Chapter # 5

CHEMICAL BONDING

You will learn in this chapter about:

- Why do atoms form chemical bonds?
- Ionic bond.
- Characteristics of ionic compounds.
- Covalent bond.
- Characteristic of covalent compounds.
- Electro negativity values.
- Ionic character of covalent bond.
- Co-ordinate covalent bond or Dative covalent bond.
- Difference between covalent and Co-ordinate covalent bonds.
- Metallic bonding.
- Inter-molecular forces.
- Hydrogen bonding.

5.1 INTRODUCTION

As we examine the world around us, we find that it is generally made up of compounds and mixtures. Rocks, coal, soil, air, trees, animals, all are JH formed by combination of atoms. Substances composed of single atoms are very rare. Examples are, argon (Ar) in the atmosphere and helium (He), mixed with natural gas. Clearly, there must be some force that holds atoms together in a molecule or crystals, otherwise the atoms would simply fly apart, and no compound could exist. The force which holds atoms together in a molecule or a crystal is called a chemical bond.

5.2 FORMATION OF CHEMICAL BONDS.

Before the discovery of electrical structure of atoms, the nature of the forces holding the atoms together in a molecule or crystal was a mystery. Now it is believed that these forces are electrical in nature and that the chemical reactions that occur between atoms involve change in their electronic structures. The electrons in the outer most shell of an atom are called the valence electrons. In the formation of chemical compounds from the elements, the valence electrons are either transferred from the outer shell of one atom to the outer shell of another atom or shared between them. This produces a chemical bond. When an atom of one element chemically combines with the atom of another element, both atoms usually attain a stable outer shell, consisting of eight electrons (octet). Only hydrogen H_2 and helium (He) atom have the stable outer shell of two electrons (Duplet). This is in accordance with the general rule that all processes tend to move towards the state of maximum stability. Generally, a stable molecule occurs, when the total energy of the combined atoms, is less than the total energy of the individual atoms.

5.3 TYPES OF CHEMICAL BONDS.

The first explanations of the nature of chemical bonds were advanced by W.Kossel (a German scientist) and G. N. Lewis (an American chemist) in 1916; they proposed two major types of chemicals bonds.

1. The ionic or electrovalent bond (by the transfer of one or more electrons from one atom to another, to form ions).
2. The covalent bond (a bond that results when atoms share electrons). Other types of bond include metallic bonds and hydrogen bonding.

REMEMBER

Chemical bonding also plays a role in determining the state of matter. At room temperature, water is liquid, carbon dioxide (CO_2) is a gas and table salt, sodium chloride (NaCl) is solid, because of difference in chemical bonding.

5.4 IONIC BOND OR ELECTROVALENT BOND

In this type of combination, there is a complete transfer of one or more electrons from one atom to another. The atom that transfers electrons gets positive charge and the atom that gains electrons gets negative charge. The strong electrostatic force acting between positive (+ve) and negative (-ve) ions, holds them together. The attraction that binds oppositely charged ions together is termed electrovalent bond or ionic bond.

For illustration let us consider the combination of sodium (Na) and chlorine (CI) atoms to form common salt, sodium chloride (NaCl). In this combination, an atom of Sodium (Na) transfers one outer most shell electron and becomes positive For illustration, let us consider the combination of sodium (Na) and sodium ion (Na^+) and an atom of chlorine gains that one electron to complete its octet and becomes chloride negative ion (Cl^-).

$NA_{2, 8, 1}$ --------→ $Na^+_{2, 8} + e^-$
$Cl_{2, 8, 7} + e^-$ ------→ $CI_{2, 8, 8}$

$Na^+ + Cl^{1-}$ ------→ Na^+ CI^- or $Na^{1+}Cl^{-1}$

Fig 5.1 Formation of Sodium Chloride

The attraction that binds (Na^+) and (CI^-) ions together is called electrovalent bond and the compound (NaCl) is called electrovalent compound or ionic compound.

IMPORTANT TO NOTE

The electrovalent bond is also known as ionic bond because electrovalent compounds when dissolved in water or melted produce ions and therefore, conducts electricity.

Another example of electrovalent bond, is the formation of magnesium oxide (MgO) from magnesium and oxygen. The magnesium (Mg) atom has two electrons in its valence shell. If these two electrons are lost, the (Mg) will become a di-positive ion (Mg^{+2}), and the oxygen (O) atom gains these two electrons to complete its octet, and becomes di-negative ion (O^{-2}). These two oppositely charged ions form electrovalent bond and the compound (MgO) is called electrovalent compound.

5.4.1 Characteristics of Ionic Compounds:

1. In ionic bond, it is impossible to say that any two ions are bonded to each other to produce molecule, but in the crystals of ionic compounds, the appositively charged ions are mutually surrounded by each other in orderly' arrangement. Thus ionic compounds are solids at room temperature.
2. Ionic compounds have high melting and boiling points because of the strong electrovalent bonds existing between the ions.
3. Ionic solids, do not conduct electricity as the ions are not free to move. Once an ionic compound is melted (fused) the ions are free to move and conduct electricity. Similarly, solutions of ionic compounds conduct electricity.
4. Ionic compounds are usually soluble in polar solvents, i.e. solvents of high di-electric constant such as water. But ionic compounds are insoluble in non polar (organic) solvents. These solvents have low di-electric constant such

as benzene, carbon tetrachloride, etc. They are mostly inorganic compounds.

5.5 COVALENT BOND

When two or more atoms of the same element or atoms of different elements having similar electro negativities react, the transfer of electrons does not occur. In these instances, the atoms achieve inert gas (noble gas) structure by sharing of electrons. Thus, the atoms complete their outer most shell by means of sharing of unpaired electrons, and a covalent bond is formed. In covalent bond each atom has to contribute equal number of unpaired electrons. The shared pair of electrons which links the atoms in a molecule is known as covalent bond.
In covalent bond the shared electron pair is commonly expressed by single short line (—) For example, the halogens (chlorine atoms) posses an electronic configuration in which there are seven electrons in their outer most shell, and lacking only one electron in order to attain the structure of an inert gas. This structure may be attained by the halogens molecule (chlorine molecule) when both atoms share one electron for Bond formation

5.5.1 Single, Double and Triple Covalent Bonds:

1. **Single Covalent Bond:** In single covalent bond only one pair of electrons is shared by the bonded atoms, in which each atom has to share one electron. This type of bond is represented by single short line (—). For example, in the formation of (H_2) and hydrogen chloride (HCl) molecules, only one pair of electrons is shared.
2. **Double Covalent Bond:** In double covalent bond only two pairs of electrons are shared by the bonded atoms, and each atom has to share two unpaired electrons. This type of bond is represented by two short lines (=) as shown in the molecules of oxygen (O_2) and carbon dioxide (CO_2).
3. **Triple Covalent Bond:** In triple covalent bond only three pairs of electrons are shared between the bonded atoms, and each atom has to share three unpaired electrons. This type of bond is denoted by three short lines (=) as shown in the molecules of nitrogen (N_2) and ethyne (C_2H_2).

5.5.2 Characteristics of Covalent Compounds:

1. Compounds with covalent bonds are usually made up of discrete units (molecules) with a weak inter molecular forces.
2. In the solid state, there are weak Vander wall forces between the molecules. Hence covalent compounds are often gases, liquids or soft solids with low melting points. In few cases, three dimensional covalent structures are formed rather than discrete units, hence diamond and silica (SiO_2) are covalent but are very hard and. have high melting points. Usually covalent compounds have low melting and boiling points.
3. They are insulators because they do not conduct electricity.
4. Covalent compounds are usually insoluble in polar solvents like water, but soluble in organic solvents like benzene, ether, carbon tetra chloride etc.

5.5.3 Electro negativity: (E. N.)

If the covalent bond is formed between two like atoms, that molecule is called non-polar-because the electron pair is shared equally between the two atoms, as in case of (H-H), (Cl-Cl), (0=0) and (N==N) molecules. However, if the covalent bond is formed between the two dissimilar atoms as in hydrogen chloride (H-Cl) molecule, the attraction for electron pair, would not be equal, one atom will attract more than the other. Hence the electron pair will be displaced from the central position and reaches near to the chlorine atom.

This power of an atom to attract the shared pair of electrons towards, itself, is known as electro negativity.
Consequently, the chlorine atom being more electronegative tends to be partially negative and hydrogen atom would be partially positive.
These values are based upon an arbitrary scale, in which fluorine is given an arbitrary standard value of electro negativity as 4.0. It is the most electronegative element. The electro negativity values of other elements are compared with that of fluorine. Note that the non metals have higher electro negativity values than the metals. Fluorine (F) has the highest electro negativity (4.0) and the metal cesium (Cs) has the lowest electro negativity (0.7) value.

5.4 Ionic Character in a Covalent Bond:

If covalent bond is formed between two like atoms e.g. in case (H-H), the hydrogen atoms are identical, and hence the shared pair of electrons is not disturbed from the centre. This molecule is called non-polar because it is electrically neutral as well as symmetrical.
If covalent bond is formed between dissimilar atoms e.g. in case of H-Cl molecule, the shared pair of electrons move closer to one of the both atoms. In H-Cl molecule chlorine is more electronegative, hence the shared pair of electrons, is drawn nearer to chlorine than hydrogen atom. This results in partial positive charge on H atom and partial negative charge on CI atom.

$$H^{+8} \text{------------} Cl^{-8}$$

The covalent bond between H and CI in H-Cl is partially ionic or polar covalent bond, because of positive and negative charges or poles.
Covalent bonds are partially ionic, if they exist between two dissimilar atoms and heir ionic character depends upon the difference in electro negativities of bonded atom.

i. **Non-Polar Covalent Bond:** According to the scale of Linus Pauling, if the difference in the electro negativities of bonded atoms is zero, then the bond is pure covalent bond or non-polar bond. The molecule containing like atoms or the atoms of same electro negativities form a pure covalent bonds or non-polar bonds.
ii. **Polar Covalent Bond:** If the difference in the electro negativities of bonded atoms is up to 1.7 that bond is called polar covalent ¥or partially ionic in character. The elements of different electro negativities always form polar covalent bond.
iii. **Electrovalent Bond:** If the difference in the electro negativities of bonded atoms is more than 1.7 then that bond is purely ionic or electrovalent. The bond between sodium and chlorine in the common salt (NaCl) is" clearly ionic, because the difference in the electro negativities is 2.1 i.e. more than 1.7.

Polar covalent bond (HCI) and Ionic bond (NaCI)

5.6 CO-ORDINATE COVALENT BOND OR DATIVE COVALENT BOND

The coordinate covalent bond is also a special type of covalent bond, in which both electrons forming a bond are supplied by one atom only. The atom which supplies the pair of electrons for bond formation is known as "**Donor**" and the atom which accepts, is known as "**Acceptor**". An arrow (—>) is commonly used to denote the bond, in which the head of an arrow pointing towards acceptor atom. The pair of electrons possessed by donor is called "**Lone-pair**" of electrons. This type of bond is also called dative covalent bond.
Examples:
In the formation of ammonium chloride, from ammonia (NH_3) and hydrogen chloride (HCI) molecules, the nitrogen (N) atom of ammonia (NH_3) acts as donor

of an electron pair and hydrogen ion (H^+) of hydrogen chloride accept it to form (NH_4Cl) as follows.
Another example is provided by the formation of hydronium ion (HO_2) from water molecule and hydrogen ion (H^+), in which the oxygen atom of water acts as donor and hydrogen ion (H^+), as acceptor.

5.6.1 Comparison between Covalent Bond and Co-ordinate Covalent Bond:

Covalent bond
1 The covalent bond is formed by the mutual sharing of electrons between atoms.
2. The covalent bond is formed between the similar or dissimilar atoms, when the electrons are mutually shared.
3. Covalent bond may be polar or non-polar
4. Covalent bond is associated with only covalent character because there is no electron transfer.
5. The shared pair of electrons in covalent bond are denoted by short lines e.g. for single pair (-) for double pair (=) and for triple pair of electrons (≡).
6. Covalent compounds are usually insoluble in water.

Co-ordinate Covalent bond
1. The co-ordinate covalent bond is formed by one sided sharing of electrons.
2. Co-ordinate covalent bond is formed between two unlike atoms, one having an electron pair available for sharing and other must accept that electron pair.
3. Co-ordinate covalent bond is always polar, because it is not formed between like atoms and is also known as co-ionic.
4. Co-ordinate covalent bond is associated with the ionic and covalent character because of partial transfer of electrons.
5. The shared pair of electrons in co-ordinate covalent bond is denoted by an arrow (—>).
6. Where as co-ordinate covalent compounds are sparingly soluble in water.

5.7 METALLIC BONDING

We have noticed in covalent bonding, that non-metals have sufficient number of valence electrons to combine and form molecules by sharing of electron pairs with one another. For example two chlorine atoms share an electron pair in Chlorine (Cl_2) molecule.
However most of the metal atoms have less than four valence electrons (many metals have only one or two). These electrons are not confined to any particular atom, instead they move freely from one atom to another atom. Hence the atoms should be considered to be positively charged ions; therefore, metal is defined as: A substance consisting *of positively charged ions, fixed in a crystal* lattice with negatively charge electrons moving freely through the crystal. Therefore free electrons act as cohesive force which hold the atoms together and form a metallic bond.
Metallic bond is defined as the combination of electrostatic attraction between the electrons and the positive nuclei of atoms. X-rays analysis reveals that metal particles are held together in a lattice of closely packed spheres.
A survey of the observed properties of metals indicates that the nature of force, holding the atoms together in a crystal, must be unusual. Metals are ductile (easily converted into thin wires) and malleable (easily bent or hammered into sheets).
They are good conductors of both heat and electricity (in solid or liquid states), they appear shiny and lustrous. Metals Eire solids at room temperature (except mercury (Hg)). They can be mixed with metals or non-metals to yield alloys with variable properties. All of these properties are because of peculiar bonding i.e. metallic bonding.

The strength of metallic bond varies considerably among different metals. Thus it is much stronger in iron than in sodium or potassium.

Fig. 5.8 Diagrammatic representation of metallic bonding

5.7.1 Metallic Binding in Sodium: (Na)
Sodium (Na) metal has one valence electron per atom. It crystallizes in a body centered cubic structure, in which each sodium is surrounded by eight nearest neighbors'. These valence electrons are not confined to any particular atom, instead, they are free to move throughout the crystal, so the resulting bond is relatively weak, that is, why metals like sodium and potassium are soft and have relatively low melting points.

5.7.2 Metallic Bonding in Iron (Fe) and Copper (Cu):
The metals iron (Fe) and copper (Cu) are hard and have high melting points because; these metals have incomplete valence shells. Therefore, the atoms become covalently bonded to each other through their unfilled orbits. As a result strong covalent bonding between atoms extends throughout the crystals. This accounts for their hardness and high melting points.

5.7.3 Explanation of the Properties of Metals:
Since electrons in metals are free to move from one atom to the next. They are generally good conductors of electricity.
When metal is heated, the mobile electrons absorb heat energy and transfer to neighbouring electrons, this means that metals are good conductors; of heat.
The mobile electrons readily absorb light falling upon them and move to higher energy levels. When they fall back to their original position they emits radiations. This causes the metallic lustre.

5.8 INTER-MOLECULAR FORCES (Vander Waals's Forces)
There are two types of forces present in molecules, that is:

(1) Intra-molecular forces and
(2) Inter-molecular forces.

Intra-molecular forces hold atoms together in a molecule. For example, - water (H_2O) molecule consists of two hydrogen atoms and one oxygen atom joined together in a specific way, that we call covalent bonds.
Inter-molecular forces are the attractive forces between the neutral molecules, which hold them together at certain temperature. These attractive forces of neutral molecules between each other are called "Vander Waal's Forces", named after the Dutch Physicist (1837-1923), who first described them. There are three types of attractive forces between molecules (1) Dispersion forces (2) Dipole-dipole forces, induced dipole and (3) Hydrogen bonding.
Generally, inter molecular forces are much weaker than intra-molecular forces.

1. Dispersion Forces (London forces):
All particles, whether individual atoms, ions or molecules experience dispersion forces, which result from the motion of electrons around atom. For example consider atoms of noble gases e.g. He, Ne, Ar. etc. Let us examine the attractive forces in neon as an example. The distribution of ten electrons around the nucleus of neon is spherically symmetrical. But in case when two (Ne) atoms, come extremely close together. The electron clouds will repel each other. This polarizes each molecule and gives rise to an induced or temporary dipoles and as a result weak attractive forces called dispersion forces also called London forces after Fritz London (who first indentified them in 1930) are developed. The

attraction is strong when particles are close together but rapidly weakens as they move apart.

Fig 5.9 Dispersion Force

Induced or temporary dipoles in Neon molecules
Thus dispersion forces (London forces) are the weak attractive forces between temporarily polarized atoms (or molecules) caused by the varying positions of the electrons during their motion about the nuclei.
London forces are generally small as their energies are in the range of 1-10 KJ/mol.

2. Dipole-Dipole forces:
Dipole-dipole forces are forces that act between polar molecules that possess dipole moments. A Dipole-dipole force, is an attractive inter-molecular force resulting from the inter action of the positive end of one molecule with the negative end of other.
Dipole-Dipole forces are generally stronger than dispersion forces
Consider the (H-Cl) molecule, as chlorine has greater electronegativity than hydrogen, a partial negative charge on chlorine atom, and a partial positive charge on hydrogen atom is developed. The (H^{+8}—CI^{-8}) has a permanent dipole moment. The electrostatic attraction of positive end of one (HC1) molecule for the negative end of another constitutes attractive forces in addition to dispersion forces. As a result polar (HQ) boils at (- 85°C) but non-polar (F^2) boils at (-188°C) though both have nearly same molecular weights (36.46 a.m.u for HC1 and 38.a.m.u for F_2).

3. Hydrogen Bonding:
When non-metal atoms of high electronegativity like those tof F, O and N, are linked to hydrogen, their exist a force of attraction between positive hydrogen atom of one molecule and negative oxygen, nitrogen or fluorine atom of another. This force is so strong enough to cause two or more molecules to associate in larger clusters, as for example, $(H_2O)_x$ and (NH_3)
This attraction between positive hydrogen and negative oxygen or free, is called hydrogen bond. This attraction or "hydrogen bond" can have about 5% to 10% of the strength of covalent bond. Hydrogen bond is
denoted by dotted lines ( ).
Hydrogen bondings differ from the word "bond* since it is a force of attraction between positive hydrogen atom of one molecule and negative oxygen atom of another molecule. That is, it is an intermolecular force, not an intra-molecular force as in the common use of word bond. For this, hydrogen bonding is particularly strong type of dipole-dipole inter-action.
Hydrogen bondings have an important affect on the properties of water and ice. Hydrogen bonding is also very important in proteins and nucleic acids and therefore in life processes.

SUMMARY

1. Chemical bonds hold groups of atoms together to form molecules or solids. Bonding occurs when a group of atoms can lower its total energy by combining.
2. Bonds can be classified into two main types; ionic bond and covalent bond.
3. Electronegativity is defined as the relative ability of an atom in a molecule to attract the shared electrons in a bond to itself. The Electronegativity difference of the bonded atoms determines the polarity of that bond.

4. A covalent bond may be single bond, a double bond or a triple bond, depending on whether the two atoms share one, two or three pairs of electrons.
5. The co-ordinate covalent bond is formed only when an atom with an unshared pair of electrons in its valence shell donates a pair of electrons to another atom or ion that needs a pair of electrons to acquire a stable electronic configuration.
6. Covalent molecules are partially ionic, if they exist between two unlike atoms and their ionic character depends upon difference in electro-negativities of bonded atoms.
7. Metallic bond is the electrostatic attraction between positive ions and electrons of the atoms.
8. A hydrogen bond is a dipole-dipole attractive force that exists between two polar molecules, containing a hydrogen atom covalently bonded to an atom of F, O or N. Water is best example of hydrogen bonding.

EXERCISE

1. Fill in the blanks:

i. ________ covalent molecule is electrically neutral as well as symmetrical.
ii. The power of an atom to attract the shared pair of electrons towards Itself is called_______
iii. ________Compounds are usually made up of discrete units; with weak inter molecular forces.
iv. NaCl is an _______ compound.
v. If electronegativity difference of bonded atoms is more then 1.7, the bond is________
vi. The electrostatic attraction between positive ions and the electrons of the atoms is called _______ bond.
vii. The forces which hold atoms together in a molecule are called_______
viii. The attraction between the positive hydrogen and negative F, O or N is called _______ bonding.
ix. CO_2 is a _______ molecule.
x. The atom which accepts a lone pair of electron is called_______

2. Tick the correct answer:

i. The force which holds atoms together in a molecule or crystal is called:
   a. Ionic bond.
   b. Covalent bond.
   c. Co-ordinate covalent bond.
   d. Chemical bond.
ii. The bond which is formed by the transfer of one or more electrons from one atom to another atom is called:
   a. Ionic bond.
   b. Covalent bond.
   c. Co-ordinate covalent bond.
   d. Chemical bond.
iii. The bond which is formed by the mutual sharing of electrons between the atoms, is called:
   a. Ionic bond.
   b. Covalent bond.
   c. Co-ordinate covalent bond.
   d. Chemical bond.
iv. The bond which is formed by one sided sharing of pair of electrons is called:
   a. Ionic bond.

b. Covalent bond.
c. Co-ordinate covalent bond.
d. Chemical bond.

v. The bond in MgO is:
a. Electro-valent bond.
b. Covalent bond.
c. Co-ordinate covalent bond.
d. Chemical bond.

vi. The shared pair of electrons which links the atoms in a molecule is known as:
a. Electro-valent bond.
b. Covalent bond.
c. Co-ordinate covalent bond.
d. Chemical bond.

vii. Double covalent bond is denoted by:
a. Single short line.
b. Two short lines.
c. Three short lines.
d. None of these.

viii. The, atom which supplies the pair of electrons for bond formation is known as:
a. Acceptor
b. Donor.
c. Receiver, ISM
d. None of these.

ix. Co-ordinate covalent bond is always formed between the two:
a. Like atoms.
b. Unlike atoms.
c. Similar atom.
d. Like and unlike atoms.

x. The shared pair of electrons in a co-ordinate covalent bond is denoted by:
a. A single line.
b. Double line.
c. An equal sign.
d. An arrow.

Write answer of the following questions:

i. Define chemical bond? Discuss how atoms unite and change into molecules?
ii. What are the valence electrons of an atom? How many valence electrons does a nitrogen atom possess?
iii. What happens to electrons, when elements combine?
iv. What part of the atom is involved in the formation of chemical bond?
v. Explain with examples? How elements are united by electro-valent bond?
vi. What common properties are shown by ionic compounds?
vii. What is meant by covalent bond? Write electronic formulas of any two covalent molecules? What is single, double and triple covalent bond?
viii. Draw the electronic formulae for the following covalent molecules?
a. H_2
b. O_2
c. N_2
d. C_2H_2
e. CO_2

ix. Classify the following bonds as ionic or covalent. For those bonds that are covalent indicate whether they are polar or non-polar.

a. H_2
b. H-Cl
c. NaCl
d. $CaCO_3$
e. HC≡CH
f. O=O

x. What are types of chemical bondings?
xi. Account for the fact that some covalent bonds are polar while others are non-polar.
xii. What is co-ordinate covalent bond? Explain with examples?
xiii. Define the term covalent bond? How does a covalent bond differ from co-ordinate covalent bond?
xiv. Explain electronegativity.
xv. Explain in your own words Pauling (E.N.) table? Explain its usefulness in predicting the relative ionic and covalent character of a given compound?
xvi. Give the characteristics of covalent compounds?
xvii. What do you understand by ionic character of covalent bond? Under what conditions are the following formed
    a. Polar covalent bond.
    b. Non-polar covalent bond.
    c. Ionic bond.
xviii. Define the term metal? And describe metallic bond?
xix. Explain the following properties of metals?
    a. Lustre.
    b. Conductivity.
    c. Malleability.
    d. Ductility.
xx. Why are some metals, such as sodium is soft, while other are hard?
xxi. Explain the origin of dipole-dipole forces between the molecules? Give an example?
xxii. What do you mean by dispersion forces? Why they are also called London forces?
xxiii. What is hydrogen bonding? What type of forces, either intra-molecular or inter-molecular forces are present in hydrogen bonding?