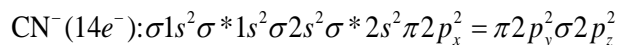


## <H1>Additional Objective Questions

### <H2> Single Correct Choice Type

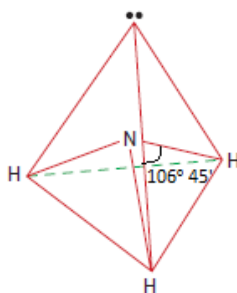
1.(B) The molecular electronic configuration of  $\text{CN}^-$  is:



$$\text{B.O} = \frac{N_b - N_a}{2} = \frac{10 - 4}{2} = 3$$

Since, there is no unpaired electron, thus,  $\text{CN}^-$  is diamagnetic in nature.

2.(C) The structure of ammonia is



Due to the lp-bp repulsion in  $\text{NH}_3$ , the bond angle in it is less than  $109.5^\circ$ , that's is  $106^\circ 45'$ .

3.(A) Lattice enthalpy of  $\text{Al}_2\text{O}_3$  is calculated from Born-Haber cycle using enthalpy of formation of  $\text{Al}(\text{g})$ , first, second and third ionization energies of  $\text{Al}(\text{g})$ , enthalpy of formation of  $\text{O}(\text{g})$ , first and second electron gain enthalpies of  $\text{O}(\text{g})$  and enthalpy of formation of  $\text{Al}_2\text{O}_3(\text{s})$ .

4.(C) The hybridization of  $\text{IF}_4^+$  is

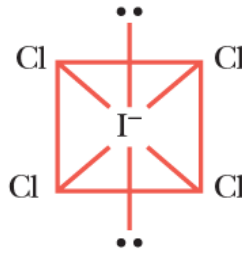
$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[7 + 4 + 0 - 1]}{2} = \frac{10}{2} = 5(sp^3d)$$

Therefore, the shape of  $\text{IF}_4^+$  is irregular tetrahedral.

5.(1)  $[\text{ICl}_4]^-$  ion will have  $sp^3d^2$  hybridization.

$$\begin{aligned} \text{Hyb.} &= \frac{[V + M - C + A]}{2} \\ &= \frac{[7 + 4 - 0 + 1]}{2} = \frac{12}{2} = 6(sp^3d^2) \end{aligned}$$

$V = \text{Valence shell electrons}$
$M = \text{Monovalent atom}$
$C = \text{Cationic charge}$
$A = \text{Anionic charge}$



6.(B)  $\text{NO}_2^+ - 180^\circ$ ,  $\text{NO}_2^- - 115^\circ$ ,  $\text{NO}_2 - 132^\circ$ ,  $\text{NO}_3^- - 120^\circ$

7.(B)  $\text{O}_2$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2 \pi^* 2p_x^1 \pi^* 2p_y^1$

Bond order of  $\text{O}_2 = 1/2 (10 - 6) = 2$

$\text{O}_2^+$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2 \pi^* 2p_x^1$

Bond order =  $(1/2) (10 - 5) = 5/2 = 2.5$

$\text{O}_2^-$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2 \pi^* 2p_x^2 \pi^* 2p_y^1$

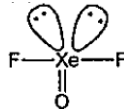
Bond order of  $\text{O}_2^- = 1/2 (10 - 7) = 3/2 = 1.5$ .

$\text{O}_2^{2-}$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2 \pi^* 2p_x^2 \pi^* 2p_y^2$

Bond order of  $\text{O}_2^{2-} = 1/2 (10 - 8) = 1$

Therefore, bond order is  $\text{O}_2^{2-} < \text{O}_2^- < \text{O}_2 < \text{O}_2^+$

8.(C) The structure of  $\text{XeOF}_2$  molecule is T-shaped.



$H = 3 + (1/2) (8 - 4) = 3 + 2 = 5 = sp^3d$  hybridization.

9.(3) According to MO theory,

$$\text{Li}_2^+(5e^-) = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^1 \Rightarrow \text{bond order} = \frac{3-2}{2} = 0.5$$

$$\text{Li}_2^-(7e^-) = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^1 \Rightarrow \text{bond order} = \frac{4-3}{2} = 0.5$$

As both are having bond order 0.5, hence, both are stable.

10.(C) The MO electronic configuration of  $\text{C}_2$  is

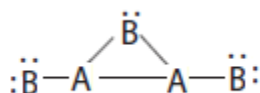
$$\text{C}_2(12e^-): \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2$$

Four electrons are present in  $2\pi$  molecular orbitals; that is why double bond contains both  $\pi$  bonds. Molecule is diamagnetic.

**11.(D)** The MO electronic configuration of  $N_2^+$  is  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2 \sigma 2p_z^1$ .

Since, one electron is removed from  $\sigma 2p_z$  orbital, therefore, it left with 2  $\pi$  bonds and half  $\sigma$  bond.

**12.(A)** A has three electrons in the outermost orbit, it needs five more electrons to complete its octet and B has six electrons so it would need two more electrons to complete the octet. So, each A will form two bonds with two B atoms, and single electron remaining on each A will share form bond with the same B atom such that one B atom is shared between the two A atoms. The structure is



**13.(B)** Bond order is inversely proportional to bond length:

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

$NO^{2+}$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^1 \sigma 2p_z^2$

Bond order is 2.5.

$NO^+$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2$

Bond order is 3.

$NO$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^1$

Bond order is 2.5.

$NO^-$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^1 \pi^* 2p_y^1$

Bond order is 2.0.

So,  $NO^+$  species has the shortest bond length.

**14. (D)**  $O_2 = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^1 = \pi^* 2p_y^1$

$$\text{Bond order} = \frac{[N_a - N_b]}{2} = \frac{10 - 6}{2} = 2 \text{ and paramagnetic.}$$

$N_2^{2-} = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2 \sigma 2p_z^2 \pi^* 2p_x^2$

$$\text{Bond order} = \frac{[N_a - N_b]}{2} = \frac{10 - 6}{2} = 2 \text{ and diamagnetic.}$$

$O_2^{2-} = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^2 = \pi^* 2p_y^2$

$$\text{Bond order} = \frac{[N_a - N_b]}{2} = \frac{10 - 8}{2} = 1 \text{ and diamagnetic.}$$

$C_2^{2-} = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2$

$$\text{Bond order} = \frac{[N_a - N_b]}{2} = \frac{10 - 4}{2} = 3 \text{ and diamagnetic.}$$

**15.(A)** Given that the dipole is 0.38 D =  $0.38 \times 3.34 \times 10^{-30}$  C m =  $1.26 \times 10^{-30}$  C m.

If the bond was 100% ionic, the magnitude of the dipole moment will be

$$|\mu| = qr = 1.6 \times 10^{-19} \times 1.61 \times 10^{-10} \text{ C m} = 2.57 \times 10^{-29} \text{ C m}$$

Therefore, the fraction of charge covalently bonded is

$$\frac{1.26 \times 10^{-30}}{2.57 \times 10^{-29}} = 0.05$$

Therefore, % fractional charge = 5%

16.(D) The structures of the given molecules are

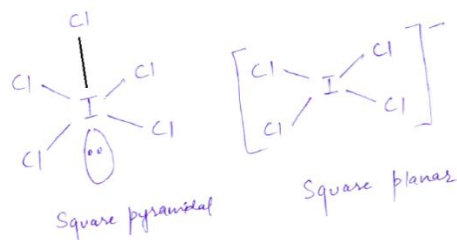
$$\text{Hyb.} = \frac{[V + M - C + A]}{2}$$

$$[\text{Hyb.}]_{\text{ICl}_4^-} = \frac{[7 + 4 - 0 + 1]}{2} = \frac{12}{2} = 6 \text{ or } sp^3d^2$$

$$[\text{Hyb.}]_{\text{ICl}_5} = \frac{[7 + 5 - 0 + 0]}{2} = \frac{12}{2} = 6 \text{ or } sp^3d^2$$

$V$  = Valence shell electrons  
 $M$  = Monovalent atom  
 $C$  = Cationic charge  
 $A$  = Anionic charge

Both have same hybridization, that is, same geometry, but due to different number of monovalent atoms, structure is different,  $\text{ICl}_5$  have square pyramidal shape and  $\text{ICl}_4^-$  have square planar shape.

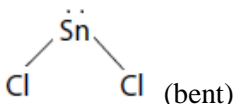


17.(C) Fajans' rule: Greater the polarization of anion in a molecule, more covalent the molecule will be.

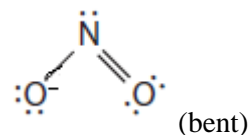
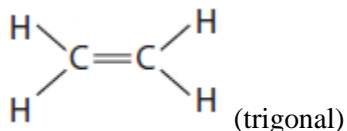
18.(A) Bond order of NO is 2.5 and paramagnetic in nature and  $\text{NO}^+$  has bond order of 3 and diamagnetic in nature. Whereas  $\text{N}_2$  and  $\text{O}_2$  are paramagnetic in nature.

19.(A) Bond strength follows the order single bond < double bond < triple bond.

20.(A)  $\text{O}=\text{C}=\text{O}$  (linear)

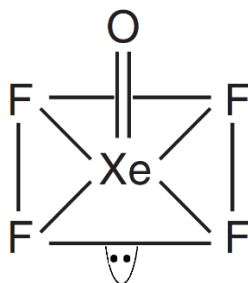


$\text{Cl}-\text{Hg}-\text{Cl}$  (linear)

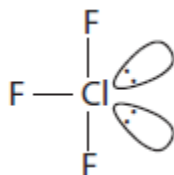


21.(A) Hybridization of Xe in  $\text{XeOF}_4$  is  $sp^3d^2$

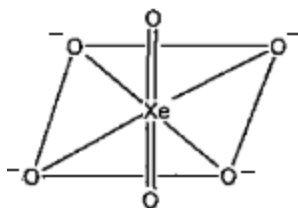
⇒ 5σ bonds + 1 lone pair (l.p).



**22.(C)**  $\text{ClF}_3$  has T-shaped geometry with two lone pairs, so it does not have zero dipole moment.



**23.(B)** Perxenate ion is  $[\text{XeO}_6]^{4-}$ . The structure is



**24.(D)** The MO electronic configuration of NO:  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^1$   
 NO has one unpaired electron in antibonding molecular orbital.

**25.(A)**  $\text{CH}_4$  – Net dipole moment is zero.

$\text{NF}_3$  – Resultant dipole moment toward nucleus of nitrogen.

$\text{NH}_3$  – Resultant dipole moment toward lone pair.

$\text{H}_2\text{O}$  – Resultant dipole moment  $2 \times H_A \times \cos 104.5^\circ$ .

**26.(C)** Ionization potential of sodium is less, electron gain enthalpy of Br is more. So, the type of bond formed is ionic bond.

27. (D) Larger the anion, greater is its polarizability (because the hold on the electron cloud by the nucleus decreases).

28. (A) The hybridization of  $\text{XeO}_3$

$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[8 + 0 + 0 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

The shape of  $\text{XeO}_3$  is pyramidal.

29. (C) Each water molecule is surrounded by 4 other water molecules that are bonded by hydrogen bonds.

30. (C) In  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ , electrovalent, covalent, and coordinate covalent bonds are present.

31. (A) If by taking up an electron, bond order increases so, anion formed will be more stable than its parent species.

$$\text{C}_2^- = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2 \sigma 2p_z^1$$

$$\text{B.O} = \frac{N_b - N_a}{2} = \frac{9 - 4}{2} = 2.5$$

$$\text{F}_2^- = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^2 = \pi^* 2p_y^2 \sigma^* 2p_z^1$$

$$\text{B.O} = \frac{N_b - N_a}{2} = \frac{10 - 9}{2} = 0.5$$

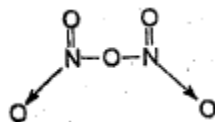
$$\text{NO}^- = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^1 = \pi^* 2p_y^1$$

$$\text{B.O} = \frac{10 - 6}{2} = 2.0$$

$$\text{O}_2^- = \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^2 = \pi^* 2p_y^1$$

$$\text{B.O} = \frac{10 - 7}{2} = 1.5$$

32. (B) The structure of  $\text{N}_2\text{O}_5$  is



33. (B) The MO electronic configuration of CO is:

$$\text{CO}(14e^-): \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2 \sigma 2p_z^2$$

Among the given species, only CO molecule is diamagnetic due to absence of any unpaired electrons.

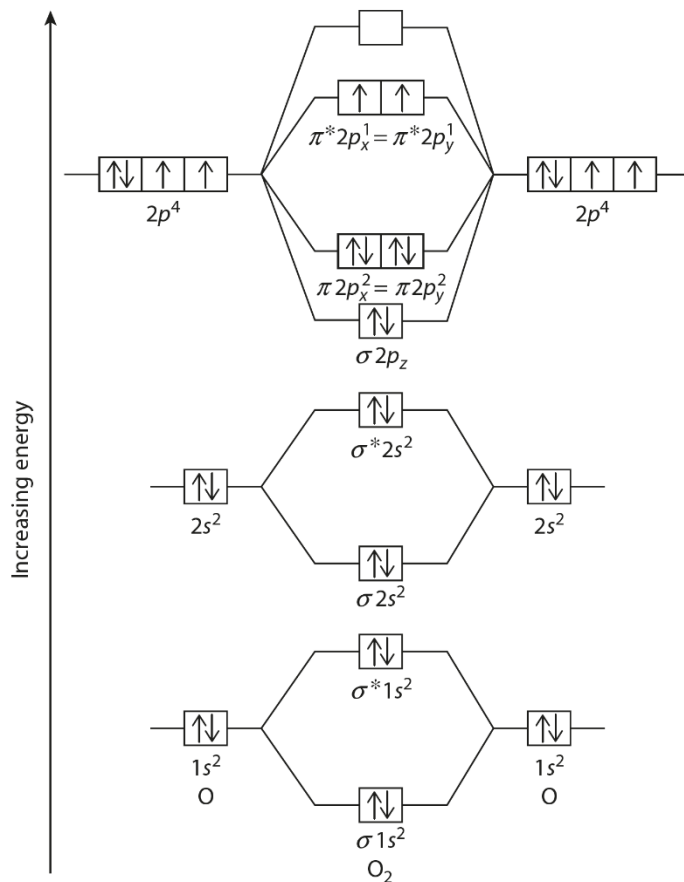
34. (B) The MO electronic configuration of  $\text{O}_2$ :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2 \pi^* 2p_x^1 \pi^* 2p_y^1$

Hence, it is paramagnetic due to two unpaired electrons.

**35.(B)**  $\text{CH}_3^-$  is  $sp^3$ -hybridized with 3 bond pairs and 1 lone pair. Thus, the shape is trigonal pyramidal like  $\text{NH}_3$ .

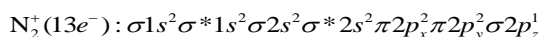
**36.(C)**  $\text{H-O-O-H}$ . Here, O-H bonds is polar, O-H bonds is polar, O-O bond is non-polar.

**37.(C)** Molecular orbital diagram of  $\text{O}_2$  is



After addition of one electron to the oxygen molecule, the electrons go to  $\pi^* 2p_x$  orbital.

**38.(D)** The MO electronic configuration of  $\text{N}_2^+$  is

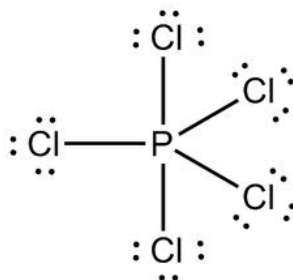


In, there is one unpaired electron present in bonding molecular orbital; rest all have paired.

**39.(A)** In benzene, due to the resonance C-C bond length is in between double bond and single bond. However, if we consider the six carbon atoms to be linked by three single and three double bonds, then the circumference of benzene ring is  $3 \times (\text{C}=\text{C}) + 3 \times (\text{C}-\text{C}) = 3 \times 1.33 + 3 \times 1.54 = 8.61, \text{ \AA}$ .

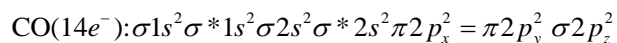
40.(C) Due to overlapping of  $p$ - $p$  orbitals, bond exists between two fluorine atoms to form fluorine molecule.

41.(D) P has 10 electrons in its valence shell in  $\text{PCl}_5$ .

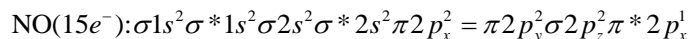


42.(B)  $s$  and  $p$  orbitals hybridize to form  $2sp$  hybrid orbitals oriented at an angle of  $180^\circ$ .

43.(B) The MO electronic configuration of CO is



$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{10 - 4}{2} = 3.0$$



$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{10 - 5}{2} = 2.5$$

44.(A) As the electronegativity of the atom increases, the strength of the ionic bond increases because electronegative atom attracts the electron pair toward itself. If electronegativity is more, strength of ionic bond is also more.

45.(B) We know

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

Greater the delocalization of  $\pi e^-$  (-ve charge) shorter the bond length.

46.(D) Hydrogen bond is the weakest bond.

47.(D)  $\text{N}_2 > \text{O}_2$ ;  $\text{N}_2 > \text{N}_2^+ > \text{NO}^+ > \text{NO}^-$

48.(A) We know

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

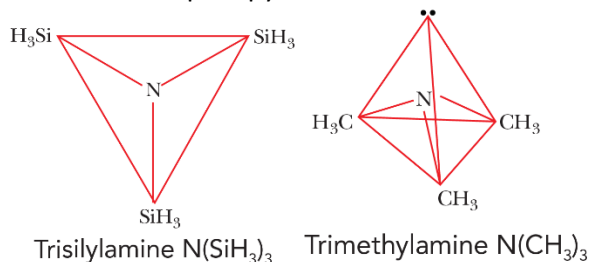


$$109.5^\circ < 120^\circ < 180^\circ$$

49.(B)  $\text{SO}_2$  is a bent molecule, and has a net dipole moment; hence, it is polar.

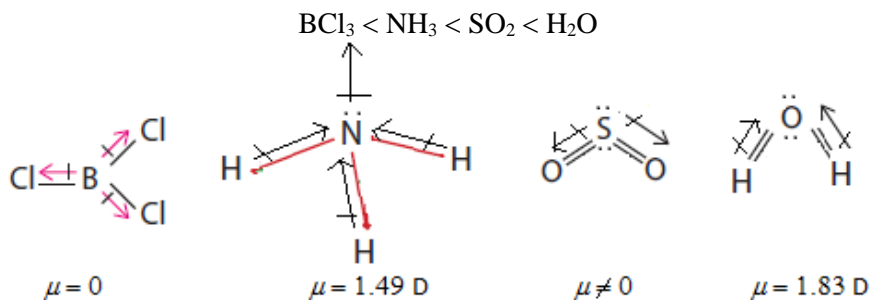
### Solution

50.(A) In both the structures, nitrogen is  $sp^3$  hybridized with three bond pair and one lone pair. In  $(\text{SiH}_3)_3\text{N}$ , due to the vacant  $d$ -orbitals of Si, a  $d\pi - p\pi$  back bond formed as a result, the shape of trisilylamine is planar and it is less basic than trimethyl amine, which does not show back bonding due to absence of vacant orbital and hence its shape is pyramidal.



51.(C)  $\text{AgCl}_2^-$  is  $sp$ -hybridized because it is linear in shape, and silver complexes are always linear in shape and silver is  $sp$ -hybridized.

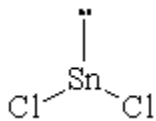
52.(B) Dipole moment is defined as the product of the magnitude of charge on any one of the atoms and the distance between them.



The dipole moment of  $\text{BCl}_3$  molecule is zero. The bond dipoles of three B–Cl bonds give a net sum of zero because the resultant of any two is equal and opposite to the third (parallelogram law of forces) as shown above. In other molecules dipole moment shows increasing trend due to the presence of central atom that is electronegative in nature. As the electronegativity of the central metal atom increases ( $\text{N} < \text{S} < \text{O}$ ), dipole moment of the resulting molecule also increases. Hence, this shows the above trend.

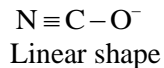
### <H2> Multiple Correct Choice Type

1.(A, C, D) The shape of  $\text{SnCl}_2$ :

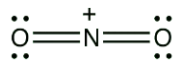


Bent or V-shape

The shape of  $\text{NCO}^-$ :

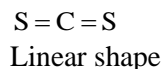


The shape of  $\text{NO}_2^+$ :



Linear shape

The shape of  $\text{CS}_2$ :



**2.(A, C)**  $\text{CO}_2$  is isostructural with  $\text{HgCl}_2$  and  $\text{C}_2\text{H}_2$ .  
All these are linear in structure and process  $sp$  hybridization.

**3.(A, B)**

**4.(A, C)** The hybridization of  $\text{IF}_5$  is

$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[7 + 5 + 0 - 0]}{2} = \frac{12}{2} = 6(sp^3d^2)$$

Structure is square pyramidal, therefore, it is polar.

The hybridization of  $\text{XeF}_4$  is

$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[8 + 4 + 0 - 0]}{2} = \frac{12}{2} = 6(sp^3d^2)$$

Structure is square planar therefore, it is non-polar.

The hybridization of  $\text{XeOF}_4$  is

$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[8 + 4 + 0 - 0]}{2} = \frac{12}{2} = 6(sp^3d^2)$$

Structure is square pyramidal, therefore, it is polar.

The hybridization of  $\text{ICl}_4^-$  is

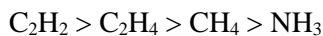
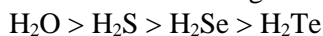
$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[7 + 4 + 1 - 0]}{2} = \frac{12}{2} = 6(sp^3d^2)$$

Structure is square planar, therefore, it is polar.

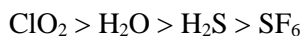
**5.(B, D)** If the electronegativity difference between the two atoms is more than 1.9, the bond is taken as ionic. Here, **(B)** and **(D)** have an electronegativity difference greater than 1.9.

**6.(A, C)**  $\text{CH}_4$  and  $\text{XeO}_3$  both are in  $sp^3$  hybridization. So, the hybrid orbitals of both central atom have the same  $s$ -character.

**7.(A, B, D)** The correct order of bond angle is:



$$180^\circ \quad 120^\circ \quad 109^\circ \quad 107^\circ$$



**8.(A, D)**  $\text{BF}_3$  has a regular trigonal planar structure and  $\text{BF}_4^-$  has a regular tetrahedral structure. Both  $\text{NF}_3$  and  $\text{PF}_3$  have pyramidal structure.

**9.(C, D)**

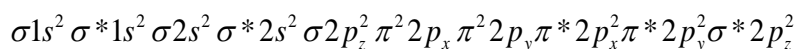


Non-bonding combinations of atomic orbitals.

**10.(A, B, C)** Due to the coordinate bond in carbon monoxide, the oxygen atom is partial +ve charge and carbon is partial -ve charge, therefore, carbonyl have more dipole moment than carbon monoxide. Since, the dipole moment depend upon the product of charge and distance between two poles, therefore, the dipole moment of  $\text{CH}_3\text{Cl}$  is greater than  $\text{CH}_3\text{F}$ .

**11.(A, B, D)** Because in pyridine,  $\text{BX}_3$ , make an adduct, therefore, the bond energy of B-X is more in  $\text{BX}_3$  than in  $\text{BX}_3$  pyridine. Alkaline earth metal form simple oxide when heated in presence of excess oxygen. Magnesium cannot be extracted by the electrolysis of fused  $\text{Mg}(\text{OH})_2$ .

**12.(A, B)** The MO electronic configuration of  $\text{F}_2^{2-}$  :



$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{10 - 10}{2} = 0$$

$\text{Ar}_2$  also has zero bond order because it is a noble gas; it exists only monoatomic in nature.

**13.(A, B, D)** The hybridization of  $\text{ClO}_4^-$  is

$$\text{Hyd.} = \frac{[V + M + A - C]}{2} = \frac{[7 + 0 + 1 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

The hybridization of  $\text{SO}_4^{2-}$  is

$$\text{Hyd.} = \frac{[6 + 0 + 2 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

The hybridization of  $\text{SiO}_4^{4-}$  is

$$\text{Hyd.} = \frac{[4 + 0 + 4 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

The hybridization of  $[\text{PCl}_4]^+$  is

$$\text{Hyd.} = \frac{[5 + 4 + 0 - 1]}{2} = \frac{8}{2} = 4(sp^3)$$

The hybridization of  $\text{SnCl}_4$  is

$$\text{Hyd.} = \frac{[4 + 4 + 0 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

The hybridization of  $\text{H}_2\text{O}(\text{s})$  is

$$\text{Hyd.} = \frac{[6 + 2 + 0 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

The hybridization of  $\text{FCIO}_3$  is

$$\text{Hyd.} = \frac{[7 + 1 + 0 - 0]}{2} = \frac{8}{2} = 4(sp^3)$$

Out of the given species, all species in option (A), (B) and (D) have  $sp^3$  hybridization and tetrahedral geometry.

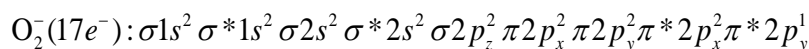
<H2> Assertion–Reasoning Type

1.(C) Because in  $sp^3$  hybridization there are no vacant orbitals as well as unpaired electrons to form  $\pi$  bond.

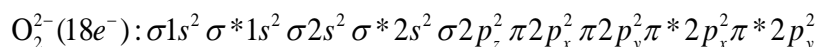
2.(C) As hydrogen bonding in ammonia is less compared to that in water and also nitrogen is less electronegative than oxygen, so  $NH_3$  has low melting point than water.

3.(B)  $BF_3$  molecule is planar because it undergoes  $sp^2$  hybridization and  $NF_3$  molecule is pyramidal in shape because in this the central atom, nitrogen undergoes  $sp^3$  hybridization.

4.(A) The MO electronic configuration of  $O_2^-$  is



The MO electronic configuration of  $O_2^-$  is



Due to the presence of unpaired electrons in superoxide, the magnetic moment of its larger than peroxide, which does not have any unpaired electron.

5.(D) *o*-Nitrophenol has intramolecular hydrogen bonding which weaker than intermolecular hydrogen bonding in *p*-nitrophenol.

6.(C)  $\sigma$  Bond is strong while  $\pi$  bond is weak because in  $\sigma$  bond formation overlapping of orbitals takes place along their internuclear axis, giving rise to maximum electron density on the axis. So  $\sigma$  bond is strong.

7.(B) Due to  $sp^3d$  hybridization the fluorine atomic orbitals overlap with the sulphur atomic orbitals of the sulphur atom to form  $SF_4$  molecule.

Sulphur in ground state:  $\uparrow\downarrow \quad \uparrow\downarrow \uparrow\uparrow \quad \square \square \square \square \square$

Sulphur in excited state:  $\uparrow\downarrow \quad \uparrow\uparrow\uparrow \quad \uparrow \square \square \square \square$

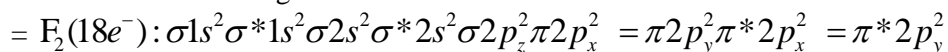
Hybridization:  $\uparrow\downarrow \quad \uparrow\times \uparrow\times \uparrow\times \quad \uparrow\times \square \square \square \square$

Total ten electrons on sulphur atom.

The shape of  $SF_4$  is see-saw shaped



8.(A) The MO electronic configuration of  $F_2$  is



$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{10 - 8}{2} = 1.0$$

**9.(D)** Except for ammonia the boiling point increases as we move down the group. The abnormally high boiling point of  $\text{NH}_3$  is due to the association of molecules through hydrogen bonding and boiling point increases down the group due to increase in van der Waals forces.

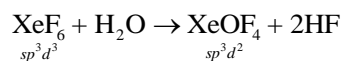
$\text{PH}_3$	<	$\text{AsH}_3$	<	$\text{NH}_3$	<	$\text{SbH}_3$	<	$\text{BiH}_3$
$-87.5^\circ\text{C}$		$-62.45^\circ\text{C}$		$-34.5^\circ\text{C}$		$-18.45^\circ\text{C}$		$+16.8^\circ\text{C}$

**10.(A)** Concept based.

**11.(D)**  $\pi(2p_x)$  has one nodal plane whereas  $\pi^*(2p_x)$  has two nodal planes. Bonding  $\pi$ -molecular orbitals have one nodal plane whereas antibonding  $\pi$ -molecular orbitals have two nodal planes.

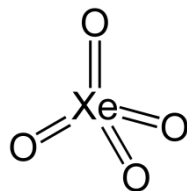
### <H2>Comprehension Type

**1.(D)** The reaction involved is



**2.(B)**  $\text{XeF}_6 + 3\text{H}_2\text{O} \rightarrow \text{XeO}_3 \text{ (B)} + 6\text{HF}$

The structure of  $\text{XeO}_3$  is tetrahedral



**3.(B)** In  $\text{XeF}_6$  **(A)**, the bond angle is  $90^\circ$ .

In  $\text{XeO}_3$  **(B)**, the bond angle is  $103^\circ$ .

**4.(D)** The shape of  $\text{PdCl}_4^{2-}$  is square planar.

**5.(C)**  $\text{NH}_3$  and  $\text{ClO}_3^-$  have one lone pair of electrons.

**6.(D)** The shape of  $\text{ICl}_4^-$  and  $\text{PtCl}_4^{2-}$  is square planar.

**7.(B)** The hybridization of  $\text{XeF}_3^+$  is

$$\text{Hyd.} = \frac{[8 + 3 + 0 - 1]}{2} = \frac{10}{2} = 5(sp^3d)$$

$AB_5 \rightarrow AB_2L_2 \rightarrow$  Two lone pairs

The hybridization of  $SF_3^+$  is

$$\text{Hyd.} = \frac{[6 + 3 + 0 - 1]}{2} = \frac{8}{2} = 4(sp^3)$$

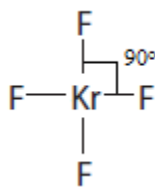
$AB_4 \rightarrow AB_3L \rightarrow$  One lone pair

The hybridization of  $CF_3^+$  is

$$\text{Hyd.} = \frac{[4 + 3 + 0 - 1]}{2} = \frac{6}{2} = 3(sp^2)$$

$AB_3 \rightarrow$  Zero lone pair

**8.(C)** The structure of  $KrF_4$  is



**9.(D)** In  $OBr_2$ , the bond angle is more than  $OCl_2$  because Br is less electronegative element than chlorine.

## <H2>Integer Answer Type

**1.(2)** Let the molecule is X, then the MO electronic configuration is

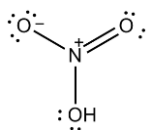
$$X_2(16e^-): \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^1 = \pi^* 2p_y^1$$

Hence, the unpaired electrons is 2.

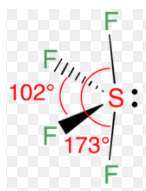
**2.0**

**3.(5)** The geometry of molecule is pentagonal pyramidal.

**4.(7)** The structure of  $HNO_3$  is



**5.(0)** The structure of  $SF_4$  is



6.(6) The electronic configuration of  $Zn^{2+}$  is  $[Ar]3d^{10}$

$Cd^{2+}$  is  $[Kr]4d^{10}$ ;  $Hg^{2+}$  is  $[Xe]4f^{14}5d^{10}$ ;  $Cu^+$  is  $[Ar]3d^{10}$ ;  $Ag^+$  is  $[Kr]4d^{10}$ ;  $Au^+$  is  $[Xe]4f^{14}5d^{10}$ . Hence, all species have pseudo inert gas configuration.

7.(4) One  $CN^-$  have two  $\pi$ -bonds, therefore, two  $CN^-$  will have four  $\pi$ -bonds.

8.(4) Water molecule can make maximum of 4 hydrogen bonds.

### <H2> Matrix–Match Type

1. A  $\rightarrow$  (r); B  $\rightarrow$  (s); C  $\rightarrow$  (q); D  $\rightarrow$  (p)

Using the formula  $H = S + (1/2)(E - V \pm C)$

(A)  $AsF_6^-$ :  $H = 6 + (1/2)(5 - 6 + 1) = 6 = sp^3d^2$

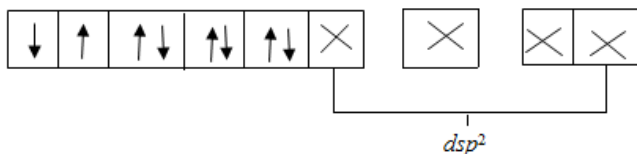
(B)  $TeCl_4$ :  $H = 4 + (1/2)(6 - 4) = 5 = sp^3d$

(C)  $SiO_2$ :  $H = 4$  (as O is divalent) +  $(1/2)(4 - 4) = 4 = sp^3$

(D) Ground state configuration of Ni =  $3d^8 4s^2$

Electronic configuration of  $Ni^{2+} = 3d^8$

The hybridization in  $[Ni(en)_2]^{2+}$



2. A  $\rightarrow$  (p); B  $\rightarrow$  (q); C  $\rightarrow$  (p); D  $\rightarrow$  (s)

(A)  $IO_2F_2^-$ :  $H = 4 + (1/2)(7 - 6 + 1) = 4 + 1 = 5 = sp^3d$

(B)  $F_2SeO$ :  $H = 3 + (1/2)(6 - 4) = 4 = sp^3$

(C)  $ClOF_3$ :  $H = 4 + (1/2)(7 - 5) = 4 + 1 = 5 = sp^3d$

(D)  $XeF_5^+$ :  $H = 5 + (1/2)(8 - 5 - 1) = 6 = sp^3d^2$

3. A  $\rightarrow$  (q, r); B  $\rightarrow$  (p, s); C  $\rightarrow$  (p, s); D  $\rightarrow$  (q, r)

$O_2^-$ : Bond order is 1.5 – paramagnetic.

CO: Bond order is 3 – diamagnetic.

$NO^+$ : Bond order is 3 – diamagnetic.

$He_2^+$ : Bond order is 0.5 – paramagnetic.

4. A  $\rightarrow$  (s); B  $\rightarrow$  (p, s); C  $\rightarrow$  (q, r); D  $\rightarrow$  (q, r)

$$\text{Bond order} = \frac{1}{2}[2 - 1] = \frac{1}{2}$$

Since  $\text{He}_2^+$  has one unpaired electron, thus  $\text{He}_2^+$  is paramagnetic in nature and it also has fractional bond order.

(A)  $\text{O}_2^{2-}$ : The electronic configuration of O is  $1s^2 2s^2 2p^4$ . The MO formed during combination of two O atoms occupy 18 electrons because of the presence of two extra electrons, leading to MO configuration of:  $\text{O}_2^{2-}: \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_x^2, \pi 2p_z^2, \pi 2p_y^2, \pi^* 2p_z^2, \pi^* 2p_y^2$

Since  $\text{O}_2^{2-}$  does not have unpaired electron,  $\text{O}_2^{2-}$  is diamagnetic.

(B) CO: It contains total 14 electrons: 6 electrons from carbon and 8 electrons from oxygen. It is isoelectronic with  $\text{N}_2$  because  $\text{N}_2$  also contains total 14 electrons.

MO configuration of CO:  $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_x^2, \pi 2p_z^2, \pi 2p_y^2$

Since CO does not have unpaired electron, it is diamagnetic.

(C) NO: Electronic configuration of N and O are  $1s^2 2s^2 2p^3$  and  $1s^2 2s^2 2p^4$ , respectively. The MO formed during combination of N and O atoms occupy 15 electrons leading to

MO configuration of NO:  $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_x^2, \pi 2p_z^2, \pi 2p_y^2, \pi^* 2p_z^1$

$$\text{Bond order of NO: } \frac{1}{2}[10 - 5] = \frac{5}{2}$$

Since NO has one unpaired electron, NO is paramagnetic, and it also has fractional bond order.

(D)  $\text{He}_2^+$ : It contains total 3 electrons.

MO configuration of  $\text{He}_2^+$ :  $\sigma 1s^2, \sigma^* 1s^1$

**5. A  $\rightarrow$  (p, q); B  $\rightarrow$  (p, q, r); C  $\rightarrow$  (p, q, r, s); D  $\rightarrow$  (p, q, r, s)**

(A) Both  $\text{PCl}_3\text{F}_2$  and  $\text{PCl}_2\text{F}_3$  have  $sp^3d$  hybridization with trigonal bipyramidal shape.

(B) Both  $\text{BF}_3$  and  $\text{BCl}_3$  have  $sp^2$  shape with trigonal planar shape and also the dipole moments are zero in both the cases (dipole moment cancels each other).

(C) Both  $\text{CO}_2$  and  $\text{CN}_2^{2-}$  have  $sp$  hybridization with linear geometry, zero dipole moment, and also have same number of electrons, that is, 22.

(D)  $\text{C}_6\text{H}_6$  known as benzene and  $\text{B}_3\text{N}_3\text{H}_6$  known as borazine are planar molecules with aromatic character; B and C atoms undergo  $sp^2$  hybridization, have equal dipole moments and total number of electrons is 30.