

CHEMICAL BONDING..

Chemical Bond

Atom → do not exist independently (except noble gas)

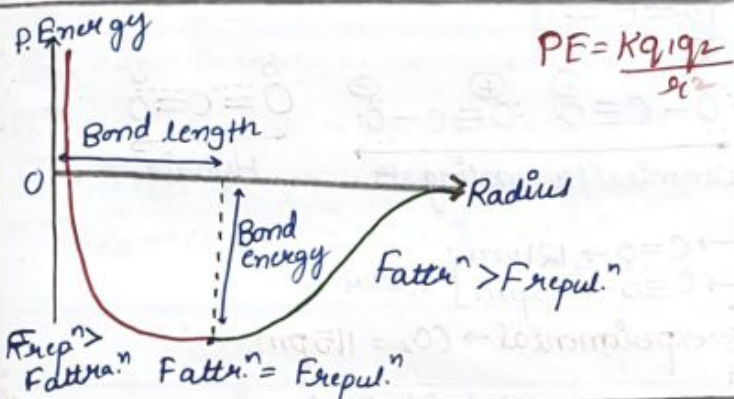
- Group of atom exist → **Molecule**.
- Attractive force which holds atom, ion together → **Chemical Bond**.
- Attraction ↑, energy ↓, Stability ↑.

Cause of Chemical Bond →

Tendency to acquire min. energy (potential energy)



Repulⁿ (+) ↔ (+) } attraction
 Attraction (-) ↔ (-)



Types of Chemical Bond →

Strong Bond

Weak Bond

(200-400 KJ/mol)

(2-40 KJ/mol)

Interatomic bond.

Intermolecular bond.

- Covalent Bond
- Ionic Bond

- VanderWall F.O.A.
- Hydrogen Bond

- Metallic (x)
- Coordinate (x)

Covalent Bond.

- Sharing of valence e^- takes place to complete its octate/duet.
- Bond formed b/w non-metal - non-metal.

→ Dots → Valence e^- .

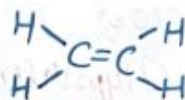
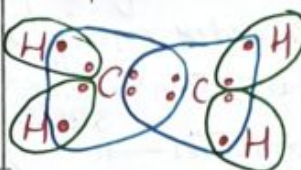
→ Structure c/a → Lewis Structure.

e.g. - Cl_2 , H_2O , CCl_4 , CO_2

→ Bond pair + lone pair = 4 (Octate Rule)

C_2H_4

C_2H_2



Theories of Covalent Bond.

① Lewis Theory

atom → Kernel (Nucleus + inner e^-)
 → Outermost Shell (max 8 e^- occupy)

These 8 e^- occupy corners of a cube, which surround Kernel.



Octet Rule

electronic theory of Covalent bond

→ 'atoms combine to complete its octate in outermost orbit'.

Sharing of e^- .

Cl_2 , N_2 , O_2

Complete transfer of e^- .

$NaCl$, $CaCl_2$, MgO .

Lewis dot Structure: →

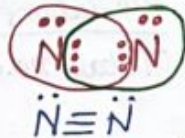
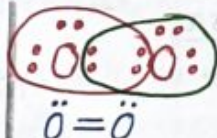
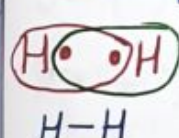
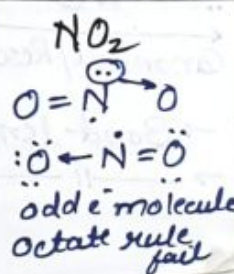
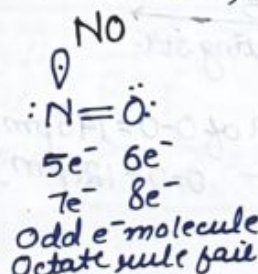
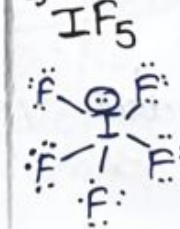
→ Identify Central atom (less no., least EN, high atomic no., large size, forms multiple bond).

Li Be B C N O F
 1 2 3 4 5 6 7

(3-(8-3)) P S Cl (1-7)
 (3-5) (2-(8-2))
 (2-5)

1 Bond = 3 L.P.
 2 B = 2 L.P.
 3 B = 1 L.P.
 ♥ ♥ ♥

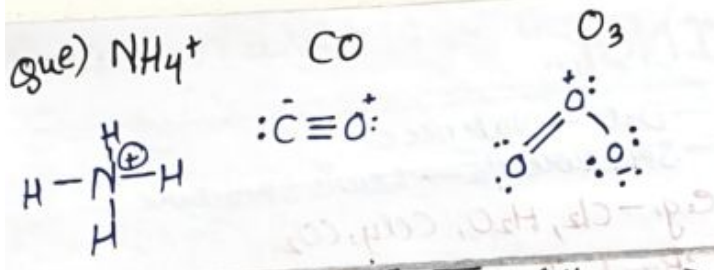
Que)



Single Covalent Bond (2e⁻)

Double Covalent Bond (4e⁻)

Triple Covalent Bond (6e⁻)

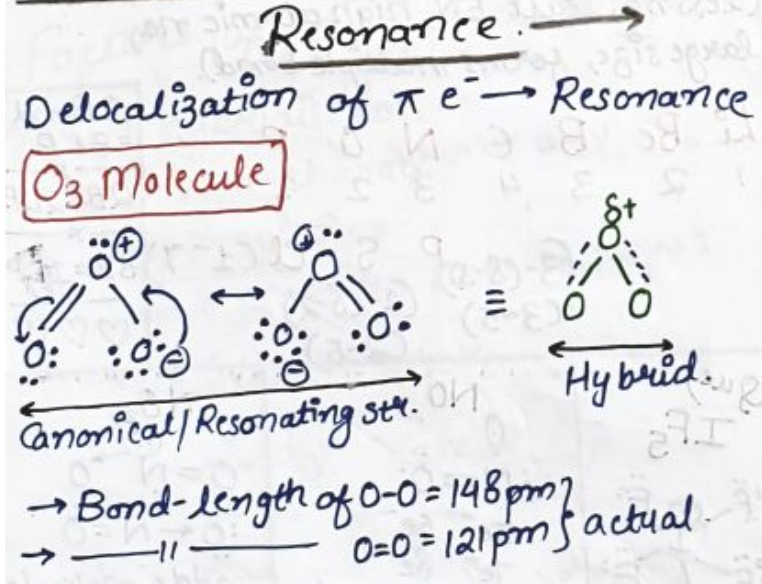
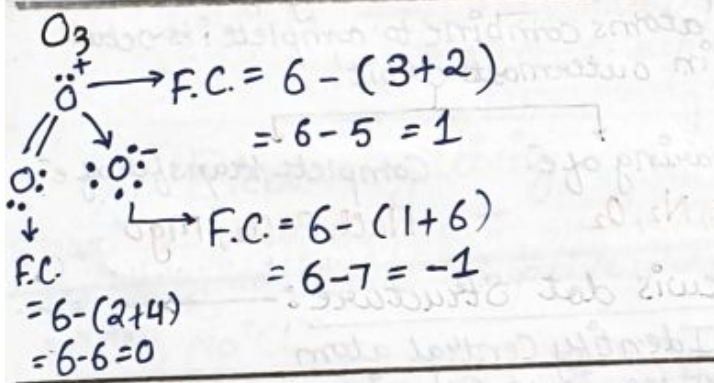


- Drawback of Lewis Theory →
- Could not explain structure of compound
 - $\text{Bond length, Order, energy etc}$
 - **HypoValent**. (less than $8e^-$ in valence shell)
 - **HyperValent** (more than $8e^-$ in valence shell)
 - Odd e^- molecule. ($\text{NO}, \text{NO}_2, \text{ClO}_2$)
 - $\text{LiCl}, \text{BeH}_2, \text{BCl}_3, \text{BF}_3, \text{AlCl}_3$, (**HypoValent**)
 - $\text{PCl}_5, \text{H}_2\text{SO}_4, \text{IF}_7$.. (**expanded Octate, HyperValent**)
 - Compounds of noble gases are also formed. ($\text{XeF}_2, \text{KrF}_2, \text{XeOF}_2, \text{XeO}_3, \text{XeOF}_4, \text{XeF}_8$)

Formal Charge

Charge on Individual atom in a compound → formal Charge

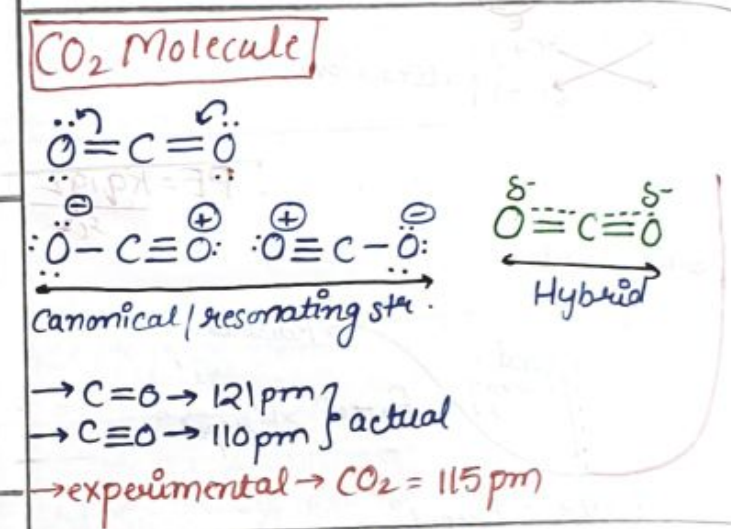
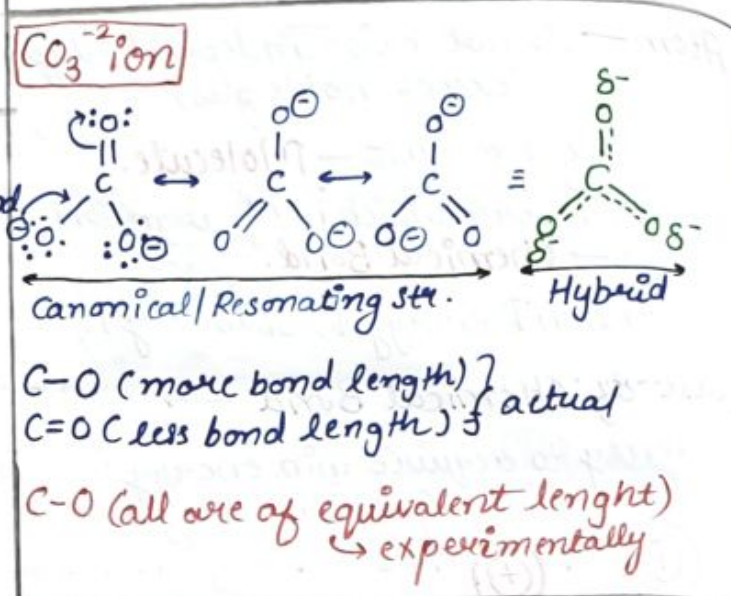
F.C. = Valence electron - (bonds + dots)



Experiment in $\text{O}_3 \rightarrow 128 \text{ pm}$ (explained by resonance)

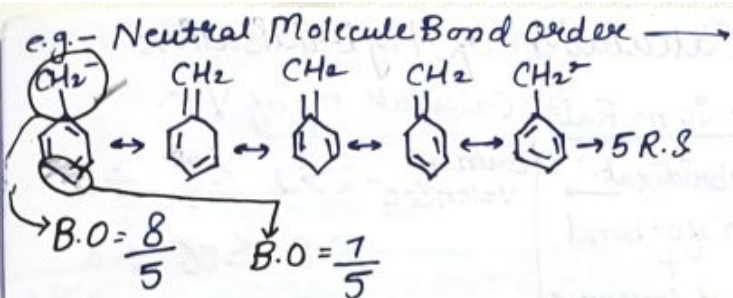
↓ Length of Single Bond ↓

Length of Double Bond ↑



- Resonance Hybrid = Real
- Canonical str. = Hypothetical
- Resonance hybrid has less energy than single canonical structure, So, resonance hybrid stabilizes the molecule.

Bond Order	
Neutral molecule	Charged Species
① Bond Order →	① Bond Order →
Single Bond = 1	= Valency of SA + Charge of comp
Double $\text{--}\text{--}$ = 2	na of SA
Triple $\text{--}\text{--}\text{--}$ = 3	
Bond order is resonance structure	
B.O. = $\frac{\text{Total no. of Bonds b/w 2 atoms in all str.}}{\text{Total no. of R.S.}}$	



e.g. - PO_4^{3-}	SO_4^{2-}	ClO_4^- (Charged Specific)
$\text{B.O.} = \frac{2 + (-3)}{4}$	$\text{B.O.} = \frac{2 + (-2)}{4}$	$\text{B.O.} = \frac{2 + (-1)}{4}$
$\boxed{1.25}$	$\boxed{1.5}$	$\boxed{1.75}$

Bond Order \propto Bond Strength \propto Bond energy
 $\rightarrow \propto \text{Stability} \propto \frac{1}{\text{Bond Length}}$

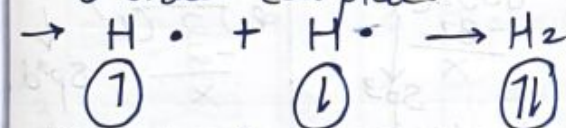
$\text{B.O.} = +\text{ve integer} = \text{Compound} \rightarrow \text{Stable}$
 $\text{B.O.} = 0 / -\text{ve} = \text{Compound} \rightarrow \text{Unstable}$
 $" = \text{fracti.} (>1) \rightarrow \text{Compound Shows Isomerism}$
 $" = \frac{1}{2} \rightarrow \text{Unstable, but exist} \rightarrow \text{H}_2^+, \text{He}_2^+$

Valence Bond Theory

Overlapping Theory

Hybridisation

\rightarrow Covalent bond is formed, when atomic orbitals complete.



\rightarrow Overlapping orbital \rightarrow Half filled, and opposite spin.

\rightarrow Extent of Overlap should be $< 50\%$

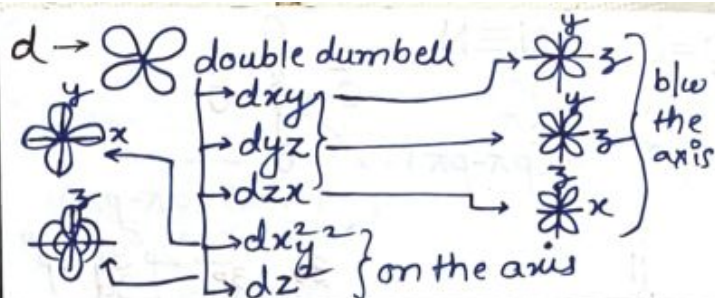
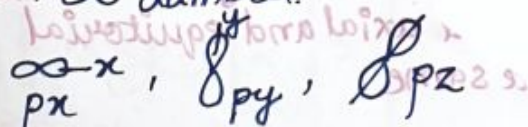
$\rightarrow > 50\% = \text{Repulsion Occurs.}$

\rightarrow Extent of Overlap \propto Strength of Bond
 $\propto \text{Stability} \propto \text{Bond energy} \propto \frac{1}{\text{Bond length}}$

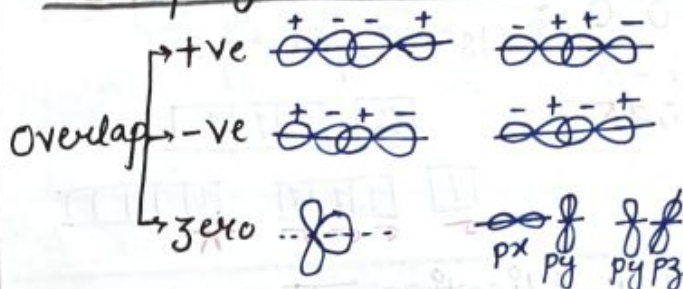
Atomic Orbital

$\text{S} \rightarrow \bigcirc$ spherical, non-directional.

$\text{P} \rightarrow \infty$ dumbbell.

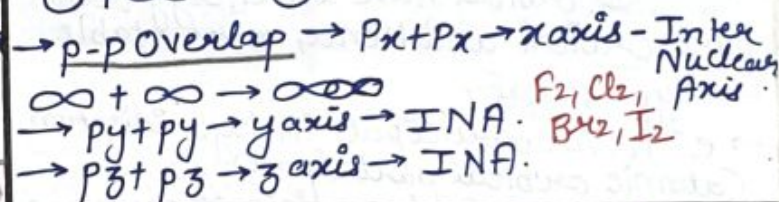
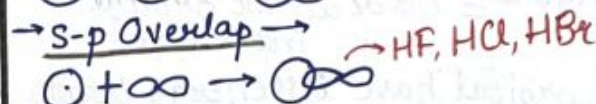
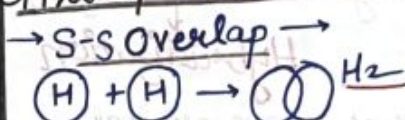


Overlap of atomic Orbital

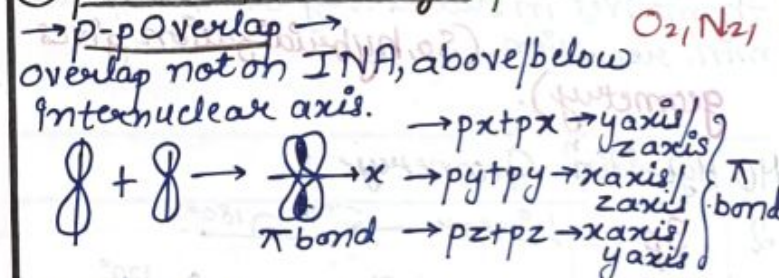


Overlap

① Axial / Head on Overlap (Sigma Bond)



② parallel / Sideways (pi bond) π



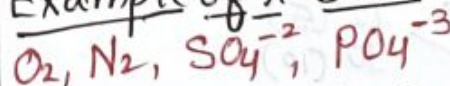
$\rightarrow \text{p-d Overlap} \rightarrow$



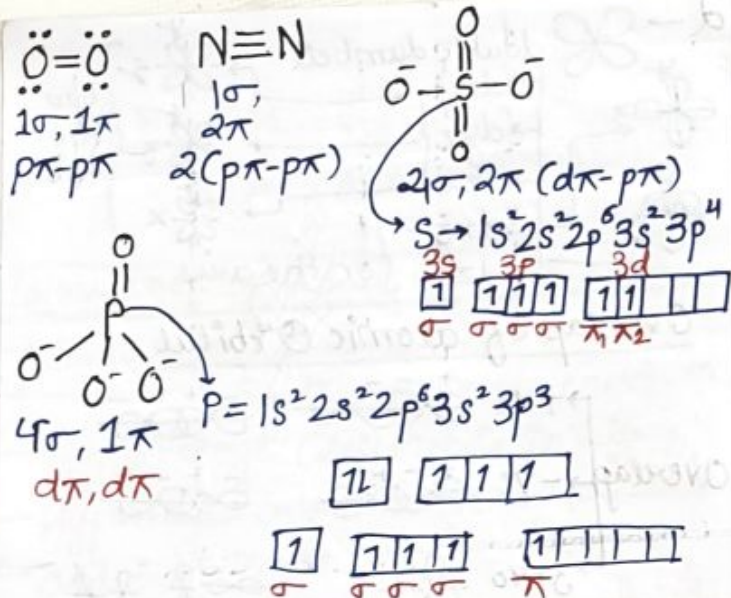
$\rightarrow \text{d-d Overlap} \rightarrow$



Example of π Bond \rightarrow



Calculate no. of π - π and π - σ bonds.



Calculation of Hybridisation

Steric no. Rule
 Hybridization \rightarrow
 $= \text{no. of } \sigma \text{ bond} + \text{no. of lone pair.}$

Calculate no. of Ve^-
 Sum of Valence e^- $> 2 \leq 8 \div 2$
 $> 8 \leq 58 \div 8$
 $> 56 \div 18$
 no. of lone pair of Central atom
 = Quotient - Surrounding atom

Hybridisation Theory

\rightarrow for Bond formation
 \rightarrow before overlapping, intermixing of orbitals occurs.

\rightarrow no. of hybrid orbital = no. of atomic orbital intermixed.

\rightarrow Hybrid orbital have less energy than pure orbital and hence, more stable.

\rightarrow Hypothetical.

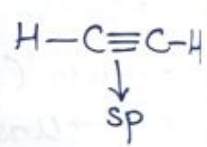
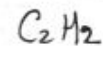
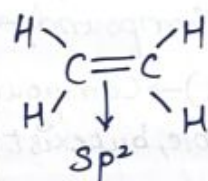
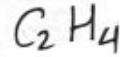
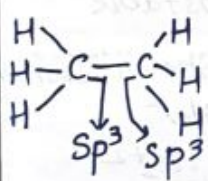
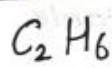
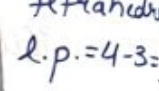
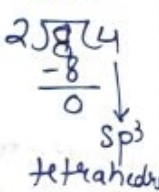
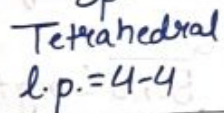
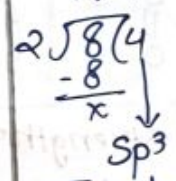
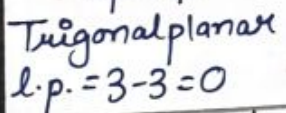
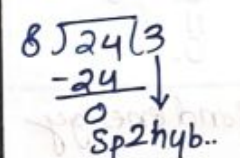
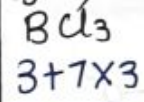
$\rightarrow e^-$ never participate in hybridisation, Atomic orbital participate.

\rightarrow Hybrid orbital participate arrange themselves in such a way that, they have min. repulsion. (So, hybridization gives geometry).

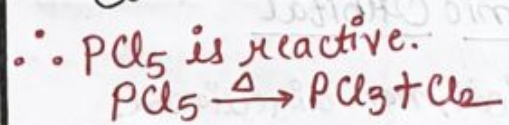
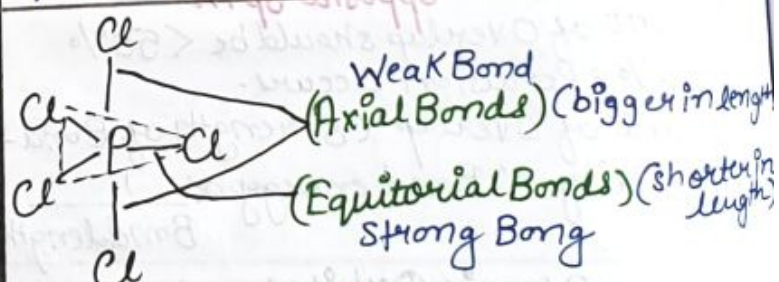
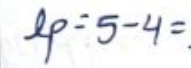
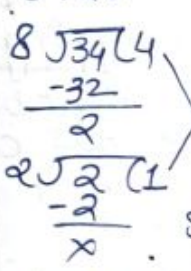
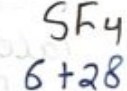
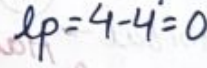
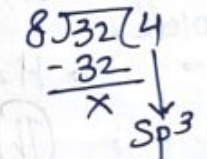
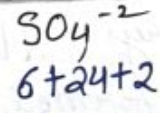
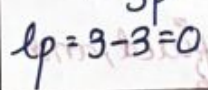
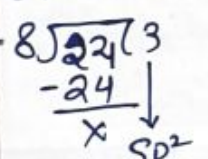
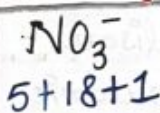
HO	Hybridisation	Geometry
2	sp	Linear $\rightarrow 180^\circ$
3	sp^2	Trigonal Planar $\rightarrow 120^\circ$
4	sp^3	Tetrahedral $\rightarrow 109.5^\circ$
5	sp^3d	Trigonal bipyramidal $\rightarrow 90^\circ(6), 120^\circ(3)$
6	sp^3d^2	Octahedral $\rightarrow 90^\circ(12)$
7	sp^3d^3	Pentagonal bipyramidal $\rightarrow 72^\circ(5), 90^\circ(10)$



Que)

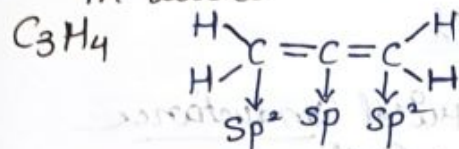


(-) charge \rightarrow addition
 (+) charge \rightarrow subtract



$\rightarrow PF_5 \rightarrow F \rightarrow$ more electronegative atom.
 and in PF_5 all axial and equatorial bonds are same.

Calculate hybridization of all Carbon in allene \rightarrow



VSEPR Theory (Valance Shell Electron Pair Repulsion Theory).

<u>CH₄</u>	<u>NH₃</u>	<u>H₂O</u>
Ve = 8	8	8
Q = 4	4	4
sp^3	sp^3	sp^3
Tetrahedral	Tetrahedral	Tetrahedral
109.5°	109.5°	109.5°
Theoretical	\rightarrow lp-bp \rightarrow repuls. ⁿ	\rightarrow lp-lp repuls. ⁿ
109.5°	107°	104.5°
experiment		
lp = 0	lp = 1	lp = 2

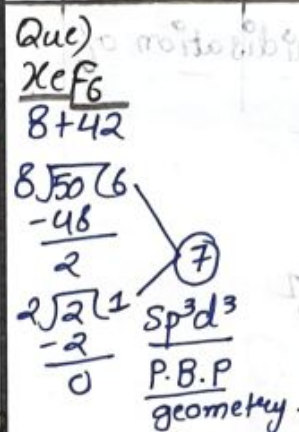
VSEPR says

lp-lp repulsion > lp-bp repuls.ⁿ > bp-bp repuls.ⁿ

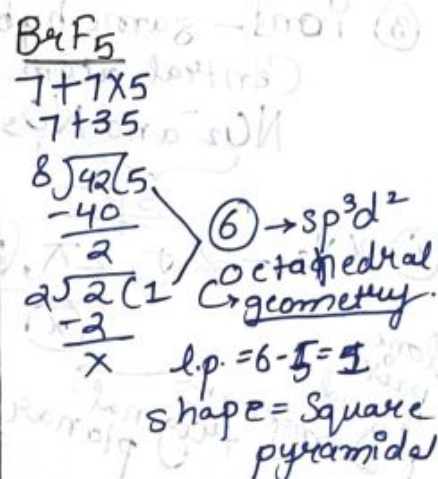
Geometry \rightarrow Structure which depends on lone pair and bond pair both

Shape \rightarrow Structure depends only on bond pair. (lp chupa Rdekho)

6	sp^3d^2		6	0	Octahedral	Octahedral
6	sp^3d^2		5	1	— —	Square pyramidal
6	sp^3d^2		4	2	— —	Square planar
7	sp^3d^3		7	0	pentagonal bipyramidal	pentagonal bipyramidal
7	sp^3d^3		6	1	— —	Capped octahedral / distorted octahedral



l.p. = 7 - 1 = 6
Shape = Capped Octahedral



NEET SLAYER



H	Hybridizat ⁿ	B.P.	L.P.	Geometry	Shape
2	sp —	2	0	Linear	Linear
3	sp^2 >	3	0	T. planar	T. planar
3	sp^2 > <	2	1	T. planar	bent/ angular
4	sp^3 <	4	0	Tetrahedral	Tetrahedral
4	sp^3 < <	3	1	Tetrahedral	pyramidal
4	sp^3 < < <	2	2	Tetrahedral	bent/ angular.
5	sp^3d < < < <	5	0	Trigonal bipyramidal	Trigonal bipyramidal
5	sp^3d < < < < <	4	1	— —	See-Saw
5	sp^3d < < < < < <	3	2	— —	T-shape
5	sp^3d < < < < < < <	2	3	— —	Linear

ELECTROVALENT/IONIC BOND

- Not possible b/w similar atoms. This type of bonding requires 2 atoms of diff. nature. One atom should have tendency to lose electrons. i.e. (electropositive) and other atom should be (electronegative) in nature.

Example - IA and VIIA group elements form strong ionic compound.

• Force of attractⁿ is equal in all direction. So, ionic bond is non-directional.

Energy released during formation of 1 mole crystal lattice is called lattice energy.

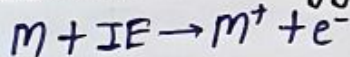
• Ionic compounds do not have a molecular formula, they only have empirical formula.

Key Note { • More the distance b/w 2 elements in the periodic table, more will be ionic character of bond.

• Total no. of electrons lost or gained is called **electrovalency**.

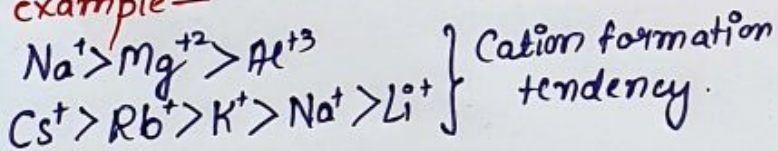
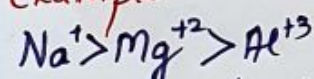
Factors favouring formation of Ionic Bonds →

i) Ionisation Energy → (IE)

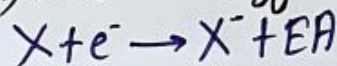


Less Ionisation Energy → Greater tendency to form Cation.

example -

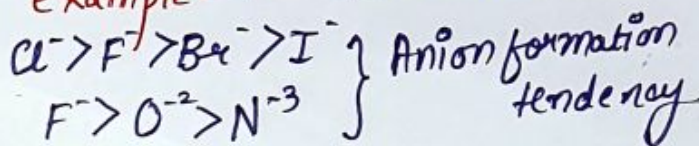


2) Electron Affinity →



High e^- affinity → Greater tendency to form anions

example -



LATTICE ENERGY

Factors affecting L.E. →

i) Lattice energy $\propto \frac{1}{r}$

$r = r^+ + r^-$ = interionic distance

ii) L. Enthalpy $\propto Z^+ \cdot Z^-$

Z^+ → charge on cation in terms of electronic charge.

Z^- → charge on anion in terms of electronic charge.

iii) Size of Cation ↑ CN ↑ LE ↑

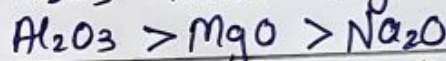
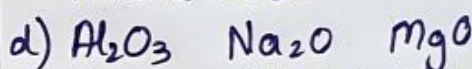
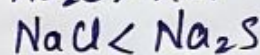
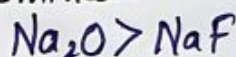
Lattice energy of Factor

a) $NaCl > KCl$ (size)

b) $NaCl < MgO$ (size, charge)

c) $NaCl < MgCl_2$ (size)

In size and charge, charge will dominate.



Characteristics of Ionic compounds

i) physical state: Hard solid

ii) Boiling point and melting point: ↑gh due to strong electrostatic force of attraction

iii) Electrical Conductivity:

Depends on ionic mobility. In solid state there are no free ions, so they are bad conductors of electricity.

In fused state / aqueous solⁿ free ions are present so they are good conductors of electricity.

Solid state < Fused State < Aqueous Solⁿ
(Conductivity Order)

IV) Ionic Reactions.

Ionic Compounds - Ionic Rxⁿ - Fast
Covalent Compounds - Molecular Rxⁿ

✓) Solubility:

Ionic compounds are soluble in polar solvent like H_2O , HF etc.

[Condition of solubility of ionic comp. in water is $(\text{Hydrat. energy} > \text{Lattice energy})$]

Solvation or Hydration

Whenever any compound generally ionic or polar solvent covalent is dissolved in polar solvent or in water, then, diff. ions of compound will get separated and will get surrounded by polar solvent molecules.

This process is **Solvation or Hydration**.

Energy released is solvation energy or Hydration energy.

Factors affecting Solvation or Hydration energy →

$$SE \propto Z^+ Z^- \propto \left[\frac{1}{r^+} + \frac{1}{r^-} \right] \propto \left[1 - \frac{1}{\epsilon_r} \right] \rightarrow \text{dielectric const.}$$

(nature of solvent)

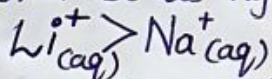
• Greater the polarity, greater will be ϵ_r .

$$\epsilon_r \uparrow \rightarrow \frac{1}{\epsilon_r} \downarrow, 1 - \left(\frac{1}{\epsilon_r} \right) \uparrow \rightarrow SE \uparrow$$

Applications of Hydration energy →

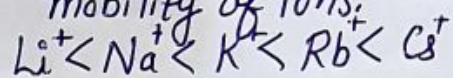
• Size of Hydrated ions:

Greater the hydration of the ion greater will be its hydrated radii.



• Mobility of the ion:

more is the hydration, smaller will be mobility of ions.



• Electrical Conductance: is related to mobility so follows the same order

Covalent Character in Ionic Compounds (Fajan's Rule)

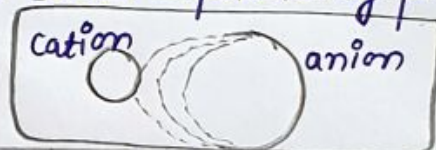
No bond is 100% ionic in nature, It has some % of covalent character which is explained on basis of Fajan's rule.

• In binary comp. AB, if electro-negativity diff. b/w elements A and B is equal to 1.7, comp. AB is 50% ionic.

• Comp. AB is covalent comp. (exceptionally HF is covalent compound even though electronegativity difference is 1.7)

Polarising Power →

phenomenon of deformalⁿ of anion by a cation is known as **polarisation** and ability of cation to polarize a nearby anion is called as polarizing power of cation



Polarisability - Tendency of anion to get distorted or polarised by cation



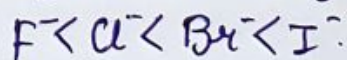
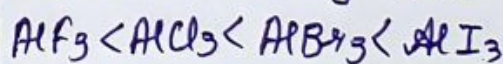
Polarising related to Cation →

• polarising power \propto charge on cation $\propto \frac{1}{\text{size}}$ \propto polarisability related to anion.

• polarisability \propto charge \propto size.

• polarising power increases, covalent character increases.

example - $NaCl$, $MgCl_2$, $AlCl_3$ polarisation increases $Na^+ < Mg^{+2} < Al^{+3}$



• Covalent character \uparrow w/ M_{se} , with M_{se} in size of halide ion

Fajan pointed that \uparrow polarization of anion in molecule, more is covalent character in it

Neet Slayer..♥

More distortion of anion, more will be polarisation then covalent character increases.

• Fajan's Rules →

i) Size of Cation —

Size of cation $\propto \frac{1}{\text{polarisation}}$

ii) Size of anion —

Size of anion $\propto \text{polarisation}$

example — LiF LiCl LiBr LiI

— Size of anion $\uparrow \uparrow$

— polarisation $\uparrow \uparrow$

— Covalent character $\uparrow \uparrow$

iii) Charge on cation —

Charge on cation $\propto \text{polarisation}$

example — $\text{Na}^+ \text{Mg}^{2+} \text{Al}^{3+}$

— charge of cation $\uparrow \uparrow$

— polarization $\uparrow \uparrow$

— Covalent character $\uparrow \uparrow$

iv) Charge on anion —

Charge on anion $\propto \text{polarisation}$

example — $\text{F}^- \text{O}^{2-} \text{N}^{3-}$

— charge on anion $\uparrow \uparrow$

— polarisation $\uparrow \uparrow$

— Covalent character $\uparrow \uparrow$

Application and exceptions of Fajan's Rule —

• Colour of some comp. can be explained on basis of polarisation of their bigger (-ve) ions.

• Variation of melting point.

(melting point of covalent comp < melting point of ionic comp)

$\text{BeCl}_2, \text{MgCl}_2, \text{CaCl}_2, \text{SrCl}_2, \text{BaCl}_2$

→ M.pt. $\uparrow \uparrow$

• Ionic character $\uparrow \uparrow$, melting point $\uparrow \uparrow$, Size of cation $\uparrow \uparrow$, Size of anion is constant.

$\text{CaF}_2, \text{CaCl}_2, \text{CaBr}_2, \text{CaI}_2 \rightarrow \text{M.pt.} \downarrow \downarrow$

• Covalent character $\uparrow \uparrow$, melting point $\downarrow \downarrow$, Size of anion $\uparrow \uparrow$, Size of cation is constant

Molecular Orbital Theory →

Given by Hund and Mulliken

Basic Postulates →

• The atomic orbitals of comparable energies and proper symmetry of 2 atoms are mixed up to form molecular orbitals.

• Total no. of MO = Total no. of A.O

• 2 A.O. combine to form 2 M.Os. One is antibonding M.O. and other is bonding M.O. The energy of bonding MO is less than antibonding MO.

• The filling up of electrons in M.Os must follow Aufbau's principle, Pauli's exclusion principle and Hund's rule

Electrons are filled in the increasing energy of the MO which is in order —

$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

Increasing energy (for electrons > 14)

$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x = \pi 2p_y) < \sigma 2p_z < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

Increasing energy (for electron ≤ 14)

Next Slayer

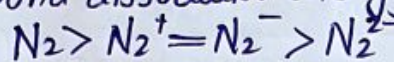
• Bond order \propto Stability of molecule

Bond order \propto Bond energy $\propto \frac{1}{\text{Bond length}}$

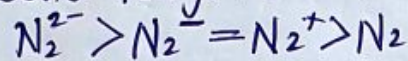
$N_b \rightarrow$ No. of electrons in bonding M.O.

$N_a \rightarrow$ No. of electrons in anti bonding M.O.

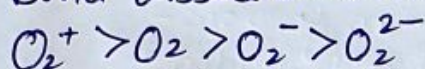
Bond dissociation energy order —



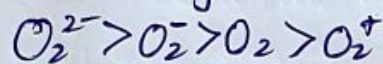
Bond length order —



Bond dissociation energy order —



Bond length order —



• Stability Of Molecules:—

If N_b is the no. of electrons occupying bonding orbitals and N_a the no. of electrons occupying antibonding orbitals, then—

A positive bond order is $N_b > N_a$ means a stable molecule. While, a negative ($N_b < N_a$) or zero ($N_a = N_b$) bond order means an unstable molecule.

Electron configuration / Bond order of simple diatomic molecules.

The electronic configuration and the bond order in case of simple diatomic molecules can be obtained by filling the molecular orbitals by applying Aufbau's principle and Hund's rule etc.

Bond order — $\frac{N_b - N_a}{2}$

Bond order \propto Bond energy \propto

Stability of a molecule $\propto \frac{1}{\text{bond length}}$

• Bond order of $N_2 = \frac{10-4}{2} = 3$

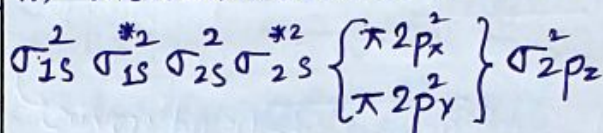
As bond order increases, stability of molecule increase

Correct order — $N_2^{2-} < N_2^{2+} < N_2^- < N_2^+ < N_2$

Bonding in some diatomic molecules and ions—

Nitrogen Molecule (N_2)

Total no. of electrons = 14 and filling in molecular orbitals we have,

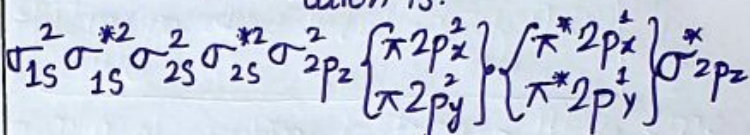


$$\text{Bond Order} = \frac{(N_b - N_a)}{2} = \frac{10 - 4}{2} = 3$$

• It is diamagnetic

Oxygen molecule (O_2) →

Total no. of e^- = 16 and electronic configuration is.



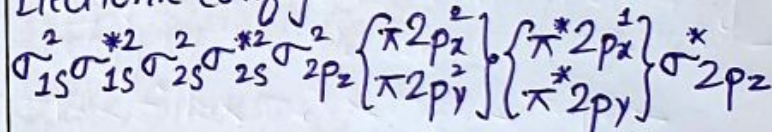
$$\text{Bond Order} = \frac{(N_b - N_a)}{2} = \frac{10 - 6}{2} = 2$$

→ 2 unpaired electrons — paramagnetic

O_2^+ ion →

Total no. of e^- = (16-1) = 15

Electronic Configuration —

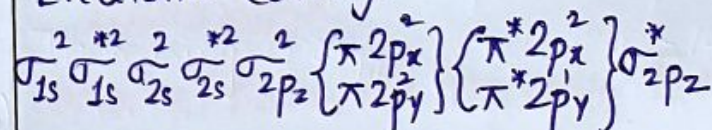


$$\text{Bond Order} = \frac{10 - 5}{2} = 2.5 - \text{paramagnetic}$$

O_2^- Superoxide ion.

• Total no. of e^- = (16+1) = 17.

Electronic Configuration —



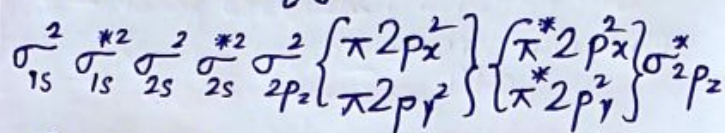
$$\text{Bond Order} = \frac{10 - 7}{2} = 1.5$$

— paramagnetic.

Peroxide ion (O_2^{2-}) \rightarrow

Total no. of $e^- = (16 \times 2) = 18$

Electronic Configuration -

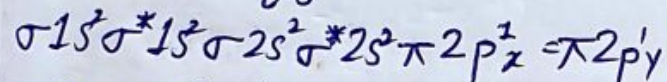


Bond Order = $\frac{10 - 8}{2} = 1$ - diamagnetic

Boron Molecule (B_2) \rightarrow

No. of $e^- = 10$

Electronic Configuration -



\rightarrow 2 unpaired electrons \rightarrow paramagnetic

Dipole Moment..

It is equal to the product of magnitude of charge and distance b/w the centres of (+ve) and (-ve) charges of a dipole.

$$\mu = Q \times d$$

- Vector Quantity.

$$\mu_{\text{net}} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2\cos\theta}$$

μ increases with increase in $\Delta E.N.$,
i.e. $HI < HBr < HCl < HF$

$$\% \text{ Ionic Character} = \frac{\mu_{\text{expt}}}{\mu_{\text{theory}}} \times 100$$

Application of $\mu \rightarrow$

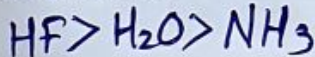
- Can predict shape of molecule
- distinguish b/w cis and trans isomers, o-m-p-isomers.

Note \rightarrow

① Dipole Moment \propto electronegativity difference.



② Dipole Moment \propto No. of lone pair of e^-



③ For Homonuclear molecules and molecules having normal shape, like linear, trigonal, tetrahedral possess zero dipole moment.
eg - Cl_2, N_2, O_2, CO_2

④ Molecules having distorted shapes like angular, pyramidal show dipole moment.
eg - H_2O, SO_2

⑤ $NH_3, H_2O, NF_3, SF_4, H_2S \rightarrow$ unsymmetrical molecular geometry.

Hydrogen Bond ..

partially positive charged hydrogen atom forms a bond with the other more electronegative atom.

This bond is Hydrogen Bond and it's weaker than covalent bond.

Types of Hydrogen Bond \rightarrow

① Intramolecular H-Bonding:
This type of Ortho H-Bonding occurs when polar H and electronegative atom are present in the same molecule

② Intermolecular H-Bonding:
Formed b/w 2 diff. molecules of same or different compounds.

NEET SLAYER & !

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21 hours of hardwork..

padhai Karo, All the very Best.

