Exercise-1

Marked Questions may have for Revision Questions.

OBJECTIVE QUESTIONS

Section (A) : Introduction to Bonding, Octet rule and Ionic bond

A-1. Sol. The conditions required for the formation of an ionic bond are :

(i) ionization enthalpy [M(g) \rightarrow M+(g) + e⁻] of electropositive element must be low.

(ii) negative value of electron gain enthalpy [X (g) + $e^- \rightarrow X^-(g)$] of electronegative element should be high.

- A-2. Sol. NaCl is ionic crystal so it is formed by Na⁺ and Cl⁻ ions.
- A-3. Sol. Cs has lowest IE1 amongst the metals and F has higher electron affinity. So Cs and F form most ionic compound.
- A-4. Sol. Electronegativity difference between two combining elements must be greater than 1.7 for ionic compound and it is the essential condition for the formation of ionic compounds. It is ionic because electronegativity difference between two combining elements is 1.8.
- **A-5.** Sol. K on account of lower IE1st can easily form K+ ion losing one electron.
- A-6. Sol. (2) The ease of formation of ionic compounds i.e. stability to form ionic compounds increases as net ionization energy of electropositive element decreases. Hence, the correct order is Na+ > Mg²⁺ > Al³⁺.

A-7. Sol. Lattice energy
$$\propto \frac{1}{\text{size of cation}} \text{LE} \propto \frac{1}{r_+ + r_-}$$

A-8. Sol. Lattice energy $\propto \frac{(\text{Interionic distance})^2}{(\text{Interionic distance})^2}$, $\text{LE}\alpha = \frac{1}{\text{size of anion}}$

Section (B) : Fajan's rule

- **B-1.** Sol. According to Fajan's rule : Covalent character in ionic compound is directly proportional to polararibility of anion and polarising power of cation.
- **B-2.** Sol. According to Fajans rule increase in oxidation state of Pb⁴⁺ increase the polarising power of cation and thus increases the polarisation of Br– ions.
- **B-3.** Sol. According to Fajan's rule as the size of cation increases their polarising power decreases and thus the covalent character decreases.
- **B-4.** Sol. As charge on cations increases, their polarising power increase and thus covalent character increase.

$$\overset{\scriptscriptstyle +}{\text{LiCl}} < \overset{\scriptscriptstyle 2+}{\text{Be}} \overset{\scriptscriptstyle 3+}{\text{Cl}}_2 < \overset{\scriptscriptstyle 3+}{\text{BC}} \overset{\scriptscriptstyle 4+}{\text{C}} \overset{\scriptscriptstyle 4+}{\text{Cl}}_4$$

B-5. Sol. Sn⁴⁺ has highest polarising power amongst Na⁺, Pb²⁺ and Al³⁺ because of smaller size and higher charge.

So SnCl₄ is most covalent and thus have least melting point.

- **B-6.** Sol. As F⁻ has lowest polarisability on account of smallest size among O²⁻, N³⁻ and C⁴⁻, it causes less polarisation and, therefore, has lowest covalent character. Hence AIF₃ is the most ionic.
- **B-7.** Sol. According to Fajan's rule bigger the anion more will be the polarisability of anion. As a result, there will be more polarisation leading to increased covalent character i.e. lesser ionic character in the compound, AlI₃.
- B-8. Sol. CuCl exists as Cu⁺ and Cl⁻. Cu⁺ has 18 electrons in valence shell.

Section (C) : Covalent Bond and Coordinate Bond

C-1. Sol. The maximum covalency of an element is equal to the actual number of s and p-electrons in the outermost shell, when formal charge is zero on it.

$$\begin{bmatrix} H \\ H \\ \sigma \\ H \\ \sigma \\ H \end{bmatrix}^{+} CI^{-}$$
; N – H bond is formed by sharing of one electron each by H and N.

Sol.

C-3.

- **C-4.** Sol. In SF₆, PCI₅ and IF₇ the valence shell has 12, 10 and 14 electrons. As all contain more than 8 electrons in their valence shell they are example of super octet molecules.
- **C-5.** Sol. In N₂ molecule each nitrogen atom contributes three electrons so total number of electrons are 6.
- **C-6.** Sol. In BCl₃ and PCl₅, B and P contain 6 and 10 electrons respectively in their valence shell. Therefore they violate octet rule.
- **C-7. Sol.** Covalent bond is formed when electronegativity difference of two atom is equal to 1.7 or less than 1.7.

$$\begin{array}{c} H \\ | \\ H - N : \longrightarrow H^{+} \\ | \\ \vdots \end{array}$$

- C-9. Sol. H
- **C-10.** Sol. The species in which central atoms has higher valencies than their normal valencies are called as hypervalent species.

Section (D) : Formal charge and Resonance



D-2. Sol. Resonance is due to delocalization of π electron.

D-3. Sol. (IV) structure is not allowed because the atoms are in different positions i.e. in cyclic form.

D-4. Sol. The nuber of e⁻ pair is same in resonating structure.

D-5._ Sol.

$$\therefore$$
 Bond order = 1.5

 \therefore Bond order = 1.5.

D-6. Sol. (1) $CO_2=2$ (2) $CH_3COO^-=1.5$ (3) CO=3 (4*) $CO_3^{2-}=1.33$

D-7. Sol. (D) Position of atoms are different.
 (E) Has unpaired electrons and this is not possible since the molecule is diamagnetic.

Section (E) : Valence bond theory

- E-1. Hint : It is fact.
- E-2. Sol. CaC_2 exists as Ca_2^+ and $C_2^{2-}\left[C \stackrel{\longrightarrow}{=} C\right]^{2-}$. N = C $C = C \stackrel{C = N}{\subset C = N}$; 9 σ and 9 π bonds.
- **E-4.** Sol. $H_{\overline{\sigma}} C \frac{\overline{\pi}}{\overline{\sigma}} C_{\overline{\sigma}} H$



- E-5. Sol.
- **E-6. Sol.** (1) and (2) have negative overlap while (C) has positive overlap. Thus (C) will show effective overlapping.
- **E-7.** Sol. (1) σ bond is formed by axial over lapping.
 - (2) p-orbital have both axial and side ways over lapping

Section (F) : VSEPR Theory



F-2.



(trigonal planar)

Sol.

(2) (trigonal planar)



(trigonal planar) (trigonal planar)

Section (G) : Hybridisation and Bond angle

- G-1. Sol. As the p-orbital in hybrid orbital increases than % p-character increases.
- **G-2.** Sol. (a) Electronic configuration of boron in ground state is $1s^22s^22p^1$.



BCI₃: sp² hybridisation

(b) Electronic configuration of nitrogen in ground state is $1s^22s^22p^3$.





(c) Electronic configuration of phosphorus in ground state is 1s²2s²2p⁶3s²3p³.



PCl₃: sp³ hybridisation

(d) Electronic configuration of boron in ground state is $1s^22s^2$.







G-4.

Sol. Steric number = 3 + 1 = 4; so the hybridization is sp₃.



G-5. Sol. sp³d

- **G-6** Sol. (2) Steric number = 0 + 3 = 3; so sp₂ hybridisation
- **G-7.** Sol. Diamond sp^3 ; graphite = sp^2 ; Acetylene = sp
- **G-8.** Sol. Steric number = 2 + 2 = 4; so sp³ hybridisation

$$O = C = C = C = O$$
$$| \qquad | \qquad |$$
$$sp \qquad sp \qquad sp \qquad sp$$

- G-9. Sol.
- **G-10.** Sol. (1) both are $sp^{3}d$ (2) both are sp3d(3) [CIF₂O]⁺ is sp^{3} but [CIF₄O]⁻ is $sp^{3}d^{2}$







G-12. Sol. Atomic size arguments can be used for these species. Larger outer atoms result in larger angles due to steric repulsion.



Phosphorus is the most electronegative of the central atoms. Consequently, it exerts the strongest pull on shared electrons, concentrating these electrons near P and increasing bonding pair-bonding pair repulsions-hence, the largest angle in PI₃. Sb, the least electronegative central atoms, has the opposite effect : Shared electrons are attracted away from Sb, reducing repulsions between the Sb–I bonds. The consequence is that the effect of the lone pair is greatest in SbI₃, which has the smallest angle.

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G-14. Sol. P_4 bond angle = 60°
PH<sub>3</sub> bond angle 90° (Drago's rule)
H<sub>2</sub>O bond angle = 104.5°
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Sol. (109.5) (107)As number of lone pairs of electrons increases, the repulsion between them increases and, therefore H - N - H bond angle decreses.

G-16. Sol. (1) $H_2O = 104.5^{\circ}$; (2) $CO_2 = 180^{\circ}$; (3) $NH_3 = 107^{\circ}$; (4) $CH_4 = 109.5^{\circ}$

G-17. Sol. As size of terminal atoms increases the bond angle increases due to steric repulsion.

Section (H) : Electron deficient bond and Back bonding

H-1._

H-5.

Sol.





Decrease in B–F bond length is due to delocalised $p\pi$ – $p\pi$ bonding between filled p-orbital of F atom and vacant p-orbital of B atom.

- **H-7.** Sol. BF₃ has $p\pi$ - $p\pi$ back bonds
- **H-8.** Sol. Because of $p\pi$ -d π delocalisation of lone pair of electrons present on N atoms (SiH₃) N is planar.

Section (I) : Molecular Orbital Theory (MOT)

- **I-1. Sol.** One bonding molecular orbital and one antibinding molecular orbital is formed as a result of linear combination of atomic orbitals.
- **I-2. Sol.** The electron density is zero in the nodal plane during the formation of a molecular orbital from atomic orbitals of the same atom.
- **I-3.** Sol. He₂ : $(\sigma 1s)_2 (\sigma^* 1s)_2$; bond order = $\overline{2} (2-2) = 0$, He₂ molecule is, therefore, unstable and does not exists.

I-4. Sol.
$$H_{2^-}$$
 bond order = $\frac{2-1}{2} = \frac{1}{2}$.

CHEMICAL BONDING

			10 – 5		10 – 5		9-4	9	9-4
I-5.	Sol. = 2.5.	O ₂ + : B.O. =	2 = 2.5 ; 1	NO : B.O	.= 2 = 2.5	; NO ²⁺ : B.O. =	<u>2</u> = 2.5 ;	CN : B.O. =	2
I-6.	Sol. bond o	B_2 bond orde order \propto 1/bond	er = 1 ; C₂bond d length.	order =	2; F2 bond or	der = 1 ; O₂⁻ bo	nd order = 1.5	5	
I-7.	Sol. N ₂ is 4 * repre	(σ1s)² (σ*1s l. sents antibon	s)² (σ2s)² (σ*2 ding molecular	2s)² (π2μ orbitals.	o2x = π2p2y)	(σ2pz) ² ;numl	ber of anti bor	nding electro	ns in
I-8.	Sol. The box N_2^+ : The box O_2 : (σ^2 The box O_2^- : (σ^2 The box NO^+ de π^2p^2y The box NO dell), (π^*2p The box bond of	N ₂ : $(\sigma 1s)^2$ (and order of N $(\sigma 1s)^2$ ($\sigma^* 1s$) and order of N $(s)^2$ ($\sigma^* 1s$) ² ($\sigma^* 1s$) ² ($\sigma^* 1s$) ² ($\sigma^* 1s$) ² and order of C privative of O ₂ and order of N rivative of O ₂ ($\sigma^* 1$) ¹ and order of N and order of N and order of N and order of N	$(\sigma^*1s)^2 (\sigma^2s)^2$ $(\sigma^*1s)^2 (\sigma^2s)^2$ $(\sigma^2s)^2 (\sigma^*2s)^2$ $(\sigma^2s)^2 (\sigma^2s)^2$ $(\sigma^2s)^2 (\sigma^2s)^2$ $(\sigma^2s)^2 (\sigma^2s)^2$ $(\sigma^2s)^2 (\sigma^2s)^2$ $(\sigma^2s)^2 (\sigma^2s)^2$ $(\sigma^$	$(\sigma^*2s)^2$ = 3.) ² ($\pi 2p2$) = 2.5. $(\sigma 2pz)^2$ = 2. $(\sigma 2pz)^2$ = 1.5. onic with) = 3. ic with C) = 2.5.	$(\pi 2p2x = \pi 2p2)$ $(x = \pi 2p2y) (o)$ $(\pi 2p2x = \pi 2p2)$ $(\pi 2p2x = \pi 2p2)$ $(\pi 2p2x = \pi 2p2)$ $O_{2^{2^{+}}}$; so $(\sigma 1)$ $O_{2^{2^{+}}}$; so $(\sigma 1)$ $O_{2^{+}}$; $(\sigma 1s)^{2}$ (σ^{*}) ation energy.	2y) $(\sigma 2pz)^2$ $(2pz)^1$ 2y) $(\pi^*2px1 = 1)^2$ $(\sigma^*2px2 = 1)^2$ $(\sigma^*1s)^2$ $(\sigma^*2px2 = 1)^2$ $(\sigma^*1s)^2$ $(\sigma^*2px^2)^2$ $(\sigma^*2px^2)^2$	π*2p1y) = π*2p1y) 2s)² (σ*2s)² σ*2s)² (σ2pz) [;]	(σ2pz)² (π2p ² (π2p2x = π	o2x = c2p2y
I-9.	Sol. (2) Bor (3) Bor (4) Bor	(1) Bond ord nd order of F2 nd order of O2 nd order of C2	ler of N ₂ ⁺ = 2.5 = 1 ; The bond = 2 ; The bond = 2 ; The bond	; The bo l order of l order of d order o	$\begin{array}{l} \text{for } \text{for }$	0 ₂ + 1/2(10 - 5)	= 2.5.		
I-10.	Sol. So (σ1 The bo	OF is deriva s)₂ (σ*1s)₂ (α ond order of C	tive of O₂ and i σ2s)₂ (σ*2s)₂ (σ DF 1/2(10 – 7)	soelectro σ2pz)2 (1 = 1.5.	onic with O₂–. π2p2x = π2p2	y)(π*2px2 = π	:*2p1y)		
I-11.	Sol. So $(\sigma 1)$ $(2) O_2^2$ electro $(3) CN^2$ and no (4) CO and no	(1) NO ⁻ deri s) ² (σ^* 1s) ² (σ^* -: (σ 1s) ² (σ^* ns. - is derivative unpaired election is derivative of unpaired election	vative of O_2 and $\sigma 2s)^2$ ($\sigma^* 2s)^2$ ($\sigma^* 1s)^2$ ($\sigma 2s)^2$ (of and isoelect ctron. of and isoelectr ctron.	d isoelec σ2p _z) ² (π σ*2s) ² (ronic with	tronic with O ₂ . $\tau 2p^{2_{x}} = \pi 2p^{2_{y}}$) $(\sigma 2p_{z})^{2} (\pi 2p^{2_{x}})$ th N ₂ : $(\sigma 1s)^{2}$ (σN_{2} : $(\sigma 1s)^{2}$ (σ)	$(\pi^* 2p_{x^1} = \pi^* 2p_{y^1})$ = $\pi 2p^{2_y}$) ($\pi^* 2$ $\sigma^* 1s$) ² ($\sigma 2s$)($\sigma 2s$) ² ($\sigma 2s$) ²	$(\sigma^{*2}s)^{2}$ and 2 unpatrix $(\sigma^{*2}s)^{2} = \pi^{*2}p^{2}y)^{2}$ $(\sigma^{*2}s)^{2} (\pi^{2}p^{2})^{2}$ $(\sigma^{*2}s)^{2} (\pi^{2}p^{2})^{2}$	aired electror and no unp $f_x = \pi 2p^{2_y}$) (o $f_x = \pi 2p^{2_y}$) (o	ns. aired 52pz) ² 52pz) ²
I-12. Sol.		O 2	O ₂ - O ₂	^{2–} O ₂ +					

Bond Order 2 1.5 1 2.5

I-13.

Sol.		Bond order	Unpaired electron
	O ₂ +	2.5	1
	NO	2.5	1
	N ₂ +	2.5	1

I-14. Sol. Molecule is unstable if number of bonding molecular orbital electrons (Nb) is less than the number of antibonding molecular orbital electrons (Na). i.e., molecule has negative bond order.
;fn ca/kh vk.kfod d{kdksa esa bys])V^akWuksa dh la[:k (Nb)] izfrca/kh vk.kfod d{kdksa eas bys])V^akWuksa dh la[:k (Nb)] izfrca/kh vk.kfod d{kdksa eas bys])V^akWuksa dh la[:k (Nb)]

Section (J) : Dipole Moment

- J-1. Sol. On account of higher electronegativity of fluorine.
- J-2. Sol. On account of larger difference in the electronegativity values of oxygen and hydrogen.
- **J-3.** Sol. Dipole moment \propto bond angle

P.

J-4. Sol.

 $\mu \neq 0$; SiF₄, BF₃ and PF₅ are symmetrical molecules thus $\mu = 0$.

J-5. Sol. 0

Symmetrical molecule has zero dipole moment.



J-6. Sol.

; Hence hybridisation of CI is sp³ and molecule is polar.

J-7. Sol.

NF₃ μ = 0.80 × 10–30 Cm.

NH₃ µ = 4.90 × 10−30 Cm.

J-8. Sol. H_2O , $\mu = 6.17 \times 10^{-30}$ Cm; NH_3 , $\mu = 4.90 \times 10^{-30}$ Cm;

NF₃, $\mu = 0.80 \times 10^{-30}$ Cm; CH₄, $\mu = zero$.

; Tetrahedral symmetrical structure and, therefore, $\mu = 0$.



 $\mu = 0$

J-10. Sol.

Sol.

J-9.

J-11. Sol. Dipole moment = $4.8 \times 10^{-10} \times 1.275 \times 10^{-8} = 4.8 \times 1.275 \times 10^{-18} = 4.8 \times 1.275 \text{ D}$ % ionic character = $\frac{1.03 \times 100}{1.275 \times 4.8} \approx 17\%$

Section (K) : Hydrogen bonding

- **K-1.** Sol. In pure phosphoric acid the PO₄³⁻ groups are bonded together through many hydrogen bonds.
- K-2. Sol. HF is least volatile (i.e. high boiling point) due to strong intermolecular hydrogen bonding.



K-3. Sol. It has intramolecular H-bonding

- **K-4.** Sol. In CH₄, H₂Se and H₂S the central atoms are not more electronegative ; hence do not form hydrogen bonds with itself and other molecule. In N₂H₄ the nitrogen is more electronegative and thus is able to form hydrogen bond with itself and other molecules like water.
- **K-5. Sol.** Inter molecular hydrogen bond is present in NH₃, H₂O and HF. H₂S does not form inter molecular hydrogen bonding because the electronnegativity of sulphur is low.
- K-6. Sol. HI does not form hydrogen bonding owing to less electronegativity value of iodine.



K-8. Sol. Due to polar nature of H₂O there is association of water molecules giving a liquid state of abnormally high boiling point.

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K-7.

Section (L) : Intermolecular forces

- L-1. Sol. All are non-polar molecules but Xe has the greatest Van der Waal's forces because it has highest molar mass (i.e. highest London dispersion force).
- **L-2. Sol.** Among ICI, H₂S and CO, each of the molecule has a dipole-dipole attraction in addition to London dispersion force. Only Br₂ does not have dipole-dipole attraction and London disperion force is most significant in determining the melting and boiling point.
- **L-3.** Sol. London disperion force depends on molar mass. The strength of London disperion force, therefore, increases with increasing molar mass. As a consequence, Cl₂ is gas, Br₂ is liquid and l₂ is solid.
- **L-4.** Sol. Graphite has layered structure. Layers are held by van der Waal's forces and distance between two layers is 340 pm. Each layer is composed of planar hexagonal rings of carbon atoms. C C bond length within the layer is 141.5 pm Each carbon atom in hexagonal ring undergoes sp² hybridisation and make three sigma bonds with three neighbouring carbon atoms. Fourth electron forms a π bond. The electrons are delocalised over the whole sheet.
- L-5. Sol. van der Waal's forces ∝ molecular weight.So AgBr will have maximum v.
- **L-6. Sol.** London forces are extremely short range in action and the weakest of all attractive forces. The order of strength of bonds/ forces is ionic bond > covalent bond > hydrogen bond > london force.
- **L-7. Sol.** The boiling points of H₂O, SO₂ and CIF₃ can be explained on the basis of dipole-dipole attraction between their respective molecules. London dispersion force exists among the non-polar molecules like H₂, O₂, Br₂ etc. in solid or liquid states. Even in atoms in molecules which have no permanent dipole, instantaneous dipoles will arise as a result of momentary inbalances in electron distribution.

Section (M) : Metallic bond

- **M-1. Sol.** The strength of metallic bonds depends upon the number of mobile electron(s) per atom. Sodium has only one mobile electron per atom where as iron has 8 mobile electrons per atom.
- M-2._ Sol. Electron sea model of metallic bonding.

Exercise-2

Marked Questions may have for Revision Questions.

OBJECTIVE QUESTIONS

1. Sol. When two atoms combine to form a stable molecule, there is attraction between the mobile electrons and kernel resulting into liberation of energy.

- **2. Sol.** A stable molecule has lower potential energy which results from attraction between the mobile electrons and kernel leading to the liberation of energy.
- **3. Sol.** Low ionisation potential of one atom favours the formation of cation and high electron affinity of the other atom favours the formation of anion.
- 4. Sol. As oxidation state of central atom increases, the polarising power increases and, therefore, covalent character of bond increases. Hence S F in SF₂ has least covalent character owing to lowest polarising power of S²⁺.
- **5. Sol.** According to Fajan's rules polarisation of bond \propto Size of cation Polarisation of bond \propto covalent character of bond.
- 6. Sol. As anion is same, the covalent character in ionic compound will depend upon the size and charge on the cation. According to Fajan's rule, PbO is most ionic because Pb²⁺ has less polarising power where as Mn₂O₇ is most covalent because Mn⁷⁺ has higher polarising power.
- 7. Sol. (1) BrF₅ contains 10 electrons inplace of eight. (2) SF₆ contains 12 electrons inplace of eight.
 (3) IF₇ contains 14 electrons inplace of eight.
- 8. Sol. Factual according to Fajan's Rule.
- **9. Sol.** As size of cation decreases, the polarising power increases and therefore, the tendency to attract more number of water molecules increase. So correct order is Na > K > Rb > Cs.



; it has only $\sigma \& \pi$ bonds. (;g dsoy $\sigma vk \le \pi ca/k \le krkg \le$)





Charge on cation



13. Sol. $N \equiv N \rightarrow$

Sol.

11.

- **14. Sol.** Formal charge on each B-atom = 3 4 = -1. Formal charge on each N-atom = 5 - 4 = +1.
- **15. Sol.** Formal charge $= 5 \frac{1}{2}(8) = +1$

Total number of bonds between atoms

16. Sol. Bond order = Total number of resonating structure = $\overline{4}$ = 1.25

$$\mathsf{H} - \mathsf{N}^{\scriptscriptstyle +} - \mathsf{N}^{\scriptscriptstyle +} = \mathsf{N}^{\scriptscriptstyle 2-}$$

- **17. Sol.** (II) ; there is positive charge on two adjacent nitrogen atoms. This leads to repulsion and thus increases the energy of the molecule.
- **18. Sol.** C–C : 1.54 Å ; C = C : 1.34 Å ; C = C : 1.20 Å.





31. Sol. In the hydrides of 16th group elements H–E–H bond angle decreases down the group as electronegativity of the central element decreases. Further it is suggested that in case of H₂S, sp³

hybridisation does not occur and S–H bond is formed by the overlapping of p-atomic orbitals of sulphur with s-orbitals of hydrogen atoms. The p-atomic orbitals are right angle to one another.

32.

Sol.

33. Sol. (1) According to VSEPR theory as electronegativity of central atom decreases, bond angle decreases. So bond angle of $H_2O > H_2S > H_2$ Se > H_2Te

(2) $C_2H_2 > C_2H_4 > CH_4 > NH_3$. In NH₃ there is bp-lp repulsion so bond angle decreases to 107° from 109.5°.

sp sp² sp³ sp³

180° 120° 109.5° 107°

(3) SF₆ < NH₃ < H₂O < OF₂ in this case bond angle of NH₃ is highest because Ip - Ip repulsion is absent in. it.

(4)
$$CIO_2 > H_2O > H_2S > SF_6$$

Note : It is supposed that in H₂S the hybrid orbitals do not participate in bonding but pure p-atomic orbitals participate in bonding.

34.

(1)

Sol.

Bond angle is 180° because of sp hybridisation of nitrogen.

- $(3) \qquad \begin{array}{c} 0 \\ 111^{\circ} \\ 111^{\circ} \\ 0 \\ 0 \\ 134^{\circ} \\ \end{array}$
- 35. Sol. (2) B₂H₆ is a electron deficient compound as boron has only six electron in its valence shell
 (3) AICl₃ is a electron deficient compound as aluminium has only six electron in its valence shell.

36. Sol. (3) $O_2 : (\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma^* 2s)^2 (\sigma^* 2s)^2 (\pi 2pz)^2 (\pi 2p2x = \pi 2p2y) (\pi^* 2px1 = \pi^* 2p1y)$

37.

42.

43.

Sol.

 O_2^- : $(\sigma_1^*s)^2 (\sigma_2^*s)^2 (\sigma_2^*s)^2$

Sol. M.O for
$$C_2 = \sigma_1 s^2 < \sigma_1 * s^2 < \sigma_2 s^2 < \sigma_2 * s^2 < Homo$$

It is important to note that double bond in C_2 consists of both pi bonds because of the presence of four

electrons in two pi molecular orbitals $C_{2^{2-}} \left[C \xrightarrow{\frac{\pi}{6}} C \right]^{-1}$

38. Sol. Bond energy \propto Bond order.

> $\frac{0-4}{2} = 3.0 \text{ ; } N_{2^{-}} \text{ : } B.O. = \frac{10-5}{2} = 2.5 \text{ ; } N_{2^{+}} \text{ : } B.O. = \frac{9-4}{2} = 2.5. \text{ ; } N_{2^{2-}} \text{ : } B.O. = \frac{10-6}{2}$ 10 - 4= 2.0. N₂ : B.O. =

- 39. Sol. Bond strength α bond order. $O_2^{2^2} = 1.0$; $O_2 = 2.0$; $O_2^{-} = 1.5$; $O_2^{+} = 2.5$. So, the correct order of bond strength is $O_2^+ > O_2 > O_2^- > O_2^{2-}$
- Sol. **Nitrogen molecule** (N_2) : $(\sigma_1s)^2$ $(\sigma_2s)^2$ $(\sigma_2s)^2$ $(\sigma_2s)^2$ $(\pi_2p^2x = \pi_2p^2y)$ $(\sigma_2pz)^2$ 40. The bond order of N₂ is 1/2(10 - 4) = 3. It contains one sigma and two π bonds.
- 2 0Number of electrons in HeH⁺ = 2 + 0 = 0; σ 1s² σ * 1s². So, B.O. = 2 = 1 41. Sol.

(En = 2.0)

; So dipole moment of NH₃ is greater than NF₃.

 $(\mu = 4.95)$

En = 4.0)

 $(\mu = 6.02)$

(iii)

(iv) $(SiH_3)_2O = 0.24 D$ (Bond angle $144.1 \pm 0.9^\circ$; $p\pi$ -d π bonding SiH_3OCH_3 = 1.166 D; bond angle decrease. (Source : R.Verma)

Orthonitrophenol has lower boiling point (i.e. more volatile) because it exists as discrete molecules than its para-derivative, where association of molecules takes place using intermolecular H-bonding.

52. Sol. When ice is formed from liquid water, The tetrahedral structure around each oxygen atom with two regular bonds to hydrogen and two hydrogen bonds to other molecules requires a very open structure with large spaces between ice molecules.

 $D_{H_2O(s)} < D_{H_2O(l)}$. $V_{H_2O(s)} > V_{H_2O(l)}$

53. Sol. Due to intermolecular H-bonding NH₃ molecule come close to each other and exist in liquid state.

- **56. Sol.** $NH_3 = 238.5 \text{ K}$; $PH_3 = 185.5$; $H_2O = 373 \text{ K}$; $AsH_3 = 210.6 \text{ K}$; $SbH_3 = 254.6 \text{ K}$.
- **57. Sol.** ICI being polar in nature has higher dipole-dipole intreraction whereas Br₂ being non-polar has only weaker London dispersion force.
- **58. Sol.** Benzene (C₆H₆) being non-polar has only london dispersion forces on account of unsymmetrical distribution of electron at any instant of time producing dipoles.
- **59.** Sol. Boiling point of SbH₃ is greater than NH₃. The higher boiling point of SbH₃ is attributed to higher van der Waal forces because of its higher molecular weight.

60.

Sol. Steric no. of $IO_2F_2^- = 4 + 1 = 5$, sp³d; Steric no. of $CIOF_3 = 4 + 1 = 5$, sp³d;

Steric no. of $F_2SeO = 3 + 1 = 4$, sp^3 ; Steric no. of $XeF_{5^+} = 5 + 1 = 6$, sp^3d^2

- **61. Sol.** F Bond dipoles of N—F bonds are counter balanced to some extent by the dipole moment of lone pair of electron acting in opposite direction. This reduces both the dipole moment and its donor powers.
- 62. Sol. Assertion and Reason both are correct statement and Reason is the correct statement of Assertion, e.g., NO₂⁺ and I₃⁻ have different hybridisation but on account of stability they have linear shape as given below.

63. Sol. Both are correct and Reason is the correct explanation of Assertion. In $[Al(H_2O)_6]^{3+}$, aluminium is in +3 oxidation state. So,

- **64. Sol.** In NO₃⁻, N is sp² hybridised , therefore , NO₃⁻ ion has trigonal planar geometry. PO₃⁻ units are linked together through P O P bonds to form either linear or cyclic structures. In these polymeric structures, the P–atom is sp³ hybridised and thus each PO₃⁻ unit has tetrahedral shape.
- 66. Sol. Both are correct statements but Reason is not the true explanation of Assertion.
- **67. Sol.** Both statements are true but **Reason** is not the true explanation of **Assertion**. As in these species the central atoms have more than an octet of electrons i.e., expanded octet)
- **68.** Sol. It is true σ is stronger while π is a weak bond but atoms rotate freely about π bond is false.
- **69. Sol.** Both S-1 and S-2 are true statements but are different and reason is not true explanation of assertion.

High electronegativity of atoms attached to carbon atom indirectly results into the contraction of the size of p-orbitals. As carbon atomic size is smaller and thus it has effective overlapping.

- **70.** Sol. Bond order \uparrow , Bond length \downarrow , Bond energy \uparrow .
- **71.** Sol. N_2 molecule is di-magnetic because N_2 molecule does not contain unpaired electron.
- 72. Sol. Assertion is true, Reason is true; Reason is the correct explanation of Assertion.

No such hydrogen bonding is present in ethers.

3. Sol. (1) CF₄ is tetrahedral where is T-shaped.

, O = C = O both linear according to VSEPR.

- (3) BF₃ is trigonal planar and PCl₃ is tetrahedral.
- (4) PF_5 is trigonal bipyramidal and IF_5 is square pyramidal.

Sol. 4.

(2)

- sp³

5. Sol. The correct order of bond angle (smallest first) is $H_2S < NH_3 < SiH_4 < BF_3$ 92.6° < 107° < 109°28' < 120°

 $N = \frac{sp^2}{O}$

(3) bent, $\mu \neq 0$ bent, $\mu \neq 0$

6. Sol. NO and NO⁺ are derivative of O₂.

NO(isoelectronic with O_{2^+}) : $(\sigma_1 s)^2 (\sigma_1 s)^2 (\sigma_2 s)^2 (\sigma_2 s)^2 (\sigma_2 p_z)^2 (\pi_2 p_z^2 = \pi_2 p_y^2) (\pi_2 p_x^2 = \pi_2 p_y^2)$ Bond order = 1/2(10 - 5) = 2.5.

 $\begin{aligned} \text{NO}^{*}(\text{isoelectronic with } O_{2}^{2+}) : (\sigma 1s)^{2} \ (\sigma^{*}1s)^{2} \ (\sigma^{*}2s)^{2} \ (\sigma^{*}2s)^{2} \ (\sigma 2p_{z})^{2} \ (\pi 2p^{2}_{x} = \pi 2p^{2}_{y}) \ (\pi^{*}2p_{x} = \pi^{*}2p_{y}) \\ \text{Bond order} = 1/2(10-4) = 3. \end{aligned}$

Bond order \propto 1/bond length.

So NO⁺ has shorter bond length.

7. Sol.

For planar BO₃ groups, the B–O bond length is usually close to 1.36 Å but for tetrahedral BO₄ groups the length increases to about 1.48 Å. This suggests that in the planar grouping π -boding involving lone pairs of electrons from the oxygen atoms occurs; this π -boding is necessarily lost in the tetrahedral group, in which a lone pair from the extra oxygen atom occupies the previously empty orbital on the boron atom.

9.

13.

Sol.

	$CH_3 - CO - CH_3$	$\overset{sp^3}{CH_3}-\overset{sp^2}{COOH}$	$CH_3^{sp^3} - C \equiv N;$	$CH_3 - CONH_2$
Sol.	(1) Acetone	(2) Acetic acid	(3) Acetonitrile	(4) Acetamide

10. Sol. The value of lattice energy depends on the charges present on the two ions and the distance between them.

Lattice energy $\propto r_+ + r_-$ where $(r_+ + r_-) =$ Inter-ionic Distance.

Lattice energy $\alpha Z_+, Z_-$

 $Z_* \Rightarrow$ charge on cation in terms electronic charge.

 $Z_{-} \Rightarrow$ charge on anion in terms electronic charge.

- 11. Sol. In case of halides of alkali metals, melting point decreases going down the group because lattice enthalpies decreases as size of alkali metal increases. But LiCl has lower melting point in comparison to NaCl due to covalent nature. Thus, NaCl is expected to have the highest melting point among given halides.
- 12. Sol. According to VSEPR theory

 $H_2 \rightarrow (1s)^2$, $\sigma^* (1s)^0$, no unpaired electron.

 $H_{2^+} \rightarrow \sigma (1s)^2$, $\sigma^* (1s)^0$, one unpaired electron.

 $H_{2^-} \rightarrow \sigma (1s)^2$, $\sigma^* (1s)^1$, one unpaired electron.

- **14.** Sol. Calcium carbide is ionic carbide having [: $C \equiv C$:]²⁻
- **15. Sol.** The electronic configuration of oxygen atom is $1s^2$, $2s^2$, $2p^4$. Each oxygen atom has 8 electrons. Hence in each $O_2^{2^-}$ ion there are 16 + 2 = 18 electrons. The electronic configuration of $O_2^{2^-}$ ion is

Here KK represents non-bonding molecular orbital of 1s orbital. O_2^{2-} contains no unpaired electrons. The electronic configuration of B_2 molecule is

 $\begin{array}{l} \mathsf{KK}\;\sigma(2s)^2\;\sigma^*(2s)^2\;\sigma(2p_z)^2\;\pi(2p_x)^1\;\pi(2p_z)^1\\ \text{it contains 2 unpaired electrons. The electronic configuration of N_2^+ ion is}\\ \mathsf{KK}\;\sigma(2s)^2\;\sigma^*(2s)^2\;\sigma(2p_z)^2\;\pi(2p_x)^2\;\sigma(2p_y)^1\\ \text{it contains one unpaired electron.}\\ \text{The electronic configuration of O_2 molecule is} \end{array}$

KK $\sigma(2s)^2 \sigma^*(2s)^2 \pi(2p_x)^2 \pi(2p_y)^2 \sigma(2p_z)^2 \pi^*(2p_x)^1 \pi^*(2p_y)^1$

$$\mathbf{Ca}^{2+} \left[: \mathbf{C} \stackrel{1\sigma}{\underset{2\pi}{\equiv}} \mathbf{C} : \right]^{2-}$$

17.

Sol.

It contains 2 unpaired electron.

16. Sol. The elements of group 14 show an oxidation state of +4 and +2. The compounds showing an oxidation state of +4 forms covalent compound and have tetrahedral structures. e.g. SnCl₄, PbCl₄, SiCl₄, etc. whereas those which show +2 oxidation state are ionic in nature and behave as reducing agent. e.g. SnCl₂, PbCl₂ etc. Further as we move down the group the tendency of the element to form covalent compound decreases but the tendency to form ionic compound increases.

18. Sol. Because of highest electronegativity of F, hydrogen bonding in F – H - - - - F is strongest.

19. Sol. High charge and small size of the cations increases polarisation. As the size of the given cations decreases as

$$K^+ > Ca^{2+} > Mg^{2+} > Be^{2+}$$

Hence, polarising power decreases as $K^+ < Ca^{2+} < Mg^{2+} < Be^{2+}$

20. Sol. Molecular orbital configuration of

$$O_2 \qquad \Rightarrow \qquad \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^1 \pi^* 2p_y^1$$

⇒ Paramagnetic
$$\frac{10-6}{2}$$

Bond order =
$$2 = 2$$

 $O_{2^{+}} \quad \Rightarrow \quad \sigma 1s^{2}\sigma^{*} \ 1s^{2}\sigma \ 2s^{2}\sigma^{*} \ 2s^{2}\sigma \ 2p^{2}_{z} \ \pi 2p_{x}^{2} \ \pi 2p_{y}^{2} \ \pi^{*} \ 2p^{1}_{x}$

Bond order = 2 = 2.5

 $N_2 \implies \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p^{2_x} \pi 2p^{2_y} \sigma 2p^{2_z}$

 \Rightarrow

 \Rightarrow

10 - 4

Bond order = 2 = 3

 $N_{2^{+}} \quad \Rightarrow \quad \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p^{2_x} \pi 2p^{2_y} \sigma 2p^{1_z}$

$$9 - 4$$

Bond order = 2 = 2.5

 $C_2 \qquad \Rightarrow \qquad \sigma 1s^2 \ \sigma^* \ 1s^2 \ \sigma 2s^2 \ \sigma^* \ 2s^2 \ \pi 2p^{2_x} \ \pi 2p^{2_y}$

$$\Rightarrow$$
 diamagnetic

2 = 2Bond order = σ1s² σ* 1s² σ2s² σ* 2s² π2p²_x π2p¹_y C_2^+ \Rightarrow \Rightarrow Paramagnetic 7 – 4 Bond order = 2 = 1.5 \Rightarrow σ1s² σ* 1s²σ 2s² σ* 2s²σ 2p²_z π2p²_x π2p²_y π* 2p¹_x NO \Rightarrow Paramagnetic 10-5 Bond order = 2 = 2.5 \Rightarrow NO⁺ $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p^2_z \pi 2p^2_x \pi 2p^2_y$ \Rightarrow diamagnetic 10 - 42 = 3Bond order =

8 - 4

21.

Sol.

Bond order =
$$\frac{10-4}{2} = 3.$$

 CN^- is derivative of N_2 ; CN^- (isoelectronic with N_2) $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p^{2_x} \pi 2p^{2_y} \sigma 2p^{2_z}$

NO⁺ is derivative of O₂; NO⁺ (isoelectronic with O₂²⁺) σ 1s² σ * 1s² σ 2s² σ * 2s² σ 2p²_z π 2p²_x π 2p²_y

Bond order =
$$\frac{10-4}{2}$$
 = 3

22.

Decrease in B – F bond length which results in the higher bond dissociation energy of B – F in BF₃ is due to delocalised $p\pi$ – $p\pi$ bonding between filled p-orbital of F atom and vacant p-orbital of B atom.

23. Sol.

Sol.

Bond Order			Bond Order			
		10 – 5			10-7	
(1)	O ₂ +	2 = 2.5	(2)	O ₂ -	2 = 1.5	
		10-8			10-4	
(3)	O ₂ ²⁻	2 = 1	(4)	O ₂ ²⁺	$\frac{1}{2} = 3$	

24. Sol. Covalent character in ionic compounds is governed by Fazan's Rule. AlCl₃ will show Maximum covalent character on account of higher polarising power of Al³⁺ because of its having higher positive charge and smaller size.

25.

Sol. NO₂⁺ Number of electron pairs = 2 Number of bond pairs = 2

→ sp³

Number of lone pair = 0

So, the species is tetrahedral with sp³ hybridisation.

The structure is pentagonal bipyramid having sp³d³ hybridisation as given below : 26. Sol.

;

$$F_{b} \xrightarrow{90^{\circ}} I \xrightarrow{72^{\circ}} F_{b}$$

$$F_{b} \xrightarrow{F_{a}} F_{b}$$

 $F_{b} - I - F_{b} = 72^{\circ}$ (5 number)

F⊳ – I –	Fa =	90°	(10	number)	

 $F_b - I$ bond length = 1.858 ± 0.004 Å

 $F_a - I$ bond length = 1.786 ± 0.007 Å.

27. Calcium carbide exists as Ca^{2+} and C_{2}^{2-} . According to the molecular orbital model, C_{2}^{2-} should Sol. have a bond order of 3 (configuration $\pi_u^2 \pi_u^2 \sigma_{9}^2$). MOT configuration suggests that it contains one σ and two

;

 π -bonds $\left[:C_{\frac{\pi}{2}} C:\right]^{2^{-}}$

28.

Sol.

36.

Page | 26

Sol. It is a simple & popular fact.

;g ,d lkek[]; o fo[;kr dkjd g{A

37.

- **Sol.** Dipole-dipole interaction(Hydrogen bonding) $\propto \overline{r^3}$
- **38. Sol.** London dispersion forces increase from He to Xe because molecular mass increases. So boiling point will also increase from He to Xe.

1

39.

Sol. $NO_{2}^{-} = sp^{2}$ $NO_{3}^{-} = sp^{2}$ $NO_{2}^{-} = sp^{2}$

 $NO_2^+ = sp$

40.

Sol. There is extensive intermolecular hydrogen bonding in the condensed phase.

41.

- Sol. (a) $ZnO + Na_2O \rightarrow Na_2 ZnO_2$ Acidic oxide Basic oxide (b) $ZnO + CO_2 - \rightarrow ZnCO_3$ Basic oxide Acidic oxide So ZnO behave like acid in equation (a) and base in equation (b)
- 42. Sol. No of electron in CO = 6 + 8 = 14 (i) CO $\rightarrow \sigma 1S^2$, $\sigma^* 1S^2$, $\sigma 2S^2$, $\sigma^* 2S^2$, $\left[\pi 2P_x^2 = \pi 2P_y^2\right]\sigma 2P_z^2$ All electrons are paired so diamagnetic (ii) O₂ $\rightarrow \sigma 1S^2$, $\sigma^* 1S^2$, $\sigma 2S^2$, $\sigma^* 2S^2$, $\sigma 2P_z^2 \left[\pi 2P_x^2 = \pi 2P_y^2\right]$, $\left[\pi^* 2P_x^1 = \pi^* 2P_y^1\right]$ Unpaired electron = 2 (Paramagnetic) (iii) B₂ $\rightarrow \sigma 1S^2$, $\sigma^* 1S^2$, $\sigma 2S^2$, $\sigma^* 2S^2 \left[\pi 2P_x^1 = \pi 2P_y^1\right]$ (Paramagnetic) (iv) NO $\rightarrow \sigma 1S^2$, $\sigma^* 1S^2$, $\sigma 2S^2$, $\sigma^* 2S^2$, $\sigma 2P_z^2 \left[\pi 2P_x^2 = \pi 2P_y^2\right]$, $\left[\pi^* 2P_x^1 = \pi^* 2P_y^0\right]$ (Paramagnetic)

ONLINE QUESTION

7. Sol. Bonding molecular orbital results in increased electron density between nuclei due to constructive interference of combining electron waves.

8. Sol.

sp³d² hybridized with one position occupied by lone pair.

- 9. Sol. Fact.
- **10. Sol.** H₄P₂O₆ has P–P linkage O O || || HO–P–P–OH

- **11. Sol.** 1, 3 and 4 suggests that valency of X is +3. So, formula of chloride is XCl₃.
- 12. Sol. $1D = 10^{-18}$ esu cm $\delta = \frac{0.38 \times 10^{-18}}{1.617 \times 10^{-8} \times 4.8 \times 10^{-10}} = 0.0485 \ 0.05$
- **13. Sol.** CIF₃, XeOF₂ & ^{XeF₃+} are sp³d hybridized with 2 lone pair e's, hence all have (T-shape) identical shape.

$$\begin{array}{cccc} & & & & & & \\ \hline F-C \vdash F & F-Xe \vdash F & F-Xe^{+} - F \\ I & I \\ F & O & F \end{array}$$

14. Sol. Instantaneous dipole-induced dipole forces are most responsible in allowing xenon gas to liquify.

15.

Sol. Fact

PART - II : JEE (ADVANCED) / IIT-JEE PROBLEMS (PREVIOUS YEARS)

* Marked Questions may have more than one correct option.

As both have same shapes they are also isostructural (same hybridisation and no lone pair of electrons).

2. Sol. The dipole moment is vector sum of the bond dipoles. The order of dipole moment is $CH_3CI > CH_2CI_2 > CHCI_3 > CCI_4$.

 $CCI_4 = 0$, $CHCI_3 = 1.03$ D, $CH_2CI_2 = 1.62$ D, $CH_3CI = 1.86$ D

$$H_{2}C \xrightarrow{\pi} CH - C \equiv N$$

$$sp^{2} \xrightarrow{\sigma} sp^{2} \qquad sp sp$$

3. Sol. In From left to right.

Note : Hybrid orbitals always form σ -bonds due to overlapping on their axis.

4. Sol. According to VSEPR theory, total number of electron pairs = 6.

total number of bond pairs = 5.

so total number of lone pair = 1.

There is one Xe—O double bond. The π -electrons of double bond create more repulsion than single covalent bond. To minimize the repulsions the lone pair and Xe—O double bond are trans to each other in octahedral geometry.

$$\begin{array}{c} F \\ F \\ F \\ s p^{3} d^{2} 0 \end{array} F$$

5. Sol. It is peroxodisulphuric acid, the Marshall's acid and contains one – O – O – linkage as given in structure

- 6. Sol. Molecular orbital electronic configuration is $\sigma_{1s^2} \sigma_{1s^2} \sigma_{2s^2} \sigma_{2s^2} \sigma_{2p_z^2} \sigma_{2p_z^2} \sigma_{2p_y^2} \sigma_{2p_y$
- 7. Sol. According to VSEPR theory

The number of electron pairs around I = 5

The number of bond pairs around I = 2

The number of lone pairs around I = 3

To have minimum repulsions among lone pair and bond pairs, the lone pairs acquire the three equatorial positions of trigonal bipyramidal as given in the structure,

ClO₃⁻, XeF₄ and SF₄ have 1, 2 and 1 lone pair of electrons respectively.

Sol. (A) In two dimensional sheet silicates, three oxygen atoms of [SiO₄]^{4–} are shared with adjacent tetrahedral [SiO₄]^{4–} units,

9.

10.

- (B) One oxygen atom is shared between two adjacent tetrahedra. [SiO₄]⁴⁻.
- (C) All four oxygen atoms are shared between adjacent tetrahedra, [SiO4]4-.
- (D) Two oxygen atoms are shared between adjacent tetrahedra, [SiO4]^{4–} forming a linear chain.
- Bond order = $\frac{N_b N_a}{2}$ Sol. $CO = {\sigma_{1s^2} \sigma_{1s^2}^* \sigma_{2s^2}^* \sigma_{2s^2}^* \pi_{2p_x^2} \pi_{2p_y^2}^* \sigma_{2p_y^2}^* \sigma_{2p_y^2$ Bond order of CO = 10 - 4/2 = 3 $NO^{-} \equiv {\sigma_{1s^{2}} \sigma_{1s^{2}}^{*} \sigma_{2s^{2}}^{*} \sigma_{2s^{2}}^{*} \sigma_{2p_{z}^{2}}^{*} \pi_{2p_{x}^{2}}^{*} \pi_{2p_{y}^{2}}^{*} \pi_{2p_{x}^{1}}^{*} \pi_{2p_{y}^{1}}^{*} \pi_{2p_{y}^{1}}^{*}}$ Bond order of NO⁻ = 10 - 6 / 2 = 2 $NO^{+} \equiv {}^{\sigma_{1s^{2}} \sigma_{1s^{2}} \sigma_{2s^{2}} \sigma_{2s^{2}} \sigma_{2s^{2}} \sigma_{2p_{z}^{2}} \pi_{2p_{x}^{2}} \pi_{2p_{y}^{2}}}$ Bond order of NO⁺ = 10 - 4 / 2 = 3 $\mathbf{CN}^{-} \equiv {}^{\sigma_{1s^{2}} \sigma_{1s^{2}} \sigma_{1s^{2}} \sigma_{2s^{2}} \sigma_{2s^{2}} \sigma_{2s^{2}} \pi_{2p_{x}^{2}} \pi_{2p_{y}^{2}} \sigma_{2p_{z}^{2}}}$ Bond order of $CN^{-} = 10 - 4 / 2 = 3$ $N_{2} \equiv \sigma_{1s^{2}} \sigma_{1s^{2}} \sigma_{2s^{2}} \sigma_{2s^{2}} \sigma_{2s^{2}} \sigma_{2p_{x}^{2}} \sigma_{2p_{y}^{2}} \sigma_{2p_{z}^{2}}$
- Bond order of $N_2 = 10 4 / 2 = 3$.
- Sol. KO_2 exists as $K^+ \& O_2^-$. In O_2^- , superoxide ion there are total number of electrons = 16 + 1 = 17. MOT configuration = $\sigma_{1s^2} \sigma_{1s^2} \sigma_{2s^2} \sigma_{2s^2} \sigma_{2p_z^2} \sigma_{2p_z^2} \sigma_{2p_x^2} \sigma_{2p_y^2} \sigma_{2p_x^2} \sigma_{2p_x$ $O_{2^{-}}$ has one unpaired electron in antibonding $\pi^{*}2p_{y^{1}}$. So it is paramagnetic.
- 11. Steric number = 4; thus sp^3 hybridisation in P₄. As each phosphorus is sp^3 , so Sol. % p character will be = $4 \times 100 = 75$.
- 12. Sol. Due to small size of B, it is very difficult to remove the electrons from boron to form ionic bond as it will require very high energy. On the other hand, due to its very small size having high polarising power causes greater polarisation and eventually significant covalent character according to Fajan's rule.
- 13. Sol. Statement-1: Germanium is a semiconductor, where the energy gap between adjacent bands is sufficiently small for thermal energy to be able to promote a small number of electrons from the full valence band to the empty conduction band. This leaves both bands partially filled, so the material can conduct electricity.

Statement-2: Incorrect statement.

14. Sol. B2; total number of electrons = 10. The MOT electron configuration violating the Hund's rule will be thus :

$$\sigma_{1s}^{\ 2} \ \sigma_{1s}^{* \ 2} \ \sigma_{2s}^{2} \ \sigma_{2s}^{* \ 2} \ \pi_{2p_x}^{\ 2} \ = \pi_{2p_y}^{\ 0} \ \sigma_{2p_z}^{\ 0}$$

$$\frac{6-4}{2}$$

So, bond order = 2 = 1As all electrons are paired, the molecule is diamagnetic.

CHEMICAL BONDING

15. Sol. XeO₂F₂ has trigonal bipyramidal geometry. Due to presence of lone pair on equitorial position, the shape is see-saw.

If 2s-2p mixing is not operative, the increasing order of Molecular orbitals will be :

16. Sol.

 $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2$

Considering this $Be_2 \& B_2$ become diamagnetic, so does N_2 . Only C_2 would be paramagnatic with electronic configuration as above

17. Ans. (C)

Sol.

 $-\, {\scriptstyle \rightarrow}\,$ It is d–d axial overlap in same phase, so d–d σ bonding.

 $-\,{\scriptstyle\rightarrow}\,$ It is p & d lateral overlap in same phase, so it is p–d π bonding.

 $-\,{\scriptstyle\rightarrow}\,$ It is p and d lateral overlap in opposite phase, so it is p–d π antibonding.

- \rightarrow It is d–d axial overlap in opposite phase, so it is d–d σ antibonding.


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square planar
(All Xe-F bonds
```


are equal)

[C=C bond is different from C-H bond]

tetrahedral (All Si-F bonds are equal)

CI 10. Ċĺ

Sol. CI–O–CI bond angle is more due to large size of CI and F–O–F bond angle is least due to bent rule. CI-O-CI CI ds cM+s vkdkj ds dkj.k vf/kd gS rFkk F-O-F csUV fu;e ls de gSA

11. Sol. $O_{2^{-}}$ has one unpaired electron in $\pi^{*}2p_{y}$ while other do not have unpaired electrons. (explained by MOT)

Sol.

12. Sol. BF₃ has triangular planar with no lone pair.

- 13. Sol. Strongest H-bonding is in F H ----- O as F is most electronegative element hence creating good partial postive charge on H. While O is good donor of electron as compare to F due to use electronegativity.
- 14. Sol. It is fact.
- **15.** Sol. Oxygen has unpaired in $\pi^* 2p_x$ and $\pi^* 2p_y$ which is explain its paramagnetic behaviour. (Explained by MOT)

17. Sol.

18. Sol. Diborane : (B₂H₆) :

Diamond has sp³ hybrid carbon and graphite has sp² hybrid carbon.

19. Sol. NH_{4^+} , $H_{3}O^+$ and OF_2 all have sp^3 hybridisation and their shapes are tetrahedral, trigonal pyramidal and bent respectively.

It have both covalent as well as coordinate bond.

21.

20.

Sol.

Sol.

 $\begin{array}{ccc} & \operatorname{NH}_3 & \operatorname{NH}_4^+ \\ (\operatorname{sp}^3) & _{+} \operatorname{H}^+ & \longrightarrow (\operatorname{sp}^3) \end{array}$ Trigonal pyramidal Tetrahedral

(Bond angle $\approx 107^{\circ}$) (Bond angle $\approx 109.5^{\circ}$)

22. Sol. Polarisability of anion increase on increasing its size.

- 23. Sol. In PCI₅, P has more change then P in PCI₃ therefore CI is more polarized in PCI₅.
- **24. Sol.** For the formation of ionic compound atom should have low ionization energy to form cation while other should have high electron affinity to form anion.
- **26. Sol.** Covalent character ionic compound is explained by Fajan rule. Covalent character increase on decreasing the size of cation and increasing charge on cation.

28.

30.

Sol.

29. Sol. $O_2 \rightarrow [KK] \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^1 \pi^* 2p_y^1$

 $O_{2^{-}} \rightarrow \text{ [KK] } \sigma 2s^2 \ \sigma^* 2s^2 \ \sigma 2p_{z^2} \ \pi 2p_{x^2} \ \pi 2p_{y^2} \ \pi^* 2p_{x^2} \ \pi^* 2p_{y^1}$

As in O₂ HOMO is π^*2p_x and π^*2p_y and they have one electron each so next electron can be added to thus orbitals.

PART - II : PRACTICE QUESTIONS

Practice Questions: 20-50 depending on chapter length.

1. Sol. Element X has atomic number 7 so its electronic configuration will be 1s² 2s² 2p³, so it's lewis

```
diagram would be
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2. Sol. The correct Lewis diagram is :O::C::O:

9. Sol. s-orbital +
$$p$$
-orbital $\rightarrow 0$
12. Sol. (1) sp (2) p (bent) (3) sp (4) sp (4) sp
13.

Sol. (square planar)

14.

15.

Sol.

19. Sol. (1) Tetrahedral and see-saw shaped.

- (2) Both are sp³ hybridised and trigonal pyramid.
- (3) Both are sp³ hybridised and tetrahedral.
- (4) Both are sp³d² hybridised and octahedral.

(II) H_2O_2 is covalent compound.

(III) $(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\pi 2p_x^1 = \pi 2p_y^1) (\sigma p_z)^0$

The highest occupied molecular orbital is π molecular orbital.

(IV) The repulsive interaction of electron pairs decreases in the order :

lone pair (ℓp) - lone pair (ℓp) > lone pair (ℓp) - bond pair (bp) > bond pair (bp) - bond pair (bp)