GENERAL, ORGANIC, AND BIOLOGICAL

CHEMISTRY

Concise

Practical

Integrated

Fourth Edition

FROST

DEAL

Periodic Table of the Elements

Main-group elements: Alkali metals (Group 1A), Alkaline earth metals (Group 2A), Groups 3A–8A; Halogens (Group 7A), Noble gases (Group 8A)

Transition elements: Groups 3B–2B

Period number	Group 1A (1)	Group 2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	Group 3A (13)	Group 4A (14)	Group 5A (15)	Group 6A (16)	Group 7A (17)	Group 8A (18)
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57* La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 208.9	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89† Ac (227)	104 Rf (267)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (281)	111 Rg (272)	112 Cn (285)	113 Nh (286)	114 Fl (289)	115 Mc (289)	116 Lv (293)	117 Ts (294)	118 Og (294)

*Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.26	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
†Actinides	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Metals | Metalloids | Nonmetals

Atomic Masses of the Elements

Name	Symbol	Atomic Number	Atomic Mass[a]	Name	Symbol	Atomic Number	Atomic Mass[a]
Actinium	Ac	89	(227)	Mendelevium	Md	101	(258)
Aluminum	Al	13	26.98	Mercury	Hg	80	200.6
Americium	Am	95	(243)	Molybdenum	Mo	42	95.94
Antimony	Sb	51	121.8	Moscovium	Mc	115	(289)
Argon	Ar	18	39.95	Neodymium	Nd	60	144.2
Arsenic	As	33	74.92	Neon	Ne	10	20.18
Astatine	At	85	(210)	Neptunium	Np	93	(237)
Barium	Ba	56	137.3	Nickel	Ni	28	58.69
Berkelium	Bk	97	(247)	Nihonium	Nh	113	(286)
Beryllium	Be	4	9.012	Niobium	Nb	41	92.91
Bismuth	Bi	83	208.9	Nitrogen	N	7	14.01
Bohrium	Bh	107	(264)	Nobelium	No	102	(259)
Boron	B	5	10.81	Oganesson	Og	118	(294)
Bromine	Br	35	79.90	Osmium	Os	76	190.2
Cadmium	Cd	48	112.4	Oxygen	O	8	16.00
Calcium	Ca	20	40.08	Palladium	Pd	46	106.4
Californium	Cf	98	(251)	Phosphorus	P	15	30.97
Carbon	C	6	12.01	Platinum	Pt	78	195.1
Cerium	Ce	58	140.1	Plutonium	Pu	94	(244)
Cesium	Cs	55	132.9	Polonium	Po	84	(209)
Chlorine	Cl	17	35.45	Potassium	K	19	39.10
Chromium	Cr	24	52.00	Praseodymium	Pr	59	140.9
Cobalt	Co	27	58.93	Promethium	Pm	61	(145)
Copernicium	Cn	112	(285)	Protactinium	Pa	91	231.0
Copper	Cu	29	63.55	Radium	Ra	88	(226)
Curium	Cm	96	(247)	Radon	Rn	86	(222)
Darmstadtium	Ds	110	(281)	Rhenium	Re	75	186.2
Dubnium	Db	105	(262)	Rhodium	Rh	45	102.9
Dysprosium	Dy	66	162.5	Roentgenium	Rg	111	(272)
Einsteinium	Es	99	(252)	Rubidium	Rb	37	85.47
Erbium	Er	68	167.26	Ruthenium	Ru	44	101.1
Europium	Eu	63	152.0	Rutherfordium	Rf	104	(267)
Fermium	Fm	100	(257)	Samarium	Sm	62	150.4
Flerovium	Fl	114	(289)	Scandium	Sc	21	44.96
Fluorine	F	9	19.00	Seaborgium	Sg	106	(266)
Francium	Fr	87	(223)	Selenium	Se	34	78.96
Gadolinium	Gd	64	157.3	Silicon	Si	14	28.09
Gallium	Ga	31	69.72	Silver	Ag	47	107.9
Germanium	Ge	32	72.63	Sodium	Na	11	22.99
Gold	Au	79	197.0	Strontium	Sr	38	87.62
Hafnium	Hf	72	178.5	Sulfur	S	16	32.07
Hassium	Hs	108	(277)	Tantalum	Ta	73	180.9
Helium	He	2	4.003	Technetium	Tc	43	(98)
Holmium	Ho	67	164.9	Tellurium	Te	52	127.6
Hydrogen	H	1	1.008	Tennessine	Ts	117	(294)
Indium	In	49	114.8	Terbium	Tb	65	158.9
Iodine	I	53	126.9	Thallium	Tl	81	204.4
Iridium	Ir	77	192.2	Thorium	Th	90	232.0
Iron	Fe	26	55.85	Thulium	Tm	69	168.9
Krypton	Kr	36	83.80	Tin	Sn	50	118.7
Lanthanum	La	57	138.9	Titanium	Ti	22	47.87
Lawrencium	Lr	103	(262)	Tungsten	W	74	183.8
Lead	Pb	82	207.2	Uranium	U	92	238.0
Lithium	Li	3	6.941	Vanadium	V	23	50.94
Livermorium	Lv	116	(293)	Xenon	Xe	54	131.3
Lutetium	Lu	71	175.0	Ytterbium	Yb	70	173.0
Magnesium	Mg	12	24.31	Yttrium	Y	39	88.91
Manganese	Mn	25	54.94	Zinc	Zn	30	65.38
Meitnerium	Mt	109	(268)	Zirconium	Zr	40	91.22

[a]Values in parentheses are the mass number of the most stable isotope.

This page intentionally left blank

CONCISE | PRACTICAL | INTEGRATED

GENERAL, ORGANIC, AND BIOLOGICAL

CHEMISTRY

FOURTH EDITION

Laura Frost
Florida Gulf Coast University

Todd Deal
Georgia Southern University

Pearson

Senior Courseware Portfolio Manager: Jessica Moro
Director of Portfolio Management: Jeanne Zalesky
Managing Producer: Kristen Flathman
Courseware Senior Analyst: Mary Ann Murray
Courseware Director, Content Development: Barbara Yien
Courseware Editorial Assistant: Amanda Davis
Content Producer: Susan McNally
Rich Media Content Producer: Paula Iborra
Full-Service Vendor: CSC
Full-Service Project Manager: Mary Tindle, SPi-Global
Copyeditor: Joanna Dinsmore
Compositor: SPi-Global
Art Coordinator: Lachina Creative, Inc.
Design Manager: Mark Ong
Interior Designer: Tamara Newnam
Cover Designer: Tamara Newnam
Rights & Permissions Project Manager: Eric Schrader
Rights & Permissions Manager: Ben Ferrini
Photo Researcher: SPi-Global
Manufacturing Buyer: Stacey Weinberger
Director of Field Marketing: Tim Galligan
Director of Product Marketing: Allison Rona
Field Marketing Manager: Chris Barker
Product Marketing Manager: Elizabeth Bell
Director MasteringChemistry Content Development: Amir Said
MasteringChemistry Senior Content Producer: Margaret Trombley
MasteringChemistry Content Producer: Meaghan Fallano

Cover Photo Credit: utima/123RF (front); Anna Kucherova/123RF (back cover)

Library of Congress Cataloging-in-Publication Data
Names: Frost, Laura D., author. | Deal, S. Todd, author.
Title: General, organic, and biological chemistry / Laura Frost (Florida Gulf Coast University), Todd Deal (Georgia Southern University).
Description: Fourth edition. | Hoboken, NJ : Pearson, [2020] | Includes index.
Identifiers: LCCN 2018055566| ISBN 9780134988696 (hardcover) | ISBN 9780134999500 (loose leaf)
Subjects: LCSH: Chemistry--Textbooks. | Biochemistry--Textbooks.
Classification: LCC QD251.3 .F76 2020 | DDC 540--dc23 LC record available at https://lccn.loc.gov/2018055566

Pearson

www.pearson.com

ISBN 10: 0-13-498869-8;
ISBN 13: 978-0-13-498869-6
(Student edition)

About the **Authors**

LAURA FROST is a Professor of Chemistry at Florida Gulf Coast University and Director of the Whitaker Center for Science, Technology, Engineering, and Mathematics (STEM) Education. She received a Bachelor's degree in chemistry from Kutztown University and her Ph.D. in chemistry with a biophysical focus from the University of Pennsylvania. She has been teaching chemistry in higher education for over 20 years and continues to teach chemistry to students in the health professions.

Professor Frost is actively engaged in teaching and learning in all STEM subjects, particularly chemistry, and uses a guided inquiry approach in her classes. She is very involved in the scholarship of teaching and learning and has demonstrated that the use of inquiry-based activities increases student learning in her one-semester chemistry course for health professions.

Dr. Frost is a member of the American Chemical Society and its Chemical Education division. A member of the University System of Georgia—Faculty Hall of Fame, Dr. Frost has been honored on more than one occasion for her teaching excellence. Her writing has also been recognized, and in 2014, she was elected to the governing council of the Textbook and Academic Authors Association.

Dr. Frost is engaged in STEM education reform through evidence-based teaching and learning at all levels (K–16), as well as through community outreach. She was principal investigator of a successful NSF grant that provides professional development for faculty in evidence-based practices in STEM. She has spoken and attended numerous conferences and workshops on this topic.

DEDICATION

This book is dedicated to my family, Baxter, Iris, and Baxter, whose constant support sustains me, and to the students and faculty who continue to inspire me. **—Laura Frost**

TODD DEAL received his B.S. degree in chemistry in 1986 from Georgia Southern College (now University) in Statesboro, Georgia, and his Ph.D. in chemistry in 1990 from The Ohio State University. He returned to his undergraduate alma mater in 1992 and has served as a faculty member and administrative leader for more than 25 years.

An award-winning educator, Dr. Deal has received numerous awards for excellence in teaching and in service, and was named Professor of the Year by the students of Georgia Southern. Professor Deal has also been recognized by Project Kaleidoscope for his innovative teaching in the sciences.

Dr. Deal is a member of American Chemical Society and its Chemical Education division. He is also a member of the Omicron Delta Kappa National Leadership Honor Society, the International Leadership Association, and the Association of Leadership Educators. Dr. Deal's teaching interests span from chemistry and the hard sciences to the soft, human sciences of leadership and leadership development.

DEDICATION

I dedicate this book to my loving wife, Karen, and to our daughters, Abbie and Anna. Thank you for believing in me. And to my students who inspired me to help them learn; this book is written for you. **—Todd Deal**

Brief Contents

Contents

Preface

To the Student

How does the body regulate carbon dioxide levels? Why are some pharmaceuticals injected and others taken orally? The key to understanding the answers to these questions starts with chemistry.

General, Organic, and Biological Chemistry was written especially for students interested in pursuing a health science career like nursing, nutrition, dental hygiene, or respiratory therapy. Yet this textbook has applications for all students interested in discovering the concepts of chemistry in everyday situations. Throughout the text, you will find that we have integrated the concepts of general, organic, and biological chemistry to create a seamless framework to help you relate chemistry to your life.

One of our goals in writing this book is to help you become better problem solvers so that you can critically assess situations at your workplace, in the news, and in your world. In this edition, we have kept the problem-solving strategies while increasing their depth to encourage greater understanding.

As you explore the pages of this book, you will encounter materials that

- apply chemistry to your life
- apply chemistry to health careers that interest you
- encourage you to develop problem-solving skills
- help you to work with and learn from your fellow students
- demonstrate how to be successful in this chemistry course and other courses.

As you read this book, you will notice that the language is less formal. Wherever possible, we relate the chemical concepts to objects in everyday life to help you understand chemistry. We also provide several study strategies with this edition, including materials for you to engage with before, during, and after class. Cognitive research in learning tells us that new ideas stick with us better if they are related to things that we already know and if we practice retrieving this information from our memories.

New to This Edition

The Fourth Edition continues to strengthen our strategy of integrating concepts from general, organic, and biological chemistry to give students a focused introduction to the fundamental and relevant connections between chemistry and life. With an emphasis on developing problem-solving skills, guiding the students' reading through Inquiry Questions, and helping students retain information through iterative retrieval practice, this text empowers students to solve problems in applied contexts relating to health and biochemistry.

- Each chapter now begins with a Learning Tip relevant to the chapter content. As students acquire the Learning Tips throughout the book, they can become more independent learners. The Learning Tips are rooted in the cognitive science literature and are supported with a list of references in the Credits section.
- While the order of chapters remains the same, some sections within the chapters have been reorganized for the better flow of concepts. In Chapter 1, we moved Section 1.6 (How Matter Changes) up to follow Section 1.2, before we begin significant figures and unit conversion. Chapter 7 now begins with Gases and Gas Laws prior to discussing attractive forces and the physical properties of liquids and solids.
- This edition includes an increased number of problems stamped with the health icon to highlight health applications of the chemistry content, making the book more relevant for students in the health professions.
- The biochemistry applications offer even more depth than the Third Edition, providing new content on drug solubility and delivery, peptides in celiac disease, common viral diseases, and CRISPR.
- We created new Practicing the Concepts videos for this edition. Each chapter now has two supporting videos. The videos, which run from 3 to 5 minutes, feature author Todd Deal. In the videos, he reviews a big idea or concept from the chapter, then helps students deepen their knowledge and develop their skills. Carefully developed visuals portray concepts vividly, and a pause-and-predict stopping point gives the student a chance for a meaningful concept check.
- Every chapter has been revised, including the sample problems and practice problems. To support areas of chapters with expanded coverage, we added new practice problems.

To the Instructor

Actively Engaging Your Students with Discovering the Concepts

From Laura Frost

Each chapter in *General, Organic, and Biological Chemistry* contains at least two guided inquiry activities, called Discovering the Concepts, at the beginning of some sections. These activities are offered to engage students in groups during class as they construct an understanding of the content in cooperation with their peers. Active learning strategies that include Discovering the Concepts have been shown to increase engagement, learning, and retention (see Freeman et al., 2014). All the information necessary to answer the questions in the activity is included at the beginning of the activity, so students should not need to use other parts of their textbook. An outline of some key points regarding the use of these activities follows.

Facilitation Faculty can facilitate the use of these activities by managing class time, guiding students to the correct answers instead of giving them the correct answers, and interjecting information during the group work where appropriate to guide student learning. It is also important for the instructor to be familiar with the activity in advance, to anticipate where students might struggle with the questioning.

Organizing Groups Faculty new to active learning often have questions about organizing student groups: Should I assign students to particular groups? Should I let students pick their own groups? Should I rotate the groups during the term? Because this course is an introductory course with few (if any) prerequisites, random assignment on the first day of class can suffice. However, I encourage diversification by gender, ethnicity, and problem-solving strategies whenever possible. With few exceptions, I have found that students become comfortable with their group members and almost insist on staying in the same group. That being said, I have met many colleagues who use active learning strategies who do rotate groups. This too can be a successful approach, with the caveat that instructors must inform students of their intention to rotate groups well in advance. Some faculty members rotate groups after an exam, some rotate them more frequently, and others keep the same groups for the entire semester. More information on group work can be found in CBE Life Sciences Evidence-Based Teaching Guide on Group Work (Wilson et al., 2018).

Skill Development Encourage students to review and extend their understanding by completing problems outside of the classroom. The practice problems at the end of each textbook section, and the additional problems and challenge problems at the end of the chapters, are ideal for this purpose.

Group Accountability To develop skill in written communication, it is essential that a record be kept of the students' activities during the class period by a student acting as group recorder. Other roles such as group manager, presenter, and technician should also be considered to keep the groups on task. An instructor may choose to grade some, all, or none of the activity, or assign participation points during the class period. However you choose to do it, group accountability should be incorporated into your grading scheme in some form.

Individual Accountability Some form of individual understanding of the activity should also be a part of the assessment. This can be done through quizzing, which also encourages skill development and retrieval practice. I give a short quiz at the beginning of a class based on the activity from the previous class. Quizzing can be done online, as a clicker quiz (works well in large classes), or on paper. The quiz should be reviewed immediately after students complete it, offering some teachable moments for student understanding.

To be efficient, the quizzing process should take no more than 10 minutes of the class period. Quizzing also has the effect of increasing student attendance and decreasing student tardiness if done at the beginning of the class. Individual accountability can also be assessed through monitoring homework, either online or on paper.

Metacognition To bring their learning full circle, students must reflect on their own learning. I provide groups with a feedback sheet that is to be completed at the end of class by one of the members of the group. The purpose of this reporting is to have students assess what they did and did not learn during the class period. The wording can be general, or changed to reflect the contents of the activity, but should provide the instructor with feedback regarding the effectiveness of the activity. The following general items are suggested for inclusion in the feedback sheet:

- List the primary topic(s) that you learned during today's activity.
- Is there anything that is still not clear regarding today's activity? Please be specific.

Many instructors are concerned with covering the required material in a course, and wonder whether an active classroom will slow the pace of coverage. I offer two thoughts related to this very real concern:

1. Because the activities help students develop problem-solving and critical-thinking skills, students are better able to apply the learned knowledge to related topics not explicitly taught in class. For example, identifying functional groups does not mean that the instructor must show the students all common functional groups in class (as may be done in a lecture). Once the concept of organic families and functional groups is introduced, students should be able to use a table to identify other functional groups, thereby maximizing instructional time in the classroom.
2. Just because a lecturer covers a topic in the classroom does not mean that a student learned it. In fact, faculty members are regularly amazed at the lack of knowledge retention from one course to the next.

I invite you to try Discovering the Concepts. If you find that these activities are helpful in engaging and reaching students, you may want to explore the full set of activities at pearsoncustomlibrary.com for use with the textbook.

I have been field-testing and revising these activities with students since 2006. My own course-based research has shown that students who use these activities perform better on the final exam and have a deeper understanding of the material (as measured by the learning level of the questions they are capable of answering) than students in my courses prior to my use of guided inquiry activities (Frost, 2010). I welcome your comments and questions at lfrost@fgcu.edu regarding their use.

Guiding Student Reading Using the Inquiry Questions

You may have noticed the Inquiry Question ? at the beginning of each section, which encapsulates the main learning goal for that section of text. Although these questions may seem obvious to the instructor, they can provide powerful guidance for the novice learner if you point them out as study strategy. Students can use these questions to guide their reading. I recommend suggesting that as they read, they jot down notes that might help to answer the Inquiry Question by the end of that section. This form of self-explanation is very useful for student retention and understanding. Research shows that when students who are less structured in their studies and think more concretely use embedded questions to guide their reading, they can retain information longer than students who don't use embedded questions (Callendar & McDaniel, 2007).

Chapter Organization and Revision

Throughout the text, we integrated general, organic, and biological chemistry topics using relevant examples and applications to solidify concepts. This text intentionally contains only 12 chapters, allowing all chapters to be covered in a single semester. Each chapter builds upon conceptual understanding and skills learned from previous chapters, providing students with an efficient path through the content and a clear context for how all of the topics connect to one another.

In this edition, at least a quarter of the chapter problems are new or have been modified. We also made extensive chapter revisions as discussed below.

1 Chemistry Basics—Matter and Measurement

- Former Section 1.6, How Matter Changes, has been moved to appear earlier in the chapter, since physical changes and chemical reactions naturally flow from understanding the nature of matter.
- The periodic table has been updated to include all elements to 118.
- The significant figures discussion in Section 1.4 now uses readings on a syringe and a digital thermometer in the discussion of measuring.
- The content on nutrition labels has been updated to reflect recent changes to nutrition labeling.

2 Atoms and Radioactivity

- The more common unit *millirem* is used to describe the biological effects of radiation.
- The Integrating Chemistry feature on "Radioisotopes and Radiation in Cancer Treatment" is expanded to include proton therapy and neutron capture therapy.

3 Compounds—How Elements Combine

- A new Solving a Problem feature, "Drawing Lewis Structures," has been added to Section 3.4.
- Table 3.7 has been updated with examples of molecules showing the preferred bonding patterns for carbon.

4 Introduction to Organic Compounds

- The chapter has been updated throughout to add a focus on the structural aspects of organic chemistry.
- The Integrating Chemistry feature "Pharmaceuticals Are Organic Compounds," has been completely rewritten with a focus on opioids.
- The Integrating Chemistry feature "Fatty Acids in Our Diets" has been updated to include recent research findings.
- The discussion of chiral molecules has been rewritten to focus on limonene.

5 Chemical Reactions

- Expanded thermodynamics coverage in Section 5.1 now includes the terms enthalpy, entropy, as well as Gibbs free energy.
- Section 5.5 now includes content on hydrolyzable and nonhydrolyzable lipids.
- The analogy used to explain equilibrium reactions has been updated for clarity.

6 Carbohydrates—Life's Sweet Molecules

- New Table 6.1 provides Fischer projections and names for D-monosaccharides containing three to six carbons, with the highest numbered chiral center highlighted.
- New Figure 6.16 consolidates the storage and structural polysaccharides into a single figure, so students can more readily compare and contrast the structures.

7 States of Matter and Their Attractive Forces: Gas Laws, Solubility, and Applications to the Cell Membrane

- The revised chapter title highlights the new conceptual balance of the chapter.
- Gas laws have moved to the beginning of the chapter.
- Section 7.1 has been expanded to include the kinetic molecular theory of gases.
- In Section 7.2, the introduction of attractive forces has been streamlined, using boiling point and vapor pressure as concrete examples.
- In Section 7.3, the content on solubility has been updated to include the terms *hydrophobic* and *hydrophilic* when referring to aqueous solutions.

8 Solution Chemistry—Sugar and Water Do Mix

- The new chapter title asks the student to look to the chapter for answers about solutions.
- The "Unique Behavior of Water" section has been updated to include coverage of specific heat.
- A reference to osmolarity in Section 8.6 connects the new concentration units to ones familiar to the student.

9 Acids, Bases, and Buffers in the Body

- Chapter 9 now has more pharmacologically relevant problems.
- Section 9.7 now discusses the relationship between pH and pK_a in terms of drug solubility and diffusion.

- A new Integrating Chemistry feature focuses on the role the kidneys and liver play in regulating CO_2 levels.
- The Henderson-Hasselbalch equation has been added to Section 9.7 as a tool for determining the ratio of conjugate base to acid.

10 Proteins—Workers of the Cell

- The discussion of amino acids as weak acids is positioned in Section 10.1 to reinforce the relationship between pH and pK_a from Chapter 9.
- A new Integrating Chemistry feature describes the role of peptides in celiac disease.

11 Nucleic Acids—Big Molecules with a Big Role

- In new Figure 11.11, the genetic code is presented using a genetic wheel instead of a table.
- A new Integrating Chemistry feature takes a closer look at common viruses.
- A new discussion of CRISPR has been added to Section 11.8.

12 Food as Fuel—An Overview of Metabolism

- New problems shift the focus toward problem solving.
- Updated art better represents the protein structures in the electron transport chain.
- Section 12.8 includes updated information on protein metabolism.

Acknowledgments

We have learned much since the first edition was published. Faculty and reviewer feedback has allowed us to enhance the Fourth Edition with some much-needed coverage while keeping the book length reasonable for a one-semester course.

The editorial staff at Pearson has been exceptional. We are extremely grateful for the assistance of Mary Ann Murray, Senior Analyst, Content Development, whose fresh eyes on the content allowed for clarity from the student perspective. Her years of textbook development were apparent. We also want to welcome back Jessica Moro to the project, Senior Courseware Portfolio Analyst, whose patience and understanding are much appreciated. We greatly appreciate the efforts of Susan McNally, Content Producer, and Mary Tindle, Project Manager, who have gone through much of the material with a fine-tooth comb, making sure that author comments were interpreted correctly by production. Thanks also to Eric Schrader, Senior Manager, Rights and Permissions, and Ben Ferini, Manager, Rights and Permissions, for their efforts on the book. We want to thank other members of the production team, including Joanna Stein, Project Manager, and Joanna Dinsmore, copyeditor. They have been very patient with us throughout the production process.

Laura Frost would like to also thank Adam Jaworski, Senior Vice President, Portfolio Management–Science, for his continued support of this project, his shared vision that an actively engaged classroom can enhance student understanding of chemistry, and his support for the inclusion of the inquiry activities. She would also like to thank Jeanne Zalesky, Director of Portfolio Management, whose strong effort on the second edition continues to allow this textbook to thrive. She also recognizes the continued mentorship and friendship of Karen Timberlake in the area of GOB chemistry writing and publishing.

Todd Deal would like, once again, to express a special appreciation to Jim Smith, our original editor, whose enthusiasm for the integrated strategy used in this project and belief in us as authors provided the foundation upon which this text is built.

This text reflects the contributions of many professors who took the time to review and edit the manuscript and provided outstanding comments, help, and suggestions. We are grateful for your contributions.

In addition, this project could not have been completed without the support of several exceptional colleagues in the Department of Chemistry at Georgia Southern University and Florida Gulf Coast University, who have taught using this text and reviewed materials, offering many comments and corrections.

If you would like to share your experience using this textbook, as either a student or faculty member, or if you have questions regarding its content, we would love to hear from you.

Laura Frost
lfrost@fgcu.edu

Todd Deal
stdeal@georgiasouthern.edu

Accuracy Reviewer

Melody Jewell, *South Dakota State University*

Reviewers

Fourth Edition

Cindy Ault, *University of Jamestown*
Allison Babij, *Ivy Tech Community College*
Karen Glover, *Clarke University*
Sarah Pierce, *Cumberland University*
Ria Ramoutar, *Georgia Southern University*
Duane Smith, *Nicholls State University*
Dennis Viernes, *University of Mary*

Previous Edition Reviewers

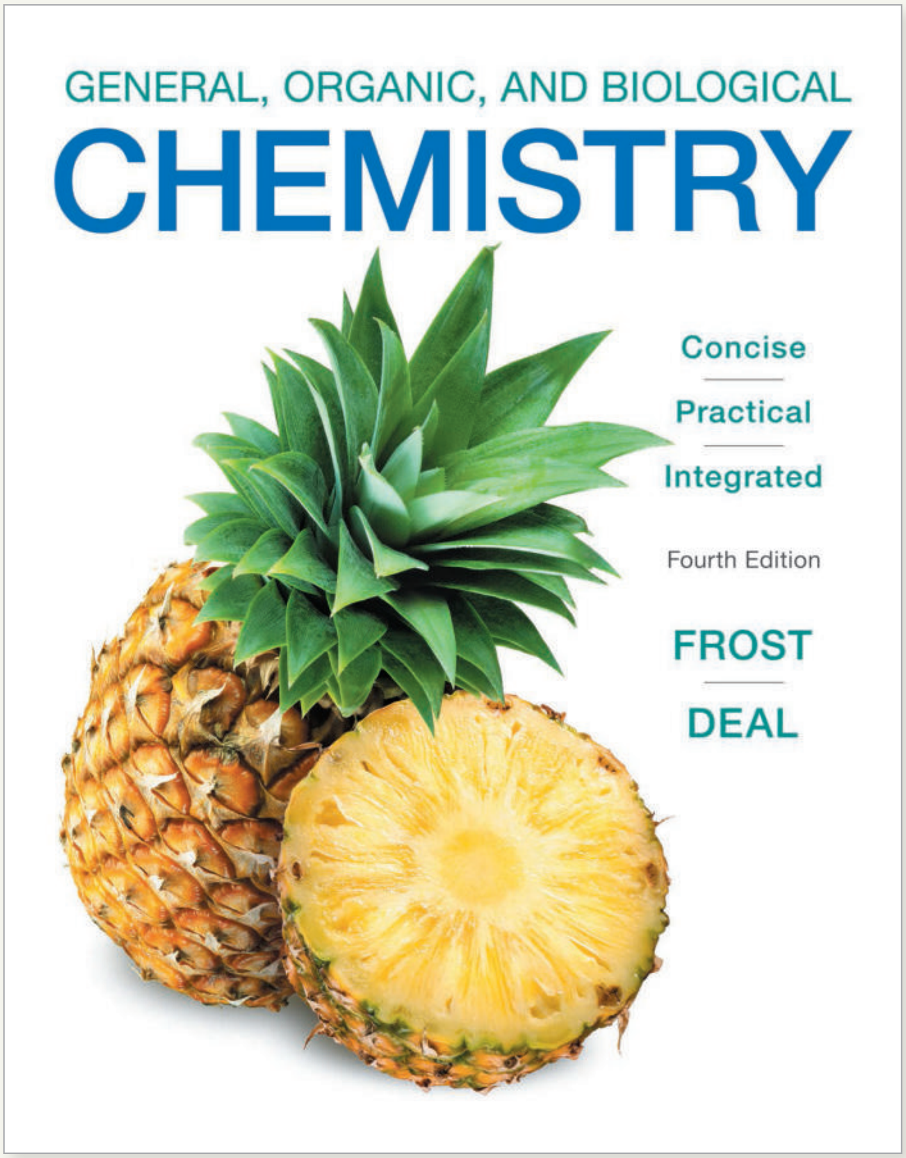
Anthony Amaro, *University of Jamestown*
Sara Egbert, *Walla Walla Community College*
Lisa Sharpe Elles, *Washburn University*
Tom Huxford, *San Diego State University*
Ria Ramoutar, *Georgia Southern University*
Tanea Reed, *Eastern Kentucky University*
John Singer, *Jackson College*
James Stickler, *Allegany College of Maryland*
Amy Taketomo, *Hartnell College*

Resources in Print and Online

Supplement	Available in Print	Available Online	Instructor or Student Supplement	Description
Mastering Chemistry (ISBN: 0-13-517528-3 / 978-0-13-517528-6)		X	Instructor and Student Supplement	The Mastering platform delivers engaging, dynamic learning opportunities—focused on your course objectives and responsive to each student's progress—that are proven to help students absorb course material and understand difficult concepts. Practicing the Concepts videos with pause-and-predict quizzes in Mastering Chemistry bring chemistry to life. These 3- to 5-minute videos feature coauthor Todd Deal introducing key topics in general, organic, and biological chemistry that students find difficult. Students are asked to solve a problem while they watch the video content. Mastering also offers Learning Catalytics questions that directly relate to the content of the text. Learning Catalytics is a "bring your own device" student engagement, assessment, and classroom intelligence system.
Instructor Solutions Manual (ISBN: 0-13-559332-8 / 978-0-13-559332-5)		X	Instructor Supplement	This Solutions Manual provides detailed solutions to all in-chapter as well as end-of-chapter exercises in the text.
Test Bank		X	Instructor Supplement	This Test Bank contains over 600 multiple-choice, true/false, and matching questions. It is available in the TestGen program, in Word format, and is included in the item library of Mastering Chemistry.
Instructor Resources (ISBN: 0-13-559331-X / 978-0-13-559331-8)		X	Instructor Supplement	This provides an integrated collection of online resources to help instructors make efficient and effective use of their time. Includes all artwork from the text, including figures and tables in PDF format for high-resolution printing, as well as four pre-built PowerPoint™ presentations. The first presentation contains the images embedded within PowerPoint slides. The second includes a complete lecture outline that is modifiable by the user. Also available are PowerPoint slides of the parent text "in-chapter" sample exercises. Also includes electronic files of the Instructor's Resource Manual, as well as the Test Bank. Access resources through http://www.pearsonhighered.com/.
Study Guide (ISBN: 0-13-416051-7 / 978-0-13-416051-1)	X		Student Supplement	This manual for students contains complete solutions to the odd-numbered end-of-chapter problems in the text.
Laboratory Manual (ISBN: 0-32-181925-X / 978-0-32-181925-3)	X		Student Supplement	Written by T. Deal, the lab manual continues the strategy of integration of concepts to help students understand chemistry. Several of the experiments included in the lab manual are original works developed by Professor Deal and his students in support of the integrated strategy. Most experiments are designed around a question, which is intended to engage students and to demonstrate the applicability of chemistry concepts to real-world problems. Many of the experiments highlight concepts from multiple chapters of the text, once again building on the strategy of integration.
Guided Inquiry Activities (ISBN: 0-13-559337-9 / 978-0-13-559337-0)	X		Instructor and Student Supplement	Guided Inquiry Activities, authored by L. Frost, are available through the Mastering Instructor Resource Area. These activities are designed for in-class use by groups of students with facilitation by the instructor. Students explore information, develop chemical concepts, and apply the concepts to further examples.

An integrated and applied approach to *General, Organic, and Biological Chemistry*

With the **Fourth Edition** of ***General, Organic, and Biological Chemistry***, authors Laura Frost and Todd Deal apply their knowledge and experience in the science of learning to this focused, concise text. Practical connections and applications are highlighted, showing both allied-health and non-science majors how to use their understanding of chemistry in future health professions and in their everyday lives. Enhanced digital tools in **Mastering Chemistry** and embedded in the Pearson eText guide students through all stages of the course, providing support when and where students need it.



Pearson

Apply the science of learning to the way students learn

DISCOVERING THE CONCEPTS

INQUIRY ACTIVITY—Reaction Energy Diagrams

Information

The diagrams shown are called reaction energy diagrams and graphically show the progress of two different chemical reactions on the *x*-axis and the amount of energy required as the reaction moves forward on the *y*-axis. The *activation energy* is the amount of energy required to get the reactants into position for collision with enough energy so that they react. Reactions with larger activation energies occur more slowly than reactions with smaller activation energies.

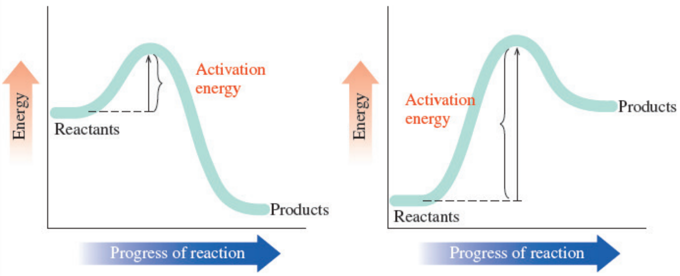


Diagram A

Diagram B

Questions

1. Which reaction has the larger activation energy?
2. Based on the diagrams, which reaction can form products more quickly?
3. A *catalyst* speeds up a chemical reaction by lowering the activation energy. Sketch diagram A and draw a second line on the same diagram for the reaction when a

p. 189

UPDATED! Discovering the Concepts can be used to engage students in groups during class as they construct an understanding of the content in cooperation with their peers.

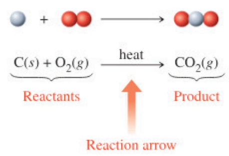
Activity: Balancing Chemical Equations

Learning Objectives

Balance a simple chemical equation

Information

Scientists use chemical formulas as a short-hand to represent substances made from the elements. Similarly, a chemical equation is used to represent how substances change during a chemical reaction.



Scheme 1. Basic parts of a chemical equation.

Questions

1. a. Fill in the following table for Scheme 1.

Element	Number in Reactants	Number in Product
Carbon		
Oxygen		

b. How is the number of each element in the reactants of a chemical equation related to the number of each element in the product?

Guided Inquiry Activities, authored by Laura Frost, engage students with more of the topics in the textbook through exploration, concept development, and application. These Activities are available in the Mastering Instructor Resources.

Relate chemistry to students' future careers

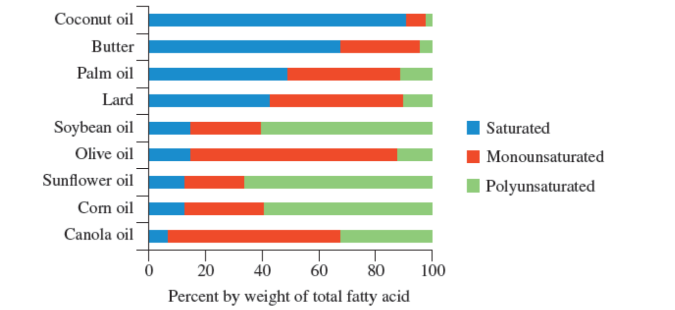
INTEGRATING Chemistry

Find out how the fatty acids in coconut oil are different from those in other oils.

Fatty Acids in Our Diets

Fats are important in a balanced diet because they play important roles as insulators and protective coverings for internal organs and nerve fibers. Mono- and polyunsaturated fats are part of a healthy diet. The Food and Drug Administration (FDA) recommends that a maximum of 30% of the calories in a normal diet come from such fatty acid-containing compounds. The FDA also recommends that the majority of our fat intake come from foods containing a higher percentage of mono- and polyunsaturated fatty acids. **TABLE 4.6** shows the fatty acid composition of some common fats. Highly saturated oils like coconut and palm oils have found uses as natural substitutes for hydrogenated oils, which are chemically saturated. More about hydrogenated oils is found in Chapter 5.

TABLE 4.6 Fatty Acid Composition of Common Dietary Fats



p. 148

EXPANDED! Topics focused on health science emphasize that good health and science literacy are critical for everyone and appear throughout the text. Examples include two new Integrating Chemistry features covering common viral diseases and gluten sensitivity.

Health-Related Problems

are integrated throughout each chapter and are tied to real-life applications from allied-health fields, helping to promote critical thinking skills and to connect the chemistry learned with their future professions.

Additionally, **clinical examples** throughout the text pay particular attention to topics such as acid-base and biochemistry.

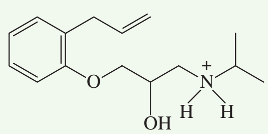
Practice Problems

9.53 The antihypertensive medication alprenolol is shown as an acid below. The pK_a for the acid is 9.6.

a. Which form is charged: acid, conjugate base, or both?

b. Which form (acid, conjugate base, or both) will predominate in a stomach with a pH between 1 and 3?

c. Which form, acid or conjugate base, will be able to more easily diffuse through a cell membrane?



Alprenolol, an antihypertensive, $pK_a = 9.6$

9.54 The anticonvulsant medication valproic acid is shown as an acid below. The pK_a for the acid is 4.8.

a. Which form is charged: acid, conjugate base, or both?

b. Which form (acid, conjugate base, or both) will predominate in the first part of the small intestine (jejunum), where the pH is between 6 and 7?

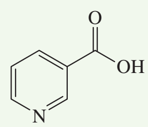
c. Which form, acid or conjugate base, will be able to more easily diffuse through a cell membrane?

9.55 Consider the vitamin niacin, which has a pK_a value of 4.85.

a. Will the acid or the conjugate base predominate at the following pH values: 3.00, 4.85, and 7.40?

b. At each pH, is that form charged or uncharged?

c. Calculate the ratio of [c. base]/[acid] present at each pH.



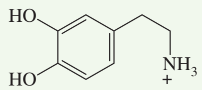
Niacin, $pK_a = 4.85$

9.56 Consider the neurotransmitter dopamine, which has a pK_a value of 8.90.

a. Will the acid or the conjugate base predominate at the following pH values: 7.45, 8.90, 13.40?

b. At each pH, is that form charged or uncharged?

c. Calculate the ratio of [c. base]/[acid] present at each pH.



Dopamine, acid form, $pK_a = 8.90$

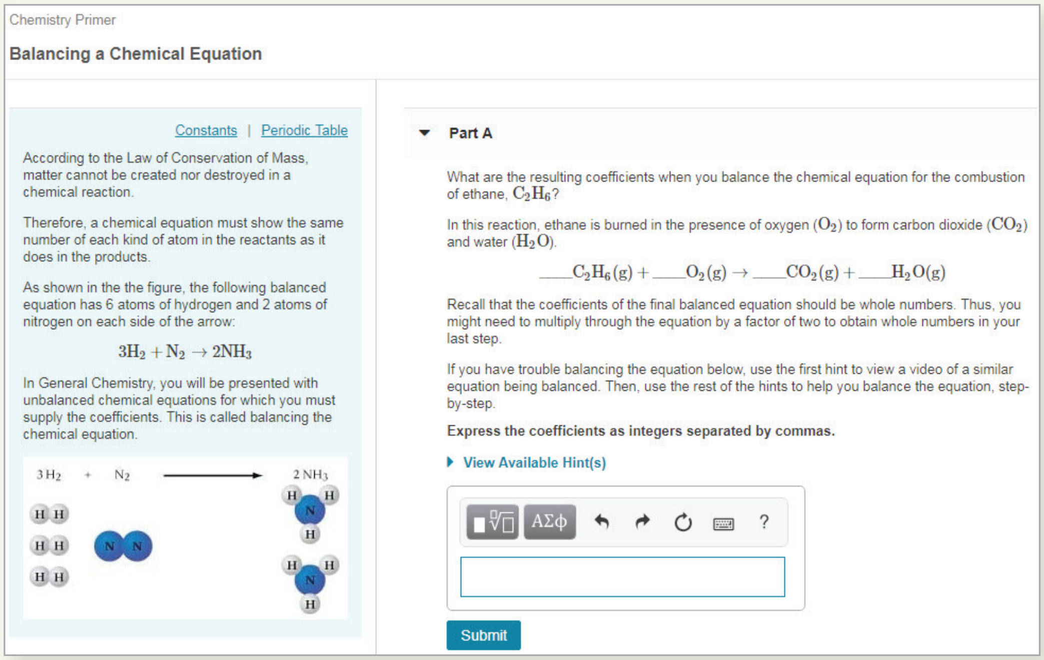
9.57 Procaine and lidocaine are used to numb the gums during dental procedures. Procaine has a pK_a of 9.1, whereas lidocaine has a pK_a of 7.9. Which do you think will relieve pain faster in the gums? Explain.

p. 379

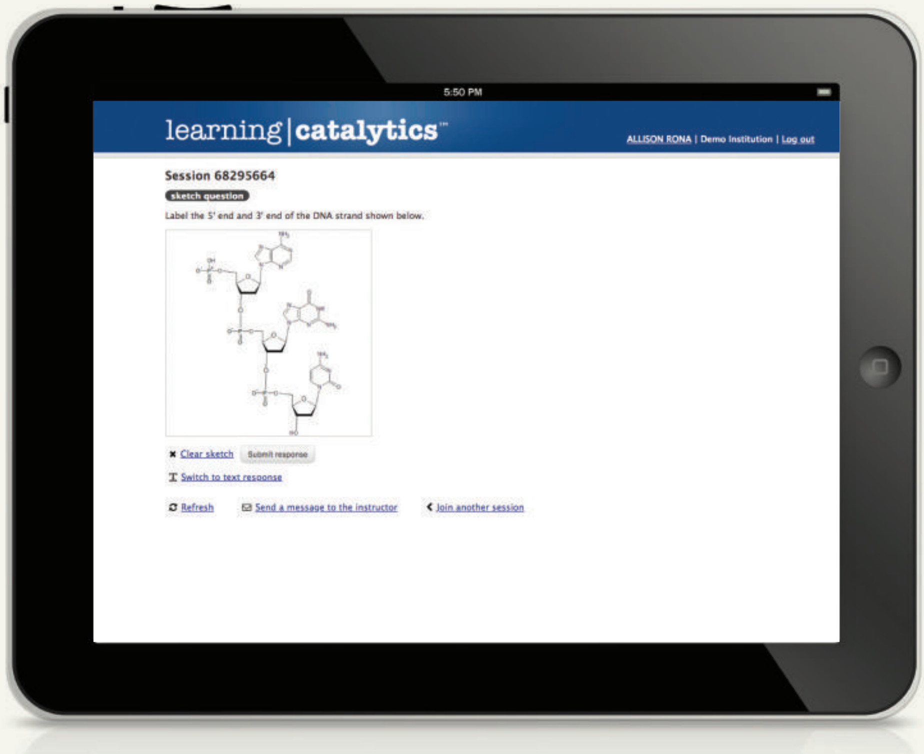
Reach every student . . .

EXPANDED! 11 new Practicing the Concepts videos explain major concepts that students struggle with in the chapter. The videos are 3–5 minutes long and are narrated by author Todd Deal, with video assessment questions written for use in Learning Catalytics, the Pearson eText, and assignable in Mastering Chemistry.

The Chemistry Primer in Mastering Chemistry helps students remediate their chemistry math skills and prepare for their first college chemistry course. Scaled to students' needs, remediation is only suggested to students that perform poorly on an initial assessment. Remediation includes tutorials, wrong-answer specific feedback, video instruction, and stepwise scaffolding to build students' abilities.

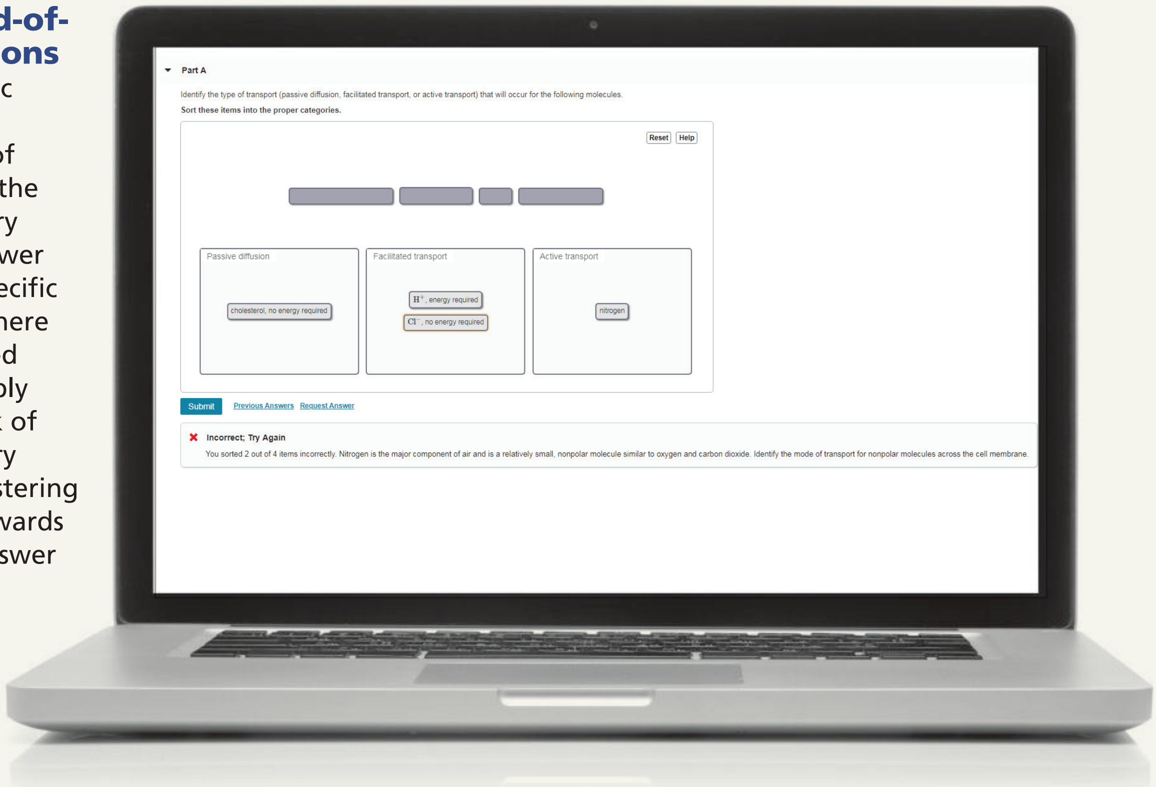


with Mastering Chemistry



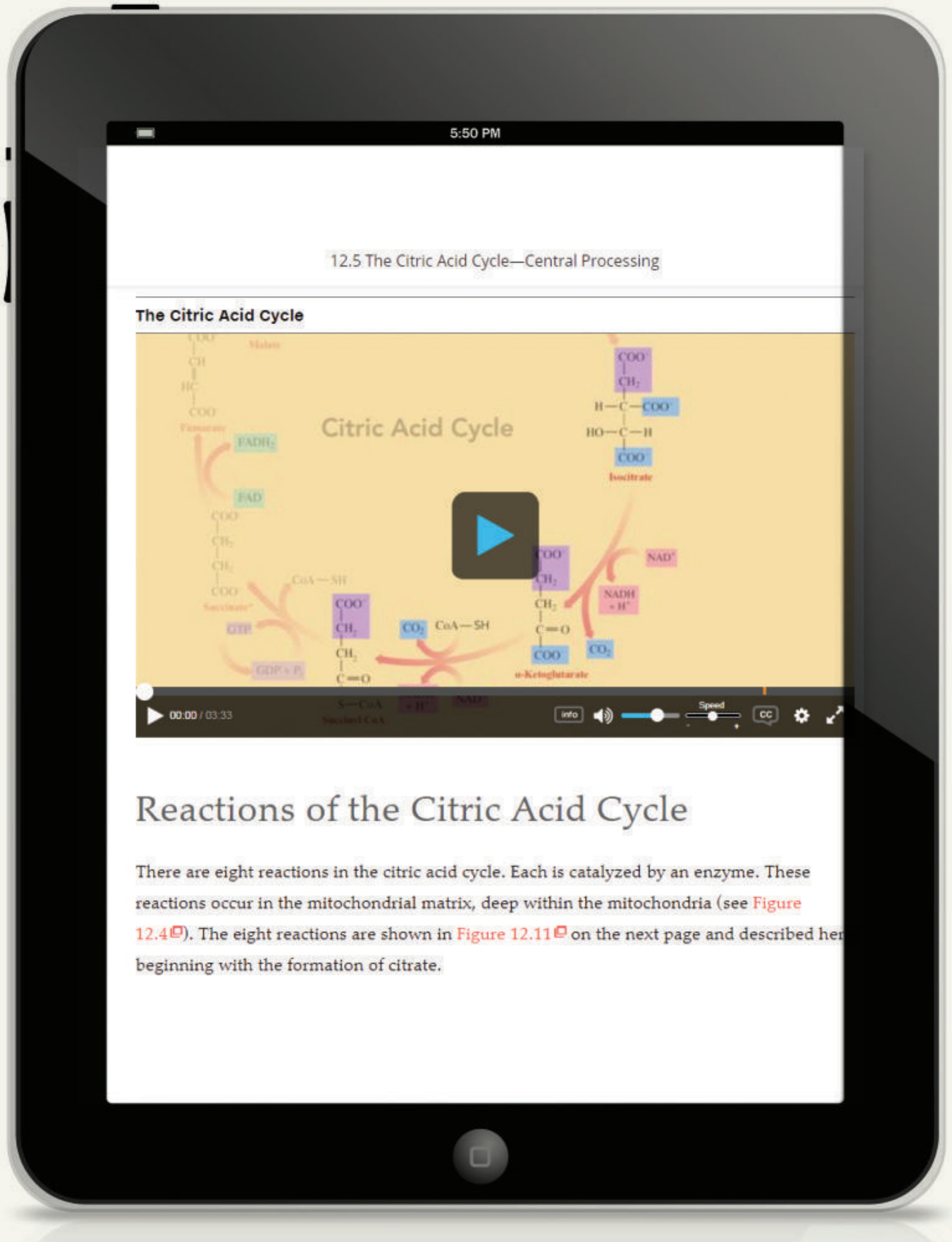
With Learning Catalytics, you'll hear from every student when it matters most. You pose a variety of questions that help students recall ideas, apply concepts, and develop critical-thinking skills. Your students respond using their own smartphones, tablets, or laptops. You can monitor responses with real-time analytics and find out what your students do—and don't—understand, to help students stay motivated and engaged.

ENHANCED! End-of-chapter questions with answer-specific feedback use data gathered from all of the students using the Mastering Chemistry to offer wrong-answer feedback that is specific to each student, where and when they need it. Rather than simply providing feedback of the "right/wrong/try again" variety, Mastering guides students towards the correct final answer without giving the answer away.

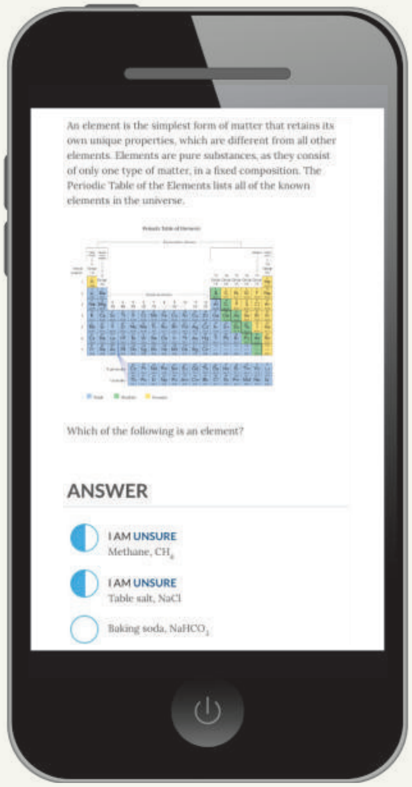


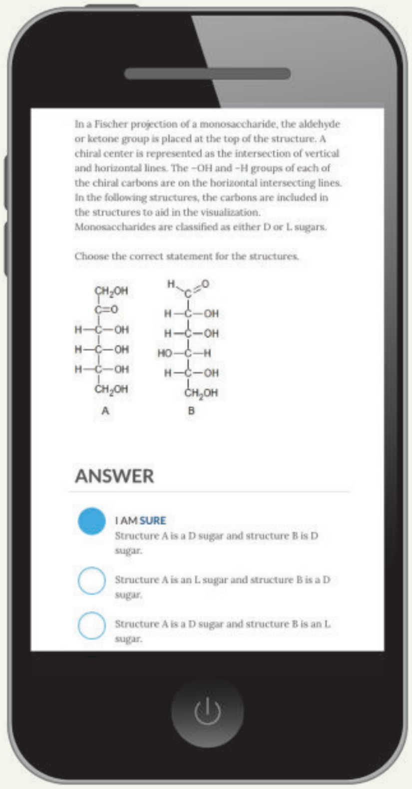
Give students anytime, anywhere access with Pearson eText

Pearson eText is a simple-to-use, mobile-optimized, personalized reading experience available within Mastering. It allows students to easily highlight, take notes, and review key vocabulary all in one place—even when offline. Seamlessly integrated videos, rich media, and embedded interactives engage students and give them access to the help they need, when they need it. Pearson eText is available within Mastering when packaged with a new book, as an upgrade students can purchase online, or can be adopted separately as your main course material.



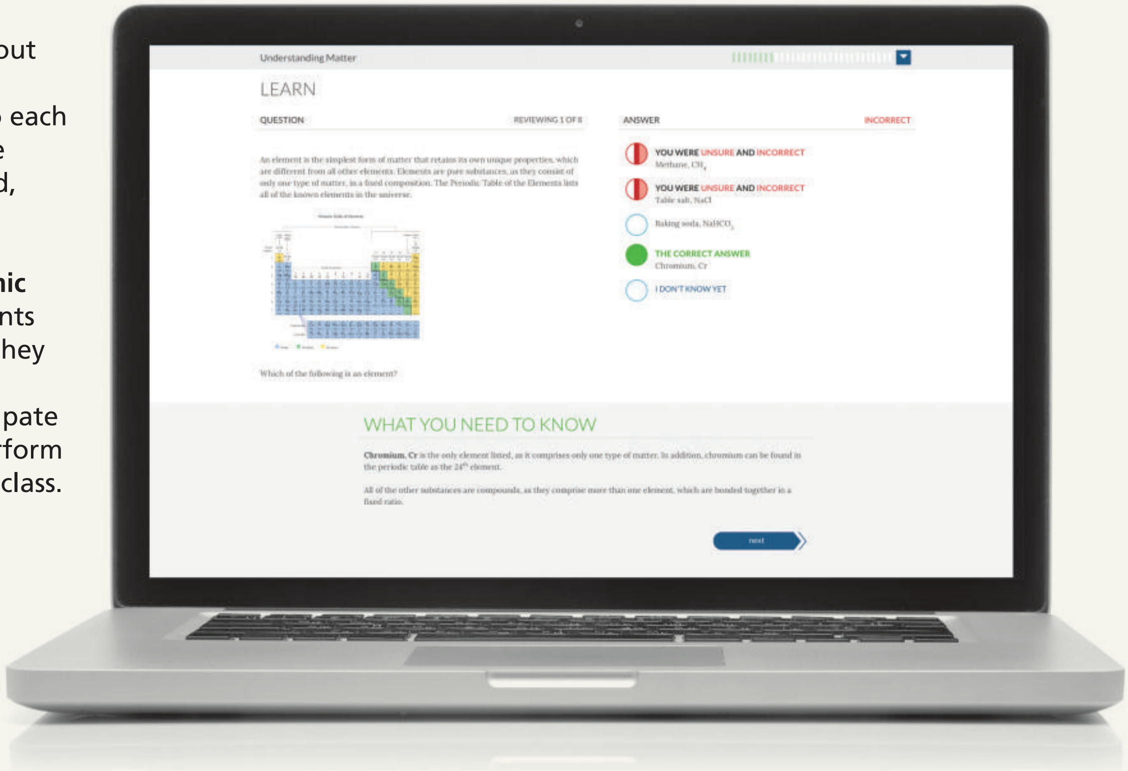
Improve learning with Dynamic Study Modules





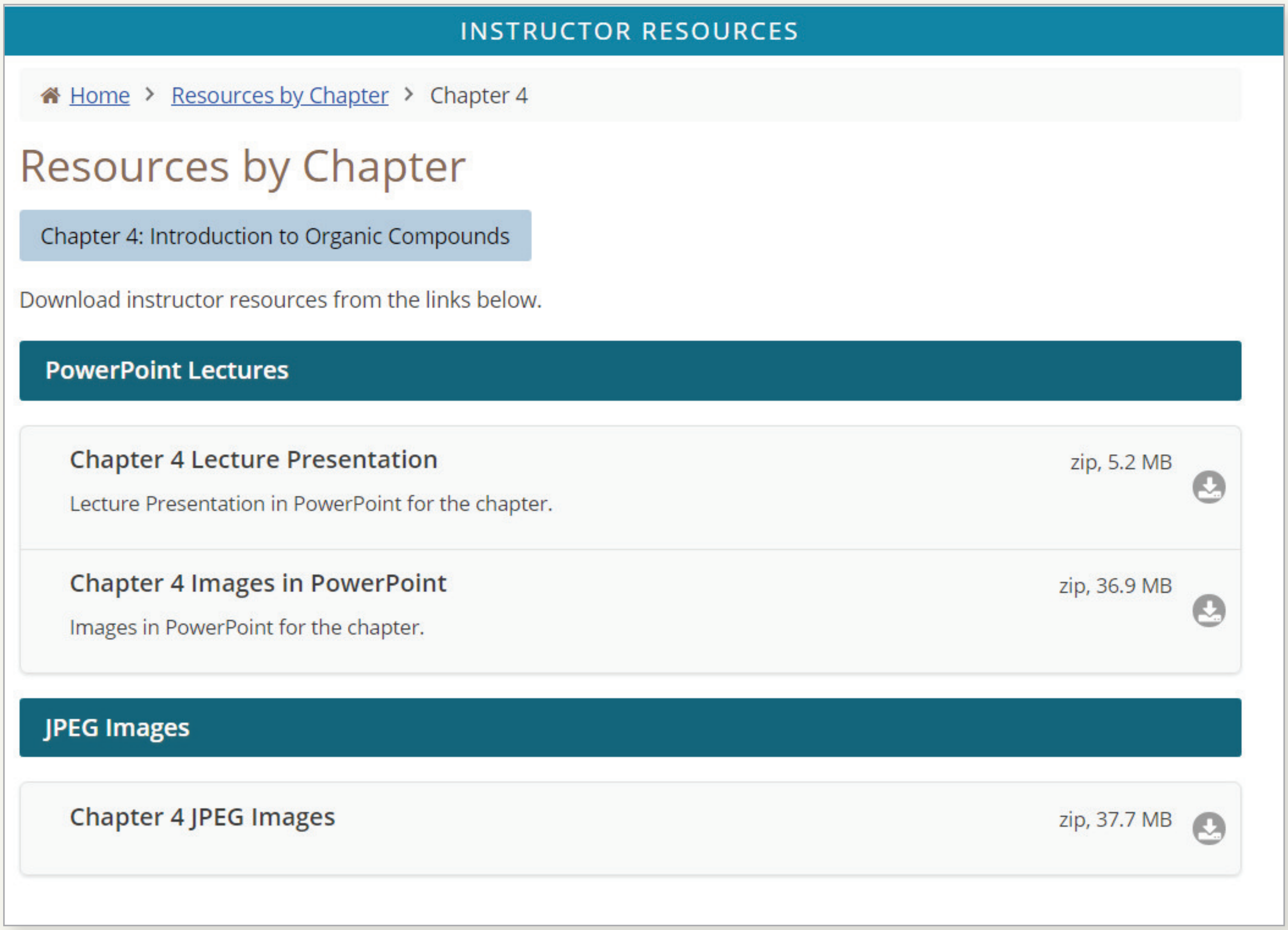
Dynamic Study Modules in Mastering Chemistry help students study effectively—and at their own pace—by keeping them motivated and engaged. The assignable modules rely on the latest research in cognitive science, using methods—such as adaptivity, gamification, and intermittent rewards—to stimulate learning and improve retention.

Each module poses a series of questions about a course topic. These question sets adapt to each student's performance and offer personalized, targeted feedback to help them master key concepts. With **Dynamic Study Modules,** students build the confidence they need to deepen their understanding, participate meaningfully, and perform better—in and out of class.



Instructor support you can rely on

General, Organic, and Biological Chemistry **includes instructor support materials in the Instructor Resources area in Mastering Chemistry. Resources include customizable PowerPoint lecture presentations and all images in JPEG format.**



1 Chemistry Basics—Matter and Measurement

We are constantly measuring things: our body height and weight, the air temperature, spaces for furniture. Even medications must be measured exactly before a dose can be given. Chemists measure matter and its changes. To introduce you to the study of chemistry and measurements, we begin with Chapter 1.

CHAPTER OUTLINE

LEARNING TIP Guide your Reading

As you read, take notes in your own words on information that answers the inquiry question that appears in the margin at the beginning of each section. Research shows that using these questions as a prompt to explain ideas in your own words helps with remembering the information later when you need it on a test.

DID YOU KNOW that everything you do every day involves chemistry? Yes, everything. From the water and shampoo in your shower, to the food you ate for breakfast, to the electronics that power your cell phone, to the therapeutic drugs for treating diseases, to the sunscreen lotion that protects your skin, and even the clothes that you wear—all of these somehow involve chemistry. Learning chemistry is really learning about how chemistry impacts our everyday lives and even provides life itself. Chemistry helps us understand concepts as diverse as how our bodies function to all of the modern conveniences that make our lives easier. So, come explore with us. It will be challenging, but fun. We promise!

All of the "stuff" that we just mentioned is composed of something that chemists call matter. **Matter** can be defined as anything that takes up space (scientists call this *volume*) and weighs something (scientists call this *mass*). From the smallest tablet dispensed by a pharmacist to the shampoo in a bottle to the air in a balloon, each of these has mass and takes up some amount of space and is a form of matter.

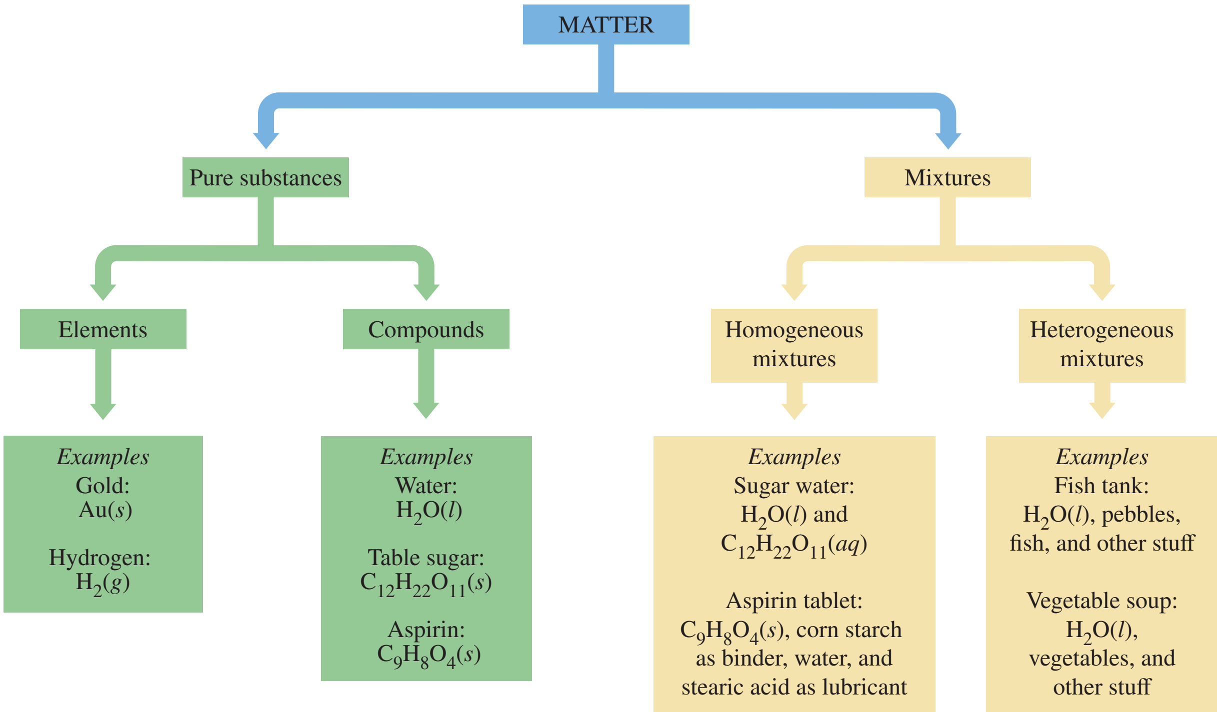
DISCOVERING THE CONCEPTS

INQUIRY ACTIVITY—Classifying Matter

Information

Chemistry is the study of matter and its changes. There are two main types of matter: *pure substances* and *mixtures* of substances. A mixture that is mostly water is called an *aqueous solution.* Matter can exist in several different *states,* the three most common being solid, liquid, or gas.

Types of Matter Flow Chart with Examples, Their Chemical Formulas, and States



Elements can exist as individual *atoms* (for example, neon [Ne]), or as pairs of atoms (for example, hydrogen, [H_2]), forming *molecules*. Atoms of different elements can also combine forming compounds.

Questions

1. Consider the examples on the flow chart. How does the formula of an element differ from that of a compound?
2. Can an element be a pure substance? Can a compound be a pure substance? How did you decide?
3. Using the information given, how might you define a pure substance? How does a pure substance differ from a mixture?
4. Based on the information in the data set, complete the following table.

Matter	Element or Compound	Atom or Molecule
He		
O_2		
CH_2O (formaldehyde)		
CH_3COOH (vinegar)		

5. As a group, devise a definition for a compound.
6. Describe the difference between a homogeneous and a heterogeneous mixture.
7. Would you classify the following matter as element, compound, or mixture? If you classify it as a mixture, classify it as homogeneous or heterogeneous.
   a. table salt (NaCl) b. nickel (Ni)
   c. chocolate chip cookie dough d. air
8. What do you think the labels (s), (*l*), (*g*), and (*aq*) on the formulas in the figure mean?

1.1 Classifying Matter: Pure Substance or Mixture

Homo sapiens is the genus and species name for the modern day human. Biologists classify all living organisms by assigning a genus and species name. In chemistry, we classify matter into different types based on both physical and chemical characteristics.

The flow chart shown in **FIGURE 1.1** gives you an idea of how we classify matter. This flow chart is your guide to classifying matter as chemists do.

As two examples of matter, consider blood and a diamond. If you are a blood donor, you may have heard blood discussed in terms of whole blood, plasma, white blood cells, and red blood cells. These terms give you a hint that blood is a mixture containing many components. In contrast, if you have ever bought a diamond, you know that one of the measures used to grade these gemstones is clarity, which is a measure of purity. The clearer (more pure) a diamond is, the more it costs. So, how do the complexity of a sample of blood and the purity of a diamond contribute to how they are classified?

Looking at the flow chart in Figure 1.1 from the top, we see that all matter is divided into two large categories or classifications—pure substances and mixtures. Let's consider these one at a time.

What's an Inquiry Question?

Inquiry Questions are designed to focus your reading on the main concepts by section. As you read each section, see if you can answer its Inquiry Question in your own words.

1.1 Inquiry Question

How is matter classified?

Pure Substances

Diamonds are composed only of carbon. Because a diamond is composed of a single substance, carbon, it is not a mixture but rather a pure substance. A **pure substance** is matter that is made up of only one type of substance and can be represented with one chemical formula.

Note from the flow chart that pure substances can be one of two types—elements or compounds. Substances containing only one type of atom are considered elements, with

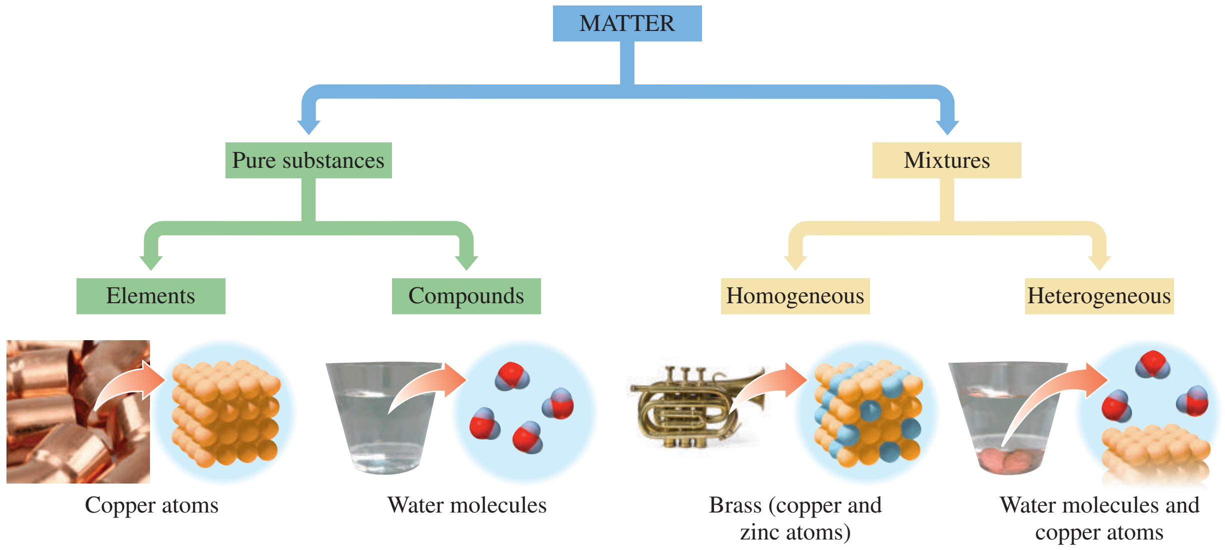


FIGURE 1.1 Flow chart used for classifying matter. Matter can be broadly classified as a pure substance or a mixture. Pure substances can be further classified as elements or compounds, and mixtures can be further classified as homogeneous or heterogeneous. The red balls on the water molecules represent oxygen, and the gray balls represent hydrogen.

diamonds being one such example. An **element** is the simplest type of matter because it is made up of only one type of atom. An **atom** is the smallest unit of matter that can exist and keep its chemically unique characteristics. Substances that contain more than one type of atom, but contain these elements in a fixed composition, are also pure substances. Pure water, for example, contains only water and has the chemical formula H_2O. Water that is distilled contains only the substance water, H_2O, and is therefore a pure substance. Most water on the planet, however, is not pure water (H_2O), but has other substances dissolved in it so is a mixture. Sea water is very salty; spring water and even purified water contain dissolved solids that make the water taste better to us.

Because pure water has more than one type of atom—it contains both hydrogen and oxygen—it is not an element but is instead a compound. A **compound** is a pure substance that is made of two or more elements that are chemically joined together.

Blood is a mixture of many components, whereas a diamond is composed of a single substance, the element carbon. A specialist in blood banking separates and stores blood.

Mixtures

As noted earlier, blood contains red blood cells, white blood cells, plasma (which is mostly water), and several other components. Blood is an example of a **mixture**, a combination of two or more substances. One of the defining characteristics of a mixture is that it can be separated into its different components. A specialist in blood banking is trained to separate blood into red blood cells and plasma for storage and use in different medical procedures at a blood bank.

Next, we consider different types of mixtures. Air contains nitrogen, oxygen, carbon dioxide, argon, and small amounts of other gases. Every time you take a breath, you inhale the same substances in the same ratio as another person breathing the same air. Air is classifed as a **homogeneous mixture** because its composition is the same throughout. In contrast, trail mix is a **heterogenous mixture** because its composition is not uniform, but varies throughout. One handful of trail mix may contain raisins, chocolate chips, and a peanut while another handful contains sunflower seeds, almonds, and dried banana. No two handfuls contain the same substances in the same amount.

Pure substances are easier to characterize, but chemists do work with and study mixtures. The air we breathe, the food we eat, the concrete we walk on are all mixtures. Most matter we encounter every day is a mixture. However, due to their varied composition, mixtures are not easily characterized.

Sample Problem 1.1 Mixture Versus Pure Substance

Classify each of the following substances as a mixture or a pure substance:

a. cake batter
b. the helium gas inside a balloon

Solution

a. Cake batter is a mixture of substances such as flour, butter, sugar, and other ingredients.
b. The helium gas in the balloon is a pure substance because it contains only one substance, the element He.

Sample Problem 1.2 Classifying Mixtures

Classify each of the following mixtures as homogeneous or heterogeneous. Briefly justify your answer.

a. olive oil
b. rocky road ice cream

Solution

a. Olive oil is a homogeneous mixture. It has the same composition throughout.
b. Rocky road ice cream is a heterogeneous mixture. Each spoonful contains different amounts of nuts, marshmallow, and ice cream.

Practice Problems

Find the answers to the odd-numbered Practice Problems at the end of the chapter.

1.1 Classify each of the following substances as a pure substance or a mixture:

a. iron

b. gelato

c. cinnamon

d. spring water

1.2 Classify each of the following substances as a pure substance or a mixture:

a. well water

b. ocean water

c. cement

d. a tropical smoothie

1.3 Classify each of the following mixtures as homogeneous or heterogeneous. Briefly justify your answer.

a. a bowl of vegetable soup

b. mouthwash

c. an unopened can of cola

d. a dinner salad

1.4 Classify each of the following mixtures as homogeneous or heterogeneous. Briefly justify your answer.

a. a bottle of sports drink

b. a blueberry pancake

c. gasoline

d. a box of raisin bran

DISCOVERING THE CONCEPTS

INQUIRY ACTIVITY—The Periodic Table

All *elements* are listed individually on the *periodic table of the elements.* There is a periodic table on the inside front cover of your textbook and an alphabetical listing of all the elements on the facing page. Refer to this as you answer the questions in this activity.

The horizontal rows on the periodic table are referred to as *periods*, and the vertical columns are referred to as *groups.* The groups are numbered across the top of the periodic table, and the periods are numbered down the left side of the table. The table is organized with metals on the left and nonmetals on the right, with a staircase dividing line between the two. Find this staircase in the periodic table on the inside front cover.

Water's chemical formula is H_2O, meaning that it contains two hydrogen atoms and one oxygen atom. Chemical formulas show the type and number of each element present in a compound.

Questions

1. In which group on the periodic table are the following elements found?
   a. sodium b. oxygen c. calcium d. carbon
2. In which period are the following elements found?
   a. hydrogen b. nitrogen c. sulfur d. phosphorus
3. Provide names for the following elements and identify them as metals or nonmetals:
   a. Cu b. Na c. Cl d. C e. K f. P
4. Look at the periodic table of the elements. About how many elements are there?
5. In each of the chemical formulas below, identify the number of atoms of each element and provide the name of each element:
   a. $C_6H_{12}O_6$ (dextrose) b. NaOH (lye, found in drain cleaners)
   c. $NaHCO_3$ (baking soda) d. $C_{15}H_{21}NO_2$ (Demerol, a painkiller)
6. Are most of the elements on the periodic table metals or nonmetals? The Earth's biomass is made up mostly of carbon, hydrogen, and oxygen. Are these elements metals or nonmetals?

1.2 Elements, Compounds, and the Periodic Table

1.2 Inquiry Question

How is the periodic table organized?

If a pure substance can be an element or a compound, how do we distinguish which is which? The tool to guide us is the **periodic table of the elements**. At its most basic level, the periodic table is a listing of all the elements found on Earth.

You may have previously encountered the periodic table, most likely hanging on the wall in a science classroom. But have you ever tried to decipher it? This strangely shaped chart contains an amazing collection of data, both in the way that it is organized and in the actual information listed. We will make extensive use of the periodic table and the information it contains. Take a look at the periodic table in **FIGURE 1.2**.

The periodic table consists of many small blocks in a simple layout. Each has a letter or two in its center and numbers above and below these letters. The letters are known as the **chemical symbol** and represent the name of each element. For many of the elements, the symbols are derived directly from the name of the element—for example, lithium = Li, carbon = C, hydrogen = H, oxygen = O, and so on. A few of the elements have symbols that do not match the first few letters of their name—for example, sodium = Na and gold = Au. These elemental symbols are derived from the Latin names for the element, natrium and aurum, respectively. Because the elements form the

Periodic Table of the Elements

Main-group elements: Group 1A (1) Alkali metals; Group 2A (2) Alkaline earth metals; Groups 3A–8A (13–18), with Group 7A (17) Halogens and Group 8A (18) Noble gases. Transition elements: Groups 3B–2B (3–12).

Period number	Group 1A (1)	Group 2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	Group 3A (13)	Group 4A (14)	Group 5A (15)	Group 6A (16)	Group 7A (17)	Group 8A (18)
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57* La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 208.9	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89† Ac (227)	104 Rf (267)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (281)	111 Rg (272)	112 Cn (285)	113 Nh (286)	114 Fl (289)	115 Mc (289)	116 Lv (293)	117 Ts (294)	118 Og (294)

*Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.26	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
†Actinides	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Metals | Metalloids | Nonmetals

FIGURE 1.2 **The periodic table of the elements.**

basic vocabulary of chemistry, it is important to know the names and symbols of the more common elements, especially those that are found in living things (carbon, hydrogen, oxygen, nitrogen [N], phosphorus [P], and sulfur [S]).

The blocks on the periodic table are organized in specific arrangements. Each block holds a different element. A vertical column of blocks is known as a **group** of elements. (Some chemists refer to the vertical columns as families; we will use the term groups in this textbook.) The elements in a group have similar chemical behaviors. Each group has a number and letter designation. The groups with A designations (1A–8A) are known as the main-group elements, and the groups with B designations are the transition elements. This system of numbering the groups is most common in North America and will be used throughout this textbook. The system using the numbers 1 through 18 for the columns has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) and is also in common use. Several of the groups have special names to designate the members of the group as shown at the top of Figure 1.2.

A horizontal row of blocks is known as a **period**. The periods are numbered from 1 to 7, with sections of Periods 6 and 7 set apart at the bottom of the periodic table. The bold, descending staircase beginning at the element boron (#5) and running diagonally down and to the right separates the metals from the nonmetals. Elements bordered by the line, with the exception of aluminum (Al), are metalloids. Metalloids contain some properties of metals and some of nonmetals.

Metals and nonmetals have different properties. Aluminum can be made into sheets or blocks and is a metal. Carbon is more brittle, and we find it as a powder or as the "lead" in a pencil. Carbon is a nonmetal.

Elements in Nutrition

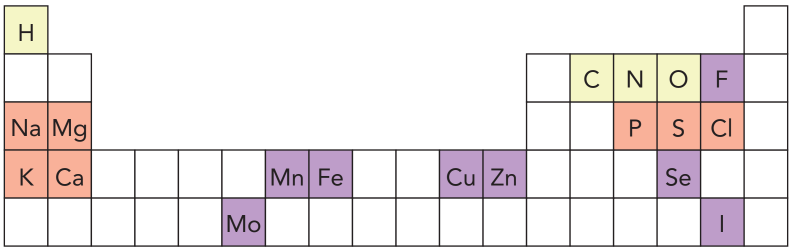
INTEGRATING Chemistry

Find out which elements the body uses.

Which elements are found in the food we eat? The elements found most commonly in food are the nonmetals carbon, hydrogen, oxygen, and nitrogen, or the CHON elements. These four elements also make up most of the biological molecules on Earth. Along with the CHON elements, the U.S. Food and Nutrition Board at the National Academies has also identified several other elements on the periodic table as being essential for human health. These elements are classified as macronutrients if required in quantities greater than 100 milligrams per day or as micronutrients if required in trace quantities, less than 100 milligrams per day.

Macronutrients include several elements—sodium, magnesium, potassium, calcium, and chlorine—that are found in the body as charged particles and are essential for the transport of electrical signals through cells. Micronutrients are often found concentrated in particular organs. For example, iodine is found in the thyroid, fluorine in the teeth, and zinc mainly in muscle and bone. Red blood cells contain the protein hemoglobin, which contains iron. **TABLE 1.1** summarizes the uses and locations of the micronutrients.

Chromium, once considered a micronutrient, has recently been suggested to be nonessential because removal of this element from the diet has been shown to have no adverse health effects and no known molecule in the body uses chromium.



Most common elements in living things

Macronutrients

Micronutrients

—Continued next page

Continued—

TABLE 1.1 Micronutrients

Element	Use in Body	Main Location in Body
Copper (Cu)	Used in several enzymes that are essential for human health and development	Liver
Fluorine (F)	Mineralizes teeth and bones	Teeth and bone
Iodine (I)	Found in thyroid hormones	Thyroid
Iron (Fe)	Found in hemoglobin, a protein that transfers oxygen from lungs to tissues	Red blood cells
Manganese (Mn)	Found in enzymes that act as antioxidants, enzymes involved in bone development and wound healing	Pituitary gland, bones, and connective tissue
Molybdenum (Mo)	Used with enzymes that control oxidation processes	Liver and kidney
Selenium (Se)	Used in proteins that regulate reproduction, thyroid, DNA synthesis, and oxidative damage	Skeletal muscle, thyroid, immune system
Zinc (Zn)	Used in cell growth and division, and in immune system function; in males, zinc is vital to fertility	Throughout the body, particularly in bones and muscle

Sample Problem 1.3 Using the Periodic Table

Use the periodic table to answer the following questions:

a. What is the elemental symbol for oxygen?
b. To what group does oxygen belong?
c. What period is oxygen in?

Solution

a. Oxygen's symbol is O.
b. Oxygen is located in Group 6A.
c. Oxygen is in Period 2.

Compounds

How do we classify substances like water (H_2O) or table salt (NaCl) that contain more than one element from the periodic table? Remember that we said a pure substance containing two or more elements chemically combined is classified as a compound. Therefore, both water and table salt are compounds. Compounds combine elements in specific ratios. How and why elements combine in these ratios are discussed in Chapter 3.

Sample Problem 1.4 Classifying Pure Substances

Classify each of the following as an element or compound and explain your classification:

a. carbon dioxide (the gas we breathe out)
b. xenon (a gas sometimes used as an anesthetic)

Solution

a. Carbon dioxide is a compound. From its name, you can tell that this substance contains carbon and "something else" (even if you did not know that oxide is a name often used for oxygen when it is in compounds). Because carbon dioxide contains two different elements, it is a compound.
b. Xenon is an element. Look for xenon on the periodic table. You will find it in group 8A (18). This gas is becoming more common as an anesthetic agent because unlike other anesthetics—such as nitrous oxide, which depletes the ozone layer of the atmosphere—it has minimal effects on the environment.

Before we continue, let's take a quick look at the representations we use for compounds. Why does the representation for water (H_2O) contain the subscript 2 after hydrogen, but the one we use for table salt (NaCl) does not have a subscript? These representations, known as **chemical formulas**, show that water is composed of two atoms of hydrogen and one atom of oxygen and table salt is composed of one sodium atom and one chlorine atom. The subscript tells us how many atoms of the preceding element are present in the compound. The absence of a subscript on oxygen, sodium, and chlorine in the formulas for water and sodium chloride is understood to mean *one* of each element. A compound's chemical formula identifies which element and how many atoms of that element are present in a compound.

Sample Problem 1.5 Elements in Compounds

Identify and give the number of atoms of each element in the following compounds:

a. $C_2H_4O_2$, acetic acid, a component of vinegar
b. H_3PO_4, phosphoric acid

Solution

a. Acetic acid contains 2 atoms of carbon, 4 atoms of hydrogen, and 2 atoms of oxygen.
b. Phosphoric acid contains 3 atoms of hydrogen, 1 atom of phosphorus, and 4 atoms of oxygen.

Practice Problems

1.5 Classify each of the following as an element or compound and explain your classification:

a. aluminum, used in soft drink cans and foil
b. sodium hypochlorite, found in bleach
c. hydrogen, fuel of the sun
d. potassium chloride, a table-salt substitute

1.6 Classify each of the following as an element or a compound:

a. Fe b. $CaCl_2$ c. Si d. KI

1.7 Use the periodic table to supply the missing information in the following chart:

Name	Elemental Symbol	Group	Period	Metal or Nonmetal
Fluorine				
		1A	2	
	Cl			
		4A	2	Nonmetal

—*Continued next page*

Continued—

1.8 Use the periodic table to supply the missing information in the following chart:

Name	Elemental Symbol	Group	Period	Metal or Nonmetal
Sodium				
		2A	4	
	P			
		5A	2	Nonmetal

1.9 Identify and give the number of atoms of each element in the following compounds:

a. $LiCO_3$, lithium carbonate, used in the treatment of bipolar disorder

b. H_2O_2, hydrogen peroxide, a disinfectant

c. $C_8H_8O_3$, vanillin, the primary flavor compound in vanilla beans

1.10 Identify and give the number of atoms of each element in the following compounds:

a. $MgSO_4$, magnesium sulfate, commonly called Epsom salts

b. $NaHCO_3$, sodium bicarbonate, commonly called baking soda

c. $C_{14}H_{18}N_2O_5$, aspartame, an artificial sweetener

1.3 How Matter Changes

? 1.3 Inquiry Question

How are physical and chemical changes in matter different, and how can we represent them?

On a hot summer day, you add several ice cubes to your glass of tea and come back later to find that the ice has disappeared. Now your glass seems to have two liquids—the tea at the bottom and a new, clear liquid on top. You may have experienced the popular experiment where a clear, vinegar-smelling liquid is poured into a papier-mâché volcano and then a foamy "lava" flows out. Each of these instances represents a change in matter, but each is a different type of change: One is a physical change, and the other is a chemical reaction.

Physical Change

The ice in a glass of tea is water in the solid state. As it melts, the ice changes from a solid to a liquid (forming a clear liquid on top of your tea). The identity of the water did not change—both ice and liquid water are still H_2O (see **FIGURE 1.3**). In fact, if you add enough heat to the liquid water, it will eventually become steam—a gas—but it is still water (H_2O). A change in the state of matter—from a solid to a liquid or a liquid to a gas, for example—represents a physical change of matter. In a **physical change**, the form of the matter is changed, but its identity or chemical formula remains the same.

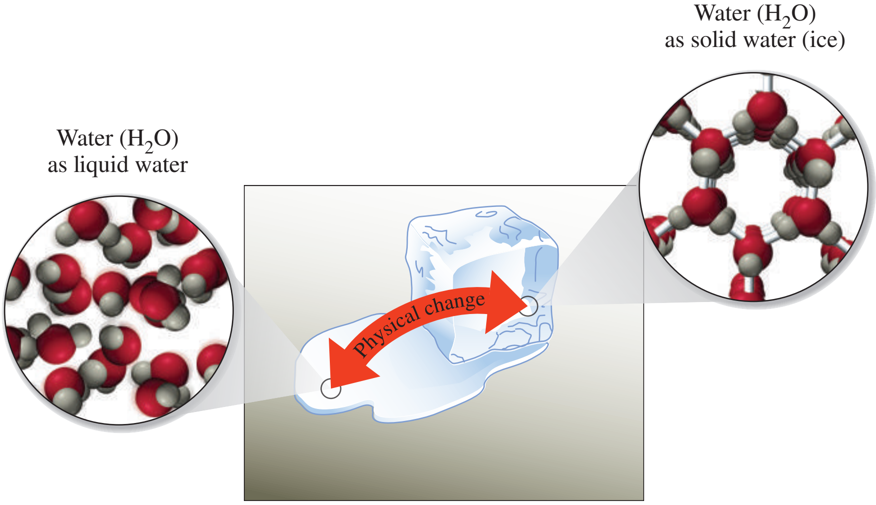


FIGURE 1.3 Physical change of water from solid to liquid. As water changes from solid to liquid, the arrangement of the molecules becomes less orderly and their motion increases. The red balls represent oxygen atoms and the gray balls represent hydrogen atoms in the water molecules.

Chemical Reaction

The volcano in our example was most likely filled with baking soda before the vinegar was poured into it. When the vinegar and baking soda combine, water and carbon dioxide bubbles form. Chemically, neither vinegar nor baking soda is the same as water or carbon dioxide. So when these two substances combined, new substances were formed. Unlike a physical change, a chemical change results in a change in the chemical identity of the substance (or substances) involved. When a substance (or substances) undergoes such a change, we more commonly refer to it as a **chemical reaction**. **TABLE 1.2** gives some examples of physical changes and chemical reactions.

TABLE 1.2 Comparing Physical Changes and Chemical Reactions

Physical Change
Sawing a log in half
An ice cube melting
Molten iron solidifying to form a nail
Mixing oil, water, and eggs to make mayonnaise
Mixing cooked spaghetti and sauce to make a meal
Chemical Reaction
Burning a log in your fireplace
Hydrogen and oxygen combining to form water
A nail rusting
Mayonnaise becoming spoiled
A spaghetti supper being metabolized to provide energy for a morning run

Sample Problem 1.6 Physical Change or Chemical Reaction

Determine whether each of the following is a physical change or a chemical reaction:

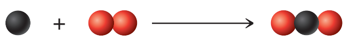
a. a plant using carbon dioxide and water to make sugar
b. water vapor condensing on a cool evening to form dew on the grass

Solution

a. This is a chemical reaction. Carbon dioxide and water are chemically different substances from sugar.
b. This is a physical change. Water vapor is a gas that condenses to form dew, which is simply liquid water.

Chemical Equations

Have you ever wondered how the big pile of charcoal that you start with in an outdoor grill is reduced to a small pile of gray dust when the fire finally dies out? When charcoal burns with oxygen, or combusts, a chemical reaction is occurring. To show what happens to the element carbon (the main component of charcoal) as it burns, we can write a **chemical equation** like the one shown here.



$$\underbrace{C(s) + O_2(g)}_{\text{Reactants}} \underbrace{\xrightarrow{\text{heat}}}_{\text{Reaction arrow}} \underbrace{CO_2(g)}_{\text{Product}}$$

The number of atoms of each element in the reactants must equal the number in the products.

The carbon in charcoal burns with oxygen, forming carbon dioxide.

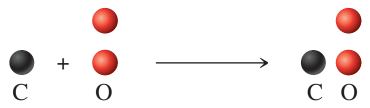
A chemical equation is the chemist's way of writing a sentence about what happens in a chemical reaction. It is meant to be read from left to right. This chemical equation explains that solid carbon (from charcoal) and oxygen gas (from air) react to form carbon dioxide gas when heated. In this example carbon and oxygen are the **reactants**, and carbon dioxide is the **product**. Reactants are found on the left side of a chemical equation and products are found on the right side. The reaction arrow means "react to form."

Special reaction conditions are often written as words or symbols above the reaction arrow. Typical examples include heat, light, or the name of a catalyst or other substance required for the reaction to occur.

The labels in parentheses after each substance (called state labels) indicate its physical state—(s)olid, (l)iquid, or (g)as. A fourth label, (aq), meaning aqueous or dissolved in water, will be discussed later. So, when solid charcoal is burned, it combines with oxygen in the air and forms carbon dioxide, a gas.

Balancing Chemical Equations

The reaction of carbon with oxygen clearly illustrates the *parts* of a chemical reaction. Notice something else. There is one atom of carbon and there are two atoms of oxygen on the reactant side and the same number of each on the product side even though a chemical change took place. The fact that the number of atoms of each element in the reactants equals the number of atoms of each element in the products illustrates the **law of conservation of mass**. Matter can neither be created nor destroyed. Matter only changes form, so the amounts of matter on the reactant side and the product side must be equal.



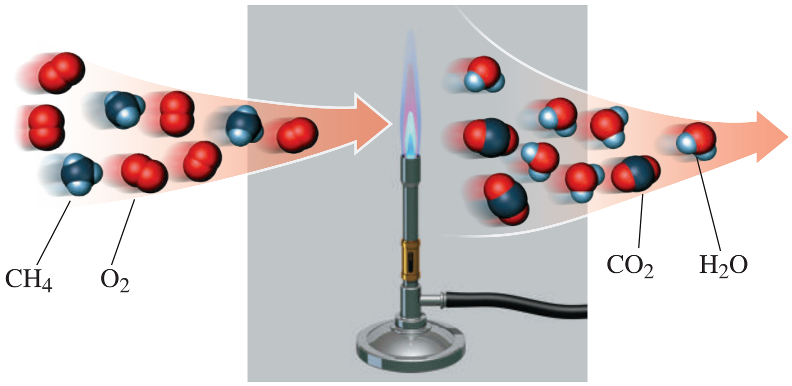
The number of atoms of each element in the reactants must equal the number in the products.

For any chemical equation that we write, the number of atoms of each element must be the same on both sides of the equation, or balanced. We can balance chemical equations when necessary by adding a number, called a **coefficient**, in front of the chemical formula for a substance in the chemical equation. This is easiest to understand by working some examples.

Solving a Problem

Balancing a Chemical Equation

Balance the chemical equation for the combustion of natural gas, also called methane. This occurs when oxygen (O_2) reacts with methane (CH_4), producing carbon dioxide and water. The unbalanced equation is written as follows:



$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Once the chemical equation is written, follow steps 1 to 3.

STEP 1 Examine the original equation. Is it balanced? If not, proceed to step 2.

Element	Number in Reactants	Number in Products
H	4	2
C	1	1
O	2	3

There are different numbers of atoms in the reactants and products, so no, it is not balanced.

STEP 2 Balance the equation one element at a time by adding coefficients. Balance elements that appear only once on a side first. For our example, we save oxygen for last because it appears in both CO_2 and H_2O in the products. The number of carbon is balanced, but the number of H is not. Writing the coefficient 2 in front of H_2O will give us 4 Hs ($2 \times H_2O$) in the products. This addition of 2 in front of H_2O likewise changes the total number of oxygen atoms on the product side to 4. Two oxygens in $2H_2O$ and 2 in CO_2. Then, writing a 2 in front of the O_2 in the reactants gives 4 oxygen atoms ($2 \times O_2$) on the reactant side.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

STEP 3 Check to see if the equation is balanced. The coefficients should represent the smallest possible set of numbers (not all divisible by another number).

Element	Number in Reactants	Number in Products
H	4	4
C	1	1
O	4	4

Now the number of atoms of each element in the reactants equals that in the products, so the equation is balanced.

Sample Problem 1.7 Balancing Chemical Equations

Add coefficients to balance the chemical equations shown.

a. $H_2(g) + O_2(g) \rightarrow H_2O(g)$

b. $C(s) + SO_2(g) \rightarrow CS_2(l) + CO(g)$

Solution

a.

STEP 1 Examine.

Element	Number in Reactants	Number in Products
H	2	2
O	2	1

STEP 2 Balance. The element not balanced is oxygen. If we write a 2 in front of the H_2O in the product, we balance the oxygen, but now the hydrogen is doubled to 4. The hydrogen can then be balanced by writing a 2 in front of the hydrogen in the reactants.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

STEP 3 Check.

Element	Number in Reactants	Number in Products
H	4	4
O	2	2

The equation written in step 2 is balanced.

b.

STEP 1 Examine.

Element	Number in Reactants	Number in Products
C	1	2
S	1	2
O	2	1

—*Continued next page*

Continued—

STEP 2 Balance. None of the elements are balanced. If we start by looking at sulfur, which appears only once on both sides, and we write a 2 in front of the SO_2 in the reactant to balance the sulfur, we get a total of 4 oxygens on the reactant side. This can be balanced by writing a 4 in front of the CO in the products. Now we have 5 carbons in the product, which can be balanced by writing a 5 in front of the C on the reactant side.

$$5C(s) + 2SO_2(g) \rightarrow CS_2(l) + 4CO(g)$$

STEP 3 Check.

Element	Number in Reactants	Number in Products
C	5	5
S	2	2
O	4	4

The equation written in step 2 is balanced.

Practice Problems

1.11 Determine whether each of the following is a physical change or a chemical reaction:

a. A copper penny turns green.

b. Sugar is melted to make candy.

c. An antacid tablet is dropped into water, and bubbles of carbon dioxide form.

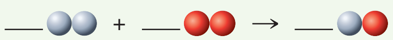
1.12 Determine whether each of the following is a physical change or a chemical reaction:

a. Cream, sugar, and flavorings are mixed and frozen to make ice cream.

b. Milk spoils and becomes sour.

c. A puddle of spilled nail polish remover "disappears."

1.13 Balance each of the following reactions by adding coefficients in the blanks. For the purpose of this exercise, place a 1 where appropriate. Usually, if no coefficient appears, "1" is implied.

a. 

b. 

1.14 Balance each of the following reactions by adding coefficients in the blanks. For the purpose of this exercise, place a 1 where appropriate. Usually, if no coefficient appears, "1" is implied.

a. 

b. 

1.15 Add coefficients to balance the chemical equations shown.

a. $Al(s) + Cl_2(g) \rightarrow AlCl_3(s)$

b. $HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$

c. $SiO_2(s) + C(s) \rightarrow SiC(s) + CO(g)$

1.16 Add coefficients to balance the chemical equations shown.

a. $NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + NO(g)$

b. $NH_3(g) + O_2(g) \rightarrow N_2(g) + H_2O(l)$

c. $C_7H_{16}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

1.4 Math Counts

The proper tools can make a job easier. Mathematics gives us some tools that can make understanding chemistry easier. This section covers some basic concepts from mathematics that are applicable to the field of chemistry. As you move through the textbook, you will encounter these math concepts again. You can always come back to this section to review them.

1.4 Inquiry Question

What basic math concepts are important to chemistry?

Units, Prefixes, and Conversion Factors

In science, health care, and business, we often make a measurement or report a value. For measurements to be consistently and easily compared, a defined set of standards or a measurement system is required. The ***Système International d'Unités* (SI)** or International System of Units used throughout the world by scientists and health care professionals alike is the modern day version of the **metric system.** Because of this, SI units and therefore metric units will be used throughout this text, so it is important that you are familiar with them.

The SI system is based on a series of standard units for each type of measurable quantity like mass and volume. The standard SI unit for mass is the **kilogram (kg)** while the standard metric unit for mass is the **gram (g),** which is a thousand times smaller than the kilogram. The standard SI unit for volume is the **liter (L),** and the standard SI unit for length is the **meter (m).** Although these measures may seem unfamiliar now, a kilogram is about 2.2 pounds, a gram is about the mass of a paper clip, a liter is a little more than a quart, and a meter is slightly more than 3 feet.

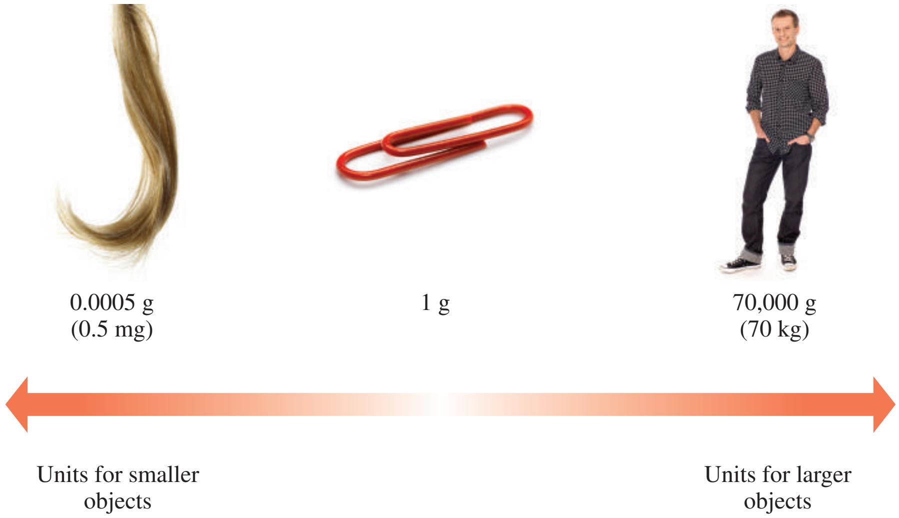
Imagine how many paper clips it would take to equal your body mass. This would be a huge number and would be awkward to work with. Or think of trying to get the mass of one human hair by comparing it to a paper clip. It would take a very large number of hairs to equal the mass of a paper clip.

To deal efficiently with quantities that are much larger or much smaller than one another, the SI system employs a set of prefixes that can be applied to the base unit, changing the meaning of the unit by a power of ten. The power-of-ten concept is similar to our currency in that 10 pennies equal 1 dime and 10 dimes equal 1 dollar. **TABLE 1.3** lists the common SI and metric prefixes, as well as the abbreviation and value for each.

For example, because the average person has a mass much greater than a paper clip, we apply the prefix kilo to the base unit of mass, the gram. "Kilo" means "1000 times greater than," so a *kilo*gram is 1000 g. Thus, an average person's mass would be expressed as 70 kg, which is a much more manageable number than 70,000 g. Note that while the kilogram is the standard SI unit of mass, it is 1000 times larger than the base unit (the gram), which has no prefix.

TABLE 1.3 SI and Metric Prefixes

Prefix	Abbreviation	Relationship to Base Unit
giga	G	× 1,000,000,000
mega	M	× 1,000,000
kilo	k	× 1000
base unit (has no prefix)		× 1 (gram, liter, meter)
deci	d	÷ 10
centi	c	÷ 100
milli	m	÷ 1000
micro	μ or mc	÷ 1,000,000
nano	n	÷ 1,000,000,000



Prefixes in the SI and metric system help us to scale objects.

We can use Table 1.3 to obtain equivalency relationships between different units. The numbers in the right-hand column provide the relationship between any unit and the base unit. For example, when moving from deciliters to liters, the last column says we should divide the base unit by 10; therefore, the equivalency is 1 dL = 0.1 L. Quantities that can be related to each other by an equal sign are called **equivalent units**. By multiplying both sides by 10, we can avoid working with decimals and say that 10 dL = 1 L.

Mastering Video
Practicing the Concepts
Conversion Factors Are Relationships

Such equivalencies can be used as **conversion factors** to convert one unit to another using one or more of these factors. For example, when the equivalency stated previously, 10 dL = 1 L, is written as a conversion factor, it can take the following forms:

$$\frac{10\ \text{dL}}{1\ \text{L}} \quad \text{or} \quad \frac{1\ \text{L}}{10\ \text{dL}}$$

These conversion factors are read as, "There are 10 deciliters in 1 liter" and "1 liter contains 10 deciliters," respectively. Conversion factors allow you to convert a quantity from one unit to the equivalent quantity in a different unit. This use of converting units to an equivalent unit is also called **dimensional analysis**. Let's try an example.

Solving a Problem

Converting Units

How many grams of vitamin C are in a tablet containing 100 mg?

STEP 1 Determine the unit on your final answer. In this case, the unit will be grams.

STEP 2 Establish the given information. In this case, you are given the value of 100 mg in the problem. You also have an equivalency relationship (see Table 1.3) showing that 1 mg = 0.001 g or 1000 mg = 1 g, which gives the following conversion factors:

$$\frac{1000\ \text{mg}}{1\ \text{g}} \quad \text{or} \quad \frac{1\ \text{g}}{1000\ \text{mg}}$$

STEP 3 Decide how to set up the problem. Which conversion factor should we use? The given units must cancel out (appearing in both a numerator and denominator), leaving the desired unit in the answer. This problem gives a number of milligrams and asks for the number of grams. We must *convert* from milligrams to grams. To do this, set up an equation using a conversion factor with the desired unit in the numerator. The form of this equation is as follows:

$$\text{Given unit} \times \underbrace{\frac{\text{desired unit}}{\text{given unit}}}_{\text{This is the conversion factor.}} \leftarrow \text{Units of answer in numerator}$$

For our problem, the equation is as follows:

$$100\ \text{mg} \times \frac{1\ \text{g}}{1000\ \text{mg}} =$$

STEP 4 Solve the problem. When the calculation is carried out, the given unit in the denominator of the conversion factor will cancel the given unit from the problem.

$$100\ \cancel{\text{mg}} \times \frac{1\ \text{g}}{1000\ \cancel{\text{mg}}} = \frac{100}{1000} = 0.1\ \text{g}$$

Sample Problem 1.8 Unit Conversions

Using the appropriate conversion factor, convert each of the following quantities to the indicated unit:

a. 10,000 cg = ______ g   b. 5 g/mL = ______ mg/dL

Solution

a.

STEP 1 Determine the unit on your final answer. In this case, the unit will be grams.

STEP 2 Establish the given information. In this case, you have a value of 10,000 cg given. You also have an equivalency relationship (Table 1.3) showing that 1 g = 100 cg.

$$\frac{100\ \text{cg}}{1\ \text{g}} \text{ or } \frac{1\ \text{g}}{100\ \text{cg}}$$

STEP 3 & 4 Decide how to set up the problem and then solve the problem. The answer must be in grams, so grams must be in the numerator of our conversion factor. The equation is set up and solved like this:

$$10{,}000\ \cancel{\text{cg}} \times \frac{1\ \text{g}}{100\ \cancel{\text{cg}}} = 100\ \text{g}$$

b.

STEP 1 Determine the unit on your final answer. In this case, the unit will be mg/dL.

STEP 2 Establish the given information. In this case, you have a value of 5 g/mL given. You have equivalency relationships (see Table 1.3) showing that 1 L = 1000 mL, 1 L = 10 dL, and 1 g = 1000 mg.

STEP 3 Decide how to set up the problem. In this problem you will need to convert from g/mL to mg/dL. This will require two conversion factors. You will want to set up the equivalencies determined in step 2 as conversion factors so that the mg unit is in the numerator and the dL unit is in the denominator, or

$$\frac{1000\ \text{mg}}{1\ \text{g}}, \frac{1\ \text{L}}{10\ \text{dL}}, \frac{1000\ \text{mL}}{1\ \text{L}}$$

—*Continued next page*

Mastering Video
Practicing the Concepts
Using Multiple Conversion Factors

Continued—

STEP 4 Solve the problem. All the units that are not part of the answer should be organized so that the unwanted units cancel each other out, leaving only the desired unit(s) for the answer. The equation can be set up like this:

$$\frac{5\ \cancel{g}}{1\ \cancel{mL}} \times \frac{1000\ mg}{1\ \cancel{g}} \times \frac{1\ \cancel{L}}{10\ dL} \times \frac{1000\ \cancel{mL}}{1\ \cancel{L}} = \frac{5{,}000{,}000\ mg}{10\ dL} = 500{,}000\ \frac{mg}{dL}$$

Significant Figures

We can report the measure of this syringe as 4.35 mL, estimating the last digit, 5.

Before proceeding further with any calculations, it is important that we understand that measuring matter relies on the precision of the instruments that we use to measure it. For example, a bathroom scale only measures to the closest half-pound, so we wouldn't consider recording a person's weight as 150.2356 lb! It is important that we report calculated answers reasonably, or as scientists refer, to the correct number of significant figures.

Suppose you are asked to read the number of mLs in the syringe shown at the right. Measurements on a syringe are made by reading across the top ring of the plunger (closest to the liquid). What would you report? To how many decimals?

Because the syringe shows marks (calibrations) every two-tenths (0.2) mL, it looks like the there is definitely more than 4.2 mL but perhaps not quite 4.4 mL. You might report your measure as 4.35 mL since the plunger looks like it is sitting closer to 4.4 than to 4.2. By reporting a value of 4.35, we are reporting that we are sure that there is between 4.3 and 4.4 mL in the syringe.

We report this measured value to three numbers, or three significant figures. **Significant figures** are the digits in a value known with certainty *plus* an estimated digit. When reading a measurement from a nondigital device like a syringe, there is always a level of uncertainty in the measurement. Significant figures allow us to report the certainty of a measurement. Digital devices like a digital thermometer automatically show us the number of significant figures in the digital readout. In any measurement, all nonzero numbers are considered significant.

This digital thermometer shows three significant figures.

Zeros tend to be the "troublemakers" because their significance depends on their position in a number. For example, when a person's weight is reported as 70 kg, the number has *only one* significant figure, so all we know about the person's weight is that it is between 65 grams and 75 kg. However, if a person's weight is reported as 70.0 kg, the last zero *is* significant and this measurement tell us the person weighs between 69.9 and 70.1 kg. When reading measurements from nondigital devices, one should always estimate the last reported number just as we did with the syringe example.

If a zero in a number without a decimal point is significant, its significance can still be shown. To indicate that the zero in a measurement of 10 inches is significant, simply put in a decimal point after the zero; 10. inches means that both digits are significant. Otherwise, to indicate significant zeros in large numbers, draw a line above the last significant zero; $10{,}0\overline{0}0$ has four significant figures.

It should be noted that numbers used in defined conversion factors and counted items have more certainty. For example, if you are counting drops per minute (gtt/min) and you counted 20 of these, you counted exactly 20. Your certainty is infinite (the drop rate was exactly 20) and there is not a limited number of significant figures in such a case. Significant figures tend to not be considered for these numbers in calculations. Only measured numbers determine the number of significant figures given in a final answer.

The rules for counting significant figures in measurements are summarized in **TABLE 1.4**.

TABLE 1.4 Counting Significant Figures in Measurements

Rule	Measurement	Number of Significant Figures
1. A digit is significant if it is		
a. not a zero	41 g 15.3 m	2 3
b. a zero between nonzero digits	101 L 6.071 kg	3 4
c. a zero at the end of a number with a decimal point	20. g 9.800 °C	2 4
2. A zero is not significant if it is		
a. at the beginning of a number with a decimal point	0.03 L 0.00024 g	1 2
b. in a large number without a decimal point	12,000 km 3,450,000 m	2 3

Sample Problem 1.9 Determining Significant Figures

Give the number of significant figures in the following measurements:

a. A digital scale reads 56.7 kg.
b. A pharmacist dispenses 30 pills.
c. A newspaper reports that about 500 people attended the event.
d. A baby's temperature is measured to be 100.° F.

Solution

a. Three significant figures; see rule 1a in Table 1.4.
b. Infinite number of significant figures; the pharmacist counted the number of pills, so the measurement is exact.
c. One significant figure; see rule 2b in Table 1.4.
d. Three significant figures; the presence of the decimal point indicates that the zeros are not just placeholders but are a part of the measurement. See rule 1c in Table 1.4.

Calculating Numbers and Rounding

Next, let's look at how we consider significant figures in calculations. Suppose you ask the person next to you their age, and they tell you that they are 27. How many days old is the person? To calculate the number of days old we could multiply 27 by 365.24 days (the precise number of days in a year) on our calculator, giving us an answer of 9,861.48 days.

Consider the answer from the calculator. We know the person's age to the nearest year, but by multiplying that value by the precise number of days in a year, we now know the age to within a tenth of a day! Somehow that doesn't seem logical. Manipulating measurements (which have some built-in uncertainty) with arithmetic cannot *increase* the certainty of the measurement. A calculator does not know the certainty in your measurement and will often make an attempt at more certainty than makes sense, so the person doing the calculation must be prepared to round the answer correctly.

How many of the resulting numbers from our previous calculation of the person's age are certain and, therefore, meaningful? Keeping all of the numbers results in an answer with six significant figures, yet we started with a number (27) that has only two.

In general, your answer can be no more certain than the least certain number that you started with. The rules for determining the number of significant figures in a calculated answer depend on the arithmetic operation being used.

Rules for Significant Figures in Calculations

- **Addition and Subtraction.** Answers should match the least number of decimal places in the measured numbers.

1.002	⟵	three decimal places
2.5	⟵	one decimal place
+10.16	⟵	two decimal places
13.662		report as 13.7

  Because one of the numbers contains only one decimal place, the final answer must be rounded to one decimal place.

- **Multiplication and Division.** Answers should match the least number of significant figures in the measured numbers. For our age-in-days calculation, the answer can have only two significant figures because we were given the person's age as 27.

$$27\ \cancel{\text{years}}\ \text{old} \times \frac{365.24\ \text{days}}{1\ \cancel{\text{year}}} = 9{,}861.48\ \text{days old; report as } 9{,}900\ \text{days old}$$

Rules for Rounding Numbers

- If the leftmost digit to be dropped (the one following your last retained significant figure) is 4 or less, simply remove it and the remaining digits.
- If the leftmost digit to be dropped is 5 or greater, increase the last retained digit by 1 and remove all other digits.
- If rounding a large number with no decimal point, substitute zeros for numbers that are not significant.
- When conducting multiple-step calculations, do not round answers until the end of the calculation. *Rounding at each step introduces rounding errors and produces inaccurate answers.*

Sample Problem 1.10 Rounding Calculated Numbers

Complete each of the following calculations and round the answer to the correct number of significant figures:

a. $75 - 21.3 =$
b. $8.24 \times 3.0 =$
c. $0.005 + 1.23 =$
d. $12.4 \div 0.06 =$

Solution

a. Answer is 54. The calculator result of 53.7 is rounded so that the answer has no decimal places because 75 has no numbers after the decimal point.
b. Answer is 25. The calculator result of 24.72 is rounded to two significant figures because 3.0 has only two significant figures.
c. Answer is 1.24. The calculator result of 1.235 is rounded to the least number of decimal places, which is two.
d. Answer is 200. The calculator result of 206.66666 is rounded to one significant figure as 200, which has the same number of significant figures as 0.06.

Scientific Notation

In chemistry, we sometimes use extremely large and extremely small numbers in the same calculation. Suppose that we wanted to figure out how many sheets of paper we could stack in a cabinet that is 3 m high. Each sheet of paper has a thickness of one-tenth of a millimeter or 0.0001 m. The number of sheets of paper would be

$$\frac{1 \text{ sheet of paper}}{0.0001 \cancel{m}} \times \frac{3 \cancel{m}}{1 \text{ cabinet}} = \frac{30{,}000 \text{ sheets of paper}}{\text{cabinet}}$$

This is a lot of pieces of paper! We can represent this with this large number. Yet it is more easily written using scientific notation. To explore scientific notation, consider that 30,000 pieces of paper is $3 \times 10 \times 10 \times 10 \times 10$ (try it on your calculator). In scientific notation, we can express the four times we multiplied by 10 as an exponent, showing it as 10^4. The scientific notation for 30,000 is written as 3×10^4.

A number smaller than 1, like the 0.0001 m for each sheet of paper, can be determined by dividing 1 by $10 \times 10 \times 10 \times 10$ or $1/10^4$. Dividing by a number is equivalent to multiplying by its reciprocal. So, in exponential form, $1/10^4$ can be written as 1×10^{-4}. The general form for scientific notation is

$$C \times 10^n$$

where C is the **coefficient** and is a number that is at least 1 but less than 10, and n is the exponent telling us the number of tens places that apply.

A positive exponent tells us that the actual number is greater than 1, and a negative exponent tells us the number is between 0 and 1. In scientific notation, only significant figures are shown in the coefficient (any place-holding zeros are not part of the notation). For example, the scientific notation of the number 0.00650 is 6.50×10^{-3} because the zeros between the decimal point and the 6 are placeholders and the zero after the 5 is significant.

We could rewrite an equivalent calculation in scientific notation as

$$\frac{1 \text{ sheet of paper}}{1 \times 10^{-4} \cancel{m}} \times \frac{3 \cancel{m}}{1 \text{ cabinet}} = \frac{3 \times 10^4 \text{ sheets of paper}}{\text{cabinet}}$$

TABLE 1.5 shows how numbers and scientific notation are related for a wide range of numbers.

Scientific notation can be displayed on your calculator as long as you know which buttons to press. Most scientific calculators have a button labeled [EE], [EXP], or [SCI], which displays the coefficient (C) and the power-of-ten exponent (n). It is worth exploring your calculator to be sure that you can input scientific notation correctly. To input 3×10^4, try pressing 3 [EE] or [EXP] 4. To input 1×10^{-4}, try pressing 1 [EE] or [EXP] [+/−] 4. With a scientific calculator (which requires a reverse input), press 1 [EE] or [EXP] 4 [+/−].

1 piece of paper = 0.0001 m

TABLE 1.5 Relating Numbers to Scientific Notation

Number	Meaning	Scientific Notation
1,000,000	$10 \times 10 \times 10 \times 10 \times 10 \times 10$	1×10^6
100,000	$10 \times 10 \times 10 \times 10 \times 10$	1×10^5
10,000	$10 \times 10 \times 10 \times 10$	1×10^4
1000	$10 \times 10 \times 10$	1×10^3
100	10×10	1×10^2
10	10	1×10^1
1	1	1×10^0
0.1	1/10	1×10^{-1}
0.01	$1/(10 \times 10)$	1×10^{-2}
0.001	$1/(10 \times 10 \times 10)$	1×10^{-3}
0.0001	$1/(10 \times 10 \times 10 \times 10)$	1×10^{-4}
0.00001	$1/(10 \times 10 \times 10 \times 10 \times 10)$	1×10^{-5}
0.000001	$1/(10 \times 10 \times 10 \times 10 \times 10 \times 10)$	1×10^{-6}

Sample Problem 1.11 Expressing Numbers in Scientific Notation

Express the following numbers in scientific notation:

a. 820,000

b. 0.00096

Solution

a. To determine the correct scientific notation for a number greater than 1, count the number of places to the right of the first number, 8. In this case, there are 5 places to the right of 8. This becomes the exponent on the number 10 in scientific notation. Any other nonzero numbers go after a decimal point. The answer is 8.2×10^5.

b. To determine the correct scientific notation for a number less than 1, count the number of places to the left of the first number, 9, and add 1. In this case, there are 3 places to the left of the 9, and adding 1 would give a total of 4. The negative of this number becomes the exponent on the number 10 in scientific notation. Any other nonzero numbers go after the decimal point. The answer is 9.6×10^{-4}.

Sample Problem 1.12 Changing Numbers from Scientific Notation to Decimal Form

Express the following numbers in decimal form:

a. 6.5×10^4

b. 8.34×10^{-6}

Solution

a. The number in the exponent tells you how many places are beyond the decimal place in the scientific notation when converting from scientific notation to decimal form. Notice that the number of significant figures does not change when the format of the number changes. In this problem, the number 5 will occupy one of the places, so 3 more zeros after the 5 are necessary. The answer is 65,000.

b. Negative exponents indicate that this number is smaller than 1 but greater than zero. The number in the exponent tells you how many places should appear to the left of the decimal place when converting from scientific to decimal form. The number 8 will occupy one of these places, so 5 more zeros should appear in the final answer to the right of the decimal place. All digits to the right of the decimal place in the scientific notation should appear in the final answer. The answer is 0.00000834.

Percent

If your chemistry class has 40 students enrolled and 20 students got an A on the first test, you might comment that one-half, or 50%, of the class got an A. Percent, represented by the symbol %, means the part out of 100 total, or hundredths. So, 50% means 50 out of a total of 100 or 50 hundredths (0.50), which reduces to $^1/_2$, the same fraction as our original example using grades. Percent allows us to directly compare two sets of numbers that have different total sizes.

$$\text{Percent (\%)} = \frac{\text{part}}{\text{whole}} \times 100$$

A fraction can be converted to a percent by dividing the numerator by the denominator, multiplying by 100, and adding a percent sign. A decimal number can be converted to a percent by multiplying by 100 and adding a percent sign.

Sample Problem 1.13 Expressing Fractions and Decimals as a Percent

a. Express 1/5 as a percent.

b. Express 0.35 as a percent.

Solution

a. For a fraction, divide the numbers of the fraction, multiply by 100, and add a percent sign.

$$1/5 = 0.20 \times 100 = 20\%$$

b. For a decimal, multiply by 100 and add the percent sign.

$$0.35 \times 100 = 35\%$$

Sample Problem 1.14 Expressing Percents as Fractions and Decimals

a. Express 25% as a decimal.

b. Express 25% as a fraction.

Solution

a. To express 25% as a decimal, divide the value by 100.

$$\frac{25}{100} = 0.25$$

b. Percent means parts out of a total of 100 parts, so the fraction is

$$\frac{25}{100}, \text{ which reduces to } \frac{1}{4}$$

Sample Problem 1.15 Determining the Percent of a Number

a. What is 40.0% of 275?

b. What is 80.0% of 65?

Solution

To determine the percent of a number, change the percent to a decimal and multiply by the number.

a. $40.0\% = 0.400$
$0.400 \times 275 = 110.$

b. $80.0\% = 0.800$
$0.800 \times 65 = 52$

Sample Problem 1.16 Determining the Percent That One Number Is of Another Number

a. What percent of 1500 is 30?

b. Thirty-five is what percent of 140?

Solution

To determine what percent one number is of a second number, a fraction is created to represent parts out of the whole. This gives a decimal answer that can then be converted to a percent by multiplying by 100.

a. $\frac{30}{1500} = 0.02$
$0.02 \times 100 = 2\%$

b. $\frac{35}{140} = 0.25$
$0.25 \times 100 = 25\%$

Solving a Problem

Calculating Percent

Suppose you go out to a nice lunch with three other friends and the final bill is \$125.86. If you decide to split the bill evenly and leave an 18% tip, how much would each person contribute?

STEP 1 Determine the units on your final answer. In this case, the final answer is in dollars.

STEP 2 Establish the given information. We know the total bill, \$125.86. We also know the percent we are going to tip, 18%.

STEP 3 Decide how to set up the problem. To determine 18% of the total bill, you express the percent as a decimal and multiply by the total bill. The total amount that will be spent is the sum of the bill plus the 18% tip.

$$18\% = 0.18$$
$$0.18 \times \$125.86 \text{ total bill} = \$22.65$$
$$\text{Total bill} = \$125.86 + \$22.65 = \$148.51$$

STEP 4 Solve the problem. Since the total bill is \$148.51, you divide this number by the four of you.

$$\$148.51/4 = \$37.13$$

Note that you must round to the nearest penny since that is your smallest unit of measure.

Practice Problems

1.17 Give the number of significant figures in each of the following measurements:
a. 25 min b. 44.30 °C
c. 0.037010 m d. 800. L

1.18 Give the number of significant figures in each of the following measurements:
a. 0.000068 g b. 100 °F
c. 9,237,200 years d. 25.00 m

1.19 Complete each of the following calculations and give your answer to the correct number of significant figures:
a. $100 \times 23 =$ b. $97.5 - 43.02 =$
c. $8064/0.0360 =$ d. $54.00 + 78 =$

1.20 Complete each of the following calculations and give your answer to the correct number of significant figures:
a. $340/96 =$ b. $0.305 + 43.0 =$
c. $0.0065 \times 11 =$ d. $19.029 - 0.00801 =$

1.21 A typical aspirin tablet contains 325 mg of the active ingredient acetylsalicylic acid. Convert to grams.

1.22 On average, an adult's lung volume is 5 L. Convert to mL.

1.23 Write scientific notation for the following numbers:
a. 203,000,000 b. 12.4 c. 0.0000000278

1.24 Write scientific notation for the following numbers:
a. 1,400,000 b. 0.079 c. 0.00000354

1.25 Express the following numbers in a decimal form:
a. 1.56×10^{-5} b. 2.8×10^{5} c. 9.0×10^{-2}

1.26 Express the following numbers in a decimal form:
a. 7.4×10^{-2} b. 3.75×10^{3} c. 1.19×10^{-8}

1.27 Write scientific notation for the following numbers:
a. a platelet dose of 500,000,000,000
b. a red blood cell count of 5,500,000

1.28 Express the following numbers in decimal form:
a. an asthma medication of 1.00×10^{-8} g
b. vitamin D dose of 5.0×10^{-6} g

1.29 Express the following numbers as a percent. Report two significant figures in your answer.
a. 1/4 b. 3/8 c. 66/100

1.30 Express the following numbers as a percent. Report two significant figures in your answer.
a. 0.58 b. 0.36 c. 0.125

1.31 Express the following numbers in decimal form:
a. 4.5% b. 13.0%
c. 66% d. 78%

1.32 Express the following numbers as a fraction:
a. 20% b. 75%
c. 40% d. 12%

1.33 Determine the number from the percent given:
a. What is 25% of 80? b. What is 0.9% of 1000?
c. What is 5.0% of 750? d. What is 66.6% of 200.?

1.34 Determine the percent from the numbers given here.
a. Fifty is what percent of 125?
b. Six is what percent of 600?
c. What percent of 300 is 15?
d. What percent of 400 is 30?

1.5 Matter: The "Stuff" of Chemistry

Chemists study matter that we cannot see, even with a high-powered microscope. Even though we cannot see this matter, we can see and measure its properties in bulk. Here we discuss some of those properties and characteristics of matter.

1.5 Inquiry Question

What properties of matter can be measured?

Mass

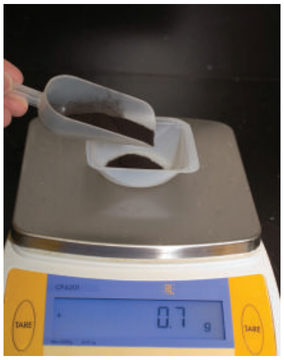
Consider that anything that takes up space can also be placed on a scale and weighed—that is, it has mass. So, matter can be more completely defined as anything that takes up space and has mass. **Mass** is a measure of the amount of material in an object. As we saw in Section 1.4, a common unit used to measure the mass of a substance is the gram (g).

A raisin or a paper clip has a mass of about 1 g. The term *weight* may be more familiar to you than the term *mass*. These terms are related, but they do not mean the same thing. The **weight** of an object is determined by the pull of gravity on the object, and that force changes depending on location. For instance, astronauts weigh less on the moon than on Earth. This is because the gravitational pull of the moon is less than that of Earth. The astronauts' mass did not change, but their weight did. The SI unit for weight, the force on an object of a given mass due to gravity, is the newton (N).

When you step on the bathroom scale, the scale does not read your weight in newtons. Are you determining your mass or your weight? How are mass and weight related to each other?

A scale or balance is used to measure mass. When balances are manufactured, they are set at the factory so that a 1-kg (1000-g) object has the same mass on all of the balances. This is known as "calibrating" the balance and accounts for the gravitational pull of Earth. After calibrating, the mass and the weight of the 1-kg object are the same. If the object and balance were transported to the moon, the balance would have to be re-calibrated to account for the lesser gravitational pull of the moon if it were to read the mass on the moon as 1 kg.

As long as an object is weighed in roughly the same location on Earth's surface, its mass and weight will have the same measured value. To avoid confusion, this book will, from this point forward, refer to weighable quantities using the term *mass*.



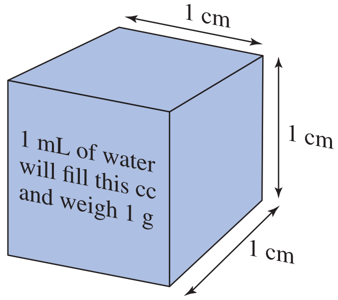
Measuring mass in the lab

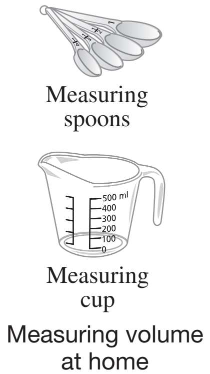

Measuring mass at home

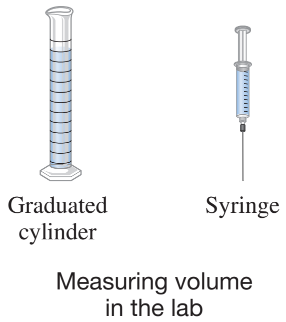
Volume

Another way of saying that matter "takes up space" is to say that it has **volume**. Large soft drink bottles are often called "2-liter" bottles. The volume of the soft drink in the bottle is 2 L (1 L = 1.057 quart). Similarly, as you blow up a balloon, it gets bigger. The volume of air in the balloon increases. So, volume is a three-dimensional measure of the amount of space occupied by matter.

FIGURE 1.4 The relationship between volume and length. One cubic centimeter (abbreviated cc or cm^3) equals 1 mL of volume. For water, this volume weighs 1 g.







When cooking or baking at home, you might use a measuring cup or a measuring spoon to measure volume. In the lab, volumes are routinely measured with a graduated cylinder or a pipet. The unit of volume typically used in the lab is the milliliter (mL). In a clinical setting, volumes are often measured with calibrated syringes. The typical unit in the clinical setting is the cubic centimeter (cc, or cm^3). One milliliter equals one cubic centimeter (see **FIGURE 1.4**). A teaspoon of cough syrup contains about 5 mL or 5 cc.

Density

Regardless of its mass or volume, most types of wood will float on water while a piece of metal will sink. This observation implies that the mass of wood that fits into a certain space is less than that for metal. This property of matter is called density. **Density** (d) is a comparison (also called a ratio) of a substance's mass (m) to its volume (V). This can be displayed mathematically as

$$d = \frac{m}{V}$$

One gram of water has a volume of one milliliter, so the density of water is 1.00 g/mL. A piece of wood will float on water because it is less dense than water, while a piece of metal will sink because it is more dense than water (see **FIGURE 1.5**).

We can think of density as a measure of the packing of matter. As another example, the density of our bones can change even though the size of our bones doesn't change. As we age, minerals are lost from our bones faster than they are replenished. The bones' strength and integrity can become compromised. In this case, the mass of the bone lessens while the size or volume stays the same, making bones less dense. This can lead to the condition known as *osteoporosis*.

FIGURE 1.5 Visualizing density. Objects that are less dense than water will float while objects more dense than water will sink.

Sample Problem 1.17 Calculating Density

A brick stamped "14-karat gold" has a mass of 1100 g and a volume of 85 mL. Calculate its density. Will the brick float or sink when placed in water?

Solution

You are given the mass and volume and asked to calculate the density in g/mL. Insert the given values into the density equation.

$$d = \frac{m}{V}$$

$$d = \frac{1100 \text{ g}}{85 \text{ mL}} = 13 \text{ g/mL}$$

The final answer is reported to two significant figures because the minimum number of significant figures measured in the problem is two. The gold brick is more dense than water (which has a density of 1.0 g/mL), so it will sink.

Because the density of a substance is constant at a given temperature, we can use density values as conversion factors to determine either the mass or the volume of a substance.

Sample Problem 1.18 Problem Solving with Density

Table sugar has a density of 1.29 g/mL. Calculate the mass in grams of 1.00 teaspoon of sugar (1.00 teaspoon has a volume of 4.93 mL).

Solution

These problems are worked similarly to the unit-conversion problems solved in the previous section.

STEP 1 Determine the units on your final answer. In this case, the units will be the mass in grams.

STEP 2 Establish the given information. In this case, the density is given and you will solve the density equation for mass. You are also reminded of the volume equivalency that 1.00 tsp = 4.93 mL.

STEP 3 Decide how to set up the problem. You are given the density and the volume. To solve for mass, the density equation must be rearranged algebraically.

$$d = \frac{m}{V}$$

To rearrange the density equation, multiply both sides by V so the mass is now the variable that can be solved for.

$$d \times V = \frac{m}{\cancel{V}} \times \cancel{V}$$

$$m = d \times V$$

STEP 4 Solve the problem. The values and conversion factor are plugged in and the mL units cancel out, leaving g as the unit on the final answer.

$$m = \frac{1.29 \text{ g}}{1 \cancel{\text{mL}}} \times \frac{4.93 \cancel{\text{mL}}}{1.00 \cancel{\text{tsp}}} \times 1.00 \cancel{\text{tsp}} = 6.36 \text{ g}$$

Because the density and mass contain three significant figures, the answer must be given to three significant figures.

Specific Gravity

Which is more dense, water or honey? Both contain water, but if you pour honey into water (or tea, which is mostly water), it drops to the bottom. From this experiment, we can observe that honey must be more dense than liquid water. Honey, although a liquid, possesses the property of density because it has mass and takes up space.

For liquids, density often is measured with respect to water. The density of water is 1.00 g/mL at 4 °C and close to that value at body temperature. The ratio of the density of a substance to the density of water is called **specific gravity** (sp gr).

$$\text{Specific gravity} = \frac{\text{density of sample}}{\text{density of water}}$$

To calculate the specific gravity of a liquid, the density of the liquid and the density of water must have the same units. Specific gravity is a unitless quantity because it is a ratio of two densities.

The specific gravity of urine is routinely measured as an indicator of kidney function. If a person's kidneys are functioning normally, the density of urine will be similar to that of water. Normal specific gravity values are between 1.005 and 1.030. These urine values also indicate whether a person is over or underhydrated. Specific gravity values closer to 1.000 indicate overhydration, while values above 1.025 can signify dehydration. For liquids, specific gravity is measured with a simple instrument called a *refractometer* (see **FIGURE 1.6**).

FIGURE 1.6 A refractometer measures specific gravity. Measuring the specific gravity of urine using a refractometer (in lab tech's hand) can indicate kidney malfunction and hydration level.