



# CHEMISTRY

AN ATOMS FIRST APPROACH

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SECOND EDITION

# Periodic Table of the Elements

Alkali earth metals		Transition metals										Halogens		Noble gases			
1A	2A											7A	8A				
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H 1.008	He 4.003	Li 6.941	Be 9.012	B 10.81	C 12.01	N 14.01	O 16.00	F 19.00	Ne 20.18	Na 22.99	Mg 24.31	Al 26.98	Si 28.09	P 30.97	S 32.07	Cl 35.45	Ar 39.95
Alkali metals		metals										nonmetals					
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K 39.10	Ca 40.08	Sc 44.96	Ti 47.88	V 50.94	Cr 52.00	Mn 54.94	Fe 55.85	Co 58.93	Ni 58.69	Cu 63.55	Zn 65.38	Ga 69.72	Ge 72.59	As 74.92	Se 78.96	Br 79.90	Kr 83.80
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb 85.47	Sr 87.62	Y 88.91	Zr 91.22	Nb 92.91	Mo 95.94	Tc (98)	Ru 101.1	Rh 102.9	Pd 106.4	Ag 107.9	Cd 112.4	In 114.8	Sn 118.7	Sb 121.8	Te 127.6	I 126.9	Xe 131.3
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs 132.9	Ba 137.3	La* 138.9	Hf 178.5	Ta 180.9	W 183.9	Re 186.2	Os 190.2	Ir 192.2	Pt 195.1	Au 197.0	Hg 200.6	Tl 204.4	Pb 207.2	Bi 209.0	Po (209)	At (210)	Rn (222)
87	88	89	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118
Fr (223)	Ra (226)	Ac† (227)	Rf (261)	Db (262)	Sg (263)	Bh (264)	Hs (265)	Mt (268)	Ds (271)	Rg (272)	Cn (285)	Uut (289)	Fl (289)	Uup (293)	Lv (293)	Uus (293)	Uuo (293)
		*Lanthanides										†Actinides					
		58	59	60	61	62	63	64	65	66	67	68	69	70	71		
		Ce 140.1	Pr 140.9	Nd 144.2	Pm (145)	Sm 150.4	Eu 152.0	Gd 157.3	Tb 158.9	Dy 162.5	Ho 164.9	Er 167.3	Tm 168.9	Yb 173.0	Lu 175.0		
		90	91	92	93	94	95	96	97	98	99	100	101	102	103		
		Th 232.0	Pa (231)	U 238.0	Np (237)	Pu (244)	Am (243)	Cm (247)	Bk (247)	Cf (251)	Es (252)	Fm (257)	Md (258)	No (259)	Lr (260)		

Group numbers 1–18 represent the system recommended by the International Union of Pure and Applied Chemistry.

# Table of Atomic Masses\*

Element	Symbol	Atomic Number	Atomic Mass	Element	Symbol	Atomic Number	Atomic Mass	Element	Symbol	Atomic Number	Atomic Mass
Actinium	Ac	89	[227] <sup>§</sup>	Germanium	Ge	32	72.59	Potassium	K	19	39.10
Aluminum	Al	13	26.98	Gold	Au	79	197.0	Praseodymium	Pr	59	140.9
Americium	Am	95	[243]	Hafnium	Hf	72	178.5	Promethium	Pm	61	[145]
Antimony	Sb	51	121.8	Hassium	Hs	108	[265]	Protactinium	Pa	91	[231]
Argon	Ar	18	39.95	Helium	He	2	4.003	Radium	Ra	88	226
Arsenic	As	33	74.92	Holmium	Ho	67	164.9	Radon	Rn	86	[222]
Astatine	At	85	[210]	Hydrogen	H	1	1.008	Rhenium	Re	75	186.2
Barium	Ba	56	137.3	Indium	In	49	114.8	Rhodium	Rh	45	102.9
Berkelium	Bk	97	[247]	Iodine	I	53	126.9	Roentgenium	Rg	111	[272]
Beryllium	Be	4	9.012	Iridium	Ir	77	192.2	Rubidium	Rb	37	85.47
Bismuth	Bi	83	209.0	Iron	Fe	26	55.85	Ruthenium	Ru	44	101.1
Bohrium	Bh	107	[264]	Krypton	Kr	36	83.80	Rutherfordium	Rf	104	[261]
Boron	B	5	10.81	Lanthanum	La	57	138.9	Samarium	Sm	62	150.4
Bromine	Br	35	79.90	Lawrencium	Lr	103	[260]	Scandium	Sc	21	44.96
Cadmium	Cd	48	112.4	Lead	Pb	82	207.2	Seaborgium	Sg	106	[263]
Calcium	Ca	20	40.08	Livermorium	Lv	116	[293]	Selenium	Se	34	78.96
Californium	Cf	98	[251]	Lithium	Li	3	6.9419	Silicon	Si	14	28.09
Carbon	C	6	12.01	Lutetium	Lu	71	175.0	Silver	Ag	47	107.9
Cerium	Ce	58	140.1	Magnesium	Mg	12	24.31	Sodium	Na	11	22.99
Cesium	Cs	55	132.90	Manganese	Mn	25	54.94	Strontium	Sr	38	87.62
Chlorine	Cl	17	35.45	Meitnerium	Mt	109	[268]	Sulfur	S	16	32.07
Chromium	Cr	24	52.00	Mendelevium	Md	101	[258]	Tantalum	Ta	73	180.9
Cobalt	Co	27	58.93	Mercury	Hg	80	200.6	Technetium	Tc	43	[98]
Copernicium	Cn	112	[285]	Molybdenum	Mo	42	95.94	Tellurium	Te	52	127.6
Copper	Cu	29	63.55	Neodymium	Nd	60	144.2	Terbium	Tb	65	158.9
Curium	Cm	96	[247]	Neon	Ne	10	20.18	Thallium	Tl	81	204.4
Darmstadtium	Ds	110	[271]	Neptunium	Np	93	[237]	Thorium	Th	90	232.0
Dubnium	Db	105	[262]	Nickel	Ni	28	58.69	Thulium	Tm	69	168.9
Dysprosium	Dy	66	162.5	Niobium	Nb	41	92.91	Tin	Sn	50	118.7
Einsteinium	Es	99	[252]	Nitrogen	N	7	14.01	Titanium	Ti	22	47.88
Erbium	Er	68	167.3	Nobelium	No	102	[259]	Tungsten	W	74	183.9
Europium	Eu	63	152.0	Osmium	Os	76	190.2	Uranium	U	92	238.0
Fermium	Fm	100	[257]	Oxygen	O	8	16.00	Vanadium	V	23	50.94
Flerovium	Fl	114	[289]	Palladium	Pd	46	106.4	Xenon	Xe	54	131.3
Fluorine	F	9	19.00	Phosphorus	P	15	30.97	Ytterbium	Yb	70	173.0
Francium	Fr	87	[223]	Platinum	Pt	78	195.1	Yttrium	Y	39	88.91
Gadolinium	Gd	64	157.3	Plutonium	Pu	94	[244]	Zinc	Zn	30	65.38
Gallium	Ga	31	69.72	Polonium	Po	84	[209]	Zirconium	Zr	40	91.22

\*The values given here are to four significant figures where possible. <sup>§</sup>A value given in brackets denotes the mass of the longest-lived isotope.

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# Chemistry

AN ATOMS FIRST APPROACH

Second Edition

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**Chemistry: An Atoms First Approach,  
Second Edition**

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Photo Researcher: Sharon Donahue

Copy Editor: Graphic World

Text Designer: Ellen Pettengell Design

Cover Designer: Irene Morris

Cover Image: Laurie Knight/Nikon Instruments

Compositor: Graphic World

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WCN: 02-200-203

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Library of Congress Control Number: 2014942489

Student Edition:

ISBN: 978-1-305-07924-3

Loose-leaf Edition:

ISBN: 978-1-305-63267-7

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# To the Professor

## Features of *Chemistry: An Atoms First Approach*

Conceptual learning and problem solving have been fundamental to our approach in *Chemistry* through eight successful editions. Our philosophy is to help students learn to think like chemists so they can apply the process of problem solving to all aspects of their lives. In recent years a significant number of instructors has decided to engage their students by covering atoms, bonding, and structures of molecules at the beginning of their general chemistry courses, an approach often called “atoms first.” In order to provide proper context within the reorganized topics, we wrote the first edition of *Chemistry: An Atoms First Approach*. In that text, we covered atoms, molecules, energy, gases, liquids, and solids before stoichiometry. While we believe that instructors using the “atoms first” approach want their students to learn the theory of atomic structure and bonding early, feedback from users of the first edition indicated strongly that they want stoichiometry covered earlier in the text to ensure that students have sufficient background to complete basic measurement and stoichiometry experiments early in the course. We have responded to these suggestions in producing the second edition of *Chemistry: An Atoms First Approach*.

We found that users of the first edition strongly support our approach to general chemistry. We have always written with a student-first approach. That is, in writing every page we ask ourselves: “How can we explain the material in a way that will be most clear to the students?” We always develop concepts in accord with the scientific method. That is, we always consider the observed properties of substances first. We then look for common threads among these properties (formulate laws). Finally we help the students understand why and how the theories of chemistry developed. In describing theories we always make clear that models are works in progress. We expect to find areas where the models fail and, in fact, that this occurrence often leads to the greatest progress in our understanding of how nature operates.

One of the main goals of our treatment of chemistry is to help students learn to be effective problem solvers. We want to go beyond memorized steps to help students think their way through the problems. To do this we take a “think like a chemist” approach. In solving problems we ask students several questions to guide them through the process: Where are we going?, What do we know?, and How do we get there? Our goal is to foster creative, concept-based problem solving,

which will serve the students in their lives and careers beyond the general chemistry course.

Over the years, thousands of students and instructors who have used our books have found that these approaches work. Thus in the second edition of the text *Chemistry: An Atoms First Approach* we have returned to our strength: a text that effectively explains chemistry to the students and helps them to learn to be creative problem solvers. The most significant difference from our widely used *Chemistry* textbook is that we present atoms, bonding, and molecules at the beginning of the text and have made sure that the rest of the topics flow smoothly from this starting point, while moving stoichiometry to a position earlier in the text.

To strengthen the atoms first approach we have emphasized at every opportunity throughout the text the importance and advantages of thinking about chemistry from an atomic/molecular perspective.

## What’s New

Considerable effort went into making the second edition of *Chemistry: An Atoms First Approach* even more focused on the atom. To maintain this emphasis, we added new content (with an atoms first focus) throughout the textbook by revising the core text and adding margin notes, critical thinking questions, and boxed features:

- **Core Text**—Throughout the textbook, various chapters were updated with new content that “speaks” to the atoms first approach. Many introductions were updated to help students focus on an atomic/molecular point of view before delving into the chapter. Many sections were also revised as needed to help enhance the atoms first emphasis throughout the textbook.
- **Margin Notes**—New margin notes were added throughout the textbook to provide insight and expand upon principles being taught from an atoms first approach.
- **Critical Thinking Questions**—New Critical Thinking questions were added throughout the textbook to help students think about what they learned from an atomic/molecular point of view.
- **Connecting to Atoms**—A new boxed feature, “Connecting to Atoms,” contains summaries that cover key atoms first concepts at point-of-use. They have visuals and text to help students understand how an atomic/molecular approach clarifies chemical concepts.

- **ChemWork Problems**—These multiconcept problems (and additional ones) are found interactively online with the same type of assistance a student would get from an instructor.

In addition to the new material described above, we continue to present atoms, bonding, and molecules at the beginning of the text. The biggest change in the TOC involved moving stoichiometry to a position early in the text to where users prefer it. Additional changes from the first edition of *Chemistry: An Atoms First Approach* to the second edition are listed below:

- **Stoichiometry**—In the first edition of *Chemistry: An Atoms First Approach* we delayed stoichiometry until later in the text. However, users reported that this caused difficulty with the laboratory schedule. Therefore in the second edition we cover stoichiometry much earlier (Chapter 5). The new stoichiometry chapter provides a complete, coherent coverage of chemical stoichiometry.
- **Chemical Reactions and Solution Stoichiometry**—In keeping with the movement of stoichiometry to Chapter 5, we have presented chemical reactions and solution stoichiometry in Chapter 6.
- **Naming Simple Compounds**—The section on naming compounds was moved to the end of Chapter 3 (Bonding). Here we are taking advantage of an atoms first approach. Naming compounds makes more sense after bonding has been discussed.
- **Learning to Solve Problems**—The material on learning to solve problems has been moved to a position much earlier in the text (now in Chapter 5). We also provided a brief introduction to problem solving in Chapter R. It is never too early to start students thinking about solving problems in a conceptual manner.

## Hallmarks of *Chemistry: An Atoms First Approach*

- *Chemistry: An Atoms First Approach* contains numerous discussions, illustrations, and exercises aimed at *overcoming misconceptions*. It has become increasingly clear from our own teaching experience that students often struggle with chemistry because they misunderstand many of the fundamental concepts. In this text, we have gone to great lengths to provide illustrations and explanations aimed at giving students a more accurate picture of the fundamental ideas of chemistry. In particular, we have attempted to represent the microscopic world of chemistry so that students have a picture in their minds of “what the atoms and molecules are doing.” The art program along with the animations emphasize this goal. We have also placed a larger emphasis on the qualitative understanding of concepts before quantitative problems are considered. Because using
- an algorithm to correctly solve a problem often masks misunderstanding—when students assume they understand the material because they got the right “answer”—it is important to probe their understanding in other ways. In this vein, the text includes a number of *Active Learning Questions* at the end of each chapter that are intended for group discussion. It is our experience that students often learn the most when they teach each other. Students are forced to recognize their own lack of understanding when they try and fail to explain a concept to another student.
- With a strong problem-solving orientation, this text talks to students about how to approach and solve chemical problems. We emphasize a thoughtful, logical approach rather than simply memorizing procedures. This approach is thoroughly described in Section 5.3 (Learning to Solve Problems), which promotes the importance of thoughtful, creative problem solving. This section emphasizes to students that thinking through a problem produces more long-term, meaningful learning that can be applied to “real life” than memorizing steps that apply only to a particular type of problem. To help students adopt this way of thinking we have organized the problem-solving process in terms of:
- Where are we going?
  - What do we know?
  - How do we get there?
  - Reality check, which prompts students to check whether their answer makes sense
- As we proceed in the text, we gradually shift more responsibility to the students to think through the examples so that they do not become overly dependent on our help.
- One of the characteristics of this text is an innovative method for dealing with acid–base equilibria, the material the typical student finds most difficult and frustrating. The key to this approach involves first deciding what species are present in solution, then thinking about the chemical properties of these species. This method provides a general framework for approaching all types of solution equilibria.
- The text contains almost 250 *Examples*, with more given in the text discussions, to illustrate general problem-solving strategies. When a specific strategy is presented, it is summarized in a Problem-Solving Strategy box and the *Example* that follows it reinforces the use of the strategy to solve the problem. In general, we emphasize the use of conceptual understanding to solve problems rather than an algorithm-based approach. This approach is strongly reinforced by the inclusion of 204 *Interactive Examples*, which encourage students to thoughtfully consider the examples step-by-step.
  - We have presented a thorough *treatment of reactions* that occur in solution, including acid–base reactions. This material appears in Chapter 6, “Types of Chemical Reactions and Solution Stoichiometry,” directly after the chapter on

chemical stoichiometry, to emphasize the connection between solution reactions and chemical reactions in general. Chapter 6 also includes oxidation–reduction reactions and balancing by oxidation state, because a large number of interesting and important chemical reactions involve redox processes.

- *Descriptive chemistry* and chemical principles are thoroughly integrated in this text. Chemical models may appear sterile and confusing without the observations that stimulated their invention. On the other hand, facts without organizing principles may seem overwhelming. A combination of observation and models can make chemistry both interesting and understandable. In the chapter on the chemistry of the elements we have used tables and charts to show how properties and models correlate. Descriptive chemistry is presented in a variety of ways—as applications of principles in separate sections, in *Examples* and exercises, in photographs, and in *Chemical Connections*.
- Throughout the book a strong *emphasis on models* prevails. Coverage includes how they are constructed, how they are tested, and what we learn when they inevitably fail. Models are developed naturally, with pertinent observation always presented first to show why a particular model was invented.
- *Chemical Connections* boxes present applications of chemistry in various fields and in our daily lives. Margin notes in the *Instructor's Annotated Edition* also highlight many more *Chemical Connections* available on the student Web site.
- We offer end-of-chapter exercises for every type of student and for every kind of homework assignment: questions that promote group learning, exercises that reinforce student understanding, and problems that present the ultimate challenge with increased rigor and by integrating multiple concepts. To further encourage this approach we have included a selection of ChemWork Problems in the text. These multiconcept problems (and additional ones) are found interactively online with the same type of assistance a student would get from an instructor. We have also included biochemistry problems to make the connection for students in the course who are not chemistry majors.
- Judging from the favorable comments of instructors and students who have used our books, the text seems to work very well in a variety of courses. We are especially pleased that *readability* is cited as a key strength when students are asked to assess our textbooks.

## Supporting Materials

Please visit <http://www.cengage.com/chemistry/zumdahl/atomsfirst2e> for information about student and instructor resources for this book and about custom versions.

## Acknowledgments

This book represents the efforts of many talented and dedicated people. We particularly want to thank Charles Hartford, Product Manager, for his vision and oversight of the project. Charlie's knowledge of the field and enthusiasm has contributed immensely to the success of our textbooks. We also want to thank Teresa Trego, Senior Content Project Manager, who did an excellent job of coordinating the production of a very complex project. We also especially appreciate the dedication of Tom Martin, Content Developer, who managed the development process in a very supportive and organized manner. He contributed in many important ways to the successful completion of this edition, keeping the details in order and managing many different people with grace and good humor. We are also grateful for the excellent work of Brendan Killion, Media Developer. In addition, Morgan Carney, Product Assistant, did an outstanding job of communicating with us and coordinating the print supplements.

Don DeCoste, University of Illinois at Urbana–Champaign, provided invaluable contributions to this text. In particular he helped us focus on emphasizing to students the importance of thinking from an atomic/molecular point of view. We are especially grateful to Tom Hummel, University of Illinois at Urbana–Champaign, who managed the end-of-chapter problems and the solutions manuals. We greatly appreciate the outstanding work of Gretchen Adams, University of Illinois at Urbana–Champaign, who created the *Interactive Examples*. We are deeply indebted to these three colleagues who contribute in so many ways to our various projects. We are also very grateful to Jim Hall, University of Massachusetts–Lowell, for creating the excellent laboratory manual for the text. We are extremely appreciative of Jim's creative and accurate work over many years on many of our projects. We owe a large debt of appreciation to Professor Douglas Mulford of Emory University for providing a great deal of useful insight as we prepared the second edition. We followed many of Professor Mulford's suggestions as we revised the text.

Thanks to others who supplied valuable assistance for this text: Cate Ricard Barr, Art Director; Sharon Donahue, Photo Researcher; Janet Del Mundo, Marketing Manager; Lorraine Towle, Marketing Coordinator; Christina Brown, Digital Project Manager; and John Sarantakis, Permissions Project Manager.

We are especially thankful to all of the reviewers who participated in different aspects of the development process from reviewing chapters to providing feedback on the development of new features. We sincerely appreciate all of these suggestions.



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# To the Student

As you jump into the study of chemistry we hope that you will find our text helpful and interesting. Our job is to present the concepts and ideas of chemistry in a way you can understand. We hope to encourage you in your studies and to help you learn to solve problems in ways you can apply in all areas of your professional and personal lives.

Our main goal is to help you learn to become a truly creative problem solver. Our world badly needs people who can “think outside the box.” Our focus is to help you learn to think like a chemist. Why would you want to do that? Chemists are great problem solvers. They use logic, trial and error, and intuition—along with lots of patience—to work through complex problems. Chemists make mistakes, as we all do in our lives. The important thing that a chemist does is to learn from the mistakes and to try again. This “can do” attitude is useful in all careers.

In this book we develop the concepts in a natural way: The observations come first and then we develop models to explain the observed behavior. Models help us to understand and explain our world. They are central to scientific thinking. Models are very useful, but they also have limitations, which we will point out. By understanding the basic concepts in chemistry we lay the foundation for solving problems.

Our main goal is to help you learn a thoughtful method of problem solving. True learning is more than memorizing facts. Truly educated people use their factual knowledge as a starting point—a basis for creative problem solving. Our strategy for solving problems is explained in Section 5.3. To solve a problem we ask ourselves questions, which help us think through the problem. We let the problem guide us to the solution. This process can be applied to all types of problems in all areas of life.

As you study the text, use the *Examples* and the problem-solving strategies to help you. The strategies are boxed to

highlight them for you, and the *Examples* show how these strategies are applied.

After you have read and studied each chapter of the text you’ll need to practice your problem-solving skills. To do this we have provided plenty of review questions and end-of-chapter exercises. Your instructor may assign these on paper or online; in either case, you’ll want to work with your fellow students. One of the most effective ways to learn chemistry is through the exchange of ideas that comes from helping one another. The online homework assignments will give you instant feedback, and, in print, we have provided answers to some of the exercises in the back of the text. In all cases, your main goal is not just to get the correct answer, but to understand the process for getting the answer. Memorizing solutions for specific problems is not a very good way to prepare for an exam (or to solve problems in the real world!).

To become a great problem solver you’ll need these skills:

1. Look within the problem for the solution. (Let the problem guide you.)
2. Use the concepts you have learned along with a systematic, logical approach to find the solution.
3. Solve the problem by asking questions and learn to trust yourself to think it out.

You will make mistakes, but the important thing is to learn from these errors. The only way to gain confidence is to practice, practice, practice and to use your mistakes to find your weaknesses. Be patient with yourself and work hard to understand rather than simply memorize.

We hope you’ll have an interesting and successful year learning to think like a chemist!

*Steve and Susan Zumdahl*

# A GUIDE TO *Chemistry, An Atoms First Approach* SECOND EDITION

**Connecting To Atoms** This new boxed feature contains atoms first summaries that cover key atoms first concepts at point-of-use.

**Chemical Connections** Interesting applications of modern chemistry show students the relevance of chemistry to the real world.

**Connecting To Atoms** is a new boxed feature that contains summaries that cover key atoms first concepts at point-of-use. They have visuals and text to help students understand how an atomic/molecular approach clarifies chemical concepts. Examples of topics include Atoms, Ions, and Isotopes: A Pictorial Summary, Lewis Structures and the Periodic Table, Polar Molecules—It's All About Symmetry, and Entropy and Positional Probability.

### CONNECTING TO ATOMS 1-1 | Zero-, First-, and Second-Order Reactions and the Collision Model

In Section 11.4 we derived integrated rate laws for zero-, first-, and second-order reactions, and in Section 11.5 we discussed mechanisms for such reactions. Let's consider what each of these types of reactions would "look" like at the molecular level.

Let's consider the generic reaction  $A+B \rightarrow AB$ , a which we can represent as  $\bullet + \bullet \rightarrow \bullet\bullet$ . How could various possible mechanisms lead to different rate laws for this reaction?

**Possible mechanism that yields a second order rate law**  
As we discussed in Section 11.6, the molecules must collide in order to react. The most obvious rate law, then, would seem to be second order since we need two molecules to collide. A simple mechanism for a second-order reaction could be:

- $\bullet + \bullet \rightarrow \bullet\bullet$  (slow)
- $\bullet\bullet \rightarrow \bullet\bullet$  (fast)

Note that the first step is the rate-determining step, making the rate =  $k[A][B]$ , or second order.

**Possible mechanism that yields a zero order rate law**  
Zero-order reactions are often the result of catalysis. For example, suppose the reaction occurs on the surface of a metal catalyst, such as platinum (Pt). Once the surface is completely covered with AB molecules, increasing the concentration of AB would have no effect on the rate since only the AB molecules on the surface can react.

**Collision Model**  
In Section 11.6, you learned about the collision model, which tells us that in order for a reaction to occur, molecules must collide. We can see that even with this simple model, we can develop different mechanisms that give rise to different rate laws. As we have discussed, however, while a collision is necessary, it is not the only factor required to result in a reaction. The molecules must be oriented correctly (which is one way in which a catalyst works to speed up a reaction), and activation energy must be overcome (which is why reactions proceed at a higher rate with increased temperature). To better understand how reactions proceed and what factors are important in the rate of a reaction, we need to take a molecular-level approach to consider how the molecules are interacting.

### CONNECTING TO ATOMS 1-1 | Atoms, Ions, and Isotopes: A Pictorial Summary

**Atomic Number** (number of protons)  
**Mass Number** (total number of protons and neutrons)  
**Neutron Number** (number of neutrons)

**Isotopes** (same number protons (same atomic number), Different number neutrons (different mass number))

**Ions** (same number protons (same atomic number), Different number electrons (different mass number))

**Cation** (Positive charge, Number of protons > Number of electrons)

**Anion** (Negative charge, Number of electrons > Number of protons)

### CHEMICAL CONNECTIONS 1-2 | Berzelius, Selenium, and Silicon

Jöns Jakob Berzelius was probably the best experimental chemist of his generation and, given the crudeness of his laboratory equipment, may be the best of all time. Unlike Lavoisier, who could afford to buy the best laboratory equipment available, Berzelius worked with minimal equipment in very plain surroundings. One of Berzelius's students described the Swedish chemist's workshop: "The laboratory consisted of two ordinary rooms with the very simplest arrangements; there were neither furnaces nor hoods, neither water system nor gas. Against the wall, stood some closets with the chemicals, in the middle the mercury trough and the blast lamp table. Beside this was the sink consisting of a stone water holder with a stop-cock and a pot standing under it. (Next door in the kitchen) stood a small heating furnace."

In these simple facilities Berzelius performed more than 2000 experiments over a 19-year period to determine accurate atomic masses for the 50 elements then known. His success can be seen from the data in the table at left. These remarkably accurate values attest to his experimental skills and patience.

Besides his table of atomic masses, Berzelius made many other major contributions to chemistry. The most important of these was the invention of a simple set of symbols for the elements along with a system for writing the formulas of compounds to replace the awkward symbolic representations of the alchemists. Although some chemists, including Dalton, objected to the new system, it was gradually adopted and forms the basis of the system we use today.

In addition to these accomplishments, Berzelius discovered the elements cerium, thorium, selenium, and silicon. Of these elements, selenium and silicon are particularly important in today's world. Berzelius discovered selenium in 1817 in connection with his studies of sulfuric acid. For years selenium's toxicity has been known, but only recently have we become aware that it may have a positive effect on human health. Studies have shown that trace amounts of selenium in the diet may protect people from heart disease and cancer. One study based on data from 27 countries showed an inverse relationship between the cancer death rate and the selenium content of soil in a particular region (low cancer death rate in areas with high selenium content). Another research paper reported an inverse relationship between the selenium content of the blood and the incidence of breast cancer in women. A study reported in 1998 used the toenail clippings of 25,737 men to show that selenium seems to protect against prostate cancer. Selenium is also found in the heart muscle and may play an important role in proper heart function. Because of these and other studies, selenium's reputation has improved, and many scientists are now studying its function in the human body.

Silicon is the second most abundant element in the earth's crust, exceeded only by oxygen. As we will see in Chapter 8, compounds involving silicon bonded to oxygen make up most of the earth's sand, rock, and soil. Berzelius prepared silicon in its pure form in 1823 by heating silicon tetrachloride (SiCl<sub>4</sub>) with potassium metal. Today, silicon forms the basis for the modern microelectronics industry centered near San Francisco in a place that has come to be known as "Silicon Valley." The technology of the silicon chip (see Figure) with its printed circuits has transformed computers from room-sized monoliths to desktop and notebook-sized units with trouble-free "solid-state" circuitry.

Element	Berzelius's Value	Current Value
Chlorine	35.41	35.45
Copper	63.00	63.55
Hydrogen	1.00	1.01
Nitrogen	207.12	207.2
Nitrogen	14.05	14.01
Oxygen	16.00	16.00
Potassium	39.19	39.10
Silver	108.21	107.87
Sulfur	12.18	12.17

### CHEMICAL CONNECTIONS 1-1 | Enzymes: Nature's Catalysts

The most impressive examples of homogeneous catalysis occur in nature, where the complex reactions necessary for plant and animal life are made possible by enzymes. Enzymes are large molecules specifically tailored to facilitate a given type of reaction. Usually enzymes are proteins, an important class of biomolecules constructed from amino acids that have the general structure:

$$\begin{array}{c} \text{H} \\ | \\ \text{H}-\text{C}-\text{R} \\ | \\ \text{H} \end{array} \begin{array}{c} \text{O} \\ || \\ \text{C} \\ | \\ \text{OH} \end{array}$$

where R represents any one of 20 different substituents. These amino acid molecules can be "hooked together" to form a polymer (a word meaning "many parts") called a protein. The general structure of a protein can be represented as follows:

Since specific proteins are needed by the human body, the proteins in food must be broken into their constituent amino acids, which are then used to construct new proteins in the body's cells. The reaction in which a protein is broken down into amino acids is called hydrolysis. Note that in this reaction a water molecule reacts with a protein molecule to produce an amino acid and a new protein containing one less amino acid. Without the enzymes found in human cells, this reaction would be much too slow to be useful. One of these enzymes is carboxypeptidase A, a zinc-containing protein (Fig. 11-18).

Carboxypeptidase A captures the protein to be acted on (called the substrate) in a special groove and positions the substrate so that the end is in the active site, where the catalysis occurs (Fig. 11-19). Note that the Zn<sup>2+</sup> ion bonds to the oxygen of the C=O (carbonyl) group. This polarizes the electron density in the carbonyl group, allowing the neighboring

C—N bond to be broken much more easily. When the reaction is completed, the remaining portion of the substrate protein and the newly formed amino acid are released by the enzyme.

The process just described for carboxypeptidase A is characteristic of the behavior of other enzymes. Enzyme catalysis can be represented by the series of reactions shown below:

$$E + S \rightleftharpoons E \cdot S \rightleftharpoons E + P$$

where E represents the enzyme, S represents the substrate, E · S represents the enzyme-substrate complex, and P represents the products. The enzyme and substrate form a complex, where the reaction occurs. The enzyme then releases the product and is ready to repeat the process. The most amazing thing about enzymes is their efficiency. Because an enzyme plays its catalytic role over and over and very rapidly, only a tiny amount of enzyme is required. This makes the isolation of enzymes for study quite difficult.

**Figure 11-19** | Protein-substrate interaction. The substrate is shown in black and red, with the red representing the terminal amino acid. Blue indicates side chains from the enzyme that help bind the substrate.

**Chemical Connections** describe current applications of chemistry. These special-interest boxes cover such topics as the invention of Post-it Notes, farming the wind, and the use of iron metal to clean up contaminated groundwater. Additional *Chemical Connections* are available on the student Web site.

**Conceptual Understanding** Conceptual learning and problem solving are fundamental to the approach of **Chemistry**. The text gives students the tools to become critical thinkers: to ask questions, to apply rules and models, and to evaluate the outcome.

*“The first principles of the universe are atoms and empty space; everything else is merely thought to exist.”*

—DEMOCRITUS

The authors' **emphasis on modeling** (or chemical theories) throughout the text addresses the problem of rote memorization by helping students better understand and appreciate the process of scientific thinking. By stressing the limitations and uses of scientific models, the authors show students how chemists think and work.

#### 4-1 | Molecular Structure: The VSEPR Model

The structures of molecules play a very important role in determining their chemical properties. As we will see later, this is particularly important for biological molecules; a slight change in the structure of a large biomolecule can completely destroy its usefulness to a cell or may even change the cell from a normal one to a cancerous one. Many accurate methods now exist for determining **molecular structure**, the three-dimensional arrangement of the atoms in a molecule. These methods must be used if

#### Critical Thinking

We now have evidence that electron energy levels in the atoms are quantized. Some of this evidence is discussed in this chapter. What if energy levels in atoms were not quantized? What are some differences we would notice?

The text includes a number of open-ended **Critical Thinking** questions that emphasize the importance of conceptual learning. These questions are particularly useful for generating group discussion.

#### LET'S REVIEW | A Summary of the Hydrogen Atom

- In the quantum (wave) mechanical model, the electron is viewed as a standing wave. This representation leads to a series of wave functions (orbitals) that describe the possible energies and spatial distributions available to the electron.
- In agreement with the Heisenberg uncertainty principle, the model cannot specify the detailed electron motions. Instead, the square of the wave function represents the probability distribution of the electron in that orbital. This allows us to picture orbitals in terms of probability distributions, or electron density maps.
- The size of an orbital is arbitrarily defined as the surface that contains 90% of the total electron probability.
- The hydrogen atom has many types of orbitals. In the ground state, the single electron resides in the 1s orbital. The electron can be excited to higher-energy orbitals if energy is put into the atom.

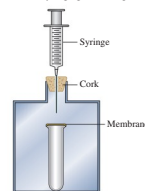
**Let's Review** boxes help students organize their thinking about the crucial chemical concepts that they encounter.

The text includes a number of **Active Learning Questions** at the end of each chapter that are intended for group discussion, as students often learn the most when they teach each other.

#### Active Learning Questions

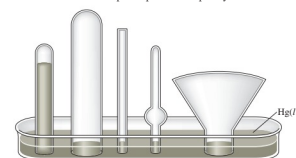
These questions are designed to be used by groups of students in class.

1. Consider the following apparatus: a test tube covered with a nonpermeable elastic membrane inside a container that is closed with a cork. A syringe goes through the cork.

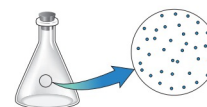


- As you push down on the syringe, how does the membrane covering the test tube change?
  - You stop pushing the syringe but continue to hold it down. In a few seconds, what happens to the membrane?
2. Figure 8-2 shows a picture of a barometer. Which of the following statements is the best explanation of how this barometer works?
- Air pressure outside the tube causes the mercury to move in the tube until the air pressure inside and outside the tube is equal.
  - Air pressure inside the tube causes the mercury to move in the tube until the air pressure inside and outside the tube is equal.
  - Air pressure outside the tube counterbalances the weight of the mercury in the tube.

- Capillary action of the mercury causes the mercury to go up the tube.
  - The vacuum that is formed at the top of the tube holds up the mercury.
- Justify your choice, and for the choices you did not pick, explain what is wrong with them. Pictures help!
3. The barometer below shows the level of mercury at a given atmospheric pressure. Fill all the other barometers with mercury for that same atmospheric pressure. Explain your answer.



4. As you increase the temperature of a gas in a sealed, rigid container, what happens to the density of the gas? Would the results be the same if you did the same experiment in a container with a piston at constant pressure? (See Fig. 8-17.)
5. A diagram in a chemistry book shows a magnified view of a flask of air as follows:



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**Problem Solving** This text talks to the student about how to approach and solve chemical problems, as one of the main goals of general chemistry is to help students become creative problem solvers. The authors emphasize a thoughtful, logical approach rather than simply memorizing procedures.

### 5-3 | Learning to Solve Problems

One of the great rewards of studying chemistry is to become a good problem-solver. Being able to solve complex problems is a talent that will serve you well in all walks of life. It is our purpose in this text to help you learn to solve problems in a flexible, creative way based on understanding the fundamental ideas of chemistry. We call this approach **conceptual problem solving**.

The ultimate goal is to be able to solve new problems (that is, problems you have not seen before) on your own. In this text we will provide problems and offer solutions by explaining how to think about the problems. While the answers to these problems are important, it is perhaps even more important to understand the process—the thinking necessary to get the answer. Although at first we will be solving the problem for you, do not take a passive role. While studying the solution, it is crucial that you interactively think through the problem with us. Do not skip the discussion and jump to the answer. Usually, the solution will involve asking a series of questions. Make sure that you understand each step in the process. This active approach should apply to problems outside of chemistry as well. For example, imagine riding with someone in a car to an unfamiliar destination. If your goal is simply to have the other person get you to that destination, you will probably not pay much attention to how to get there (passive), and if you have to find this same place in the future on your own, you will probably not be able to do it. If, however, your goal is to learn how to get there, you would pay attention to distances, signs, and turns (active). This is how you should read the solutions in the text (and the text in general).

While actively studying our solutions to problems is helpful, at some point you will need to know how to think through these problems on your own. If we help you too much as you solve a problem, you won't really learn effectively. If we always "drive," you won't interact as meaningfully with the material. Eventually you need to learn to drive yourself. We will provide more help at the beginning of the text and less as we proceed to later chapters.

There are two fundamentally different ways you might use to approach a problem. One way emphasizes memorization. We might call this the "pigeonholing method." In this approach, the first step is to label the problem—to decide in which pigeonhole it fits. The pigeonholing method requires that we provide you with a set of steps that you memorize and store in the appropriate slot for each different problem you encounter. The difficulty with this method is that it requires a new pigeonhole each time a problem is changed by even a small amount.

Consider the driving analogy again. Suppose you have memorized how to drive from your house to the grocery store. Do you know how to drive back from the grocery store to your house? Not necessarily. If you have only memorized the directions and do not understand fundamental principles such as "I traveled north to get to the store, so my house is south of the store," you may find yourself stranded. In a more complicated example, suppose you know how to get from your house to the store (and back) and from your house to the library (and back). Can you get from the library to the store without having to go back home? Probably not if you have only memorized directions and you do not have a "big picture" of where your house, the store, and the library are relative to one another.



▲ Pigeon holes can be used for sorting and classifying objects like mail.

In **Chapter 5**, "Stoichiometry," the authors dedicate a section, **Learning to Solve Problems**, that emphasizes the importance of problem solving. This section helps students understand that thinking their way through a problem produces more long-term, meaningful learning than simply memorizing steps, which are soon forgotten.

Chapters 1–8 introduce a series of questions into the in-chapter **Examples** to engage students in the process of problem solving, such as **Where are we going?** and **How do we get there?** This more active approach helps students think their way through the solution to the problem.

#### EXAMPLE 5-1 The Average Mass of an Element



▲ Copper nugget.

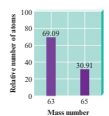


Figure 5-3 | Mass spectrum of natural copper.

When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 5-3 are obtained. Use these data to compute the average mass of natural copper. (The mass values for  $^{63}\text{Cu}$  and  $^{65}\text{Cu}$  are 62.93 u and 64.93 u, respectively.)

**Solution**  
*Where are we going?*  
To calculate the average mass of natural copper

*What do we know?*

- $^{63}\text{Cu}$  mass = 62.93 u
- $^{65}\text{Cu}$  mass = 64.93 u

*How do we get there?*  
As shown by the graph, of every 100 atoms of natural copper, 69.09 are  $^{63}\text{Cu}$  and 30.91 are  $^{65}\text{Cu}$ . Thus the mass of 100 atoms of natural copper is

$$(69.09 \text{ atoms}) \left( \frac{62.93 \text{ u}}{\text{atom}} \right) + (30.91 \text{ atoms}) \left( \frac{64.93 \text{ u}}{\text{atom}} \right) = 6355 \text{ u}$$

The average mass of a copper atom is

$$\frac{6355 \text{ u}}{100 \text{ atoms}} = 63.55 \text{ u/atom}$$

This mass value is used in doing calculations involving the reactions of copper and is the value given in the table inside the front cover of this book.

#### INTERACTIVE EXAMPLE 5-3

##### Determining Moles of Atoms

Aluminum (Al) is a metal with a high strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. Compute both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.



▲ (left) Pure aluminum. (right) Aluminum alloys are used for many products used in our kitchens.

##### Solution

*Where are we going?*

To calculate the moles and number of atoms in a sample of Al

*What do we know?*

- Sample contains 10.0 g of Al
- Mass of 1 mole ( $6.022 \times 10^{23}$  atoms) of Al = 26.93 g

*How do we get there?*

We can calculate the number of moles of Al in a 10.0-g sample as follows:

$$10.0 \text{ g-Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g-Al}} = 0.371 \text{ mol Al atoms}$$

The number of atoms in 10.0 g (0.371 mole) of aluminum is

$$0.371 \text{ mol-Al} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol-Al}} = 2.23 \times 10^{23} \text{ atoms}$$

**REALITY CHECK** | One mole of Al has a mass of 26.98 g and contains  $6.022 \times 10^{23}$  atoms. Our sample is 10.0 g, which is roughly 1/3 of 26.98. Thus the calculated amount should be on the order of 1/3 of  $6 \times 10^{23}$ , which it is.

See Exercise 5-46

**Interactive Examples** engage students in the problem-solving process by requiring them to think through the example step-by-step rather than simply scanning the written example in the text as many students do.

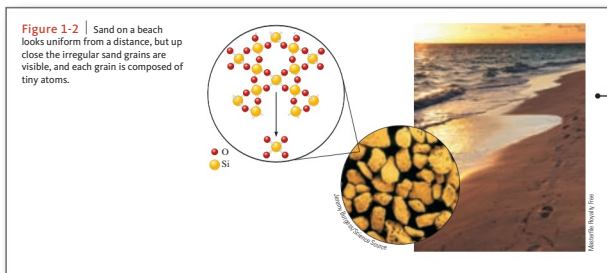
#### PROBLEM-SOLVING STRATEGY

##### Steps to Apply the VSEPR Model

1. Draw the Lewis structure for the molecule.
2. Count the electron pairs and arrange them in the way that minimizes repulsion (that is, put the pairs as far apart as possible).
3. Determine the positions of the atoms from the way the electron pairs are shared.
4. Determine the name of the molecular structure from the positions of the atoms.

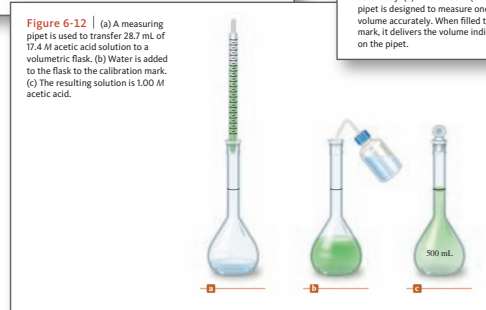
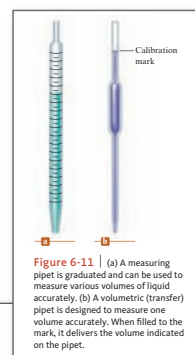
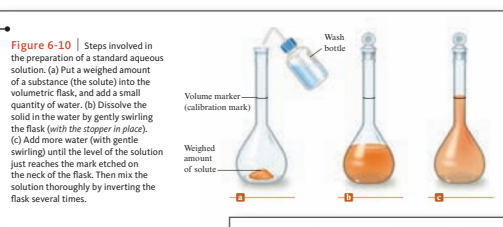
**Problem-Solving Strategy** boxes focus students' attention on the very important process of problem solving.

**Dynamic Art Program** Most of the glassware, orbitals, graphs, flowcharts, and molecules have been redrawn to better serve visual learners and enhance the textbook.



The art program emphasizes molecular-level interactions that help students visualize the “micro/macro” connection.

Realistic drawings of glassware and instrumentation found in the lab help students make real connections.



**Table 9-8** | Intramolecular (Bonding) Forces

Type of Bonding	Nature of Attraction	Model	Range of Energies (kJ/mol)	Examples
Ionic	Interactions among ions (cations and anions)		400–4,000	LiF
Covalent	Atoms sharing electrons		150–1,100	H <sub>2</sub>
Metallic	Atoms sharing electrons in a nondirectional way (ions in a “sea of electrons”)		75–100	Au

**Table 9-9** | Intermolecular Forces

Type of Interaction	Nature of Attraction	Model	Range of Energies (kJ/mol)	Examples
Ion-dipole	Attraction between charge on anion and end of dipole with opposite partial charge		35–400	
Dipole-dipole	Attraction between opposite partially charged ends of polar molecules		5–30	
Hydrogen bond	Attraction between partially positively charged H atom attached to a highly electronegative atom and a lone pair on another atom		10–40	
London dispersion	Attraction between an instantaneous dipole and a dipole induced in a neighboring atom or molecule		<1–40	

Tables containing visuals help students understand and compare intramolecular and intermolecular forces between ions, atoms, and molecules.

**FOR REVIEW**

**Key Terms**

- Section 14-1  
common ion effect
- Section 14-2  
buffered solution  
Henderson-Hasselbalch equation
- Section 14-3  
buffering capacity
- Section 14-4  
pH curve (titration curve)  
millimole (mmol)  
equivalence point  
(stoichiometric point)
- Section 14-5  
acid-base indicator  
phenolphthalein

**Buffered Solutions**

- Contains a weak acid (HA) and its salt (NaA) or a weak base (B) and its salt (BHCl)
- Resists a change in its pH when H<sup>+</sup> or OH<sup>-</sup> is added
- For a buffered solution containing HA and A<sup>-</sup>
  - The Henderson-Hasselbalch equation is useful:  
$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$
  - The capacity of the buffered solution depends on the amounts of HA and A<sup>-</sup> present
  - The most efficient buffering occurs when the  $\frac{[\text{A}^-]}{[\text{HA}]}$  ratio is close to 1
- Buffering works because the amounts of HA (which reacts with added OH<sup>-</sup>) and A<sup>-</sup> (which reacts with added H<sup>+</sup>) are large enough that the  $\frac{[\text{A}^-]}{[\text{HA}]}$  ratio does not change significantly when strong acids or bases are added

**Acid-Base Titrations**

- The progress of a titration is represented by plotting the pH of the solution versus the volume of added titrant; the resulting graph is called a pH curve or titration curve
- Strong acid-strong base titrations show a sharp change in pH near the equivalence point
- The shape of the pH curve for a strong base-strong acid titration before the equivalence point is quite different from the shape of the pH curve for a strong base-weak acid titration
  - The strong base-weak acid pH curve shows the effects of buffering before the equivalence point
  - For a strong base-weak acid titration, the pH is greater than 7 at the equivalence point because of the basic properties of A<sup>-</sup>
- Indicators are sometimes used to mark the equivalence point of an acid-base titration
  - The end point is where the indicator changes color
  - The goal is to have the end point and the equivalence point be as close as possible

**Review Questions** Answers to the Review Questions can be found on the Student Web site (accessible from [www.cengagebrain.com](http://www.cengagebrain.com)).

1. What is meant by the presence of a common ion?  
How does the presence of a common ion affect an equilibrium such as  
$$\text{HNO}_2(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{NO}_2^-(\text{aq})$$
  
What is an acid-base solution called that contains a common ion?
2. Define a buffer solution. What makes up a buffer solution? How do buffers absorb added H<sup>+</sup> or OH<sup>-</sup> with little pH change?

It is necessary that the concentrations of the weak acid and the weak base in a buffered solution be equal? Explain. What is the pH of a buffer when the weak acid and conjugate base concentrations are equal?  
A buffer generally contains a weak acid and its weak conjugate base, or a weak base and its weak conjugate acid, in water. You can solve for the pH by setting up the equilibrium problem using the  $K_a$  reaction of the weak acid or the  $K_b$  reaction of the weak base.

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Each chapter has a **For Review** section to reinforce key concepts and includes review questions for students to practice independently.

**Active Learning Questions** are designed to promote discussion among groups of students in class.

51a Chapter 1 Chemical Foundations

**Active Learning Questions**

These questions are designed to be used by groups of students in class.

1. Paracelsus, a sixteenth-century alchemist and healer, adopted as his slogan: "The patients are your textbook, the sickbed is your study." Is this view consistent with using the scientific method?
2. What is wrong with the following statement? "The results of the experiment do not agree with the theory. Something must be wrong with the experiment."
3. Which of the following is true about an individual atom? Explain.
  - a. An individual atom should be considered to be a solid.
  - b. An individual atom should be considered to be a liquid.
  - c. An individual atom should be considered to be a gas.
  - d. The state of the atom depends on which element it is.
  - e. An individual atom cannot be considered to be a solid, liquid, or gas.
 Justify your choice, and for choices you did not pick, explain what is wrong with them.
4. These questions concern the work of J. J. Thomson.
  - a. From Thomson's work, which particles do you think he would feel are most important for the formation of compounds (chemical changes), and why?
  - b. Of the remaining two subatomic particles, which do you place second in importance for forming compounds, and why?
  - c. Propose three models that explain Thomson's findings and evaluate them. To be complete you should include Thomson's findings.
5. Which of the following explain how an ion is formed? Explain your answer.
  - a. adding or subtracting protons to/from an atom
  - b. adding or subtracting neutrons to/from an atom
  - c. adding or subtracting electrons to/from an atom
6. You have a chemical in a sealed glass container filled with air. The setup is sitting on a balance as shown below. The chemical is ignited by means of a magnifying glass focusing sunlight on the reactant. After the chemical has completely burned, which of the following is true? Explain your answer.



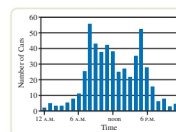
- a. The balance will read less than 250.0 g.
- b. The balance will read 250.0 g.
- c. The balance will read greater than 250.0 g.
- d. The scale's reading cannot be determined without knowing the identity of the chemical.

7. You may have noticed that when water boils, you can see bubbles that rise to the surface of the water. Which of the following is inside these bubbles? Explain.
  - a. air
  - b. hydrogen and oxygen gas
  - c. oxygen gas
  - d. water vapor
  - e. carbon dioxide gas
8. One of the best indications of a useful theory is that it raises more questions for further experimentation than it originally answered. Does this apply to Dalton's atomic theory? Give examples.
9. Dalton assumed that all atoms of the same element were identical in all their properties. Explain why this assumption is not valid.
10. Which (if any) of the following can be determined by knowing the number of protons in a neutral element? Explain your answer.
  - a. the number of neutrons in the neutral element
  - b. the number of electrons in the neutral element
  - c. the name of the element

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the *Student Solutions Manual*.

**Questions**

11. The difference between a *law* and a *theory* is the difference between *what* and *why*. Explain.
12. As part of a science project, you study traffic patterns in your city at an intersection in the middle of downtown. You set up a device that counts the cars passing through this intersection for a 24-hour period during a weekday. The graph of hourly traffic looks like this.



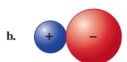
- a. At what time(s) does the highest number of cars pass through the intersection?
- b. At what time(s) does the lowest number of cars pass through the intersection?
- c. Briefly describe the trend in numbers of cars over the course of the day.

# Comprehensive End-of-Chapter Practice and Review

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the *Student Solutions Manual*.

## Questions

17. Compare and contrast the bonding found in the  $\text{H}_2(\text{g})$  and  $\text{HF}(\text{g})$  molecules with that found in  $\text{NaF}(\text{s})$ .
18. The following electrostatic potential diagrams represent  $\text{H}_2$ ,  $\text{HCl}$ , or  $\text{NaCl}$ . Label each and explain your choices.



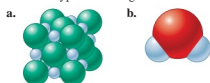
19. Describe the type of bonding that exists in the  $\text{Cl}_2(\text{g})$  molecule. How does this type of bonding differ from that found in the  $\text{HCl}(\text{g})$  molecule? How is it similar?

20. Some of the important properties of ionic compounds are as follows:

- low electrical conductivity as solids and high conductivity in solution or when molten
- relatively high melting and boiling points
- brittleness
- solubility in polar solvents

How does the concept of ionic bonding discussed in this chapter account for these properties?

21. Label the type of bonding for each of the following.



22. Distinguish between the following terms.

- molecule versus ion
- covalent bonding versus ionic bonding

**Questions** are homework problems directed at concepts within the chapter and in general don't require calculation.

There are numerous **Exercises** to reinforce students' understanding of each section. These problems are paired and organized by topic so that instructors can review them in class and assign them for homework.

## 197a Chapter 4 Molecular Structure and Orbitals

- The bond angle in  $\text{SO}_2$  should be similar to the bond angle in  $\text{CS}_2$  or  $\text{SCl}_2$ .
  - Of the compounds  $\text{CF}_4$ ,  $\text{KrF}_4$ , and  $\text{SeF}_4$ , only  $\text{SeF}_4$  exhibits an overall dipole moment (is polar).
  - Central atoms in a molecule adopt a geometry of the bonded atoms and lone pairs about the central atom in order to maximize electron repulsions.
10. Give one example of a compound having a linear molecular structure that has an overall dipole moment (is polar) and one example that does not have an overall dipole moment (is nonpolar). Do the same for molecules that have trigonal planar and tetrahedral molecular structures.
11. In the hybrid orbital model, compare and contrast  $\sigma$  bonds with  $\pi$  bonds. What orbitals form the  $\sigma$  bonds and what orbitals form the  $\pi$  bonds? Assume the  $z$ -axis is the internuclear axis.
12. Write the name of each of the following molecular structures.



13. Give the expected hybridization for the molecular structures illustrated in the previous question.
14. In the molecular orbital model, compare and contrast  $\sigma$  bonds with  $\pi$  bonds. What orbitals form the  $\sigma$  bonds and what orbitals form the  $\pi$  bonds? Assume the  $z$ -axis is the internuclear axis.
15. Why are  $d$  orbitals sometimes used to form hybrid orbitals? Which period of elements does not use  $d$  orbitals for hybridization? If necessary, which  $d$  orbitals ( $3d$ ,  $4d$ ,  $5d$ , or  $6d$ ) would sulfur use to form hybrid orbitals requiring  $d$  atomic orbitals? Answer the same question for arsenic and for iodine.

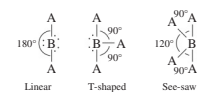
16. The atoms in a single bond can rotate about the internuclear axis without breaking the bond. The atoms in a double and triple bond cannot rotate about the internuclear axis unless the bond is broken. Why?
17. Compare and contrast bonding molecular orbitals with antibonding molecular orbitals.
18. What modification to the molecular orbital model was made from the experimental evidence that  $\text{B}_2$  is paramagnetic?
19. Why does the molecular orbital model do a better job in explaining the bonding in  $\text{NO}^+$  and  $\text{NO}$  than the hybrid orbital model?
20. The three  $\text{NO}$  bonds in  $\text{NO}_3^-$  are all equivalent in length and strength. How is this explained even though any valid Lewis structure for  $\text{NO}_3^-$  has one double bond and two single bonds to nitrogen?

## Exercises

In this section, similar exercises are paired.

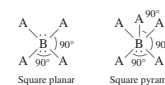
### Molecular Structure and Polarity

21. Predict the molecular structure (including bond angles) for each of the following.
- $\text{SeO}_3$
  - $\text{SeO}_2$
22. Predict the molecular structure (including bond angles) for each of the following.
- $\text{PCl}_3$
  - $\text{SCl}_2$
  - $\text{SiF}_4$
23. Predict the molecular structure and bond angles for each molecule or ion in Exercises 81 and 87 from Chapter 3.
24. Predict the molecular structure and bond angles for each molecule or ion in Exercises 82 and 88 from Chapter 3.
25. There are several molecular structures based on the trigonal bipyramidal geometry (see Table 4-3). Three such structures are



Which of the compounds/ions  $\text{BrF}_3^-$ ,  $\text{ClF}_3$ ,  $\text{XeF}_4$ ,  $\text{SF}_4$ ,  $\text{PF}_5$ ,  $\text{ClF}_4$ , and  $\text{SF}_6$  have these molecular structures?

26. Two variations of the octahedral geometry (see Table 4-1) are illustrated below.



Which of the compounds/ions  $\text{BrF}_3^-$ ,  $\text{ClF}_3$ ,  $\text{XeF}_4$ ,  $\text{SF}_4$ ,  $\text{PF}_5$ ,  $\text{ClF}_4$ , and  $\text{SF}_6$  have these molecular structures?



**Wealth of End-of-Chapter Problems** The text offers an unparalleled variety of end-of-chapter content with problems that increase in rigor and integrate multiple concepts.

New **ChemWork** end-of-chapter multiconcept problems, with many additional problems, are found interactively online with the same type of assistance a student would get from an instructor. These problems are available to assign online for more practice.

**ChemWork Problems**

These multiconcept problems (and additional ones) are found interactively online with the same type of assistance a student would get from an instructor.

60. Complete the following table, including the mass number and the atomic number with the symbol for the isotope.

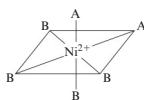
Number of Protons	Number of Neutrons	Symbol
9	10	
13	14	
53	74	
34	45	
16	16	

61. Complete the following table.

Atoms	Number of Protons	Number of Neutrons
${}^4_2\text{He}$		
${}^{20}_{10}\text{Ne}$		
${}^{48}_{22}\text{Ti}$		
${}^{190}_{78}\text{Os}$		
${}^{59}_{27}\text{Co}$		

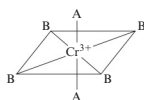
**Challenge Problems**

89. Consider the following complex ion, where A and B represent ligands.



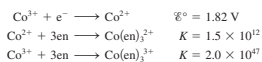
The complex is known to be diamagnetic. Do A and B produce very similar or very different crystal fields? Explain.

90. Consider the pseudo-octahedral complex ion of  $\text{Cr}^{3+}$ , where A and B represent ligands.



Ligand A produces a stronger crystal field than ligand B. Draw an appropriate crystal field diagram for this complex ion (assume the A ligands are on the z-axis).

91. Consider the following data:



where en = ethylenediamine.

- Calculate  $\mathcal{E}^\circ$  for the half-reaction  $\text{Co}(\text{en})_3^{3+} + e^- \longrightarrow \text{Co}(\text{en})_3^{2+}$
  - Based on your answer to part a, which is the stronger oxidizing agent,  $\text{Co}^{3+}$  or  $\text{Co}(\text{en})_3^{3+}$ ?
  - Use the crystal field model to rationalize the result in part b.
92. Henry Taube, 1983 Nobel Prize winner in chemistry, has studied the mechanisms of the oxidation-reduction reactions of transition metal complexes. In one experiment he and his students studied the following reaction:
- $$\text{Cr}(\text{H}_2\text{O})_6^{2+}(\text{aq}) + \text{Co}(\text{NH}_3)_6\text{Cl}^{2+}(\text{aq}) \longrightarrow \text{Cr}(\text{III}) \text{ complexes} + \text{Co}(\text{II}) \text{ complexes}$$

**Challenge Problems** take students one step further and challenge them more rigorously than the Additional Exercises.

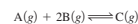
**Integrative Problems** combine concepts from multiple chapters.

**Integrative Problems**

These problems require the integration of multiple concepts to find the solutions.

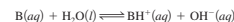
118. Some nonelectrolyte solute (molar mass = 142 g/mol) was dissolved in 150. mL of a solvent (density = 0.879 g/cm<sup>3</sup>). The elevated boiling point of the solution was 355.4 K. What mass of solute was dissolved in the solvent? For the solvent, the enthalpy of vaporization is 33.90 kJ/mol, the entropy of vaporization is 95.95 J/K · mol, and the boiling-point elevation constant is 2.5 K · kg/mol.

119. For the equilibrium



the initial concentrations are  $[\text{A}] = [\text{B}] = [\text{C}] = 0.100 \text{ atm}$ . Once equilibrium has been established, it is found that  $[\text{C}] = 0.040 \text{ atm}$ . What is  $\Delta G^\circ$  for this reaction at 25°C?

120. What is the pH of a 0.125-M solution of the weak base B if  $\Delta H^\circ = -28.0 \text{ kJ}$  and  $\Delta S^\circ = -175 \text{ J/K}$  for the following equilibrium reaction at 25°C?



**Marathon Problem**

This problem is designed to incorporate several concepts and techniques into one situation.

116. Consider the following reaction:



At 25°C, the following two experiments were run, yielding the following data:

Experiment 1:  $[\text{Y}]_0 = 3.0 \text{ M}$

$[\text{CH}_3\text{X}]$ (mol/L)	Time (h)
$7.08 \times 10^{-3}$	1.0
$4.52 \times 10^{-3}$	1.5
$2.23 \times 10^{-3}$	2.3
$4.76 \times 10^{-4}$	4.0
$8.44 \times 10^{-5}$	5.7
$2.75 \times 10^{-5}$	7.0

**Marathon Problems** also combine concepts from multiple chapters; they are the most challenging problems in the end-of-chapter material.

# About the Authors



**Steven S. Zumdahl** earned a B.S. in Chemistry from Wheaton College (IL) and a Ph.D. from the University of Illinois, Urbana–Champaign. He has been a faculty member at the University of Colorado–Boulder, Parkland College (IL), and the University of Illinois at Urbana–Champaign (UIUC), where he is Professor Emeritus. He has received numerous awards, including the National Catalyst Award for Excellence in Chemical Education, the University of Illinois Teaching Award, the UIUC Liberal Arts and Sciences Award for Excellence in Teaching, UIUC Liberal Arts and Sciences Advising Award, and the School of Chemical Sciences Teaching Award (five times). He is the author of several chemistry textbooks. In his leisure time he enjoys traveling and collecting classic cars.

**Susan A. Zumdahl** earned a B.S. and M.A. in Chemistry at California State University–Fullerton. She has taught science and mathematics at all levels, including middle school, high school, community college, and university. At the University of Illinois at Urbana–Champaign, she developed a program for increasing the retention of minorities and women in science and engineering. This program focused on using active learning and peer teaching to encourage students to excel in the sciences. She has coordinated and led workshops and programs for science teachers from elementary through college levels. These programs encourage and support active learning and creative techniques for teaching science. For several years she was director of an Institute for Chemical Education (ICE) field center in Southern California, and she has authored several chemistry textbooks. Susan spearheaded the development of a sophisticated Web-based electronic homework system for teaching chemistry. She enjoys traveling, classic cars, and gardening in her spare time—when she is not playing with her grandchildren.





R E V I E W

# Measurement and Calculations in Chemistry

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**R-1** Units of Measurement

**R-2** Uncertainty in Measurement

**Precision and Accuracy**

**R-3** Significant Figures and Calculations

**R-4** Learning to Solve Problems Systematically

**R-5** Dimensional Analysis

**R-6** Temperature

**R-7** Density

**R-8** Classification of Matter

**R-9** Energy

**R-10** The Mole

*Glassware used to measure volumes of liquids. (Alexander Rath/Shutterstock.com)*

**M**aking observations is fundamental to all science. These observations can be qualitative or quantitative. A quantitative observation is called a **measurement**, which always has two parts: a **number** and a scale (called a **unit**). Both parts must be present for a measurement to be meaningful. A qualitative observation does not involve a number. Examples of qualitative observations are “the substance is blue” and “the sun is very hot.”

In this chapter we will discuss measurements in detail and explain the various properties of the numbers and units associated with measurements. This material should be familiar to you from previous science courses, but we include it here to provide a review of these topics that are critical to the operations of chemistry.

## R-1 | Units of Measurement



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▲ Soda is commonly sold in 2-liter bottles—an example of the use of SI units in everyday life.

❗ A unit such as volume that is based on a fundamental unit is called a “derived unit.”

In our study of chemistry we will use measurements of mass, length, time, temperature, electric current, and the amount of a substance, among others. Scientists recognized long ago that standard systems of units had to be adopted if measurements were to be useful. If every scientist had a different set of units, complete chaos would result. Unfortunately, different standards were adopted in different parts of the world. The two major systems are the *English system* used in the United States and the *metric system* used by most of the rest of the industrialized world. This duality causes a good deal of trouble; for example, parts as simple as bolts are not interchangeable between machines built according to the two systems. As a result, the United States has begun to adopt the metric system.

Most scientists in all countries have for many years used the metric system. In 1960, an international agreement set up a system of units called the *International System* (*le Système International* in French), or the **SI system**. This system is based on the metric system and units derived from the metric system. The fundamental SI units are listed in Table R-1. We will discuss how to manipulate these units later in this chapter. ◀

Because the fundamental units are not always convenient (expressing the mass of a pin in kilograms is awkward), prefixes are used to change the size of the unit. These are listed in Table R-2. Some common objects and their measurements in SI units are listed in Table R-3.

One physical quantity that is very important in chemistry is *volume*, which is not a fundamental SI unit but is derived from length. ◀ ❗ A cube that measures 1 meter (m) on each edge is represented in Fig. R-1. This cube has a volume of  $(1 \text{ m})^3 = 1 \text{ m}^3$ . Because there are 10 decimeters (dm) in a meter, the volume of this cube is  $(1 \text{ m})^3 = (10 \text{ dm})^3 = 1000 \text{ dm}^3$ . A cubic decimeter, that is  $(1 \text{ dm})^3$ , is commonly called a *liter* (L), which is a unit of volume slightly larger than a quart. As shown in Fig. R-1,

**Table R-1** | The Fundamental SI Units

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

**Table R-2** | The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	$10^{18}$
peta	P	1,000,000,000,000,000	$10^{15}$
tera	T	1,000,000,000,000	$10^{12}$
giga	G	1,000,000,000	$10^9$
mega	M	1,000,000	$10^6$
kilo	k	1,000	$10^3$
hecto	h	100	$10^2$
deka	da	10	$10^1$
—	—	1	$10^0$
deci	d	0.1	$10^{-1}$
centi	c	0.01	$10^{-2}$
milli	m	0.001	$10^{-3}$
micro	$\mu$	0.000001	$10^{-6}$
nano	n	0.000000001	$10^{-9}$
pico	p	0.000000000001	$10^{-12}$
femto	f	0.000000000000001	$10^{-15}$
atto	a	0.000000000000000001	$10^{-18}$

\*See Appendix 1.1 if you need a review of exponential notation.

**Table R-3** | Some Examples of Commonly Used Units

Unit	Example
<b>Length</b>	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
<b>Mass</b>	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
<b>Volume</b>	A 12-oz can of soda has a volume of about 360 mL.

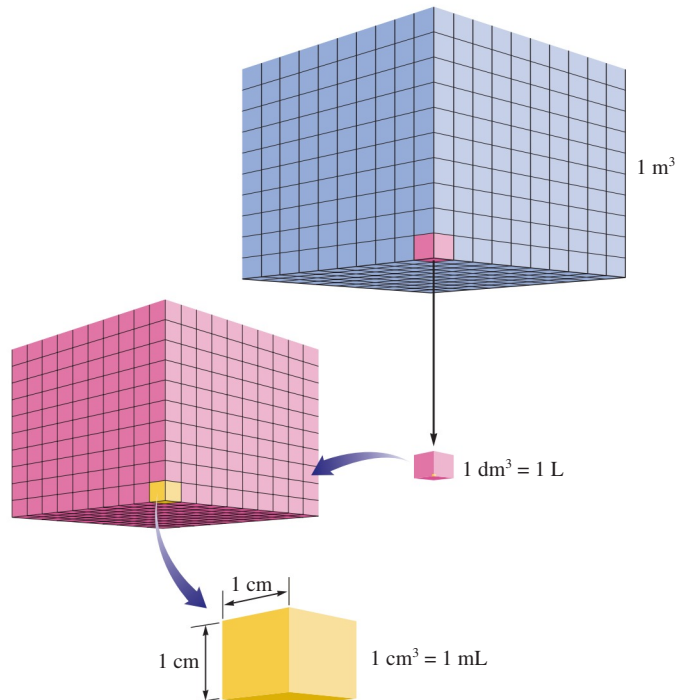
1000 liters are contained in a cube with a volume of 1 cubic meter. Similarly, since 1 decimeter equals 10 centimeters (cm), the liter can be divided into 1000 cubes each with a volume of 1 cubic centimeter:

$$1 \text{ liter} = (1 \text{ dm})^3 = (10 \text{ cm})^3 = 1000 \text{ cm}^3$$

Also, since  $1 \text{ cm}^3 = 1 \text{ milliliter (mL)}$ ,

$$1 \text{ liter} = 1000 \text{ cm}^3 = 1000 \text{ mL}$$

Thus 1 liter contains 1000 cubic centimeters, or 1000 milliliters.



**Figure R-1** | The largest cube has sides 1 m in length and a volume of  $1 \text{ m}^3$ . The middle-sized cube has sides 1 dm in length and a volume of  $1 \text{ dm}^3$ , or 1 L. The smallest cube has sides 1 cm in length and a volume of  $1 \text{ cm}^3$ , or 1 mL.

## CHEMICAL CONNECTIONS R-1 | *Critical Units!*

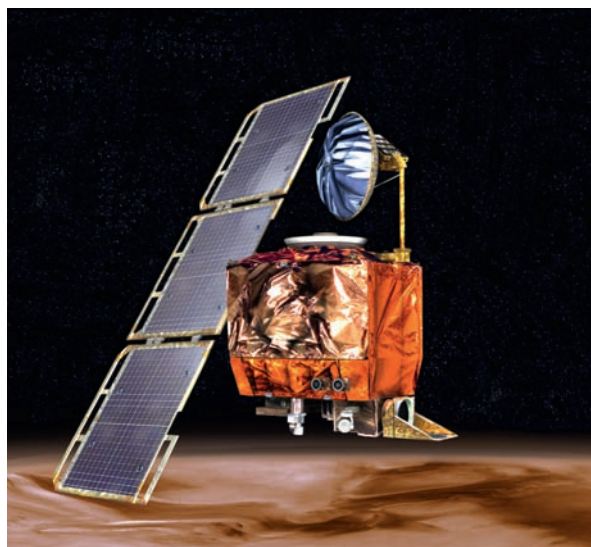
How important are conversions from one unit to another? If you ask the National Aeronautics and Space Administration (NASA), very important! In 1999, NASA lost a \$125 million Mars Climate Orbiter because of a failure to convert from English to metric units.

The problem arose because two teams working on the Mars mission were using different sets of units. NASA's scientists at the Jet Propulsion Laboratory in Pasadena, California, assumed that the thrust data for the rockets on the Orbiter they received from Lockheed Martin Astronautics in Denver, which built the spacecraft, were in metric units. In reality, the units were English. As a result the Orbiter dipped 100 kilometers lower into the Mars atmosphere than planned, and the friction from the atmosphere caused the craft to burn up.

NASA's mistake refueled the controversy over whether Congress should require the United States to switch to the metric system. About 95% of the world now uses the metric system, and the United States is slowly switching from English to metric. For example, the automobile industry has adopted metric fasteners, and we buy our soda in two-liter bottles.

Units can be very important. In fact, they can mean the difference between life and death on some occasions. In 1983,

for example, a Canadian jetliner almost ran out of fuel when someone pumped 22,300 pounds of fuel into the aircraft instead of 22,300 kilograms. Remember to watch your units!



Artist's conception of the lost Mars Climate Orbiter.

Chemical laboratory work frequently requires measurement of the volumes of liquids. Several devices for the accurate determination of liquid volume are shown in Fig. R-2.

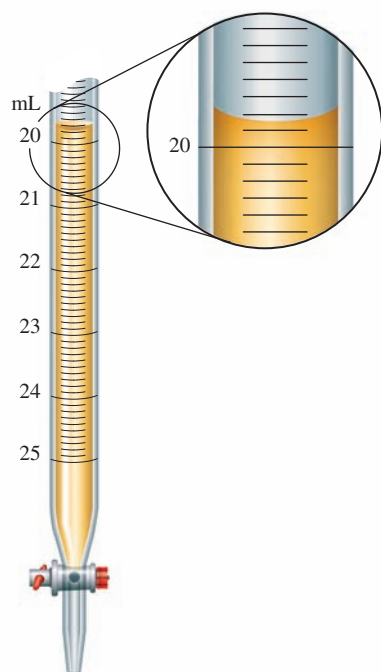
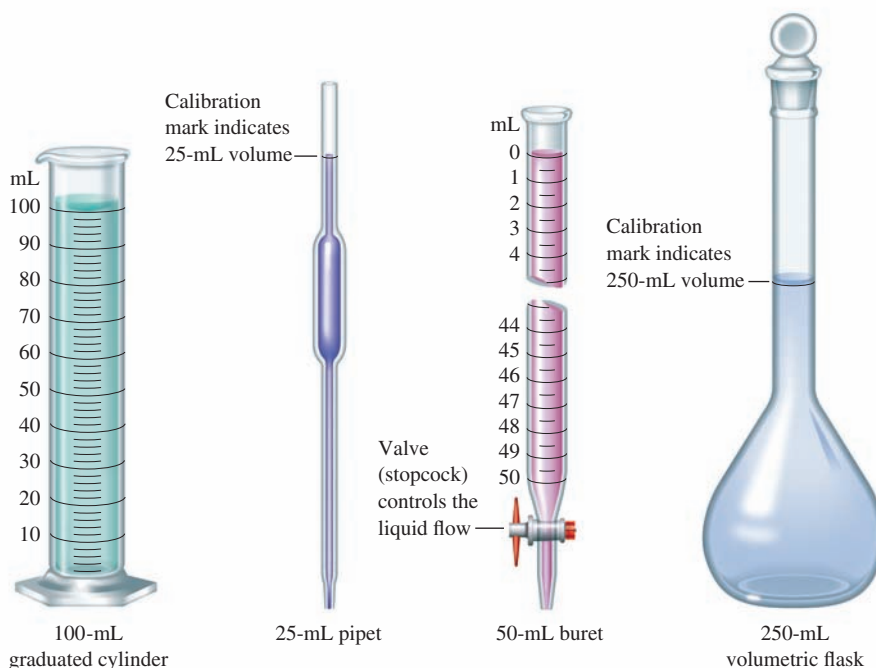
An important point concerning measurements is the relationship between mass and weight. Although these terms are sometimes used interchangeably, they are *not* the same. **Mass** is a measure of the resistance of an object to a change in its state of motion. Mass is measured by the force necessary to give an object a certain acceleration. On the earth we use the force that gravity exerts on an object to measure its mass. We call this force the object's **weight**. Since weight is the response of mass to gravity, it varies with the strength of the gravitational field. Therefore, your body mass is the same on the earth or on the moon, but your weight would be much less on the moon than on the earth because of the moon's smaller gravitational field.

Because weighing something on a chemical balance involves comparing the mass of that object to a standard mass, the terms *weight* and *mass* are sometimes used interchangeably, although this is incorrect.

## R-2 | Uncertainty in Measurement

The number associated with a measurement is obtained using some measuring device. For example, consider the measurement of the volume of a liquid using a buret (shown in Fig. R-3 with the scale greatly magnified). Notice that the meniscus of the liquid occurs at about 20.15 milliliters. This means that about 20.15 mL of liquid has been

**Figure R-2** | Common types of laboratory equipment used to measure liquid volume.



**Figure R-3** | Measurement of volume using a buret. The volume is read at the bottom of the liquid curve (called the meniscus).

**i** A measurement always has some degree of uncertainty.

delivered from the buret (if the initial position of the liquid meniscus was 0.00 mL). Note that we must estimate the last number of the volume reading by interpolating between the 0.1-mL marks. Since the last number is estimated, its value may be different if another person makes the same measurement. If five different people read the same volume, the results might be as follows:

Person	Results of Measurement
1	20.15 mL
2	20.14 mL
3	20.16 mL
4	20.17 mL
5	20.16 mL

These results show that the first three numbers (20.1) remain the same regardless of who makes the measurement; these are called *certain* digits. However, the digit to the right of the 1 must be estimated and therefore varies; it is called an *uncertain* digit. We customarily report a measurement by recording all the certain digits plus the *first* uncertain digit. In our example it would not make any sense to try to record the volume of thousandths of a milliliter, because the value for hundredths of a milliliter must be estimated when using the buret.

It is very important to realize that a *measurement always has some degree of uncertainty*. The uncertainty of a measurement depends on the precision of the measuring device. For example, using a bathroom scale, you might estimate the mass of a grapefruit to be approximately 1.5 pounds. Weighing the same grapefruit on a highly precise balance might produce a result of 1.476 pounds. In the first case, the uncertainty occurs in the tenths of a pound place; in the second case, the uncertainty occurs in the thousandths of a pound place. **i** Suppose we weigh two similar grapefruits on the two devices and obtain the following results:

	Bathroom Scale	Balance
Grapefruit 1	1.5 lb	1.476 lb
Grapefruit 2	1.5 lb	1.518 lb



**i** Uncertainty in measurement is discussed in more detail in Appendix 1-5.

Do the two grapefruits have the same mass? The answer depends on which set of results you consider. Thus a conclusion based on a series of measurements depends on the certainty of those measurements. For this reason, it is important to indicate the uncertainty in any measurement. **i** This is done by always recording the certain digits and the first uncertain digit (the estimated number). These numbers are called the **significant figures** of a measurement.

The convention of significant figures automatically indicates something about the uncertainty in a measurement. The uncertainty in the last number (the estimated number) is usually assumed to be  $\pm 1$  unless otherwise indicated. For example, the measurement 1.86 kilograms can be taken to mean  $1.86 \pm 0.01$  kilograms.

## INTERACTIVE EXAMPLE R-1

### Uncertainty in Measurement

In analyzing a sample of polluted water, a chemist measured out a 25.00-mL water sample with a pipet (see Fig. R-2). At another point in the analysis, the chemist used a graduated cylinder (see Fig. R-2) to measure 25 mL of a solution. What is the difference between the measurements 25.00 mL and 25 mL?

#### Solution

Even though the two volume measurements appear to be equal, they really convey different information. The quantity 25 mL means that the volume is between 24 mL and 26 mL, whereas the quantity 25.00 mL means that the volume is between 24.99 mL and 25.01 mL. The pipet measures volume with much greater precision than does the graduated cylinder.

See Question R-21

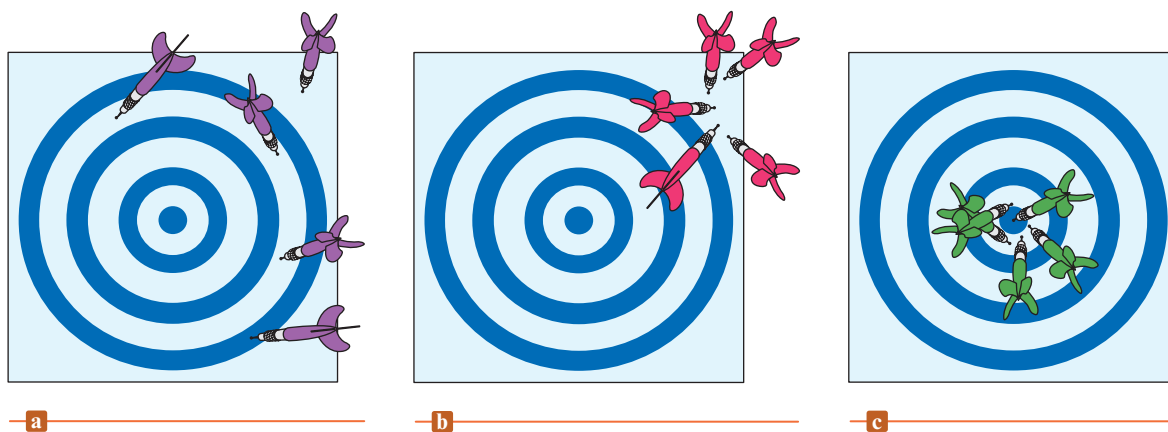
When making a measurement, it is important to record the results to the appropriate number of significant figures. For example, if a certain buret can be read to  $\pm 0.01$  mL, you should record a reading of twenty-five milliliters as 25.00 mL, not 25 mL. This way, at some later time when you are using your results to do calculations, the uncertainty in the measurement will be known to you.

### Precision and Accuracy

Two terms often used to describe the reliability of measurements are *precision* and *accuracy*. Although these words are frequently used interchangeably in everyday life, they have different meanings in the scientific context. **Accuracy** refers to the agreement of a particular value with the true value. **Precision** refers to the degree of agreement among several measurements of the same quantity. Precision reflects the *reproducibility* of a given type of measurement. The difference between these terms is illustrated by the results of three different dart throws shown in Fig. R-4.

Two different types of errors are illustrated in Fig. R-4. A **random error** (also called an *indeterminate error*) means that a measurement has an equal probability of being high or low. This type of error occurs in estimating the value of the last digit of a measurement. The second type of error is called **systematic error** (or *determinate error*). This type of error occurs in the same direction each time; it is either always high or always low. Fig. R-4(a) indicates large random errors (poor technique). Fig. R-4(b) indicates small random errors but a large systematic error, and Fig. R-4(c) indicates small random errors and no systematic error.

In quantitative work, precision is often used as an indication of accuracy; we assume that the *average* of a series of precise measurements (which should “average out” the random errors because of their equal probability of being high or low) is accurate, or close to the “true” value. However, this assumption is valid only if systematic errors are



**Figure R-4** | The results of several dart throws show the difference between precise and accurate. (a) Neither accurate nor precise (large random errors). (b) Precise but not accurate (small random errors, large systematic error). (c) Bull's-eye! Both precise and accurate (small random errors, no systematic error).

absent. Suppose we weigh a piece of brass five times on a very precise balance and obtain the following results:

Weighing	Result
1	2.486 g
2	2.487 g
3	2.485 g
4	2.484 g
5	2.488 g

Normally, we would assume that the true mass of the piece of brass is very close to 2.486 grams, which is the average of the five results:

$$\frac{2.486 \text{ g} + 2.487 \text{ g} + 2.485 \text{ g} + 2.484 \text{ g} + 2.488 \text{ g}}{5} = 2.486 \text{ g}$$

However, if the balance has a defect causing it to give a result that is consistently 1.000 gram too high (a systematic error of +1.000 gram), then the measured value of 2.486 grams would be seriously in error. The point here is that high precision among several measurements is an indication of accuracy *only* if systematic errors are absent.

## EXAMPLE R-2

### Precision and Accuracy

To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret (see Fig. R-2) and then read the volume delivered. Following are the results of five trials:

Trial	Volume Shown by Graduated Cylinder	Volume Shown by the Buret
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	25 mL	26.54 mL

Is the graduated cylinder accurate?

**i** Precision is an indication of accuracy only if there are no systematic errors.

### Solution

The results of the trials show very good precision (for a graduated cylinder). **i** The student has good technique. However, note that the average value measured using the buret is significantly different from 25 mL. Thus this graduated cylinder is not very accurate. It produces a systematic error (in this case, the indicated result is low for each measurement)

See Question R-2

## R-3 | Significant Figures and Calculations

Calculating the final result for an experiment usually involves adding, subtracting, multiplying, or dividing the results of various types of measurements. Since it is very important that the uncertainty in the final result is known correctly, we have developed rules for counting the significant figures in each number and for determining the correct number of significant figures in the final result.

**i** Leading zeros are never significant figures.

**i** Captive zeros are always significant figures.

**i** Trailing zeros are sometimes significant figures.

**i** Exact numbers never limit the number of significant figures in a calculation.

**i** Exponential notation is reviewed in Appendix 1-1.

### Rules for Counting Significant Figures

- 1. Nonzero integers.** Nonzero integers always count as significant figures.
- 2. Zeros.** There are three classes of zeros:
  - a. Leading zeros** are zeros that *precede* all the nonzero digits. These do not count as significant figures. In the number 0.0025, the three zeros simply indicate the position of the decimal point. This number has only two significant figures. **i**
  - b. Captive zeros** are zeros *between* nonzero digits. These always count as significant figures. The number 1.008 has four significant figures. **i**
  - c. Trailing zeros** are zeros at the *right end* of the number. They are significant only if the number contains a decimal point. The number 100 has only one significant figure, whereas the number  $1.00 \times 10^2$  has three significant figures. The number one hundred written as 100. also has three significant figures. **i**
- 3. Exact numbers.** Many times calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an infinite number of significant figures. Other examples of exact numbers are the 2 in  $2\pi r$  (the circumference of a circle) and the 4 and the 3 in  $\frac{4}{3}\pi r^3$  (the volume of a sphere). Exact numbers also can arise from definitions. For example, one inch is defined as *exactly* 2.54 centimeters. Thus, in the statement  $1 \text{ in} = 2.54 \text{ cm}$ , neither the 2.54 nor the 1 limits the number of significant figures when used in a calculation. **i**

Note that the number  $1.00 \times 10^2$  above is written in **exponential notation**. **i** This type of notation has at least two advantages: the number of significant figures can be easily indicated, and fewer zeros are needed to write a very large or very small number. For example, the number 0.000060 is much more conveniently represented as  $6.0 \times 10^{-5}$ . (The number has two significant figures.)

### INTERACTIVE EXAMPLE R-3

#### Significant Figures

Give the number of significant figures for each of the following results.

- a.** A student's extraction procedure on tea yields 0.0105 g of caffeine.
- b.** A chemist records a mass of 0.050080 g in an analysis.
- c.** In an experiment a span of time is determined to be  $8.050 \times 10^{-3}$  s.

### Solution

- The number contains three significant figures. The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant.
- The number contains five significant figures. The leading zeros (to the left of the 5) are not significant. The captive zeros between the 5 and the 8 are significant, and the trailing zero to the right of the 8 is significant because the number contains a decimal point.
- This number has four significant figures. Both zeros are significant.

See Exercises R-15 through R-18

To this point we have learned to count the significant figures in a given number. Next, we must consider how uncertainty accumulates as calculations are carried out. The detailed analysis of the accumulation of uncertainties depends on the type of calculation involved and can be complex. However, in this textbook we will employ the following simple rules that have been developed for determining the appropriate number of significant figures in the result of a calculation.

### Rules for Significant Figures in Mathematical Operations

- For multiplication or division, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation. For example, consider the calculation

$$\begin{array}{ccc}
 4.56 \times 1.4 = 6.38 & \xrightarrow{\text{Corrected}} & 6.4 \\
 \uparrow & & \uparrow \\
 \text{Limiting term has} & & \text{Two significant} \\
 \text{two significant} & & \text{figures} \\
 \text{figures} & & 
 \end{array}$$

The product should have only two significant figures, since 1.4 has two significant figures.

- For addition or subtraction, the result has the same number of decimal places as the least precise measurement used in the calculation. For example, consider the sum

$$\begin{array}{r}
 12.11 \\
 18.0 \\
 \hline
 1.013 \\
 31.123
 \end{array}
 \xrightarrow{\text{Corrected}}
 \begin{array}{r}
 31.1 \\
 \uparrow \\
 \text{One decimal place}
 \end{array}$$

← Limiting term has one decimal place

The correct result is 31.1, since 18.0 has only one decimal place.

**i** For multiplication and division: significant figures are counted. For addition and subtraction: decimal places are counted.

Note that for multiplication and division, significant figures are counted. For addition and subtraction, the decimal places are counted. **i**

In most calculations you will need to round numbers to obtain the correct number of significant figures. The following rules should be applied when rounding.

### Rules for Rounding

- In a series of calculations, carry the extra digits through to the final result, *then* round.
- If the digit to be removed
  - is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.
  - is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4. **i**

**i** Rule 2 is consistent with the operation of electronic calculators.