

Chemical Bonding

1. OCTET RULE (LEWIS-KOSSEL RULE)

Noble gases have 8 electron in their outermost shell (complete octet) and outermost configuration is ns^2p^6 . Every atom has a tendency to complete its octet by losing or gaining or by sharing electron.

1.1 Exceptions of Octet Rule

- (i) **Electron-Deficient Compounds** : Compounds having less than 8 electrons in the outermost shell of the central atom. e.g. BF_3 , $BeCl_2$, $AlCl_3$.
- (ii) **Expansion of Octet** : The compounds having more than 8 electrons in the outermost shell of the central atom. e.g. PCl_5 , SF_6
- (iii) **Transition Metal Ions** : Transition metal ions have 9 to 18 electrons in their outermost shell, e.g. Cr^{+3} , Mn^{+2} , Cu^+
- (iv) H and Li atoms attain configuration like He by gaining and losing one electron, respectively, e.g. H^{-1} , Li^{+1} .
- (v) **Odd Electron Bond** : Compounds having unpaired electron are also exceptions to the octet rule. The molecules having odd electrons are NO_2 , NO , ClO_2 , O_2^{-1} etc.

Ex.1 Octet rule is not followed in :

[1] Ti^{+3} [2] Na^{+1} [3] Ca^{+2} [4] K^{+1} **Ans. [1]**

Sol. Ti^{+3} has 9 electron in its outermost shell.

Ex.2 Which of the following compounds has the central atom with incomplete octet ?

[1] NH_4^{+1} [2] BCl_3 [3] CCl_4 [4] PCl_3 **Ans. [2]**

Sol. B has 6 electron in BCl_3 .

Ex.3 An element forms cation by losing electron. The configuration of the cation may be like that of :

[1] inert gas [2] pseudo inert gas [3] inert pair [4] all above **Ans. [4]**

Sol. Inert gas – 8 electron in outermost shell. Pseudo inert gas – 18 electron in outermost shell.

1.2 Inert Pair Configuration

The heavy elements of IIIA, IVA and VA groups form cations having charge equal to group number and two less than group number, e.g. Tl^{+1} , In^{+1} , Sn^{+2} , Pb^{+2} . In these elements due to more penetration power of s orbitals, the atom generally does not lose e^- present in ns orbital, so only np electrons are lost by the atom. This effect is called "inert pair effect".

2. IONIC BOND

- (i) Atoms get established by ion formation in formation of an ionic bond. One atom forms a cation by losing electrons and other forms anion by accepting electrons.
- (ii) The octets of the atoms are completed by transfer of electrons.
- (iii) Cations and anions are bonded by electrostatic force of attraction in ionic bond.
- (iv) Ionic bond is formed between metals (cations) and nonmetals (anions).
- (v) During the formation of an ionic bond, cation can attain one of the following configurations.

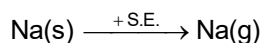
(a) **Inert Configuration** : I A and II A group metals (octet configuration) Na^{+1} , Mg^{+2} , K^{+1} , Rb^{+1} attain s^2p^6 inert configuration.

(b) **Pseudo Inert Configuration** : Metals of I B and II B groups – Cu^{+} , Zn^{+2} , Ag^{+1} , Au^{+1} , Cd^{+2} , Hg^{+2} .

(c) Transition metals, e.g. Ti^{+3} (9), Ti^{+2} (10).

(vi) **Energies involved in ionic bond formation** :

(a) Sublimation Energy – The energy required to convert solid metal to gaseous state.



(b) Ionisation Energy – The energy required to remove an electron from the outermost shell of metal in gaseous state.

(c) Dissociation Energy – The energy required to convert the nonmetal molecule to atoms.



(d) Electron Affinity – The energy required during formation of an anion by addition of an electron to a nonmetal in gaseous state.

(e) Lattice Energy –

(i) The energy released during the formation of ionic bond.

(ii) Cations and anions form crystal lattice of ionic crystal in space by electrostatic force of attraction.

(iii) Lattice Energy $\propto \frac{1}{r^{+} + r^{-}}$

(f) Ionic compound is formed when the energy required (Sublimation energy + Ionisation energy + Dissociation energy) is less than energy released (Electron affinity + Lattice energy) So total energy change

$$-\Delta H = S + I + \frac{D}{2} - E - U$$

(Borne-Haber Equation)

2.1 Factors Affecting Formation of Ionic Bond

- (a) The sublimation energy and ionisation potential of a metal should be minimum so that it may easily form a cation.
- (b) The dissociation energy and electron affinity of a nonmetal should be high so that it may easily form an anion.
- (c) There should be a strong electrostatic force of attraction between cation and anion.
- (d) Metals have low ionisation potential and nonmetals have high electron affinity, so ionic bond is formed between metals and nonmetals.

2.2 Characteristic of Ionic Compounds

(i) **Solid and Crystalline Structure** :

Ionic compounds do not show molecular structure.

Ionic compounds have definite crystal structures.

CsCl has body centered cubic (BCC) structure.

(ii) **Melting and Boiling Point** :

Ionic compounds have high melting and boiling points.

(iii) **Hardness** :

Ionic compounds are solid with brittle nature.

(iv) **Conductivity** :

Ions are not free in solid state so ionic compounds are bad conductor of electric current in solid state but in solution and fused state electric current passes through and the ionic compound becomes good conductor of electricity

(v) **Solubility of Ionic Compounds** :

Ionic compounds are soluble in solvents which have high dielectric constant like water.

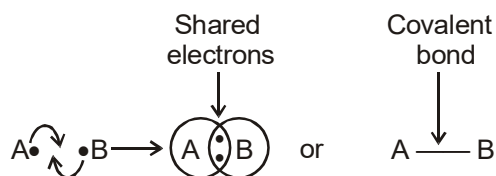
(vi) Solvation of Hydration Energy :

The energy released by the attraction of ions and solvent molecules is called solvation energy and when solvent is water this is called hydration energy.

$$\text{Solubility in water} = \text{Lattice energy} < \text{Hydration energy}$$

3. COVALENT BOND

A covalent bond is formed between two nonmetal atoms. e.g. in H–H, O=O and N≡N.



Formation of a covalent bond

(a) Nonpolar Covalent Bonds :

Covalent bond between two identical atoms (e.g. H–H, O=O) electronegativity of both atom is same.

(b) Polar Covalent Bond :

When covalent bond is formed between two different atoms. e.g. HCl, HF

3.1 Characteristics of Covalent Compounds**1. Physical State :**

Covalent compounds generally occur in gaseous and liquid states, but sometimes in solid state also.

(a) Diamond :

- (i) In the structure of diamond, every carbon atom is attached to four other carbon atoms, tetrahedrally.
- (ii) Density of diamond is high.
- (iii) Every carbon atom is sp^3 hybridised and forms four C–C σ bond by sp^3 – sp^3 overlapping
- (iv) Electrons are not free, so diamond is bad conductor of electricity.
- (v) Oxidation number of C in diamond is zero

(b) Graphite :

- (i) In the structure of graphite, every carbon atom is attached to three other carbon atoms forming hexagonal layers.
- (ii) Every carbon atom is sp^2 hybridised and forms three σ bonds by sp^2 – sp^2 overlapping.
- (iii) Every carbon atom has one free electron. So graphite is a good conductor of electricity
- (iv) C–C bond length is 1.42 Å in plane.

2. Boiling and Melting Points :

Melting and boiling points of covalent compounds are generally low due to weak van der Waals forces between molecules.

3. Conductivity :

Covalent compounds are bad conductors of electricity.

Ex.4 F_2 and Cl_2 are gases, Br_2 is liquid and I_2 is solid why ?

Sol. Intermolecular distance is high in gaseous state, low in solid state and intermediate in liquid state. When

molecular weight increases, van der Waals forces also increase. So F_2 and Cl_2 are gases due to weak van der Waals forces and I_2 is solid.

Ex.5 $CHCl_3$ is liquid but CHI_3 is solid. Why ?

Sol. Due to increase in molecular weight, intermolecular force in CHI_3 is greater than in $CHCl_3$.

Ex.6 CO_2 is gas but SiO_2 is solid. Why ?

Sol. CO_2 has only weak van der Waals forces, but in SiO_2 , one silicon atom is surrounded by four oxygen atoms tetrahedrally forming macromolecular structure. So SiO_2 is solid.

Ex.7 H_2O is liquid but H_2S is a gas. Why ?

Sol. H_2O is liquid due to strong H-bonding.

Ex.8 NH_3 can be liquefied easily but PH_3 is not. Why ?

Sol. H-bonding present in NH_3 , while PH_3 is nonpolar due to same value of electronegativity of P and H.

4. DIPOLE MOMENT

Polarity of bond is measured by dipole moment. The product of charge (e) and internuclear distance (d) between two poles of polar molecule is called dipole moment (μ). The unit of dipole moment is Debye (D). $\mu = e.d$.

$$\text{Percentage of ionic character} = \frac{\mu(\text{obs.})}{\mu(\text{cal.})} \times 100 \quad [1D = 10^{-18} \text{ e.s.u cm}]$$

(i) Dipole moment of nonpolar compounds is zero.

(ii) When the electronegativity difference of atoms in compounds increases dipole moment also increases.

(iii) If the structure of a compound is symmetric like linear, trigonal planar & Tetrahedral the dipole moment will be zero. e.g. BF_3 , BCl_3 , CF_4 , CCl_4 , CH_4 etc.

(iv) Structure of $CHCl_3$, CH_2Cl_2 , $CHCl_3$, CH_3F is tetrahedral, but due to different bonds, dipole moment will not be zero.

Ex.9 What will be the dipole moment of HCl if $d = 1.34 \text{ \AA}$ and charge = $4.8 \times 10^{-10} \text{ esu}$? Calculate ionic character of the bond when observed value of μ is 1.08D.

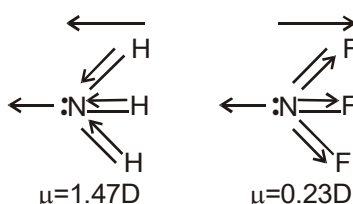
Sol.

$$\begin{aligned} \mu_{(\text{cal.})} &= e.d \\ &= 4.8 \times 10^{-10} \times 1.34 \times 10^{-8} \\ &= 6.4 \times 10^{-18} \text{ esu cm} \end{aligned}$$

$$\text{Percentage of ionic character} = \frac{1.08 \times 10^{-18}}{6.40 \times 10^{-18}} \times 100 = 16.88\%$$

Ex.10 Dipole moment of NH_3 is greater than that of NF_3 .

Sol. N-H bond moment and lone pair moment in NH_3 add up because they are in same direction. In NF_3 lone pair moment and N-F bond moment are in opposite direction. So dipole moment will decrease.



Dipole Moments of some Compounds

Ex.11 Which of the following compounds should have zero dipole moment.

[1] CO [2] CCl₄ [3] NH₃ [4] CH₃Cl **Ans. [2]**

Sol. Tetrahedral geometry of CCl₄ in which four C–Cl bond moments balance each other.

Ex.12 Which of the following compounds should have higher dipole moment than the remaining three?

[1] HF [2] H₂O [3] NH₃ [4] NF₃ **Ans. [2]**

Sol. Higher dipole moment of H₂O is due to H-bonding.

5. BOND PAIR AND LONE PAIR

(i) In covalent compounds, the number of bond pairs is equal to the number of shared pairs of electron.

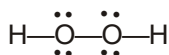
$$\text{Number of bond pairs} = \text{Number of bonds}$$

(ii) The electron pair, which does not participate in bond formation, is called lone pair.

(iii) :N≡N: (N₂) 2 lone pairs and 3 bond pairs.

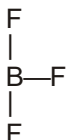
(iv) :Cl–Cl: (Cl₂) 6 lone pairs and 1 bond pairs.

Ex.13 What are the numbers of lone pairs and bond pairs in H₂O₂ molecule ?



Sol. Lone pair = 4, Bond pair = 3

Ex.14 What are the numbers of lone pairs and bond pairs on B in BF₃ ?

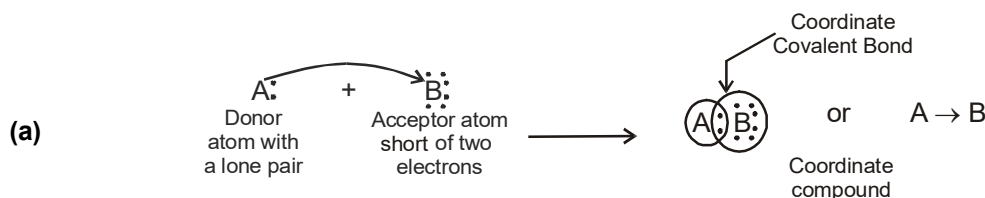


Sol. Lone Pair = 0, Bond Pair = 3

6. COORDINATE BOND OR DATIVE BOND OR COORDINATE COVALENT BOND

(i) Coordinate bond is formed by unequal sharing of electron pair between two atoms, one of which, called the donor atom, provides the electron pair and the other atom, called the acceptor atom, receives that electron pair.

(ii) The donor should have complete octet with a lone pair of electrons and acceptor atom should have incomplete octet and a vacant orbital to accept the electron pair.



It is also known as **Dative Bond**, according to Scientist, **Menzie**.

(iii) Formation of covalent bond is essential before formation of coordinate bond.

(iv) This bond is different from a covalent bond, in the latter one electron provided by each of the bonding atoms.

(v) Coordinate bond is also different from ionic bond because electrons are not completely transferred.

(vi) In this bond, electrons are donated only in the form of pairs.

(vii) Electron pair is provided by only one atom and shared equally by both the atoms. So according to Sugden it is a polar bond.

(viii) Once a coordinate bond is formed, there remains no difference between the coordinate and covalent bonds. So the property of coordinate compounds are similar to those of covalent compounds.

Ex.15 Which of the following compounds has coordinate bond ?

[1] Methyl cyanide [2] Methyl isocyanate [3] Methyl cyanate [4] None of these

Ans. [3]

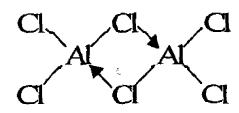
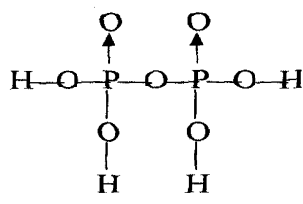
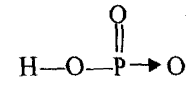
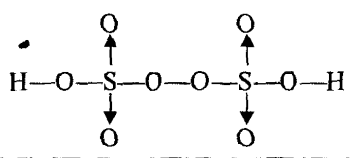
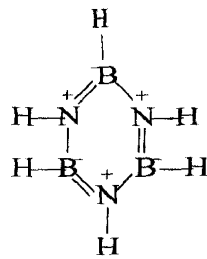
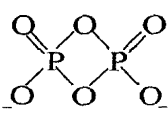
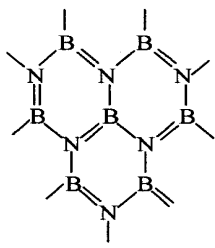
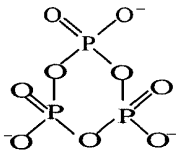
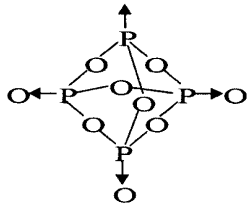
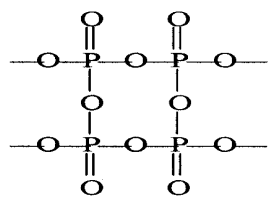
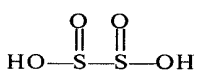
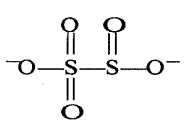
Sol. $\text{CH}_3\text{C}\equiv\text{N}\rightarrow\text{O}$

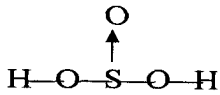
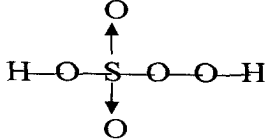
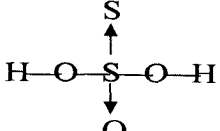
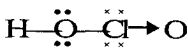
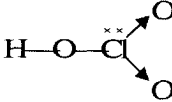
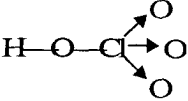
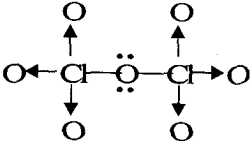
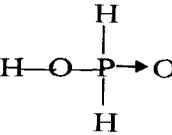
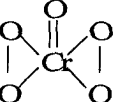
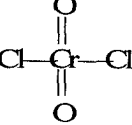
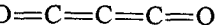
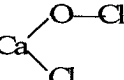
STRUCTURES \ FORMULAE OF IMPORTANT COMPOUNDS

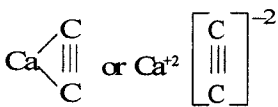
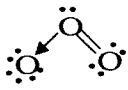
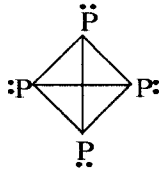
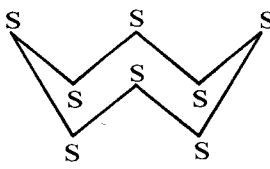
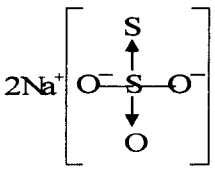
Compound	Structure	Electron dot / Cross Formula
1. Sodium Sulphide (Na_2S)	$(\text{Na}^+)_2\text{S}^{2-}$	$\text{Na}^+ \quad :\ddot{\text{S}}:^- \quad \text{Na}^+$
2. Calcium Hydride (CaH_2)	$\text{Ca}^{++}(\text{H}^-)_2$	$\text{H}^- \quad \text{Ca}^{++} \quad \text{H}^-$
3. Hydrogen Peroxide (H_2O_2)	$\text{H}-\text{O}-\text{O}-\text{H}$	$\text{H}:\ddot{\text{O}}:\ddot{\text{O}}:\text{H}$
4. Phosphene (PH_3)	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{P}-\text{H} \end{array}$	$\begin{array}{c} \text{H} \\ \times \\ \text{H}:\ddot{\text{P}}:\text{H} \end{array}$
5. Carbon mone-oxide (CO)	$\text{C}\equiv\text{O}$	$:\text{C} \quad :: \quad \text{O}:$
6. Carbon di-oxide (CO_2)	$\text{O}=\text{C}=\text{O}$	$\ddot{\text{O}} :: \text{C} :: \ddot{\text{O}}$
7. Hydrogen Cyanide (HCN)	$\text{H}-\text{C}\equiv\text{N}$	$\text{H}:\text{C}::\text{N}$
8. Potassium Isocyanide (KNC)	$\text{K}^+ \text{N}\equiv\text{C}$	$\text{K}^+ \left[\text{N}::\text{C} \right]^-$
9. Sodium Hydroxide (NaOH)	$\text{Na}^+ \text{O}^-\text{H}$	$\text{Na}^+ \left[:\ddot{\text{O}}:\text{H} \right]^-$
10. Nitrous oxide (N_2O)	$\text{N}\equiv\text{N}\rightarrow\text{O}$	$:\text{N}::\text{N}:\ddot{\text{O}}:$
11. Nitric oxide (NO or N_2O_2)	$\text{O}=\text{N}-\text{N}=\text{O}$	$\ddot{\text{O}} :: \text{N}:\text{N}::\ddot{\text{O}}$
12. Dinitrogen tri oxide (N_2O_3)	$\begin{array}{c} \text{O} \\ \uparrow \\ \text{O}=\text{N}-\text{N}=\text{O} \end{array}$	$\begin{array}{c} :\ddot{\text{O}}: \\ \ddot{\text{O}}::\text{N}:\text{N}::\ddot{\text{O}} \end{array}$
13. Dinitrogen tetraoxide (N_2O_4)	$\begin{array}{c} \text{O} \quad \text{O} \\ \uparrow \quad \uparrow \\ \text{O}=\text{N}-\text{N}=\text{O} \end{array}$	$\begin{array}{c} :\ddot{\text{O}}:\ddot{\text{O}}: \\ \ddot{\text{O}}::\text{N}:\text{N}::\ddot{\text{O}} \end{array}$
14. Nitrogen Penta oxide (N_2O_5)	$\begin{array}{c} \text{O} \quad \quad \text{O} \\ \uparrow \quad \quad \uparrow \\ \text{O}=\text{N}-\text{O}-\text{N}=\text{O} \end{array}$	$\begin{array}{c} :\ddot{\text{O}}:\ddot{\text{O}}:\ddot{\text{O}}: \\ \ddot{\text{O}}::\text{N}:\text{O}:\text{N}::\ddot{\text{O}} \end{array}$
15. Nitrous acid (HNO_2)	$\text{H}-\text{O}-\text{N}=\text{O}$	$\text{H}:\ddot{\text{O}}:\text{N}::\ddot{\text{O}}$
16. Nitric acid (HNO_3)	$\begin{array}{c} \text{O} \\ \uparrow \\ \text{H}-\text{O}-\text{N}=\text{O} \end{array}$	$\begin{array}{c} :\ddot{\text{O}}: \\ \text{H}:\ddot{\text{O}}:\text{N}::\ddot{\text{O}} \end{array}$
17. Nitrate Ion (NO_3^-)	$\begin{array}{c} \text{O} \\ \\ \text{O}\leftarrow\text{N}-\text{O}^- \end{array}$	$\left[\begin{array}{c} :\ddot{\text{O}}: \\ \ddot{\text{O}}:\text{N}:\ddot{\text{O}}: \end{array} \right]^-$

Compound	Structure	Electron dot / Cross Formula
18. Ammonium Chloride (NH ₄ Cl)	$\left[\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array} \right] \text{Cl}^-$	$\left[\begin{array}{c} \text{H} \\ \cdot \\ \cdot \\ \text{H} \cdot \text{N} \cdot \text{H} \\ \cdot \\ \cdot \\ \text{H} \end{array} \right]^+ \quad \text{Cl}^-$
19. Carbonate Ion (CO ₃ ²⁻)	$\begin{array}{c} \text{O} \\ \\ \text{O}-\text{C}-\text{O}^- \end{array}$	$\left[\begin{array}{c} \text{:O:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \text{ C } \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \end{array} \right]^{2-}$
20. Sulphur di-oxide (SO ₂)	$\text{O} \leftarrow \text{S} = \text{O}$	$\text{:}\ddot{\text{O}}\text{:} \times \text{S} \times \text{:}\ddot{\text{O}}\text{:}$
21. Sulphur Tri-oxide (SO ₃)	$\begin{array}{c} \text{O} \\ \uparrow \\ \text{O} \leftarrow \text{S} = \text{O} \end{array}$	$\begin{array}{c} \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \times \text{S} \times \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \end{array}$
22. Sulphite Ion (SO ₃ ²⁻)	$\begin{array}{c} \text{O} \\ \uparrow \\ \text{O}-\text{S}-\text{O}^- \end{array}$	$\left[\begin{array}{c} \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \times \text{S} \times \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \end{array} \right]^{2-}$
23. Sulphate Ion (SO ₄ ²⁻)	$\begin{array}{c} \text{O} \\ \uparrow \\ \text{O}-\text{S}-\text{O}^- \\ \downarrow \\ \text{O} \end{array}$	$\left[\begin{array}{c} \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \times \text{S} \times \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^{2-}$
24. Thio-sulphate Ion (S ₂ O ₃ ²⁻)	$\begin{array}{c} \text{S} \\ \uparrow \\ \text{O}-\text{S}-\text{O}^- \\ \downarrow \\ \text{O} \end{array}$	$\left[\begin{array}{c} \times \times \text{S} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \times \text{S} \times \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^{2-}$
25. Phosphoric Acid (H ₃ PO ₄)	$\begin{array}{c} \text{H} \\ \\ \text{O} \\ \\ \text{H}-\text{O}-\text{P}-\text{O}-\text{H} \\ \\ \text{O} \end{array}$	$\begin{array}{c} \text{H} \\ \cdot \\ \cdot \\ \text{H} \cdot \text{:}\ddot{\text{O}}\text{:} \times \text{P} \times \text{:}\ddot{\text{O}}\text{:} \cdot \text{H} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \end{array}$
26. Phosphorous acid (H ₃ PO ₃)	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{O}-\text{P}-\text{O}-\text{H} \\ \\ \text{O} \end{array}$	$\begin{array}{c} \text{H} \\ \cdot \\ \cdot \\ \text{H} \cdot \text{:}\ddot{\text{O}}\text{:} \times \text{P} \times \text{:}\ddot{\text{O}}\text{:} \cdot \text{H} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \end{array}$
27. Phosphate Ion (PO ₄ ³⁻)	$\begin{array}{c} \text{O} \\ \uparrow \\ \text{O}-\text{P}-\text{O}^- \\ \downarrow \\ \text{O} \end{array}$	$\left[\begin{array}{c} \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \times \text{P} \times \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^{3-}$
28. Potassium Permanganate (KMnO ₄)	$\text{K}^+ \left[\begin{array}{c} \text{O} \\ \uparrow \\ \text{O} \leftarrow \text{Mn} - \text{O} \\ \downarrow \\ \text{O} \end{array} \right]$	$\left[\begin{array}{c} \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \times \text{Mn} \times \text{:}\ddot{\text{O}}\text{:} \\ \cdot \\ \cdot \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^{3-}$

Compound	Structure
29. Carbon di-sulphide (CS ₂)	S=C=S
30. Nitronium Ion (NO ₂ ⁺)	O=N ⁺ =O
31. Nitrite Ion (NO ₂ ⁻)	O=N-O ⁻
32. Hydronium Ion (H ₃ O ⁺)	$\begin{array}{c} \text{H} \quad \text{H} \\ \diagdown \quad \diagup \\ \text{O}^+ \\ \diagup \\ \text{H} \end{array}$
33. Sulphuryl chloride (SO ₂ Cl ₂)	$\begin{array}{c} \text{Cl} \\ \\ \text{O} \leftarrow \text{S} \rightarrow \text{O} \\ \\ \text{Cl} \end{array}$
34. Potassium chlorate (KClO ₃)	$\text{K}^+ \left[\begin{array}{c} \text{O} \\ \uparrow \\ \text{O} \leftarrow \text{Cl} - \text{O} \end{array} \right]^-$
35. Potassium per-chlorate (KClO ₄)	$\text{K}^+ \left[\begin{array}{c} \text{O} \\ \uparrow \\ \text{O} \leftarrow \text{Cl} - \text{O} \\ \downarrow \\ \text{O} \end{array} \right]^-$
36. Potassium dichromate (K ₂ Cr ₂ O ₇)	$(\text{K}^+)_2 \left[\begin{array}{c} \text{O} \quad \text{O} \\ \uparrow \quad \uparrow \\ \text{O} - \text{Cr} - \text{O} - \text{Cr} - \text{O} \\ \downarrow \quad \downarrow \\ \text{O} \quad \text{O} \end{array} \right]^{2-}$
37. Ferry - cyanide Ion [Fe(CN) ₆] ⁻³	$\left[\begin{array}{ccc} \text{N} \equiv \text{C}^- & & \text{C} \equiv \text{N} \\ \swarrow & & \searrow \\ \text{N} \equiv \text{C}^- & \text{Fe}^{++} & \text{C} \equiv \text{N} \\ \swarrow & & \searrow \\ \text{N} \equiv \text{C}^- & & \text{C} \equiv \text{N} \end{array} \right]^{-3}$
38. Fero cyanide Ion [Fe(CN) ₆] ⁻⁴	$\left[\begin{array}{ccc} \text{N} \equiv \text{C}^- & & \text{C} \equiv \text{N} \\ \swarrow & & \searrow \\ \text{N} \equiv \text{C}^- & \text{Fe}^{++} & \text{C} \equiv \text{N} \\ \swarrow & & \searrow \\ \text{N} \equiv \text{C}^- & & \text{C} \equiv \text{N} \end{array} \right]^{-4}$

Compound	Structure
39. Aluminium chloride (Al_2Cl_6)	
40. Pyro phosphoric acid ($H_4P_2O_7$)	
41. Hypophosphate acid (HPO_3)	
42. Per-di-sulphuric acid ($H_2S_2O_8$)	
<p>Borazine (Inorganic benzene)</p> 	<p>Dimeta phosphate ion</p> 
<p>Boron nitride (Inorganic graphite)</p> 	<p>Trimeta phosphate ion</p> 
<p>Phosphorous penta oxide</p> 	<p>Tetra meta phosphate ion</p> 
<p>$H_2S_2O_4$ (Dithionous acid)</p> 	<p>$S_2O_5^{-2}$ (Pyrosulphurous ion)</p> 

Compound's Name	Formula	Structure
Sulphurous acid	H_2SO_3	
Persulphuric acid Peroxo monosulphuric acid (caro's acid)	H_2SO_5	
Thiosulphuric acid	$H_2S_2O_3$	
Chlorous acid	$HClO_2$	
Chloric acid	$HClO_3$	
Perchloric acid	$HClO_4$	
Perchloric anhydride	Cl_2O_7	
Hypophosphorous acid	H_3PO_2	
Perchromate	CrO_5	
Chromyl Chloride	CrO_2Cl_2	
Carbon suboxide	C_3O_2	
Bleaching Powder	$CaOCl_2$	

Compound's Name	Formula	Structure
Calcium carbide	CaC_2	
Ozone	O_3	
Ground state of phosphorous	P_4	
Ground state of sulphur	S_8	
Nitrosyl choride (Tilden reagent)	NOCl	$\text{Cl}-\text{N}=\text{O}$
Sodium thiosulphate	$\text{Na}_2\text{S}_2\text{O}_3$	

6.1 Characteristics of Coordinate Compounds

- (i) These compounds show partial ionic character due to polar nature.
- (ii) Melting and boiling points of these compounds are less than those of ionic compounds and more than those of covalent compounds.
- (iii) **Solubility** : These are partially soluble in water (polar solvent) and soluble in nonpolar solvents.
- (iv) **Conductivity** : Coordinate compounds are bad conductor of electricity due to absence of ions.
- (v) **Stability** : These are stable due to polar nature. When bonding molecules are independently stable, the coordinate bond is less stable, e.g. in $\text{H}_3\text{N} \rightarrow \text{BF}_3$ molecule, the coordinate bond is less stable because NH_3 and BF_3 are themselves stable.

6.2 Compounds Having Ionic, Covalent and Coordinate Bonds

Some compounds have all the three types of bonds, i.e. ionic, covalent and coordinate bond, e.g.

- (i) NH_4Cl – Ionic bond between NH_4^+ and Cl^- and covalent bonds and a coordinate bond are present in NH_4^+ .

7. IONIC CHARACTER IN COVALENT COMPOUNDS

- (i) When electronegativity difference of two atoms is less than 0.8, the covalent bond is nonpolar, e.g. PH_3 .
- (ii) When electronegativity difference of atoms is 0.1 to 0.8, the bond between them is nonpolar covalent bond. When electronegativity difference of atoms is 0.8 to 1.6, the bond between them is polar covalent bond. Then electronegativity difference is 1.7 to 3.2, the bond is ionic.

$$\text{Ionic character} \propto \text{Electronegativity difference of bonded atoms}$$

- (iii) Pauling used dipole moment for depicting percentage of polarity and ionic character of the bond. According to Pauling, a bond can never be 100% ionic.
- (iv) When electronegativity difference between two atoms is 2.1, there is 50% ionic character in the bond.
- (v) When electronegativity difference is zero (identical atoms), the bond will be 100% covalent.

According to Haney and Smith, the percentage of ionic character of a polar covalent bond can be calculated with help of the following expression.

$$\% \text{ Ionic character} = (0.16\Delta + 0.035\Delta^2) \times 100$$

Where Δ = Electronegativity difference between bonded atoms.

Ex.16 Atomic number of A, B and C are Z, Z + 1 and Z + 2, respectively, and C is an alkali metal. What should be the nature of bond between A and C ?

[1] Ionic [2] Coordinate [3] Covalent [4] Hydrogen bond **Ans. [1]**

Sol. (Z + 2) stands for alkali metal, (Z + 1) for inert gas and Z for halogen. Since A is halogen and C is alkali metal, the nature of bond will be ionic.

Ex.17 Three electrons are present in the outermost shell of an element A and six electrons in that of B. What compound will be formed by A and B ?

[1] A_2B_3 [2] AB_2 [3] A_6B_6 [4] All of the above **Ans. [1]**

Sol. A has 3 electrons in outermost shell, like Al. B has 6 electrons in outermost shell, like O, their compound will be Al_2O_3 , i.e. A_2B_3 .

Ex.18 Compound having dative and covalent bond, is :

[1] SO_2 [2] NH_4Cl [3] $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ [4] AgCl **Ans. [1]**

Sol. All the three types of bonds are present in NH_4Cl and CuSO_4 , ionic bond in AgCl and both dative and covalent bonds in SO_2 .

Ex.19 Compound formed by the elements having atomic number 6 and 16, will be :

[1] Ionic [2] Covalent [3] Coordinate [4] Hydrogen bonded **Ans.[2]**

Sol. Atomic number 6 = Carbon ; Atomic number 16 = Sulphur

The compound will be CS_2 (covalent)

Ex.20 Diamond and graphite are :

[1] isotone [2] isomers [3] isotope [4] allotropes **Ans.[4]**

Sol. Diamond and graphite are allotropic forms of carbon.

Ex.21 Which of the following compounds has ionic and covalent bonds ?

[1] NaCN [2] CO_2 [3] NaCl [4] N_2

Sol. $\text{Na}^+ \text{C}^{\ominus} \equiv \text{N}$

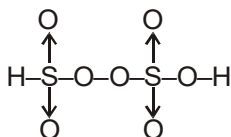
Ex.22 Inorganic graphite is :

- [1] boron nitrate [2] boron nitride (BN) [3] boron carbonate [4] None of the above **Ans. [2]**

Ex.23 In disulphuric acid ($H_2S_2O_7$), the number of covalent bonds and coordinate bonds, respectively, are

- [1] 6 and 4 [2] 6 and 6 [3] 4 and 4 [4] 4 and 6 **Ans. [1]**

Sol. Structure of $H_2S_2O_7$ is



Ex.24 When electronegativity difference of B and Cl is 1.0, the % ionic character in the bond will be :

- [1] 10% [2] 19.5% [3] 30% [4] 45.5% **Ans. [2]**

Sol. $\Delta = 1.0$

$$\begin{aligned} \% \text{ ionic character} &= (0.16 \times 1 + 0.035 \times 1^2) \times 100 \\ &= 0.195 \times 100 = 19.5\% \end{aligned}$$

Ex.25 What should be the percentage ionic character in CsF when electronegativity difference is 3.3 ?

- [1] 90.9% [2] 0.09% [3] 93.3% [4] 95.7% **Ans. [1]**

Sol. % ionic character

$$\begin{aligned} &= (0.16 \times 3.3 + 0.035 \times 3.3^2) \times 100 \\ &= 90.9\% \end{aligned}$$

Ex.26 What is the increasing order of ionic character in H_2O , H_2S and H_2Se ?

Sol. Electronegativity difference is maximum in H and O. So $H_2Se < H_2S < H_2O$

Ex.27 What is the increasing order of ionic character in OF_2 and Cl_2O ?

Sol. Electronegativity difference in F–O is greater than in O–Cl. $\therefore Cl_2O < OF_2$

8. COVALENT CHARACTER IN IONIC COMPOUNDS (FAJAN'S RULE)

Some ionic compounds show that covalent character by polarization of anion

(i) **Size of Cation** : Size of cation must be small

$$\text{Covalent character} \propto \frac{1}{\text{Size of cation}}$$

$$\text{ionic character} \propto \text{Size of cation}$$

(ii) **Size of Anion** : Size of anion must be large

$$\text{Covalent character} \propto \text{Size of anion}$$

$$\text{Ionic character} \propto \frac{1}{\text{Size of anion}}$$

(iii) **Charge on Cation and Anion** :

$$\text{Covalent character} \propto \text{Charge on cation and anion}$$

(iv) Inert and Pseudo Inert Structures :

Configuration of cations must be equal to pseudo inert type $[ns^0(n-1)d^{10}]$ e.g. Cu^+ , Ag^+ , Au^+ , Zn^{2+} , Cd^{2+} , Hg^{2+} , Pb^{+4} etc.

Ex.28 AgCl is white, AgBr is yellow and AgI is deep yellow, why ?

Sol. I^- is largest, so polarization of I^- by Ag^+ is maximum and the shared electrons are easily transferred to higher energy level in visible light. So AgI is deep yellow. But in chloride, electrons are transferred easily. So AgCl is colourless similarly.

(i) $SnCl_2$ is white but SnI_2 is yellow.

(ii) $PbCl_2$ is white but PbI_2 is yellow.

(iii) $HgCl_2$ is white but HgI_2 is red.

(iv) $SbCl_3$ is white but SbI_3 is black.

8.1 Applications of Fajan's Rule

Due to polarization of anion compounds gives a specific colour.

8.1.1 Ionic Potential

(i) It is represented by ϕ and is only for cation.

$$\text{Ionic potential} = \frac{\text{Charge on cation}}{\text{Size of cation}}$$

(ii) When the size of cation is small and the charge is high, the value of ϕ is high and the polarizing power of cation is also high.

(iii) In a period, radius decreases but charge increases. So ϕ increases and the covalent character of the bond increases.

$$Li^+ < Be^{+2} < B^{+3} < C^{+4}$$

8.1.2 Application of Ionic Potential**(a) Nature of Bond :**

(i) When $\sqrt{\phi} < 2.2$, the bond is ionic

Example – NaCl, KCl, RbCl, $BaCl_2$, $CaCl_2$, $SnCl_2$, etc.

(ii) If $\sqrt{\phi} > 2.2$, the bond is covalent.

Example – LiCl, $BeCl_2$, BCl_3 , etc,

Covalent character $\propto \sqrt{\phi}$

(b) Acidic and Basic Nature of Metal Oxides :

(i) If $\sqrt{\phi} < 2.2$, the oxide is basic.

Example – Oxides of groups I A and II A.

$$\text{Basic character of oxides} \propto \frac{1}{\sqrt{\phi}}$$

(ii) If $\sqrt{\phi} = 2.2$ to 3.2, the oxide is amphoteric.

Example – Oxides of groups III A and IV A

(iii) If $\sqrt{\phi} > 3.2$ then oxide will be acidic

Example – Oxides of V A, VI A and VII A

(c) Boiling and Melting Points of Compounds :

In a group, $\sqrt{\phi}$ decreases and the ionic character also increases. Therefore, melting points and boiling points increase.

(d) Nature of Metal Halides :

When $\sqrt{\phi}$ of metal < 2.2 , the ionic character and conductivity increase and when $\sqrt{\phi} > 2.2$, the covalent character increases and conductivity decreases.

$$\text{Formation of complex ion} \propto \sqrt{\phi}$$

Ex.29 Which of the following compounds has minimum melting point ?

[1] PbCl_2 [2] SnCl_2 [3] NaCl [4] SnF_4 **Ans. [4]**

Sol. Covalent character \propto Charge on cation

Ex.30 Which of the following compounds has maximum melting point ?

[1] CaF_2 [2] CaCl_2 [3] CaBr_2 [4] CaI_2 **Ans. [1]**

Sol. F^- is smallest, so CaF_2 is most ionic. Thus, melting point and boiling point are maximum.

$$\text{Ionic character} \propto \frac{1}{\text{Size of anion}}$$

Ex.31 The order of ionic size is $\text{Na}^+ > \text{Mg}^{+2}$ and $\text{S}^{-2} > \text{Cl}^-$. Which of these compounds is least soluble in polar solvents ?

[1] MgS [2] Na_2S [3] NaCl [4] MgCl_2 **Ans. [1]**

Sol. Mg^{+2} (small cation) and S^{-2} (large anion), therefore, MgS is covalent, so least soluble in polar solvents.

Ex.32 Melting point of SnCl_2 is 606° while melting point of SnCl_4 is 114°C . Why ?

Sol. SnCl_4 is covalent due to small size of Sn^{+4} .

Ex.33 Which of the following compounds has least ionic character ?

[1] NaCl [2] BaCl_2 [3] KCl [4] HgCl_2 **Ans. [4]**

Sol. Hg^{+2} is pseudo inert ion, but others are inert ions.

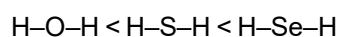
9. BOND LENGTH

The distance between nuclei of two bonded atoms is called bond length.

Factors effecting Bond Length

1. Radius of Bonded Atoms :

Bond length increases with increase in radii of the bonded atoms.

**2. Difference Between Electronegativity of Two Bonded Atoms :**

Bond length can be determined by the relation given by **Shoemaker and Stevenson**.

$$d_{\text{A-B}} = r_{\text{A}} + r_{\text{B}} - 0.09 (\Delta)$$

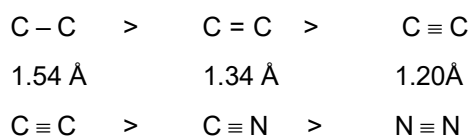
$$d_{\text{A-B}} = \text{Bond length} \quad r_{\text{A}} = \text{Radius of A}$$

$$r_{\text{B}} = \text{Radius of B}$$

$$\Delta = \text{Difference in electronegativity of A and B.}$$

3. Bond Order :

Bond length of single bond is greater than that of the double bond and bond length of double bond is greater than that of triple bond.



4. **Hybridization :**

When s character in hybrid orbitals increases, bond length decreases, because s orbital is attracted by the nucleus with greater force.

$$\text{Bond length} \propto \frac{1}{\text{s character in hybrid orbital}}$$

or bond length \propto p character in hybrid orbital

Ex.34 Which of the following statements is not true ?

[1] N–N bond length in N_2 is less than in NH_2-NH_2

[2] O–O bond length in O_3 is less than in H_2O_2

[3] Bond length in H–Cl is greater than in H–F

[4] Bond length in H_2S is greater than in H_2Se

Ans. [4]

Sol. Electronegativity of S is greater than that of Se. So bond length in H_2S is less than bond length in H_2Se .

Ex.35 Correct order of bond length of N–H in NH_3 , and bond length of N–F in NF_3 is :

[1] N – H = N – F

[2] N–F < N–H

[3] N–F > N–H

[4] None of these

Ans. [3]

Sol. The electronegativity difference of N–F is less than that in N–H. So N–H bond is more polar.

10. **BOND ANGLE**

Bond angle between two adjacent bonds in a molecule is called bond angle.

Factors effecting Bond Angle

1. **Type of Hybridization :**

Bond angle depends on hybridization of an atom. Increasing order of bond angle is as follows :

$$sp^3 < sp^2 < sp$$

Hybridization	sp^3	sp^2	sp
Bond angle	$109^{\circ}28'$	120°	180°

$$\text{Bond angle} \propto \text{s – character in hybrid orbitals}$$

$$\text{Bond angle} \propto \frac{1}{\text{p – character in hybrid orbitals}}$$

2. **Lone Pair of Electrons :**

Bond angle decreases with increase in the number of lone pairs on the central atom.

$$\text{Bond angle} \propto \frac{1}{\text{Number of lone pairs}}$$

3. **Electronegativity of Central Atom :**

Bond angle will be high if the electronegativity value of the central atom is high.

$$\text{Bond angle} \propto \text{Electronegativity of central atom}$$

Molecule	H ₂ O	H ₂ S	H ₂ Se
Bond angle	105°	92°	37.5°

11. BOND ENERGY

The energy released during formation of one mole gaseous substance from two neutral gaseous species is called bond energy of the bond formed.

(a) Dissociation Energy :

The energy required to dissociate the bond into two neutral atoms is called dissociation energy.

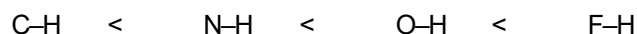
(b) Bond Energy :

More than one bond is present in a polyatomic molecule. So, the average of bond dissociation energies of different bonds is called bond energy.

11.1 Factors Affecting Bond Energy

(1) Atomic Radius :

Bond energy decreases with increase in atomic size because overlapping decreases.



$$\text{Bond energy} \propto \frac{1}{\text{atomic radius}}$$

(2) Bond length :

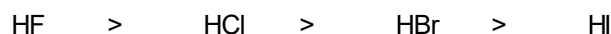
Bond energy increase with decrease in bond length because of more effective orbital overlap.



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

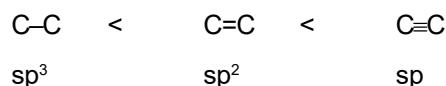
(3) Polarity of the Bond :

Bond energy increases with polarity of bond and polarity depends on difference in electronegativity.



(4) Hybridization :

Bond energy increases with increase in s-character or decrease in p-character in hybrid orbital.

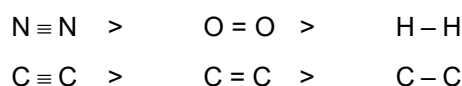


$$\text{Bond energy} \propto \text{s-character in hybrid orbital}$$

$$\text{Bond energy} \propto \frac{1}{\text{p-character in hybrid orbital}}$$

(5) Bond Order :

Bond energy increases with bond order.



Ex.36 Highest bond energy will be.

[1] F – F

[2] Cl – Cl

[3] I – F

[4] Br – Br

Ans. [2]

Sol. Due to small size of F atoms, electron-electron (l.p.) repulsion will be effective, so bond energy of F_2 will be less than Cl_2 .

Ex.37 NF_3 is more stable than NCl_3 . Why ?

Sol. NF_3 is more polar than NCl_3 . So bond energy of will be higher.

Ex.38 The bond energy of I_2 and Cl_2 is 150.6 and 247.8 kJ./mole. Calculate the bond energy of ICl bond.

Sol. $I_2 = 150.6$

$Cl_2 = 247.8$

$I = 75.3$

$Cl = 123.9$

$ICl = 75.3 + 123.9 = 199.2$ KJ/mole.

12. VALENCE SHELL ELECTRON PAIR REPULSION THEORY (VSEPR)

☆ This theory was put forward by Sidgwick and Powell in 1940 and was further improved by Nyholm and Gillespie in 1957

☆ The electron pairs surrounding the central atom repel one another and move so far apart from one another that there are no further repulsions between them. As a result, the molecule has minimum energy and maximum stability.

☆ The basic assumptions of the VSEPR theory

(1) The shape of a molecule containing only two atoms is always linear.

(2) For molecules containing 3 or more atoms, one of the atoms is called the central atom to which other atoms are linked.



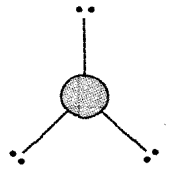
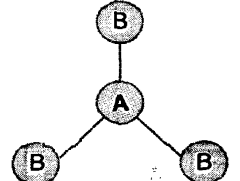
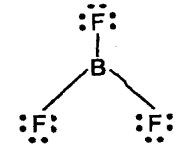
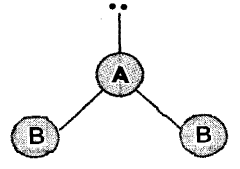
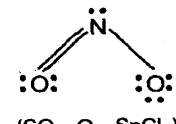
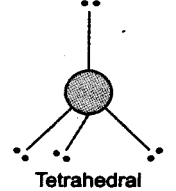
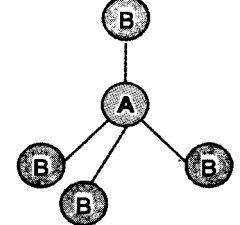
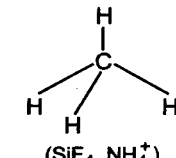
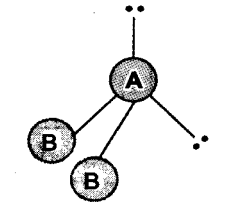
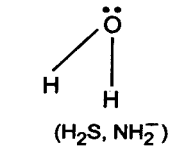
(3) If the central atom is linked to similar atoms and is surrounded by bond pairs of electrons only, the repulsions between them are similar. As a result, the shape of the molecule is *symmetrical* and the molecule is said to have a *regular geometry*.

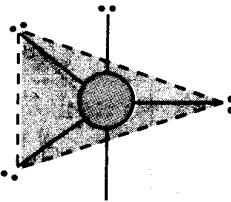
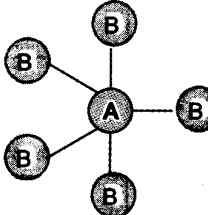
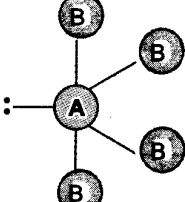
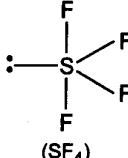
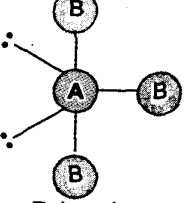
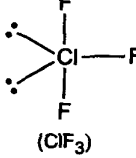
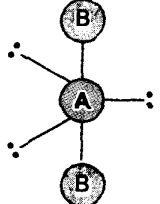
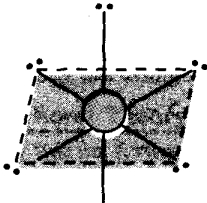
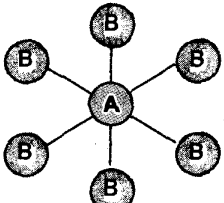
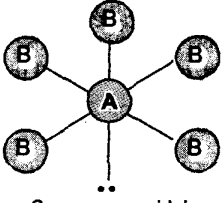
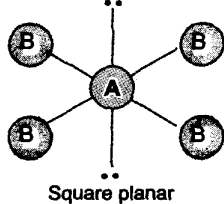
(4) If the central atom is linked to different atoms or is surrounded by bond pairs as well as lone pairs of electrons, the repulsions between them are different. As a result, the molecule has an *irregular* or *distorted geometry*. The order of repulsions between electron pairs is as follows :

Lone pair—Lone pair > Lone pair-Bond pair > Bond pair-Bond pair.

(5) The exact shape of the molecule depends upon the total number of electron pairs present around the central atom.

Table : Shapes (Geometry) fo molecules containing bond pairs only or bond pairs and lone pairs

Total No. of electron pairs	Geometry of the electron pairs	Bond pairs	Lone pairs	Geometry (Shape) of the molecule	Illustrative Example (Other Examples)
		(Molecular Formula)			
2	 <p>Linear</p>	2	0	 <p>Linear</p>	$\ddot{\text{O}} = \text{C} = \ddot{\text{O}}$ (BeF ₂ , BeCl ₂ , HgCl ₂)
3	 <p>Triangular planar</p>	3	0	 <p>Triangular planar</p>	 <p>(AlCl₃, SO₃)</p>
		(AB ₃)			
		2	1	 <p>Bent (V-shape)</p>	 <p>(SO₂, O₃, SnCl₂)</p>
		(AB ₂ L)			
4	 <p>Tetrahedral</p>	4	0	 <p>Tetrahedral</p>	 <p>(SiF₄, NH₄⁺)</p>
		(AB ₄)			
				3	1
(AB ₃ L)					
		2	2	 <p>Bent</p>	 <p>(H₂S, NH₂⁻)</p>
(AB ₂ L ₂)					

5	 <p>Trigonal bipyramidal</p>	5	0	 <p>Trigonal bipyramidal</p>	<p>PCl₅ (PF₅, SbCl₅, AsF₅)</p>
		4	1	 <p>See saw</p>	 <p>(SF₄)</p>
		3	2	 <p>T-shaped</p>	 <p>(ClF₃)</p>
		2	3	 <p>Linear</p>	<p>XeF₂ (I₃⁻, Br₃⁻, ICl₂⁻)</p>
6	 <p>Octahedral</p>	6	0	 <p>Octahedral</p>	<p>SF₆</p>
		5	1	 <p>Square pyramidal</p>	<p>ClF₅ (IF₅, BrF₅)</p>
		4	2	 <p>Square planar</p>	<p>XeF₄</p>

☆ **Calculation of total number of electron pairs, bond pairs and lone pairs and predicting the shapes of the molecules and ions.**

(i) Total no. of electron pairs around the central atom

$$= \frac{1}{2}(\text{No. of valence electrons of central atom} + \text{No. of atoms linked to central atom by single bonds})$$

For negative ions, add number of electrons equal to the units of negative charge on the ion to the valence electrons of the central atom.

For positive ions, subtract number of electrons equal to the units of positive charge on the ion from the valence electrons of the central atom.

(ii) No. of bond pairs (shared pairs) = No. of atoms linked to central atom by single bonds.

(iii) No. of lone pairs = Total no. of electron pairs - No. of shared pairs.

13. VALENCE BOND THEORY

This theory was first put forward by **Heitler and London** in 1927.

(A) Covalent bond is formed between atoms by sharing of unpaired electrons of the outermost shell in order to attain greater stability.

(B) Overlapping

(i) Overlapping does not take place in ionic bond, only takes place in covalent bond.

(ii) The orbitals of unpaired electrons overlap in the formation of covalent bond.

(iii) When overlapping is coaxial, a σ bond is formed and when overlapping is collateral, a π bond is formed.

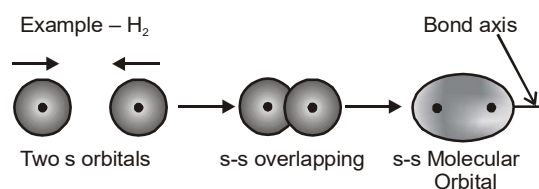
(iv) A σ bond is stronger than a π bond.

(v) A π bond is never formed alone as it is very weak. It is formed after a σ bond is formed.

(vi) s-Orbital always forms σ bond.

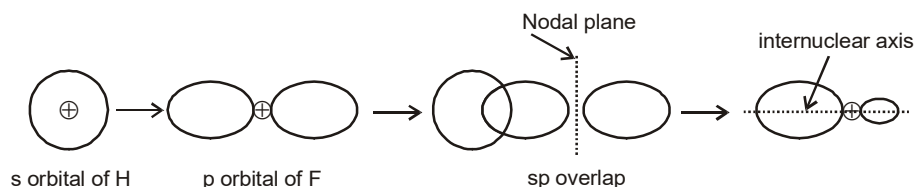
13.1 Types of Overlapping

(i) **s-s Overlapping** – When two s orbitals of two atoms each having an unpaired electron, participate in overlapping, it is called s-s overlapping.

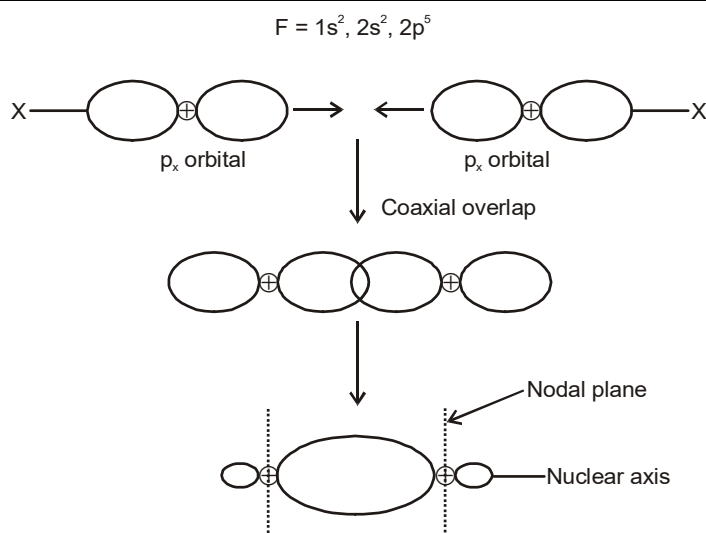


Formation of H_2 by s-s overlapping

(ii) **s-p Overlapping** – When s orbital of an atom and p orbital of the other atom, each having one unpaired electron, participate in overlapping, a sausage-shaped σ molecular orbital is formed.



(iii) **p-p Overlapping** – Two p orbitals can form either a σ bond or a π bond by different type of overlapping. When overlapping of p orbital is coaxial, a σ bond is formed and when it is collateral, a π bond is formed.



The order of bond strength is as follows

1. p-p (coaxial) > s-p > s-s
2. p-p (coaxial) > p-p- (collateral)

Ex.39 Four bonds are never formed between two atoms. Why ?

Sol. There are three axes (x, y and z) perpendicular to each other. A maximum of two π bonds can be formed by collateral overlapping and one σ bond is formed by coaxial overlapping. Thus, highest bond order between two atoms can be three and never four.

Ex.40 p-p Overlapping is not possible in :

- [1] Cl_2 [2] O_2 [3] N_2 [4] H_2 **Ans. [4]**

Sol. p-Orbital is not present in H atom.

Ex.41 A σ bond is formed in O_2 by overlapping of :

- [1] $2p_y - 2p_z$ [2] $2p_y - 2s$ [3] $2p_z - 2s$ [4] $2p_z - 2p_z$ **Ans. [4]**

Sol. Extent of $p_y - p_z$ and $p_y - s$ overlapping is zero.

Ex.42 Which of the following orbitals will form a weak bond by the overlapping ?

- [1] $sp^2 - sp^2$ [2] $sp^2 - s$ [3] $s - s$ [4] $p - p$ (coaxial) **Ans. [3]**

Sol. Due to non directional character of s orbital, bond formed by s-s orbitals will be weak.

Ex.43 Which of the following overlapping will form a π bond ?

- [1] $sp - p$ [2] $sp^2 - s$ [3] $sp^3 - sp^3$ [4] None of these **Ans. [4]**

Sol. Hybrid orbitals always form a σ bond.

14. HYBRIDIZATION

- (1) Hybridization occurs in the central atom of a compound, e.g. in CH_4, H_2O .
- (2) Generally, the orbitals of the outermost shell of the central atom participate in hybridization.
- (3) In hybridization, all types of orbitals (empty and half-filled and fully-filled) can participate.
- (4) If the orbitals participating in hybridization have paired electrons, the new orbitals formed will also have paired electrons.
- (5) The σ bonds formed by hybrid orbitals are strong and their orientation is different in space.
- (6) All hybrid orbitals are generally identical having one larger lobe and one smaller lobe.

$$\text{Total no. of hybrid orbitals} = \text{No. of } \sigma \text{ bonds} + \text{No. of lone pairs}$$

- (7) In hybridization, the orbitals of the outermost shell of central atom having almost equal energy, combine and form new hybrid orbitals of identical energy.
- (8) The number of hybrid orbitals formed will be equal to the number of orbitals participating in hybridization e.g. 2, 3 and 4 hybrid orbitals are formed in sp , sp^2 and sp^3 hybridization, respectively.
- (9) Orbitals having unpaired electrons overlap for pairing and overlapping occurs in the orbitals having paired electrons.
- (10) Bond pair is formed by overlapping of hybrid orbitals. So lone pair or bond pair is present in a hybrid orbital.
- (11) The hybridization of central atom of a molecule can be determined by the following formula.

$$\text{Total no. of hybrid orbitals} = \text{No. of } \sigma \text{ bonds} + \text{No. of lone pairs}$$

- (12) Hybrid orbitals have directional property, so more the directional property of hybrid orbital, stronger will be the bond formed : $sp < sp^2 < sp^3 < sp^3d < sp^3d^2 < sp^3d^3$

14.1 Types of Hybridisation

sp^3 Hybridisation :

- (i) In this hybridisation, one s orbital and three p orbitals join up and form four sp^3 hybrid orbitals of same energy.
- (ii) The geometry of the compound will be tetrahedral and bond angle will be about $109^\circ 28'$. e.g. CH_4 , CCl_4 & Diamond

sp^2 Hybridisation :

- (i) One s and two p orbitals join up and form three new sp^2 hybrid orbitals.
- (ii) The geometry of the molecule will be trigonal and bond angle will be 120° .

Examples – BF_3 , BCl_3 , BH_3 , C_2H_4 and graphite.

sp Hybridisation :

- (i) In sp hybridisation, one s orbital and one p orbital join up and form two sp hybrid orbitals of equal energy.
- (ii) Geometry of along sp hybrid orbital will be linear and bond angle will be 180° .
- (iii) Two σ bonds will be formed by sp hybridised atom.
- (iv) The dipole moment will be zero.

Examples – BeF_2 , $BeCl_2$, C_2H_2 , CO_2 etc.

sp^3d Hybridisation :

One s, three p and one d orbitals join up and form five sp^3d hybrid orbitals of same energy.

Example – PCl_5 , ClF_3 , SF_4 , XeF_2 , $TeCl_4$, ICl_2^+ etc.

sp^3d^2 Hybridisation :

One s, three p and two d orbitals combine and form six sp^3d^2 hybrid orbitals of same energy.

Example – SF_6 , ClF_5 , XeF_4 , IF_4 , SiF_6^{2-} , ICl_4^- .

sp^3d^3 Hybridisation :

One s, three p and three d orbitals join up and form seven new sp^3d^3 hybrid orbitals of same energy.

Example – IF_7 , SeF_6 , etc.

Ex.44 Which of the following molecules is bent ?

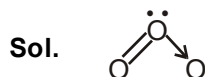
[1] CO_2

[2] O_3

[3] N_2O

[4] None of these

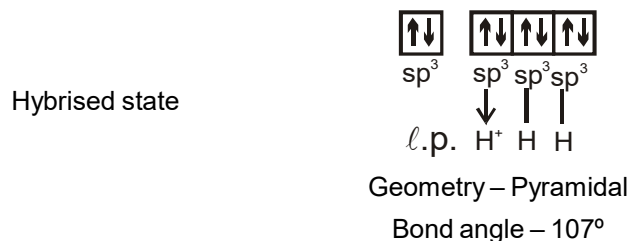
Ans. [2]



Ex.45 What is the type of hybridisation in H_3O^+ ?

Sol. O atom is sp^3 hybridised and has one lone pair. Its geometry is pyramidal.

Ground state $O^8 = 1s^2, 2s^2 2p^4$



Ex.46 Which of the following species is planar ?

- [1] NH_4^+ [2] BF_4^- [3] XeF_4 [4] $SiCl_4$ **Ans. [3]**

Sol. Types of hybridisation are sp^3 in NH_4^+ , BF_4^- and $SiCl_4$ and sp^3d^2 in XeF_4 .

Ex.47 Which of the following ions has tetrahedral geometry ?

- [1] Na^+ [2] NH_4^+ [3] Mg^{+2} [4] CO_3^{-2} **Ans. [2]**

Sol. NH_4^+ is sp^3 hybridisation so structure will be tetrahedral.

Ex.48 What should be the geometry of the molecule when there are three bond pairs and one lone pair on the central atom ?

- [1] Tetrahedral [2] Pyramidal [3] Trigonal [4] Octahedral **Ans. [2]**

Sol. For example, NH_3 has three N–H bond pair and one lone pair. Its geometry is trigonal pyramidal.

Ex.49 Which of the following molecules has sp^2 hybridisation on N atom ?

- [1] NH_3 [2] $NOCl$ [3] N_2H_4 [4] All of the above **Ans. [2]**

Sol. Nitrogen atom in $Cl-\ddot{N}=\ddot{O}$ has one lone pair, two σ bond pairs and one π bond pair. So, N atom is in a state of sp^2 hybridisation.

Ex.50 If hybridisation on Co atom is sp^3d^2 , the geometry of $[Co(NH_3)_6]^{+3}$ will be similar to that of :

- [1] SiF_6^{-2} [2] SF_6 [3] both 1 and 2 [4] None of these **Ans. [3]**

Sol. S is in a state of sp^3d^2 hybridisation in SiF_6^{-2} and SF_6 and the geometry of sp^3d^2 and d^2sp^3 will be octahedral.

14.2 Hybridisation in Transition Metal Ions

In sp^3d hybridisation, d orbital of the outermost orbit participates. But in dsp^3 hybridisation, the d orbital of inner orbit ($n - 1$), which is vacant in transition elements, participates.

14.3 Examples on sp hybridisation

Example	σ bond	l.p.e.	Hybridisation	Bond angle	Shape
BeH ₂	2	-	sp	180°	Linear
BeF ₂	2	-	sp	180°	Linear
BeCl ₂	2	-	sp	180°	Linear
BeBr ₂	2	-	sp	180°	Linear
BeI ₂	2	-	sp	180°	Linear
CO ₂	2	-	sp	180°	Linear
CO	1	1	sp	180°	Linear
NO ₂ ⁺	2	-	sp	180°	Linear
C ₂ H ₂	2	-	sp	180°	Linear
HCN	2	-	sp	180°	Linear
ZnCl ₂	2	-	sp	180°	Linear
HgCl ₂	2	-	sp	180°	Linear
CdCl ₂	2	-	sp	180°	Linear
R-Mg-X	2	-	sp	180°	Linear
[Ag(CN) ₂] ⁻	2	-	sp	180°	Linear
[Cu(CN) ₂] ⁻	2	-	sp	180°	Linear
N ₂ O	2	-	sp	180°	Linear
N ₃ ⁻	2	-	sp	180°	Linear

Examples on sp^2 hybridisation

Example	σ bond	l.p.e.	Hybridisation	Bond angle	Shape
BH ₃	3	-	sp ²	120°	Trigonal planar
BF ₃	3	-	sp ²	120°	Trigonal planar
BCl ₃	3	-	sp ²	120°	Trigonal planar
CH ₃ (radical)	3	-	sp ²	120°	Trigonal planar
CH ₂ =CH ₂	3,3	-	sp ²	120°	Trigonal planar
Graphite	3	-	sp ²	120°	Trigonal planar
HNO ₃	3	-	sp ²	120°	Trigonal planar
NO ₃ ⁻	3	-	sp ²	120°	Trigonal planar
HNO ₂	2	1	sp ²	<120°	Angular (V)
SO ₂	2	1	sp ²	<120°	Angular (V)
SO ₃	3	-	sp ²	120°	Trigonal planar
HCO ₃ ⁻	3	-	sp ²	120°	Trigonal planar
CO ₃ ⁻²	3	-	sp ²	120°	Trigonal planar
SnCl ₂	2	1	sp ²	<120°	Angular (V)
SnBr ₂	2	1	sp ²	<120°	Angular (V)
SnI ₂	2	1	sp ²	<120°	Angular (V)
AlCl ₃	3	-	sp ²	120°	Trigonal planar
GaCl ₃	3	-	sp ²	120°	Trigonal planar
PbCl ₂	2	1	sp ²	<120°	Angular (V)

Examples on sp^3 Hybridisation

Example	σ bond	l.p.e.	Hybridisation	Shape
CH_4	4	-	sp^3	Tetrahedron
CCl_4	4	-	sp^3	Tetrahedron
CBr_4	4	-	sp^3	Tetrahedron
PCl_4^+	4	-	sp^3	Tetrahedron
NH_4^+	4	-	sp^3	Tetrahedron
BF_4^-	4	-	sp^3	Tetrahedron
AlF_4^-	4	-	sp^3	Tetrahedron
GaF_4^-	4	-	sp^3	Tetrahedron
BeF_4^{2-}	4	-	sp^3	Tetrahedron
MgF_4^{2-}	4	-	sp^3	Tetrahedron
$Ni(CO)_4$	4	-	sp^3	Tetrahedron
$Zn(CN)_4^{2-}$	4	-	sp^3	Tetrahedron
$Cd(CN)_4^{2-}$	4	-	sp^3	Tetrahedron
$Hg(CN)_4^{2-}$	4	-	sp^3	Tetrahedron
NF_3	3	1	sp^3	Pyramidal
NCl_3	3	1	sp^3	Pyramidal
$N(CH_3)_3$	3	1	sp^3	Pyramidal
PF_3	3	1	sp^3	Pyramidal
PCl_3	3	1	sp^3	Pyramidal
$AsCl_3$	3	1	sp^3	Pyramidal
$SbCl_3$	3	1	sp^3	Pyramidal
$BiCl_3$	3	1	sp^3	Pyramidal
NH_3	3	1	sp^3	Pyramidal
$:CH_3$	3	1	sp^3	Pyramidal
H_3O^+	3	1	sp^3	Pyramidal
SO_3^{2-}	3	1	sp^3	Pyramidal
ClO_3	3	1	sp^3	Pyramidal
XeO_3	3	1	sp^3	Pyramidal
H_2O	2	2	sp^3	Angular (V)
H_2S	2	2	sp^3	Angular (V)
NH_2^-	2	2	sp^3	Angular (V)
OF_2	2	2	sp^3	Angular (V)
Cl_2O	2	2	sp^3	Angular (V)
SCl_2	2	2	sp^3	Angular (V)
Diamond	4	-	sp^3	Tetrahedron
SiO_2	4	-	sp^3	Tetrahedron
SiC	4	-	sp^3	Tetrahedron

EXAMPLES ON sp^3d HYBRIDISATION

Example	σ bond	L.p.e.	Hybridisation	Bond angle	Shape
PF_5	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
PCl_5	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
AsF_5	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
$AsCl_5$	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
$SbCl_5$	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
$BiCl_5$	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
SbF_5	5	-	sp^3d	120° & 90°	Trigonal bipyramidal
SF_4	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
SeF_4	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
TeF_4	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
PoF_4	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
PF_4^-	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
AsF_4^-	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
SbF_4^-	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
SCl_4	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
$SeCl_4$	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
$TeCl_4$	4	1	sp^3d	$<180^\circ, <90^\circ, <120^\circ$	Irregular tetrahedron
ClF_3	3	2	sp^3d	87.6°	T-shape
BrF_3	3	2	sp^3d	87.6°	T-shape
IF_3	3	2	sp^3d	87.6°	T-shape
ICl_3	3	2	sp^3d	87.6°	T-shape
I_3^-	2	3	sp^3d	180°	Linear
Br_3^-	2	3	sp^3d	180°	Linear
ICl_2^-	2	3	sp^3d	180°	Linear
IBr_2^-	2	3	sp^3d	180°	Linear
ClF_2^-	2	3	sp^3d	180°	Linear
XeF_2	2	3	sp^3d	180°	Linear

15. MOLECULAR ORBITAL THEORY

- (a) There is another approach to chemical bonding known as molecular orbital theory (MOT) developed by Mulliken (1932) and Hund. Which explains the bonding characteristics in a better way.
- (b) The molecular orbital theory considers the entire molecule as a unit with all the electrons moving under the influence of all the nuclei present in the molecular.
- (c) This approach recognizes that each electron belongs to the molecule as a whole and may move within the entire molecule.

15.1 Important Features of M.O.T.

- (i) Like an A.O. (atomic orbital) which is around the nucleus of an atom there are M.O. (Molecular orbital) which are around the nuclei of a molecule.
- (ii) The molecular orbitals are entirely different from the atomic orbitals from which they are formed.
- (iii) The molecular orbitals possess different energy levels like atomic orbitals in an isolated atom.
- (iv) The shape of molecular orbitals are dependent upon the shapes of atomic orbitals from which they are formed.
- (v) Molecular orbitals are arranged in order of increasing energy just like atomic orbitals.

- (vi) The number of molecular orbitals formed is equal to the number of atomic orbitals combining in bond formation.
- (vii) Like atomic orbitals, the filling of electron in molecular orbitals is governed by the three principles such as Aufbau principle, Hund's rule and Pauli's exclusion principle.

15.2 Conditions for atomic orbitals to form M.O.

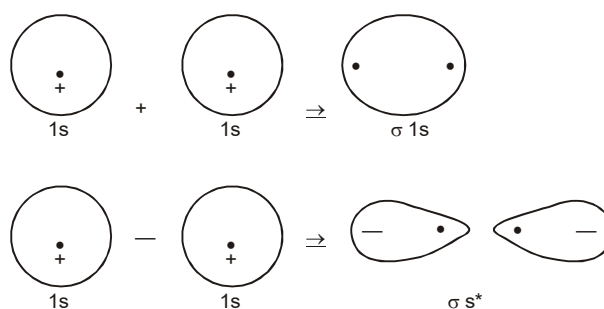
- (i) The combining A.O. should be a comparable energy
- (ii) The combining atomic orbitals must overlap to a large extent greater the overlap, stable is the molecule formed.

15.3 Different between A.O. and M.O.

A.O. (Atomic orbital)	M.O. (Molecular Orbital)
1. An e ⁻ in an A.O. is under the influence of only one nucleus.	An e ⁻ in a M.O. is under the influence of nuclei of two or more atoms of a molecule.
2. Their existence is because of inherent property of the atoms	These are formed by the combination of atomic orbitals of comparable energies
3. They are less stable than bonding M.O. Which are more stable than antibonding M.O.	They are less or more stable than A.O.
4. They have simple shapes	They have complex shapes
5. They are represented by s, p, d, f	They are represented by σ , σ^* , π , π^* etc.

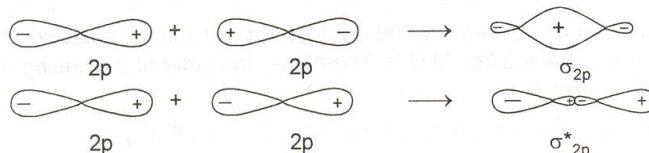
15.4 Formation of Bonding and Antibonding Molecular orbitals (LCAO method)

- (i) An electron in an atom is described by a wave function, ψ called an atomic orbitals. Similarly, the behavior of an electron in a molecule is described by a molecular wave function called molecular orbitals.
- (ii) Quantum mechanics shows that linear combination of two functions gives, not one but two combination and hence two molecule orbitals; a bonding orbital and an antibonding orbital.
- (iii) Such an orbital. Obviously cannot lead to the formation of a chemical bond and is therefore, termed as an antibonding M.O. combination of atomic orbitals-sigma (σ) and Pi (π) molecular orbitals.
- (iv) In the M.O. theory, orbitals are identified as σ or π depending upon the type of the symmetry of the molecular orbital.
- (v) A sigma (σ) M.O. is one that has cylindrical symmetry around the internuclear axis, i.e. it does not show any change of sigma on rotation through 180° about the axis.
- (vi) It can be said that a sigma M.O. has no nodal plane (in which the electron density is zero) along the inter nuclear axis.

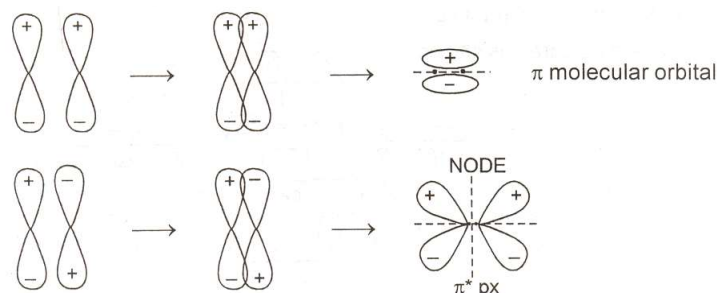


Formation of σ 1s and σ^* is bonding anti bonding M.O.

- (vii) The bonding orbital is designated simply as σ orbital and antibonding as σ^* .
- (viii) Such a sigma (σ) orbital is also formed when any two p-atomic orbitals overlap in end-on (along their axes) position.

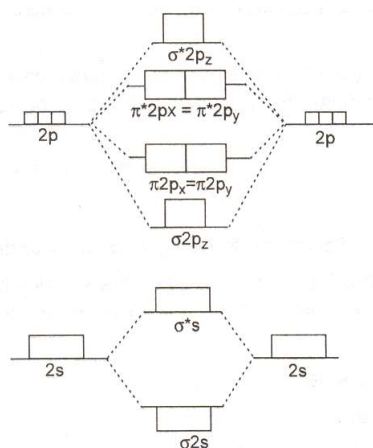


- (ix) Since the energy of a molecular orbital is directly related to the number of nodal planes, the π -orbital is more energetic than the σ -orbital. This explains why a π -bond is weaker than a σ -bond.



15.5 Relative energies of M.O. and filling of e⁻

- (i) Two general criteria, which determine the energy of the molecular orbitals are
- Initial energy of the atomic orbitals
 - The extent of the overlap of the atomic orbitals
- (ii) It is obvious that molecular orbitals formed from lower energy atomic orbitals have lower energy than the molecular orbitals formed from higher energy atomic orbitals.
- (iii) As the σ overlap is much more effective than π -overlap, σ p-molecular orbital is of lowest energy, even though originally all the three π -orbitals are of equal energy.
- (iv) The relative energies of the M.O. are obtained experimentally from spectroscopic data.
- (v) The sequence in the order of increasing energy is given below i.e., for less than N_2 .
- $$1s < \sigma^*1s < \sigma 2s < \sigma^*2s < \pi 2p_x = \pi 2p_y < \sigma 2p_z < \pi^* 2p_x = \pi^* 2p_y < \sigma^* 2p_z$$
- (vi) The sequence in the order of increasing energy is given below i.e., for more than N_2 .
- $$1s < \sigma^*1s < \sigma 2s < \sigma^*2s < \sigma 2p_z < \pi 2p_x = \pi 2p_y < \pi^* 2p_x = \pi^* 2p_y < \sigma^* 2p_z$$



M.O. Energy level diagram for O_2 , F_2 and Ne .

- (vi) Using preceding figures for respective molecules as a guide, we can build up the electronic structure of simple diatomic molecules.
- (vii) The filling of electrons in various M.O. follows similar rules to those for filling A.O. similar M.O. belongs to entire molecule, the total number of electrons are taken into consideration and accommodated in molecular orbitals according to the following rules.
- Aufbau principle : The M.O. of lower energy is filled up first.
 - Pauli exclusion principles. Each M.O. can accommodate a maximum of two electrons having opposite spins.
 - Hund's rule. Whenever two equal energy M.O. are available electrons prefer to remain unpaired. This is particularly significant for π -orbitals because π -bonding and π -antibonding M.O. have equal energies.

15.6 Stability of Molecules**(A) Stability of molecules in terms of bonding and antibonding electrons :**

(i) If $N_b > N_a$, the molecule is stable.

(ii) If $N_b < N_a$, the molecule is unstable.

(iii) Even if $N_b = N_a$, the molecule is unstable. This is due to the fact that antibonding effect is somewhat stronger than the bonding effect.

(B) Stability of molecules in terms of bond order :

(a) The relative stability of a molecule is further evaluated by a parameter known as bond order.

(b) It can be defined as number of covalent bonds formed in molecule.

(c) It is given by one half of the difference between the number of electrons in bonding orbitals and those in antibonding orbitals.

(d) Bond order = $\frac{1}{2}$ [No. of electrons in bonding orbitals – No. of electrons in antibonding orbitals]

$$= \frac{1}{2} [N_b - N_a]$$

(e) The bond order of 1, 2 and 3 corresponds to single, double and triple bonds respectively. It may be mentioned that according to MO theory, even a fractional bond order is possible.

(f) Stability of molecule \propto bond order

(g) Dissociation energy \propto bond order

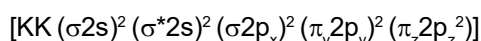
(h) bond length $\propto \frac{1}{\text{bond order}}$

(i) If all the electrons in a molecule are paired the substance is diamagnetic on the other hand if there are unpaired electron in the molecules, the substance is paramagnetic. More the number of unpaired electron in the molecule greater is the paramagnetism of the substance.

15.7 M.O. Configuration

(A) Homonuclear diatomic species : These types of ions have two identical atoms linked together of A_2 type.

N_2 : (a) The electron configuration of N is $1s^2, 2s^2, 2p^3$. In N_2 molecule, there are 14 electrons of which 4 are in K shell and are nonbonding. Each atom contributes four atomic orbitals and thus, the molecule consists of 8 M.O. Then ten electron are allotted to M.O. as follows



where KK denotes the closed K shell structure $(\sigma 1s)^2 (\sigma^* 1s)^2$

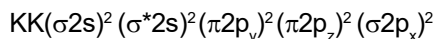
(b) Bond order $\rightarrow \frac{1}{2} (8 - 2) = 3$

(c) Thus N_2 molecule has three bonds, one sigma and two pi.

(B) Heteronuclear diatomic molecules

(a) CO : (a) We know that total number of valence electrons = $8 + 6 = 14$

The distribution of electrons in various molecular orbitals of CO molecule is done in the following manner



(b) Bond order = $\frac{1}{2} (8 - 2) = 3$

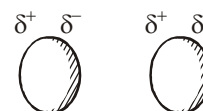
(c) It represents one sigma bond and two π -bonds (d) The molecule is stable and diamagnetic in nature.

16. SECONDARY FORCES :**16.1 Vander waal's forces**

(I) These are weak; non directional, non valence force of attraction among neutral species.

(II) These are electrical in nature, due to induced polarity caused by temporary displacement of electrons towards one end of the inert atoms, becoming a temporary dipole.

(III) This temporary dipole in one molecule can induce opposite dipoles in surrounding molecule. due to displacement of electrons, one end becomes -ve and other +ve. These partially charged ends, induce surrounding molecules accordingly.



(IV) Strength of vander waal force depends on ease of distortion of electron cloud.

Vander waal force \propto size of atom or molecule \propto atomic wt. or molecular weight.

(V) Therefore m.p. and b.p. of noble gas elements (inert gas atom) and halogens increases down the group.

(VI) These force originate from three different causes -

(a) **Keesom force or dipole dipole force** - one dipole molecule orient the other dipole molecule, to bring opposite ends close to each other for attraction. This is called orientation effect. Example - HCl, H₂O, NH₃ etc.

(b) **Debye force or dipole induce dipole force** :- Forces which results from the interaction of a permanent dipole and induced dipole are called Debye force.

eg. When non-polar substance like benzene come in contact with polar molecules like NH₃, induced dipole moment in benzene appears (Induction effect).

(c) **London force or Instantaneous dipole instantaneous induce dipole attraction** :-

Due to vibration or moment of atom electron cloud and nuclei shifts temporarily towards opposite ends, leads to attraction between them, eg. In atoms of noble gases.

16.2 Hydrogen Bond

(1) Hydrogen bond is formed in covalent compounds and not in ionic compounds.

(2) H-bond is not formed in nonpolar covalent compounds.

(3) H-bond is the attraction between H atom and strongly electronegative atom.

16.2.1 Intermolecular Hydrogen Bond : Bond between two nearest molecules.

(a) **Homointermolecular H-Bond** : This type of H-bond is formed between same molecules. For example, H₂O, HF, NH₃, ROH, CH₃COOH, etc.

(b) **Heterointermolecular H-Bond** : The type of H-bond is formed between molecules of different compounds.

16.2..2 Intramolecular Hydrogen Bond : Bond in a single molecule e.g. Nitrophenol

(i) Due to higher electronegative in nature of O, electron density transfers from electron-rich O^{δ-} to electron deficient H^{δ+} bonded to the electronegative atom of the other molecule.

(ii) H-bond is formed between two molecules, so this is called intermolecular H-bond.

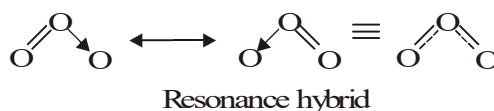
(iii) One H₂O molecule can form a maximum of four H-bonds.

17. RESONANCE

(a) The concept of resonance was introduced by Heisenberg (1920), and later developed by Pauling and Ingold, to explain the properties of certain molecules.

(b) It has been found that the observed properties of certain compounds cannot be satisfactorily explained by writing a single Lewis structure. The molecule is then supposed to have many structures, each of which can explain most of the properties of the molecule but none can explain all the properties of the molecules. The actual structure is in between of all these contributing structures and is called resonance hybrid and the different individual structures are called resonating structures or canonical forms. This phenomenon is called resonance.

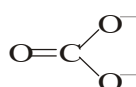
(c) Let us discuss resonance in ozone, according to its resonance structure it should have one single bond (O—O = 1.48Å) but experiments show that both the bonds are same which can be proved by its resonance hybrid as shown below.



To calculate bond order in the polyatomic molecule or ion use following formula :

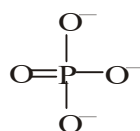
$$\text{Bond order} = \frac{\text{Total number of bonds in a molecule}}{\text{Resonating Structures}}$$

e.g.



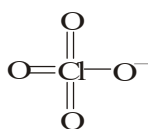
C — O

$$\text{Bond order} = \frac{4}{3} = 1.33$$



P — O

$$\text{Bond order} = \frac{5}{4} = 1.25$$



$$\text{Bond order} = \frac{7}{4} = 1.75$$

18. METALLIC BOND

- The constituent particles of metallic solids are metal atoms which are held together by metallic bond.
- In order to explain the nature of metallic bond Lorentz proposed a simple theory known as electron gas model or electron sea model.
- A metal atom is supposed to consist of two parts, valence electrons and the remaining part (the nucleus and the inner shells) called kernel.
- The kernels of metal atoms occupy the lattice sites while the space between the kernel is occupied by valence electrons.
- Due to small ionisation energy the valence electrons or metal atoms are not held by the nucleus firmly. Therefore, the electrons leave the field of influence of one kernel and come under the influence of another kernel. Thus the electrons are not localised but are mobile.
- The simultaneous attraction between the kernels and the mobile electrons which hold the kernel together is known as metallic bond.

19. SOME USEFUL FORMULAE**Ionic character :**

- \propto Melting and boiling point
- \propto Solubility (Polar solvent)
- \propto Conductivity (Fused state)
- \propto Size of cation
- $\propto \frac{1}{\text{Size of anion}}$
- $\propto \frac{1}{\text{Charge on cation \& anion}}$
- \propto Inert configuration
- $\propto \frac{1}{\text{Pseudo inert configuration}}$
- \propto Difference in electronegativity of the covalently bonded atoms
- \propto Dipole moment

Covalent character :

- $\propto \frac{1}{\text{Melting \& Boiling point}}$
- $\propto \frac{1}{\text{Solubility}}$
- $\propto \frac{1}{\text{Conductivity}}$
- $\propto \frac{1}{\text{Hardness}}$
- $\propto \frac{1}{\text{Size of cation}}$
- \propto Size of anion
- \propto Charge an ions
- $\propto \frac{1}{\text{Inert configuration}}$
- \propto Pseudo inert configuration
- $\propto \frac{1}{\text{Dipole moment}}$

Bond length :

- \propto Size of bonded atoms
- $\propto \frac{1}{\text{Difference of electronegativity of bonded atoms}}$
- $\propto \frac{1}{s \text{ Character in hybrid orbital}}$
- $\propto p \text{ Character in hybrid orbital}$
- $\propto \frac{1}{\text{Resonance}}$

Bond angle :

- $\propto \frac{1}{\text{Number of lone pair}}$
- \propto Electronegativity of central atom
- $\propto \frac{1}{\text{Electronegativity of terminal atom}}$
- $\propto s \text{ character in hybrid orbitals}$

Bond energy :

- $\propto \frac{1}{\text{Atomic size}}$
- $\propto \frac{1}{\text{Bond length}}$
- $\propto s \text{ character in hybrid orbitals}$
- \propto difference of electronegativity values of bonded atoms
- $\propto \frac{1}{\text{Lone pair present on bonded atom}}$

