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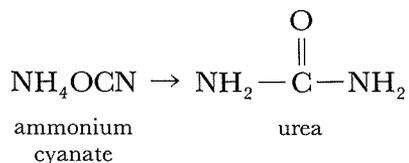
The Special Role of Carbon

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- Key Concepts**
- 21-1** The special talents of carbon: polymerization, cross-linking, multiple bonds, and delocalization.
 - 21-2** The chemistry of the neighbors of carbon: boron, nitrogen, silicon.
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Organic chemistry just now is enough to drive one mad. It gives me the impression of a primeval tropical forest, full of the most remarkable things, a monstrous and boundless thicket, with no way of escape, into which one may well dread to enter.

Friedrich Wöhler (1835)

The term *organic*, applied to chemistry in Friedrich Wöhler's time, signified "living." Organic chemistry then dealt with the various compounds that were associated with living organisms. Most chemists drew a sharp line between organic and inorganic compounds; they ascribed some special type of "life force" to the former. But in 1828 Wöhler demolished the idea of a vital force in the most direct way possible, by showing that an undoubtedly biological molecule, urea, could be obtained in the laboratory merely by heating an undoubtedly inorganic, nonbiological salt, ammonium cyanate:



The synthesis of other organic substances followed rapidly. The term *organic chemistry* gradually came to mean the chemistry of the compounds of carbon, so when chemists in the twentieth century wanted to talk specifically about processes of living organisms, they had to invent another word, *biochemistry*. It is true that the chemistry of life is a subclass of the chemistry of carbon compounds, and it is worth reflecting why this should be so.

Wöhler's complaint about how complicated organic chemistry was can be echoed today, but at an entirely different level of knowledge. We have answered many of the questions that baffled Wöhler, but now we are asking questions that he could never have raised. Charles Darwin once remarked, "It is the merest rubbish, asking about the origin of life; one might as well ask about the origin of matter." It is a measure of how far we have come that scientists today are busy asking both questions, and devising experiments to answer them. One of the goals of the Viking unmanned Martian landers was to search for life on that planet. The results, although not absolutely definitive, were so negative that few people expect future probes to reveal life on Mars. It is interesting that a tacit assumption was made in all the Viking experiments that life would be *carbon-based*. After the fact, no one suggests that this assumption was to be blamed for the negative results, or that Mars is teeming with silicon-based or nitrogen-based creatures. Carbon seems to have a special role in life, and the first part of this chapter will examine why this is so. The rest of the chapter will be an overview of two very large areas of chemistry—organic chemistry and biochemistry. Some of the chapters of this book deal with fundamental and essential techniques, and should be studied intensively. Others, such as this one, are designed to give you a general impression of an area of chemistry. As you read this chapter, try to understand and appreciate, rather than memorize.

21-1 THE SPECIAL TALENTS OF CARBON

The chemistry of carbon puts that of all other elements to shame. The American Chemical Society has maintained a register of chemical compounds mentioned in the literature since 1965. By the middle of 1978 there were $4\frac{1}{4}$ million different chemical substances in the register. Of these, 4 million were compounds based on a carbon backbone, and the remaining quarter million were about evenly divided between alloys and inorganic compounds. Hence, excluding metallic alloys, we know of 32 times as many organic compounds as inorganic! So many carbon compounds exist because carbon can link with itself as no other element can, to make straight chains and branched chains. Some of these compounds are shown in Figure 21-1. Chains made by the repetition of a subunit are called **polymers**, and the repeated unit is called a **monomer**.

Hydrocarbons are compounds containing only C and H atoms. The simplest hydrocarbons are linear polymers of the subunit $-\text{CH}_2-$, with the ends of the polymer terminated by hydrogen atoms. Other hydrocarbons have branched chains or rings of connected atoms. Butane is a tetramer (four subunits), and is a gas used for heating and cooking. Five- to 12-carbon polymers are gasolines; heptane (Figure 21-1) is one example. Kerosene is a mixture of molecules with 12 to 16 carbon atoms, and lubricating oils and paraffin wax are mixtures of chains with 17 and more carbons. Polyethylene

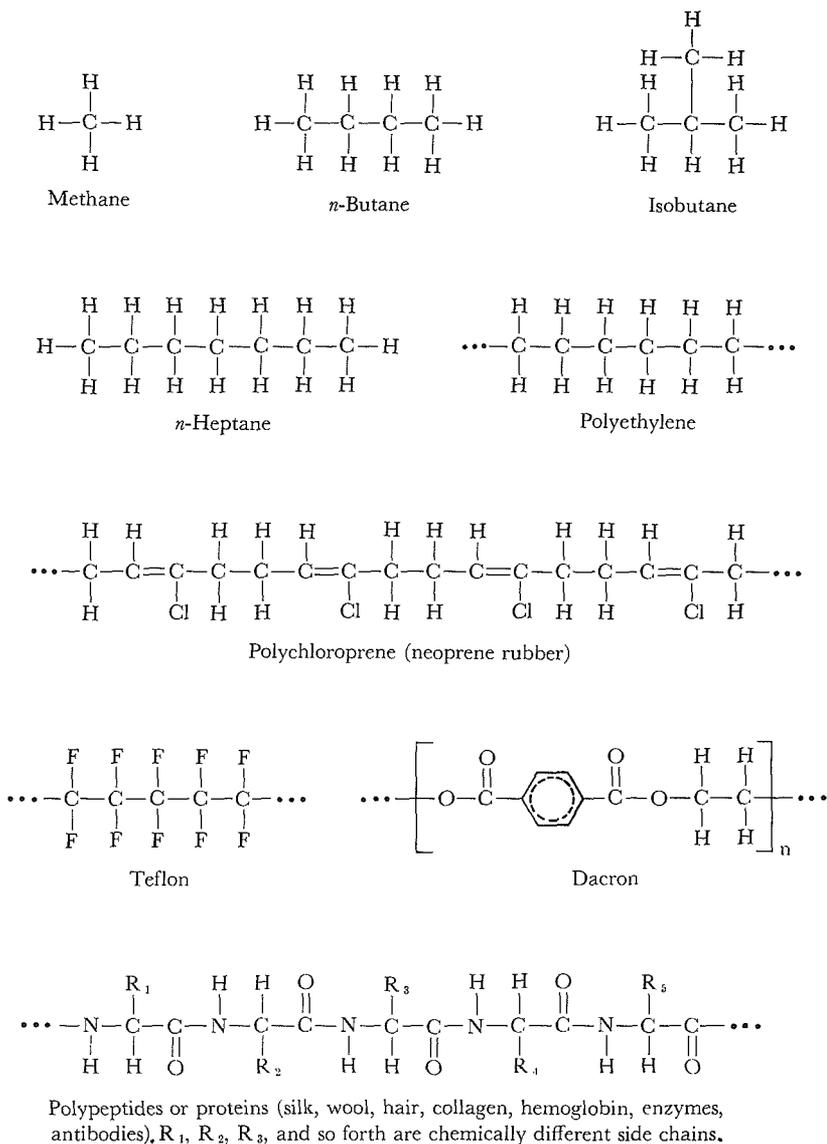


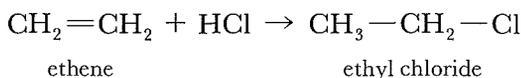
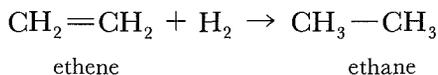
Figure 21-1

Natural and synthetic chains of carbon atoms, with nitrogen and oxygen. The first two rows are hydrocarbons of increasing chain length from methane, through the commercial heating gases (butane) and gasolines (heptane), to polyethylene plastic. The double bond at every fourth carbon connection in polychloroprene is typical of natural and synthetic rubbers. Dacron shows two kinds of multiple bonding: $\text{C}=\text{O}$ double bonds of the familiar π^b type, and delocalized benzenelike bonding. Polypeptide chains are cross-linked one to another as in Figure 21-2.

plastic has 5000 to 50,000 $\text{—CH}_2\text{—}$ monomer units per chain. There are many other organic chains, with more atoms than just C and H. Neoprene rubber, Teflon, and Dacron (Figure 21-1) are synthetic polymers, and the polypeptide chain shown at the bottom of Figure 21-1 is the polymer from which all proteins are built.

Because carbon can make as many as four bonds, branched and cross-linked chains can be built. Isobutane (Figure 21-1) is a branched-chain isomer of C_4H_{10} . Figure 21-2 shows silk and its synthetic analogue, nylon. Both are constructed from parallel, covalently bonded chains that are cross-linked into a sheet by hydrogen bonds. Bakelite and Melmac are hard, inflexible plastics because their monomers are covalently linked in three dimensions.

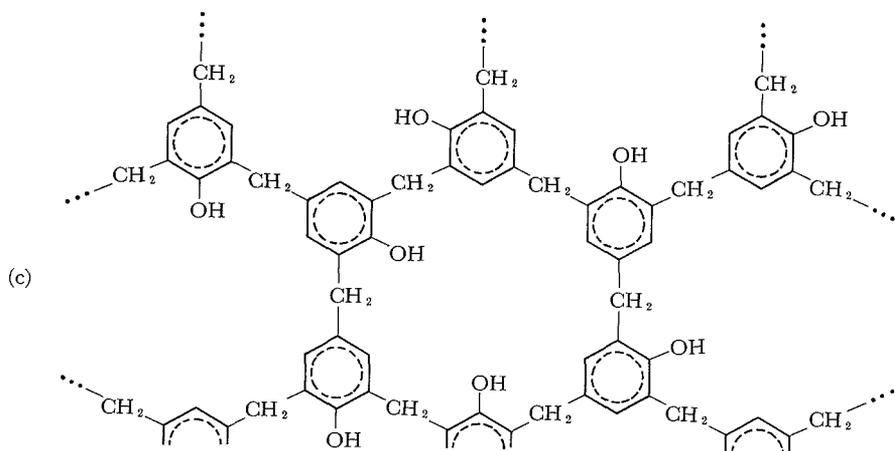
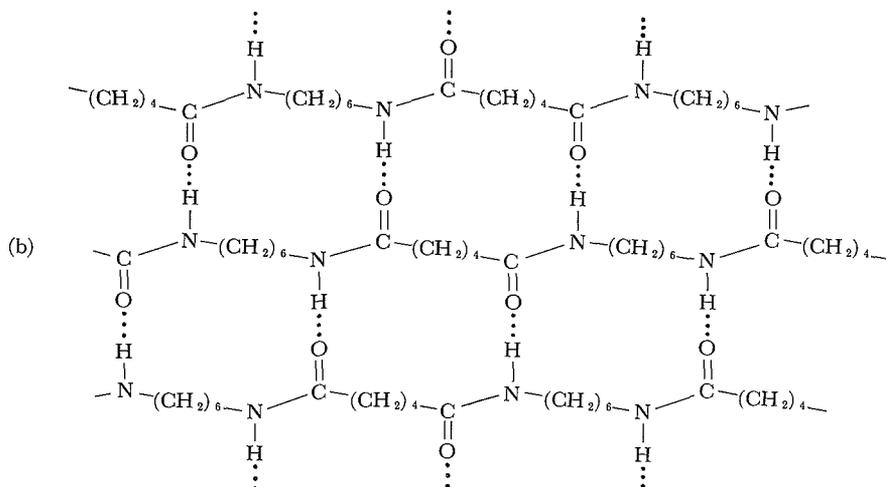
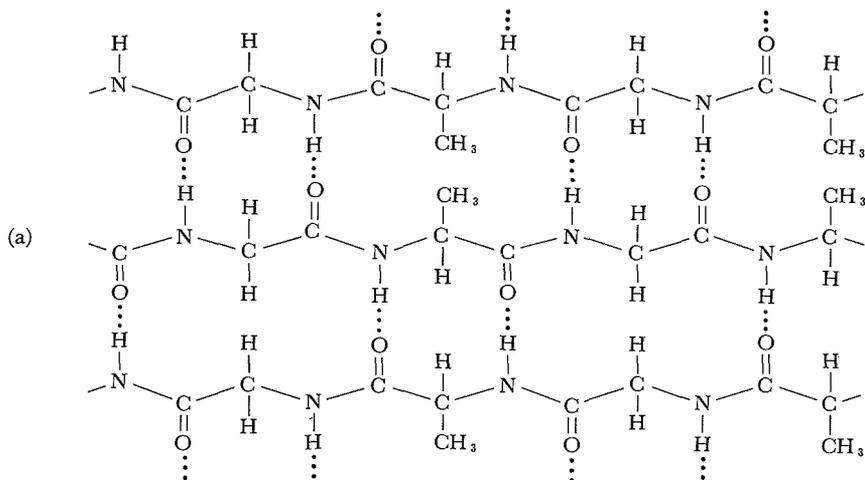
The other distinguishing feature of carbon is its ability to make double bonds with itself and with other elements, and to do so in the middle of these chains. Neoprene rubber (Figure 21-1) has such double bonds between carbon atoms. Dacron has double bonds between C and O, and it also has the delocalized multiple bonding that we saw in Chapter 13 for benzene. Figure 21-3 depicts some other examples of double bonds in carbon compounds. Since the double bond can often be converted to a single bond by adding an atom at each end of the bond, such double-bond compounds are called **unsaturated**:

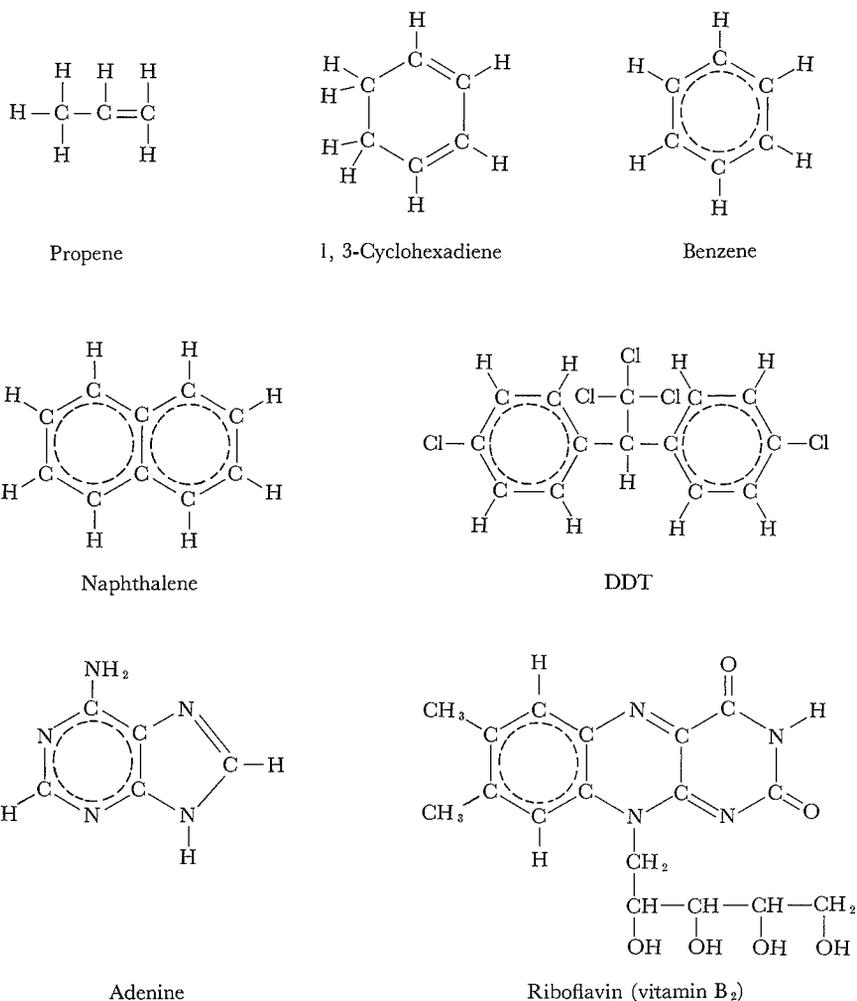


(*Ethene* is the systematic name for C_2H_4 ; its common name is *ethylene*.) Compounds with rings of atoms having delocalized, benzenelike multiple bonds

Figure 21-2

Three varieties of natural and synthetic polymers. (a) Silk, made from polypeptide chains. ► The chains are cross-linked into sheets by hydrogen bonds. (b) Nylon 66 is closely patterned after silk. It was invented, in 1935, by W. H. Carothers at E. I. du Pont de Nemours & Co., Inc. It has hydrogen bonding similar to silk, but at longer intervals down the chains. In both fibers, the fiber axis is horizontal in the figure and parallel to the covalently bonded chains. (c) Bakelite is one of the earliest synthetic plastics, having been invented, in 1909, by L. H. Baekeland, an American chemist who also contributed to the chemistry of photography. Bakelite is one member of a class of phenol–formaldehyde resins that are strong and hard because of their three-dimensional network of covalent bonds.



**Figure 21-3**

Examples of double bonds and delocalized bonds in organic compounds. Adenine, an essential component of the genetic polymer DNA (deoxyribonucleic acid) and of the energy-storing molecule ATP (adenosine triphosphate), is a pentamer of HCN. It has been prepared from HCN under conditions simulating those of earth in the early stages of the evolution of life. Dashed circles represent delocalized bonds of the type encountered for benzene in Chapter 13.

are called **aromatic** compounds. Dacron (Figure 21-1) and naphthalene, DDT, adenine, and riboflavin (Figure 21-3) all have aromatic components. Adenine and riboflavin also show that carbon can make double bonds to nitrogen, and that nitrogen can participate in a delocalized, aromatic ring.

Much of organic chemistry involves the special properties of aromatic ring systems. Aromatic molecules and transition-metal complexes are the two main classes of compounds in which the energy required to excite an electron falls in the visible part of the spectrum. Hence, these compounds are involved in dyes of all descriptions, and in mechanisms for trapping and transferring photon energy.

The four distinguishing features of organic compounds can be summarized as follows:

1. Long-chain polymers with C—C bonds
2. Branched and cross-linked chains
3. Double and triple bonds
4. Delocalized aromatic bonds

How many of these characteristics are exhibited by the immediate neighbors of carbon—B, N, and Si? What can carbon do that these elements cannot do, and why is this so? What particular combination of electrons and orbitals makes carbon so versatile?

21-2 THE CHEMISTRY OF THE NEIGHBORS OF CARBON

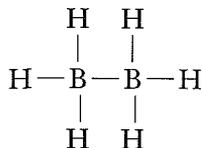
Boron, carbon, and nitrogen all are second-period elements of similar size. They differ in the number of valence electrons that they possess: three electrons in B, four in C, and five in N. Silicon, a third-period element, is like carbon in having four valence electrons, but they are one major energy level farther from the nucleus and have a principal quantum number of 3 instead of 2. Below the valence electrons, Si has 10 inner-orbital electrons: 2 with principal quantum number 1, and 8 with quantum number 2. In contrast, B, C, and N have only 2 electrons below their valence orbitals. All the differences in chemical properties among B, C, N, and Si that will concern us in this chapter come from these two factors: the number of valence electrons and the number of electrons in completed inner orbitals.

Boron has three valence electrons and four valence orbitals per atom. It commonly uses three orbitals in sp^2 hybridization in compounds such as BF_3 . Carbon has four valence electrons and four orbitals. Except when involved in multiple bonds, it uses sp^3 hybrid orbitals. Nitrogen has five electrons and four orbitals. It typically makes three bonds to other atoms in tetrahedral configurations, and the fourth sp^3 atomic orbital is occupied by the lone electron pair (Section 13-3). Both carbon and nitrogen can make double and triple bonds involving the π overlap bond discussed in Section 13-4. The bond length for both elements decreases by 13% in a double bond and 22% in a triple bond. The atoms are held more tightly because of the electrons in π^b molecular orbitals derived from overlapping $2p$ atomic orbitals. Conversely, the overlap of these orbitals is too small for significant

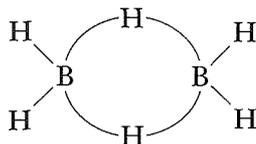
bonding unless the atoms are closer together. This is the reason why Si and other elements in the third period of the table and beyond cannot form multiple bonds. Silicon has 10 inner-orbital electrons instead of 2 as in C and N. The repulsion between these inner-orbital electrons does not permit two Si atoms to come close enough for p -orbital overlap and the formation of a double bond. Although chemists are actively trying to synthesize compounds with Si=Si and Si=C double bonds, none have yet been prepared. With one or two exceptions, double and triple bonds are confined to elements in the second period of the periodic table, with no more than two inner-orbital electrons per atom. Exceptions such as S=O, P=O, and Si=O use overlap between p and d orbitals, as we shall see later in this section with Si.

Boron

We can understand why boron is not a good candidate for carbonlike chemistry by looking at the series of boron hydrides. The hydride BH_3 exists only as a short-lived decomposition product of higher hydrides. Other hydrides are known: B_2H_6 , B_4H_{10} , B_5H_9 , B_5H_{11} , B_6H_{10} , B_6H_{12} , B_8H_{12} , B_9H_{15} , and $\text{B}_{10}\text{H}_{14}$. The simplest boron hydride, B_2H_6 , has eight atoms and only 12 valence electrons. If it were to have an ethanelike structure,



it would need 14 electrons for the seven covalent bonds. But diborane has only 12 valence electrons: It is an *electron-deficient* compound. Its true structure is



Each B atom has two normal two-center covalent B—H bonds, which use a total of eight electrons. The remaining four electrons are used in two *three-center* B—H—B bonds, in which each of the three atoms contributes an orbital to the bonding molecular orbital. This concept of the three-center bond is enough to explain the structures of all boron hydrides. It also explains why boron cannot do the things that carbon can do.

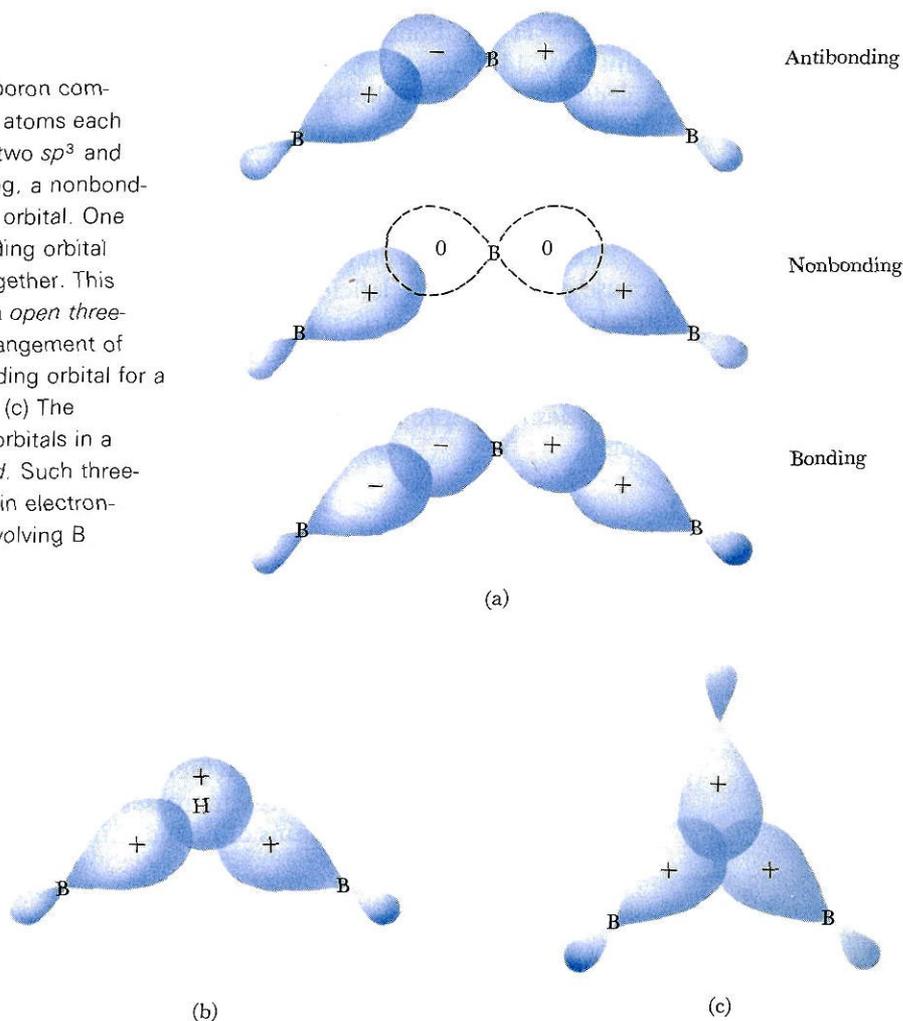
For most compounds in which the number of valence electrons is at least as great as the number of valence orbitals, the two-atom chemical bond model is sufficient, and we need consider only two atoms at a time. However,

as we learned in the discussion of benzene (Section 13-5), localized molecular orbitals are only an approximation to reality. Sometimes we must construct delocalized molecular orbitals from atomic orbitals contributed by several, or occasionally all, of the atoms in a molecule. In benzene the C—H and the C—C σ bonds can be dealt with individually, but the six p orbitals must be considered together.

To explain the behavior of boron, the smallest unit of bonding that we can consider is sometimes *three atoms*. Three atomic orbitals, one from each atom, can combine to make three molecular orbitals: one bonding, one antibonding, and one nonbonding—the last with virtually the same energy as the original atomic orbitals (Figure 21-4). We have seen nonbonding

Figure 21-4

Three-center orbitals in boron compounds. (a) Three boron atoms each can donate one orbital (two sp^3 and one p) to make a bonding, a nonbonding, and an antibonding orbital. One electron pair in the bonding orbital holds all three atoms together. This arrangement is called an *open three-center bond*. (b) The arrangement of atomic orbitals in a bonding orbital for a B—H—B bridge bond. (c) The arrangement of atomic orbitals in a *closed three-center bond*. Such three-center bonds are found in electron-deficient compounds involving B and Al.



orbitals before. In HF (Figures 12-11 and 12-12), the $2p_x$ and $2p_y$ orbitals of F are nonbonding, as are the d_{xy} , d_{yz} , and d_{xz} orbitals of the metal in an octahedral coordination complex.

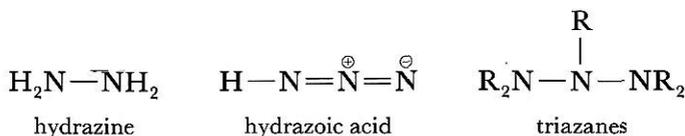
Two electrons in one of these *bonding* three-center molecular orbitals can hold three atoms together. This economy in bonding helps to compensate for the electron deficiency in boron. However, it also forces a cramped geometry on its compounds that makes boron unsuitable as a rival for carbon. Vast molecular networks can be constructed from the straight- and branched-chain carbon hydrides (hydrocarbons), in which the atoms are connected two at a time. In contrast, the boron hydrides, in which the atoms are connected three at a time, build structures whose boron frameworks are fragments of an icosahedron (Figure 21-5a). The hydride B_4H_{10} is a small fragment of the icosahedron (Figure 21-5b). It has six normal two-center bonds between B and H, one two-center B—B bond, and four three-center B—H—B bonds. Each of these bonds requires one electron pair. In this way, 14 atoms are held together by using 26 atomic orbitals but only 22 electrons. The hydride B_9H_{15} is three-fourths of a complete icosahedron (Figure 21-5c). In this compound, 24 atoms are held together with 51 atomic orbitals and only 42 bonding electrons. The complete B_{12} icosahedron is found in crystalline boron. The manner in which such three-center bonds are used in the larger boron hydrides is shown for B_5H_{11} in Figure 21-6.

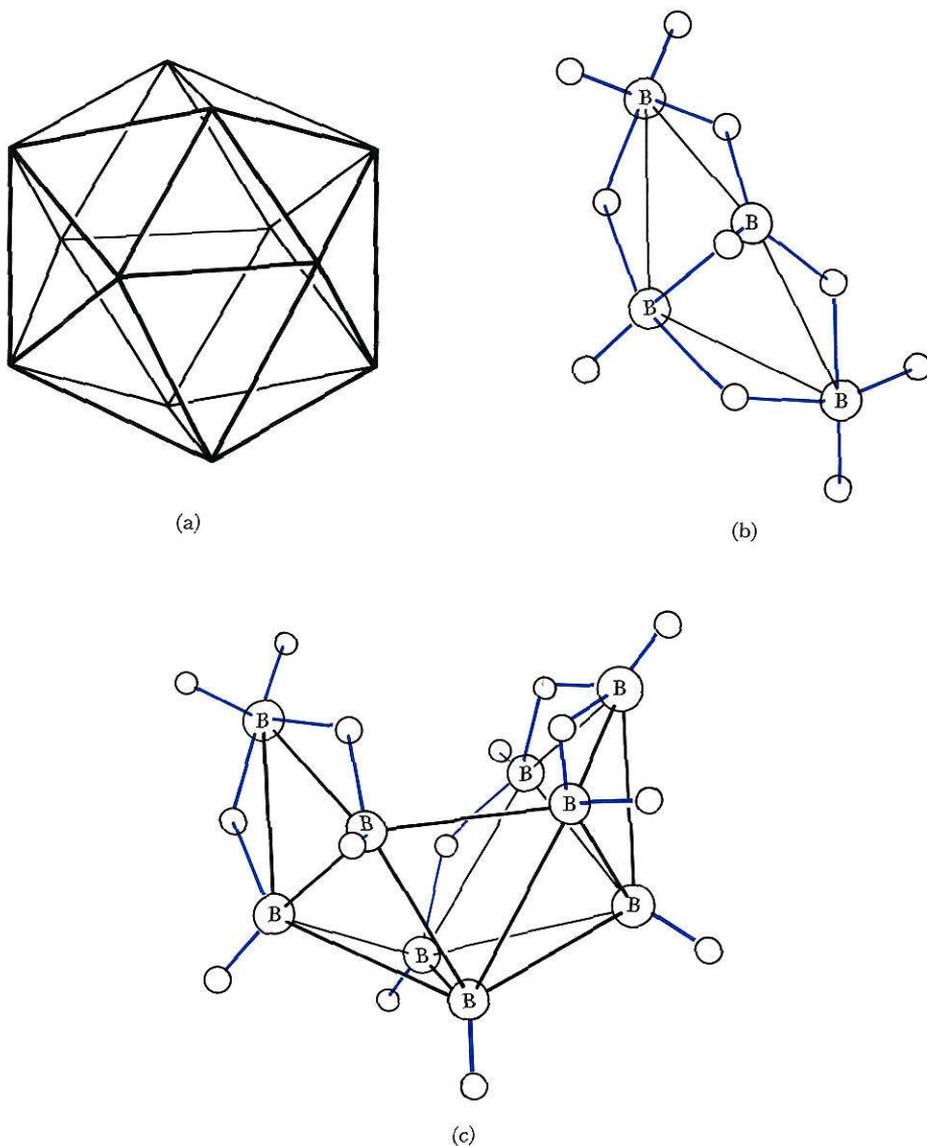
In conclusion, boron is an unsuitable candidate for organic chemistry because of its electron deficiency, which leads to three-center bonding and a tendency for boron structures to close in upon themselves. Even worse, the geometrical arrangement produced makes it impossible for p orbitals to lie parallel on adjacent atoms and to form π bonds. In terms of approaching the desirable properties of carbon, boron comes close, but not close enough.

Nitrogen

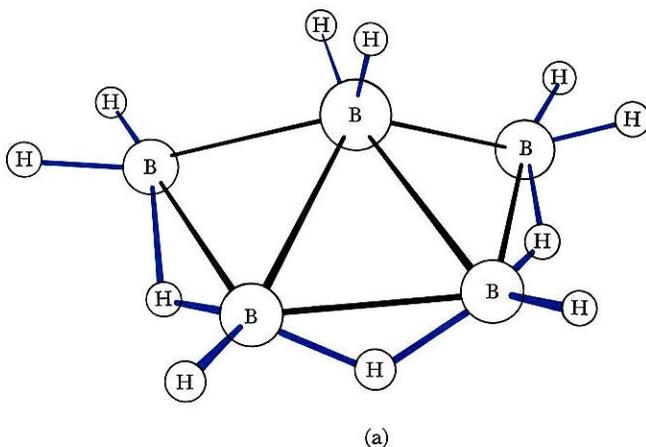
Nitrogen, like carbon, can make double and triple bonds to itself and to other first- and second-period atoms. But nitrogen suffers from a defect opposite to that of boron; it has too many electrons. Repulsions between lone electron pairs on neighboring nitrogen atoms make the N—N single-bond energy only 161 kJ mole⁻¹, in comparison with 348 kJ mole⁻¹ for a C—C bond. In the C—N bond, in which one of these repelling lone electron pairs is absent, the bond energy increases to 292 kJ mole⁻¹.

Some compounds with chains of linked nitrogen atoms exist:

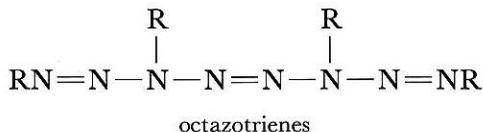
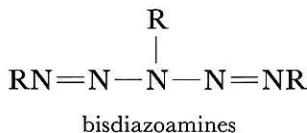
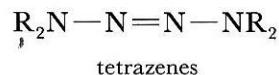
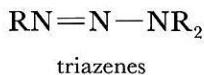
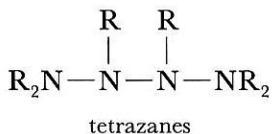
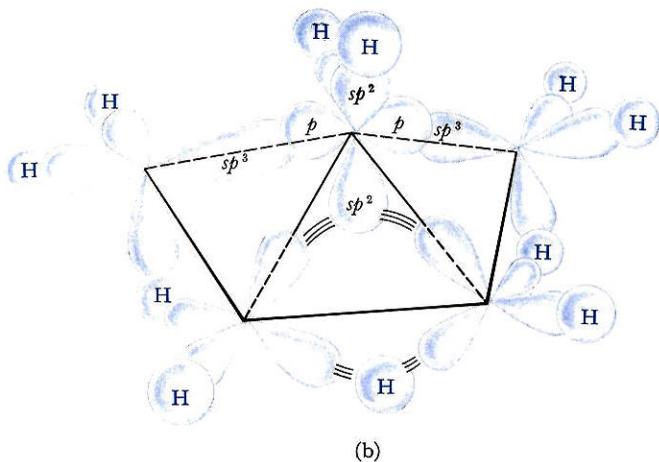


**Figure 21-5**

(a) The icosahedron is the boron framework for almost all boron hydrides. An icosahedron has 12 vertices and 20 equilateral triangular faces. (b) Tetraborane-10, B_4H_{10} , has its four boron atoms outlining two faces of the icosahedron. Bonds are marked in color. Six of the hydrogen atoms make normal two-center covalent bonds to boron; the others participate in four B—H—B bridges. The two central boron atoms are joined by a conventional two-center bond. (c) Enneaborane-15, B_9H_{15} , has a framework that is derived from the icosahedron by removing any three adjacent boron vertices that do not form an equilateral triangle. Ten hydrogen atoms make two-center covalent bonds to the boron atoms; the other five hydrogen atoms participate in B—H—B bridges.

**Figure 21-6**

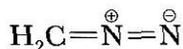
Structure and bonding orbitals in pentaborane-11, B_5H_{11} . Each of the boron atoms has sp^3 hybridization except the central one, which has one unhybridized p orbital and three sp^2 orbitals. The closed three-center bond uses two sp^3 orbitals and one sp^2 orbital. The open three-center bond involving the central B atom uses two sp^3 and one p , as in Figure 21-4a. The entire molecule uses 31 atomic orbitals but only 26 electrons.



Hydrazine is used as a rocket fuel. Hydrazoic acid is extremely explosive and toxic. It is sometimes used in detonators for explosives. The higher *hydronitrogens*, as these compounds are called by analogy with the hydrocarbons, can seldom be prepared in the simplest forms, with hydrogen atoms for the R's shown in the preceding structures. Those sufficiently stable even to exist have phenyl groups (benzene rings) for R, or methyl or ethyl groups (CH_3- or CH_3CH_2-). They are extremely unstable, and most are explosively so. They decompose rapidly under all conditions. Or as one scientist has said, "They stand on the edge of existence."

An important factor in the instability of nitrogen chains is the unusual stability of the triple bond in the $\text{N}\equiv\text{N}$ molecule. The N_2 triple bond, whose bond energy is 946 kJ mole^{-1} , is *six times* as strong as the $\text{N}-\text{N}$ single bond, whereas the $\text{C}\equiv\text{C}$ triple bond in acetylene is only 2.3 times the strength of the $\text{C}-\text{C}$ single bond. A long nitrogen chain is far less stable than the system remaining after the chain breaks into a series of N_2 molecules.

Nitrogen participates in chains and rings with carbon and, like carbon, forms double bonds. Diazomethane,



is one of the most versatile and useful reagents in organic chemistry, despite the fact that it is highly toxic, dangerously explosive, and cannot be stored without decomposition. A molecule with two or more adjacent nitrogen atoms is rarely stable.

Silicon

The critical difference between Si and C is the greater number of inner-orbital electrons in Si, and the consequent inability of two silicon atoms to come close enough together for double and triple bonds. Silicon forms *silanes* analogous to the alkane hydrocarbons to be discussed in Section 21-3. Silanes have the general formula $\text{Si}_n\text{H}_{2n+2}$. The longest of these chains that has been prepared is only hexasilane (Figure 21-7). These silanes, like the hydronitrogens, are dangerously reactive. The smallest silanes are stable in a vacuum, but all are spontaneously inflammable in air, and all react explosively with halogens. They are powerful reducing agents.

The silanes are so unstable and susceptible to oxidation because the $\text{Si}-\text{O}$ bond is much more stable than the $\text{Si}-\text{Si}$ bond: 369 kJ mole^{-1} versus 177 kJ mole^{-1} . In contrast, with carbon the $\text{C}-\text{O}$ and $\text{C}-\text{C}$ bond energies are almost the same: 351 kJ mole^{-1} and 348 kJ mole^{-1} (Table 21-1). Hydrocarbons are oxidized much less easily than are the silanes. Although the reaction

Table 21-1

Relative Bonding Abilities of Carbon and Its Neighbors

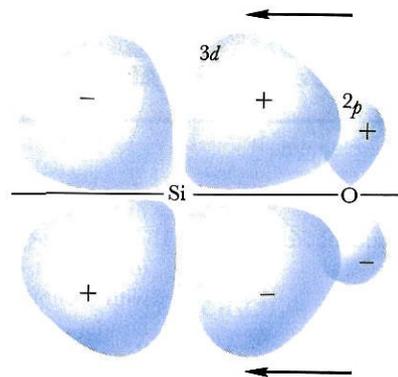
Element, R	B	C	N	Si
Valence electrons	3	4	5	4
Usual coordination	Threefold (sp^2) (or fourfold with 3-center bonds)	Fourfold (sp^3)	Fourfold (sp^3) (including lone pair)	Fourfold (sp^3)
Single bond energies (kJ mole ⁻¹)				
R—R		348	161	177
R—C		348	292	290
R—N		292	161	—
R—O		351	~230	369
R—H		413	391	295
Electron-to-orbital ratio and bonding behavior	Electron-deficient; 3-center bonds (in multiple-B compounds)	Electron match; 2-center bonds	Electron surfeit; 2-center bonds. Lone-pair repulsions	Electron match; 2-center bonds
Linkage of like atoms	Icosahedral shells	Extensive chains	Chains, limited extent	Chains, limited extent
Double and triple bonds, π bonding	π bonding impossible with 3-center bonds in icosahedral framework	Good π overlap in double and triple bonds. Entire bonding capacity <i>cannot</i> be satisfied with one other like atom. Builds networks	Good π overlap in double and triple bonds. Entire bonding capacity can be satisfied with one other like atom. Builds N ₂ molecules	Double and triple bonds impos- sible. Builds networks, princi- pally using the stable Si—O bond rather than the less stable Si—Si bond

must be ignited by heat, and continues under ordinary conditions only because the heat released by the reaction keeps the reactants at a high temperature.

Part of this difference in oxidation of C and Si compounds arises because the Si—Si bond is weaker than the C—C bond. This is to be expected from the greater size of Si. The bonding electrons are farther from each nucleus, and the bond is not as strong. The same effect gives Si a lower ionization energy than C and makes it less electronegative (Table 9-1). But an even more important factor in the difference in the behavior of C and Si

Figure 21-8

The strength of the Si—O bond is due to an unusual partial double-bond character. One of the filled $2p$ lone-pair orbitals of oxygen shares its electrons with an empty $3d$ orbital of Si that has similar energy. For this reason, the Si—O bond energy is 369 kJ mole^{-1} ; whereas the comparable silicon bond with C, which lacks the lone electron pairs, is only 290 kJ mole^{-1} .



is the anomalously high strength of the Si—O bond. In carbon, the empty $3d$ orbitals have a much higher energy than the filled lone-pair $2p$ orbitals of oxygen. There is no interaction between them. However, in silicon the added nuclear charge lowers the energy of the empty $3d$ atomic orbitals closer to the energy of the oxygen $2p$ orbitals. Oxygen can then share part of its lone-pair electrons with Si (Figure 21-8) in a back bonding similar to the $L \rightarrow M(\pi)$ and $M \rightarrow L(\pi)$ sharing in coordination complexes discussed in Section 20-3. Since the d_{xy} -type orbitals of Si extend farther toward O than an Si p orbital of a π bond, Si and O need not come as close as if they formed a $p\pi-p\pi$ double bond. The result of this sharing of oxygen lone electron pairs is that, although the Si—Si bond is 171 kJ mole^{-1} weaker than the C—C bond, the Si—O bond is 18 kJ mole^{-1} stronger than the C—O bond.

These results suggest that compounds in which Si atoms are linked by bridging oxygen atoms might be stable. This is so, and these compounds are the silicones. As shown in Figure 21-7, silicones can exist as straight chains, as rings, or as “ladder” compounds with two parallel linked chains. The silicones are extremely inert compounds. The silanes are much *more* reactive than the hydrocarbons; the silicones are much *less* reactive.

Comparison of Boron, Nitrogen, and Silicon

Each of the neighboring elements of carbon is unable to do the things that make carbon so important: to build long, stable chains with branching, cross-linking, and double bonds, and rings with delocalized electrons. The relative behavior of these elements is summarized in Table 21-1. Boron is forced into an unfavorable geometry by its deficiency of electrons and cannot overlap p orbitals to make double bonds. Although N can occasionally replace C in carbon rings and chains, and can form double bonds as easily as carbon can, long chains of nitrogen atoms are unstable. Silicon

is hampered by the weakness of its Si—Si bond in comparison with the Si—O bond and by its inability to make double bonds.

Carbon, then, is the fortunate combination of a small atom that has as many valence electrons as valence orbitals, and a bond to itself that is as strong as a bond to oxygen. Science fiction writers have long speculated on totally alien extraterrestrial life based on nonaqueous chemistry and an element other than carbon. Silicon has been the favorite element, and Mars has been the favorite homeland for rock-metabolizing, silicone-putty-fleshed monsters. But the more we learn about what carbon compounds do in terrestrial living creatures, the less easy it is to imagine silicon compounds performing even remotely similar roles. Carbon *is* special, and its properties can be duplicated by no other element.

21-3 SATURATED HYDROCARBONS OR ALKANES

Compounds that contain only carbon and hydrogen are, as we have seen, called **hydrocarbons**, and those in which all carbon atoms form four single bonds to other atoms are called **saturated hydrocarbons**, **paraffins**, or **alkanes**. The word *paraffin* originates from the Greek word for “little reactivity,” and the chemical properties of these paraffins are in marked contrast to those of the silanes and hydronitrogens.

Examples of several alkanes are given in Table 21-2. The first four have common names; those with 5 through 19 carbon atoms are commonly described by a Greek prefix indicating the total number of carbon atoms, and the standard suffix *-ane*. If there are more than 19 carbon atoms the chemical formula is usually employed as the name. Each carbon atom has four tetrahedrally oriented bonds either to another carbon atom or to a hydrogen.

The general chemical formula for the noncyclic alkanes is C_nH_{2n+2} . Alkanes exhibit a regular increase in melting point and boiling point with increasing molecular weight. Methane, ethane, propane, and butane are gases; pentane through $C_{20}H_{42}$ are liquids; and $C_{21}H_{44}$ and heavier compounds are waxy solids.

If there are four or more carbon atoms, there can be more than one way of connecting the carbons. Consequently, structural isomers can exist, in which the same numbers of each kind of atom are present, but the atoms are connected in different ways. The five isomers of hexane have the following carbon skeletons and systematic names (CH_3- is called the *methyl group*):

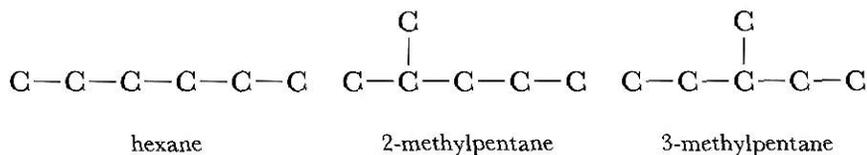
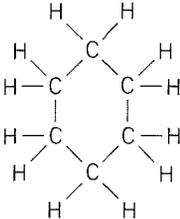
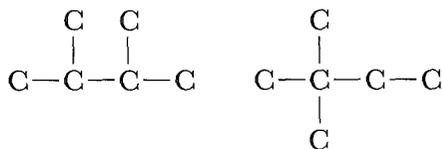


Table 21-2

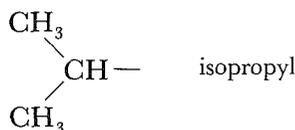
Some Common Saturated Hydrocarbons

Formula	Common name	Systematic name
CH_4	Methane	Methane
CH_3-CH_3	Ethane	Ethane
$\text{CH}_3-\text{CH}_2-\text{CH}_3$	Propane	Propane
$\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{CH}_3$	<i>n</i> -Butane	Butane
$\begin{array}{c} \text{CH}_3-\text{CH}-\text{CH}_3 \\ \\ \text{CH}_3 \end{array}$	Isobutane	Methylpropane
$\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{CH}_3$	<i>n</i> -Pentane	Pentane
$\begin{array}{c} \text{CH}_3-\text{CH}_2-\text{CH}-\text{CH}_3 \\ \\ \text{CH}_3 \end{array}$	Isopentane	Methylbutane
$\begin{array}{c} \text{CH}_3 \\ \\ \text{CH}_3-\text{C}-\text{CH}_3 \\ \\ \text{CH}_3 \end{array}$	Neopentane	Dimethylpropane
$\begin{array}{c} \text{CH}_3 \quad \quad \text{CH}_3 \\ \quad \quad \\ \text{CH}_3-\text{C}-\text{CH}_2-\text{CH}-\text{CH}_3 \\ \\ \text{CH}_3 \end{array}$	Isooctane	2,2,4-Trimethylpentane
	Cyclohexane	Cyclohexane



2,3-dimethylbutane 2,2-dimethylbutane

The old labels of *normal-* (or *n-*) for straight chain, *iso-* for a branched chain, and *neo-* for a third isomer rapidly become confusing as the number of carbon atoms increases, and the systematic nomenclature of the right column in Table 21-2 must be used. With the systematic nomenclature, the compound is given the name corresponding to the *longest carbon chain* that can be traced through the molecule. The molecule is stretched along this longest chain, and carbon atoms are counted beginning with the end that has the nearest branch point. The side chains are then identified, and located by giving the number of the carbon to which they are attached on the main chain. Hydrocarbon side chains are named by analogy with the hydrocarbons: CH_3- , methyl; CH_3CH_2- , ethyl; $\text{CH}_3\text{CH}_2\text{CH}_2-$, propyl; and



Thus, neopentane in systematic nomenclature is dimethylpropane, and not trimethylethane or even tetramethylmethane, because the longest continuous carbon chain has three carbon atoms as in propane.

Example 1

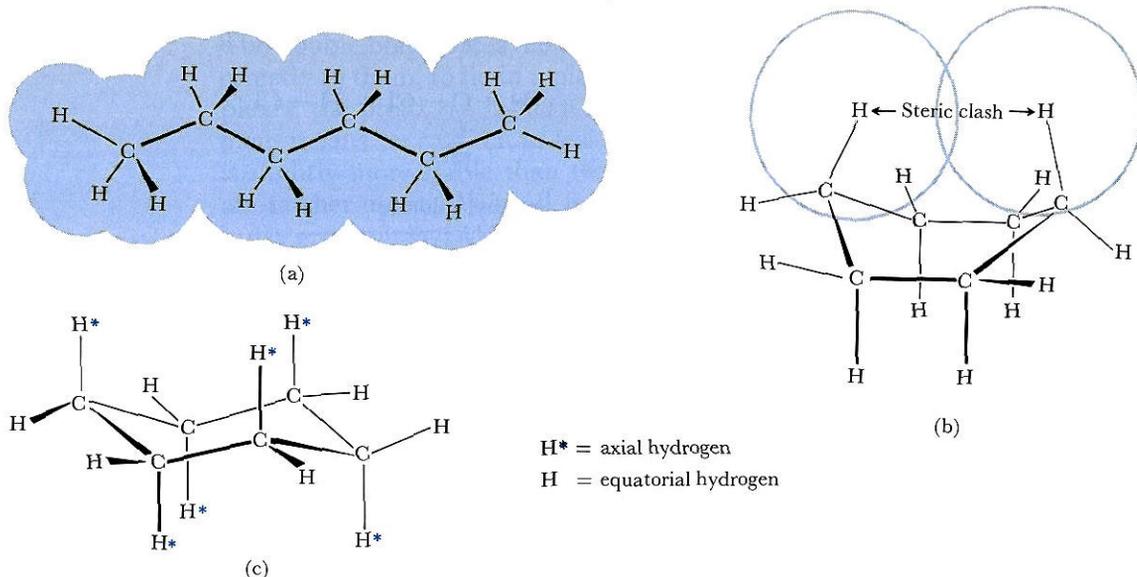
Draw the carbon skeletons and give the systematic names for the nine isomers of C_7H_{16} .

Solution

You will find one heptane (7-carbon straight chain), two substituted hexanes (6-carbon chain with side branches), five pentanes, and one butane. The last compound is 2,2,3-trimethylbutane.

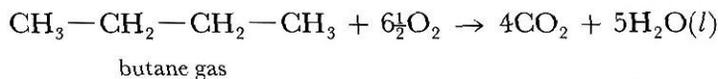
Example 2

Give the systematic name of the following compound:

**Figure 21-9**

n-Hexane, C_6H_{14} , and cyclohexane, C_6H_{12} . (a) The straight-chain hydrocarbon *n*-hexane. (b) The boat configuration of cyclohexane. The two top hydrogen atoms at the prow and stern of the boat are too close; this configuration is less stable than (c), the chair form. This same type of ring occurs in hexose sugars, which also adopt the chair form. These drawings emphasize the tetrahedral bond geometry around each carbon atom. (Adapted from R. E. Dickerson and I. Geis, *Chemistry, Matter, and the Universe*, W. A. Benjamin, Menlo Park, Calif., 1976.)

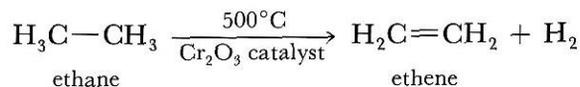
Combustion makes the alkanes useful as fuels:



$$\Delta H_{298}^{\circ} = -2878 \text{ kJ}$$

Propane and butane gas, gasolines, and kerosenes all are alkanes whose value lies in their combustibility.

Dehydrogenation is the removal of atoms of hydrogen and the creation of double or triple bonds. This process usually occurs at high temperatures and in the presence of a catalyst such as Cr_2O_3 :

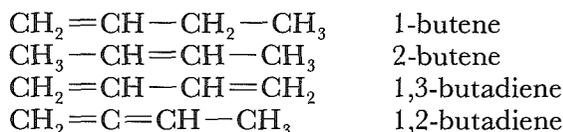


These dehydrogenated products are called *alkenes* or *olefins*. We shall discuss them further in the next section.

Halogenation is the reaction of a hydrocarbon with F_2 , Cl_2 , or Br_2 (I_2 is too inert under ordinary conditions) and the replacement of one or more H atoms by halogen atoms:

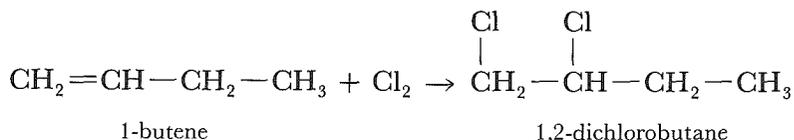
The double bond forces the two central C atoms, and the C and H attached directly to them, to lie in a plane. As with the isomers of coordination complexes, the prefix *cis*- indicates adjacent positioning of similar groups, and *trans*- means "across" or, at least, not adjacent. The *trans*-2-butene molecule is slightly more stable than the *cis* form because its bulky methyl groups are farther apart. We shall find that **steric hindrance**, or the bumping of bulky groups, plays a significant role in determining the structures of organic and biological molecules.

In the longer paraffins, dehydrogenation leads to a mixture of several products with the double bond in different places. The straight-chain isomer of butane, *n*-butane, can be dehydrogenated to two structural isomers of butene with one double bond, and two isomers of butadiene with two double bonds:

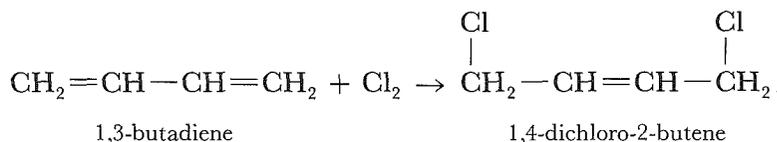


The numbers 1, 2, and 3 locate the positions of the double bonds.

Addition reactions can occur at the double bonds with H_2 , HCl , or Cl_2 . For example,



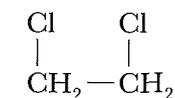
The corresponding 1,3-butadiene addition reaction is peculiar; the addition takes place at the extreme ends of the two double bonds, in what appears to be a simultaneous (concerted) process. One double bond disappears in the reaction, and the other moves to the center of the molecule:



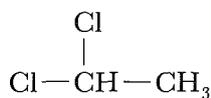
This unusual behavior occurs because the double bonds in the 1,3-butadiene molecule are delocalized. Such an alternating arrangement of double and single bonds ($-\text{C}=\text{C}-\text{C}=\text{C}-$) is called a **conjugated system**. When such conjugated double bonds occur in flat, closed rings with all atoms in a plane, we call the compounds **aromatic**. (See Section 21-6.)

21-5 DERIVATIVES OF HYDROCARBONS: FUNCTIONAL GROUPS

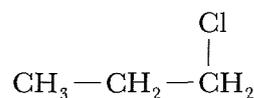
In a **chlorination** reaction, one or more hydrogen atoms can be replaced by Cl, and many isomers are possible. Some examples with their systematic names, are



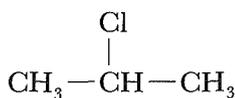
1,2-dichloroethane



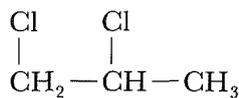
1,1-dichloroethane



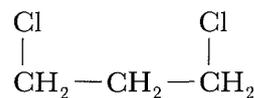
1-chloropropane



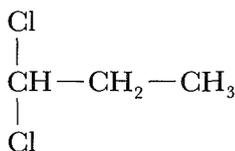
2-chloropropane



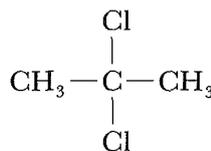
1,2-dichloropropane



1,3-dichloropropane



1,1-dichloropropane



2,2-dichloropropane

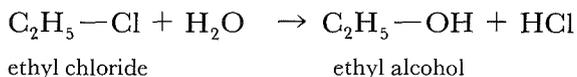
Example 3

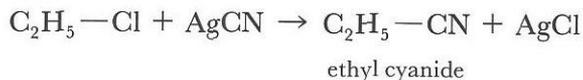
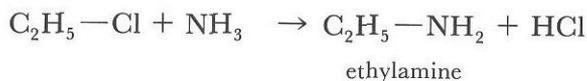
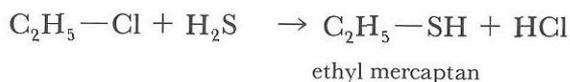
How many different isomers are there of trichloropropane, and what are they?

Solution

Five. 1,2,3; 1,2,2; 1,1,3; 1,1,2; 1,1,1.

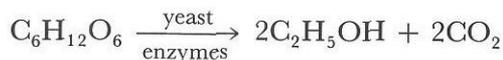
These chlorinated hydrocarbons are the starting materials for the preparation of many classes of compounds that cannot be prepared directly from the hydrocarbons. Their chemical reactivity lies in the C—Cl bond, and the rest of the molecules act as a unit in many reactions. Therefore, it is convenient to think of the hydrocarbon part of the molecules as a **radical** attached to a **functional group**. Ethyl chloride, $\text{CH}_3\text{CH}_2\text{—Cl}$, behaves chemically like the combination of an ethyl radical, $\text{CH}_3\text{CH}_2\text{—}$ or $\text{C}_2\text{H}_5\text{—}$, and a chloride group, —Cl . Many **replacement reactions** can occur, given the proper temperatures and catalysts:





In subsequent reactions of these products, the ethyl group usually remains intact, while chemical activity takes place at the bond between the ethyl radical and the functional group.

Several common functional groups are listed in Table 21-3 and are shown in three-dimensional skeletal models in Figures 21-10 and 21-11. The **alcohols** are good solvents for organic materials, and the lower-molecular-weight alcohols are soluble in water. Methanol, or “wood alcohol,” is a toxic alcohol that produces blindness and death when ingested. It attacks the nervous system by dissolving fatty material at nerve endings. The less toxic ethanol, or “grain alcohol,” is the end product of energy extraction in anaerobic (non-oxygen-using) organisms such as yeasts:



Methanol and ethanol are employed in vast quantities both as solvents and as raw materials for chemical syntheses. Methanol is synthesized commer-

Figure 21-10

Models of hydrocarbon derivatives, showing typical functional groups. (a) Methyl alcohol, with the —OH group; (b) acetaldehyde (named as a derivative of acetic acid, CH_3COOH), with the —CHO aldehyde group; (c) dimethyl ether, with the —O— ether bridge; and (d) dimethyl ketone, or acetone, with the ketone linkage.

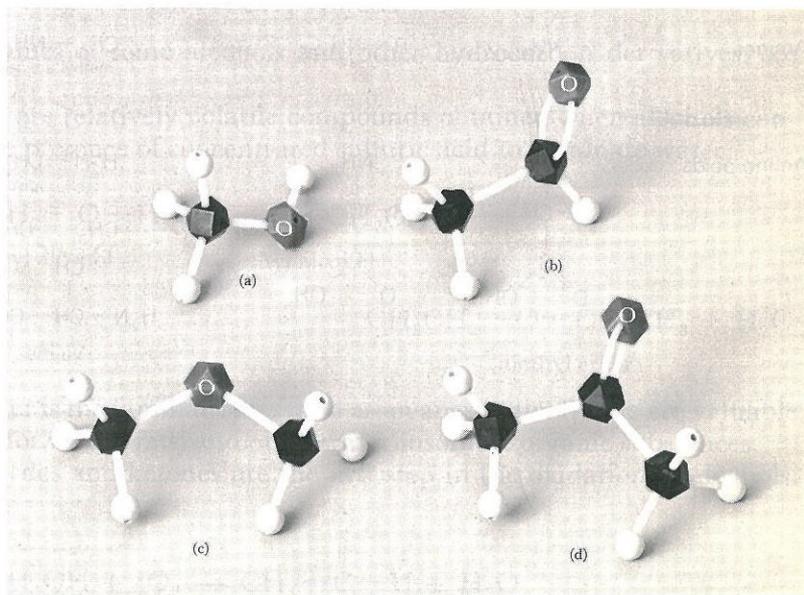


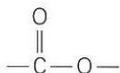
Table 21-3

Hydrocarbon Derivatives and Functional Groups

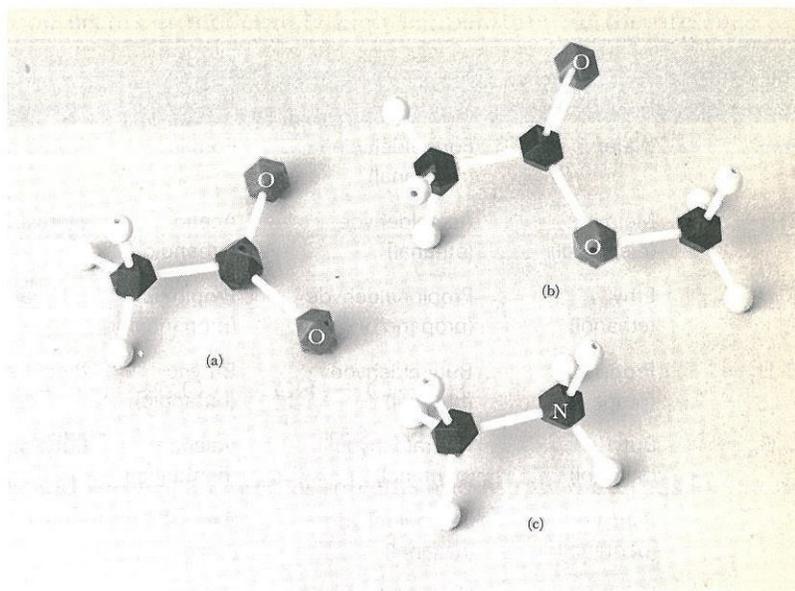
Derivative	Functional group	General formula	Examples
Halides	—Cl, —Br	R—Cl	$\text{CH}_3\text{—CH}_2\text{—Cl}$ ethyl chloride (chloroethane) $\text{Cl—CH}_2\text{—CH}_2\text{—Cl}$ 1,2-dichloroethane
Alcohols	—OH	R—OH	$\text{CH}_3\text{—OH}$ methanol $\text{CH}_3\text{—CH}_2\text{—OH}$ ethanol
Ethers	—O—	$\text{R}_1\text{—O—R}_2$	$\text{CH}_3\text{—O—CH}_3$ dimethyl ether $\text{CH}_3\text{—O—CH}_2\text{—CH}_3$ methyl ethyl ether
Ketones	$\begin{array}{c} \text{O} \\ \\ \text{—C—} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{R}_1\text{—C—R}_2 \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—CH}_3 \end{array}$ dimethyl ketone or acetone
Aldehydes	$\begin{array}{c} \text{O} \\ \\ \text{—C—H} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{R—C—H} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{H—C—H} \end{array}$ formaldehyde $\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—H} \end{array}$ acetaldehyde
Acids	$\begin{array}{c} \text{O} \\ \\ \text{—C—OH} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{R—C—OH} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{H—C—OH} \end{array}$ formic acid $\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—OH} \end{array}$ acetic acid
Esters	$\begin{array}{c} \text{O} \\ \\ \text{—C—O—} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{R}_1\text{—C—O—R}_2 \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{—C—O—CH}_2\text{—CH}_3 \end{array}$ ethyl acetate
Amines	—NH ₂	R—NH ₂	$\text{CH}_3\text{—NH}_2$ methyl amine $(\text{CH}_3)_2\text{—NH}$ dimethyl amine
Amino acids	$\begin{array}{c} \text{NH}_2 \\ \\ \text{—CH} \\ \\ \text{C} \\ // \quad \backslash \\ \text{O} \quad \text{OH} \end{array}$	$\begin{array}{c} \text{NH}_2 \\ \\ \text{R—C—H} \\ \\ \text{C} \\ // \quad \backslash \\ \text{O} \quad \text{OH} \end{array}$	$\begin{array}{c} \text{H}_2\text{N—CH}_2\text{—COOH} \\ \text{glycine} \end{array}$ $\begin{array}{c} \text{CH}_3 \\ \\ \text{H}_2\text{N—CH—COOH} \\ \text{alanine} \end{array}$ $\begin{array}{c} \text{CH}_3 \quad \text{CH}_3 \\ \backslash \quad / \\ \text{CH} \\ \\ \text{H}_2\text{N—CH—COOH} \\ \text{valine} \end{array}$ $\begin{array}{c} \text{O} \quad \text{OH} \\ // \quad \backslash \\ \text{C} \\ \\ \text{CH}_2 \\ \\ \text{H}_2\text{N—CH—COOH} \\ \text{aspartic acid} \end{array}$

Figure 21-11

Organic acids and bases, and their derivatives. (a) Acetic acid, shown with its carboxyl group ionized. (b) Methyl acetate, with the characteristic



ester linkage. (c) Methylamine, with the amine $-\text{NH}_2$ group. The model shows the amine in its ionic form, $-\text{NH}_3^+$.



cially from carbon dioxide and hydrogen:

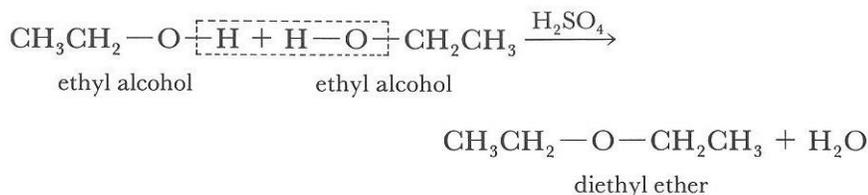


and ethanol is produced from ethene:



(For the names of some alcohols and other hydrocarbon derivatives, see Table 21-4.)

Ethers are relatively volatile compounds obtained when alcohols condense in the presence of concentrated sulfuric acid to eliminate water:



Diethyl ether is the familiar ether used as an anesthetic. Ethers are valuable as solvents for waxes, fats, and other water-insoluble organic substances.

Aldehydes and **ketones** are the first step in the oxidation of alcohols:

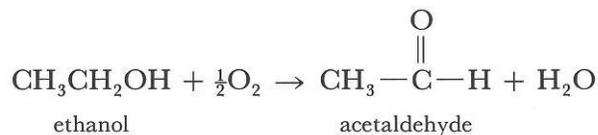


Table 21-4**Names of Hydrocarbon Derivatives^a**

R	Alcohols R—OH	Aldehydes R—CHO	Acids R—COOH	Esters CH₃—COO—R	Esters R—COO—CH₃
H—	Water	Formaldehyde (methanal)	Formic	Acetic acid	Methyl formate
CH ₃ —	Methyl (methanol)	Acetaldehyde (ethanal)	Acetic (ethanoic)	Methyl acetate	Methyl acetate (methyl ethanoate)
C ₂ H ₅ —	Ethyl (ethanol)	Propionaldehyde (propanal)	Propionic (propanoic)	Ethyl acetate	Methyl propionate (methyl propanoate)
C ₃ H ₇ —	Propyl (propanol)	Butyraldehyde (butanal)	Butyric (butanoic)	Propyl acetate	Methyl butyrate (methyl butanoate)
C ₄ H ₉ —	Butyl (butanol)	Valeraldehyde (pentanal)	Valeric (pentanoic)	Butyl acetate	Methyl valerate (methyl pentanoate)
C ₅ H ₁₁ —	Pentyl (pentanol)	Caproaldehyde (hexanal)	Caproic (hexanoic)	Pentyl acetate	Methyl caproate (methyl hexanoate)
C ₆ H ₁₃ —	Hexyl (hexanol)	Heptaldehyde (heptanal)	Heptanoic (heptanoic)	Hexyl acetate	Methyl heptylate (methyl heptanoate)
C ₇ H ₁₅ —	Heptyl (heptanol)	Octaldehyde (octanal)	Caprylic (octanoic)	Heptyl acetate	Methyl caprylate (methyl octanoate)
C ₈ H ₁₇ —	Octyl (octanol)	Pelargonic aldehyde (nonanal)	Pelargonic (nonanoic)	Octyl acetate	Methyl pelargonate (methyl nonanoate)
C ₁₁ H ₂₃ —	Undecyl (undecanol)	Lauric aldehyde (dodecanal)	Lauric	Undecyl acetate	Methyl laurate
C ₁₅ H ₃₁ —	Pentadecyl (pentadecanol)	Palmitic aldehyde (hexadecanal)	Palmitic	Pentadecyl acetate	Methyl palmitate
C ₁₇ H ₃₅ —	Heptadecyl (heptadecanol)	Stearic aldehyde (octadecanal)	Stearic	Heptadecyl acetate	Methyl stearate
CH ₃ (CH ₂) ₇ CH=CH(CH ₂) ₇ — (cis isomer)			Oleic		Methyl oleate

^aThe problem of names in organic chemistry is formidable. There are two parallel systems: the common names, and systematic names agreed upon by the International Union of Pure and Applied Chemistry (IUPAC).

Common names are generally shorter and more convenient, but are only labels. From the systematic name you usually can determine most of the molecule's structure. Systematic names are given in parentheses in this table.

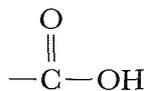
Common names are based on two series: those of the alkanes and those of the acids. The alkane series begins with arbitrary names but quickly shifts to the Greek prefixes indicating the number of carbon atoms: methyl, ethyl, propyl, butyl, pentyl, hexyl, and so forth. Unfortunately, the acid series retains its nonnumerical names, which usually reflect the source of the material.

Note that the numerical prefixes for acids are one place out of step with the alcohols because the carbon atom of the carboxyl group is included in the counting. Thus, C₅H₁₁COOH is *hexanoic* acid and not *pentanoic*.

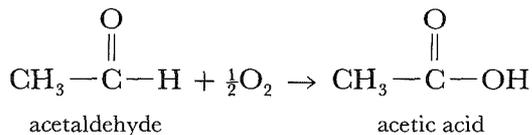
Aldehydes and the carbon-linked part of esters use the acid nomenclature. Alcohols, ethers, ketones, amines, and the oxygen-linked part of esters use the alkane nomenclature.

You should know the names through C₄, and should understand the principles of systematic nomenclature beyond this point.

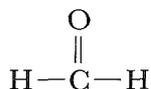
This reaction occurs at moderately high temperatures in the presence of a catalyst such as finely divided silver, or a mixture of powdered iron and molybdenum oxide. The second step in oxidation leads to a **carboxylic acid**, an acid with the carboxyl group,



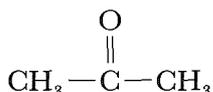
For example,



Aldehydes and ketones are used as solvents and as raw materials for chemical syntheses. Formaldehyde,

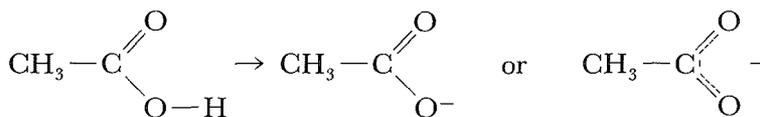


is the starting point for phenyl-formaldehyde resins such as Bakelite. Acetone,

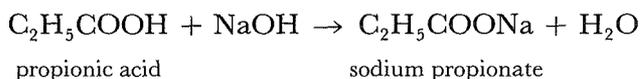


is one of the most common laboratory solvents.

The carboxylic acids are relatively weak acids; they dissociate to a limited extent in aqueous solution. When the carboxyl group does dissociate, the negative charge is spread over both oxygen atoms. The three *p* orbitals on the two oxygen atoms and the carbon atom connecting them are combined into one delocalized molecular orbital:



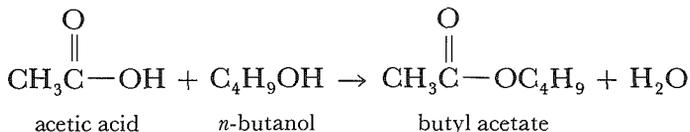
Both carbon-oxygen bonds in the ionized carboxyl group have the *same length*. The negative charge is spread over all three atoms. (The middle structure in the preceding equation can be considered as one of the two resonance structures contributing to the true carboxyl ion structure. What would the other resonance structure look like?) With metal hydroxides and carbonates, the carboxylic acids react as any other acid would to make salts:



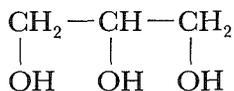
Sodium propionate is dissociated in aqueous solution into sodium ions and propionate ions, and is obtained as a salt only on drying.

Formic acid, HCOOH , is the main irritant in insect stings. Acetic acid, CH_3COOH , is the acid in vinegar. The acids from butyric (C_4) to heptanoic (C_7) have acrid odors that are encountered in rancid butter and strong cheese.

Esters are made by allowing acids and alcohols to react:

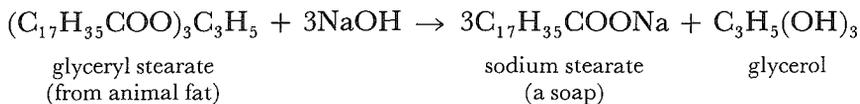


Although they are named as though they were salts, esters are not ionized. Many are volatile liquids with pleasing, fruity odors. Butyl acetate gives bananas their odor and therefore is called banana oil. Ethyl butyrate, $\text{C}_3\text{H}_7\text{COOC}_2\text{H}_5$, has the odor of pineapples, and octyl acetate, $\text{CH}_3\text{COOC}_8\text{H}_{17}$, the odor of oranges. Oils such as linseed, cottonseed, and olive oil, and fats such as butter, lard, and tallow, are esters of the trihydroxyl alcohol glycerol,



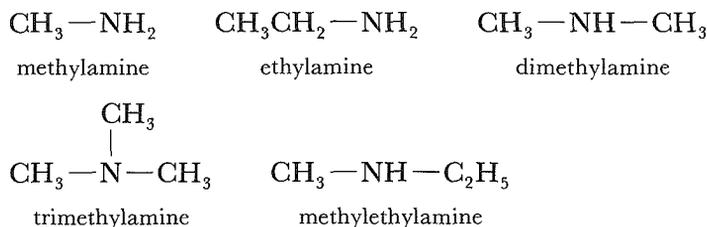
with large molecular weight acids such as palmitic, $\text{C}_{15}\text{H}_{31}\text{COOH}$; stearic, $\text{C}_{17}\text{H}_{35}\text{COOH}$; and oleic, $\text{C}_{17}\text{H}_{33}\text{COOH}$.

Soluble soaps are the alkali metal salts of these fatty acids, obtained by treating animal fats with alkali metal hydroxides, especially NaOH :



In aqueous solution, a soap molecule has a hydrocarbon end and a charged end. Soaps "lift" dirt into solution by surrounding a small amount of grease with many molecules; all their hydrocarbon tails point in toward the grease and their carboxyl groups point out. The soap molecules thus "package" the grease in droplets or **micelles** that can be taken up into the solution and washed away.

The most common organic bases are called **amines** and can be thought of as derivatives of ammonia:



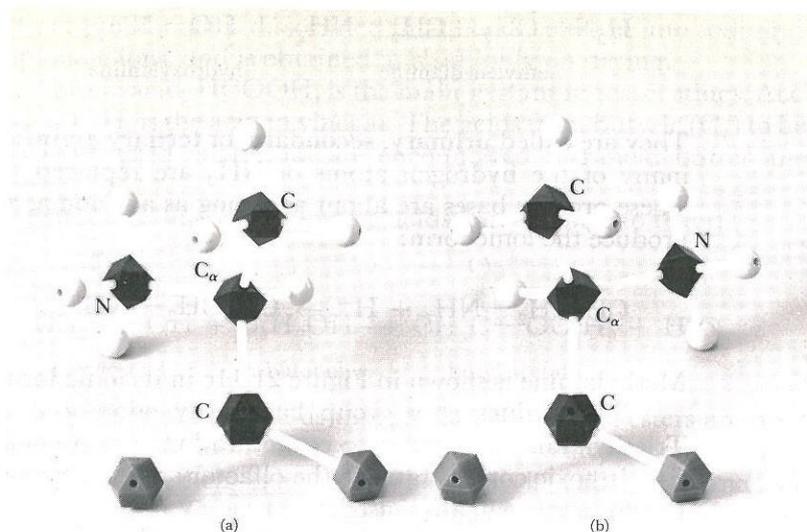
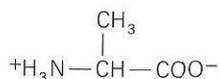
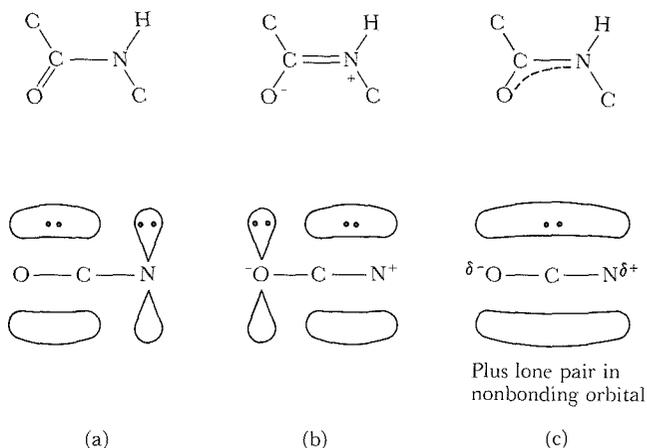


Figure 21-12 The two optical isomers of an amino acid, showing the side chain branching in opposite directions from the central carbon atom, which is called the α carbon. (a) L-Alanine, and (b) D-alanine. Both the carboxyl group and the amine group are shown in their ionized form



The proteins of all living organisms are built from only L-amino acids. Their mirror images, D-amino acids, are found in small amounts in bacterial cell walls and in antibiotics produced by some microorganisms. One of the problems in explaining the evolution of life is accounting for this asymmetry of the components of living organisms. Any carbon atom that has four different atoms bonded to it will have two different possible configurations which are mirror images of one another. Such a carbon atom is called an asymmetric carbon.

Once the peptide bond is made, electrons in the $\text{C}=\text{O}$ double bond become delocalized and the peptide $\text{C}-\text{N}$ bond acquires a partial double bond character. The peptide unit (Figure 21-13) is thus forced to remain planar. This unit is the cornerstone of all protein structure and is one of the most important examples of delocalization of π bonding in chemical systems.

**Figure 21-13**

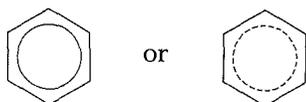
A structure for the peptide bond between amino acids in proteins can be drawn so that a double bond connects C and O, and the peptide bond itself is a single C—N bond (a). Another structure can be drawn with a single bond from C to O and a double peptide bond (b). This structure places charges on O and N, and is therefore less favorable. The true situation can be represented by combining p atomic orbitals from O, C, and N to make bonding, nonbonding, and antibonding delocalized molecular orbitals. The delocalized bonding orbital extends over all three atoms (c), and is therefore more stable. The nonbonding orbital is not shown. The extra stability of the delocalized electrons more than compensates for the slight charge separation at O and N. The partial double bond character of the the C—N peptide bond prevents rotation around the bond and keeps the peptide unit planar.

21-6 AROMATIC COMPOUNDS

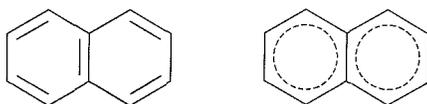
Aromatic compounds are ring molecules with delocalized electrons. The simplest of these is benzene, C_6H_6 . The delocalized electrons give aromatic compounds the special properties that differentiate them from the **aliphatic** compounds we have been examining so far. The benzene ring is commonly written as one of the Kekulé structures,



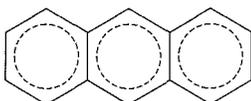
although a better representation of delocalization is



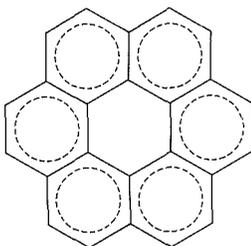
The delocalization can extend over more than one adjacent ring as in naphthalene:



and anthracene:



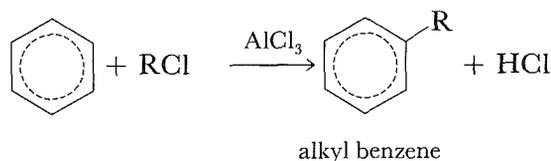
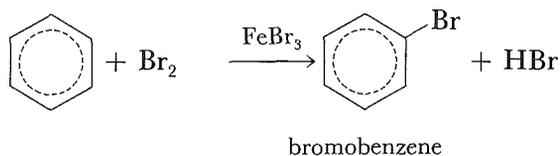
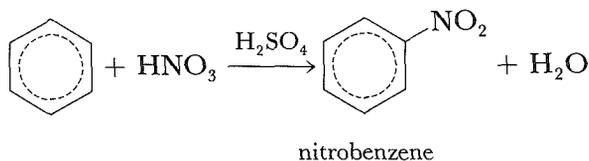
Coronene has seven adjoining rings:



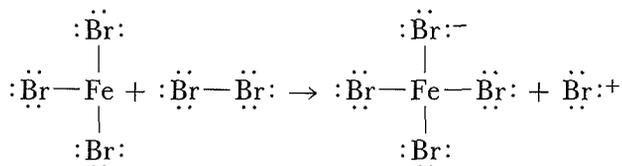
The ultimate limit of this process is graphite, with its sheets of hexagonal rings and delocalization over the entire sheet. Because of these delocalized electrons, graphite is a good conductor of electricity, whereas diamond is not. In a sense, graphite is a two-dimensional metal whose electron mobility is restricted to the individual stacked sheets.

Benzene is surprisingly unreactive in comparison to alkenes such as butene. In its lack of reactivity it is more like the saturated alkanes. It does not undergo addition reactions at a double bond; if it did so, the extent of the delocalization of electrons would be reduced. Because of this delocalization, benzene is 166 kJ mole^{-1} more stable than we would expect a compound with three single and three double bonds to be (Figure 15-8). In general, the larger the region in a molecule over which delocalization occurs, the more stable the molecule is.

Rather than addition, the typical reaction of aromatic rings is **substitution**:



(In writing structures of this sort, it is common practice to omit the H's bonded to the ring carbons; each apex of a hexagonal ring represents a C to which one H is attached.) In the first of these reactions, sulfuric acid aids the reaction by converting HNO_3 to NO_2^+ , the species that attacks the benzene ring. Sulfuric acid also acts as a dehydrating agent to remove the water formed as a product. The compounds FeBr_3 and AlCl_3 are **catalysts**. To see why they are necessary we must look at the mechanism of the reaction. Aromatic rings are particularly susceptible to attack by **electrophilic groups**, or Lewis acids, which have a strong affinity for electron pairs. In the bromination reaction, Br_2 is not electrophilic; in the absence of the FeBr_3 catalyst, no reaction takes place within a reasonable time. However, FeBr_3 itself has an attraction for another Br^- ion with its electron pair and will tear a Br_2 molecule into Br^- and Br^+ ions:



The electrophilic Br^+ then attacks the aromatic ring and attracts an electron pair to make a C—Br bond. This intermediate compound is unstable and dissociates either by ejecting the Br^+ to make the starting material again, or by ejecting a hydrogen ion to make the end product, bromobenzene:

black. The substances catalyzed, called the **substrates**, bind to the surface of the catalyst. If chemical groups on the surface of the catalyst weaken a bond in a substrate, the cleavage of the substrate becomes easier. This is what happens with platinum catalysis. The finely divided black powder of Pt is a more efficient catalyst than a block of Pt only because it has much more surface exposed.

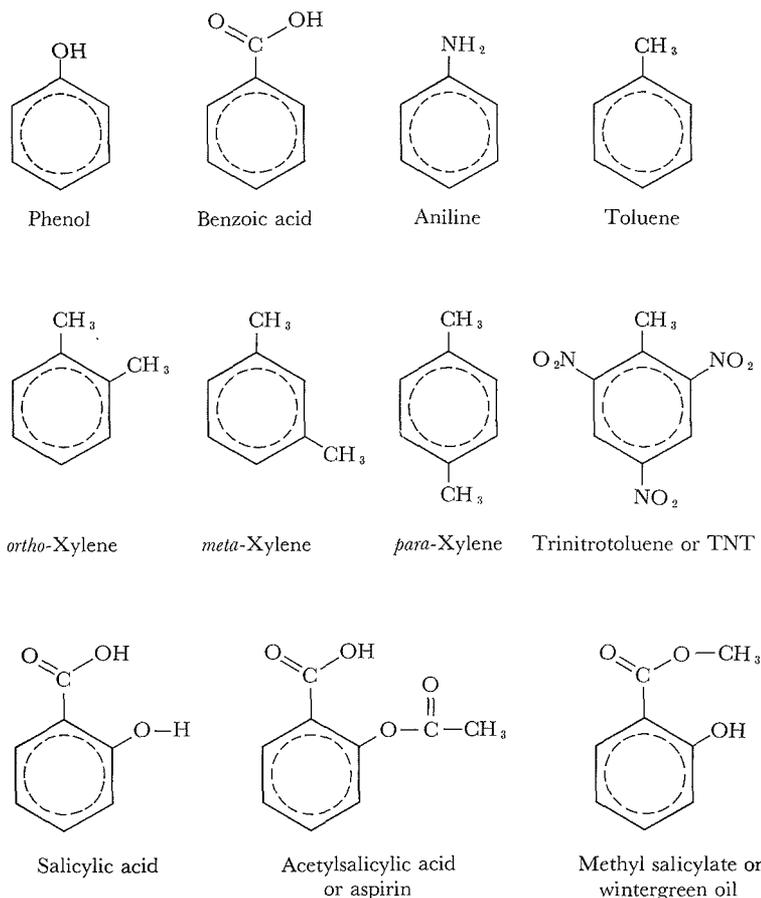
The compound AlCl_3 plays a catalytic role in the alkylation reaction to produce benzene derivatives with alkyl side chains. An important class of biological catalysts are the protein molecules called **enzymes**. These molecules have regions on their surface, called **active sites**, where catalysis occurs. Transition metals are frequently bound to the enzymes at their active sites, and are essential participants in catalysis. We shall look at an example of enzymatic catalysis in Section 21-10.

Several of the derivatives of benzene are shown in Figure 21-14. Phenol is weakly acidic, unlike the alcohols of which it appears to be an aromatic analogue. This ability of phenol and its derivatives to lose the hydroxyl proton arises because electrons of the oxygen become partially involved in the delocalization. The bond from ring to oxygen attains a partial double bond character, and hydrogen, robbed of some of its bonding electron pair, dissociates easily. However, the acidity of the phenols is generally less than that of the carboxylic acids.

For the same reason, aniline is a weaker base than ammonia or the aliphatic amines. The nitrogen lone electron pair that would have attracted a proton is partially involved with the aromatic ring and is less able to attract the proton and to ionize the molecule.

21-7 AROMATIC COMPOUNDS AND THE ABSORPTION OF LIGHT

Aromatic ring compounds with delocalized electrons, like transition-metal complexes with d orbitals, frequently have energy levels close enough together to absorb visible light. Hence, these two classes of compounds are often brightly colored. When a photon of energy is absorbed, one electron in a π^b bonding orbital (Figure 13-26) is promoted to the lowest π^* antibonding molecular orbital. Therefore, this absorption is called a $\pi \rightarrow \pi^*$ transition. In benzene and in naphthalene the levels are too far apart for the absorption to be in the visible spectrum, and these compounds are colorless. But if two nitro groups are added to make 1,3-dinitronaphthalene, the electronic-level spacing falls below $25,000 \text{ cm}^{-1}$ and the compound appears pale yellow. (Table 20-3 with its spectral and complementary colors will be helpful throughout this section.) This phenomenon happens because the delocalized electron system has been enlarged to include the two nitro groups, and the energy levels (and the spacings between them)

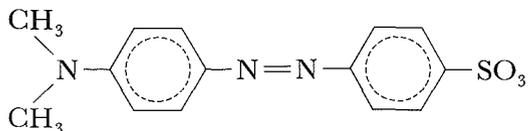
**Figure 21-14**

Some representative derivatives of benzene. Salicylic acid can form an ester in two ways: by using either its acid group in methyl salicylate or its hydroxyl group in aspirin. Unlike alcohols, phenols are acids, although they are usually far weaker acids than the carboxylic acids. Aromatic amines such as aniline are weaker bases than aliphatic amines. *Ortho*-, *meta*-, and *para*- (frequently abbreviated *o*-, *m*-, and *p*-) denote relative positions of groups attached to the benzene ring, as for the three isomers of xylene.

have fallen accordingly. The effect is continued in Martius Yellow, a common dye for wool and silk. An added hydroxyl group enlarges the conjugated system even more, and the energy of a $\pi \rightarrow \pi^*$ transition decreases. The color of the compound changes to yellow-orange. The light actually absorbed by the three compounds is ultraviolet for naphthalene, violet for 1,3-dinitronaphthalene, and blue for Martius Yellow.

In the acidic form of *p*-nitrophenol, the negative charge on the oxygen atom is lacking. It is not as easy to involve the oxygen lone electron-pairs in delocalization; therefore, the energy level of the first excited electronic state is not lowered as much. Absorption occurs with a maximum just inside the ultraviolet, at 320 nm, and the compound appears a pale greenish-yellow. Phenolphthalein, which is colorless in acid and pink in base, is a more complicated molecule that works by exactly the same principle.

A particularly good way of expanding a delocalized system is illustrated by the azo dyes, which have two aromatic rings bridged by the —N=N— group. Methyl orange, another acid–base indicator (Figure 5-6), is an azo dye:



It is red in acid and yellow in base. (In which conditions are its electronic energy levels more widely spaced? Can you figure out why?)

An important example of delocalization and energy absorption is chlorophyll, which was discussed in the Postscript to Chapter 20. The aromatic ring surrounding the Mg^{2+} ion is an extended delocalized system derived from porphyrin (Figure 20-19). The electronic energy levels are such that one absorption occurs in the violet, at 430 nm, and a second in the red, at 690 nm, (Figure 20-22). When light is absorbed by chlorophyll molecules, the energy excites an electron to a higher energy level, thereby enabling it to reduce the Fe^{3+} ions in *ferredoxin*, which is a protein of molecular weight 13,000 that has two iron atoms coordinated to sulfur. The reoxidation of ferredoxin supplies the energy to drive other reactions that eventually lead to the splitting of water, the reduction of carbon dioxide, and ultimately the synthesis of glucose, $C_6H_{12}O_6$.

21-8 CARBOHYDRATES

The sugar glucose, produced in the leaves of green plants, is a **carbohydrate**. The name *carbohydrate* comes from an early misconception about the structures of these compounds. The formula for glucose, $C_6H_{12}O_6$, can be written as $(C \cdot H_2O)_6$. Substances whose formulas could be represented by equal amounts of carbon and water were called carbohydrates.

The glucose molecule is polymerized in chains of thousands of monomer units in plants to make cellulose, and in a slightly different way to make starch. A close relative of glucose, *N*-acetylglucosamine (NAG), is polymerized to form chitin, the material from which the shells of insects are made. NAG and a close variant, *N*-acetylmuramic acid (NAM), are copolymerized in alternating sequence in chains that make up part of the walls of bacterial cells. Glucose is decomposed in a stepwise fashion to pro-

duce the energy that a living organism requires. Excess glucose is carried in the bloodstream to the liver and is converted into the animal starch glycogen, which is reconverted to glucose when needed. Glucose, cellulose, starch, and glycogen all are carbohydrates.

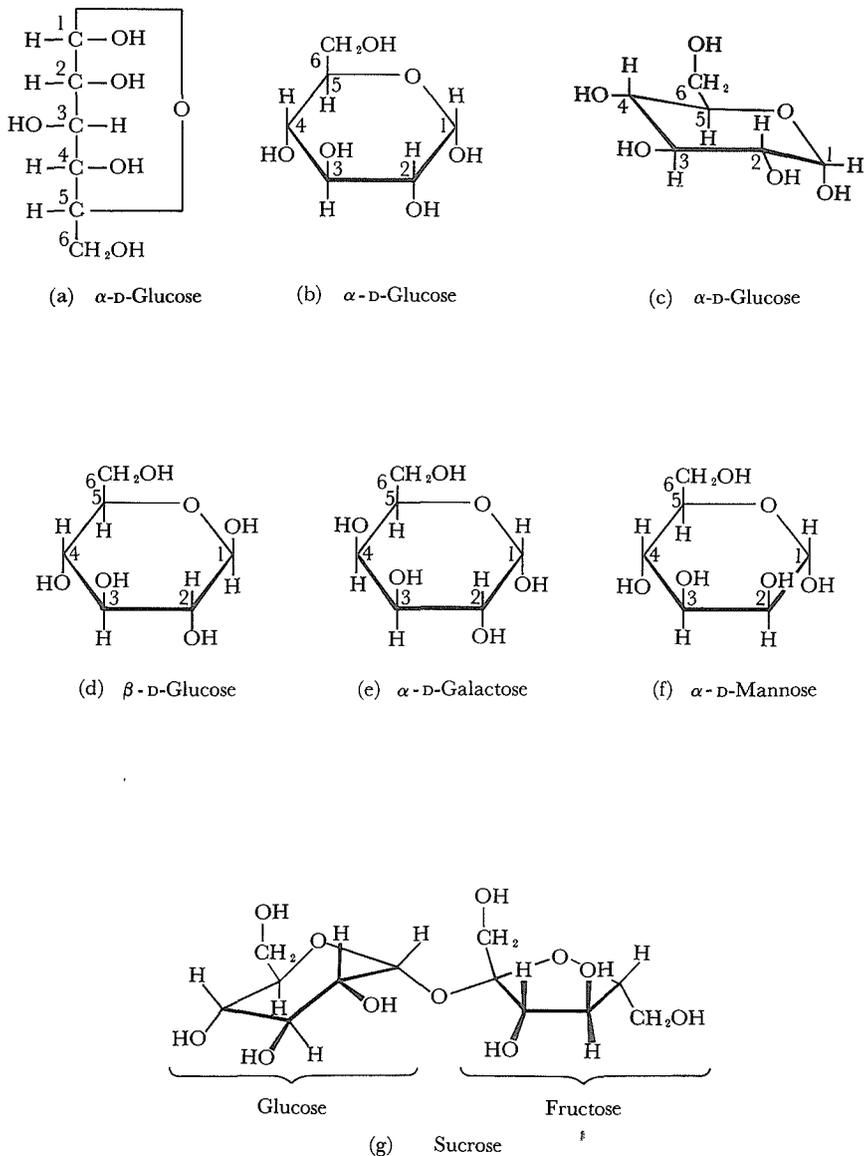
Carbohydrates, in the form of starch, are the primary sources of energy from foods. To obtain this energy, we either eat the grains in which the starch is stored, or feed the grains to animals and let them synthesize meat protein before we eat *them*. In either case, the energy that we obtain ultimately originates from starch, the polymerized product of photosynthesis. We encounter cellulose fibers in cotton and linen, and in the artificial products cellulose acetate and rayon. The shelter over our heads probably is cellulose in the form of wood. This book is a processed cellulose called paper. Even our money, having ceased to be made from noble metals, is well on its way to becoming notarized cellulose. In this section we shall look very briefly at what carbohydrates are and how they are used.

The most fundamental unit of a carbohydrate is a **monosaccharide**, or simple sugar. ~~S~~ Sugars can have three, four, five, or six carbon atoms, in which case they are called **trioses**, **tetroses**, **pentoses**, or **hexoses**. We shall look only at hexoses, and especially at the most common one, D-glucose. The structure of D-glucose is depicted in Figure 21-15, parts a–c. Figure 21-15a shows the numbering of the six carbon atoms, and the Fischer convention of writing formulas to indicate the structure around an asymmetric carbon.

An **asymmetric carbon atom** is one that is bonded to four different groups, as are carbon atoms 1 through 5 in glucose. As we saw for the α carbon of an amino acid (Figure 21-12), each such asymmetric carbon atom has two different arrangements of the four groups, which are related by a mirror reflection. With five asymmetric carbon atoms, and two different configurations around each, there are a total of $2^5 = 32$ different isomers of the hexose sugars.

By the Fischer convention, the bonds to the right and left in Figure 21-15a lead from the central atom to atoms that lie above the plane of the page. Bonds extending up or down from the central atom go to atoms below the plane of the diagram. A change in configuration at any asymmetric carbon atom in the hexose is produced by exchanging the —H and —OH groups right for left in the Fischer diagram. This asymmetry is easier to see in the flat hexagon representation of the same molecule in Figure 21-15b. The actual shape of the molecule, with its tetrahedral geometry at the carbon atoms, is depicted more accurately by Figure 21-15c. Glucose has the chair conformation, which we first saw with cyclohexane, rather than the boat form.

In the 32 isomers of hexose that arise from the 32 possible interchanges of arrangement at carbon atoms 1 through 5, the positions —H and —OH at carbon atom 1 are indicated by the prefixes α and β . In α -hexoses the hydroxyl group points down as in part b or c of Figure 21-15; in β -hexoses it points up as in Figure 21-15d. A complete mirror reversal of a hexose at all five asymmetric carbon atoms simultaneously produces an L-hexose from

**Figure 21-15**

(a) α -D-Glucose in the Fischer representation, (b) in the flat hexagon diagram, and (c) in a form that most closely represents its actual shape. (d) β -D-Glucose, which differs from the α form only at carbon 1. (e) α -D-Galactose, which differs from glucose at carbon 4. β -D-Galactose is produced by exchanging —H and —OH at carbon 1. (f) α -D-Mannose, which differs from α -D-glucose only at carbon 2. (g) Sucrose, a dimer of α -D-glucose and β -D-fructose.

a D-hexose. Therefore, for each type of hexose there are four variants: α -D, α -L, β -D, and β -L. There must be $32/4 = 8$ different named types of hexose. However, only three of these occur naturally: glucose, galactose, and mannose. These three sugars differ at carbon atoms 2 and 4, and are compared in Figures 21-15d, e, and f. Galactose occurs in the milk sugar lactose, and mannose is a plant product (named for the Biblical *manna*). However, the most common hexose by far is glucose.

Of the hexose sugars with a five-membered ring, the most common is *fructose*. Fructose occurs naturally in honey and fruit (hence its name), and is combined with glucose as the common table sugar *sucrose* Figure 21-15g.

Polysaccharides

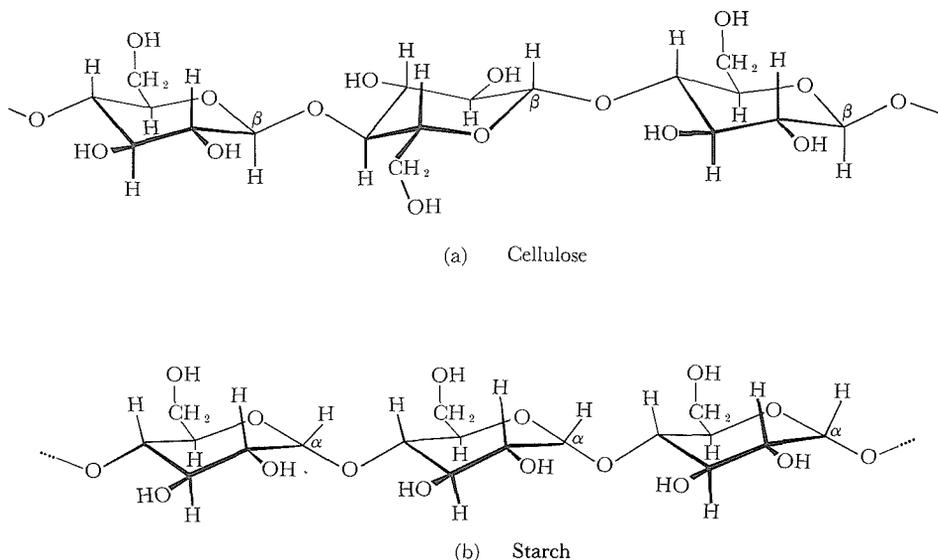
Cellulose is the structural fiber in trees and plants. It is found in wood, cotton, and linen, and, in a modified form, in paper. Cellulose is a polymer of β -D-glucose, with a typical chain length of about 3000 monomer units. The connection from one β -glucose to another, shown in Figure 21-16a, is called a **β -glucoside link**.

The hydroxyl groups of glucose can form esters; the treatment of cellulose with acetic anhydride, acetic acid, and a small amount of sulfuric acid produces the derivative cellulose acetate. The chains are broken to a length of 200–300 monomers, and an average of two acetate groups attach to each monomer. Cellulose acetate is the material for photographic film backing; it is also dissolved in acetone and extruded through fine holes in a metal cup to form threads of rayon.

Cellulose is not a source of food. With the exception of termites, and ruminants such as cows, both of which carry cellulose-digesting microorganisms in their stomachs, animals are incapable of breaking the β -glucosidic bond. The cleavage is an enzymatically catalyzed process, and we lack the enzymes. In 1967, a process was developed for degrading cellulose to produce an artificial flour that, although usable in baking like starch flour, had no nutritive value. It was touted briefly as a dieting aid, but rapidly sank into obscurity. (*Life* magazine referred to it as “non-food,” and suggested that its inventor be paid in non-money.) However, it has been suggested quite seriously that if man could in some way learn to live in a symbiotic relationship with cellulose-digesting microorganisms in his intestines as ruminants do, his food problems would be resolved for many centuries.*

*This is one of those superficially attractive suggestions with possibly disastrous social consequences. It is an illustration of the unpleasant fact that a blind application of science and technology to isolated problems often creates more problems than it solves. What would be the most likely consequences if cellulose suddenly became an apparently unlimited source of cheap food? We can list a few of the results.

1. A rapid diminution of interest in the crisis of overpopulation, and an upsurge in the total population of the planet. (footnote continued on page 830)

**Figure 21-16**

(a) Cellulose, a polymer of β -D-glucose. (b) Starch, a polymer of α -D-glucose.

Starch also is a polymer of glucose, but with the α linkage of Figure 21-16b. Starch is the standard storage medium for glucose to be used as a food supply in plants, and is our chief source of trapped solar energy. It is stored in plant stalks, leaves, roots, seeds, and grains. All organisms possess the enzymes necessary to digest starch. The first step in fermentation, whether it takes place in the stomach or in the brewer's vat, is the breakdown of starch to glucose. A piece of bread held in your mouth eventually will taste sweet because the enzymes in the saliva can digest bread starch to sugar.

Polymers of hexose derivatives are the structural materials in insect shells (chitin) and bacterial cell walls. In insect chitin, a hexose derivative called *N*-acetylglucosamine is polymerized without cross-linking. One layer

2. Great changes in standards and modes of living in the face of the population increases made possible by "unlimited" cheap food.

3. Deforestation of large areas of the world where people are starving. This deforestation would lead to flooding, erosion of topsoil, crop failures of conventional foods, and probable starvation again.

The introduction of edible cellulose, *by itself*, would be as shortsighted and disastrous a step as finding an efficient way of diffusing the smog of Los Angeles over the farms of Iowa. It would not avoid the eventual crisis; it would merely delay its arrival. Our planet is so carefully structured that almost any major technological or scientific change introduced without thinking is likely to lead to trouble. Who could have convinced Henry Ford that his mass transportation machines would eventually become a curse on the landscape?

in the walls of bacteria is a polymer of hexose derivatives, cross-linked with short chains of four amino acids for strength. We and all other higher organisms have evolved an enzyme, lysozyme, to protect us by lysing or dissolving this polysaccharide wall structure in invading bacteria. Lysozyme is found in most external secretions such as sweat and tears. One of the few places where D-amino acids exist in nature is in the walls of certain bacteria. One view is that they have been placed there by the bacteria simply to keep them out of the way, but it also has been speculated that they might have evolved in the structure of the wall as a defensive measure against attack by enzymes (not lysozymes, which attack the β -glucoside link) that operate most effectively against the common L-amino acids.

21-9 PROTEINS AND ENZYMES

A protein is a folded polymer of amino acids. Such a polymer is shown at the bottom of Figure 21-1, and a model of a single amino acid appears in Figure 21-12. Enzymes are one class of proteins, and perhaps the most glamorous class. They are approximately globular molecules with molecular weights from 10,000 to several million and diameters of 20 Å and more. They serve as the catalysts that control biological reactions. Other globular protein molecules such as myoglobin and hemoglobin are carriers and storage units for molecular oxygen (Figures 20-25, 20-26). The cytochromes are oxidation–reduction proteins that serve as intermediate links in the extraction of energy from foods (Figure 20-23). The gamma globulins are antibody molecules with a molecular weight of 160,000. They attach themselves to viruses, bacteria, or other foreign bodies, and precipitate them from body fluids to protect their host. All of these proteins are **globular proteins**.

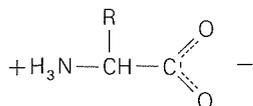
The other large class of proteins is the **fibrous proteins**. These are mainly structural materials. **Keratin**, found in skin, hair, wool, nails, and beaks, is a fibrous protein. **Collagen** in tendons, the underlayers of skin, and the cornea of the eye is another type of fibrous protein, as are silks and many kinds of insect fibers. Proteins in combination with carbohydrates, and with lipids (long-chain fats and fatty acids), are the structural materials of all living organisms.

The chief distinction between a protein chain and a chain of polyethylene or Dacron is that not all side chains in a protein are alike. In fibrous proteins, it is the repetition of the sequence of side groups that gives a particular fibrous protein—silk or hair or collagen—its special mechanical properties. Globular proteins are even more intricate. These molecules typically have 100 to 500 amino acids polymerized in one long chain, and the complete sequence of side groups is the same in every molecule of the same globular protein. The side groups can be hydrocarbonlike, acidic, basic, or neutral but polar. Both the folding of the protein chain to make a compact globular molecule and the chemical behavior of the molecule once

Table 21-5**Side Groups of the Twenty Common Amino Acids**

Symbol	Name	Side group
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All of the amino acids given except Pro can be represented by the formula



in which the side groups, —R, are listed below. The entire molecule is shown for Pro.

Acidic side groups

Asp	Aspartic acid	—CH ₂ COOH
Glu	Glutamic acid	—CH ₂ CH ₂ COOH

Basic side groups

Lys	Lysine	—CH ₂ CH ₂ CH ₂ CH ₂ NH ₂
Arg	Arginine	—CH ₂ CH ₂ CH ₂ NH—C(NH)NH ₂
His	Histidine	—CH ₂ —C ₅ H ₄ N

Uncharged but polar side groups

Asn	Asparagine	—CH ₂ —CO—NH ₂
Gln	Glutamine	—CH ₂ CH ₂ —CO—NH ₂
Ser	Serine	—CH ₂ OH
Thr	Threonine	—CH(OH)—CH ₃
Gly	Glycine	—H

Sulfur-containing side groups

Cys	Cysteine	—CH ₂ SH
Met	Methionine	—CH ₂ CH ₂ —S—CH ₃

Aliphatic side groups

Ala	Alanine	—CH ₃
-----	---------	------------------

it is folded are determined by the kind and sequence of amino acid side groups.

Only 20 different kinds of amino acid side groups are ordinarily found in living organisms. These side groups are shown in Table 21-5. Some of them are hydrocarbonlike, such as Val, Leu, Ile, and Phe. The hydro-

Table 21-5 (continued)

Side Groups of the Twenty Common Amino Acids

Symbol	Name	Side group
Val	Valine	$\begin{array}{c} \text{—CH—CH}_3 \\ \\ \text{CH}_3 \end{array}$
Leu	Leucine	$\begin{array}{c} \text{—CH}_2\text{—CH—CH}_3 \\ \\ \text{CH}_3 \end{array}$
Ile	Isoleucine	$\begin{array}{c} \text{—CH—CH}_2\text{—CH}_3 \\ \\ \text{CH}_3 \end{array}$
Pro	Proline (entire amino acid shown)	$\begin{array}{c} \text{CH}_2 \\ / \quad \backslash \\ \text{CH}_2 \quad \text{CH}_2 \\ \quad \backslash \\ \text{+H}_2\text{N—CH—COO}^- \end{array}$
<u>Aromatic side groups</u>		
Phe	Phenylalanine	$\text{—CH}_2\text{—} \langle \text{benzene ring} \rangle$
Tyr	Tyrosine	$\text{—CH}_2\text{—} \langle \text{benzene ring} \rangle \text{—OH}$
Trp	Tryptophan	$\text{—CH}_2\text{—} \langle \text{indole ring} \rangle$

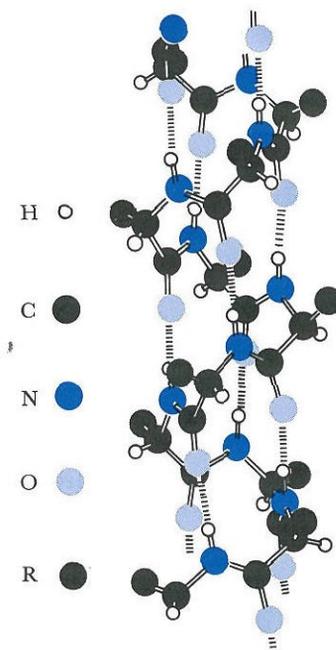
phobic groups are more stable if they can be removed from an aqueous environment. Hence a protein chain in aqueous solution tends to fold into a molecule with these groups *inside*. Other side groups are charged: The acids Asp and Glu are ionized and negatively charged, and the bases Lys and Arg are positively charged at pH 7. Others, such as Asn, Gln, and Ser, although uncharged, are compatible with an aqueous environment. One of the most important factors in determining how a protein chain will fold into a globular molecule is the stability that results if hydrophobic groups

are buried on the inside of the molecule and charged groups are outside. Although either optical isomer shown in Figure 21-12 might seem equally probable, all of the amino acids in proteins are L-amino acids (Figure 21-12a).

A protein chain is particularly stable when coiled into a right-handed α helix (Figure 21-17). In this structure, the side groups point away from the axis of the helix, and the C=O groups of one turn of the helix are bonded to the H—N groups of the next higher turn by a hydrogen bond. Hydrogen bonds form between especially electronegative atoms, such as F or O, and hydrogen atoms with a slight local excess of partial positive charge. Such bonds are mainly electrostatic and depend on the two atoms' being able to approach closely. Because they are small, O and F can make such bonds; the larger Cl ordinarily cannot. Hydrogen bonds are quite common in proteins because of the presence of the carbonyl oxygen and amine hydrogen on the polypeptide chain. As Figure 21-13 shows, the partial double bond character of the C—N peptide bond not only keeps the linkage planar, it also makes the oxygen atom slightly negative and the nitrogen atom with its hydrogen atom slightly positive. These are favorable conditions for hydrogen bonding.

Figure 21-17

The α helix, a type of protein chain folding found in both fibrous and globular proteins. The α helix was proposed first by L. Pauling and R. B. Corey from model-building experiments based on the bond lengths and bond angles determined in x-ray analyses of individual amino acids and polymers of two and three amino acids. The structure since has been discovered in hair and wool, in skin keratin, and in globular proteins such as myoglobin and hemoglobin.



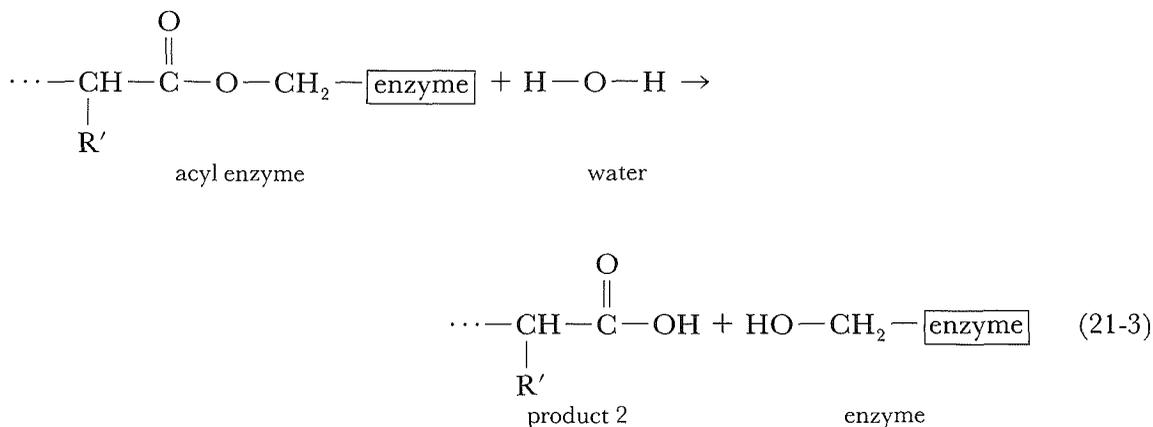
In hair, wool, and other keratins, α helices are twisted into threads, strands, and cables to make the fibers that we can see and manipulate. In silk, the chains are stretched full length rather than in an α helix, and are cross-connected by hydrogen bonds into the sheets illustrated in Figure 21-2a. In the globular proteins, the strands can be neither fully extended nor fully α -helical; there must be some bending back and forth to keep the molecule compact. In myoglobin (Figure 20-25), the 153 amino acids in the protein chain are coiled into eight lengths of α helix (lettered A through H), which then are folded back and forth to make a compact molecule. Helices E and F form a pocket into which the heme group fits, and the oxygen molecule binds to the iron of this heme. Hemoglobin is constructed along similar lines and has four such myoglobinlike units (Figure 20-26). The small protein cytochrome *c* (Figure 20-23) has less room for α helices. Its 103 amino acids are spun around its heme group like a cocoon, leaving only one edge exposed. In larger enzymes such as trypsin (223 amino acids) and carboxypeptidase (307 amino acids), there are regions in the center of the molecule where the protein chain zigzags back and forth in several parallel and antiparallel strands held together by hydrogen bonds very much as in silk.

The purpose of an enzyme is to provide a surface to which its *substrates* (the molecules acted upon) can bind, and to facilitate the formation or rupture of bonds in these molecules. The site on the surface of the enzyme where these activities take place is called the **active site**. The enzyme has two functions: **recognition** and **catalysis**. If it indiscriminately bound every molecule that came along, it would spend only a small proportion of its time catalyzing the reaction that it was supposed to catalyze. Conversely, even if it bound the right molecules, it would be useless if it could not assist in making or breaking the proper bonds. Enzymes recognize their true substrates by having at their active site properly positioned amino acid side chains that can interact with the substrate molecule by way of charge interaction, hydrogen bonding, or the attraction of hydrophobic groups. This selection of molecules that an enzyme will or will not bind is called its **specificity**.

Once bound, the substrate is subject to attack by groups on the enzyme. Many enzymes involved in bond-breaking reactions use metals such as Zn, Mg, Mn, or Fe. Sometimes one part of the substrate will coordinate to the metal; in other cases, the metal draws electrons from the substrate and weakens a bond. Both are illustrated in the catalytic action of trypsin discussed in the next section.

Only recently have we known the molecular structures of proteins. The first x-ray analysis of a protein, that of myoglobin, was completed in 1959. That of the first enzyme, lysozyme, was accomplished in 1964. Research on larger enzymes, electron carriers, and antibodies is progressing rapidly. We now know the detailed molecular framework of more than 90 proteins. Biochemistry, in these areas, merges imperceptibly with its sister field of molecular biology.

essence of catalysis is that the catalyst provides an easier, and hence faster, pathway for reaction, but comes through the process unscathed at the end. The enzyme is restored in a second step by the introduction of a molecule of water:

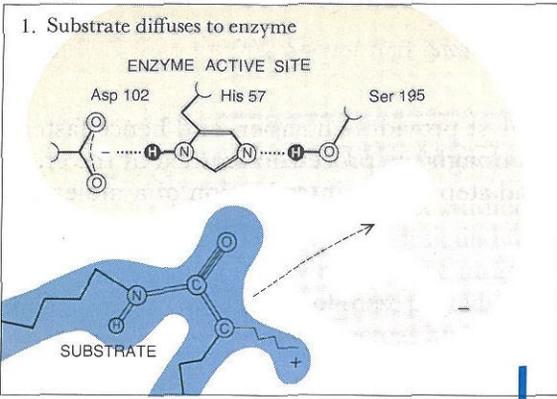


The left half of the original substrate molecule falls away, the serine side chain on the enzyme is restored, and the enzyme is ready to repeat the process. Notice that equation 21-3 is the reverse of equation 21-2, but with a water molecule substituting for the right half of the substrate chain.

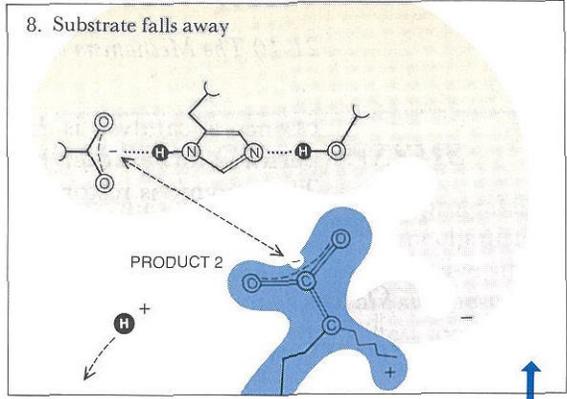
Nothing happens in enzymatic catalysis that could not take place without the catalyst. A water molecule *could* come up to the protein chain in equation 21-1 and split it apart, donating its —OH to the left half and its H— to the right half of the separated chain. The activation-energy barrier to this direct reaction would be formidable, however, so the reaction itself would be very slow. The two-step reaction assisted by the serine side chain on the enzyme does not require as high an energy for any intermediate form, hence the reaction takes place more quickly. Since the effect of the enzyme is a lowering of the activation-energy barrier, the reverse reaction is catalyzed to the same extent. The same ultimate equilibrium conditions are reached with or without the catalyst; the enzyme only makes the approach to equilibrium more rapid.

The part that the enzyme plays in making the reaction easier is better seen in the eight-step diagram of Figure 21-18. The first four steps, acylation of the enzyme, correspond to equation 21-2, and the last four, deacylation, to equation 21-3. Besides the key serine side chain at the active site, the serine proteases also have histidine and aspartic acid side chains that are directly involved in the catalytic mechanism. Before the substrate chain binds, the serine is hydrogen bonded to the histidine, which in turn is hydrogen bonded at the other side of its five-membered ring to the aspartic acid (step 1). In step 2, the protein chain to be cut binds to the active site of the enzyme, which is especially shaped to receive it. In step 3, the serine oxygen

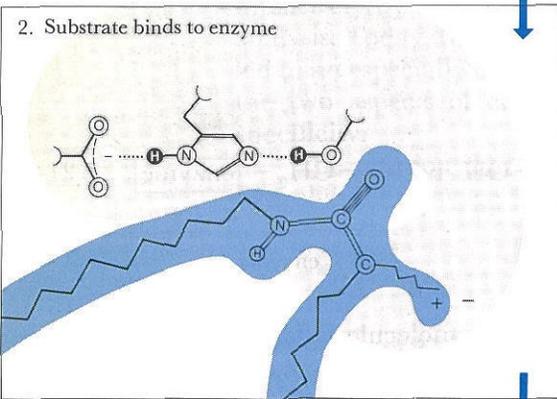
1. Substrate diffuses to enzyme



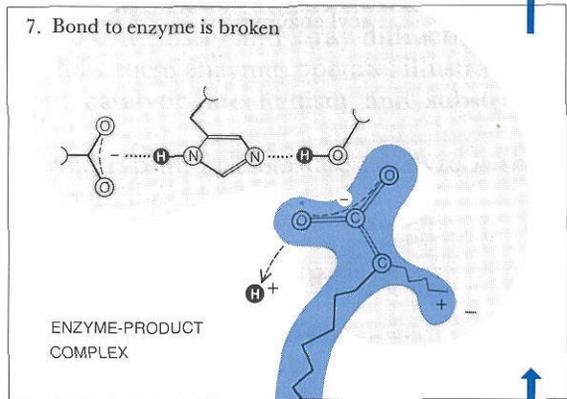
8. Substrate falls away



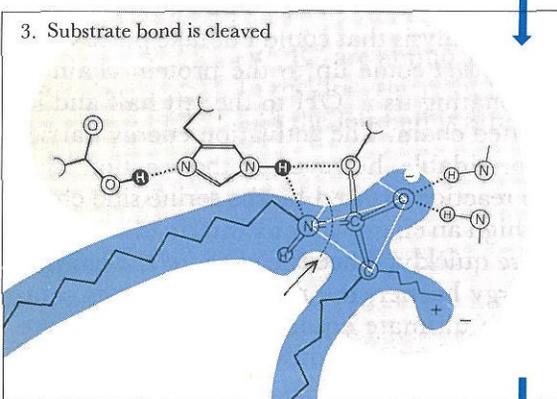
2. Substrate binds to enzyme



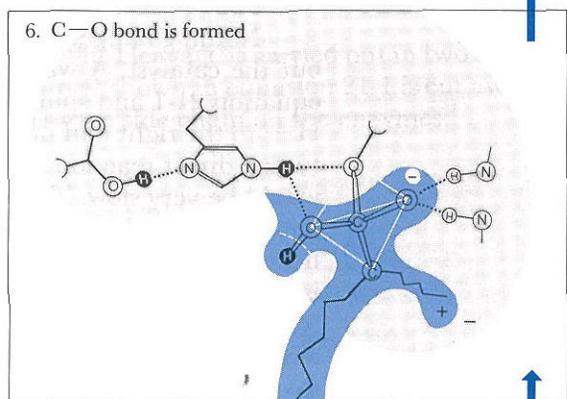
7. Bond to enzyme is broken



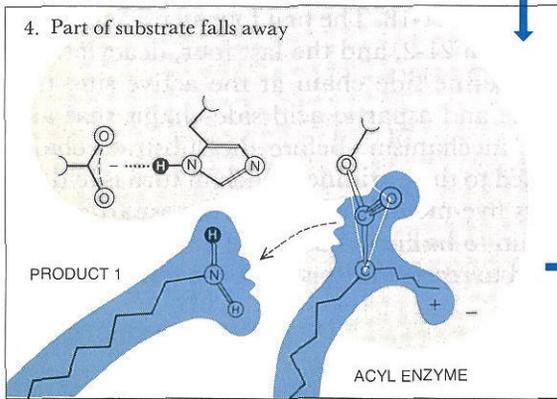
3. Substrate bond is cleaved



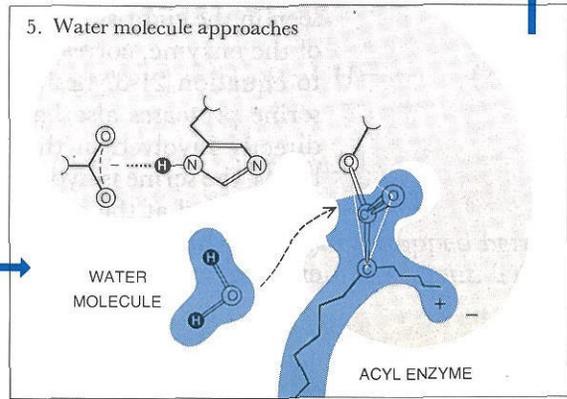
6. C—O bond is formed



4. Part of substrate falls away



5. Water molecule approaches



has passed its hydrogen to the histidine nitrogen, and formed a bond with the carboxyl carbon of the substrate chain. This carbon now has a tetrahedral arrangement of bonds to four other atoms. The carbonyl double bond to oxygen at the right has become a single bond, and the extra electron pair has been pushed onto the carbonyl oxygen, giving it a negative charge. The enzyme helps to stabilize this negative charge by means of two amide hydrogen atoms from the enzyme protein chain. (Remember that the NH in a protein chain has a partial positive charge.) Step 3 is called the **tetrahedral intermediate**.

At the same time that the bond between serine oxygen and carbonyl carbon is being formed, the bond between the carbonyl carbon and the amide nitrogen is being weakened, and this weakening is assisted by the nearby presence of the formerly serine hydrogen atom, now on the histidine nitrogen. As the peptide N—C bond breaks, this hydrogen attaches to the N to make a completed —NH₂ group on the end of the departing chain, labeled product 1 in step 4. Half of the substrate chain has now fallen away, and the other half is attached to the serine side chain of the enzyme. Bonding around the carbonyl carbon atom once again is trigonal and planar, with a C=O double bond.

Steps 5 through 8 (equation 21-3) are just the reverse of steps 1 through 4, with the water molecule replacing product 1. The water molecule donates its —OH to a new tetrahedral intermediate (step 6) and its H— to the histidine, which then passes it on to the serine (step 7) as the acyl bond is broken. The second half of the substrate chain, or product 2, is then free to fall away, and the enzyme is ready for another cycle (step 8).

So far nothing has been said about the specificity of the enzymes. If trypsin, chymotrypsin, and elastase use identical catalytic mechanisms, then how do they differ from one another? The answer is that they are selective about the kind of side chain next to which they cleave a peptide bond. The critical side chain is marked R' in equations 21-1 through 21-3, just before the carbonyl group of the bond to be cut. Each of the three enzymes has a pocket on its surface into which this side chain is inserted when the substrate binds. The specificity pocket in trypsin is long and deep, with a negative charge from an ionized aspartic acid at the bottom (Figure 21-19a). Hence trypsin favors cleavage next to the positively charged basic side chains of lysine or arginine. In chymotrypsin the specificity pocket is wider (Figure 21-19b) and is lined entirely with hydrophobic side chains, so chymotrypsin favors cleavage next to a bulky aromatic side chain such as phenylalanine or tryptophan. In elastase, the specificity pocket is quite shallow (Figure 21-19c). Elastase is less selective than the other two enzymes, but tends to favor cleavage next to small side chains such as alanine or valine.

◀ **Figure 21-18**

Catalytic mechanism of the enzymes trypsin, chymotrypsin, and elastase. See text for explanation. (Drawing by Irving Geis, from R. M. Stroud, "A Family of Protein-Cutting Proteins," *Scientific American*, July 1974. Copyright Stroud, Dickerson, and Geis.)

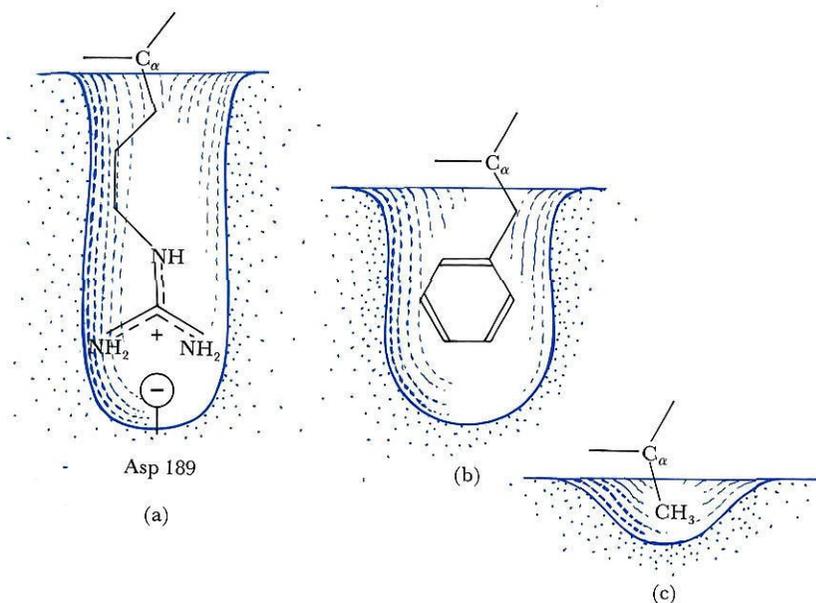


Figure 21-19

Specificity pockets of trypsin (a), chymotrypsin (b), and elastase (c). The size of each pocket, and the nature of the side chains lining it, determine what kind of amino acid chain it will hold best. This in turn determines at which position along a substrate chain cleavage will occur. (From Dickerson and Geis, *Chemistry, Matter, and the Universe*, W. A. Benjamin, Menlo Park, Calif., 1976.)

Hence both the catalytic mechanism and the substrate specificity can be explained by the folding of the polypeptide chain of the enzyme and the positioning of side chains along it. The actual folding of the protein chain in trypsin can be seen in Figure 21-20. The enzyme is built from one continuous polypeptide chain of 223 amino acids. (The conventional chain numbering has gaps and insertions to harmonize with chymotrypsin and elastase.) The molecule is roughly spherical and 45 Å in diameter, with a bowl-shaped depression or dimple on one side for the active site. The essential aspartic acid, histidine, and serine in the active site are shown by black stippling in Figure 21-20. The protein chain to be cut is shown in color with black outlining, and the arrow indicates the bond to be cleaved. The dashed lines at both ends of the substrate indicate that the chain stretches on for a considerable length in both directions. The specificity pocket for side chain R' is indicated in light-colored stippling, and since the molecule illustrated is trypsin, an arginine side chain is shown inserted into the pocket, attracted by the negative charge on aspartic acid 189 at the bottom of the pocket.

When enzyme chemists of only a generation ago spoke of the “mechanism” of enzymatic catalysis, they could only do so figuratively, even though they had a literal meaning in mind. Now we can diagram a mechanism that is every bit as literal and mechanical as the mechanism of a combination



Figure 21-20

Main-chain skeleton of the trypsin molecule. The α -carbon atoms are shown by shaded spheres, with certain of them given residue numbers for identification. For simplicity, the connecting $-\text{CO}-\text{NH}-$ amide groups are represented only by straight lines. A portion of the polypeptide chain substrate appears in dark color with black outline. The specificity pocket is shaded, with an asparagine side chain from the substrate inserted. The catalytically important Asp, His, and Ser are poised for cleavage of the peptide bond, as marked by an arrow. (Drawing by Irving Geis, from R. M. Stroud, "A Family of Protein-Cutting Proteins," *Scientific American*, July 1974. Copyright Stroud, Dickerson, and Geis.)

lock. Although we may not yet know the mechanism of a particular enzyme of interest, there is no longer anything mysterious about the basis of enzymatic catalysis, or the principles involved. Enzymes are superb examples of molecular engineering.

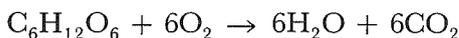
21-11 ENERGY AND METABOLISM IN LIVING SYSTEMS

To mountaineers who take their hobby seriously, a particularly challenging operation is a “dynamic traverse.” This is a traverse across a difficult piece of terrain where, at each instant, the climber is in an unstable situation, and where he is prevented from a disastrous fall only by his momentum. In a sense, every living organism is engaged continually in a dynamic traverse. One of the most important generalizations in science, the second law of thermodynamics, states that in any process taking place in a closed system (i.e., the object studied plus its entire environment with which it exchanges matter or energy), the disorder of the system as a whole *increases* (Chapter 16). A living organism is an intricate chemical machine, evolved to a high level of complexity. It is faced constantly with the dilemma that every chemical reaction that takes place inside it increases the disorder and reduces its complexity. A constant supply of energy is needed from outside sources, not only to provide the power to do physical work, but also to keep down the level of disorder, or *entropy*, within. If this supply of energy fails, then death and the breakdown of the chemical machine is only a matter of time: a day for the fast-living shrew, a few weeks for man. Just as momentum saves the climber from falling, so a constant influx of energy keeps the living machine from collapsing. The degradation of high-energy fuels and the extraction of their energy is called **metabolism**. In this section we shall trace briefly the outlines of the metabolism that is common to all oxygen-using living organisms. Since glucose is so important a metabolite, we shall use it as an illustration.

The Combustion of Glucose

If 180.16 g of glucose, or 1 mole, are burned, 2816 kJ of heat are liberated. Such an uncontrolled combustion is wasteful; only a small fraction, if any, of the energy stored in glucose is put to good use. It is more efficient to feed glucose to horses and use them to pull a load than it is to burn glucose and operate a locomotive with it. This is because in the horse’s metabolism glucose is broken down in a series of small steps. The energy released at each step is stored in the chemical bonds of a special molecule, adenosine triphosphate (ATP), and is available for use in other chemical reactions that make muscles do work. Combustion in the horse is controlled and efficient; combustion in the locomotive is less controlled and wasteful.

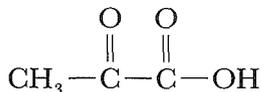
The standard free energy change in the combustion of glucose,



is $\Delta G^0 = -2870 \text{ kJ mole}^{-1}$. This represents the chemical drive available from glucose combustion. Of this, 2816 kJ come from the heat of combustion ($\Delta H^0 = -2816 \text{ kJ mole}^{-1}$) and 54 kJ come from the increased disorder of the products ($\Delta S^0 = +181 \text{ J K}^{-1} \text{ mole}^{-1}$, and $-T \Delta S^0 = -54 \text{ kJ mole}^{-1}$). In the horse, 40% of the liberated free energy is saved by using it to synthesize 38 molecules of ATP for every molecule of glucose. It is this orderly breakdown process that we shall examine.

The Three-Step Process in Metabolic Oxidation

There are three parts to the combustion process in oxygen-using organisms. In the first part, **anaerobic fermentation** or **glycolysis**, all foods, no matter what their chemical nature, are degraded to pyruvic acid,

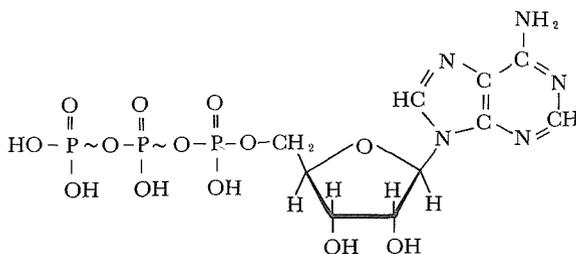


Little energy is obtained from this process. Its main purpose is to reduce everything to a common set of chemicals and prepare for the real energy-producing steps. In the second part, called the **citric acid cycle**, pyruvate is oxidized to CO_2 , and the hydrogen atoms from pyruvate are transferred to the carrier molecules NAD^+ (nicotinamide adenine dinucleotide) and FAD (flavin adenine dinucleotide). Again, only a small amount of free energy is stored as ATP in this cycle. Its principal purpose is to break up the $1142 \text{ kJ mole}^{-1}$ of free energy in pyruvate into four smaller and more easily handled packages, of around 220 kJ mole^{-1} each, in the form of 4 moles of reduced carrier molecules. The third part of the process, the **respiratory chain**, accepts these reduced carrier molecules. It reoxidizes them, uses the hydrogen atoms obtained during the oxidation to reduce O_2 to water, and uses the free energy obtained to synthesize ATP.

We can see two objectives to this three-part machinery: to reduce the thousands of different possible foods to a common set of chemical reactions as rapidly as possible, and to break the inconveniently large packages of free energy into several smaller ones that can be handled by the machinery for synthesizing ATP. Let us now look more closely at each of these three parts.

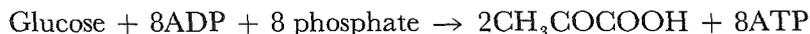
Step 1: Glycolysis

The first part of this combustion process does not require oxygen. It is common to all living organisms, and is known as anaerobic fermentation or glycolysis (“glucose breakdown”). If O_2 is present, the end product is pyruvic acid, as we have stated. But in other organisms that do not use

**Figure 21-21**

The structure of adenosine triphosphate (ATP). The bonds marked by wavy lines in the phosphate groups liberate an unusually large amount of energy when they are cleaved by hydrolysis, and are the means by which chemical energy is stored in the ATP molecule.

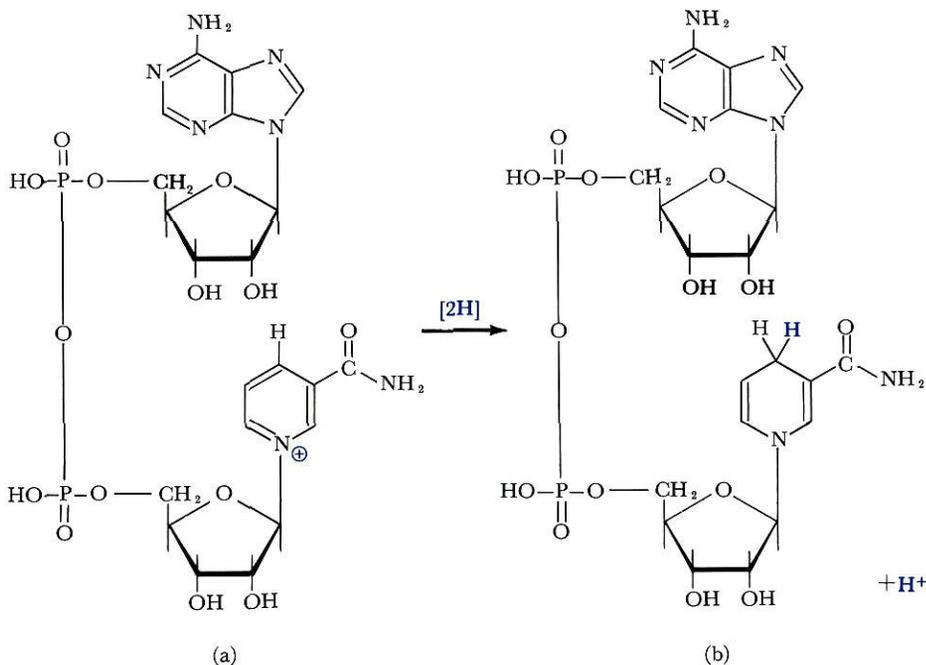
principal means of energy storage in any living cell. Every time a molecule of glucose is degraded biochemically to two molecules of pyruvate, eight molecules of ATP are formed from 8ADP:



This results in the storage of $8 \times 30.5 = 244 \text{ kJ mole}^{-1}$ of free energy. The enzymes that control all of the steps of the breakdown ensure that the energy released at a step of this process is used to synthesize an ATP molecule rather than wasted as heat.

There are two other carriers that we should examine before going on to the citric acid cycle. One is nicotinamide adenine dinucleotide, whose structure is shown in Figure 21-22. It resembles ATP in having an adenine group, ribose, and phosphate. However, the essential part is a nicotine ring that can be reduced and oxidized. This molecule is a redox carrier. When a metabolite is oxidized at one step in the citric acid cycle, the oxidized form of nicotinamide adenine dinucleotide, NAD^+ , can accept two H atoms and be reduced to NADH and H^+ . The other important carrier is FAD or flavin adenine dinucleotide, which is reduced to FADH_2 . Both of these carriers feed into the last production line of the energy storage factory, the terminal oxidation chain or respiratory chain. This is a four-step pathway, involving the cytochrome enzymes, in which the reduced electron carriers, NADH and FADH_2 , are reoxidized. In this process, oxygen is reduced to water, and the energy released is stored as ATP. Each time a reduced carrier molecule is reoxidized, the energy released by this oxidation is conserved by the synthesis of several molecules of ATP.

Many of the essential vitamins are the semifinished components for energy carriers such as these. Small amounts of these vitamins must be supplied from outside sources because we have lost the ability to synthesize them ourselves. For example, the chemical group that is oxidized and reduced in FAD is a flavin. Some organisms can synthesize flavins; perhaps

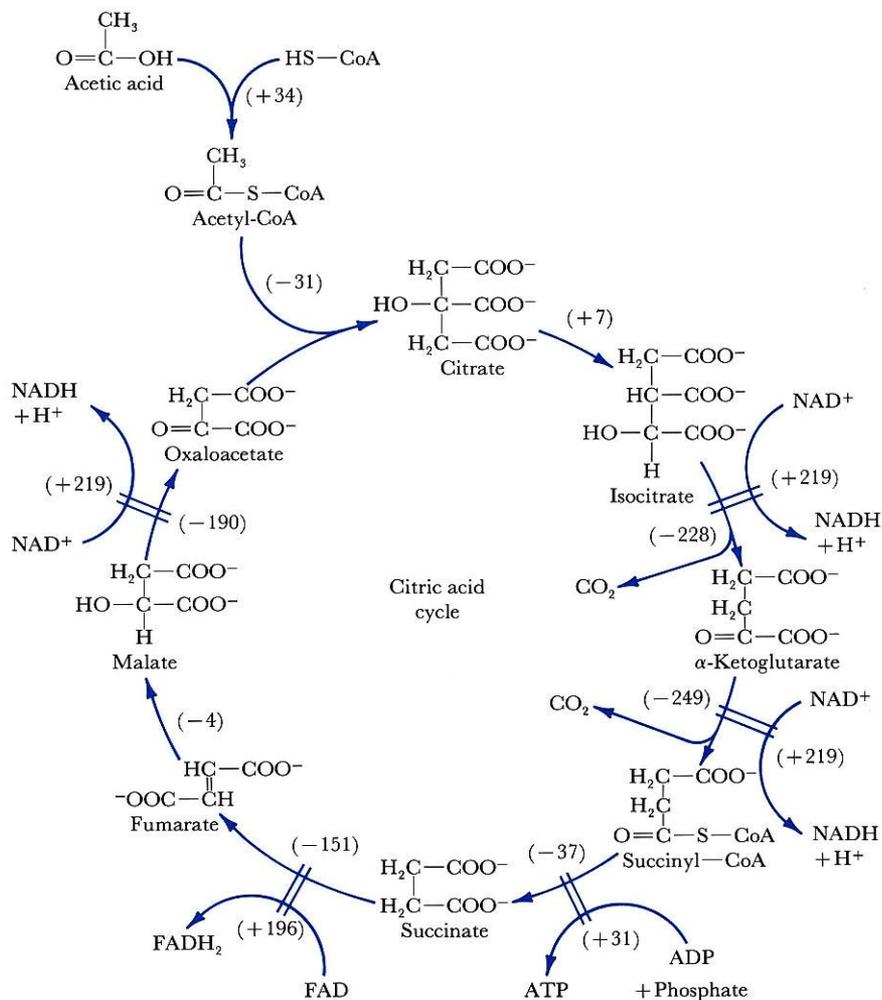
**Figure 21-22**

The structure of nicotinamide adenine dinucleotide (a) oxidized, NAD⁺, and (b) reduced, NADH. Note the similarity between the top half of the molecule as drawn here and ATP. Reduced NADH is a carrier molecule that can pass on its stored energy in chemical syntheses in a similar way as reduced ferredoxin does in photosynthesis.

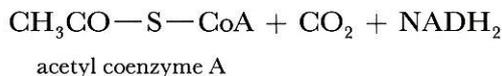
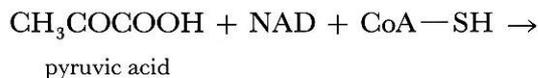
we or our ancestor species also could synthesize them at one time, but we cannot do so now. To keep the energy-transfer system operating and to replace the FAD molecules as they are gradually worn out, we need small amounts of *riboflavin*, or vitamin B₂, supplied from external sources (Figure 21-3). Similarly, the nicotinamide ring of NAD⁺ is made from nicotinic acid, or *niacin*, which we must obtain in small quantities in our diet.

Step 2: Citric Acid Cycle

Now we are ready to look at the second part of the three-part energy-retrieval process of metabolism. The main role of this part, the citric acid cycle, is to convert the more than 1050 kJ mole⁻¹ of free energy contained in pyruvic acid molecules into four packages of 220 kJ mole⁻¹, which are in the form of reduced NADH and FADH₂. The “primer” step before the cycle begins is the combination of pyruvic acid with a molecule called reduced coenzyme A (CoA—SH) to form acetyl coenzyme A. Acetyl coenzyme A is the raw material for the citric acid cycle, which is diagrammed in Figure 21-23:

**Figure 21-23**

The citric acid cycle, also called the Krebs cycle or the tricarboxylic acid cycle. The numbers in parentheses are the standard free energies in kilojoules for the reactions shown. A double bar indicates that the oxidation in the cycle and the reduction of a carrier molecule are made to occur together by an enzyme. The reduced carrier molecules feed into the terminal oxidation chain, where they are oxidized again, and O_2 is reduced to H_2O .



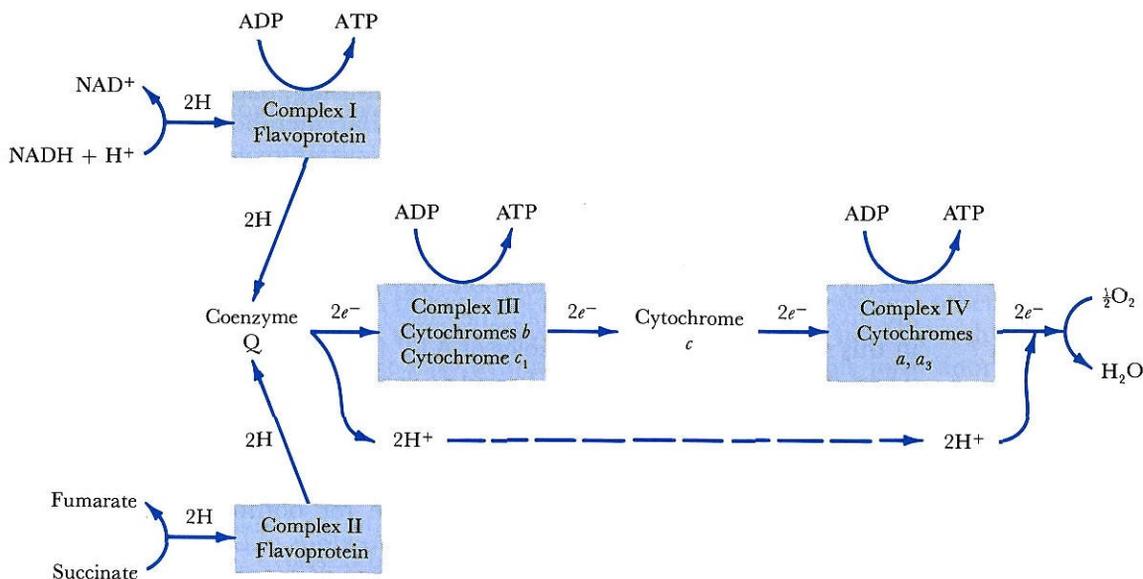
In the cycle, the two-carbon acetate group is combined first with four-carbon oxaloacetate to make the six-carbon citrate ion. Then citrate is degraded in seven steps to release two of its carbon atoms as CO_2 and to restore the oxaloacetate again. Each of the steps in the citric acid cycle is either an oxidation (isocitrate to α -ketoglutarate, malate to oxaloacetate) or a rearrangement in preparation for the next oxidation (citrate to isocitrate). In four of the oxidation steps, the liberated energy is saved by using it to reduce a carrier molecule: NAD^+ or FAD .

Step 3: Terminal Oxidation or Respiratory Chain

Each NADH molecule, no matter what its origin, funnels into the third part of the metabolic process, the terminal oxidation chain or respiratory chain, and produces three molecules of ATP . Each FADH_2 arrives midway through the same chain and produces only two ATP . The respiratory chain is a series of flavin-containing proteins and cytochromes (Figure 20-23), along which the hydrogen atoms and electrons from NADH and FADH_2 are passed, until they ultimately reduce O_2 to H_2O . The components of the respiratory chain are shown in Figure 21-24. When NADH is reoxidized, the two hydrogens are used to reduce a flavoprotein, and the free energy is saved by synthesizing a molecule of ATP from ADP and phosphate. The flavoprotein is reoxidized by reducing a small organic quinone molecule, known as ubiquinone or coenzyme Q . At this point, the electrons and protons of the reducing hydrogen atoms go their separate ways. The electrons are used to reduce the iron atom in a cytochrome b from Fe^{3+} to Fe^{2+} , and the protons go into solution. Cytochrome b reduces a cytochrome c_1 , and the electrons go from c_1 to c to a to a_3 , all of these being iron-containing cytochrome molecules related to the molecule shown in Figure 20-23. Two more molecules of ATP are synthesized along the way. Finally, the electrons from cytochrome a_3 and the protons from solution recombine to reduce an atom of oxygen to H_2O .

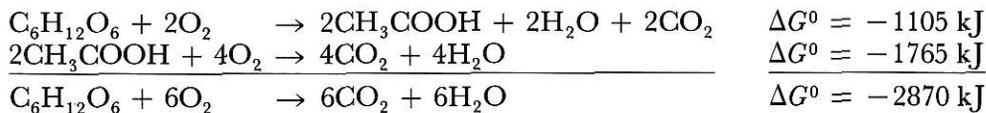
When succinate is oxidized to fumarate in the citric acid cycle (Figure 21-23), the two H atoms are given to a molecule of FAD , which is actually bound to the enzyme succinate dehydrogenase. From there the hydrogen atoms are passed to the same pool of coenzyme Q molecules used by NADH , and the process continues, but with only two ATP produced.

With every NADH molecule reduced during the citric acid cycle being “worth” three ATP , and every FADH_2 worth two, you can keep score on energy storage during one turn of the cycle. During the degradation of 1 mole of acetyl coenzyme A, 883 kJ of free energy are released, and 12 ATP -equivalents are saved: $12 \times 30.5 \text{ kJ} = 366 \text{ kJ}$. In the total degradation of 1 mole of glucose—glycolysis to 2 moles of pyruvate, conversion to 2 moles of acetate, and two turns of the citric acid cycle—the balance sheet looks like this (for simplicity, the conversion of pyruvate to acetyl CoA has been included in glycolysis):

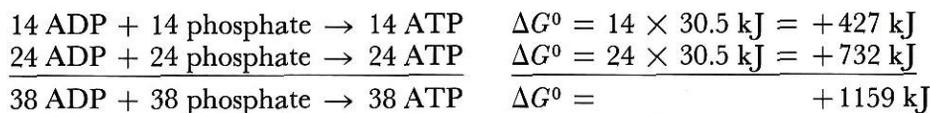
**Figure 21-24**

The terminal respiratory chain. The components are organized on the inner surface of the inner mitochondrial membrane into four macromolecular complexes, containing cytochromes, flavoproteins, and other non-heme iron proteins. Coenzyme Q, or ubiquinone, and cytochrome *c* act as conduits from one complex to the next. Reduction is by transfer of hydrogen atoms until coenzyme Q is reached, after which it occurs by transfer of electrons, with protons going into solution. Electrons and protons are reunited at the end when oxygen is reduced to water. Free energy is stored as ATP at three of the four complexes.

Degradation



Synthesis



The overall efficiency of energy conversion from glucose as food to ATP stored in the muscles is $1159/2870 = 0.40$.

Anaerobic fermentation or glycolysis, the citric acid cycle, and the respiratory chain are part of the common heritage of all life on earth above the level of bacteria. Some aerobic, or oxygen-using, bacteria also have this same machinery to oxidize glucose or similar metabolite completely to carbon dioxide and water. Other anaerobic, or non-oxygen-using, bacteria make do only with fermentation: ingesting glucose or other energy-rich molecules, breaking them down into smaller molecules such as propionic acid, acetic acid, or ethanol, and using the relatively small quantity of free energy that is released. The anaerobic way of life is wasteful; when yeasts are given a plentiful supply of air they use the entire machinery described in this section to produce 38 moles of ATP from each mole of glucose, but when deprived of oxygen and forced to ferment glucose anaerobically to ethanol, they obtain only 2 moles of ATP per mole of glucose.

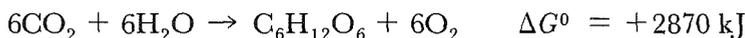
If O_2 respiration yields 19 times as much energy per gram of food as anaerobic fermentation, then why should *any* life form exist in an anaerobic lifestyle? The bacteria that are responsible for botulism in improperly canned foods, or gangrene in wounds, are two kinds of *Clostridia*, which not only cannot use oxygen, but are poisoned by it. Why are they not oxygen-breathers like the rest of us?

The answer probably lies in the way that life evolved on earth. The primitive earth is believed to have had a reducing atmosphere, with gases such as H_2 , CH_4 , NH_3 , H_2O , and H_2S , but little or no free O_2 . Under these reducing conditions, organic molecules that were formed by nonbiological means would not be degraded by oxidation as they are today, but would accumulate over the millennia. The first life forms were probably scavengers, evolving from this chemical "soup" in the oceans, and obtaining energy by degrading naturally occurring high-free-energy compounds. The *Clostridia* and their relatives today most likely are living fossils: descendants of these early anaerobic fermenters that retreated into the odd anaerobic pockets of the world when the atmosphere as a whole accumulated large quantities of free O_2 and acquired an oxidizing character.

What caused the atmosphere to change so drastically? The changeover appears to be a by-product of the invention of a new kind of energy trapping, photosynthesis, which gave its possessors an enormous advantage over the purely fermentative energy-scavengers. The organisms that developed this new talent could use the energy of sunlight to make their own high-free-energy molecules, rather than depending on what they could find in their surroundings. They became the ancestors of all green plants. Living organisms today are divided metabolically into two categories: those that can make their own food from sunlight, and those that cannot. Since the creatures in the second category subsist by eating those in the first, energy storage by means of photosynthesis is the driving source that makes all life on earth possible.

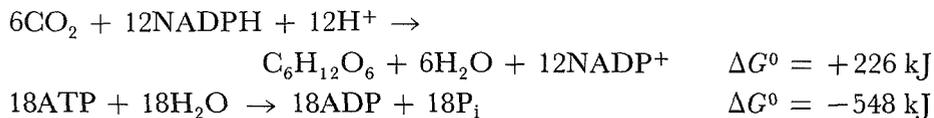
Winding the Mainspring of Life: Photosynthesis

The overall reaction of green-plant photosynthesis is the reverse of that of combustion of glucose:



Water is split apart as the source of hydrogen atoms to reduce CO_2 to glucose, and the unwanted oxygen gas is released into the atmosphere. The energy for this highly nonspontaneous process comes from sunlight. The earliest kinds of bacterial photosynthesis used substances other than water as sources of reducing hydrogens— H_2S , organic matter, or hydrogen gas itself—but the easy availability of water made this form enormously successful, and it is used today by all algae and green plants. The simplest organisms that have O_2 -liberating photosynthesis are the blue-green algae, which should more properly be called by their modern name of *cyanobacteria*, since they are indeed bacteria that have learned the trick of making their own food with CO_2 , H_2O , and sunlight.

Green-plant photosynthesis can be divided into two separate processes: the **photo** reactions and the **synthesis** reactions, or as they are more commonly known, the **light** reactions and the **dark** reactions. In the dark reactions, CO_2 is reduced to glucose using hydrogen atoms from NADPH (NADP⁺ is NAD⁺ with a phosphate on one ribose —OH), with ATP as driving energy:

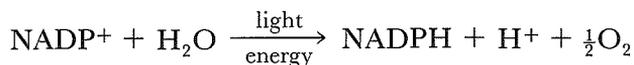


(P_i is a shorthand notation for inorganic phosphate.) The first reaction by itself would be nonspontaneous by 226 kJ per mole of glucose, but the second reaction provides energy to spare, giving the process a net spontaneous drive of 322 kJ. The dark reactions are indifferent as to the source of the NADPH and ATP to keep them running. Although the source in green plants today is sunlight via the light reactions, it may be that the dark reactions are older, and originally were driven by NADPH and ATP from a different source. The machinery for the dark reactions is known as the **Calvin–Benson cycle**, and is similar in logic to the citric acid cycle. Carbon dioxide is combined first with a carrier molecule, ribulose diphosphate. After a series of steps, some of which look like glycolysis run in reverse, part of the organic material is ejected as glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, while the remainder is used to make ribulose diphosphate for the next turn of the cycle.

The driving energy for the Calvin–Benson cycle comes from the light reactions. The light-absorbing elements are molecules of chlorophyll (Fig-

ure 20-21), a conjugated ring of carbon atoms with delocalized electrons surrounding a magnesium atom. One type of chlorophyll molecule sits at the photocenter or trap where chemical reaction actually occurs, but other chlorophylls and related conjugated molecules surround the photocenter and act as “antennas,” absorbing photons of light and passing electronic excitation on to the photocenter molecules.

The light energy absorbed by chlorophyll is used to drive an inherently unlikely reaction: reduction of NADP^+ using *water* as the reducing agent:



Oxygen gas is released into the atmosphere as a by-product. Energy also is stored by making ATP, and this ATP and NADPH provide the fuel for the dark reactions.

A more detailed comparison of glycolysis, respiration, and photosynthesis, and the various enzymes that control them, suggests that the energy-managing machinery of life evolved on our planet in a series of steps:

1. Anaerobic fermentation, especially glycolysis. The first living organisms were scavengers, obtaining energy by breaking down high-free-energy molecules that were not biologically formed into smaller molecules, without oxidizing them. During this distant era, the atmosphere of our planet was reducing in character, with no free oxygen.

2. Dark reactions for glucose synthesis, leading eventually to the Calvin–Benson cycle. Any organism that could use external energy sources to make its own high-free-energy molecules for later use would be at an enormous advantage over its less talented brothers. The first external energy source need not have been solar radiation.

3. Light reactions for production of ATP and NADPH to drive the dark reactions. Sunlight was the most widespread and abundant energy source on the primitive earth, as it is today. The by-product of the light reactions as carried out by green plants is free O_2 , and this, poured into the atmosphere for more than a billion years, gradually changed our atmosphere from reducing to what we know today, with 20% free oxygen.

4. Oxygen respiration. With free O_2 available, it was perhaps inevitable that some line of organisms would learn how to use it to extract far more energy from a given amount of food. The advantages of combustion with O_2 were so great that the vast majority of life forms—plant and animal—use O_2 respiration today.

As far as we can learn, the planet itself was formed approximately 4.6 billion years ago, and simple fermenting one-celled life forms were already present 3.5 billion years ago. They may have been photosynthetic as early as 3.1 billion years ago, but geological evidence about the oxidation state of sedimentary iron deposits indicates that the atmosphere only became oxidizing around 1.8 to 1.4 billion years ago. Multicelled life forms, which

perhaps were dependent on the energy surplus that only O₂ respiration could provide, made their appearance roughly between 1000 and 700 million years ago, and the pattern for the later evolution of higher organisms was set. The most revolutionary single step, after the evolution of life itself, was the tapping of an extraplanetary energy source, the sun. This step turned life from a thinly spread scavenger of naturally occurring high-free-energy molecules into a force capable of transforming the surface of the planet and ultimately leaving it altogether. Whether, from a planetary standpoint, life proves to be a beneficial or a pathological phenomenon remains to be seen.

Summary

We have come a long way in this chapter from speculating about the relative chemistries of B, C, N, and Si. Carbon undoubtedly has a special role, conferred upon it by its adequate supply of electrons for orbitals, absence of repelling lone electron pairs, and ability to make double and triple bonds. The simple **alkanes**, or single-bonded compounds of hydrogen and carbon, illustrate the diversity of compounds that carbon can build because of its ability to make long, stable chains. The **alkyl halides** are the bridges from the relatively unreactive alkanes to the wealth of hydrocarbon derivatives: **alcohols**, **ethers**, **aldehydes**, **ketones**, **esters**, **acids**, **amines**, **amino acids**, and others that have not been discussed in this chapter. The ability of carbon to make double and triple bonds has been seen in the **alkenes** and **alkynes**, and is especially important in **conjugated** and **aromatic** molecules.

The aromatic compounds are special because a certain number of their electrons are delocalized, or not confined to the region between any two bonded atoms. In general, the greater the delocalization, the more stable the molecule will be. Benzene is 166 kJ mole⁻¹ more stable than a calculation of the bond energies of a Kekulé structure would lead one to expect. Moreover, the first unfilled orbitals in many delocalized-electron molecules are lower and closer to the filled orbitals than in other molecules. The energy required to excite an electron often falls in the visible part of the spectrum, making these compounds colored. The ability to absorb visible light is crucial in molecules of chlorophyll, because the energy of the light absorbed is used to synthesize glucose. Without photosynthesis, life on this planet would have remained a rare and irrelevant scavenger, breaking down once again those organic molecules that had been synthesized naturally by non-biological means.

The molecule that is synthesized as an energy-storage agent in photosynthesis is **glucose**, a **carbohydrate**. Besides being storehouses for chemical energy, carbohydrates are important structural materials in plants: wood, cotton, woody stalk tissue in softer plants. Glucose is polymerized into **cellulose**, which is the basis of the structural carbohydrates, and cannot be

redigested, and **starch**, which is stored in seeds, grains, and roots, and is destined to be degraded later for its glucose.

When needed as an energy source, the glucose is not burned in one step. Instead, it is degraded in a series of more than 25 separate steps. During many of these steps the energy released is saved by synthesizing molecules at ATP. **Anaerobic fermentation** or **glycolysis** provides the initial breakdown of glucose to pyruvate, and the **citric acid cycle** completes the oxidation of carbon to CO_2 . The hydrogens are transferred to carrier molecules, NAD^+ and FAD . These are reoxidized in the **respiratory chain**, where energy is stored as more ATP, and the hydrogen atoms are used to reduce O_2 to H_2O .

Photosynthesis is the effective reverse of the foregoing processes. The **dark reactions** use NADPH and ATP to reduce CO_2 to glucose, and the **light reactions** employ the energy of absorbed photons to generate the necessary NADPH and ATP.

Proteins are polymers of amino acids. The **fibrous** proteins have structural roles in hair, skin, nails, muscle, and tendons. In these structures the protein chains are coiled into cables or cross-connected by hydrogen bonds into sheets. The **globular** proteins include enzymes, carriers, and antibodies. In these the chains may be in helices or sheets, but these substructures are folded back on themselves to form compact, discrete molecules.

The most important class of globular proteins is made up of the biological catalysts or **enzymes**. These are characterized by a **catalytic mechanism**, by which they hasten the approach of a given reaction toward thermodynamic equilibrium, and by a **substrate specificity**, by which they choose between potential substrate molecules, acting on some and refusing to accept others. The region on the surface of the enzyme where catalysis occurs is called the **active site**. The mechanism of catalysis may involve charged groups, electron or hydrogen atom donors and acceptors, and metal atoms at the active site. Enzymatic specificity arises from the shape of the surface of the enzyme, and interactions with the substrate such as hydrogen bonding, charge interaction, and hydrophobic attractions. The enzyme and its substrate are chemically tailored for one another.

Trypsin, chymotrypsin, and elastase are three related enzymes that share the same catalytic mechanism but have different substrate specificities. They act on a protein chain substrate by forming an acyl intermediate between a serine on the enzyme and half the substrate chain, and then cleaving away this half of the chain with a water molecule. The specificity arises from a pocket adjacent to the active site, into which a substrate side chain fits.

A living organism as found on earth (and this is the only kind we know of now) is a complex collection of carbon-based molecules, tailored by evolution for survival, and using solar energy directly or indirectly as a means of driving otherwise unfavorable reactions and keeping the entropy within the organism low. As long as this can be maintained, the organism survives. When the machinery ultimately breaks down, the individual organism

goes over into that state of low energy and high entropy that traditionally is known as death.

Self-Study Questions

1. What are hydrocarbons, and what size-range of hydrocarbon molecules is found in oils, paraffin waxes, gasolines, heating gas, and kerosene?
2. Illustrate branched chains in (a) hydrocarbons and (b) a synthetic fiber. How does cross-linking affect the physical properties of the fiber or plastic?
3. Why are hydrocarbons with double bonds between carbons referred to as “unsaturated” hydrocarbons?
4. Why is the chemical bonding in boron compounds so different from that for carbon?
5. In what way does the chemical bonding in B_2H_6 differ from that in C_2H_6 ? How is this evidenced in the geometry of the two molecules?
6. What is a three-center bond in boron compounds? How many electrons are used, and how many atoms are held together?
7. Does the three-center bond violate the principle, enunciated in Chapters 12 and 13, that a combination of a given number of atomic orbitals leads to the *same number* of molecular orbitals? In a three-center bond, how many molecular orbitals remain unfilled by electrons?
8. The carbon hydrides, or hydrocarbons, are straight-chain, cyclic, or branched-chain molecules with carbon atom backbones. How does the geometry of the boron hydrides differ from this pattern?
9. Why is it difficult for nitrogen to form long $N-N-N-N-N$ chains analogous to those of carbon? How is this a consequence of the electronic structures of the atoms?
10. To what extent do nitrogen hydrides exist? What is their single most striking property?
11. How does the bond energy of N_2 contribute to the instability of long-chain nitrogen compounds?
12. Why is silicon different in chemical behavior from carbon? Why is SiO_2 a high-melting solid, whereas CO_2 is a gas?
13. What is the backbone chain structure in silicones? In silanes? How do these two classes compare in reactivity with hydrocarbons?
14. Why can't silicon form double bonds with other silicon atoms? Then how can it form a double bond with oxygen?
15. In view of the difficulty of forming long chains of nitrogen atoms, why can nitrogen fit so easily into carbon chains?
16. What is the distinction between alkanes, alkenes, and alkynes?
17. How are branched-chain alkanes given systematic names? How is the framework for naming the alkane chosen? How are the branching side chains identified and located?

18. What are structural isomers? Show with diagrams that there are only nine structural isomers of heptane.
19. Write the structures of the following alkane radicals: methyl, ethyl, isopropyl.
20. Why are cyclopropane and cyclobutane termed *strained* molecules? Why is cyclohexane not strained?
21. Why is the chair conformation of cyclohexane more stable than the boat form?
22. How are alkene molecules constrained, in a way that alkanes are not?
23. What do the numbers in 1-butene and 2-butene signify? Why is there no 3-butene?
24. How many geometrical isomers are there of 1-butene? Of 2-butene? Draw them and give their proper systematic names.
25. What is a diene?
26. What types of chemical reactions do alkanes most commonly undergo? What are the principal reactions of alkenes?
27. What are functional groups in organic molecules? In what ways do they determine the chemical behavior of the molecules?
28. How can alcohols be prepared from halogenated hydrocarbons? How else are alcohols obtained?
29. What is an ether, and how is it made from alcohols?
30. How are aldehydes made from alcohols? How are ketones made?
31. What compounds are produced when aldehydes are oxidized? Can ketones be oxidized under similar conditions?
32. How does delocalization of electrons influence the structure of a carboxylate ion? Where is the negative charge after a carboxylic acid has ionized?
33. What is a fatty acid?
34. What are the salts of fatty acids and alkali metal cations called? What makes such salts useful to us, and how do they work?
35. How do soaps operate as cleaning agents?
36. How are fats related to fatty acids? What is the alcohol that is esterified in fats?
37. Why are organic amines bases? How can they be bases if they have no hydroxyl groups to dissociate? What inorganic base do they most resemble?
38. What are primary, secondary, and tertiary amines? Illustrate with methyl groups. Why can there be quaternary ammonium ions, but not quaternary amines?
39. What is an amino acid? In what sense is it both an acid and a base? What is the zwitterion form of an amino acid, and under what conditions is it present?
40. What is meant by an asymmetric carbon atom in an amino acid?
41. How are amino acids combined to form a protein molecule? What molecule is released during this combining process?

42. What are the two stereoisomers of an amino acid called? Which one is found almost exclusively in living organisms?
43. What is a conjugated molecule? An aromatic molecule? In what sense are they examples of electron delocalization?
44. What effect does delocalization of electrons have on the spacings of electronic energy levels? Why does it follow that many of these molecules are colored?
45. How do certain conjugated molecules function as “antennas” for light, and what is done with the absorbed light energy?
46. Why do aromatic molecules not undergo ethenelike addition reactions across double bonds? What reactions occur instead?
47. In what sense can graphite be considered a relative of benzene? What substance bears the same logical relationship to ethane?
48. Why were sugars given the name “carbohydrates,” and why is this a misleading name?
49. What is the molecular formula of glucose? What is its molecular weight?
50. What is the difference between α - and β -D-glucose? Why does an equilibrium mixture exist in aqueous solution?
51. What is the difference between the intersubunit bonds in cellulose and starch? Why does no confusion arise in a plant regarding making and breaking of these two different kinds of bonds? What enzymes are involved?
52. What use is made by a plant of the two polymers, starch and cellulose? What substances play the corresponding roles in the human body?
53. How can cows and termites digest cellulose, when they cannot make the necessary enzyme, cellulase?
54. What is the difference between globular and fibrous proteins? What are the main functions of each?
55. How many different amino acid side chains are encountered in most living organisms? Give an example of one that is (a) positively charged; (b) negatively charged; (c) hydrophobic and aliphatic; (d) hydrophobic and aromatic; (e) uncharged but polar.
56. What is an α helix? Is it found in fibrous, or globular, proteins? What makes it a particularly stable structure for a protein chain to adopt?
57. How do hydrogen bonds influence the structure of a globular protein? How do hydrophobic side chains influence its structure?
58. Which atom in an amino acid is responsible for the existence of optical isomers? How many optical isomers are there? How many ordinarily are found in proteins of living organisms?
59. Is there any inherent reason why life on earth should be based on L-amino acids exclusively, rather than D-amino acids? Would there be any conceivable advantage to the use of either L- or D-amino acids, as opposed to a mixture of the two?
60. What is the active site of an enzyme? Is it involved with enzymatic catalysis, or with substrate specificity?

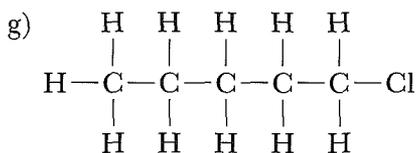
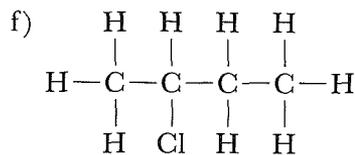
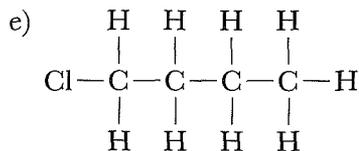
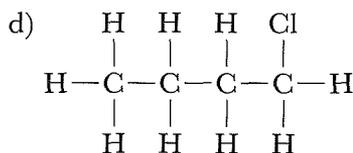
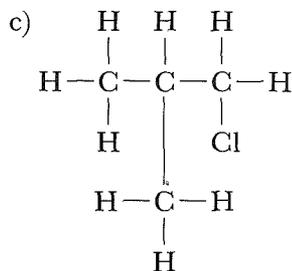
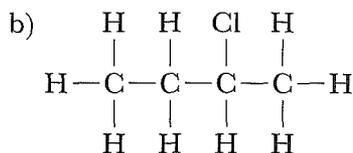
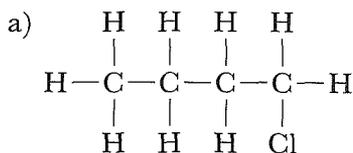
61. How do enzymes recognize their proper substrates and reject other molecules?
62. What is the acyl enzyme in the trypsin cleavage of a protein chain?
63. What is the tetrahedral intermediate in protein digestion by trypsin or chymotrypsin? How many tetrahedral intermediates are formed in one complete cycle of catalysis?
64. Where does the hydrogen atom come from that is added to the amino end of the severed substrate protein chain during protein digestion by trypsin? Where does the hydrogen come from to restore the enzyme to its initial state at the end of one cycle of catalysis?
65. How does the enzyme help to stabilize the tetrahedral intermediate during cleavage by trypsin?
66. How do trypsin, chymotrypsin, and elastase discriminate between their respective substrates?
67. What parts do histidine and aspartic acid play at the active site during trypsin cleavage?
68. What are NAD^+ , NADP^+ , and FAD , and what biological roles do they play? In what sense are they carriers of free energy, and how much does each carry? How does their energy compare with the amount of free energy carried by a mole of ATP ?
69. In what sense are NAD^+ , NADP^+ , and FAD carriers of reducing power? Can ATP also function as such a carrier?
70. What is a vitamin? What vitamins are necessary for the synthesis of NAD^+ ? Of FAD ? Why is no vitamin necessary for the synthesis of ATP ?
71. Is a molecule that is a vitamin for one organism necessarily a vitamin for every other organism? Why, or why not?
72. How much free energy from combustion of a mole of glucose comes from enthalpy, and how much from entropy?
73. Why is the combustion of glucose in a cat more efficient than combustion of the same glucose in a Bunsen burner?
74. What is glycolysis? In what sense is it anaerobic? What is the end product of glycolysis in human muscles deprived of oxygen? What is the end product if ample oxygen is present?
75. In what sense is glycolysis probably a chemical fossil?
76. How does anaerobic fermentation compare with aerobic respiration, in terms of free energy produced per gram of glucose?
77. What is the citric acid cycle? What metabolite goes into the cycle, and what carbon compounds come out? What happens to the hydrogen atoms?
78. Where do the hydrogen atoms from the citric acid cycle eventually go? How is energy stored in this process? How much energy is stored per 2 moles of hydrogen atoms?
79. Why do we suspect that the citric acid cycle and respiratory chain are younger than the glycolytic pathway?
80. How is the respiratory chain linked to the citric acid cycle? What molecules are oxidized by the respiratory chain? What molecules are

- reduced by the respiratory chain at its other end? How is energy stored along the chain?
81. Why does the reoxidation of FADH_2 produce less ATP than reoxidation of NADH?
 82. How many moles of ATP are synthesized during the breakdown of a mole of glucose? How much free energy is stored in this way? How much free energy is released by oxidizing the glucose, and what percentage is saved via ATP?
 83. Which component of energy metabolism of higher organisms is common to all life forms? When the component operates without the other components, how many moles of ATP result from 1 mole of glucose? How many times more productive is the entire system?
 84. What are the two metabolic components of photosynthesis? Which is believed to be older, and what basis is there for such a hypothesis?
 85. What gross effect has green-plant photosynthesis had on our planetary atmosphere?

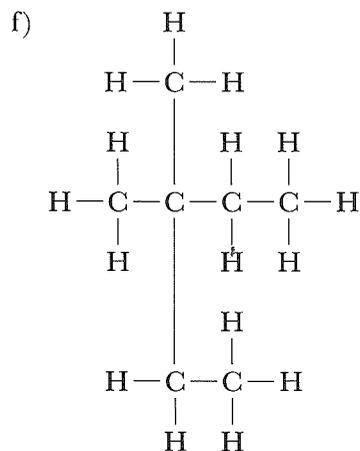
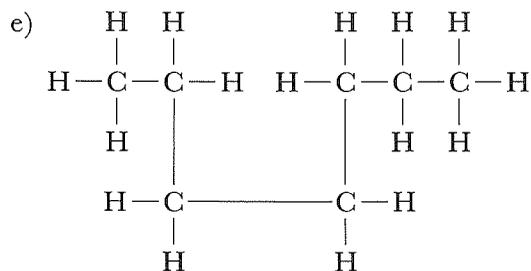
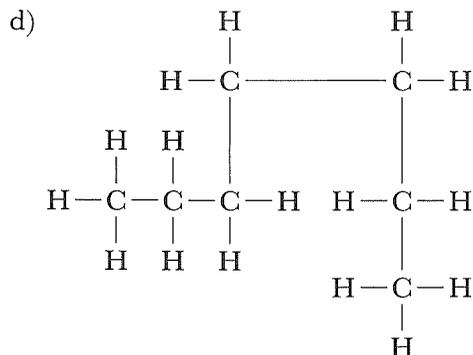
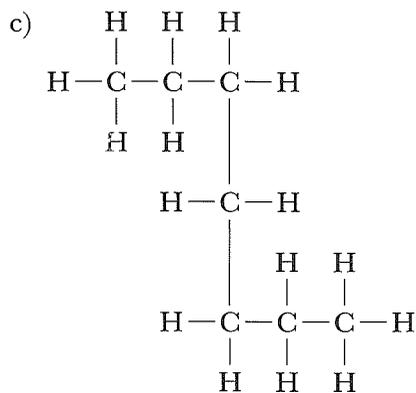
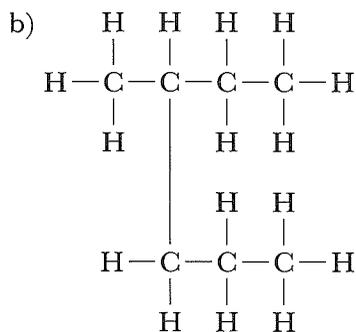
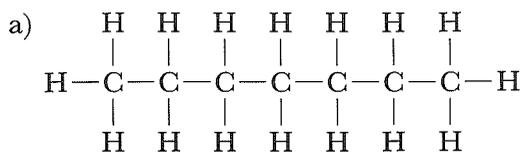
Problems

Molecular geometry

1. Which of the following isomers are identical?



2. Consider the following six molecules:



Of molecules b–f, which are identical with *n*-heptane (molecule a)? What are the systematic names for all the molecules that are not identical with a?

Isomers and nomenclature

3. Draw all the possible structures for compounds having the following molecular formulas. Give their systematic names.

- C_3H_8
- C_3H_4
- C_4H_8
- C_3H_5Cl

4. Draw all the possible structural isomers of C_5H_{10} . Give the systematic name for each isomer.

Isomers

5. Which of the following compounds are isomers of one another?

- $CH_3CH_2CH_2OH$
- $CH_3CHClCH_3$
- $CH_3CH_2CH_3$
- $$\begin{array}{c} O \\ || \\ CH_3CCH_2CH_3 \end{array}$$
- $CH_3CH_2CH_2Cl$

Geometry and bonding

6. Sketch the structure of each of the following molecules and indicate the hybridization around each carbon atom.

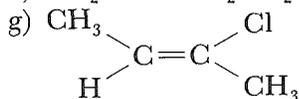
- C_2Cl_4
- CBr_4
- C_2Cl_2
- CH_2Cl_2
- C_2F_6

Systematic nomenclature

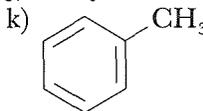
7. What is the systematic name of each of the following compounds?

- $CH_3CH_2CH=CH_2$
- $CH_3CH_2C\equiv CH$
- $CH_2=CF_2$
- $(CH_3)_2C=CHCH_3$

- $CH_3CH=CCL_2$
- $CH_2=CHCH_2CH_2CH=CH_2$



- $CH_2=CH-CBr=CH-CH_3$
- $(CH_3)_2C=CHCH_2CH(CH_3)_2$
- $CH_3CHClC\equiv CCH_3$

**Nomenclature**

8. Give the systematic names of the following substances:

- $CH_3CH_2CH_2COOCH_3$
- $CH_3CHOHCH_2CH_2CH_2CH_2CH_3$
- $CH_3CH_2CH_2OCH_2CH_3$
- $CH_3COCH(CH_3)_2$

Structural formulas

9. Draw the structural formula for each of the following compounds:

- trans*-2-hexene
- cis*-2,3-dichloro-2-butene
- 1-methylcyclopentene
- trans*-1,2-dibromocyclohexane
- 4-ethyl-1-octene
- 3-hexyne
- cis*-diiodoethylene
- 2-methyl-2-butene
- 2-bromo-1,3-butadiene

Hydrocarbon structure

10. Write the structural formulas of

- five different, simple alkanes
- five different, simple alkenes (only *one* double bond)

- c) five different, simple alkynes (only *one* triple bond)
 d) five different, simple cycloalkanes (only *one* ring)

Show that they conform to the general formulas C_nH_{2n+2} , C_nH_{2n} , C_nH_{2n-2} , and C_nH_{2n} , respectively.

Isomers of hydrocarbons

11. A hydrocarbon has a molecular weight of approximately 60 and contains 17.2% hydrogen. What is the molecular formula of the hydrocarbon? Write all of the structural isomers that have this formula.
12. A hydrocarbon has a molecular weight of 56.0 and contains 85.7% carbon. What is its molecular formula? What structural and geometric isomers could it have?

Alcohol structures

13. Write all the structural formulas of the isomeric alcohols with molecular composition $C_6H_{13}OH$. Give each its proper systematic name.

Reactions of alkenes

14. Outline a scheme for synthesizing a butanol from 1-butene. What intermediate compound might be formed? Would the product be 1-butanol or 2-butanol, and what is the principle that tells you which?

Isomers and structure

15. How many structural isomers are there of $C_6H_{14}O$ that are ethers? What are their systematic names? What other kinds of compounds in addition to ethers can be structural isomers of $C_6H_{14}O$?

16. Name the following substances. What kind of chemical compound are they? Are c and d isomers?

- a) $C_6H_5OC_2H_5$
 b) $(CH_3)_2CHOCH(CH_3)_2$
 c) $CH_3OCH_2C_2CH_3$
 d) $CH_3CH_2CH_2OCH_3$

17. Name the following substances. What kind of chemical compound are they? Are c and d isomers?

- a) $C_6H_5COOC_2H_5$
 b) $C_6H_5COOC_6H_5$
 c) $CH_3COOCH_2CH_2CH_3$
 d) $CH_3CH_2CH_2COOCH_3$

Oxidation of alcohols

18. What compounds are obtained by subjecting 1-propanol to moderate (i.e., nondestructive) oxidizing conditions? The intermediate and the final product in this reaction have quite different vapor pressures at room temperature. Explain this in terms of molecular structure. How could this help you to design experimental conditions to maximize the yield of either intermediate or final product?

Molecular structures

19. Draw the structural formula for each of the following:
- a) sodium propionate
 b) *m*-bromobenzoic acid
 c) ethyl benzoate
 d) isobutyryl chloride
 e) methyl formate
 f) diethylamine
 g) tri-*n*-propyl amine
 h) benzyl amine
 i) *m*-bromoaniline
 j) tetraethylammonium hydroxide
 k) alanine

Suggested Reading

- N. L. Allinger and J. Allinger, *Structures of Organic Molecules*, Prentice-Hall, Englewood Cliffs, N.J., 1965.
- “The Biosphere,” *Scientific American*, September 1970. An issue devoted to a common theme: the chemistry of our planet and of life.
- R. Breslow, “The Nature of Aromatic Molecules,” *Scientific American*, August 1972.
- R. E. Dickerson and I. Geis, *Chemistry, Matter, and the Universe*, W. A. Benjamin, Menlo Park, Calif., 1976. Chapters 18–26 provide an elementary introduction to organic and biochemistry that is the next logical step in difficulty beyond this chapter.
- R. E. Dickerson and I. Geis, *The Structure and Action of Proteins*, W. A. Benjamin, Menlo Park, Calif., 1969. Elementary introduction, abundantly illustrated.
- E. Frieden, “The Chemical Elements of Life,” *Scientific American*, July 1972. Discussion of the essential elements for life, and the roles that they play.
- W. Herz, *The Shape of Carbon Compounds*, W. A. Benjamin, Menlo Park, Calif., 1963. Elementary introduction to organic chemistry, with emphasis on reaction mechanisms.
- R. C. Johnson, *Introductory Descriptive Chemistry*, W. A. Benjamin, Menlo Park, Calif., 1966. Useful information on boron and nitrogen chemistry.
- L. Stryer, *Biochemistry*, W. H. Freeman, San Francisco, 1975. Excellent introductory textbook.