Chemical bonding

Most important question answer

Q1Write the favourable factors for the .formation of ionic bond.

Ans : 1. Low Ionization Enthalpy of Metal
2. High Electron Affinity of Non-metal
3. High Lattice Energy
4. Large Electronegativity Difference
5. Low Electron Gain Enthalpy (Negative Value)

Q2. **Although geometries of NH_3 and H_2O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.**

Ans:

- Both NH_3 and H_2O have sp^3 hybridized central atoms and form distorted tetrahedral geometries.
- However:
  - NH_3 has 1 lone pair and 3 bond pairs
  - H_2O has 2 lone pairs and 2 bond pairs
- Lone pair–lone pair repulsion > lone pair–bond pair repulsion > bond pair–bond pair repulsion
- Hence, in H_2O, there is greater repulsion due to 2 lone pairs, which compresses the bond angle more than in NH_3.

Q3. **Discuss the shape of the following molecules using the VSEPR model**
Molecules: $BeCl_2$, BCl_3, $SiCl_4$, AsF_5, H_2S, PH_3

Ans: (i) $BeCl_2$

- Central atom: Be
- Valence electrons of Be: 2
- Bond pairs: 2 (with 2 Cl atoms)
- Lone pairs: 0
- Geometry: Linear
- Bond angle: 180°

(ii) BCl_3

- Central atom: B
- Valence electrons of B: 3
- Bond pairs: 3
- Lone pairs: 0
- Geometry: Trigonal planar

(iii) $SiCl_4$

- Central atom: Si
- Valence electrons of Si: 4
- Bond pairs: 4
- Lone pairs: 0
- Geometry: Tetrahedral

(iv) AsF_5

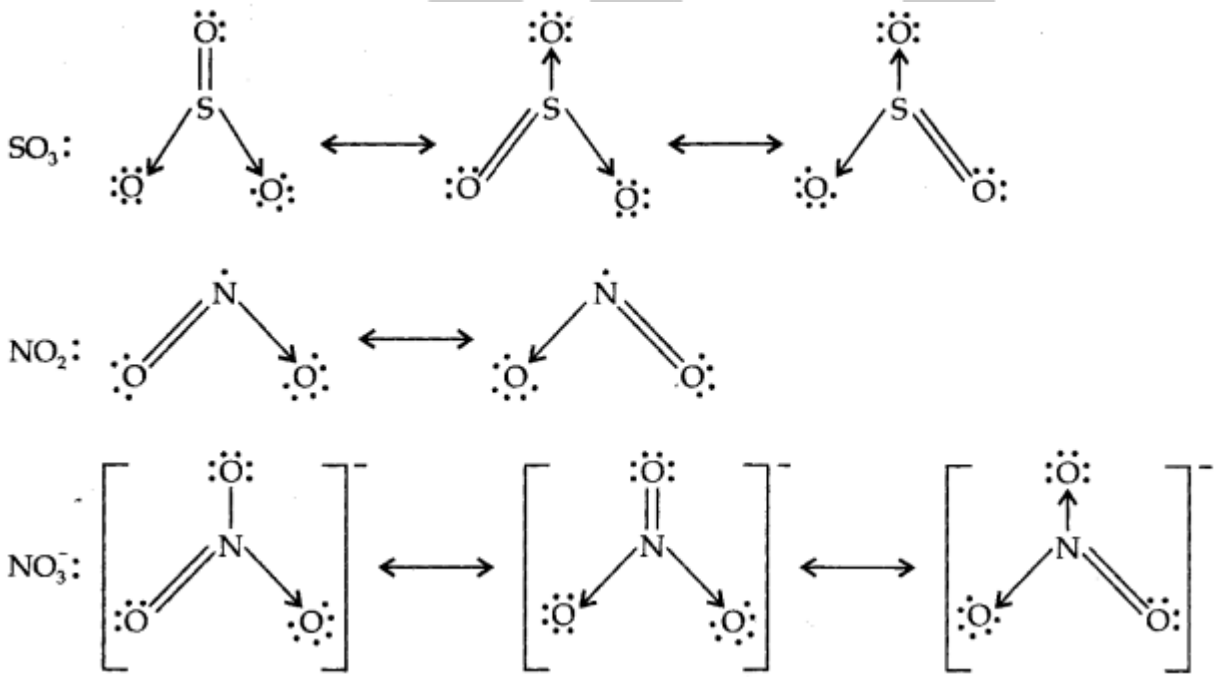
- Central atom: As
- Valence electrons of As: 5
- Bond pairs: 5
- Lone pairs: 0

- Geometry: Trigonal bipyramidal

(v) H_2S
- Central atom: S
- Valence electrons of S: 6
- Bond pairs: 2 (with 2 H atoms)
- Lone pairs: 2
- Geometry: Bent or V-shaped

(vi) PH_3
- Central atom: P
- Valence electrons of P: 5
- Bond pairs: 3
- Lone pairs: 1
- Geometry: Trigonal pyramidal

Q4. Write the resonance structures for SO_3, NO_2 and $NO3^-$
Ans :



Q5. Write the significance/applications of dipole moment.
Ans : **Significance / Applications:**

1. **Determination of Molecular Shape:**
    - Dipole moment helps in predicting the shape of molecules.
    - For example, CO_2 is linear with zero dipole moment, whereas H_2O is bent with a net dipole moment.
2. **Predicting Bond Polarity and Molecular Polarity:**
    - It indicates how electrons are shared between atoms.
    - Molecules with zero dipole moment are nonpolar; nonzero dipole moment indicates polarity.
3. **Distinguishing Between Isomers:**

- Isomers may have different dipole moments even if they have the same molecular formula.

4. **Determining Ionic Character:**
   - Higher dipole moment usually indicates greater ionic character of a bond.
5. **Studying Molecular Interactions:**
   - Dipole moments influence intermolecular forces like hydrogen bonding, dipole-dipole interactions.
6. **Used in Spectroscopy and Solvent Selection:**
   - Dipole moment data is used in IR spectroscopy and in selecting solvents for reactions.

Q6 Define electronegativity. How does it differ from electron gain enthalpy ?

Ans : Definition:

Electronegativity is the ability of an atom in a molecule to attract the shared pair of electrons towards itself in a chemical bond.

Difference between Electronegativity and Electron Gain Enthalpy:

Aspect	Electronegativity	Electron Gain Enthalpy
Definition	Ability of an atom in a molecule to attract bonded electrons	Energy change when an atom in the gaseous state gains an electron
Nature	Dimensionless quantity (relative scale)	Energy (kJ/mol) — can be positive or negative
Depends on	Bonded state of the atom (chemical environment)	Isolated gaseous atom
Measurement	No direct measurement; calculated from bond energies	Measured experimentally as energy change
Application	Predicting bond polarity and molecular structure	Predicting ease of anion

Q7.Distinguish between a sigma (σ) bond and a pi (π) bond

Property	Sigma (σ) Bond	Pi (π) Bond
Formation	By head-on (axial) overlap of atomic orbitals (s-s, s-p, or p-p)	By sidewise (lateral) overlap of parallel p orbitals
Electron density	Concentrated along the internuclear axis	Electron density above and below the internuclear axis
Bond strength	Generally stronger	Weaker than σ bonds
Rotation	Allows free rotation about the bond axis	Restricts rotation due to side overlap
Presence	Present in all single bonds	Present in double and triple bonds along with σ bond
Number per bond	One per bond	One (double bond), two (triple bond)

Q8 Explain the formation of H_2 molecule on the basis of valence bond theory

Ans:

- Each H atom has one 1s orbital with one electron.
- When two H atoms come close, their 1s orbitals overlap head-on forming a σ (sigma) bond.
- The electrons pair up with opposite spins, occupying the bonding molecular orbital.
- This results in a stable H_2 molecule with lower energy than two separate H atoms.

- The bond length is such that the energy is minimized.

Q9. Important conditions required for linear combination of atomic orbitals (LCAO) to form molecular orbitals

Ans:

1. Same or nearly same energy levels of atomic orbitals.
2. Proper symmetry: Overlapping orbitals must have the same symmetry about the internuclear axis.
3. Effective overlap: Orbitals must overlap in space sufficiently.
4. Atomic orbitals must be from atoms that are close enough for interaction.

Q10. **Compare the relative stability and magnetic properties of O_2, O_2^- (superoxide), and O_2^{2-} (peroxide)**

Ans:

Species	Bond Order	Stability	Magnetic Property
O_2	2	More stable	Paramagnetic (2 unpaired e^-)
O_2^- (Superoxide)	1.5	Less stable than O_2	Paramagnetic (1 unpaired e^-)
O_2^{2-} (Peroxide)	1	Least stable	Diamagnetic (all paired e^-)

Q11. **Define hydrogen bond. Is it weaker or stronger than van der Waals forces?**

Ans :

- **Hydrogen bond**: A special type of dipole-dipole interaction where a **hydrogen atom** is covalently bonded to a highly electronegative atom (F, O, or N) and interacts with a lone pair on another electronegative atom.
- **Stronger than van der Waals forces** but weaker than covalent/ionic bonds.
- Example: Hydrogen bonding in water (H_2O).

Q12. What is meant by bond order? Calculate the bond order of N_2, O_2, O_2^+, and O_2^-

Ans:

$$\text{Bond order} = \frac{1}{2}(N_b - N_a)$$

$$\text{E.C of } N_2 = 1s^2\, 2s^2\, 2p_x^1 2p_y^1 2p_z^1$$

(*i*) M.O. configuration of $N_2 = [\sigma 1s]^2[\sigma^* 1s]^2[\sigma 2s]^2[\sigma^* 2s]^2[\pi 2p_x]^2[\pi 2p_y]^2[\sigma 2p_z]^2$

$$\text{Bond order (B.O.)} = \frac{1}{2}(N_b - N_a)$$

$$= \frac{1}{2}[8-2] = 3$$

(*ii*) M.O. configuration of $O_2 = [\sigma 1s]^2[\sigma^* 1s]^2[\sigma 2s]^2[\sigma^* 2s]^2[\sigma 2p_z]^2$

$$\text{B.O} = \frac{1}{2}[N_b - N_a]$$

$$= \frac{1}{2}[8-4] = 2$$

(*iii*) M.O. configuration of $O_2^+ = KK[\sigma 2s]^2[\sigma^* 2s]^2[\sigma 2p_z]^2[\pi 2p_x]^2[\pi 2p_y]^2[\pi^* 2p_x]^1$

$$\text{B.O.} = \frac{1}{2}[8-3] = 2.5$$

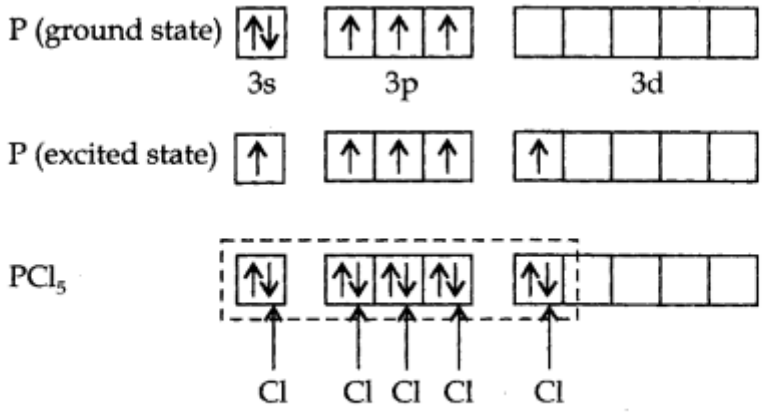
(*iv*) M.O. configuration of O_2^-

$$= KK[\sigma 2s]^2[\sigma^* 2s]^2[\sigma 2p_z]^2[\pi 2p_x]^2[\pi 2p_y]^2[\pi^* 2p_x]^2[\pi^* 2p_y]^1$$

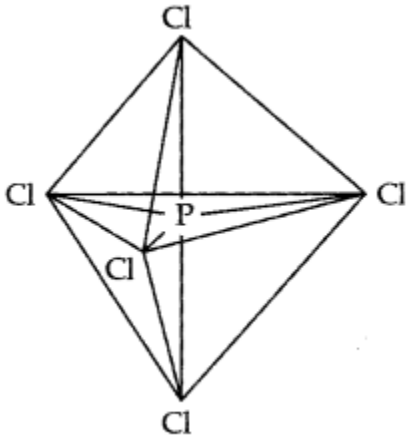
$$\text{B.O.} = \frac{1}{2}[8-5] = 1.5$$

Q13. Describe the hybridisation in case of PCl5. Why are the axial bonds longer as compared to equatorial bonds?

Ans :



sp^3 hybrid orbitals filled by electron pairs donated by five Cl atoms



Trigonal bipyramidal geometry of PCl_5 molecule

Q14. ClF_3 is T-shaped but BF_3 is planar. Explain.

Ans :

- **BF_3**: Boron has 3 valence electrons and forms 3 bonds with fluorine, no lone pairs on B → **Trigonal planar** shape (120° bond angles).
- **ClF_3**: Chlorine has 7 valence electrons; forms 3 bonds with F and has 2 lone pairs → total 5 electron pairs → **Trigonal bipyramidal electron geometry** but with 2 lone pairs occupying equatorial positions, the molecular shape becomes **T-shaped**.
-

Q15. $N(SiH_3)_3$ and $N(CH_3)_3$ are not isostructural. Give reason.

Ans :

- $N(SiH_3)_3$: Silicon has vacant 3d orbitals → can expand its octet → less steric hindrance → molecule remains pyramidal.
- $N(CH_3)_3$: Carbon does not have d orbitals → more steric hindrance from methyl groups → molecule tends to be **planar** due to lone pair repulsion.
- So, they differ in structure because of the availability of d orbitals in Si and steric factors.

Q16. Draw molecular orbital diagram for N_2^+ molecule.

Ans :

- N_2 has 10 valence electrons; N_2^+ has 9 electrons.
- Fill MOs ($\sigma 2s$, $\sigma^* 2s$, $\pi 2p$, $\sigma 2p$, $\pi^* 2p$, $\sigma^* 2p$) with 9 electrons.
- Bond order = (bonding - antibonding)/2 = (6 - 3)/2 = 1.5.
- N_2^+ is **paramagnetic** because of one unpaired electron.

Q17. HCl is a covalent compound but it ionises in solution?

- HCl is covalent in the gaseous state or pure form.

- In aqueous solution, HCl dissociates into H^+ and Cl^- due to **high polarity** of the H-Cl bond and water's polarity.
- Hence it behaves as a strong acid and ionizes in solution.

Q18. The molecule of CO_2 is linear whereas that of $SnCl_2$ is angular. Why?
Ans :

- **CO_2**: Central atom carbon has 2 double bonds, no lone pairs → **linear** (180°).
- **$SnCl_2$**: Central atom Sn has 1 lone pair and 2 bond pairs → bent/angular shape due to lone pair repulsion.

Q19 Arrange the following in order of property indicated
Ans :
(i) **O_2, O_2^+, O_2^-, O_2^{2-} (increasing stability)**

- $O_2^{2-} < O_2^- < O_2 < O_2^+$ (since bond order increases, stability increases)

(ii) **LiCl, NaCl, KCl, RbCl (increasing covalent character)**

- RbCl < KCl < NaCl < LiCl (smaller cation size → more covalent)

(iii) **NO_2, NO_2^+, NO_2^- (decreasing bond angle)**

- $NO_2^+ > NO_2 > NO_2^-$ (charge affects repulsion and bond angle)

(iv) **H–F, H–Cl, H–Br, H–I (increasing bond dissociation enthalpy)**

- H–I < H–Br < H–Cl < H–F (bond strength decreases down the group)

Q20. Arrange the following in order of property indicated
Ans :
(i) **H_2O, NH_3, H_2S, HF (increasing polar character)**

- $H_2S < NH_3 < H_2O < HF$

(ii) **HF, HCl, HBr, HI (decreasing dipole moment)**

- HF > HCl > HBr > HI

(iii) **NO_3^-, NO_2^-, NO (decreasing 's' character of hybridization)**

- $NO_3^- > NO_2^- > NO$ (more s character → more electronegative character)

(iv) **$BeCl_2$, BCl_3, CCl_4, PCl_3 (increasing bond angle)**

- $PCl_3 < CCl_4 < BCl_3 < BeCl_2$

Q21. **Give reasons for the following:**
(a) NH_3 has higher boiling point than PH_3.

- NH_3 exhibits **hydrogen bonding** (N-H bonds), which is strong intermolecular force, increasing boiling point.
- PH_3 does not show hydrogen bonding (P-H bonds are less polar), so weaker Van der Waals forces dominate → lower boiling point.

(b) Ionic compounds do not conduct electricity in solid state.

- In solid ionic compounds, ions are fixed in the lattice and **cannot move freely**.
- Electrical conductivity requires **mobile charged particles** (ions or electrons), which are absent in solid state.

(c) LiCl is more covalent than KCl.

- Li^+ ion is **smaller and highly polarizing** → distorts electron cloud of Cl^- → more covalent character (Fajan's rules).
- K^+ is larger, less polarizing → more ionic character.

(d) NH_3 is more polar than NF_3.

- NH_3: N-H bond dipoles point **towards N** (more electronegative) and lone pair adds to net dipole → molecule is polar.
- NF_3: N-F bond dipoles point **away from N**, opposing lone pair dipole → overall molecular dipole is less → less polar than NH_3.

(e) H_2O has bent structure.

- Oxygen has 2 lone pairs and 2 bonded pairs → according to VSEPR theory, lone pairs repel more → bent shape with bond angle ~104.5°.

(a) Define bond dissociation enthalpy. How is it related to bond order?

Ans :

- **Bond dissociation enthalpy (BDE):** Energy required to break one mole of a particular bond in the gaseous state to give separated atoms.
- **Relation with bond order:** Higher bond order → stronger bond → higher BDE. Lower bond order → weaker bond → lower BDE.

(b) Explain why N_2 has greater bond dissociation enthalpy than N_2^+, while O_2 has lesser bond dissociation enthalpy than O_2^+?

Ans :

- **N_2 vs N_2^+:** Removing one electron from N_2^+ reduces antibonding electrons → bond order decreases from 3 (N_2) to 2.5 (N_2^+) → N_2 has higher BDE.
- **O_2 vs O_2^+:** Removing one electron from O_2^+ removes electron from antibonding orbital → bond order increases from 2 (O_2) to 2.5 (O_2^+) → O_2^+ has higher BDE.

Q22. Applications of dipole moment

Ans :

1. Predicting molecular polarity.
2. Determining molecular geometry.
3. Differentiating isomers.
4. Understanding solubility and intermolecular forces.

Q23. **Assign reasons:**

(i) NH_3 is freely soluble in water, while PH_3 is not.

- NH_3 forms **strong hydrogen bonds** with water → highly soluble.
- PH_3 does not form H-bonds → **low solubility**.

(ii) B_2 is paramagnetic while C_2 is not.

- **B_2** has **two unpaired electrons** in π orbitals → **paramagnetic**.
- **C_2** has all electrons paired → **diamagnetic**.

Q24. **o-Nitrophenol is steam volatile while p-Nitrophenol is not. Explain.**

- **o-Nitrophenol** forms **intramolecular H-bonds** → less intermolecular attraction → low boiling point → steam volatile.
- **p-Nitrophenol** forms **intermolecular H-bonds** → high boiling point → not steam volatile.

Q25. **Define bond enthalpy. Why is F_2's bond enthalpy less than Cl_2's?**

- **Bond enthalpy**: Energy required to break one mole of a bond in gaseous molecules.
- F_2 has **low bond enthalpy** because:
  - Small atomic size → **electron repulsion** between lone pairs.
  - **Weaker bond** than Cl_2.

BY: NK MISHRA