

## CHEMICAL BONDING

1. What is "Octet rule"? Give two examples of the compounds which do not obey Octet rule.

**Octet Rule:** the tendency of atom to attain the maximum of eight electrons in the valence shell is called octet rule. Many compounds do not obey this rule. For example: BF<sub>3</sub>, SF<sub>6</sub>, PCl<sub>5</sub> etc

2. Bond distance is the compromise distance between two atoms.

When the atoms approaches to each other to form a bond, they attract each other and the energy of the system is lowered. When they reach to certain distance their forces of attraction are maximum. On further coming close, they start repelling each other and thus the energy of the system increases. Now the two atoms try to remain at compromising distance whereas the energy of the system is minimum. This bond distance is known as compromising distance between two atoms.

## Atomic Size

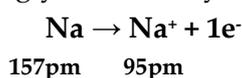
3. Why the radius of Na<sup>+</sup> ion is less than the atomic radius of Na atom?

OR

Why cationic radius is smaller than parent atom?

A cation is formed by the loss of one or more electron from a neutral atom. The size of cation is smaller than its parent atom because:

- By the loss of electron, shell may loss.
- Loss electron may result in imbalance in proton electron ratio. Hence electrons are strongly attracted by nucleus.



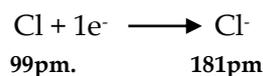
4. The size chlorine atom is smaller than Cl<sup>-</sup> ion. Justify it.

OR

Why the ionic radius of anion increases from 99 pm to 181 pm?

The ionic radii of negative ions are larger than size of parent atom due to following reason

- Repulsion between valence shell electrons and incoming electron. As a result, valence shell expands to minimize repulsion.
- Imbalance of proton-electron ratio. As a result, hold of nucleus on valence electron decreases hence decreases. For example



5. **Why an anionic radius is greater than atomic radius?**

The radius of the positive ion is less than anion because the effective nuclear charge on the ion increases. Greater the positive charge on the ion, smaller the radius. Similarly, the radius of anion becomes more than the. Neutral atom because increasing electron makes the outermost shell to expand.

6. **Define ionic and covalent radii.**

**Ionic Radii:** The radius of an ion when it is supposed to be spherical is called ionic radius. e.g radius of Cl<sup>-</sup> ion = 181pm

**Covalent Radius:** the half of the distance between centers of two covalently bonded same atoms is called covalent radius. e.g covalent radii of carbon atom is 77pm.

**Ionization Energy**

7. **Why is the second ionization energy greater than first one?**



When second electron is removed from the unipositive ion, the nuclear attraction on valence electrons increases. As a result, valence electrons tightly bounded to the nucleus so greater amount of energy is required to remove second electron and value of second ionization energy is greater than first ionization energy.

8. **How does ionization energy vary in a group of periodic table?**

**Along Group:** Ionization energy decreases down the group due to

- Increase in atomic radius
- Decrease in nuclear attraction
- Increase in shielding effect

As a result, removal of electron becomes easier hence ionization energy decreases.

9. **How does ionization energy vary in periodic table?**

**Along Group:** Ionization energy decreases down the group due to

- Increase in atomic radius
- Decrease in nuclear attraction
- Increase in shielding effect

As a result, removal of electron becomes easier hence ionization energy decreases.

**Along Periods:** Ionization energy increases across the period due to

- Decrease in atomic radius

- Increase in nuclear attraction
- Shielding effect remains same

As a result, removal of electron becomes difficult hence ionization energy increases.

### Electron Affinity

10. Define electron affinity. And give an example.

**Electron affinity:** The amount of energy released or absorbed when an electron is added in the valence shell of an isolated gaseous atom to form uni-negative ion is called electron affinity.



### Electronegativity

11. Define electronegativity. Give its trend in the periodic table.

**Electronegativity:** The tendency of an atom to attract the shared pair of electrons towards itself is called electronegativity.

#### Periodic Trends:

**Along Group:** Electronegativity decreases down the group due to

- Increase in atomic radius
- Decrease in nuclear attraction
- Increase in shielding effect

**Along Periods:** Electronegativity increases across the period due to

- Decrease in atomic radius
- Increase in nuclear attraction
- Shielding effect remains same

### Ionic Bond

12. How the nature of a chemical bond is predicted with the help of electronegativity values of two bonded atoms?

When the electronegativity difference between two bonded atoms is 1.7 or more than that, then the **ionic bond** is formed.

If the electronegativity difference between two bonded atoms is ranges from 0.7 -1.7, then **polar covalent bond** is formed.

If the electronegativity difference between two bonded atoms is less than 0.7, then **non-polar covalent bond** is formed.

**13. No bond in chemistry is 100% ionic. Justify it.**

No bond is 100% ionic because the maximum difference of electro negativity between two bonded atoms is 3.3 which is for CsF and its percentage ionic character is 92%. It means that the maximum value for percentage ionic character is 92% so the above statement is justified.

**Coordinate Covalent Bond**

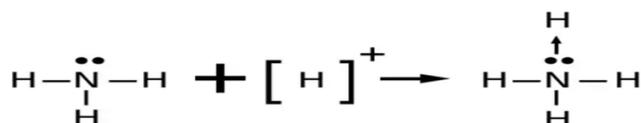
**14. How a coordinate covalent bond differs from covalent bond?**

<b>Covalent Bond:</b> The bond which is formed by mutual sharing of electron pairs between two bonded atoms is called covalent bond.	<b>Coordinate Covalent Bond:</b> The covalent bond in which shared electron pair is donated by only one bonded electron is called coordinate covalent bond.
In covalent bond, each bonded atom shares one electron.	In coordinate covalent bond, the shared electron pair is donated by only one atom other atom does not donate electron.
$\text{H} \cdot \cdot \text{Cl} : \quad \text{H} - \text{Cl} :$	$\text{H} - \overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{N}}} - \text{H} + [\text{H}]^+ \rightarrow \text{H} - \overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{N}}} - \text{H}$

**15. Define co-ordinate covalent bond and give an example.**

A covalent bond formed by the donation of an electron pair by one of the bonded atom to other is called co-ordinate covalent bond or Dative bond.

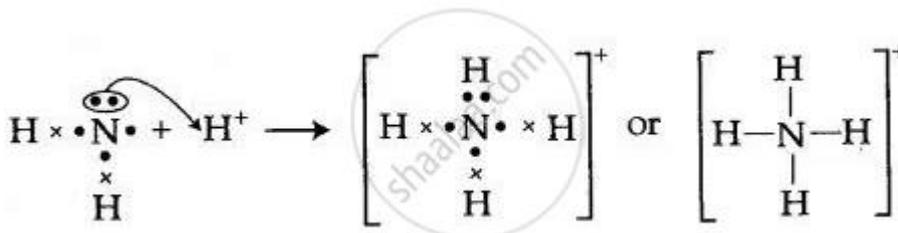
For example.



**16. Why the distinction between coordinate covalent bond and covalent bond vanishes after bond formation in  $\text{NH}_4^+$ .**

**Bond Formation in  $\text{NH}_4^+$ :**

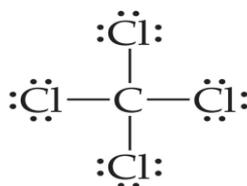
The coordinate covalent bond is formed with same species  $\text{H}^+$  as three of them already bonded. Every bond has 25% coordinate covalent bond character and 75% covalent bond character.



17. Write the Lewis structures for the following compounds:

(a) HCN

(b) CCl<sub>4</sub>



### VSEPR

18. Write two main postulates of VSEPR theory?

- Both bond pairs and lone pairs of electrons are involved in determining the geometry of molecules.
- The electron pairs are arranged at maximum distance apart from each other to minimize the repulsion.

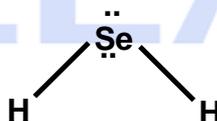
19. Explain the geometry of H<sub>2</sub>Se molecule.

H<sub>2</sub>Se is a type of AB<sub>2</sub> molecule, it has

Bond Pairs: 2

Lone Pairs: 2

Its molecular geometry is bent/ angular and its arrangement of atoms show that it has tetrahedral shape.

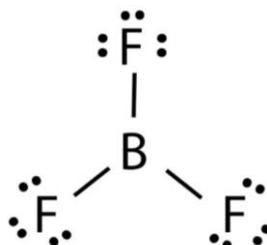


20. What is the basic assumption of VSEPR theory, and discuss the shape of a molecule (BF<sub>3</sub>) containing three electron pairs?

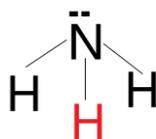
Basic Assumptions of VSEPR:

- Both bond pairs and lone pairs of electrons are involved in determining the geometry of molecules.
- The electron pairs are arranged at maximum distance apart from each other to minimize the repulsion.

Shape of BF<sub>3</sub>:



21. State the geometry of  $\text{NH}_3$  molecule on the basis of VSEPR theory.



In  $\text{NH}_3$ , N is surrounded by four electron pairs. Three are bond pair and one is lone pair. Due to presence of lone pair, tetrahedral geometry changes to trigonal pyramidal with bond angle  $107.5^\circ$

### VBT/Hybridization

22. How Sigma and pi bonds are formed.

Sigma bond is formed by linear overlapping of two half filled atomic orbitals and electron density is symmetrical along the bond axis. Whereas in case of pi-bond, the electron density is not symmetrical and it consists of two regions above and below the bond axis. That's why pi-bond is more diffused than sigma bond.

23. Explain the structure of  $\text{CH}_4$  on the basis of  $\text{sp}^3$  hybridization.

Sigma bond is formed by linear overlapping of two half filled atomic orbitals and electron density is symmetrical along the bond axis. Whereas in case of pi-bond, the electron density is not symmetrical and it consists of two regions above and below the bond axis. That's why pi-bond is more diffused than sigma bond.

24. The bond angles of  $\text{H}_2\text{O}$  and  $\text{NH}_3$  are not  $109.5^\circ$  like that of  $\text{CH}_4$ . Although O-atom and N-atom are  $\text{SP}^3$  hybridized like C-in  $\text{CH}_4$ , give reason.

Like  $\text{CH}_4$ , the molecule of  $\text{H}_2\text{O}$  and  $\text{NH}_3$  are also  $\text{AB}_4$  type molecule. Carbon, oxygen and nitrogen atoms undergo  $\text{sp}^3$  hybridization.  $\text{CH}_4$  is perfectly tetrahedral with bond angle  $109.5^\circ$ . in case of ammonia, there are three bond pairs and one lone pair. Lone pair – bond pair repulsion is greater than bond pair-bond pair repulsion. Due to this reason, angle reduces to  $107.5^\circ$ . in case of  $\text{H}_2\text{O}$ , there are two lone pairs on oxygen. Due to this increased repulsion of two lone pairs, the angle further reduces to  $104.5^\circ$ .

25. The bond angles of  $\text{H}_2\text{O}$  and  $\text{NH}_3$  are not  $109.5^\circ$  like that of  $\text{CH}_4$  although oxygen and nitrogen atoms are  $\text{SP}^3$  - hybridized. Why?

There are four electron pairs in both  $\text{NH}_3$  and  $\text{H}_2\text{O}$ .

In  $\text{NH}_3$ , there are three bonding electron pairs and one lone pair. Lone pair exerts a repulsion on bonding pair due to which bond angle reduces to  $107.5^\circ$ .

Whereas in  $\text{H}_2\text{O}$ , there are two bond pairs and two lone pairs on oxygen atom which greater repulsion and hence angle reduces from  $109.5^\circ$  to  $107.5^\circ$ .

**26.  $\pi$ -bonds are more diffused than  $\delta$ -bonds. Give reason.**

Sigma bond is formed by linear overlap of half-filled orbitals. In sigma bond, the electronic charge density lies along the inter-nuclear axis and strongly attracted by both nuclei.

Pi bond is formed by parallel overlap of half-filled atomic orbitals. In pi-bond, the electronic charge density lies above and below the nuclear axis and is weakly attracted by nuclei.

Less nuclear attraction causes a more diffusion of pi-bond than sigma bond.

**27. Sigma bond is stronger than pi-bond. Explain.**

In sigma bond, the electron density lies very close to the line joining the two nuclei. So nuclei are tightly bound. Electrons of the pi bond are above and below the joining line. They are more diffused (spread) so cannot hold the two nuclei with a greater force.

**28. Differentiate between sigma bond and pi bond.**

<b>Sigma Bond</b>	<b>Pi Bond</b>
The bond formed by linear or head to head overlapping of partially filled orbitals	The bond formed by parallel or side by side overlapping of partially filled orbitals
In sigma bond, electron density is present between two nuclei.	In pi bond, electron density is above and below the bond axis.
It is comparatively inert bond	It is comparatively reactive.
It is a strong bond	It is a weak bond
Atoms are always bonded with sigma bond first.	Pi bond is formed when sigma bond is already present in atoms.
It is represented by $\sigma$	It is represented by $\pi$

**Molecular Orbital Theory (MOT)**

**29. Differentiate between bonding and anti-bonding molecular orbitals?**

**Bonding Molecular Orbitals:** the atomic orbitals of bonding atoms overlap to form same number of molecular orbitals. The molecular orbitals which have lower energy than parent atomic orbitals are called bonding molecular orbitals.

**Anti-Bonding Molecular Orbitals:** The molecular orbitals which have high energy than parent atomic orbitals are called Anti-Bonding Molecular Orbitals.

**30. Why the energy of anti-bonding molecular orbital is higher than corresponding bonding molecular orbitals?**

In bonding molecular orbitals BMO, the electron density is maximum on the joining line of two nuclei, so the nucleus-nucleus repulsion is minimum and the system is stable. In anti-

bonding molecular orbitals ABMO, the electron density is not between the two nuclei but on right and left of bonded atoms.

In ABMO, the electronic clouds are on the right and left of two nuclei. In this way, the nucleus-nucleus repulsion is there and the system is unstable.

**31. Explain that Helium does not exist in diamagnetic state.**

The electronic configuration is

He:  $1s^2$

Each Helium atom contributes two electrons. Two electrons enter bonding molecular orbital  $\sigma(1s)$  orbitals and the remaining two electrons go to antibonding  $\sigma^*(1s)$  molecular orbitals.

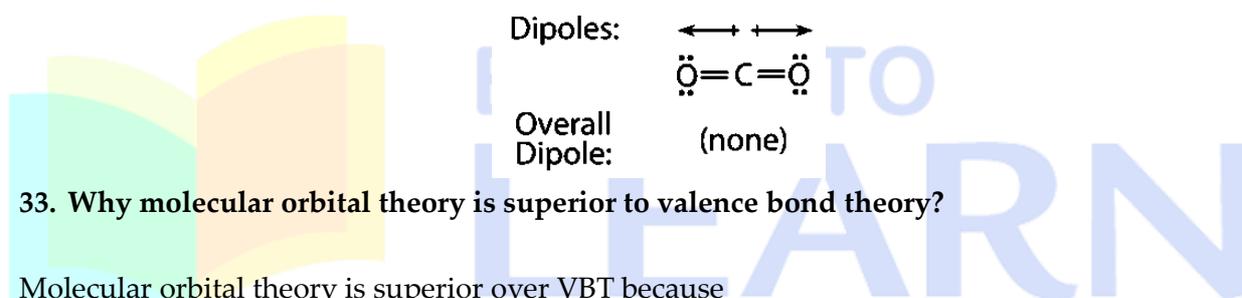
The bond order for  $He_2$  is zero

$$\frac{2-2}{2} = 0 \quad \text{Thus } He_2 \text{ molecule is not formed.}$$

**32. Define bond order. Calculate bond order of hydrogen molecule.**

**Bond Order:** the half of the difference between no. of electrons in bonding molecular orbital and anti-bonding molecular orbital is called bond order.

**$H_2$  Molecule:**



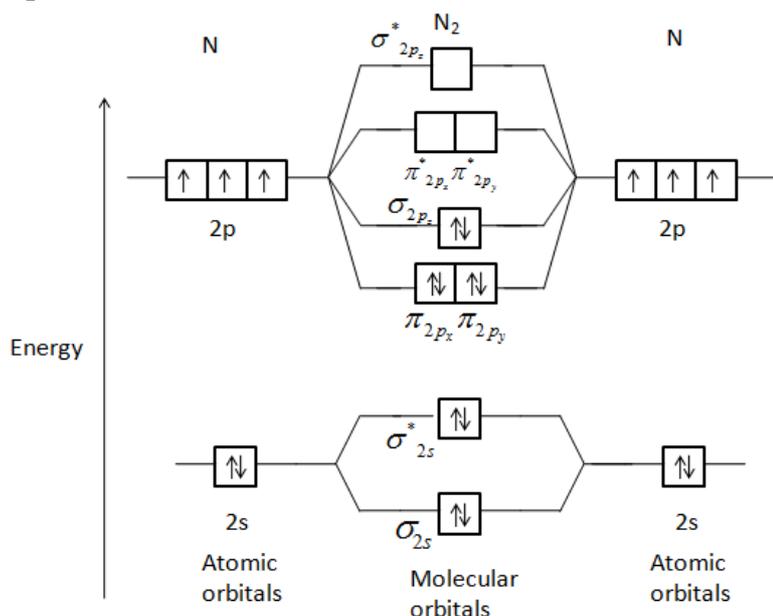
**33. Why molecular orbital theory is superior to valence bond theory?**

Molecular orbital theory is superior over VBT because

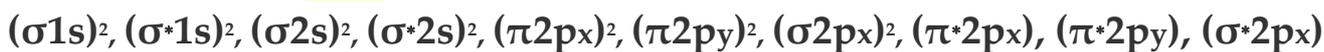
- It tells us the reason of no bond formation between noble gases.
- It tells us the number of bonds whether sigma or pi in  $N_2$  and  $O_2$ .
- It tells us about the paramagnetic and diamagnetic nature of a substance.

But VBT does not give such answers.

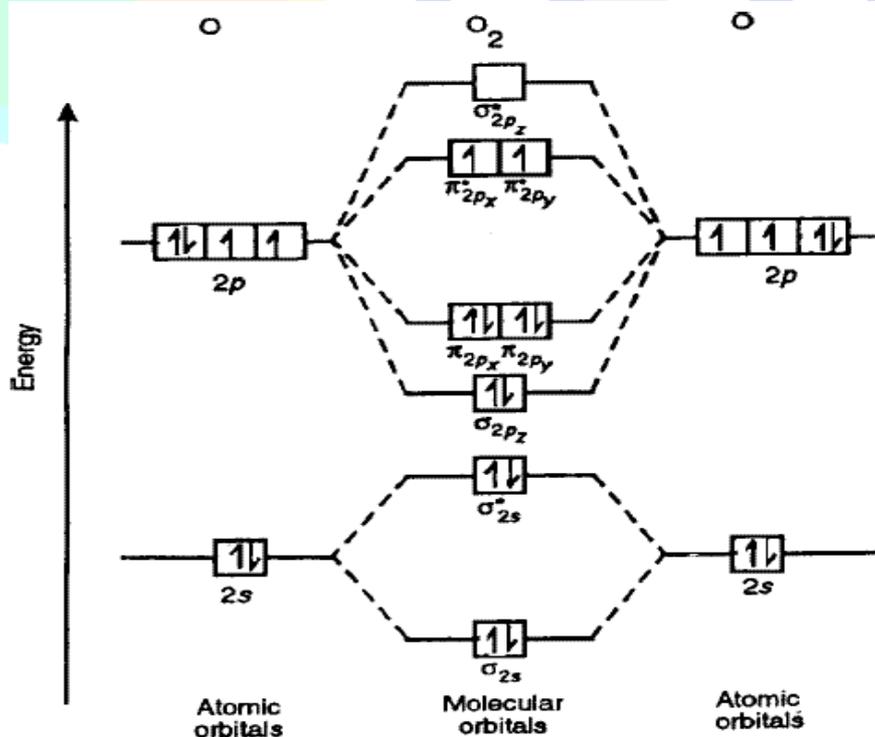
34. Represent the molecular orbitals of N<sub>2</sub> molecule in the increasing order of energy.



Relative Energy Level:



35. Sketch the molecular orbital picture of O<sub>2</sub>.



36. Differentiate between atomic orbital and molecular orbital.

**Atomic orbitals:** the region around the nucleus of an atom in which probability of finding the electrons is maximum is called atomic orbital.

It has no sharp boundaries and can have different shapes.

**Molecular orbitals:** the orbitals which have more than one nuclei present in its electronic cloud is called molecular orbitals.

### Bond Energy

**37. Write the names of factors affecting the strength of a bond?**

The following factors affect the strength of bond.

- Size of bonded atoms
- Electronegativity difference
- Bond length
- Bond polarity

**38. Define bond energy and give one example?**

**Bond Energy:** the average amount of energy required to break all the bonds of a particular type in one mole of a substance is called bond energy. **OR**

The average amount of energy required to break Avogadro's number of bonds of particular type is called bond energy.

Bond energy of C-C bond =  $348 \text{ KJmol}^{-1}$

**39. Why the abnormality of bond length and bond strength in HI is less. Why the abnormality of bond length and bond strength in HI is less prominent than that of HCl?**

**OR**

**Why the abnormality of bond strength in HI is less prominent than that of HCl.**

The abnormality of bond length and bond strength is shown below.

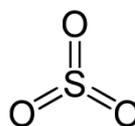
HI Molecule	HCl Molecule
Calculated Bond Energy = $291 \text{ KJ/mol}$	Calculated Bond Energy = $336 \text{ KJ/mol}$
Observed Bond Energy = $299 \text{ KJ/mol}$	Observed Bond Energy = $431 \text{ KJ/mol}$
Difference = $8 \text{ KJ/mol}$	Difference = $95 \text{ KJ/mol}$

The difference is due to the greater electro negativity of Chlorine as compared to Iodine. As a result, HCl has greater ionic character than HI due to which HCl has short bond length and greater bond energy.

## Dipole Moment

### 40. Why the dipole moment of SO<sub>2</sub> is 1.61 D but that of SO<sub>3</sub> is zero?

In SO<sub>2</sub> both S=O bonds are polar. SO<sub>2</sub> molecule has an angular structure. This structure is not symmetrical. In this case individual bond moments do not cancel each others. Therefore it has dipole moment. But in case of SO<sub>3</sub> this is a symmetrical molecule individual bond moment cancel each others so it has zero dipole moment.



### 41. The dipole moment of CO<sub>2</sub> and CS<sub>2</sub> are zero, but that of SO<sub>2</sub> is 1.61D. justify it.

CO<sub>2</sub> and CS<sub>2</sub> have linear structures which are symmetrical. In these structures individual dipole moment cancels each other. Hence, dipole moment becomes zero. However, SO<sub>2</sub> has bent or angular structure in which individual dipole moments do not cancel each other. Therefore SO<sub>2</sub> has dipole moment.

### 42. Define "Dipole moment Give its various units.

**Dipole Moment:** The product of magnitude of charge and the distance between their center is called dipole moment.

**Mathematical Form:** if 'q' is the charge and 'r' is the distance between them, then

$$\mu = q \times r$$

**Units:** SI unit of dipole moment is mC (meter coulomb)

Generally smaller unit is used for dipole moment called Debye (D).

$$1\text{D} = 3.336 \times 10^{-30} \text{ mC}$$

### 43. What is expression to calculate the dipole moment of a molecule?

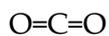
**Dipole Moment:** The product of magnitude of charge and the distance between their center is called dipole moment.

**Mathematical Expression:** if 'q' is the charge and 'r' is the distance between them, then

$$\mu = q \times r$$

### 44. Why CO is polar and CO<sub>2</sub> is non-polar?

CO has a dipole moment while CO<sub>2</sub> does not have any dipole moment. The reason is that CO<sub>2</sub> has the linear structure, where the dipole being equal and opposite, cancel out each other's effect.



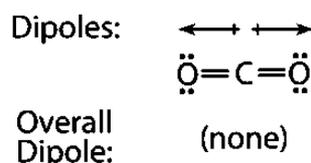
$$\mu = 0.0\text{D}$$



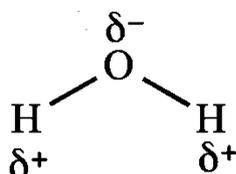
$$\mu = 1.2\text{D}$$

### 45. Dipole moment of CO<sub>2</sub> is zero While that of H<sub>2</sub>O is 1.85D Explain.

In case of CO<sub>2</sub>, the individual bond moment cancel each other due to linear structure of CO<sub>2</sub> so, it has zero dipole moment.



While in water individual bond moment do not cancel each other. Due to angular structure it has some dipole moment value.

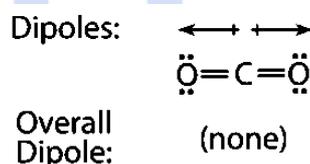


**46. The dipole moments of CO<sub>2</sub> and CS<sub>2</sub> are zero but that of SO<sub>2</sub> is 1.61 D.**

CO<sub>2</sub> and CS<sub>2</sub> have linear structures which are symmetrical. In these structures individual dipole moment cancels each other. Hence, dipole moment becomes zero. However, SO<sub>2</sub> has bent or angular structure in which individual dipole moments do not cancel each other. Therefore, SO<sub>2</sub> has dipole moment.

**47. Why dipole moment of CO<sub>2</sub> is zero?**

Carbon has four electrons in the outermost orbital. It makes two sigma bonds and two pi bonds with oxygen atoms. It means that it has two double bonds. Carbon has no lone pair in CO<sub>2</sub>. So CO<sub>2</sub> is AB<sub>2</sub> type molecule having the linear structure.



Two dipoles cancel the effect of each other and it behaves as non-polar molecule.

**48. How the percentage ionic character of a covalent bond is determined by dipole moment?**

From the experimentally determined dipole moments, the percentage ionic character in a bond can be calculated. For this purpose, we should know the actual dipole moment  $\mu_{obs}$  of the molecule and actual bond length. The dipole moment of 100% ionic compound is represented as  $\mu_{ionic}$ .

$$\% \text{age ionic character} = \frac{\mu_{obs}}{\mu_{ionic}} \times 100$$

**49. Why BF<sub>3</sub> is non-polar but SO<sub>2</sub> is polar?**

BF<sub>3</sub> is a non-polar molecule because it has symmetrical triangular planar geometry having zero dipole moment. This is all due to the cancellation of individual bond moments.

Whereas triatomic molecule SO<sub>2</sub> is bent like H<sub>2</sub>O. dipole moment of SO<sub>2</sub> is 1.16D which ruled out its linear structure. Thus due to presence of dipole moment SO<sub>2</sub> is polar.

## Properties of Compounds

### 50. Rate of reaction of ionic compounds is faster than covalent compound. Why?

The reactions of covalent compounds are slow because they involve bond breakage and bond formation in a chemical reaction. The molecules undergo a chemical change as a whole.

### 51. How the type of bonding affects 'solubility of compounds'?

Usually compounds have ionic bonding or covalent bonding. The solubility of ionic or covalent compounds depend on the nature of their bonding.

Solubility of any compounds depends on "like dissolve like" rule.

Ionic compounds and polar compounds are soluble in ionic solvents like water and non-polar covalent compounds are soluble in non-polar solvents like benzene, ether, carbon tetrachloride.

### 52. Why the melting points, boiling points, heat of Sublimation and heat of vaporization of electrovalent compounds are higher as compared with those of covalent compounds?

In ionic compounds ions are held together by strong electrostatic forces of attractions. A lot of energy is required to break these forces. Hence melting and boiling points of ionic compounds are very high.

While, in covalent compounds intermolecular forces are present in the form of dipole-dipole forces and hydrogen bonding. These forces are easy to break than ionic forces. Hence, covalent compounds have low melting and boiling points as compared to ionic compounds.

### 53. Ionic compounds do not show the phenomena of isomerism. Why?

The compounds which have same molecular formula but different structural formula show the phenomena of isomerism. The phenomena of isomerism is shown by covalent compounds. Ionic compounds have formula unit not have molecular formula that's why ionic compounds do not show isomerism.

## LONG QUESTION

1. Describe postulates of valence shell electron pair repulsion theory.
2. Explain the molecular orbital structures of following molecules on the basis of the MOT:  
(i)  $N_2$  (Nitrogen) (ii)  $O_2$  (oxygen)
3. Write the main points of valence shell electron pair repulsion theory and explain the structure of ammonia on the basis of this theory.
4. Describe the bonding in  $O_2$  according to molecular orbital theory and explain its paramagnetic property.

5. Explain paramagnetic behaviour of  $O_2$  on the basis of MOT and prove MOT is superior to other theories.
6. Explain structure of  $CH_4$  on the basis of hybridization theory.
7. Define electron affinity. Name the factors affecting it. How does it vary in the periodic table?
8. Write main postulates of molecular orbital theory
9. Write four postulates of valence shell electron pair repulsion theory .
10. What is hybridization? Explain  $sp^2$  hybridization with example.
11. Draw the shapes of following molecules according to VSEPR theory.  
i)  $BeCl_2$     ii)  $BF_3$     iii)  $H_2O$     iv)  $NH_3$
12. Give four postulates of Valence Shell Electron Pair Repulsion Theory
13. Explain ionization energy by giving one example? Also discuss its periodic trends.
14. Define hybridization. Explain the geometry of ethene on the basis of  $sp^2$  hybridization.
15. Explain postulates of molecular orbital theory?
16. Write down the main postulates of VSEPR theory and discuss the structure of  $NH_3$  with reference of this theory.
17. Give the defects of Bohr's atomic model.
18. Explain important points of molecular orbital theory and draw structure of nitrogen ( $N_2$ ) molecule according to this theory.
19. Explain bonding in  $O_2$  according to molecular orbital theory.
20. What is orbital hybridization? Explain hybridization of carbon in  $CH_4$ .
21. Explain paramagnetic nature of  $O_2$  according to molecular orbital theory.