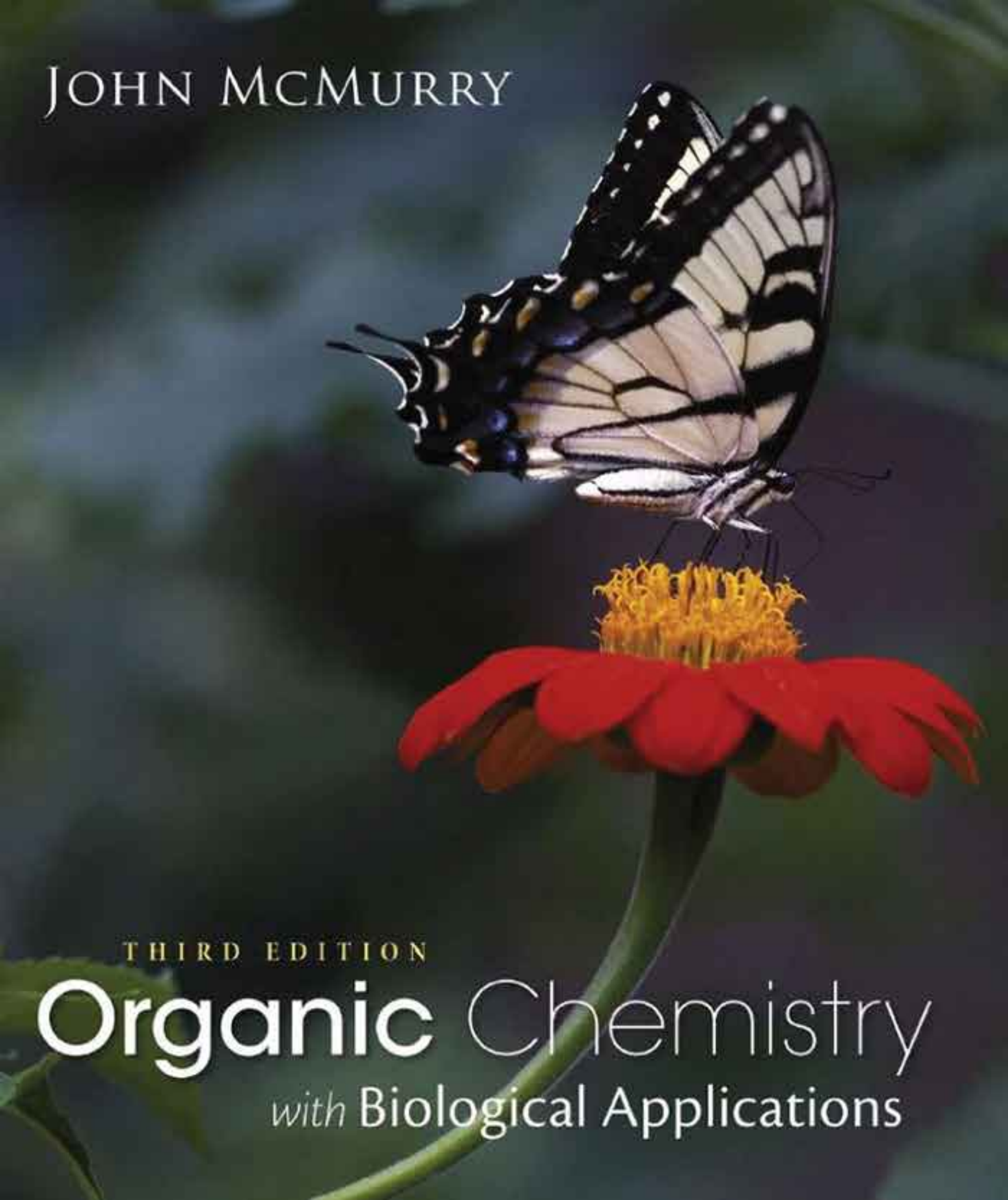
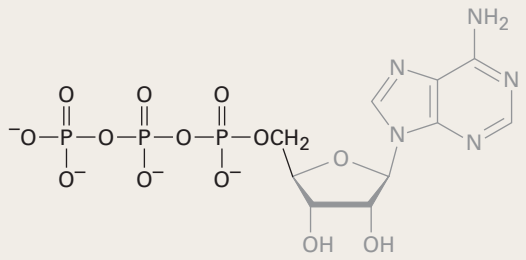

JOHN MCMURRY
THIRD EDITION
Organic Chemistry
with Biological Applications

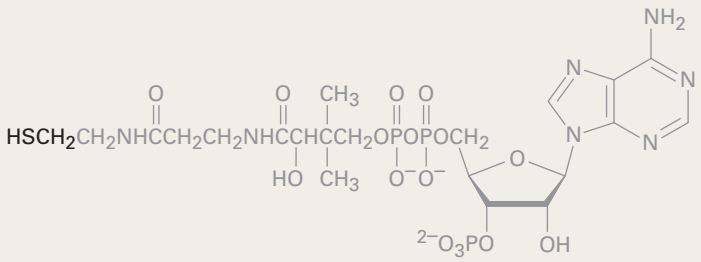
Structures of Common Coenzymes

The reactive parts of the molecules are darkened, while nonreactive parts are ghosted.

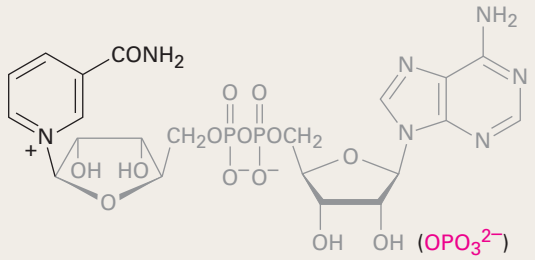
Adenosine triphosphate—ATP (phosphorylation)



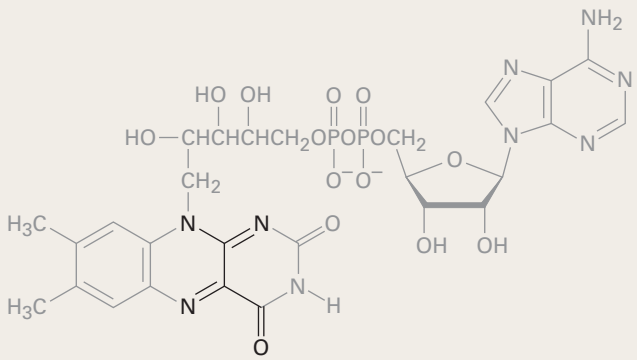
Coenzyme A (acyl transfer)



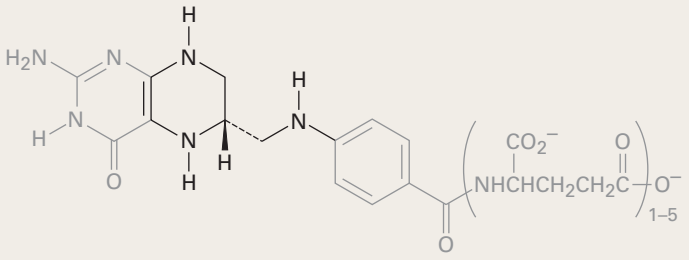
Nicotinamide adenine dinucleotide—NAD^+ (oxidation/reduction)
($NADP^+$)



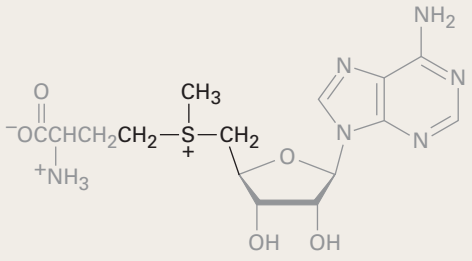
Flavin adenine dinucleotide—FAD (oxidation/reduction)



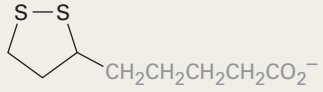
Tetrahydrofolate (transfer of C_1 units)



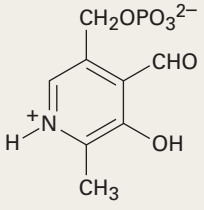
***S*-Adenosylmethionine (methyl transfer)**



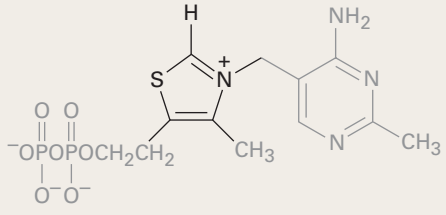
Lipoic acid (acyl transfer)



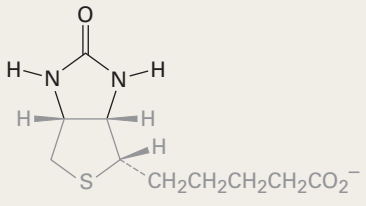
Pyridoxal phosphate (amino acid metabolism)



Thiamin diphosphate (decarboxylation)



Biotin (carboxylation)



Cornell University

Department of Chemistry
and Chemical Biology

John McMurry
Professor of Chemistry and
Chemical Biology

Baker Laboratory
Ithaca, New York 14853-4203

Dear Colleague:

All of us who teach organic chemistry know that most of the students in our courses, even the chemistry majors, are interested primarily in the life sciences rather than in pure chemistry. Because we are teaching so many future biologists, biochemists, and doctors rather than younger versions of ourselves, more and more of us are questioning why we continue to teach the way we do. Why do we spend so much time discussing the details of reactions that are of interest to research chemists but have little connection to biology? Why don't we instead spend more time discussing the organic chemistry of living organisms?

There is still much to be said for teaching organic chemistry in the traditional way, but it is also true that until now there has been no real alternative for those instructors who want to teach somewhat differently. And that is why I wrote **Organic Chemistry with Biological Applications 3e**. As chemical biology continues to gain in prominence, I suspect that more and more faculty will be changing their teaching accordingly.

Make no mistake: this is still a textbook on organic chemistry. But my guiding principle in deciding what to include and what to leave out has been to focus almost exclusively on those reactions that have a direct counterpart in biological chemistry. The space saved by leaving out nonbiological reactions has been put to good use, for almost every reaction discussed is followed by a biological example and approximately 25% of the book is devoted entirely to biomolecules and the organic chemistry of their biotransformations. In addition, **Organic Chemistry with Biological Applications 3e** is nearly 200 pages shorter than standard texts, making it possible for faculty to cover the entire book in a typical two-semester course.

Organic Chemistry with Biological Applications 3e is different from any other text; I believe that it is ideal for today's students.

Sincerely,

John McMurry

All royalties from *Organic Chemistry with Biological Applications* will be donated to the Cystic Fibrosis (CF) Foundation. This book and donation are dedicated to the author's eldest son and to the thousands of others who daily fight this disease. To learn more about CF and the programs and services provided by the CF Foundation, please visit **http://www.cff.org**.

Organic Chemistry

WITH BIOLOGICAL APPLICATIONS

3rd Edition

Organic Chemistry

WITH BIOLOGICAL APPLICATIONS

John McMurry

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e27 Synthetic Polymers | 925

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PREFACE

I've taught organic chemistry many times for many years. Like most faculty, I began by trying to show 19-year-old students the logic and beauty of the subject, thinking that they would find it as fascinating as I did. It didn't take long, though, before I realized what a disconnect there was between my own interests and expectations and those of my students. Some students did develop a real appreciation for the subject, but most seemed to worry primarily about getting into medical school. And why not? If a student has a clear career goal, why shouldn't that person focus his or her efforts toward meeting that goal?

All of us who teach organic chemistry know that the large majority of our students—90% or more, and including many chemistry majors—are interested primarily in medicine, biology, and other life sciences rather than in pure chemistry. But if we are primarily teaching future physicians, biologists, biochemists, and others in the life sciences (not to mention the occasional lawyer, politician, or business person), why do we continue to teach the way we do? Why do our textbooks and lectures spend so much time discussing details of topics that interest professional chemists but have no connection to biology? Wouldn't the limited amount of time we have be better spent paying more attention to the organic chemistry of living organisms and less to the organic chemistry of the research laboratory? Wouldn't it better serve our students if we helped them reach *their* goals rather than reach goals we set for them? I believe so, and I have written this book, *Organic Chemistry with Biological Applications*, third edition, to encourage others who might also be thinking that the time has come to do things a bit differently.

This is, first and foremost, a textbook on organic chemistry. Look through it and you'll find that almost all the standard topics are here, although the treatment of some has been attenuated to save space. Nevertheless, my guiding principle in writing this text has been to put a greater emphasis on those organic reactions and topics that are relevant to biological chemistry than on those that are not.

Organic chemistry, which began historically as the chemistry of living organisms, is now shifting back in that direction, judging from the increasing amount of biologically oriented research done in many chemistry departments and from the renaming of many departments to include chemical biology. Shouldn't our teaching reflect that shift?

Organization of the Text

Four distinct groups of chapters are apparent in this text. The first group (Chapters 1–6 and 10–11) covers the traditional principles of organic chemistry and spectroscopy that are essential for building further understanding.

The second group (Chapters 7–9 and 12–18) covers the common organic reactions found in all texts. As each laboratory reaction is discussed, however, a biological example is also shown to make the material more interesting and meaningful to students. For instance, trans fatty acids are described at the same time that catalytic hydrogenation is discussed (Section 8-5); biological methylations with *S*-adenosylmethionine are covered with S_N2 reactions (Section 12-11); and biological reductions with NADH are introduced along with laboratory $NaBH_4$ reductions (Section 13-3).

The third group of chapters (19–24) is unique to this text in its depth of coverage. These chapters deal exclusively with the main classes of biomolecules—amino acids and proteins, carbohydrates, lipids, and nucleic acids—and show how thoroughly organic chemistry permeates biological chemistry. Following an introduction to each class, major metabolic pathways for that class are discussed from the perspective of mechanistic organic chemistry.

And finally, for those faculty who want additional coverage of natural products, polymers, and pericyclic reactions, the book ends with a fourth group of chapters (25–27) devoted to those topics. This final group is available in both electronic and hard-copy formats at the request of the adopter.

What's New

Text content has been revised substantially for this 3rd edition as a result of user feedback. Most noticeably, two new chapters have been made available for those who want them: Chapter 26 on Pericyclic Reactions and Chapter 27 on Synthetic Polymers. Other changes include:

- Every chapter ends with a brief *Something Extra* essay that has been repositioned to follow immediately after the last text section where it is more likely to be noticed and read.
- The problems at the ends of chapters are now organized by topic to make it easier for students to find questions on specific subjects.
- New problems have been added in every chapter, 164 in all.
- Text references to all numbered **FIGURES** and **TABLES** are called out in color to help students move more easily between text and art.
- All figure captions have a boldfaced title, and the captions themselves use colored text to make it easier to focus on specific features in the figure art.

New topics in this 3rd edition include:

- A new *Something Extra*, "Organic Foods: Risk versus Benefit," in Chapter 1
- A new *Something Extra*, "Alkaloids: From Cocaine to Dental Anesthetics," in Chapter 2
- New coverage of bridged bicyclic molecules in Section 4-9
- New coverage of mercury-catalyzed alkyne hydration in Section 8-15
- New coverage of aromatic fluorination and fluorinated drugs in Section 9-6
- New coverage of alcohol to alkyl fluoride conversions in Section 12-3
- A new Section 12-5, "Organometallic Coupling Reactions," covering both organocopper reactions and the palladium-catalyzed Suzuki–Miyaura reaction

- A new *Something Extra*, "Naturally Occurring Organohalides," in Chapter 12
- New coverage of epoxide cleavage by nucleophiles in Section 13-10
- A new Section 13-11, "Crown Ethers and Ionophores"
- New coverage of hydrates of α-keto acids in Section 14-5
- A new *Something Extra*, "Barbiturates," in Chapter 17
- Threonine catabolism deleted from Section 20-4
- New coverage of Kiliani–Fischer carbohydrate chain extension and Wohl degradation in Section 21-6
- A new Section 23-7, "Prostaglandins and Other Eicosanoids"
- A new *Something Extra*, "Statin Drugs," in Chapter 23
- A new electronic Chapter 26, "Orbitals and Organic Chemistry: Pericyclic Reactions"
- A new electronic Chapter 27, "Synthetic Polymers"

I believe that there is more than enough standard organic chemistry in this book, and that the coverage of biological chemistry far surpasses that found in any other text. My hope is that all the students we teach, including those who worry about medical school, will come to agree that there is also logic and beauty here.

Features

Reaction Mechanisms

The innovative vertical presentation of reaction mechanisms that has become a hallmark of all my texts is retained in *Organic Chemistry with Biological Applications*, third edition. Mechanisms in this format have the reaction steps printed vertically, while the changes taking place in each step are explained next to the reaction arrows. With this format, students can see what is occurring at each step in a reaction without having to jump back and forth between structures and text. See Figure 14.4 for a chemical example and Figure 22.8 for a biological example.

Visualization of Biological Reactions

One of the most important goals of this book is to demystify biological chemistry—to show students how the mechanisms of biological reactions are the same as those of laboratory organic reactions. Toward this end, and to let students visualize more easily the changes that occur during reactions of large biomolecules, I use an innovative method for focusing attention on the reacting parts in large molecules by "ghosting" the nonreacting parts. See Figure 13.4 for an example.

Other Features

- "Why do we have to learn this?" I've been asked this question by students so many times that I thought I should answer it in writing. Thus, every chapter begins with a short introduction called "Why This Chapter?" that provides an up-front answer to the question, explaining why the material about to be covered is important and how the organic chemistry in each chapter relates to biological chemistry.

- Worked Examples in each chapter are titled to give students a frame of reference. Each Worked Example includes a Strategy and worked-out Solution, followed by Problems for students to try on their own.
- A *Something Extra* is provided in each chapter following the final text section to relate real-world concepts to students' lives. New topics in this edition include Organic Foods: Risk versus Benefit (Chapter 1), Alkaloids: From Cocaine to Dental Anesthetics (Chapter 2), Naturally Occurring Organohalides (Chapter 12), Barbiturates (Chapter 17), and Statin Drugs (Chapter 23).
- Visualizing Chemistry problems at the end of each chapter offer students an opportunity to see chemistry in a different way by visualizing whole molecules rather than simply interpreting structural formulas.
- The Summary and Key Word list at the end of each chapter helps students focus on the key concepts in that chapter.
- The Summary of Reactions at the end of specific chapters brings together the key reactions from those chapters into a single complete list.
- An overview entitled "A Preview of Carbonyl Chemistry" following Chapter 13 highlights the idea that studying organic chemistry involves both summarizing past ideas and looking ahead to new ones.
- Current IUPAC nomenclature rules are used in this text. Recognizing that these rules have not been universally adopted in the United States, the small differences between new and old rules are also discussed.

Alternate Edition

Hybrid version with access (24 months) to OWLv2 with MindTap Reader ISBN: 978-1-285-86784-7

A briefer, paperbound version of *Organic Chemistry with Biological Applications,* third edition, does not contain the end-of-chapter problems, which can be assigned in OWL, the online homework and learning system for this book. Access to OWLv2 and MindTap Reader eBook is included with the Hybrid version. MindTap Reader is the full version of the text, with all end-of-chapter questions and problem sets.

Supporting Materials for Students and Instructors

Please visit **www.cengage.com/chemistry/mcmurry/ocba3e** for information about student and instructor resources for this text.

Acknowledgments

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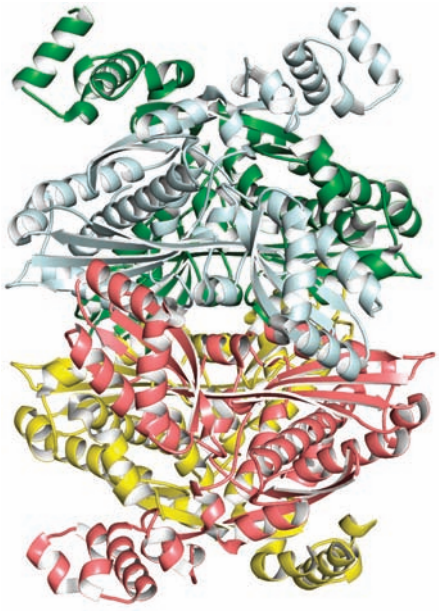
Structure and Bonding

1



A model of the enzyme HMG-CoA reductase, which catalyzes a crucial step in the body's synthesis of cholesterol.

CONTENTS

WHY THIS CHAPTER? We'll ease into the study of organic chemistry by first reviewing some ideas about atoms, bonds, and molecular geometry that you may recall from your general chemistry course. Much of the material in this chapter and the next is likely to be familiar to you, but it's nevertheless a good idea to make sure you understand it before going on.

A scientific revolution is now taking place—a revolution that will give us safer and more effective medicines, cure our genetic diseases, increase our life spans, and improve the quality of our lives. The revolution is based in understanding the structure, regulation, and function of the approximately 21,000 genes in the human body, and it relies on organic chemistry as the enabling science. It is our fundamental chemical understanding of biological processes at the molecular level that has made the revolution possible and that continues to drive it. Anyone who wants to understand or be a part of the remarkable advances now occurring in medicine and the biological sciences must first understand organic chemistry.

As an example of how organic and biological chemistry together are affecting modern medicine, look at coronary heart disease—the buildup of cholesterol-containing plaques on the walls of arteries, leading to restricted blood flow and eventual heart attack. Coronary heart disease is the leading cause of death for both men and women older than age 20, and it's estimated that up to one-third of women and one-half of men will develop the disease at some point in their lives.

The onset of coronary heart disease is directly correlated with blood cholesterol levels, and the first step in disease prevention is to lower those levels. It turns out that only about 25% of our blood cholesterol comes from what we eat; the remaining 75% (about 1000 mg each day) is made, or *biosynthesized,* by our bodies from dietary fats and carbohydrates. Thus, any effective plan for

lowering our cholesterol level means limiting the amount that our bodies biosynthesize, which in turn means understanding and controlling the chemical reactions that make up the metabolic pathway for cholesterol biosynthesis.

Now look at **FIGURE 1.1**. Although the figure probably looks unintelligible at this point, don't worry; before long it will make perfectly good sense. What's shown in Figure 1.1 is the biological conversion of a compound called 3-hydroxy-3-methylglutaryl coenzyme A (HMG-CoA) to mevalonate, a crucial step in the pathway by which our bodies synthesize cholesterol. Also shown in the figure is an X-ray crystal structure of the active site in the HMG-CoA reductase enzyme that catalyzes the reaction, along with a molecule of the drug atorvastatin (sold under the trade name Lipitor), which binds to the enzyme and stops it from functioning. With the enzyme thus inactivated, cholesterol biosynthesis is prevented.

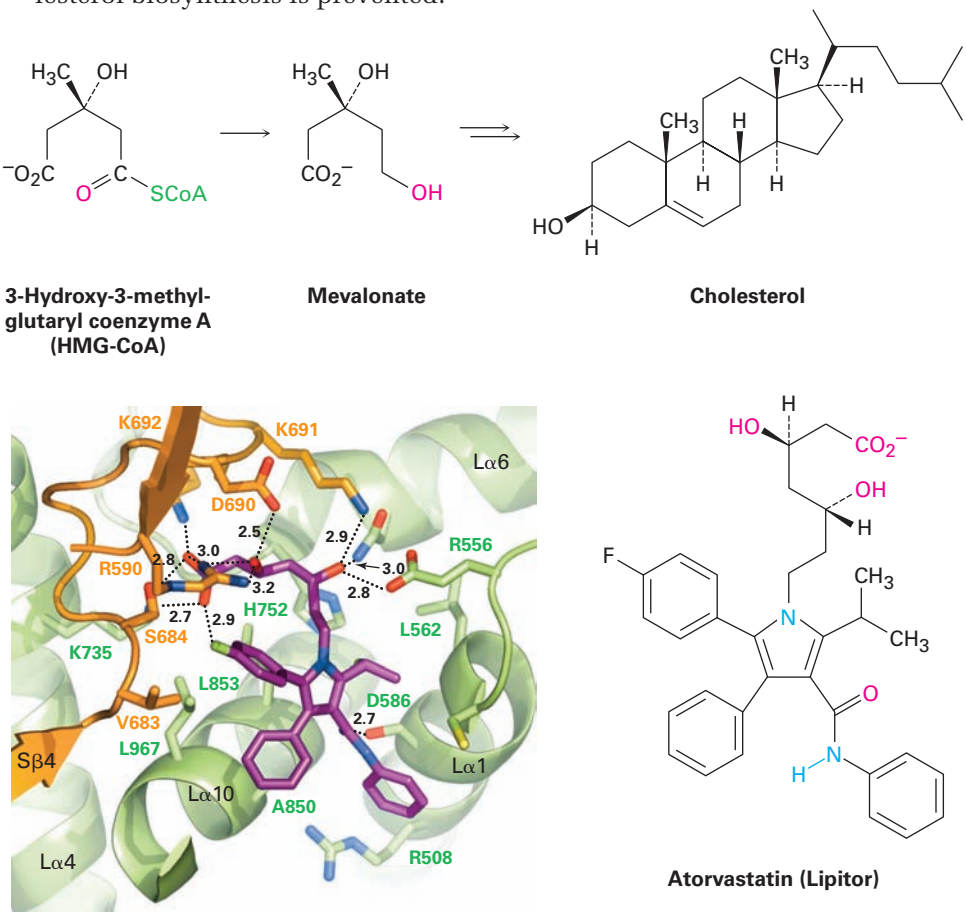


FIGURE 1.1 How does atorvastatin control cholesterol biosynthesis? The metabolic conversion of 3-hydroxy-3-methylglutaryl coenzyme A (HMG-CoA) to mevalonate is a crucial step in the body's pathway for biosynthesizing cholesterol. An X-ray crystal structure of the active site in the HMG-CoA reductase enzyme that catalyzes the reaction is shown, along with a molecule of atorvastatin (Lipitor) that is bound in the active site and stops the enzyme from functioning. With the enzyme thus inactivated, cholesterol biosynthesis is prevented.

Atorvastatin is one of a widely prescribed class of drugs called *statins,* which reduce a person's risk of coronary heart disease by lowering the level of cholesterol in his or her blood. Taken together, the statins—atorvastatin (Lipitor), simvastatin (Zocor), rosuvastatin (Crestor), pravastatin (Pravachol), lovastatin (Mevacor), and several others—are the most widely prescribed drugs in the world, with global sales of $29 billion annually.

The statins function by blocking the HMG-CoA reductase enzyme and preventing it from converting HMG-CoA to mevalonate, thereby limiting the body's biosynthesis of cholesterol. As a result, blood cholesterol levels drop and coronary heart disease becomes less likely. It sounds simple, but it would

be impossible without detailed knowledge of the steps in the pathway for cholesterol biosynthesis, the enzymes that catalyze those steps, and how precisely shaped organic molecules can be designed to block those steps. Organic chemistry is what makes it all happen.

Historically, the term **organic chemistry** dates to the late 1700s, when it was used to mean the chemistry of compounds found in living organisms. Little was known about chemistry at that time, and the behavior of the "organic" substances isolated from plants and animals seemed different from that of the "inorganic" substances found in minerals. Organic compounds were generally low-melting solids and were usually more difficult to isolate, purify, and work with than high-melting inorganic compounds. By the mid-1800s, however, it was clear that there was no fundamental difference between organic and inorganic compounds: the same principles explain the behaviors of all substances, regardless of origin or complexity. The only distinguishing characteristic of organic chemicals is that *all contain the element carbon*.

But why is carbon special? Why, of the more than 70 million presently known chemical compounds, do more than 99% of them contain carbon? The answers to these questions come from carbon's electronic structure and its consequent position in the periodic table **(FIGURE 1.2)**. As a group 4A element, carbon can share four valence electrons and form four strong covalent bonds. Furthermore, carbon atoms can bond to one another, forming long chains and rings. Carbon, alone of all elements, is able to form an immense diversity of compounds, from the simple to the staggeringly complex—from methane, with one carbon atom, to DNA, which can have more than *100 million* carbons.

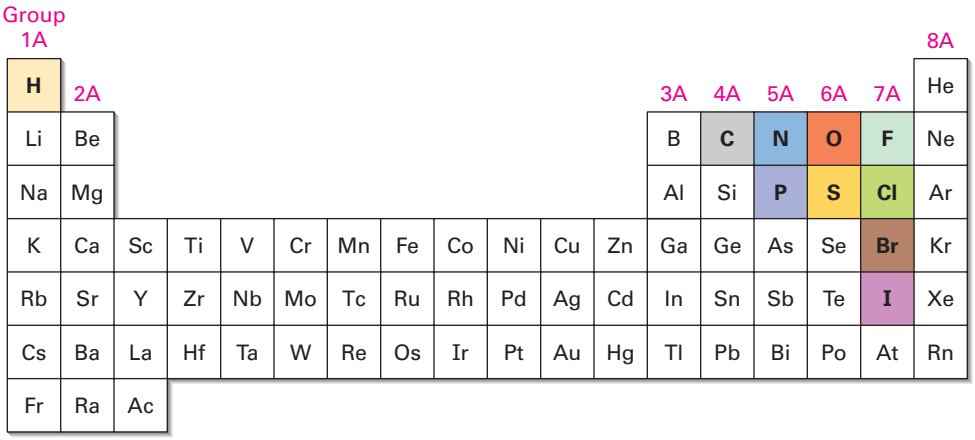


FIGURE 1.2 Elements commonly found in organic compounds. Carbon, hydrogen, and other elements commonly found in organic compounds are shown in the colors typically used to represent them.

Not all carbon compounds are derived from living organisms of course. Modern chemists have developed a remarkably sophisticated ability to design and synthesize new organic compounds in the laboratory—medicines, dyes, polymers, and a host of other substances. Organic chemistry touches the lives of everyone; its study can be a fascinating undertaking.

1-1 Atomic Structure: The Nucleus

As you might remember from your general chemistry course, an atom consists of a dense, positively charged *nucleus* surrounded at a relatively large distance by negatively charged *electrons* **(FIGURE 1.3)**. The nucleus consists

of subatomic particles called *neutrons,* which are electrically neutral, and *protons,* which are positively charged. Because an atom is neutral overall, the number of positive protons in the nucleus and the number of negative electrons surrounding the nucleus are the same.

Although extremely small—about 10^{-14} to 10^{-15} meter (m) in diameter—the nucleus nevertheless contains essentially all the mass of the atom. Electrons have negligible mass and circulate around the nucleus at a distance of approximately 10^{-10} m. Thus, the diameter of a typical atom is about 2×10^{-10} m, or 200 picometers (pm), where 1 pm = 10^{-12} m. To give you an idea of how small this is, a thin pencil line is about 3 million carbon atoms wide. (Many organic chemists and biochemists, particularly those in the United States, still use the unit *angstrom* (Å) to express atomic distances, where 1 Å = 100 pm = 10^{-10} m, but we'll stay with the SI unit picometer in this book.)

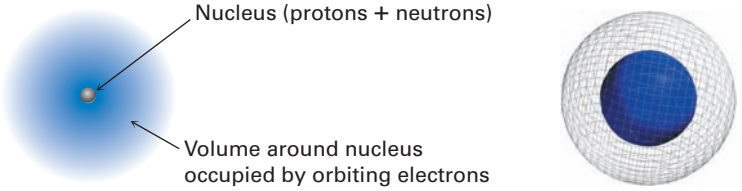


FIGURE 1.3 Schematic view of an atom. The dense, positively charged nucleus contains most of the atom's mass and is surrounded by negatively charged electrons. The three-dimensional view on the right shows calculated electron-density surfaces. Electron density increases steadily toward the nucleus and is 40 times greater at the **blue solid surface** than at the **gray mesh surface**.

A specific atom is described by its *atomic number (Z),* which gives the number of protons (or electrons) it contains, and its *mass number (A),* which gives the total number of protons plus neutrons in its nucleus. All the atoms of a given element have the same atomic number—1 for hydrogen, 6 for carbon, 15 for phosphorus, and so on—but they can have different mass numbers depending on how many neutrons they contain. Atoms with the same atomic number but different mass numbers are called **isotopes**. The weighted average mass in unified atomic mass units (u) of an element's naturally occurring isotopes is called the element's *atomic weight*—1.008 u for hydrogen, 12.011 u for carbon, 30.974 u for phosphorus, and so on.

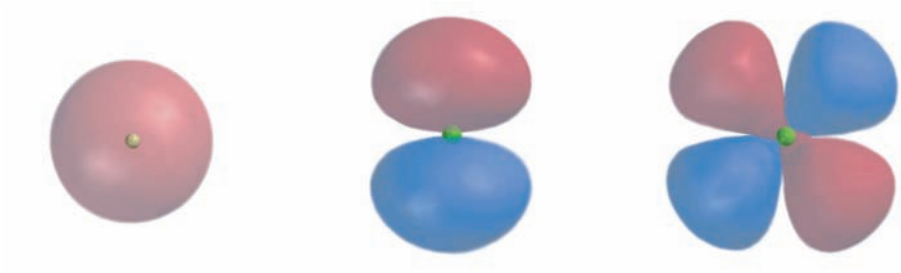
1-2 Atomic Structure: Orbitals

How are the electrons distributed in an atom? According to the quantum mechanical model, the behavior of a specific electron in an atom can be described by a mathematical expression called a *wave equation*—the same sort of expression used to describe the motion of waves in a fluid. The solution to a wave equation is called a *wave function,* or **orbital**, and is denoted by the Greek letter psi, ψ.

When the square of the wave function, ψ^2, is plotted in three-dimensional space, an orbital describes the volume of space around a nucleus that an electron is most likely to occupy. You might therefore think of an orbital as looking like a photograph of the electron taken at a slow shutter speed. In such a photo, the orbital would appear as a blurry cloud, indicating the region of space around the nucleus where the electron has been. This electron cloud doesn't have a sharp boundary, but for practical purposes we can set the limits

by saying that an orbital represents the space where an electron spends 90% to 95% of its time.

What do orbitals look like? There are four different kinds of orbitals, denoted *s*, *p*, *d*, and *f*, each with a different shape. Of the four, we'll be concerned primarily with *s* and *p* orbitals because these are the most common in organic and biological chemistry. An *s* orbital is spherical, with the nucleus at its center; a *p* orbital is dumbbell-shaped; and four of the five *d* orbitals are cloverleaf-shaped, as shown in **FIGURE 1.4**. The fifth *d* orbital is shaped like an elongated dumbbell with a doughnut around its middle.



An *s* orbital **A *p* orbital** **A *d* orbital**

FIGURE 1.4 Representations of *s*, *p*, and *d* orbitals. An *s* orbital is spherical, a *p* orbital is dumbbell-shaped, and four of the five *d* orbitals are cloverleaf-shaped. **Different lobes** of *p* orbitals are often drawn for convenience as teardrops, but their true shape is more like that of a doorknob, as indicated.

The orbitals in an atom are organized into different layers around the nucleus, or **electron shells**, of successively larger size and energy. Different shells contain different numbers and kinds of orbitals, and each orbital within a shell can be occupied by a maximum of two electrons. The first shell contains only a single *s* orbital, denoted 1*s*, and thus holds only 2 electrons. The second shell contains one 2*s* orbital and three 2*p* orbitals and thus holds a total of 8 electrons. The third shell contains a 3*s* orbital, three 3*p* orbitals, and five 3*d* orbitals, for a total capacity of 18 electrons. These orbital groupings and their energy levels are shown in **FIGURE 1.5**.

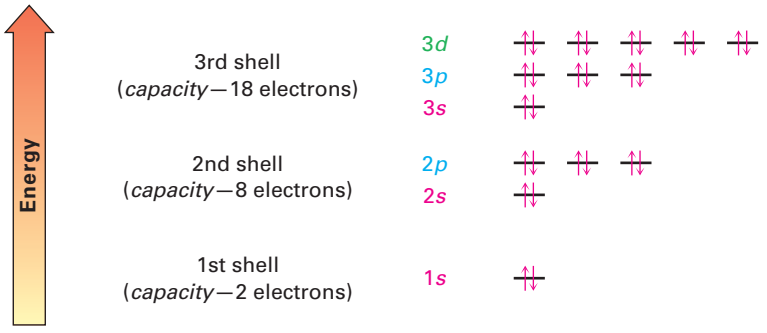
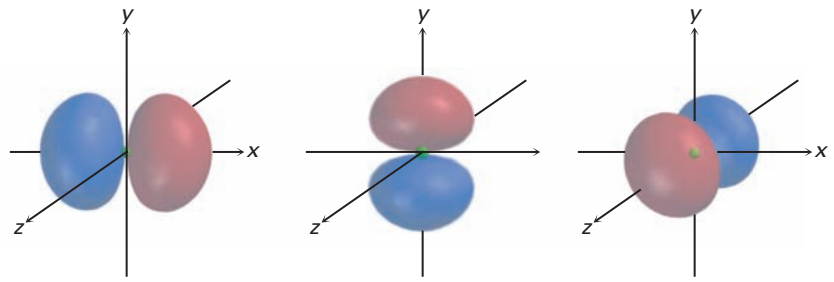


FIGURE 1.5 Energy levels of electrons in an atom. The first shell holds a maximum of 2 electrons in one **1*s*** orbital; the second shell holds a maximum of 8 electrons in one **2*s*** and three **2*p*** orbitals; the third shell holds a maximum of 18 electrons in one **3*s***, three **3*p***, and five **3*d*** orbitals; and so on. The two electrons in each orbital are represented by up and down arrows, ↑↓. Although not shown, the energy level of the 4*s* orbital falls between 3*p* and 3*d*.

The three different *p* orbitals within a given shell are oriented in space along mutually perpendicular directions, denoted p_x, p_y, and p_z. As shown in **FIGURE 1.6**, the two lobes of each *p* orbital are separated by a region of zero electron density called a **node**. Furthermore, the two orbital regions separated by the node have different algebraic signs, + and −, in the wave function, as represented by the different colors in Figure 1.6. As we'll see in Section 1-11, the algebraic signs of the different orbital lobes have important consequences with respect to chemical bonding and chemical reactivity.



A $2p_x$ orbital A $2p_y$ orbital A $2p_z$ orbital

FIGURE 1.6 Shapes of the 2*p* orbitals. Each of the three mutually perpendicular, dumbbell-shaped orbitals has two lobes separated by a node. The **two lobes** have different algebraic signs in the corresponding wave function, as indicated by the different colors.

1-3 Atomic Structure: Electron Configurations

The lowest-energy arrangement, or **ground-state electron configuration**, of an atom is a listing of the orbitals occupied by its electrons. We can predict this arrangement by following three rules:

Rule 1

The lowest-energy orbitals fill up first, according to the order $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d$, a statement called the *aufbau principle.* Note that the $4s$ orbital lies between the $3p$ and $3d$ orbitals in energy.

Rule 2

Electrons act in some ways as if they were spinning around an axis, in somewhat the same way that the earth spins. This spin can have two orientations, denoted as up (↑) and down (↓). Only two electrons can occupy an orbital, and they must be of opposite spin, a statement called the *Pauli exclusion principle*.

Rule 3

If two or more empty orbitals of equal energy are available, one electron occupies each with spins parallel until all orbitals are half-full, a statement called *Hund's rule*.

Some examples of how these rules apply are shown in TABLE 1.1. Hydrogen, for instance, has only one electron, which must occupy the lowest-energy orbital. Thus, hydrogen has a $1s$ ground-state configuration. Carbon has six electrons and the ground-state configuration $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1$, and so forth. Note that a superscript is used to represent the number of electrons in a particular orbital.

TABLE 1.1 Ground-State Electron Configurations of Some Elements

Element	Atomic number	Configuration	Element	Atomic number	Configuration
Hydrogen	1	1*s* ↑	Phosphorus	15	3*p* ↑ ↑ ↑
					3*s* ↑↓
Carbon	6	2*p* ↑ ↑ —			2*p* ↑↓ ↑↓ ↑↓
		2*s* ↑↓			2*s* ↑↓
		1*s* ↑↓			1*s* ↑↓

PROBLEM 1.1

Give the ground-state electron configuration for each of the following elements:
(a) Oxygen **(b)** Phosphorus **(c)** Sulfur

PROBLEM 1.2

How many electrons does each of the following biological trace elements have in its outermost electron shell?
(a) Magnesium **(b)** Cobalt **(c)** Selenium

1-4 Development of Chemical Bonding Theory

By the mid-1800s, the new science of chemistry was developing rapidly and chemists had begun to probe the forces holding atoms together in compounds. In 1858, August Kekulé and Archibald Couper independently proposed that, in all its compounds, carbon is *tetravalent*—it always forms four bonds when it joins other elements to form stable compounds. Furthermore, said Kekulé, carbon atoms can bond to one another to form extended chains of linked atoms.

Shortly after the tetravalent nature of carbon was proposed, extensions to the Kekulé–Couper theory were made when the possibility of *multiple* bonding between atoms was suggested. Emil Erlenmeyer proposed a carbon–carbon triple bond for acetylene, and Alexander Crum Brown proposed a carbon–carbon double bond for ethylene. In 1865, Kekulé provided another major advance when he suggested that carbon chains can double back on themselves to form *rings* of atoms.

Although Kekulé and Couper were correct in describing the tetravalent nature of carbon, chemistry was still viewed in a two-dimensional way until 1874. In that year, Jacobus van't Hoff and Joseph Le Bel added a third dimension to our ideas about organic compounds. They proposed that the four bonds of carbon are not oriented randomly but have specific spatial directions. Van't Hoff went even further and suggested that the four atoms to which carbon is bonded sit at the corners of a regular tetrahedron, with carbon in the center.

A representation of a tetrahedral carbon atom is shown in FIGURE 1.7. Note the conventions used to show three-dimensionality: solid lines represent bonds in the plane of the page, the heavy wedged line represents a bond coming out of the page toward the viewer, and the dashed line represents a bond receding back behind the page away from the viewer. These representations will be used throughout this text.

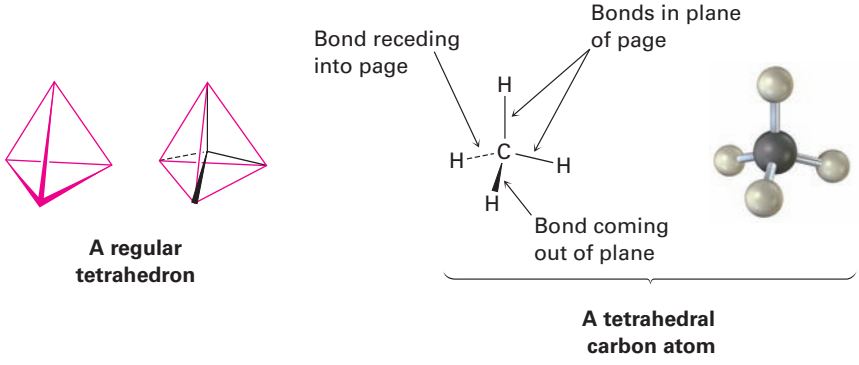


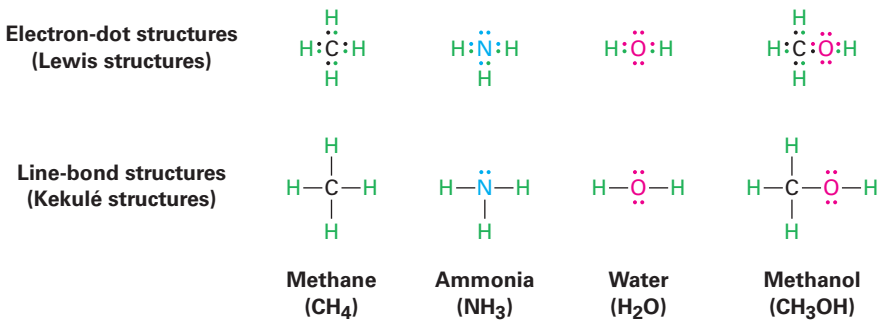
FIGURE 1.7 A representation of van't Hoff's tetrahedral carbon atom. The solid lines represent bonds in the plane of the paper, the heavy wedged line represents a bond coming out of the plane of the page, and the dashed line represents a bond going back behind the plane of the page.

Why, though, do atoms bond together, and how can bonds be described electronically? The *why* question is relatively easy to answer: atoms bond together because the compound that results is more stable and lower in energy than the separate atoms. Energy (usually as heat) is always released and flows *out of* the chemical system when a chemical bond forms. Conversely, energy must always be put *into* the system to break a chemical bond. Making bonds always releases energy, and breaking bonds always absorbs energy. The *how* question is more difficult. To answer it, we need to know more about the electronic properties of atoms.

We know through observation that eight electrons (an electron *octet*) in an atom's outermost shell, or **valence shell**, impart special stability to the noble-gas elements in group 8A of the periodic table: Ne (2 + *8*); Ar (2 + 8 + *8*); Kr (2 + 8 + 18 + *8*). We also know that the chemistry of main-group elements is governed by their tendency to take on the electron configuration of the nearest noble gas. The alkali metals in group 1A, for example, achieve a noble-gas configuration by losing the single *s* electron from their valence shell to form a cation, while the halogens in group 7A achieve a noble-gas configuration by gaining a *p* electron to fill their valence shell and form an anion. The resultant ions are held together in compounds like $Na^+ Cl^-$ by an electrostatic attraction of unlike charges that we call an **ionic bond**.

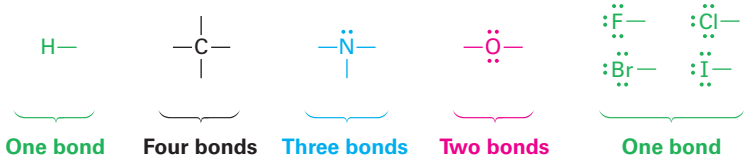
But how do elements closer to the middle of the periodic table form bonds? Look at methane, CH_4, the main constituent of natural gas, for example. The bonding in methane is not ionic because it would take too much energy for carbon ($1s^2\ 2s^2\ 2p^2$) to either gain or lose *four* electrons to achieve a noble-gas configuration. Instead, carbon bonds to other atoms, not by gaining or losing electrons, but by *sharing* them. Such a shared-electron bond, first proposed in 1916 by G. N. Lewis, is called a **covalent bond**. The neutral collection of atoms held together by covalent bonds is called a **molecule**.

A simple way of indicating the covalent bonds in molecules is to use what are called *Lewis structures,* or **electron-dot structures**, in which the valence-shell electrons of an atom are represented as dots. Thus, hydrogen has one dot representing its 1*s* electron, carbon has four dots ($2s^2\ 2p^2$), oxygen has six dots ($2s^2\ 2p^4$), and so on. A stable molecule results whenever a noble-gas configuration is achieved for all the atoms—eight dots (an octet) for main-group atoms or two dots for hydrogen. Simpler still is the use of *Kekulé structures,* or **line-bond structures**, in which two-electron covalent bonds are indicated as lines drawn between atoms.

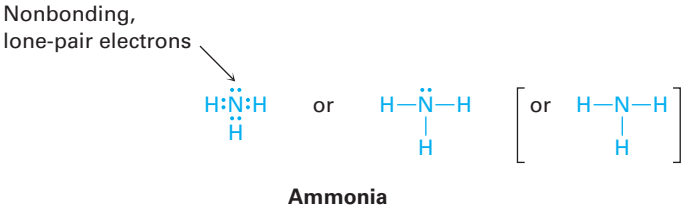


The number of covalent bonds an atom forms depends on how many additional valence electrons it needs to reach a noble-gas configuration. Hydrogen

has one valence electron ($1s$) and needs one more to reach the helium configuration ($1s^2$), so it forms one bond. Carbon has four valence electrons ($2s^2\ 2p^2$) and needs four more to reach the neon configuration ($2s^2\ 2p^6$), so it forms four bonds. Nitrogen has five valence electrons ($2s^2\ 2p^3$), needs three more, and forms three bonds; oxygen has six valence electrons ($2s^2\ 2p^4$), needs two more, and forms two bonds; and the halogens have seven valence electrons, need one more, and form one bond.



Valence electrons that are not used for bonding are called **lone-pair electrons**, or *nonbonding electrons.* The nitrogen atom in ammonia (NH_3), for instance, shares six valence electrons in three covalent bonds and has its remaining two valence electrons in a nonbonding lone pair. As a time-saving shorthand, nonbonding electrons are often omitted when drawing line-bond structures, but you still have to keep them in mind since they're often crucial in chemical reactions.



Ammonia

Predicting the Number of Bonds Formed by Atoms in Molecules — WORKED EXAMPLE 1.1

How many hydrogen atoms does phosphorus bond to in phosphine, $PH_?$?

Strategy

Identify the periodic group of phosphorus, and tell from that how many electrons (bonds) are needed to make an octet.

Solution

Phosphorus, like nitrogen, is in group 5A of the periodic table and has five valence electrons. It thus needs to share three more electrons to make an octet and therefore bonds to three hydrogen atoms, giving PH_3.

Drawing Electron-Dot and Line-Bond Structures — WORKED EXAMPLE 1.2

Draw both electron-dot and line-bond structures for chloromethane, CH_3Cl.

Strategy

Remember that a bond—that is, a pair of shared electrons—is represented as a line between atoms.