2

The Nature of Molecules

Concept Outline

2.1 Atoms are nature's building material.

Atoms. All substances are composed of tiny particles called atoms, each a positively charged nucleus around which orbit negative electrons.

Electrons Determine the Chemical Behavior of Atoms. Electrons orbit the nucleus of an atom; the closer an electron's orbit to the nucleus, the lower its energy level.

2.2 The atoms of living things are among the smallest.

Kinds of Atoms. Of the 92 naturally occurring elements, only 11 occur in organisms in significant amounts.

2.3 Chemical bonds hold molecules together.

Ionic Bonds Form Crystals. Atoms are linked together into molecules, joined by chemical bonds that result from forces like the attraction of opposite charges or the sharing of electrons.

Covalent Bonds Build Stable Molecules. Chemical bonds formed by the sharing of electrons can be very strong, and require much energy to break.

2.4 Water is the cradle of life.

Chemistry of Water. Water forms weak chemical associations that are responsible for much of the organization of living chemistry.

Water Atoms Act Like Tiny Magnets. Because electrons are shared unequally by the hydrogen and oxygen atoms of water, a partial charge separation occurs. Each water atom acquires a positive and negative pole and is said to be "polar."

Water Clings to Polar Molecules. Because the opposite partial charges of polar molecules attract one another, water tends to cling to itself and other polar molecules and to exclude nonpolar molecules.

Water Ionizes. Because its covalent bonds occasionally break, water contains a low concentration of hydrogen (H^+) and hydroxide (OH^-) ions, the fragments of broken water molecules.

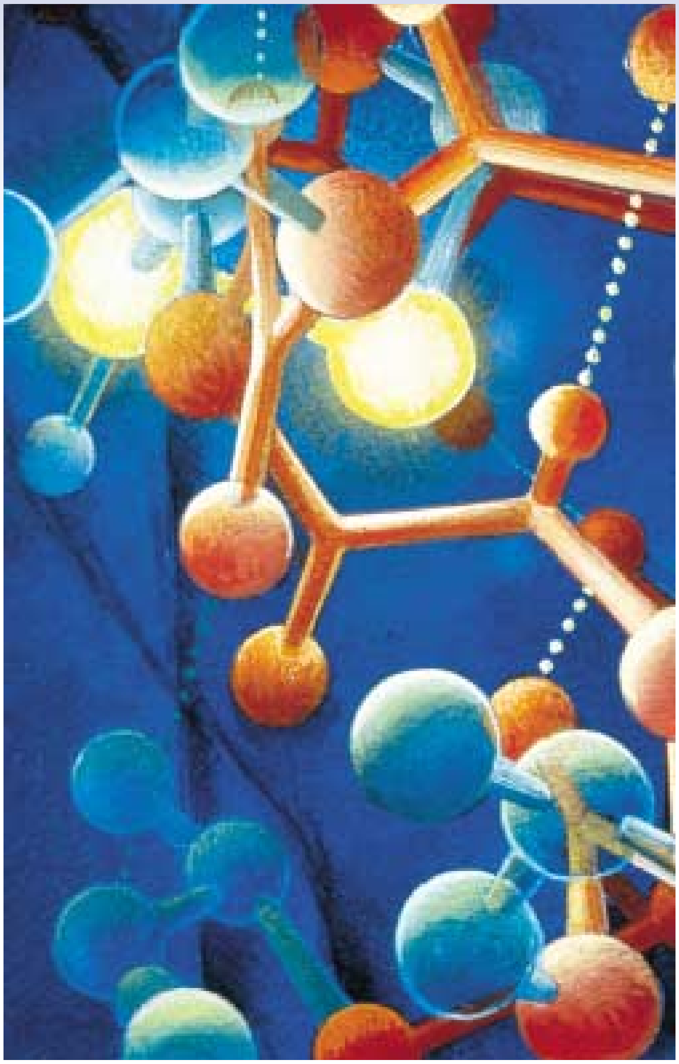


FIGURE 2.1
Cells are made of molecules. Specific, often simple, combinations of atoms yield an astonishing diversity of molecules within the cell, each with unique functional characteristics.

About 10 to 20 billion years ago, an enormous explosion likely marked the beginning of the universe. With this explosion began the process of evolution, which eventually led to the origin and diversification of life on earth. When viewed from the perspective of 20 billion years, life within our solar system is a recent development, but to understand the origin of life, we need to consider events that took place much earlier. The same processes that led to the evolution of life were responsible for the evolution of molecules (figure 2.1). Thus, our study of life on earth begins with physics and chemistry. As chemical machines ourselves, we must understand chemistry to begin to understand our origins.

2.1 Atoms are nature's building material.

Atoms

Any substance in the universe that has mass (see below) and occupies space is defined as **matter.** All matter is composed of extremely small particles called **atoms.** Because of their size, atoms are difficult to study. Not until early in this century did scientists carry out the first experiments suggesting what an atom is like.

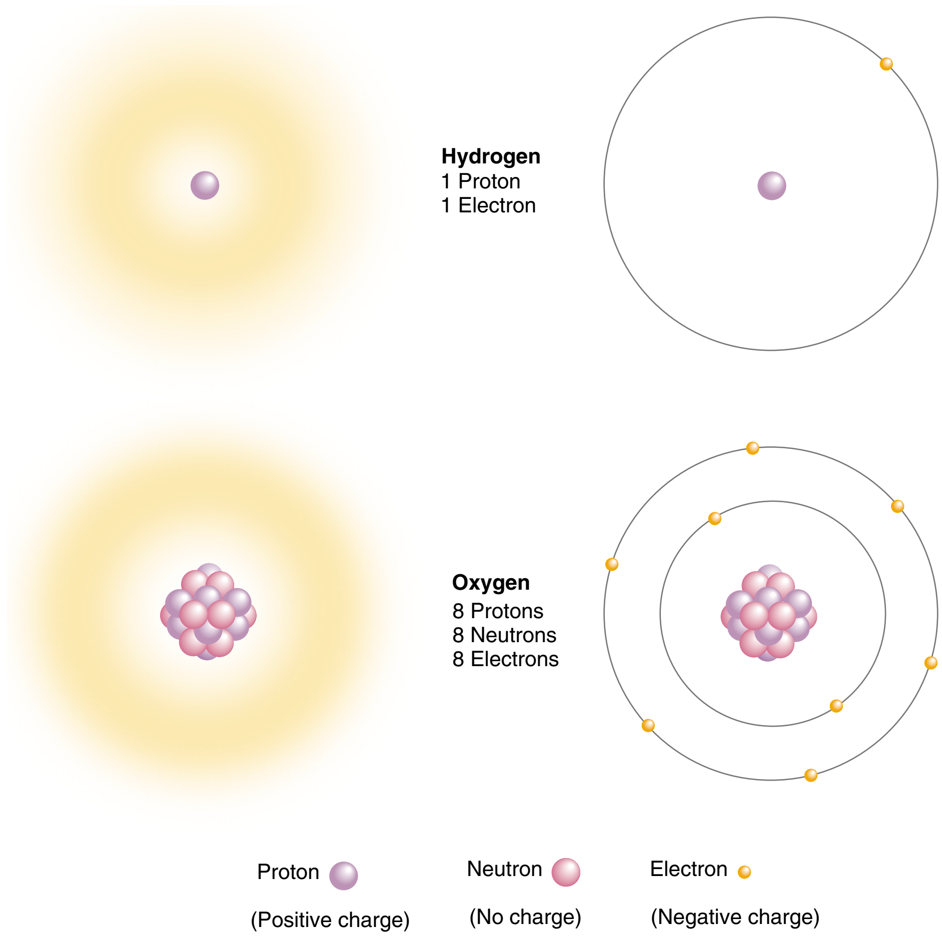


FIGURE 2.2
Basic structure of atoms. All atoms have a nucleus consisting of protons and neutrons, except hydrogen, the smallest atom, which has only one proton and no neutrons in its nucleus. Oxygen, for example, has eight protons and eight neutrons in its nucleus. Electrons spin around the nucleus a far distance away from the nucleus.

The Structure of Atoms

Objects as small as atoms can be "seen" only indirectly, by using very complex technology such as tunneling microcopy. We now know a great deal about the complexities of atomic structure, but the simple view put forth in 1913 by the Danish physicist Niels Bohr provides a good starting point. Bohr proposed that every atom possesses an orbiting cloud of tiny subatomic particles called **electrons** whizzing around a core like the planets of a miniature solar system. At the center of each atom is a small, very dense nucleus formed of two other kinds of subatomic particles, **protons** and **neutrons** (figure 2.2).

Within the nucleus, the cluster of protons and neutrons is held together by a force that works only over short subatomic distances. Each proton carries a positive (+) charge, and each electron carries a negative (–) charge. Typically an atom has one electron for each proton. The number of protons (the atom's **atomic number**) determines the chemical character of the atom, because it dictates the number of electrons orbiting the nucleus which are available for chemical activity. Neutrons, as their name implies, possess no charge.

Atomic Mass

The terms *mass* and *weight* are often used interchangeably, but they have slightly different meanings. *Mass* refers to the amount of a substance, while *weight* refers to the force gravity exerts on a substance. Hence, an object has the same *mass* whether it is on the earth or the moon, but its *weight* will be greater on the earth because the earth's gravitational force is greater than the moon's. The **atomic mass** of an atom is equal to the sum of the masses of its protons and neutrons. Atoms that occur naturally on earth contain from 1 to 92 protons and up to 146 neutrons.

The mass of atoms and subatomic particles is measured in units called *daltons.* To give you an idea of just how small these units are, note that it takes 602 million million billion (6.02×10^{23}) daltons to make 1 gram! A proton weighs approximately 1 dalton (actually 1.009 daltons), as does a neutron (1.007 daltons). In contrast, electrons weigh only $\frac{1}{1840}$ of a dalton, so their contribution to the overall mass of an atom is negligible.

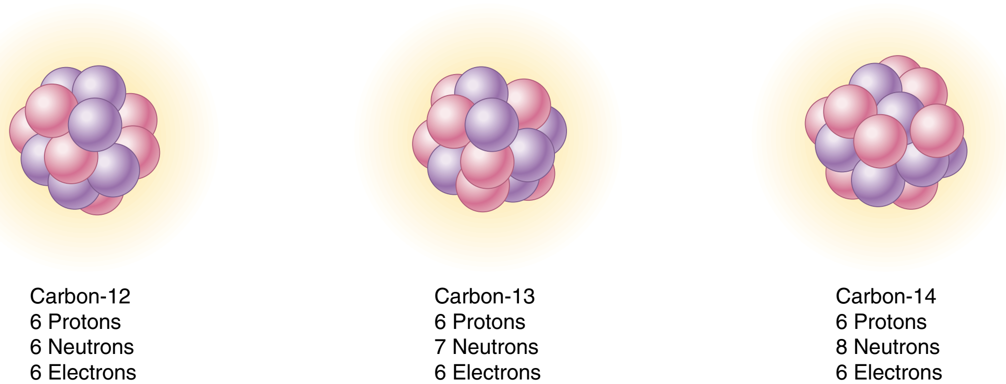


FIGURE 2.3
The three most abundant isotopes of carbon. Isotopes of a particular atom have different numbers of neutrons.

Isotopes

Atoms with the same atomic number (that is, the same number of protons) have the same chemical properties and are said to belong to the same **element.** Formally speaking, an element is any substance that cannot be broken down to any other substance by ordinary chemical means. However, while all atoms of an element have the same number of protons, they may not all have the same number of neutrons. Atoms of an element that possess different numbers of neutrons are called **isotopes** of that element. Most elements in nature exist as mixtures of different isotopes. Carbon (C), for example, has three isotopes, all containing six protons (figure 2.3). Over 99% of the carbon found in nature exists as an isotope with six neutrons. Because its total mass is 12 daltons (6 from protons plus 6 from neutrons), this isotope is referred to as carbon-12, and symbolized ^{12}C. Most of the rest of the naturally occurring carbon is carbon-13, an isotope with seven neutrons. The rarest carbon isotope is carbon-14, with eight neutrons. Unlike the other two isotopes, carbon-14 is unstable: its nucleus tends to break up into elements with lower atomic numbers. This nuclear breakup, which emits a significant amount of energy, is called radioactive decay, and isotopes that decay in this fashion are **radioactive isotopes.**

Some radioactive isotopes are more unstable than others and therefore decay more readily. For any given isotope, however, the rate of decay is constant. This rate is usually expressed as the **half-life,** the time it takes for one half of the atoms in a sample to decay. Carbon-14, for example, has a half-life of about 5600 years. A sample of carbon containing 1 gram of carbon-14 today would contain 0.5 gram of carbon-14 after 5600 years, 0.25 gram 11,200 years from now, 0.125 gram 16,800 years from now, and so on. By determining the ratios of the different isotopes of carbon and other elements in biological samples and in rocks, scientists are able to accurately determine when these materials formed.

While there are many useful applications of radioactivity, there are also harmful side effects that must be considered in any planned use of radioactive substances. Radioactive substances emit energetic subatomic particles that have the potential to severely damage living cells, producing mutations in their genes, and, at high doses, cell death. Consequently, exposure to radiation is now very carefully controlled and regulated. Scientists who work with radioactivity (basic researchers as well as applied scientists such as X-ray technologists) wear radiation-sensitive badges to monitor the total amount of radioactivity to which they are exposed. Each month the badges are collected and scrutinized. Thus, employees whose work places them in danger of excessive radioactive exposure are equipped with an "early warning system."

Electrons

The positive charges in the nucleus of an atom are counterbalanced by negatively charged electrons orbiting at varying distances around the nucleus. Thus, atoms with the same number of protons and electrons are electrically neutral, having no net charge.

Electrons are maintained in their orbits by their attraction to the positively charged nucleus. Sometimes other forces overcome this attraction and an atom loses one or more electrons. In other cases, atoms may gain additional electrons. Atoms in which the number of electrons does not equal the number of protons are known as **ions,** and they carry a net electrical charge. An atom that has more protons than electrons has a net positive charge and is called a **cation.** For example, an atom of sodium (Na) that has lost one electron becomes a sodium ion (Na^+), with a charge of +1. An atom that has fewer protons than electrons carries a net negative charge and is called an **anion.** A chlorine atom (Cl) that has gained one electron becomes a chloride ion (Cl^-), with a charge of –1.

An atom consists of a nucleus of protons and neutrons surrounded by a cloud of electrons. The number of its electrons largely determines the chemical properties of an atom. Atoms that have the same number of protons but different numbers of neutrons are called isotopes. Isotopes of an atom differ in atomic mass but have similar chemical properties.

Electrons Determine the Chemical Behavior of Atoms

The key to the chemical behavior of an atom lies in the arrangement of its electrons in their orbits. It is convenient to visualize individual electrons as following discrete circular orbits around a central nucleus, as in the Bohr model of the atom. However, such a simple picture is not realistic. It is not possible to precisely locate the position of any individual electron precisely at any given time. In fact, a particular electron can be anywhere at a given instant, from close to the nucleus to infinitely far away from it.

However, a particular electron is more likely to be located in some positions than in others. The area around a nucleus where an electron is most likely to be found is called the orbital of that electron (figure 2.4). Some electron orbitals near the nucleus are spherical (*s* orbitals), while others are dumbbell-shaped (*p* orbitals). Still other orbitals, more distant from the nucleus, may have different shapes. Regardless of its shape, no orbital may contain more than two electrons.

Almost all of the volume of an atom is empty space, because the electrons are quite far from the nucleus relative to its size. If the nucleus of an atom were the size of an apple, the orbit of the nearest electron would be more than 1600 meters away. Consequently, the nuclei of two atoms never come close enough in nature to interact with each other. It is for this reason that an atom's electrons, not its protons or neutrons, determine its chemical behavior. This also explains why the isotopes of an element, all of which have the same arrangement of electrons, behave the same way chemically.

Energy within the Atom

All atoms possess energy, defined as the ability to do work. Because electrons are attracted to the positively charged nucleus, it takes work to keep them in orbit, just as it takes work to hold a grapefruit in your hand against the pull of gravity. The grapefruit is said to possess *potential energy*, the ability to do work, because of its position; if you were to release it, the grapefruit would fall and its energy would be reduced. Conversely, if you were to move the grapefruit to the top of a building, you would increase its potential energy. Similarly, electrons have potential energy of position. To oppose the attraction of the nucleus and move the electron to a more distant orbital requires an input of energy and results in an electron with greater potential energy. This is how chlorophyll captures energy from light during photosynthesis (chapter 10)—the light excites electrons in the chlorophyll. Moving an electron closer to the nucleus has the opposite effect: energy is released, usually as heat, and the electron ends up with less potential energy (figure 2.5).

A given atom can possess only certain discrete amounts of energy. Like the potential energy of a grapefruit on a step of a staircase, the potential energy contributed by the position of an electron in an atom can have only certain values.

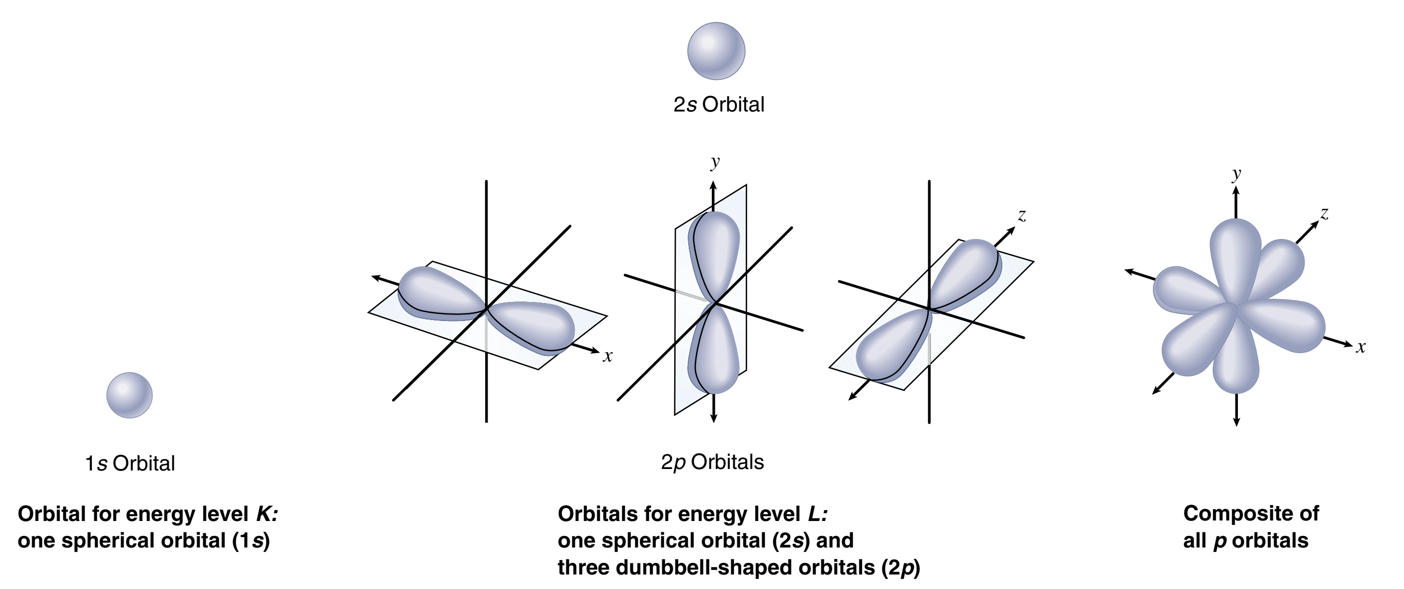


FIGURE 2.4
Electron orbitals. The lowest energy level or electron shell, which is nearest the nucleus, is level *K*. It is occupied by a single *s* orbital, referred to as 1*s*. The next highest energy level, *L*, is occupied by four orbitals: one *s* orbital (referred to as the 2*s* orbital) and three *p* orbitals (each referred to as a 2*p* orbital). The four *L*-level orbitals compactly fill the space around the nucleus, like two pyramids set base-to-base.

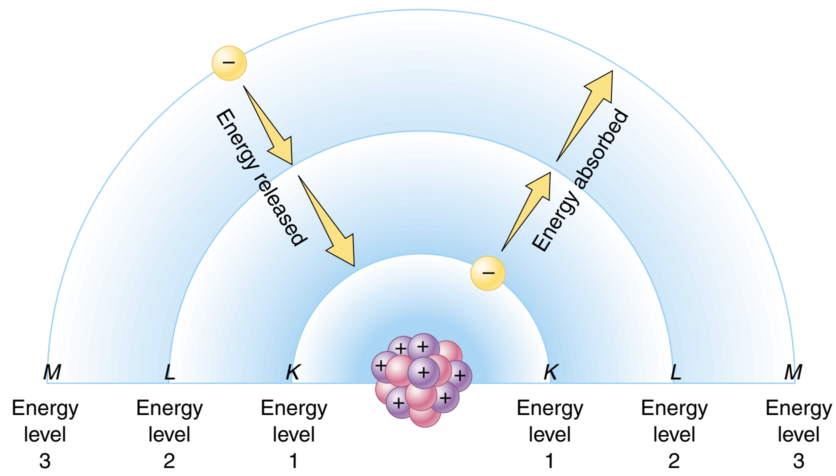


FIGURE 2.5
Atomic energy levels. When an electron absorbs energy, it moves to higher energy levels farther from the nucleus. When an electron releases energy, it falls to lower energy levels closer to the nucleus.

Every atom exhibits a ladder of potential energy values, rather than a continuous spectrum of possibilities, a discrete set of orbits at particular distances from the nucleus.

During some chemical reactions, electrons are transferred from one atom to another. In such reactions, the loss of an electron is called **oxidation,** and the gain of an electron is called **reduction** (figure 2.6). It is important to realize that when an electron is transferred in this way, it keeps its energy of position. In organisms, chemical energy is stored in high-energy electrons that are transferred from one atom to another in reactions involving oxidation and reduction.

Because the amount of energy an electron possesses is related to its distance from the nucleus, electrons that are the same distance from the nucleus have the same energy, even if they occupy different orbitals. Such electrons are said to occupy the same **energy level.** In a schematic diagram of an atom (figure 2.7), the nucleus is represented as a small circle and the electron energy levels are drawn as concentric rings, with the energy level increasing with distance from the nucleus. Be careful not to confuse energy levels, which are drawn as rings to indicate an electron's *energy*, with orbitals, which have a variety of three-dimensional shapes and indicate an electron's most likely *location.*

Electrons orbit a nucleus in paths called orbitals. No orbital can contain more than two electrons, but many orbitals may be the same distance from the nucleus and, thus, contain electrons of the same energy.

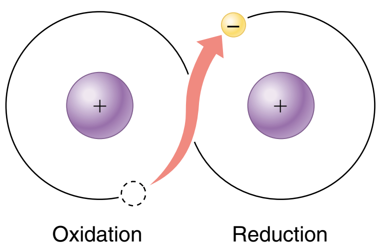


FIGURE 2.6
Oxidation and reduction. Oxidation is the loss of an electron; reduction is the gain of an electron.

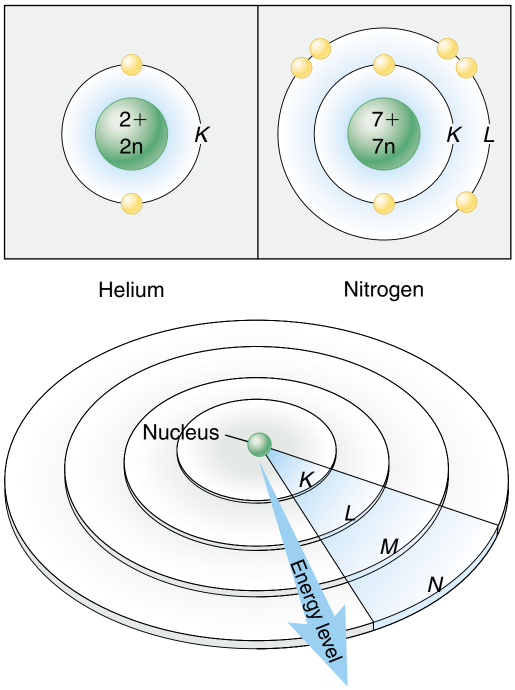


FIGURE 2.7
Electron energy levels for helium and nitrogen. Gold balls represent the electrons. Each concentric circle represents a different distance from the nucleus and, thus, a different electron energy level.

Kinds of Atoms

There are 92 naturally occurring elements, each with a different number of protons and a different arrangement of electrons. When the nineteenth-century Russian chemist Dmitri Mendeleev arranged the known elements in a table according to their atomic mass (figure 2.8), he discovered one of the great generalizations in all of science. Mendeleev found that the elements in the table exhibited a pattern of chemical properties that repeated itself in groups of eight elements. This periodically repeating pattern lent the table its name: the periodic table of elements.

The Periodic Table

The eight-element periodicity that Mendeleev found is based on the interactions of the electrons in the outer energy levels of the different elements. These electrons are called **valence electrons** and their interactions are the basis for the differing chemical properties of the elements. For most of the atoms important to life, an outer energy level can contain no more than eight electrons; the chemical behavior of an element reflects how many of the eight positions are filled. Elements possessing all eight electrons in their outer energy level (two for helium) are **inert,** or nonreactive; they include helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). In sharp contrast, elements with seven electrons (one fewer than the maximum number of eight) in their outer energy level, such as fluorine (F), chlorine (Cl), and bromine (Br), are highly reactive. They tend to gain the extra electron needed to fill the energy level. Elements with only one electron in their outer energy level, such as lithium (Li), sodium (Na), and potassium (K), are also very reactive; they tend to lose the single electron in their outer level.

Mendeleev's periodic table thus leads to a useful generalization, the **octet rule** (Latin *octo,* "eight") or **rule of eight:** atoms tend to establish completely full outer energy levels. Most chemical behavior can be predicted quite accurately from this simple rule, combined with the tendency of atoms to balance positive and negative charges.

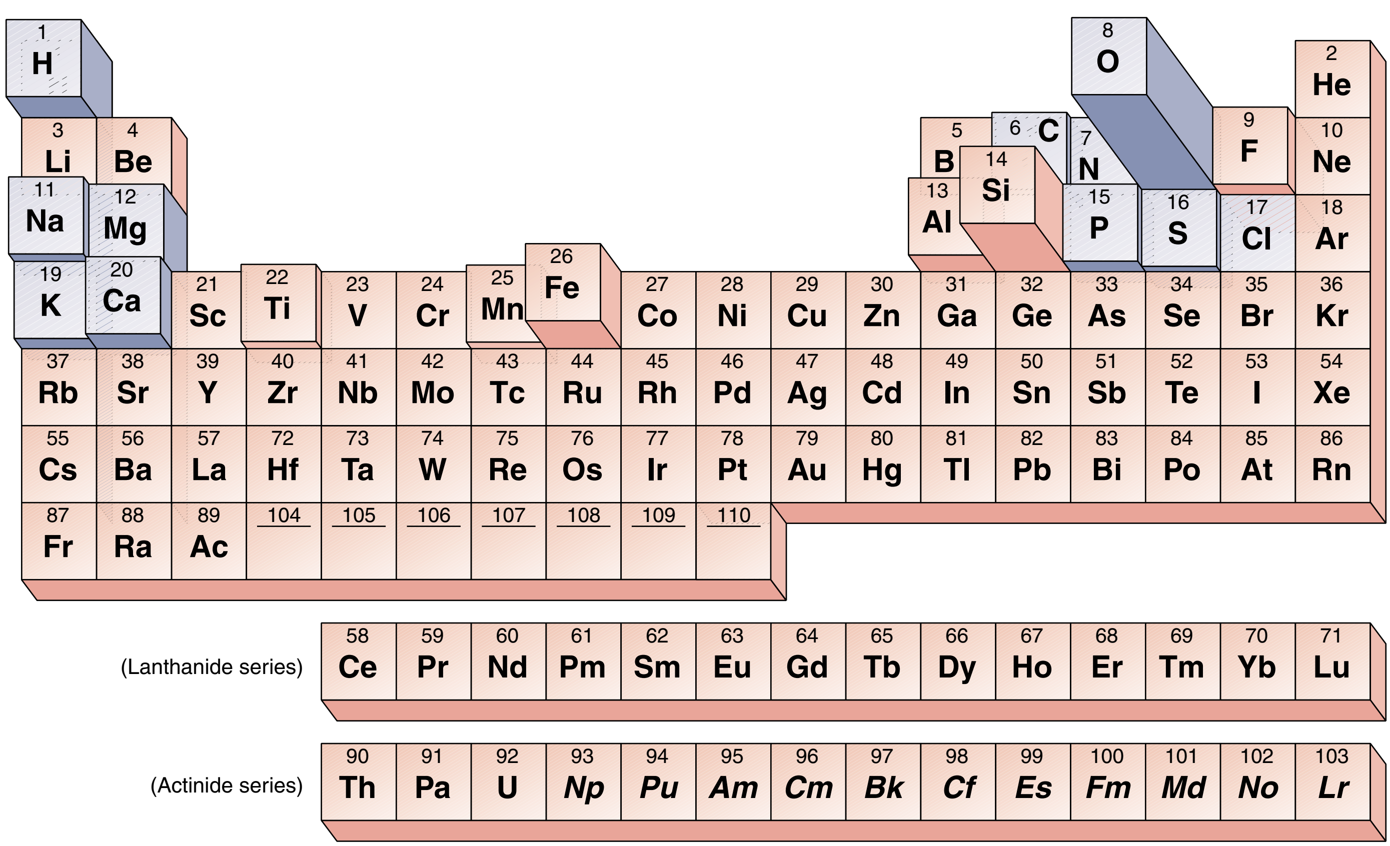


FIGURE 2.8
Periodic table of the elements. In this representation, the frequency of elements that occur in the earth's crust is indicated by the height of the block. Elements found in significant amounts in living organisms are shaded in blue.

Table 2.1 The Most Common Elements on Earth and Their Distribution in the Human Body					
Element	**Symbol**	**Atomic Number**	**Approximate Percent of Earth's Crust by Weight**	**Percent of Human Body by Weight**	**Importance or Function**
Oxygen	O	8	46.6	65.0	Required for cellular respiration; component of water
Silicon	Si	14	27.7	Trace	
Aluminum	Al	13	6.5	Trace	
Iron	Fe	26	5.0	Trace	Critical component of hemoglobin in the blood
Calcium	Ca	20	3.6	1.5	Component of bones and teeth; triggers muscle contraction
Sodium	Na	11	2.8	0.2	Principal positive ion outside cells; important in nerve function
Potassium	K	19	2.6	0.4	Principal positive ion inside cells; important in nerve function
Magnesium	Mg	12	2.1	0.1	Critical component of many energy-transferring enzymes
Hydrogen	H	1	0.14	9.5	Electron carrier; component of water and most organic molecules
Manganese	Mn	25	0.1	Trace	
Fluorine	F	9	0.07	Trace	
Phosphorus	P	15	0.07	1.0	Backbone of nucleic acids; important in energy transfer
Carbon	C	6	0.03	18.5	Backbone of organic molecules
Sulfur	S	16	0.03	0.3	Component of most proteins
Chlorine	Cl	17	0.01	0.2	Principal negative ion outside cells
Vanadium	V	23	0.01	Trace	
Chromium	Cr	24	0.01	Trace	
Copper	Cu	29	0.01	Trace	Key component of many enzymes
Nitrogen	N	7	Trace	3.3	Component of all proteins and nucleic acids
Boron	B	5	Trace	Trace	
Cobalt	Co	27	Trace	Trace	
Zinc	Zn	30	Trace	Trace	Key component of some enzymes
Selenium	Se	34	Trace	Trace	
Molybdenum	Mo	42	Trace	Trace	Key component of many enzymes
Tin	Sn	50	Trace	Trace	
Iodine	I	53	Trace	Trace	Component of thyroid hormone

Distribution of the Elements

Of the 92 naturally occurring elements on earth, only 11 are found in organisms in more than trace amounts (0.01% or higher). These 11 elements have atomic numbers less than 21 and, thus, have low atomic masses. Table 2.1 lists the levels of various elements in the human body; their levels in other organisms are similar. Inspection of this table suggests that the distribution of elements in living systems is by no means accidental. The most common elements inside organisms are not the elements that are most abundant in the earth's crust. For example, silicon, aluminum, and iron constitute 39.2% of the earth's crust, but they exist in trace amounts in the human body. On the other hand, carbon atoms make up 18.5% of the human body but only 0.03% of the earth's crust.

Ninety-two elements occur naturally on earth; only eleven of them are found in significant amounts in living organisms. Four of them—oxygen, hydrogen, carbon, nitrogen—constitute 96.3% of the weight of your body.

2.3 Chemical bonds hold molecules together.

Ionic Bonds Form Crystals

A group of atoms held together by energy in a stable association is called a **molecule.** When a molecule contains atoms of more than one element, it is called a **compound.** The atoms in a molecule are joined by **chemical bonds;** these bonds can result when atoms with opposite charges attract (ionic bonds), when two atoms share one or more pairs of electrons (covalent bonds), or when atoms interact in other ways. We will start by examining **ionic bonds,** which form when atoms with opposite electrical charges (ions) attract.

A Closer Look at Table Salt

Common table salt, sodium chloride (NaCl), is a lattice of ions in which the atoms are held together by ionic bonds (figure 2.9). Sodium has 11 electrons: 2 in the inner energy level, 8 in the next level, and 1 in the outer (valence) level. The valence electron is unpaired (free) and has a strong tendency to join with another electron. A stable configuration can be achieved if the valence electron is lost to another atom that also has an unpaired electron. The loss of this electron results in the formation of a positively charged sodium ion, Na^+.

The chlorine atom has 17 electrons: 2 in the inner energy level, 8 in the next level, and 7 in the outer level. Hence, one of the orbitals in the outer energy level has an unpaired electron. The addition of another electron to the outer level fills that level and causes a negatively charged chloride ion, Cl^-, to form.

When placed together, metallic sodium and gaseous chlorine react swiftly and explosively, as the sodium atoms donate electrons to chlorine, forming Na^+ and Cl^- ions. Because opposite charges attract, the Na^+ and Cl^- remain associated in an **ionic compound,** NaCl, which is electrically neutral. However, the electrical attractive force holding NaCl together is not directed specifically between particular Na^+ and Cl^- ions, and no discrete sodium chloride molecules form. Instead, the force exists between any one ion and all neighboring ions of the opposite charge, and the ions aggregate in a crystal matrix with a precise geometry. Such aggregations are what we know as salt crystals. If a salt such as NaCl is placed in water, the electrical attraction of the water molecules, for reasons we will point out later in this chapter, disrupts the forces holding the ions in their crystal matrix, causing the salt to dissolve into a roughly equal mixture of free Na^+ and Cl^- ions.

An ionic bond is an attraction between ions of opposite charge in an ionic compound. Such bonds are not formed between particular ions in the compound; rather, they exist between an ion and all of the oppositely charged ions in its immediate vicinity.

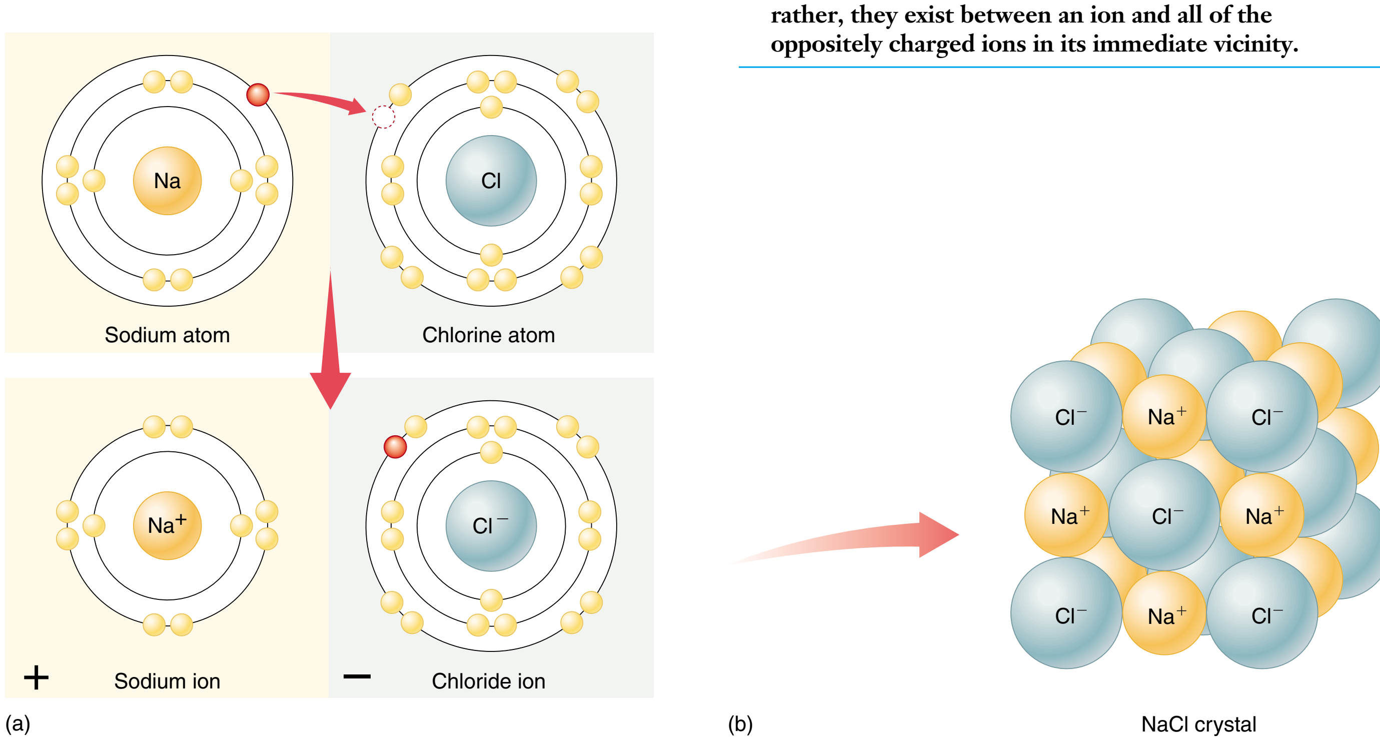


FIGURE 2.9
The formation of ionic bonds by sodium chloride. (*a*) When a sodium atom donates an electron to a chlorine atom, the sodium atom becomes a positively charged sodium ion, and the chlorine atom becomes a negatively charged chloride ion. (*b*) Sodium chloride forms a highly regular lattice of alternating sodium ions and chloride ions.

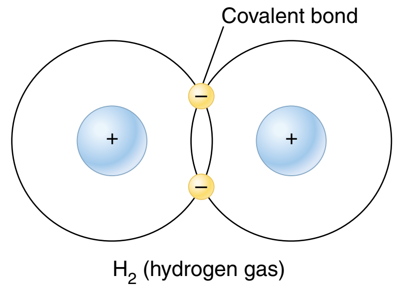
Covalent Bonds Build Stable Molecules

Covalent bonds form when two atoms share one or more pairs of valence electrons. Consider hydrogen (H) as an example. Each hydrogen atom has an unpaired electron and an unfilled outer energy level; for these reasons the hydrogen atom is unstable. When two hydrogen atoms are close to each other, however, each atom's electron can orbit both nuclei. In effect, the nuclei are able to share their electrons. The result is a diatomic (two-atom) molecule of hydrogen gas (figure 2.10).

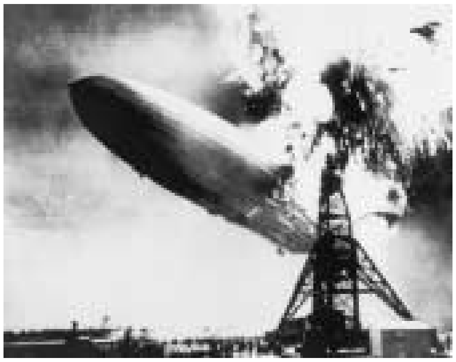
The molecule formed by the two hydrogen atoms is stable for three reasons:

1. **It has no net charge.** The diatomic molecule formed as a result of this sharing of electrons is not charged, because it still contains two protons and two electrons.
2. **The octet rule is satisfied.** Each of the two hydrogen atoms can be considered to have two orbiting electrons in its outer energy level. This satisfies the octet rule, because each shared electron orbits both nuclei and is included in the outer energy level of *both* atoms.
3. **It has no free electrons.** The bonds between the two atoms also pair the two free electrons.

Unlike ionic bonds, covalent bonds are formed between two specific atoms, giving rise to true, discrete molecules. While ionic bonds can form regular crystals, the more specific associations made possible by covalent bonds allow the formation of complex molecular structures.



(a)



(b)

FIGURE 2.10
Hydrogen gas. (*a*) Hydrogen gas is a diatomic molecule composed of two hydrogen atoms, each sharing its electron with the other. (*b*) The flash of fire that consumed the *Hindenburg* occurred when the hydrogen gas that was used to inflate the dirigible combined explosively with oxygen gas in the air to form water.

Covalent Bonds Can Be Very Strong

The strength of a covalent bond depends on the number of shared electrons. Thus **double bonds,** which satisfy the octet rule by allowing two atoms to share *two* pairs of electrons, are stronger than **single bonds,** in which only one electron pair is shared. This means more chemical energy is required to break a double bond than a single bond. The strongest covalent bonds are **triple bonds,** such as those that link the two nitrogen atoms of nitrogen gas molecules. Covalent bonds are represented in chemical formulations as lines connecting atomic symbols, where each line between two bonded atoms represents the sharing of one pair of electrons. The **structural formulas** of hydrogen gas and oxygen gas are H—H and O=O, respectively, while their **molecular formulas** are H_2 and O_2.

Molecules with Several Covalent Bonds

Molecules often consist of more than two atoms. One reason that larger molecules may be formed is that a given atom is able to share electrons with more than one other atom. An atom that requires two, three, or four additional electrons to fill its outer energy level completely may acquire them by sharing its electrons with two or more other atoms.

For example, the carbon atom (C) contains six electrons, four of which are in its outer energy level. To satisfy the octet rule, a carbon atom must gain access to four additional electrons; that is, it must form four covalent bonds. Because four covalent bonds may form in many ways, carbon atoms are found in many different kinds of molecules.

Chemical Reactions

The formation and breaking of chemical bonds, the essence of chemistry, is called a **chemical reaction.** All chemical reactions involve the shifting of atoms from one molecule or ionic compound to another, without any change in the number or identity of the atoms. For convenience, we refer to the original molecules before the reaction starts as **reactants,** and the molecules resulting from the chemical reaction as **products.** For example:

$$\underset{\text{reactants}}{A—B + C—D} \rightarrow \underset{\text{products}}{A—C + B + D}$$

The extent to which chemical reactions occur is influenced by several important factors:

1. **Temperature.** Heating up the reactants increases the rate of a reaction (as long as the temperature isn't so high as to destroy the molecules).
2. **Concentration of reactants and products.** Reactions proceed more quickly when more reactants are available. An accumulation of products typically speeds reactions in the reverse direction.
3. **Catalysts.** A catalyst is a substance that increases the rate of a reaction. It doesn't alter the reaction's equilibrium between reactants and products, but it does shorten the time needed to reach equilibrium, often dramatically. In organisms, proteins called enzymes catalyze almost every chemical reaction.

A covalent bond is a stable chemical bond formed when two atoms share one or more pairs of electrons.

2.4 Water is the cradle of life.

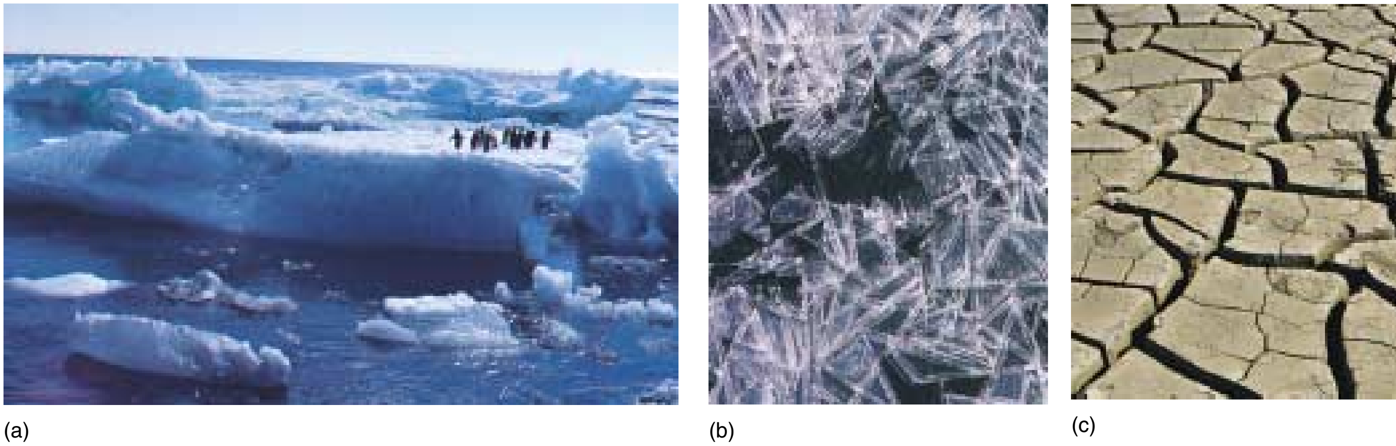


FIGURE 2.11
Water takes many forms. As a liquid, water fills our rivers and runs down over the land to the sea. (*a*) The iceberg on which the penguins are holding their meeting was formed in Antarctica from a huge block of ice that broke away into the ocean water. (*b*) When water cools below 0°C, it forms beautiful crystals, familiar to us as snow and ice. However, water is not always plentiful. (*c*) At Badwater, in Death Valley, California, there is no hint of water except for the broken patterns of dried mud.

Chemistry of Water

Of all the molecules that are common on earth, only **water** exists as a liquid at the relatively low temperatures that prevail on the earth's surface, three-fourths of which is covered by liquid water (figure 2.11). When life was originating, water provided a medium in which other molecules could move around and interact without being held in place by strong covalent or ionic bonds. Life evolved as a result of these interactions, and it is still inextricably tied to water. Life began in water and evolved there for 3 billion years before spreading to land. About two-thirds of any organism's body is composed of water, and no organism can grow or reproduce in any but a water-rich environment. It is no accident that tropical rain forests are bursting with life, while dry deserts appear almost lifeless except when water becomes temporarily plentiful, such as after a rainstorm.

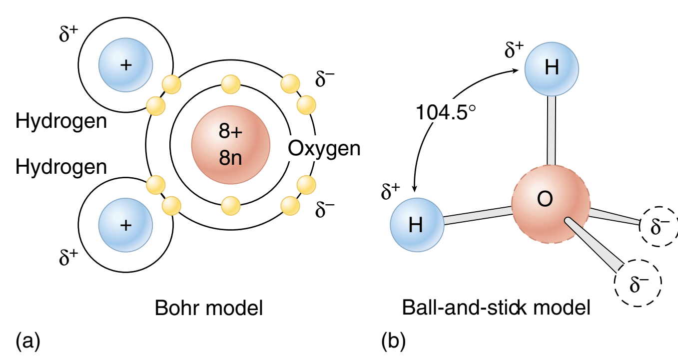


FIGURE 2.12
Water has a simple molecular structure. (*a*) Each molecule is composed of one oxygen atom and two hydrogen atoms. The oxygen atom shares one electron with each hydrogen atom. (*b*) The greater electronegativity of the oxygen atom makes the water molecule polar: water carries two partial negative charges (δ^-) near the oxygen atom and two partial positive charges (δ^+), one on each hydrogen atom.

The Atomic Structure of Water

Water has a simple atomic structure. It consists of an oxygen atom bound to two hydrogen atoms by two single covalent bonds (figure 2.12*a*). The resulting molecule is stable: it satisfies the octet rule, has no unpaired electrons, and carries no net electrical charge.

The single most outstanding chemical property of water is its ability to form weak chemical associations with only 5 to 10% of the strength of covalent bonds. This property, which derives directly from the structure of water, is responsible for much of the organization of living chemistry.

The chemistry of life is water chemistry. The way in which life first evolved was determined in large part by the chemical properties of the liquid water in which that evolution occurred.

Water Atoms Act Like Tiny Magnets

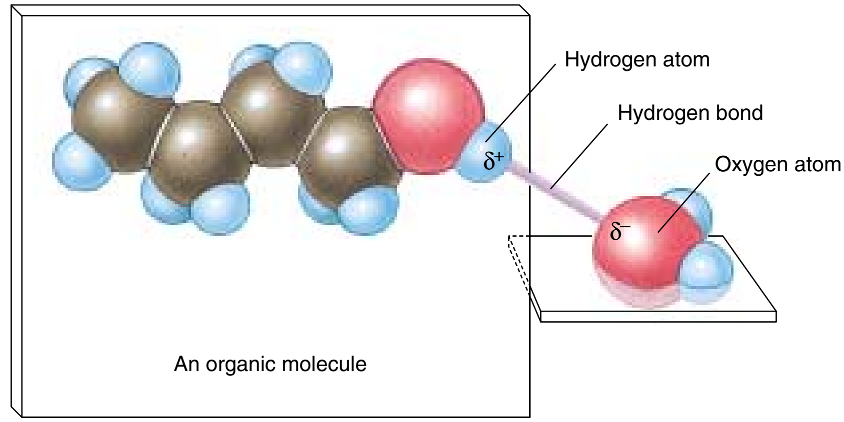


FIGURE 2.13
Structure of a hydrogen bond.

Both the oxygen and the hydrogen atoms attract the electrons they share in the covalent bonds of a water molecule; this attraction is called **electronegativity.** However, the oxygen atom is more electronegative than the hydrogen atoms, so it attracts the electrons more strongly than do the hydrogen atoms. As a result, the shared electrons in a water molecule are far more likely to be found near the oxygen nucleus than near the hydrogen nuclei. This stronger attraction for electrons gives the oxygen atom two partial negative charges (δ^-), as though the electron cloud were denser near the oxygen atom than around the hydrogen atoms. Because the water molecule as a whole is electrically neutral, each hydrogen atom carries a partial positive charge (δ^+). The Greek letter delta (δ) signifies a partial charge, much weaker than the full unit charge of an ion.

What would you expect the shape of a water molecule to be? Each of water's two covalent bonds has a partial charge at each end, δ^- at the oxygen end and δ^+ at the hydrogen end. The most stable arrangement of these charges is a *tetrahedron*, in which the two negative and two positive charges are approximately equidistant from one another (figure 2.12*b*). The oxygen atom lies at the center of the tetrahedron, the hydrogen atoms occupy two of the apexes, and the partial negative charges occupy the other two apexes. This results in a bond angle of 104.5° between the two covalent oxygen-hydrogen bonds. (In a regular tetrahedron, the bond angles would be 109.5°; in water, the partial negative charges occupy more space than the hydrogen atoms, and, therefore, they compress the oxygen-hydrogen bond angle slightly.)

The water molecule, thus, has distinct "ends," each with a partial charge, like the two poles of a magnet. (These partial charges are much less than the unit charges of ions, however.) Molecules that exhibit charge separation are called **polar molecules** because of their magnet-like poles, and water is one of the most polar molecules known. *The polarity of water underlies its chemistry and the chemistry of life.*

Polar molecules interact with one another, as the δ^- of one molecule is attracted to the δ^+ of another. Because many of these interactions involve hydrogen atoms, they are called **hydrogen bonds** (figure 2.13). Each hydrogen bond is individually very weak and transient, lasting on average only $\frac{1}{100{,}000{,}000{,}000}$ second (10^{-11} sec). However, the cumulative effects of large numbers of these bonds can be enormous. Water forms an abundance of hydrogen bonds, which are responsible for many of its important physical properties (table 2.2).

The water molecule is very polar, with ends that exhibit partial positive and negative charges. Opposite charges attract, forming weak linkages called hydrogen bonds.

Table 2.2 The Properties of Water

Property	Explanation	Example of Benefit to Life
Cohesion	Hydrogen bonds hold water molecules together	Leaves pull water upward from the roots; seeds swell and germinate
High specific heat	Hydrogen bonds absorb heat when they break, and release heat when they form, minimizing temperature changes	Water stabilizes the temperature of organisms and the environment
High heat of vaporization	Many hydrogen bonds must be broken for water to evaporate	Evaporation of water cools body surfaces
Lower density of ice	Water molecules in an ice crystal are spaced relatively far apart because of hydrogen bonding	Because ice is less dense than water, lakes do not freeze solid
High polarity	Polar water molecules are attracted to ions and polar compounds, making them soluble	Many kinds of molecules can move freely in cells, permitting a diverse array of chemical reactions

Water Clings to Polar Molecules

The polarity of water causes it to be attracted to other polar molecules. When the other molecules are also water, the attraction is referred to as **cohesion.** When the other molecules are of a different substance, the attraction is called **adhesion.** It is because water is cohesive that it is a liquid, and not a gas, at moderate temperatures.

The cohesion of liquid water is also responsible for its **surface tension.** Small insects can walk on water (figure 2.14) because at the air-water interface all of the hydrogen bonds in water face downward, causing the molecules of the water surface to cling together. Water is adhesive to any substance with which it can form hydrogen bonds. That is why substances containing polar molecules get "wet" when they are immersed in water, while those that are composed of nonpolar molecules (such as oils) do not.

The attraction of water to substances like glass with surface electrical charges is responsible for capillary action: if a glass tube with a narrow diameter is lowered into a beaker of water, water will rise in the tube above the level of the water in the beaker, because the adhesion of water to the glass surface, drawing it upward, is stronger than the force of gravity, drawing it down. The narrower the tube, the greater the electrostatic forces between the water and the glass, and the higher the water rises (figure 2.15).

FIGURE 2.14
Cohesion. Some insects, such as this water strider, literally walk on water. In this photograph you can see the dimpling the insect's feet make on the water as its weight bears down on the surface. Because the surface tension of the water is greater than the force that one foot brings to bear, the strider glides atop the surface of the water rather than sinking.

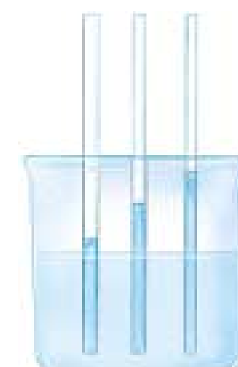


FIGURE 2.15
Capillary action. Capillary action causes the water within a narrow tube to rise above the surrounding water; the adhesion of the water to the glass surface, which draws water upward, is stronger than the force of gravity, which tends to draw it down. The narrower the tube, the greater the surface area available for adhesion for a given volume of water, and the higher the water rises in the tube.

Water Stores Heat

Water moderates temperature through two properties: its high specific heat and its high heat of vaporization. The temperature of any substance is a measure of how rapidly its individual molecules are moving. Because of the many hydrogen bonds that water molecules form with one another, a large input of thermal energy is required to break these bonds before the individual water molecules can begin moving about more freely and so have a higher temperature. Therefore, water is said to have a high **specific heat,** which is defined as the amount of heat that must be absorbed or lost by 1 gram of a substance to change its temperature by 1 degree Celsius (°C). Specific heat measures the extent to which a substance resists changing its temperature when it absorbs or loses heat. Because polar substances tend to form hydrogen bonds, and energy is needed to break these bonds, the more polar a substance is, the higher is its specific heat. The specific heat of water (1 calorie/gram/°C) is twice that of most carbon compounds and nine times that of iron. Only ammonia, which is more polar than water and forms very strong hydrogen bonds, has a higher specific heat than water (1.23 calories/gram/°C). Still, only 20% of the hydrogen bonds are broken as water heats from 0° to 100°C.

Because of its high specific heat, water heats up more slowly than almost any other compound and holds its temperature longer when heat is no longer applied. This characteristic enables organisms, which have a high water content, to maintain a relatively constant internal temperature. The heat generated by the chemical reactions inside cells would destroy the cells, if it were not for the high specific heat of the water within them.

A considerable amount of heat energy (586 calories) is required to change 1 gram of liquid water into a gas. Hence, water also has a high **heat of vaporization.** Because the transition of water from a liquid to a gas requires the input of energy to break its many hydrogen bonds, the evaporation of water from a surface causes cooling of that surface. Many organisms dispose of excess body heat by evaporative cooling; for example, humans and many other vertebrates sweat.

At low temperatures, water molecules are locked into a crystal-like lattice of hydrogen bonds, forming the solid we call ice (figure 2.16). Interestingly, ice is less dense than liquid water because the hydrogen bonds in ice space the water molecules relatively far apart. This unusual feature enables icebergs to float. Were it otherwise, ice would cover nearly all bodies of water, with only shallow surface melting annually.

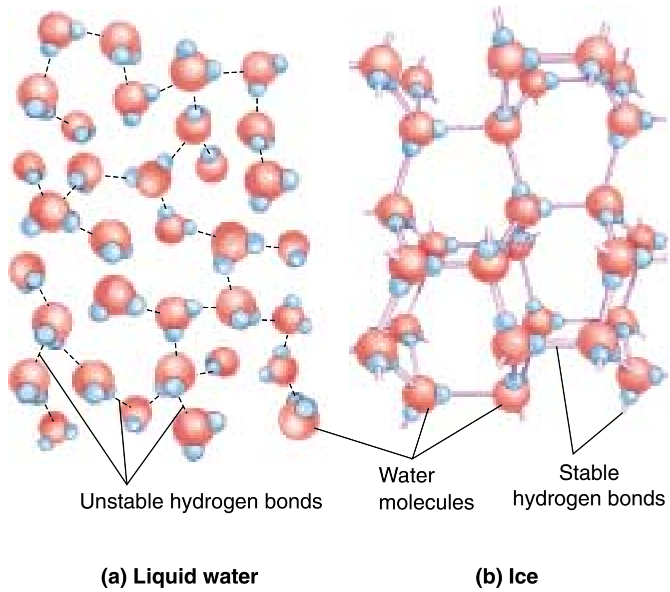


FIGURE 2.16
The role of hydrogen bonds in an ice crystal. (*a*) In liquid water, hydrogen bonds are not stable and constantly break and re-form. (*b*) When water cools below 0°C, the hydrogen bonds are more stable, and a regular crystalline structure forms in which the four partial charges of one water molecule interact with the opposite charges of other water molecules. Because water forms a crystal latticework, ice is less dense than liquid water and floats. If it did not, inland bodies of water far from the earth's equator might never fully thaw.

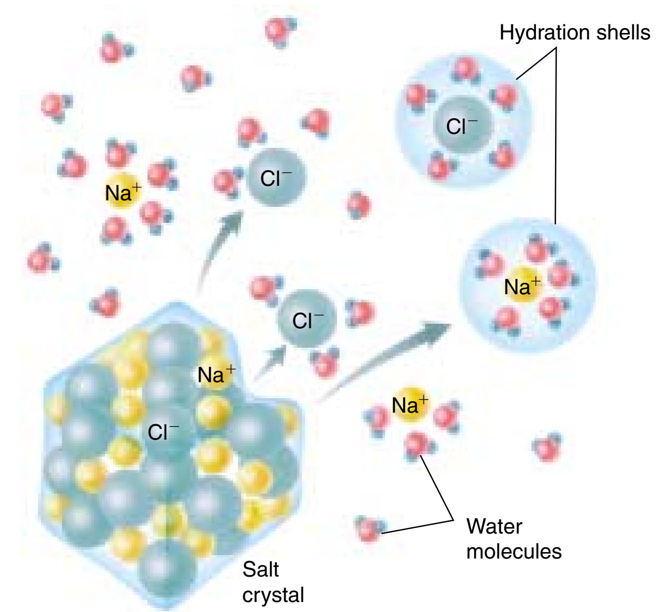


FIGURE 2.17
Why salt dissolves in water. When a crystal of table salt dissolves in water, individual Na^+ and Cl^- ions break away from the salt lattice and become surrounded by water molecules. Water molecules orient around Cl^- ions so that their partial positive poles face toward the negative Cl^- ion; water molecules surrounding Na^+ ions orient in the opposite way, with their partial negative poles facing the positive Na^+ ion. Surrounded by hydration shells, Na^+ and Cl^- ions never reenter the salt lattice.

Water Is a Powerful Solvent

Water is an effective solvent because of its ability to form hydrogen bonds. Water molecules gather closely around any substance that bears an electrical charge, whether that substance carries a full charge (ion) or a charge separation (polar molecule). For example, sucrose (table sugar) is composed of molecules that contain slightly polar hydroxyl (OH) groups. A sugar crystal dissolves rapidly in water because water molecules can form hydrogen bonds with individual hydroxyl groups of the sucrose molecules. Therefore, sucrose is said to be *soluble* in water. Every time a sucrose molecule dissociates or breaks away from the crystal, water molecules surround it in a cloud, forming a **hydration shell** and preventing it from associating with other sucrose molecules. Hydration shells also form around ions such as Na^+ and Cl^- (figure 2.17).

Water Organizes Nonpolar Molecules

Water molecules always tend to form the maximum possible number of hydrogen bonds. When nonpolar molecules such as oils, which do not form hydrogen bonds, are placed in water, the water molecules act to exclude them. The nonpolar molecules are forced into association with one another, thus minimizing their disruption of the hydrogen bonding of water. In effect, they shrink from contact with water and for this reason they are referred to as **hydrophobic** (Greek *hydros*, "water" and *phobos*, "fearing"). In contrast, polar molecules, which readily form hydrogen bonds with water, are said to be **hydrophilic** ("water-loving").

The tendency of nonpolar molecules to aggregate in water is known as **hydrophobic exclusion.** By forcing the hydrophobic portions of molecules together, water causes these molecules to assume particular shapes. Different molecular shapes have evolved by alteration of the location and strength of nonpolar regions. As you will see, much of the evolution of life reflects changes in molecular shape that can be induced in just this way.

Water molecules, which are very polar, cling to one another, so that it takes considerable energy to separate them. Water also clings to other polar molecules, causing them to be soluble in water solution, but water tends to exclude nonpolar molecules.

Water Ionizes

The covalent bonds within a water molecule sometimes break spontaneously. In pure water at 25°C, only 1 out of every 550 million water molecules undergoes this process. When it happens, one of the protons (hydrogen atom nuclei) dissociates from the molecule. Because the dissociated proton lacks the negatively charged electron it was sharing in the covalent bond with oxygen, its own positive charge is no longer counterbalanced, and it becomes a positively charged ion, H^+. The rest of the dissociated water molecule, which has retained the shared electron from the covalent bond, is negatively charged and forms a **hydroxide ion** (OH^-). This process of spontaneous ion formation is called **ionization:**

$$\underset{\text{water}}{H_2O} \rightarrow \underset{\text{hydroxide ion}}{OH^-} + \underset{\text{hydrogenion (proton)}}{H^+}$$

At 25°C, a liter of water contains $\frac{1}{10{,}000{,}000}$ (or 10^{-7}) mole of H^+ ions. (A **mole** is defined as the weight in grams that corresponds to the summed atomic masses of all of the atoms in a molecule. In the case ofH^+, the atomic mass is 1, and a mole of H^+ ions would weigh 1 gram. One mole of any substance always contains 6.02×10^{23} molecules of the substance.) Therefore, the **molar concentration** of hydrogen ions (represented as $[H^+]$) in pure water is 10^{-7} mole/liter. Actually, the hydrogen ion usually associates with another water molecule to form a hydronium (H_3O^+) ion.

pH

A more convenient way to express the hydrogen ion concentration of a solution is to use the **pH scale** (figure 2.18). This scale defines pH as the negative logarithm of the hydrogen ion concentration in the solution:

$$pH = -\log [H^+]$$

Because the logarithm of the hydrogen ion concentration is simply the exponent of the molar concentration of H^+, the pH equals the exponent times –1. Thus, pure water, with an $[H^+]$ of 10^{-7} mole/liter, has a pH of 7. Recall that for every H^+ ion formed when water dissociates, an OH^- ion is also formed, meaning that the dissociation of water produces H^+ and OH^- in equal amounts. Therefore, a pH value of 7 indicates neutrality—a balance between H^+ and OH^-—on the pH scale.

Note that the pH scale is *logarithmic*, which means that a difference of 1 on the scale represents a tenfold change in hydrogen ion concentration. This means that a solution with a pH of 4 has *10 times* the concentration of H^+ than is present in one with a pH of 5.

Acids. Any substance that dissociates in water to increase the concentration of H^+ ions is called an acid. Acidic solutions have pH values below 7. The stronger an acid is, the more H^+ ions it produces and the lower its pH. For example, hydrochloric acid (HCl), which is abundant in your stomach, ionizes completely in water. This means that 10^{-1} mole per liter of HCl will dissociate to form 10^{-1} mole per liter of H^+ ions, giving the solution a pH of 1. The pH of champagne, which bubbles because of the carbonic acid dissolved in it, is about 2.

Bases. A substance that combines with H^+ ions when dissolved in water is called a base. By combining with H^+ ions, a base lowers the H^+ ion concentration in the solution. Basic (or alkaline) solutions, therefore, have pH values above 7. Very strong bases, such as sodium hydroxide (NaOH), have pH values of 12 or more.

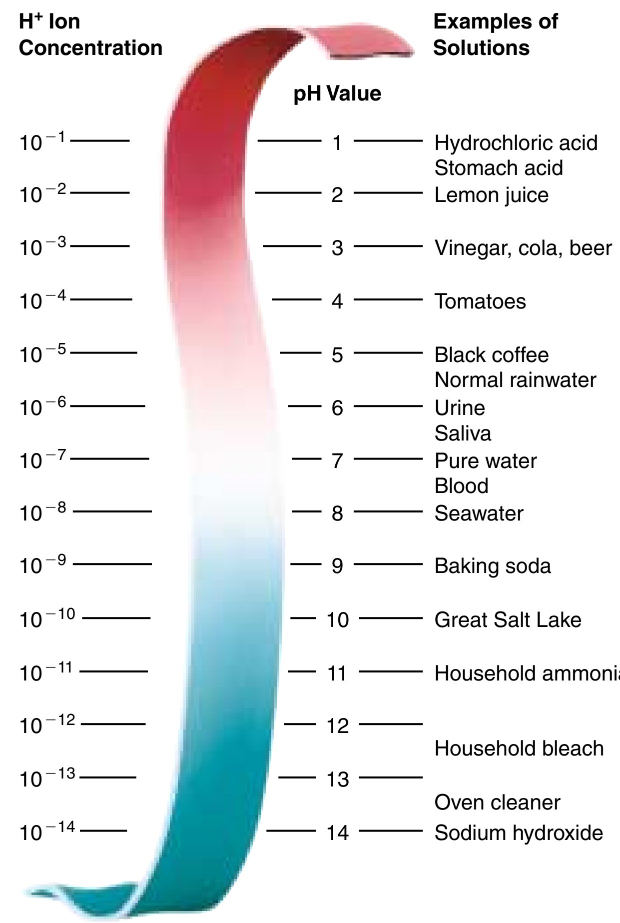


FIGURE 2.18
The pH scale. The pH value of a solution indicates its concentration of hydrogen ions. Solutions with a pH less than 7 are acidic, while those with a pH greater than 7 are basic. The scale is logarithmic, so that a pH change of 1 means a tenfold change in the concentration of hydrogen ions. Thus, lemon juice is 100 times more acidic than tomato juice, and seawater is 10 times more basic than pure water, which has a pH of 7.

Buffers

The pH inside almost all living cells, and in the fluid surrounding cells in multicellular organisms, is fairly close to 7. Most of the biological catalysts (enzymes) in living systems are extremely sensitive to pH; often even a small change in pH will alter their shape, thereby disrupting their activities and rendering them useless. For this reason it is important that a cell maintain a constant pH level.

Yet the chemical reactions of life constantly produce acids and bases within cells. Furthermore, many animals eat substances that are acidic or basic; cola, for example, is a strong (although dilute) acidic solution. Despite such variations in the concentrations of H^+ and OH^-, the pH of an organism is kept at a relatively constant level by buffers (figure 2.19).

A **buffer** is a substance that acts as a reservoir for hydrogen ions, donating them to the solution when their concentration falls and taking them from the solution when their concentration rises. What sort of substance will act in this way? Within organisms, most buffers consist of pairs of substances, one an acid and the other a base. The key buffer in human blood is an acid-base pair consisting of carbonic acid (acid) and bicarbonate (base). These two substances interact in a pair of reversible reactions. First, carbon dioxide (CO_2) and H_2O join to form carbonic acid (H_2CO_3), which in a second reaction dissociates to yield bicarbonate ion (HCO_3^-) and H^+ (figure 2.20). If some acid or other substance adds H^+ ions to the blood, the HCO_3^- ions act as a base and remove the excess H^+ ions by forming H_2CO_3. Similarly, if a basic substance removes H^+ ions from the blood, H_2CO_3 dissociates, releasing more H^+ ions into the blood. The forward and reverse reactions that interconvert H_2CO_3 and HCO_3^- thus stabilize the blood's pH.

The reaction of carbon dioxide and water to form carbonic acid is important because it permits carbon, essential to life, to enter water from the air. As we will discuss in chapter 4, biologists believe that life first evolved in the early oceans. These oceans were rich in carbon because of the reaction of carbon dioxide with water.

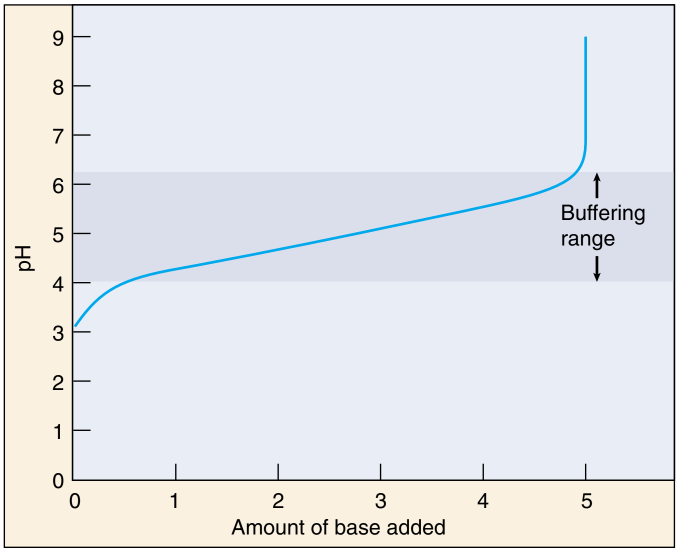


FIGURE 2.19
Buffers minimize changes in pH. Adding a base to a solution neutralizes some of the acid present, and so raises the pH. Thus, as the curve moves to the right, reflecting more and more base, it also rises to higher pH values. What a buffer does is to make the curve rise or fall very slowly over a portion of the pH scale, called the "buffering range" of that buffer.

In a condition called blood acidosis, human blood, which normally has a pH of about 7.4, drops 0.2 to 0.4 points on the pH scale. This condition is fatal if not treated immediately. The reverse condition, blood alkalosis, involves an increase in blood pH of a similar magnitude and is just as serious.

The pH of a solution is the negative logarithm of the H^+ ion concentration in the solution. Thus, low pH values indicate high H^+ concentrations (acidic solutions), and high pH values indicate low H^+ concentrations (basic solutions). Even small changes in pH can be harmful to life.

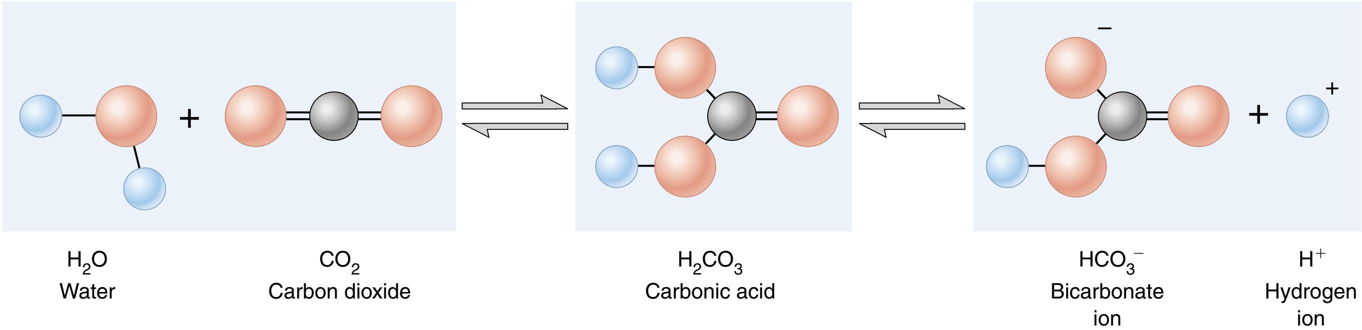


FIGURE 2.20
Buffer formation. Carbon dioxide and water combine chemically to form carbonic acid (H_2CO_3). The acid then dissociates in water, freeing H^+ ions. This reaction makes carbonated beverages acidic, and produced the carbon-rich early oceans that cradled life.

Chapter 2

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Summary — Questions — Media Resources

2.1 Atoms are nature's building material.

- The smallest stable particles of matter are protons, neutrons, and electrons, which associate to form atoms.
- The core, or nucleus, of an atom consists of protons and neutrons; the electrons orbit around the nucleus in a cloud. The farther an electron is from the nucleus, the faster it moves and the more energy it possesses.
- The chemical behavior of an atom is largely determined by the distribution of its electrons and in particular by the number of electrons in its outermost (highest) energy level. There is a strong tendency for atoms to have a completely filled outer level; electrons are lost, gained, or shared until this condition is reached.

1. An atom of nitrogen has 7 protons and 7 neutrons. What is its atomic number? What is its atomic mass? How many electrons does it have?

2. How do the isotopes of a single element differ from each other?

3. The half-life of radium-226 is 1620 years. If a sample of material contains 16 milligrams of radium-226, how much will it contain in 1620 years? How much will it contain in 3240 years? How long will it take for the sample to contain 1 milligram of radium-226?

- Atomic Structure

- Basic Chemistry
- Atoms

2.2 The atoms of living things are among the smallest.

- More than 95% of the weight of an organism consists of oxygen, hydrogen, carbon, and nitrogen, all of which form strong covalent bonds with one another.

4. What is the octet rule, and how does it affect the chemical behavior of atoms?

2.3 Chemical bonds hold molecules together.

- Ionic bonds form when electrons transfer from one atom to another, and the resulting oppositely charged ions attract one another.
- Covalent bonds form when two atoms share electrons. They are responsible for the formation of most biologically important molecules.

5. What is the difference between an ionic bond and a covalent bond? Give an example of each.

- Bonds
- Ionic Bonds

- Bonds

2.4 Water is the cradle of life.

- The chemistry of life is the chemistry of water (H_2O). The central oxygen atom in water attracts the electrons it shares with the two hydrogen atoms. This charge separation makes water a polar molecule.
- A hydrogen bond is formed between the partial positive charge of a hydrogen atom in one molecule and the partial negative charge of another atom, either in another molecule or in a different portion of the same molecule.
- Water is cohesive and adhesive, has a great capacity for storing heat, is a good solvent for other polar molecules, and tends to exclude nonpolar molecules.
- The H^+ concentration in a solution is expressed by the pH scale, in which pH equals the negative logarithm of the H^+ concentration.

6. What types of atoms participate in the formation of hydrogen bonds? How do hydrogen bonds contribute to water's high specific heat?

7. What types of molecules are hydrophobic? What types are hydrophilic? Why do these two types of molecules behave differently in water?

8. What is the pH of a solution that has a hydrogen ion concentration of 10^{-3} mole/liter? Would such a solution be acidic or basic?

- Water
- ph Scale