# CHEMICAL BONDING

As we know that atoms rarely occur in the free state at ordinary temperature. A cluster formed may be either a molecule or in ion, due to the attraction of atom. So, a chemical bond is defined as the attractive force that holds the two or more atoms to gather in a molecule or an ion. It is also known as valence bond.

## CLASSIFICATION OF CHEMICAL BOND.

The chemical bond may be classified into two parts.

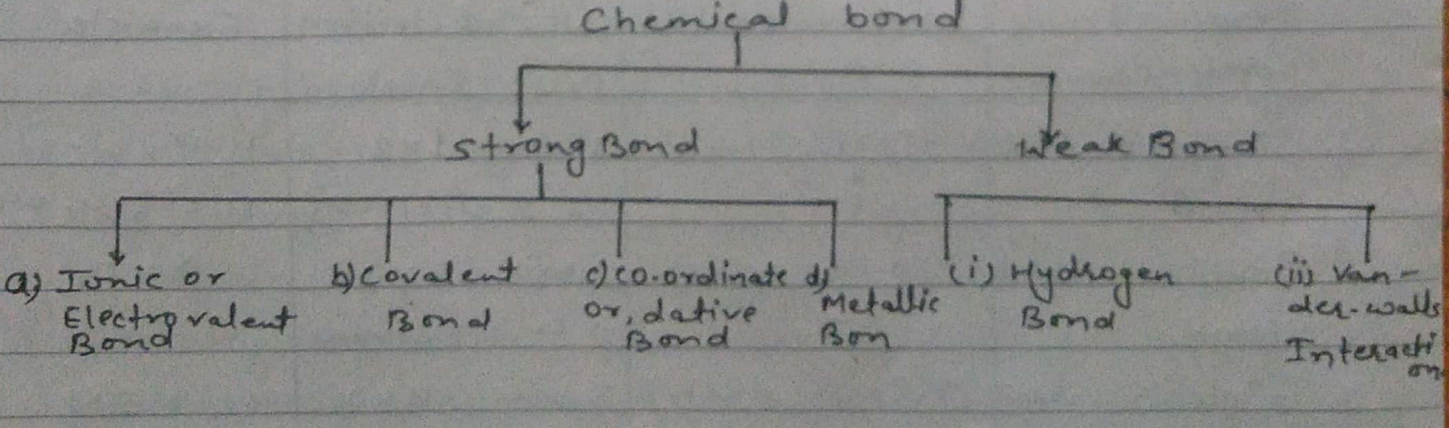
1. strong chemical bond
2. Weak chemical bond

Now, strong chemical bond is those type of bond in which the force of attraction is greater. It is classified as in four parts.

a) Ionic or electrovalent bond
b) Covalent bond
c) Co-ordinate or dative bond
d) Metallic bond.

Now, Weak chemical bond is those type of bond in which the force of attraction is less. It is classified as in two parts.

i) Hydrogen bond
ii) Van-der-walls interaction.



# COVALENT BOND

The attractive force between the two nuclei for the shared paired electron that hold the atoms together is called covalent bond and the compounds containings covalent bond are called covalent compound. It is also known as electron pair bond. It was given by G.N. Lewis, in 1916.

According to lewis concept, when two atoms forming a covalent bond between them, then each of the atoms attains the stable configuration of nearest inert gas by completing its octet or duplets. Since a covalent bond between the atoms results by the interaction of their electrons which common to both the atoms. So it is also known as atomic bond. Covalent bonds are classified into three types —

1. Single covalent bond
2. Double covalent bond
3. Triple covalent bond

Double and Triple covalent bonds are called multiple covalent. The covalent bond formed by the sharing electrons pair according to their number, which leads the formation of single, double and Triple covalent bond.

| Types of covalent bond. | Reprensented By | No. of electron Pairs involved | Examples |
|---|---|---|---|
| 1. Single covalent bond | Single dash (—) | 1 pair = 1×2 = 2 electrons | H-H, Cl-Cl<br>H-Cl, H-OH |
| 2. Double covalent bond | Double dash (=) | 2 pair = 2×2 = 4 electrons | O=O |
| 3. Triple covalent bond | Triple dash (≡) | 3 pair = 3×2 = 6 electrons | N≡N |

Factors favouring the formation of covalent compound: Conditions for the formation of covalent compound —

⇒ The following factor leads the formation of covalent compound.

1. High ionisation Energy :→ The higher ionisation energy lead the formation of co-valent bond, because it is in capable of losing electrons to form cations.
2. Equal electron affinities:→ for covalent bond two atoms must be equal attraction for electrons.
3. High nuclear charge and small internuclear distance:— High charge on the bonding nuclei and smaller internuclear distance favour the formation of covalent bond.
4. Numbers of valence electron :→ Each of the two atoms should have 5, 6, or 7 valence electron while H-atom has only one electron so that the both the atoms achives the stable octet by sharing 3, 2, or 1 electron pair respectively.
5. Equal electronegativity :→ When the electronegativity of both the atom is equal, sharing of electron pair(s) occurs and covalent bond is established.

VALENCE BOND THEORY OF COVALENT BOND

In 1927, Heitler and London proposed a theory related to covalent bond which is called valence Bond Theory (VBT). According to this theory,

1. An atomic orbital of the outermost shell of an atom containing one electron has a tendency to overlap of the another atomic orbital of another atom containing one electron of opposite spin.
2. The overlap of two atomic orbitals gives rise to a single bond orbital which is localised orbital and it is

occupied by the both electrons.

3. The two electrons that occupy the bond orbital have opposite spins.
4. Electron occupies the entire bond orbital, i.e, the electron pair present in the bond orbital now belongs to each of the two atomic orbitals.
5. As a result of overlapping there is maximum electron density. A large part of bonding force of covalent bond results from the electro static attraction between the nuclei and the accumulated electron clouds between them.

EXTENSION OF VALENCE BOND THEORY

The above theory was extended by Pauling and Slater in 1931. According to the extension of VB theory,

1. The formation of a covalent bond is always accompanied by the evolution of energy. The energy released as a result of overlapping of the orbitals stabilies the system.

NOTE :→ 1. The amount of energy released per mole at the time of overlapping of orbital to form the covalent bond, is called bond energy or Stabilisation energy.

2. The energy required to break the bond in a molecule is called bond dissociation energy.

3. The equilibrium distance at which the two atomic nuclei are now held is called bond length.

2. Overlapping takes place only between those valence Shell orbital which have unpaired electron.

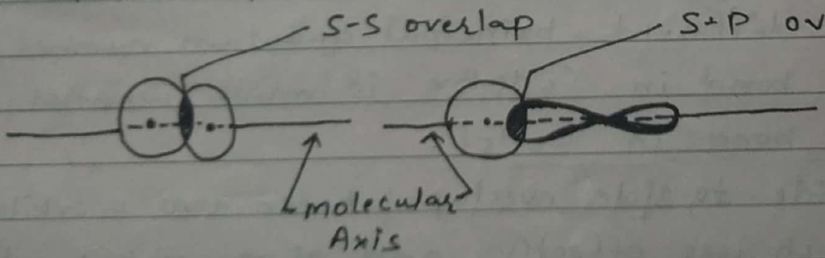
3. Between the two orbitals of the samer energy or stability, the orbital which is non sperical forms stronger bond than which is spherically symmetrical.
4. A spherically symmetrical orbital does not show any preference in direction while non sperical orbitals tend to form a bond in the direction of maximum electron density within the orbital.

TYPES OF OVERLAP OF ATOMIC ORBITAL.

:) There are three types of overlap of atomic orbital between the two atoms of covalent bond.

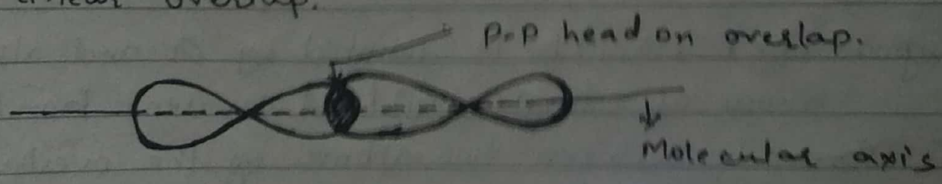
a) S-S overlap :→ In this type of overlap half filled s-orbital of one atom overlaps with the half filled of the other atom.

b) S-P overlap :→ In this type of overlap, half filled s-orbital of one atom overlaps with the half filled p-orbital of the other atom.

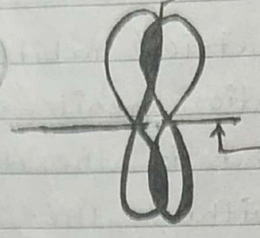



c) P-P overlap :→ In this type of overlap, single filled p-orbital of one atom overlaps with half filled p-orbital of the other atom. P-P overlap may be overlap in any of following two ways —

i) P-P head to head overlap:→ It is also known as head on, end on, or end to end or linear overlap.



(iii) P-P side to side overlap :→ It is also known as side wise, side way or lateral overlap. In this type orbital axis should be parallel.




STRENGTH OF COVALENT BOND

1. The strength of a covalent bond is proportional to the extent of overlapping between atomic orbital. Stronger bond has a shorter bond length.
2. Bonding between two s-orbital is weak particularly when two orbitals are of different energy.
3. A stronger bond is given by head to head overlap of two p-orbitals which are of same energy. The bond is much weaker, if the orbitals haves different principal quantum number, increases. C-Br bond in $C_6H_5Br$ is much weaker than C-F bond in $C_6H_5F$.
4. The side-to-side overlap between two p-orbitals in much less effective and gives weak bond.

On the basis of concept of overlapping of atomic orbital, the covalent bond is divided into two parts.

a) Sigma bond or, σ-bond
b) Pi bond or π-bond

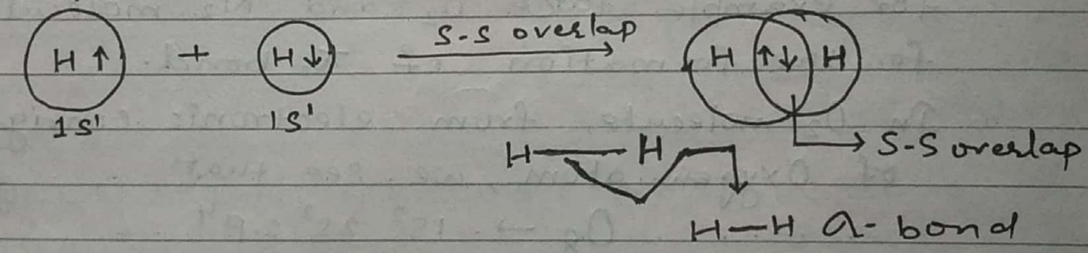
Sigma bond :→ It is denoted by σ and also known as single bond. A covalent bond is formed between two atoms by the overlap of their half filled atomic orbital along the line joining the nuclei of ~~H₂~~ both the atoms, then it is called Sigma bond. It means that the covalent bond is formed by the linear overlap or head to head overlap of the half filled atomic orbital of the two atoms.

Characteristics of Sigma bond (σ- bond) —

1. The boundary surface of σ-bond takes up an avoid shape.
2. The electron cloud of this bond is symmetrical about the bond axis.
3. This bond has two electrons which have opposite spins. These electrons are attracted equally by the nuclei of both atom

formation of σ- Bond in some molecules —

1. $H_2$ - molecule

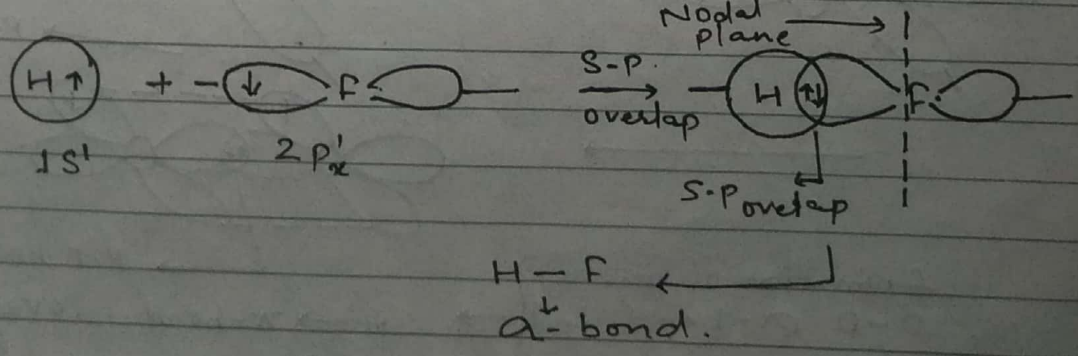



2. HF molecule

$H_1 \rightarrow 1s^1$ $\boxed{\uparrow}$ S - orbital

$F_9 \rightarrow 1s^2\ 2s^2\ 2p^5$

P-orbital $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$



Pi- Bond :→ It is denoted by $\pi$. It is always formed due to multiple bond. A covalent bond which is formed between two atoms by the overlap of their single filled p-orbitals along a line perpendicular to their nuclear axis, is called Pi-($\pi$)- Bond. It means that $\pi$ bond is obtained by the side-to-side overlap of half-filled p-orbitals of the two atoms.

Characteristics of $\pi$- bond —

1. This bond has one or only one nodal plane which contains the nuclear axis and divides it into two sausage like halves.
2. This bond has an increased electron density in the inter-nuclear region, though not on the bond axis.

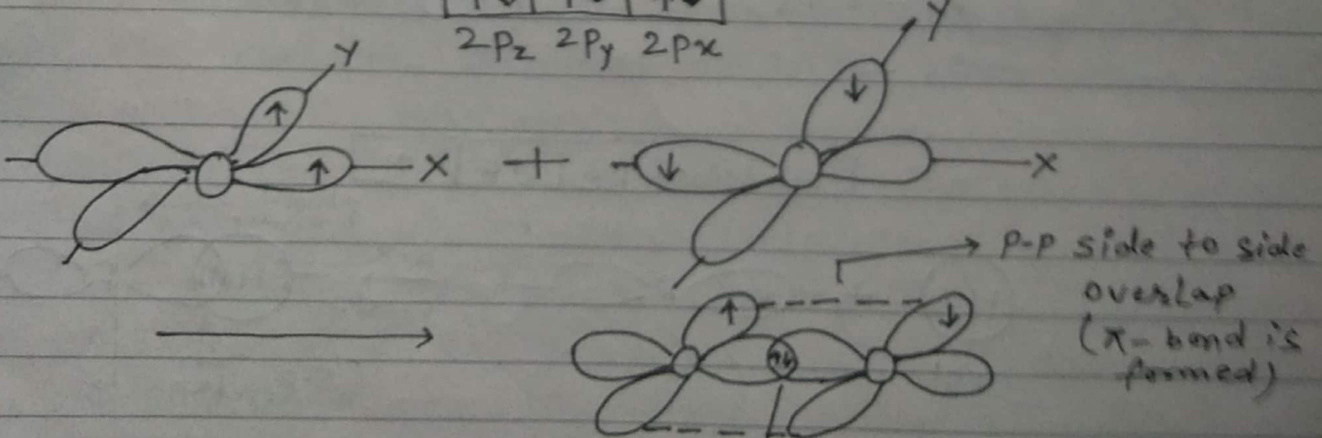
for example take $O_2$ and $N_2$ molecules, for the formation of $\pi$- bond.

In $O_2$ molecule, from electronic configuration of Oxygen atom, we see that

$$O_8 \rightarrow 1s^2\ 2s^2\ 2p^4$$

Consider the P-orbital, as its component,

| ↑↓ | ↑↓ | ↑↓ |
|---|---|---|
| $2p_z$ | $2p_y$ | $2p_x$ |

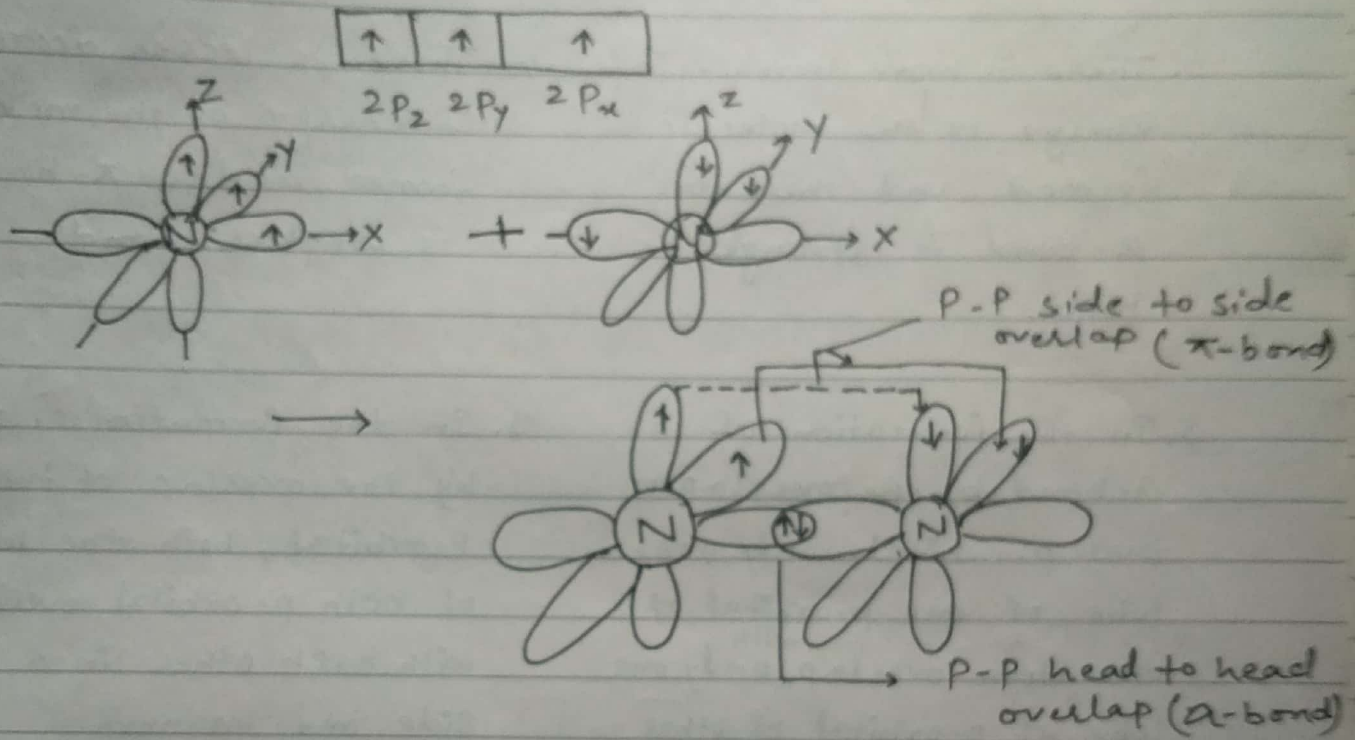



formation of one O-O σ-bond and one O-O π bond → $O = O$ (π-bond, σ- bond)

Now Consider in $N_2$ molecule, from electronic configuration of N-atom, we see that

$$N_7 \longrightarrow 1s^2\ 2s^2\ 2p^3$$

Consider the p orbital as its component.




formation of one N-N σ-bond and two N-N π bond.

π-bond.
N ≡ N
π-bond
σ-bond

NOTE:→ The double bond contains one π-bond and one σ-bond while in Triple bond contains two π-bond and one σ-bond. Whereas in single bond only σ-bond is obtained.

1. Single bond — (σ-bond)
2. Double bond = π-bond, σ-bond
3. Triple bond ≡ π-bond, σ-bond, π-bond

# DISTINGUISH BETWEEN SIGMA AND PI - BOND.

| SIGMA - BOND | PI - BOND |
|---|---|
| 1. It is obtained by the head to head overlap. | 1. It is obtained by the side to side overlap. |
| 2. There is more decrease in energy of the molecule formed and there fore σ-bond is stronger bond. | 2. There is lesser decrease in energy of the molecule formed and so π-bond is weaker bond. |
| 3. In the formation of σ-bond by the overlap of two p-orbitals only one lobe of one p-orbital of one atom overlap with only one lobe of p-orbital of other atom in head on manner. | 3. In the formation of π-bond by the overlap of two p-orbital, both the lobe of both p-orbital overlap with each other in a side one manner. |
| 4. The probability of finding the electrons between the two nuclei is maximum, since they are attracted equally by both the nuclei. so it is stronger bond. | 4. The probability of finding electrons between the two nuclei is poor and so it is poor bond. |
| 5. The electron density of this bond is distributed symmetrically about the nuclear axis. | 5. The electron density of this bond is distributed unsymmetrically about the nuclear axis. |
| 6. There can be free rotation of the atoms relative to one another round the bond axis of σ-bond. | 6. There is no possibility of free rotation of the atoms relative to one another round the bond axis of π-bond. |