SOLUTIONS

Module - 1 / JEE-2022

Chemistry	Chapter - 4B	Chemical Bonding - II

EXERCISE-A

1.(C) In KCN, ionic bond is present between K^+ and CN^- and covalent bonds are present between carbon and nitrogen $C \equiv N$.

2.(D) Octlet of $BeCl_2$ is not complete.

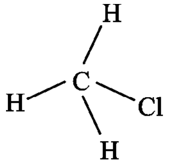
3.(A) X^+Y^-. Because Electropositive elements forms cation and electronegative elements forms anion.

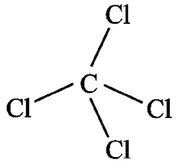
4.(A) H_2 $H-H$ **5.(C)** $N \equiv N$ $:N\vdots\vdots N:$

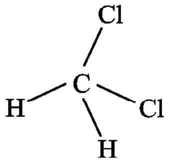
6.(C) Dipole moment is vector quantity.
In trigonal planar geometry (for sp^2 hybridisation), the vector sum of two bond moments is equal and oposite to the dipole moment of third bond. **7.(A)**

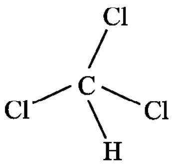
8.(B) (i) Dipole moment is vector quantity. When vector sum of all dipoles in molecule will be zero, then molecule will not have net dipole moment.

(ii) For net dipole moment to be equal to zero, all the atoms attached to central atom must be identical and geometry must be regular.


Methyl chloride


Carbon tetrachloride


Methylene chloride


Chloroform

∴ Carbon tetrachloride having regular geometry and identical atoms attached to bonds has zero dipole moment.

9.(B) In regular tetrahedral structure, dipole moment of one bond is cancelled by opposite dipole moment of the other bonds.

10.(D) Li^+ : 76 pm ; Na^+ : 102 pm ; Mg^{2+} : 72 pm ; Be^{2+} : 31 pm

11.(B) Hybridisation is process of redistribution of energy among orbitals of almost same energy level. The new orbitals formed are of equal energy and the number of new orbitals formed is equal to number of participating orbitals. The new orbitals present are at maximum distance from each other, so as to minimize the repulsion between electrons.

12.(D) In covalent bonds between two identical non-metal atoms share the pair of electrons equally between them, e. g F_2, O_2, N_2.

13.(C)

14.(C) From PH_3, NH_3 and SbH_3 are all sp^3 hybridised. Their central atom has both bond pair as well as lone pair electrons. The lone pair occupy the fourth orbital. CH_3^+ has only three pairs of electrons so it is sp^2 hybridised.

15.(A) Compound having sp hybridisation will bave linear shape.

∴ CO_2 or $(O = C = O)$ has C in sp hybrid state and has linear shape.

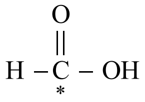
16.(B) Dipole moment of compound having regular geometry and same type of atoms is zro. It is vector quantity. The zero dipole moment of BF_3 is due to its symmetrical (triangular planar) structure. The three fluorine atoms lie at the corners of an equilateral triangle with boron at the centre.

Thus the vectorial addition of the dipole moments of the three bonds gives a net sum of zero because the resultant of any two dipole monents is equal and opposite to the third. The dipole moment of NH_3 is 1.46 D indicating its unsymmetrical structure. The dipole moment of CH_2Cl_2 (the molecule uses sp^3 hybridisation but not symmetric) is 1.57 D

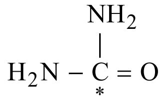
17.(A) Molecule having sp^3 hybridisation and one lone pair of electron will have pyramidal structure.

(i) CO_3^{2-} and NO_3^{2-} have tetrahedron structure.

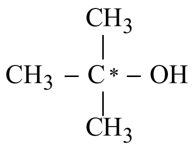
(ii) In PCl_3, P is sp^3 hybridised and has one lone pair of electrons, hence pyramidal in shape.

18.(C) **(A)** 

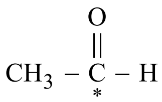
$H-\overset{\overset{O}{\|}}{\underset{*}{C}}-OH$

(B) 

$H_2N-\overset{\overset{NH_2}{|}}{\underset{*}{C}}=O$

(C) 

$CH_3-\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{C^*}}-OH$

(D) 

$CH_3-\overset{\overset{O}{\|}}{\underset{*}{C}}-H$

(A) $3\sigma, 1\pi$ **(B)** $3\sigma, 1\pi$ **(C)** 4σ **(D)** $3\sigma, 1\pi$

$\therefore$ $(CH_3)_3$ COH has 4σ bonds and sp^3 hybridisation.

19.(A)

20.(B) Similar atoms are linked to each other and thus there is no polarity.

21.(C) (i) Non metallic oxides are more covalent (or less ionic) as compared to metallic oxides.

(ii) Higher the polarising power of cation (higher oxidation state of similar size cation's) more will be covalent character

(i) P_2O_5 will be more covalent than other metallic oxides.

(ii) Oxidation state of Mn is +7 in Mn_2O_7, oxidation state of Cr in CrO_3 is +6 and oxidation state of Mn is +2 in MnO.

$\therefore$ MnO is most ionic.

P_2O_5, being a non-metallic oxide will definitely be more covalent than the other metallic oxides. Further, we know that higher the polarising power of the cation (highter oxidation state of the similar size cations) more will be the covalent charachter. Hence Mn is in +7 O.S in Mn_2O_7, Cr is + 6 in CrO_3 and Mn is + 2 in MnO. So MnO is the most ionic and Mn_2O_7 is the most covalent.

22.(D) O_2 – Oxygen (z = 8) has molecular obtial configuration of O_2.

$(\sigma 1s^2)\ (\sigma^* 1s^2)\ (\sigma 2s^2)\ (\sigma^* 2s^2)\ (\sigma 2p_z^2)\ (\pi 2p_x^2)\ (\pi 2p_y^2)(\pi^* 2\pi_x^1)\ (\pi^* 2\pi_y^1)$ i.e., 2 unpaired electrons and 14 paired.

23.(C) Structure of a molecule can be ascertained by knowing the number of hybrid bonds in the molecule. Thus

In NF_3, N is sp^3 hybridized as 4 orbitals are involved in bonding.

In NO_3^-, N is sp^2 hybridized as 3 orbitals are involved in bonding.

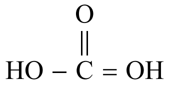
In BF_3, B is sp^2 hybridized as 3 orbitals are involved in bonding.

In H_3O^+, O is sp^3 hybridized as 4 orbitals are involved in bonding.

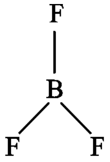
Thus, isostructural pairs are $[NF_3, H_3O^+]$ and $[NO_3^-, BF_3]$

24.(A) (i) Polar bond is formed when electronegativity difference between bonded molecule is more than 1.7.

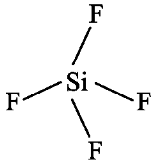
(ii) 4σ bond – sp^3 hybridisation, 3σ and 1π bond – sp^2 hybridsation and 2σ and 2π bonds – sp hybridisation.



$HO-\overset{\overset{O}{\|}}{C}=OH$ C is sp^2 hybridised and molecule is polar.

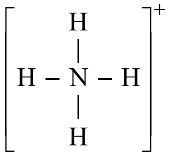


B is sp^2 hybridised and molecule is non polar.



Si is sp^3 hybridised and molecule is non polar.

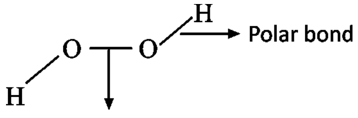
25(C) (i)



$\left[H-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{N}}-H\right]^+$ $[Cl]^-$: It has ionic and polar covalent bond.

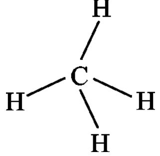
(ii) $H-C\equiv N$: It has polar covalent bond.

(iii)



It has plar, non polar both types of covalent bond.

(iv)



It has non polar covalent bonds only.

26.(D) Critical temperature of water is higher than O_2 because H_2O molecule has dipole moment which is due to its V-shape.

27.(B)

28.(A) Central O-atom cannot have more than 8 electrons (octet).

29.(C) LiCl is a covalent compound since due to the large size of the anion (Cl^-) its effective nuclear charge lessens and its valence shells are held less tightly towards its nucleus.

30.(D) (i) Bond length $\propto \dfrac{1}{\text{Bond order}}$

(ii) Bond order is calculate by either the help of molecular orbital theory or by resonance.

(i) Bond order of CO as calculated by molecular orbital theory = 3

(ii) Bond order of $CO_2 = \dfrac{\text{Number of bonds in all possible sides}}{\text{Number of resonating structure}} = \dfrac{4}{2} = 2$

(iii) Bond order in CO_3^{2-} (by resonance method) $= \dfrac{4}{3} = 1.33$

$\therefore$ Order of bond of C – O is $CO < CO_2 < CO_3^{2-}$

31.(A) In H_2S, S has sp^3 hybridisation and 2 lone pair of electrons is H_2S.

$\therefore$ It has angular geometry and non-zero value of dipole moment.

32.(D) The structure of species can be predicted on the basis of hybridisation which in turn can be known by knowing the number of hybrid orbitals in that species.

For SF_4 : S is sp^3 hybridised in SF_4. Thus SF_4 has 5 hybrid orbitals of which only four are used by F, leaving one lone pair of electrons on sulphur (valence electron)

For CF_4 : C is sp^3 hybridised in CF_4. Since all the four orbitals of carbon are involved in bond formation, no lone pair is present on C having four valence elecrtrons.

For XeF_4 : Xe is sp^2d^2 hybridised in XeF_4. Of the six hybrid orbitals, four form bond with F, leaving behind two pair of electrons on Xe.

33.(B) For NO_2^+ : sp hybridisation For NO_3^- : sp^2 hybridisation For NO_4^+ : sp^3 hybridisation

34.(B) Hybridisation of N in NH^3 is sp^3 Hybridisation of Pt in $[PtC_4]^{2-}$ is dsp^2

Hybridisation of P in PCl_5 is sp^3d Hybridisation of B in BCl_3 is sp^2

35.(A) $H_3N \rightarrow BF_3$ where both N, B, are attaining tetrahedral geometry.

36.(B) Write configuration of all species. Half filled and fully filled orbitals are more stable as compared to nearly half filled and nearly fully filled orbitals.

$Li^- = 1s^2, 2s^2 : Be^- = 1s^2, 2s^2, 2p^1$ $B^- = 1s^2, 2s^2, 2p^2 ; C^- \; 1s^2, 2s^2, 2p^3$

$\therefore$ Be^- will be least stable.

37.(A) Isoelectronic species have same number of electrons and isostructural species have same hybridisation at central atom.

NO_3^- ; Number of $e^- = 7 + 8 \times 1 = 32$, hybridisation of N is sp^3

CO_3^{2-} ; Number of $e^- = 6 + 8 \times 3 + 2 = 32$, hybridisation of C is sp^3

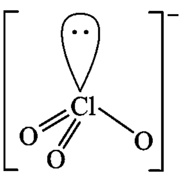
ClO_3^- ; Number of $e^- = 17 + 8 \times 3 + 1 = 42$, hybridisation of Cl is sp^3

SO_3 ; Number of $e^- = 16 + 8 \times 3 = 36$, hybridisation of S is sp^2

$\therefore$ NO_3^- and CO_3^{2-} are isostructural and isoelectronic.

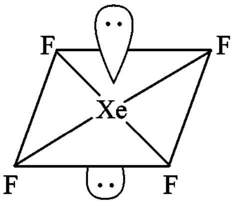
38.(D)

ClO_3^-
(1 lone pair)



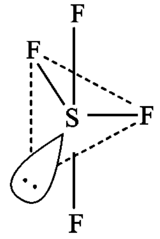
Pyramidal

XeF_4
(2 lone pair)



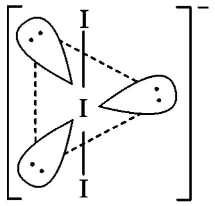
Square planar

SF_4
(1 lone pair)



Irregular tetrahedral

I_3^-
(3 lone pair)



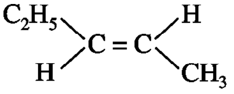
Linear

39.(AC) CO_2, $HgCl_2$ and C_2H_2 have linear structure (sp hybridization) while $SnCl_2$ is trigonal planar (sp^2 hybridisation). NO_2 has angular structure (V-shape).

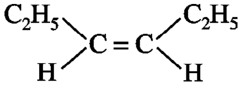
40.(BCD) $[O = N = O]^+ ; [N \equiv C - O]^- ; S = C = S$

It can be seen from the structure above that CS_2 being sp hybridized has a linear shape and other two molecules are isoelectronic to CS_2, so they are also linear.

41.(BC) Alkanes don’t have dipole moment because electronegativity difference between carbon and hydrogen in zero. These alkenes are not symmetrical and so they have dipole moment.



trans 2-pentene



cis 3-hexene

42.(BD) Isoelectronic species have same number of electrons.

Number of e^- in $CH_3^+ = 6 + 3 - 1 = 8$ Number of e^- in $H_3O^+ = 3 + 8 - 1 = 10$

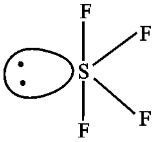
Number of e^- in $NH_3^- = 7 + 3 = 10$ Number of e^- in $CH_3– = 6 + 3 + 1 = 10$

$\therefore$ H_3O^+, NH_3 and CH_3^- are isoelectronic

43.(B) Boron, in BF_3, is sp^2 hybridised leading to trigonal planar shape

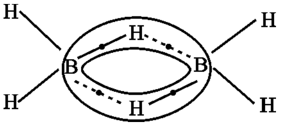
44.(C)

45.(A) SF_4 is sp^3d hybridised and here two different FXF angles are present.



46.(B)

47.(C)



48.(C) Hybridisation of N in NH_2^- is sp^3 and has shape similar to H_2O

49.(D) Hybridisation of S is sp^3 in SO_4^{2-}

Hybridisation of Cl is sp^3 in ClO_4^-

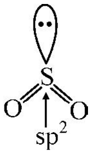
50.(A) NH_3 is the most basic compound out of NH_3, PH_3, AsH_3, SbH_3.

Basic strength decreases from NH_3 to SbH_3 $NH_3 > PH_3 > AsH_3 > SbH_3$.

51. (C) Higher the charge ⇒ higher force of attraction between the ions ⇒ more difficult to break the lattice ⇒ Ans has to be BaO or MgO and clearly MgO will have more packed structure as compared to BaO (as there is large difference in size of Ba and O)

52. (B) Higher the charge on cation, more is the covalent character is.

53. (B)



$O = C = O$ sp

$N \equiv N \rightarrow O$ sp

$C \equiv O$ sp

EXERCISE - B

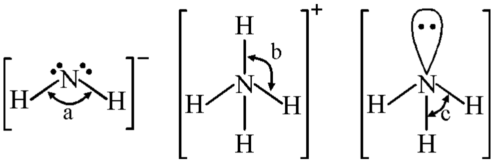
1. (A) Intramolecular H – bonding is possible in II, which will lead to its solubility ; ⇒ ‘II’ will be present more in vapour form as compared to ‘I’. (Read consequences of Intramolecular H-bonding).

2. (C) CH_4 will have perfect tetrahedal geometry.

$CHCl_3$ will have tetrahedral geometry (but not perfect, value to the presence of different type of atoms around carbon)

SF_4 has bent-see-saw shape (based on pentagonal bipyramidal geometry)

3. (B)



$a \approx 104.5°$; $b = 109.5°$; $c \approx 107°$ $\Rightarrow$ $(a < c < b)$

4. (C) As Bond strength decrease, size of atoms increase $\Rightarrow$ $Cl_2 > F_2 > I_2$

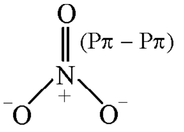
Note : F_2 is an exception. It happens due to very small size of the F atoms, which leads to repulsions between electrons (lone pairs) of the the two F atoms making F - F bond and thus weakening it.

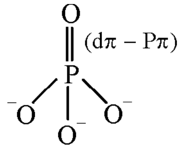
5. (B) HI has stronger london dipersion force than HBr due to its large molar mass. This results in high boiling point of HI wrt to HBr. Same is applicable to HBr compared to HCl.

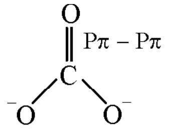
HF is an anomaly here due to presence of hydrogen bonding. (HF > HI > HBr > HCl)

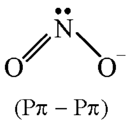
Note : London dispersion involes the force the creation of an instantaneous dipole on an atom due to uneven distribution of electrons of the atom at any particular instant. The more the electrons an atom has, the greater the strength of the instantaneous dipole, so the london force would be stronger on compounds with heavier elements.

6. (B)









(check your self)

$N \equiv s^2p^3$

$[N^+] = s^2p^2 \equiv s^2p_y^1p_x^1$

$\equiv s^1p_y^1p_x^1p_z^1$

$\equiv (sp^2)^1(sp^2)^1(sp^2)^1p_y^1$

$P = s^2p_y^1p_x^1p_z^1$

$P \equiv s^1p_x^1p_y^1p_z^1d_4^1$

$P \equiv (sp^3)^1(sp^3)^1(sp^3)^1d_A^1$

7.(B) It forms hydrogen bonds with water

8.(D) (i) Number of valence electrons of the two oxygen in O_2^{2-} ion = 6 × 2 + 2 = 14. ∴ Number of unpaired electrons = O, hence diamagnetic.

(ii) Number of valence electrons of all atoms in $O_3 = 6 \times 3 = 18$.
Thus, it also does not have any unpaired electron, hence diamagnetic.

(iii) Number of valence electrons of all atom in $N_2O = 2 \times 5 + 6 = 16$. Hence, here also all electrons are paired.

(iv) Number of valence electrons of all atoms in $O_2^- = 2 \times 6 + 1 = 13$.
Thus it has one unpaired electron, hence paramagnetic.

9.(AC) The outer most shell of C, N and O has 4, 5 and 6 electrons respectively. Thus CN^- and NO^+ each has 10 electrons to accommodate in the molecular orbitals. So their bond order is same. O_2^- has 13 and CN^+ has 8 electrons in outermost orbits.

10.(A)

11.(C)

Species	**B.O.**
O_2^-	1.5
O_2^{2-}	1.0
O_2^{2+}	3
O_2^+	2.5

$$B.O \propto \frac{1}{\text{bond length}}$$

12.(A) Isolectronic species have same number of electrons.

Electrons in CO = 6 + 8 = 14 Electrons in CN^- = 6 + 7 + 1 = 14

Electrons in O_2^- = 8 + 8 + 1 = 17 Electrons in O_2^+ = 8 + 8 – 1 = 15

∴ CO and CN^- are isoelectronic

13.(D) Hydrogen bonding is formed in the compound having F or O or N attached to hydrogen atom.

∵ HCl does not have F or N or O ∴ It does not form hydrogen bond.

14.(A) The electronic configureation of each species according to molecular orbital theory.

NO (7 + 8 = 15e) : $(\sigma 1s^2)\,(\sigma^* 1s^2)\,(\sigma 2s^2)\,(\sigma^* 2s^2)\,(\sigma 2p_z^{\,2})\,(\pi 2p_x^{\,2})\,(\pi 2p_y^{\,2})\,(\pi^* 2\pi_x^{\,1})$ unpaired electron

CO (6 + 8 = 14e) : $(\sigma 1s^2)\,(\sigma^* 1s^2)\,(\sigma^* 2s^2)\,(\pi 2p_x^{\,2})\,(\pi 2p_x^{\,2})\,(\pi 2p_y^{\,2})\,(\sigma 2p_z^{\,2})$

CN^- (6 + 7 + 1 = 14e) : similar to that of CO

O2(8 + 8 = 16e) : $\sigma 1p^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2\,(\sigma p_z^{\,2})\,(\pi 2p_x^{\,2})\,(\pi 2p_y^{\,2})\,(\sigma^* 2p_x^{\,1})\,(\sigma^* 2p_y^{\,1})$

15.(C) Greater the difference between electronegativities of two covalently bonded atoms more will be strength of hydrogen bond.

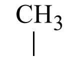
∴ F – H .... F bond is strongest due to largest difference in electronegativity of atoms and smallest size of F atom.

16.(A) Compounds having F or O or N attached to H form hydrogen bond. The hydrogen bond between F and H is strongest due to highest difference in electronegativity between F and H.

CH_3CH_2OH
Ethanol

$C_2H_5 - O - C_2H_5$
Diethylether

CH_3CH_2Cl
Ethylchloride

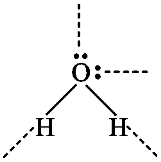


$$CH_3 - \overset{\overset{\displaystyle CH_3}{|}}{N} - CH_3$$

Trimethylamine

∴ Ethanol having H attached to O atom will form hydrogen bonds. Rest of the compounds do not form hydrogen bonds.

17.(B) H_2O molecule can form four hydrogen bonds per molecule, two via lone pairs and two via hydrogen atoms.



18.(B) Calcium cabide is an ionic compound $(Ca^{2+}C^{2-})$ which produces acetylene on reacting with water. Thus structure of C^{2-} is $[C \equiv C]^{2-}$. It has one σ and two π bonds.

19.(C) In KO_2^-, O_2 has one unpaired electron. As for O_2, there are two such unpaired electrons. In other cases, electrons are paired.

20.(A) Each of the species has 14 electrons which are distributed in MOs as :

$\sigma 1s^2, \sigma^* 1s2, \sigma^* 2s^2, \pi 2p_x^{\,2}, \pi 2p_y^{\,2}, \sigma 2p_z^{\,2}$

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

21.(C) $N_2(7 + 7 = 14) = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_y^{\,2}, \pi 2p_y^{\,2}, \sigma 2p_z^{\,2}$

$F_2\,(9 + 9 = 18) = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^{\,2}, \pi 2p_x^{\,2}, \pi 2p_y^{\,2}, \pi^* 2p_x^{\,2}, \pi^* 2p_y^{\,2}$

$O_2^-\,(8 + 8 + 1 = 17) = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^{\,2}, \pi 2p_x^{\,2}, \pi 2p_y^{\,2}, \pi^* 2p_x^{\,2}, \pi^* p_y^{\,1}$

$O_2^{2-}\,(8 + 8 + 2 = 18) = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^{\,2}, \pi 2p_x^{\,2}, \pi 2p_y^{\,2}, \pi^* 2p_x^{\,2}, \pi^* p_y^{\,2}$

∴ O_2^- is the only species having unpaired electron.

22.(B) O_2 : Bond order = $= \frac{10-6}{2} = 2$ (two unpaired electrons in antibonding molecular orbital)

O_2^+: Bond order $= \frac{10-5}{2} = 2.5$ (One unpaired electron in antibonding molecular orbital). Hence O_2 as well as O_2^+ is paramagnetic, and bond order of O_2^+ is greater than that O_2.

23.(C) Argon being an inert gas process only Vander Waal's interaction in solid state.

24.(A) He exhibit Vander Waal's forces while HCl, NH_3 and H_2O exhibit dipole-dipole interaction. Vander Waal's forces are the weakest interaction.

25.(C) Vander Waal's forces of interaction are much weaker than metallic, ionic and covalent bonds.

26.(D) Pure covalent molecules in crystals are held closely by Vander Waal's forces.

27.(D) Polarizability $\propto$ Size of the species
Down the group size increases and Xenon being the largest possesses the highest polarizability.

28.(C) I_2 and CCl_4 both are non polar in nature. Thus can only exhibit London dispersion forces.

29.(C) NaCl upon dissolving in water dissociate into Na^+ and Cl^- ions. These ions are surrounded by H_2O molecules with ion-dipole interaction between them.

30.(B) Non-polar molecules such as O_2, N_2 and noble gases exhibit London dispersion forces.

31.(D) H-bond is a type of weak electrostatic interaction between an electronegative atom and a hydrogen atom connected to another electronegative atom.

32.(A) In ice each water molecule is linked to other four water molecules through hydrogen bonding due to which density of ice is less than that of water heating ice at 0°C to 4°C, it gets converted to liquid water thus volume decreases.

MISCELLANEOUS EXERCISE

Chemical Bonding - II

1.(D) For N_2 & below :

$\sigma ls \sigma^* ls \sigma 2s \sigma^* 2s \pi 2P_x = \pi 2P_y \sigma 2P_z \pi^* 2P_x = \pi^* 2P_y \sigma^* 2P_z$

B.O of $C_2 = \frac{n_b - n_a}{2} = \frac{8-4}{2} = 2$; B.O of $C_2^{2\ominus} = \frac{n_b - n_a}{2} = \frac{10-4}{2} = 3$

B. O of $B_2 = \frac{n_b - n_a}{2} \frac{6-4}{2} = 1$; B.O of $B_2^{\oplus} = \frac{n_b - n_a}{2} = \frac{5-4}{2} = \frac{1}{2}$

B. O of $N_2 = \frac{n_b - n_a}{2} = \frac{10-4}{2} = 3$; B. O of $N_2^{\oplus} = \frac{n_b - n_a}{2} = \frac{9-4}{2} = \frac{5}{2}$

for O_2 & above : $\sigma ls \sigma^* 13 \sigma 2s \sigma^* 2s \sigma 2P_z \pi 2P_x = \pi 2P_y \pi^* 2P_x = \pi^* 2P_y \sigma^* 2P_z$

B.O of $O_2 = \frac{n_b - n_a}{2} = \frac{10-6}{2} = 2$; B.O of $O_2^{\ominus} = \frac{n_b - n_a}{2} = \frac{10-7}{2} = \frac{3}{2}$

Higher the bond order shorter the bond length

So answer is (D) B.O of $O_2 >$ B.O of $O_2^{\ominus}$

$\therefore$ Bond length of $O_2^- >$ Bond length of O_2.

2.(A) From MOT, we can write :

for B_2 : $\sigma ls^2 \sigma^* ls^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2P_x^2 = \pi 2P_y$ {as Hund's rule is violated} ; So B.O $= \frac{n_b - n_a}{2} = \frac{6-4}{2} = 1$

and B_2 has no unpaired e^-, so it is diamagnetic.

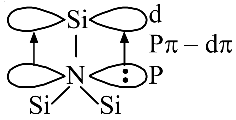
3.(C) For peroxide ion $\left(O_2^-\right)$, from MOT we can write :

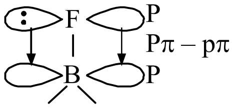
for $O_2^{2\ominus}$: $\sigma ls^2 \sigma^* ls^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2P_z^2 \pi 2P_x^2 = \pi 2P_x^2 \pi^* 2P_x^2 = \pi^* 2P_y^2$

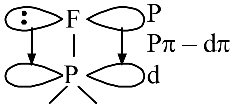
B.O $= \frac{n_b - n_a}{2} = \frac{10-8}{2} = 1$

4.(A) $\pi^* 2P_x$ has two nodal planes being an antibonding molecular orbital.

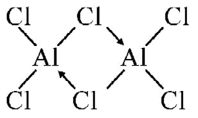
5.(BCD) In trisilyamine N is sp^2 hybridized, with lone pair of e^- in unhybridised P orbital oriented parallel with the vacant orbital of silicon, so a side wise overlap between filled unhybrid P orbital of N and vocant of orbital of Si results in back bonding. Similar phenomenon is observed in B-F and P-F bond.

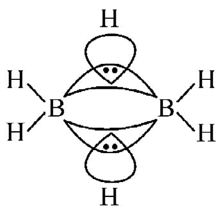






6.(AC) $AlCl_3$ and BH_3 exist as dimer, so as to reduce their Lewis acidic character.





Banana Bond {3c – 2e bond}

7.(ABCD)

PCl$_5$ exists as $[PCl_4]^{\oplus}$ $[PCl_6]^{\ominus}$ in solid state N$_2$O$_5$ exists as $[NO_2]^{\oplus}[NO_3]^{\ominus}$ in solid state.

$[PCl_4]^{\oplus}$ $\Rightarrow$ tetrahedral (sp^3) $[NO_3]^{\ominus}$ $\Rightarrow$ octahedral(sp^{3}d^2)

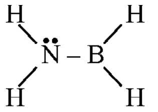
where as PCl$_5$ in gaseous phase is sp$_3$d, Trigonal bipyeamidal

$[NO_2]^{\oplus}$ $\Rightarrow$ Inear (sp) $[NO_3]^{\ominus}$ $\Rightarrow$ Trigonal planar (sp^2)

where as N$_2$O$_5$ in gaseous phase each N is sp2, Trigonal planar.

8.(ABCD)

In case of filled P orbital of an otom is oriented parallel with vacant P orbital of an atom, then Pπ - Pπ bond is possible.



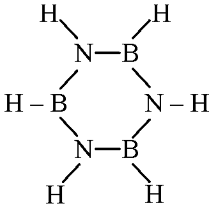
$:C = \ddot{O}:$

$H - \overset{\oplus}{C} = \ddot{O}:$

filled p orbital of N is parallel with vacant P orbital of B

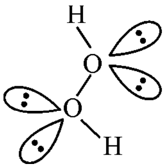
filled p orbital of O is parallel with vacant P orbital of C

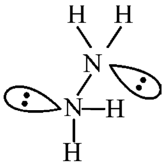
filled p orbital of O is parallel with vacant P orbital of C



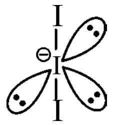
filled P orbital of N is parallel with vacant P orbital of S.

9.(AB)





10.(C)

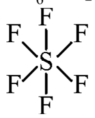


No. of lone pair in $I_3^{\ominus}$ is 3.

11.(ABC)(A), (B), (C) are planar. In (s) Two π bonds are ⊥ to each other, so the H atoms of and first and last carbon are out of plane.

12.(D) As discused in solution of Q. 11, carbons atoms are in angle of 120° and H atoms are perpendicular to that of C atoms

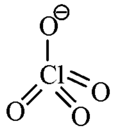
13.(B) SF$_6$ is sp^{3}d^2 octahedral

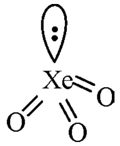


So angle is 90^c

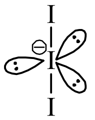
$\% \text{ d} - \text{character} = \frac{2}{6} \times 100 = 33\%$

14.(B)

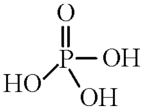




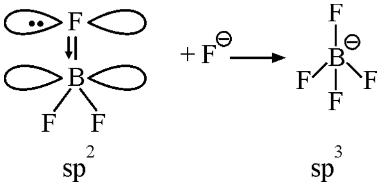
15.(D) $I_3^{\ominus}$ has 3 lone pairs.



16.(B) Higher the extent of H-bonding stronger the intermolecular force and higher viscosity



17.(ACD)



18.(ABCD) $PF_3 > PH_3$ F is more electronegative than O

$Cl_2O > F_2O$ F is more electronegative than O

$(SiH_3)_2O > (CH_3)_2O$ because pπ-dπ bond between, O and Si.

$NCl_3 > NF_3$ F is more electronegative than N.

Higher the electronegativity of atom joined with the central atom, lesser the repulsion between bond pairs of e^-, lesser the bond angle.

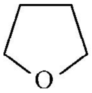
19.(C) $H_3Si - \ddot{N} = C = S \longrightarrow H_3Si \leftrightarrows \underset{180^\circ}{N} = C = S$

a possible back bonding between p orbital of N & vacant d orbital of Si.

20.(AC) In solid state PCl_5 exists as $[PCl_4]^+[PCl_6]^-$.

So answer is $sp^3(A), sp^3d^2(C)$

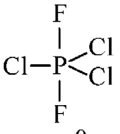
21.(BCD) H-bonding is development of attraction between $H^{\delta+}$ and $X^{\delta-}$ {on electronegative element carrying delta negative}.
Et – OH, H – OH, Et – NH_2 and H – F shows H – bonding.

- No – H bond between Et_2O and  (THF)
- H-bonding between Et–OH and H–OH
- H-bonding between Et–NH_2 and Et_2O
- H-bonding between Et_2O and HF

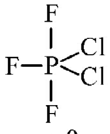
22.(AB) Higher the extent of intermolecular H-bonding, Higher the viscosity, solubility in water and boiling point.

23.(D) In P_4, each P is sp^3 hybridized, so P character is 75%

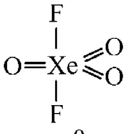
24.(BD)



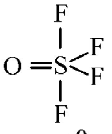
μ = 0



μ ≠ 0

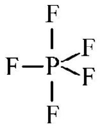


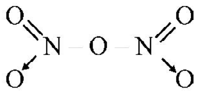
μ = 0



μ ≠ 0

25.(BD) In BF_3 and NO_3^- due to resonance bond lengths are same





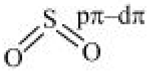
In PF_5 three equitorial bonds are equivalent and two axial bonds are equivalent.

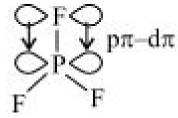
26.(A) I, II are more stable contributing resonating structure

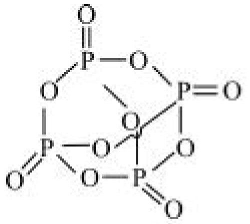
27.(D) In PF_3 possible back bonding makes the P–F bond aquire double bond character.

28.(B) In vinyl chloride lone pair of Cl is involved in resonance with C = C, so C – Cl bond acquires double bond character.

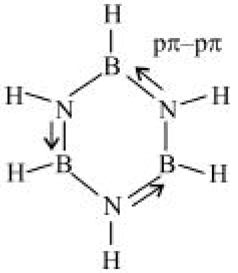
29.(D)







P = O is dπ-pπ



30.(D) Due to absence of lone pair on central atom due to back bonding.

31.(A) Dipole moment of $CH_3 - Cl$ is more than that of $CH_3 - F$ due to longer C – Cl bond length.
So answer is (A).

32. (A) → (p) (B) → (s) (C) → (r) (D) → (q)

33. (A) → (r) (B) → (p) (C) → (q) (D) → (s)

Higher the bond order shorter the bond length.

34. (A) → (p, q, r, t) (B) → (q, r, s, t) (C) → (p, q, r) (D) → (p, q, r, s) (E) → (q, r, s, t)

35. (A) → (p, t) (B) → (p, q, t) (C) → (s, t) (D) → (r, t) (E) → (p, t)

36. (A) → (r) (B) → (s) (C) → (p) (D) → (q) (E) → (t)

37. (A) → (r) (B) → (s) (C) → (q) (D) → (t) (E) → (p)

38. (A) → (r) (B) → (s) (C) → (p) (D) → (q)

39. (A) → (r) (B) → (p) (C) → (s) (D) → (q)

40. (A) → (q) (B) → (r) (C) → (p) (D) → (s)

41.(6) Linear species : BeF_2, ICl_2^-, I_3^-, XeF_2, NO_2^+, CO_2

Bent species : ICl_2^+, $SnCl_2$, H_2S, NO_2^-, ClO_2^-, BrF_2^+

42.(5) Bond length $\alpha \dfrac{1}{\text{Bond order}}$

Transformation	Change in bond order	Change in bond length
$CO \longrightarrow CO^+$	3.0 to 3.5	decreases
$NO \longrightarrow NO^+$	2.5 to 3.0	decreases
$NO_2 \longrightarrow NO_2^+$	1.5 to 2.0	decreases
$O_2 \longrightarrow O_2^+$	2.0 to 2.5	decreases
$H_2 \longrightarrow H_2^-$	1.0 to 0.5	increases
$N_2 \longrightarrow N_2^+$	3.0 to 2.5	increases
$B_2 \longrightarrow B_2^+$	1.0 to 0.5	increases
$C_2 \longrightarrow C_2^{2-}$	2.0 to 3.0	decreases
$C_2 \longrightarrow C_2^+$	2.0 to 1.5	increases

43.(5)

Species	Structure	polar/non-polar
BeFCl	Unsymmetrical linear	polar
BeF_4^{2-}	Tetrahedral	Non-polar
BrF_3	T-shape	polar
PCl_3F_2	Trigonal bipyramidal	Non-polar
XeO_3	Pyramidal	polar
XeF_6	Monocapped octahedron	polar
XeF_4	Square planar	Non-polar
XeO_4	Tetrahedral	Non-polar
I_3^-	Linear	Non-polar