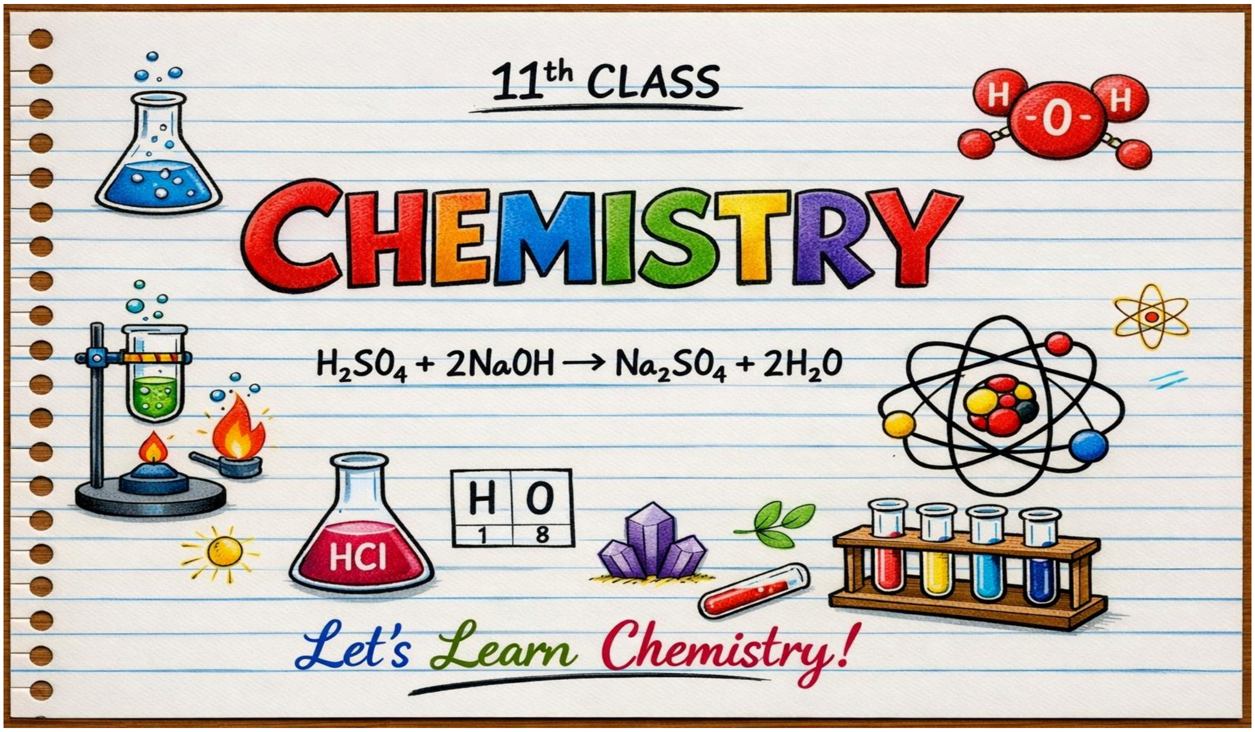


# Class: 11th

# Subject: Chemistry

# Unit 3: CHEMICAL BONDING

## ❖ Important MCQs:

**1. Chemical bond is defined as:**

(a) Force that repels atoms

(b) Force that holds atoms, molecules or ions together ✅

(c) Force that destroys molecules

(d) Force between electrons only

**2. Properties of a substance mainly depend on:**

(a) Temperature only

(b) Type of chemical bond present in the substance ✅

(c) Mass of atoms only

(d) Number of neutrons

**3. Van der Waals forces are considered:**

(a) Strong chemical bonds

(b) Ionic bonds

(c) Weak intermolecular forces ✅

(d) Metallic bonds

**4. According to Lewis concept, atoms form bonds to:**

(a) Increase instability

(b) Achieve noble gas configuration ✅

(c) Lose electrons permanently

(d) Become ions only

**5. Ionic bond is formed by:**

(a) Sharing of electrons

(b) Complete transfer of electrons from one atom to another ✅

(c) Sharing of protons

(d) Sharing of neutrons

**6. Ionic bond generally forms between atoms having:**

(a) Similar electronegativity

(b) Same ionization energy

(c) Large difference in electronegativity ✅

(d) Same atomic size

**7. Sodium (Na) forms $Na^+$ ion because it:**

(a) Gains an electron

(b) Loses one electron to achieve noble gas configuration ✅

(c) Shares electrons

(d) Gains protons

**8. Chlorine forms $Cl^-$ ion because it:**

(a) Loses an electron

(b) Gains one electron to complete its octet ✅

(c) Shares neutrons

(d) Becomes neutral

**9. In NaCl, the bond between $Na^+$ and $Cl^-$ is due to:**

(a) Electron sharing

(b) Magnetic attraction

(c) Electrostatic attraction between oppositely charged ions ✅

(d) Nuclear force

**10. According to Lewis theory, atoms complete their outer shell by:**

(a) Gaining protons

(b) Losing neutrons

(c) Gaining, losing, or sharing electrons to achieve octet ✅

(d) Increasing atomic number

**11. A dative bond is formed when:**

(a) Both atoms contribute one electron each

(b) One atom donates a lone pair of electrons to another atom ✅

(c) Electrons are completely transferred

(d) No electrons are involved

**12. In a coordinate covalent bond, the shared electron pair is:**

(a) Contributed by both atoms equally

(b) Donated by one atom only ✅

(c) Lost during bonding

(d) Not shared at all

**13. In $H_2O$ molecule, oxygen has:**

(a) No lone pairs

(b) One lone pair

(c) Two lone pairs ✅

(d) Three lone pairs

**14. In the formation of $H_3O^+$ ion, the coordinate bond is formed between:**

(a) H and H

(b) O and $H^+$ ✅

(c) O and O

(d) H and $O^-$

**15. The proton ($H^+$) acts as:**

(a) Electron donor

(b) Electron acceptor ✅

(c) Neutral atom

(d) Lone pair donor

**16. In carbon monoxide (CO), the coordinate bond is formed when:**

(a) Carbon donates electrons to oxygen

(b) Oxygen donates a lone pair to carbon ✅

(c) Both atoms lose electrons

(d) No electrons are shared

**17. Carbon in CO completes its octet by:**

(a) Losing electrons

(b) Gaining electrons

(c) Accepting a lone pair from oxygen ✅

(d) Forming ionic bond only

**18. Ozone ($O_3$) contains:**

(a) Only single bonds

(b) One double bond and one coordinate covalent bond ✅

(c) Only ionic bonds

(d) Only metallic bonds

**19. In ozone ($O_3$), the central oxygen atom carries:**

(a) No charge

(b) Positive charge ✅

(c) Negative charge

(d) Double positive charge

**20. In $O_3$, the oxygen atom accepting lone pair carries:**

(a) Positive charge

(b) Negative charge ✅

(c) Neutral charge

(d) No charge

**21. Polyatomic ions are:**

(a) Single atoms with charge

(b) Ions composed of more than one atom with a net charge ✅

(c) Neutral molecules

(d) Only metals

**22. Which of the following is a positively charged polyatomic ion?**

(a) $CO_3^{2-}$

(b) $SO_4^{2-}$

(c) $NH_4^+$ ✅

(d) $NO_3^-$

**23. Most polyatomic ions carry:**

(a) Positive charge

(b) Negative charge ✅

(c) No charge

(d) Variable charge only

**24. Expanded octet means:**

(a) Less than 8 electrons

(b) Exactly 8 electrons

(c) More than 8 electrons around central atom ✅

(d) No electrons

**25. Expanded octet occurs due to involvement of:**

(a) s orbitals only

(b) p orbitals only

(c) d orbitals in bonding ✅

(d) f orbitals only

**26. In $SO_4^{2-}$ ion, sulfur has:**

(a) 8 electrons

(b) 10 electrons

(c) 12 electrons in valence shell ✅

(d) 6 electrons

**27. Sulfur exceeds octet in $SO_4^{2-}$ because:**

(a) It cannot bond

(b) It uses d orbitals for additional electrons ✅

(c) It loses electrons

(d) It becomes neutral

**28. Expanded octet explains:**

(a) Fixed oxidation state

(b) Variable oxidation states of sulfur ✅

(c) No bonding ability

(d) Only ionic bonding

**29. In $I_3^-$ ion, the central iodine atom has:**

(a) 8 electrons

(b) 10 electrons in valence shell ✅

(c) 6 electrons

(d) 12 electrons

**30. The central iodine atom in $I_3^-$ shows:**

(a) Octet deficiency

(b) Expanded octet ✅

(c) No bonding

(d) Only covalent bonding without lone pairs

**31. Intermolecular forces are:**

(a) Stronger than chemical bonds

(b) Weaker than chemical bonds ✅

(c) Equal to ionic bonds

(d) Only present in solids

**32. Intermolecular forces exist between:**

(a) Atoms only

(b) Molecules only

(c) All kinds of atoms and molecules when close together ✅

(d) Ions only

**33. Intermolecular forces are also called:**

(a) Ionic forces

(b) Covalent forces

(c) van der Waals forces ✅

(d) Nuclear forces

**34. Intermolecular forces mainly affect:**

(a) Chemical properties

(b) Nuclear properties

(c) Physical properties of substances ✅

(d) Atomic number

**35. Which of the following is NOT a type of intermolecular force?**

(a) London dispersion forces

(b) Dipole-dipole forces

(c) Hydrogen bonding

(d) Ionic bonding ✅

**36. London dispersion forces are also known as:**

(a) Permanent dipole forces

(b) Instantaneous dipole-induced dipole forces ✅

(c) Ionic forces

(d) Metallic forces

**37. Hydrogen bonding is:**

(a) Strong chemical bond

(b) Weak intermolecular force ✅

(c) Ionic bond

(d) Metallic bond

**38. Bond energy is defined as:**

(a) Energy required to form bonds

(b) Energy required to break bonds in one mole of substance ✅

(c) Energy of electrons only

(d) Nuclear energy

**39. Unit of bond energy is:**

(a) Joule

(b) kJ $mol^{-1}$ ✅

(c) Watt

(d) Pascal

**40. Bond energy is a measure of:**

(a) Size of atom

(b) Strength of bond ✅

(c) Mass of atom

(d) Volume of molecule

**41. Bond strength increases with:**

(a) Decrease in electronegativity difference

(b) Increase in electronegativity difference ✅

(c) Increase in atomic size only

(d) Decrease in electrons

**42. Among H–X bonds, the highest bond energy is for:**

(a) HI

(b) HBr

(c) HCl

(d) HF ✅

**43. Bond energy decreases from HF to HI because:**

(a) Atomic size decreases

(b) Electronegativity difference decreases ✅

(c) Number of electrons increases

(d) Pressure increases

**44. Multiple bonds are:**

(a) Weaker than single bonds

(b) Equal to single bonds

(c) Stronger than single bonds ✅

(d) Non-existent

**45. Bond length is defined as:**

(a) Distance between electrons

(b) Distance between nuclei of bonded atoms ✅

(c) Distance between protons

(d) Distance between ions

**46. Bond length increases with:**

(a) Decrease in atomic size

(b) Increase in atomic size ✅

(c) Increase in electronegativity

(d) Increase in bond strength

**47. Bond length decreases when:**

(a) Atomic size increases

(b) Electronegativity difference increases ✅

(c) Electrons decrease

(d) Temperature increases

**48. Ionic character in a covalent bond leads to:**

(a) Increase in bond length

(b) Decrease in bond length ✅

(c) No change

(d) Breaking of bond

**49. The strongest type of chemical bond is:**

(a) Metallic bond

(b) Covalent bond

(c) Ionic bond ✅

(d) Hydrogen bond

**50. Which of the following has the lowest bond energy?**

(a) Ionic bond

(b) Covalent bond

(c) Hydrogen bond

(d) London dispersion forces ✅

**51. According to Valence Bond Theory (VBT), a covalent bond is formed when:**

(a) Electrons are completely transferred

(b) Half-filled orbitals of similar energy overlap ✅

(c) Protons are shared

(d) Neutrons overlap

**52. According to VBT, the strength of a covalent bond depends on:**

(a) Number of neutrons

(b) Extent of overlap of atomic orbitals ✅

(c) Atomic number only

(d) Shape of nucleus

**53. Sigma (σ) bond is formed by:**

(a) Sidewise overlap of orbitals

(b) Linear (head-on) overlap along the nuclear axis ✅

(c) Overlap of filled orbitals only

(d) Transfer of electrons

**54. Which type of overlap occurs in HCl molecule according to VBT?**

(a) s–s overlap

(b) p–p overlap

(c) s–p overlap ✅

(d) d–d overlap

**55. In $O_2$ molecule, the double bond consists of:**

(a) Two σ bonds

(b) One σ and one π bond ✅

(c) Two π bonds

(d) Only one σ bond

**56. Hybridization is the process in which:**

(a) Electrons are transferred

(b) Atomic orbitals mix to form new equivalent orbitals ✅

(c) Nuclei combine

(d) Protons are shared

**57. In hybridization, the resulting orbitals have:**

(a) Different energies and shapes

(b) Same energy and same shape ✅

(c) Same shape but different energies

(d) No definite shape

**58. In sp hybridization, the number of hybrid orbitals formed is:**

(a) One

(b) Two ✅

(c) Three

(d) Four

**59. The geometry of sp hybridized orbitals is:**

(a) Trigonal planar

(b) Tetrahedral

(c) Linear (180°) ✅

(d) Bent

**60. Which molecule shows sp hybridization?**

(a) $CH_4$

(b) $BF_3$

(c) $BeCl_2$ ✅

(d) $NH_3$

**61. In $sp^2$ hybridization, the number of hybrid orbitals formed is:**

(a) Two

(b) Three ✅

(c) Four

(d) One

**62. The geometry of $sp^2$ hybridized molecules is:**

(a) Linear

(b) Tetrahedral

(c) Trigonal planar (120°) ✅

(d) Bent

**63. $BF_3$ molecule has:**

(a) sp hybridization

(b) $sp^2$ hybridization with trigonal planar shape ✅

(c) $sp^3$ hybridization

(d) No hybridization

**64. In $sp^3$ hybridization, the number of hybrid orbitals formed is:**

(a) Two

(b) Three

(c) Four ✅

(d) Five

**65. $CH_4$ molecule has:**

(a) Linear geometry

(b) Trigonal planar geometry

(c) Tetrahedral geometry due to $sp^3$ hybridization ✅

(d) Bent structure

**66. VSEPR model explains the:**

(a) Bond energies

(b) Shapes of molecules based on electron pair repulsions ✅

(c) Atomic mass

(d) Nuclear forces

**67. According to VSEPR theory, molecular shape depends on:**

(a) Number of protons

(b) Arrangement of electron pairs around central atom ✅

(c) Atomic number

(d) Neutron density

**68. Electron pairs arrange themselves to:**

(a) Minimize distance

(b) Maximize repulsions

(c) Remain at maximum distance to minimize repulsions ✅

(d) Form ionic bonds

**69. Which of the following participates in determining molecular geometry?**

(a) Bond pairs only

(b) Lone pairs only

(c) Both lone pairs and bond pairs ✅

(d) Neutrons

**70. The order of repulsion between electron pairs is:**

(a) bp-bp > lp-bp > lp-lp

(b) lp-lp > lp-bp > bp-bp ✅

(c) lp-bp > lp-lp > bp-bp

(d) bp-bp > lp-lp > lp-bp

**71. Lone pair electrons exert:**

(a) Less repulsion than bond pairs

(b) Equal repulsion as bond pairs

(c) Greater repulsion than bond pairs ✅

(d) No repulsion

**72. A double bond in VSEPR theory is treated as:**

(a) Two separate pairs

(b) One electron pair region ✅

(c) No electron pair

(d) Three electron pairs

**73. The central atom in a molecule is usually:**

(a) Most electronegative atom

(b) Least electronegative atom ✅

(c) Largest atom only

(d) Always hydrogen

**74. The geometry of $AB_2$ type molecules is:**

(a) Bent

(b) Linear (180°) ✅

(c) Tetrahedral

(d) Trigonal planar

**75. The geometry of $AB_3$ type molecules is:**

(a) Linear

(b) Bent

(c) Trigonal planar (120°) ✅

(d) Tetrahedral

**76. In $SnCl_2$, the shape is:**

(a) Linear

(b) Trigonal planar

(c) V-shaped (bent) due to lone pair ✅

(d) Tetrahedral

**77. The shape of $SO_3$ molecule is:**

(a) Bent

(b) Trigonal planar (120°) ✅

(c) Linear

(d) Tetrahedral

**78. In $SO_2$ molecule, the shape is:**

(a) Linear

(b) Trigonal planar

(c) V-shaped due to lone pair ✅

(d) Tetrahedral

**79. The geometry of $AB_4$ type molecules is:**

(a) Linear

(b) Trigonal planar

(c) Tetrahedral (109.5°) ✅

(d) Bent

**80. In $NH_3$ molecule, the shape is:**

(a) Tetrahedral

(b) Trigonal planar

(c) Trigonal pyramidal due to one lone pair ✅

(d) Linear

**81. The shape of $H_2O$ molecule is:**

(a) Linear

(b) Trigonal planar

(c) V-shaped (angular) due to lone pairs ✅

(d) Tetrahedral

**82. The bond angle in $H_2O$ molecule is approximately:**

(a) 109.5°

(b) 120°

(c) 104.5° ✅

(d) 90°

**83. The bond angle in $H_2O$ is reduced due to:**

(a) Bond pair–bond pair repulsion

(b) Lone pair–lone pair repulsion being greater ✅

(c) Absence of electrons

(d) Equal repulsion

**84. The shape of $H_3O^+$ ion is:**

(a) Linear

(b) Trigonal planar

(c) Trigonal pyramidal ✅

(d) Tetrahedral

**85. The shape of $NH_2^-$ (amide ion) is:**

(a) Linear

(b) Trigonal planar

(c) V-shaped (bent) ✅

(d) Tetrahedral

**86. The geometry of $AB_5$ type molecules is:**

(a) Tetrahedral

(b) Trigonal bipyramidal ✅

(c) Octahedral

(d) Linear

**87. In $PCl_5$ molecule, the bond angles are:**

(a) 109.5° only

(b) 120° and 90° ✅

(c) 180° only

(d) 104.5°

**88. In $I_3^-$ ion, the actual shape is:**

(a) Bent

(b) Trigonal planar

(c) Linear (180°) ✅

(d) Tetrahedral

**89. The geometry of $AB_6$ type molecules is:**

(a) Trigonal planar

(b) Tetrahedral

(c) Octahedral (90° angles) ✅

(d) Linear

**90. $XeF_4$ molecule has:**

(a) Tetrahedral shape

(b) Trigonal pyramidal shape

(c) Square planar shape due to two lone pairs ✅

(d) Linear shape

**91. Molecular Orbital Theory (MOT) is based on:**

(a) Classical bonding concept

(b) Quantum mechanical treatment of atoms and molecules ✅

(c) Ionic interaction only

(d) Nuclear forces

**92. In MOT, atomic orbitals combine to form:**

(a) Hybrid orbitals

(b) Molecular orbitals extending over the entire molecule ✅

(c) Lone pairs

(d) Protons

**93. The bonding molecular orbital is formed by:**

(a) Out-of-phase combination

(b) In-phase combination of atomic orbitals ✅

(c) No overlap

(d) Electron loss

**94. Antibonding molecular orbitals are formed by:**

(a) In-phase overlap

(b) Out-of-phase overlap leading to higher energy orbitals ✅

(c) Electron transfer

(d) Ionic attraction

**95. Bond order is calculated as:**

(a) Bonding electrons + antibonding electrons

(b) (Bonding electrons − antibonding electrons) / 2 ✅

(c) Total electrons / 2

(d) Difference × 2

**96. A bond order of zero indicates:**

(a) Strong bond

(b) Weak bond

(c) No bond formation ✅

(d) Ionic bond

**97. In $H_2$ molecule, the molecular orbital formed is:**

(a) π bond

(b) σ bond from s–s overlap ✅

(c) π* orbital

(d) Ionic bond

**98. Head-on overlap of p orbitals results in:**

(a) π bonding orbital

(b) σ bonding and σ* antibonding orbitals ✅

(c) Lone pairs

(d) Hybrid orbitals

**99. Sideways overlap of p orbitals produces:**

(a) σ bonds only

(b) π bonding and π* antibonding orbitals ✅

(c) Ionic bonds

(d) No bonding

**100. Oxygen molecule ($O_2$) shows paramagnetism due to:**

(a) All electrons paired

(b) Presence of unpaired electrons in antibonding π* orbitals ✅

(c) Ionic nature

(d) Absence of bonding

## ❖ Important Short Questions

**1. What is a chemical bond? Define it.**

**Ans:** A chemical bond is the force that holds two or more atoms, molecules, or ions together to form a stable substance. It determines the physical and chemical properties of substances.

**Example:** NaCl (sodium chloride)

**2. How is an ionic bond formed? Explain briefly.**

**Ans:** An ionic bond is formed by the complete transfer of electrons from an atom with low ionization energy (metal) to an atom with high electron affinity (non-metal), resulting in oppositely charged ions held by electrostatic attraction.

**Example:** NaCl ($Na^+$ and $Cl^-$)

**3. Explain covalent bonding in terms of the Lewis concept.**

**Ans:** According to Lewis concept, covalent bonds are formed when atoms share one or more pairs of electrons to complete their outermost shell and achieve a noble gas configuration (octet rule).

**Example:** $H_2$, $Cl_2$

**4. What is a dative (coordinate covalent) bond? Give an example.**

**Ans:** A dative bond is a covalent bond in which both electrons of the shared pair are donated by one atom to an electron-deficient atom.

**Example:** $H_3O^+$ (formed from $H_2O + H^+$)

**5. How is a coordinate bond formed in CO molecule?**

**Ans:** In CO, oxygen donates a lone pair of electrons to carbon to complete carbon's octet, forming a coordinate covalent bond along with a normal covalent bond.

**Example:** CO molecule

**6. Explain the formation of coordinate bond in ozone ($O_3$).**

**Ans:** In ozone, one oxygen atom donates a lone pair to another oxygen atom, forming a coordinate bond. The central oxygen becomes positively charged and the acceptor oxygen becomes negatively charged.

**Example:** $O_3$ molecule

**7. What is an expanded octet? Give an example.**

**Ans:** An expanded octet occurs when the central atom has more than eight electrons in its valence shell due to involvement of d-orbitals.

**Example:** $SO_4^{2-}$ (sulfur has 12 electrons)

**8. What are polyatomic ions? Give two examples.**

**Ans:** Polyatomic ions are ions composed of more than one atom bonded together and carrying an overall charge.

**Examples:** $NH_4^+$, $SO_4^{2-}$

**9. How does sulfur expand its octet in $SO_4^{2-}$ ion?**

**Ans:** Sulfur expands its octet by utilizing its vacant d-orbitals and forming multiple bonds with oxygen atoms, allowing it to accommodate more than eight electrons.

**Example:** $SO_4^{2-}$ ion (12 valence electrons around S)

**10. Describe the structure of the tri-iodide ion ($I_3^-$).**

**Ans:** In $I_3^-$, the central iodine atom has five electron pairs (2 bond pairs and 3 lone pairs). Due to VSEPR theory, the lone pairs occupy equatorial positions, giving a linear shape with a bond angle of 180°.

**Example:** $I_3^-$ ion (linear structure)

**11. What are intermolecular forces? Define them.**

**Ans:** Intermolecular forces are the forces of attraction between molecules that hold molecules together in liquids and solids. They are weaker than chemical bonds.

**Example:** Forces between water molecules in $H_2O$.

**2. Define van der Waals forces.**

**Ans:** Van der Waals forces are weak intermolecular forces that include dipole–dipole interactions and London dispersion forces acting between molecules.

**Example:** Attraction between noble gas atoms like Ar.

**13. What are London dispersion forces?**

**Ans:** London dispersion forces are weak forces caused by temporary dipoles formed due to random movement of electrons in atoms or nonpolar molecules.

**Example:** Forces between $I_2$ molecules.

**14. What are dipole–dipole forces?**

**Ans:** Dipole–dipole forces are attractive forces between polar molecules where the positive end of one molecule is attracted to the negative end of another.

**Example:** HCl–HCl interactions.

**15. What is hydrogen bonding? Give an example.**

**Ans:** Hydrogen bonding is a strong dipole–dipole interaction that occurs when hydrogen is bonded to highly electronegative atoms like N, O, or F.

**Example:** Hydrogen bonding in $H_2O$ molecules.

**16. What is bond energy and what is its unit?**

**Ans:** Bond energy is the amount of energy required to break one mole of a particular bond in the gaseous state.

Unit: kJ/mol

**Example:** H–H bond energy in $H_2$.

**17. What is meant by bond length?**

**Ans:** Bond length is the average distance between the nuclei of two bonded atoms in a molecule.

**Example:** H–H bond length in $H_2$ molecule.

**18. How does electronegativity difference affect bond energy?**

**Ans:** Greater electronegativity difference increases bond polarity, resulting in stronger attraction between atoms and thus higher bond energy.

**Example:** H–F bond has higher bond energy than H–H bond.

**19. How does atomic size affect bond length?**

**Ans:** Larger atomic size leads to longer bond length because the nuclei are farther apart. Smaller atoms form shorter bonds.

**Example:** C–C < C–Cl bond lengths.

**20. Differentiate between chemical bonds and intermolecular forces.**

**Ans: Chemical bonds** are strong forces that hold atoms together within a molecule (intramolecular), while **intermolecular forces** are weaker forces between molecules.

**Example:** Covalent bond in $H_2$ vs hydrogen bonding between $H_2O$ molecules.

**21. What does Valence Bond Theory (VBT) explain?**

**Ans:** Valence Bond Theory explains the formation of covalent bonds by the overlap of atomic orbitals and the pairing of electrons.

**Example:** Formation of $H_2$ molecule through 1s–1s overlap.

**22. How is a sigma (σ) bond formed?**

**Ans:** A sigma bond is formed by head-on (axial) overlap of atomic orbitals along the internuclear axis.

**Example:** Overlap of two s orbitals in $H_2$.

**23. How is a pi (π) bond formed?**

**Ans:** A pi bond is formed by sideways (lateral) overlap of parallel p orbitals above and below the internuclear axis.

**Example:** π bond in ethene ($C_2H_4$).

**24. What is hybridization? Define it.**

**Ans:** Hybridization is the mixing of atomic orbitals of similar energies to form new equivalent hybrid orbitals of the same energy and shape.

**Example:** $sp^3$ hybridization in $CH_4$.

**25. What is the basic principle of the VSEPR model?**

**Ans:** The VSEPR model states that electron pairs around a central atom repel each other and arrange themselves as far apart as possible to minimize repulsion, determining molecular shape.

**Example:** Bent shape of $H_2O$ due to lone pair repulsion.

## ❖ Important Long Questions:

### 🌟 Q1. Explain Atomic Orbital Hybridization in detail. Describe sp, $sp^2$, and $sp^3$ hybridization with suitable examples and geometries.

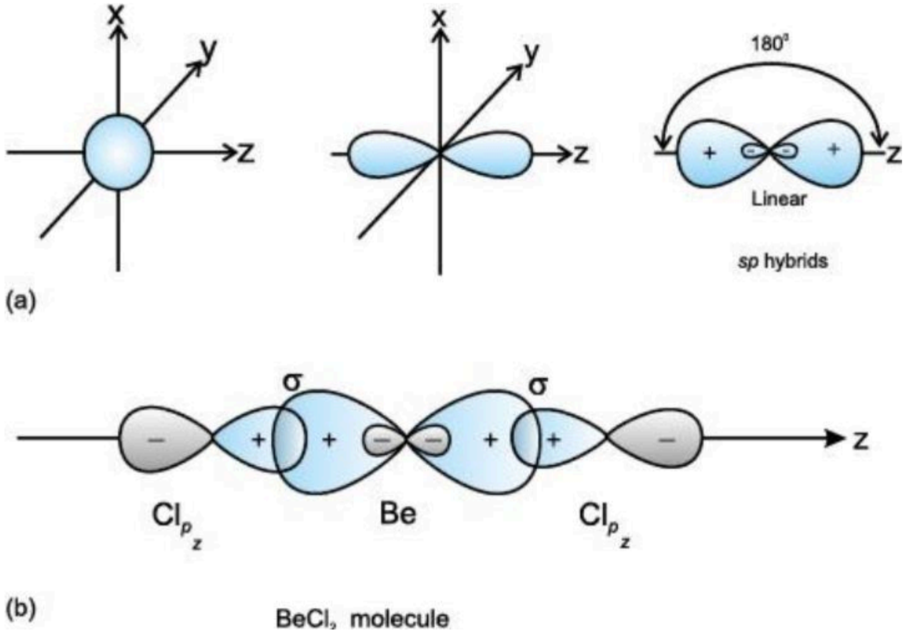
### ❖ Definition of Hybridization

Hybridization is the process in which atomic orbitals of nearly equal energy (such as s and p orbitals) mix together to form a new set of equivalent orbitals called hybrid orbitals. These hybrid orbitals have the same energy, same shape, and are oriented in specific directions to minimize repulsion.

Hybridization explains the geometry and shape of molecules.

### ✦ Types of Hybridization

### 1. sp Hybridization



**Explanation:**

- One s orbital and one p orbital combine to form two sp hybrid orbitals.
- These orbitals are arranged in a linear geometry.

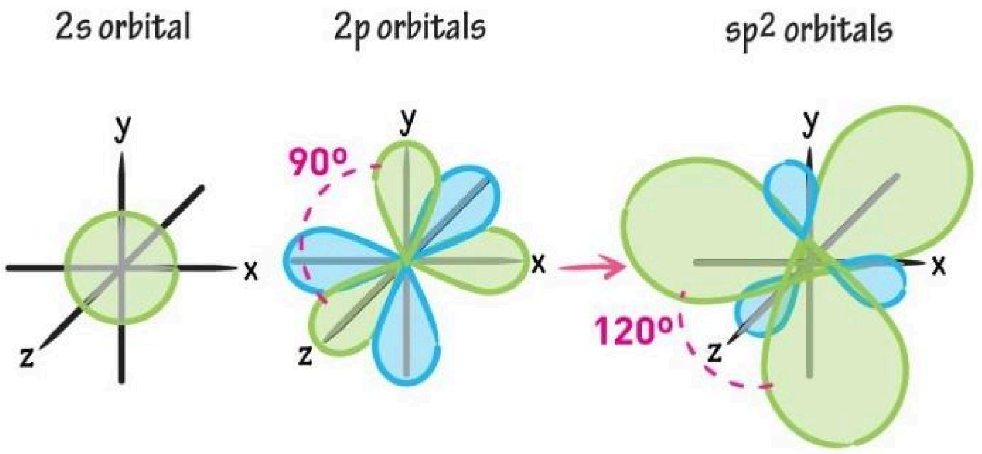
**Key Features:**

- Number of hybrid orbitals: 2
- **Geometry**: Linear
- **Bond angle**: 180°

**Example:** $BeCl_2$

- Beryllium uses sp hybridization.
- Two sp orbitals overlap with chlorine atoms forming two sigma bonds.

## 2. $sp^2$ Hybridization



**Explanation:**

- One s orbital and two p orbitals mix to form three $sp^2$ hybrid orbitals.
- These are arranged in a trigonal planar geometry.

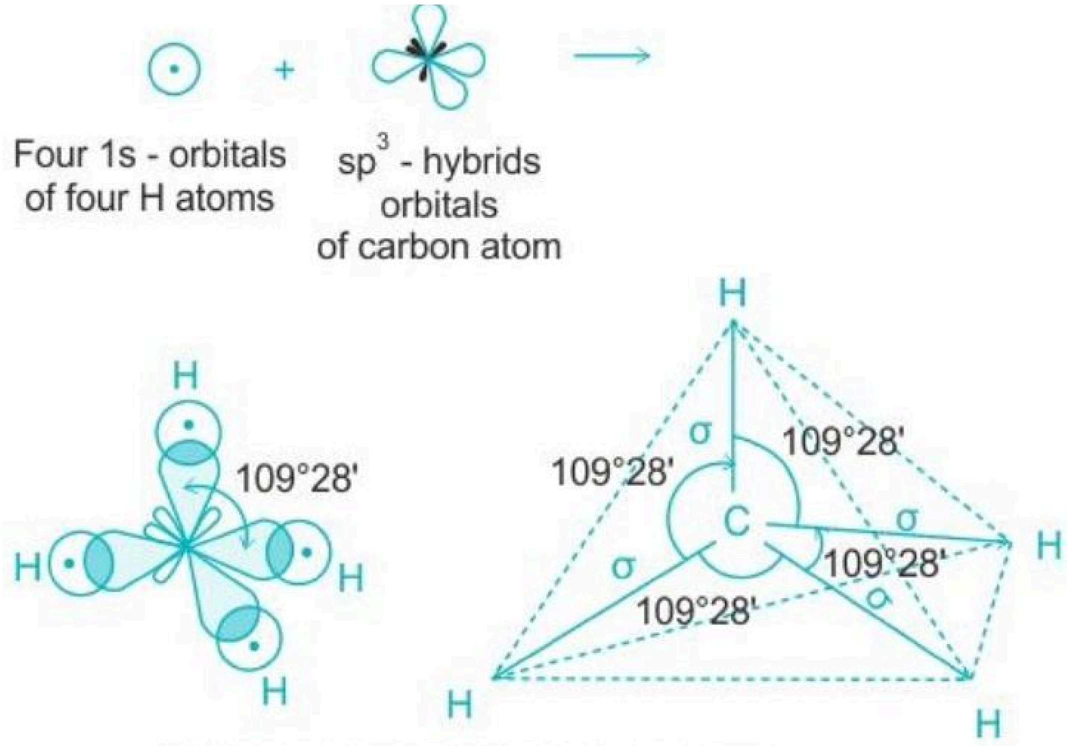
**Key Features:**

- Number of hybrid orbitals: 3
- **Geometry:** Trigonal planar
- **Bond angle:** 120°

**Example: $BF_3$**

- Boron undergoes $sp^2$ hybridization.
- Three $sp^2$ orbitals form sigma bonds with fluorine atoms.

## 3. $sp^3$ Hybridization



Tetrahedral geometry of $CH_4$ molecule

**Explanation:**

- One s orbital and three p orbitals combine to form four $sp^3$ hybrid orbitals.
- These are arranged in a tetrahedral geometry.

**Key Features:**

- Number of hybrid orbitals: 4
- **Geometry:** Tetrahedral

- **Bond angle:** 109.5°

**Example: $CH_4$ (Methane)**

- Carbon undergoes $sp^3$ hybridization.
- Each hybrid orbital overlaps with hydrogen 1s orbitals forming four sigma bonds.

### Comparison Table:

| Hybridization | Orbitals Mixed | Geometry | Bond Angle | Example |
|---|---|---|---|---|
| sp | 1s + 1p | Linear | 180° | $BeCl_2$ |
| $sp^2$ | 1s + 2p | Trigonal planar | 120° | $BF_3$ |
| $sp^3$ | 1s + 3p | Tetrahedral | 109.5° | $CH_4$ |

### ✦ Summary:

Hybridization is the mixing of atomic orbitals to form new hybrid orbitals that determine molecular shape and bond angles.

- sp hybridization → linear (180°) → $BeCl_2$
- $sp^2$ hybridization → trigonal planar (120°) → $BF_3$
- $sp^3$ hybridization → tetrahedral (109.5°) → $CH_4$

Hybridization helps explain the 3D structure of molecules, bond angles, and bonding behavior in a clear way.

### 🌟 Q2. Explain VSEPR theory in detail and describe the shapes of the following molecules with reasons: $H_2O$, $NH_3$, $PCl_5$, $SF_6$, and $I_3^-$.

## ❖ Introduction / Definition of VSEPR Theory

VSEPR (Valence Shell Electron Pair Repulsion) theory states that electron pairs around a central atom repel each other and arrange themselves as far apart as possible to minimize repulsion. This arrangement determines the shape of the molecule.

**Electron pairs include:**

- Bond pairs (shared electrons between atoms)
- Lone pairs (non-bonding electrons on central atom)

### Key Principles of VSEPR Theory

- Electron pairs repel each other due to like charges.
- **Repulsion order:**
- Lone pair–lone pair > lone pair–bond pair > bond pair–bond pair
- Lone pairs occupy more space than bond pairs.
- Molecular shape depends on the number of electron pairs around the central atom.
- Lone pairs compress bond angles.
- Shapes of Given Molecules

### 1. $H_2O$ (Water)

- Central atom: Oxygen
- Valence electrons = 6
- Forms 2 bonds with H → 2 bond pairs
- Remaining electrons = 2 lone pairs

- Total electron pairs = 4

**Electron geometry:** Tetrahedral

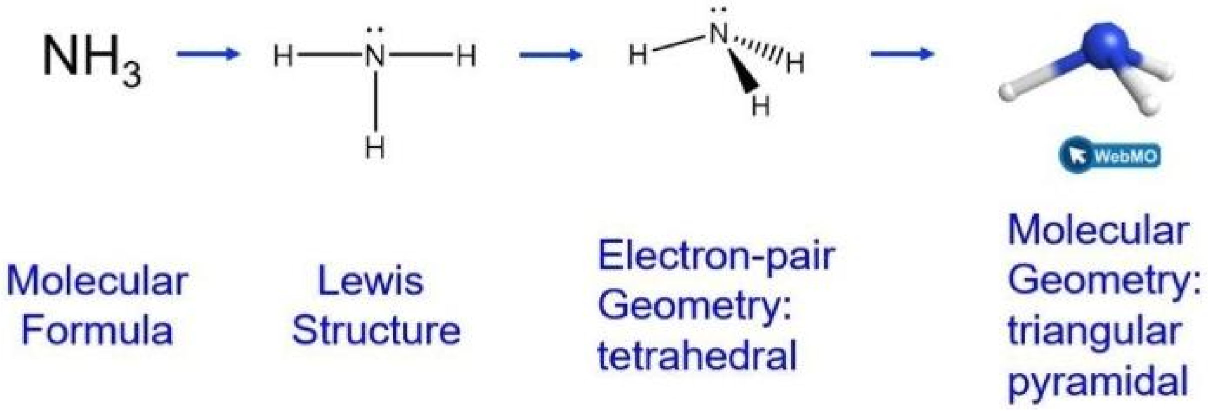
Molecular shape: Bent (V-shaped)

Bond angle: 104.5°

**Explanation:**

Although electron geometry is tetrahedral, the presence of two lone pairs causes strong repulsion, which pushes the bond pairs closer together and reduces the bond angle from 109.5° to 104.5°.

## 2. $NH_3$ (Ammonia)



- Central atom: Nitrogen
- Valence electrons = 5
- Forms 3 bonds with H → 3 bond pairs
- Remaining = 1 lone pair
- Total electron pairs = 4

**Electron geometry:** Tetrahedral

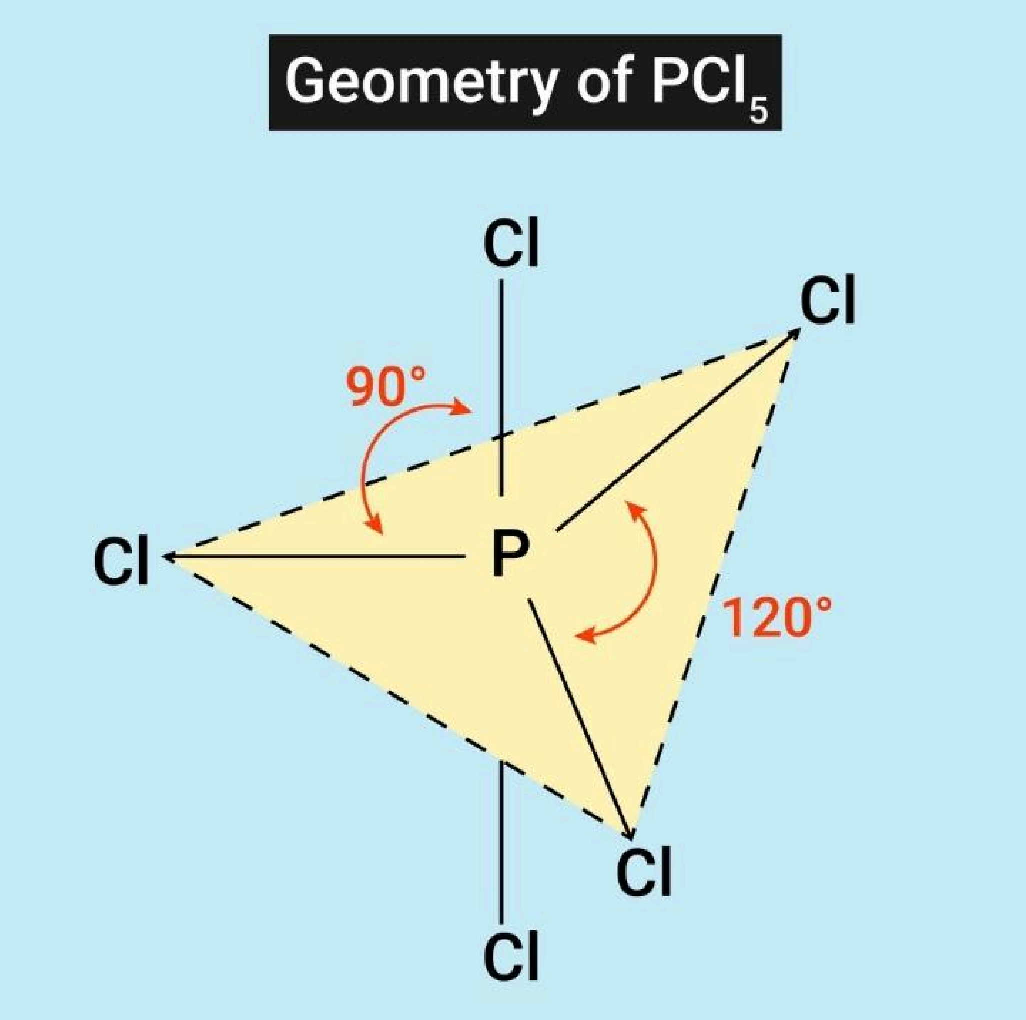
**Molecular shape:** Trigonal pyramidal

**Bond angle:** ~107°

**Explanation:**

One lone pair repels bonding pairs more strongly, compressing the bond angle slightly from the ideal tetrahedral angle (109.5°) to about 107°.

### 3. $PCl_5$ (Phosphorus Pentachloride)



- Central atom: Phosphorus
- Valence electrons = 5
- Forms 5 bonds ⟶ 5 bond pairs
- No lone pairs

**Geometry:** Trigonal bipyramidal

**Bond angles:**

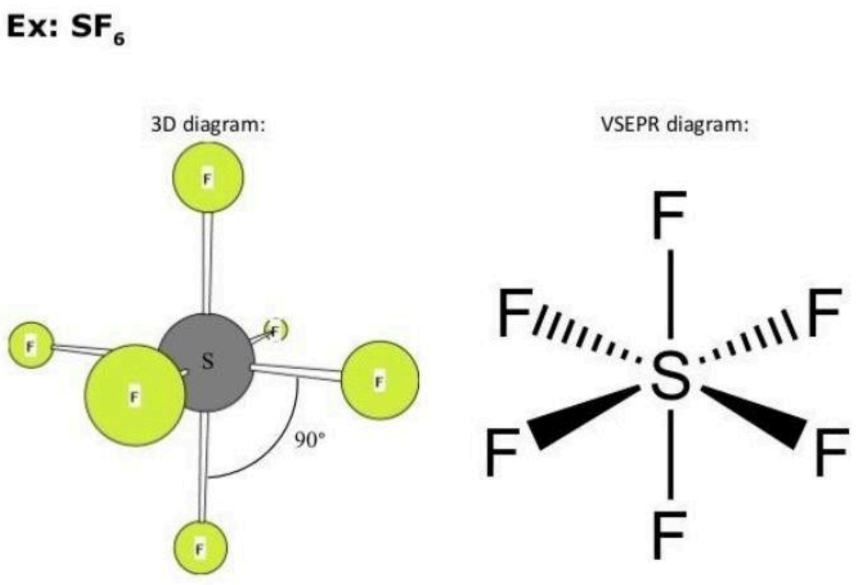
120° (equatorial–equatorial)

90° (axial–equatorial)

**Explanation:**

Since there are no lone pairs, electron pairs arrange themselves symmetrically, giving a trigonal bipyramidal shape without distortion.

### 4. $SF_6$ (Sulfur Hexafluoride)



- Central atom: Sulfur
- Valence electrons = 6
- Forms 6 bonds → 6 bond pairs
- No lone pairs

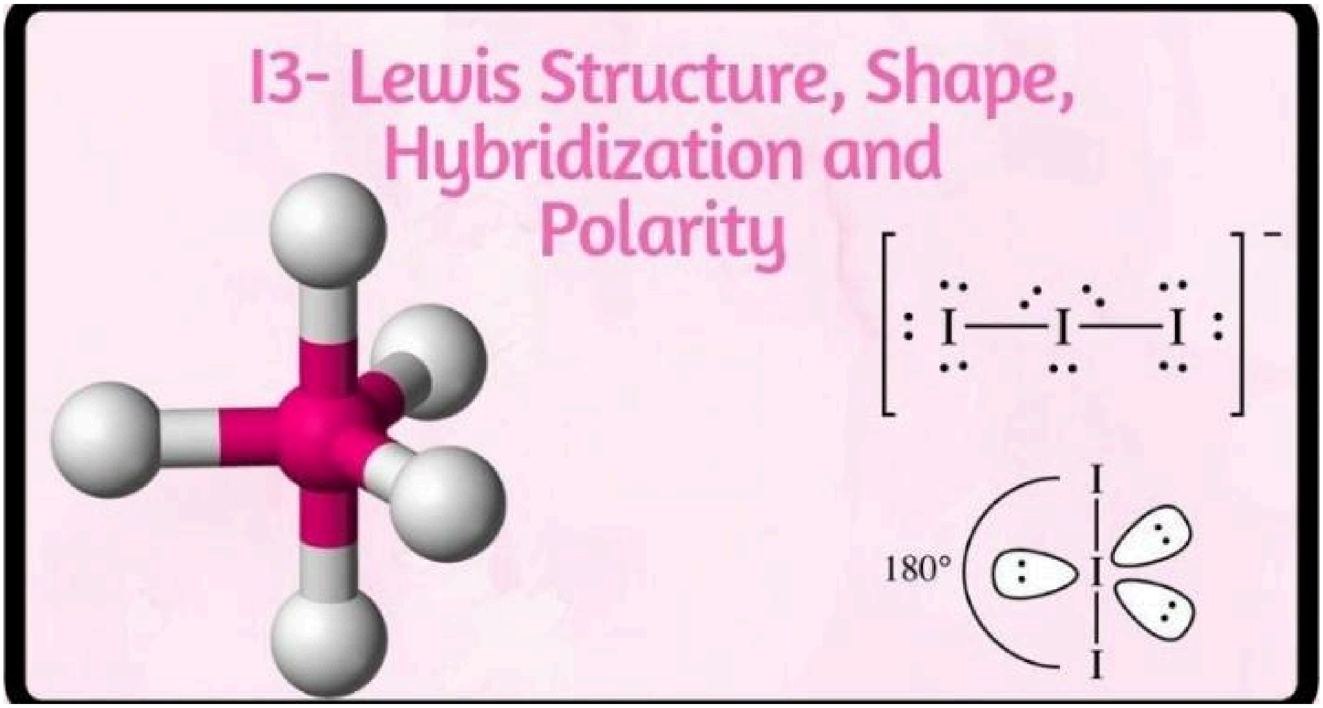
**Geometry:** Octahedral

**Bond angles:** 90°

**Explanation:**

Six bonding pairs arrange symmetrically around the sulfur atom, minimizing repulsion equally in all directions, resulting in an octahedral structure.

### 5. $I_3^-$ (Triiodide Ion)



- Central atom: Iodine
- Total electron pairs = 5
- 2 bond pairs
- 3 lone pairs

**Electron geometry:** Trigonal bipyramidal

**Molecular shape:** Linear

**Bond angle:** 180°

**Explanation:**

Three lone pairs occupy equatorial positions to minimize repulsion. The two iodine atoms remain opposite each other, producing a linear shape with a bond angle of 180°.

✦ **Summary:**

VSEPR theory explains molecular shapes based on electron pair repulsions.

- Lone pairs exert greater repulsion than bond pairs.
- Electron pairs arrange themselves to maximize distance.
- Molecular shape depends on both bonding and non-bonding electron pairs.

**Final Shapes:**

$H_2O$ ⟶ Bent (104.5°)

$NH_3$ ⟶ Trigonal pyramidal (~107°)

$PCl_5$ ⟶ Trigonal bipyramidal

$SF_6$ ⟶ Octahedral

$I_3^-$ ⟶ Linear (180°)

### 🌟 Q3. Explain Molecular Orbital Theory (MOT) with its postulates. Draw and explain molecular orbital diagrams of $H_2$ and $N_2$, and calculate their bond order.

**❖ Introduction:**

Molecular Orbital Theory (MOT) explains bonding in molecules by combining atomic orbitals to form molecular orbitals that belong to the whole molecule. Electrons are not localized between two atoms but are delocalized over the entire molecule.

### Postulates of MOT

1. Atomic orbitals of atoms combine to form molecular orbitals.

2. When two atomic orbitals combine, they form two molecular orbitals:

- Bonding Molecular Orbital (lower energy)
- Antibonding Molecular Orbital (higher energy)

3. Bonding orbitals are formed by constructive overlap, antibonding by destructive overlap.

4. Electrons fill molecular orbitals according to Aufbau principle, Pauli exclusion principle, and Hund's rule.

5. Each molecular orbital can hold a maximum of two electrons with opposite spins.

6. Bond order determines stability of a molecule.

**Bond order formula:**

Bond Order = (Number of bonding electrons − Number of antibonding electrons) / 2

## Molecular Orbital Diagram of $H_2$

**Formation:**

Two hydrogen atoms combine:

$1s + 1s \longrightarrow \sigma(1s)$ (bonding) + $\sigma^*(1s)$ (antibonding)

**Electron Filling:**

Total electrons in $H_2$ = 2

Both electrons go into bonding orbital:

$\sigma(1s)^2$

**Bond Order Calculation:**

Bond Order = (2 − 0) / 2

Bond Order = 1

**Conclusion:**

$H_2$ has a single covalent bond

It is stable

It is diamagnetic (all electrons paired)

## Molecular Orbital Diagram of $N_2$

**Electronic Configuration of Nitrogen:**

$1s^2 2s^2 2p^3$

Total valence electrons in $N_2$ = 10

**MO Energy Order for $N_2$:**

$\sigma2s < \sigma2s < \pi2p\_y = \pi2p\_z < \sigma2p\_x < \pi2p\_y = \pi2p\_z < \sigma2p\_x$

**Electron Filling:**

$\sigma(2s)^2$

$\sigma^*(2s)^2$

$\pi(2p\_y)^2$

$\pi(2p\_z)^2$

$\sigma(2p\_x)^2$

**Bonding Electrons:**

$\sigma(2s) \longrightarrow 2$

π(2p) ⟶ 4

σ(2p) ⟶ 2

Total bonding electrons = 8

**Antibonding Electrons:**

σ*(2s) ⟶ 2

Total antibonding electrons = 2

**Bond Order Calculation:**

Bond Order = (8 − 2) / 2

Bond Order = 3

**Conclusion:**

- $N_2$ has a triple bond
- It is highly stable
- It is diamagnetic (all electrons paired)

✦ **Summary:**

- MOT explains bonding using molecular orbitals formed from atomic orbitals.
- Bonding orbitals stabilize, antibonding orbitals destabilize molecules.
- Bond order indicates stability: higher bond order = stronger bond.

**Results:**

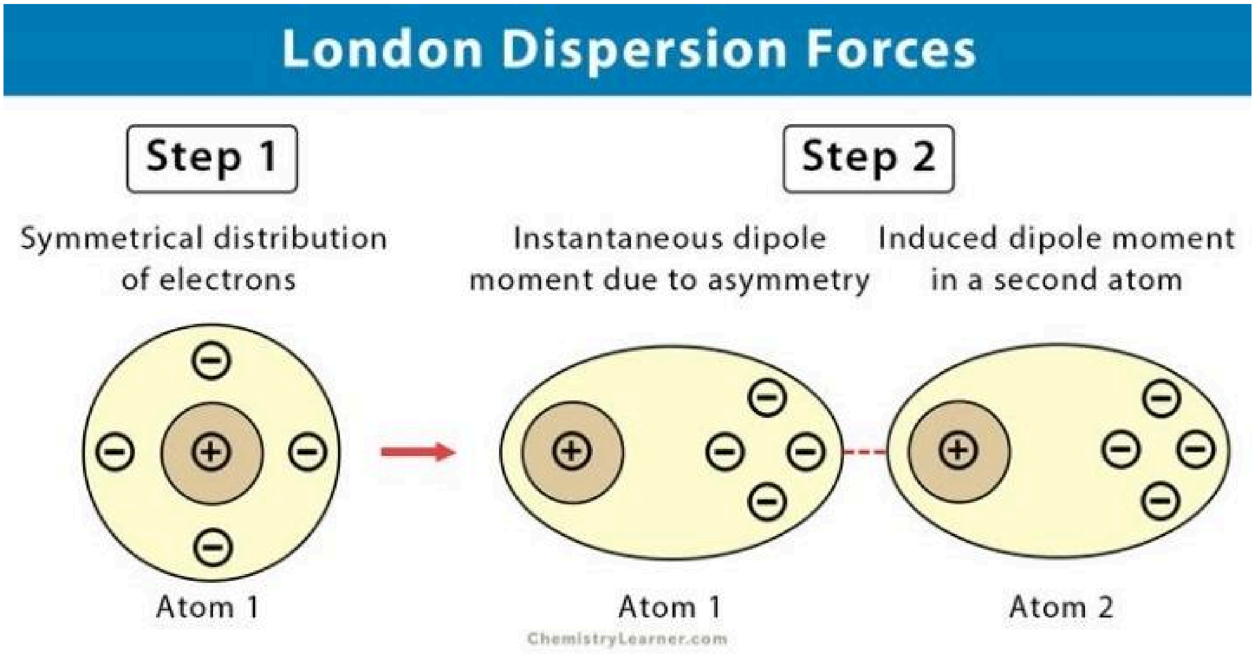
- $H_2$ ⟶ Bond order = 1 (single bond)
- $N_2$ ⟶ Bond order = 3 (triple bond)

## 🌟 Q4. Compare and explain different types of intermolecular forces including London dispersion forces, dipole–dipole forces, and hydrogen bonding with examples. Also discuss their effect on physical properties.

### ❖ Introduction

Intermolecular forces are the attractive forces between molecules. These forces are weaker than chemical bonds but strongly influence physical properties such as boiling point, melting point, volatility, and viscosity.

### 1. London Dispersion Forces (LDF)

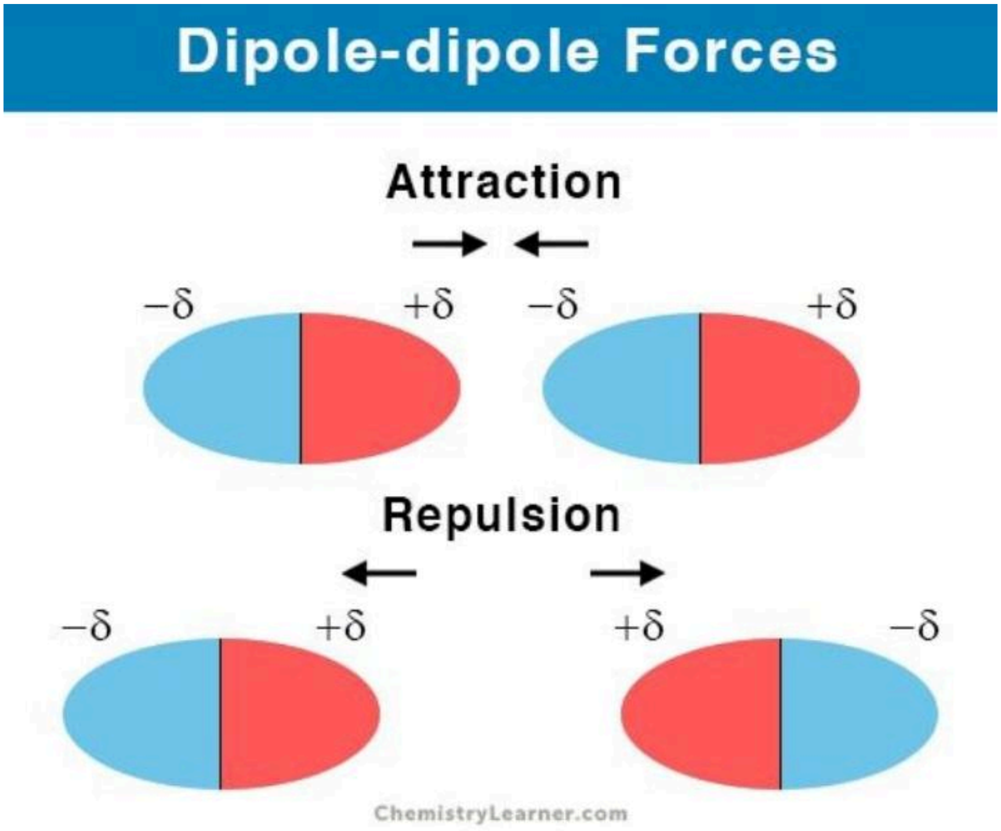


London dispersion forces arise due to temporary (instantaneous) dipoles formed by the random motion of electrons. At any moment, the electron cloud may become uneven, creating a temporary dipole that induces another dipole in a nearby molecule.

These forces are present in all atoms and molecules, especially in nonpolar substances. Their strength increases with increasing molecular size, mass, and number of electrons.

**Examples:** noble gases like helium and argon, and nonpolar molecules such as iodine ($I_2$) and methane ($CH_4$).

### 2. Dipole–Dipole Forces

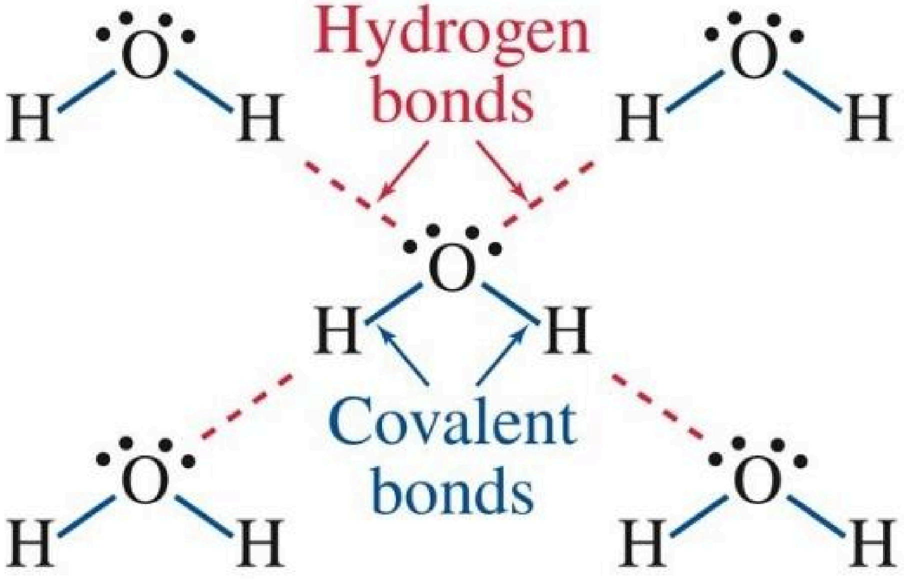


Dipole–dipole forces occur between polar molecules that have permanent dipoles. The positive end of one molecule is attracted to the negative end of another molecule.

These forces are stronger than London dispersion forces but weaker than hydrogen bonding. Their strength depends on the polarity of the molecule.

**Examples:** hydrogen chloride (HCl), sulfur dioxide ($SO_2$), and chloromethane ($CH_3Cl$).

### 3. Hydrogen Bonding



Hydrogen bonding is a special and relatively strong type of intermolecular force that occurs when hydrogen is covalently bonded to highly electronegative atoms such as nitrogen, oxygen, or fluorine and is attracted to a lone pair on another such atom.

It is highly directional and much stronger than other intermolecular forces.

**Examples:** water ($H_2O$), ammonia ($NH_3$), and hydrogen fluoride (HF).

### Comparison and Explanation of Effects on Physical Properties

The strength of intermolecular forces directly affects physical properties:

**Boiling Point:**

- Substances with stronger intermolecular forces require more energy to separate molecules, so they have higher boiling points. For example, water has a much higher boiling point than methane because of hydrogen bonding.

**Melting Point:**

- Strong intermolecular forces hold molecules tightly in a solid structure, increasing the melting point. Hydrogen-bonded substances generally have higher melting points.

**Volatility:**

- Volatility refers to how easily a substance evaporates. Weak intermolecular forces lead to high volatility because molecules can escape easily. For example, methane is more volatile than water.

**Viscosity:**

- Viscosity is resistance to flow. Strong intermolecular forces increase viscosity. Liquids with hydrogen bonding, such as glycerol or water, flow more slowly than liquids with weaker forces.

### ✦ Summary:

London dispersion forces are the weakest and exist in all substances, dipole–dipole forces occur in polar molecules, and hydrogen bonding is the strongest among intermolecular forces and occurs when hydrogen is bonded to N, O, or F.

As the strength of intermolecular forces increases, boiling point and melting point increase, while volatility decreases and viscosity increases. These forces play a key role in determining the physical behavior of substances.

🌟 Q5. Explain Valence Bond Theory (VBT) and describe the formation of sigma (σ) and pi (π) bonds. Also differentiate between sigma and pi bonds with examples.

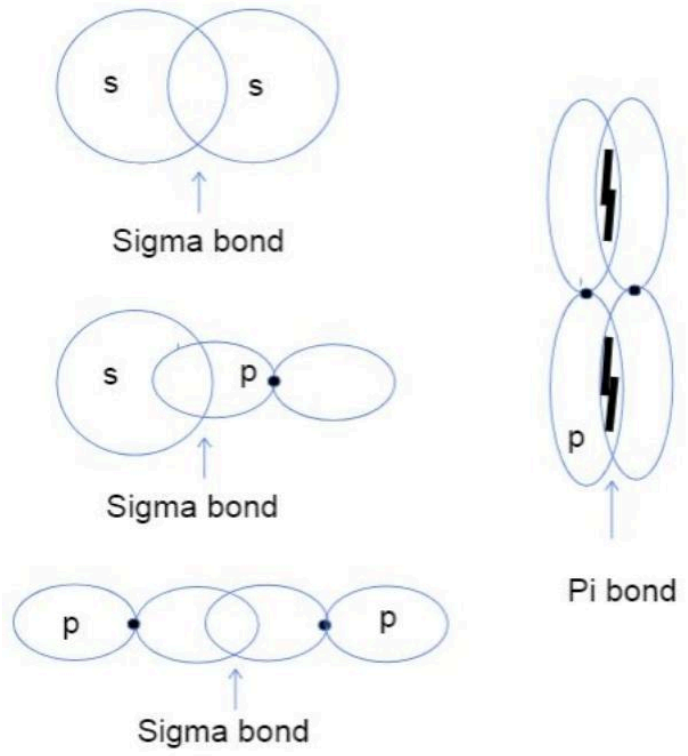
### Valence Bond Theory (VBT)

Valence Bond Theory explains that a chemical bond is formed when two atomic orbitals overlap and the electrons in these orbitals pair up with opposite spins. The overlapping orbitals contain electrons that are shared between the two atoms, which results in bond formation.

According to VBT, the strength of a bond depends on the extent of overlap between atomic orbitals. Greater overlap leads to a stronger bond. The bonded atoms try to achieve a stable electronic configuration similar to noble gases.

VBT also explains that orbitals involved in bonding must have similar energy and proper orientation to overlap effectively.

## Formation of Sigma (σ) Bond



A sigma bond is formed by head-on (axial) overlap of atomic orbitals. This overlap occurs along the internuclear axis (the line joining the nuclei of the two atoms).

Sigma bonds can form by overlap of:

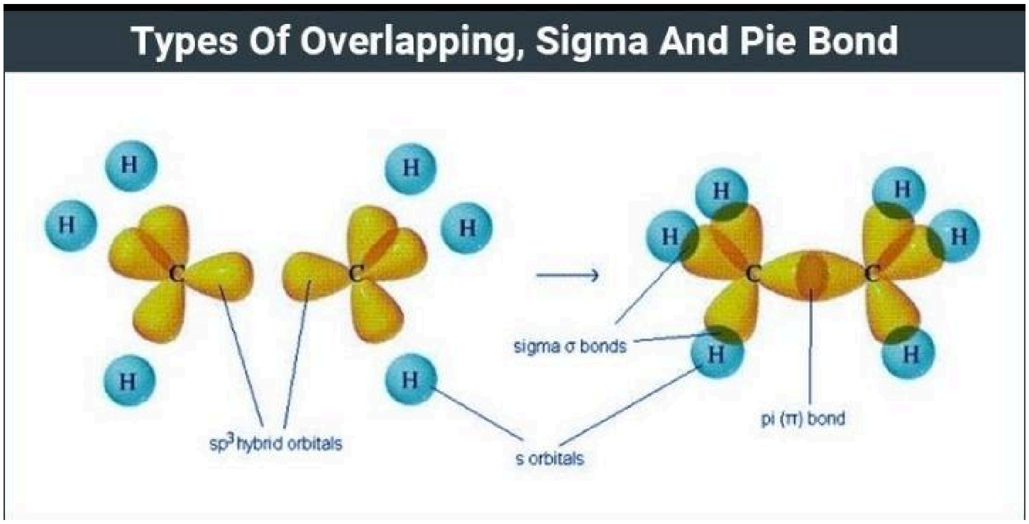
- s–s orbitals
- s–p orbitals
- p–p orbitals (end-to-end)

In a sigma bond, the electron density is concentrated directly between the nuclei, making it a strong bond.

**Examples:**

- $H_2$ molecule (s–s overlap)
- HCl molecule (s–p overlap)
- Single bonds in ethane (C–C σ bond)

## Formation of Pi (π) Bond



A pi bond is formed by sidewise (lateral) overlap of two parallel p orbitals. This overlap occurs above and below the internuclear axis.

In a pi bond, the electron density is distributed above and below the plane of the nuclei, not directly along the axis. Pi bonds are weaker than sigma bonds because the overlap is less effective.

**Examples:**

- Double bond in ethene (C=C consists of one σ and one π bond)
- Triple bond in ethyne (C≡C consists of one σ and two π bonds)

### Difference Between Sigma (σ) and Pi (π) Bonds

A sigma bond is formed by head-on overlap of orbitals, while a pi bond is formed by sidewise overlap of orbitals. Sigma bonds have electron density concentrated along the internuclear axis, whereas pi bonds have electron density above and below the axis.

Sigma bonds are generally stronger due to effective overlap, while pi bonds are comparatively weaker. A sigma bond allows free rotation of atoms around the bond axis, but a pi bond restricts rotation because rotation would break the overlap.

Sigma bonds can exist independently in single bonds, while pi bonds always exist along with a sigma bond in double or triple bonds.

✦ **Summary:**

Valence Bond Theory states that chemical bonds form due to orbital overlap and electron pairing. Sigma bonds are formed by head-on overlap and are strong, while pi bonds are formed by sidewise overlap and are weaker. In multiple bonds, one sigma bond is always present along with one or two pi bonds, which together determine the structure and stability of molecules.

# 📘 EXERCISE

## Q1. Four choice are given for each question. Select the correct choice.

**I. Chemical bond formation takes place when**

(a) Force of attraction is equal to the force of repulsion

(b) Force of repulsion is greater than force of attraction

(c) Force of attraction overcomes force of repulsion ✅

(d) None of these

**II. An ionic compound $A^+ B^-$ is most likely to be formed when**

(a) Ionization energy of A is high and electron affinity of B is low

(b) Ionization energy of A is low and electron affinity of B is high ✅

(c) Both the ionization energy of A and electron affinity of B are high

(d) Both the ionization energy of A and electron affinity of B are low

**III. Which of the following molecules has zero dipole moment?**

(a) $NH_3$

(b) $CHCl_3$

(c) $H_2O$

(d) $BF_3$ ✅

**IV. The number of σ and π bonds in $N_2$ molecule are**

(a) One σ and one π bond

(b) One σ and two π bonds ✅

(c) Three σ bonds only

(d) Two σ and one π

**V. Which of the following species has unpaired electrons in antibonding molecular orbitals?**

(a) $O_2^{2+}$

(b) $N_2^{2-}$

(c) B

(d) $F_2$ ✅

**VI. The shape of $ICl_3$ according to VSEPR model is**

(a) Tetrahedral

(b) Trigonal planar

(c) Trigonal bipyramidal

(d) T-shape ✅

**VII. Which of the following molecules has a net dipole moment?**

(a) $CO_2$

(b) $CS_2$

(c) $SO_2$ ✅

(d) $CCl_4$

**VIII. How many electrons are present in the valence shell of P in $PO_4^{3-}$?**

(a) 8

(b) 10

(c) 12 ✅

(d) 14

**IX. How many extra electrons than a normal octet are present in the valence shell of I in $ICl_3$?**

(a) 2

(b) 3

(c) 4 ✅

(d) 5

**X. What is the type and shape of $[ICl_4]^-$ according to VSEPR model?**

(a) $AB_4$ tetrahedral

(b) $AB_4$ pyramidal

(c) $AB_5$ trigonal bipyramidal

(d) $AB_6$ square planar ✅

**XI. Which of the following molecules has $sp^3$ hybridization and tetrahedral geometry?**

(a) $BF_3$

(b) $SO_2$

(c) $CCl_4$ ✅

(d) $PCl_5$

**XII. Which of the following species contains a dative bond?**

(a) $CH_4$

(b) NaCl

(c) $NH_4^+$ ✅

(d) $O_2$

## Q.2 Attempt the following short-answer questions:

### a. Define the following:

**(i) What is Dipole?**

A dipole is formed when there is a separation of positive and negative charges in a molecule due to difference in electronegativity.

**Example:** HCl ($H\delta^+$ — $Cl\delta^-$)

**(ii) What is Bond Order?**

Bond order is the number of bonds between two atoms in a molecule.

Bond Order = (Bonding electrons − Antibonding electrons) / 2

**Example:** $O_2$ has bond order = 2

**(iii) What is Permanent Dipole–Dipole Force?**

It is the force of attraction between polar molecules due to permanent positive and negative ends.

**Example:** HCl molecules attract each other.

**(iv) What is London Dispersion Force?**

It is a weak force caused by temporary dipoles due to random movement of electrons.

**Example:** $CH_4$, noble gases

## b) Draw the Lewis (electron dot) structures for the following species

### (i) Draw the Lewis structure of HCN

**Structure:**

H — C ≡ N

- Nitrogen has one lone pair

### (ii) Draw the Lewis structure of $NCl_3$

**Structure:**

Cl — N — Cl

|

Cl

- Nitrogen has one lone pair

### (iii) Draw the Lewis structure of CO

**Structure:**

C ≡ O

- One lone pair on each atom
- Contains coordinate bond

### (iv) Draw the Lewis structure of $O_3$ (Ozone)

Structure (Resonance):

O = O — O ↔ O — O = O

- One single and one double bond
- Charges present on atoms

**(v) Draw the Lewis structure of $NO_2$**

**Structure:**

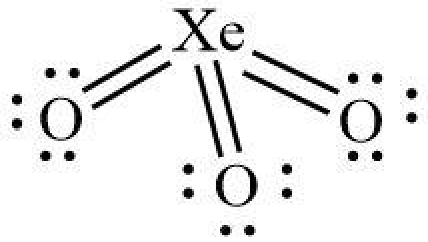
O — N = O

- One unpaired electron
- Shows resonance

**Final Tip (Exam Point of View)**

- Always show lone pairs in structures
- Mention resonance where needed
- For $NO_2$, mention free radical
- For CO, mention coordinate bond

### c. Xenon is a noble gas (Group 18); Xenon trioxide ($XeO_3$) has the following structure



**(i) By counting electron pairs, explain why $XeO_3$ has this shape**

Xenon (Xe) belongs to Group 18 and has 8 valence electrons.

In $XeO_3$:

- Xe forms bonds with 3 oxygen atoms ⟶ 3 bond pairs
- Remaining electrons on Xe form 1 lone pair

👉 Total electron pairs around Xe = 4 (3 bond pairs + 1 lone pair)

**According to VSEPR Theory:**

4 electron pairs arrange themselves in tetrahedral geometry to minimize repulsion.

**However**, one of these is a lone pair, which occupies more space than bond pairs.

👉 **Therefore**, the actual molecular shape becomes trigonal pyramidal instead of tetrahedral.

**Reason (Important for Exam):**

**Repulsion order:**

- Lone pair–bond pair > bond pair–bond pair
- The lone pair pushes the three bonded oxygen atoms downward
- This causes distortion in shape and reduces bond angles

**(ii) Structure with partial charges and dipole moment**

**Partial Charges:**

- Oxygen is more electronegative than xenon
- So each oxygen atom carries partial negative charge ($\delta^-$)
- Xenon carries partial positive charge ($\delta^+$)

**Dipole Moment Explanation:**

- Each Xe–O bond has a dipole directed from Xe to O
- Because the molecule is not symmetrical (due to lone pair), these dipoles do not cancel

👉 **Therefore,** $XeO_3$ has a net dipole moment and is a polar molecule

✦ **Summary:**

- Total electron pairs = 4 ⟶ $sp^3$ hybridization
- Electron geometry = tetrahedral
- Molecular shape = trigonal pyramidal
- Lone pair causes distortion
- Oxygen = $\delta^-$, Xenon = $\delta^+$
- Dipoles do not cancel ⟶ molecule is polar

✅ **Exam Tip (Very Important):**

**Always write:**

- Electron pair count (3 BP + 1 LP)
- Geometry vs shape
- Lone pair effect
- Polarity reason

### d. Explain the difference between the formation of o and π bonds in terms of VBT.

**Answer:**

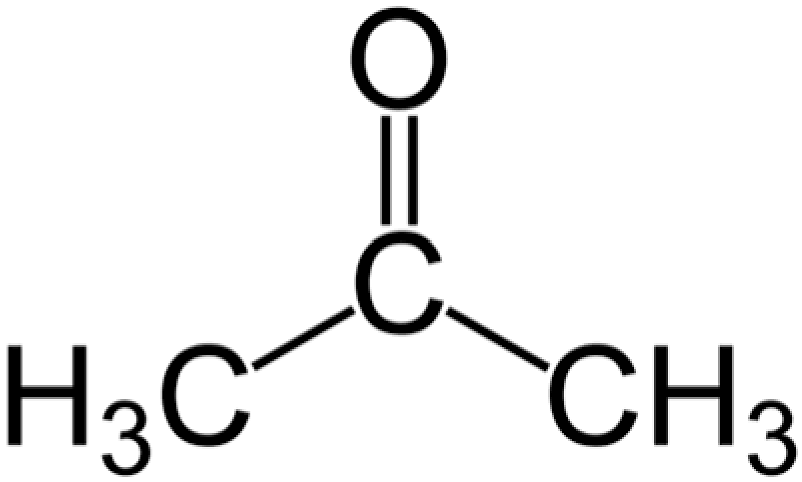
A σ (sigma) bond is formed by the head-on (axial) overlap of atomic orbitals such as s–s, s–p, or p–p along the internuclear axis. The electron density is concentrated directly between the nuclei.

A π (pi) bond is formed by the sidewise (lateral) overlap of two parallel unhybridized p orbitals. The electron density is distributed above and below the internuclear axis.

**In summary:**

σ bonds are formed by head-on overlap along the axis, whereas π bonds are formed by sidewise overlap of parallel p orbitals.

### e. The structure of propanone (acetone) is:



**i. Show how the central carbon atom forms σ and π bonds through hybridization.**

**Answer:**

In propanone, the central carbon (the carbonyl carbon) is $sp^2$-hybridized.

- It forms three σ-bonds using its $sp^2$ hybrid orbitals:
- One σ-bond with the oxygen atom
- One σ-bond with the left $CH_3$ group
- One σ-bond with the right $CH_3$ group

The remaining unhybridized p-orbital on the central carbon overlaps sideways with a p-orbital on oxygen to form a π-bond.

✦ **Summary:**

The carbonyl carbon in propanone is $sp^2$-hybridized, forming three σ-bonds with $sp^2$ orbitals and one π-bond through p-orbital overlap.

**ii. Can propanone make a hydrogen bond with water when both are intermixed?**

**Answer:**

**Yes**. Propanone can form hydrogen bonds with water, but only as a hydrogen-bond acceptor, not a donor.

The oxygen atom in propanone has lone pairs, which can form hydrogen bonds with the hydrogen atoms of water.

**However,** propanone has no hydrogen attached to oxygen or nitrogen, so it cannot donate a hydrogen bond.

✦ **Summary:**

Propanone forms hydrogen bonds with water because its oxygen atom can accept hydrogen bonds from water molecules.

**f. Predict the shapes of sulfate ($[SO_4]^{2-}$), borohydride ($[BH_4]^-$) and tri-iodide ($[I_3]^-$) ions according to the VSEPR model.**

**Answer:**

**Sulfate ion ($[SO_4]^{2-}$):**

- Central sulfur atom has four bonding pairs and no lone pair ⟶ Tetrahedral shape.

**Borohydride ion ($[BH_4]^-$):**

- Central boron atom has four bonding pairs and no lone pair ⟶ Tetrahedral shape.

**Tri-iodide ion ($[I_3]^-$):**

- Central iodine atom has two bonding pairs and three lone pairs ⟶ Linear shape (due to repulsion of lone pairs).

**In summary:**

- $[SO_4]^{2-}$ ⟶ Tetrahedral
- $[BH_4]^-$ ⟶ Tetrahedral

- $[I_3]^-$ ⟶ Linear

**g. Sketch the molecular orbital pictures of π(2p) and π*(2p).**

**Answer:**

1. π(2p) bonding molecular orbital (side-by-side overlap, same phase):

```
(+)       (+)

  p         p

   \       /

    \     /

     \   /

      \ /

      / \

     /   \

    /     \

   /       \

 (-)       (-)
```

- The lobes of the p orbitals overlap side-by-side in the same phase.
- Electron density is concentrated above and below the internuclear axis.
- This is a bonding molecular orbital (π).

### 2. π*(2p) antibonding molecular orbital (side-by-side overlap, opposite phase):

```
 (+)      (-)

  p        p

   \      /

    \    /

     \  /

      \/

      /\

     /  \

    /    \

   /      \

 (-)      (+)
```

- The p orbitals overlap side-by-side out of phase.
- A node (zero electron density) exists between the nuclei.
- This is an antibonding molecular orbital (π*).

### ✦ Summary:

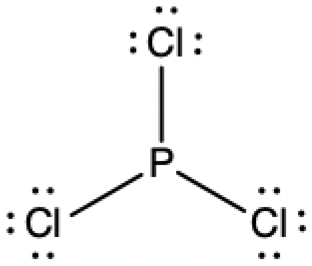
- **π(2p):** constructive overlap ⟶ bonding orbital
- **π*(2p):** destructive overlap ⟶ antibonding orbital

Electron density in π is between and around the axis, while π* has a node between nuclei.

### h. Sketch the hybrid orbitals and bond formation in $PCl_3$, $SiCl_4$ and $NH_4^+$.

**Answer:**

### 1. $PCl_3$ (Phosphorus trichloride)



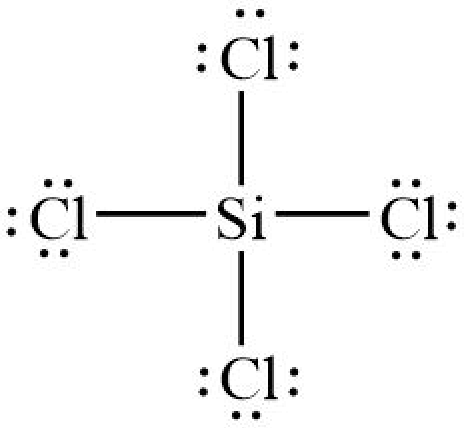
Phosphorus has 5 valence electrons. In $PCl_3$, it forms three σ bonds with three chlorine atoms and retains one lone pair. To achieve bonding, one s orbital and three p orbitals of phosphorus mix to form four $sp^3$ hybrid orbitals.

**Out of these four $sp^3$ hybrid orbitals:**

- Three hybrid orbitals overlap with p orbitals of chlorine atoms to form three σ bonds.
- One hybrid orbital contains a lone pair.

- Due to the presence of one lone pair and three bond pairs, the electron pair geometry is tetrahedral, but the molecular shape becomes trigonal pyramidal.

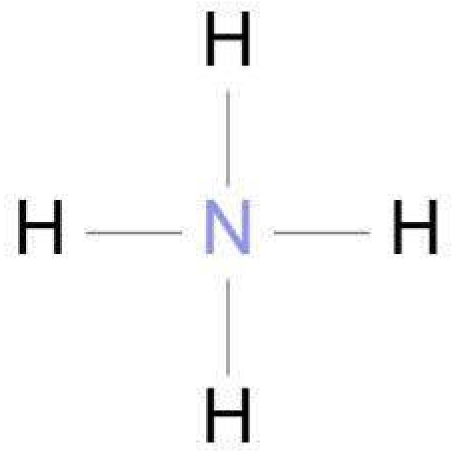
### 2. $SiCl_4$ (Silicon tetrachloride)



Silicon has 4 valence electrons. It forms four σ bonds with four chlorine atoms. For bonding, one s orbital and three p orbitals of silicon hybridize to form four $sp^3$ hybrid orbitals.

- Each $sp^3$ hybrid orbital overlaps with a chlorine p orbital to form a σ bond.
- There are no lone pairs on silicon.
- With four bond pairs and no lone pairs, the arrangement is tetrahedral, and the molecular shape is also tetrahedral.

### 3. $NH_4^+$ (Ammonium ion)



Nitrogen has 5 valence electrons. In $NH_4^+$, it forms four σ bonds with hydrogen atoms. One lone pair of nitrogen is donated to form an additional bond due to the positive charge, resulting in four bonding pairs and no lone pair.

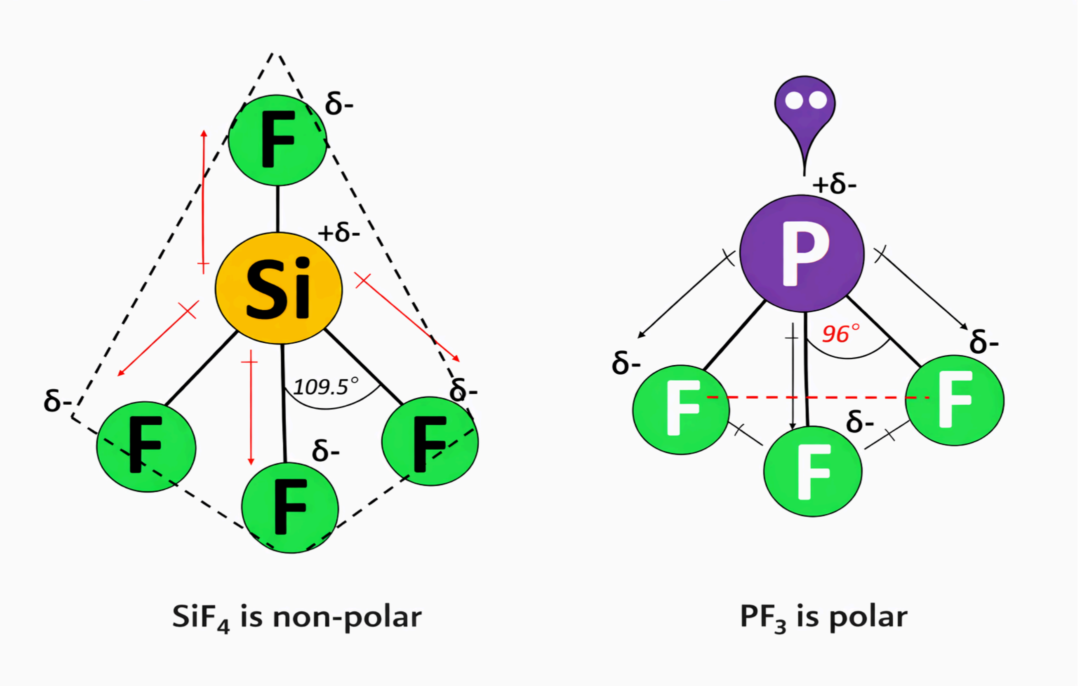
Nitrogen undergoes $sp^3$ hybridization, forming four equivalent hybrid orbitals. Each hybrid orbital overlaps with the 1s orbital of hydrogen to form σ bonds.

Because there are four bond pairs and no lone pairs, the geometry is tetrahedral.

✦ **Summary:**

- **$PCl_3$:** $sp^3$ hybridization, 3 σ bonds + 1 lone pair ⟶ trigonal pyramidal
- **$SiCl_4$:** $sp^3$ hybridization, 4 σ bonds, no lone pair ⟶ tetrahedral
- **$NH_4^+$:** $sp^3$ hybridization, 4 σ bonds, no lone pair ⟶ tetrahedral

### i. The structures of $PF_3$ and $SiF_4$ are given. Redraw these with partial charges and state which is polar and which is non-polar.



**Answer:**

**$PF_3$ (Phosphorus trifluoride)**

- Phosphorus is less electronegative than fluorine, so F atoms carry δ− and P carries δ+.
- $PF_3$ has a trigonal pyramidal shape due to one lone pair on phosphorus.

Because of the asymmetrical shape and presence of a lone pair, the bond dipoles do not cancel.

➡️ **$PF_3$ is polar**

**$SiF_4$ (Silicon tetrafluoride)**

- Silicon is less electronegative than fluorine, so F atoms carry δ− and Si carries δ+.
- $SiF_4$ has a tetrahedral and symmetrical shape.

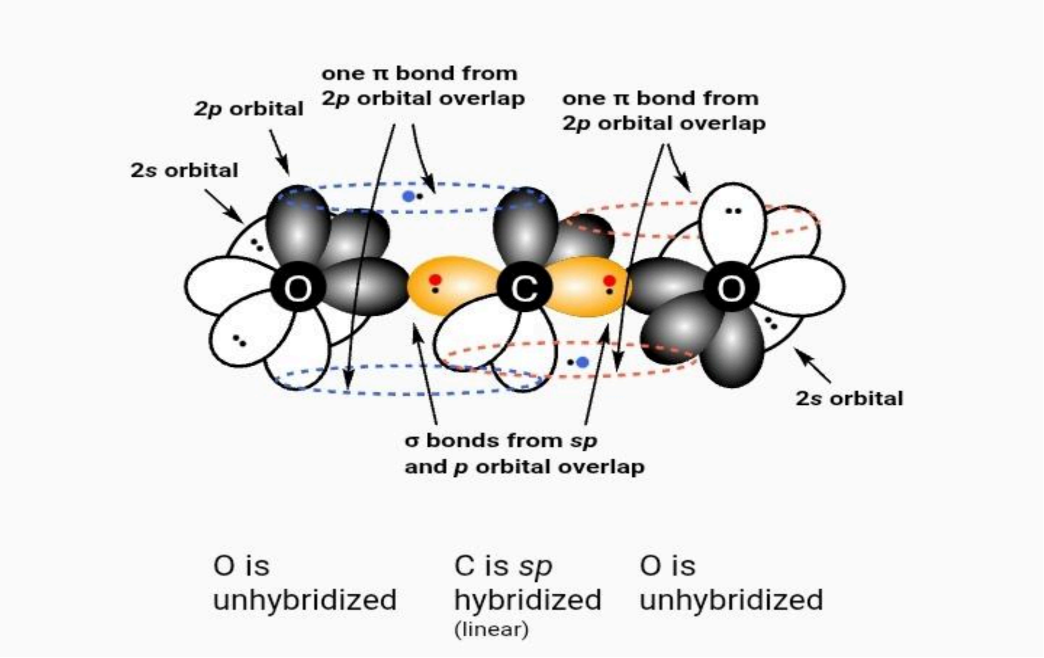
Due to the symmetrical tetrahedral arrangement, all bond dipoles cancel each other.

➡️ **$SiF_4$ is non-polar**

**Final Answer:**

- $PF_3$ ⟶ Polar (due to trigonal pyramidal shape and lone pair)
- $SiF_4$ ⟶ Non-polar (due to symmetrical tetrahedral shape)

### j. Draw the orbital structure of $CO_2$ in terms of Valence Bond Theory (VBT).



**Answer:**

In $CO_2$, the central carbon atom is sp hybridized. One 2s and one 2p orbital of carbon mix to form two sp hybrid orbitals arranged linearly at 180°. The remaining two unhybridized 2p orbitals are perpendicular to each other and participate in π bonding.

**Bond formation:**

- Two σ bonds are formed by head-on overlap of sp orbitals of carbon with p orbitals of oxygen.
- Two π bonds are formed by sidewise overlap of unhybridized p orbitals of carbon with p orbitals of oxygen.

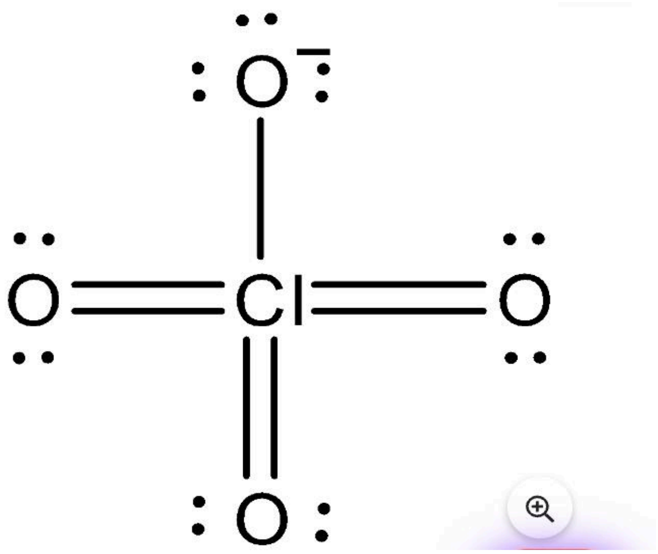
**Conclusion:**

- Carbon is sp hybridized
- Geometry is linear (180°)
- Each C=O double bond consists of 1 σ bond and 1 π bond formed by orbital overlap in VBT.

### k. Draw the Lewis structures and state whether the following ions involve expanded octets:

**i) $ClO_4^-$ (Perchlorate ion)**

**Lewis structure:**



**Central atom: Cl**

Cl is bonded to 4 oxygen atoms (resonance structures exist)

**Expanded octet:** Yes

- Chlorine has more than 8 electrons around it.

**ii) $ICl_4^-$ (Tetrachloroiodate ion)**

**Lewis structure:**

Cl

|

Cl — I — Cl

|

Cl

(2 lone pairs on I)

Central iodine has 4 bonding pairs and 2 lone pairs.

**Expanded octet:** Yes

- Iodine has more than 8 electrons.

**iii) $NH_4^+$ (Ammonium ion)**

**Lewis structure:**

H

|

H — N — H

|

H

Nitrogen forms 4 single bonds with H atoms.

**Expanded octet:** ❌ No

- Nitrogen has exactly 8 electrons (octet satisfied).

**iv) $I_3^-$ (Triiodide ion)**

**Lewis structure:**

I — I — I

(central I has 3 lone pairs)

Central iodine has 2 bonding pairs and 3 lone pairs.

**Expanded octet:** Yes

- Central iodine has more than 8 electrons.

**Final Conclusion:**

- $ClO_4^-$ ⟶ Expanded octet: Yes
- $ICl_4^-$ ⟶ Expanded octet: Yes
- $NH_4^+$ ⟶ Expanded octet: No
- $I_3^-$ ⟶ Expanded octet: Yes

### L. The bond between K and Cl is ionic but that between Si and Cl is polar covalent. Explain why.

**Answer:**

**K and Cl:**

Potassium (K) has low ionization energy and readily loses one electron to form $K^+$, while chlorine (Cl) has high electronegativity and gains that electron to form $Cl^-$. Due to the large difference in electronegativity and complete transfer of electrons, the bond formed is ionic (electrovalent).

**Si and Cl:**

Silicon (Si) and chlorine (Cl) have a smaller electronegativity difference. Instead of complete transfer, they share electrons unequally, with chlorine attracting the shared electron pair more strongly. This results in a polar covalent bond.

**Conclusion:**

- K–Cl ⟶ ionic bond (electron transfer due to large electronegativity difference)
- Si–Cl ⟶ polar covalent bond (unequal sharing of electrons due to moderate electronegativity difference)

### m. $SO_2$ is a polar molecule but $SO_3$ is not. Justify.

**Answer:**

**$SO_2$ (Sulfur dioxide):**

- The central sulfur atom has one lone pair and two bonding pairs. According to VSEPR theory, the molecule has a bent (V-shaped) geometry. Due to this asymmetrical shape, the dipole moments of the S–O bonds do not cancel each other. Hence, $SO_2$ is polar.

**$SO_3$ (Sulfur trioxide):**

- The central sulfur atom has no lone pair and three bonding pairs. The molecule has a trigonal planar geometry with symmetrical arrangement of the three S–O bonds. The bond dipoles cancel each other due to symmetry. Hence, $SO_3$ is non-polar.

**Conclusion:**

- $SO_2$ ⟶ Polar (bent shape, dipoles do not cancel)
- $SO_3$ ⟶ Non-polar (trigonal planar shape, dipoles cancel)

## n. Which of $O_2^{2+}$ and $O_2^{2-}$ would be paramagnetic? Give reason in the light of Molecular Orbital Theory (MOT).

**Answer:**

Using Molecular Orbital Theory (MOT), paramagnetism depends on the presence of unpaired electrons in molecular orbitals.

**$O_2^{2+}$ (dioxygen dication):**

- Removal of two electrons from $O_2$ removes electrons from antibonding π* orbitals. All remaining electrons become paired.

⟶ Diamagnetic (no unpaired electrons)

**$O_2^{2-}$ (peroxide ion):**

- Addition of two electrons to $O_2$ fills the antibonding π* orbitals, pairing all electrons.

⟶ Diamagnetic (no unpaired electrons)

**Conclusion:**

Neither $O_2^{2+}$ nor $O_2^{2-}$ is paramagnetic; both are diamagnetic because all electrons are paired in their molecular orbitals.

## o. Which of the following bonds would be most polar?

(i) C–Cl

(ii) Si–F

(iii) Se–F

**Answer: (ii) Si–F**

**Explanation:**

Bond polarity depends on the difference in electronegativity (ΔEN) between the two bonded atoms. A larger electronegativity difference results in a more polar bond.

**Electronegativity values (approx.):**

- C ≈ 2.5
- Si ≈ 1.8
- Se ≈ 2.4
- Cl ≈ 3.0
- F ≈ 4.0

**Electronegativity differences:**

C–Cl ⟶ $|3.0 - 2.5| = 0.5$

Si–F ⟶ $|4.0 - 1.8| = 2.2$

Se–F ⟶ $|4.0 - 2.4| = 1.6$

Since Si–F has the largest electronegativity difference, it is the most polar bond among the given options.

**Conclusion:**

👉 Si–F is the most polar bond.

## P. Compare the bond energies of single, double, and triple bonds between the same two atoms (e.g., H–H, O=O, N≡N). Explain the trend in terms of the number of shared electrons.

**Answer:**

For the same pair of atoms, bond energy increases with bond order:

Single bond < Double bond < Triple bond

**Examples:**

H–H (single) < O=O (double) < N≡N (triple)

**Explanation:**

- A single bond involves the sharing of one pair of electrons (2 electrons).
- A double bond involves the sharing of two pairs of electrons (4 electrons).
- A triple bond involves the sharing of three pairs of electrons (6 electrons).

**As the number of shared electron pairs increases:**

- The attraction between the nuclei and shared electrons increases
- The bond length decreases
- The bond becomes stronger, hence bond energy increases

**Conclusion:**

More shared electrons ⟶ stronger attraction ⟶ higher bond energy

⇒ Triple bonds are strongest, followed by double bonds, then single bonds.

# DESCRIPTIVE QUESTIONS

## 🌟 Q.3 How the bonding in the following molecules can be explained with respect to Valence Bond Theory (VBT)?

### i) $Cl_2$ (Chlorine molecule)

Each chlorine atom has 7 valence electrons and requires one more electron to complete its octet.

In $Cl_2$, each chlorine atom contributes one unpaired electron in a 3p orbital.

- These two half-filled 3p orbitals overlap head-on to form a σ bond.
- A shared pair of electrons is formed between the two chlorine atoms.

👉 **Result:** One σ bond formed by p–p overlap.

### ii) $O_2$ (Oxygen molecule)

Each oxygen atom has 6 valence electrons. To complete the octet, each forms two bonds.

- One σ bond is formed by head-on overlap of $sp^2$ (or p–p) orbitals.

- One π bond is formed by sidewise overlap of unhybridized p orbitals.

👉 **Result:** A double bond (O=O) consisting of 1 σ bond + 1 π bond.

### iii) $N_2$ (Nitrogen molecule)

Each nitrogen atom has 5 valence electrons.

- One σ bond is formed by head-on overlap of sp (or p–p) orbitals.
- Two π bonds are formed by sidewise overlap of two pairs of unhybridized p orbitals.

👉 **Result:** A triple bond (N≡N) consisting of 1 σ bond + 2 π bonds.

### iv) HF (Hydrogen fluoride)

Hydrogen has one electron ($1s^1$) and fluorine has 7 valence electrons.

- The 1s orbital of hydrogen overlaps head-on with a half-filled 2p orbital of fluorine.
- This forms a σ bond.

👉 **Result:** A single σ bond formed by s–p overlap.

### v) $H_2S$ (Hydrogen sulfide)

Sulfur has 6 valence electrons and forms two bonds with hydrogen atoms.

- Sulfur undergoes $sp^3$ hybridization (one s + three p orbitals).

- Two $sp^3$ hybrid orbitals contain lone pairs, and two contain unpaired electrons.
- Each unpaired electron overlaps with the 1s orbital of hydrogen to form σ bonds.

👉 **Result:** Two σ bonds and two lone pairs on sulfur; molecular shape is bent.

✦ **Summary:**

| **Molecule** | **Type of Bonding (VBT)** |
|---|---|
| **$Cl_2$** | Single σ bond (p–p overlap) |
| **$O_2$** | Double bond (1 σ + 1 π) |
| **$N_2$** | Triple bond (1 σ + 2 π) |
| **HF** | Single σ bond (s–p overlap) |
| **$H_2S$** | Two σ bonds ($sp^3$ hybridization on S) |

🌟 **Q.4 What are the postulates of VSEPR model? Discuss the structures of the following species with reference to this theory.**

## Postulates of VSEPR Model

1. The shape of a molecule depends on the number of electron pairs (bond pairs and lone pairs) around the central atom.
2. Electron pairs repel each other and arrange themselves to minimize repulsion.
3. Repulsion order: lone pair–lone pair > lone pair–bond pair > bond pair–bond pair.

4. Multiple bonds are treated as a single electron domain but exert slightly greater repulsion than single bonds.
5. The actual shape of the molecule depends on the positions of atoms (bond pairs), not lone pairs.

## Structures of the given species

### i) $CH_4$ (Methane)

Carbon has four bonding pairs and no lone pairs. To minimize repulsion, the four electron pairs arrange themselves in a tetrahedral geometry. All bond angles are approximately 109.5°.

### ii) $NH_3$ (Ammonia)

Nitrogen has three bonding pairs and one lone pair. The electron pairs adopt a tetrahedral arrangement, but due to the presence of a lone pair (which repels more strongly), the shape becomes trigonal pyramidal. The bond angle is slightly less than 109.5°, about 107°.

### iii) $H_3O^+$ (Hydronium ion)

Oxygen has three bonding pairs and one lone pair. The electron pair geometry is tetrahedral, but the molecular shape is trigonal pyramidal due to the lone pair. The bond angle is around 107°.

### iv) $PCl_5$ (Phosphorus pentachloride)

Phosphorus has five bonding pairs and no lone pairs. To minimize repulsion, the electron pairs arrange themselves in a trigonal bipyramidal geometry, with bond angles of 90° and 120°.

### v) $SO_2$ (Sulfur dioxide)

Sulfur has two bonding pairs and one lone pair. The electron pair geometry is trigonal planar, but due to the lone pair, the molecular shape becomes bent (V-shaped). The bond angle is slightly less than 120°, around 119°.

### vi) $SF_6$ (Sulfur hexafluoride)

Sulfur has six bonding pairs and no lone pairs. The electron pairs arrange themselves in an octahedral geometry, with bond angles of 90°.

### Conclusion:

According to VSEPR theory, molecular shapes are determined by minimizing electron pair repulsions, and the presence of lone pairs distorts ideal geometries, leading to shapes like trigonal pyramidal and bent structures.

## 🌟 Q.5 Explain the orbital hybridization for $CH_4$, $NH_3$, $BF_3$ and $BeCl_2$.

### 1) $CH_4$ (Methane)

Carbon has ground state configuration: $1s^2\ 2s^2\ 2p^2$.

One 2s orbital and three 2p orbitals mix to form four equivalent $sp^3$ hybrid orbitals.

- Each $sp^3$ orbital overlaps with the 1s orbital of hydrogen to form σ bonds.
- Geometry is tetrahedral with bond angle 109.5°.

### 2) $NH_3$ (Ammonia)

Nitrogen has five valence electrons. It undergoes $sp^3$ hybridization.

- One s and three p orbitals form four $sp^3$ hybrid orbitals.
- Three $sp^3$ orbitals form σ bonds with hydrogen atoms.
- One $sp^3$ orbital contains a lone pair.
- Due to lone pair–bond pair repulsion, the shape is trigonal pyramidal with bond angle ~107°.

### 3) $BF_3$ (Boron trifluoride)

Boron has three valence electrons. It undergoes $sp^2$ hybridization.

- One s and two p orbitals mix to form three $sp^2$ hybrid orbitals.
- These orbitals form σ bonds with three fluorine atoms.
- One p orbital remains unhybridized and is empty.
- Geometry is trigonal planar with bond angle 120°.

### 4) $BeCl_2$ (Beryllium chloride)

Beryllium has two valence electrons. It undergoes sp hybridization.

- One s and one p orbital mix to form two sp hybrid orbitals.

- Each sp orbital forms a σ bond with a chlorine atom.
- No lone pairs are present on Be.
- Geometry is linear with bond angle 180°.

✦ **Summary:**

- $CH_4$ ⟶ $sp^3$ hybridization ⟶ tetrahedral
- $NH_3$ ⟶ $sp^3$ hybridization ⟶ trigonal pyramidal
- $BF_3$ ⟶ $sp^2$ hybridization ⟶ trigonal planar
- $BeCl_2$ ⟶ sp hybridization ⟶ linear

### 🌟 Q.6 Draw the molecular orbital diagrams of the following molecules. Calculate their bond orders.

#### I) $H_2$

MO formation:

- σ(1s) (bonding orbital)
- σ*(1s) (antibonding orbital)

**Electron filling:**

σ(1s): ⬆️⬇️

σ*(1s): empty

Bond order = (bonding electrons − antibonding electrons) ÷ 2

= (2 − 0) ÷ 2 = 1

**Answer:** Bond order = 1 ⟶ single bond, stable molecule

## II) $He_2$

MO formation:

- σ(1s)
- σ*(1s)

Electron filling:

- σ(1s): ⬆️⬆️⬇️
- σ*(1s): ⬆️⬇️

Bond order = (2 − 2) ÷ 2 = 0

**Answer:** Bond order = 0 → no bond formation, molecule is unstable

## III) $N_2$

Total electrons = 14

MO energy order: σ(1s), σ*(1s), σ(2s), σ*(2s), π(2p), σ(2p)

Electron filling:

σ(1s): ⬆️⬇️

σ*(1s): ⬆️⬇️

σ(2s): ⬆️⬇️

σ*(2s): ⬆️⬇️

π(2p): ⬆️⬇️ ⬆️⬇️

σ(2p): ⬆️⬇️

Bonding electrons = 10

Antibonding electrons = 4

Bond order = (10 − 4) ÷ 2 = 3

**Answer:** Bond order = 3 ⟶ triple bond, very strong and stable molecule

### IV) $O_2$

Total electrons = 16

MO energy order: σ(1s), σ*(1s), σ(2s), σ*(2s), σ(2p), π(2p), π*(2p)

Electron filling:

σ(1s): ⬆️⬇️

σ*(1s): ⬆️⬇️

σ(2s): ⬆️⬇️

σ*(2s): ⬆️⬇️

σ(2p): ⬆️⬇️

π(2p): ⬆️⬇️ ⬆️⬇️

π*(2p): ⬆️⬆️

Bonding electrons = 10

Antibonding electrons = 6

Bond order = (10 − 6) ÷ 2 = 2

**Answer:** Bond order = 2 ⟶ double bond

$O_2$ is paramagnetic due to two unpaired electrons in π* orbitals

### 🌟 Q.7 Discuss the formation of $F_2$ molecule in the light of Lewis concept, VBT, and MOT.

### 1) Lewis Concept

Fluorine atom has 7 valence electrons ($2s^2\ 2p^5$). Each F atom needs one more electron to complete its octet.

- Each fluorine shares one electron with the other.
- A single covalent bond is formed.
- Each fluorine achieves a stable octet.
- Each fluorine also has three lone pairs.

Lewis structure: F — F

(each F has 3 lone pairs)

**Conclusion:**

$F_2$ is formed by sharing one pair of electrons ⟶ single covalent bond.

### 2) Valence Bond Theory (VBT)

**According to VBT:**

- Each fluorine atom has one unpaired electron in a 2p orbital.

- When two F atoms approach, their half-filled p orbitals overlap end-to-end.
- This overlap forms a σ (sigma) bond.

**Key points:**

- The bond is formed by p–p overlap.
- The bond is a σ bond (head-on overlap).
- No hybridization is involved in simple VBT description.

**Conclusion:**

$F_2$ molecule is formed by the overlap of half-filled 2p orbitals of two fluorine atoms forming a σ bond.

### 3) Molecular Orbital Theory (MOT)

**In MOT:**

- Atomic orbitals combine to form molecular orbitals.
- For $F_2$, the relevant orbitals are 2p orbitals.

MO configuration (valence electrons only):

- σ(2s), σ*(2s), σ(2p), π(2p), π*(2p), σ*(2p)

$F_2$ has 14 valence electrons (7 + 7).

Electron filling results in:

- All bonding orbitals filled
- Some antibonding orbitals also filled

Bond order calculation: Bond order = (bonding electrons − antibonding electrons) ÷ 2

= (10 − 8) ÷ 2 = 1

**Also:**

All electrons are paired ⟶ diamagnetic molecule

**Conclusion:**

$F_2$ has bond order 1 (single bond) and is diamagnetic according to MOT.

✦ **Summary:**

**Lewis Concept:** Single covalent bond formed by sharing one pair of electrons to complete octet.

**VBT:** σ bond formed by overlap of half-filled 2p orbitals.

**MOT:** Bond order = 1; all electrons paired ⟶ diamagnetic molecule.

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