TYPES OF BONDING IN A COMPOUND & INTERMOLECULAR FORCES Section - 7

From the *difference in electronegativities* of the atoms forming a bond, it can be predicted *what type of bond is formed* by the atoms. When the difference is zero or very small, the bond is essentially *covalent*, and the electron density is more or less equally share d by two atoms. When the difference is large, one atom more-or-less completely withdraws the electron pair to it (one having higher electronegativity) and the bond is ionic. A bond *AB* (between two atoms *A* and *B*) is Ionic if difference in electronegativities of *A* and *B* is greater than 1.7 and it is Covalent, if it is less than 1.7.

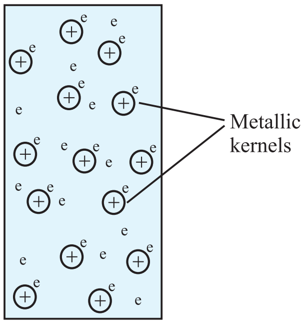
- Metals form Ionic bonds by losing electrons. However some metals form covalent bonds also and some don't form covalent bonds at all. The tendency of normal metals increases to form covalent bond as we proceed from left to right across a period in the *Periodic Table,* e.g., Ce and Ba form only Ionic bonds, Pb and Bi form both Ionic as well as Covalent bonds. Also metals rarely form bonds with other metals.
- Non-metals form Ionic bonds by gaining electrons. All non-metals can also form Covalent bonds as well. The tendency to form Ionic bond increases in going from left to right across a period and in going down the group, e.g., C forms only Covalent bond, F forms both Covalent as well as Ionic bonds. O forms O^{2-} in few compounds whereas S forms S^{2-} in many compounds.

Note : Most compounds of metals with non-metals are ionic and most compounds of non-metals with non–metals are covalent.

- Some non-metals such as C, Si, Ge, N, P, O are very reluctant to form negative ions. C, Si, Ge form only. Covalent bonds whereas N, P, O can form both Ionic and Covalent bonds.

Note : If these atoms form Ionic bonds, they also form Covalent bond with some other atom in the same compound. For example: In KOH, O forms Ionic bond with K but forms Covalent bond with H.

Metallic Bond :



Metals have low ionisation energies and hence valence electrons in them are weakly bounded to the positive charged nucleus or *kernel* (*kernel* means nucleus and the electrons other than the valence electrons). In metals the number of valence electrons is less than the number of orbitals to hold them. This makes possible for an electron to pass from one another through overlapping orbitals. Consider the case of Li, which has one s valence electron. In the solid crystal, each Li atom is close to eight other atoms, so that its orbitals will overlap with all of them at the same time and this picture is continued throughout the crystal. The electrons are shared by all the atoms, i.e., they are *delocalised* throughout the crystal.

So metal is network of ions (kernels) in a sea of electrons, with the electron sea acting as the binding force which holds the metal ions (kernels) together.

The valence electrons are not associated with any particular atom and can therefore move here and there. If the two ends of piece of a metal are connected to the poles of a battery, electrons will move towards the positive pole of the battery where they will flow out of the metal into the battery, being replaced by fresh electrons from the negative pole of the battery. In this manner an electric current can be conducted by the metals.

Intermolecular Forces

Intermolecular forces are the sum of the forces of attraction and repulsion between interacting particles (atoms and molecules). Attractive intermolecular forces are known as van der Waals forces, in honour of Dutch scientist Johannes van der Waals (1837-1923), who explained the deviation of real gases from the ideal behaviour through these forces (You will study about this in 'States of Matter'). These forces are other than due to covalent bonds, the hydrogen bonds, or the electrostatic interaction of ions with one another or with neutral molecules or charged molecules. The term includes:

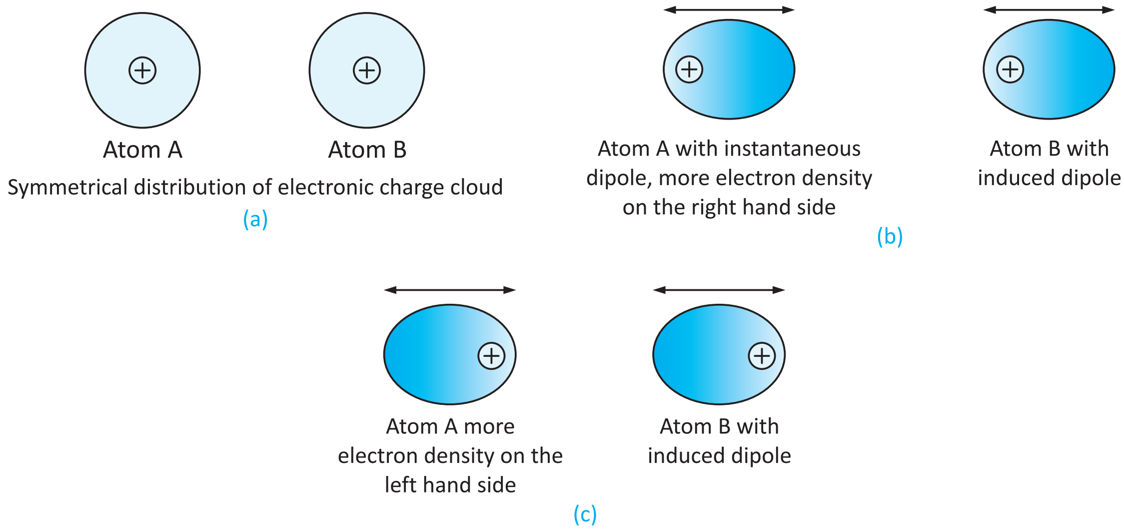
- force between two instantaneously induced dipoles (London dispersion force)
- force between two permanent dipoles
- force between a permanent dipole and a corresponding induced dipole

Note that attractive forces between an ion and a dipole are known as ion-dipole forces and these are not van der Waals forces.

Dispersion Forces or London Forces

Atoms and nonpolar molecules are electrically symmetrical and have no dipole moment because their electronic charge cloud is symmetrically distributed. But a dipole may develop momentarily even in such atoms and molecules. This can be understood as follows.Suppose we have two atoms 'A' and 'B' in the close vicinity of each other [Figure (a)]. It may so happen that momentarily electronic charge distribution in one of the atoms, say 'A', becomes unsymmetrical *i.e.,* the charge cloud is more on one side than the other [Figure (b and c)]. This results in the development of instantaneous dipole on the atom 'A' for a very short time. This instantaneous or transient dipole distorts the electron density of the other atom 'B', which is close to it and as a consequence a dipole is induced in the atom 'B'.

The temporary dipoles of atom 'A' and 'B' attract each other. Similarly temporary dipoles are induced in molecules also. This force of attraction was first proposed by the German physicist Fritz London, and for this reason force of attraction between two temporary dipoles is known as London force. Another name for this force is dispersion force.



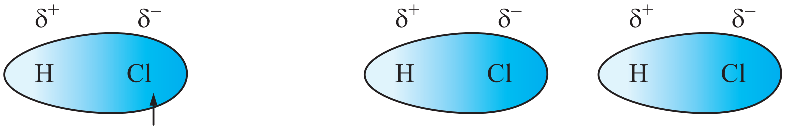
Dispersion forces or London forces between atoms

These forces are always attractive and interaction energy is inversely proportional to the sixth power of the distance between two interacting particles (i.e., $1/r^6$ where r is the distance between two particles). These forces are important only at short distances (~500 pm) and their magnitude depends on the polarisability of the particle.

Dipole - Dipole Forces

Dipole-dipole forces act between the molecules possessing permanent dipole. Ends of the dipoles possess “partial charges” and these charges are shown by Greek letter *delta* (δ). Partial charges are always less than the unit electronic charge (1.6×10^{-19} C). The polar molecules interact with neighbouring molecules. Figure (a) shows electron cloud distribution in the dipole of hydrogen chloride and Figure (b) shows dipole-dipole interaction between two HCl molecules. This interaction is stronger than the London forces but is weaker than ion-ion interaction because only partial charges are involved.

The attractive force decreases with the increase of distance between the dipoles. As in the above case here also, the interaction energy is inversely proportional to distance between polar molecules.



More charge density towards chlorine

(a)

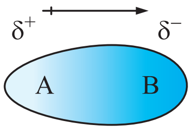
(b)

(a) Distribution of electron cloud in HCl a polar molecule, (b) Dipole-dipole interaction between two HCl molecules

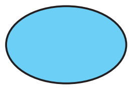
Dipole-dipole interaction energy between stationary polar molecules (as in solids) is proportional to $1/r^3$ and that between rotating polar molecules is proportional to $1/r^6$, where r is the distance between polar molecules. Besides dipole dipole interaction, polar molecules can interact by London forces also. Thus cumulative effect is that the total of intermolecular forces in polar molecules increase.

Dipole-Induced Dipole Forces

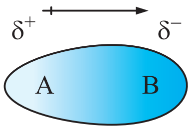
This type of attractive forces operate between the polar molecules having permanent dipole and the molecules lacking permanent dipole. Permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming its electronic cloud. Thus an induced dipole is developed in the other molecule. In this case also interaction energy is proportional to $1/r^6$ where r is the distance between two molecules. Induced dipole moment depends upon the dipole moment present in the permanent dipole and the polarisability of the electrically neutral molecule. We have already learnt in Unit 4 that molecules of larger size can be easily polarized. High polarisability increases the strength of attractive interactions.



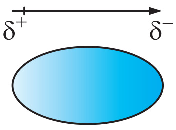
Permanent dipole
(a polar molecule)



A non-polar molecule



Permanent dipole
(a polar molecule)



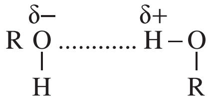
Induced dipole in
a non-polar molecule

In this case also cumulative effect of dispersion forces and dipole-induced dipole interactions exists.

Hydrogen Bonding :

Molecules that contain F–H, O–H, N–H bonds (e.g., Water, Alcohols, Organic acids, Amines, Aromatic alcohols, etc.) show a strong tendency to associate, i.e., to link up to form larger molecules. This feature exists in solid form and as well as in solutions in certain solvents. In all of such compounds, OH, or N–H or F–H, bond is highly polar due to large difference in electronegativity. The electrostatic attraction between such molecules should be quite strong. The *positive end* of one molecule attracts and is strongly attracted by the *negative end* of the neighbouring molecule. In this manner a large number of molecules are associated to form a cluster of molecules. Since in each case the hydrogen atom is responsible for the formation of linkages, this is known as *Hydrogen bond* or H-bond. It is impossible for hydrogen to form a second covalent bond so the additional linkage is shown by a *dotted line*. Hydrogen bonds are always of type : –A–H....B– , where A and B may be atoms of O, F, N. Hydrogen bonds are comparatively weak, with bond energies of 10-100 kJ/mol, but they are widespread and have important effect on many physical properties of many Organic and Inorganic compounds.

$$\overset{\delta+}{H}-\overset{\delta-}{F}\cdots\cdots\cdots\overset{\delta+}{H}-\overset{\delta-}{F}$$



$$R\,\underset{\underset{H}{|}}{\overset{\delta-}{O}}\cdots\cdots\cdots\overset{\delta+}{H}-\underset{\underset{R}{|}}{O}$$

Hydrogen Bonding is of two types :

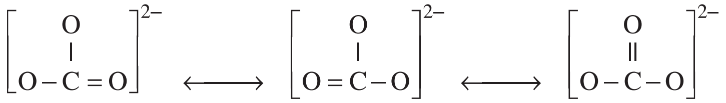
(i) Intermolecular H-bonding **(ii)** Intramolecular H-bonding

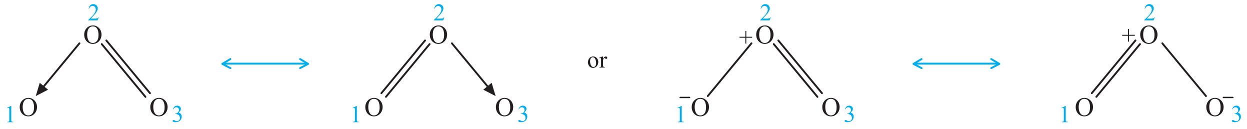
You will learn more about Hydrogen Bonding in the Chapter Chemical Bonding-II.

RESONANCE Section 8

There are certain molecules whose properties cannot be explained by a single structure.

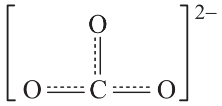
For example : The molecules of CO_3^{2-} ion, three Lewis dot structures are possible. Similarly Lewis structure for ozone (O_3) shows that one of O atom is bounded to a O atom by a double bond and by a coordinate covalent bond to the other O atom and there are two possible structures for it.



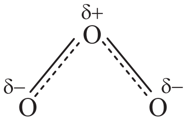


The length of three bonds in CO_3^{2-} ions and two bonds in O_3 is expected to be different. However from spectroscopic analysis, it is observed that the lengths of bonds are equal and lie somewhere between that of a single and double bond. Hence a particular structure is not sufficient to account for the observed facts.

The actual structure of CO_3^{2-} ion and O_3 molecule has neither of these Lewis structures. The real structure is an average (midway between) of these structures. This phenomenon of representing the actual structure in such a manner is called as Resonance. The individual structures are known as *Canonical forms* or *Resonance structures*. The actual structure is called as Resonance hybrid.



Resonance Hybrid of
Carbonate ion



Resonance Hybrid of
Ozone