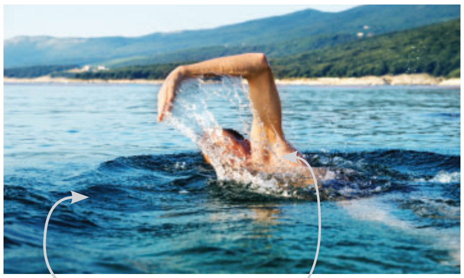
1.1 The Fundamental Chemistry of Life

As you learn about all living organisms, you will study everything from the microscopic subatomic level all the way to the macroscopic level of the organism. You will begin to understand how everything is uniquely adapted to provide a specific structure and function. You will soon see that the properties of life stem from a hierarchical arrangement of chemical parts.

Matter makes up everything in the universe, including all living organisms. Matter is composed of elements. An element is a pure substance that cannot be broken down into simpler substances using ordinary chemical or physical techniques. The smallest particle of an element is an atom. Elements differ from one another in their atomic structure. Atoms often bind to each other chemically in fixed numbers and ratios to form molecules. For example, the oxygen gas we breathe is formed from the chemical combination of two oxygen atoms. A chemical compound is a stable combination of different elements that are held together by chemical bonds. In the study of biology, you will encounter a great variety of chemical compounds.

The chemistry of life inside a cell is complex in terms of the sizes and shapes of molecules and their functions in chemical reactions. Even so, all organic (carbon-containing) compounds in living organisms are composed primarily of carbon, C; hydrogen, H; and oxygen, O. As well, they often include nitrogen, N. There are about 21 other elements found in living organisms, but these four elements make up 96 % of the weight of a living organism (**Table 1**). Most of the other 4 % is composed of only seven other elements: calcium, Ca; phosphorus, P; potassium, K; sulfur, S; sodium, Na; chlorine, Cl; and magnesium, Mg. These elements often occur as ions or in inorganic compounds within living organisms. The rest of the elements that are required by organisms are found in such small amounts (< 0.1 %) that they are called trace elements. Iodine, I, and iron, Fe, are examples of trace elements. A deficiency in any trace element can lead to health problems.

Table 1 Percentage Composition of Selected Elements in Living and Non-living Things



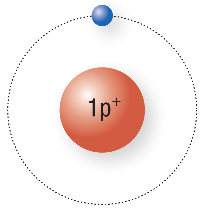
Seawater	%	Human	%	Pumpkin	%	Earth's crust	%
oxygen	88.3	oxygen	85.0	oxygen	65.0	oxygen	46.6
hydrogen	11.0	hydrogen	10.7	carbon	18.5	silicon	27.7
chlorine	1.9	carbon	3.3	hydrogen	9.5	aluminum	8.1
sodium	1.1	potassium	0.34	nitrogen	3.3	iron	5.0
magnesium	0.1	nitrogen	0.16	calcium	2.0	calcium	3.6
sulfur	0.09	phosphorus	0.05	phosphorus	1.1	sodium	2.8
potassium	0.04	calcium	0.02	potassium	0.35	potassium	2.6
calcium	0.04	magnesium	0.01	sulfur	0.25	magnesium	2.1
carbon	0.003	iron	0.008	sodium	0.15	other elements	1.5
silicon	0.0029	sodium	0.001	chlorine	0.15		
nitrogen	0.0015	zinc	0.0002	magnesium	0.05		
strontium	0.0008	copper	0.0001	iron	0.004		
				iodine	0.0004		

Atomic Structure

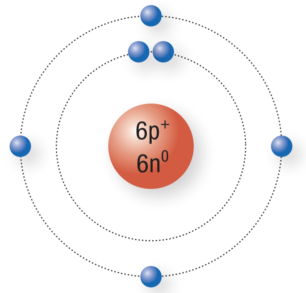
Elements consist of individual atoms: the smallest units that retain the chemical and physical properties of a particular element. An atom is composed of three subatomic particles: protons, neutrons, and electrons. The number of protons in an atom defines its elemental identity (**Figure 1**). Protons and neutrons are located in the nucleus. Electrons are located in the region surrounding the nucleus. Protons have a positive charge, neutrons have no charge, and electrons have a negative charge. An atom has no net charge because the number of protons is equal to the number of electrons.

The atomic number of an element is equal to the number of protons in the element. For example, carbon has an atomic number of 6, so all carbon atoms have 6 protons in their nucleus. A carbon atom also has 6 electrons. The number of neutrons in the nucleus can vary, as you will learn later in this section.

The mass number of an atom is the total number of protons and neutrons in the nucleus. Electrons are not included in this number because the mass of an electron is negligible compared with the mass of a proton or neutron. The mass of an atom is therefore determined by the number of protons and neutrons it contains. **Table 2** lists the atomic numbers and mass numbers of the most common elements in living organisms. The atomic symbol of an element is sometimes shown with the element's atomic number and mass number (**Figure 2**).



(a) hydrogen



(b) carbon

Figure 1 The basic atomic structure of (a) hydrogen and (b) carbon

mass number — $^{39}_{19}K$ — atomic number

Figure 2 Potassium, K, has 19 protons and 20 neutrons in its nucleus.

Table 2 Atomic Number and Mass Number of the Most Common Elements in Living Organisms

Element	Symbol	Atomic number	Mass number of the most common form
hydrogen	H	1	1
carbon	C	6	12
nitrogen	N	7	14
oxygen	O	8	16
sodium	Na	11	23
magnesium	Mg	12	24
phosphorus	P	15	31
sulfur	S	16	32
chlorine	Cl	17	35
potassium	K	19	39
calcium	Ca	20	40
iron	Fe	26	56
iodine	I	53	127

Isotopes and Radioisotopes

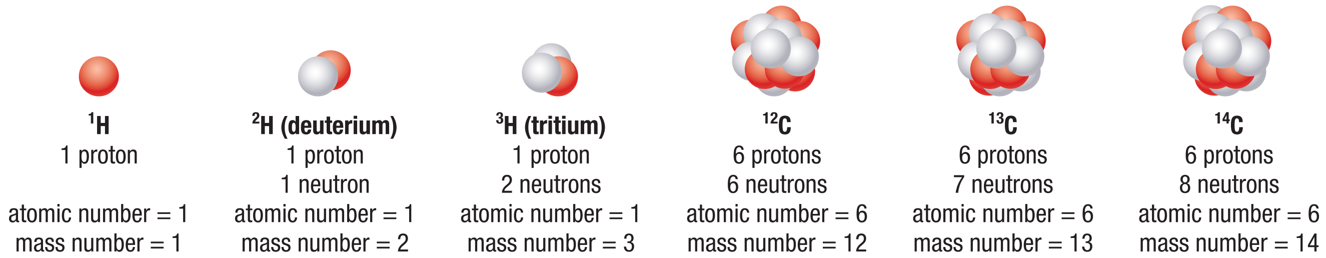
All atoms of the same element have the same number of protons, but they may have different numbers of neutrons in the nucleus. This means that atoms with the same atomic number can have different atomic masses. **Isotopes** are different forms of the same element, with different atomic masses. Because isotopes of the same element have the same number of protons and electrons, they behave exactly the same in a chemical reaction.

isotope a form of an element that differs in its number of neutrons

RADIOISOTOPES AND RADIOACTIVE TRACERS

radioisotope a radioactive isotope of an element

The nuclei of some isotopes of an element are unstable and tend to break down, or decay, giving off particles of matter that can be detected as radioactivity. The decay process transforms an unstable, radioactive isotope—called a **radioisotope**—into an atom of another element. For example, both hydrogen and carbon have three isotopes, which all behave the same in a chemical reaction (**Figure 3**). A carbon isotope with the atomic mass 12 (called carbon-12 or ^{12}C) has 6 neutrons, 6 electrons, and 6 protons. Carbon-12 accounts for 99 % of all the carbon in nature. The isotope ^{13}C has 7 neutrons, 6 electrons, and 6 protons. Like ^{12}C, ^{13}C is a stable isotope. A third isotope, ^{14}C, has 8 neutrons, 6 electrons, and 6 protons. ^{14}C is an unstable radioisotope of carbon. It decays, giving off particles and energy. As it decays, one neutron splits into a high-energy electron and a proton. The isotope then has 7 neutrons, 7 electrons, and 7 protons. This is characteristic of the most common form of the element nitrogen. Thus, the decay of ^{14}C transforms the carbon atom into ^{14}N, a nitrogen atom.



(a) nuclei of the different isotopes of hydrogen

(b) nuclei of the different isotopes of carbon

Figure 3 Comparison of the nuclei of different isotopes of (a) hydrogen and (b) carbon

Radioactive decay continues at a steady rate, with a constant proportion of radioisotope atoms breaking down during a given time interval. The rate of decay of a radioisotope is independent of chemical reactions or environmental conditions, such as temperature or pressure. The radiation from decaying isotopes may damage molecules in living cells, thus harming the organism. However, some radioisotopes are useful in geological and biological research because of their steady rate of decay. They provide scientists with information about the age of organic materials, rocks, and fossils. As well, radioisotopes have numerous medical applications and uses in instrumentation to elucidate the structures of unknown compounds.

Radioisotopes generally behave the same way in cells as non-radioactive isotopes of the same element. However, because radioisotopes give off a radioactive signal as they decay, they are easily detectable in a cell. Radioactive tracers are radioisotopes that are used to follow a specific chemical through a chemical reaction. Using the particles emitted by a radioisotope as a signal, scientists and doctors can trace the path of the radioisotope as it moves through the cells to different locations in the body. In this way, radioisotopes have found many applications in biological, chemical, and medical research. Scientist Melvin Calvin, a pioneer in the study of photosynthesis, used ^{14}C-labelled molecules to determine the sequence of reactions in photosynthesis. Radioisotopes are used to study many biochemical reactions and to perform basic techniques, such as DNA sequencing. Since most biological compounds contain carbon and hydrogen, ^{14}C and ^{3}H (tritium) are commonly used as tracers in biological research.

Radioisotopes are also used in the relatively new field of nuclear medicine to help with the diagnosis and treatment of diseases. For example, the thyroid gland produces hormones that affect growth and metabolism. This gland, located in front of the trachea, is the only organ of the body that actively absorbs iodine. If a patient's symptoms indicate an abnormal level of thyroid hormone output, the physician may inject a small amount of radioactive iodine-131 into the patient and then use a photographic device to scan the thyroid gland. The radioactivity produces an image that is similar to an X-ray, which helps to identify the possible causes of the condition.

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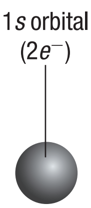
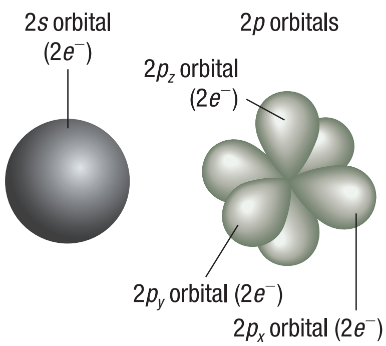
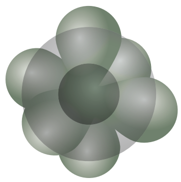
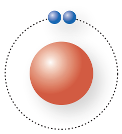
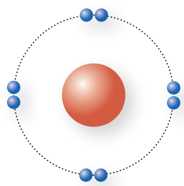
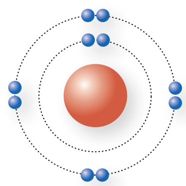
Electron Arrangements

The arrangement of electrons determines the chemical properties of an atom, because only electrons are usually directly involved in a chemical reaction. Recall that the number of electrons in an atom is equal to the number of protons in the nucleus. Since the electrons carry a negative charge that is exactly equal in magnitude but opposite to the positive charge of the protons in the nucleus, the atom is electrically neutral. Electrons move around the atomic nucleus in specific regions, called orbitals. An **orbital** is a region of space that one or two electrons can occupy. Even though one or two electrons may occupy a given orbital, the most stable and balanced condition occurs when the orbital contains two electrons. Electron orbitals are grouped into energy levels, which are sometimes called energy shells. These energy shells are numbered 1, 2, 3, and so on, indicating their relative distance from the nucleus. The lowest energy shell of an atom is closest to the nucleus. A maximum of two electrons can occupy the lowest energy shell. The second and third energy shells hold up to eight and 18 electrons, respectively.

orbital a region of space that is occupied by electrons located around the nucleus of an atom

The first electron shell, the 1*s* orbital, is a single spherical orbital (**Table 3**). Hydrogen, for example, has only a 1*s* electron orbital, containing one electron. Similarly, helium has only a 1*s* electron orbital, but it contains two electrons. Atoms with more than two electrons have higher energy levels. The shell at the second energy level consists of a 2*s* orbital and three 2*p* orbitals. The 2*s* orbital is spherical in shape. Each 2*p* orbital shape looks like two balloons tied together. The three 2*p* orbitals bisect in the centre at right angles to each other, giving the orbitals their overall shape. The orbital that is occupied by an electron is what determines the energy level of the electron. The farther away the electron is from the nucleus, the greater its energy. The balloon-like 2*p* orbitals contain electrons that are farther away from the nucleus than the electrons in the 2*s* orbital, and thus hold the electrons with a higher energy level. In large atoms, some higher-energy electrons occupy *d* and *f* orbitals, which have even more complex shapes.

Table 3 Types of Electron Orbitals

	1*s* orbital	2*s* and 2*p* orbital	Neon: 1*s*, 2*s*, 2*p*
Electron orbitals			
Electron-shell diagrams			

Valence electrons are the electrons in an atom's outermost energy shell, or valence shell. Atoms with an outermost energy shell that is not completely filled with electrons tend to be chemically reactive atoms. Atoms with a completely filled outermost energy level are chemically inactive, or inert. For example, hydrogen has a single electron in the 1*s* shell, its outermost shell. Hydrogen is a highly reactive element. The helium atom is chemically inert because it has two electrons in its 1*s* shell,

valence electron an electron in the outermost energy level or shell of an atom

Table 4 Valence Electron Shell of Common Biological Elements and Helium

Atomic number	Element	Lewis dot diagram	Valence shell
1	hydrogen	$\mathrm{H}\cdot$	$1s^1$
6	carbon	$\cdot\dot{\underset{\cdot}{\mathrm{C}}}\cdot$	$2s^2\,2p^2$
7	nitrogen	$\cdot\ddot{\underset{\cdot}{\mathrm{N}}}\cdot$	$2s^2\,2p^3$
8	oxygen	$\cdot\ddot{\underset{\cdot}{\mathrm{O}}}:$	$2s^2\,2p^4$
2	helium	$\mathrm{He}:$	$1s^2$

its outermost shell. (**Table 4**). Neon has all eight positions in its outer shell occupied. It, too, is an inert and highly stable element (Table 3).

Since an unfilled valence shell is less stable than a filled valence shell, atoms with an incomplete outer shell have a strong tendency to interact with other atoms. They may gain, lose, or share enough electrons to complete their outermost shell. All the elements in living organisms have unfilled outermost shells (Table 4) and can therefore participate in chemical reactions with other atoms.

Some atoms with outer energy shells that are almost empty or almost full gain or lose electrons and form stable charged ions. For example, sodium has two electrons in its first energy shell, eight in its second energy shell, and only one valence electron in its third energy shell. This outermost single electron is weakly held and readily lost to another atom. The sodium atom is left with a completely filled outer shell, becoming a positively charged ion, Na^+. In contrast, chlorine has seven electrons in its outermost shell and readily accepts another electron to fill its outer shell completely and become a negatively charged ion, Cl^-.

Atoms can also become more stable when they share electrons in such a way that their valence orbitals are filled. This sharing of electrons in the valence shells of atoms creates what are called hybridized electron orbitals. In a hybridized electron orbital, there is a direct overlap of the valence electron orbitals of the two atoms, so the orbital is a combination of two different orbitals. Sharing electrons is the most common way for atoms to bond and form biological molecules. The sharing of electrons by C, H, O, and N underlies the formation of countless chemical bonds that hold biological molecules together.

Chemical Bonds

Atoms of inert elements, such as helium, neon, and argon, occur naturally in single-atom form. Atoms of reactive elements, however, combine with each other to form compounds. These atoms form a stable attraction to one another called a chemical bond. Four types of chemical bonds are important in biological molecules: ionic bonds, covalent bonds, and two types of intermolecular forces. You will learn about these bonds in this section. You will also learn about polar molecules, which influence how biological molecules interact.

Ionic Bonds

ionic bond a bond that results from the attraction between two oppositely charged atoms or molecules

cation an ion that has a positive charge

anion an ion that has a negative charge

An **ionic bond** forms between atoms that have lost or gained electrons to become charged (**Figure 4**). Atoms that have lost or gained electrons are called ions. Ions of opposite charge—one positive and the other negative—are strongly attracted to one another, and this attraction leads to an ionic bond. A positively charged ion is called a **cation**, and a negatively charged ion is called an **anion**. Ions are very strongly attracted to water molecules. As a result, ionic compounds tend to dissociate and dissolve in water, forming hydrated ions.

Figure 4 Sodium, Na, and chlorine, Cl, atoms form an ionic bond to become NaCl. The sodium atom loses an electron, and the chlorine atom gains an electron. Sodium chloride crystals are cubic in shape.

Covalent Bonds

Covalent bonds form when atoms share one or more pairs of valence electrons. The formation of molecular hydrogen, H_2, from two hydrogen atoms is the simplest example of the sharing mechanism (**Figure 5**). In H_2, two hydrogen atoms share two electrons and, as a result, both fill their valence electron shell. The strength of a covalent bond depends on the **electronegativity** of the atoms involved in the sharing of the electron pair. Electronegativity is the measure of an atom's attraction for additional electrons.

In molecular diagrams, a dash or a pair of dots represents a single pair of shared electrons in a covalent bond. For example, a molecule of hydrogen is represented as H–H or H:H. The number of covalent bonds that an atom can form is usually equal to the number of additional electrons needed to fill its valence shell. The shared orbitals that form covalent bonds extend between atoms at specific angles and directions, giving covalently bound molecules distinct three-dimensional forms.

A carbon atom has four electrons in its valence shell and can therefore bond with four hydrogen atoms to form the compound methane, CH_4 (**Figure 6**). Each hydrogen atom shares its one electron with the carbon atom, completely filling the carbon atom's valence shell. The carbon atom shares its four valence electrons with the four hydrogen atoms, completely filling the hydrogen atoms' valence shells. The bonds between the carbon atom and hydrogen atoms in the compound methane are arranged in four symmetrical orbitals, so the hydrogen atoms are 109.5° from each other. The overall arrangement of this molecule is tetrahedral. The number and tetrahedral arrangement of the bonds around a carbon atom allow carbon atoms to link together in more complicated biological compounds (**Table 5**).

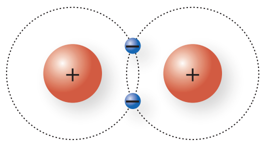


Figure 5 Covalent bond between two hydrogen atoms. The atoms share their electrons to fill the valence shell of each atom completely, creating a stable molecule of hydrogen, H_2.

electronegativity the measure of an atom's attraction to shared electrons

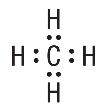
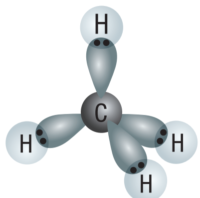
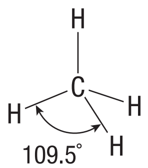
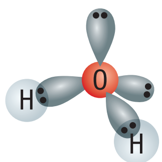
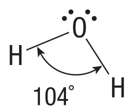


Figure 6 Lewis dot diagram showing the structure of methane, CH_4

Table 5 Orbital and VSEPR models for Methane, Carbon Dioxide, and Water

Name and formula (shape)	Orbital model	VSEPR model	Bonding valence electron pairs	Non-bonding valence electron pairs
methane, CH_4 (tetrahedral)			4	0
water, H_2O (angular)			2	2

The electrons attempt to move as far away from one another as possible, creating an angle between the bonds. You can predict the arrangement of the bond angles around an atom using the valence shell electron pair repulsion (VSEPR) theory developed by Canadian chemist Ronald J. Gillespie. The VSEPR theory states that because electrons are negatively charged, valence electron pairs repel one another and move as far apart from one another as possible. For example, in a water molecule, there are two valence electron spaces available on the oxygen atom (Table 5). The valence electrons from two hydrogen atoms fill these spaces, forming the molecule H_2O. The VSEPR theory also takes into account non-bonding electron pairs of an atom. Oxygen has two electron pairs or four non-bonding electrons in its valence shell. The negative charge of these non-bonding electrons repels the pairs that make up the O–H bonds, so the O–H bonds arrange themselves at an angle of 104° from each other.

Polar Molecules

polar covalent bond a bond between two atoms, made up of unequally shared electrons

Although all covalent bonds involve the sharing of valence electrons, they differ widely in the degree of sharing that takes place. The more electronegative an atom is, the more strongly it attracts electrons. Electronegativity is influenced by the atomic number and the distance between the valence electrons and the nucleus of an atom. Electronegativity increases as the distance between the electrons and the nucleus decreases. Oxygen, for example, has a very high electronegativity because two additional electrons can occupy the valence orbitals that are very close to the oxygen's nucleus. In contrast, hydrogen is less electronegative because, although an additional electron can also get close to its nucleus, the very small size of the nucleus provides a much weaker force of attraction.

The unequal sharing of electrons between two atoms with different electronegativity results in a **polar covalent bond**. The difference in electronegativity between atoms in common bonds is shown in **Table 6**. Remember that most biological compounds contain carbon, hydrogen, and oxygen, and many also contain nitrogen. Due to their relatively high electronegativities, both oxygen and nitrogen form polar bonds with atoms of most other elements (**Table 7**). Carbon and hydrogen differ in electronegativity by only 0.4. Such a small difference results in bonds that can be considered non-polar.

Table 6 Electronegativity Differences of Selected Bonds

Bond	Electronegativity difference (ΔEN)
H—H	0.0
C—H	0.4
C—N	0.5
N—H	0.8
C—O	0.9
O—H	1.2

Table 7 Common Polar and Non-polar Bonds in Biological Molecules

Polar bonds	Non-polar bonds
C—O	C—C
O—H	C=C
C=O	C—H
N—H	

The atom that attracts the valence electrons more strongly carries a partial negative charge, which results in the other atom carrying a partial positive charge. Because the atoms carry partial charges, the whole molecule may have a non-uniform charge distribution. This is the **polarity** of the molecule. For example, the oxygen atom in water forms a covalent bond with two hydrogen atoms. The oxygen atom attracts the shared electrons much more strongly than the hydrogen atoms (**Figure 7(a)**). Therefore, the bonds between oxygen and hydrogen are strongly polar. The more electronegative oxygen pulls the electrons closer, so the oxygen atom has a partial negative charge. Because the water molecule has an asymmetrical shape, with the oxygen atom positioned at a bend in the molecule, the partial charges on the atoms create an unequal distribution of charges on the molecule as a whole. In contrast, molecules such as carbon dioxide, CO_2, have an equal distribution of charges (**Figure 7(b)**). Carbon dioxide contains two double bonds with oxygen, making its three-dimensional arrangement appear linear. The bonds are polar, but the symmetry of the molecule results in a balanced distribution of the charges. Overall, the molecule is non-polar.

polarity partial positive or negative charge at the ends of a molecule

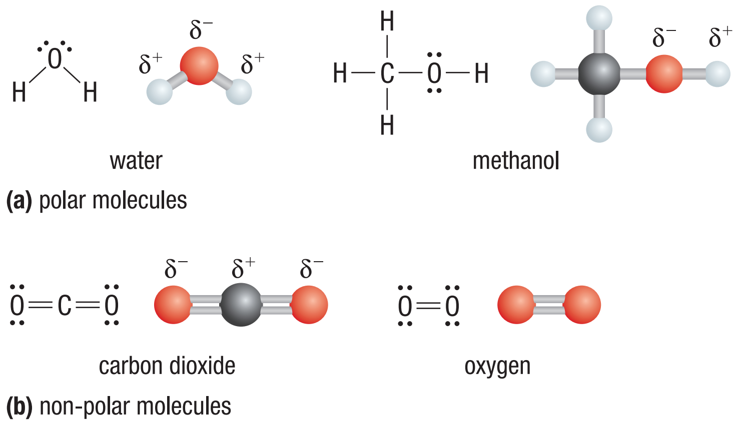


Figure 7 Comparison of (a) polar and (b) non-polar biological molecules. The lowercase Greek letter delta (δ) is used to indicate a partial charge. The symbol δ^+ indicates partially positive atoms, and the symbol δ^- indicates partially negative atoms.

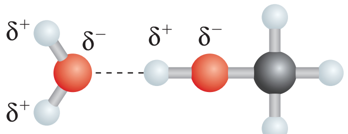


Figure 8 Methanol, CH_3OH, is pulled toward water molecules. The partial positive charge on the hydrogen atom in CH_3OH is pulled by the partial negative charge in the oxygen atom in water.

Polar molecules attract and align themselves to other polar molecules and tend to be soluble in water. Polar molecules, including water molecules, tend to exclude non-polar molecules, such as oils and fats. As a result, non-polar molecules have very low solubility in polar liquids. **Figure 8** shows how a molecule of methanol will align itself with a water molecule. The partial negative charge of water's oxygen atom is attracted to the end of the methanol molecule that has a partial positive charge.

Intermolecular Forces

In addition to the intramolecular forces that exist within a molecule, there are forces of attraction between molecules, or **intermolecular forces**. Intermolecular forces are known as **van der Waals forces**. They are extremely important because they influence the physical properties, such as the solubility, melting point, and brittleness, of a substance. Van der Waals forces act between similar molecules, as well as between different types of molecules. The strength of these forces is dependent on the size, shape, and polarity of molecules.

intermolecular force the force of attraction between two molecules

van der Waals forces very weak attractions between two molecules, or parts of molecules, when they are close together

HYDROGEN BONDS

A hydrogen atom that is covalently bonded to a strongly electronegative atom in one molecule, such as oxygen or nitrogen, can become attracted to a strongly electronegative atom in a different molecule. The resulting attractive force is called a **hydrogen bond** (**Figure 9**). Hydrogen bonding is evident in the attraction of the hydrogen atom in CH_3OH and the partially negative oxygen atom in H_2O (Figure 8). Hydrogen bonds may form between atoms in the same or different molecules.

hydrogen bond the attractive force between a partially positively charged hydrogen atom and a partially negatively charged atom in another molecule

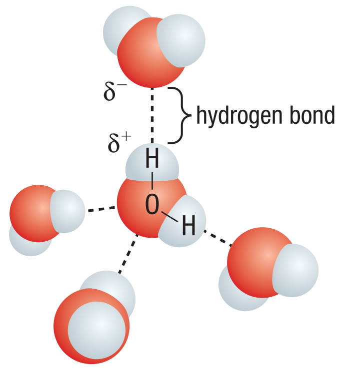


Figure 9 Each water molecule forms hydrogen bonds with up to four neighbouring molecules.

Hydrogen bonds are the strongest and most biologically significant form of van der Waals forces. Individual hydrogen bonds are weak compared with ionic and covalent bonds, but they can be very significant when they occur in large numbers. Although most of the strongest bonds in living organisms are covalent, weaker hydrogen bonds are crucial to the function of cells and cellular processes. The large size of biological molecules offers many opportunities for hydrogen bonding, both within and between molecules and with surrounding water molecules. A large number of hydrogen bonds are collectively strong, and they lend stability to the three-dimensional structure of large molecules, such as proteins.

The hydrogen bonds that exist between water molecules are responsible for many of the properties that make water a uniquely important molecule for all living organisms. Some of these properties include a very high heat capacity; high melting and boiling points; and cohesion, adhesion, and surface tension. Trees depend on cohesion to help transport water through xylem tissue up from their roots. The adhesion of water to the xylem cell walls of a plant helps to counteract the downward pull of gravity. Hydrogen bonds also give water an unusually high surface tension, causing it to behave as if it were coated with an invisible film (**Figure 10**).

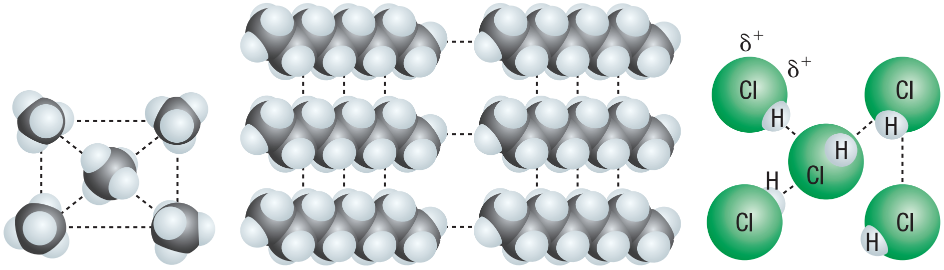
Figure 10 The hydrogen bonds in water give it an unusually high surface tension.

The weaker attractive force of hydrogen bonds makes them much easier to break than covalent or ionic bonds, especially when there is an increase in temperature, which increases the movement of the molecules. Hydrogen bonds begin to break extensively as temperatures rise above 45 °C. At 100 °C, hydrogen bonds in water are rapidly overcome.

OTHER VAN DER WAALS FORCES

Other van der Waal forces are even weaker and result from the momentary attractions of the electrons of one molecule to the nuclei of another molecule. These forces develop between all molecules, but they are only significant where other, stronger bonds are not prominent. Such is the case between non-polar molecules or between regions of slightly positive and slightly negative charges within a single molecule. Although an individual bond that has formed as a result of van der Waals forces is weak and transient, the formation of many bonds of this type can stabilize the shape of a large molecule, such as a fat.

The size and shape of a molecule influences the number and total strength of van der Waals forces of attraction. A larger molecule has larger forces of attraction. For example, small methane molecules, CH_4, are gases at room temperature because the very weak van der Waals forces are unable to hold the molecules together. In contrast, much larger non-polar octane molecules, C_8H_{18}, are liquid at room temperature because of the cumulative effect of many more van der Waals forces between the larger molecules (**Figure 11(a)** and **(b)**, next page). Also, linear molecules can align more easily with other molecules, and therefore the van der Waal forces are stronger. Polar molecules, such as hydrogen chloride, HCl, can also experience an attractive force between the positive and negative ends of the interacting molecules (**Figure 11(c)**). These influences of size and shape also apply to interactions that involve hydrogen bonds. The effect is very significant in long linear cellulose molecules, which have numerous OH functional groups and are able to form very strong solid fibres. Globular shaped molecules such as starches have fewer accessible atoms for van der Waals forces and therefore tend to form less rigid solids.



(a) methane, $CH_4(g)$ **(b)** octane, $C_8H_{18}(l)$ **(c)** hydrogen chloride, HCl

Figure 11 Non-polar molecules, such as (a) methane and (b) octane, can experience small attractive van der Waals forces, which hold the molecules together. (c) Polar molecules, such as hydrogen chloride, also experience van der Waals forces between their positively and negatively charged regions.

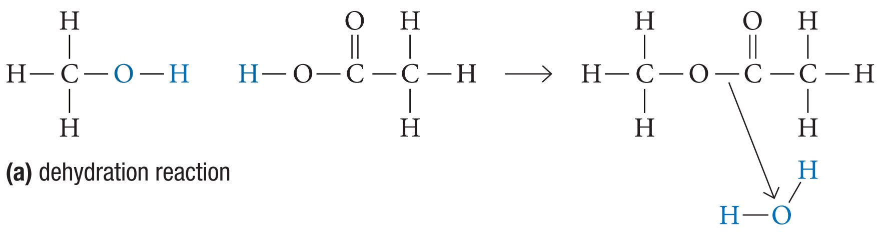
Chemical Reactions

Thousands of different chemical reactions occur inside cells, but they all have one thing in common. All chemical reactions involve the breaking and formation of chemical bonds, thereby changing the arrangements of atoms and ions. In the simple reaction between hydrogen gas, H_2, and oxygen gas, O_2, bonds within the molecules are broken and then new bonds are formed between oxygen and hydrogen to form water.

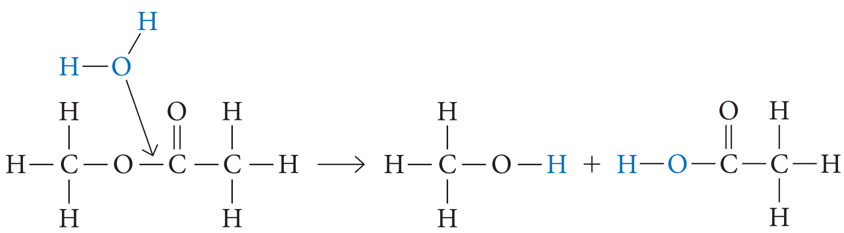
$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

There are four major types of chemical reactions that are common in biological processes: dehydration, hydrolysis, neutralization, and redox reactions. **Dehydration reactions** (also called condensation reactions) consist of the removal of an —OH and an —H from two reactant molecules. The —OH and —H form a water molecule, while the two reactant molecules are joined together (**Figure 12(a)**). Dehydration reactions are the most common method used by cells to join smaller molecules and assemble extremely large macromolecules, such as complex carbohydrates and proteins.

dehydration reaction a chemical reaction in which subunits of a larger molecule are joined by the removal of water; also called a condensation reaction



(a) dehydration reaction



(b) hydrolysis reaction

Figure 12 (a) Dehydration is the removal of –OH and –H from two reactant molecules to create water and a new bond. (b) Hydrolysis is the cleaving of a bond by the addition of water (as –OH and –H), splitting a larger molecule.

Hydrolysis reactions are the reverse of dehydration reactions. Water acts as a reactant to split or "lyse" a larger molecule. In living organisms, hydrolysis breaks down large molecules into smaller subunits. A bond in the reactant molecule is broken, and the –OH and –H from a split water are attached, resulting in two products (**Figure 12(b)**).

Neutralization reactions occur between acids and bases to produce salts. Water is also often produced in these reactions. An example of a neutralization reaction is the reaction between hydrochloric acid and sodium hydroxide. The products of these two reactants are water and sodium chloride (common table salt).

Redox reactions are the fourth type of chemical reaction. During a redox reaction (named for "reduction" and "oxidation"), electrons are lost from one atom and gained by another atom.

oxidation

$$Xe^- + Y \longrightarrow X + Ye^-$$

reduction

As you will learn, entire atoms may be transferred during a redox reaction. The term **oxidation** refers to the loss of electrons. The result is an oxidized molecule or atom. The oxidation of one molecule is always linked to the reduction of another molecule. The term **reduction** refers to the gain of electrons. In a redox reaction, the oxidizing agent is the molecule or atom being reduced. The reducing agent is the molecule or atom being oxidized. Note that the term "oxidation" also refers to the transfer of entire hydrogen atoms (and their electrons) from less electronegative atoms to more electronegative atoms, as shown in the combustion of methane below.

oxidation

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + \text{energy} + 2\,H_2O$$

reduction

During oxidation reactions in biological systems, the electrons involved are more strongly attracted to the oxidizing agent. For example, during the oxidation of methane, CH_4, the electrons in the original C–H bonds are strongly attracted to oxygen. As the reaction proceeds, the weaker forces of attraction between the C and H atoms are overcome and the atoms separate. They are then pulled toward and form strong bonds with the oxygen atoms. The products now contain much stronger C=O bonds and O–H bonds, compared with the C–H bonds in methane. Thus, redox reactions involve electrons moving from where they are weakly held to where they are more strongly held. As you will learn in Chapter 3, redox reactions are responsible for most of the energy transfer within cells.

hydrolysis reaction a chemical reaction in which water is used as a reactant to split a larger molecule into smaller subunits

neutralization reaction a reaction in which an acid and a base combine to create a salt and water

redox reaction an electron transfer reaction

oxidation a reaction in which a molecule loses electrons

reduction a reaction in which a molecule gains electrons

LEARNING **TIP**

Redox Reactions

A helpful pneumonic for remembering the roles of oxidation and reduction is "LEO the lion says GER":

Loss of Electrons = Oxidation

Gain of Electrons = Reduction

1.1 Review

Summary

- Isotopes are different forms of the same element, with different numbers of neutrons. A radioisotope is an unstable isotope that decays to release particles.
- Valence electrons are the electrons in the outermost electron shell of an atom.
- There are four types of chemical bonds in biochemistry: ionic, covalent, and hydrogen bonds, and weak van der Waals forces.
- Electronegativity is a measure of an atom's attraction for electrons. Differences in electronegativity result in bond polarity.
- Intermolecular forces are attractive forces between molecules.
- Dehydration is the removal of –OH and –H from two reactant molecules to form a larger molecule and water.
- Hydrolysis occurs when a bond in a large molecule is broken, and water is added to the resulting subunits.
- Oxidation is the loss of electrons, and reduction is the gain of electrons. The oxidation of one molecule or atom is always linked to the reduction of another molecule or atom. This is called a redox reaction.

Questions

1. One atom has 6 protons and a mass number of 13. Another atom has 6 protons and a mass number of 15. K/U
   (a) Identify each of the atoms.
   (b) Explain why there is a difference in the mass numbers.
2. (a) List the three common isotopes of hydrogen.
   (b) What are radioisotopes?
   (c) How are radioisotopes used in scientific research and medicine? K/U
3. An atom has eight electrons, and six of these electrons are valence electrons. K/U C
   (a) Draw an electron shell model for this atom. Which shells are occupied?
   (b) What is the name of the element?
4. How do bonding arrangements in a molecule affect the shape of the molecule? K/U
5. Compare ionic bonds with covalent bonds. K/U T/I
6. How can the atomic composition and shape of a molecule affect its polarity? K/U
7. (a) What effect do the polarity, size, and shape of a molecule have on the physical properties of the molecule?
   (b) How do these factors influence intermolecular forces? K/U T/I
8. How do polar covalent bonds and non-polar covalent bonds differ? K/U
9. In a bond between nitrogen and hydrogen (N—H), which atom will the electrons be closer to? Explain your reasoning. T/I
10. Oxygen plays a major role in biological molecules. Explain how oxygen plays a role in polarity, bond shape, and redox reactions. K/U
11. (a) In what ways do hydrogen bonds produce attractive forces between molecules? Include a labelled diagram to illustrate your answer.
    (b) How do hydrogen bonds influence the physical properties of water? T/I C
12. Describe dehydration and hydrolysis. How are these two types of reactions related? Draw a labelled diagram to support your answer. K/U C
13. (a) Describe reduction and oxidation.
    (b) Can a reduction reaction occur independently of an oxidation reaction, or vice versa? Why or why not? K/U