oxygen and chlorine, bromine, or iodine (any halogen except fluorine), in which case oxygen is written last.
2. *If both elements are in the same group, the one closer to the bottom of the table is named first.*
3. *The name of the second element is given an* -ide *ending.*
4. *Greek prefixes* (Table 2.6) indicate the number of atoms of each element. (Exception: The prefix *mono-* is never used with the first element.) When the prefix ends in *a* or *o* and the name of the second element begins with *a* vowel, the *a* or *o* of the prefix is often dropped.

The following examples illustrate these rules:

Cl_2O	dichlorine monoxide	NF_3	nitrogen trifluoride
N_2O_4	dinitrogen tetroxide	P_4S_{10}	tetraphosphorus decasulfide

Rule 4 is necessary because we cannot predict formulas for most molecular substances the way we can for ionic compounds. Molecular compounds that contain hydrogen and one other element are an important exception, however. These compounds can be treated as if they were neutral substances containing H^+ ions and anions. Thus, you can predict that the substance named hydrogen chloride has the formula HCl, containing one H^+ to balance the charge of one Cl^-. (The name *hydrogen chloride* is used only for the pure compound; water solutions of HCl are called hydrochloric acid. The distinction, which is important, will be explained in Section 4.1.) Similarly, the formula for hydrogen sulfide is H_2S because two H^+ ions are needed to balance the charge on S^{2-}.

Sample Exercise 2.14

Relating the Names and Formulas of Binary Molecular Compounds

Name the compounds (**a**) SO_2, (**b**) PCl_5, (**c**) Cl_2O_3.

SOLUTION

The compounds consist entirely of nonmetals, so they are molecular rather than ionic. Using the prefixes in Table 2.6 , we have (**a**) sulfur dioxide, (**b**) phosphorus pentachloride, (**c**) dichlorine trioxide.

Practice Exercise

Give the name for each of the following binary compounds of carbon: (**a**) CS_2, (**b**) CO, (**c**) C_3O_2, (**d**) CBr_4, (**e**) CF.

Self-Assessment Exercise

2.38 What is the name of $NaClO_2$?

(**a**) Sodium chlorine oxide

(**b**) Sodium chlorite

(**c**) Sodium chlorate

Exercises

2.39 Give the chemical formula for (**a**) chromate ion (**b**) bromide ion (**c**) nitrite ion (**d**) sulphite ion (**e**) permanganate ion.

2.40 Give the names and charges of the cation and anion in each of the following compounds: (**a**) CaO, (**b**) Na_2SO_4, (**c**) $KClO_4$, (**d**) $Fe(NO_3)_2$, (**e**) $Cr(OH)_3$.

2.41 Name the following ionic compounds: (**a**) Li_2O, (**b**) $FeCl_3$, (**c**) NaClO, (**d**) $CaSO_3$, (**e**) $Cu(OH)_2$, (**f**) $Fe(NO_3)_2$, (**g**) $Ca(CH_3COO)_2$, (**h**) $Cr_2(CO_3)_3$, (**i**) K_2CrO_4, (**j**) $(NH_4)_2SO_4$.

2.42 Write the chemical formulas for the following compounds: (**a**) aluminum hydroxide, (**b**) potassium sulfate, (**c**) copper(I) oxide, (**d**) zinc nitrate, (**e**) mercury(II) bromide, (**f**) iron(III) carbonate, (**g**) sodium hypobromite.

2.43 Give the name or chemical formula, as appropriate, for each of the following acids: (**a**) $HBrO_3$, (**b**) HBr, (**c**) H_3PO_4, (**d**) hypochlorous acid, (**e**) iodic acid, (**f**) sulfurous acid.

2.44 Give the name or chemical formula, as appropriate, for each of the following binary molecular substances: (**a**) SF_6, (**b**) IF_5, (**c**) XeO_3, (**d**) dinitrogen tetroxide, (**e**) hydrogen cyanide, (**f**) tetraphosphorus hexasulfide.

2.45 Write the chemical formula for each substance mentioned in the following word descriptions (use the front inside cover to find the symbols for the elements you do not know). (**a**) Zinc carbonate can be heated to form zinc oxide and carbon dioxide. (**b**) On treatment with hydrofluoric acid, silicon dioxide forms silicon tetrafluoride and water. (**c**) Sulfur dioxide reacts with water to form sulfurous acid. (**d**) The substance phosphorus trihydride, commonly called phosphine, is a toxic gas. (**e**) Perchloric acid reacts with cadmium to form cadmium(II) perchlorate. (**f**) Vanadium(III) bromide is a colored solid.

Answers to Self-Assessment Exercises

2.37 (b)

2.9 Some Simple Organic Compounds

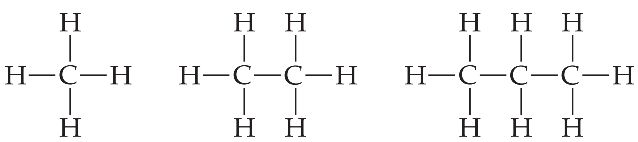
Organic compounds were believed to require living cells for their synthesis, either plant or animal in origin. Now there are many ways in which we may synthesize organic compound from inorganic materials. They represent over 99% of all known compounds and each one has carbon as a central element. By the end of this section, you should be able to

- Recognize and name simple alkanes and alcohols

The study of compounds of carbon is called **organic chemistry**, and as noted earlier, compounds that contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements, are called *organic compounds*. Organic compounds are a very important part of chemistry, far outnumbering all other types of chemical substances. We will examine organic compounds in a systematic way later on, but you will encounter many examples of them throughout the text. Here we present a brief introduction to some of the simplest organic compounds and the ways in which they are named.

Alkanes

Compounds that contain only carbon and hydrogen are called **hydrocarbons**. In the simplest class of hydrocarbons, **alkanes**, each carbon is bonded to four other atoms. The three smallest alkanes are methane (CH_4), ethane (C_2H_6), and propane (C_3H_8). The structural formulas of these three alkanes are as follows:

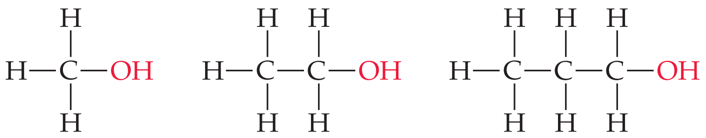


Methane   Ethane   Propane

Although hydrocarbons are binary molecular compounds, they are not named like the binary inorganic compounds discussed in Section 2.8. Instead, each alkane has a name that ends in *-ane*. The alkane with four carbons is called *butane*. For alkanes with five or more carbons, the names are derived from prefixes like those in Table 2.6. An alkane with eight carbon atoms, for example, is *octane* (C_8H_{18}), where the *octa-* prefix for eight is combined with the *-ane* ending for an alkane.

Some Derivatives of Alkanes

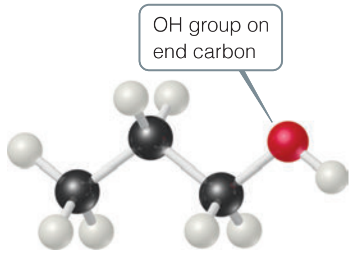
Other classes of organic compounds are obtained when one or more hydrogen atoms in an alkane are replaced with *functional groups*, which are specific groups of atoms. An **alcohol**, for example, is obtained by replacing an H atom of an alkane with an —OH group. The name of the alcohol is derived from that of the alkane by adding an *-ol* ending:



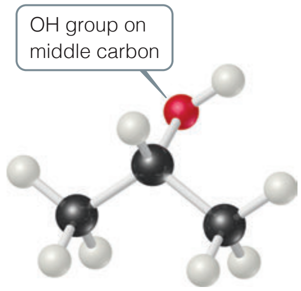
Methanol Ethanol 1-Propanol

Alcohols have properties that are very different from those of the alkanes from which the alcohols are obtained. For example, methane, ethane, and propane are all colorless gases under normal conditions, whereas methanol, ethanol, and propanol are colorless liquids. We will discuss the reasons for these differences in Chapter 11.

The prefix "1" in the name 1-propanol indicates that the replacement of H with OH has occurred at one of the "outer" carbon atoms rather than the "middle" carbon atom. A different compound, called 2-propanol is obtained when the OH functional group is attached to the middle carbon atom (**Figure 2.24**).



1-Propanol

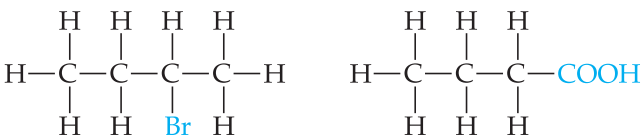


2-Propanol

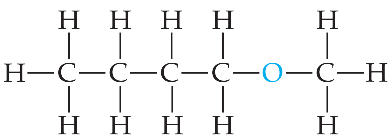
▲ **Figure 2.24 The two forms (isomers) of propanol.**

Compounds with the same molecular formula but different arrangements of atoms are called **isomers**. For example, 1-propanol and 2-propanol are *structural isomers*, compounds that have the same molecular formula but different structural formulas. There are many different kinds of isomers, as we will discover later in this book.

As already noted, many different functional groups can replace one or more of the hydrogens on an alkane—for example, one or more of the halogens, or a special grouping of carbon and oxygen atoms, such as the carboxylic acid group, —COOH. Here are a few examples of functional groups you will be encountering in the chapters that lie ahead (the functional group is highlighted in blue):



2-Bromobutane Butyric acid



Methyl butyl ether

Much of the richness of organic chemistry is possible because organic compounds can form long chains of carbon–carbon bonds. The series of alkanes that begins with methane, ethane, and propane and the series of alcohols that begins with methanol, ethanol, and propanol can both be extended for as long as we desire, in principle. The properties of alkanes and alcohols change as the chains get longer. Octanes, which are alkanes with eight carbon atoms, are liquids under normal conditions. If the alkane series is extended to tens of thousands of carbon atoms, we obtain *polyethylene*, a solid substance that is used to make thousands of plastic products, such as plastic bags, food containers, and laboratory equipment.

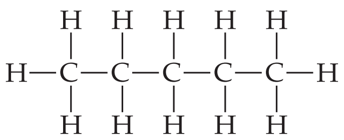
Sample Exercise 2.15

Writing Structural and Molecular Formulas for Hydrocarbons

Assuming the carbon atoms in *pentane* are in a linear chain, write (**a**) the structural formula and (**b**) the molecular formula for this alkane.

SOLUTION

(**a**) Alkanes contain only carbon and hydrogen, and each carbon is attached to four other atoms. The name *pentane* contains the prefix *penta*- for five (Table 2.6), and we are told that the carbons are in a linear chain. If we then add enough hydrogen atoms to make four bonds to each carbon, we obtain the structural formula



This form of pentane is often called *n*-pentane, where the *n*- stands for "normal" because all five carbon atoms are in one line in the structural formula.

(**b**) Once the structural formula is written, we determine the molecular formula by counting the atoms present. Thus, *n*-pentane has the molecular formula C_5H_{12}.

Practice Exercise

(**a**) What is the molecular formula of hexane, the alkane with six carbons? (**b**) What are the name and molecular formula of an alcohol derived from hexane?

STRATEGIES FOR SUCCESS How to Take a Test

At about this time in your study of chemistry, you are likely to face your first examination. The best way to prepare is to study, do homework diligently, and get help from the instructor on any material that is unclear or confusing. (See the advice for learning and studying chemistry presented in the preface of the book.) We present here some general guidelines for taking tests.

Depending on the nature of your course, the exam could consist of a variety of different types of questions.

1. **Multiple-choice questions** In large-enrollment courses, the most common kind of test question is the multiple-choice question. Many of the practice exercise problems in this book are written in this format to give you practice at this style of question. When faced with this type of problem, the first thing to realize is that the instructor has written the question so that at first glance all the answers appear to be correct. Thus, you should not jump to the conclusion that because one of the choices looks correct, it must be correct.

   If a multiple-choice question involves a calculation, do the calculation, check your work, and *only then* compare your answer with the choices. Keep in mind, though, that your instructor has anticipated the most common errors you might make in solving a given problem and has probably listed the incorrect answers resulting from those errors. Always double-check your reasoning and use dimensional analysis to arrive at the correct numeric answer and the correct units.

   In multiple-choice questions that do not involve calculations, if you are not sure of the correct choice, eliminate all the choices you know for sure to be incorrect. The reasoning you use in eliminating incorrect choices may offer insight into which of the remaining choices is correct.

2. **Calculations in which you must show your work** In questions of this kind, you may receive partial credit even if you do not arrive at the correct answer, depending on whether the instructor can follow your line of reasoning. It is important, therefore, to be neat and organized in your calculations. Pay particular attention to what information is given and to what your unknown is. Think about how you can get from the given information to your unknown.

   You may want to write a few words or a diagram on the test paper to indicate your approach. Then write out your calculations as neatly as you can. Show the units for every number you write down, and use dimensional analysis as much as you can, showing how units cancel.

3. **Questions requiring drawings** Questions of this kind will come later in the course, but it is useful to talk about them here. (You should review this box before each exam to remind yourself of good exam-taking practices.) Be sure to label your drawing as completely as possible.

Self-Assessment Exercise

2.46 Do isomers have the same molecular formula?

(a) Yes

(b) No

Exercises

2.47 (**a**) What is a hydrocarbon? (**b**) Pentane is the alkane with a chain of five carbon atoms. Write a structural formula for this compound and determine its molecular and empirical formulas.

2.48 (**a**) What is a functional group? (**b**) What functional group characterizes an alcohol? (**c**) Write a structural formula for 1-pentanol, the alcohol derived from pentane by making a substitution on one of the carbon atoms.

2.49 Chloropropane is derived from propane by substituting Cl for H on one of the carbon atoms. (**a**) Draw the structural formulas for the two isomers of chloropropane. (**b**) Suggest names for these two compounds.

Answers to Self-Assessment Exercises

2.45 (a)

Chapter Summary and Key Terms

THE ATOMIC THEORY OF MATTER; THE DISCOVERY OF ATOMIC STRUCTURE (SECTION 2.1 AND 2.2) **Atoms** are the basic building blocks of matter. They are the smallest units of an element that can combine with other elements. Atoms are composed of even smaller particles, called **subatomic particles.** Some of these subatomic particles are charged and follow the usual behavior of charged particles: Particles with the same charge repel one another, whereas particles with opposite charges are attracted to one another.

We considered some of the important experiments that led to the discovery and characterization of subatomic particles. Thomson's experiments on the behavior of **cathode rays** in magnetic and electric fields led to the discovery of the electron and allowed its charge-to-mass ratio to be measured. Millikan's oil-drop experiment determined the charge of the electron. Becquerel's discovery of **radioactivity**, the spontaneous emission of radiation by atoms, gave further evidence that the atom has a substructure. Rutherford's studies of how α particles scatter when passing through thin metal foils led to the **nuclear model** of the atom, showing that the atom has a dense, positively charged **nucleus**.

THE MODERN VIEW OF ATOMIC STRUCTURE (SECTION 2.3) Atoms have a nucleus that contains **protons** and **neutrons**; **electrons** move in the space around the nucleus. The magnitude of the charge of the electron, 1.602×10^{-19} C, is called the **electronic charge**. The charges of particles are usually represented as multiples of this charge—an electron has a 1− charge, and a proton has a 1+ charge. The masses of atoms are usually expressed in terms of **atomic mass units** ($1 \text{ u} = 1.66054 \times 10^{-24}$ g).

Elements can be classified by **atomic number**, the number of protons in the nucleus of an atom. All atoms of a given element have the same atomic number. The **mass number** of an atom is the sum of the numbers of protons and neutrons. Atoms of the same element that differ in mass number are known as **isotopes**.

ATOMIC WEIGHTS (SECTION 2.4) The atomic mass scale is defined by assigning a mass of exactly 12 u to a ^{12}C atom. The **atomic weight** (average atomic mass) of an element can be calculated from the relative abundances and masses of that element's isotopes. The **mass spectrometer** provides the most direct and accurate means of experimentally measuring atomic (and molecular) weights.

THE PERIODIC TABLE (SECTION 2.5) The **periodic table** is an arrangement of the elements in order of increasing atomic number. Elements with similar properties are placed in vertical columns. The elements in a column are known as a **group**. The elements in a horizontal row are known as a **period**. The **metallic elements** (**metals**), which comprise the majority of the elements, dominate the left side and the middle of the table; the **nonmetallic elements (nonmetals)** are located on the upper right side. Many of the elements that lie along the line that separates metals from nonmetals are **metalloids**.

MOLECULES AND MOLECULAR COMPOUNDS (SECTION 2.6) Atoms can combine to form **molecules**. Compounds composed of molecules (**molecular compounds**) usually contain only nonmetallic elements. A molecule that contains two atoms is called a **diatomic molecule**. The composition of a substance is given by its **chemical formula**. A molecular substance can be represented by its **empirical formula**, which gives the relative numbers of atoms of each kind, but is usually represented by its **molecular formula**, which gives the actual numbers of each type of atom in a molecule. **Structural formulas** show the order in which the atoms in a molecule are connected. **Ball-and-stick models** and **space-filling models** convey additional information about the shapes of molecules.

IONS AND IONIC COMPOUNDS (SECTION 2.7) Atoms can either gain or lose electrons, forming charged particles called **ions**. Metals tend to lose electrons, becoming positively charged ions (**cations**). Nonmetals tend to gain electrons, forming negatively charged ions (**anions**). Because **ionic compounds** are electrically neutral, containing both cations and anions, they usually contain both metallic and nonmetallic elements. Atoms that are joined together, as in a molecule, but carry a net charge are called **polyatomic ions**. The chemical formulas used for ionic compounds are empirical formulas, which can be written readily if the charges of the ions are known. The total positive charge of the cations in an ionic compound must equal the total negative charge of the anions.

NAMING INORGANIC COMPOUNDS (SECTION 2.8) The set of rules for naming chemical compounds is called **chemical nomenclature**. We studied the systematic rules used for naming three classes of inorganic substances: ionic compounds, acids, and binary molecular compounds.

In naming an ionic compound, the cation is named first and then the anion. Cations formed from metal atoms have the same name as the metal. If the metal can form cations of differing charges, the charge is given using Roman numerals. Monatomic anions have names ending in *-ide*. Polyatomic anions containing oxygen and another element (**oxyanions**) have names ending in *-ate* or *-ite*. In naming binary molecular compounds, Greek prefixes are used to denote the number of each element in the molecular formula, and the element farthest to the left in the periodic table (closest to metallic elements) is generally written first.

SOME SIMPLE ORGANIC COMPOUNDS (SECTION 2.9) **Organic chemistry** is the study of compounds that contain carbon. The simplest class of organic molecules is the **hydrocarbons**, which contain only carbon and hydrogen. Hydrocarbons in which each carbon atom is attached to four other atoms are called **alkanes**. Alkanes have names that end in *-ane*, such as methane and ethane. Other organic compounds are formed when an H atom of a hydrocarbon is replaced with a functional group. An **alcohol**, for example, is a compound in which an H atom of a hydrocarbon is replaced by an OH functional group. Alcohols have names that end in *-ol*, such as methanol and ethanol. Compounds with the same molecular formula but different bonding arrangements of their constituent atoms are called **isomers**.

Learning Outcomes After studying this chapter, you should be able to:

- List the basic postulates of Dalton's atomic theory. (Section 2.1) *Related Exercises: 2.2, 2.3, 2.60, 2.61*
- Describe the key experiments that led to the discovery of electrons and to the nuclear model of the atom. (Section 2.2) *Related Exercises: 2.5, 2.6, 2.62, 2.63*
- Describe the structure of the atom in terms of protons, neutrons, and electrons and express the relative electrical charges and masses of these subatomic particles. (Section 2.3) *Related Exercises: 2.10, 2.65*
- Use chemical symbols together with atomic number and mass number to express the subatomic composition of isotopes. (Section 2.3) *Related Exercises: 2.11, 2.12, 2.14, 2.33*
- Calculate the atomic weight of an element from the masses of individual atoms and a knowledge of natural abundances. (Section 2.4) *Related Exercises: 2.18, 2.20, 2.71, 2.72*
- Describe how elements are organized in the periodic table by atomic number and by similarities in chemical behavior, giving rise to periods and groups. (Section 2.5) *Related Exercises: 2.23, 2.76*
- Identify the locations of metals and nonmetals in the periodic table. (Section 2.5) *Related Exercises: 2.22, 2.75*
- Distinguish between molecular substances and ionic substances in terms of their composition. (Section 2.6 and 2.7) *Related Exercises: 2.32, 2.36, 2.37, 2.87*
- Distinguish between empirical formulas and molecular formulas. (Section 2.6) *Related Exercises: 2.25, 2.26, 2.27, 2.77*
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules. (Section 2.6) *Related Exercises: 2.29, 2.81*
- Explain how ions are formed by the gain or loss of electrons and use the periodic table to predict the charges of common ions. (Section 2.7) *Related Exercises: 2.34, 2.82, 2.83*
- Write the empirical formulas of ionic compounds, given the charges of their component ions. (Section 2.7) *Related Exercises: 2.35, 2.36, 2.37, 2.85*
- Write the name of an ionic compound given its chemical formula or write the chemical formula given its name. (Section 2.8) *Related Exercises: 2.41, 2.42, 2.90, 2.91*
- Name or write chemical formulas for binary inorganic compounds and for acids. (Section 2.8) *Related Exercises: 2.43, 2.44, 2.92, 2.93*
- Identify organic compounds and name simple alkanes and alcohols. (Section 2.9) *Related Exercises: 2.47, 2.48, 2.49, 2.96*

Key Equations

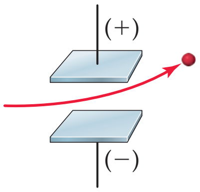
- $$\text{Atomic weight} = \sum_{\substack{\text{over} \\ \text{all isotopes}}} [(\text{isotope mass}) \times (\text{fractional isotope abundance})] \quad [2.1]$$
  Calculating atomic weight as a fractionally weighted average of isotopic masses.

Exercises

Visualizing Concepts

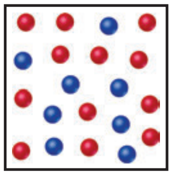
These exercises are intended to probe your understanding of key concepts rather than your ability to utilize formulas and perform calculations. Exercises with red numbers have answers in the back of the book.

2.50 A charged particle moves between two electrically charged plates, as shown here.

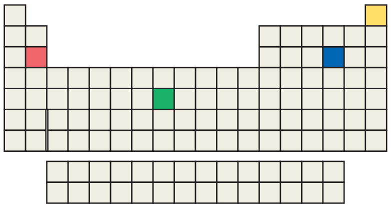


(**a**) What is the sign of the electrical charge on the particle? (**b**) As the charge on the plates is increased, would you expect the bending to increase, decrease, or stay the same? (**c**) As the mass of the particle is increased while the speed of the particles remains the same, would you expect the bending to increase, decrease, or stay the same? [Section 2.2]

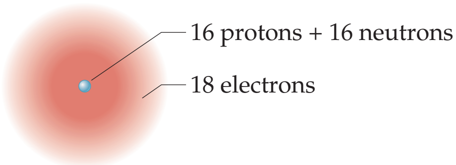
2.51 The following diagram is a representation of 20 atoms of a fictitious element, which we will call nevadium (Nv). The red spheres are ^{293}Nv, and the blue spheres are ^{295}Nv. (**a**) Assuming that this sample is a statistically representative sample of the element, calculate the percent abundance of each element. (**b**) If the mass of ^{293}Nv is 293.15 u and that of ^{295}Nv is 295.15 u, what is the atomic weight of Nv? [Section 2.4]



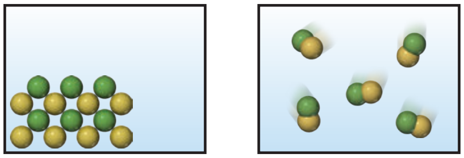
2.52 Four of the boxes in the following periodic table are colored. Which of these are metals and which are nonmetals? Which one is an alkaline earth metal? Which one is a noble gas? [Section 2.5]



2.53 Does the following drawing represent a neutral atom or an ion? Write its complete chemical symbol, including mass number, atomic number, and net charge (if any). [Sections 2.3 and 2.7]

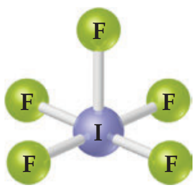


2.54 Which of the following diagrams most likely represents an ionic compound, and which represents a molecular one? Explain your choice. [Sections 2.6 and 2.7]

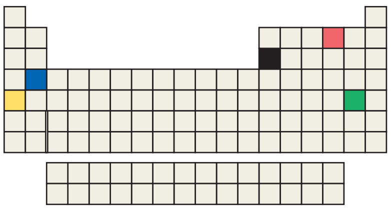


(i) (ii)

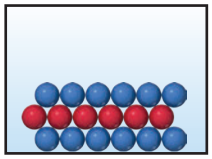
2.55 Write the chemical formula for the following compound. Is the compound ionic or molecular? Name the compound. [Sections 2.6 and 2.8]



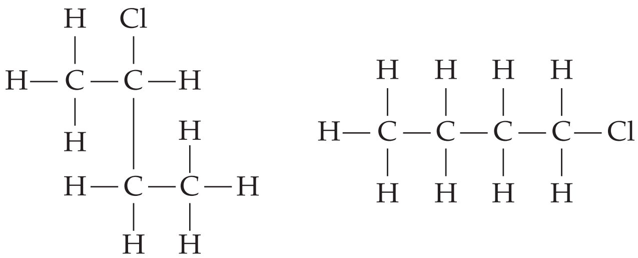
2.56 Five of the boxes in the following periodic table are colored. Predict the charge on the ion associated with each of these elements. [Section 2.7]



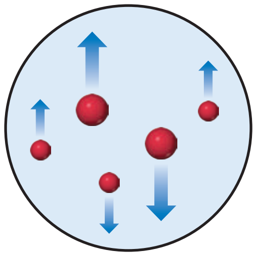
2.57 The following diagram represents an ionic compound in which the red spheres represent cations and the blue spheres represent anions. Which of the following formulas is consistent with the drawing? KBr, K_2SO_4, $Ca(NO_3)_2$, $Fe_2(SO_4)_3$. Name the compound. [Sections 2.7 and 2.8]



2.58 Are these two compounds isomers? Explain. [Section 2.9]



2.59 In the Millikan oil-drop experiment (see Figure 2.4), the tiny oil drops are observed through the viewing lens as rising, stationary, or falling, as shown here. (**a**) What causes their rate of fall to vary from their rate in the absence of an electric field? (**b**) Why do some drops move upward? [Section 2.2]



The following exercises are divided into sections that deal with specific topics in the chapter. The exercises are grouped in pairs, with the answers given in the back of the book to the odd-numbered exercises, as indicated by the red exercise numbers. Those exercises whose numbers appear in brackets are more challenging than the nonbracketed exercises.

The Atomic Theory of Matter and the Discovery of Atomic Structure (Sections 2.1 and 2.2)

2.60 Sodium reacts with oxygen in air to form two compounds: sodium oxide and sodium peroxide. In forming sodium oxide, 23.0 g of sodium combines with 8.0 g of hydrogen. In forming sodium peroxide, 23.0 g of sodium combines with 16.0 g of oxygen. (**a**) What are the mass ratios of oxygen in the two compounds? (**b**) What fundamental law does this experiment demonstrate?

2.61 In a series of experiments, a chemist prepared three different compounds that contain only iodine and fluorine and determined the mass of each element in each compound:

Compound	Mass of Iodine (g)	Mass of Fluorine (g)
1	4.75	3.56
2	7.64	3.43
3	9.41	9.86

(**a**) Calculate the mass of fluorine per gram of iodine in each compound. (**b**) How do the numbers in part (a) support the atomic theory?

2.62 Discovering which of the three subatomic particles proved to be the most difficult—the proton, neutron, or electron? Why?

2.63 Millikan determined the charge on the electron by studying the static charges on oil drops falling in an electric field (Figure 2.4). A student carried out this experiment using several oil drops for her measurements and calculated the charges on the drops. She obtained the following data:

Droplet	Calculated Charge (C)
A	1.60×10^{-19}
B	3.15×10^{-19}
C	4.81×10^{-19}
D	6.31×10^{-19}

(**a**) What is the significance of the fact that the droplets carried different charges? (**b**) What conclusion can the student draw from these data regarding the charge of the electron? (**c**) What value (and to how many significant figures) should she report for the electronic charge?

The Modern View of Atomic Structure; Atomic Weights (Sections 2.3 and 2.4)

2.64 The radius of an atom of copper (Cu) is about 140 pm. (**a**) Express this distance in millimeters (mm). (**b**) How many Cu atoms would have to be placed side by side to span a distance of 5.0 mm? (**c**) If you assume that the Cu atom is a sphere, what is the volume in cm^3 of a single atom?

2.65 Determine whether each of the following statements is true or false. If false, correct the statement to make it true: (**a**) The nucleus has most of the mass and comprises most of the volume of an atom. (**b**) Every atom of a given element has the same number of protons. (**c**) The number of electrons in an atom equals the number of neutrons in the atom. (**d**) The protons in the nucleus of the helium atom are held together by a force called the strong nuclear force.

2.66 Consider an atom of ^{58}Ni. (**a**) How many protons, neutrons, and electrons does this atom contain? (**b**) What is the symbol of the ion obtained by removing two electrons from ^{58}Ni? (**c**) What is the symbol for the isotope of ^{58}Ni that possesses 33 neutrons?

2.67 (**a**) Which two of the following are isotopes of the same element: $^{106}_{46}X$, $^{107}_{46}X$, $^{107}_{47}X$? (**b**) What is the identity of the element whose isotopes you have selected?

2.68 Each of the following isotopes is used in medicine. Indicate the number of protons and neutrons in each isotope: (**a**) samarium-153, (**b**) lutetium-177, (**c**) bismuth-213, (**d**) molybdenum-99, (**e**) lead-212, (**f**) caesium-131.

2.69 Fill in the gaps in the following table, assuming each column represents a neutral atom.

Symbol	^{89}Y				
Protons		78			89
Neutrons			123		
Electrons			81	50	
Mass no.		195		119	227

2.70 One way in which Earth's evolution as a planet can be understood is by measuring the amounts of certain isotopes in rocks. One quantity recently measured is the ratio of ^{129}Xe to ^{130}Xe in some minerals. In what way do these two isotopes differ from one another? In what respects are they the same?

2.71 (**a**) What is the mass in u of a carbon-12 atom? (**b**) Why is the atomic weight of carbon reported as 12.011 in the table of elements and the periodic table in the front inside cover of this text?

2.72 Bromine has two naturally occurring isotopes, bromine-79 (atomic mass = 78.9183 u; abundance = 50.69%) and bromine-81 (atomic mass = 80.9163 u; abundance = 49.31%). Calculate the atomic weight of bromine.

2.73 Consider the mass spectrometer shown in Figure 2.10. Determine whether each of the following statements is true or false. If false, correct the statement to make it true: (**a**) The paths of neutral (uncharged) atoms are not affected by the magnet. (**b**) The height of each peak in the mass spectrum is inversely proportional to the mass of that isotope. (**c**) For a given element, the number of peaks in the spectrum is equal to the number of naturally occurring isotopes of that element.

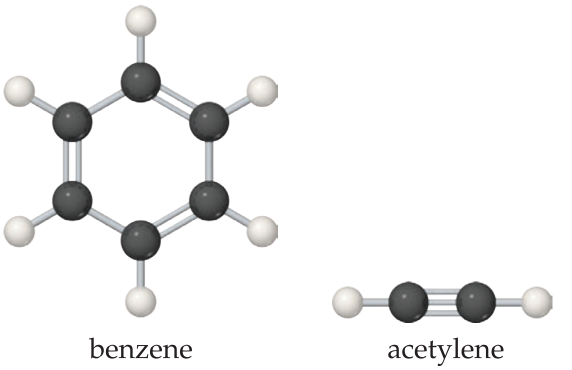
2.74 Mass spectrometry is more often applied to molecules than to atoms. We will see in Chapter 3 that the *molecular weight* of a molecule is the sum of the atomic weights of the atoms in the molecule. The mass spectrum of H_2 is taken under conditions that prevent decomposition into H atoms. The two naturally occurring isotopes of hydrogen are ^{1}H (atomic mass = 1.00783 u; abundance 99.9885%) and ^{2}H (atomic mass = 2.01410 u; abundance 0.0115%). (**a**) How many peaks will the mass spectrum have? (**b**) Give the relative atomic masses of each of these peaks. (**c**) Which peak will be the largest, and which the smallest?

The Periodic Table, Molecules, Molecular Compounds, Ions, and Ionic Compounds (Sections 2.5, 2.6, and 2.7)

2.75 Locate each of the following elements in the periodic table; give its name and atomic number, and indicate whether it is a metal, metalloid, or nonmetal: (**a**) Hg, (**b**) At, (**c**) Mo, (**d**) W, (**e**) Sn, (**f**) V, (**g**) K.

2.76 For each of the following elements, write its chemical symbol, determine the name of the group to which it belongs (Table 2.3), and indicate whether it is a metal, metalloid, or nonmetal: (**a**) polonium, (**b**) strontium, (**c**) neon, (**d**) rubidium, (**e**) bromine.

2.77 Ball-and-stick representations of benzene, a colorless liquid often used in organic chemistry reactions, and acetylene, a gas used as a fuel for high-temperature welding, are shown here. (**a**) Determine the molecular formula of each. (**b**) Determine the empirical formula of each.

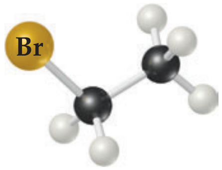


2.78 Two substances have the same molecular and empirical formulas. Does this mean that they must be the same compound?

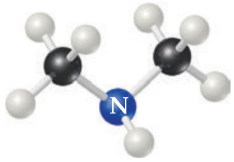
2.79 Determine the molecular and empirical formulas of the following: (**a**) the organic solvent *benzene*, which has six carbon atoms and six hydrogen atoms; (**b**) the compound *silicon tetrachloride*, which has a silicon atom and four chlorine atoms and is used in the manufacture of computer chips; (**c**) the reactive substance *diborane*, which has two boron atoms and six hydrogen atoms; (**d**) the sugar called *glucose,* which has six carbon atoms, twelve hydrogen atoms, and six oxygen atoms.

2.80 How many of the indicated atoms are represented by each chemical formula: (**a**) carbon atoms in $C_4H_9COOCH_3$, (**b**) oxygen atoms in $Ca(ClO_3)_2$, (**c**) hydrogen atoms in $(NH_4)_2HPO_4$?

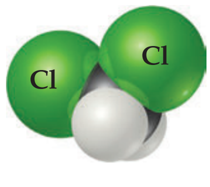
2.81 Write the molecular and structural formulas for the compounds represented by the following models:



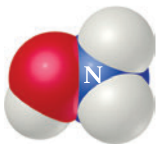
(a)



(b)



(c)



(d)

2.82 Fill in the gaps in the following table:

Symbol	$^{133}Cs^+$			
Protons		35	15	
Neutrons		46	16	30
Electrons			18	20
Net charge		1−		5+

2.83 Using the periodic table, predict the charge of the most stable ion of the following elements: (**a**) Li, (**b**) Ba, (**c**) Po, (**d**) I, (**e**) Sb.

2.84 The most common charge associated with selenium is 2−. Indicate the chemical formulas you would expect for compounds formed between selenium and (**a**) barium, (**b**) lithium, (**c**) aluminum.

2.85 Predict the chemical formulas of the compounds formed by the following pairs of ions: (**a**) Cr^{3+} and CN^-, (**b**) Mn^{2+} and ClO_4^-, (**c**) Na^+ and $Cr_2O_7^{2-}$, (**d**) Cd^{2+} and CO_3^{2-}, (**e**) Ti^{4+} and O^{2-}.

2.86 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	Na^+	Ca^{2+}	Fe^{2+}	Al^{3+}
O^{2-}	Na_2O			
NO_3^-				
SO_4^{2-}				
AsO_4^{3-}				

2.87 Predict whether each of the following compounds is molecular or ionic: (**a**) BI_3 (**b**) $N(CH_3)_3$ (**c**) $Zr(NO_3)_2$ (**d**) N_2H_4 (**e**) $OsCO_3$ (**f**) H_2SO_4 (**g**) HgS (**h**) IOH.

Naming Inorganic Compounds; Some Simple Organic Compounds (Sections 2.8 and 2.9)

2.88 Selenium, an element required nutritionally in trace quantities, forms compounds analogous to sulfur. Name the following ions: (**a**) SeO_4^{2-}, (**b**) Se^{2-}, (**c**) HSe^-, (**d**) $HSeO_3^-$.

2.89 Give the names and charges of the cation and anion in each of the following compounds: (**a**) CuS, (**b**) Ag_2SO_4, (**c**) $Al(ClO_3)_3$, (**d**) $Co(OH)_2$, (**e**) $PbCO_3$.

2.90 Name the following ionic compounds: (**a**) KCN, (**b**) $NaBrO_2$, (**c**) $Sr(OH)_2$, (**d**) CoTe, (**e**) $Fe_2(CO_3)_3$, (**f**) $Cr(NO_3)_3$, (**g**) $(NH_4)_2SO_3$, (**h**) NaH_2PO_4, (**i**) $KMnO_4$, (**j**) $Ag_2Cr_2O_7$.

2.91 Give the chemical formula for each of the following ionic compounds: (**a**) sodium phosphate, (**b**) zinc nitrate, (**c**) barium bromate, (**d**) iron(II) perchlorate, (**e**) cobalt(II) hydrogen carbonate, (**f**) chromium(III) acetate, (**g**) potassium dichromate.

2.92 Provide the name or chemical formula, as appropriate, for each of the following acids: (**a**) hydroiodic acid, (**b**) chloric acid, (**c**) nitrous acid, (**d**) H_2CO_3, (**e**) $HClO_4$, (**f**) CH_3COOH.

2.93 The oxides of nitrogen are very important components in urban air pollution. Name each of the following compounds: (**a**) N_2O, (**b**) NO, (**c**) NO_2, (**d**) N_2O_5, (**e**) N_2O_4.

2.94 Assume that you encounter the following sentences in your reading. What is the chemical formula for each substance mentioned? (**a**) Sodium hydrogen carbonate is used as a deodorant. (**b**) Calcium hypochlorite is used in some bleaching solutions. (**c**) Hydrogen cyanide is a very poisonous gas. (**d**) Magnesium hydroxide is used as a cathartic. (**e**) Tin(II) fluoride has been used as a fluoride additive in toothpastes. (**f**) When cadmium sulfide is treated with sulfuric acid, fumes of hydrogen sulfide are given off.

2.95 (**a**) What is meant by the term *isomer*? (**b**) Among the four alkanes, ethane, propane, butane, and pentane, which is capable of existing in isomeric forms?

2.96 Consider the following organic substances: ethylethanoate, ethylmethylether, hexanol, and propanone. (**a**) Which of these molecules contains three carbons? (**b**) Which of these molecules contain a C = O group?

2.97 Draw the structural formulas for four structural isomers of C_4H_9Br.

Additional Exercises

These exercises are not divided by category, although they are roughly in the order of the topics in the chapter.

2.98 Suppose a scientist repeats the Millikan oil-drop experiment but reports the charges on the drops using an unusual (and imaginary) unit called the *warmomb* (wa). The scientist obtains the following data for four of the drops:

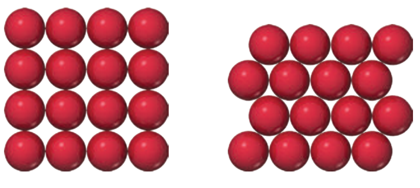
Droplet	Calculated Charge (wa)
A	3.84×10^{-8}
B	4.80×10^{-8}
C	2.88×10^{-8}
D	8.64×10^{-8}

(**a**) If all the droplets were the same size, which would fall most slowly through the apparatus? (**b**) From these data, what is the best choice for the charge of the electron in warmombs? (**c**) Based on your answer to part (b), how many electrons are there on each of the droplets? (**d**) What is the conversion factor between warmombs and coulombs?

2.99 The natural abundance of ^{3}He is 0.000137%. (**a**) How many protons, neutrons, and electrons are in an atom of ^{3}He? (**b**) Based on the sum of the masses of their subatomic particles, which is expected to be more massive, an atom of ^{3}He or an atom of ^{3}H (which is also called *tritium*)? (**c**) Based on your answer to part (b), what would need to be the precision of a mass spectrometer that is able to differentiate between peaks that are due to $^3He^+$ and $^3H^+$?

2.100 A cube of gold that is 1.00 cm on a side has a mass of 19.3 g. A single gold atom has a mass of 197.0 u. (**a**) How many gold atoms are in the cube? (**b**) From the information given, estimate the diameter in Å of a single gold atom. (**c**) What assumptions did you make in arriving at your answer for part (**b**)?

2.101 The diameter of a rubidium atom is 495 pm We will consider two different ways of placing the atoms on a surface. In arrangement A, all the atoms are lined up with one another to form a square grid. Arrangement B is called a *close-packed* arrangement because the atoms sit in the "depressions" formed by the previous row of atoms:



A B

(**a**) Using arrangement A, how many Rb atoms could be placed on a square surface that is 1.0 cm on a side? (**b**) How many Rb atoms could be placed on a square surface that is 1.0 cm on a side, using arrangement B? (**c**) By what factor has the number of atoms on the surface increased in going to arrangement B from arrangement A? If extended to three dimensions, which arrangement would lead to a greater density for Rb metal?

2.102 (**a**) Assuming the dimensions of the nucleus and atom shown in Figure 2.9, what fraction of the *volume* of the atom is taken up by the nucleus? (**b**) Using the mass of the proton from Table 2.1 and assuming its diameter is 1.0×10^{-15} m, calculate the density of a proton in g/cm^3.

2.103 Identify the element represented by each of the following symbols and give the number of protons and neutrons in each: (**a**) $^{11}_{5}X$ (**b**) $^{75}_{33}X$ (**c**) $^{86}_{36}X$ (**d**) $^{67}_{30}X$.

2.104 The nucleus of ^{6}Li is a powerful absorber of neutrons. It exists in the naturally occurring metal to the extent of 7.5%. In the era of nuclear deterrence, large quantities of lithium were processed to remove ^{6}Li for use in hydrogen bomb production. The lithium metal remaining after removal of ^{6}Li was sold on the market. (**a**) What are the compositions of the nuclei of ^{6}Li and ^{7}Li? (**b**) The atomic masses of ^{6}Li and ^{7}Li are 6.015122 and 7.016004 u, respectively. A sample of lithium depleted in the lighter isotope was found on analysis to contain 1.442% ^{6}Li. What is the average atomic weight of this sample of the metal?

2.105 The element argon has three naturally occurring isotopes, with 18, 20, and 22 neutrons in the nucleus, respectively. (**a**) Write the full chemical symbols for these three isotopes. (**b**) Describe the similarities and differences between the three kinds of atoms of argon.

2.106 The element chromium (Cr) consists of four naturally occurring isotopes with atomic masses 49.9460, 51.9405, 52.9407, and 53.9389 u. The relative abundances of these four isotopes are 4.3, 83.8, 9.5, and 2.4%, respectively. From these data, calculate the atomic weight of chromium.

2.107 Copper (Cu) consists of two naturally occurring isotopes with masses of 62.9296 and 64.9278 u. (**a**) How many protons and neutrons are in the nucleus of each isotope? Write the complete atomic symbol for each, showing the atomic number and mass number. (**b**) The average atomic mass of Cu is 63.55 u. Calculate the abundance of each isotope.

2.108 Using a suitable reference such as the *CRC Handbook of Chemistry and Physics* or http://www.webelements.com, look up the following information for nickel: (**a**) the number of known isotopes, (**b**) the atomic masses (in u), (**c**) the natural abundances of the five most abundant isotopes.

2.109 There are two different isotopes of bromine atoms. Under normal conditions, elemental bromine consists of Br_2 molecules, and the mass of a Br_2 molecule is the sum of the masses of the two atoms in the molecule. The mass spectrum of Br_2 consists of three peaks:

Mass (u)	Relative Size
157.836	0.2569
159.834	0.4999
161.832	0.2431

(**a**) What is the origin of each peak (of what isotopes does each consist)? (**b**) What is the mass of each isotope? (**c**) Determine the average molecular mass of a Br_2 molecule. (**d**) Determine the average atomic mass of a bromine atom. (**e**) Calculate the abundances of the two isotopes.

2.110 It is common in mass spectrometry to assume that the mass of a cation is the same as that of its parent atom. (**a**) Using data in Table 2.1, determine the number of significant figures that must be reported before the difference in masses of ^{1}H and $^1H^+$ is significant. (**b**) What percentage of the mass of an ^{1}H atom does the electron represent?

2.111 From the following list of elements—Mg, Li, Tl, Pb, Se, Cl, Xe, Si, C—pick the one that best fits each description. Use each element only once: (**a**) an alkali metal, (**b**) an alkaline earth metal, (**c**) a noble gas, (**d**) a halogen, (**e**) a metalloidin Group 14, (**f**) a nonmetal listed in Group 14, (**g**) a metal that forms a 3+ ion, (**h**) a nonmetal that forms a 2− ion, (**i**) an element that is used as radiation shielding.

2.112 The first atoms of seaborgium (Sg) were identified in 1974. The longest-lived isotope of Sg has a mass number of 266. (**a**) How many protons, electrons, and neutrons are in an ^{266}Sg atom? (**b**) Atoms of Sg are very unstable, and it is therefore difficult to study this element's properties. Based on the position of Sg in the periodic table, what element should it most closely resemble in its chemical properties?

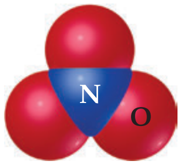
2.113 The explosion of an atomic bomb releases many radioactive isotopes, including strontium-90. Considering the location of strontium in the periodic table, suggest a reason for the fact that this isotope is particularly dangerous for human health.

2.114 A U.S. 1-cent coin (a penny) has a diameter of 19 mm and a thickness of 1.5 mm. Assume the coin is made of pure copper, whose density and approximate market price are 8.9 g/cm^3 and \$2.40 per pound, respectively. Calculate the value of the copper in the coin, assuming its thickness is uniform.

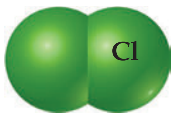
2.115 The U.S. Mint produces a dollar coin called the American Silver Eagle that is made of nearly pure silver. This coin has a diameter of 41 mm and a thickness of 2.5 mm. The density and approximate market price of silver are 10.5 g/cm^3 and

\$0.51 per gram, respectively. Calculate the value of the silver in the coin, assuming its thickness is uniform.

2.116 From the molecular structures shown here, identify the one that corresponds to each of the following species: (**a**) chlorine gas; (**b**) propane; (**c**) nitrate ion; (**d**) sulfur trioxide; (**e**) methyl chloride, CH_3Cl.



(i)



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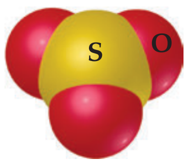
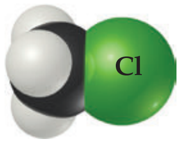


(iii)



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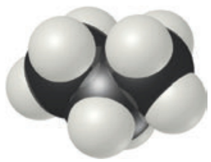


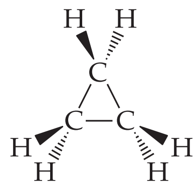
(v)

2.117 Name each of the following chlorides. Assuming that the compounds are ionic, what charge is associated with the metallic element in each case? (**a**) AgCl, (**b**) $TiCl_4$, (**c**) $IrC1_3$, (**d**) LiCl.

2.118 Fill in the blanks in the following table:

Cation	Anion	Formula	Name
			Sodium carbonate
Ni^{2+}	CH_3COO-		
		$Cu(ClO_4)_2$	
Ca^{2+}	F^-		
		$NaMnO_4$	
			Zinc sulfide
		Mg_3N_2	

2.119 Cyclopropane is an interesting hydrocarbon. Instead of having three carbons in a row, the three carbons form a ring, as shown in this perspective drawing (see Figure 2.16 for a prior example of this kind of drawing):



Cyclopropane was at one time used as an anesthetic, but its use was discontinued, in part because it is highly flammable. (**a**) What is the empirical formula of cyclopropane? How does it differ from that of propane? (**b**) The three carbon atoms are necessarily in a plane. What do the different wedges mean? (**c**) What change would you make to the structure shown to illustrate chlorocyclopropane? Are there isomers of chlorocyclopropane?

2.120 Elements in the same group of the periodic table often form oxyanions with the same general formula. The anions are also named in a similar fashion. Based on these observations, suggest a chemical formula or name, as appropriate, for each of the following ions: (**a**) BrO_4^-, (**b**) SeO_3^{2-}, (**c**) arsenate ion, (**d**) hydrogen tellurate ion.

2.121 Carbonic acid occurs in carbonated beverages. When allowed to react with lithium hydroxide, it produces lithium carbonate. Lithium carbonate is used to treat depression and bipolar disorder. Write chemical formulas for carbonic acid, lithium hydroxide, and lithium carbonate.

2.122 Give the chemical names of each of the following familiar compounds: (**a**) NaCl (table salt), (**b**) $NaHCO_3$ (baking soda), (**c**) NaOCl (in many bleaches), (**d**) NaOH (caustic soda), (**e**) $(NH_4)_2CO_3$ (smelling salts), (**f**) $CaSO_4$ (plaster of Paris).

2.123 Many familiar substances have common, unsystematic names. For each of the following, give the correct systematic name: (**a**) saltpeter, KNO_3; (**b**) soda ash, Na_2CO_3; (**c**) lime, CaO; (**d**) muriatic acid, HCl; (**e**) Epsom salts, $MgSO_4$; (**f**) milk of magnesia, $Mg(OH)_2$.

2.124 Because many ions and compounds have very similar names, there is great potential for confusing them. Write the correct chemical formulas to distinguish between (**a**) sodium carbonate and sodium bicarbonate, (**b**) potassium peroxide and potassium oxide, (**c**) calcium sulfide and calcium sulfate, (**d**) manganese (II) oxide and manganese (III) oxide, (**e**) hydride ion and hydroxide ion, (**f**) magnesium nitride and magnesium nitrite, (**g**) silver nitrate and silver nitrite, (**h**) cuprous oxide and cupric oxide.

2.125 In what part of the atom does the strong nuclear force operate?

WHAT'S AHEAD

3 CHEMICAL REACTIONS AND STOICHIOMETRY

3.1 | The Conservation of Mass, Chemical Equations, and Stoichiometry

Much of the energy that we use in our daily lives, including transportation, comes from chemical reactions. In the internal combustion engine, fuel and air react to produce energy, along with gaseous byproducts (mostly carbon dioxide and water) that exit the exhaust pipes of vehicles, an issue that is central to discussions surrounding climate change. Suppose we wanted to know how many molecules of CO_2 are produced by burning a certain amount of fuel. The tools of chemistry we will learn in this section enable us to accurately determine how many molecules of hydrocarbon and oxygen are consumed and how many molecules of byproducts are generated.

When you finish this section, you should be able to:

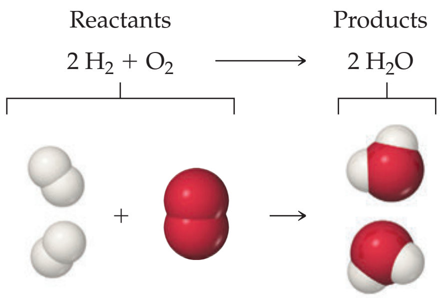
- Explain the law of conservation of mass in terms of reactants and products in a chemical equation
- Balance chemical equations by writing out the chemical formulas (and appropriate coefficients) of the reactants and products in chemical reactions

Stoichiometry (pronounced stoy-key-OM-uh-tree, roughly meaning "element measure" in Greek) is the area of study that examines the quantities of substances consumed and produced in chemical reactions. Chemists and chemical engineers use stoichiometry every day in running the reactions of the worldwide chemical industry.

Stoichiometry is built on an understanding of atomic weights (Section 2.4), chemical formulas, and the **law of conservation of mass** (Section 2.1). This important principle tells us that: *Atoms are neither created nor destroyed during a chemical reaction*. The changes that occur during any reaction merely rearrange the atoms. The same collection of atoms is present both before and after the reaction.

We represent chemical reactions by **chemical equations**. When the gas hydrogen (H_2) burns, for example, it reacts with oxygen (O_2) in the air to form water, H_2O. We write the chemical equation for this reaction as

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O \qquad [3.1]$$



▲ **Figure 3.1 A balanced chemical equation.**

We read the + sign as "reacts with" and the arrow as "produces." The chemical formulas on the left side of the arrow represent the starting substances, called **reactants**. The chemical formulas on the right side of the arrow represent substances produced in the reaction, called **products**. The numbers in front of the formulas, called coefficients, indicate the relative numbers of molecules of each kind involved in the reaction. (As in algebraic equations, *the coefficient 1 is usually not written*).

Because atoms are neither created nor destroyed in any chemical reaction, a balanced chemical equation must have an equal number of atoms of each element on each side of the arrow. On the right side of Equation 3.1, for example, there are two molecules of H_2O, each composed of two atoms of hydrogen and one atom of oxygen (**Figure 3.1**). Thus, $2\,H_2O$ (read "two molecules of water") contains $2 \times 2 = 4$ H atoms and $2 \times 1 = 2$ O atoms. Notice *that the number of atoms is obtained by multiplying each subscript in a chemical formula by the coefficient for the formula*. Because there are four H atoms and two O atoms on each side of the equation, the equation is balanced.

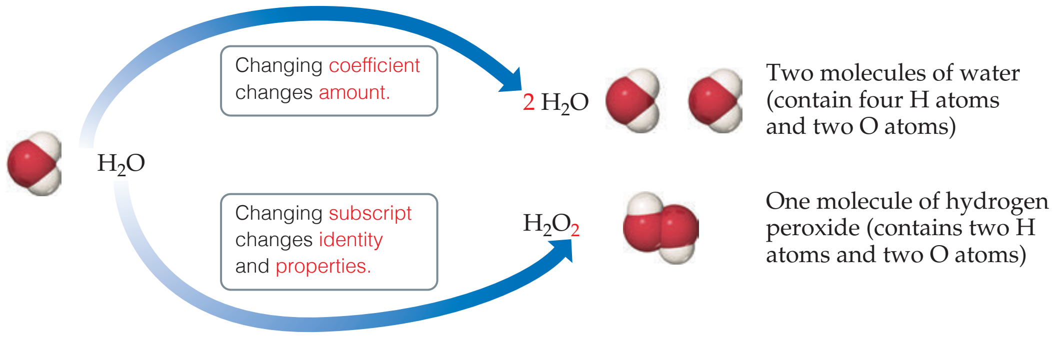
How to Balance Chemical Equations

Chemists write chemical equations to identify the reactants and products in a reaction. To determine the amount of product that can be made, or the amount of a reactant that is required, the equation needs to be *balanced* using stoichiometry.

To construct a **balanced chemical equation**, we start by writing the formulas for the reactants on the left-hand side of the arrow and the products on the right-hand side. We balance the equation by determining the coefficients that provide equal numbers of each type of atom on both sides of the equation. For most purposes, a balanced equation should contain the smallest possible whole-number coefficients.

In balancing an equation, you need to understand the difference between coefficients and subscripts. As **Figure 3.2** illustrates, changing a subscript in a formula—from H_2O to H_2O_2, for example—changes the identity of the substance. The substance H_2O_2, hydrogen peroxide, is quite different from the substance H_2O, water. *Never change subscripts when balancing an equation*. In contrast, placing a coefficient in front of a formula changes only the amount of the substance and not its identity. Thus, $2\,H_2O$ means two molecules of water, $3\,H_2O$ means three molecules of water, and so forth.

▶ Figure 3.2 The difference between changing subscripts and changing coefficients in chemical equations.



A Step-by-Step Example of Balancing a Chemical Equation

To illustrate the process of balancing an equation, consider the reaction that occurs when methane (CH_4), the principal component of natural gas, burns in air to produce carbon dioxide gas (CO_2) and water vapor (H_2O) (**Figure 3.3**). Both products contain oxygen atoms that come from O_2 in the air. Thus, O_2 is a reactant, and the unbalanced equation is:

$$CH_4 + O_2 \longrightarrow CO_2 + H_2O \quad \text{(unbalanced)} \qquad [3.2]$$

It is usually best to first balance those elements that occur in the fewest chemical formulas in the equation. In our example, C appears in only one reactant, CH_4, and one product, CO_2. The same is true for H (CH_4 and H_2O). Notice, however, that O appears in one reactant (O_2) and two products (CO_2 and H_2O). So, let's begin with C. Because one molecule of CH_4 contains the same number of C atoms as one molecule of CO_2, the coefficients for these substances must be the same in the balanced equation. Therefore, we start by choosing the coefficient 1 (unwritten) for both CH_4 and CO_2.

Next we focus on H. On the left side of the equation we have CH_4, which has four H atoms, whereas on the right side of the equation we have H_2O, containing two H atoms. To balance the H atoms in the equation we place the coefficient 2 in front of H_2O. Now there are four H atoms on each side of the equation:

$$CH_4 + O_2 \longrightarrow CO_2 + 2\,H_2O \quad \text{(unbalanced)} \qquad [3.3]$$

While the equation is now balanced with respect to hydrogen and carbon, it is not yet balanced for oxygen: there are 2 O atoms on the left-hand side, and a total of 4 O atoms on the right-hand side. Adding the coefficient 2 in front of O_2 balances the equation by giving four O atoms on each side:

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O \quad \text{(balanced)} \qquad [3.4]$$

We can provide even more information in a chemical equation: the physical state of the reactants, products, and more details about what conditions are required for the reaction to proceed. We use the symbols (*g*), (*l*), (*s*), and (*aq*) for substances that are gases, liquids, solids, and dissolved in aqueous (water) solution, respectively. Thus, Equation 3.4 is fully written as:

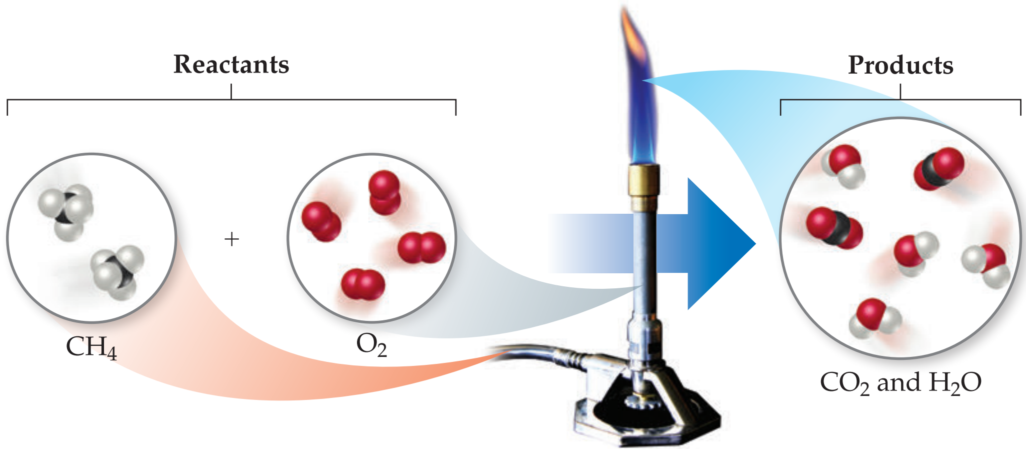
$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g) \qquad [3.5]$$

Symbols that represent the conditions under which the reaction proceeds can appear above or below the reaction arrow. One example that we will encounter later in this chapter involves the symbol Δ (Greek uppercase delta); a delta above the reaction arrow indicates the addition of heat.

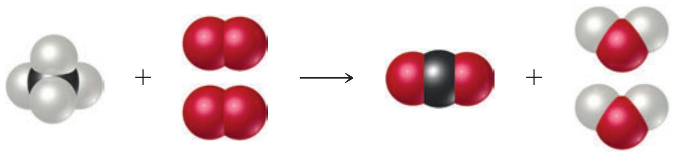
For Equation 3.5, Figures 3.3 and 3.4 provide molecular views of the reaction as it would happen in a Bunsen burner and the balanced reaction, respectively.

Go Figure

In the molecular level views shown in the figure, how many C, H, and O atoms are present as reactants? Are the same number of each type of atom present as products?



▲ **Figure 3.3 Methane reacts with oxygen in a Bunsen burner.**



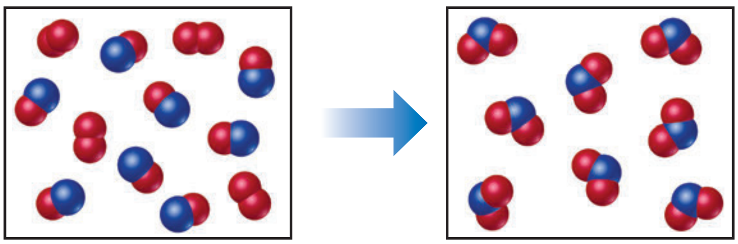
$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$$

$$\underbrace{CH_4 + 2\,O_2}_{(1\,C,\ 4\,H,\ 4\,O)} \qquad \underbrace{CO_2 + 2\,H_2O}_{(1\,C,\ 4\,H,\ 4\,O)}$$

◀ **Figure 3.4 Balanced chemical equation for the combustion of CH_4.**

Sample Exercise 3.1
Interpreting and Balancing Chemical Equations

The following diagram represents a chemical reaction in which the red spheres are oxygen atoms and the blue spheres are nitrogen atoms. (**a**) Write the chemical formulas for the reactants and products. (**b**) Write a balanced equation for the reaction. (**c**) Is the diagram consistent with the law of conservation of mass?



SOLUTION

(**a**) The left box, which represents reactants, contains two kinds of molecules, those composed of two oxygen atoms (O_2) and those composed of one nitrogen atom and one oxygen atom (NO). The right box, which represents products, contains only one kind of molecule, which is composed of one nitrogen atom and two oxygen atoms (NO_2).

(**b**) The unbalanced chemical equation is

$$O_2 + NO \longrightarrow NO_2 \quad \text{(unbalanced)}$$

An inventory of atoms on each side of the equation shows that there are one N and three O on the left side of the arrow and one N and two O on the right. To balance O, we must increase the number of O atoms on the right while keeping the coefficients for NO and NO_2 equal. Sometimes a trial-and-error approach is required; we need to go back and forth several times from one side of an equation to the other, changing coefficients first on one side of the equation and then the other until it is balanced. In our present case, let's start by increasing the number of O atoms on the right side of the equation by placing the coefficient 2 in front of NO_2:

$$O_2 + NO \longrightarrow 2\,NO_2 \quad \text{(unbalanced)}$$

Now the equation gives two N atoms and four O atoms on the right, so we go back to the left side. Placing the coefficient 2 in front of NO balances both N and O:

$$\underbrace{O_2 + 2\,NO}_{(2\,N,\ 4\,O)} \longrightarrow \underbrace{2\,NO_2}_{(2\,N,\ 4\,O)} \quad \text{(balanced)}$$

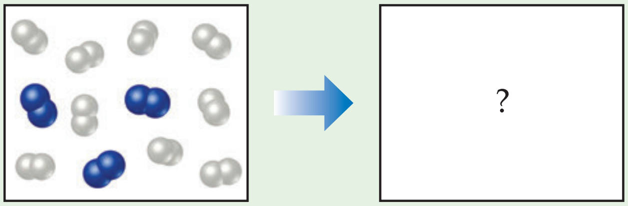
Continued

(**c**) The reactants box contains four O_2 and eight NO. Thus, the molecular ratio is one O_2 for each two NO, as required by the balanced equation. The products box contains eight NO_2, which means the number of NO_2 product molecules equals the number of NO reactant molecules, as the balanced equation requires.

There are eight N atoms in the eight NO molecules in the reactants box. There are also $4 \times 2 = 8$ O atoms in the O_2 molecules and 8 O atoms in the NO molecules, giving a total of 16 O atoms. In the products box, we find eight NO_2 molecules, which contain eight N atoms and $8 \times 2 = 16$ O atoms. Because there are equal numbers of N and O atoms in the two boxes, the drawing is consistent with the law of conservation of mass.

Practice Exercise

In the following diagram, the white spheres represent hydrogen atoms and the blue spheres represent nitrogen atoms.



The two reactants combine to form a single product, ammonia, NH_3, which is not shown. Write a balanced chemical equation for the reaction. Based on the equation and the contents of the left (reactants) box, how many NH_3 molecules should be shown in the right (products) box?
(**a**) 2 (**b**) 3 (**c**) 4 (**d**) 6 (**e**) 9

Sample Exercise 3.2
Balancing Chemical Equations

Balance the equation

$$Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$$

SOLUTION

Begin by counting each kind of atom on the two sides of the arrow. There are one Na, one O, and two H on the left side, and one Na, one O, and three H on the right. The Na and O atoms are balanced, but the number of H atoms is not. To increase the number of H atoms on the left, let's try placing the coefficient 2 in front of H_2O:

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

Although beginning this way does not balance H, it does increase the number of reactant H atoms, which we need to do. (Also, adding the coefficient 2 on H_2O unbalances O, but we will take care of that after we balance H.) Now that we have 2 H_2O on the left, we balance H by putting the coefficient 2 in front of NaOH:

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

Balancing H in this way brings O into balance, but now Na is unbalanced, with one Na on the left and two on the right. To rebalance Na, we put the coefficient 2 in front of the reactant:

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

We now have two Na atoms, four H atoms, and two O atoms on each side. The equation is balanced.

Comment Notice that we moved back and forth, placing a coefficient in front of H_2O, then NaOH, and finally Na. In balancing equations, we often find ourselves following this pattern of moving back and forth from one side of the arrow to the other, placing coefficients first in front of a formula on one side and then in front of a formula on the other side until the equation is balanced. You can always tell if you have balanced your equation correctly by checking that the number of atoms of each element is the same on the two sides of the arrow, and that you've chosen the smallest set of coefficients that balances the equation.

Practice Exercise

Balance these equations by providing the missing coefficients:

(**a**) __ $Fe(s)$ + __ $O_2(g) \longrightarrow$ __ $Fe_2O_3(s)$

(**b**) __ $Al(s)$ + __ $HCl(aq) \longrightarrow$ __ $AlCl_3(aq)$ + __ $H_2(g)$

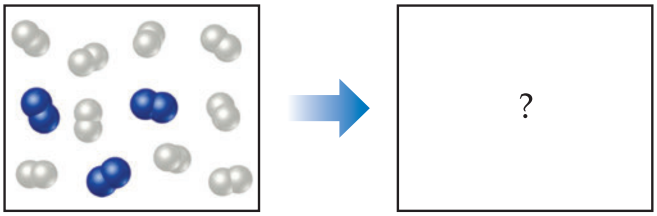
(**c**) __ $CaCO_3(s)$ + __ $HCl(aq) \longrightarrow$ __ $CaCl_2(aq)$ + __ $CO_2(g)$ + __ $H_2O(l)$

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.1 How many atoms of oxygen are represented by the notation 3 $Mg(OH)_2$?
(a) 1 **(b)** 2 **(c)** 3 **(d)** 5 **(e)** 6

3.2 In the following diagram, the white spheres represent hydrogen atoms and the blue spheres represent nitrogen atoms.



The two reactants combine to form a single product, ammonia, NH_3, which is not shown. Write a balanced chemical equation for the reaction. Based on the equation and the contents of the left (reactants) box, how many NH_3 molecules should be shown in the right (products) box?

(a) 2 **(b)** 3 **(c)** 4 **(d)** 6 **(e)** 9

3.3 The unbalanced equation for the reaction between methane and bromine is

$$\_\_CH_4(g) + \_\_Br_2(l) \longrightarrow \_\_CBr_4(s) + \_\_HBr(g)$$

Once this equation is balanced with the smallest possible integers, what is the value of the coefficient in front of bromine, Br_2?

(a) 1 **(b)** 2 **(c)** 3 **(d)** 4 **(e)** 6

Exercises

3.4 Write "true" or "false" for each statement. (**a**) We balance chemical equations as we do because energy must be conserved. (**b**) If the reaction $2\,O_3(g) \rightarrow 3\,O_2(g)$ goes to completion and all O_3 is converted to O_2, then the mass of O_3 at the beginning of the reaction must be the same as the mass of O_2 at the end of the reaction. (**c**) You can balance the "water-splitting" reaction $H_2O(l) \rightarrow H_2(g) + O_2(g)$ by writing it this way: $H_2O_2(l) \rightarrow H_2(g) + O_2(g)$.

3.5 Balance the following equations:

(a) $SiCl_4(l) + H_2O(l) \longrightarrow Si(OH)_4(s) + HCl(aq)$

(b) $CO_2(g) + H_2O \longrightarrow C_6H_{12}O_6(s) + O_2(g)$

(c) $Al(OH)_3(s) + H_2SO_4(l) \longrightarrow Al_2(SO_4)_3(s) + H_2O(l)$

(d) $H_3PO_4(aq) \longrightarrow H_4P_2O_7(aq) + H_2O(l)$

3.6 Balance the following equations:

(a) $CaS(s) + H_2O(l) \longrightarrow Ca(HS)_2(aq) + Ca(OH)_2(aq)$

(b) $NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$

(c) $FeCl_3(s) + Na_2CO_3(aq) \longrightarrow Fe_2(CO_3)_3(s) + NaCl(aq)$

(d) $FeS_2(s) + O_2(g) \longrightarrow Fe_2O_3(s) + SO_2(g)$

3.7 Write balanced chemical equations corresponding to each of the following descriptions: (**a**) Potassium cyanide reacts with an aqueous solution of sulfuric acid to form hydrogen cyanide gas. (**b**) When an aqueous solution of ammonium nitrite (NH_4NO_2) reacts with an aqueous solution of potassium hydroxide, ammonia gas, water and metal nitrate is formed. (**c**) When hydrogen gas is passed over solid hot iron(III) oxide, the resulting reaction produces iron and gaseous water. (**d**) When liquid ethanoic acid (CH_3COOH) is combusted, carbon dioxide and water are formed.

Answers to Self-Assessment Exercises

3.1 (e) **3.2** (d) **3.3** (d)

3.2 | Simple Patterns of Chemical Reactivity: Combination, Decomposition, and Combustion

One of the great triumphs of chemistry over the last hundred years is the development of fertilizers that enable us to feed the world. Ammonia, NH_3, is one of the principle chemicals farmers use to increase crop yield. The industrial process that are used to convert the elements nitrogen and hydrogen into ammonia is one of the most important chemical reactions in the world.

In this section, we will learn about broad classes of chemical reactions. When you finish this section, you should be able to:

- Recognize chemical reactions that are combination, decomposition, or combustion reactions
- Predict the products of these reactions
- Balance chemical equations for these reactions

Combination and Decomposition Reactions

In **combination reactions**, two or more substances react to form one product, according to

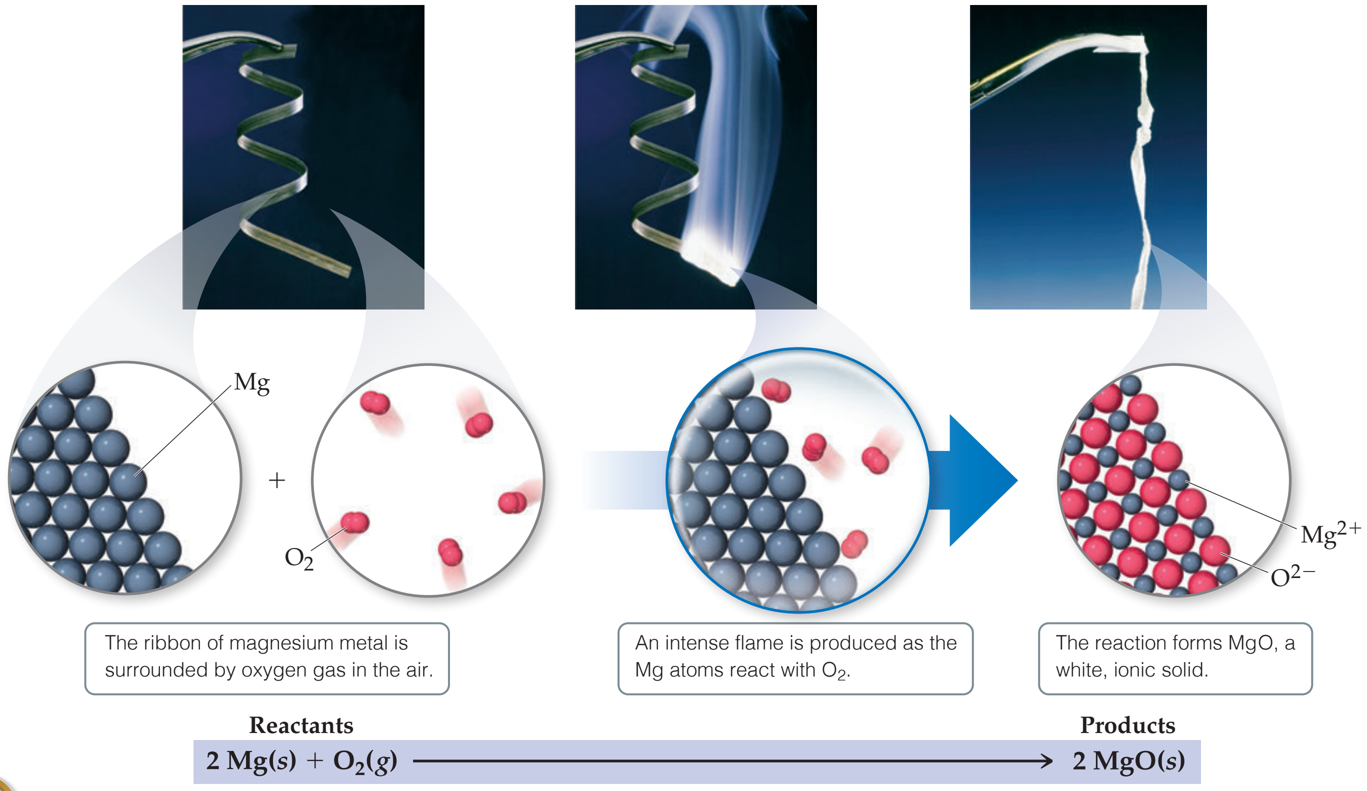
$$A + B \longrightarrow C \qquad [3.6]$$

For example, magnesium metal burns brilliantly in air to produce magnesium oxide. This reaction is used to produce the bright white flame generated by flares and some fireworks (Figure 3.5):

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s) \qquad [3.7]$$

A combination reaction between a metal and a nonmetal, as in Equation 3.7, produces an ionic solid. Recall that the formula of an ionic compound can be determined from the charges of its ions (Section 2.7). When magnesium reacts with oxygen, the magnesium loses electrons and forms the magnesium ion, Mg^{2+}. The oxygen gains electrons and forms the oxide ion, O^{2-}. Thus, the reaction product is MgO.

You should be able to recognize when a reaction is a combination reaction, and to predict the products when the reactants are a metal and a nonmetal.



▲ **Figure 3.5 Combustion of magnesium metal in air, a combination reaction.**

Other examples of combination reactions include the formation of small gaseous molecules such as CO_2 and NH_3 from their elements; and the reaction of calcium oxide with water to produce calcium hydroxide:

$$C(s) + O_2(g) \longrightarrow CO_2(g) \quad [3.8]$$

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g) \quad [3.9]$$

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq) \quad [3.10]$$

In a **decomposition reaction**, a single substance undergoes a reaction to produce two or more products:

$$C \longrightarrow A + B \quad [3.11]$$

An example of a decomposition reaction is when a metal carbonate decomposes to a metal oxide and carbon dioxide upon heating:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g) \quad [3.12]$$

Industrially, the decomposition of calcium carbonate at high temperatures is quite important. Limestone and seashells, natural sources of calcium carbonate, are heated to prepare calcium oxide, known as "quicklime" or "lime." Tens of millions of tons of CaO are used each year in making glass, in metallurgy to isolate metals from ores, in cement, and in steel manufacturing to remove impurities.

Another important example of a decomposition reaction is the decomposition of sodium azide, NaN_3, to form sodium metal and nitrogen gas:

$$2\,NaN_3(s) \longrightarrow 2\,Na(s) + 3\,N_2(g) \quad [3.13]$$

This reaction is what you find in automobile airbags (**Figure 3.6**). Approximately 100 g of NaN_3, upon physical impact, will explosively produce about 50 L of nitrogen gas.

▲ **Figure 3.6 Decomposition of sodium azide, $NaN_3(s)$, produces $N_2(g)$ that inflates air bags in automobiles.**

Combustion Reactions

Combustion reactions are rapid reactions that produce a flame. Most combustion reactions involve O_2 from air as a reactant. The combustion of hydrocarbons (compounds that contain only carbon and hydrogen) in air, illustrated in Equation 3.5, is a major energy-producing process in our world.

Hydrocarbons combusted in air react with O_2 to form CO_2 and H_2O. The number of molecules of O_2 required, as well as the number of product molecules formed, depend on the composition of the hydrocarbon, which acts as the fuel in the reaction. For example, the combustion of propane (C_3H_8, **Figure 3.7**), a gas used for cooking and home heating, is described by the chemical equation

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g) \quad [3.14]$$

Go Figure

Does this reaction produce or consume thermal energy (heat)?

▲ **Figure 3.7 Propane burning in air.** Liquid propane in the tank, C_3H_8, vaporizes and mixes with air as it escapes through the nozzle. The combustion reaction of C_3H_8 and O_2 produces a blue flame.

TABLE 3.1 Combination and Decomposition Reactions

Combination Reactions	
$A + B \longrightarrow C$	Two or more reactants combine to form a single product. Many elements react with one another in this fashion to form compounds.
$C(s) + O_2(g) \longrightarrow CO_2(g)$	
$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$	
$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$	
Decomposition Reactions	
$C \longrightarrow A + B$	A single reactant breaks apart to form two or more substances. Many compounds react this way when heated.
$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$	
$PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$	
$Cu(OH)_2(s) \longrightarrow CuO(s) + H_2O(g)$	

The state of the water in this reaction is listed as gas here, since the propane flame burns at a high temperature; but depending on conditions, the water molecules that are produced could be in the gas or liquid phase.

Millions of compounds are made only of carbon, hydrogen, and oxygen. Notable classes of such molecules are sugars and alcohols. Combustion of these oxygen-containing derivatives of hydrocarbons in air also produces CO_2, H_2O and energy. Many of the substances that function as energy sources in metabolism, such as the sugar glucose ($C_6H_{12}O_6$), react with O_2 to ultimately form CO_2 and H_2O. In our bodies, however, the reactions take place in a series of intermediate steps that occur at body temperature. These reactions that involve intermediate steps are called *oxidation reactions* rather than combustion reactions.

Sample Exercise 3.3

Writing Balanced Equations for Combination and Decomposition Reactions

Write a balanced equation for (**a**) the combination reaction between lithium metal and fluorine gas and (**b**) the decomposition reaction that occurs when solid barium carbonate is heated (two products form, a solid and a gas).

SOLUTION

(**a**) With the exception of mercury, all metals are solids at room temperature. Fluorine occurs as a diatomic molecule. Thus, the reactants are Li(*s*) and $F_2(g)$. The product will be composed of a metal and a nonmetal, so we expect it to be an ionic solid. Lithium ions have a 1+ charge, Li^+, whereas fluoride ions have a 1− charge, F^-. Thus, the chemical formula for the product is LiF. The balanced chemical equation is

$$2\,Li(s) + F_2(g) \longrightarrow 2\,LiF(s)$$

(**b**) The chemical formula for barium carbonate is $BaCO_3$. As mentioned, many metal carbonates decompose to metal oxides and carbon dioxide when heated. In Equation 3.7, for example, $CaCO_3$ decomposes to form CaO and CO_2. Thus, we expect $BaCO_3$ to decompose to BaO and CO_2. Barium and calcium are both in Group 2 in the periodic table, which further suggests they react in the same way:

$$BaCO_3(s) \longrightarrow BaO(s) + CO_2(g)$$

Practice Exercise

Which of the following reactions is the balanced equation that represents the decomposition reaction that occurs when silver(I) oxide is heated?
(**a**) $AgO(s) \longrightarrow Ag(s) + O(g)$ (**b**) $2\,AgO(s) \longrightarrow 2\,Ag(s) + O_2(g)$
(**c**) $Ag_2O(s) \longrightarrow 2\,Ag(s) + O(g)$ (**d**) $2\,Ag_2O(s) \longrightarrow 4\,Ag(s) + O_2(g)$
(**e**) $Ag_2O(s) \longrightarrow 2\,Ag(s) + O_2(g)$

Sample Exercise 3.4

Writing Balanced Equations for Combustion Reactions

Write the balanced equation for the reaction that occurs when methanol, $CH_3OH(l)$, is burned in air.

SOLUTION

When any compound containing C, H, and O is combusted, it reacts with the $O_2(g)$ in air to produce $CO_2(g)$ and $H_2O(g)$. Thus, the unbalanced equation is

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

The C atoms are balanced, one on each side of the arrow. Because CH_3OH has four H atoms, we place the coefficient 2 in front of H_2O to balance the H atoms:

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Adding this coefficient balances H but gives four O atoms in the products. Because there are only three O atoms in the reactants, we are not finished. We can place the coefficient $\frac{3}{2}$ in front of O_2 to give four O atoms in the reactants ($\frac{3}{2} \times 2 = 3$ O atoms in $\frac{3}{2}\,O_2$):

$$CH_3OH(l) + \tfrac{3}{2}\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Although this equation is balanced, it is not in its most conventional form because it contains a fractional coefficient. However, multiplying through by 2 removes the fraction and keeps the equation balanced:

$$2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

Practice Exercise

Write the balanced equation for the reaction that occurs when ethylene glycol, $C_2H_4(OH)_2$, burns in air.

(**a**) $C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(**b**) $2\,C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$
(**c**) $C_2H_4(OH)_2(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(**d**) $C_2H_4(OH)_2(l) + 5\,O(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(**e**) $4\,C_2H_4(OH)_2(l) + 10\,O_2(g) \longrightarrow 8\,CO_2(g) + 12\,H_2O(g)$

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.8 When Na and S undergo a combination reaction, what is the chemical formula of the product?

(a) NaS **(c)** NaS_2 **(e)** Na_3S_2

(b) Na_2S **(d)** Na_2S_3

3.9 Which of the following reactions is the balanced equation for the decomposition of silver(I) oxide?

(a) $AgO(s) \longrightarrow Ag(s) + O(g)$

(b) $2\,AgO(s) \longrightarrow 2\,Ag(s) + O_2(g)$

(c) $Ag_2O(s) \longrightarrow 2\,Ag(s) + O(g)$

(d) $2\,Ag_2O(s) \longrightarrow 4\,Ag(s) + O_2(g)$

(e) $Ag_2O(s) \longrightarrow 2\,Ag(s) + O_2(g)$

3.10 Write the balanced equation for the reaction that occurs when ethylene glycol, $C_2H_4(OH)_2$, burns in air.

(a) $C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$

(b) $2\,C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$

(c) $C_2H_4(OH)_2(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$

(d) $C_2H_4(OH)_2(l) + 5\,O(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$

(e) $4\,C_2H_4(OH)_2(l) + 10\,O_2(g) \longrightarrow 8\,CO_2(g) + 12\,H_2O(g)$

Exercises

3.11 (**a**) When the metallic element lithium combines with the nonmetallic element chlorine, $Cl_2(g)$, what is the chemical formula of the product? (**b**) Is the product a solid, liquid, or gas at room temperature? (**c**) In the balanced chemical equation for this reaction, what is the coefficient in front of the product if the coefficient in front of $Cl_2(g)$ is 1?

3.12 Write a balanced chemical equation for the reaction that occurs when (**a**) $Mg(s)$ reacts with $Cl_2(g)$; (**b**) barium carbonate decomposes into barium oxide and carbon dioxide gas when heated; (**c**) the hydrocarbon styrene, $C_8H_8(l)$, is combusted in air; (**d**) dimethylether, $CH_3OCH_3(g)$, is combusted in air.

3.13 Balance the following equations and indicate whether they are combination, decomposition, or combustion reactions:

(a) $C_7H_{16}(s) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$

(b) $Li_3N(s) + BN(s) \longrightarrow Li_3BN_2(s)$

(c) $Zn(OH)_2(s) \longrightarrow ZnO(s) + H_2O(l)$

(d) $Ag_2O(s) \longrightarrow Ag(s) + O_2(g)$

Answers to Self-Assessment Exercises

3.8 (b) **3.9** (d) **3.10** (b)

3.3 Formula Weights and Elemental Compositions of Substances

Sulfuric acid, $H_2SO_4(l)$, is a common laboratory chemical that is used by the metric ton in many reactions in the chemical industry. We can see from the molecular model on the left that one molecule of H_2SO_4 contains one sulfur atom (yellow), four oxygen atoms (red), and two hydrogen atoms (white). But in the lab, we dispense milliliters of liquid H_2SO_4, or more commonly, milliliters of its aqueous solution, $H_2SO_4(aq)$. How can we connect the chemical equations that we write, which represent individual molecules, to the reactions we do in the lab, where quantities are measured in grams or milliliters? That is the topic we explore in this section.

When you finish this section, you should be able to:

- Calculate the formula weight of a substance from its empirical formula or it's molecular weight from its molecular formula
- Calculate the elemental composition of a substance from the mass percentages of the elements that make up the substance

Formula and Molecular Weights

The **formula weight** (FW) of a substance is the sum of the atomic weights (AW) of the atoms in the chemical formula of the substance. Using the atomic weights from the periodic table, we find, for example, that the formula weight of sulfuric acid, H_2SO_4, is 98.1 u (atomic mass units):

$$\begin{aligned}\text{FW of } H_2SO_4 &= 2(\text{AW of H}) + (\text{AW of S}) + 4(\text{AW of O})\\ &= 2(1.0\text{ amu}) + (32.1\text{ amu}) + 4(16.0\text{ amu})\\ &= 98.1\text{ amu}\end{aligned}$$

For convenience, we have rounded off the atomic weight to one decimal place, a practice we will follow in most calculations in this book.

If the chemical formula is the chemical symbol of an element, such as Na, the formula weight equals the atomic weight of the element (for Na, this would be 23.1 amu). If the chemical formula is that of a molecule, like H_2SO_4, the formula weight can also be called the **molecular weight** (MW).

Not all substances, though, are molecules. For instance, ionic substances such as calcium chloride exist as three-dimensional arrays of ions (see Figure 2.18). In these cases, the empirical formula is used as the formula unit, and the formula weight is the sum of the atomic weights of the atoms in the empirical formula. For example, the formula unit of $CaCl_2$ consists of one Ca^{2+} ion and two Cl^- ions. Thus, the formula weight of $CaCl_2$ is

$$\text{FW of } CaCl_2 = (\text{AW Ca}) + 2\,(\text{AW Cl}) = 40.1\text{ amu} + 2(35.5\text{ amu}) = 111.1\text{ amu}$$

Elemental Compositions of Substances

Let's say you are a forensic chemist, working in a crime lab. Your colleagues find a mysterious white powder at a crime scene. Is it salt, sugar, methamphetamine, cocaine or something else?

One way to determine the identity of a substance is to measure its **elemental composition** and compare it to the calculated elemental compositions of possible candidate substances. We do these calculations based on the masses of each element in the compound:

$$\begin{matrix}\text{\% mass composition}\\ \text{of element}\end{matrix} = \frac{\begin{pmatrix}\text{number of atoms}\\ \text{of element}\end{pmatrix}\begin{pmatrix}\text{AW}\\ \text{of element}\end{pmatrix}}{\text{FW of substance}} \times 100\% \qquad [3.15]$$

The sum of all the mass percentages of each element in the compound must add up to 100%.

As an example, let's calculate the mass percentage of sulfur in sulfuric acid. Based on atoms we can see that for each H_2SO_4 molecule, one atom out of seven is sulfur. But that does not mean that 1/7 of the mass of the compound is sulfur, since the atoms weigh different amounts:

$$\% \text{ S in } H_2SO_4 = \frac{(1)(32.1 \text{ amu})}{98.1 \text{ amu}} \times 100\% = 32.7\%$$

So we find that almost a third of the mass of a given quantity of pure H_2SO_4 is due to sulfur.

Sample Exercise 3.5

Calculating Formula Weights

Calculate the formula weight of (**a**) sucrose, $C_{12}H_{22}O_{11}$ (table sugar); and (**b**) calcium nitrate, $Ca(NO_3)_2$.

SOLUTION

(**a**) By adding the atomic weights of the atoms in sucrose, we find the formula weight to be 342.0 amu:

$$\begin{aligned} 12 \text{ C atoms} &= 12(12.0 \text{ amu}) = 144.0 \text{ amu} \\ 22 \text{ H atoms} &= 22(1.0 \text{ amu}) = 22.0 \text{ amu} \\ 11 \text{ O atoms} &= 11(16.0 \text{ amu}) = \underline{176.0 \text{ amu}} \\ & \qquad\qquad\qquad\qquad 342.0 \text{ amu} \end{aligned}$$

(**b**) If a chemical formula has parentheses, the subscript outside the parentheses is a multiplier for all atoms inside. Thus, for $Ca(NO_3)_2$ we have

$$\begin{aligned} 1 \text{ Ca atom} &= 1(40.1 \text{ amu}) = 40.1 \text{ amu} \\ 2 \text{ N atoms} &= 2(14.0 \text{ amu}) = 28.0 \text{ amu} \\ 6 \text{ O atoms} &= 6(16.0 \text{ amu}) = \underline{96.0 \text{ amu}} \\ & \qquad\qquad\qquad\qquad 164.1 \text{ amu} \end{aligned}$$

▶ Practice Exercise

Calculate the formula weight of (**a**) $Al(OH)_3$, (**b**) CH_3OH, and (**c**) TaON.

STRATEGIES FOR SUCCESS Problem Solving

Practice is the key to success in solving problems. As you practice, you can improve your skills by following these steps:

1. **Analyze the problem.** Read the problem carefully. What does it say? Draw a picture or diagram that will help you to visualize the problem. Write down both the data you are given and the quantity you need to obtain (the unknown).
2. **Develop a plan for solving the problem.** Consider a possible path between the given information and the unknown. What principles or equations relate the known data to the unknown? Recognize that some data may not be given explicitly in the problem; you may be expected to know certain quantities or look them up in tables (such as atomic weights). Recognize also that your plan may involve either a single step or a series of steps with intermediate answers.
3. **Solve the problem.** Use the known information and suitable equations or relationships to solve for the unknown. Dimensional analysis (Section 1.7) is a useful tool for solving a great number of problems. Be careful with significant figures, signs, and units.
4. **Check the solution.** Read the problem again to make sure you have found all the solutions asked for in the problem. Does your answer make sense? That is, is the answer outrageously large or small or is it in the ballpark? Finally, are the units and significant figures correct?

Sample Exercise 3.6

Calculating Percentage Composition

Calculate the percentage of carbon, hydrogen, and oxygen (by mass) in $C_{12}H_{22}O_{11}$.

SOLUTION

We'll use the steps outlined in the Strategies For Success: Problem Solving feature to answer the question.

Analyze We are given a chemical formula and asked to calculate the percentage by mass of each element.

Plan We use Equation 3.10, obtaining our atomic weights from a periodic table. We know the denominator in Equation 3.10, the formula weight of $C_{12}H_{22}O_{11}$, from Sample Exercise 3.5. We must use that value in three calculations, one for each element.

Solve

$$\%C = \frac{(12)(12.0\ \cancel{u})}{342.0\ \cancel{u}} \times 100\% = 42.1\%$$

$$\%H = \frac{(22)(1.0\ \cancel{u})}{342.0\ \cancel{u}} \times 100\% = 6.4\%$$

$$\%O = \frac{(11)(16.0\ \cancel{u})}{342.0\ \cancel{u}} \times 100\% = 51.5\%$$

Check Our calculated percentages must add up to 100%, which they do. We could have used more significant figures for our atomic weights, giving more significant figures for our percentage composition, but we have adhered to our suggested guideline of rounding atomic weights to one digit beyond the decimal point.

▶ Practice Exercise

What is the percentage of nitrogen, by mass, in calcium nitrate?
(**a**) 8.54% (**b**) 17.1% (**c**) 13.7% (**d**) 24.4% (**e**) 82.9%

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.14 Which of the following is the correct formula weight for calcium phosphate?

(a) 310.2 amu **(b)** 135.1 amu
(c) 182.2 amu **(d)** 278.2 amu
(e) 175.1 amu

3.15 What is the percentage of nitrogen, by mass, in calcium nitrate?

(a) 8.54% **(b)** 17.1%
(c) 13.7% **(d)** 24.4%
(e) 82.9%

3.16 A mysterious white powder found at a crime scene is analyzed and contains 66.8 ± 0.5% carbon by mass. One of the investigators hypothesizes that the substance is cocaine ($C_{17}H_{21}NO_4$). What percent carbon, by mass, is in cocaine?

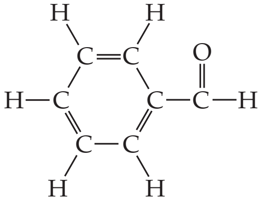
(a) 39.5% **(b)** 64.3%
(c) 67.3% **(d)** 70.6%
(e) 72.5%

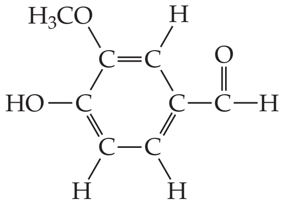
Exercises

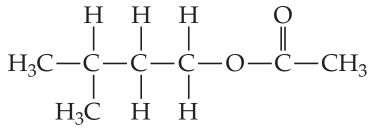
3.17 Determine the formula weights of each of the following compounds: (**a**) lead (IV) chloride; (**b**) copper(II) oxide; (**c**) iodic acid, HIO_3; (**d**) sodium perchlorate, $NaClO_4$; (**e**) indium nitride, (**f**) phosphorus pentoxide, P_4O_{10}; (**g**) boron trichloride.

3.18 Calculate the percentage by mass of oxygen in the following compounds: (**a**) vanillin, $C_8H_8O_3$; (**b**) isopropyl alcohol, C_3H_8O; (**c**) acetaminophen, $C_8H_9NO_2$; (**d**) cyclopropanone, C_3H_4O; (**e**) dioxin, $C_{12}H_4Cl_4O_2$; (**f**) penicillin, $C_{16}H_{18}N_2O_4S$.

3.19 Based on the following structural formulas, calculate the percentage of carbon by mass present in each compound:

(a) 
Benzaldehyde (almond fragrance)

(b) 
Vanillin (vanilla flavor)

(c) 
Isopentyl acetate (banana flavor)

Answers to Self-Assessment Exercises

3.14 (a) **3.15** (b) **3.16** (c)

3.4 Avogadro's Number and the Mole; Molar Mass

Even the smallest samples in the laboratory contain enormous numbers of atoms, ions or molecules.For example, a teaspoon (about 5 mL) of water contains 2×10^{23} water molecules, a number so large it defies comprehension. Chemists, therefore, have devised a way to conveniently count such enormous numbers.

When you finish this section, you should be able to:

- Explain the *mole* and the origin of *Avogadro's number*
- Calculate the *molar mass* for a compound and relate this to its formula weight
- Convert between grams, molecules, and moles of a substance

The Mole and Avogadro's Number

In everyday life, we use counting units such as dozen (12 objects), score (20 objects), gross (144 objects) or ream (500 objects).In chemistry, the counting unit for numbers of atoms, ions, or molecules in a laboratory-size sample is the *mole*, abbreviated mol.One **mole** is the amount of matter that contains as many objects as the number of atoms in exactly 12 g of isotopically pure ^{12}C. From experiments, scientists have determined this number to be 6.0221415×10^{23}, which we usually round to 6.02×10^{23}. Scientists call this value **Avogadro's number**, N_A, in honor of the Italian scientist Amedeo Avogadro (1776–1856), and it is often cited with units of reciprocal moles, $\mathbf{6.02 \times 10^{23}\ mol^{-1}}$.* The unit (read as either "inverse mole" or "per mole") reminds us that there are 6.02×10^{23} objects per one mole.A mole of atoms, a mole of molecules, or a mole of anything else all contain Avogadro's number of objects:

$$1\ \text{mol}\ ^{12}\text{C atoms} = 6.02 \times 10^{23}\ ^{12}\text{C atoms}$$
$$1\ \text{mol}\ H_2O\ \text{molecules} = 6.02 \times 10^{23}\ H_2O\ \text{molecules}$$
$$1\ \text{mol}\ NO_3^-\ \text{ions} = 6.02 \times 10^{23}\ NO_3^-\ \text{ions}$$

Avogadro's number is so large that it is difficult to imagine.Spreading 6.02×10^{23} marbles over the Earth's surface would produce a layer about 3 miles thick.Avogadro's number of pennies placed side by side in a straight line would encircle the Earth 300 trillion(3×10^{14}) times.

Molar Mass

A dozen is the same number, 12, whether we have a dozen eggs or a dozen elephants. Clearly, however, a dozen eggs does not have the same mass as a dozen elephants. Similarly, a mole is always the same number (6.02×10^{23}), but 1 mol samples of different substances have *different masses*.

Compare, for example, 1 mol of ^{12}C and 1 mol of ^{24}Mg. A single ^{12}C atom has a mass of 12 amu, whereas a single ^{24}Mg atom is twice as massive, 24 amu (to two significant figures).Because a mole of anything always contains the same number of particles, a mole of ^{24}Mg must be twice as massive as a mole of ^{12}C. Because 1 mol of ^{12}C has a mass of 12 g (by definition), 1 mol of ^{24}Mg must have a mass of 24 g.This example illustrates a general rule relating the mass of an atom the mass of Avogadro's number (1 mol) of these atoms: *The atomic weight of an element in atomic mass units is numerically equal to the mass in grams of 1 mol of that element.*

For example, Cl has an atomic weight of 35.5 amu; therefore, 1 mol of Cl atoms has a mass of 35.5 g.

Au has an atomic weight of 197 amu; therefore, 1 mol of Au atoms has a mass of 197 g.

The same relationship holds for the formula weight or molecular weight of a substance and the mass of 1 mol of that substance:

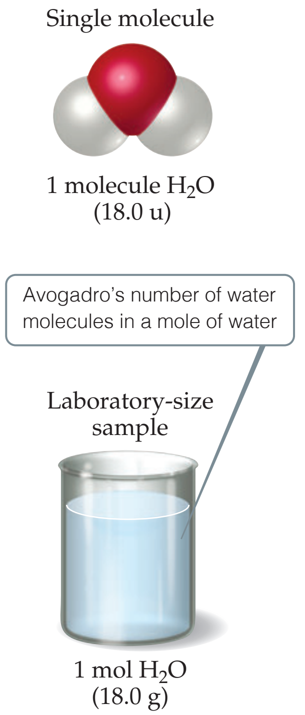
H_2O has a molecular weight of 18.0 amu; therefore, 1 mol of H_2O has a mass of 18.0 g

NaCl has a formula weight of 58.5 amu; therefore 1 mol of NaCl has a mass of 58.5 g.

The mass in grams of one mole of a substance is called the **molar mass** of the substance.The units of molar mass are g/mol, or g mol^{-1}.*The molar mass in grams per mole of any substance is numerically equal to its formula weight in atomic mass units*.For NaCl, for

Go Figure

What number do you get if you divide the mass of 1 mol of water by the mass of 1 molecule of water?

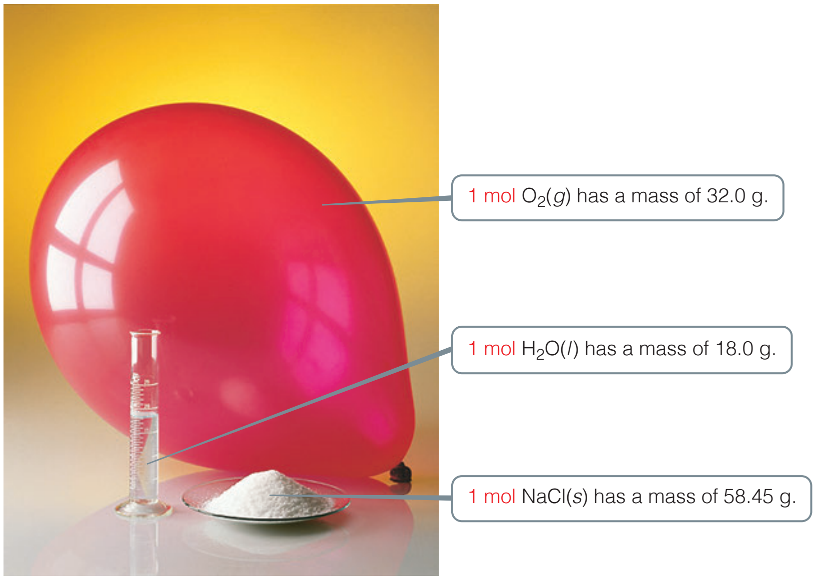


▲ **Figure 3.8 Comparing the mass of 1 molecule and 1 mol of H_2O.** Both masses have the same number but different units (atomic mass units and grams). Expressing both masses in grams indicates their huge difference: 1 molecule of H_2O has a mass of 2.99×10^{-23} g, whereas 1 mol H_2O has a mass of 18.0 g.

TABLE 3.2 Mole Relationships

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	6.02×10^{23} N atoms
Molecular nitrogen or "dinitrogen"	N_2	28.0	28.0	6.02×10^{23} N_2 molecules $2(6.02 \times 10^{23})$ N atoms
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	6.02×10^{23} $BaCl_2$ formula units 6.02×10^{23} Ba^{2+} ions $2(6.02 \times 10^{23})$ Cl^- ions

[a]Recall that the mass of an electron is more than 1800 times smaller than the masses of the proton and the neutron; thus, ions and atoms have essentially the same mass.



▲ **Figure 3.9 One mole each of a solid (NaCl), a liquid (H_2O), and a gas (O_2).** In each case, the mass in grams of 1 mol—that is, the molar mass—is numerically equal to the formula weight in atomic mass units. Each of these samples contains 6.02×10^{23} formula units.

instance, the formula weight is 58.5 amu and therefore its molar mass is 58.5 g/mol.Mole relationships for several substances are shown in Table 3.1, and **Figures 3.8** and **3.9** illustrate 1 mol quantities of common substances.

The entries in **Table 3.2** for N and N_2 point out the importance of stating the chemical form of a substance when using the mole concept. For instance, suppose you read that 1 mol of nitrogen is used to make ammonia.You might interpret this statement to mean that 1 mol of nitrogen atoms was used (14.0 g). Unless otherwise stated, however, what is meant is 1 mol of nitrogen molecules, N_2 (28.0 g), was used, because N_2 is the naturally occurring form of the element. To avoid ambiguity, it is important to explicitly state the chemical form being discussed.Using the chemical formula—N or N_2, for instance—avoids any confusion.

Converting Between Masses, Moles, and Atoms/Molecules/Ions

Now we are ready to learn how to relate chemical equations to the amounts of chemicals used and produced in reactions.We will use the mole concept with dimensional analysis (Section 1.7) to help us (**Figure 3.10**).

Go Figure What units would you put under "molar mass" and "Avogadro's number" on this diagram?

▲ **Figure 3.10 Procedure for interconverting mass and number of formula units.** The number of moles of the substance is central to the calculation. Thus, the mole concept can be thought of as the bridge between the mass of a sample in grams and the number of formula units contained in the sample.

For example, let's calculate how many copper atoms are in an old copper penny. Such a penny has a mass of about 3 grams, and let's assume for simplicity that the penny is pure copper:

$$\begin{aligned}\text{Number of Cu atoms} &= (3\text{ g Cu})\left(\frac{1\text{ mol Cu}}{63.5\text{ g Cu}}\right)\left(\frac{6.02\times10^{23}\text{ Cu atoms}}{1\text{ mol Cu}}\right)\\ &= 3\times10^{22}\text{ Cu atoms}\end{aligned}$$

We have rounded the answer to one significant figure because we only used one significant figure for the mass of the penny. Notice how dimensional analysis provides a straightforward route from grams to number of atoms.The molar mass and Avogadro's number are used as conversion factors to convert grams to moles and then moles to atoms.Also notice that our answer is a very large number: this makes sense, since there are enormous numbers of atoms in a macroscopic sample we can pick up with our hands.Any time that you calculate the number of atoms, ions, or molecules in a laboratory-scale sample, you should expect the number to be very large.However, if you were to calculate number of moles in a laboratory sample, the number may not be that large, and in fact might be less than 1.That is true for the moles of copper in our old penny:

$$\begin{aligned}\text{Moles of Cu} &= (3\text{ g Cu})\left(\frac{1\text{ mol Cu}}{63.5\text{ g Cu}}\right)\\ &= 5\times10^{-2}\text{ mol Cu}\end{aligned}$$

CHEMISTRY AND LIFE Glucose Monitoring

Our body converts most of the food we eat into glucose. After digestion, glucose is delivered to cells via the blood. Cells need glucose to live, and the hormone insulin must be present in order for glucose to enter the cells. Normally, the body adjusts the concentration of insulin automatically, in concert with the glucose concentration after eating, so that normal blood glucose levels are 70–120 mg/dL. However, in a diabetic person, either little or no insulin is produced (Type 1 diabetes) or insulin is produced but the cells cannot take it up properly (Type 2 diabetes). In either case the blood glucose levels are higher than they are in a normal person. A person who has not eaten for 8 hours or more is diagnosed as diabetic if his or her glucose level is 126 mg/dL or higher.

Glucose meters work by the introduction of blood from a person, usually by a prick of the finger, onto a small strip of paper that contains chemicals that react with glucose. Insertion of the strip into a small battery-operated reader gives the glucose concentration (**Figure 3.11**). The mechanism of the readout varies from one monitor to another—it may be a measurement of a small electrical current or a measurement of light produced in a chemical reaction. Depending on the reading on any given day, a diabetic person may need to receive an injection of insulin or simply limit his or her intake of sugar-rich foods for a while.

▲ **Figure 3.11 Glucose meter.**

Sample Exercise 3.7
Estimating Numbers of Atoms

Without using a calculator, arrange these samples in order of increasing numbers of carbon atoms: 12 g ^{12}C, 1 mol C_2H_2, 9×10^{23} molecules of CO_2.

SOLUTION

Analyze We are given amounts of three substances expressed in grams, moles, and number of molecules and asked to arrange the samples in order of increasing numbers of C atoms.

Plan To determine the number of C atoms in each sample, we must convert 12 g ^{12}C, 1 mol C_2H_2, and 9×10^{23} molecules CO_2 to numbers of C atoms. To make these conversions, we use the definition of mole and Avogadro's number.

Solve One mole is defined as the amount of matter that contains as many units of the matter as there are C atoms in exactly 12 g of ^{12}C. Thus, 12 g of ^{12}C contains 1 mol of C atoms $= 6.02 \times 10^{23}$ C atoms.

One mol of C_2H_2 contains 6.02×10^{23} C_2H_2 molecules. Because there are two C atoms in each molecule, this sample contains 12.04×10^{23} C atoms.

Because each CO_2 molecule contains one C atom, the CO_2 sample contains 9×10^{23} C atoms.

Hence, the order is 12 g ^{12}C (6×10^{23} C atoms) $<$ 9×10^{23} CO_2 molecules (9×10^{23} C atoms) $<$ 1 mol C_2H_2 (12×10^{23} C atoms).

Check We can check our results by comparing numbers of moles of C atoms in the samples because the number of moles is proportional to the number of atoms. Thus, 12 g of ^{12}C is 1 mol C, 1 mol of C_2H_2 contains 2 mol C, and 9×10^{23} molecules of CO_2 contain 1.5 mol C, giving the same order as stated previously.

Practice Exercise

Which of the following samples contains the fewest sodium atoms? (**a**) 1 mol sodium oxide (**b**) 45 g sodium fluoride (**c**) 50 g sodium chloride (**d**) 1 mol sodium nitrate

Sample Exercise 3.8
Converting Grams to Moles

Calculate the number of moles of glucose ($C_6H_{12}O_6$) in a 5.380 g sample.

SOLUTION

Analyze We are given the number of grams of a substance and its chemical formula and asked to calculate the number of moles.

Plan The molar mass of a substance provides the factor for converting grams to moles. The molar mass of $C_6H_{12}O_6$ is 180.0 g/mol (Sample Exercise 3.9).

Solve Using 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$ to write the appropriate conversion factor, we have

$$\text{Moles } C_6H_{12}O_6 = (5.380\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)$$
$$= 0.02989 \text{ mol } C_6H_{12}O_6$$

Check Because 5.380 g is less than the molar mass, an answer less than 1 mol is reasonable. The unit mol is appropriate. The original data had four significant figures, so our answer has four significant figures.

Practice Exercise

How many moles of water are in 1.00 L of water, whose density is 1.00 g/mL?

Sample Exercise 3.9
Converting Moles to Grams

Calculate the mass, in grams, of 0.433 mol of calcium nitrate.

SOLUTION

Analyze We are given the number of moles and the name of a substance and asked to calculate the number of grams in the substance.

Plan To convert moles to grams, we need the molar mass, which we can calculate using the chemical formula and atomic weights.

Solve Because the calcium ion is Ca^{2+} and the nitrate ion is NO_3^-, the chemical formula for calcium nitrate is $Ca(NO_3)_2$. Adding the atomic weights of the elements in the compound gives a formula weight of 164.1 u. Using 1 mol $Ca(NO_3)_2$ = 164.1 g $Ca(NO_3)_2$ to write the appropriate conversion factor, we have

$$\text{Grams } Ca(NO_3)_2 = (0.433\ \cancel{\text{mol } Ca(NO_3)_2})\left(\frac{164.1 \text{ g } Ca(NO_3)_2}{1\ \cancel{\text{mol } Ca(NO_3)_2}}\right)$$
$$= 71.1 \text{ g } Ca(NO_3)_2$$

Check The number of moles is less than 1, so the number of grams must be less than the molar mass, 164.1 g. Using rounded numbers to estimate, we have $0.5 \times 150 = 75$ g, which means the magnitude of our answer is reasonable. Both the units (g) and the number of significant figures (3) are correct.

Practice Exercise

What is the mass, in grams, of **(a)** 6.33 mol of $NaHCO_3$ and **(b)** 3.0×10^{-5} mol of sulfuric acid?

Sample Exercise 3.10

Calculating Numbers of Molecules and Atoms from Mass

(a) How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$?
(b) How many oxygen atoms are in this sample?

SOLUTION

Analyze We are given the number of grams and the chemical formula of a substance and asked to calculate **(a)** the number of molecules and **(b)** the number of O atoms in the substance.

Plan **(a)** The strategy for determining the number of molecules in a given quantity of a substance is summarized in Figure 3.12. We must convert 5.23 g to moles of $C_6H_{12}O_6$ and then convert moles to molecules of $C_6H_{12}O_6$. The first conversion uses the molar mass of $C_6H_{12}O_6$, 180.0 g/mol, and the second conversion uses Avogadro's number.

(b) To determine the number of O atoms, we use the fact that there are six O atoms in each $C_6H_{12}O_6$ molecule. Thus, multiplying the number of molecules we calculated in (a) by the factor (6 atoms O/1 molecule $C_6H_{12}O_6$) gives the number of O atoms.

Solve

(a) Convert grams $C_6H_{12}O_6$ to molecules $C_6H_{12}O_6$.

$$\text{Molecules } C_6H_{12}O_6 = (5.23\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6.02 \times 10^{23}\ \text{molecules } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)$$
$$= 1.75 \times 10^{22}\ \text{molecules } C_6H_{12}O_6$$

(b) Convert molecules $C_6H_{12}O_6$ to atoms O.

$$\text{Atoms O} = (1.75 \times 10^{22}\ \cancel{\text{molecules } C_6H_{12}O_6})\left(\frac{6\ \text{atoms O}}{\cancel{\text{molecule } C_6H_{12}O_6}}\right)$$
$$= 1.05 \times 10^{23}\ \text{atoms O}$$

Check

(a) Because the mass we began with is less than a mole, there should be fewer than 6.02×10^{23} molecules in the sample, which means the magnitude of our answer is reasonable. A ballpark estimate of the answer comes reasonably close to the answer we derived in this exercise: $5/200 = 2.5 \times 10^{-2}$ mol; $(2.5 \times 10^{-2})(6 \times 10^{23}) = 15 \times 10^{21} = 1.5 \times 10^{22}$ molecules. The units (molecules) and the number of significant figures (three) are appropriate.

(b) The answer is six times as large as the answer to part (a), exactly what it should be. The number of significant figures (three) and the units (atoms O) are correct.

Practice Exercise

How many chlorine atoms are in 12.2 g of CCl_4?
(a) 4.77×10^{22}
(b) 7.34×10^{24}
(c) 1.91×10^{23}
(d) 2.07×10^{23}

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.20 Which of the following samples contains the fewest sodium atoms?
- **(a)** 1.0 mol sodium oxide
- **(b)** 45 g sodium fluoride
- **(c)** 50 g sodium chloride
- **(d)** 1.0 mol sodium nitrate

3.21 A sample of an ionic compound containing iron and chlorine is analyzed and found to have a molar mass of 126.8 g/mol. What is the charge of the iron in this compound?
(a) 1+ **(b)** 2+ **(c)** 3+ **(d)** 4+

3.22 How many chlorine atoms are in 12.2 g of CCl_4?
- **(a)** 4.77×10^{22}
- **(b)** 7.34×10^{24}
- **(c)** 1.91×10^{23}
- **(d)** 2.07×10^{23}

Exercises

3.23 (**a**) Write "true" or "false" for each statement. (**a**) A mole of ducks contain a mole of feathers. (**b**) A mole of ammonia gas has a mass of 17.0 g. (**c**) The mass of 1 ammonia molecule is 17.0 g. (**d**) A mole of $MgSO_4(s)$ contains 4 moles of oxygen atoms.

3.24 Without doing any detailed calculations (but using a periodic table to give atomic weights), rank the following samples in order of increasing numbers of atoms: 0.5 mol BCl_3 molecules, 197 g gold, 6.0×10^{23} CCl_4 molecules.

3.25 Calculate the following quantities:

(**a**) mass, in grams, of 0.105 mol sucrose ($C_{12}H_{22}O_{11}$)

(**b**) moles of $Zn(NO_3)_2$ in 143.50 g of this substance

(**c**) number of molecules in 1.0×10^{-6} mol CH_3CH_2OH

(**d**) number of N atoms in 0.410 mol NH_3

3.26 (**a**) What is the mass, in grams, of 2.50×10^{-3} mol of ammonium phosphate?

(**b**) How many moles of chloride ions are in 0.2550 g of aluminum chloride?

(**c**) What is the mass, in grams, of 7.70×10^{20} molecules of caffeine, $C_8H_{10}N_4O_2$?

(**d**) What is the molar mass of cholesterol if 0.00105 mol has a mass of 0.406 g?

3.27 The molecular formula of saccharin, an artificial sweetener, is $C_7H_5NO_3S$. (**a**) What is the molar mass of saccharin? (**b**) How many moles of sachharin are in 2.00 mg of this substance? (**c**) How many molecules are in 2.00 mg of this substance? (**d**) How many C atoms are present in 2.00 mg of saccharin?

3.28 A sample of glucose, $C_6H_{12}O_6$, contains 1.250×10^{21} carbon atoms. (**a**) How many atoms of hydrogen does it contain? (**b**) How many molecules of glucose does it contain? (**c**) How many moles of glucose does it contain? (**d**) What is the mass of this sample in grams?

3.29 The allowable concentration level of vinyl chloride, C_2H_3Cl, in the atmosphere in a chemical plant is 2.0×10^{-6} g/L. How many moles of vinyl chloride in each liter does this represent? How many molecules per liter?

Answers to Self-Assessment Exercises

3.20 (c) 3.21 (b) 3.22 (c)

3.5 Formula Weights and Elemental Compositions of Substances

The empirical formula for a substance tells us the relative number of atoms of each element in the substance. (Section 2.6) The empirical formula H_2O shows that water contains two H atoms for each O atom. This ratio also applies on the molar level: 1 mol of

H_2O contains 2 mol of H atoms and 1 mol of O atoms. Conversely, *the ratio of the numbers of moles of all elements in a compound gives the subscripts in the compound's empirical formula*. Thus, the mole concept provides a way of calculating empirical formulas from experimental data.

When you finish this section, you should be able to:

- Calculate the empirical formula of a compound from the mass percentages of the elements that make up the compound.
- Calculate the molecular formula of a compound from its molar mass and empirical formula.

Mercury and chlorine combine to form a compound that is measured to be 74.0% mercury and 26.0% chlorine by mass. Thus, if we had a 100.0 g sample of the compound, it would contain 74.0 g of mercury and 26.0 g of chlorine. (Samples of any size can be used in problems of this type, but we will generally use 100.0 g to simplify the calculation of mass from percentage.) Using atomic weights to get molar masses, we calculate the number of moles of each element in the sample:

$$(74.0\ \cancel{\text{g Hg}})\left(\frac{1\ \text{mol Hg}}{200.6\ \cancel{\text{g Hg}}}\right) = 0.369\ \text{mol Hg}$$

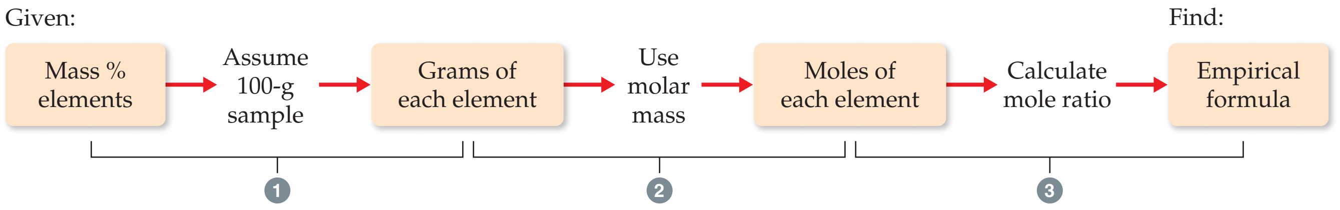
$$(26.0\ \cancel{\text{g Cl}})\left(\frac{1\ \text{mol Cl}}{35.5\ \cancel{\text{g Cl}}}\right) = 0.732\ \text{mol Cl}$$

We then divide the larger number of moles by the smaller number to obtain the Cl:Hg mole ratio:

$$\frac{\text{moles of Cl}}{\text{moles of Hg}} = \frac{0.732\ \text{mol Cl}}{0.369\ \text{mol Hg}} = \frac{1.98\ \text{mol Cl}}{1\ \text{mol Hg}}$$

Because of experimental errors, calculated values for a mole ratio may not be whole numbers, as in the calculation here. The number 1.98 is very close to 2, however, and so we can confidently conclude that the empirical formula for the compound is $HgCl_2$. The empirical formula is correct because its subscripts are the smallest integers that express the *ratio* of atoms present in the compound.

The general procedure for determining empirical formulas is outlined in **Figure 3.12**



▲ **Figure 3.12 Procedure for calculating an empirical formula from percentage composition.**

Sample Exercise 3.11
Calculating an Empirical Formula

Ascorbic acid (vitamin C) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?

SOLUTION

Analyze We are to determine the empirical formula of a compound from the mass percentages of its elements.

Plan The strategy for determining the empirical formula involves the three steps given in Figure 3.13.

Continued

Solve

(**1**) For simplicity we assume we have exactly 100 g of material, although any other mass could also be used.

Mass % elements → Assume 100-g sample → Grams of each element → Use molar mass → Moles of each element → Calculate mole ratio → Empirical formula

In 100.00 g of ascorbic acid we have 40.92 g C, 4.58 g H, and 54.50 g O.

(**2**) Next we calculate the number of moles of each element. We use atomic masses with four significant figures to match the precision of our experimental masses.

Mass % elements → Assume 100-g sample → Grams of each element → Use molar mass → Moles of each element → Calculate mole ratio → Empirical formula

$$\text{Moles C} = (40.92\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}}\right) = 3.407\ \text{mol C}$$

$$\text{Moles H} = (4.58\ \cancel{\text{g H}})\left(\frac{1\ \text{mol C}}{1.008\ \cancel{\text{g H}}}\right) = 4.54\ \text{mol H}$$

$$\text{Moles O} = (54.50\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}}\right) = 3.406\ \text{mol O}$$

(**3**) We determine the simplest whole-number ratio of moles by dividing each number of moles by the smallest number of moles.

Mass % elements → Assume 100-g sample → Grams of each element → Use molar mass → Moles of each element → Calculate mole ratio → Empirical formula

$$\text{C:}\frac{3.407}{3.406} = 1.000 \quad \text{H:}\frac{4.54}{3.406} = 1.33 \quad \text{O:}\frac{3.406}{3.406} = 1.000$$

The ratio for H is too far from 1 to attribute the difference to experimental error; in fact, it is quite close to $1\frac{1}{3}$. This suggests we should multiply the ratios by 3 to obtain whole numbers:

$$\text{C : H : O} = (3 \times 1 : 3 \times 1.33 : 3 \times 1) = (3 : 4 : 3)$$

Thus, the empirical formula is $C_3H_4O_3$.

Check It is reassuring that the subscripts are moderate-size whole numbers. Also, calculating the percentage composition of $C_3H_4O_3$ gives values very close to the original percentages.

Practice Exercise

A 2.144-g sample of phosgene, a compound used as a chemical warfare agent during World War I, contains 0.260 g of carbon, 0.347 g of oxygen, and 1.537 g of chlorine. What is the empirical formula of this substance?
(**a**) CO_2Cl_6 (**b**) $COCl_2$ (**c**) $C_{0.022}O_{0.022}Cl_{0.044}$ (**d**) C_2OCl_2

Molecular Formulas from Empirical Formulas

For molecular substances, the empirical formula and the molecular formula are often different. For example, benzene has a molecular formula of C_6H_6, but it's empirical formula CH is the same as that of the gas acetylene, whose molecular formula is C_2H_2. Knowledge of the empirical formula is not sufficient to differentiate these two very different compounds. Fortunately, we can obtain the molecular formula for any compound from its empirical formula if we know either the molecular weight of the compound, which can be measured by a variety of methods, including mass spectrometry (link to Section 2.4). *The subscripts in the molecular formula of a substance are always whole-number multiples of the subscripts in its empirical formula*. (Section 2.6) This multiple can be found by dividing the molecular weight by the empirical formula weight:

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}}$$

For example, the empirical formula of ascorbic acid (vitamin C) is $C_3H_4O_3$.. This means the empirical formula weight is 3(12.0 amu) + 4(1.0 amu) + 3(16.0 amu) = 88.0 amu. The experimentally determined molecular weight is 176 amu. Thus, we find the whole-number multiple that converts the empirical formula to the molecular formula by dividing

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}} = \frac{176\text{ amu}}{88.0\text{ amu}} = 2$$

Consequently, we multiply the subscripts in the empirical formula by this multiple, giving the molecular formula: $C_6H_8O_6$.

Sample Exercise 3.12

Determining a Molecular Formula

Mesitylene, a hydrocarbon found in crude oil, has an empirical formula of C_3H_4 and an experimentally determined molecular weight of 121 amu. What is its molecular formula?

SOLUTION

Analyze We are given an empirical formula and a molecular weight of a compound and asked to determine its molecular formula.

Plan The subscripts in a compound's molecular formula are whole-number multiples of the subscripts in its empirical formula. We find the appropriate multiple by using Equation 3.11.

Solve The formula weight of the empirical formula C_3H_4 is

$$3(12.0\text{ amu}) + 4(1.0\text{ amu}) = 40.0\text{ amu}$$

Next, we use this value in Equation 3.11:

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}} = \frac{121}{40.0} = 3.03$$

Only whole-number ratios make physical sense because molecules contain whole atoms. The 3.03 in this case could result from a small experimental error in the molecular weight. We therefore multiply each subscript in the empirical formula by 3 to give the molecular formula: C_9H_{12}.

Check We can have confidence in the result because dividing molecular weight by empirical formula weight yields nearly a whole number.

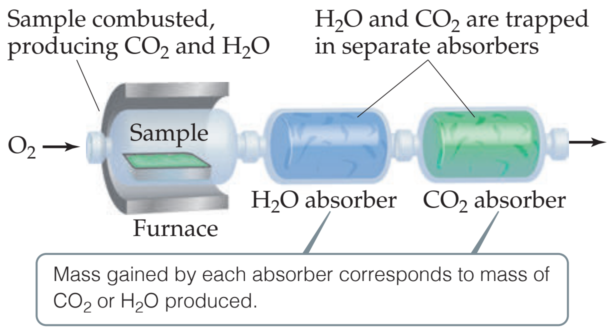
Practice Exercise

Cyclohexane, a commonly used organic solvent, is 85.6% C and 14.4% H by mass with a molar mass of 84.2 g/mol. What is its molecular formula?
(**a**) C_6H (**b**) CH_2 (**c**) C_5H_{24} (**d**) C_6H_{12} (**e**) C_4H_8

Combustion Analysis

One technique for determining empirical formulas in the laboratory is *combustion analysis*, commonly used for compounds containing principally carbon and hydrogen.

When a compound containing carbon and hydrogen is completely combusted in an apparatus such as that shown in **Figure 3.13**, the carbon is converted to CO_2 and the hydrogen is converted to H_2O. (Section 3.2). From the masses of CO_2 and H_2O we can calculate the number of moles of C and H in the original sample and thereby the empirical formula. If a third element is present in the compound, its mass can be determined by subtracting the measured masses of C and H from the original sample mass.



◀ **Figure 3.13 Apparatus for combustion analysis.**

Sample Exercise 3.13

Determining an Empirical Formula by Combustion Analysis

2-propanol, sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of 2-propanol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of 2-propanol.

SOLUTION

Analyze We are told that 2-propanol contains C, H, and O atoms and are given the quantities of CO_2 and H_2O produced when a given quantity of the alcohol is combusted. We must determine the empirical formula for 2-propanol, a task that requires us to calculate the number of moles of C, H, and O in the sample.

Plan We can use the mole concept to calculate grams of C in the CO_2 and grams of H in the H_2O—the masses of C and H in the alcohol before combustion. The mass of O in the compound equals the mass of the original sample minus the sum of the C and H masses. Once we have the C, H, and O masses, we can proceed as in Sample Exercise 3.13.

Solve Because all of the carbon in the sample is converted to CO_2, we can use dimensional analysis and the following steps to calculate the mass C in the sample.

Mass CO_2 produced → Molar mass CO_2 44.0 g/mol → Moles CO_2 produced → 1 C atom per CO_2 molecule → Moles of C original sample → Molar mass C 12.0 g/mol → Mass C in original sample

Using the values given in this example, the mass of C is

$$\text{Grams C} = (0.561\ \text{g}\ \cancel{CO_2})\left(\frac{1\ \cancel{\text{mol}\ CO_2}}{44.0\ \cancel{\text{g}\ CO_2}}\right)\left(\frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol}\ CO_2}}\right)\left(\frac{12.0\ \text{g C}}{1\ \cancel{\text{mol C}}}\right)$$

$$= 0.153\ \text{g C}$$

Because all of the hydrogen in the sample is converted to H_2O, we can use dimensional analysis and the following steps to calculate the mass H in the sample. We use three significant figures for the atomic mass of H to match the significant figures in the mass of H_2O produced.

Mass H_2O produced → Molar mass H_2O 18.0 g/mol → Moles H_2O produced → 2 H atoms per H_2O molecule → Moles H in original sample → Molar mass H 1.01 g/mol → Mass H in original sample

Using the values given in this example, we find that the mass of H is

$$\text{Grams H} = (0.306\ \text{g}\ \cancel{H_2O})\left(\frac{1\ \cancel{\text{mol}\ H_2O}}{18.0\ \cancel{\text{g}\ H_2O}}\right)\left(\frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol}\ H_2O}}\right)\left(\frac{1.01\ \text{g H}}{1\ \cancel{\text{mol H}}}\right) = 0.0343\ \text{g H}$$

The mass of the sample, 0.255 g, is the sum of the masses of C, H, and O. Thus, the O mass is

$$\text{Mass of O} = \text{mass of sample} - (\text{mass of C} + \text{mass of H}) = 0.255\ \text{g} - (0.153\ \text{g} + 0.0343\ \text{g}) = 0.068\ \text{g O}$$

The number of moles of C, H, and O in the sample is therefore

$$\text{Moles C} = (0.153\ \text{g}\ \cancel{\text{C}})\left(\frac{1\ \text{mol C}}{12.0\ \text{g}\ \cancel{\text{C}}}\right) = 0.0128\ \text{mol C}$$

$$\text{Moles H} = (0.0343\ \text{g}\ \cancel{\text{H}})\left(\frac{1\ \text{mol H}}{1.01\ \text{g}\ \cancel{\text{H}}}\right) = 0.0340\ \text{mol H}$$

$$\text{Moles O} = (0.068\ \text{g}\ \cancel{\text{O}})\left(\frac{1\ \text{mol O}}{16.0\ \text{g}\ \cancel{\text{O}}}\right) = 0.0043\ \text{mol O}$$

To find the empirical formula, we must compare the relative number of moles of each element in the sample, as illustrated in Sample Exercise 3.13.

$$\text{C:}\frac{0.0128}{0.0043} = 3.0 \quad \text{H:}\frac{0.0340}{0.0043} = 7.9 \quad \text{O:}\frac{0.0043}{0.0043} = 1.0$$

The first two numbers are very close to the whole numbers 3 and 8, giving the empirical formula C_3H_8O.

Practice Exercise

The compound dioxane, which is used as a solvent in various industrial processes, is composed of C, H, and O atoms. Combustion of a 2.203-g sample of this compound produces 4.401 g CO_2 and 1.802 g H_2O. A separate experiment shows that it has a molar mass of 88.1 g/mol. Which of the following is the correct molecular formula for dioxane?
(a) C_2H_4O **(b)** $C_4H_4O_2$ **(c)** CH_2 **(d)** $C_4H_8O_2$

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.30 A 2.144-g sample of phosgene, a compound used as a chemical warfare agent during World War I, contains 0.260 g of carbon, 0.347 g of oxygen, and 1.537 g of chlorine. What is the empirical formula of this substance?

(a) CO_2

(b) $COCl_2$

(c) $C_{0.022}O_{0.022}Cl_{0.044}$

(d) C_2OCl_2

3.31 Cyclohexane, a commonly used organic solvent, is 85.6% carbon and 14.4% hydrogen by mass, with a molar mass of 84.2 g/mol. What is its molecular formula?

(a) C_6H

(b) CH_2

(c) C_5H_{24}

(d) C_6H_{12}

(e) C_4H_8

3.32 The compound dioxane, which is used as a solvent in various industrial processes, is composed of C, H, and O atoms. Combustion of a 2.203-g sample of this compound produces 4.401 g CO_2 and 1.802 g H_2O. A separate experiment shows that it has a molar mass of 88.1 g/mol. Which of the following is the correct molecular formula for dioxane?

(a) C_2H_4O

(b) $C_4H_4O_2$

(c) CH_2

(d) $C_4H_8O_2$

Exercises

3.33 Give the empirical formula of each of the following compounds if a sample contains (**a**) 0.052 mol C, 0.103 mol H, and 0.017 mol O; (**b**) 2.10 g nickel and 0.58 g oxygen; (**c**) 26.56% K, 35.41% Cr, and 38.03% O by mass.

3.34 Determine the empirical formulas of the compounds with the following compositions by mass:

(**a**) 74.0% C, 8.7% H, and 17.3% N

(**b**) 57.5% Na, 40.0% O, and 2.5% H

(**c**) 41.1% N, 11.8% H, and the remainder S

3.35 A compound whose empirical formula is XF_3 consists of 65% F by mass. What is the atomic mass of X?

3.36 What is the molecular formula of each of the following compounds?

(**a**) empirical formula CH, molar mass = 78.0 g/mol

(**b**) empirical formula OH, molar mass = 34.0 g/mol

3.37 Determine the empirical and molecular formulas of each of the following substances:

(**a**) Styrene, a compound used to make Styrofoam® cups and insulation, contains 92.3% C and 7.7% H by mass and has a molar mass of 104 g/mol.

(**b**) Caffeine, a stimulant found in coffee, contains 49.5% C, 5.15% H, 28.9% N, and 16.5% O by mass and has a molar mass of 195 g/mol.

(**c**) Monosodium glutamate (MSG), a flavor enhancer in certain foods, contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na, and has a molar mass of 169 g/mol.

3.38 (**a**) Combustion analysis of toluene, a common organic solvent, gives 5.86 mg of CO_2 and 1.37 mg of H_2O. If the compound contains only carbon and hydrogen, what is its empirical formula? (**b**) Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005-g sample of menthol is combusted, producing 0.2829 g of CO_2 and 0.1159 g of H_2O. What is the empirical formula for menthol? If menthol has a molar mass of 156 g/mol, what is its molecular formula?

3.39 Valproic acid, used to treat seizures and bipolar disorder, is composed of C, H, and O. A 0.165-g sample is combusted to produce 0.166 g of water and 0.403 g of carbon dioxide. What is the empirical formula for valproic acid? If the molar mass is 144 g/mol, what is the molecular formula?

3.40 Washing soda, a compound used to prepare hard water for washing laundry, is a hydrate, which means that a certain number of water molecules are included in the solid structure. Its formula can be written as $Na_2CO_3 \cdot xH_2O$, where x is the number of moles of H_2O per mole of Na_2CO_3. When a 2.558-g sample of washing soda is heated at 125 °C, all the water of hydration is lost, leaving 0.948 g of Na_2CO_3. What is the value of x?

Answers to Self-Assessment Exercises

3.30 (b) 3.31 (d) 3.32 (d)

3.6 | Reaction Stoichiometry

When a chemical reaction is carried out, it's vital to understand how much of each product will be produced and how much of each reactant will be consumed. To carry out chemical reactions without this knowledge can lead to unintended consequences. Maybe an expensive reactant will be wasted because much more of it was added than was needed. A reaction might generate more gas than the reaction container can hold, leading to an explosion. In some reactions, particularly those involving solids, it can be difficult to separate the desired product from excess reactants. In this section, you will learn how to calculate the quantities of reactants consumed and products produced, given a balanced chemical equation representing the reaction.

When you finish this section, you should be able to:

- Determine the number of grams (or moles) of a product formed in a chemical reaction given the number of grams (or moles) of the reactants, and vice versa

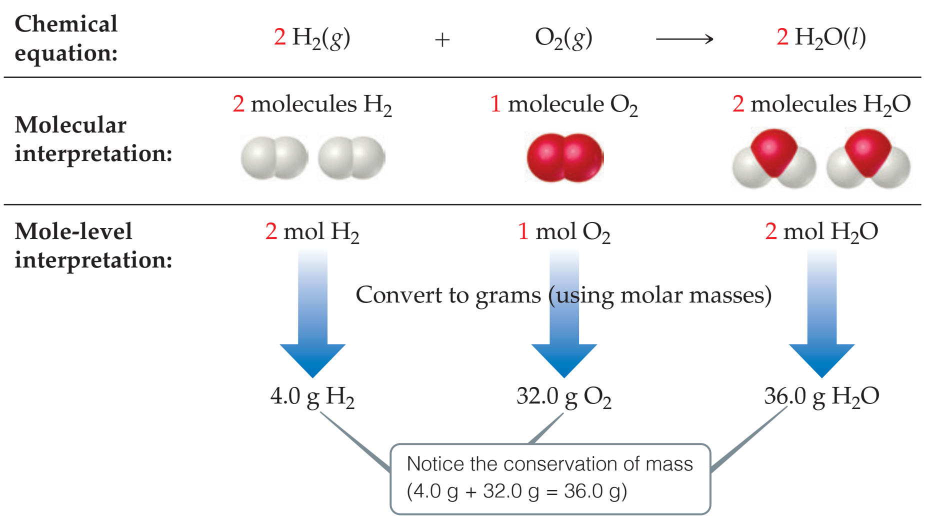
The coefficients in a chemical equation represent the relative numbers of molecules in a reaction. The mole concept allows us to convert this information to the masses of the substances in the reaction. For instance, the coefficients in the balanced equation:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l) \qquad [3.16]$$

indicate that two molecules of H_2 react with one molecule of O_2 to form two molecules of H_2O. It follows that the relative numbers of moles are identical to the relative numbers of molecules:

$$\begin{array}{ccccc} 2\,H_2(g) & + & O_2(g) & \longrightarrow & 2\,H_2O(l) \\ 2 \text{ molecules} & & 1 \text{ molecule} & & 2 \text{ molecules} \\ 2(6.02 \times 10^{23} \text{ molecules}) & & 1(6.02 \times 10^{23} \text{ molecules}) & & 2(6.02 \times 10^{23} \text{ molecules}) \\ 2 \text{ mol} & & 1 \text{ mol} & & 2 \text{ mol} \end{array}$$

We can generalize this observation to all balanced chemical equations: *The coefficients in a balanced chemical equation indicate both the relative numbers of molecules (or formula units) in the reaction and the relative numbers of moles*. **Figure 3.14** shows how this result corresponds to the law of conservation of mass.



◀ **Figure 3.14 Interpreting a balanced chemical equation quantitatively.**

The quantities 2 mol H_2, 1 mol O_2, and 2 mol H_2O given by the coefficients in Equation 3.16 are called stoichiometrically equivalent quantities. The relationship between these quantities can be represented as

$$2\text{ mol } H_2 \simeq 1\text{ mol } O_2 \simeq 2\text{ mol } H_2O$$

where the $\simeq$ symbol means "is stoichiometrically equivalent to." Stoichiometric relations such as these can be used to convert between quantities of reactants and products in a chemical reaction. For example, the number of moles of H_2O produced from 1.57 mol of O_2 is

$$\text{Moles } H_2O = (1.57\ \cancel{\text{mol } O_2})\left(\frac{2\text{ mol } H_2O}{1\ \cancel{\text{mol } O_2}}\right) = 3.14\text{ mol } H_2O$$

As an additional example, consider the combustion of butane (C_4H_{10}), the fuel in disposable lighters:

$$2\, C_4H_{10}(l) + 13\, O_2(g) \longrightarrow 8\, CO_2(g) + 10\, H_2O(g) \qquad [3.17]$$

Let's calculate the mass of CO_2 produced when 1.00 g of C_4H_{10} is burned. The coefficients in Equation 3.17 tell us how the amount of C_4H_{10} consumed is related to the amount of CO_2 produced: 2 mol $C_4H_{10} \simeq$ 8 mol CO_2. To use this stoichiometric relationship, we must convert grams of C_4H_{10} to moles using the molar mass of C_4H_{10}, 58.0 g/mol:

$$\begin{aligned}\text{Moles } C_4H_{10} &= (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\text{ mol } C_4H_{10}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right)\\ &= 1.72 \times 10^{-2}\text{ mol } C_4H_{10}\end{aligned}$$

We then use the stoichiometric factor from the balanced equation to calculate moles of CO_2:

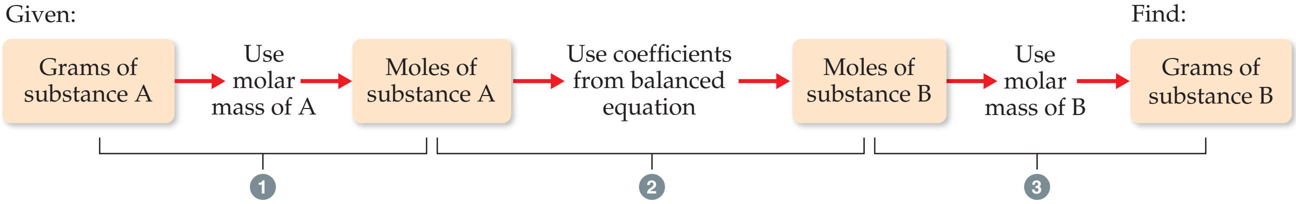
$$\begin{aligned}\text{Moles } CO_2 &= (1.72 \times 10^{-2}\ \cancel{\text{mol } C_4H_{10}})\left(\frac{8\text{ mol } CO_2}{2\ \cancel{\text{mol } C_4H_{10}}}\right)\\ &= 6.88 \times 10^{-2}\text{ mol } CO_2\end{aligned}$$

Finally, we use the molar mass of CO_2, 44.0 g/mol, to calculate the CO_2 mass in grams:

$$\begin{aligned}\text{Grams } CO_2 &= (6.88 \times 10^{-2}\ \cancel{\text{mol } CO_2})\left(\frac{44.0\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right)\\ &= 3.03\text{ g } CO_2\end{aligned}$$

This conversion sequence involves three steps, as illustrated in **Figure 3.15**. These three conversions can be combined in a single equation:

$$\begin{aligned}\text{Grams } CO_2 &= (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\ \cancel{\text{mol } C_4H_{10}}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right)\left(\frac{8\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_4H_{10}}}\right)\left(\frac{44.0\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right)\\ &= 3.03\text{ g } CO_2\end{aligned}$$



▲ **Figure 3.15 Procedure for calculating amounts of reactants consumed or products formed in a reaction.** The number of grams of a reactant consumed or product formed can be calculated in three steps, starting with the number of grams of any reactant or product.

To calculate the amount of O_2 consumed in the reaction of Equation 3.17, we again rely on the coefficients in the balanced equation for our stoichiometric factor, $2 \text{ mol } C_4H_{10} \simeq 13 \text{ mol } O_2$:

$$\text{Grams } O_2 = (1.00 \text{ g } \cancel{C_4H_{10}})\left(\frac{1 \text{ } \cancel{\text{mol } C_4H_{10}}}{58.0 \text{ } \cancel{\text{g } C_4H_{10}}}\right)\left(\frac{13 \text{ } \cancel{\text{mol } O_2}}{2 \text{ } \cancel{\text{mol } C_4H_{10}}}\right)\left(\frac{32.0 \text{ g } O_2}{1 \text{ } \cancel{\text{mol } O_2}}\right)$$
$$= 3.59 \text{ g } O_2$$

Many chemical reactions either consume or produce heat (Figure 3.7). This heat is also a stoichiometric quantity. For instance, if a reaction of a given number of reactant moles produces 100 J of energy in the form of heat, performing the reaction with twice the number of reactant moles will produce 200 J of heat. We will explore these ideas further in Chapter 5.

Sample Exercise 3.14
Calculating Amounts of Reactants and Products

Determine how many grams of water are produced in the oxidation of 1.00 g of glucose, $C_6H_{12}O_6$:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(l)$$

SOLUTION

Analyze We are given the mass of a reactant and must determine the mass of a product in the given reaction.

Plan We follow the general strategy outlined in Figure 3.16:

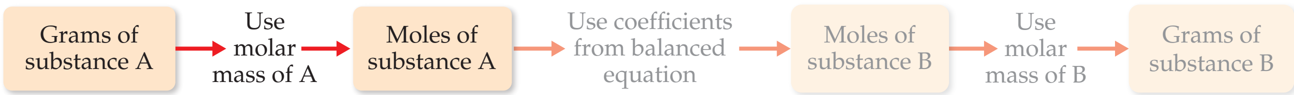
(**1**) Convert grams of $C_6H_{12}O_6$ to moles using the molar mass of $C_6H_{12}O_6$.

(**2**) Convert moles of $C_6H_{12}O_6$ to moles of H_2O using the stoichiometric relationship $1 \text{ mol } C_6H_{12}O_6 \simeq 6 \text{ mol } H_2O$.

(**3**) Convert moles of H_2O to grams using the molar mass of H_2O.

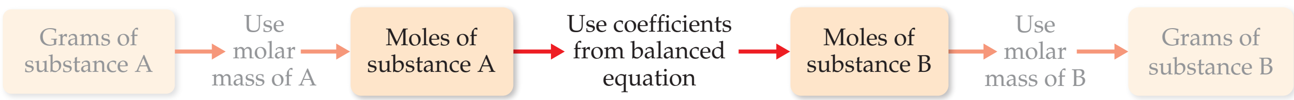
Solve

(**1**) First we convert grams of $C_6H_{12}O_6$ to moles using the molar mass of $C_6H_{12}O_6$.



$$\text{Moles } C_6H_{12}O_6 = (1.00 \text{ } \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0 \text{ } \cancel{\text{g } C_6H_{12}O_6}}\right)$$

(**2**) Next we convert moles of $C_6H_{12}O_6$ to moles of H_2O using the stoichiometric relationship $1 \text{ mol } C_6H_{12}O_6 \simeq 6 \text{ mol } H_2O$.



$$\text{Moles } H_2O = (1.00 \text{ } \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ } \cancel{\text{mol } C_6H_{12}O_6}}{180.0 \text{ } \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6 \text{ mol } H_2O}{1 \text{ } \cancel{\text{mol } C_6H_{12}O_6}}\right)$$

(**3**) Finally, we convert moles of H_2O to grams using the molar mass of H_2O.

Grams of substance A → Use molar mass of A → Moles of substance A → Use coefficients from balanced equation → Moles of substance B → Use molar mass of B → Grams of substance B

$$\text{Grams } H_2O = (1.00\text{ g } \cancel{C_6H_{12}O_6})\left(\frac{1\text{ mol } \cancel{C_6H_{12}O_6}}{180.0\text{ g } \cancel{C_6H_{12}O_6}}\right)\left(\frac{6\text{ mol } \cancel{H_2O}}{1\text{ mol } \cancel{C_6H_{12}O_6}}\right)\left(\frac{18.0\text{ g } H_2O}{1\text{ mol } \cancel{H_2O}}\right)$$

$$= 0.600\text{ g } H_2O$$

Check We can check how reasonable our result is by doing a ballpark estimate of the mass of H_2O. Because the molar mass of glucose is 180 g/mol, 1 g of glucose equals 1/180 mol. Because 1 mol of glucose yields 6 mol H_2O, we would have $6/180 = 1/30$ mol H_2O. The molar mass of water is 18 g/mol, so we have $1/30 \times 18 = 6/10 = 0.6$ g of H_2O, which agrees with the full calculation. The units, grams H_2O, are correct. The initial data had three significant figures, so three significant figures for the answer is correct.

▶ Practice Exercise

Sodium hydroxide reacts with carbon dioxide to form sodium carbonate and water:

$$2\,NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

How many grams of Na_2CO_3 can be prepared from 2.40 g of NaOH? (**a**) 3.18 g (**b**) 6.36 g (**c**) 1.20 g (**d**) 0.0300 g

Sample Exercise 3.15

Calculating Amounts of Reactants and Products

Solid lithium hydroxide is used in space vehicles to remove the carbon dioxide gas exhaled by astronauts. The hydroxide reacts with the carbon dioxide to form solid lithium carbonate and liquid water. How many grams of carbon dioxide can be absorbed by 1.00 g of lithium hydroxide?

SOLUTION

Analyze We are given a verbal description of a reaction and asked to calculate the number of grams of one reactant that reacts with 1.00 g of another.

Plan The verbal description of the reaction can be used to write a balanced equation:

$$2\,LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

We are given the mass in grams of LiOH and asked to calculate the mass in grams of CO_2. We can accomplish this with the three conversion steps in Figure 3.15. The conversion of Step 1 requires the molar mass of LiOH $(6.94 + 16.00 + 1.01 = 23.95\text{ g/mol})$. The conversion of Step 2 is based on a stoichiometric relationship from the balanced chemical equation: 2 mol LiOH $\simeq$ 1 mol CO_2. For the Step 3 conversion, we use the molar mass of $CO_2: 12.01 + 2(16.00) = 44.01$ g/mol.

Solve

$$(1.00\text{ g } \cancel{LiOH})\left(\frac{1\text{ mol } \cancel{LiOH}}{23.95\text{ g } \cancel{LiOH}}\right)\left(\frac{1\text{ mol } \cancel{CO_2}}{2\text{ mol } \cancel{LiOH}}\right)\left(\frac{44.01\text{ g } CO_2}{1\text{ mol } \cancel{CO_2}}\right)$$

$$= 0.919\text{ g } CO_2$$

Check Notice that 23.95 g LiOH/mol $\approx$ 24 g LiOH/mol, 24 g LiOH/mol $\times$ 2 mol LiOH = 48 g LiOH, and (44 g CO_2/mol)/(48 g LiOH) is slightly less than 1. Thus, the magnitude of our answer, 0.919 g CO_2, is reasonable based on the amount of starting LiOH. The number of significant figures and units are also appropriate.

▶ Practice Exercise

Propane, C_3H_8 (Figure 3.7), is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?
(**a**) 5.00 g (**b**) 0.726 g (**c**) 2.18 g (**d**) 3.63 g

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.41 In this section, we learned that when 1.00 g of butane (C_4H_{10}) reacts with 3.59 g of oxygen (O_2) it produces 3.03 g of carbon dioxide (CO_2). True or False: Is it possible using only addition and/or subtraction, to calculate the number of grams H_2O produced.

3.42 Sodium hydroxide reacts with carbon dioxide to form sodium carbonate and water:

$$2\,NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

How many grams of Na_2CO_3 can be prepared from 2.40 g of NaOH?
(a) 3.18 g
(b) 6.36 g
(c) 1.20 g
(d) 0.0300 g

3.43 Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?
(a) 5.00 g
(b) 0.726 g
(c) 2.18 g
(d) 3.63 g

3.44 Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas. How many grams of calcium hydride are needed to form 4.50 g of hydrogen?
(a) 1.11 g
(b) 2.25 g
(c) 46.9 g
(d) 93.8 g

Exercises

3.45 Hydrofluoric acid, HF(aq), cannot be stored in glass bottles because compounds called silicates in the glass are attacked by the HF(aq). Sodium silicate (Na_2SiO_3), for example, reacts as follows:

$$Na_2SiO_3(s) + 8\,HF(aq) \longrightarrow H_2SiF_6(aq) + 2\,NaF(aq) + 3\,H_2O(l)$$

(a) How many moles of HF are needed to react with 0.300 mol of Na_2SiO_3?

(b) How many grams of NaF form when 0.500 mol of HF reacts with excess Na_2SiO_3?

(c) How many grams of Na_2SiO_3 can react with 0.800 g of HF?

3.46 Several brands of antacids use $Al(OH)_3$ to react with stomach acid, which contains primarily HCl:

$$Al(OH)_3(s) + HCl(aq) \longrightarrow AlCl_3(aq) + H_2O(l)$$

(a) Balance this equation.

(b) Calculate the number of grams of HCl that can react with 0.500 g of $Al(OH)_3$.

(c) Calculate the number of grams of $AlCl_3$ and the number of grams of H_2O formed when 0.500 g of $Al(OH)_3$ reacts.

(d) Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.47 Aluminum sulfide reacts with water to form aluminum hydroxide and hydrogen sulfide. (**a**) Write the balanced chemical equation for this reaction. (**b**) How many grams of aluminum hydroxide are obtained from 14.2 g of aluminum sulfide?

3.48 Automotive air bags inflate when sodium azide, NaN_3, rapidly decomposes to its component elements:

$$2\,NaN_3(s) \longrightarrow 2\,Na(s) + 3\,N_2(g)$$

(a) How many moles of N_2 are produced by the decomposition of 1.50 mol of NaN_3?

(b) How many grams of NaN_3 are required to form 10.0 g of nitrogen gas?

(c) How many grams of NaN_3 are required to produce 10.0 ft^3 of nitrogen gas, about the size of an automotive air bag, if the gas has a density of 1.25 g/L?

3.49 A piece of aluminum foil 1.00 cm^2 and 0.550-mm thick is allowed to react with bromine to form aluminum bromide.

(**a**) How many moles of aluminum were used? (The density of aluminum is 2.699 g/cm^3.) (**b**) How many grams of aluminum bromide form, assuming the aluminum reacts completely?

3.50 The complete combustion of octane, C_8H_{18}, produces 5470 kJ of heat. Calculate how many grams of octane is required to produce 20,000 kJ of heat.

Answers to Self-Assessment Exercises

3.41 True 3.42 (a) 3.43 (d) 3.44 (c)

3.7 Limiting Reactants

Often, the reactants used in a chemical reaction are not present in precise stoichiometric amounts. For example, a natural gas-fired power plant generates electricity by producing hot gases that drive turbines, predominantly through the following combustion reaction:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g) \quad [3.18]$$

Power plants typically operate with an excess of $O_2(g)$ to achieve the maximum energy from the hydrocarbon fuel and minimize production of harmful byproducts, like carbon monoxide, that result from incomplete combustion. Consequently, the amount of CH_4 introduced determines how much CO_2 and H_2O water are produced, as well as the amount of energy released. In this section, we will learn how to do quantitative calculations for reactions where the reactants are not present in stoichiometrically equivalent quantities.

When you have completed this section, you should be able to:

- Identify limiting reactants and calculate amounts, in grams or moles, of reactants consumed and products formed in chemical reactions.
- Calculate the percent yield of a chemical reaction from the actual yield and the quantities of each reactant.

Suppose you wish to make several sandwiches using one slice of cheese and two slices of bread for each. Using Bd = bread, Ch = cheese, and Bd_2Ch = sandwich, we can represent the recipe for making a sandwich like a chemical equation:

$$2\,Bd + Ch \longrightarrow Bd_2Ch \quad [3.19]$$

If you have ten slices of bread and seven slices of cheese, you can make only five sandwiches and will have two slices of cheese left over. The amount of bread available limits the number of sandwiches.

An analogous situation occurs in chemical reactions when one reactant is used up before the others. The reaction stops as soon as any reactant is totally consumed, leaving the excess reactants as leftovers. Suppose, for example, we have a mixture of 10 mol H_2 and 7 mol O_2, which react to form water:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) \quad [3.20]$$

Because 2 mol $H_2 \triangleq$ 1 mol O_2, the number of moles of O_2 needed to react with all the H_2 is

$$\text{Moles } O_2 = (10\,\cancel{\text{mol } H_2})\left(\frac{1 \text{ mol } O_2}{2\,\cancel{\text{mol } H_2}}\right) = 5 \text{ mol } O_2$$

Because 7 mol O_2 is available at the start of the reaction, 7 mol O_2 − 5 mol O_2 = 2 mol O_2 is still present when all the H_2 is consumed.

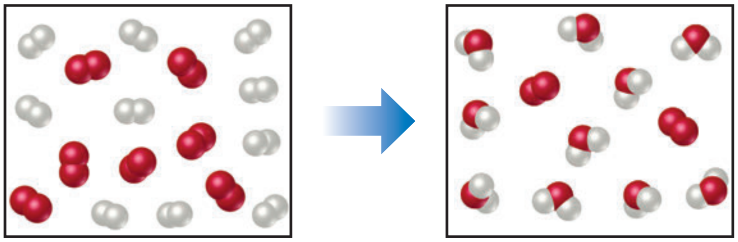
The reactant that is completely consumed in a reaction is called the **limiting reactant** because it determines, or limits, the amount of product formed. The other reactants are sometimes called *excess reactants*. In our example, shown in Figure 3.16, H_2 is the limiting reactant, which means that once all the H_2 has been consumed, the reaction stops. At that point some of the excess reactant O_2 is left over.

There are no restrictions on the starting amounts of reactants in any reaction. Indeed, many reactions are carried out using an excess of one reactant. The quantities of reactants consumed and products formed, however, are restricted by the quantity of the limiting reactant. For example, when a combustion reaction takes place in the open air, oxygen is plentiful and is therefore the excess reactant. If you run out of fuel while driving, the car stops because the fuel is the limiting reactant in the combustion reaction that moves the car. Before we leave the example illustrated in **Figure 3.16**, let's summarize the data:

	$2\,H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\,H_2O(g)$
Before reaction:	10 mol		7 mol		0 mol
Change (reaction):	−10 mol		−5 mol		+10 mol
After reaction:	0 mol		2 mol		10 mol

Go Figure If the amount of H_2 is doubled, how many moles of H_2O would have formed?

Before reaction | After reaction



$10\ H_2$ and $7\ O_2$ | $10\ H_2O$ and $2\ O_2$ (no H_2 molecules)

▲ **Figure 3.16 Limiting reactant.** Because H_2 is completely consumed, it is the limiting reactant. Because some O_2 is left over after the reaction is complete, it is the excess reactant. The amount of H_2O formed depends on the amount of limiting reactant, H_2.

The second line in the table (Change) summarizes the amounts of reactants consumed (where this consumption is indicated by the minus signs) and the amount of the product formed (indicated by the plus sign). These quantities are restricted by the quantity of the limiting reactant and depend on the coefficients in the balanced equation. The mole ratio H_2:O_2:H_2O = 10:5:10 is a multiple of the ratio of the coefficients in the balanced equation, 2:1:2. The after quantities are found by adding the before and change quantities for each column. The amount of the limiting reactant (H_2) must be zero at the end of the reaction. What remains is 2 mol O_2 (excess reactant) and 10 mol H_2O (product).

Sample Exercise 3.16
Calculating the Amount of Product Formed from a Limiting Reactant

The most important commercial process for converting N_2 from the air into nitrogen-containing compounds is based on the reaction of N_2 and H_2 to form ammonia (NH_3):

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$$

How many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?

SOLUTION

Analyze We are asked to calculate the number of moles of product, NH_3, given the quantities of each reactant, N_2 and H_2, available in a reaction. This is a limiting reactant problem.

Plan If we assume one reactant is completely consumed, we can calculate how much of the second reactant is needed. By comparing the calculated quantity of the second reactant with the amount available, we can determine which reactant is limiting. We then proceed with the calculation, using the quantity of the limiting reactant.

Solve

The number of moles of H_2 needed for complete consumption of 3.0 mol of N_2 is

$$\text{Moles } H_2 = (3.0\ \cancel{\text{mol } N_2})\left(\frac{3\text{ mol } H_2}{1\ \cancel{\text{mol } N_2}}\right) = 9.0\text{ mol } H_2$$

Because only 6.0 mol H_2 is available, we will run out of H_2 before the N_2 is gone, which tells us that H_2 is the limiting reactant. Therefore, we use the quantity of H_2 to calculate the quantity of NH_3 produced:

$$\text{Moles } NH_3 = (6.0\ \cancel{\text{mol } H_2})\left(\frac{2\text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}}\right) = 4.0\text{ mol } NH_3$$

Notice that we can calculate not only the number of moles of NH_3 formed but also the number of moles of each reactant remaining after the reaction. Notice also that although the initial (before) number of moles of H_2 is greater than the final (after) number of moles of N_2, H_2 is nevertheless the limiting reactant because of its larger coefficient in the balanced equation.

Check Examine the change row of the summary table to see that the mole ratio of reactants consumed and product formed, 2:6:4, is a multiple of the coefficients in the balanced equation, 1:3:2. We confirm that H_2 is the limiting reactant because it is completely consumed in the reaction, leaving 0 mol at the end. Because 6.0 mol H_2 has two significant figures, our answer has two significant figures.

Comment It is useful to summarize the reaction data in a table:

	$N_2(g)$ +	$3\,H_2(g)$ ⟶	$2\,NH_3(g)$
Before reaction:	3.0 mol	6.0 mol	0 mol
Change (reaction):	−2.0 mol	−6.0 mol	+4.0 mol
After reaction:	1.0 mol	0 mol	4.0 mol

Practice Exercise

When 24 mol of methanol and 15 mol of oxygen combine in the combustion reaction $2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$, what is the excess reactant and how many moles of it remains at the end of the reaction?
(a) 9 mol $CH_3OH(l)$ **(b)** 10 mol $CO_2(g)$ **(c)** 10 mol $CH_3OH(l)$ **(d)** 14 mol $CH_3OH(l)$ **(e)** 1 mol $O_2(g)$

Sample Exercise 3.17
Calculating the Amount of Product Formed from a Limiting Reactant

The reaction

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

is used to produce electricity in a hydrogen fuel cell. Suppose a fuel cell contains 150 g of $H_2(g)$ and 1500 g of $O_2(g)$ (each measured to two significant figures). How many grams of water can form?

SOLUTION

Analyze We are asked to calculate the amount of a product, given the amounts of two reactants, so this is a limiting reactant problem.

Plan To identify the limiting reactant, we can calculate the number of moles of each reactant and compare their ratio with the ratio of coefficients in the balanced equation. We then use the quantity of the limiting reactant to calculate the mass of water that forms.

Solve From the balanced equation, we have the stoichiometric relations

$$2\text{ mol } H_2 \simeq 1\text{ mol } O_2 \simeq 2\text{ mol } H_2O$$

Using the molar mass of each substance, we calculate the number of moles of each reactant:

$$\text{Moles } H_2 = (150\,\cancel{\text{g } H_2})\left(\frac{1\text{ mol } H_2}{2.02\,\cancel{\text{g } H_2}}\right) = 74\text{ mol } H_2$$

$$\text{Moles } O_2 = (1500\,\cancel{\text{g } O_2})\left(\frac{1\text{ mol } O_2}{32.0\,\cancel{\text{g } O_2}}\right) = 47\text{ mol } O_2$$

The coefficients in the balanced equation indicate that the reaction requires 2 mol of H_2 for every 1 mol of O_2. Therefore, for all the O_2 to completely react, we would need $2 \times 47 = 94$ mol of H_2. Since there are only 74 mol of H_2, all of the O_2 cannot react, so it is the excess reactant, and H_2 must be the limiting reactant. (Notice that the limiting reactant is not necessarily the one present in the lowest amount.)

We use the given quantity of H_2 (the limiting reactant) to calculate the quantity of water formed. We could begin this calculation with the given H_2 mass, 150 g, but we can save a step by starting with the moles of H_2, 74 mol, we just calculated:

$$\text{Grams } H_2O = (74\,\cancel{\text{mol } H_2})\left(\frac{2\,\cancel{\text{mol } H_2O}}{2\,\cancel{\text{mol } H_2}}\right)\left(\frac{18.0\text{ g } H_2O}{1\,\cancel{\text{mol } H_2O}}\right)$$
$$= 1.3 \times 10^2\text{ g } H_2O$$

Check The magnitude of the answer seems reasonable based on the amounts of the reactants. The units are correct, and the number of significant figures (two) corresponds to those in the values given in the problem statement.

Comment The quantity of the limiting reactant, H_2, can also be used to determine the quantity of O_2 used:

$$\text{Grams } O_2 = (74\,\cancel{\text{mol } H_2})\left(\frac{1\,\cancel{\text{mol } O_2}}{2\,\cancel{\text{mol } H_2}}\right)\left(\frac{32.0\text{ g } O_2}{1\,\cancel{\text{mol } O_2}}\right)$$
$$= 1.2 \times 10^3\text{ g } O_2$$

The mass of O_2 remaining at the end of the reaction equals the starting amount minus the amount consumed:

$$1500\text{ g} - 1200\text{ g} = 300\text{ g}.$$

Practice Exercise

Molten gallium reacts with arsenic to form the semiconductor, gallium arsenide, GaAs, used in light–emitting diodes and solar cells:

$$Ga(l) + As(s) \longrightarrow GaAs(s)$$

If 4.00 g of gallium is reacted with 5.50 g of arsenic, how many grams of the excess reactant are left at the end of the reaction?
(a) 1.20 g As **(b)** 1.50 g As **(c)** 4.30 g As **(d)** 8.30 g Ga

Theoretical and Percent Yields

The quantity of product calculated to form when all the limiting reactant is consumed is called the **theoretical yield**. The amount of product actually obtained, called the *actual yield*, is almost always less than (and can never be greater than) the theoretical yield. There are many reasons for this difference. Some of the reactants may not react, for example, or they may react in a way different from that desired (side reactions). In addition, it is not always possible to recover all of the product from the reaction mixture. The **percent yield** of a reaction is the actual yield divided by the theoretical yield, multiplied by 100 to convert to percent:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$