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.

HYDROGEN BOND

- Bond formed when the -ve end of one molecule attracts the +ve end of other
1. Intermolecular: H^- Bonding occur within one single molecule.
2. Intermolecular: H Bonding between two different molecules of same or diff. compounds.

CO-VALENT 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.

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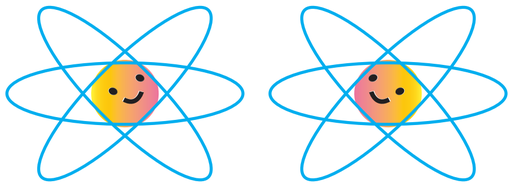


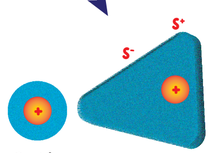
TYPE OF CO-VALENT BOND

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

FAJAN'S RULE

No compounds is 100% ionic or 100% covalent

- Covalent Nature $\propto$ Size of anion
- Covalent Nature $\propto \frac{1}{\text{Size of cation}}$



Cation Polarized anion

Polarization of anion by cation

THEORIES OF CO-VALENT BOND

KOSEL LEWIS APPROCH

KOSSEL LEWIS APPROACH

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

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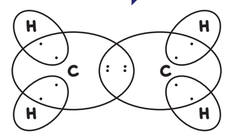
Octet Rule

LEWIS SYMBOLS

Valence e- are represented by dots around the element.

H: Na: :$\ddot{Cl}$: ·$\ddot{O}$· . :N:



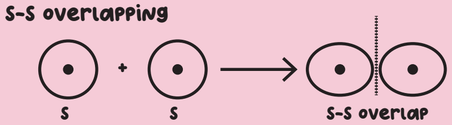
VALENCE BOND THEORY VBT

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

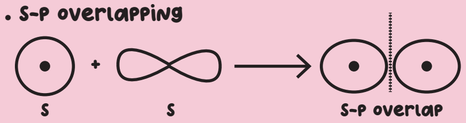
Type of overlapping

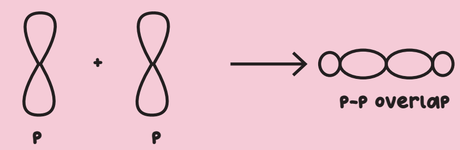
Sigma(σ) Sidewise overlapping

- S-S overlapping

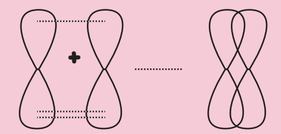


- S-P overlapping





Pi(π) Sidewise overlapping



Decreasing order of repulsive interaction:

lp - lp > lp - bp > bp . bp

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

HYBRIDISATION

Concept of mixing atomic orbital to form new hybrid

MOLECULAR ORBITAL THEORY

- MOT are formed the combinate of atomic orbitals.
- No. of molecular orbitals = No. of atomic orbital combined.

MOT

- Bonding molecular orbitals
- Anti bonding atomic orbitals

ELECTRONIC CONFIGURATION

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$(\sigma 1S < \sigma^*1S < \sigma 2S < \sigma^*2S < \sigma 2P_z < (\pi 2P_x = \pi 2P_x) < (\pi^*2P_x = \pi^*2P_y) < \sigma^*2P_z$

(For O_2, F_2, Ne_2)

$(\sigma 1S < \sigma^*1S < \sigma 2S < \sigma^*2S < (\pi 2P_x = \pi 2P_y) < \sigma 2P_z) < (\pi^*2P_x = \pi^*2P_y) < \sigma^*2P_z$

(For B_2, C_2, N_2)

BOND PARAMETERS

FORMAL CHARGE

$$F.C. = r - l - \frac{1}{2}B$$

- Bond length: Equilibrium distance between the nuclei of two bonded atom

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

BOND ORDER & BOND ANGLE

- No. of Bond between the two atoms
- Angle between the orbitals containing bonding e^- pair around central atom.

BOND ENTHALPY

Amount of energy required to break one mole of bonds.

DIPOLE MOMENT

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

M = charge × Distance of seperation

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_2O_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

Formation of Molecular Orbitals