CHEMISTRY

Unit 3: Chemical Bonding

Complete MCQs • Short Questions • Long Questions

Compiled & Presented by

taleemzone.com

SECTION 1

Multiple Choice Questions

47 MCQs with correct answers, reasons & solutions

Exercise MCQs

1. Chemical bond formation takes place when:

(a) Force of attraction are 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

✓ **Correct Answer: (c) Force of attraction overcomes force of repulsion**

2. An ionic compound A+B- is most likely to be formed when:

(a) The ionization energy of A is high and electron affinity of B is low (b) The 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

✓ **Correct Answer: (b) The ionization energy of A is low and electron affinity of B is high**

3. Which of the following molecules has zero dipole moment?

(a) NH3 (b) CHCl3 (c) H2O (d) BF3

✓ **Correct Answer: (d) BF3**

4. The numbers of sigma and pi bonds in the N2 molecule are:

(a) one sigma and one pi bonds (b) one sigma and two pi bonds (c) three sigma bonds only (d) two sigma and one pi

✓ **Correct Answer: (b) one sigma and two pi bonds**

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

(a) O2(2+) (b) N2(2-) (c) B2 (d) F2

✓ **Correct Answer: (a) O2(2+)**

6. The shape of ICl5 according to the VSEPR model is:

(a) Tetrahedral (b) Trigonal planar (c) Trigonal bipyramidal (d) T-shape

✓ **Correct Answer: (d) T-shape**

7. Which of the following molecules has a dipole moment?

(a) CO2 (b) CS2 (c) SO2 (d) CCl4

✓ **Correct Answer: (c) SO2**

8. How many electrons are present in the valence shell of P in PO4(3-)?

(a) 8 (b) 10 (c) 12 (d) 14

✓ **Correct Answer: (c) 12**

9. How many extra electrons than a normal octet are there in the valence shell of I in ICl5?

(a) 2 (b) 3 (c) 4 (d) 5

✓ **Correct Answer: (a) 2**

10. What is the type of [ICl4]- according to the VSEPR model?

(a) AB4 tetrahedral (b) AB4 pyramidal (c) AB5 trigonal bipyramidal (d) AB6 square planar

✓ **Correct Answer: (a) AB4 tetrahedral**

11. Which of the following molecules has a central atom with sp3 hybridization and a tetrahedral electron pair geometry?

(a) BF3 (b) SO2 (c) CCl4 (d) PCl5

✓ **Correct Answer: (c) CCl4**

12. Which of the following species contains a dative bond?

(a) CH4 (b) NaCl (c) NH4+ (d) O2

✓ **Correct Answer: (c) NH4+**

SLO Based MCQs - Chemical Bond

13. In which of the following molecules, ionic bond is found?

(a) NH3 (b) HF (c) CsI (d) CaC2

✓ **Correct Answer: (c) CsI**

14. Which of the following has all the three characters i.e. ionic, covalent and coordinate covalent?

(a) H2O (b) KBr (c) NH3 (d) NH4Cl

✓ **Correct Answer: (d) NH4Cl**

15. Which one is a polar molecule?

(a) CO2 (b) HCl (c) BF3 (d) CCl4

✓ **Correct Answer: (b) HCl**

16. Which one is a non-polar molecule?

(a) CaCl2 (b) CaC2 (c) CS2 (d) H2S

✓ **Correct Answer: (c) CS2**

17. Which of the following shows incorrect bond polarity?

(a) H(d+) - F(d-) (b) H(d+) - Br(d-) (c) C(d+) - Cl(d-) (d) C(d+) - C(d+)

✓ **Correct Answer: (d) C(d+) - C(d+)**

18. Which of the following molecules has zero dipole moment?

(a) NH_3 (b) $CHCl_3$ (c) H_2O (d) BF_3

✓ **Correct Answer: (d) BF_3**

19. The most electronegative atom is:

(a) N (b) C (c) O (d) F

✓ **Correct Answer: (d) F**

20. The weakest bond is:

(a) ionic (b) polar covalent (c) non-polar covalent (d) hydrogen bond

✓ **Correct Answer: (d) hydrogen bond**

21. Which of the following elements has incomplete octet after making the bond?

(a) C in CH_4 (b) O in H_2O (c) Be in $BeCl_2$ (d) N in NH_3

✓ **Correct Answer: (c) Be in $BeCl_2$**

22. Octet rule is not obeyed by which of the following compounds during its formation?

(a) H_2O (b) NaCl (c) PF_5 (d) NH_3

✓ **Correct Answer: (c) PF_5**

23. Covalent bonds are:

(a) Rigid and directional (b) Non rigid and directional (c) Rigid and non directional (d) Non rigid and non directional

✓ **Correct Answer: (a) Rigid and directional**

Electronegativity Difference

24. If the electronegativity difference between bonded atoms is more than 0.5 but less than 1.7, the bond will be:

(a) Ionic (b) Polar covalent (c) Non polar covalent (d) Coordinate covalent

✓ **Correct Answer: (b) Polar covalent**

25. Tendency of an atom to attract the shared pair of electrons is called:

(a) Ionization Energy (b) Electron Affinity (c) Electronegativity (d) Electropositivity

✓ **Correct Answer: (c) Electronegativity**

Bond Length and Energy

26. The bond distance between H-H is:

(a) 436.45 pm (b) 74.5 pm (c) 154 pm (d) 133 pm

✓ **Correct Answer: (b) 74.5 pm**

27. The bond formation energy of a compound is:

(a) less than bond dissociation energy (b) greater than bond dissociation energy (c) equal to bond dissociation energy (d) inversely proportional to bond dissociation energy

✓ **Correct Answer: (c) equal to bond dissociation energy**

28. The actual bond length in a polar covalent compound is:

(a) lesser than expected (b) greater than expected (c) equal to the expected (d) exactly half of expected

✓ **Correct Answer: (a) lesser than expected**

29. The expected bond energy of HCl is lesser than the actual, this is because:

(a) Size of hydrogen is very small (b) HCl is non-polar compound (c) HCl is a polar compound (d) There exists hydrogen bonding

✓ **Correct Answer: (c) HCl is a polar compound**

Dipole Moment

30. The molecule with greatest dipole moment is:

(a) H2O (b) H2S (c) NH3 (d) HF

✓ **Correct Answer: (a) H2O**

31. Which will have zero dipole moment?

(a) H2O (b) H2S (c) BF3 (d) CHCl3

✓ **Correct Answer: (c) BF3**

32. The SI Unit for dipole moment is:

(a) Debye (b) Nm (c) Nm4 (d) mC

✓ **Correct Answer: (d) mC**

33. One Debye is equal to:

(a) 3.336 x 10^-30 mC (b) 9.1 x 10^31 mC (c) 1.66 x 10^24 mC (d) 6.06 x 10^23 mC

✓ **Correct Answer: (a) 3.336 x 10^-30 mC**

34. Dipole moment is the product of:

(a) Charge x displacement (b) Charge x distance (c) Newton x displacement (d) Charge x mass

✓ **Correct Answer: (b) Charge x distance**

35. Dipole moment of CO2 is:

(a) 1.8D (b) 1.94D (c) 1.0D (d) zero

✓ **Correct Answer: (d) zero**

VSEPR Theory

36. The molecule with linear structure is:

(a) H2O (b) H2S (c) CO2 (d) BF3

✓ **Correct Answer: (c) CO2**

37. The number of lone pair of electrons in ammonium ion is:

(a) one (b) three (c) two (d) zero

✓ **Correct Answer: (d) zero**

38. Which one is phosphonium ion?

(a) HPO4(-1) (b) NH2(-1) (c) PH4(+1) (d) PH3

✓ **Correct Answer: (c) PH4(+1)**

39. What is molecular geometry of SO4(2-) ion?

(a) Tetrahedral (b) Trigonal planar (c) Trigonal pyramidal (d) Linear

✓ **Correct Answer: (a) Tetrahedral**

40. Which one is not AB4 type molecule according to VSEPR theory?

(a) SO2 (b) SO4(2-) (c) PH3 (d) H2S

✓ **Correct Answer: (a) SO2**

Hybridization

41. sp3 Hybridization is associated with structure:

(a) linear (b) trigonal (c) tetrahedral (d) octahedral

✓ **Correct Answer: (c) tetrahedral**

42. Which type of hybridization of C occurs in CO2?

(a) sp3 (b) sp2 (c) sp (d) dsp2

✓ **Correct Answer: (c) sp**

43. The bond angle in ethene molecule is:

(a) 180 degrees (b) 109.5 degrees (c) 120 degrees (d) 107.5 degrees

✓ **Correct Answer: (c) 120 degrees**

44. Which of the following molecules show sp hybridization of central atom?

(a) BeCl2 (b) NH3 (c) PH3 (d) C2H4

✓ **Correct Answer: (a) BeCl2**

45. What is bond angle in NF3?

(a) 109.5 degrees (b) 107.5 degrees (c) 104.5 degrees (d) 102 degrees

✓ **Correct Answer: (d) 102 degrees**

46. Which of the following hybridizations is found in ethyne?

(a) sp2 (b) sp (c) sp3 (d) No hybridization

✓ **Correct Answer: (b) sp**

Molecular Orbital Theory (MOT)

47. Which one of the following will show paramagnetic property?

(a) N2 (b) O2 (c) O2(2-) (d) O2(2+)

✓ **Correct Answer: (b) O2**

48. Which of the following statements is not correct regarding bonding molecular orbitals?

(a) Bonding molecular orbitals possess less energy than atomic orbitals from which they are formed (b) Bonding molecular orbitals have low electron density between the two nuclei (c) Every electron in the bonding molecular orbitals contributes to the attraction between atoms (d) Bonding molecular orbitals are formed when the electron waves undergo constructive interference

✓ **Correct Answer: (b) Bonding molecular orbitals have low electron density between the two nuclei**

49. Which of the following molecules has unpaired electrons in antibonding molecular orbitals?

(a) O2(2+) (b) N2(2-) (c) B2 (d) F2

✓ **Correct Answer: (a) O2(2+)**

50. Highest bond order is found in:

(a) O2 (b) O2(-1) (c) O2(+2) (d) O2(+1)

✓ **Correct Answer: (c) O2(+2)**

51. Paramagnetic specie is:

(a) O2(2-) (b) N2 (c) O2(2-) (d) N2(2-)

✓ **Correct Answer: (d) N2(2-)**

52. Which theory explains the paramagnetic behaviour exhibited by O2 molecule?

(a) Band theory (b) VBT (c) VSEPR theory (d) MOT

✓ **Correct Answer: (d) MOT**

53. The number of bonds in nitrogen molecules is:

(a) One sigma and one pi (b) One sigma and two pi (c) Three sigma only (d) Two sigma and two pi

✓ **Correct Answer: (b) One sigma and two pi**

Answer Key

Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	c	15	b	29	c	43	c
2	b	16	c	30	a	44	a
3	d	17	d	31	c	45	d
4	b	18	d	32	d	46	b
5	a	19	d	33	a	47	b
6	d	20	d	34	b	48	b
7	c	21	c	35	d	49	a
8	c	22	c	36	c	50	c
9	a	23	a	37	d	51	d
10	a	24	b	38	c	52	d
11	c	25	c	39	a	53	b
12	c	26	b	40	a	-	-
13	c	27	c	41	c	-	-
14	d	28	a	42	c	-	-

Concept	Key Point
Ionic Bond	Complete transfer of electrons; formed between metal and non-metal
Covalent Bond	Sharing of electrons; directional and rigid
Dative Bond	Both electrons donated by one atom
Sigma Bond	Head-to-head overlap; stronger
Pi Bond	Sideways overlap; weaker
Dipole Moment	Measure of molecular polarity (u = q x d)
VSEPR Theory	Electron pairs repel to minimize repulsion
Hybridization	Mixing of orbitals to form equivalent orbitals
Bond Order	(Bonding e- minus Antibonding e-) / 2

Shape	Examples	Bond Angle
Linear	BeCl2, CO2, I3-	180 degrees
Trigonal Planar	BF3, SO3	120 degrees
Tetrahedral	CH4, CCl4, NH4+	109.5 degrees
Trigonal Pyramidal	NH3, PH3, PCl3	less than 109.5 degrees
V-shaped (Bent)	H2O, SO2, H2S	less than 109.5 degrees
Trigonal Bipyramidal	PCl5	90, 120 degrees
Octahedral	SF6, XeF4	90 degrees

Hybridization	Geometry	Examples
sp	Linear	BeCl2, CO2, C2H2
sp2	Trigonal Planar	BF3, C2H4, SO3
sp3	Tetrahedral	CH4, NH3, H2O, CCl4

Species	Unpaired Electrons	Magnetic Nature
O2	2	Paramagnetic
O2(2-)	0	Diamagnetic
N2	0	Diamagnetic
B2	2	Paramagnetic

SECTION 2

Short Questions

Concise answers, equations & comparison tables

Exercise Short Question Answers

Q1. Define the following: (i) Dipole (ii) Bond Order (iii) Permanent Dipole-Permanent Dipole Force (iv) London Dispersion Force

(i) Dipole

A molecule with a partial positive charge on one part and a partial negative charge on the other part is called a dipole. Examples: HF, HCl, HBr.

(ii) Bond Order

The number of bonds formed between two atoms after the atomic orbitals overlap is called bond order. It is half the difference between the number of bonding electrons (a) and antibonding electrons (b).

```
Bond Order = (a - b) / 2
```

(iii) Permanent Dipole-Permanent Dipole Force

The negative end of one permanent dipole attracts the positive end of another permanent dipole molecule. This force of attraction is called permanent dipole-permanent dipole force.

```
H-Cl ... H-Cl ... H-Cl (d+ ... d- attraction chain)
```

(iv) London Dispersion Force

The momentary force of attraction between an instantaneous dipole molecule and an induced dipole molecule is called London dispersion force. It is present in all types of molecules, whether polar or non-polar.

Q2. Draw the Lewis (electron dot) structures for HCN, NCl3, CO, O3 and NO2.

HCN: H-C triple bond N (linear). NCl3: N bonded singly to three Cl atoms with one lone pair on N (pyramidal). CO: C triple bond O with one lone pair on each atom. O3: central O double bonded to one O and single/dative bonded to the other, bent shape. NO2: central N double bonded to one O and single bonded to the other with one unpaired electron, bent shape.

Q3. Xenon trioxide (XeO3): explain its shape by counting electron pairs, and show partial charges and the dipole direction.

XeO3 has a trigonal pyramidal shape due to lone pair-bond pair repulsion. Xe has 8 valence electrons and forms three bonds with oxygen atoms (3 bond pairs) plus one lone pair, giving 4 electron pairs (tetrahedral electron geometry). With one lone pair, the actual molecular shape becomes trigonal pyramidal, with the dipole pointing from the electropositive Xe towards the more electronegative oxygen atoms and the lone pair.

Q4. Explain the difference between the formation of sigma and pi bonds.

Sigma Bond	Pi Bond
Formed by head-to-head overlap along the internuclear axis	Formed by lateral (sideways) overlap of unhybridized p-orbitals
Electron density maximum around the internuclear axis	Electron density located above and below the internuclear axis
Stronger bond	Weaker bond
Formed by s-s, s-p, or p-p overlap	Formed only by p-p overlap
Single bond always contains one sigma bond	Present in double and triple bonds

Q5. Propanone (acetone): (i) show how the central carbon forms sigma and pi bonds through hybridization (ii) can it hydrogen bond with water?

(i) Hybridization of the carbonyl carbon

The central (carbonyl) carbon atom in acetone is sp2 hybridized. Three sp2 hybrid orbitals form three sigma bonds (two with CH3 groups and one with the O atom); the unhybridized p orbital forms a pi bond with oxygen (C=O). The geometry around the carbonyl carbon is trigonal planar.

(ii) Hydrogen bonding with water

Yes, propanone can form hydrogen bonds with water. The oxygen atom of the carbonyl group (C=O) has lone pairs and acts as a hydrogen bond acceptor, allowing water molecules to hydrogen bond with it.

Q6. Predict the shapes of sulfate (SO4 2-), borate (BH4-) and tri-iodide (I3-) ions using VSEPR.

Ion	Shape	Reason
SO4(2-)	Tetrahedral	Central S has 4 bond pairs, 0 lone pairs
BH4-	Tetrahedral	Central B has 4 bond pairs, 0 lone pairs
I3-	Linear	Central I has 2 bond pairs, 3 lone pairs

Q7. Sketch the molecular orbital pictures of pi(2p) and pi*(2p).

pi(2p) Bonding MO - formed by sideways overlap of p orbitals; has a nodal plane on the nuclear axis (zero electron density there); electron density lies above and below the axis; lower energy than the parent atomic orbitals. pi*(2p) Anti-bonding MO - formed by sideways overlap of opposite-sign lobes; has least electron density in the inter-nuclear region and an additional nodal plane between the nuclei; higher energy than the parent atomic orbitals.

Q8. Sketch the hybrid orbitals and bond formation in PCl3, SiCl4 and NH4+.

Species	Hybridization	Bonding	Shape
PCl3	sp3	3 sigma bonds, 1 lone pair on P	Trigonal pyramidal
SiCl4	sp3	4 sigma bonds with 4 Cl atoms	Tetrahedral
NH4+	sp3	4 sigma bonds with 4 H atoms, no lone pair	Tetrahedral

Q9. PF3 and SiF4: show partial charges and state which is polar and which is non-polar.

Molecule	Charge Distribution	Polarity	Shape
PF3	delta- on F, delta+ on P	Polar (u = 0.985 D)	Trigonal pyramidal, dipoles do not cancel
SiF4	delta- on F, delta+ on Si	Non-polar (u = 0 D)	Tetrahedral, dipoles cancel

Q10. Draw the orbital structure of the CO2 molecule in terms of VBT.

CO2 is a linear molecule with a bond angle of 180 degrees. The central carbon atom is sp hybridized. Two sp hybrid orbitals form sigma bonds with the two oxygen atoms, while two unhybridized p orbitals form pi bonds with the oxygen atoms.

```
O = C = O (linear, 180 degrees)
```

Q11. Draw Lewis structures for ClO4-, ICl4-, NH4+ and I3- and state whether they involve expanded octets.

Ion	Electrons Around Central Atom	Expanded Octet?
ClO4-	12 (Cl, 3rd period, has empty d orbitals)	Yes
ICl4-	10 (I, 5th period)	Yes
NH4+	8	No, obeys octet rule
I3-	10 (I, 5th period)	Yes

Q12. The bond between K and Cl is ionic but that between Si and Cl is polar covalent. Explain why.

Bond	Electronegativity Difference	Type of Bond
K-Cl	2.34 (greater than 1.8)	Ionic
Si-Cl	1.26 (0.4 to 1.8)	Polar Covalent

The K-Cl bond is ionic because the electronegativity difference exceeds 1.8. The Si-Cl bond is polar covalent because the difference falls between 0.4 and 1.8.

Q13. SO2 is a polar molecule but SO3 is not. Justify.

Molecule	Geometry	Dipole Cancellation	Result
SO2	Bent (V-shaped)	Bond dipoles do not cancel	Polar
SO3	Trigonal planar	Bond dipoles cancel	Non-polar (net dipole = 0)

Q14. Which of O2 and O2 2- would be paramagnetic? Give reason using MOT.

Species	Unpaired Electrons	Magnetic Nature
O2	2 (in pi2py and pi2pz)	Paramagnetic
O2(2-)	0	Diamagnetic

In the O2 molecular orbital diagram there are two unpaired electrons in the pi(2py) and pi(2pz) orbitals, making O2 paramagnetic. In O2(2-), the two extra electrons completely fill the pi* orbitals, removing the unpaired electrons and making it diamagnetic.

Q15. Which bond is most polar: C-Cl, Si-F or Se-F?

Bond	Electronegativity Difference	Polarity
C-Cl	0.66	Slightly polar
Si-F	2.1	Ionic
Se-F	1.45	Most polar (of the covalent options)

Se-F is the most polar covalent bond among these.

Q16. Compare the bond energies of single, double and triple bonds between the same two atoms.

```
Triple Bond > Double Bond > Single Bond (bond strength order)
```

Single bond = 1 shared pair; double bond = 2 shared pairs; triple bond = 3 shared pairs. As the number of shared electron pairs increases, nuclear attraction becomes stronger, bond length becomes shorter, and bond energy becomes higher. Example: N triple-bond N > O double-bond O > F-F.

SLO Based Short Question Answers - Chemical Bond

Q17. Define ionic bond.

According to Lewis theory, an ionic bond is formed by the complete transfer of electrons from an atom with low ionization energy to another atom with high electron affinity.

Q18. What is a dative bond?

A dative bond is formed between two atoms when the shared pair of electrons is donated by only one of the bonding atoms.

Q19. What is an ionic bond? Explain with an example.

An ionic bond is formed by the complete transfer of one or more electrons from one atom to another, producing oppositely charged ions, and occurs between metals and non-metals. Example: in NaCl, sodium donates one electron to chlorine, forming Na+ and Cl- ions held together by electrostatic attraction.

Q20. Why are ionic compounds usually solid and have high melting points?

Ionic compounds form a regular lattice of alternating positive and negative ions. The strong electrostatic forces between oppositely charged ions require a large amount of energy to break, so ionic compounds are usually hard solids with high melting and boiling points.

Q21. Explain why ionic compounds conduct electricity in molten or aqueous states but not in solid state.

In the solid state, ions are fixed in position and cannot move freely. When melted or dissolved in water, the ions become free to move, allowing them to conduct electricity. Mobile ions are essential for electrical conductivity in ionic substances.

Q22. How is lattice energy related to ionic bond strength?

Lattice energy is directly related to the strength of an ionic bond - greater lattice energy means a stronger ionic bond. It depends on ion size and charge; smaller, highly charged ions form lattices with higher lattice energies.

Q23. Why are covalent compounds generally poor conductors of electricity?

Covalent compounds do not contain free ions or electrons that can carry electric current, since atoms are held together by shared electron pairs in neutral molecules. So, in solid, liquid or aqueous form, covalent compounds are poor conductors.

Q24. How does a double or triple covalent bond differ from a single bond?

Bond Type	Shared Electron Pairs	Strength	Length
Single Bond	1 pair	Weakest	Longest
Double Bond	2 pairs	Stronger	Shorter
Triple Bond	3 pairs	Strongest	Shortest

Example: oxygen forms a double bond (O2), nitrogen forms a triple bond (N2).

Q25. What is bond order and how is it related to bond strength?

Bond order refers to the number of chemical bonds between a pair of atoms; a higher bond order means stronger and shorter bonds. N2 bond order = 3 (strongest), O2 bond order = 2 (stronger), F2 bond order = 1 (weakest).

Q26. Explain with an example how bond polarity arises in a covalent bond.

Bond polarity arises when two atoms with different electronegativities share electrons unequally; the more electronegative atom develops a partial negative charge. Example: in HCl, chlorine is more electronegative than hydrogen, making the bond polar.

```
H(delta+) ------ Cl(delta-)
```

Q27. What is a coordinate covalent bond? How is it different from a regular covalent bond?

A coordinate covalent bond is formed when one atom donates both electrons shared in the bond, whereas in a regular covalent bond each atom contributes one electron. Example: formation of NH4+, where NH3 donates a lone pair to H+.

```
H3N: + H+ → [H3N→H]+
```

Q28. What conditions are necessary for the formation of coordinate bonds?

One atom must have a lone pair of electrons (the donor) and the other atom must have an empty orbital (the acceptor). The donor provides both electrons to form the shared pair, resulting in a stable coordinate covalent bond.

Q29. What is metallic bonding? How is it different from ionic or covalent bonding?

Metallic bonding is the attraction between positively charged metal ions and a "sea" of delocalized electrons. Unlike ionic or covalent bonds, electrons are not shared or transferred between specific atoms; this delocalization gives metals their unique properties.

Q30. Why are metals good conductors of electricity?

In metals, valence electrons are free to move throughout the structure. These mobile electrons carry electric current easily when a voltage is applied, making metals excellent conductors in both solid and liquid states.

Dipole Moment

Q31. Define Dipole moment.

Dipole moment is a quantitative measurement of the polarity of a bond or a molecule. For diatomic molecules like HF, HCl, HBr, HI and NO, it is directed from the positive end (delta+) to the negative end (delta-). It is measured in Debye units (D); a higher dipole moment indicates greater polarity.

Q32. Why is the dipole moment of CCl4 zero?

CCl4 has four C-Cl bonds that are individually polar due to the large electronegativity difference between C and Cl, but the bond moments are directed such that they cancel each other out in the perfectly tetrahedral geometry. The net dipole moment is zero, making CCl4 a non-polar molecule.

Q33. Why does BeCl2 have zero dipole moment?

BeCl2 is a linear molecule with two identical Cl atoms on either side of the central Be atom at 180 degrees. The individual Be-Cl bond moments are equal in magnitude and opposite in direction, so they cancel out, giving zero dipole moment.

Q34. Why does BF3 have zero dipole moment?

BF3 has a symmetrical trigonal planar structure. As a rule, when a molecule has identical ligands arranged in a regular geometry, the individual bond dipoles cancel each other, giving an overall dipole moment of zero (non-polar).

Q35. What is bond polarity and how is it determined?

Bond polarity arises when atoms with different electronegativities form a covalent bond; the more electronegative atom develops a partial negative charge while the other becomes partially positive. The greater the electronegativity difference, the more polar the bond.

Q36. Define dipole moment and what does it indicate?

Dipole moment is a measure of the polarity of a molecule; it is the product of the charge and the distance between the charges. A higher dipole moment indicates a more polar molecule. It is represented by the Greek letter mu and measured in Debye units (D).

Q37. Why is water a polar molecule but carbon dioxide non-polar?

Molecule	Shape	Dipole Cancellation	Polarity
H2O	Bent (V-shaped)	Do not cancel	Polar
CO2	Linear	Cancel each other	Non-polar

Water's bent shape (due to lone pairs on oxygen) means its bond dipoles do not cancel, while linear CO2 has its two C=O bond dipoles cancel out exactly.

Expanded Octets

Q38. Explain how the octet expands in SO4(2-).

```
Valence electrons around S = 2x(double bond electrons) + 2x(single bond electrons) = 2(4) + 2(2) = 12
```

Sulfur has 12 electrons in its valence shell in SO4(2-), exceeding the octet by 4 electrons.

Q39. Explain how the octet expands in I3-.

```
Valence electrons around I = 2x(single bond electrons) + 2x(lone pairs) = 2(2) + 2(3) = 10
```

The central iodine atom in I3- has 10 electrons in its valence shell, so the octet expands by 2 electrons.

Q40. Define polyatomic ions with examples.

Polyatomic ions are ions composed of more than one type of atom. Their formal charge is the net charge calculated from the number of electrons in their valence shells after bond formation. Examples: carbonate (CO3 2-), sulfate (SO4 2-), nitrate (NO3-), ammonium (NH4+, positively charged).

Difference of Electronegativity

Q41. How can the nature of a bond be determined by electronegativity?

Electronegativity Difference	Type of Bond
Less than 0.4	Pure Covalent Bond
0.4 to 1.8	Polar Covalent Bond
Greater than 1.8	Ionic Bond

Bond Energy and Bond Length

Q42. How does bond length change with the type of bond?

```
Single Bond > Double Bond > Triple Bond (bond length order)
```

Bond length is the distance between the nuclei of two bonded atoms and decreases as the number of shared electron pairs increases. Example: C-C > C=C > C triple-bond C.

Q43. How are bond energy and bond length related to each other?

```
Bond Energy is proportional to 1 / Bond Length
```

There is an inverse relationship between bond energy and bond length. A shorter bond is generally stronger with higher bond energy; as bond length increases, nuclear-electron attraction decreases, weakening the bond.

VSEPR Theory

Q44. What is VSEPR theory? State its basic principle.

Valence Shell Electron Pair Repulsion (VSEPR) theory states that electron pairs around the central atom repel each other and arrange themselves as far apart as possible, which determines molecular shape by minimizing repulsion.

Q45. What is the shape of methane (CH4) and why?

Methane has a tetrahedral shape with bond angles of about 109.5 degrees, because the four bonding pairs of electrons repel each other equally and arrange themselves as far apart as possible around the central carbon.

Q46. Why is the shape of water (H2O) bent instead of linear?

Water has two lone pairs and two bonding pairs on the central oxygen atom. Lone pairs repel more strongly than bonding pairs, distorting the shape into a bent geometry with a bond angle of about 104.5 degrees.

Q47. Why does CO2 have a linear shape?

CO2 has two double bonds on opposite sides of the carbon atom and no lone pairs on the central atom. The repulsion between the two electron regions results in a linear arrangement with a bond angle of 180 degrees.

Hybridization

Q48. What is hybridization and why does it occur?

Hybridization is the mixing of atomic orbitals to form new hybrid orbitals suitable for bonding; it explains molecular shapes and bond angles better than unhybridized orbitals. Example: in CH4, carbon uses sp3 hybridization to form four equivalent bonds (sigma bonds + lone pairs = 4).

Q49. What is the geometry and hybridization of NH3?

In NH3, nitrogen is sp3 hybridized, having three sigma bonds and one lone pair. The molecule has a trigonal pyramidal geometry with a bond angle of about 107 degrees, slightly less than tetrahedral due to lone pair repulsion.

Molecular Orbital Theory (MOT)

Q50. Why does the He2 molecule not exist?

The electron configuration of He is 1s2. To form He2, the 1s orbitals of two He atoms combine to form bonding (sigma1s) and antibonding (sigma*1s) orbitals. Of the four electrons, two occupy the bonding orbital and two occupy the antibonding orbital.

```
Bond Order = (2 - 2) / 2 = 0
```

Hence, the He2 molecule is not formed, since its bond order is zero.

Q51. Define Bond order.

The number of bonds formed between two atoms after atomic orbital overlap is called bond order; it is half the difference between the number of bonding electrons and antibonding electrons.

```
Bond Order = (electrons in B.M.O - electrons in A.B.M.O) / 2

For H2: Bond Order = (2 - 0) / 2 = 1
```

Summary of Periodic Trends

Concept	Description
Ionic Bond	Complete transfer of electrons, formed between metals and non-metals
Covalent Bond	Mutual sharing of electrons, formed between non-metals
Dative Bond	Both electrons donated by one atom
Metallic Bond	Attraction between metal ions and delocalized electrons
Sigma Bond	Head-to-head overlap, stronger
Pi Bond	Sideways overlap, weaker
VSEPR Theory	Electron pairs repel to minimize repulsion
Hybridization	Mixing of orbitals to form equivalent orbitals
Bond Order	Number of bonds between two atoms
Dipole Moment	Measure of molecular polarity

SECTION 3

Long Questions

Detailed explanations, Quick Checks & summary tables

Q1. Define chemical bond. Explain types of bonds according to the Lewis concept.

A chemical bond is the force that holds together two or more atoms, molecules or ions. The properties of a substance depend on the type of chemical bond between its atoms. The term "chemical bond" includes intra-molecular forces: 1) Ionic bond 2) Covalent bond 3) Dative bond 4) Metallic bond.

Lewis Concept - atoms make bonds to complete their outermost shells and attain a noble-gas-like configuration, mostly through formation of an octet in the valence shell.

1. Ionic Bond (Electrovalent Bond)

The ionic bond is formed by complete transfer of electrons from an atom with low ionization energy to another atom with high electron affinity.

Atom	Configuration	After Bond Formation
Na (11)	2, 8, 1	Na+ (2, 8) - like Ne (10)
Cl (17)	2, 8, 7	Cl- (2, 8, 8) - like Ar (18)

The oppositely charged Na+ and Cl- ions are held together by a strong ionic bond in the NaCl crystal. Note: a similar bond type is expected between Group 1/2 metals and Group 16/17 non-metals.

2. Covalent Bond

A covalent bond is formed by the mutual sharing of electrons between two atoms; by sharing electrons, each atom completes its valence shell and attains the nearest noble gas configuration. There are two types: Non-Polar Covalent Bond and Polar Covalent Bond.

(i) Non-Polar Bond - formed due to equal sharing of electrons between two atoms of the same electronegativity, e.g. Cl2, H2, F2, Br2, I2; the bonded atoms remain electrically neutral.

(ii) Polar Covalent Bond - formed due to unequal sharing of electrons between two atoms, e.g. H-Cl, H-Br, H2O.

3. Dative Bond (Co-ordinate Covalent Bond)

A dative bond is formed between two atoms when the shared pair of electrons is donated by only one of the bonding atoms.

(i) Hydronium Ion (H3O+) - H2O has two covalent bonds and two lone pairs on oxygen. The proton (H+) is electron-deficient. Oxygen donates one of its lone pairs to H+, forming a dative bond.

(ii) Carbon Monoxide (CO) - oxygen shares two electrons with carbon to make a normal double covalent bond; carbon still needs one more pair to complete its octet, so oxygen donates one of its lone pairs to form a dative bond.

(iii) Ozone (O3) - formed by reaction between an oxygen molecule and an oxygen atom (free radical). One oxygen atom donates a pair of electrons to make a coordinate covalent bond; the central atom ends up with a total of three bonds (double covalent plus coordinate covalent).

Quick Check 3.1

(a) Draw the Lewis structures of N2 and CS2 molecules.

N2: N triple-bonded to N (N#N). CS2: S=C=S, linear structure with two C=S double bonds.

(b) How many electrons are in the valence shell of B in BF3? Can it accept a lone pair of electrons?

There are six electrons in the valence shell of B in BF3. Yes, it can accept a lone pair of electrons because boron is electron-deficient and can accept a lone pair to complete its octet.

Expanded Octet

(i) Expanded Octets in Sulfur - sulfur can expand its octet because it has empty 3d orbitals available for bonding.

```
Valence electrons around I in I3- = 2x(single bond electrons) + 2x(lone pairs) = 2(2) + 2(3) = 10
```

(ii) Expanded Octets in I3- Ion - the central iodine atom in I3- has 10 electrons in the valence shell, so the octet expands by two electrons.

Quick Check 3.2

(a) Can elements of period 2 have an expanded octet? Explain why or why not.

No, period 2 elements cannot have an expanded octet, because their valence shell contains only 2s and 2p orbitals, which can hold a maximum of eight electrons; the 3d orbitals needed for expansion are not present.

(b) Predict and explain the expanded octets in PO4(3-) and ClO3-.

For PO4(3-): the central P atom has 10 electrons, so the octet expands by two electrons. For ClO3-: the central Cl atom has 10 electrons, so the octet expands by two electrons.

Q3. How can the type of bond be determined by electronegativity?

The difference in electronegativity between two bonded atoms gives an approximate measure of bond polarity and indicates the type of bond formed.

Electronegativity Difference	Type of Bond
Less than 0.4	Pure Covalent (Non-polar)
0.4 to 1.8	Polar Covalent
Above 1.8	Ionic

Quick Check 3.3

(a) Predict whether the bonds in HF, KI, CaF2, ICl and Br2 are non-polar covalent, polar covalent, or ionic.

TABLE: Compound | Electronegativity Difference | Nature of Bond

HF | 1.8 | Highly Polar

KI | 1.7 | Highly Polar, Ionic

CaF2 | 3.0 | Ionic

ICl | 0.5 | Polar

Br2 | 0 | Non-Polar

ENDTABLE

Q4. Define dipole moment. Explain the dipole moment of H2O, SO2, H2S, BeCl2, CCl4 and BF3.

Dipole moment is a quantitative measurement of the polarity of a bond or a molecule. A molecule with a delta+ charge on one part and a delta- charge on the other is called a dipole and is said to have a dipole moment. For diatomic molecules like HF, HCl, HBr, HI and NO, the dipole moment is directed from the positive end (delta+) to the negative end (delta-). It is measured in Debye (D); a higher dipole moment indicates greater polarity.

Molecule	Dipole Moment (D)	Shape	Polarity
H2O	1.85	V-shaped (Angular)	Polar
SO2	1.61	V-shaped (Angular)	Polar
H2S	Small	Non-linear	Slightly Polar
BeCl2	0	Linear	Non-polar
CCl4	0	Tetrahedral	Non-polar
BF3	0	Trigonal Planar	Non-polar

H2O - dipole moment 1.85 D; the molecule is non-linear (V-shaped) - a linear H-O-H would have zero dipole moment, so the non-zero value proves the V-shape. SO2 - dipole moment 1.61 D, oriented oppositely compared to water. H2S - non-linear; individual bonds are polar but do not cancel each other.

BeCl2 - linear molecule (180 degrees); individual bond moments cancel out, net dipole moment = 0. CCl4 - perfectly tetrahedral; four C-Cl bond moments cancel each other, net dipole moment = 0. BF3 - trigonal planar symmetrical structure, dipole moment = 0.

Quick Check 3.4

(a) Why does CO have a dipole moment but CO2 does not?

CO has a dipole moment of 0.122 D - carbon is partially positive and oxygen partially negative due to the electronegativity difference. CO2 has no dipole moment - its linear structure has equal and opposite C=O dipoles that cancel each other exactly.

(b) Are the individual bonds in CCl4 polar? What about the overall molecule?

CCl4 has four C-Cl bonds that are individually polar due to the large electronegativity difference between C and Cl. However, the bond moments are directed such that they cancel each other in the perfectly tetrahedral geometry, so the net dipole moment of the CCl4 molecule is zero, making it non-polar overall.

(c) Are HF, CH2Cl2, O2 and H2S polar or non-polar?

TABLE: Molecule | Polarity | Reason

HF | Polar | Electronegativity difference between H and F

CH2Cl2 | Polar | Tetrahedral geometry, dipoles don't cancel

O2 | Non-polar | Homoatomic molecule, no electronegativity difference

H2S | Polar | Bent shape and electronegativity difference

ENDTABLE

Q5. Define intermolecular forces. Give their types.

The forces of attraction between two molecules in a compound are called intermolecular forces. They are binding forces like chemical bonds but much weaker; they exist between all atoms and molecules when sufficiently close, and are collectively called van der Waals forces.

Types of Intermolecular Forces: 1) Instantaneous dipole-induced dipole forces (London dispersion forces) 2) Permanent dipole-permanent dipole forces 3) Hydrogen bonding.

Q6. Define bond energy and bond length. Give their relation with electronegativity.

Bond Energy - the average amount of energy required to break all bonds of a particular type in one mole of a substance, measured in kJ/mol. Factors affecting it: 1) electronegativity difference of bonded atoms 2) size of atoms.

Single Bond	kJ/mol	Multiple Bond	kJ/mol
H-H	432	C=C	614
H-F	565	C triple-C	839
H-Cl	427	O=O	495
H-Br	363	C=O	745
H-I	295	N triple-N	941
C-C	347	C=N	615
C-N	305	N=O	607
C-O	358	-	-
N-H	391	-	-
O-H	467	-	-

Key points - bond energy rises with increasing electronegativity difference (e.g. HF 565 kJ/mol > HI 295 kJ/mol); multiple bonds have greater energy than single bonds: C triple-C > C=C > C-C.

Bond Length - the distance between the nuclei of two atoms forming a covalent bond, measured by electron diffraction, X-ray diffraction, or spectral studies. Factors: 1) electronegativity 2) size of atoms 3) nature of the bond (single/double/triple).

Bond	Bond Length (pm)	Bond	Bond Length (pm)
H-H	74	Si-F	155
H-Br	144	C-F	135
C-C	154	C-Cl	180
C=C	133	C-Br	196
C triple-C	120	C-I	214
C=O	122	Si-H	146
B-F	130	-	-

Key points - electronegativity causes shortening of bonds; bond length increases with increasing atomic size (e.g. C-Cl 180 pm > C-F 135 pm because Cl is larger).

Quick Check 3.5

(a) HI is a stronger acid and reducing agent than HF, which is weaker. Explain.

HI is a stronger acid because the H-I bond is weaker and releases hydrogen more easily than HF. The H-F bond is stronger due to the greater electronegativity difference, making HF a weak acid.

(b) Why is acetylene (HC triple-bond CH) more stable than ethene (H2C=CH2)?

Acetylene has a triple bond in which three bond pairs are shared, giving stronger nuclear attraction and a shorter, more stable bond. Ethene's double bond shares only two bond pairs, giving weaker nuclear attraction, so it is less stable than acetylene's triple bond.

Q7. Give a comparison among ionic, covalent, metallic bonds and intermolecular forces.

Bond Type	Bond Energy (kJ/mol)
Ionic Bond in NaCl	760
O-H bond in water	464
Hydrogen Bonding	20-50
Permanent Dipole-Permanent Dipole Force	5-20
Van der Waals Forces	1-20

Type	Strength	Examples
Ionic Bond	Strongest (760 kJ/mol)	NaCl
Covalent Bond	Strong (464 kJ/mol)	O-H bond in water
Metallic Bond	Medium (100-150 kJ/mol)	Metals
Hydrogen Bond	Weak (20-25 kJ/mol)	H2O, HF, NH3
Dipole-Dipole Forces	Very Weak (5-20 kJ/mol)	Polar molecules
London Dispersion Forces	Weakest (1-20 kJ/mol)	All molecules

Q8. Explain the Valence Bond Theory (VBT) and the bonding in H2, HCl, Cl2, H2S, O2, HF and N2.

VSEPR predicts molecular shapes but does not explain why bonds form; VBT explains both bond formation and molecular shape.

Postulates of VBT - 1) Formation of Covalent Bond: a covalent bond forms when half-filled orbitals of similar energy on two atoms overlap. 2) Strength of Bond: greater orbital overlap gives a stronger bond. 3) Direction of Covalent Bond: covalent bonds are directional, determined by the shape and mode of overlap.

Sigma Bond Formation

A sigma bond forms by the overlap of two half-filled atomic orbitals along the internuclear axis.

(i) s-s Overlap - H2: each H atom has a half-filled s orbital; after bonding, electrons pair up and electron density is symmetrical around the two nuclei.

(ii) s-p Overlap - HCl: the half-filled s orbital of H overlaps with the half-filled p orbital of Cl; electron density is higher near Cl due to its higher electronegativity, making HCl polar.

(iii) p-p Overlap - Cl2: p orbitals of two Cl atoms overlap along the nuclear axis; electron density is symmetrical since both atoms have equal electronegativity.

(iv) H2S - two 3p orbitals of sulfur, each with one electron, overlap with the 1s orbitals of two hydrogen atoms, giving a V-shaped molecule with a bond angle of about 92 degrees.

Pi Bond Formation

(v) O2 - each oxygen atom has two unpaired electrons in perpendicular p orbitals; the px orbitals overlap end-to-end to form a sigma bond, while the py orbitals overlap sideways to form a pi bond, giving a double bond (one sigma, one pi).

(vi) HF - the half-filled 1s orbital of H overlaps with the half-filled p orbital of F, forming a sigma bond.

(vii) N2 - one half-filled 2px orbital overlaps linearly to form a sigma bond, while the two half-filled 2py and 2pz orbitals undergo sideways overlap to form two pi bonds, giving a triple bond (one sigma, two pi).

Quick Check 3.6

(a) Explain orbital overlap in H2O and N2.

H2O - two half-filled s orbitals of hydrogen overlap with two half-filled p orbitals (py, pz) of oxygen, forming a V-shaped molecule. N2 - three unpaired electrons on each nitrogen atom in perpendicular p orbitals; the px orbitals overlap end-to-end (sigma bond) while py and pz overlap sideways (two pi bonds), giving a triple bond.

(b) Draw the orbital overlap for F2 and HF.

F2 - p orbitals of two fluorine atoms overlap to form a sigma bond, with electron density symmetrical around both nuclei. HF - the half-filled s orbital of hydrogen overlaps with the half-filled p orbital of fluorine, with electron density higher near fluorine.

Q9. Define hybridization. Explain the structure of BeCl2 with reference to sp hybridization.

Hybridization is a process in which atomic orbitals of slightly different energies and shapes mix together to form a new set of equivalent orbitals of the same energy and shape.

Basic rules: 1) hybridization mixes orbitals within a single atom (or ion) 2) only orbitals of comparable energy can mix 3) the number of mixing orbitals equals the number of resulting hybrid orbitals.

Types of hybridization: sp (linear, 180 degrees), sp2 (trigonal planar, 120 degrees), sp3 (tetrahedral, 109.5 degrees).

sp Hybridization - BeCl2

In sp hybridization, one s and one p orbital intermix to give two hybrid orbitals of the same energy and shape, arranged linearly (180 degrees). Examples: BeCl2, HC triple-bond CH, MgCl2.

BeCl2 - ground state of Be is 1s2 2s2 (no unpaired electrons); in the excited state it becomes 1s2 2s1 2p1 (two unpaired electrons); the hybridized state is 1s2 (sp)1 (sp)1, giving two sp hybrid orbitals lying linearly at 180 degrees. BeCl2 forms when these two sp orbitals of Be overlap with the half-filled p orbitals of two chlorine atoms.

Q10. Explain the BF3 molecule with reference to sp2 hybridization.

In sp2 hybridization, one s and two p orbitals intermix to form three sp2 hybrid orbitals arranged in a trigonal planar geometry (120 degrees). Examples: BF3, CH2=CH2, BH3, BCl3.

BF3 - boron's configuration is 1s2 2s2 2p1; in the excited state it becomes 1s2 2s1 2px1 2py1; the sp2 hybridized state is 1s2 (sp2)1 (sp2)1 (sp2)1, giving three sp2 hybrid orbitals in trigonal planar geometry. Each sp2 orbital overlaps with a 2p orbital of fluorine, giving a triangular planar structure.

Q11. Explain sp3 hybridization with the help of an example.

In sp3 hybridization, one s and three p orbitals intermix to give four hybrid orbitals of the same energy and shape, arranged tetrahedrally (109.5 degrees). Examples: CH4, NH3, H2O, NH4+.

Methane (CH4) - carbon's ground state is 1s2 2s2 2px1 2py1; in the excited state it becomes 1s2 2s1 2px1 2py1 2pz1; the sp3 hybridized state is 1s2 (sp3)1 (sp3)1 (sp3)1 (sp3)1, giving four sp3 orbitals in tetrahedral geometry. Each sp3 orbital overlaps with a 1s orbital of hydrogen, giving four C-H bonds directed towards the corners of a regular tetrahedron. Note: each sp3 hybrid orbital has two lobes, one larger and one smaller; the smaller lobe is usually omitted for simplicity.

Q12. What are the postulates of the VSEPR model? How are molecular shapes predicted?

The VSEPR model describes molecular shapes based on the electron pairs surrounding the central atom. It was presented by Sidgwick and Powell (1940). Basic assumption: a molecule adopts a shape that minimizes electronic repulsion among valence electron pairs, arranging them at the farthest possible distances.

Postulates - 1) both lone pairs and bond pairs determine molecular geometry 2) electron pairs arrange to remain at maximum distance apart 3) repulsion order: lone pair-lone pair > lone pair-bond pair > bond pair-bond pair 4) double and triple bond electron pairs count as a single pair 5) highly electronegative atoms decrease bond pair repulsion but increase lone pair repulsion.

Steps to determine molecular shape: 1) select the least electronegative atom (or the one with fewest atoms) as the central atom 2) count total electron pairs around the central atom 3) total electron pairs determine the electron pair geometry 4) the actual shape excludes lone pairs.

```
NH3 example: Valence e- of N = 5; electrons shared by 3H = 3; total = 8; total pairs = 4; lone pairs = 4 - 3(bond pairs) = 1
```

No. of Electron Pairs	No. of Lone Pairs	Electron Pair Geometry	Shape	Bond Angle	Examples
2	0	Linear	Linear	180	CO2, BeCl2
3	0	Trigonal	Trigonal	120	BCl3, SO3
3	1	Trigonal	Angular (V-shaped)	less than 120	SO2, SnCl2
4	0	Tetrahedral	Tetrahedral	109.5	CH4, CCl4
4	1	Tetrahedral	Pyramidal	less than 109.5	NH3, H3O+
4	2	Tetrahedral	Angular (V-shaped)	less than 109.5	H2O, OF2
5	0	Trigonal Bipyramidal	Trigonal Bipyramidal	120, 90	PCl5
6	0	Octahedral	Octahedral	90	SF6, XeF4

Q13. Explain AB2 and AB3 molecules with reference to VSEPR.

AB2 Type (Linear Geometry) - two electron pairs arranged at 180 degrees around the central atom. Example: BeCl2, central Be has two bond pairs and no lone pair, giving a linear shape.

AB3 Type (Trigonal Planar Geometry) - three electron pairs arranged at 120 degrees for maximum separation. Example: BF3 has a trigonal planar shape with F-B-F bond angle = 120 degrees. Example: SnCl2 - Sn has 4 valence electrons, forms 2 bonds with Cl and keeps the remaining 2 electrons as a lone pair, giving a V-shaped (angular) molecule with a bond angle less than 120 degrees.

AB3 with multiple bonds - SO3: all three regions are occupied by S=O bonds, giving a perfectly trigonal shape with each angle = 120 degrees. SO2: one corner is occupied by a lone pair and two corners by S=O double bonds, giving a V-shaped (angular) structure.

Q14. Explain AB4 molecules with reference to VSEPR theory.

AB4 Type (Tetrahedral Geometry) - four electron pairs arranged in a regular tetrahedron, bond angle = 109.5 degrees.

(i) CH4 (Methane) - four bonded electron pairs give a regular tetrahedral shape, H-C-H bond angle = 109.5 degrees. (ii) NH3 (Ammonia) - three H atoms plus one lone pair reduce the bond angle from 109.5 to 107 degrees, giving a triangular pyramidal shape.

(iii) NF3 (Nitrogen Trifluoride) - three F atoms bonded to N further reduce the bond angle to 102.5 degrees due to the high electronegativity of F. (iv) H2O (Water) - two lone pairs and two bond pairs reduce the bond angle to 104.5 degrees, giving a V-shaped (angular) structure. (v) H3O+ is tetrahedral/trigonal pyramidal in shape, while NH2- is V-shaped.

Quick Check 3.7

(a) Calculate the number of lone pairs on the central atoms in H2O, ICl3 and I3-.

H2O - central oxygen has two lone pairs. ICl3 - central iodine has two lone pairs. I3- - central iodine has three lone pairs.

(b) Calculate the number of bond pairs and lone pairs in SiH4, H2Se and PH3.

TABLE: Molecule | Bond Pairs | Lone Pairs

SiH4 | 4 | 0

H2Se | 2 | 2

PH3 | 3 | 1

ENDTABLE

(c) Predict the shapes and bond angles in SiH4, H2Se and PH3.

TABLE: Molecule | Shape | Bond Angle

SiH4 | Tetrahedral | 109.5 degrees

H2Se | Bent (V-shaped) | 91 degrees

PH3 | Trigonal Pyramidal | 93.5 degrees

ENDTABLE

(d) How does the bond angle in H2S differ from that in H2O?

The bond angle in H2O (104.5 degrees) is larger than in H2S (92 degrees) because oxygen's greater electronegativity pulls the bonding electrons closer to it than in H2S.

Q15. Explain AB5 and AB6 molecules with reference to VSEPR theory.

AB5 Type (Trigonal Bipyramidal Geometry) - five electron pairs around the central atom. PCl5: trigonal bipyramidal geometry, angles within the plane = 120 degrees, angle between axial and planar atoms = 90 degrees. I3- (tri-iodide ion): an AB5 system with two bond pairs and three lone pairs; the lone pairs occupy the trigonal-region corners, giving a linear shape with I atoms at 180 degrees.

AB6 Type (Octahedral Geometry) - six electron pairs around the central atom. SF6: octahedral geometry, all F-S-F angles = 90 degrees. XeF4: an AB6 type with four bond pairs and two lone pairs; the lone pairs sit at 180 degrees to each other, giving a square planar shape with all F-Xe-F angles = 90 degrees.

Quick Check 3.8

(a) Draw the Lewis structures for IF3 and IF5 and predict their VSEPR geometry.

IF3 - trigonal bipyramidal electron geometry with three bonding pairs and two lone pairs. IF5 - square pyramidal geometry with five bond pairs and one lone pair.

Q16. Explain the application of the VSEPR model in drug designing.

A drug is a chemical substance that prevents, diagnoses, or treats a disease. Key concepts: Ligands are bioactive molecules that interact with targets; a Receptor is usually a protein that receives and responds to a ligand; Targets include enzymes, proteins, nucleic acids and cellular pathways; Specificity means each drug interacts with a specific target to a different degree.

Molecular shape determines how a drug interacts with its biological target - only ligands with a suitable shape can fit into active sites, much like a specific key fits a specific lock. The VSEPR model helps determine the shapes of such biological systems. Example: Aspirin, an analgesic (painkiller), interacts with the enzyme COX (cyclooxygenase), binding to its active site through an acetyl group; shape compatibility is crucial for successful binding.

Q17. Give the postulates of Molecular Orbital Theory. Explain sigma and pi bond formation by MOT.

Postulates of MOT - 1) Overlapping of Orbitals: atomic orbitals of combining atoms overlap to form molecular orbitals characteristic of the whole molecule. 2) Polarization of Orbitals: atomic orbitals overlap with lobes of matching wave-function sign. 3) Types of MOs: Bonding MO (sigma or pi) has lower energy than the parent atomic orbitals (same-sign overlap); Anti-bonding MO (sigma or pi) has higher energy (opposite-sign overlap).

```
Bond Order = (electrons in B.M.O - electrons in A.B.M.O) / 2
```

s-s Overlap (H2 Molecule) - two 1s orbitals combine to give sigma1s (bonding MO, symmetrical about the nuclear axis, lower energy) and sigma*1s (anti-bonding MO, electron density away from the nuclei, higher energy).

```
Bond Order for H2 = (2 - 0) / 2 = 1
```

p-p Overlap - (a) Head-on approach: p orbitals along the same axis form sigma(2px) bonding and sigma*(2px) antibonding orbitals. (b) Sideways approach: parallel p orbitals (py or pz) form pi(2py)/pi(2pz) bonding and pi*(2py)/pi*(2pz) antibonding orbitals, with zero electron density (a nodal plane) on the nuclear axis.

Q18. Draw the molecular orbital diagrams of H2, He2, N2 and O2. Calculate their bond orders.

(i) H2 - electronic configuration: 2H(1s1) gives H2(sigma1s2).

`Bond Order = (2 - 0) / 2 = 1 → H2 has a single bond`

(ii) He2 - He configuration is 1s2; molecular orbital configuration is (sigma1s)2 (sigma*1s)2.

`Bond Order = (2 - 2) / 2 = 0 → He2 molecule is not formed`

(iii) N2 - electronic configuration of N: 1s2 2s2 2px1 2py1 2pz1. MO order: sigma2s < sigma*2s < pi2py = pi2pz < sigma2px < pi*2py = pi*2pz < sigma*2px.

`Bond Order = (6 - 0) / 2 = 3 → Triple bond between N atoms (one sigma, two pi)`

(iv) O2 - electronic configuration of O: 1s2 2s2 2p4. MO order: sigma2s < sigma*2s < sigma2px < pi2py = pi2pz < pi*2py = pi*2pz < sigma*2px.

`Bond Order = (6 - 2) / 2 = 2 → Double bond between O atoms (one sigma, one pi)`

O2 has two unpaired electrons in pi(2py) and pi(2pz), which explains its paramagnetic behaviour.

Q19. Discuss the formation of F2 in the light of the Lewis concept, VBT and MOT.

(i) Lewis Structure of F2 - each F atom has three lone pairs and shares one pair of electrons to form a single F-F covalent bond.

(ii) Structure of F2 by VBT - p orbitals of two fluorine atoms overlap along the nuclear axis; electron density is symmetrical around both nuclei since both atoms have the same electronegativity.

(iii) Formation of F2 according to MOT - electronic configuration of F: 1s2 2s2 2p5. The molecular orbital diagram gives the order: sigma2s, sigma*2s, sigma2px, pi2py = pi2pz, pi*2py = pi*2pz, sigma*2px.

`Bond Order = (6 - 4) / 2 = 1 → F2 has a single covalent bond`

Summary of Periodic Trends

Property	Ionic Bond	Covalent Bond	Metallic Bond
Formation	Transfer of electrons	Sharing of electrons	Sea of electrons
Types of Atoms	Metal + Non-metal	Non-metal + Non-metal	Metal + Metal
Electronegativity Difference	Greater than 1.8	0 - 1.8	Small
Strength	Strongest	Strong	Medium
State at Room Temp	Solid	Gas, Liquid, Solid	Solid
Conductivity	Only in molten/aqueous	Poor	Good