Element	Earth's crust	Human	Pumpkin
C	–	18	3.3
H	–	10	10.7
O	46.6	65	85
N	–	3	0.16
P	–	1.1	0.05
S	–	0.25	–
Na	2.8	0.15	0.001
K	2.6	0.35	0.34
Ca	3.6	2	0.02
Mg	2.1	0.05	0.01
Fe	5.0	0.004	0.008
Al	8.1	–	–
Si	27.7	–	–
Cl	–	0.15	–
I	–	0.0004	–
Zn	–	–	0.0002
Cu	–	–	0.0001
other elements	1.5	–	0.00005

Proportions of **different elements** in the earth's crust, the human body and a pumpkin.

C H A P T E R

4

The Atom and Chemical Bonds

Living organisms are composed of lifeless molecules. When these molecules are isolated and examined individually, they conform to all the physical and chemical laws that describe the behaviour of inanimate matter. Yet living organisms possess extraordinary attributes not shown by any random collection of molecules. The composition of living matter is strikingly different from that of the inanimate world. **Antoine Laurent Lavoisier** noted the relative chemical simplicity of the "mineral world" and contrasted it with the complexity of "plant and animal worlds"; the later he knew were composed of compounds rich in the elements carbon, oxygen, nitrogen, and phosphorus.

ELEMENTAL COMPOSITION OF LIVING MATTER

With the exception of certain metal ions, Biochemistry is, for the most part, related to the chemistry of the elements of the second and third periods of the Periodic Table. On earth, all cells, regardless of their origin (animal, plant, or microbial), contain the same elements in approximately the same proportions (Table 4–1). Thus, of the 107 known elements, only about 20 are essential for terrestrial life. Perhaps there is some logical chemical explanation for their selection.

54   FUNDAMENTALS OF BIOCHEMISTRY

Table 4–1.   **The elemental composition of living cells**

Element	Composition by weight (%)	Element	Composition by weight (%)
O	65		
C	18	Cu, Zn, Se, Mo,	
H	10	F, Cl, I, Mn,	0.70
N	3	Co, Fe	
Ca	1.5		
P	1.0		
K	0.35		
S	0.25	Li, Sr, Al, Si,	
Na	0.15	Pb, V, As, Br	Traces*
Mg	0.05		

*   Variable occurrences in cells. No known function in most cases.

Six nonmetals (O, C, H, N, P, S), which contribute almost 98% of the total mass of cells, provide the structural elements of protoplasm. From them, the functional components of cells (such as walls, membranes, genes, enzymes etc) are formed. These 6 elements all occur in the first three periods of the Periodic Table (refer Table 4–2).

Table 4–2.   **The structural elements of protoplasm***

Period	Group							
	I	II	III	IV	V	VI	VII	VIII
1	(H)							He
2	Li	Be	B	(C)	(N)	(O)	F	Ne
3	Na	Mg	Al	Si	(P)	(S)	Cl	Ar

*The six common nonmetals have been shown encircled.

The relative abundance of these 6 lighter elements in the seas, crust and atmosphere of earth does not by itself explain their utilization for life. Aluminium, for example, is more abundant than carbon, but it performs no known function essential to life. By contrast, the intrinsic chemical properties of these elements suggest their unique suitability as building blocks for life. Desirable features of structural elements apparently are as follows (Rodwell, Victor W., 1979) :

(*a*)   Small atomic radius

(*b*)   The versatility conferred by the ability to form 1–, 2–, 3– and 4– electron bonds

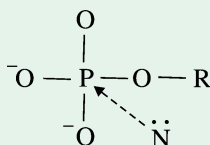(*c*)   The ability to form multiple bonds.

Carbohydrates and lipids contain C, H and O while proteins and nucleic acids contain in addition N and S, and N and P respectively. Na, K, Mg, Ca, and Cl are usually found in ionic forms whose concentrations must remain within narrow limits. The presence of Bo, Si, V, Mn, Fe, Co, Cu, Zn and Mb, though in ultratrace amounts, is essential for the functioning of enzymes. These elements are used as cofactors in the catalytic function of enzymes.

Small atoms form the tightest and most stable bonds — a distinct advantage for structural elements. **H, O, N and C are the smallest atoms capable of forming 1– , 2–, 3– and 4– electron bonds respectively.** Utilization of all possible types of electron bonds permits maximum versatility in molecular design. So also does the ability to form multiple bonds, a property confined almost entirely to P, S, and the elements of period 2. Advantages of carbon-based versus silicon-based life include the following :

1. Greater chemical stability of C—C bonds than the Si—Si bonds.
2. The ability of carbon, but not of silicon, to form multiple bonds; for example, the oxides of carbon are diffusible monomeric gases, whereas the oxides of silicon is a viscous polymer.
3. The stability of C—C bonds, but not of Si—Si bonds, to rupture by nucleophilic reagents (electron-rich elements or compounds) such as O_2, H_2O or NH_3.

Similar factors uniquely qualify P and S for utilization in energy transfer reactions. **Energy transfer is facilitated by bonds susceptible to nucleophilic attack.** A most common example is the nucleophilic attack of the 6-OH of glucose on the terminal P—O—P bond of ATP, forming ADP plus glucose-6-phosphate. P and

> **Nucleophilic attack** is the attack of an electron-rich centre upon an electron-deficient centre.

S resemble Si in that P—O—P or S—O—S bonds, like Si—Si bonds, are susceptible to nucleophilic rupture by virtue of their unoccupied third orbitals. However, unlike Si, P and S form multiple bonds (more versatile), a consequence of their smaller atomic diameters. Most energy transfer reactions in Biochemistry may be visualized as resulting from attack of a nucleophil (N) on the unoccupied third orbital of a phosphorus atom :

$$
\begin{array}{c}
O \\
| \\
{}^-O - P - O - R \\
| \\
{}^-O \qquad \ddot{N}
\end{array}
$$

The characteristic chemical and physical properties of the chemical elements of life are the same throughout the known universe. It, thus, seems probable that if life exists elsewhere, the same elements are employed for the same or similar reasons.

STRUCTURE OF AN ATOM

Each matter is composed of very small particles called **atoms**, which cannot be created, destroyed or subdivided. Atoms of the same element are similar to one another and equal in weight. Atoms of different elements have different properties and weight. Although, at one time, the atoms were conceived to be the smallest particles, subatomic particles were later recognized in due course of time. The 3 fundamental subatomic particles are: proton, neutron and electron. Besides these fundamental particles, about 35 other atomic particles are also known to exist. Many of them are, however, extremely unstable and they merely represent a bundle of energy. Some of the stable particles other than fundamental particles, are positron, photon, neutrino, graviton and antiproton. These particles are, however, of little importance in the study of Biochemistry because their

ERNEST RUTHERFORD

(LT, 1871–1937)

Ernest Rutherford grew up on a farm in New Zealand before entering Cambridge University to study under J.J. Thomson. Although he was a physicist, for his pioneering work in atomic structure Rutherford was awarded the Nobel Prize in Chemistry in 1908.

Contents
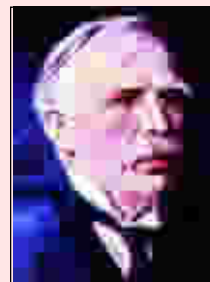
existence is rarely encountered in biological systems.

Regarding the arrangement of fundamental particles inside an atom, **Ernest Rutherford** (1911) proposed the most satisfactory model which is accepted even today with some modifications (Fig. 4–1). Accordingly, an atom is made up of a central *nucleus* containing positively-charged protons and neutral neutrons, surrounded by negatively-charged electrons which move around it (the nucleus) in discrete, successive, concentric volumes in space known as *orbits* or *shells*. The model is analogous to the sun's planetary system but differs from it in having the subatomic particles, protons and electrons as charged.
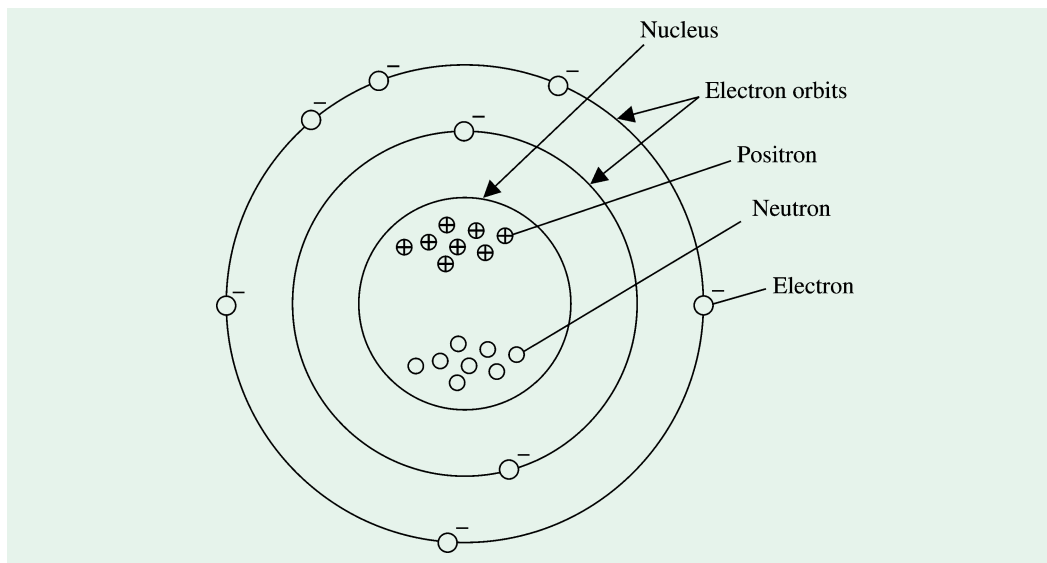
Fig. 4–1. Structure of an atom

The electron shells are numbered (from within) as 1, 2, 3, 4, 5, 6 and 7 and are designated by the letters *K, L, M, N, O, P, Q* respectively (Table 4–3). Each shell has a certain number of electrons and the maximum number of electrons for each shell is fixed. The maximum number is given by $2n^2$, where *n* is the serial number of the shell. Thus, the maximum number of electrons in *K, L, M, N, O, P, Q* shells will be 2, 8, 18, 32, 50, 72, 98 respectively. The maximum number of electrons in the outermost shell is 8 and in the penultimate shell is 18.

Table 4–3. **Distribution of electrons in shells and subshells**

Shell number (n)	1	2L		3			4 etc			
Shell name	K	L		M			N			
Subshell number	0	0	1	0	1	2	0	1	2	3
Subshell name	*s*	*s*	*p*	*s*	*p*	*d*	*s*	*p*	*d*	*f*
Number of electrons in subshells	2	2	6	2	6	10	2	6	10	14
Total number of electrons in a subshell		2		18			32			

The shells are subdivided into subshells. The number of subshells in a shell is equal to the number of the shell from within. K shell has one subshell called *s*; second L shell has two subshells *s* and *p*; the third M shell has three subshells *s, p* and *d* and fourth N subshell has four subshells *s, p, d* and *f*.

The subshells *s, p, d* and *f* can have a maximum of 2, 6, 10 and 14 electrons, respectively.

The position of electrons in the various shells and subshells are represented as follows. Major shells in which the electrons exist are indicated by the numbers 1, 2, 3 etc and the subshells designated by *s, p, d, f* etc. The superscript on *s, p, d* and *f* gives the number of electrons in the subshell. Thus, $1s^2$ indicates the presence of two electrons in the s subshell of the first major shell (*K*). Similarly, $4f^8$ indicates the presence of 8 electrons in the *f* subshell of fourth major shell (*N*).

Hydrogen atom is the simplest atom which consists of one proton, one neutron and one electron. In this atom, the lone electron is situated in the lone orbit or shell around the nucleus. In helium atom also, two electrons are situated in the single shell. However, in other elements the electrons are arranged in several shells. Thus, a **neon atom** has two shells of 2 and 8 electrons (a total of 10 electrons) and **argon atom** has three shells of 2, 8 and 8 electrons (a total of 18 electrons). These orbits or shells may never have more electrons than a certain maximum. When this maximum number of electrons is reached, the shell is said to be *saturated*. The elements whose outermost shells are saturated with electrons are relatively inert and do not participate in the chemical reactions under normal conditions. The elements in this category (He, Ne, Ar, Kr, Xe, Rd) are gases at normal temperature and they are called *noble gases* because of their inertness. The elements, whose atoms have electrons one more or one less (or even higher values) than the inert gas configuration in their outermost shell, are chemically active. This chemical activity may be interpreted as a tendency of those atoms to acquire noble configuration by accepting or losing electron(s). Acquiring the stable configuration results in lowering of energy.

The mass of an atom depends entirely upon its nucleus. A neutron has nearly the same mass as a proton. In absolute terms, each proton or neutron weighs 1.66043×10^{-24} g. The mass of an electron is negligible, about 1/1823 of the mass of a proton or neutron. The mass of a proton or neutron is called the *atomic mass unit* (*amu*). Thus, if a carbon atom contains 6 protons, 6 neutrons and 6 electrons, its amu will be equivalent to the proton + neutron, *i.e.*, 6 + 6 = 12. This amu is also the atomic weight of the element. Thus, *atomic weight* of an element may be defined as the combined weight of protons and neutrons, each taken as a unit weight. The number of protons on an atom is called the *atomic number* of the atom. Thus, the atomic number of the carbon is 6. Although the number of neutrons for a particular atom is fixed, it may vary, sometime giving rise to different species of the same atom. The different atomic species, having the same atomic number (as they have same proton number) but different atomic weight (as they have different neutron numbers) are called *isotopes*. Thus, a hydrogen atom with two neutrons (named as deuterium) is an isotope of normal hydrogen atom which contains only one neutron. The number of protons and electrons in both species of the hydrogen is same.

IONIZATION POTENTIAL

The amount of energy required to remove to most loosely-held electron from a normal (uncharged) gaseous atom is called ionization potential. It may be expresed as :

$$A(g) \xrightarrow{\text{EP}} A^+(g) + e^-$$
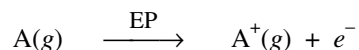
where, g denotes gaseous form of the atom A, and EP denotes the amount of energy required to remove one electron (*i.e.*, ionization potential).

The removal of electron generates a positive charge on the atom. The magnitude of ionization potential depends upon several factors. If the **radius of the atom** is large, the orbiting electrons are farther from the positively-charged nucleus and it is easier to remove them. Therefore, the ionization potential of the atoms with larger radii is smaller. The **magnitude of the positive charge on the nucleus** also affects the ionization potential. If the charge (positive) on the nucleus is increased, it becomes more difficult to remove electrons and therefore the ionization potential increases. The potential also depends upon the **number of electron orbits.** The inner shells or orbits of the electrons

act as a shield or screen between the nucleus and the outer shells. Therefore, each inner shell existing between the nucleus and the outermost shell produces a shielding effect, which decreases the force of attraction between the nucleus and the electrons of outer shells. Hence, *higher the number of inner shells, lesser the value of ionization potential*.

An atom can lose more than one electron during ionization. The amount of energy required to remove one electron from an uncharged atom is commonly called *first ionization potential*. The energy required for the removal of one electron from a charged ion, *i.e.*, a second electron from the original atom is called *second ionization potential*.

$$A^{+}(g) \xrightarrow{\text{II EP}} A^{2+}(g) + e^{-}$$
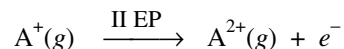
Similarly, *third ionization potential* represents the energy required to remove one electron from a 2^{+} ion, *i.e.*, third electron from the original atom.

$$A^{2+}(g) \xrightarrow{\text{III EP}} A^{3+}(g) + e^{-}$$
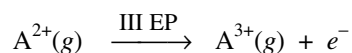
NATURE OF CHEMICAL BONDING

A perusal of the periodic table shows that besides the inert gases (He, Ne, Ar, Kr, Xe, Rd–all of them have 8 electrons in their outermost shells except He where outermost shell is also the first shell (*K*) and cannot have more than 2 electrons), there are 3 types of elements :

(*a*) those having 1 to 3 electrons in the outermost shell,

(*b*) those having 5 to 7 electrons in the outermost shell, and

(*c*) those having 4 electrons in the outermost shell.

Since the driving force for chemical combination is the tendency to acquire inert gas stable configuration, **first types of elements** will try to stabilize themselves by losing 1, 2 or 3 electrons because it is difficult to gain 7, 6 or 5 electrons than to lose 1, 2 or 3 electrons for obtaining stability. Similarly, **elements of the second type** will try to stabilize themselves by gaining 3, 2 or 1 electrons as losing 5, 6, or 7 electrons will be more difficult. Thus, the elements of the first and second type usually combine by either losing or gaining electrons.

The **elements of the third type,** of which *carbon* is representative, have to lose 4 electrons or gain 4 electrons in order to obtain inert gas configuration and if the element has balanced elecropositive and electronegative characters then either process (of losing or gaining electrons) becomes still more difficult. In such cases, the element achieves the stable configuration by sharing of these electrons with other atoms. Besides carbon, one more element is in a similar position and that is *hydrogen* which may attain stable configuration by either losing or gaining one electron. It can also do so by sharing of electrons with other elements.

TYPES OF CHEMICAL BONDING

Several atoms are assembled and held together to form thousands of molecules which participate in the building and function of physical and biological systems. A **bond** is any force which holds two atoms together. The formation of bond between two atoms is due to some redistribution or regrouping of electrons to form a more stable configuration. The regrouping of electrons in the combining atoms may take place in either of the 3 ways :

(*a*) by a *transfer* of one or more electrons from one atom to another— **electrovalent bonding**

(*b*) by a *sharing* of one or more pairs of electrons between the combining atoms— **covalent bonding**

(*c*) by a combination of the two processes of *transfer and sharing*— **coordinate bonding.**

Electrovalent or Polar or Ionic Bond

Ionic bond formation takes place between atoms of strongly electropositive and strongly electronegative elements. An element preceding an inert gas in the periodic table is strongly electronegative and the element immediately following the inert gas is strongly electropositive. For example, chlorine is electronegative while sodium is electropositive. According to W. Kossel (1916), a transfer of electron(s) takes place from the outermost shell of the electropositive atom to the outermost shell of the electronegative atom, resulting in the formation of stable positive and negative ions respectively which are held together by electrostatic forces of attraction to form a *molecule* or more precisely an *ion pair*. Thus,

$$A^e \; + \; B \longrightarrow \; A^+ \; + \; B^{-e} \quad \text{or} \quad A^+ \lll B^{-e}$$
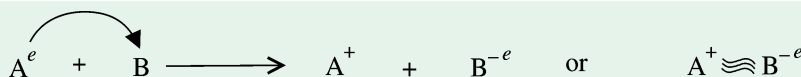
The atoms involved are electrically neutral before combining. The element A which has lost its electrons is known as electropositive whereas element B which has gained the electrons is termed electronegative element. The compound formed by electron transfer is termed as *electrovalent* by Langmuir (1919) because the resulting compound is electrolyte. It is also called *polar* since the molecule develops a positive and a negative pole. The electrovalent compounds always exist in ionic form, are hard and nonvolatile, have high melting and boiling points because of stronger nature of the bond and are soluble in polar solvents (such as water and alcohol) and because of the presence of ions conduct electricity in solution or in the fused state. The electrovalent compounds having identical electronic configuration exhibit the phenomenon of isomorphism. For example, NaF (2, 8), magnesium oxide (2, 8), calcium chloride (2, 8, 8) and potassium sulfide (2, 8, 8) are isomorphous.

Electrovalent linkage is common in inorganic compounds like NaCl (Fig. 4-2), Na_2O, CaO etc. In the case of NaCl,
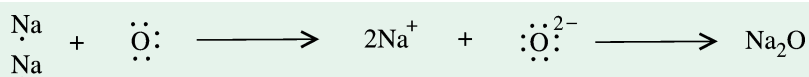

Fig. 4-2. Cubic crystals of sodium chloride. The bond involved is ionic bond

$$\text{Na} \quad + \quad \cdot\ddot{\underset{\cdot\cdot}{\text{Cl}}}\text{:} \longrightarrow \quad \text{Na}^+ \quad + \quad \text{:}\ddot{\underset{\cdot\cdot}{\text{Cl}}}\text{:}^-$$
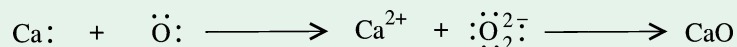
The electronic configuraion of sodium atom is $1s^2\,2s^2\,2p^6\,3s^1$ and that of chlorine is $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$. After transfer of one electron from sodium to chlorine, the sodium atom contains a configuration ($1s^2\,2s^2\,2p^6$) similar to neon, and chlorine attains a configuration ($1s^2\,2s^2\,2p^6\,3s^2\,3p^6$) which is similar to that of argon. The Na^+ can be said to be isoelectronic with neon while Cl^- is isoelectronic with argon. The bond between the ions results from the attraction between oppositely-charged ions. This linking of the ionized atoms is the formation of ionic bonds. The strength of the bond is roughly proportional to the product of the charges on ions and inversely proportional to the square of the distance between the effective centres of ions.

Ionic bonds may be formed by transfer of more than one electrons also. For example, in the case of Na_2O, one oxygen atom can form an ionic bond by accepting one electron from each of the two sodium atoms :

$$\begin{matrix} \overset{\cdot}{\text{Na}} \\ \\ \underset{\cdot}{\text{Na}} \end{matrix} \quad + \quad \ddot{\underset{\cdot\cdot}{\text{O}}}\text{:} \quad \longrightarrow \quad 2\text{Na}^+ \quad + \quad \text{:}\ddot{\underset{\cdot\cdot}{\text{O}}}\text{:}^{2-} \quad \longrightarrow \quad \text{Na}_2\text{O}$$

Oxygen can also form ionic bonds by accepting two electrons from a single atom of the element which has a low second ionization potential. For example, in the case of CaO, one calcium atom can link to one oxygen atom by transferring its two electrons :

$$\text{Ca:} \quad + \quad \ddot{\underset{\cdot\cdot}{\text{O}}}\text{:} \quad \longrightarrow \quad \text{Ca}^{2+} \quad + \quad \text{:}\ddot{\underset{\cdot\cdot_2}{\text{O}}}\text{:}^{2-} \quad \longrightarrow \quad \text{CaO}$$

The number of the charge (+ or −) or the number of electrons transferred during an ionic bond formation is called the *electrovalence number* or the *electrovalency* of the participating atoms. Thus, sodium has an electrovalency of 1 while calcium or oxygen has 2. The sign (+ or −) of this number depends upon whether the electron is donated or received. The element donating electron has a positive (+) electrovalency while that receiving electron has a negative (−) sign. Since the number of electrons transferred from an electropositive element to an elecronegative element during an ionic bond formation is equal to the number of electrons present in its outermost orbit, the electrovalency for such elements is the number of electrons present in outermost orbit. Similarly, the electrovalency of an electronegative element is the number of electrons which it can accept to achieve an inert gas configuration.

Covalent or Nonpolar Bond

Covalent bond formation, first suggested by **G.N. Lewis (1916)**, consists in sharing or holding a pair of electrons in partnership between two combining atoms, so that the pair counts towards the electronic grouping of both atoms. By this mechanism also, the stability akin to the inert gas is attained by each atom. For each pair of electrons to be shared between two atoms, each of the constituent atom contributes one electron :

$$\text{A} \cdot + \text{B} \cdot \longrightarrow \text{A : B}$$
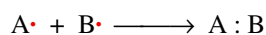
This type of linkage which is the result of equal contribution and equal sharing of electrons is known as covalent bond. The compound formed

GILBERT NEWTON LEWIS
(LT, 1875–1946)

Gilbert Newton Lewis was one of the foremost American chemists of the first half of the twentieth century. In addition to his pioneer work in describing chemical bonding through the symbolism named after him, Lewis was a driving force in the introduction of thermodynamics into the mainstream of chemistry, and he made important contributions to acid-base theory.
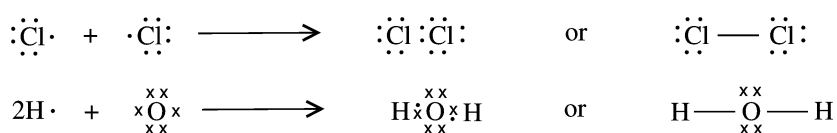
by electron sharing is termed as *covalent* or *nonpolar* by Langmuir (1919). The covalent compounds always exist in molecular form, are nonelectrolytes or nonionizable, soluble in organic solvents (such as benzene, ether, pyridine etc) and have low melting and boiling points because of weaker nature of the bond. They are usually liquids or gases and are generally soft, easily-fusible and volatile. They are nonconducting in the fused state or in solution. The covalent bond is rigid and directional and as such there is a possibility of position isomerism and stereoisomerism amongst these compounds.

Covalent linkage is common in organic compounds, although inorganic compounds also have it. In covalent compounds, one pair of shared electrons corresponds to a single bond, two pairs to double bond, three pairs of electrons or six shared electrons to triple bond. Some common examples from 3 categories are :
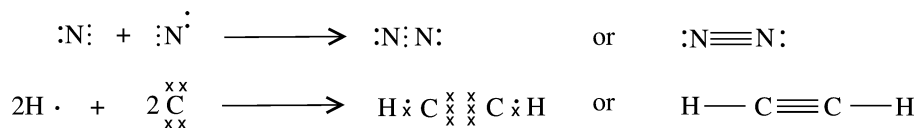
(*a*) One pair of electrons : single bond

$$:\!\ddot{C}l\cdot \;+\; \cdot\ddot{C}l\!: \;\longrightarrow\; :\!\ddot{C}l\!:\!\ddot{C}l\!: \qquad \text{or} \qquad :\!\ddot{C}l\!-\!\ddot{C}l\!:$$

$$2H\cdot \;+\; \overset{\times\times}{\underset{\times\times}{\times O\times}} \;\longrightarrow\; H\overset{\times\times}{\underset{\times\times}{\times O\times}}H \qquad \text{or} \qquad H\!-\!\overset{\times\times}{\underset{\times\times}{O}}\!-\!H$$

(*b*) Two pairs of electrons : double bond

$$\ddot{O}\!: \;+\; :\!\ddot{O} \;\longrightarrow\; \ddot{O}\!:\!\ddot{O} \qquad \text{or} \qquad \ddot{O}\!=\!\ddot{O}$$
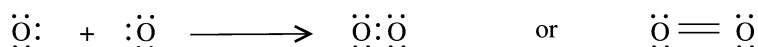
(*c*) Three pairs of electrons : triple bond

$$:\!\dot{N}\!: \;+\; :\!\dot{N}^{\cdot} \;\longrightarrow\; :\!N\!:\!N\!: \qquad \text{or} \qquad :\!N\!\equiv\!N\!:$$

$$2H\cdot \;+\; 2\overset{\times\times}{\underset{\times\times}{C}} \;\longrightarrow\; H\overset{\times}{\cdot}C\overset{\times}{\underset{\times}{}}\overset{\times}{\underset{\times}{}}C\overset{\times}{\cdot}H \quad \text{or} \quad H\!-\!C\!\equiv\!C\!-\!H$$

With the exception of hydrogen atom which by sharing one electron easily attains the helium configuration, it may be observed that only negative atoms can form covalent bonds by sharing electrons between them. Positive atoms contain an excess of electrons, which they have to lose to attain a stable inert gas structure; and therefore they form electrovalent bonds.

The force of bond formation in covalent bonds is the same as that in ionic bonds, *i.e.*, electrostatic attraction between the two atoms, although this force of attraction develops in a different manner. When two atoms, destined to link through a covalent bond, come within a definite range, the wave function of the electrons of two atoms overlap each other (Fig. 4–3). This overlapping causes the accumulation of the negative charge between the two atomic nuclei. The accumulated negative charge, in turn, attracts the nuclei (with positive charges) of the two atoms and, thus, they are held together.
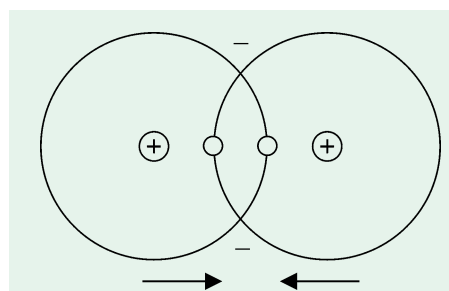
Fig. 4–3. Covalent sharing and bond formation between two atoms

A. Electron affinity or electronegativity

Electronegativity of an atom is a measure of its power to attract electrons that it is sharing in a covalent bond. Atoms of different elements have different electronegativities. Atoms with higher *atomic number* (magnitude of the positive charge on the nucleus) and less *atomic radius* (the distance of the valence electrons from the nucleus) have higher electronegativities. Electronegativity values of some elements commonly encountered are given in Table 4–4. When two elements, that differ in

Table 4–4. **The electronegativities of some elements**

Element	Electronegativity*	Element	Electronegativity*
F	4.0	Cu	1.9
O	3.5	Fe	1.8
Cl	3.0	Co	1.8
N	3.0	Ni	1.8
Br	2.8	Mo	1.8
S	2.5	Zn	1.6
C	2.5	Mn	1.5
I	2.5	Mg	1.2
Se	2.4	Ca	1.0
P	2.1	Li	1.0
H	2.1	Na	0.9
		K	0.8

*The higher the number, the more electronegative is the element.

electronegativity, form a covalent bond (for example, C and O), that bond is polarized; the shared electrons are more likely to be in the region of the more electronegative atom (O) than of the less electronegative (C). In the extreme cases of two atoms of very different electronegativity (Na and Cl, for example), one of the atoms actually gives up the electron(s) to the other atom, resulting in the formation of ions and ionic interactions such as those in solid NaCl.

B. Polar and nonpolar covalent bonds

Polar bond. Polar covalent bonds may be characterized as transition state between ionic and covalent bonds. In this case, there is neither the complete transfer of electrons from one atom to the other nor equal sharing. When a covalent bond is formed between two atoms with *different electronegativities*, the electrons involved in the bond are not shared equally. The atom with higher electronegativity pulls the bonding electrons closer to it. In other words, the electron density in the molecular orbital would be greater around the atom with higher electronegativity. The result of this displacement of the molecular orbital toward the more electronegative atom will acquire a small negative charge and the less electronegative atom will acquire a small positive charge. For example, in the case of **C—Cl bond**, chlorine is more electronegative than carbon. As such, the electron density in the molecular orbital would be higher around chlorine atom than around the carbon atom. Thus, the chlorine atom acquires a small negative charge and the carbon atom acquires a small positive charge. This may be indicated as in Fig. 4–4. Such a bond which appears to have a positive end and a negative end is said to be a **polar bond**.

$C^{\delta+}$ —— $Cl^{\delta-}$
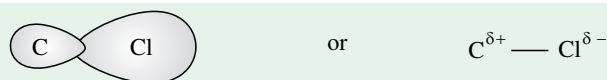
Fig. 4–4. **C—Cl bond is a polar bond.**
(The Greek letter delta, δ stands for *small*. The symbol δ^+ means a small positive charge and δ^- means a small negative charge.)

Water molecule is another good example of a polar bond. In this molecule, sharing of electrons between one oxygen and two hydrogen atoms is not equal. The electrons of hydrogen atoms are attracted more towards the nucleus of the oxygen atom. Consequently, oxygen develops a δ^- while

hydrogen develops a δ^+ charge. Yet another common example of a polar bond is H—Cl, which may be depicted as $H^{\delta+}$—$Cl^{\delta-}$ (Fig. 4—4). A polar molecule like that of HCl having two centres of charge is called a **dipole** and the molecule becomes **dipolar**. Although the ends of molecule may have partial charges, the molecule itself is neutral because of the equal number of protons and electrons.

Fig. 4–5. A dipole for HCl molecule d is the distance between the positive and negative centres and is called bond distance or bond length.

The degree of a polarity of a polar molecule (*i.e.*, a molecule having a polar covalent bond) is expressed in terms of its **dipole moment** (μ) which is equal to the product of electric charge (e) in **e.s.u.** and the distance (d) in Å between the positive and negative centres. Thus,

$$\mu = e \times d$$

As e is of the order of 10^{-10} e.s.u. and d of the order of 10^{-8} cm, the dipole moment (μ) is, therefore, of the order of 10^{-18} and this unit is known as **Debye** (D). Thus, $1D = 10^{-18}$ e.s.u. cm $= 10^{-10}$ e.s.u. Å. The μ is a vector quantity and is often indicated by an arrow parallel to the line joining the points of charge and pointing towards the negative end. Thus, $H^{\delta+}$—$Cl^{\delta-}$ can also be represented as H—Cl. Greater the value of the dipole moment of a molecule, greater is the polarity of the bond between the atoms.

As expected, polar compounds have properties intermediate between ionic and covalent compounds. Because of the dipole attraction between the atoms, polar compounds have higher melting and boiling points than nonpolar covalent compounds. These compounds are hard but not as hard as ionic compounds.

Nonpolar bond. When a covalent bond is formed between two atoms with *same electronegativities* (C—C, H—H and F—F), the electrons involved in the bond are shared equally. Electron density in the molecular orbitals bonding the two atoms together is same around each atom. There is no positive and negative end. Such a bond is said to be **nonpolar bond** (Fig. 4–6).
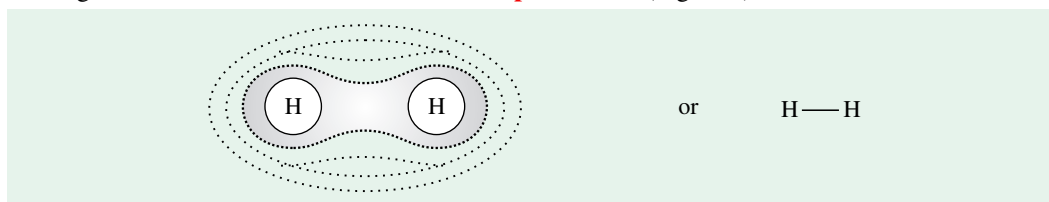
Fig. 4–6. H—H bond is a nonpolar bond.

C. Bond length or bond distance

The distance between two atomic nuclei in a covalent molecule is called the **bond distance** or **bond length.** The distance is such that the repulsion between atomic nuclei is balanced by the stability acquired through the overlapping of orbitals. The bond lengths are characteristic properties of a molecule and give information about its structure, properties etc. Bond lengths are measured by x-ray crystallography and by microwave spectroscopy. Bond lengths are generally affected by hybridization, electronegativity, steric condition etc. Some typical bond lengths are given in Table 4–5.

D. Bond angle

The 3 p orbitals of an atom are at the angle of 90° to each other. Hence, bonds formed by these should have an angle of 90°. Similarly, sp^3, sp^2 and sp orbitals in hybridized carbon atoms have an angle of 109.5°, 120° and 180° respectively. For example, in methane, where the four groups attached to sp^3 orbitals are identical, the angle is 109.5°. But if one of these groups is different then usually the angle is distorted. Thus, in bromopropane, C—C—Br bond angle is 114.2°.
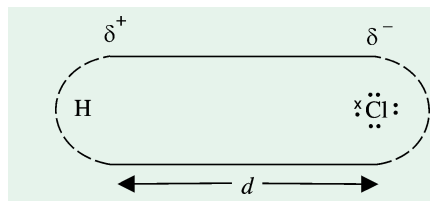
Table 4–5.	Bond lengths* of some common covalent bonds

Bond	Bond length (in Å)	Typical compound
C—C	1.54	Ethane, Propane
	1.48	Butadiene
	1.38	Butadiyne
C = C	1.34	Ethylene
	1.28	Butatriene
C ≡ C	1.20	Acetylene
C—H	1.11	Methane
	1.10	Benzene
	1.08	Acetylene
C—O	1.41	Ethanol
	1.34	Formic acid
C = O	1.20	Formaldehyde
	1.16	Carbon dioxide
C—N	1.47	Methylamine
	1.36	Formamide
C—Cl	1.78	Methyl chloride
C—Br	1.94	Methyl bromide
C—I	2.14	Methyl iodide
O—H	0.96	Methanol
N—H	1.01	Methyl amine

*Note that for most bonds, the value of bond length lies between 1 and 2 Å.

H—O—H bond angle in water and H—S—H bond angle in H_2S should be similar or 90° but the measured angle is 104.5° and 92.1°, respectively. Usually in strained molecules, the bond angles may be greatly distorted. An understanding of bond angle is essential for the study of the shape and geometry of the molecule. Bond angles in some molecules have been shown in Fig. 4–7.
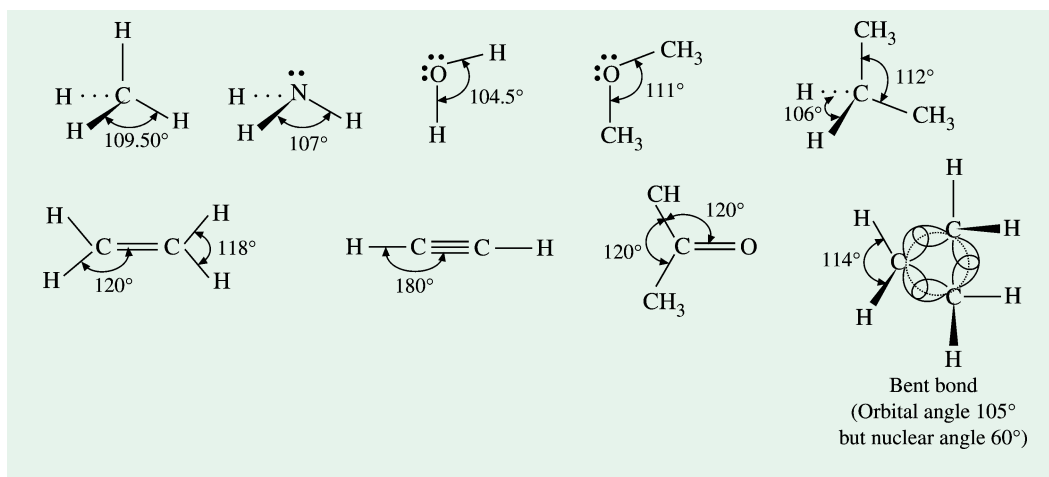
Fig. 4–7. A diagrammatic representation of bond angles in some molecules

E. Bond strength and bond energy

The energy required to break a bond into its constituents is referred to as bond energy and is a measure of **bond strength**. The greater the energy required to break a bond, the greater will be its strength. Thus, bond energies provide a useful picture of the strengths of various bonds. Bond energy depends upon the relative electronegativities of the elements involved. The number of electrons shared also influences bond strength: double bonds are stronger than single bonds, and triple bonds are stronger yet. Bond strengths are usually determined by quantitative measurements of heats of chemical reactions (calorimetry) and by spectroscopic methods. The strength of a bond is expressed as bond energy, in joules. Bond energy can be thought of as either the amount of energy required to break a bond or the amount of energy gained by the surroundings when two atoms form the bond. Some of the important bond energies are given in Table 4–6.

> **In biochemistry,** calories have often been used as units of energy– bond energy and free energy. The **joule** is the unit of energy in the International System of Units. For conversions, 1 cal is equal to 4.18 J.

Coordinate or Semipolar or Dative Bond

In a purely covalent bond, each of the electrons in a shared pair or duplet is contributed by each of the two combining atoms; in other words, we have a sharing of give-and-take character. Equal contribution and equal sharing of electrons is thus characteristic of polar bonding. Coordinate bond is also formed by mutual sharing of electrons but in this case the two electrons that are shared come from the same atom (Perkins, 1921). The shared pair of electrons is called *lone pair*. The atom which provides the pair of electrons is called the **donor** and the atom accepting this pair is called the **acceptor**. After the formation of the bond, the lone pair of electrons is held in common. This sort of bonding is called **coordinate** (Sidgwick) or **dative** (Menzies). In this mechanism, although the sharing is equitable, the contribution is one-sided, and therefore a slight polarity develops in the molecule. For this reason, this bond is also called **semipolar** (Sugden). This type of linkage is represented by an arrow, pointing away from the donor atom (or pointing toward the acceptor atom). Usually the donor is an atom which has already acquired stable electronic configuration and the acceptor is generally two short of the stable configuration.

Table 4–6.   **Strength of bonds common in biomolecules**

Type of bond	Bond dissociation energy (kJ/mol)	Type of bond	Bond dissociation energy (kJ/mol)
Single bonds		**Double bonds**	
O—H	461	C = O	712
H—H	435	C = N	615
P—O	419	C = C	611
C—H	414	P = O	502
N—H	389	**Triple bonds**	
C—O	352	N ≡ N	930
C—C	348	C ≡ N	816
S—H	339		
C—N	293	**Noncovalent bonds or interactions**	
C—S	260	Hydrogen bonds	
N—O	222	van der Waals interactions	4–0
S—S	214	Hydrophobic interactions	
		Ionic interactions	

$$:\overset{..}{\underset{..}{A}}: \; + \; \overset{..}{\underset{..}{B}}: \longrightarrow \; :\overset{..}{\underset{..}{A}}:\overset{..}{\underset{..}{B}}: \quad \text{or} \quad A \longrightarrow B \quad \text{or} \quad A^+ \longrightarrow B^-$$

Because of slight increase in electron density near about the acceptor atom, the acceptor acquires a small negative charge (δ^-) and simultaneously the donor atom acquires a small positive charge (δ^+) because of slight decrease in electron density around it. Formation of O_3, SO_2, SO_3, boron hydride–ammonium complex and hydronium ion involves the coordinate bond.

$$O_3: \quad :\overset{..}{O}::\overset{..}{O}: + \; \overset{..}{\underset{..}{O}}: \longrightarrow :\overset{..}{O}::\overset{..}{O}:\overset{..}{\underset{..}{O}}: \quad \text{or} \quad O = O^{\delta+} \longrightarrow O^{\delta-}$$

$$SO_2: \quad :\overset{..}{O}::\overset{..}{S}:\overset{..}{\underset{..}{O}}: \quad \text{or} \quad O = S^{\delta+} \longrightarrow O^{\delta-}$$

$$SO_3: \quad :\overset{..}{O}::\overset{..}{S}:\overset{..}{\underset{..}{O}}: \quad \text{or} \quad O = S^{2\delta+} \longrightarrow O^{\delta-}$$
$$\qquad \overset{..}{\underset{..}{:O:}} \qquad\qquad\qquad\qquad \downarrow$$
$$\qquad\qquad\qquad\qquad\qquad\qquad O^{\delta-}$$

Boron hydride-ammonium complex :

$$\overset{..}{\underset{..}{H}}\!:\!B \; + \; :\overset{..}{\underset{..}{N}}:H \longrightarrow H:\overset{..}{\underset{..}{B}}:\overset{..}{\underset{..}{N}}:H \quad \text{or}$$

(Boron hydride-ammonium complex structural diagrams)

$$\begin{array}{ccc} & H & H \\ & | & | \\ H\!-\!B & \longleftarrow & N\!-\!H \\ & | & | \\ & H & H \end{array}$$

Hydronium ion :

$$H\!-\!\overset{H}{\underset{..}{O}}: \; + \; H^+ \longrightarrow \left[H\!-\!\overset{H}{\underset{..}{O}} \rightarrow H^+ \right] \quad \text{or} \quad \left[H\!-\!\overset{H}{\underset{..}{O}} \rightarrow H \right]^+ \quad \text{or} \quad H_3O^+$$

Coordinate compounds exhibit characteristics similar to covalent compounds :

1. They do not ionize in water and are poor conductors of electricity.
2. They are very sparingly soluble in water but dissolve readily in organic slovents.
3. Since a coordinate linkage is semipolar, the coordinate compounds possess melting and boiling points which are higher than those of purely covalent compounds but lower than those of ionic compounds.
4. The coordinate linkage is also rigid and directional. Therefore, such compounds exhibit space isomerism.
5. The coordinate linkage is easily broken when donor and acceptor are molecules which are capable of independent existence.

Coordinate bonds are less common in organic compounds. However, *they are commonly found in complex compounds*. The two important biological molecules containing coordinate covalent bonds are **chlorophyll** (Fig. 4-8) and **heme** in which either Mg or Fe atom is linked coordinately

Fig. 4-8. Model of chlorophyll which contains coordinate covalent bond

to nitrogen atoms, respectively. In **vitamin B$_{12}$** also, a Co atom is linked to nitrogen atoms through coordinate bonds. The compound containing coordinate covalent bonds are called **coordinate complexes.**

Noncovalent Bonds or Interactions

In addition to covalent bonding, there are weaker forces of interaction that profoundly influence conformation of biomolecules and their function. These **noncovalent forces**, as they are called, play key roles in the faithful replication of DNA, the folding of proteins into intricate 3-dimensional forms, the specific recognition of substrates by enzymes, and the detection of signal molecules. *Indeed, all biological structures and processes depend on the interplay of noncovalent interactions as well as covalent ones*.

With respect to bonding, weak and strong are used to indicate the amount of energy in a bond. **Strong bonds** such as covalent bonds found in biomolecules require an average of 100 kilocalories/ mole or kcal mol^{-1}, to be cleaved and hence are stable and seldom break under physiological conditions. In contrast, **weak bonds,** such as hydrogen bonds, have energies of 2 to 7 kcal mol^{-1} and are easily broken. Weak bonds are transient; individually they form and break in small fractions of a second. The transient nature of noncovalent interactions confers flexibility on macromolecules, such as proteins and nucleic acids, that is critical to their function. Furthermore, the large number of noncovalent interactions in a single macromolecule makes it unlikely that at any given moment, all the interactions will be broken; thus macromolecular structures are stable over time. The four fundamental noncovalent bonds are : electrostatic (or ionic) bonds, hydrogen bonds, hydrophobic bonds and van der Waals forces or bonds. They differ in geometry, strength and specificity and are profoundly affected by the presence of water.

A. Electrostatic or ionic bonds

Ionic bonds are formed due to the attraction between atoms or groups, of opposite charges (+ and −). A charged group on a substrate can attract an oppositely-charged group on an enzyme. The force (*F*) of such an electrostatic attraction is given by Coulomb's law :

$$F \;=\; \frac{q_1\, q_2}{r^2\, D}$$

where, q_1 and q_2 = charges of the two atoms or groups,

r = distance between the two atoms or groups, and

D = dielectric constant of the medium.

The attraction is strongest in a vaccum where *D* is 1 and is weakest in a medium such as water where *D* is 80. The distance between oppositely-charged atoms in an optimal electrostatic attraction is about 2.8 Å. The average bond energy of ionic bonds in aqueous solution is about 5 kcal mol^{-1}. This kind of attraction is also called *saline bond, salt linkage, salt bridge* or *ion pair*.

Ionic bonding occurs in crystals and salts that are ionized when dissolved in water. For example, NaCl is a salt composed of Na^{+} (cation) and Cl^{-} (anion). Common examples of the salts of biomolecules are sodium acetate, potassium pyruvate and ethanolamine chloride (Fig. 4–7). This type of interaction even permits the bonding between two different molecules in heteroproteins (for example, in nucleoproteins, between the negatively-charged nucleic acid and the positively-charged basic proteins, esp., the histones).
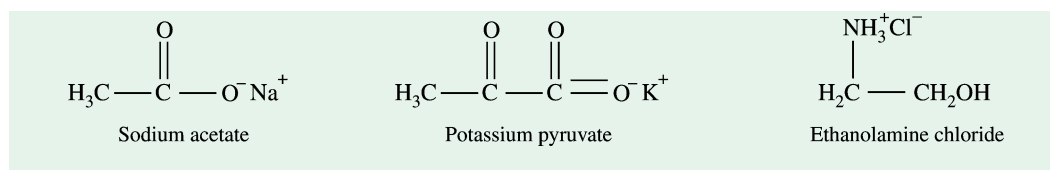

Fig. 4–7.   **Some salts of biomolecules**

B. Hydrogen bonds

Hydrogen bonds can be formed between uncharged molecules as well as charged ones. *In a hydrogen bond, a hydrogen atom is shared by two other atoms* (Fig. 4–9). The atom to which the hydrogen is more tightly linked is called the *hydrogen donor*, whereas the other atom is called as the *hydrogen acceptor*. The acceptor has a partial negative charge that attracts the hydrogen atom. In fact, a hydrogen bond can be considered as an intermediate in the transfer of a proton from an acid to a base.
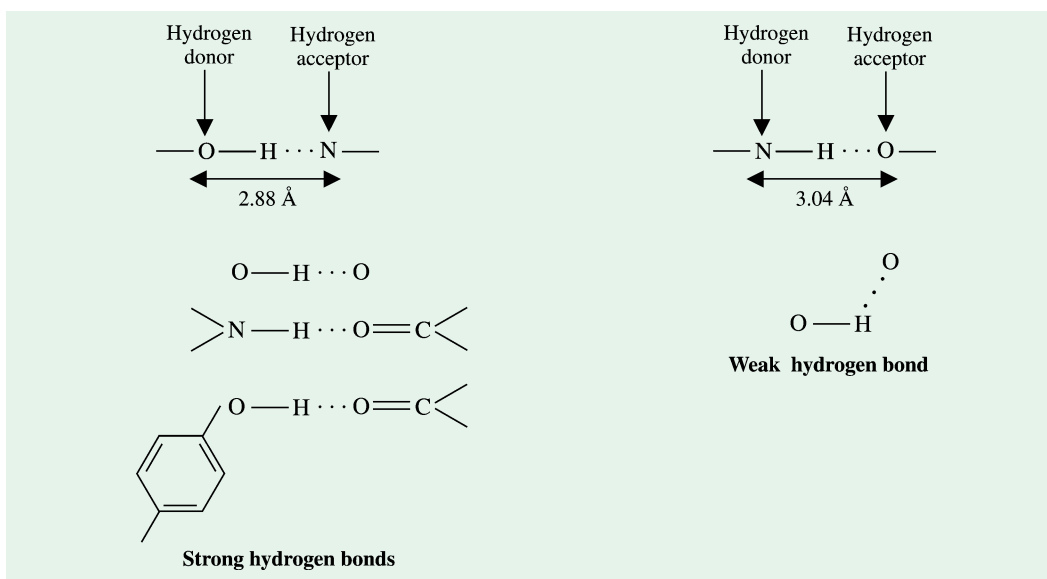

Fig. 4–9. **Structure and types of hydrogen bonds**

The donor in a hydrogen bond in biological systems is an oxygen or nitrogen atom that has a covalently-attached hydrogen atom. The acceptor is either oxygen or nitrogen. The bond lengths of different types of hydrogen bonds along with their bond energies are given in Table 4–7. A perusal of the table reveals that the length of a hydrogen bond is intermediate between that of a covalent bond and a van der Waals bond. The bond energies of hydrogen bonds range between 2 and 7 kcal/mol. Hydrogen bonds are stronger than the van der Waals but much weaker than covalent bonds. The strongest hydrogen bonds are those in which the donor, hydrogen and acceptor atoms are colinear.

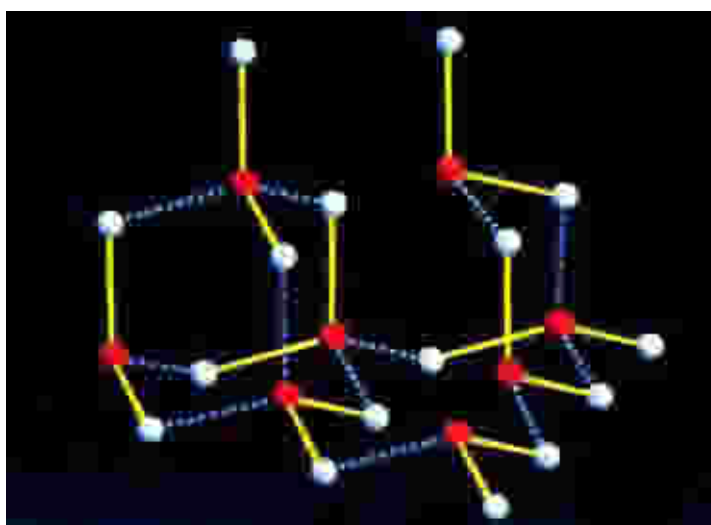Hydrogen bonds in biomolecules (water, for example; Fig. 4-10) are also

Fig. 4-10. **Hydrogen bonds seen between water molecules**

more specific than other weak bonds because they require particular complementary groups that donate or accept hydrogen. *An important feature of hydrogen bonds is that they are highly directional.* Hydrogen bonding is of 2 types : intramolecular (within a molecule) and intermolecular (between two molecules). Both types are common to many macromolecules such as proteins and DNA molecule. **Intramolecular bonding** gives rise to chelation, *i.e.*, ring formation and this normally occurs only with the formation of 5-, 6- or 7- membered rings. **Intermolecular bonding,** however, gives rise to association, thereby raising the boiling point; it also raises the surface tension and the viscosity, but lowers the dielectric constant. It may exist in compounds in the liquid or solid state and its formation is very much affected by the shape of the molecular, *i.e.*, by the steric factor.

Table 4–7. **Bond lengths and bond energies of some hydrogen bonds**

| Bond	Length (Å)	Energy (kcal mol^{-1})
$—O—H \cdots N\langle$	2.88	7
$—O—H \cdots O=$	2.70	6
$—O—H \cdots O\langle$		6
$\rangle N—H \cdots N\langle$	3.10	2 – 4
$\rangle N—H \cdots O=$	3.04	2 – 3

C. Hydrophobic or nonpolar interactions

The essentiality of hydrophilic (water-loving) properties of biomolecules is obvious. That hydrophobic (water-fearing) characteristics can be valuable may not be as readily apparent . As in the formation of micelles, hydrophobic groups of macromolecules, if in proper spatial relation, will interact (*not bond*) to the exclusion of solvent molecules (water) and thereby reside in a hydrophobic environment. On the contrary, hydrophilic groups usually remain exposed to the aqueous environment where they interact with water molecules. Hydrophobic interactions are a major driving force in the folding of macromolecules, the binding of substrates to enzymes and the formation of membranes that define the boundaries of cells and their internal compartments.

In macromolecules such as proteins, the acceptance or rejection by the aqueous environment of the hydrophilic and hydrophobic moieties respectively exerts a dominant influence on their final conformation. Here the nonpolar side chains of neutral amino acids tend to be closely associated with one another. *The relationship is nonstoichiometric*; hence no true bond may be said to exist. *This clustering together of nonpolar molecules or groups in water is called hydrophobic interaction*. The familiar sight of dispersed oil droplets coming together in water to form a single large oil drop is an analogous process.

To understand the basis of hydrophilic attractions, let us take an example wherein a single nonpolar molecule, such as hexane, is introduced into some water (Fig. 4–11). A cavity in the water is created, which temporarily disrupts some hydrogen bonds between water molecules. The dispersed water molecules then reorient themselves to form a maximum number of hydrogen bonds. The water molecules around the hexane molecule are much more ordered than elsewhere in the solution. Now consider the arrangement of two hexane molecules in water. The two possibilities are : either they sit in two small cavities (Fig. 4–11A) or in a single larger one (Fig. 4–11B). The experimental fact is that

the two hexane molecules come together and occupy a single large cavity. This association releases some of the more ordered water molecules around the separated hexanes. In fact, the basis of a hydrophilic attraction is this enhanced freedom of released water molecules. *Nonpolar solute molecules are driven together in water not primarily because they have a high affinity for each other but because water bonds strongly to itself.*

(A)   (B)
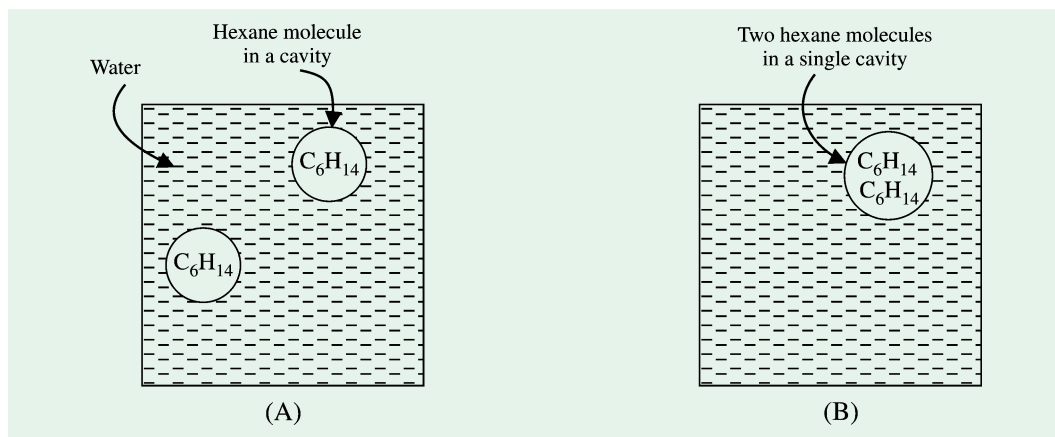
Fig. 4–11. Schematic of two molecules of hexane in a small volume of water
The hexane molecules occupy different cavities in the water structure (A), or
they occupy the same cavity, which is energetically more favoured (B).

D. van der Waals interactions

These have already been dealt with in detail in Chapter 2 (Water : The Solvent of Life) on pages 25 and 26.

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PROBLEMS

1. Rank the following in terms of expected dipole moment, and explain your choice :

$$H_2S, CCl_4, \overset{+}{H_3N} - CH_2 - COO^-, \overset{+}{H_3N} - CH_2 - CH_2 - CH_2 - COO^-$$