

TYPES OF BONDS

IONIC/ELECTROVALENT BOND

- Strong electrostatic force of attraction between positive and negative ions.
- Crystalline in nature
- High M.P and B.P
- Soluble in polar solvents. Eg: NaCl, MgCl₂ etc.

HYDROGEN BOND

- Bond formed when the -ve end of one molecule attracts the +ve end of other
- 1. Intermolecular: H-Bonding occurs 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. Eg: CH₄, H₂, Cl₂.

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THEORIES OF CO-VALENT BOND

KOSSEL LEWIS APPROACH

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

↓
Octet Rule

LEWIS SYMBOLS

Valence e⁻ are represented by dots around the element.



VALENCE BOND THEORY

VBT

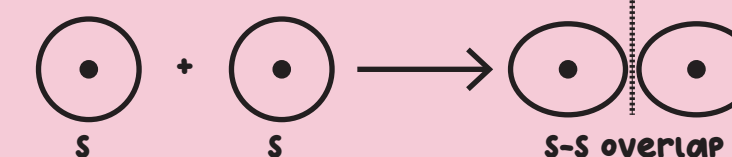
A covalent bond is formed by the overlapping of two half filled atomic orbitals.

TYPE OF OVERLAPPING

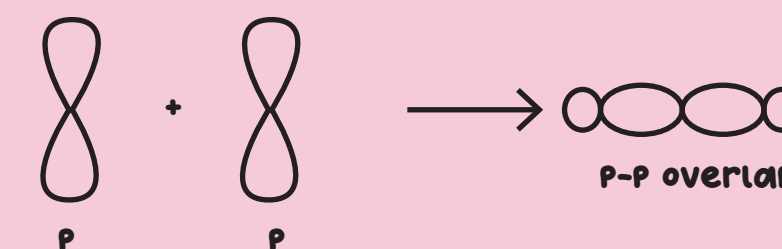
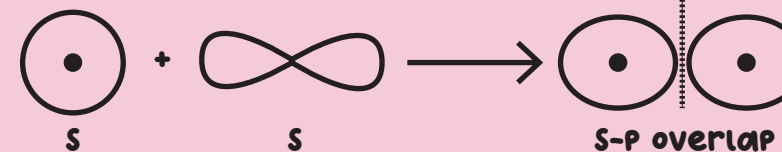
Sigma(σ)

Sidewise overlapping

S-S overlapping

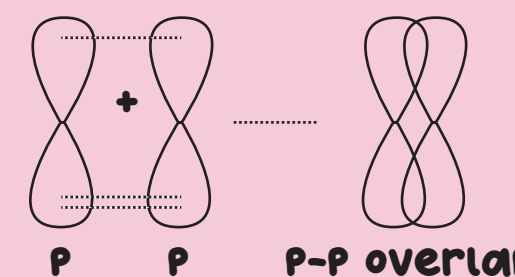


S-P overlapping



Pi(π)

Sidewise overlapping



Decreasing order of repulsive interaction:
IP - IP > IP - BP > BP . BP

HYBRIDISATION

Concept of mixing atomic orbital to form new hybrid

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

MOLECULAR ORBITAL THEORY

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

MOT

Bonding molecular orbitals

Anti bonding atomic orbitals

ELECTRONIC CONFIGURATION

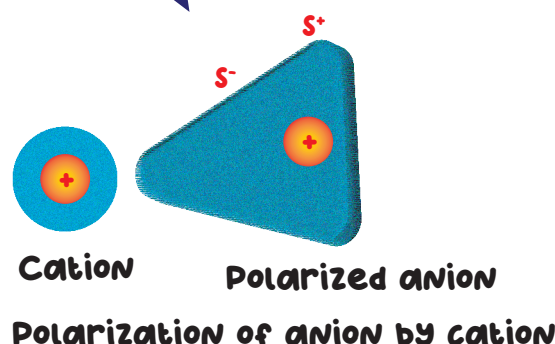
< (σ1s < σ*1s < σ2s < σ*2s < σ2p_z < (π2p_x = π2p_y) < (π*2p_x = π*2p_y) < σ*2p_z (For O₂, F₂, Ne₂)

(σ1s < σ*1s < σ2s < σ*2s < (π2p_x = π2p_y) < σ2p_z) < (π*2p_x = π*2p_y) < σ*2p_z (For B₂, C₂, N₂)

CHEMICAL BONDING AND MOLECULAR STRUCTURE

FAJAN'S RULE

- No compounds is 100% ionic or 100% covalent
- Covalent nature ∝ Size of anion
- Covalent nature ∝ $\frac{1}{\text{Size of cation}}$



FORMAL CHARGE

F.C. = $r - l - \frac{1}{2}B$
Bond length: Equilibrium distance between the nuclei of two bonded atom
bond length ∝ $\frac{1}{\text{bond order}}$

BOND PARAMETERS

BOND ORDER

No. of Bond between the two atoms

BOND ANGLE

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 Separation

TYPE OF MOLECULE	No. of Bonding Pair	No. of Lone Pair	Arrangement of e ⁻ pair	SHAPE	EXAMPLE
AB ₂ E	2	1	Trigonal planar	Bent	SO ₂ , O ₃
AB ₃ E	3	1	Tetrahedral	Trigonal pyramidal	NH ₃
AB ₃ E ₂	2	2	Tetrahedral	Bent	H ₂ O
AB ₄ E	4	1	Trigonal bi-pyramidal	See saw	SF ₄
AB ₃ E ₂	3	2	Trigonal bi-pyramidal	T-Shape	ClF ₃
AB ₅ E	5	1	Octahedral	Square pyramid	XeF ₅
AB ₄ E ₂	4	2	Octahedral	Square planar	XeF ₄

Formation of Molecular Orbitals

