From Atoms to Cells:
A Chemical Connection

CHAPTER 2

In laboratories all over the world, sophisticated technology is being developed for a wide variety of scientific applications. Refinements in molecular biology techniques now make it possible to routinely identify microorganisms, detect genetic disease, diagnose cancer, sequence the genes of organisms, break down toxic wastes, synthesize drugs and industrial products, and genetically engineer microorganisms, plants, and animals. A common thread that runs through new technologies and hundreds of traditional techniques is that, at some point, they involve chemicals and chemical reactions. In fact, if nearly any biological event is traced out to its ultimate explanation, it will invariably involve atoms, molecules, reactions, and bonding.

It is this relationship between the sciences that makes a background in chemistry necessary to biologists and microbiologists. Students with a basic chemistry background will enhance their understanding of and insight into microbial structure and function, metabolism, genetics, drug therapy, immune reactions, and infectious disease. This chapter has been organized to promote a working knowledge of atoms, molecules, bonding, solutions, pH, and biochemistry and to build foundations to later chapters. It concludes with an introduction to cells and a general comparison of procaryotic and eucaryotic cells as a preparation for chapters 4 and 5.

A molecular probe machine, called an ion microprobe, designed to analyze the isotopes found in very tiny samples of meteors, ancient rocks, and fossil samples. Chemists have used this device to determine the age of certain rocks found in Greenland (3.85 billion years old) and whether the sample may have come from a living thing.

Chapter Overview

- The understanding of living cells and processes is enhanced by a knowledge of chemistry.
- The structure and function of all matter in the universe is based on atoms.
- Atoms have unique structures and properties that allow chemical reactions to occur.
- Atoms contain protons, neutrons, and electrons in combinations to form elements.
- Living things are composed of approximately 25 different elements.
- Elements interact to form bonds that result in molecules and compounds with different characteristics than the elements that form them.
- Atoms can show variations in charge and polarity.
- Atoms and molecules undergo chemical reactions such as oxidation/reduction, ionization, and dissolution.
- The properties of carbon have been critical in forming macromolecules of life such as proteins, fats, carbohydrates, and nucleic acids.
- The nature of macromolecule structure and shape dictates its functions.
- Cells carry out fundamental activities of life, such as growth, metabolism, reproduction, synthesis, and transport, that are all essentially chemical reactions on a grand scale.

Atoms, Bonds, and Molecules: Fundamental Building Blocks

The universe is composed of an infinite variety of substances existing in the gaseous, liquid, and solid states. All such tangible materials that occupy space and have mass are called **matter.** The

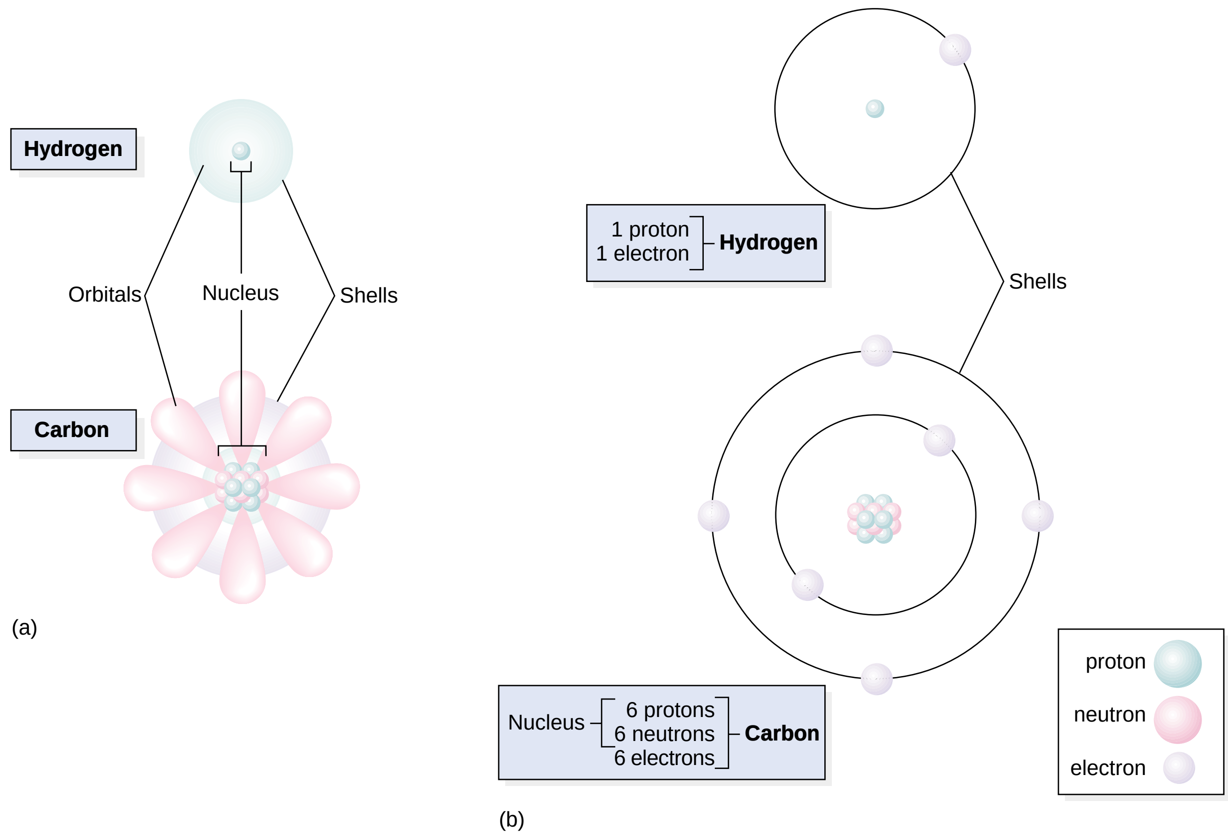


FIGURE 2.1

Models of atomic structure. (a) Three-dimensional models of hydrogen and carbon that approximate their actual structure. The nucleus is surrounded by electrons in orbitals that occur in levels called shells. Hydrogen has just one shell and one orbital. Carbon has two shells and four orbitals; the shape of the outermost orbitals is paired lobes rather than circles or spheres. **(b)** Simple models of the same atoms make it easier to show the numbers and arrangements of shells and electrons, and the numbers of protons and neutrons in the nucleus.

organization of matter—whether air, rocks, or bacteria—begins with individual building blocks called atoms. An **atom*** is defined as a tiny particle that cannot be subdivided into smaller substances without losing its properties. Even in a science dealing with very small things, an atom's minute size is striking; for example, an oxygen atom is only 0.0000000013 mm (0.0013 nm) in diameter, and one million of them in a cluster would barely be visible to the naked eye.

Although scientists have not directly observed the detailed structure of an atom, the exact composition of atoms has been well established by extensive physical analysis using sophisticated instruments. In general, an atom derives its properties from a combination of subatomic particles called **protons** (p^+), which are positively charged; **neutrons** (n^0), which have no charge (are neutral); and **electrons** (e^-), which are negatively charged. The relatively larger protons and neutrons make up a central core, or **nucleus,** that is surrounded by one or more electrons (figure 2.1). The nucleus makes up the larger mass (weight) of the atom, whereas the electron region accounts for the greater volume. To get a perspective on proportions, consider this: If an atom were the size of a football stadium, the nucleus would be about the size of a marble! The stability of atomic structure is largely maintained by: (1) the mutual attraction of the protons and electrons (opposite charges attract each other) and (2) the exact balance of proton number and electron number, which causes the opposing charges to cancel each other out. At least in theory then, isolated intact atoms do not carry a charge.

*__atom__ (at′-um) Gr. *atomos,* not cut.

DIFFERENT TYPES OF ATOMS: ELEMENTS AND THEIR PROPERTIES

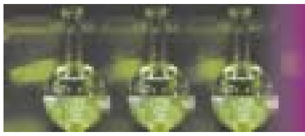
All atoms share the same fundamental structure. All protons are identical, all neutrons are identical, and all electrons are identical. But when these subatomic particles come together in specific, varied combinations, unique types of atoms called **elements** result. Each element has a characteristic atomic structure and predictable chemical behavior. To date, 92 naturally occurring elements have been described, and 18 have been produced artificially by physicists. By convention, an element is assigned a distinctive name with an abbreviated shorthand symbol. Table 2.1 lists some of the elements common to biological systems, their atomic characteristics, and some of the natural and applied roles they play.

TABLE 2.1

The Major Elements of Life and Their Primary Characteristics

Element	Atomic Symbol*	Atomic Number	Atomic Weight	Ionized Form**	Significance in Microbiology
Calcium	Ca	20	40.1	Ca^{++}	Part of outer covering of certain shelled amebas; stored within bacterial spores
Carbon	C	6	12.0	—	Principal structural component of biological molecules
		6	14.0	—	Isotope used in dating fossils
Chlorine	Cl	17	35.5	Cl^{-}	Component of disinfectants; used in water purification
Cobalt	Co	27	58.9	Co^{++}, Co^{+++}	Trace element needed by some bacteria to synthesize vitamins
		27	60	—	An emitter of gamma rays; used in food sterilization; used to treat cancer
Copper	Cu	29	63.5	Cu^{+}, Cu^{++}	Necessary to the function of some enzymes; Cu salts are used to treat fungal and worm infections
Hydrogen	H	1	1	H^{+}	Necessary component of water and many organic molecules; H_2 gas released by bacterial metabolism
		1	3	—	Tritium has 2 neutrons; radioactive; used in clinical laboratory procedures
Iodine	I	53	126.9	I^{-}	A component of antiseptics and disinfectants; contained in a reagent of the Gram stain
		53	131, 125		Radioactive isotopes for diagnosis and treatment of cancers
Iron	Fe	26	55.8	Fe^{++}, Fe^{+++}	Necessary component of respiratory enzymes; some microbes require it to produce toxin
Magnesium	Mg	12	24.3	Mg^{++}	A trace element needed for some enzymes; component of chlorophyll pigment
Manganese	Mn	25	54.9	Mn^{++}, Mn^{+++}	Trace elements for certain respiratory enzymes
Nitrogen	N	7	14.0	—	Component of all proteins and nucleic acids; the major atmospheric gas
Oxygen	O	8	16.0	—	An essential component of many organic molecules; molecule used in metabolism by many organisms
Phosphorus	P	15	31	—	A component of ATP, nucleic acids, cell membranes; stored in granules in cells
		15	32	—	Radioactive isotope used as a diagnostic and therapeutic agent
Potassium	K	19	39.1	K^{+}	Required for normal ribosome function and protein synthesis; essential for cell membrane permeability
Sodium	Na	11	23.0	Na^{+}	Necessary for transport; maintains osmotic pressure; used in food preservation
Sulfur	S	16	32.1	—	Important component of proteins; makes disulfide bonds; storage element in many bacteria
Zinc	Zn	30	65.4	Zn^{++}	An enzyme cofactor; required for protein synthesis and cell division; important in regulating DNA

**Based on the Latin name of the element. The first letter is always capitalized; if there is a second letter, it is always lowercased.*
***A dash indicates an element that is usually found in combination with other elements, rather than as an ion.*



SPOTLIGHT ON MICROBIOLOGY 2.1
Searching for Ancient Life with Isotopes

Determining the age of the earth and the historical time frame of living things has long been a priority of biologists. Much evidence comes from fossils, geologic sediments, and genetic studies, yet there has always been a need for an exacting scientific reference for tracing samples back in time, possibly even to the beginnings of the earth itself. One very precise solution to this problem comes from patterns that exist in isotopes. The isotopes of an element have the same basic chemical structure, but over billions of years, they have come to vary slightly in the number of neutrons. For example, carbon has 3 isotopes: C12, predominantly found in living things; C13, a less common form associated with nonliving matter; and C14, a radioactive isotope. All isotopes exist in relatively predictable proportions in the earth, solar system, and even universe, so that any variations from the expected ratios would indicate some other factor besides random change.

Isotope chemists use giant machines called microprobes to analyze the atomic structures in fossils and rock samples (see chapter opening photo). These amazing machines can rapidly sort and measure the types and amounts of isotopes, which reflect a sample's age and possibly its origins. The accuracy of this method is such that it can be used like an "atomic clock." It was recently used to verify the dateline for the origins of the first life forms, using 3.85 billion-year-old sediment samples from Greenland. Testing indicated that the content of C12 in the samples was substantially higher than the amount in inorganic rocks, and it was concluded that living cells must have accumulated the C12. This finding shows that the origin of life was 400 million years earlier than the previous estimates.

In a separate study, some ancient Martian meteorites were probed to determine if certain microscopic rods could be some form of microbes (figure 1.10*b*). By measuring the ratios of oxygen isotopes in carbonate ions (CO_3^{-2}), chemists were able to detect significant fluctuations in the isotopes from different parts of the same meteorite. Such differences would most likely be caused by huge variations in temperature or other extreme environments that are incompatible with life. From this evidence, they concluded that the tiny rods were not Martian microbes.

THE MAJOR ELEMENTS OF LIFE AND THEIR PRIMARY CHARACTERISTICS

The unique properties of each element result from the numbers of protons, neutrons, and electrons it contains, and each element can be identified by certain physical measurements.

Each element is assigned an **atomic number (AN)** based on the number of protons it has. The atomic number is a valuable measurement because an element's proton number does not vary, and knowing it automatically tells you the usual number of electrons (recall that a neutral atom has an equal number of protons and electrons). Another useful measurement is the **mass**[1] **number (MN),** equal to the number of protons and neutrons. If one knows the mass number and the atomic number, it is possible to determine the numbers of neutrons by subtraction. Hydrogen is a unique element because its common form has only one proton, one electron, and no neutron, making it the only element with the same atomic and mass number.

Isotopes are variant forms of the same element that differ in the number of neutrons and thus have different mass numbers. These multiple forms occur naturally in certain proportions. Carbon, for example, exists primarily as carbon 12 with 6 neutrons (MN=12); but a small amount (about 1%) is carbon 13 with 7 neutrons and carbon 14 with 8 neutrons. Although isotopes have virtually the same chemical properties, some of them have unstable nuclei that spontaneously release energy in the form of radiation. Such *radioactive isotopes* play a role in a number of research and medical applications. Because they emit detectable signs, they can be used to trace the position of key atoms or molecules in chemical reactions, they are tools in diagnosis and treatment, and they are even applied in sterilization procedures (see ionizing radiation in chapter 11). Another application of isotopes is in dating fossils and other ancient materials (Spotlight on Microbiology 2.1). An element's **atomic weight** is the average of the mass numbers of all its isotopic forms (table 2.1).

Electron Orbitals and Shells

The structure of an atom can be envisioned as a central nucleus surrounded by a "cloud" of electrons that constantly rotate about the nucleus in pathways (see figure 2.1). The pathways, called **orbitals,** are not actual objects or exact locations, but represent volumes of space in which an electron is likely to be found. Electrons occupy energy **shells,** proceeding from the lower-level energy electrons nearest the nucleus to the higher-energy electrons in the farthest orbitals.

Electrons fill the orbitals and shells in *pairs*, starting with the shell nearest the nucleus. The first shell contains one orbital and a maximum of 2 electrons; the second shell has four orbitals and up to 8 electrons; the third shell with 9 orbitals can hold up to 18 electrons; and the fourth shell with 16 orbitals contains up to 32 electrons. The number of orbitals and shells and how completely they are filled depends on the numbers of electrons, so that each element will have a unique pattern. For example, helium (AN=2) has only a filled first shell of 2 e^-; oxygen (AN=8) has a filled first shell and a partially filled second shell of 6 e^-; and magnesium (AN=12) has a filled first shell, a filled second one, and a third shell that fills only one orbital, so is nearly empty. As we will see, the chemical properties of an element are controlled mainly by the distribution of electrons in the outermost shell. Figures 2.1 and 2.2 present various simplified models of atomic structure and electron maps.

1. Mass refers to the amount of matter that a particle contains. The proton and neutron have almost exactly the same mass, which is about 1.7×10^{-24} g, or 1 dalton.

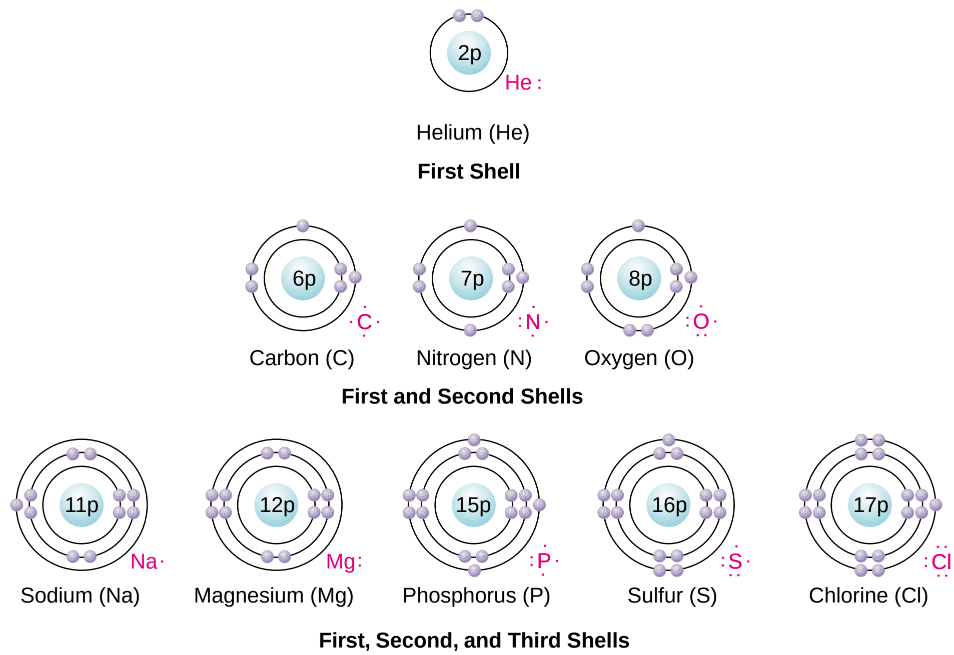


FIGURE 2.2

Electron orbitals and shells. Models of several elements show how the shells are filled by electrons as the atomic numbers increase (numbers noted inside nuclei). Electrons tend to appear in pairs, but certain elements have incompletely filled outer shells. Chemists depict elements in shorthand form (red Lewis structures) that indicate only the valence electrons, since these are the electrons involved in chemical bonds.

CHAPTER CHECKPOINTS

- Protons (p^+) and neutrons (n^0) make up the nucleus of an atom. Electrons (e^-) orbit the nucleus.
- All elements are composed of atoms but differ in the numbers of protons, neutrons, and electrons they possess.
- Elements are identified by *atomic weight*, or *mass*, or by *atomic number*.
- Isotopes are varieties of one element that contain the same number of protons but different numbers of neutrons.
- The number of electrons in an element's outermost orbital (compared with the total number possible) determines its chemical properties and reactivity.

Bonds and Molecules

Most elements do not exist naturally in pure, uncombined form but are bound together as molecules and compounds. A **molecule*** is a distinct chemical substance that results from the combination of two or more atoms. Some molecules such as oxygen (O_2) and nitrogen gas (N_2) consist of atoms of the same element. Molecules that are combinations of two or more *different* elements are termed **compounds.** Compounds such as water (H_2O) and biological molecules (proteins, sugars, fats) are the predominant substances in living systems. When atoms bind together in molecules, they lose the properties of the atom and take on the properties of the combined substance. In the same way that an atom has an atomic weight, a molecule has a molecular weight (MW), which is calculated from the sum of all of the atomic weights of the atoms it contains.

The **chemical bonds** of molecules and compounds result when two or more atoms share, donate (lose), or accept (gain) electrons (figure 2.3). The number of electrons in the outermost shell of an element is known as its **valence.*** The valence determines the degree of reactivity and the types of bonds an element can make. Elements with a filled outer orbital are relatively stable because they have no extra electrons to share with or donate to other atoms. For example, helium has one filled shell, with no tendency either to give up electrons or to take them from other elements, making it a stable, inert (nonreactive) gas. Elements with partially filled outer orbitals are less stable and are more apt to form some sort of bond. Many chemical reactions are based on the tendency of atoms with unfilled outer shells to gain greater stability by achieving, or at least approximating, a filled outer shell. For example, an atom such as oxygen that can accept 2 additional electrons will bond readily with atoms (such as hydrogen) that can share or donate electrons. We explore some additional examples of the basic types of bonding in the following section.

In addition to reactivity, the number of electrons in the outer shell also dictates the number of chemical bonds an atom can make. For instance, hydrogen can bind with one other atom, oxygen can bind with up to two other atoms, and carbon can bind with four (see figure 2.13).

COVALENT BONDS AND POLARITY: MOLECULES WITH SHARED ELECTRONS

Covalent (cooperative valence) **bonds** form between atoms with valences that suit them to sharing electrons rather than to donating or receiving them. A simple example is hydrogen gas (H_2), which consists of two hydrogen atoms. A hydrogen atom has only a single electron, but when two of them combine, each will bring its electron to orbit about both nuclei, thereby approaching a filled orbital (2 electrons) for both atoms and thus creating a **single covalent bond** (figure 2.4*a*). Covalent bonding also occurs in oxygen gas (O_2), but with a difference. Because each atom has 2 electrons to share in this molecule, the combination creates two pairs of shared

*molecule (mol′-ih-kyool) L. *molecula*, little mass.

*valence (vay′-lents) L. *valentia*, strength. A measure of atomic binding capacity.

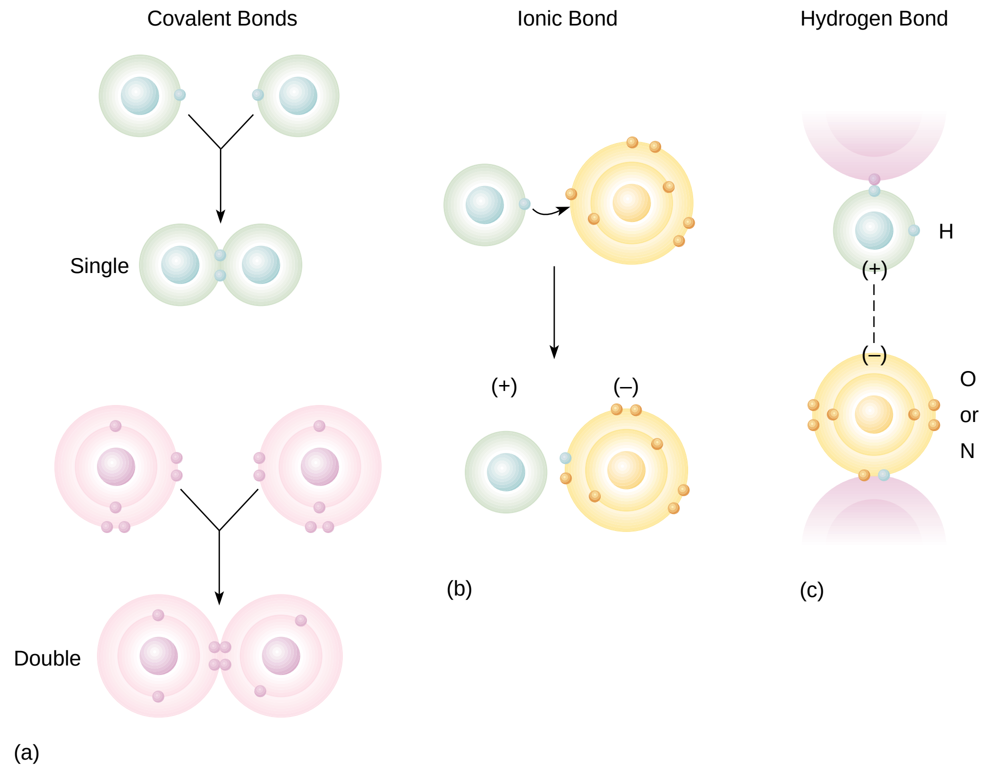


FIGURE 2.3

General representation of three types of bonding. (a) Covalent bonds, both single and double. **(b)** Ionic bond. **(c)** Hydrogen bond. Note that hydrogen bonds are represented in models and formulas by dotted lines, as shown in **(c)**.

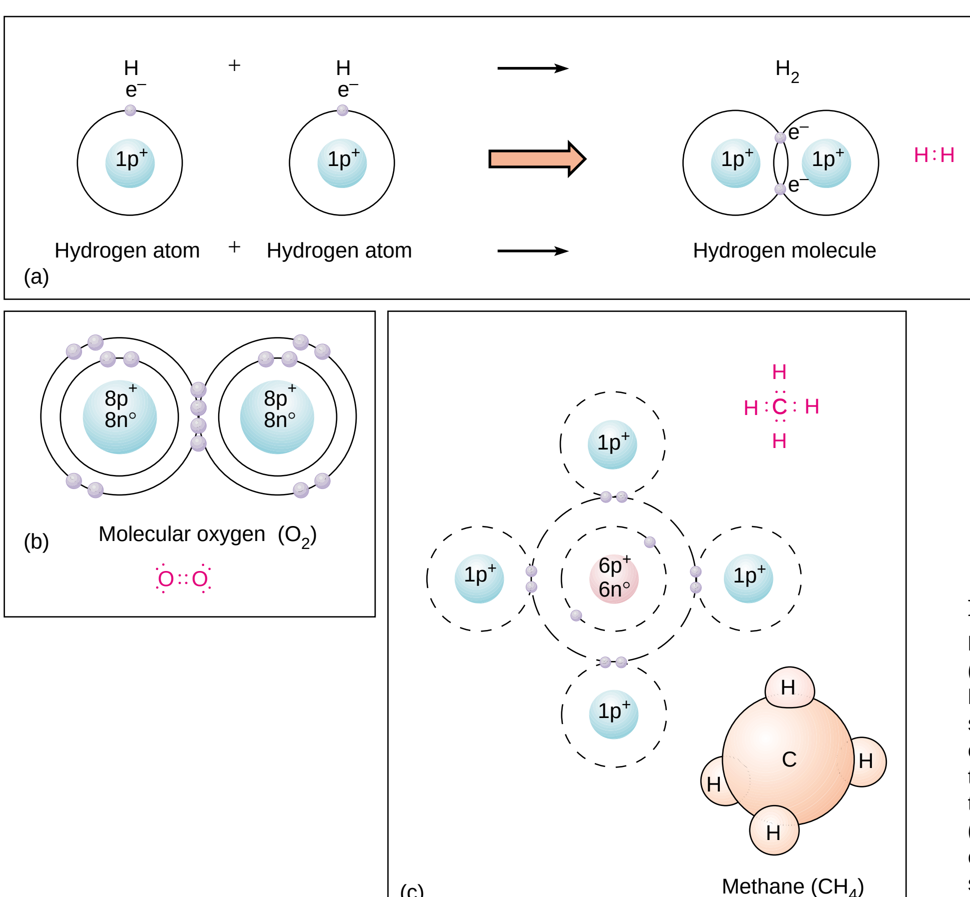


FIGURE 2.4

Examples of molecules with covalent bonding. (a) A hydrogen molecule is formed when two hydrogen atoms share their electrons and form a single bond. **(b)** In a double bond, the outer orbitals of two oxygen atoms overlap and permit the sharing of 4 electrons (one pair from each) and the saturation of the outer orbital for both. **(c)** Simple, working, and three-dimensional models of methane. Note that carbon has 4 electrons to share and hydrogens each have one, thereby completing the shells for all atoms in the compound.

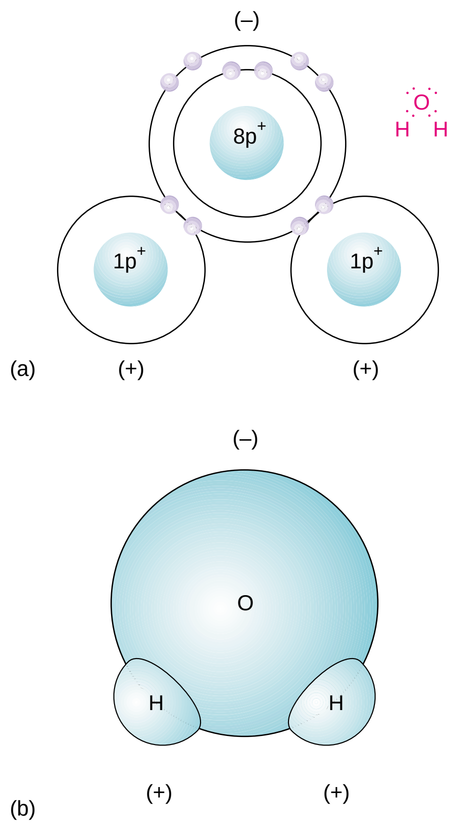


FIGURE 2.5

Polar molecule. (a) Simple models and **(b)** a three-dimensional model of a water molecule indicate the polarity, or unequal distribution, of electrical charge, which is caused by the pull of the shared electrons toward the oxygen side of the molecule.

electrons, also known as a **double covalent bond** (figure 2.4*b*). The majority of the molecules associated with living things are composed of single and double covalent bonds between the most common biological elements (carbon, hydrogen, oxygen, nitrogen, sulfur, and phosphorus), which are discussed in more depth in chapter 7. A slightly more complex pattern of covalent bonding is shown for methane gas (CH_4) in figure 2.4*c*.

Other effects of bonding result in differences in polarity. When atoms of different electronegativity form covalent bonds, the electrons are not shared equally and may be pulled more toward one atom than another. This pull causes one end of a molecule to assume a partial negative charge and the other end to assume a partial positive charge. A molecule with such an asymmetrical distribution of charges is termed **polar** and has positive and negative poles. Observe the water molecule shown in figure 2.5 and note that, because the oxygen atom is larger and has more protons than the hydrogen atoms, it will tend to draw the shared electrons with greater force toward its nucleus. This unequal force causes the oxygen part of the molecule to express a negative charge (due to the electrons' being attracted there) and the hydrogens to express a positive charge (due to the protons). The polar nature of water plays an extensive role in a number of biological reactions, which are discussed later. Polarity is a significant property of many large molecules in living systems and greatly influences both their reactivity and their structure.

When covalent bonds are formed between atoms that have the same or similar electronegativity, the electrons are shared equally between the two atoms. Because of this balanced distribution, no part of the molecule has a greater attraction for the electrons. This sort of electrically neutral molecule is termed **nonpolar.**

IONIC BONDS: ELECTRON TRANSFER AMONG ATOMS

In reactions that form **ionic bonds,** electrons are transferred completely from one atom to another and are not shared. These reactions invariably occur between atoms with valences that complement each other, meaning that one atom has an unfilled shell that will readily accept electrons and the other atom has an unfilled shell that will readily lose electrons. A striking example is the reaction that occurs between sodium (Na) and chlorine (Cl). Elemental sodium is a soft, lustrous metal so reactive that it can burn flesh, and molecular chlorine is a very poisonous yellow gas. But when the two are combined, they form sodium chloride[2] (NaCl)—the familiar nontoxic table salt—a compound with properties quite different from either parent element (figure 2.6).

How does this transformation occur? Sodium has 11 electrons (2 in shell one, 8 in shell two, and only 1 in shell three), so it is 7 short of having a complete outer shell. Chlorine has 17 electrons (2 in shell one, 8 in shell two, and 7 in shell three), making it 1 short of a complete outer shell. These two atoms are very reactive with one another, because a sodium atom will readily donate its single electron and a chlorine atom will avidly receive it. (The reaction is slightly more involved than a single sodium atom's combining with a single chloride atom (Microbits 2.2), but this complexity does not detract from the fundamental reaction as described here.) The outcome of this reaction is not many single, isolated molecules of NaCl but rather a solid crystal complex that interlinks millions of sodium and chloride ions (figure 2.6*b*).

Ionization: Formation of Charged Particles

Molecules with intact ionic bonds are electrically neutral, but they can produce charged particles when dissolved in a liquid called a solvent. This phenomenon, called **ionization,** occurs when the ionic bond is broken and the atoms dissociate (separate) into unattached, charged particles called **ions*** (figure 2.7). To illustrate what imparts a charge to ions, let us look again at the reaction between sodium and chlorine. When a sodium atom reacts with chlorine and loses one electron, the sodium is left with one more proton than electrons. This imbalance produces a positively charged sodium ion (Na^+). Chlorine, on the other hand, has gained one electron and now has one more electron than protons, producing a negatively charged ion (Cl^-). Positively charged ions are termed **cations,*** and negatively charged ions are termed **anions.*** (A good mnemonic device is to think of the "t" in cation as a plus (+) sign and the first "n" in anion as a negative (−) sign.) Substances such

2. In general, when a salt is formed, the ending of the name of the negatively charged ion is changed to *-ide*.

***ion** (eye′-on) Gr. *ion,* going.

***cation** (kat′-eye-on) An ion that migrates toward the negative pole, or cathode, of an electrical field.

***anion** (an′-eye-on) An ion that migrates toward the positive pole, or anode.

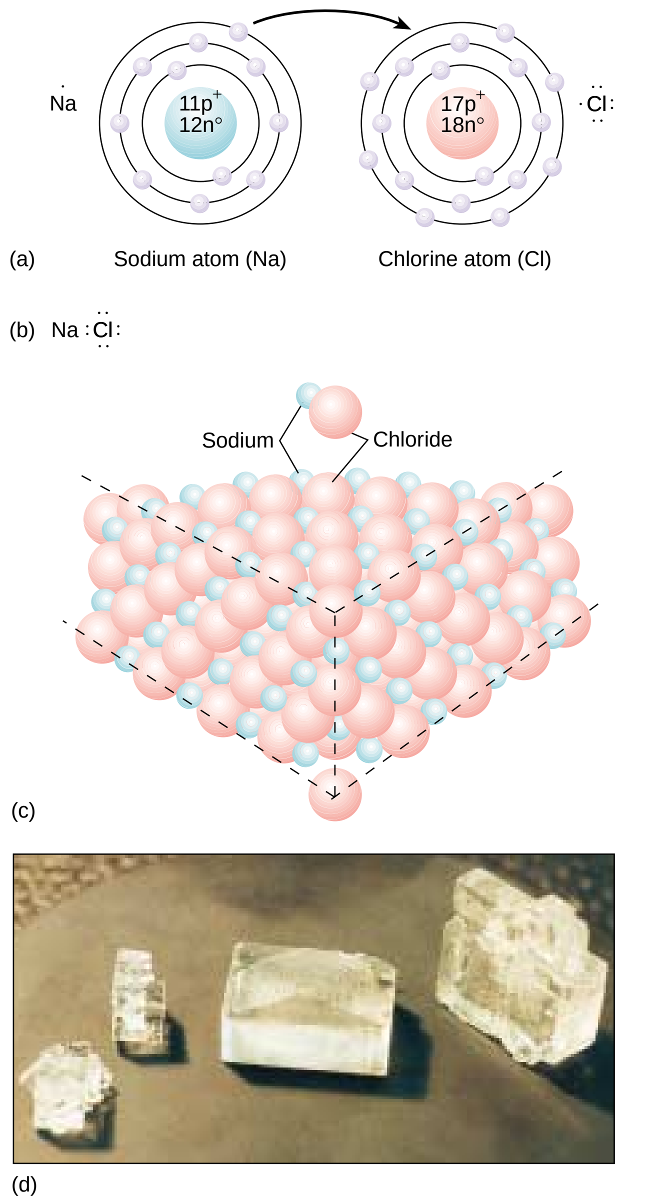


FIGURE 2.6

Ionic bonding between sodium and chlorine. (a) When the two elements are placed together, sodium loses its single outer orbital electron to chlorine, thereby filling chlorine's outer shell. **(b)** Simple model of ionic bonding. **(c)** Sodium and chloride ions form large molecules, or crystals, in which the two atoms alternate in a definite, regular, geometric pattern. **(d)** Note the cubic nature of NaCl crystals at the macroscopic level.

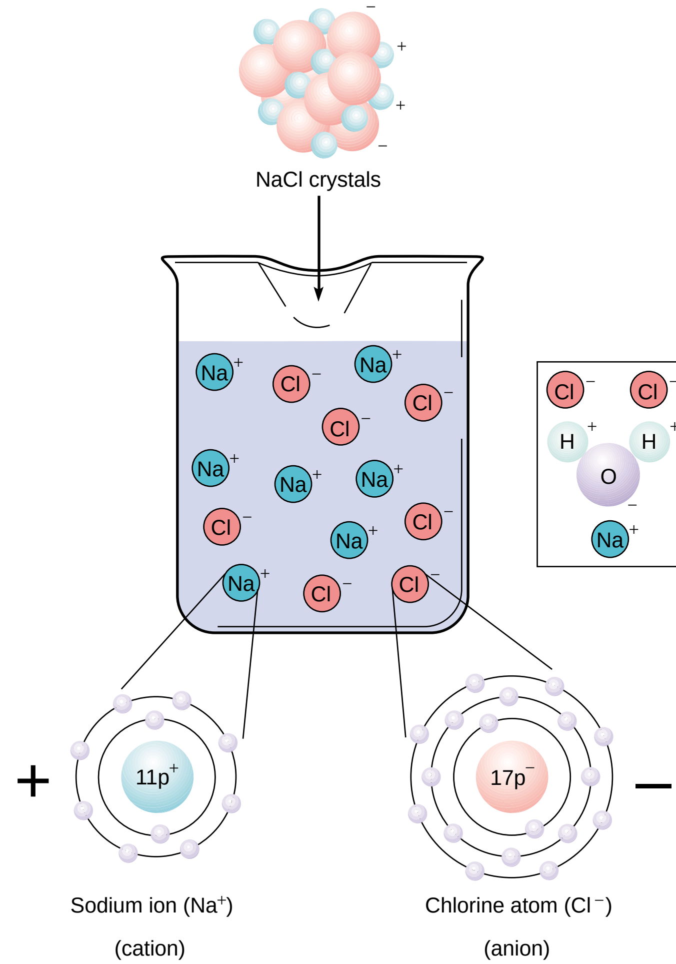


FIGURE 2.7

Ionization. When NaCl in the crystalline form is added to water, the ions are released from the crystal as separate charged particles (cations and anions) into solution. (See also figure 2.11.)

as salts, acids, and bases that release ions when dissolved in water are termed **electrolytes** because their charges enable them to conduct an electrical current. Owing to the general rule that particles of like charge repel each other and those of opposite charge attract each other, we can expect ions to interact electrostatically with other ions and polar molecules. Such interactions are important in many cellular chemical reactions, in the formation of solutions, and in the reactions microorganisms have with dyes. The transfer of electrons from one molecule to another constitutes a significant mechanism by which biological systems store and release energy (see Microbits 2.2).

Hydrogen Bonding Some types of bonding involve neither sharing, losing, nor gaining electrons but instead are due to attractive forces between nearby molecules or atoms. One such bond is a **hydrogen bond,** a weak type of bond that forms between a hydrogen covalently bonded to one molecule and an oxygen or nitrogen atom on the same molecule or on a different molecule. Because hydrogen in a covalent bond tends to be positively charged, it will attract a nearby negatively charged atom and form an easily disrupted bridge with it. This type of bonding is usually represented in molecular models with a dotted line. A simple example of hydrogen bonding occurs between water molecules (figure 2.8). More extensive hydrogen bonding is partly responsible for the structure and stability of proteins and nucleic acids (see figure 2.22*b* and 2.25).

Chemical Shorthand: Formulas, Models, and Equations

The atomic content of molecules can be represented by a few convenient **formulas.** We have already been exposed to the molecular formula, which concisely gives the atomic symbols and the number of the elements involved in subscript (CO_2, H_2O). More complex molecules such as glucose ($C_6H_{12}O_6$) can also be symbolized this

MICROBITS 2.2

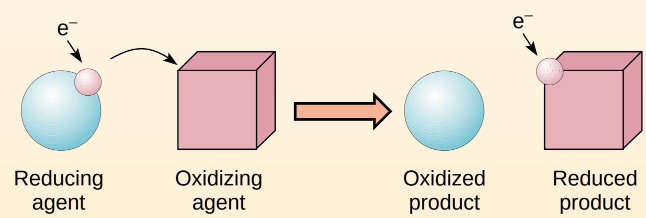
Redox: Electron Transfer and Oxidation-Reduction Reactions

The metabolic work of cells, such as synthesis, movement, and digestion, revolves around energy exchanges and transfers. The management of energy in cells is almost exclusively dependent on chemical rather than physical reactions because most cells are far too delicate to operate with heat, radiation, and other more potent forms of energy. The outer-shell electrons are readily portable and easily manipulated sources of energy. It is in fact the movement of electrons from molecule to molecule that accounts for most energy exchanges in cells. Fundamentally, then, a cell must have a supply of atoms that can gain or lose electrons if they are to carry out life processes.

The phenomenon in which electrons are transferred from one atom or molecule to another is termed an **oxidation** and **reduction** (shortened to **redox**) **reaction.** Although the term *oxidation* was originally adopted for reactions involving the addition of oxygen, the term oxidation can include any reaction causing electron release, regardless of the involvement of oxygen. By comparison, reduction is any reaction that causes an atom to receive electrons. All redox reactions occur in pairs. To analyze the phenomenon, let us again review the production of NaCl, but from a different standpoint. Although it is true that these atoms form ionic bonds, the chemical combination of the two is also a type of redox reaction.

When these two atoms react to form sodium chloride, a sodium atom gives up an electron to a chlorine atom. During this reaction, sodium is oxidized because it loses an electron, and chlorine is reduced because it gains an electron. To take this definition further, an atom or molecule, such as sodium, that can donate electrons and thereby reduce another molecule is a **reducing agent;** one that can receive extra electrons and thereby oxidize another molecule is an **oxidizing agent.** You may find this concept easier to keep straight if you think of redox agents as partners: The one that gives its electrons away is oxidized; the partner that receives the electrons is reduced. (A mnemonic device to keep track of this is *LEO* says *GER*. "Lose Electrons Oxidized; Gain Electrons Reduced."

Redox reactions are essential to many of the biochemical processes discussed in chapter 8. In cellular metabolism, electrons alone can be transferred from one molecule to another as described here, but sometimes oxidation and reduction occur with the transfer of hydrogen atoms (which are a proton and an electron) from one compound to another.



Simplified diagram of the exchange of electrons during an oxidation-reduction reaction.

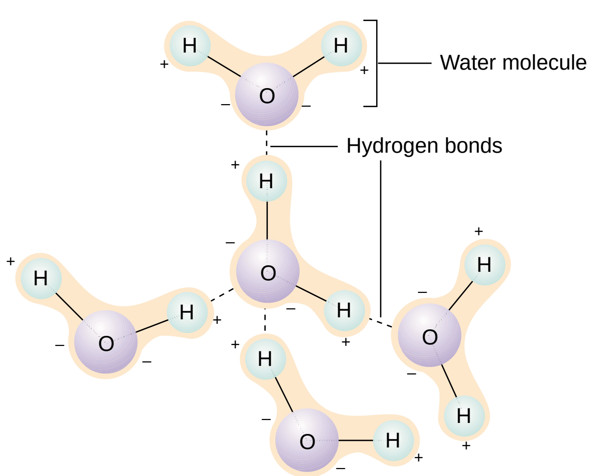


FIGURE 2.8

Hydrogen bonding in water. Because of the polarity of water molecules, the negatively charged oxygen end of one water molecule is weakly attracted to the positively charged hydrogen end of an adjacent water molecule.

way, but this formula is not unique, since fructose and galactose also share it. Molecular formulas are useful, but they only summarize the atoms in a compound; they do not show the position of bonds between atoms. For this purpose, chemists use structural formulas illustrating the relationships of the atoms and the number and types of bonds (figure 2.9). Other structural models present the three-dimensional appearance of a molecule, illustrating the orientation of atoms (differentiated by color codes) and the molecule's overall shape (figure 2.10).

The printed page tends to make molecules appear static, but this picture is far from correct, because molecules are capable of changing through chemical reactions. For ease in tracing chemical exchanges between atoms or molecules, and to derive some sense of the dynamic character of reactions, chemists use shorthand **equations** containing symbols, numbers, and arrows to simplify or summarize the major characteristics of a reaction. Molecules entering or starting a reaction are called **reactants,** and substances left by a reaction are called **products.** In most instances, summary chemical reactions do not give the details of the exchange, in order to keep the expression simple and to save space.

In a **synthesis*** **reaction,** the reactants bond together in a manner that produces an entirely new molecule (reactant A plus reactant B yields product AB). An example is the production of sulfur

*__synthesis__ (sin′-thuh-sis) Gr. *synthesis,* putting together.

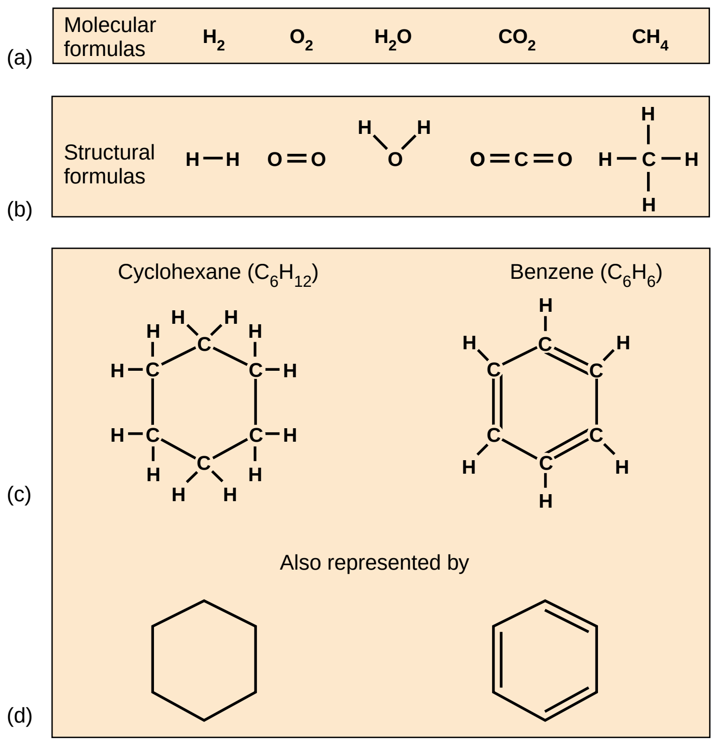


FIGURE 2.9

Comparison of molecular and structural formulas. (a) Molecular formulas provide a brief summary of the elements in a compound. **(b)** Structural formulas clarify the exact relationships of the atoms in the molecule, depicting single bonds by a single line and double bonds by two lines. **(c)** In structural formulas of organic compounds, cyclic or ringed compounds may be completely labeled, or **(d)** they may be presented in a shorthand form in which carbons are assumed to be at the angles and attached to hydrogens. See figure 2.14 for structural formulas of three sugars with the same molecular formula, $C_6H_{12}O_6$.

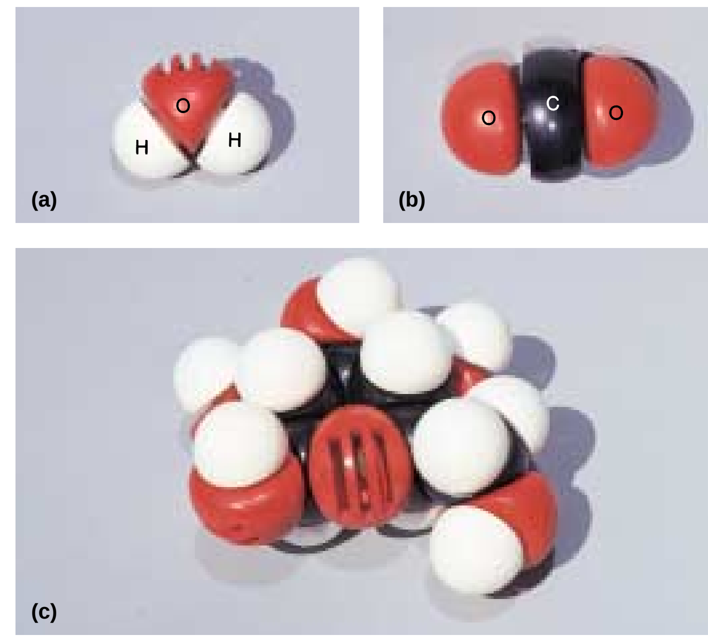


FIGURE 2.10

Three-dimensional, or space-filling, models of (a) water, (b) carbon dioxide, and (c) glucose. By convention, the red atoms are oxygen, the white ones hydrogen, and the black ones carbon.

dioxide, a by-product of burning sulfur fuels and an important component of smog:

$$S + O_2 \rightarrow SO_2$$

Some synthesis reactions are not such simple combinations. When water is synthesized, for example, the reaction does not really involve one oxygen atom combining with two hydrogen atoms, because elemental oxygen exists as O_2 and elemental hydrogen exists as H_2. A more accurate equation for this reaction is:

$$2H_2 + O_2 \rightarrow 2H_2O$$

The equation for reactions must be balanced—that is, the number of atoms on one side of the arrow must equal the number on the other side to reflect all of the participants in the reaction. To arrive at the total number of atoms in the reaction, multiply the prefix number by the subscript number; if no number is given, it is assumed to be 1.

In **decomposition reactions,** the bonds on a single reactant molecule are permanently broken to release two or more product molecules. One example is the resulting molecules when large nutrient molecules are digested into smaller units; a simpler example can be shown for the common chemical hydrogen peroxide:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

During **exchange reactions,** the reactants trade portions between each other and release products that are combinations of the two. This type of reaction occurs between acids and bases when they form water and a salt:

$$AB + XY \rightleftharpoons AX + BY$$

The reactions in biological systems can be **reversible,** meaning that reactants and products can be converted back and forth. These reversible reactions are symbolized with a double arrow, each pointing in opposite directions, as in the exchange reaction above. Whether a reaction is reversible depends on the proportions of these compounds, the difference in energy state of the reactants and products, and the presence of catalysts (substances that increase the rate of a reaction). Additional reactants coming from another reaction can also be indicated by arrows that enter or leave at the main arrow:

$$\overset{CD \;\curvearrowright\; C}{X + Y \rightarrow XYD}$$

SOLUTIONS: HOMOGENEOUS MIXTURES OF MOLECULES

A **solution** is a mixture of one or more substances called **solutes** uniformly dispersed in a dissolving medium called a **solvent.** An important characteristic of a solution is that the solute cannot be

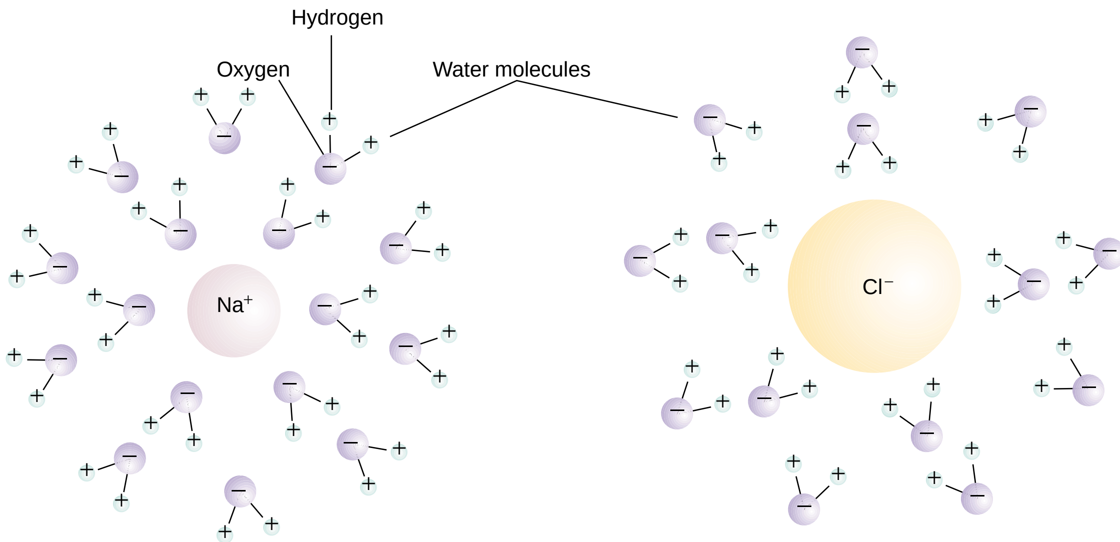


FIGURE 2.11

Hydration spheres formed around ions in solution. In this example, a sodium cation attracts the negatively charged region of water molecules, and a chloride anion attracts the positively charged region of water molecules. In both cases, the ions become covered with spherical layers of specific numbers and arrangements of water molecules.

separated by filtration or ordinary settling. The solute can be gaseous, liquid, or solid, and the solvent is usually a liquid. Examples of solutions are salt or sugar dissolved in water and iodine dissolved in alcohol. In general, a solvent will dissolve a solute only if it has similar electrical characteristics as indicated by the rule of solubility, expressed simply as "like dissolves like." For example, water is a polar molecule and will readily dissolve an ionic solute such as NaCl, yet a nonpolar solvent such as benzene will not dissolve NaCl.

Water is the most common solvent in natural systems, having several characteristics that suit it to this role. The polarity of the water molecule causes it to form hydrogen bonds with other water molecules, but it can also interact readily with charged or polar molecules. When an ionic solute such as NaCl crystals is added to water, it is **dissolved,** thereby releasing Na^+ and Cl^- into solution. Dissolution occurs because Na^+ is attracted to the negative pole of the water molecule and Cl^- is attracted to the positive pole; in this way, they are drawn away from the crystal separately into solution. As it leaves, each ion becomes **hydrated,** which means that it is surrounded by a sphere of water molecules (figure 2.11). Molecules such as salt or sugar that attract water to their surface are termed **hydrophilic.*** Nonpolar molecules, such as benzene, that repel water are considered **hydrophobic.*** A third class of molecules, such as the phospholipids in cell membranes, are considered *amphipathic** because they have both hydrophilic and hydrophobic properties.

Because most biological activities take place in aqueous (water-based) solutions, the concentration of these solutions can be very important (see chapter 7). The **concentration** of a solution expresses the amount of solute dissolved in a certain amount of solvent. It can be calculated by weight, volume, or percentage. A common way to calculate percentage of concentration is to use the weight of the solute, measured in grams (g), dissolved in a specified volume of solvent, measured in milliliters (ml). For example, dissolving 3 g of NaCl in 100 ml of water produces a 3% solution; dissolving 30 g in 100 ml produces a 30% solution; and dissolving 3 g in 1,000 ml (1 liter) produces a 0.3% solution. A solution with a small amount of solute and a relatively greater amount of solvent (0.3%) is considered dilute or weak. On the other hand, a solution containing significant percentages of solute (30%) is considered concentrated or strong.

A common way to express concentration of biological solutions is by its molar concentration, or *molarity* (M). A standard molar solution is obtained by dissolving one *mole,* defined as the molecular weight of the compound in grams, in 1 L (1,000 ml) of solution. To make a 1 M solution of sodium chloride, we would dissolve 58 g of NaCl to give 1 L of solution; a 0.1 M solution would require 5.8 g of NaCl in 1 L of solution.

*hydrophilic (hy-droh-fil′-ik) Gr. *hydros,* water, and *philos,* to love.

*hydrophobic (hy-droh-fob′-ik) Gr. *phobos,* fear.

*amphipathic (am′-fy-path′-ik) Gr. *amphi,* both.

ACIDITY, ALKALINITY, AND THE pH SCALE

Another factor with far-reaching impact on living things is the concentration of acidic or basic solutions in their environment. To understand how solutions develop acidity or basicity, we must look again at the behavior of water molecules. Hydrogens and oxygen tend to remain bonded by covalent bonds, but in certain instances, a single hydrogen can break away as the ionic form (H^+), leaving the remainder of the molecule in the form of an OH^- ion. The H^+ ion is positively charged because it is essentially a hydrogen ion that has lost its electron; the OH^- is negatively charged because it remains in possession of that electron. Ionization of water is constantly occurring, but in pure water containing no other ions,

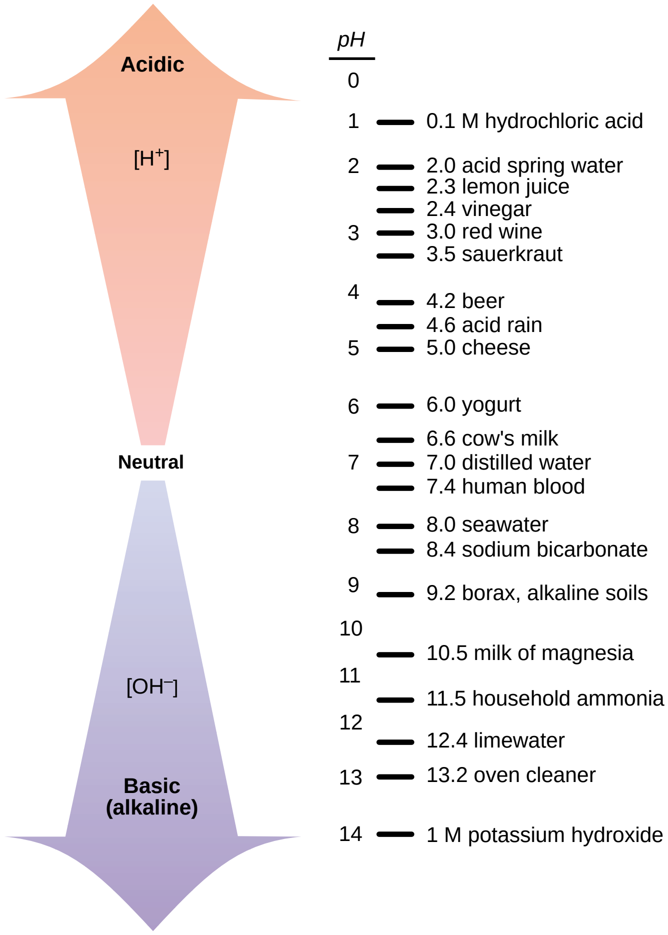


FIGURE 2.12

The pH scale. Shown are the relative degrees of acidity and basicity and the approximate pH readings for various substances.

H^+ and OH^- are produced in equal amounts, and the solution remains neutral. By one definition, a solution is considered **acidic** when a component dissolved in water (acid) releases excess hydrogen ions[3] (H^+); a solution is **basic** when a component releases excess hydroxyl ions (OH^-), so that there is no longer a balance between the two ions.

To measure the acid and base concentrations of solutions, scientists use the **pH scale,** a graduated numerical scale that ranges from 0 (the most acidic) to 14 (the most basic). This scale is a useful standard for rating relative acidity and basicity; use figure 2.12 to familiarize yourself with the pH readings of some common substances. It is not an arbitrary scale but actually a mathematical derivation based on the negative logarithm (reviewed in appendix B) of the concentration of H^+ ions in moles per liter (symbolized as $[H^+]$) in a solution, represented as:

$$\text{pH} = -\log[H^+]$$

Acidic solutions have a greater concentration of H^+ than OH^-, starting with pH 0, which contains 1.0 moles H^+/1. Each of

3. Actually, it forms a hydronium ion (H_3O^+), but for simplicity's sake, we will use the notation of H^+.

TABLE 2.2

Hydrogen Ion and Hydroxyl Ion Concentrations at a Given pH

Moles/L of Hydrogen Ions	Logarithm	pH	Moles/L of OH^-
1.0	10^{-0}	0	10^{-14}
0.1	10^{-1}	1	10^{-13}
0.01	10^{-2}	2	10^{-12}
0.001	10^{-3}	3	10^{-11}
0.0001	10^{-4}	4	10^{-10}
0.00001	10^{-5}	5	10^{-9}
0.000001	10^{-6}	6	10^{-8}
0.0000001	10^{-7}	7	10^{-7}
0.00000001	10^{-8}	8	10^{-6}
0.000000001	10^{-9}	9	10^{-5}
0.0000000001	10^{-10}	10	10^{-4}
0.00000000001	10^{-11}	11	10^{-3}
0.000000000001	10^{-12}	12	10^{-2}
0.0000000000001	10^{-13}	13	10^{-1}
0.00000000000001	10^{-14}	14	10^{-0}

the subsequent whole-number readings in the scale changes in $[H^+]$ by a tenfold reduction, so that pH 1 contains [0.1 moles H^+/1], pH 2 contains [0.01 moles H^+/1], and so on, continuing in the same manner up to pH 14, which contains [0.00000000000001 moles H^+/1]. These same concentrations can be represented more manageably by exponents: pH 2 has a $[H^+]$ of 10^{-2} moles, and pH 14 has a $[H^+]$ of 10^{-14} moles (table 2.2). It is evident that the pH units are derived from the exponent itself. Even though the basis for the pH scale is $[H^+]$, it is important to note that, as the $[H^+]$ in a solution decreases, the $[OH^-]$ increases in direct proportion. At midpoint—pH 7, or neutrality—the concentrations are exactly equal and neither predominates, this being the pH of pure water previously mentioned.

In summary, the pH scale can be used to rate or determine the degree of acidity or basicity (also called alkalinity) of a solution. On this scale, a pH below 7 is acidic, and the lower the pH, the greater the acidity; a pH above 7 is basic, and the higher the pH, the greater the basicity. Incidentally, although pHs are given here in even whole numbers, more often, a pH reading exists in decimal form; for example, pH 4.5 or 6.8 (acidic) and pH 7.4 or 10.2 (basic). Because of the damaging effects of very concentrated acids or bases, most cells operate best under neutral, weakly acidic, or weakly basic conditions (see chapter 7).

Aqueous solutions containing both acids and bases may be involved in **neutralization** reactions, which give rise to water and other neutral by-products. For example, when equal molar solutions of hydrochloric acid (HCl) and sodium hydroxide (NaOH, a base) are mixed, the reaction proceeds as follows:

$$HCl + NaOH \rightarrow H_2O + NaCl$$

Here the acid and base ionize to H^+ and OH^- ions, which form water, and other ions, Na^+ and Cl^-, which form sodium chloride. Any product other than water that arises when acids and bases react is called a **salt.** Many of the organic acids (such as lactic and

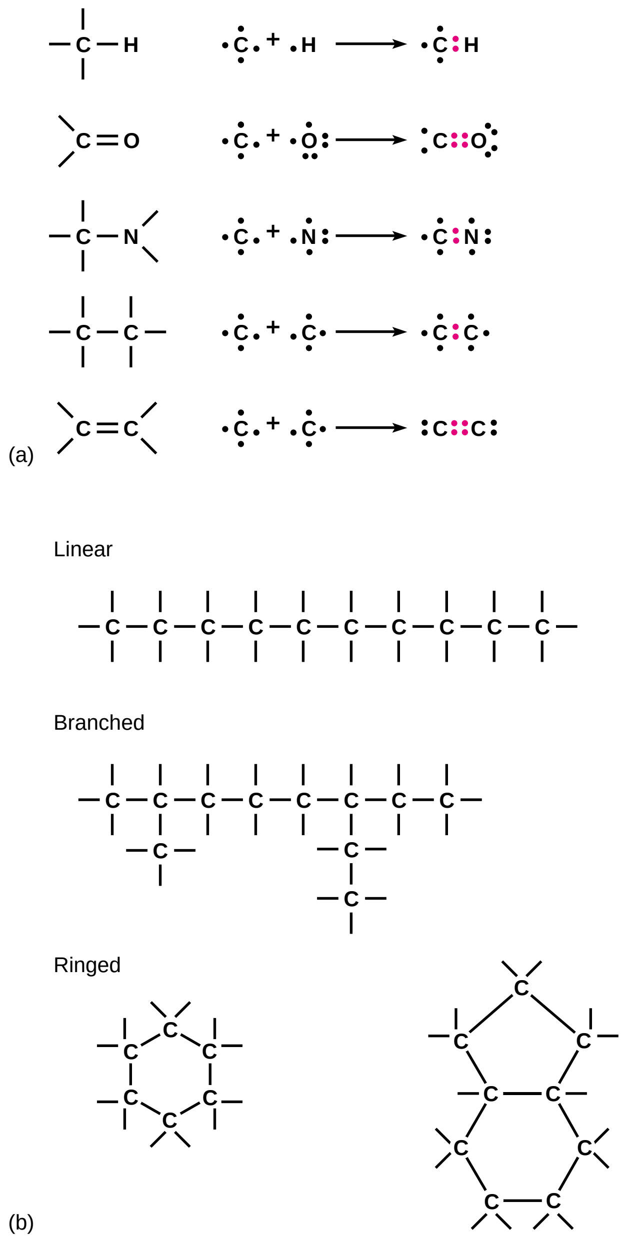


FIGURE 2.13

The versatility of bonding in carbon. In most compounds, each carbon makes a total of four bonds. **(a)** Both single and double bonds can be made with other carbons, oxygen, and nitrogen; single bonds are made with hydrogen. Simple electron models show how the electrons are shared in these bonds. **(b)** Multiple bonding of carbons can give rise to long chains, branched compounds, and ringed compounds, many of which are extraordinarily large and complex.

succinic acids) that function in **metabolism*** are available as the acid and the salt form (such as lactate, succinate), depending on the conditions in the cell (see chapter 8).

*metabolism (muh-tab′-oh-lizm) A general term referring to the totality of chemical and physical processes occurring in the cell.

*inorganic (in-or-gan′-ik) Any chemical substances that do not contain both carbon and hydrogen.

TABLE 2.3

Representative Functional Groups and Classes of Organic Compounds

Formula of Functional Group	Name	Class of Compounds
R* — O — H	Hydroxyl	Alcohols, carbohydrates
R — C(=O) — OH	Carboxyl	Fatty acids, proteins, organic acids
R — CH$_2$ — NH$_2$	Amino	Proteins, nucleic acids
R — C(=O) — O — R	Ester	Lipids
R — CH$_2$ — SH	Sulfhydryl	Cysteine (amino acid), proteins
R — C(=O) — H	Carbonyl, terminal end	Aldehydes, polysaccharides
R — C(=O) — C —	Carbonyl, internal	Ketones, polysaccharides
R — O — P(=O)(OH) — OH	Phosphate	DNA, RNA, ATP

**The R designation on a molecule is shorthand for residue, and its placement in a formula indicates that what is attached at that site varies from one compound to another.*

THE CHEMISTRY OF CARBON AND ORGANIC COMPOUNDS

So far, our main focus has been on the characteristics of atoms, ions, and small, simple substances that play diverse roles in the structure and function of living things. These substances are often lumped together in a category called **inorganic* chemicals.** Examples of inorganic chemicals include NaCl (sodium chloride), $Mg_3(PO_4)_2$ (magnesium phosphate), $CaCO_3$ (calcium carbonate), and CO_2 (carbon dioxide). In reality, however, most of the chemi-

TABLE 2.4

Macromolecules and Their Functions

Macromolecule	Description/Basic Structure	Examples/Functions
Carbohydrates		
Monosaccharides	3–7-carbon sugars	Glucose, fructose / Sugars involved in metabolic reactions; building block of disaccharides and polysaccharides
Disaccharides	Two monosaccharides	Maltose (malt sugar) / Composed of two glucoses; an important breakdown product of starch Lactose (milk sugar / Composed of glucose and galactose Sucrose (table sugar) / Composed of glucose and fructose
Polysaccharides	Chains of monosaccharides	Starch, cellulose, glycogen / Cell wall, food storage
Lipids		
Triglycerides	Fatty acids + glycerol	Fats, oils / Major component of cell membranes; storage
Phospholipids	Fatty acids + glycerol + phosphate	Membranes
Waxes	Fatty acids, alcohols	Mycolic acid / Cell wall of mycobacteria
Steroids	Ringed structure (not a polymer)	Cholesterol, ergosterol / Membranes of eucaryotes and some bacteria
Proteins	Amino acids	Enzymes; part of cell membrane, cell wall, ribosomes, antibodies / Metabolic reactions; structural components
Nucleic acids	Pentose sugar + phosphate + nitrogenous base Purines; adenine, guanine Pyrimidines: cytosine, thymine, uracil	
Deoxyribonucleic acid (DNA)	Contains deoxyribose sugar and thymine, not uracil	Chromosomes; genetic material of viruses / Inheritance
Ribonucleic acid (RNA)	Contains ribose sugar and uracil, not thymine	Ribosomes; mRNA, tRNA / Expression of genetic traits

cal reactions and structures of living things occur at the level of more complex molecules, termed **organic*** **chemicals.** These are defined as molecules that contain a basic framework of the elements carbon and hydrogen. Organic molecules vary in complexity from the simplest, methane (CH_4; figure 2.4*c*), which has a molecular weight of 16, to certain antibody molecules (produced by an immune reaction) that have a molecular weight of nearly 1,000,000 and are among the most complex molecules on earth.

The role of carbon as the fundamental element of life can best be understood if we look at its chemistry and bonding patterns. The valence of carbon makes it an ideal atomic building block to form the backbone of organic molecules; it has 4 electrons in its outer orbital to be shared with other atoms (including other carbons) through covalent bonding. As a result, it can form stable chains containing thousands of carbon atoms and still has bonding sites available for forming covalent bonds with numerous other atoms. The bonds that carbon forms are linear, branched, or ringed, and it can form four single bonds, two double bonds, or one triple bond (figure 2.13). The atoms with which carbon is most often associated in organic compounds are hydrogen, oxygen, nitrogen, sulfur, and phosphorus.

*__organic__ (or-gan′-ik) Gr. *organikos,* instrumental.

FUNCTIONAL GROUPS OF ORGANIC COMPOUNDS

One important advantage of carbon's serving as the molecular skeleton for living things is that it is free to bind with an unending array of other molecules. These special molecular groups or accessory molecules that bind to organic compounds are called **functional groups.** Functional groups help define the chemical class of certain groups of organic compounds and confer unique reactive properties on the whole molecule (table 2.3). Because each type of functional group behaves in a distinctive manner, reactions of an organic compound can be predicted by knowing the kind of functional group or groups it carries. Many synthesis, decomposition, and transfer reactions rely upon functional groups such as R—OH or $R—NH_2$. The **—R** designation on a molecule is shorthand for residue, and its placement in a formula indicates that the group attached at that site varies from one compound to another.

CHAPTER CHECKPOINTS

Covalent bonds are chemical bonds in which electrons are shared between atoms. Equally distributed electrons form nonpolar covalent bonds, whereas unequally distributed electrons form polar covalent bonds.

Ionic bonds are chemical bonds in which the outer electron shell either donates or receives electrons from another atom so that the outer shell of each atom is completely filled.

Hydrogen bonds are weak chemical bonds that form between covalently bonded hydrogens and either oxygens or nitrogens on different molecules.

Chemical energy is generated by the movement of electrons from one atom or molecule to another. Chemical equations express movement of energy in chemical reactions such as synthesis or decomposition reactions.

Solutions are mixtures of solutes and solvents that cannot be separated by filtration or settling.

The pH, ranging from a highly *acidic* solution to highly *basic* solution, refers to the concentration of hydrogen ions. It is expressed as a number from 0 to 14.

Biologists define organic molecules as those containing both carbon and hydrogen.

Carbon is the backbone of biological compounds because of its ability to form single, double, or triple covalent bonds with itself and many different elements.

Functional (R) groups are specific arrangements of organic molecules that confer distinct properties, including chemical reactivity, to organic compounds.

Macromolecules: Superstructures of Life

The compounds of life fall into the realm of **biochemistry.** Biochemicals are organic compounds produced by (or components of) living things, and they include four main families: carbohydrates, lipids, proteins, and nucleic acids (table 2.4). The compounds in these groups are assembled from smaller molecular subunits, or building blocks, and because they are often very large compounds, they are termed **macromolecules.** All macromolecules except lipids are formed by **polymerization,** a process in which repeating subunits termed **monomers*** are bound into chains of various lengths termed **polymers.*** For example, proteins (polymers) are composed of a chain of amino acids (monomers) (see figure 2.22*a*). The large size and complex, three-dimensional shape of macromolecules enables them to function as structural components, molecular messengers, energy sources, enzymes (biochemical catalysts), nutrient stores, and sources of genetic information. In the following section and in later chapters, we will consider numerous concepts relating to the roles of macromolecules in cells.

CARBOHYDRATES: SUGARS AND POLYSACCHARIDES

The term **carbohydrate** originates from the way that most members of this chemical class resemble combinations of carbon and water. Although carbohydrates can be generally represented by the formula $\mathbf{(CH_2O)_n}$, in which *n* indicates the number of units of this combination of atoms, some carbohydrates contain additional atoms of sulfur or nitrogen. In molecular configuration, the carbons form chains or rings with two or more hydroxyl groups and either an aldehyde or a ketone group, giving them the technical designation of *polyhydroxy aldehydes* or *ketones* (figure 2.14).

Carbohydrates exist in a great variety of configurations. The common term **sugar** (*saccharide*)* refers to a simple carbohydrate such as a monosaccharide or a disaccharide that has a sweet taste. A **monosaccharide** is a simple polyhydroxy aldehyde or ketone molecule containing from 3 to 7 carbons; a **disaccharide** is a combination of two monosaccharides; and a **polysaccharide** is a polymer of five or more monosaccharides bound in linear or branched chain patterns (figure 2.14). Monosaccharides and disaccharides are specified by combining a prefix that describes some characteristic of the sugar with the suffix **-ose.** For example, **hexoses** are composed of 6 carbons, and **pentoses** contain 5 carbons. **Glucose** (Gr. sweet) is the most common and universally important hexose; **fructose** is named for fruit (one of its sources); and xylose, a pentose, derives its name from the Greek word for wood. Disaccharides are named similarly: **lactose** (L. milk) is an important component of milk; **maltose** means malt sugar; and **sucrose** (Fr. sugar) is common table sugar or cane sugar.

The Nature of Carbohydrate Bonds

The subunits of disaccharides and polysaccharides are linked by means of **glycosidic bonds,** in which carbons (each is assigned a number) on adjacent sugar units are bonded to the same oxygen atom like links in a chain (figure 2.15). For example, maltose is formed when the number 1 carbon on a glucose bonds to the oxygen on the number 4 carbon on a second glucose; sucrose is formed when glucose and fructose bind oxygen between their number 1 and number 2 carbons; and lactose is formed when glucose and galactose connect by their number 1 and number 4 carbons. In order to form this bond, one carbon gives up its OH group and the other (the one contributing the oxygen to the bond) loses the H from its OH group. Because a water molecule is produced, this reaction is known as **dehydration synthesis,** a process common to most polymerization reactions (see proteins, page 47). Three polysaccharides (starch, cellulose, and glycogen) are structurally and biochemically distinct, even though all are polymers of the same monosaccharide—glucose. The basis for their differences lies primarily in the exact way the glucoses are bound together, which greatly affects the characteristics of the end product (figure 2.16). The synthesis and breakage of each type of bond requires a specialized catalyst called an enzyme (see chapter 8).

The Functions of Polysaccharides

Polysaccharides typically contribute to structural support and protection and serve as nutrient and energy stores. The cell walls in plants and many microscopic algae derive their strength and rigidity from **cellulose,** a long, fibrous polymer (figure 2.16*a*). Because of this role, cellulose is probably one of the most common organic substances on the earth, yet it is digestible only by certain bacteria, fungi, and protozoa. These microbes, called decomposers, play an essential role in breaking down and recycling plant materials (see figure 7.2). Some bacteria secrete slime layers of a glucose polymer

*monomer (mahn′-oh-mur) Gr. *mono,* one, and *meros,* part.

*polymer (pahl′-ee-mur) Gr. *poly,* many; also the root for polysaccharide and polypeptide.

*saccharide (sak′-uh-ryd) Gr. *sakcharon,* sweet.

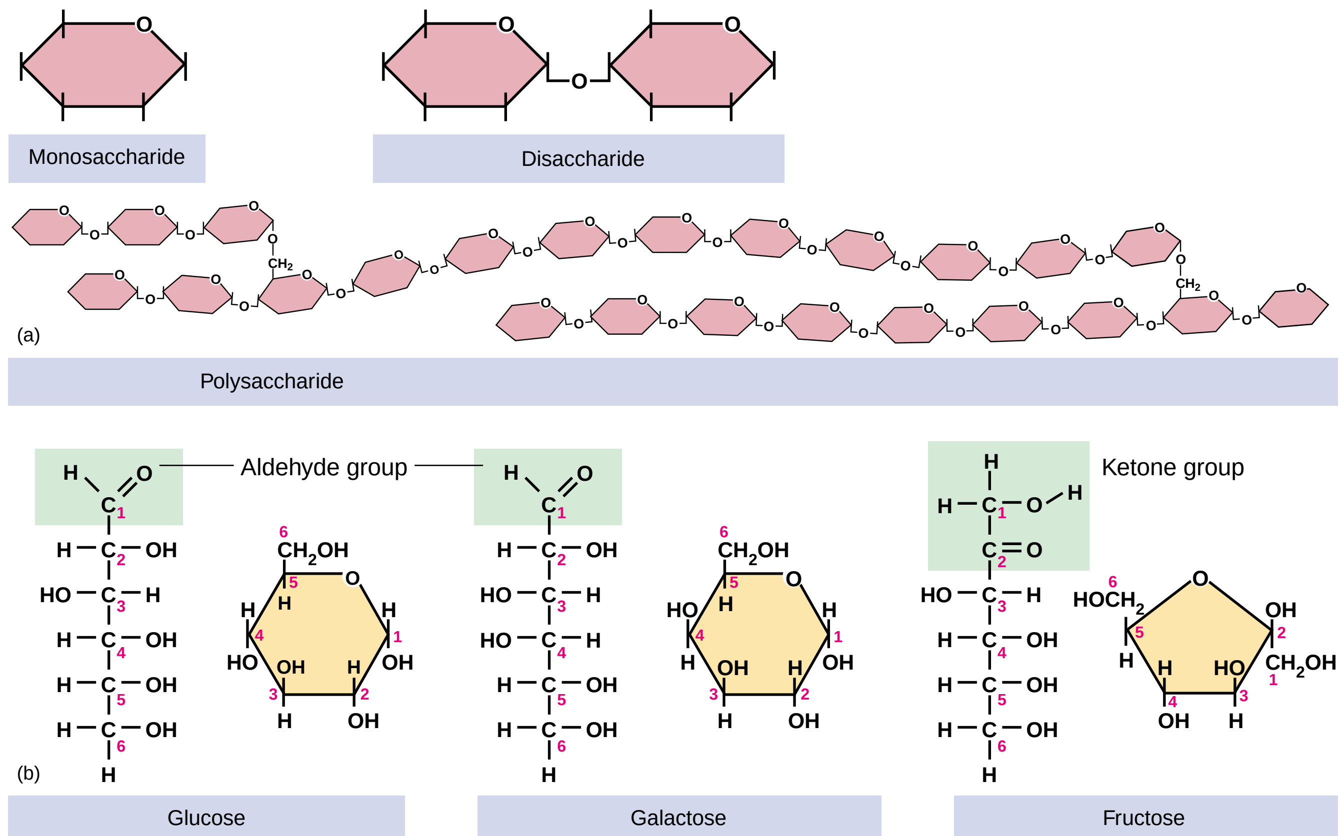


FIGURE 2.14

Common classes of carbohydrates. (a) Major saccharide groups, named for the number of sugar units each contains. **(b)** Three hexoses with the same molecular formula and different structural formulas. Both linear and ring models are given. The linear form indicates aldehyde and ketone groups, although in solution the sugars exist in the ring form. Note that the carbons are numbered so as to keep track of reactions within and between monosaccharides.

called *dextran*. This substance causes a sticky layer to develop on teeth that leads to plaque (see figure 4.12).

Other structural polysaccharides can be conjugated (chemically bonded) to amino acids, nitrogen bases, lipids, or proteins. **Agar,** an indispensable polysaccharide in preparing solid culture media, is a natural component of certain seaweeds. It is a complex polymer of galactose and sulfur-containing carbohydrates. The exoskeletons of certain fungi contain **chitin,** a polymer of glucosamine (a sugar with an amino functional group). **Peptidoglycan*** is one special class of compounds in which polysaccharides (glycans) are linked to peptide fragments (a short chain of amino acids). This molecule provides the main source of structural support to the bacterial cell wall. The cell wall of gram-negative bacteria also contains **lipopolysaccharide,** a complex of lipid and polysaccharide responsible for symptoms such as fever and shock (see chapters 4 and 13).

The outer surface of many cells has a delicate "sugar coating" composed of polysaccharides bound in various ways to proteins (the combination is called mucoprotein or glycoprotein). This structure, called the **glycocalyx,*** functions in attachment to other cells or as a site for *receptors*—surface molecules that receive and respond to external stimuli. Small sugar molecules account for the differences in human blood types, and carbohydrates are a component of large protein molecules called antibodies. Some viruses have glycoproteins on their surface with which they bind to and invade their host cells.

Polysaccharides are usually stored by cells in the form of glucose polymers such as **starch** (figure 2.16*b*) or **glycogen,** but only organisms with the appropriate digestive enzymes can break them down and use them as a nutrient source. Because a water molecule is required for breaking the bond between two glucose molecules, digestion is also termed **hydrolysis.*** Starch is the primary storage food of green plants, microscopic algae, and some fungi; glycogen (animal starch) is a stored carbohydrate for animals and certain groups of bacteria.

LIPIDS: FATS, PHOSPHOLIPIDS, AND WAXES

The term **lipid,** derived from the Greek word *lipos,* meaning fat, is not a chemical designation, but an operational term for a variety of substances that are not soluble in polar solvents such as water

*peptidoglycan (pep-tih-doh-gly′-kan).

*glycocalyx (gly″-koh-kay′-lix) Gr. *glycos,* sweet, and *calyx,* covering.

*hydrolysis (hy-drol′-uh-sis) Gr. *hydros,* water, and *lyein,* to dissolve.

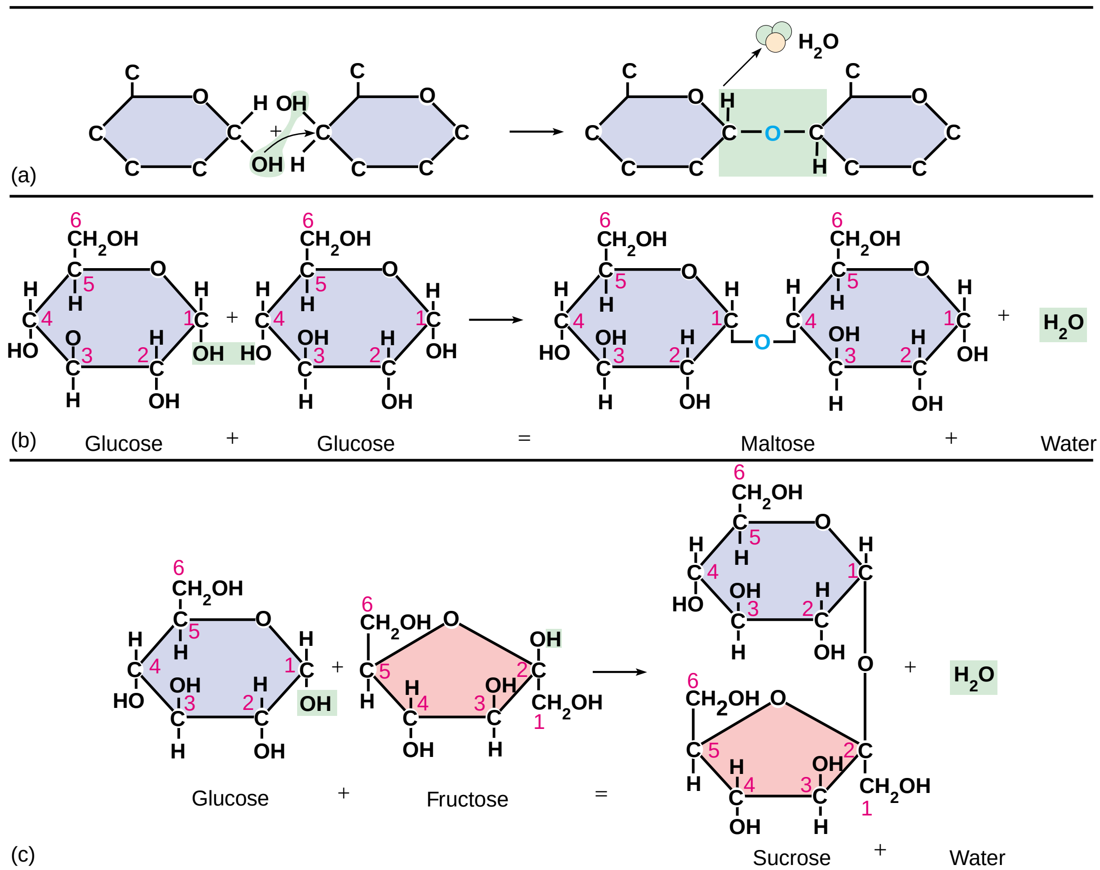
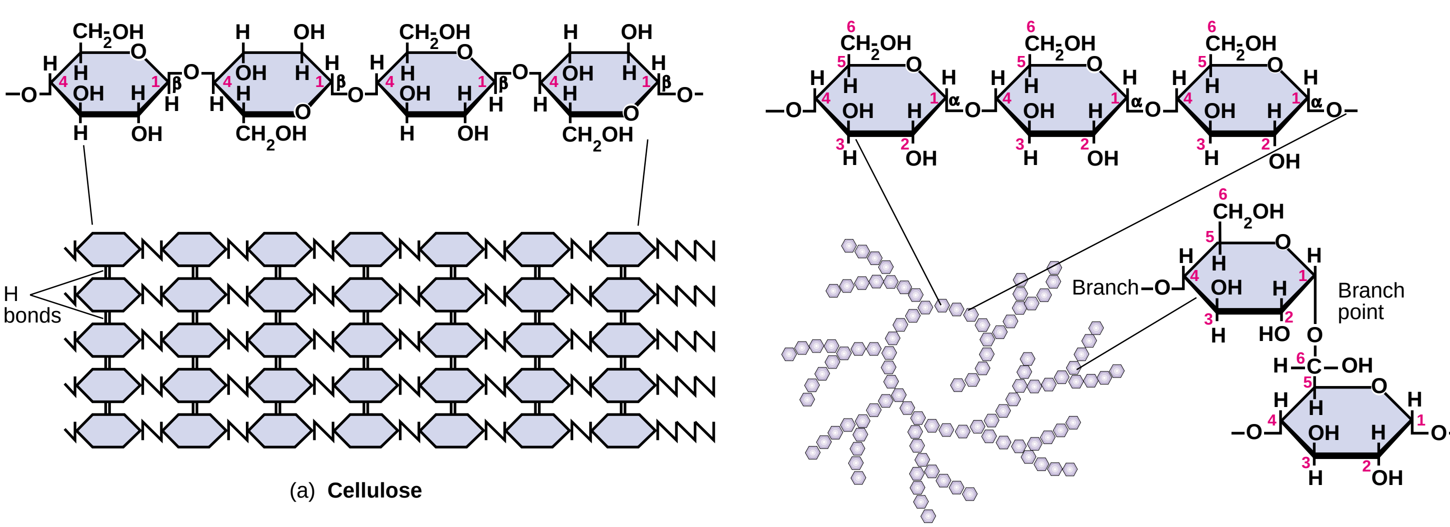


FIGURE 2.15

Glycosidic bond. (a) General scheme in the formation of a glycosidic bond by dehydration synthesis. **(b)** Formation of the 1,4 bond between two α glucoses to produce maltose and water. **(c)** Formation of the 1,2 bond between glucose and fructose to produce sucrose and water.



(a) **Cellulose**

(b) **Starch**

FIGURE 2.16

Polysaccharides. (a) Cellulose is composed of β glucose bonded in 1,4 bonds that produce linear, lengthy chains of polysaccharides that are H-bonded along their length. This is the typical structure of wood and cotton fibers. **(b)** Starch is also composed of glucose polymers, in this case α glucose. The main structure is amylose bonded in a 1,4 pattern, with side branches of amylopectin bonded by 1,6 bonds. The entire molecule is compact and granular.

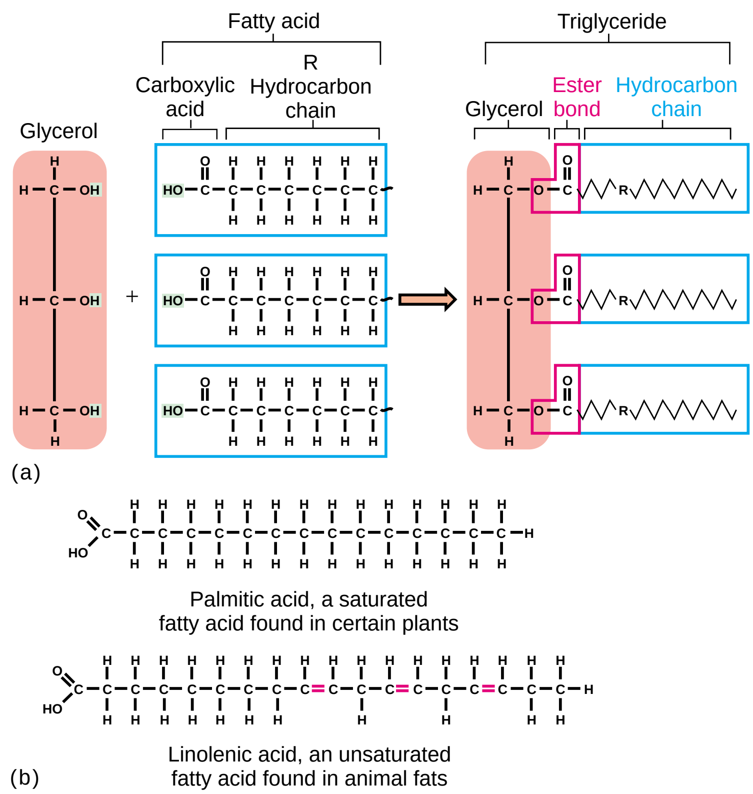


FIGURE 2.17

Synthesis and structure of a triglyceride. **(a)** Because a water molecule is released at each ester bond, this is another form of dehydration synthesis. The jagged lines and R symbol represent the hydrocarbon chains of the fatty acids, which are commonly very long. **(b)** Structural formulas for saturated and unsaturated fatty acids.

(recall that oil and water do not mix) but will dissolve in nonpolar solvents such as benzene and chloroform. This property occurs because the substances we call lipids contain relatively long or complex C—H (hydrocarbon) chains that are nonpolar and thus hydrophobic. The main groups of compounds classified as lipids are triglycerides, phospholipids, steroids, and waxes.

Important storage lipids are the **triglycerides,** a category that includes fats and oils. Triglycerides are composed of a single molecule of glycerol bound to three fatty acids (figure 2.17). **Glycerol** is a 3-carbon alcohol with three OH groups that serve as binding sites, and **fatty acids** are long-chain hydrocarbon molecules with a carboxyl group (COOH) at one end that is free to bind to the glycerol. The bond that forms between the —OH group and the —COOH is defined as an **ester bond.** The hydrocarbon portion of a fatty acid can vary in length from 4 to 24 carbons and, depending on the fat, it may be saturated or unsaturated. If all carbons in the chain are single-bonded to 2 other carbons and 2 hydrogens, the fat is saturated; if there is at least one C=C double bond in the chain, it is unsaturated. The structure of fatty acids is what gives fats and oils (liquid fats) their greasy, insoluble nature. In general, solid fats (such as beef tallow) are more saturated, and oils (or liquid fats) are more unsaturated. In most cells, triglycerides are stored in long-term concentrated form as droplets or globules. When the ester linkage is acted on by digestive enzymes called lipases, the fatty acids and glycerol are freed to be used in metabolism. Fatty acids are a superior source of energy, yielding twice as much per gram as other storage molecules (starch). Soaps are K^+ or Na^+ salts of fatty acids whose qualities make them excellent grease removers and cleaners (see chapter 11).

Membrane Lipids

A class of lipids that serves as a major structural component of cell membranes is the **phospholipids.** Although phospholipids also contain glycerol and fatty acids, they have some significant differences from triglycerides. Phospholipids contain only two fatty acids attached to the glycerol, and the third glycerol binding site holds a phosphate group. The phosphate is in turn bonded to an alcohol,[4] which varies from one phospholipid to another (figure 2.18*a*). These lipids have a hydrophilic region from the charge on the phosphoric acid–alcohol "head" of the molecule and a hydrophobic region that corresponds to the long, uncharged "tail" (formed by the fatty acids). When exposed to an aqueous solution, the charged heads are attracted to the water phase, and the nonpolar tails are repelled from the water phase (figure 2.18*b*). This property causes lipids to naturally assume single and double layers (bilayers), which contribute to their biological significance in membranes. When two single layers of polar lipids come together to form a double layer, the outer hydrophilic face of each single layer will orient itself toward the solution, and the hydrophobic portions will become immersed in the core of the bilayer. The structure of lipid bilayers confers characteristics on membranes such as selective permeability and fluid nature (Microbits 2.3).

Miscellaneous Lipids

Steroids are complex ringed compounds commonly found in cell membranes and animal hormones. The best known of these is the sterol (meaning a steroid with an OH group) called **cholesterol**

4. Alcohols are hydrocarbons containing OH groups.

FIGURE 2.18

Phosphlipids—membrane molecules. **(a)** A complex model of a single molecule of a phospholipid. The phosphate-alcohol head lends a charge to one end of the molecule; its long, trailing hydrocarbon chain is uncharged. **(b)** The behavior of phospholipids in water-based solutions causes them to become arranged **(1)** in single layers called micelles, with the charged head oriented toward the water phase and the hydrophobic nonpolar tail buried away from the water phase, or **(2)** in double-layered phospholipid systems with the hydrophobic tails sandwiched between two hydrophilic layers.

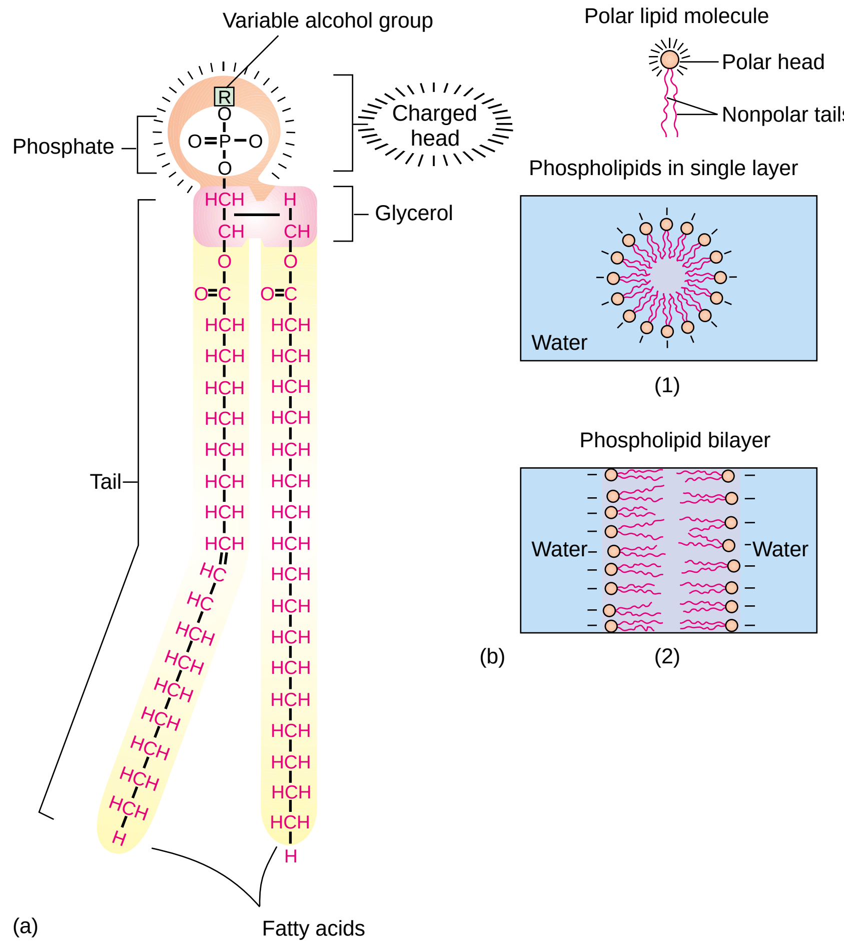
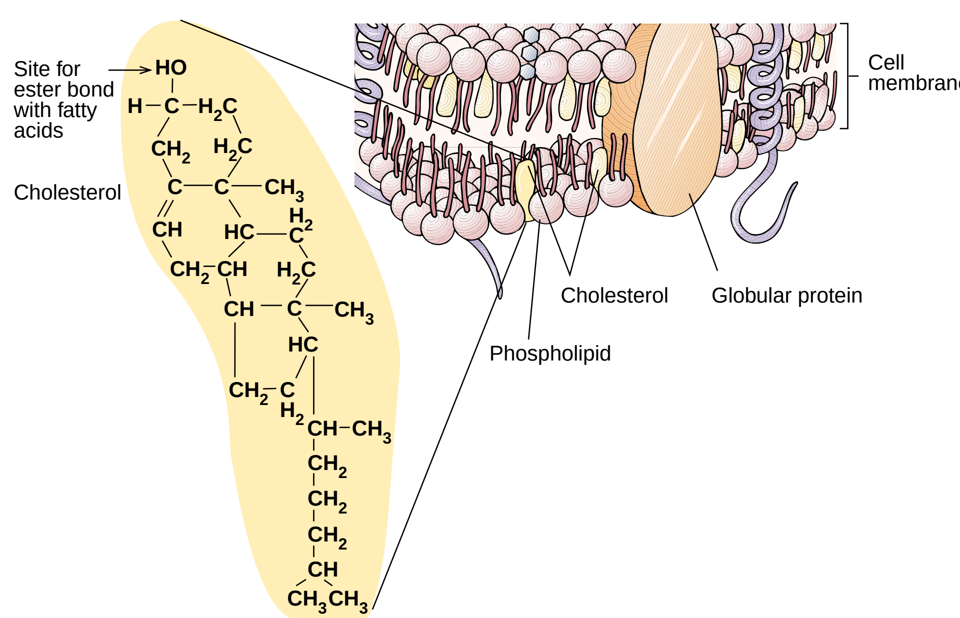


FIGURE 2.19

Formula for cholesterol, an alcoholic steroid that is inserted in some membranes. Cholesterol can become esterified with fatty acids at its OH group, imparting a polar quality similar to that of phospholipids.



MICROBITS 2.3

Membranes: Cellular Skins

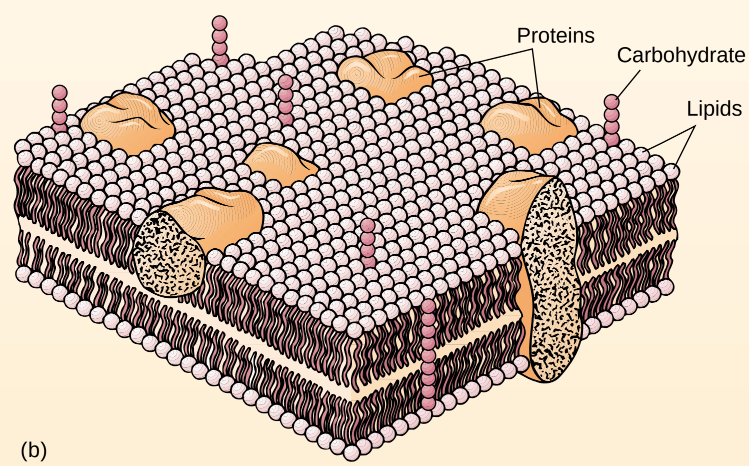
The word **membranes** appears frequently in descriptions of cells in this chapter and in chapters 4 and 5. The word itself describes any lining or covering, including such multicellular structures as the mucous membranes of the body. From the perspective of a single cell, however, a membrane is a thin, double-layered sheet composed of lipids such as phospholipids and sterols (averaging about 40% of membrane content) and protein molecules (averaging about 60%). The primary role of membranes is as a **cell membrane** that completely encases the cytoplasm. Membranes are also components of eucaryotic organelles such as nuclei, mitochondria, and chloroplasts, and they appear in internal pockets of certain procaryotic cells. Even some viruses, which are not cells at all, can have a membranous protective covering.

Cell membranes are so thin—on the average, just 0.0070 μm (7 nm) thick—that they cannot actually be seen with an optical microscope. Even at magnifications made possible by electron microscopy (500,000×), very little of the precise architecture can be visualized, and a cross-sectional view has the appearance of railroad tracks. Following detailed microscopic and chemical analysis, S. J. Singer and C. K. Nicholson proposed a simple and elegant theory for membrane structure called the **fluid mosaic*** **model.** According to this theory, a membrane is a continuous bilayer formed by lipids that are oriented with the polar lipid heads toward the outside and the nonpolar heads toward the center of the membrane. Embedded at numerous sites in this bilayer are various-sized globular proteins. Some proteins are situated only at the surface; others extend fully through the entire membrane. The configuration of the inner and outer sides of the membrane can be quite different because of the variations in protein shape and position.

Membranes are dynamic and constantly changing because the lipid phase is in motion and many proteins can migrate freely about, somewhat as icebergs do in the ocean. This fluidity is essential to such activities as engulfment of food and discharge or secretion by cells. The structure of the lipid phase provides an impenetrable barrier to many substances. This property accounts for the **selective permeability** and capacity to regulate transport of molecules. It also serves to segregate activities within the cell's cytoplasm. Membrane proteins function in receiving molecular signals (receptors), in binding and transporting nutrients, and in acting as enzymes, topics to be discussed in chapters 7 and 8.

*__mosaic__ (moh-zay′-ik) An intricate design made up of many small fragments.

(a)



(b)

Extreme magnification of **(a)** a cross section of a cell membrane, which appears as double tracks. **(b)** A generalized version of the fluid mosaic model of a cell membrane indicates a bilayer of lipids with globular proteins embedded to some degree in the lipid matrix. This structure explains many characteristics of membranes, including flexibility, solubility, permeability, and transport.

(figure 2.19). Cholesterol reinforces the structure of the cell membrane in animal cells and in an unusual group of cell-wall-deficient bacteria called the mycoplasmas (see chapter 4). The cell membranes of fungi also contain a sterol, called ergosterol. *Prostaglandins* are fatty acid derivatives found in trace amounts that function in inflammatory and allergic reactions, blood clotting, and smooth muscle contraction. Chemically, a *wax* is an ester formed between a long-chain alcohol and a saturated fatty acid. The resulting material is typically pliable and soft when warmed but hard and water-resistant when cold (paraffin, for example). Among living things, fur, feathers, fruits, leaves, human skin, and insect exoskeletons are naturally waterproofed with a coating of wax. Bacteria that cause tuberculosis and leprosy produce a wax (wax D) that repels ordinary stains and contributes to their pathogenicity.

PROTEINS: SHAPERS OF LIFE

The predominant organic molecules in cells are **proteins,** a fitting term adopted from the Greek word *proteios,* meaning first or prime. To a large extent, the structure, behavior, and unique qualities of

TABLE 2.5

Twenty Amino Acids and Their Abbreviations

Acid	Abbreviation	Characteristic of R Groups*
Alanine	Ala	NP
Arginine	Arg	+
Asparagine	Asn	P
Aspartic acid	Asp	–
Cysteine	Cys	P
Glutamic acid	Glu	–
Glutamine	Gln	P
Glycine	Gly	P
Histidine	His	+
Isoleucine	Ile	NP
Leucine	Leu	NP
Lysine	Lys	+
Methionine	Met	NP
Phenylalanine	Phe	NP
Proline	Pro	NP
Serine	Ser	P
Threonine	Thr	P
Tryptophan	Trp	NP
Tyrosine	Tyr	P
Valine	Val	NP

**NP, nonpolar; P, polar; +, positively charged; –, negatively charged.*

each living thing are a consequence of the proteins they contain. To best explain the origin of the special properties and versatility of proteins, we must examine their general structure. The building blocks of proteins are **amino acids,** which exist in 20 different naturally occurring forms (table 2.5). Various combinations of these amino acids account for the nearly infinite variety of proteins. Amino acids have a basic skeleton consisting of a carbon (called the α carbon) linked to an amino group (NH_2), a carboxyl group (COOH), a hydrogen atom (H), and a variable R group. The variations among the amino acids occur at the R group, which is different in each amino acid and imparts the unique characteristics to the molecule and to the proteins that contain it (figure 2.20). A covalent bond called a **peptide bond** forms between the amino group on one amino acid and the carboxyl group on another amino acid. As a result of peptide bond formation, it is possible to produce molecules varying in length from two amino acids to chains containing thousands of them.

Various terms are used to denote the nature of compounds containing peptide bonds. **Peptide*** usually refers to a molecule composed of short chains of amino acids, such as a dipeptide (two amino acids), a tripeptide (three), and a tetrapeptide (four) (figure 2.21). A **polypeptide** contains an unspecified number of amino acids, but usually has more than 20, and is often a smaller subunit of a protein. A protein is the largest of this class of compounds and usually contains a minimum of 50 amino acids. It is common for the terms *polypeptide* and *protein* to be used interchangeably, though not all polypeptides are large enough to be considered proteins. In chapter 9 we see that protein synthesis is not just a random connection of amino acids; it is directed by information provided in DNA.

***peptide** (pep′-tyd) Gr. *pepsis,* digestion.

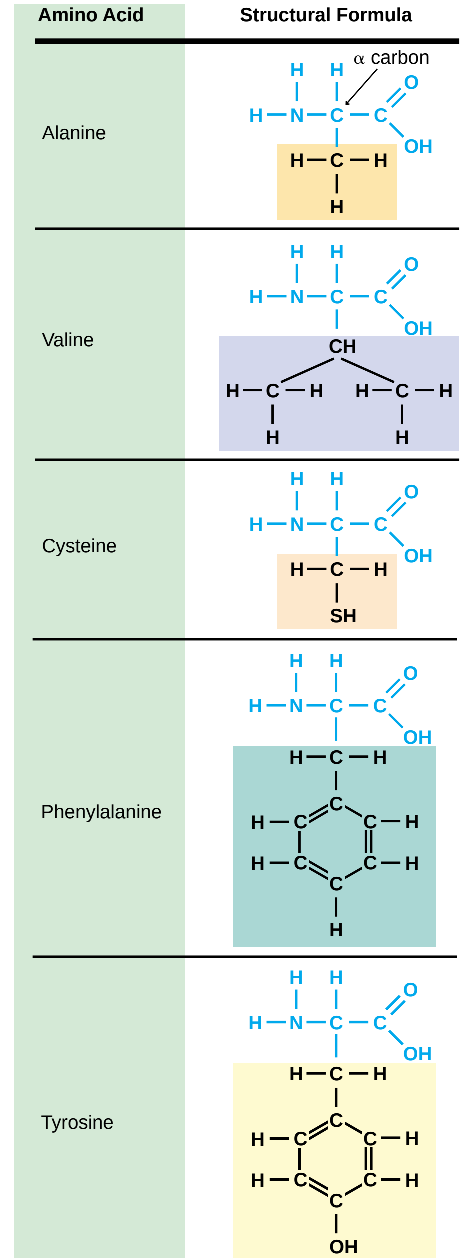


FIGURE 2.20

Structural formulas of selected amino acids. The basic structure common to all amino acids is shown in blue and the variable group, or R group, is placed in a colored box. Note the variations in structure of this reactive component.

Protein Structure and Diversity

The reason that proteins are so varied and specific is that they do not function in the form of a simple straight chain of amino acids (called the primary structure). A protein has a natural tendency to assume more complex levels of organization, called the secondary,

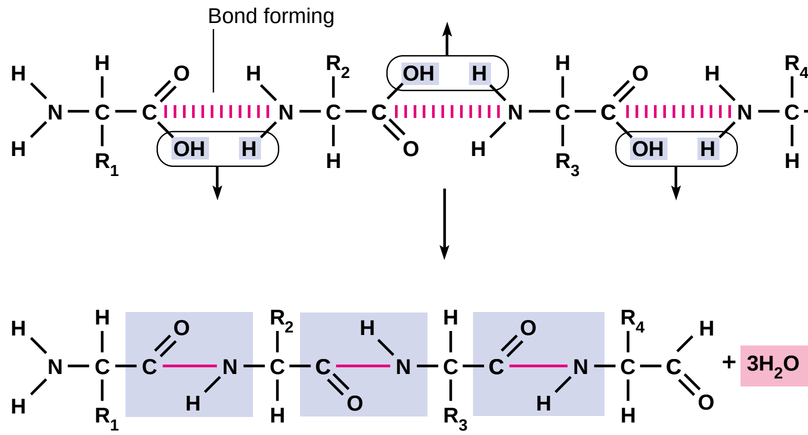


FIGURE 2.21

The formation of peptide bonds in a tetrapeptide.

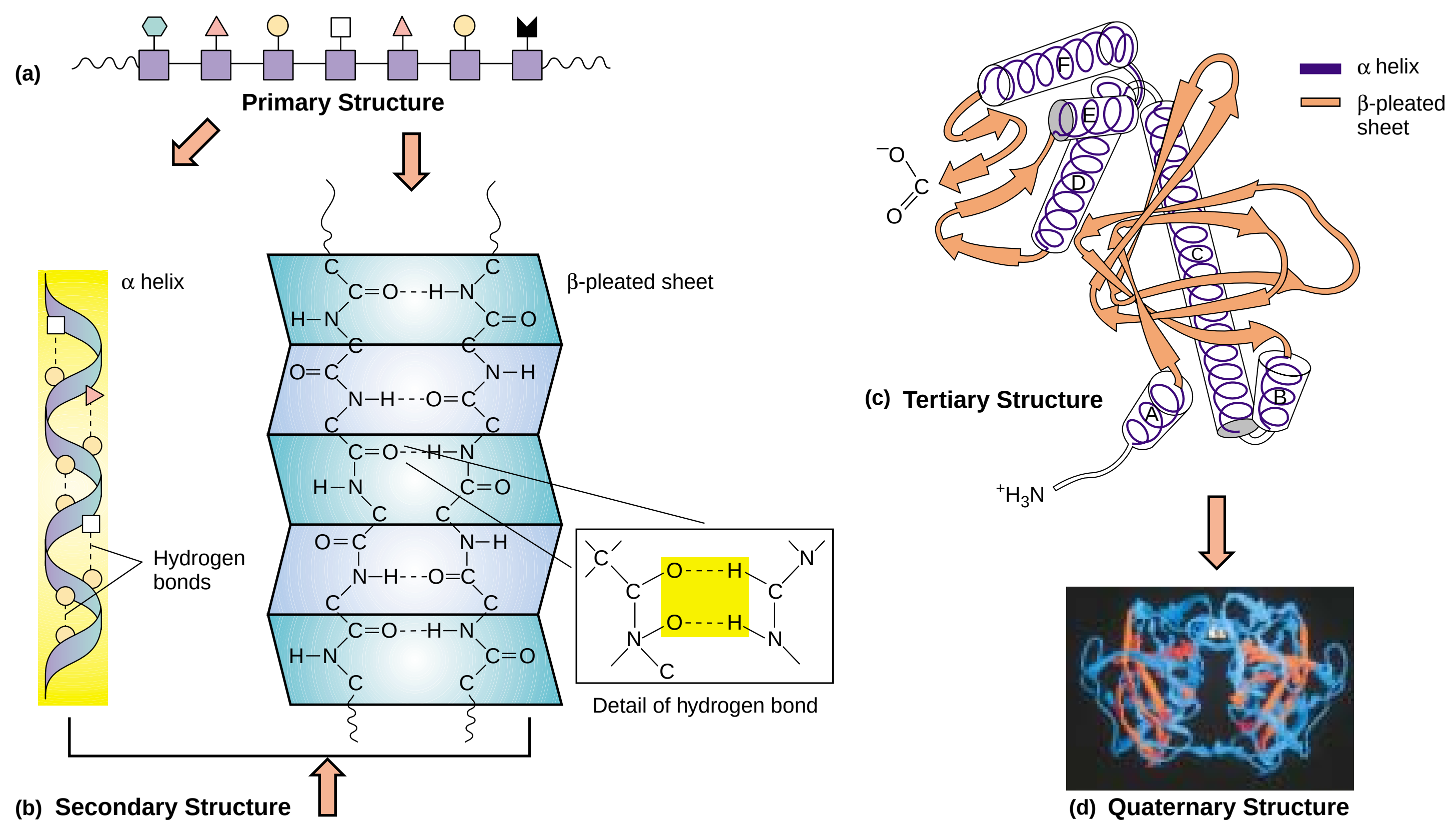


FIGURE 2.22

Stages in the formation of a functioning protein. (a) Its primary structure is a series of amino acids bound in a chain. **(b)** Its secondary structure develops when the chain forms hydrogen bonds that fold it into one of several configurations such as an alpha helix or beta-pleated sheet. Some proteins have several configurations in the same molecule. **(c)** A protein's tertiary structure is due to further folding of the molecule into a three-dimensional mass that is stabilized by hydrogen, ionic, and disulfide bonds between functional groups. The letters and arrows denote the order and direction of the folded chain. **(d)** The quaternary structure exists only in proteins that consist of more than one polypeptide chain. Shown here is a computer model of the nitrogenase iron protein, with the two polypeptide chains arranged symmetrically.

tertiary, and quaternary structures (figure 2.22). The **primary (1°) structure** is more correctly described as the type, number, and order of amino acids in the chain, which varies extensively from protein to protein. The **secondary (2°) structure** arises when various functional groups exposed on the outer surface of the molecule interact by forming hydrogen bonds. This interaction causes the amino acid chain to twist into a coiled configuration called the α *helix* or to fold into an accordion pattern called a β-*pleated sheet*. Some proteins contain both types of secondary configurations. Proteins at the secondary level undergo a third degree of torsion called the **tertiary (3°) structure** created by additional bonds between functional groups. In proteins with the sulfur-containing amino

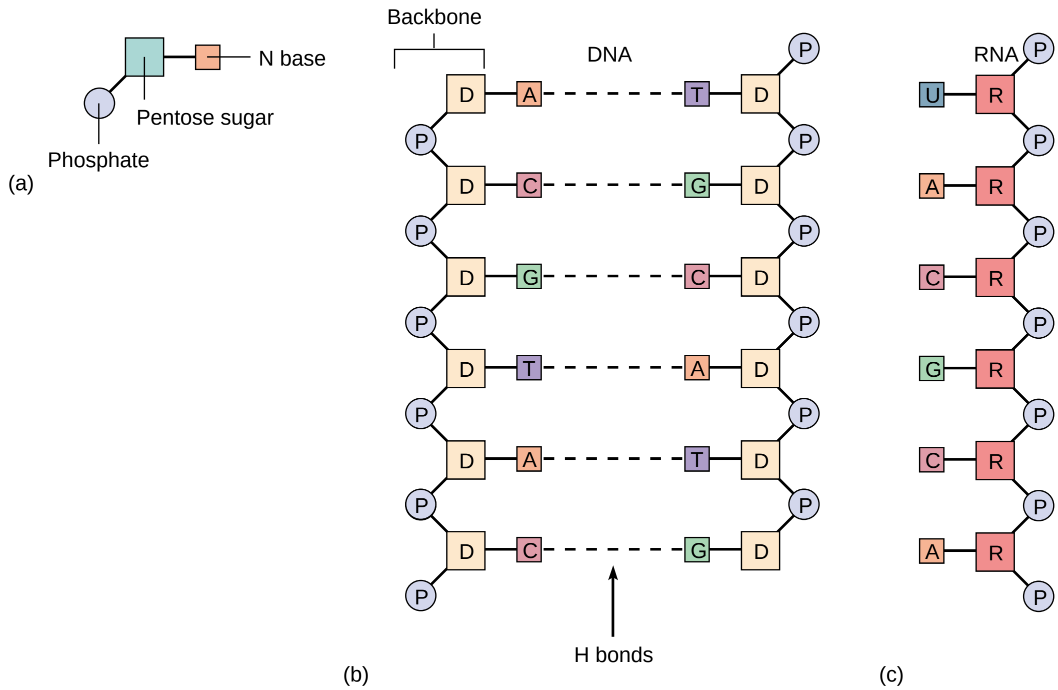


FIGURE 2.23

The general structure of nucleic acids. (a) A nucleotide, composed of a phosphate, a pentose sugar, and a nitrogen base, is the monomer of both DNA and RNA. **(b)** In DNA, the polymer is composed of alternating deoxyribose (D) and phosphate (P) with nitrogen bases (A, T, C, G) attached to the deoxyribose. Two of these polynucleotide strands are oriented so that the bases are paired across the central axis of the molecule. **(c)** In RNA, the polymer is composed of alternating ribose (R) and phosphate (P) attached to nitrogen bases (A, U, C, G), but it is only a single strand.

acid **cysteine,*** considerable tertiary stability is achieved through covalent **disulfide bonds** between sulfur atoms on two different parts of the molecule (figure 2.22*c*). Some complex proteins assume a **quaternary (4°) structure,** in which more than one polypeptide forms a large, multiunit protein. This is typical of antibodies (see chapter 15) and some enzymes that act in cell synthesis.

The most important outcome of intrachain[5] bonding and folding is that each different type of protein develops a unique shape, and its surface displays a distinctive pattern of pockets and bulges. As a result, a protein can react only with molecules that complement or fit its particular surface features like a lock and key. Such a degree of specificity can provide the functional diversity required for many thousands of different cellular activities. **Enzymes** serve as the catalysts for all chemical reactions in cells, and nearly every reaction requires a different enzyme (see chapter 8). **Antibodies** are complex glycoproteins with specific regions of attachment for bacteria, viruses, and other microorganisms; certain bacterial toxins (poisonous products) react with only one specific organ or tissue; and proteins embedded in the cell membrane have reactive sites restricted to a certain nutrient. Some proteins function as *receptors* to receive stimuli from the environment. The functional three-dimensional form of a protein is termed the *native state,* and if it is disrupted by some means, the protein is said to be *denatured*. Such agents as heat, acid, alcohol, and some disinfectants disrupt the stabilizing intrachain bonds and cause the molecule to become nonfunctional (see figure 11.4).

5. **Intra**chain means within the chain; **inter**chain would be between two chains.

*cysteine (sis′-tuh-yeen) Gr. *kystis,* sac. An amino acid first found in urine stones.

THE NUCLEIC ACIDS: A CELL COMPUTER AND ITS PROGRAMS

The nucleic acids, **deoxyribonucleic* acid (DNA)** and **ribonucleic* acid (RNA),** were originally isolated from the cell nucleus. Shortly thereafter, they were also found in other parts of nucleated cells, in cells with no nuclei (bacteria), and in viruses. The universal occurrence of nucleic acids in all known cells and viruses emphasizes their important roles as informational molecules. DNA, the master computer of cells, contains a special coded genetic program with detailed and specific instructions for each organism's heredity. It transfers the details of its program to RNA, operator molecules responsible for carrying out DNA's instructions and translating the DNA program into proteins that can perform life functions. For now, let us briefly consider the structure and some functions of DNA, RNA, and a close relative, adenosine triphosphate (ATP).

Both nucleic acids are polymers of repeating units called **nucleotides,*** each of which is composed of three smaller units: a *nitrogen base*, a *pentose* (5-carbon) sugar, and a *phosphate* (figure 2.23*a*). The nitrogen base is a cyclic compound that comes in two forms: *purines* (two rings) and *pyrimidines* (one ring). There are two types of purines—**adenine (A)** and **guanine (G)**—and three types of pyrimidines—**thymine (T)**, **cytosine (C)**, and **uracil (U)** (figure 2.24). A characteristic that differentiates DNA from RNA is that DNA contains all of the nitrogen bases except uracil, and RNA contains all of the nitrogen bases except thymine. The nitrogen base

*deoxyribonucleic (dee-ox″-ee-ry″-boh-noo-klay′-ik).

*ribonucleic (ry″-boh-noo-klay′-ik) It is easy to see why the abbreviations are used!

*nucleotide (noo′-klee-oh-tyd) From nucleus and acid.

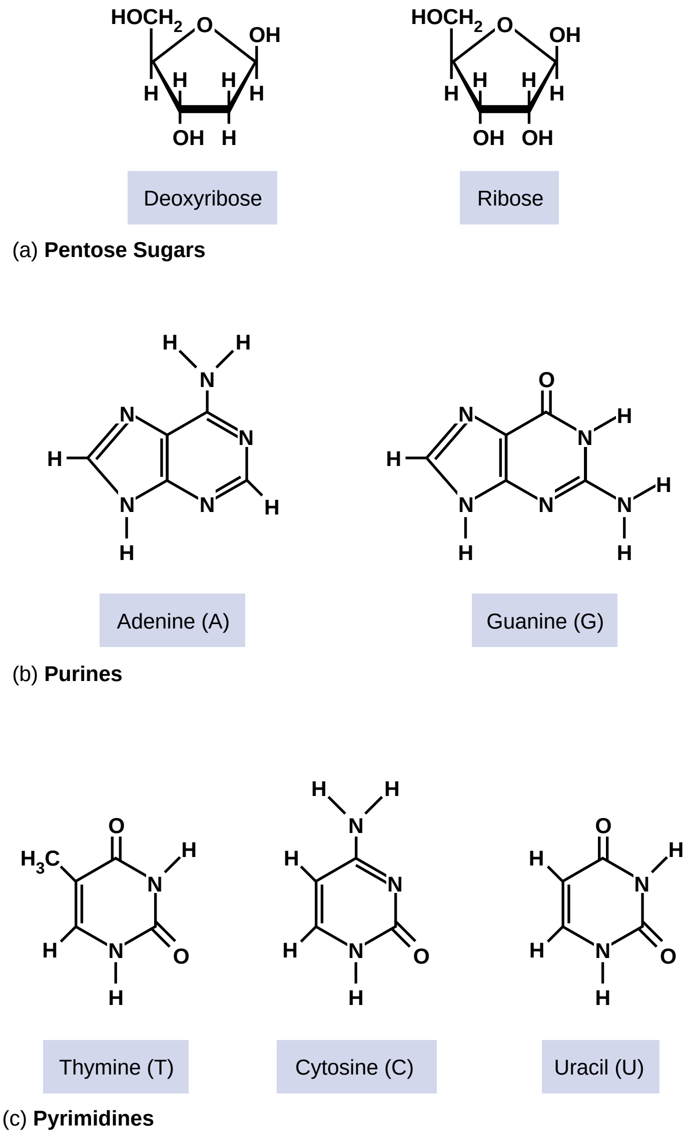


FIGURE 2.24

The sugars and nitrogen bases that make up DNA and RNA. **(a)** DNA contains deoxyribose, and RNA contains ribose. **(b)** A and G purines are found in both DNA and RNA. **(c)** C pyrimidine is found in both DNA and RNA, but T is found only in DNA, and U is found only in RNA.

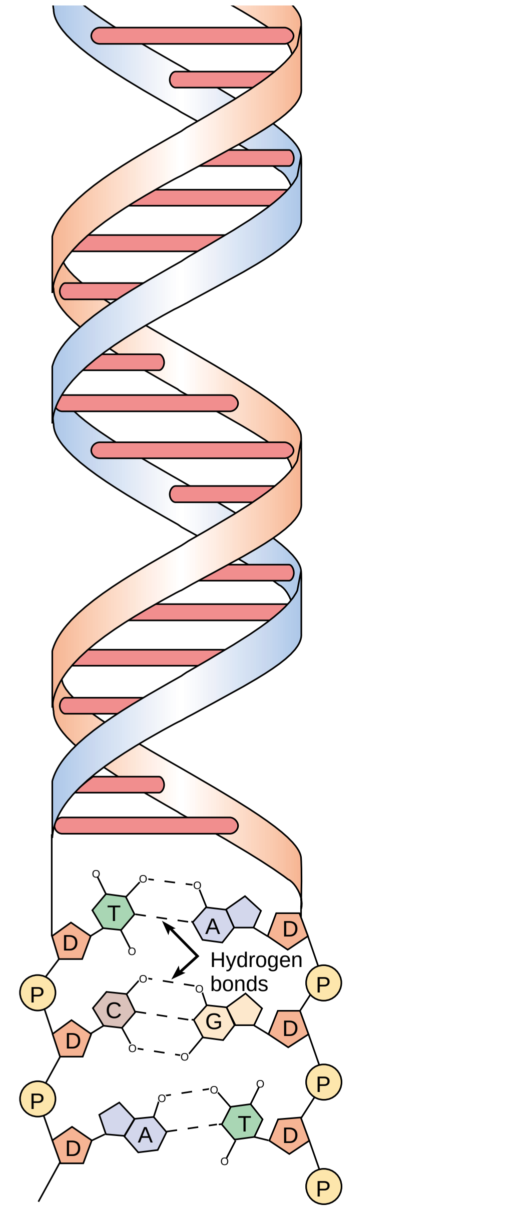


FIGURE 2.25

A structural representation of the double helix of DNA. Shown are the details of hydrogen bonds between the nitrogen bases of the two strands.

is covalently bonded to the sugar *ribose* in RNA and *deoxyribose* (because it has one less oxygen than ribose) in DNA. Phosphate (PO_4^{-3}), a derivative of phosphoric acid (H_3PO_4), provides the final covalent bridge that connects sugars in series. Thus, the backbone of a nucleic acid strand is a chain of alternating phosphate-sugar-phosphate-sugar molecules, and the nitrogen bases branch off the side of this backbone (see figure 2.23*b*,*c*).

The Double Helix of DNA

DNA is a huge molecule formed by two very long polynucleotide strands linked along their length by hydrogen bonds between complementary pairs of nitrogen bases. The pairing of the nitrogen bases occurs according to a predictable pattern: Adenine ordinarily pairs with thymine, and cytosine with guanine. The bases are attracted in this way because each pair shares oxygen, nitrogen, and hydrogen atoms exactly positioned to align perfectly for hydrogen bonds (figure 2.25).

For ease in understanding the structure of DNA, it is sometimes compared to a ladder, with the sugar-phosphate backbone representing the rails and the paired nitrogen bases representing the steps. Owing to the manner of nucleotide pairing and stacking of the bases, the actual configuration of DNA is a *double helix* that looks somewhat like a spiral staircase (see figures 2.25 and 9.4). As is true of protein, the structure of DNA is intimately related to its function. DNA molecules are usually extremely long, a feature that satisfies a requirement for storing genetic information in the sequence of base pairs the molecule contains. The hydrogen bonds between pairs can be disrupted when DNA is being copied, and the fixed complementary base pairing is essential to maintain the genetic code.

Making New DNA: Passing on the Genetic Message

The biological properties of cells and viruses are ultimately programmed by a master code composed of nucleic acids. This code is in the form of DNA in all cells and many viruses; other viruses are based on RNA alone. Regardless of the exact genetic program, both cells and viruses will continue to exist only if they can duplicate their genetic material and pass it on to subsequent generations. Figure 2.26 summarizes the main steps in this process and how it differs between cells and viruses.

During its division cycle, the cell has a mechanism for making a copy of its DNA by **replication,*** using the original strand as a pattern (figure 2.26*a*). Note that replication is guided by the double-stranded nature of DNA and the precise pairing of bases that create the master code. Replication requires the separation of the double strand into two single strands by an enzyme that helps to split the hydrogen bonds along the length of the molecule. This event exposes the base code and makes it available for copying. Free nucleotides are used to synthesize matching strands that complement the bases in the code by adhering to the pairing requirements of A–T and C–G. The end result is two separate double strands with the same order of bases as the original molecule. Viruses have a related mode of replication, but being genetic parasites, they must first enter a host cell and take over the cell's regular synthetic machinery and program it to make copies of the virus nucleic acid and, eventually, create whole viruses (figure 2.26*b*). The details of nucleic acid chemistry of cells and viruses will be covered in greater detail in chapter 9.

RNA: Organizers of Protein Synthesis

Like DNA, RNA consists of a long chain of nucleotides. However, RNA is a single strand containing ribose sugar instead of deoxyribose and uracil instead of thymine (see figure 2.23). Several functional types of RNA are formed using the DNA template through a replication-like process. The three major types of RNA are important for protein synthesis. Messenger RNA (mRNA) is a copy of a gene from DNA that provides the order and type of amino acids in a protein; transfer RNA (tRNA) is a carrier that delivers the correct amino acids for protein assembly; and ribosomal RNA (rRNA) is a major component of ribosomes (see figure 4.19). More information on these important processes is presented in chapter 9.

ATP: The Energy Molecule of Cells

A relative of RNA involved in an entirely different cell activity is **adenosine triphosphate (ATP).** ATP is a nucleotide containing adenine, ribose, and three phosphates rather than just one (figure 2.27). It belongs to a category of high-energy compounds (also including guanosine triphosphate, GTP) that give off energy when the bond is broken between the second and third (outermost) phosphate. The presence of these high-energy bonds makes it possible for ATP to release and store energy for cellular chemical reactions. Breakage of the bond of the terminal phosphate releases energy to do cellular work and also generates adenosine diphosphate (ADP). ADP can be converted back to ATP when the third phosphate is restored, thereby serving as an energy depot. Carriers for oxidation-reduction activities (nicotinamide adenine dinucleotide [NAD], for instance) are also derivatives of nucleotides (see chapter 8).

*__replication__ (reh″-plih-kay′-shun) A process of making an exact copy of something.

CHAPTER CHECKPOINTS

- Macromolecules are very large organic molecules (polymers) built up by polymerization of smaller molecular subunits (monomers).
- Carbohydrates are biological molecules whose polymers are monomers linked together by glycosidic bonds. Their main functions are protection and support (in organisms with cell walls) and also nutrient and energy stores.
- Lipids are biological molecules such as fats that are insoluble in water and contain special ester linkages. Their main functions are cell components, cell secretions, and nutrient and energy stores.
- Proteins are biological molecules whose polymers are chains of amino acid monomers linked together by peptide bonds.
- Proteins are called the "shapers of life" because of the many biological roles they play in cell structure and cell metabolism.
- Protein shape determines protein function. Shape is dictated by amino acid composition and by the pH and temperature of the protein's immediate environment.
- Nucleic acids are biological molecules whose polymers are chains of nucleotide monomers linked together by phosphate–pentose sugar covalent bonds. Double-stranded nucleic acids are linked together by hydrogen bonds. Nucleic acids are information molecules that direct cell metabolism and reproduction. Nucleotides such as ATP also serve as energy transfer molecules in cells.

Cells: Where Chemicals Come to Life

As we proceed in this chemical survey from the level of simple molecules to increasingly complex levels of macromolecules, at some point we cross a line from the realm of lifeless molecules and arrive at the fundamental unit of life called a **cell.**[6] A cell is indeed a huge aggregate of carbon, hydrogen, oxygen, nitrogen, and many other atoms, and it follows the basic laws of chemistry and physics, but it is much more. The combination of these atoms produces characteristics, reactions, and products that can only be described as **living.**

FUNDAMENTAL CHARACTERISTICS OF CELLS

The bodies of living things such as bacteria and protozoa consist of only a single cell, whereas those of animals and plants contain trillions of cells. Regardless of the organism, all cells have a few common characteristics. They tend to be spherical, polygonal, cubical, or cylindrical, and their protoplasm (internal cell contents) is encased in a cell or cytoplasmic membrane (see Microbits 2.3). They have chromosomes containing DNA and ribosomes for protein synthesis, and they are exceedingly complex in function. Aside from

6. The word *cell* was originally coined from an Old English term meaning "small room" because of the way plant cells looked to early microscopists.

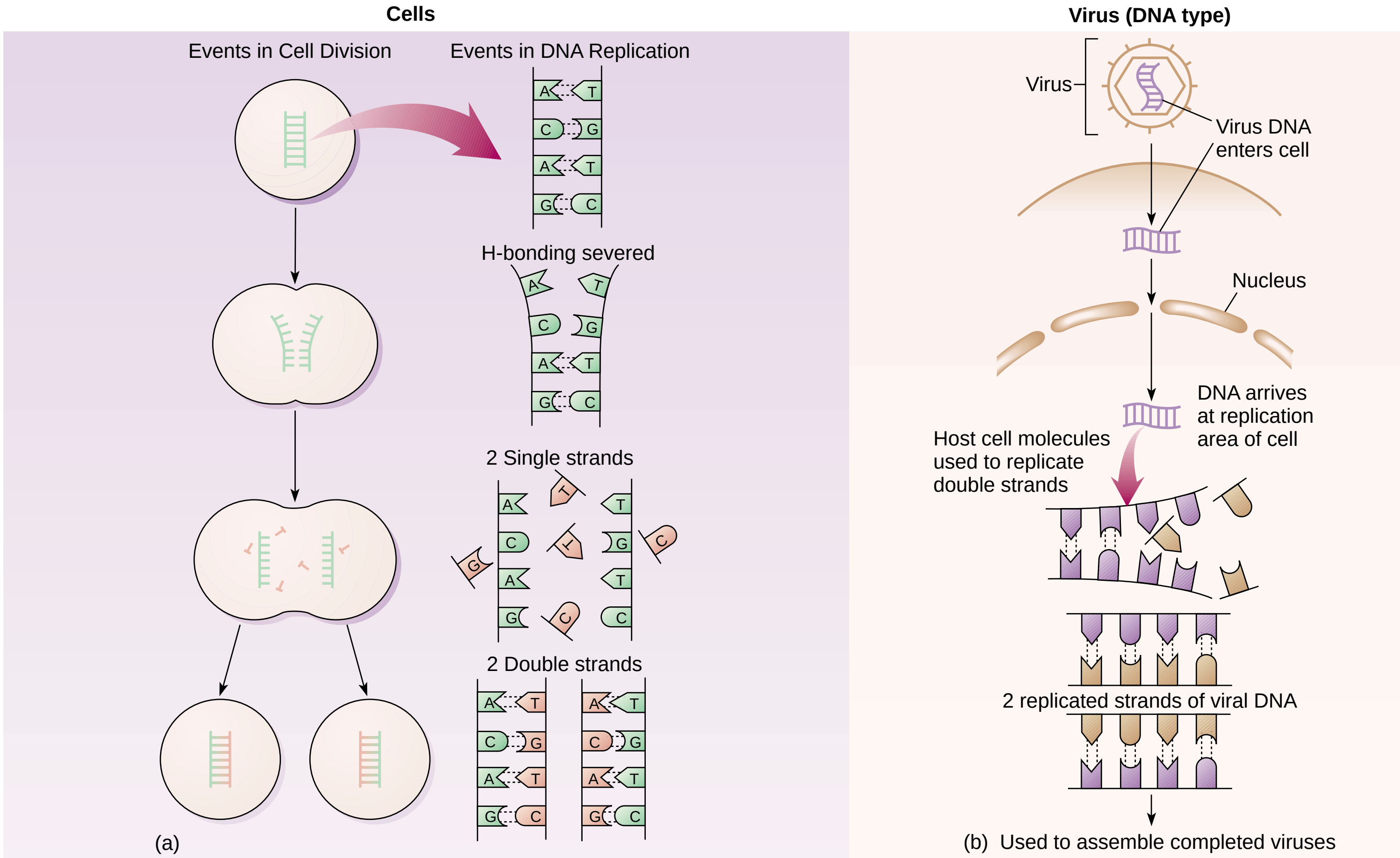


FIGURE 2.26

Simplified view of DNA replication in cells and viruses. (a) The DNA in the cell's chromosome must be duplicated as the cell is dividing. This duplication is accomplished through the separation of the double DNA strand into two single strands. New strands are then synthesized using the original strands as guides to assemble the correct complementary new bases. **(b)** In DNA viruses, the DNA is inserted into its host cell and enters the replication area. Here it is separated into two strands and copied by the machinery of the host cell. Thousands of copies of viral nucleic acid will become part of the completed viruses later released by the infected host cell.

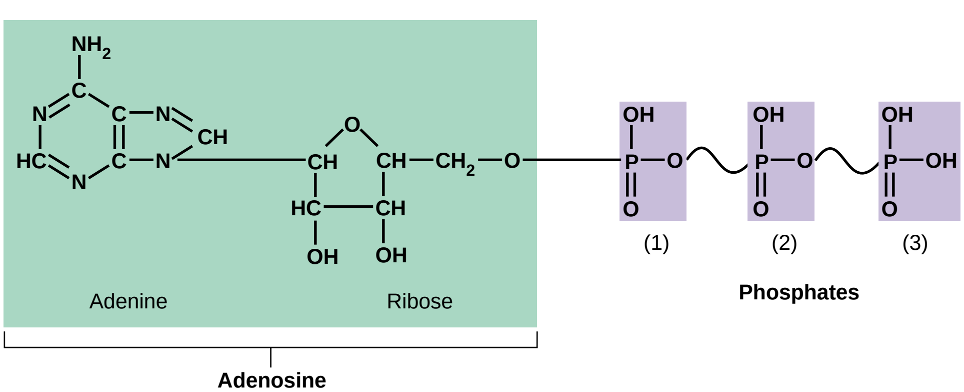


FIGURE 2.27

The structural formula of an ATP molecule, the chemical form of energy transfer in cells. The wavy lines that connect the phosphates represent bonds that release large amounts of energy when broken.

these few similarities, most cell types fall into one of two fundamentally different lines (discussed in chapter 1): the small, seemingly simple procaryotic cells and the larger, structurally more complicated eucaryotic cells.

Eucaryotic cells are found in animals, plants, fungi, and protists. They contain a number of complex internal parts called organelles that perform useful functions for the cell involving growth, nutrition, or metabolism. By convention, organelles are defined as cell components that perform specific functions and are enclosed by membranes. Organelles also partition the eucaryotic cell into smaller compartments. The most visible organelle is the nucleus, a roughly ball-shaped mass surrounded by a double membrane that contains the DNA of the cell. Other organelles include the Golgi apparatus, endoplasmic reticulum, vacuoles, and mitochondria (table 2.6).

TABLE 2.6

A General Comparison of Procaryotic and Eucaryotic Cells and Viruses*

Function or Structure	Characteristic	Procaryotic Cells	Eucaryotic Cells	Viruses**
Genetics	Nucleic acids	+	+	+
	Chromosomes	+	+	−
	True nucleus	−	+	−
	Nuclear envelope	−	+	−
Reproduction	Mitosis	−	+	−
	Production of sex cells	+/−	+	−
	Binary fission	+	+	−
Biosynthesis	Independent	+	+	−
	Golgi apparatus	−	+	−
	Endoplasmic reticulum	−	+	−
	Ribosomes	+***	+	−
Respiration	Enzymes	+	+	−
	Mitochondria	−	+	−
Photosynthesis	Pigments	+/−	+/−	−
	Chloroplasts	−	+/−	−
Motility/locomotor structures	Flagella	+/−***	+/−	−
	Cilia	−	+/−	−
Shape/protection	Cell wall	+***	+/−	−
	Capsule	+/−	+/−	−
	Spores	+/−	+/−	−
Complexity of function		+	+	+/−
Size (in general)		0.5–3 μm	2–100 μm	< 0.2 μm

**+ means most members of the group exhibit this characteristic; – means most lack it; +/– means only some members have it.*
*** Viruses cannot participate in metabolic or genetic activity outside their host cells.*
**** The procaryotic type is functionally similar to the eucaryotic, but structurally unique.*

Procaryotic cells are possessed only by the bacteria and archaea. Sometimes it may seem that procaryotes are the microbial "have-nots" because, for the sake of comparison, they are described by what they lack. They have no nucleus or other organelles. This apparent simplicity is misleading, because the fine structure of procaryotes is complex. Overall, procaryotic cells can engage in nearly every activity that eucaryotic cells can, and many can function in ways that eucaryotes cannot.

PROCESSES THAT DEFINE LIFE

To lay the groundwork for a detailed coverage of cells in chapters 4 and 5, this section provides an overview of cell structure and function and introduces the primary characteristics of life. The biological activities or properties that help define and characterize cells as living entities are: (1) growth, (2) reproduction and heredity, (3) metabolism, including cell synthesis and the release of energy, (4) movement and/or irritability, (5) the capacity to transport substances into and out of the cell, and (6) cell support, protection, and storage mechanisms. Although eucaryotic cells have specific organelles to perform these functions, procaryotic cells must rely on a few simple, multipurpose cell components. As indicated in chapter 1, viruses are not cells, are not generally considered living things, and show certain signs of life only when they invade a host cell. Table 2.6 indicates their relative simplicity compared with cells.

Reproduction: Bearing Offspring

A cell's **genome,*** its complete set of genetic material, is composed of elongate strands of DNA. The DNA is packed into discrete bodies called **chromosomes,*** In eucaryotic cells, the chromosomes are located entirely within a nuclear membrane. Procaryotic DNA occurs in a special type of circular chromosome that is not enclosed by a membrane of any sort.

Living things devote a portion of their life cycle to producing offspring that will carry on their particular genetic line for many generations. In **sexual reproduction,** offspring are produced through the union of sex cells from two parents. In **asexual**[7] **reproduction,** offspring originate through the division of a single parent cell into two daughter cells. Sexual reproduction occurs in most eucaryotes, and eucaryotic cells also reproduce asexually by several processes. One is a type of cell division called *binary fission,* a simple process in which the cell splits equally in two. Many eucaryotic cells engage in **mitosis,*** an orderly division of chromosomes that

7. *Asexual* refers to the absence of sexual union. (The prefix *a* or *an* means not or without.)

***genome** (jee′-nohm) A combination of the words *gene* and *chromosome*. Refers to an organism′s entire set of hereditary factors.

***chromosome** (kro′-moh-sohm) Gr. *chroma*, colored, and *soma*, body. The name is derived from the fact that chromosomes stain readily with dyes.

***mitosis** (my-toh′-sis) Gr. *mitos*, thread, and *osis*, a condition. Often the term *mitosis* is used synonymously with eucaryotic cell division.

usually accompanies cell division (see figure 5.6). In contrast, procaryotic cells reproduce primarily by binary fission. They have no mitotic apparatus, nor do they reproduce by typical sexual means.

Metabolism: Chemical and Physical Life Processes

Cells synthesize proteins using hundreds of tiny particles called **ribosomes.** In eucaryotes, ribosomes are dispersed throughout the cell or inserted into membranous sacs known as the **endoplasmic reticulum** (see figure 5.8). Procaryotes have smaller ribosomes scattered throughout the protoplasm, since they lack an endoplasmic reticulum. Eucaryotes generate energy by chemical reactions in the **mitochondria,** whereas procaryotes use their cell membrane for this purpose. Photosynthetic microorganisms (algae and some bacteria) trap solar energy by means of pigments and convert it to chemical energy in the cell. Algae (eucaryotes) have compact, membranous bundles called **chloroplasts,** which contain the pigment and perform the photosynthetic reactions. Photosynthetic reactions and pigments of procaryotes do not occur in chloroplasts, but in specialized areas of the cell membrane.

Irritability or Motility

All cells have the capacity to respond to chemical, mechanical, or light stimuli. This quality, called **irritability,** helps cells adapt to the environment and obtain nutrients. Although not present in all cells, true **motility,** or self-propulsion, is a notable sign of life. Eucaryotic cells move by one of the following locomotor organelles: cilia, which are short, hairlike appendages; flagella, which are longer, whiplike appendages; or pseudopods, fingerlike extensions of the cell membrane. Motile procaryotes move by means of unusual, propeller-like flagella unique to bacteria or by special fibrils that produce a gliding form of motility. They have no cilia or pseudopods.

Protection and Storage

Many cells are supported and protected by rigid **cell walls,** which prevent them from rupturing while also providing support and shape. Among eucaryotes, cell walls occur in plants, microscopic algae, and fungi, but not in animals or protozoa. The majority of procaryotes have cell walls, but they differ in composition from the eucaryotic varieties. As protection against depleted nutrient sources, many microbes store nutrients intracellularly. Eucaryotes store nutrients in membranous sacs called vacuoles, and procaryotes concentrate them in crystals called granules or inclusions.

Transport: Movement of Nutrients and Wastes

Cell survival depends on drawing nutrients from the external environment and expelling waste and other metabolic products from the internal environment. This two-directional transport is accomplished in both eucaryotes and procaryotes by the cell membrane. This membrane, described in Microbits 2.3, has a very similar structure in both eucaryotic and procaryotic cells. Eucaryotes have an additional organelle, the **Golgi apparatus,** that assists in sorting and packaging molecules for transport and removal from the cell.

Table 2.6 summarizes the differences and similarities among procaryotes, eucaryotes, and viruses. You will probably want to review this table after you have completed chapters 4 and 5.

CHAPTER CHECKPOINTS

As the atom is the fundamental unit of matter, so is the cell the fundamental unit of life.

All true cells contain biological molecules that carry out the processes that define life: metabolism and reproduction; these two basic processes are supported by the functions of irritability and motility, protection, storage, and transport.

The cell membrane is of critical importance to all cells because it controls the interchange between the cell and its environment.

CHAPTER CAPSULE WITH KEY TERMS

I. **Basic Properties of Atoms and Elements**
   A. Atomic Structure
      1. All **matter** in the universe is composed of minute particles called **atoms**—the simplest form of matter not divisible into a simpler substance by chemical means. Atoms are composed of smaller particles called **protons** (p^+), positive in charge; **neutrons** (n^0), uncharged; and **electrons** (e^-), negative in charge. Protons and neutrons comprise the **nucleus** of an atom, and electrons move about the nucleus in **orbitals** arranged in **shells.**
      2. Atoms that differ in numbers of protons, neutrons, and electrons are **elements.** Elements can be described by **mass number** (the total number of protons and neutrons in the nucleus) and **atomic number** (the number of protons alone), and each is known by a distinct name and symbol. An unreacted atom is neutral because electron charge cancels out proton charge. Elements may exist in variant forms called **isotopes,** which differ in the number of neutrons.
      3. Electrons fill the orbitals in pairs: The first orbital (shell) holds 2 e^-, the second shell can hold 8 e^- (4 pairs), and the third shell can hold 18 e^- (9 pairs). The electron number of each element dictates orbital filling; the outermost shell becomes the focus of reactivity and bonding. In general, atoms with a filled outermost shell are less reactive than atoms with unfilled ones.
   B. Chemical Bonds and Molecules
      1. Atoms interact to form **chemical bonds** and **molecules.** Member atoms of molecules may be the same element or different elements; if the member elements are different, the substance is a **compound.**
      2. The type of bond is dictated by the electron makeup (**valence**) of the outer orbitals of the atoms.
      3. With **covalent bonds,** the electrons are shared and orbit within the entire molecule. This molecule can be **polar,** due to an imbalance of charge, or **nonpolar.**
      4. With **ionic bonds,** an atom with a low number of valence electrons loses them to an atom that has a nearly filled outer orbital. When ionic bonds are broken, the atoms **ionize** into

charged particles called **ions.** Ions that have lost electrons and have a positive charge are **cations** (Na^+), and those that have gained electrons and have a negative charge are **anions** (Cl^-). Ions of the opposite charge are attracted to each other, and those with the same charges repel each other.

5. **Hydrogen bonds** are caused by weak attractive forces between covalently bonded hydrogen and negatively charged oxygen or nitrogen on the same or nearby molecules.

C. Redox Reactions

Chemicals may participate in a transfer of electrons, called an **oxidation-reduction** (redox) reaction, between pairs of atoms or molecules. **Oxidation** is a reaction in which electrons are released, and **reduction** is a reaction in which these same electrons are received. Any atom or molecule that donates electrons to another atom or molecule is a **reducing agent,** and one that picks up electrons is an **oxidizing agent.**

D. Chemical Formulas, Models, and Equations

1. **Chemical formulas** may be presented as simple molecular summaries of the atoms or expressed as structural formulas that give the details of bonding.
2. **Chemical equations** summarize a chemical reaction by showing the reactants (starting chemicals) and the products (end results), and can indicate **synthesis** reactions, **decomposition** reactions, **exchange** reactions, and reversible reactions.

E. Solutions, Acids, Bases, and pH

1. A **solution** is a combination of a solid, liquid, or gaseous chemical termed a **solute** dissolved in a liquid medium called a **solvent.** Water-soluble solutes are either charged or polar. The dissolved solute will become **hydrated** due to electrostatic attraction. Chemicals that attract water are **hydrophilic;** those that repel it are **hydrophobic.** The **concentration** of a solution is defined as the amount of solute dissolved in a given amount of solvent.
2. Acidity and basicity of solutions are represented by the **pH scale,** based on a standardized range of **hydrogen ion concentration** $[H^+]$; an **acid** is a solution that contains a concentration of $[H^+]$ greater than 0.0000001 moles/liter (10^{-7}), and a **base** is a solution that contains a concentration of $[H^+]$ below that amount. When a solution contains exactly that $[H^+]$, it is considered neutral (pH 7). The working pH scale ranges from the most acidic reading of 0 to the most basic (alkaline) reading of 14. Acidic and basic ions in aqueous solutions may interact to form water and a salt through **neutralization.**

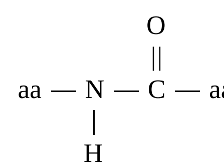
II. **Organic Chemistry**

A. **Organic compounds** contain some combination of carbon and hydrogen covalently bonded. Carbon can also bond with other carbons and with oxygen, nitrogen, and phosphorus to create the diversity of biological molecules. Organic compounds constitute the most important molecules in the structure and function of cells. **Inorganic compounds** are composed of some combination of atoms other than carbon and hydrogen.

B. **Functional groups** are special accessory molecules that bind to carbon and provide the diversity and reactivity seen in organic compounds.

C. Biochemistry and Macromolecules

1. Organic compounds that function in living things are **biochemicals.** Many of them are **macromolecules** (very large compounds). Levels of structure are **monomers** (single units) and **polymers** (long chains of single units) assembled by **polymerization.**
2. **Carbohydrates** are composed of carbon, hydrogen, and oxygen (CH_2O) and contain aldehyde or ketone groups.
3. **Monosaccharides (glucose)** are the simplest sugars. When two sugars are joined by the −OH group and carbon through **dehydration synthesis,** a **glycosidic bond** occurs.
4. **Disaccharides** are composed of two monosaccharides **(lactose, sucrose). Polysaccharides** are chains of five or more monosaccharides; examples are **cellulose, peptidoglycan, starch,** and **glycogen.** Di- and polysaccharides are digested by specific enzymes that break the bond through **hydrolysis.**
5. **Lipids** are not soluble in water and other polar solvents due to their nonpolar, hydrophobic chains. **Triglycerides,** including fats and oils, consist of a **glycerol** molecule bonded to 3 **fatty acid** molecules. They are important storage lipids. **Phospholipids** are composed of a glycerol bound to two fatty acids and a phosphoric acid–alcohol group; the molecule has a hydrophilic head and long hydrophobic fatty acid chains; they form single or double lipid layers in the presence of water and are important constituents of cell membranes.
6. **Proteins** are highly complex macromolecules assembled from 20 different subunits called **amino acids (aa).** Amino acids are combined in a certain order by **peptide bonds:**



   a. A **peptide** is a short chain of aa's; a dipeptide has two, and a tripeptide has three; a **polypeptide** is usually 20–50 aa's; a **protein** contains more than 50 aa's. Larger proteins predominate in cells.
   b. The chain of amino acids is a protein's **primary structure.** Interactions between functional groups result in additional levels of structure. Hydrogen bonds within the chain fold it first into a helix or sheet called the **secondary structure.** This folds again, forging stronger **disulfide bonds** on nearby cysteines and producing a three-dimensional **tertiary structure.** Proteins composed of two or more polypeptides exist in a **quaternary** state. Variations in shape provide specificity and give rise to the diversity in enzymes, antibodies, and structural proteins.

7. **Nucleic Acids**
   a. **Deoxyribonucleic acid (DNA)** and **ribonucleic acid (RNA)** are very complex molecules that carry, express, and pass on the genetic information of all cells and viruses. Their basic building block is a nucleotide, composed of a **nitrogen base,** a **pentose sugar,** and a **phosphate.** Nitrogen bases are the ringed compounds: **adenine (A), cytosine (C), thymine (T), guanine (G), and uracil (U);** pentose sugars may be deoxyribose or ribose. The basic design is a polynucleotide, with the sugars linking up in an alternating series with phosphates to make a backbone, and the bases branching off the sugars.
   b. DNA contains **deoxyribose** sugar, has all of the bases except uracil, and occurs as a double-stranded helix with the bases hydrogen-bonded in pairs between the helices; the pairs mate according to the pattern A–T and C–G. DNA is the master code for a cell's life processes and must

be transmitted to offspring. During cell division, it is **replicated** by separation of the double strand into two single strands, which are used as a template to form two new double strands. RNA contains **ribose** sugar, has all of the bases except thymine, and is a single-stranded molecule. It expresses the DNA code into proteins.

8. **Adenosine triphosphate (ATP)** is a nucleotide involved in the transfer and storage of energy in cells. It contains adenine, ribose, and three phosphates in a series. Splitting off the last phosphate in the triphosphate releases a packet of energy that may be used to do cell work.

III. **Introduction to Cell Structure**

A. **Cells** are huge aggregates of macromolecules organized to carry out complex processes described as **living.** All organisms consist of cells, which fall into one of two types: **procaryotic,** which are small, structurally simple cells that lack a **nucleus** and other **organelles;** and **eucaryotic,** larger cells with a nucleus and organelles, found in plants, animals, fungi, and protozoa. **Viruses** are not cells and are not generally considered living because they cannot function independently.

B. Organisms demonstrate several essential qualities of life.

1. **Growth** and **reproduction** involves producing offspring asexually (with one parent) or sexually (with two parents).
2. **Metabolism** refers to the chemical reactions in cells, including the synthesis of proteins on ribosomes and the release of energy (ATP).
3. **Motility** originates from special locomotor structures such as flagella and cilia; **irritability** is responsiveness to external stimuli.
4. **Protective** external structures include capsules and cell walls; nutrient **storage** takes place in compact intracellular masses.
5. **Transport** involves conducting nutrients into the cell and wastes out of the cell.

C. **Membranes** surround the cytoplasm and may occur internally in organelles. The cell membrane is a continuous ultrathin bilayer of lipids studded with proteins that controls cell **permeability** and transport. Proteins also serve as sites of recognition and cell communication.

MULTIPLE-CHOICE QUESTIONS

1. The smallest unit of matter with unique characteristics is
   a. an electron c. an atom
   b. a molecule d. a proton
2. The ____ charge of a proton is exactly balanced by the ____ charge of a (an) ____ .
   a. negative, positive, electron
   b. positive, neutral, neutron
   c. positive, negative, electron
   d. neutral, negative, electron
3. Electrons move around the nucleus of an atom in pathways called
   a. shells c. circles
   b. orbitals d. rings
4. Which part of an element does not vary in number?
   a. electron c. proton
   b. neutron d. all of these vary
5. The number of electrons of a neutral atom is automatically known if one knows the
   a. atomic number
   b. atomic weight
   c. number of orbitals
   d. valence
6. Elements
   a. exist in 92 natural forms
   b. have distinctively different atomic structures
   c. vary in atomic weight
   d. a and b are correct
   e. a, b, and c are correct
7. If a substance contains two or more elements of different types, it is considered
   a. a compound c. a molecule
   b. a monomer d. organic
8. Bonds in which atoms share electrons are defined as ____ bonds.
   a. hydrogen c. double
   b. ionic d. covalent
9. What kind of bond would you expect potassium to form with chlorine?
   a. ionic c. polar
   b. covalent d. nonpolar
10. When a compound carries a positive charge on one end and a negative charge on the other end, it is said to be
    a. ionized c. polar
    b. hydrophilic d. oxidized
11. Hydrogen bonds can form between ____ adjacent to each other.
    a. two hydrogen atoms
    b. two oxygen atoms
    c. a hydrogen atom and an oxygen atom
    d. negative charges
12. Ions with the same charge will be ____ each other, and ions with opposite charges will be ____ each other.
    a. repelled by, attracted to
    b. attracted to, repelled by
    c. hydrated by, dissolved by
    d. dissolved by, hydrated by
13. An atom that can donate electrons during a reaction is called
    a. an oxidizing agent
    b. a reducing agent
    c. an ionic agent
    d. an electrolyte
14. In a solution of NaCl and water, NaCl is the ____ and water is the ____ .
    a. acid, base c. solute, solvent
    b. base, acid d. solvent, solute
15. A substance that releases H^+ into a solution
    a. is a base c. is an acid
    b. is ionized d. has a high pH
16. A solution with a pH of 2 ____ than a solution with a pH of 8.
    a. has less H^+
    b. has more H^+
    c. has more OH^-
    d. is less concentrated

17. Fructose is a type of
    a. disaccharide
    b. monosaccharide
    c. polysaccharide
    d. amino acid
18. Bond formation in polysaccharides and polypeptides is accompanied by the removal of a
    a. hydrogen atom
    b. hydroxyl ion
    c. carbon atom
    d. water molecule
19. The monomer unit of polysaccharides such as starch and cellulose is
    a. fructose c. ribose
    b. glucose d. lactose
20. A phospholipid contains
    a. three fatty acids bound to glycerol
    b. three fatty acids, a glycerol, and a phosphate
    c. two fatty acids and a phosphate bound to glycerol
    d. three cholesterol molecules bound to glycerol
21. Proteins are synthesized by linking amino acids with ____ bonds.
    a. disulfide c. peptide
    b. glycosidic d. ester
22. The amino acid that accounts for disulfide bonds in the tertiary structure of proteins is
    a. tyrosine c. cysteine
    b. glycine d. serine
23. DNA is a hereditary molecule that is composed of
    a. deoxyribose, phosphate, and nitrogen bases
    b. deoxyribose, a pentose, and nucleic acids
    c. sugar, proteins, and thymine
    d. adenine, phosphate, and ribose
24. What is meant by DNA replication?
    a. duplication of the sugar-phosphate backbone
    b. matching of base pairs
    c. formation of the double helix
    d. the exact copying of the DNA code into two new molecules

CONCEPT QUESTIONS

1. How are the concepts of an atom and an element related? What causes elements to differ?
2. a. How are mass number and atomic number derived? What is the atomic weight?
   b. Using data in table 2.1, give the electron number of nitrogen, sulfur, calcium, phosphorus, and iron.
   c. What is distinctive about isotopes of elements, and why are they important?
3. a. How is the concept of molecules and compounds related?
   b. Compute the molecular weight of oxygen and methane.
4. a. Why is an isolated atom neutral?
   b. Describe the concept of the atomic nucleus, electron orbitals, and shells.
   c. What causes atoms to form chemical bonds?
   d. Why do some elements not bond readily?
   e. Draw the atomic structure of magnesium and predict what kinds of bonds it will make.
5. Distinguish between the general reactions in covalent, ionic, and hydrogen bonds.
6. a. Which kinds of elements tend to make covalent bonds?
   b. Distinguish between a single and a double bond.
   c. What is polarity?
   d. Why are some covalent molecules polar and others nonpolar?
   e. What is an important consequence of the polarity of water?
7. a. Which kinds of elements tend to make ionic bonds?
   b. Exactly what causes the charges to form on atoms in ionic bonds?
   c. Verify the proton and electron numbers for Na^+ and Cl^-.
   d. Differentiate between an anion and a cation.
   e. What kind of ion would you expect magnesium to make, on the basis of its valence?
8. Differentiate between an oxidizing agent and a reducing agent.
9. Why are hydrogen bonds relatively weak?
10. a. Compare the three basic types of chemical formulas.
    b. Review the types of chemical reactions and the general ways they can be expressed in equations.
11. a. Define solution, solvent, and solute.
    b. What properties of water make it an effective biological solvent, and how does a molecule like NaCl become dissolved in it?
    c. How is the concentration of a solution determined?
    d. What is molarity? Tell how to make a 1M solution of $Mg_3(PO_4)_2$ and a 0.1 M solution of $CaSO_4$.
12. a. What determines whether a substance is an acid or a base?
    b. Briefly outline the pH scale.
    c. How can a neutral salt be formed from acids and bases?
13. a. What atoms must be present in a molecule for it to be considered organic?
    b. What characteristics of carbon make it ideal for the formation of organic compounds?
    c. What are functional groups?
    d. Differentiate between a monomer and a polymer.
    e. How are polymers formed?
    f. Name several inorganic compounds.
14. a. What characterizes the carbohydrates?
    b. Differentiate between mono-, di-, and polysaccharides, and give examples of each.
    c. What is a glycosidic bond, and what is dehydration synthesis?
    d. What are some of the functions of polysaccharides in cells?
15. a. Draw simple structural molecules of triglycerides and phospholipids to compare their differences and similarities.
    b. What is an ester bond?
    c. How are saturated and unsaturated fatty acids different?
    d. What characteristic of phospholipids makes them essential components of cell membranes?
    e. Why is the hydrophilic end of phospholipids attracted to water?
16. a. Describe the basic structure of an amino acid.
    b. What makes the amino acids distinctive, and how many of them are there?
    c. What is a peptide bond?
    d. Differentiate between a peptide, a polypeptide, and a protein.
    e. Explain what causes the various levels of structure of a protein molecule.
    f. What functions do proteins perform in a cell?

17. a. Describe a nucleotide and a polynucleotide, and compare and contrast the general structure of DNA and RNA.
    b. Name the two purines and the three pyrimidines.
    c. Why is DNA called a double helix?
    d. What is the function of RNA?
    e. What is ATP, and what is its function in cells?

18. a. Outline the general structure of a cell, and describe the characteristics of cells that qualify them as living.
    b. Why are viruses not considered living?
    c. Compare the general characteristics of procaryotic and eucaryotic cells.
    d. What are cellular membranes, and what are their functions?
    e. Explain the fluid mosaic model of a membrane.

CRITICAL-THINKING QUESTIONS

1. The "octet rule" in chemistry helps predict the tendency of atoms to acquire or donate electrons from the outer shell. It says that those with fewer than 4 tend to donate electrons and those with more than 4 tend to accept additional electrons; those with exactly 4 can do both. Using this rule, determine what category each of the following elements falls into: N, S, C, P, O, H, Ca, Fe, and Mg. (You will need to work out the valence of the atoms.)
2. Predict the kinds of bonds that occur in ammonium (NH_3), phosphate (PO_4), disulfide (S–S), and magnesium chloride ($MgCl_2$). (Use simple models such as those in figure 2.3.)
3. Work out the following problems:
   a. What is the number of protons in helium?
   b. Will an H bond form between $H_3C—CH═O$ and H_2O? Why or why not?
   c. Draw the following molecules and determine which are polar: Cl_2, NH_3, CH_4.
   d. What is the pH of a solution with a concentration of 0.00001 moles/ml (M) of H^+?
   e. What is the pH of a solution with a concentration of 0.00001 moles/ml (M) of OH^-?
4. a. Describe how hydration spheres are formed around cations and anions.
   b. Which substances will be expected to be hydrophilic and hydrophobic, and what makes them so?
   c. Distinguish between polar and ionic compounds, using your own words.
5. In what way are carbon-based compounds like children's Tinker Toys or Lego blocks?
6. Is galactose an aldehyde or a ketone sugar?
7. a. How many water molecules are released when a triglyceride is formed?
   b. How many peptide bonds are in a tetrapeptide?
8. a. Use pipe cleaners to help understand the formation of the 2^o and 3° structures of proteins.
   b. Note the various ways that your pipe cleaner structure can be folded and the diversity of shapes that can be formed.
9. a. Looking at figure 2.25, can you see why adenine forms hydrogen bonds with thymine and why cytosine forms them with guanine?
   b. Show on paper the steps in replication of the following segment of DNA:

```
ATGTTCCCGATCGGC
|||||||||||||||
TACAAGGGCTAGCCG
```

10. A useful mnemonic (memory) device for recalling the major characteristics of life is: *G*iant *R*ats *H*ave *M*any *C*olored *T*eeth. Can you list some of the characteristics for which each letter stands?

INTERNET SEARCH TOPIC

Go to a search engine and explore the topic of isotopes and dating ancient rocks. How can isotopes be used to determine if rocks contain evidence of life?