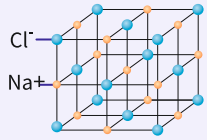
4. CHEMICAL BONDING AND MOLECULAR STRUCTURE

TYPES OF BONDS

IONIC/ELECTROVALENT BOND

- Strong electro static force of attraction between positive and negative ions.
- Crystalline in nature
- High M.P and B.P
- Soluble in Polar Solvents.
- Eg: NaCl, $MgCl_2$ etc.



NaCl structure

CO-ORDINATE BONDING

- Bond formed by one sided sharing of electrons. i.e, one atom donates a Pair of e^- while other accepts it.
- Bad conductors of electricity.
- Sparingly soluble in polar solvents but readily soluble in non-polar solvents. Eg:- NH_4+

COVALENT BOND

- Bond formed by mutual sharing of e^-.
- Low M.P. and B.P.
- Bad conductor of electricity
- Insoluble in Polar Solvents but soluble in non-polar solvent.
- Ex: CH_4, H_2, Cl_2.

TYPE OF COVALENT BOND

- Polar covalent bond. Eg: NH_3, $CHCl_3$
- Non-polar covalent bond. Eg: Cl_2, CO_2.

HYDROGEN BOND

- Bond formed when the -ve end of one molecule attracts the +ve end of H.

1. **Intramolecular**: H^- Bonding occur within one single molecule.
2. **Intermolecular**: H Bonding between two different molecules of same or different compounds.

THEORIES OF CO-VALENT BOND

KOSSEL LEWIS APPROCH

Atoms can combine either by transfer of e^- or by sharing of valence e^- in order to have a complete octet in their valence shell.

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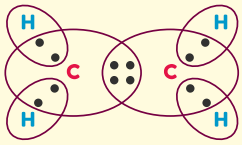
Octet Rule



LEWIS SYMBOLS

Valence e- are represented by dots around the element.

Formal Charge

$$FC = V - N - \frac{B}{2}$$

BOND PARAMETERS

Dipole Moment

Product of the magnitude of the charge and distance between centres of positive and negative charge.

μ = charge × distance between two atoms

Bond length: distance between the nuclei of two bonded atom.

$$\text{bond length} \propto \frac{1}{\text{bond order}}$$

Fajan's Rule

No compounds is 100% ionic or 100% covalent.

- Covalent nature ∝ Charge on cation
- $\text{Covalent nature} \propto \frac{1}{\text{size of cation}}$

Bond Enthalpy

Amount of energy required to break one mole of bonds.

VSEPR THEORY

The shape of a molecule depends upon the numbers of valence shell e^- (B.P or l.p) surrounding in the central atom.

Decreasing order of repulsive interaction: lp - lp > lp - bp > bp - bp

Type of Molecule	No. of Bonding pair	No. of Lone Pair	Arrangement of e^- pair	Shape	Example
AB_2E	2	1	Trigonal planer	Bent	SO_2 ,O_3
AB_3E	3	1	Tetrahedral	Trigonal Pyramidal	NH_3
AB_3E_2	2	2	Tetrahedral	Bent	H_2O
AB_4E	4	1	Trigonal bi-pyramidal	See saw	SF_4
AB_3E_2	3	2	Trigonal bi-pyramidal	T-Shape	ClF_5
AB_5E	5	1	Octahedral	Square Pyramid	XeF_5
AB_4E_2	4	2	Octahedral	Square Planner	XeF_4

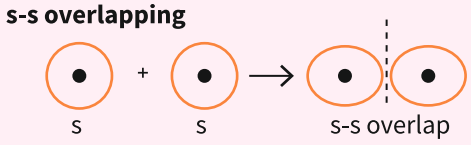
VALENCE BOND THEORY (VBT)

A covalent bond is formed by the overlapping of two half filled atomic orbitals.

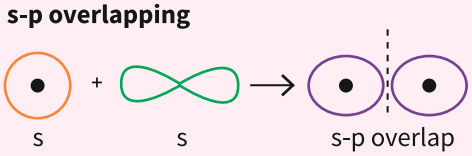
Type of overLapping

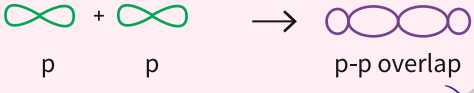
Sigma(σ) Sidewise overlapping

s-s overlapping

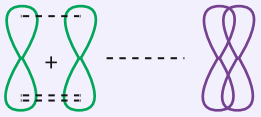


s-p overlapping





Pi(π) Sidewise overlapping



HYBRIDISATION

Concept of mixing atomic orbital to form new hybrid.

Formation of Molecular Orbitals

Atomic orbitals	Molecular orbitals	No. of nodal plane
	σ	0
	σ*	1
p_x p_x	π	1
p_x p_x	π*	2
p_z p_z	π	2
p_z p_z	π*	3

MOLECULAR ORBITAL THEORY

- MOT states that each atom tends to combine together and form molecular orbitals
- No. of molecular orbitals = no. of atomic orbital combined.

MOT

Bonding molecular orbitals | **Anti bonding atomic orbitals**

ELECTRONIC CONFIGURATION

Electron filling order upto 14 electrons

$$\sigma_{1s} < \sigma^*_{1s} < \sigma_{2s} < \sigma^*_{2s} < \pi_{2px} \equiv \pi_{2py} < \sigma_{2pz} < \pi^*_{2px} \equiv \pi^*_{2py} < \sigma^*_{2pz}$$

Electron filling order for more than 14 electrons

$$\sigma_{1s} < \sigma^*_{1s} < \sigma_{2s} < \sigma^*_{2s} < \sigma_{2pz} < \pi_{2px} \equiv \pi_{2py} < \pi^*_{2px} \equiv \pi^*_{2py}$$