CHEMICAL BONDING SHORT NOTES

Chemical Bonding

A force that acts b/w two or more atoms to hold them together as a stable molecule is known as chemical Bond.

Reason - a) To attain stability b) To follow octet Rule.

$$\text{Stability} \propto \frac{1}{\text{Energy}}$$

Characteristics of Ionic compounds

a) Hard crystalline solid
b) High MP/ BP
c) Highly soluble in polar solvent. (H_2O)

Ionic Bond : (Electrovalent Bond)

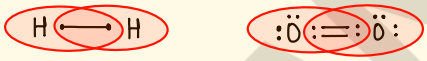
Electrostatic attraction b/w cation and anion.

Metal + Non-metal

Cation + Anion

Covalent Bond :

Bond formed by mutual sharing of electrons between combining atoms.



Bond pair - pair of electron participating in Bonding.
Lone pair - pair of electron not participating in Bonding.

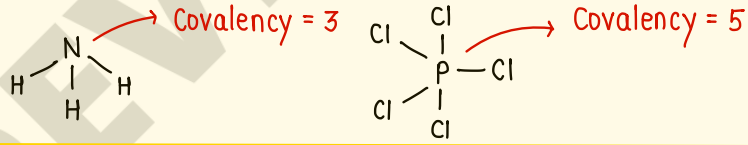
Electrovalency :

Charge on cation or anion in Ionic Bond.

NaCl Electrovalency of Na = +1
Electrovalency of Cl = -1

Covalency :

No. of Bonds formed by an atom in a covalent Bond.

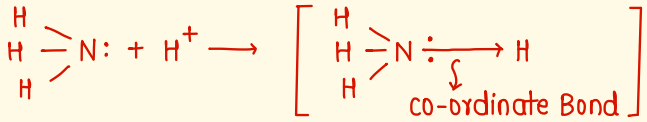


Favourable conditions for ionic Bonding :

a) Low ionisation energy of metal.
b) High electron gain enthalpy
c) High Lattice energy (Force of attraction)

Co-ordinate Bond :

Special type of covalent Bond in which Both the electrons are given by one atom and accepted By another.



Lattice energy :

Amount of energy released when one mole of crystal lattice is formed.

Lattice energy depends on -

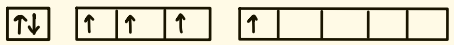
a) Charge L.E $\propto$ Charge b) Size of atoms - L.E $\propto \frac{1}{\text{Size}}$

Variable covalency :

No. of Bonds = No. of unpaired electrons

Only valence electron participate in Bonding.

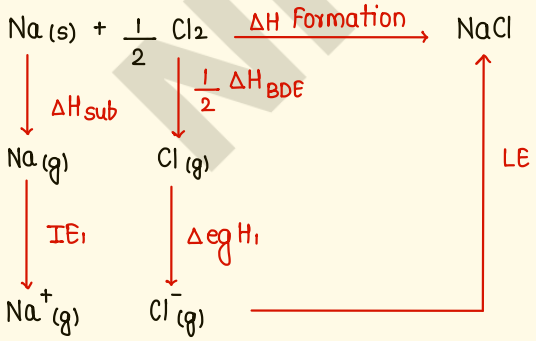
e.g. SO_2 S = [Ne] $3s^1 3p^3 3d^2$



4 unpaired electron = 4 Bond.

Born-Haber cycle :

Deals with the energy changes involved in formation of one mole of ionic crystal.



$$\Delta H_f = \Delta H_{sub} + IE_1 + \frac{1}{2}\Delta H_{BDE} + \Delta_{eg} H_1 + LE$$

Why to study Born-Haber ?

Because we cannot determine value of LE experimentally

Drawbacks of octet rule

a) electron deficient molecules (Hypovalent)
e.g. BF_3 and $BeCl_2$
b) Expansion of octet (Hypervalent)
e.g. PCl_5 and SO_3
c) odd electron species. e.g. NO_2

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CHEMICAL BONDING SHORT NOTES

Bond Parameters -

a) Bond order (BO)

Number of Bonds between two combining atoms.

H-H B·O = 1
C-C B·O = 1
C=C B·O = 2
C≡C B·O = 3

O_3 B·O = 1·5
C_6H_6 B·O = 1·5
} Due to Resonance

b) Bond length (BL)

Distance b/w nucleus of two atoms

c) Bond energy (BE)

Energy required to dissociate any Bond.

$$BL \propto \text{Size of atoms} \propto \frac{1}{\text{No. of Bonds}}$$

B·L - H-F < H-Cl < H-Br < H-I

B·L - C-C > C=C > C≡C
1·54 Å 1·34 Å 1·2 Å

$$B.E \propto \text{No. of Bonds} \propto \frac{1}{\text{Size of atom}}$$

B·E - H-F > H-Cl > H-Br > H-I

B·E - C-C < C=C < C≡C

B·E - C-C > N-N > O-O > F-F

B·E - Cl-Cl > Br-Br > F-F > I-I

B·E - S-S > O-O

d) Bond angle (Hybridisation)

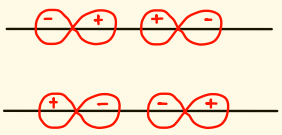
Angle b/w two adjacent Bonds.

Valence Bond theory (VBT)

Stronger the covalent Bond
↓
More is B·E
↓
Less is B·L

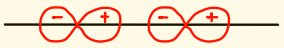
Overlapping :

a) +ve overlapping



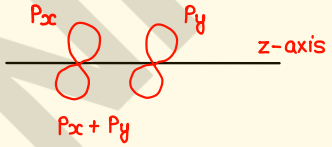
Same sign lobes overlap
Bond is Formed.

b) -ve overlapping



Opposite sign lobes overlap.
No Bond Formation.

c) Zero Overlapping.



effective overlapping zero.
No Bond Formation.

- More close is the overlapping orbital to nucleus, more extent of overlapping and stronger is bond
  1S - 1S > 2S - 2S > 3S - 3S
- More the directional nature of overlapping orbitals more is the extent of overlapping.
  2S - 2S < 2S - 2p < 2p - 2p

Axial overlapping or End to End Overlapping :

- Overlapping along inter nuclear axis.
- Sigma Bond is Formed.
- S-orbital always undergo axial overlapping.

Parallel overlapping or Side By side or Lateral overlapping or Sideways overlapping.

- Overlapping above and below inter-nuclear axis.
- π-Bond is Formed.
- First Bond Formed is always σ.

Difference between sigma and pi Bond :

Sigma Bond	Pi Bond
1) Formed by axial overlapping	1) Formed by parallel overlapping
2) Stronger Because extent of overlapping is more.	2) Weaker because extent of overlapping is less.
3) Allowed overlapping s-s, s-p, and p-p.	3) Allowed overlapping p-p, p-d and d-d.
4) Free Rotation possible.	4) Free rotation not possible

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CHEMICAL BONDING SHORT NOTES

Bond angle -

Angle b/w two Bonds.

Order of repulsion - LP-LP > LP-BP > BP-BP

Factors affecting Bond angle -

a) Lone pair

$$\text{Bond angle} \propto \frac{1}{\text{No. of lone pair}}$$

b) Electronegativity of the central atom.

Bond angle $\propto$ EN of central atom.

c) Size of side atom

Bond angle $\propto$ Size of side atom.

Hybridisation :

Overlapping of Hybrid orbitals.

Hybrid orbitals are identicals in size, shape and energy.
All types of orbitals i.e. empty, half-filled or fully filled orbital take part in hybridisation.
Hybrid orbitals orient themselves in space such that repulsion b/w them is minimum.
Order of repulsion - LP-LP > LP-BP > BP-BP

How to calculate Hybridisation

$$\text{Steric number} = \frac{1}{2}(V+M-C+A)$$

V = No. of valence e^- in central atom
M = No. of Monovalent atoms (H, F, Cl, Br, I)
C = Cationic charge
A = Anionic charge.

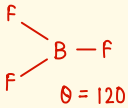
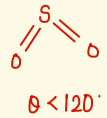
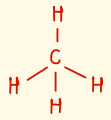
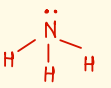
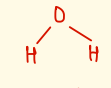
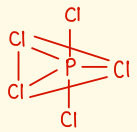
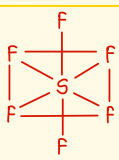
SN = No. of BP + No. of LP

Note - π Bond plays no role in hybridisation
No. of σ Bond = No. of surrounding atom.

How to find central atom

a) Least EN atom.
b) Hydrogen can never be central atom

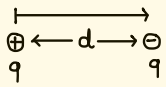
Polar molecule = $\mu_{net} \neq 0$
Non-polar molecule = $\mu_{net} = 0$
$\cos\theta \uparrow = \mu_{net} \uparrow = \theta \downarrow$

Example	Hyb	Geometry	Shape	structure
$BeCl_2$	SN = $\frac{1}{2}(2+2-0+0)$ = 2 — sp B.P = 2 L.P = 0	Linear	Linear	Cl - Be - Cl $\theta = 180^\circ$
BF_3	SN = $\frac{1}{2}(3+3-0+0)$ = 3 — sp^2 B.P = 3 L.P = 0	Trigonal planar	Trigonal planar	 $\theta = 120^\circ$
SO_2	SN = $\frac{1}{2}(6+0-0+0)$ = 3 — sp^2 B.P = 2 LP = 1	Trigonal planar	V-shape or Bent	 $\theta < 120^\circ$
CH_4	SN = $\frac{1}{2}(4+4-0+0)$ = 4 - sp^3 BP = 4 L.P = 0	Tetrahedral	Tetrahedral	 $\theta = 109^\circ 28'$
NH_3	SN = $\frac{1}{2}(5+3-0+0)$ = 4 - sp^3 BP = 3 L.P = 1	Tetrahedral	pyramidal	 $\theta < 109^\circ 28'$
H_2O	SN = $\frac{1}{2}(6+2-0+0)$ = 4 - sp^3 B.P = 2 L.P = 2	Tetrahedral	V-shape or Angular or Bent	 $\theta < 109^\circ 28'$
PCl_5	SN = $\frac{1}{2}(5+5-0+0)$ = 5 — sp^3d B.P = 5 LP = 0	Trigonal Bipyramidal	Trigonal Bipyramidal	
SF_4	SN = $\frac{1}{2}(6+4-0+0)$ = 5 — sp^3d B.P = 5 L.P = 1	Trigonal Bipyramidal	See-saw	
ClF_3	SN = $\frac{1}{2}(7+3-0+0)$ = 5 — sp^3d B.P = 3 L.P = 2	Trigonal Bipyramidal	T-shape	
XeF_2	SN = $\frac{1}{2}(8+2-0+0)$ = 5 — sp^3d B.P = 2 L.P = 3	Trigonal Bipyramidal	Linear	
SF_6	SN = $\frac{1}{2}(6+6-0+0)$ = 6 — sp^3d^2 B.P = 6 L.P = 0	Square Bipyramidal or octahedral	Square Bipyramidal or octahedral	

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CHEMICAL BONDING SHORT NOTES

Dipole moment (μ):



$\mu = q \times d$

Vector quantity: (+ve to -ve)
(less EN to More EN)

H – Cl $\quad \mu \neq 0$

Cl – Cl $\quad \mu = 0$

More the EN, more is dipole moment.

μ = H-F > H-Cl > H-Br > H-I

μ = C-Cl > C-F > C-Br > C-I

μ = CCl_4 < $CHCl_3$ < CH_2Cl_2 < CH_3Cl

μ = NH_3 > NF_3

Application of Dipole moment:

1) To determine whether a molecule is polar or non polar.
2) To Find % ionic character in covalent Bond.

$$\% \text{ ionic character} = \frac{\text{Experimental } \mu}{\text{Theoretical } \mu} \times 100$$

Fajan's Rule:

To Find covalent character in Ionic Bond.

Covalent character $\propto$ Polarisation.

- Factors affecting polarisation:

a) Size – For max. covalent character
Cation size should be small
Anion size should be large.

LiCl > NaCl > KCl > RbCl > CsCl

b) charge
For max. covalent character
Charge on cation and anion should be more.

NaCl < $MgCl_2$ < $AlCl_3$ < $SiCl_4$

c) Electronic configuration of cation.
Cations having pseudo noble gas configuration has more covalent character.

CuCl > NaCl

Drawbacks of VBT:

a) According to VBT, Oxygen molecule was Diamagnetic But in practical, it was formed to be paramagnetic

b) Fractional Bond order was not explained.

c) Stability of molecule was not decided by VBT.

- Main points of MOT:

a) In atoms, electrons are present in atomic orbitals

b) In molecules, electrons are present in molecular orbitals.

c) No. of molecular orbitals = Number of atomic orbitals combined.

When two atomic orbitals combine = 2 molecular orbitals are formed.

a) Bonding molecular orbital (BMO)
b) Anti-Bonding molecular orbital (ABMO)

MO is under the influence of two or more than two nuclei.
Electrons are Filled according to Afbau's Rule, Hund's Rule of max. multiplicity and pauli's exclusion principle.
MO are formed by Linear combination of atomic orbitals.

A and B $\quad$ LCAO

$\Psi_A \quad \Psi_B$

$\Psi_{BMO} = \Psi_A + \Psi_B$ constructive interference

$\Psi_{ABMO} = \Psi_A - \Psi_B$ Destructive interference

$\Psi_{MO} = \Psi_A \pm \Psi_B$

BMO (Bonding Molecular orbital)

Formed by constructive interference.
Lower energy and more stability.
Increases electron density in region b/w 2 nuclei
↳ may or may not have node.

$\Psi_{BMO} = \Psi_A + \Psi_B$

$\Psi^2_{BMO} = \Psi_A^2 + \Psi_B^2 + 2\Psi_A\Psi_B$

ABMO (Anti-Bonding Molecular orbital)

Formed by destructive interference.
Higher energy and less stability.
Electron density decreases b/w 2 nuclei.

$\Psi_{ABMO} = \Psi_A - \Psi_B$

$\Psi^2_{ABMO} = \Psi_A^2 + \Psi_B^2 - 2\Psi_A\Psi_B$

Formal charge

$$F.C = \text{Valence } e^- - [\text{Non-Bonding } e^-] - \frac{1}{2}(\text{Bonding } e^-)$$



$S = 6 - [3 + 2]$

$= +1$

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CHEMICAL BONDING SHORT NOTES

Types of Molecular orbital :

a) σ MO - Formed by overlapping along internuclear axis.

σ 1S	σ^* 1S	σ 2S	σ^* 2S	σ 2Pz	σ^* 2Pz
BMO	ABMO	BMO	ABMO	BMO	ABMO

b) π MO - Overlapping is above and below inter-nuclear axis

π 2Px π^* 2Px π 2Py π^* 2Py

Electronic configuration of molecules :

Case 1 : No. of e^- $\leq$ 14 (π 2Px = π 2Py) σ 2Pz

Case 2 : No. of e^- > 14 σ 2Pz (π 2Px = π 2Py)

Na = No. of e^- in ABMO

Nb = No. of e^- in BMO

$B.O = \frac{1}{2}(N_b - N_a)$

B.O > 0 $\longrightarrow$ molecule exists

B.O $\leq$ 0 $\longrightarrow$ molecule does not exist.

Stability of molecule

more the B.O = more is stability.

If B.O is same, then lower no. of e^- is ABMO, Higher is stability.

$B.O \propto B.E \propto \frac{1}{B.L}$

Diamagnetic - zero unpaired e^-

Paramagnetic - Atleast one unpaired e^-.

Trick

10 e^- $\longrightarrow$	B.O = 1
11 e^- $\longrightarrow$	B.O = 1.5
12 e^- $\longrightarrow$	B.O = 2
13 e^- $\longrightarrow$	B.O = 2.5
14 e^- $\longrightarrow$	B.O = 3
15 e^- $\longrightarrow$	B.O = 2.5
16 e^- $\longrightarrow$	B.O = 2
17 e^- $\longrightarrow$	B.O = 1.5
18 e^- $\longrightarrow$	B.O = 1

Exception - CO^+ = 13 e^-

B.O = 3.5

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