

# CHAPTER-04

## CHEMICAL BONDING

CLASS-XI

SUB-CHEMISTRY

- **Chemical Bond**

The force that holds different atoms in a molecule is called chemical bond.

- **Octet Rule**

Atoms of different elements take part in chemical combination in order to complete their octet or to attain the noble gas configuration.

- **Valence Electrons**

It is the outermost shell electron which takes part in chemical combination.

• **Facts Stated by Kossel in Relation to Chemical Bonding**

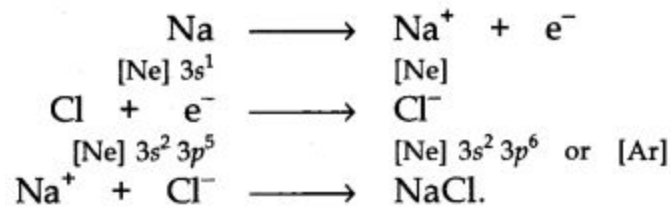
– In the periodic table, the highly electronegative halogens and the highly electro-positive alkali metals are separated by noble gases.

– Formation of an anion and cation by the halogens and alkali metals are formed by gain of electron and loss of electron respectively.

– Both the negative and positive ions acquire the noble gas configuration.

– The negative and positive ions are stabilized by electrostatic attraction

Example,



## • Modes of Chemical Combination

– By the transfer of electrons: The chemical bond which formed by the complete transfer of one or more electrons from one atom to another is termed as electrovalent bond or ionic bond.

– By sharing of electrons: The bond which is formed by the equal sharing of electrons between one or two atoms is called covalent bond. In these bonds electrons are contributed by both.

— Co-ordinate bond: When the electrons are contributed by one atom and shared by both, the bond is formed and it is known as dative bond or co-ordinate bond.

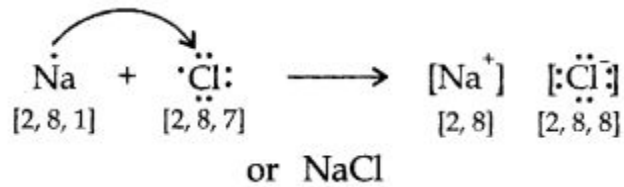
### • **Ionic or Electrovalent Bond**

Ionic or Electrovalent bond is formed by the complete transfer of electrons from one atom to another. Generally, it is formed between metals and non-metals. We can say that it is the electrostatic force of attraction which holds the oppositely charged ions together.

The compounds which is formed by ionic or electrovalent bond is known as electrovalent compounds. For Example, ,

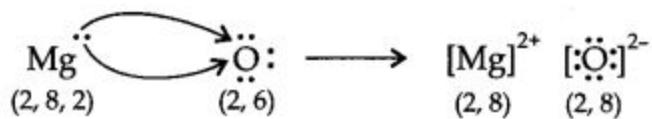
(i) NaCl is an electrovalent compound.

Formation of NaCl is given below:



Na<sup>+</sup> ion has the configuration of Ne while Cl<sup>-</sup> ion represents the configuration of Ar.

(ii) Formation of magnesium oxide from magnesium and oxygen.



Electrovalency: Electrovalency is the number of electrons lost or gained during the formation of an ionic bond or electrovalent bond.

## • Factors Affecting the Formation of Ionic Bond

(i) Ionization enthalpy: As we know that ionization enthalpy of any element is the amount of energy required to remove an electron from outermost shell of an isolated gaseous atom to convert it into cation.

Hence, lesser the ionization enthalpy, easier will be the formation of a cation and have greater chance to form an ionic bond. Due to this reason alkali metals have more tendency to form an ionic bond.

For example, in formation of  $\text{Na}^+$  ion I.E = 496 kJ/mole

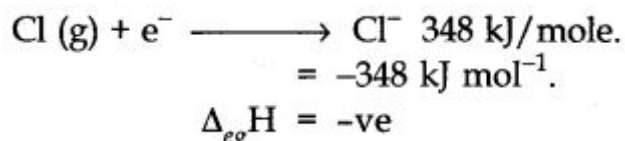
While in case of magnesium, it is 743 kJ/mole. That's why the formation of positive ion for sodium is easier than that of magnesium.

Therefore, we can conclude that lower the ionization enthalpy, greater the chances of ionic bond formation.

(ii) Electron gain enthalpy (Electron affinities): It is defined as the energy released when an isolated gaseous atom takes up an electron to form anion. Greater the negative electron gain enthalpy, easier will be the formation of anion.

Consequently, the probability of formation of ionic bond increases.

For example. Halogens possess high electron affinity. So, the formation of anion is very common in halogens.



(iii) Lattice energy or enthalpy: It is defined as the amount of energy required to separate 1 mole of ionic compound into separate oppositely charged ions.

### **Lattice energy of an ionic compound depends upon following factors:**

(i) Size of the ions: Smaller the size, greater will be the lattice energy.

(ii) Charge on the ions: Greater the magnitude of charge, greater the interionic attraction and hence higher the lattice energy.

### **• General Characteristics of ionic Compounds**

(i) Physical State: They generally exist as crystalline solids, known as crystal lattice.



Ionic compounds do not exist as single molecules like other gaseous molecules e.g., H<sub>2</sub> , N<sub>2</sub> , O<sub>2</sub> , Cl<sub>2</sub> etc.

(ii) Melting and boiling points: Since ionic compounds contain high interionic force between them, they generally have high melting and boiling points.

(iii) Solubility: They are soluble in polar solvents such as water but do not dissolve in organic solvents like benzene, CCl<sub>4</sub> etc.

(iv) Electrical conductivity: In solid state they are poor conductors of electricity but in molten state or when dissolved in water, they conduct electricity.

(v) Ionic reactions: Ionic compounds produce ions in the solution which gives

very fast reaction with oppositely charged ions.

For example,



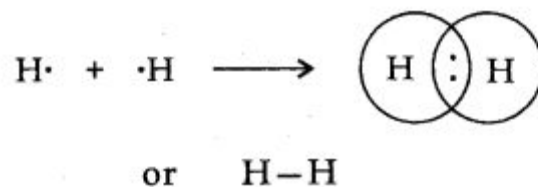
### • **Covalent Bond—Lewis-Langmuir Concept**

When the bond is formed between two or more atoms by mutual contribution and sharing of electrons, it is known as covalent bond.

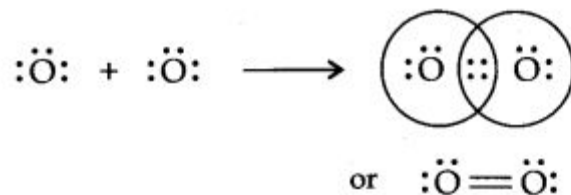
If the combining atoms are same the covalent molecule is known as homoatomic. If they are different, they are known as heteroatomic molecule.

For Example,

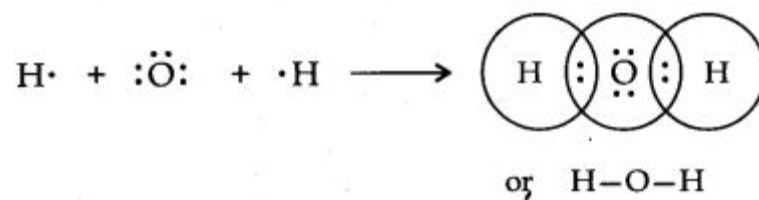
(i) Formation of hydrogen molecule.



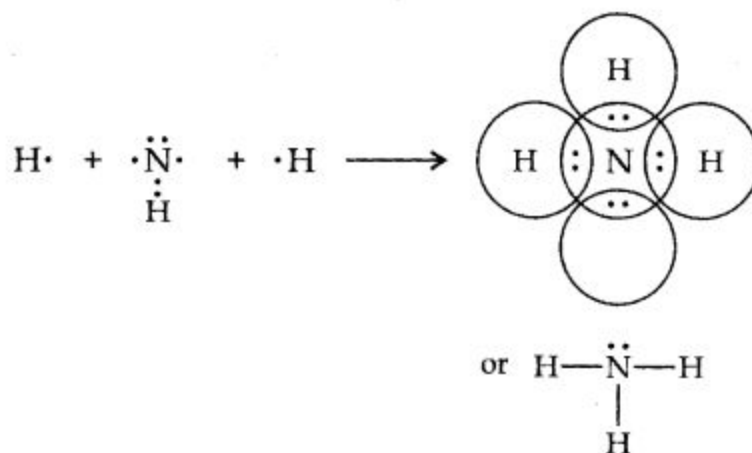
(ii) Formation of oxygen molecule.



(iii) Formation of  $\text{H}_2\text{O}$  molecule.



(iv) Formation of  $\text{NH}_3$  molecule.



## • Lewis Representation of Simple Molecules (the Lewis Structures)

The Lewis dot Structure can be written through the following steps:

- (i) Calculate the total number of valence electrons of the combining atoms.
- (ii) Each anion means addition of one electron and each cation means removal of one electron. This gives the total number of electrons to be distributed.
- (iii) By knowing the chemical symbols of the combining atoms.
- (iv) After placing shared pairs of electrons for single bond, the remaining electrons may account for either multiple bonds or as

lone pairs. It is to be noted that octet of each atom should be completed.

Lewis structures of some typical molecules and ions

Molecule/ion		Lewis Representation	
(i)	CO	$\text{:C}\ddot{\text{O}}\text{:}$	or $\text{:C}\equiv\text{O}\text{:}$
(ii)	CO <sub>2</sub>	$\ddot{\text{O}}\text{:C}\text{:}\ddot{\text{O}}$	or $\ddot{\text{O}}=\text{C}=\ddot{\text{O}}$
(iii)	O <sub>2</sub>	$\text{:O}\ddot{\text{O}}\text{:}$	or $\text{:}\ddot{\text{O}}=\ddot{\text{O}}\text{:}$
(iv)	O <sub>3</sub>	$\begin{array}{c} \ddot{\text{O}}^+ \\ \diagup \quad \diagdown \\ \text{:}\ddot{\text{O}}\text{:} \quad \text{:}\ddot{\text{O}}\text{:}^- \end{array}$	or $\begin{array}{c} \ddot{\text{O}}^+ \\ \diagup \quad \diagdown \\ \text{:}\ddot{\text{O}}\text{:} \quad \text{:}\ddot{\text{O}}\text{:}^- \end{array}$
(v)	H <sub>2</sub> O <sub>2</sub>	$\text{H}:\ddot{\text{O}}:\ddot{\text{O}}:\text{H}$	or $\text{H}-\ddot{\text{O}}-\ddot{\text{O}}-\text{H}$
(vi)	SO <sub>3</sub>	$\begin{array}{c} \ddot{\text{O}}:\text{S}:\ddot{\text{O}} \\ \quad \quad \quad \ddot{\text{O}} \end{array}$	or $\begin{array}{c} \ddot{\text{O}}=\text{S}=\ddot{\text{O}} \\ \quad \quad \quad \ddot{\text{O}} \end{array}$
(vii)	HClO <sub>4</sub> (Perchloric acid)	$\begin{array}{c} \ddot{\text{O}} \\ \text{:}\ddot{\text{H}}:\ddot{\text{O}}:\ddot{\text{Cl}}:\ddot{\text{O}}\text{:} \\ \quad \quad \quad \ddot{\text{O}} \end{array}$	or $\begin{array}{c} \ddot{\text{O}} \\ \text{:}\ddot{\text{H}}-\ddot{\text{O}}-\text{Cl}-\ddot{\text{O}}\text{:} \\ \quad \quad \quad \ddot{\text{O}} \end{array}$

(viii)	CN <sup>-</sup> (Cyanide ion)	$[\text{:C}\ddot{\text{N}}\text{:}]^-$	or	$[\text{:C}\equiv\text{N}\text{:}]^-$
(ix)	NO <sub>2</sub> <sup>-</sup> (Nitrite ion)	$[\text{:}\ddot{\text{O}}\text{:}\ddot{\text{N}}\text{:}\ddot{\text{O}}\text{:}]^-$	or	$[\text{:}\ddot{\text{O}}=\ddot{\text{N}}-\ddot{\text{O}}\text{:}]^-$
(x)	SO <sub>4</sub> <sup>2-</sup>	$\left[ \begin{array}{c} \ddot{\text{O}} \\ \text{:}\ddot{\text{O}}:\text{S}:\ddot{\text{O}}\text{:} \\ \quad \quad \quad \ddot{\text{O}} \end{array} \right]^{2-}$	or	$\left[ \begin{array}{c} \ddot{\text{O}} \\ \text{:}\ddot{\text{O}}-\text{S}-\ddot{\text{O}}\text{:} \\ \quad \quad \quad \ddot{\text{O}} \end{array} \right]^{2-}$
(xi)	O <sub>2</sub> <sup>2-</sup>	$[\text{:}\ddot{\text{O}}:\ddot{\text{O}}\text{:}]^{2-}$	or	$[\text{:}\ddot{\text{O}}-\ddot{\text{O}}\text{:}]^{2-}$

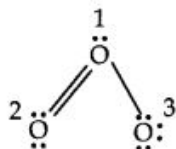
- **Formal Charge**

In polyatomic ions, the net charge is the charge on the ion as a whole and not by particular atom. However, charges can be assigned to individual atoms or ions. These are called formal charges.

It can be expressed as

$$\text{Formal charge (F.C.) on an atom in a Lewis structure} = \left[ \begin{array}{l} \text{total number of valence} \\ \text{electrons in the free} \\ \text{atom} \end{array} \right] - \left[ \begin{array}{l} \text{total number of non} \\ \text{bonding (lone pair)} \\ \text{electrons} \end{array} \right] - (1/2) \left[ \begin{array}{l} \text{total number of} \\ \text{bonding (shared)} \\ \text{electrons} \end{array} \right]$$

It does not indicate the real charge. Let us consider the Ozone molecule ( $O_3$ ). The Lewis structure of  $O_3$  may be drawn as:



The atoms have been numbered as 1, 2 and 3. The formal charge on:

The central O atom marked 1

$$= 6 - 2 - \frac{1}{2} (6) = +1$$

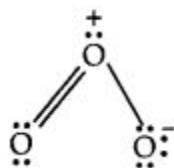
The end O atom marked 2

$$= 6 - 4 - \frac{1}{2} (4) = 0$$

The end O atom marked 3

$$= 6 - 6 - \frac{1}{2} (2) = -1$$

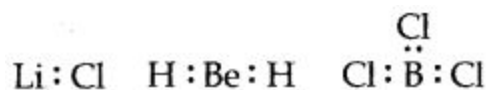
Hence, we represent  $O_3$  along with the formal charges as follows:



## • Limitations of the Octet Rule

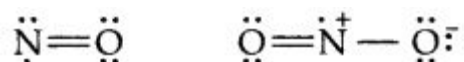
(i) The incomplete octet of the central atoms: In some covalent compounds

central atom has less than eight electrons, i.e., it has an incomplete octet. For example,



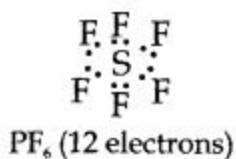
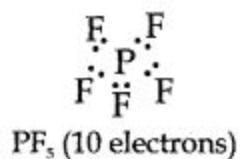
Li, Be and B have 1, 2, and 3 valence electrons only.

(ii) Odd-electron molecules: There are certain molecules which have odd number of electrons the octet rule is not applied for all the atoms.



(iii) The expanded Octet: In many compounds there are more than eight valence electrons around the central atom. It is termed as expanded octet. For Example,





## • Other Drawbacks of Octet Theory

(i) Some noble gases, also combine with oxygen and fluorine to form a number of compounds like XeF<sub>2</sub> , XeOF<sub>2</sub> etc.

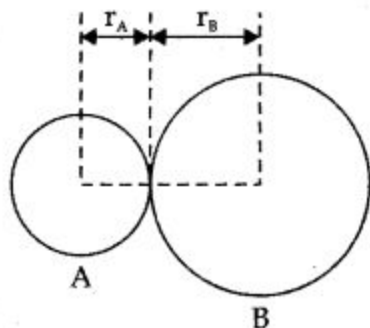
(ii) This theory does not account for the shape of the molecule.

(iii) It does not give any idea about the energy of The molecule and relative stability.

## • Bond Length

It is defined as the equilibrium distance between the centres of the nuclei of the two bonded atoms. It is expressed in terms of

A. Experimentally, it can be defined by X-ray diffraction or electron diffraction method.



*The bond length in a covalent molecule AB.  
 $R = r_A + r_B$  ( $R$  is the bond length and  $r_A$  and  $r_B$   
are the covalent radii of atoms A and B respectively)*

## • Bond Angle

It is defined as -the angle between the lines representing the orbitals containing the bonding – electrons.

It helps us in determining the shape. It can be expressed in degree. Bond angle can be experimentally determined by spectroscopic methods.

## • Bond Enthalpy

It is defined as the amount of energy required to break one mole of bonds of a particular type to separate them into gaseous atoms.

Bond Enthalpy is also known as bond dissociation enthalpy or simple bond enthalpy. Unit of bond enthalpy =  $\text{kJ mol}^{-1}$

Greater the bond enthalpy, stronger is the bond. For e.g., the H—H bond enthalpy in hydrogen is  $435.8 \text{ kJ mol}^{-1}$ .

The magnitude of bond enthalpy is also related to bond multiplicity. Greater the bond multiplicity, more will be the bond enthalpy. For e.g., bond enthalpy of C — C bond is  $347 \text{ kJ mol}^{-1}$  while that of C = C bond is  $610 \text{ kJ mol}^{-1}$ .

In polyatomic molecules, the term mean or average bond enthalpy is used.

Average Bond Enthalpies of some common Bonds

Bond	Bond enthalpy ( $\text{kJ mol}^{-1}$ )	Bond	Bond enthalpy ( $\text{kJ mol}^{-1}$ )
H-H	435.8	C-C	347
Cl-Cl	239	C=C	610
Br-Br	196	C $\equiv$ C	836
I-I	158	N-N	158
H-Cl	431.0	N=N	418
H-F	564	N $\equiv$ N	946

## • Bond Order

According to Lewis, in a covalent bond, the bond order is given by the number of bonds between two atoms in a molecule. For example,

Bond order of  $\text{H}_2$  (H - H) = 1

Bond order of  $\text{O}_2$  (O = O) = 2

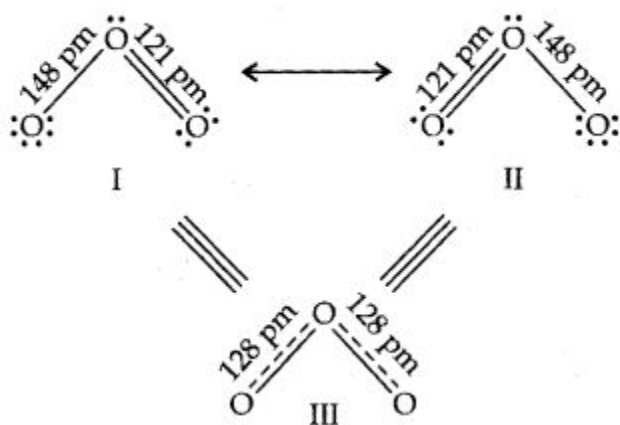
Bond order of  $\text{N}_2$  (N  $\equiv$  N) = 3

Isoelectronic molecules and ions have identical bond orders. For example,  $F_2$  and  $O_2^{2-}$  have bond order = 1.  $N_2$ ,  $CO$  and  $NO^+$  have bond order = 3. With the increase in bond order, bond enthalpy increases and bond length decreases. For example,

For  $N_2$  B.O. = 3  
Bond enthalpy = 946 kJ/mol.  
For  $O_2$ , B.O. = 2  
Bond enthalpy = 498 kJ/mol.

## • Resonance Structures

There are many molecules whose behaviour cannot be explained by a single-Lewis structure, For example, Lewis structure of Ozone represented as follows:



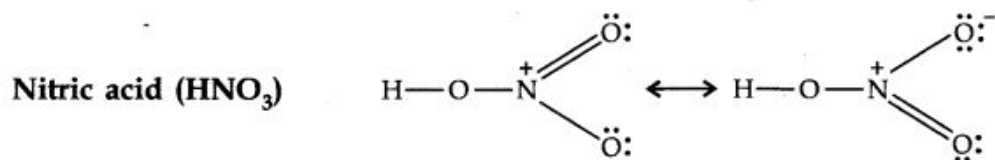
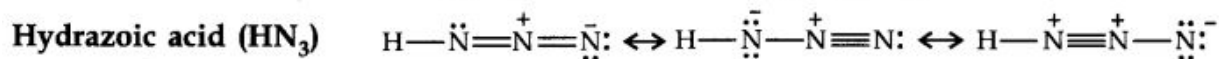
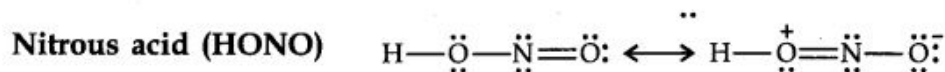
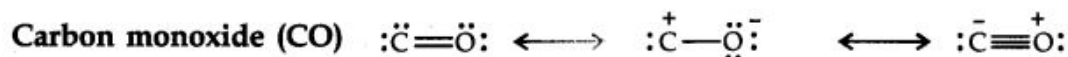
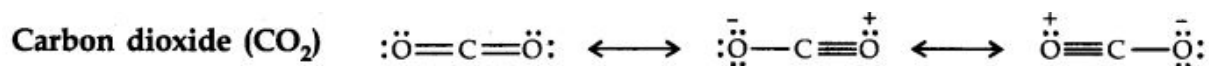
*Resonance in the  $O_3$  molecule*

*(Structures I and II represent the two canonical forms while the Structure III is the resonance hybrid)*

Thus, according to the concept of resonance, whenever a single Lewis structure cannot explain all the properties of the molecule, the molecule is then supposed to have many structures with similar energy. Positions of nuclei, bonding and nonbonding pairs of electrons are taken as the canonical structure of the hybrid which describes the molecule accurately. For  $O_3$ , the two structures shown above are canonical structures and

the III structure represents the structure of  $O_3$  more accurately. This is also called resonance hybrid.

Some resonating structures of some more molecules and ions are shown as follows:



## • Polarity of Bonds

### Polar and Non-Polar Covalent bonds

**Non-Polar Covalent bonds:** When the atoms joined by covalent bond are the same like;  $H_2$ ,  $O_2$ ,  $Cl_2$ , the shared pair of electrons is

equally attracted by two atoms and thus the shared electron pair is equidistant to both of them.

Alternatively, we can say that it lies exactly in the centre of the bonding atoms. As a result, no poles are developed and the bond is called as non-polar covalent bond. The corresponding molecules are known as non-polar molecules.

For Example,

