4. Chemical Bonding and Molecular Structure – Answers and Solutions

1. Electrovalent Bonding

1. (c) Electrovalent compounds are good conductor of heat and electricity in molten state or in aqueous solution.

2. (a) The order of the bond strength (strongest to weakest) is Ionic bond > Covalent bond > Coordinate bond > Metallic bond > Hydrogen bond.

3. (a) *Cs* is more electropositive.

4. (a) The ionic character of the molecule/atom is determined by the influence of covalent bond, as the size of cations increases, ionic character increases as per the given order of covalency i.e.

$CaCl_2 > KCl > BaCl_2 > AgCl$.

In $AgCl$, cation is smaller with least ionic character.

5. (c) $CaCl_2$ will have electrovalent bonding because calcium is electropositive metal while chlorine is electronegative so they will combine with electrovalent bond.

6. (b) As soon as the electronegativity increases, ionic bond strength increases.

7. (a) When electronegativity difference is from 1.7 to 3.0, the bond formed between them will be ionic bond.

8. (a) Lithium oxide and calcium fluoride show ionic characters.

9. (a) *Br-F* has maximum electronegativity difference compare to another compound.

10. (d) $M = 2,8,3$; $A = 2,8,7$

$M^{3+}A^{-} \equiv MA_3$.

11. (b) No. of valence electrons in *F* are 7. Thus, in order to have 6 valence electrons, it must lose one electron and will have the formula F^{+}.

12. (b) Sugar is an organic compound which is covalently bonded so in water it remains as free molecules.

13. (d) When sodium chloride is dissolved in water, the sodium ion is hydrated.

14. (c) Interionic attraction is due to opposite charges.

15. (a) Ions with low charge can easily gain or lose electrons easily. A large cation can lose electrons and a small anion can gain electrons easily.

16. (d) BecausOe of high lattice energy, they are insoluble in water.

17. (c) Directional means bonded atoms prefer specific orientations in space relative to one another. In ionic bonds there is no electron sharing, charge is uniform in all direction, so do not show directional properties.

18. (a) By definition.

19. (d) Because the electronegativity difference is less than 2.1 between *H* and *Cl* ; *HCl* is not an ionic compound.

20. (b) As stability is directly related to lattice energy and lattice energy depends on charge and size of ions. So, the order is :

$Fe_2O_3 < Cr_2O_3 < Al_2O_3 < MgO$.

21. (b) Electron deficient $-B_2H_6$

Saline $-LiH$

Electron precise $-CH_4$

Interstitial $-CrH$

Electron rich $-NH_3$.

2. Covalent Bonding

1. (b) Covalent bond shows directional properties and has a definite geometry.

2. (a) In graphite all carbon atoms are sp^2-hybridised and have covalent bond.

3. (d) $(+1) + x + 3(-2) = 0 \Rightarrow 1 + x - 6 = 0 \Rightarrow x = 6 - 1 = 5$.

4. (a) PCl_5 does not follow octet rule, it has 10 electrons in its valence shell.

5. (c) Diamond, silicon and quartz molecule bounded by covalent bond.

6. (b) The acidity of hydrides of *VI* group elements increase from top to bottom as the bond strength $X-H$ decrease from top to bottom

$H_2O < H_2S < H_2Se < H_2Te$.

7. (b) Sulphur has the second highest catenation property after carbon. Its molecule has eight atom bonded together (*i.e.* S_8).

8. (b) Polarity character is due to the difference in electronegativity of two atoms or molecule.

9. (a) The given electronic configuration shows that *A* represents noble gas because the octet is complete. *A* is neon which has 10 atomic number.

10. (b) The electronic configuration of *C* represent chlorine. Its stable form is dichlorine (Cl_2) *i.e.*, C_2.

11. (d) The electronic configuration shows that *B* represents phosphorus and *C* represents chlorine. The stable compound formed is PCl_3 *i.e.*, BC_3.

12. (b) The bond between *B* and *C* will be covalent. Both *B* and *C* are non-metal atoms. *B* represents phosphorus and *C* represent chlorine.

13. (a) The electronic configuration represents

$2s^2 2p^5$ = fluorine = most electronegative element

$3s^2 3p^5$ = chlorine

$4s^2 4p^5$ = bromine

$5s^2 5p^5$ = iodine

14. (a) Among these, NaH and CaH_2 are ionic hydrides and B_2H_6 and NH_3 are covalent hydrides.

15. (c) Silica has tendency to form long chain covalent structure such as carbon so it has giant covalent structure.

16. (a) The approach of the atoms towards one another should be accompanied by decrease of energy.

17. (a) With similar electronegativities, ionic bond cannot be formed.

18. (d) Urea is a covalent compound which does not break into ions when dissolved in water. So it does not conduct electricity.

19. (d) Due to the small size and higher ionization energy, boron forms covalent compound.

20. (d) To change the state of matter, we need to weaken the inter-molecular forces of attraction.

21. (d) Among the given species. The bond dissociation energy of $C-O$ bond is minimum in case of CO_3^{2-} by which $C-O$ bond become more weaker in CO_3^{2-} or the bond order of CO_3^{2-} (1.33) is minimum so the bond becomes weaker.

22. (d) Because of the small size of H and O atoms, the bond length is minimum for $O-H$ bond.

23. (c) Hydrogen atom has only $1s$ orbital which overlap with another $1s$ orbital of H atom to form H_2 molecule. There is sufficient overlap between spherical orbitals. Also there is no repulsion due to other electrons as it has only one electron which is being used in the σ bonding. So the bond is very strong.

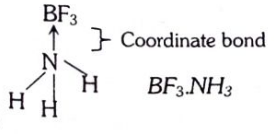
In the formation of N_2 molecule, p orbitals of N atom overlap axially (to form a σ bond) and laterally (to form two π – bonds). N atom experience repulsion due to other electrons which do not take part in bonding. But the combined effect of 3 bonds overcome the repulsion effect, so the bond dissociation energy of $N \equiv N$ is more than $H-H$.

24. (a) Residual bonding capacity is proportional to number of free valence electrons that do not remain permanently associated with a nucleus.

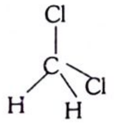
25. (b) Due to its small size and 2 electrons in s-orbital Be forms covalent compound.

3. Co-ordinate or Dative Bonding

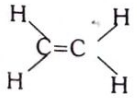
1. (a)



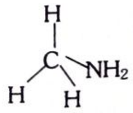
$BF_3.NH_3$



CH_2Cl_2



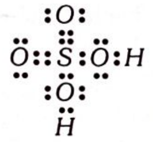
C_2H_4



CH_3NH_2

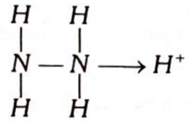
2. (d) Co-ordinate bond is a special type of covalent bond which is formed by sharing of electrons between two atoms, where both the electrons of the shared pair are contributed by one atom. Since this type of sharing of electrons exits in O_3, SO_3 and H_2SO_4. Therefore all these contain coordinate bonds.

3. (c) Lewis dot structure of sulphuric acid is



Hence, 2 dative bonds.

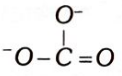
4. (a)



Structure of $N_2H_5^+$: 5 covalent bonds and 1 coordinate bond.

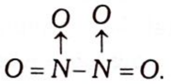
5. (d) $Na^+[N \equiv C]^-$.

6. (b)



$^-O-\overset{O^-}{\overset{|}{C}}=O$ has covalent bonds only.

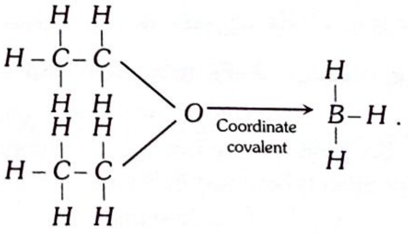
7. (c) Lewis structure of N_2O_4 is



$O=\overset{O}{\overset{\uparrow}{N}}-\overset{O}{\overset{\uparrow}{N}}=O.$

8. (b) Perchlorate ion, ClO_4^- has a tetrahedral structure with sp^3 hybridisation. Cl and O are bonded through covalent bonds and bond between Na^+ and ClO_4^- is ionic.

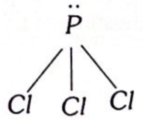
9. (b)



4. VSEPR Theory

1. (a) $H_2\ddot{\underset{\cdot\cdot}{O}} \Rightarrow$ No. of lone pairs=2

$Cl-\overset{\cdot\cdot}{S}n-Cl \Rightarrow$ No. of lone pairs=1



$\Rightarrow$ No. of lone pairs=1

$F-\overset{\cdot\cdot}{\underset{\cdot\cdot}{X}}\dot{e}-F \Rightarrow$ No. of lone pairs=3

2. (d) *l-p, b-p* repulsions are much greater than *b-p, b-p* repulsions. Bond angle of *C–N–C* in trimethylamine is 108°.

3. (a) Oxygen has 2 lone pair of electrons so its bond angle is less than expected due to $lp - lp$ and $lp - bp$ repulsion.

4. (c) In NH_3, although the central atom N is sp^3 hybridised, the bond angle is 107° (instead of 109°28'). The decrease in bond angle is due to repulsion between the lone pair of electrons on *N*-atom and bond pair of electrons between *N* and *H* atoms.

5. (a) Decrease of bond angle (from expected) can be explained by VSEPR theory.

6. (c) Bond angle of H_2O is 104.5° due to *lp-lp* repulsion.

7. (c) Count the no. of bond pairs and lone–pair around the central atom.

IF_5 : Summation of group number = 42

$$\text{Bond pairs} = \frac{42}{8} = 5 \text{ (Residue 2)}$$

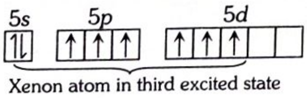
$$\text{Lone pair} = \frac{2}{2} = 1$$

5 bond pairs, 1 lone pair means the geometry is square pyramidal.

8. (c) In general, a compound with formula AB_3 is sp^2 hybridised with triangular planar geometry. However, the given compound has pyramidal structure which is possible only when it has a lone pair of electrons. Thus AB_3 has 3 bond pairs and one lone pair.

9. (a) Oxidation state of Xe in XeF_6 is +6.

In the formation of XeF_6, sp^3d^3 hybridisation occurs which gives the molecule a pentagonal bipyramidal structure. Six positions are occupied by fluorine atoms and one position is occupied by a lone pair of electrons. Due to presence of lone pair, distortion in structure takes place. The actual structure is distorted octahedral.



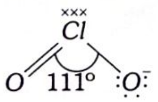
Xenon atom in third excited state

10. (d) ICl_2^- - linear

I_3^- - linear

N_3^- - linear

ClO_2^--angular due to sp^3 hybridisation of Cl – atom

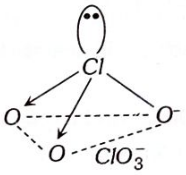


11. (a) H_2O and SO_2 both are bent.

12. (d) Cl^+ forms 4σ bonds with $4F$ atoms and left with a lone pair of electrons. So Cl^+ shows sp^3d hybridization and expected geometry is trigonal bipyramidal. Its actual shape is 'see – saw'.

13. (b) OF_2 is formed which has a bent geometry, as predicted by VSEPR theory, having 103° bond angle (less than water).

14. (b) Central atom Cl forms 3σ bonds with a lone pair of electrons. Thus it will show sp^3 hybridisation. So Trigonal pyramidal structure.



15. (d) In N – atom, number of valence electrons $= 5$

Due to the presence of one negative charge, number of valence electrons $= 5 + 1 = 6$ one O – atom forms two bond (= bond) and two *O*-atom shared with two electrons of *N*-atom

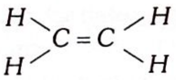
Thus, 3 *O*-atoms shared with 8 electrons of *N*-atom.

∴ Number of bond pairs (or shared pairs) $= 4$

Number of lone pairs $= 0$.

16. (c) $XeF_4 \Rightarrow 4\,bp + 2\,lp \Rightarrow$ square planar $\Rightarrow$ all bonds are equal

$BF_4^- \Rightarrow 4\,bp + 0\,lp \Rightarrow$ tetrahedral (all bonds are equal)

$C_2H_4 \Rightarrow$  $\Rightarrow C = C$

$$C_2H_4 \Rightarrow \begin{matrix} H & & & & H \\ & C & = & C & \\ H & & & & H \end{matrix} \Rightarrow C = C$$

bond is not equal to $C - H$ bond

$SiF_4 \Rightarrow 4\,bp + 0\,lp \Rightarrow$ tetrahedral (all bonds are equal)

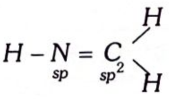
Thus, in C_2H_4 all the bonds are not equal.

17. (d) $NH_3 = 107°, PH_3 = 93°, H_2O = 104.5°$

$H_2Se = 91°, H_2S = 92.5°$.

18. (b) BCl_3 is trigonal in shape where bond angle is $120°$. In NH_3 and PCl_3 (both are having tetrahedral geometry) though the central atom has equal no. of lone pairs and bond pairs, the valence shell of *P* is relatively bigger and bond angle is lower.

19. (c) A. H_3O^+ – Trigonal pyramidal

B.  – Planar

$$H - \underset{sp}{N} = \underset{sp^2}{C}\begin{matrix} H \\ \\ H \end{matrix}$$

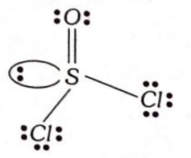
C. ClO_2^- – Angular

D. NH_4^+ – Tetrahedral

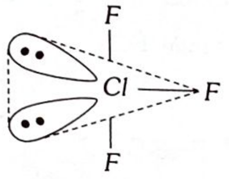
E. PCl_5 – Trigonal bipyramidal

20. (d) S is sp^3 hybridized with a lone pair of electrons. Expected geometry is tetrahedral but actual shape is trigonal pyramidal.

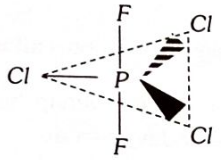
Due to a lone pair of electron the bond angles should be less than 107°. Due to small size of O atom and bond pair of electrons, angle $Cl-S-Cl$ will contract further. Thus only option (d) is possible.



21. (c) Three bond pair and two lone pair present in ClF_3 molecule



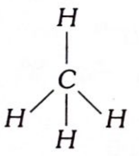
22. (a) Correct structure of PCl_3F_2 is



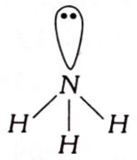
Trigonal bipyramidal

For minimum repulsion between atoms.

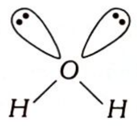
23. (b)



Tetrahedral



Pyramidal (one lone pair)



Bent (two lone pair)

Bond angle 109.5° 107.1° 104.5°

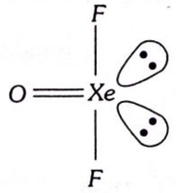
Order of repulsions $lp-bp > bp-bp$, hence greater the no. of lone pairs smaller the bond angle

(If hybridization of central atom is same).

5. Hybridisation

1. (b) PF_5 involves sp^3d hybridization and hence has trigonal bipyramidal geometry.

2. (c)

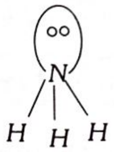


sp^3d hybridization of Xe

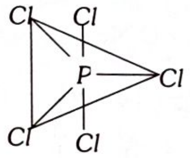
(Trigonal bipyramidal)

(T-shape)

3. (b) In the ammonia molecule N atom is sp^3 – hybridized but due to the presence of one lone pair of e^- (*i.e.* due to greater $L_p - b_p$ repulsion) it has distorted tetrahedral (or pyramidal) geometry.



4. (a) Structure of PCl_5 is



axial bonds are 90° and equatorial bonds are 120°.

5. (a) $BH_4^- \Rightarrow$ 4 bond pairs + 0 lone pair $\Rightarrow sp^3$ hybridised = tetrahedral geometry

NH_2^- = V – shape

CO_3^{2-} = triangular planar

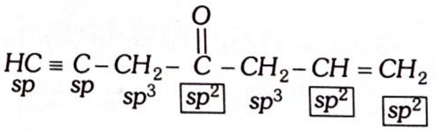
H_3O^+ = pyramidal

6. (a) As compare to pure atomic orbitals, hybrid orbitals have low energy.

7. (b) Hg has 2 electrons in its valence shell and it forms 2σ bonds in $HgCl_2$ molecule. So it's hybridization is sp type.

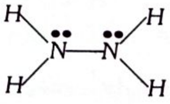
8. (c) The shape of sp^3d hybridized molecule is trigonal bipyramidal. The three planar triangular hybrid orbitals may be assumed to be formed from one $s-$ and two p-orbitals. So the remaining p_z – orbital may combine with d_{z^2} (aligned along z – axis) to form two axial hybrid orbitals that result in the formation of sp^3d hybridized orbitals.

9. (a)



$$\underset{sp}{HC}\equiv\underset{sp}{C}-\underset{sp^3}{CH_2}-\underset{\boxed{sp^2}}{\overset{\overset{O}{\|}}{C}}-\underset{sp^3}{CH_2}-\underset{\boxed{sp^2}}{CH}=\underset{\boxed{sp^2}}{CH_2}.$$

10. (c)



2 lone pair on nitrogen

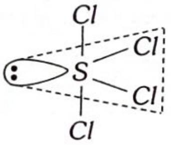
$\left.\begin{array}{l}4\ N-H \text{ bond pair}\\ 1\ N-N \text{ bond pair}\end{array}\right\}$ 5 bond pairs .

11. (a)

	sp	sp^2	sp^3
% s–character	50%	33.33%	25%

More the % s–character more the electronegativity [other factors being equal].

12. (d) $SCl_4 \Rightarrow 4\,b.p + 1\,l.p.$

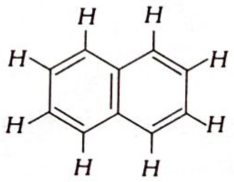


6. Overlapping- σ and π- Bonds

1. (b) π-bond is formed by lateral overlapping of unhybridised p-p orbitals.

2. (a) $\underset{\sigma}{O}\overset{\pi}{=}\underset{\sigma}{C}\overset{\pi}{=}\underset{\sigma}{C}\overset{\pi}{=}\underset{\sigma}{C}\overset{\pi}{=}O$ 4σ and 4π.

3. (c) The given compound will have the correct structure as

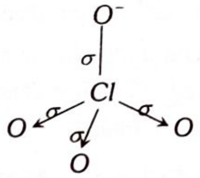


There are 5 π-bonds and $8C-H+11C-C$ σ-bonds, *i.e.,* 19 σ-bonds are present in the above molecule.

4. (b) Hybridization formula

$$=\frac{1}{2}\left[\begin{array}{l}\text{No. of valence}\\ e^-\text{ over central}\\ \text{element}\end{array} + \begin{array}{l}\text{No. of monovalent}\\ \text{surrounding}\\ \text{elements}\\ (H,F,Cl,Br,I)\end{array} \pm \text{ Charge}\right]$$

$$=\frac{1}{2}(7+0-0+1)=4$$



$\Rightarrow sp^3$ hybridization and tetrahedral structure (symmetrical)

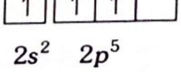
${}_8O = 1s^2 2s^2 2p^4$

So all the four sigma bonds will be sp^3-p bond.

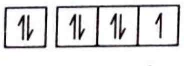
5. (d) The p_x and p_y orbitals do not overlap thus, no bond is formed.

6. (a) E.C. of ${}_5B$ in ground state $= 1s^2\,2s^2\,2p^1$

E.C. of ${}_5B$ in excited state $= 1s^2\quad 2s^1\quad 2p^2$



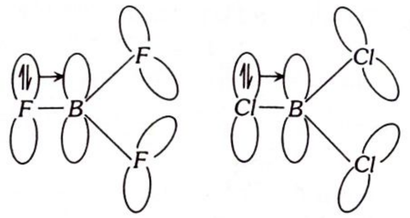
E.C. of ${}_9F$ $= 1s^2\quad 2s^2\quad 2p^5$



Only $2p_z$ orbital is vacant with boron to make a $p\pi-p\pi$ back bonding with the filled orbitals of fluorine and which exactly takes place in the formation of BF_3. So there is maximum $p\pi-p\pi$ back bonding in BF_3.

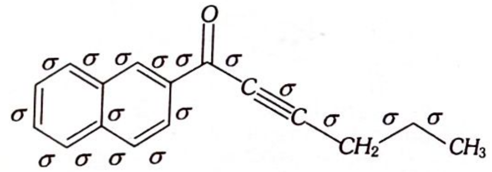
7. (b) Stronger $2p(B)-2p(F)\pi$ – bonding

$\pi p-\pi p$ back bonding



Size of Cl is more than the size of F so in case of BF_3 strong $2p(B)-2p(F)\pi$ – bonding occurs so lewis acidity of BF_3 is less than BCl_3.

8. (b)



$17, C-C$ sigma bond present in this structure.

9. (a) It has *+ve* and *–ve* overlap both simultaneous. So it leads to non–bonding overlap.

7. Dipole Moment

1. (b) Charge of $e^- = 1.6\times10^{-19}$

Dipole moment of $HBr = 1.6\times10^{-30}$

Inter atomic spacing $= 1\text{Å} = 1\times10^{-10}m$

% of ionic character in

$$HBr = \frac{\text{dipole moment of } HBr\times100}{\text{inter spacing distance}\times q}$$

$$=\frac{1.6\times10^{-30}}{1.6\times10^{-19}\times10^{-10}}\times100$$

$$=10^{-30}\times10^{29}\times100 = 10^{-1}\times100 = 0.1\times100 = 10\%.$$

2. (c) Given ionic charge $= 4.8\times10^{-10}$ e.s.u. and ionic distance $= 1A^\circ = 10^{-8}cm$. We know that dipole moment = ionic charge × ionic distance $= 4.8\times10^{-10}\times10^{-8}$ $= 4.8\times10^{-18}$ e.s.u. cm = 4.8 debye.

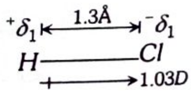
3. (a) % of ionic character

$$=\frac{\text{Experimental value of dipole moment}}{\text{Expected value of dipole moment}}$$

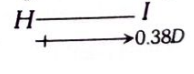
$$=\frac{1.03}{6.12}\times100 = 16.83\% \approx 17\%.$$

4. (c) Dipole moment $\mu = \delta\times$ bond length

$\delta_1\times1.3(10^{-10}) = 1.03$



$\delta_2\times1.6(10^{-10}) = 0.38$

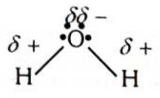


$$\frac{\delta_1}{\delta_2}=\frac{3.3\times10^{-10}}{1\times10^{-10}}$$

$\delta_1 : \delta_2 = 3.3 : 1.$

5. (a) The dipole moment of $CH_4 = 0D$, $NF_3 = 0.2D$, $NH_3 = 1.47D$ and $H_2O = 1.85D$. Therefore the correct order of the dipole moment is $CH_4 < NF_3 < NH_3 < H_2O$.

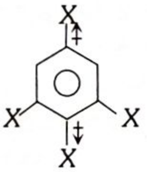
6. (c) Dipole-dipole forces act between the molecules possessing permanent dipole and ends of dipoles possess 'partial charges'. Partial charges present on ends of a dipole *e.g.*,



$\mu = 1.84\,D$

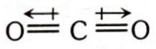
7. (d) Due to a lone pair in P atom PF_3 does not has a symmetrical structure.

8. (c) In XeF_4, the four fluorine atoms are present at the corners of a square plane and two lone pairs are present at axial positions. All the individual bond moments in such an arrangement cancel out, as a result, the molecule has zero dipole moment.



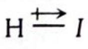
9. (a) , The bond moments at *para* positions cancel due to opposite direction. The other two groups are *meta* to each other, thus angle between the two bond moments is 120° and hence, $\mu = 1.5D$.

10. (c) CO_2 being symmetrical has zero dipole moment

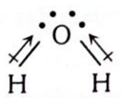


$\mu = 0$

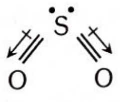
Among HI, SO_2 and H_2O dipole moment is highest for H_2O as in it the central atom contains 2 lone pairs



$\mu = 0.38D$

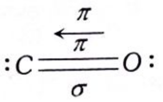


$\mu = 1.84\,D$

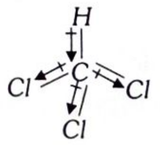


$\mu = 1.62D$

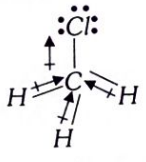
11. (a) Because oxygen is more electronegative than carbon atom so the $\sigma -$ electrons are towards oxygen atom. The co-ordinated covalent $\pi -$ bond has $\pi -$electrons which are donated by oxygen to carbon atom. So there is a drift of $\pi -$ electrons from O to C.



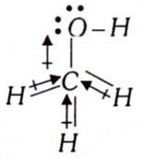
12. (c)



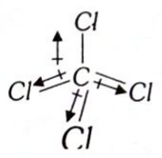
CHCl₃



CH₃Cl



CH₃OH



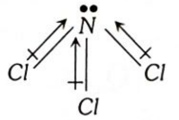
CCl₄

In $CHCl_3$ the vector summation of $3C-Cl$ bonds is very small so the net dipole moment is smaller than CH_3Cl. In CH_3OH, the electronegativity difference of O and C is large so it will have higher dipole moment. CCl_4 will have zero dipole moment because of symmetrical structure the four $C-Cl$ bond dipole moments will cancel.

So the correct order of dipole moments is :

$CCl_4 < CHCl_3 < CH_3Cl < CH_3OH$

13. (d)



$\mu \neq 0$.

14. (d) Due to high electronegativity difference between H and F atoms, HF has highest value of dipole moment.

8. Polarisation and Fajan's Rule

1. (d) Greater is the negative charge and size of anion, higher is its polarising power.

2. (a) According to Fajan's rule.

3. (a) The more is the electronegativity difference between the atoms, more will be the polarity in molecules.

4. (a) Small size and high charge of cation increases its polarising power.

5. (d) $3SCl_2(l) + 4NaF(s) \longrightarrow SF_4(g) + S_2Cl_2(l) + 4NaCl(s)$.

6. (c) Percent ionic character is given by

$$\% \text{ of ionic character} = 16(X_A - X_B) + 3.5(X_A - X_B)^2$$

From the above relation, it is clear that as soon as $(X_A - X_B)$ increases, % ionic character will also increase. Therefore, curve C show a correct path.

7. (c) Ionic character is inversely proportional to polarising power of cation. The correct order of atomic radius: B<Ga<Al

$BCl_3 < GaCl_3 < AlCl_3$.

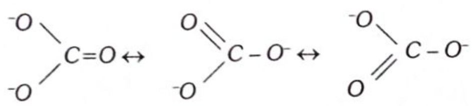
8. (d) We know that ionic characters

$$= 16[E_A - E_B] + 3.5 \times [E_A - E_B]^2$$

or ionic character = 72.24%.

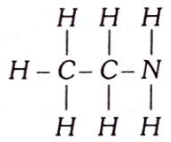
9. Resonance

1. (c) We know that carbonate ion has following resonating structures



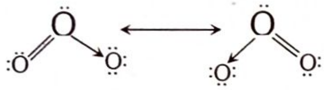
$$\text{Bond order} = \frac{\text{Total number of bonds between atoms}}{\text{Total number of resonating structure}}.$$

2. (c) All bonds are single bond. So there is no scope of resonance.



Ethyl amine.

3. (b) In $O_3, O-O$ bond length is not identical with that of molecular oxygen. It is found to be intermediate of $O-O$ and $O=O$ bond length. This arises due to resonance.



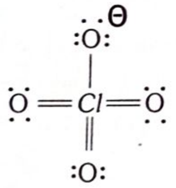
4. (b) Pi electrons are considered as mobile electrons which can take part in resonance.

5. (a) $\text{Bond order} = \frac{\text{Total number of bonds between atoms}}{\text{Total number of resonating structure}}$

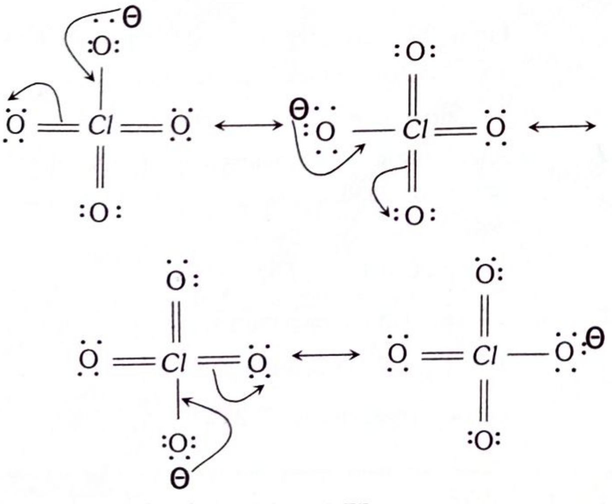
$$= \frac{5}{4} = 1.25.$$

6. (c) In resonance structure there should be the same number of electron pairs.

7. (c)

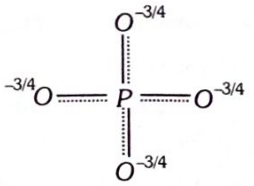


Perchlorate ion (ClO_4^-) have the following possible resonating structures :



$Cl-O$ bond order $= 7/4 = 1.75$.

8. (c) In the resonance hybrid structure of PO_4^{3-}, charge of –3 is distributed on 4 oxygen atom.



Resonance hybrid

So formal charge on each oxygen atom $= -3/4$

$= -0.75$.

10. Molecular Orbital Theory

1. (a) Molecular orbital electronic configuration of these species are :

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

$\pi 2p_z{}^2, \pi^* 2p_y{}^2 \pi^* 2p_z{}^1$

$O_2(16e) = \sigma 1s^2 \sigma^* 1s^2, \sigma 2s^2 \sigma^* 2s^2, \sigma 2p_x{}^2, \pi 2p_y{}^2,$

$\pi 2p_z{}^2 \pi^* 2p_y{}^1 \pi^* 2p_z{}^1$

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

$\pi 2p_z{}^2 \pi^* 2p_y{}^2 \pi^* 2p_z{}^2$

Hence number of antibonding electrons are 7,6,and 8 respectively.

2. (d) Molecular orbital configuration of peroxide ion, O_2^{2-} is

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

It is diamagnetic.

$$\text{Bond order} = \frac{10-8}{2} = 1.$$

3. (a) N_2 molecule has 14 electrons. The molecular orbital electronic configuration of the molecule is as :

$N_2 : KK\,(\sigma 2s)^2, (\sigma^* 2s)^2, (\pi 2p_x)^2, (\pi 2p_y)^2 \approx (\sigma 2p_z)^2$

N_2^- ion is formed when N_2 accept an electron hence, it has 15 electrons. The molecular orbital electronic configuration of the molecule is as ;

$N_2^- : KK\,(\sigma 2s)^2, (\sigma^* 2s)^2 \approx (\pi 2p_x)^2, (\pi 2p_y)^2, (\sigma 2p_z)^2,$

$(\pi^* 2p_x)^1$

Hence, this electron goes to antibonding π molecular orbital.

4. (c) $\overset{*}{\pi} 2p_y$ has two nodal planes.

5. (c) No. of electrons in $CO^+ = 13$

$CO^+ = KK'(\sigma 2s)^2 (\sigma^* 2s)^2 (\pi 2p_x)^2 (\pi 2p_y)^2 (\sigma 2p_z)^1$

This is the expected MO electronic configuration of CO^+.

But in case of CO and $CO^+, (\sigma^* 2s)$ MO is higher in energy than $(\pi 2p_x), (\pi 2p_y)$ and $(\sigma 2p_z)$. So the actual configuration are :

$$CO = KK'(\sigma 2s)^2(\pi 2p_x)^2(\pi 2p_y)^2(\sigma 2p_z)^2(\sigma^* 2s)^2$$

$$CO^+ = KK'(\sigma 2s)^2(\pi 2p_x)^2(\pi 2p_y)^2(\sigma 2p_z)^2(\sigma^* 2s)^1$$

$$N_b = 8, \quad N_a = 1$$

$$B.O. = \frac{1}{2}(N_b - N_a)$$

$$= \frac{1}{2}(8-1)$$

$$= 3.5.$$

6. (c) $H_2^+ = (\sigma 1s)^1$

Bond order $= \frac{1}{2}(1-0) = 1/2$

$H_2^- = (\sigma 1s)^2(\sigma^* 1s)^1$

Bond order $= \frac{1}{2}(2-1) = 1/2$

Both H_2^+ and H_2^- have bond order of 1/2, but H_2^- has an electron in anti-bonding orbital so it is high in energy and less stable than H_2^+.

7. (b) Higher the bond order, higher is the bond dissociation energy and shorter is the bond length.

8. (b) Substances which are weakly repelled by external magnetic field are called diamagnetic substances, *e.g.* H_2O.

9. (a) The correct increasing order of energies of molecular orbitals of N_2 is given below.

$$\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x \approx \pi 2p_y) < \sigma 2p_z$$
$$< (\pi^* 2p_x \simeq \pi^* 2p_y) < \sigma^* 2p_z.$$

10. (d) Existence of molecule, bonding nature and energy order of molecular orbitals can be explained on the basis of molecular orbital theory as follows

(i) Molecules having zero bond order never exists while molecular having non-zero bond order is either exists or expected to exist.

(ii) Higher the value of bond order, higher will be its bond strength.

Electrons present in bonding molecular orbital are known as bonding electrons (N_b) and electrons present on anti-bonding molecular orbital are known as anti-bonding electrons (N_a) and half of their difference is known as bond order *i.e.*,

(a) $Be_2 (4+4=8) = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2$

$B.O. = \frac{1}{2}$ [Number of bonding electrons (N_b) −

Number of anti-bonding electrons (N_a)] $= \frac{4-4}{2} = 0$

Here, bond order of Be_2 is zero. Thus, it does not exist.

(b) $He_2 (2+2=4) = \sigma 1s^2, \sigma 1s^2$

$$B.O. = \frac{2-2}{2} = 0$$

Here, bond order of He_2 is zero. Hence, it does not exist

$He_2^+ \ (2+2-1=3) = \sigma 1s^2, \sigma^* 1s^1$

$$B.O. = \frac{2-1}{2} = 0.5$$

Since, the bond order is not zero, this molecules is expected to exist.

(c) $N_2 \ (7+7=14) = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_x^2$

$\approx \pi 2p_y^2, \sigma 2p_z^2$

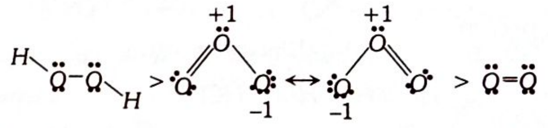
$$B.O. = \frac{10-4}{2} = 3$$

Thus, dinitrogen (N_2) molecule contain triple bond and no any molecule of second period have more than double bond. Hence, bond strength of N_2 is maximum amongst the homonuclear diatomic molecules belonging to the second period.

(d) It is incorrect. The correct order of energies of molecular orbitals in N_2 molecule is

$$\sigma 2s < \sigma^* 2s < (\pi 2p_x \simeq \pi 2p_y)$$
$$< \sigma 2p_z < \pi^* 2p_x \approx \pi^* 2p_y < \sigma^* 2p_z.$$

11. (b)



Bond order = 1 1.5 2

12. (c) (a) O_2, $B.O = 2$

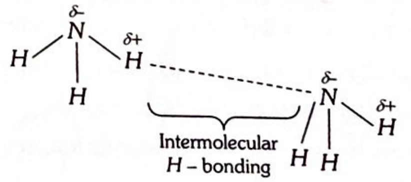
(b) F_2, $B.O = 1$

(c) O_2^+, $B.O = 2.5$

(d) F_2^-, $B.O = 0.5$.

11. Hydrogen Bonding

1. (c) It contains intermolecular hydrogen bonding.

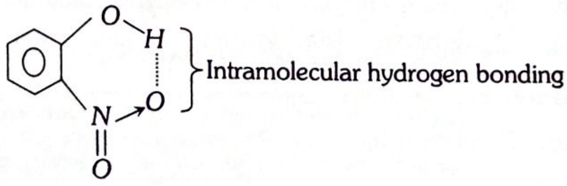


2. (a) A compound having maximum electronegative element will form strong Hydrogen bond.

3. (b) In ice cube all molecules are held by inter molecular hydrogen bond.

4. (b) Due to greater resonance stabilization.

5. (c)



o-nitrophenol

Due to chelation, $-OH$ group is not available to form hydrogen bond with water, hence it is sparingly soluble in water.

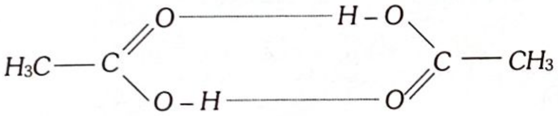
6. (b) Ethyl alcohol shows inter-molecular hydrogen bonding in its solution because of positive charged hydrogen.

7. (b) Strength of H-bond is in the order $H.....F > H.....O > H.....N$.

   But each H_2O molecule is linked to four other H_2O molecules through $H-$bonds whereas each HF molecule is linked only to two other HF molecules.

   Hence, $b.p \text{ of } H_2O > b.p \text{ of } HF > b.p. \text{ of } NH_3$.

8. (b) Benzene is a non-polar solvent (very low di-electric constant), unlike water which is a polar solvent (high di-electric constant). Due to non–polar nature of benzene it does not interfere the inter–molecular hydrogen bonding in acetic acid (unlike solvent like water).



Dimer of acetic acid

9. (b) $HF > HI > HBr > HCl$

   HF is hydrogen bonded, thus has highest boiling point, and it is liquid at or below 19°C. The remaining hydrogen halides are gaseous and their boiling points depend on the vander Waal's forces. Larger the size (or molecular mass), greater is the vander Waals forces, hence higher is the boiling point. Thus the order is

   $HI > HBr > HCl$ and overall $HF > HI > HBr > HCl.$

10. (b) There is no acidic proton in diethyl ether, hence it does not exhibit strong hydrogen bonding.

12. Types of Bonding and Forces in Solid

1. (d) Dry ice or solid CO_2 has covalent forces as intramolecular forces and Vander Waal's forces as intermolecular forces.

2. (d) Mercury has very weak interatomic forces so it remains in liquid state.

3. (c) In Xe-molecule, the type of bonding which exist is van der Waal's forces.

4. (b) Solid neon possess only dispersion forces so it has weakest bond.

5. (d) Blue vitriol is $CuSO_4.5H_2O$ and it has all types of bonds.

6. (c) London dispersion forces operate only over very short distance. The energy of interaction varies as $\frac{1}{(\text{distance between two interacting particles})^6}$.

   Large or more complex are the molecules, greater is the magnitude of London forces. This is obviously due to the fact that the large electron clouds are easily distorted or polarised.

   Hence, greater the polarisability of the interacting particles, greater is the magnitude of the interaction energy.

7. (d) Due to weak Vander Waal's forces of attraction between the molecules, the molecules of CO_2 in dry ice can be called as discrete molecules.

8. (a) Down the same group lattice energy decreases due to increasing size of cation.

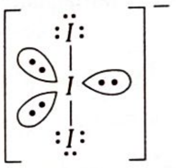
9. (b) Br_2 is a non – polar molecule. So the predominante force is Vander Waal's force in determining the melting and boiling point.

13. IIT-JEE/ AIEEE

1. (d) The value of lattice energy depends on the charges present on the two ions and distance between them. It shall be high if charges are high and ionic radii are small.

2. (b) Molten sodium chloride conducts electricity due to the presence of free ions.

3. (a) $I_3^- =$



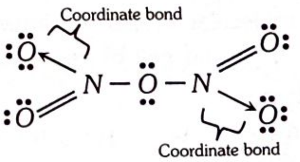
   Total number of lone pair of electron $= 9$.

4. (a) KCl is ionic compound.

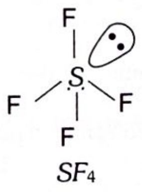
5. (c) Structure of KCN is $[K^+(C^- \equiv \dot{N})]$.

6. (c) Covalent character in ionic compounds is governed by Fazan's Rule. $AlCl_3$ will show Maximum covalent character on account of higher polarising power of Al^{3+} because of its having higher positive charge and smaller size.

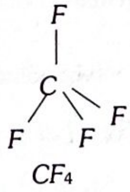
7. (a) Structure of N_2O_5.



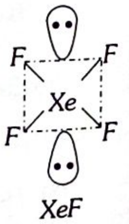
8. (d)



SF4



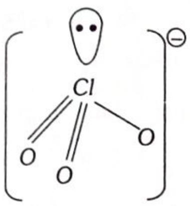
CF4

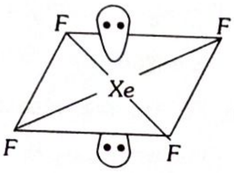


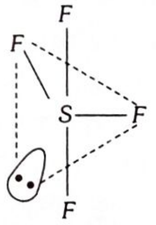
XeF

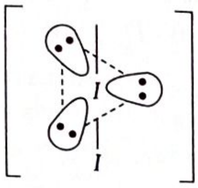
9. (c) $[NF_3$ and $H_3O^+]$ are pyramidal while $[NO_3^-$ and $BF_3]$ are planar. Hence answer (c) is correct.

10. (d)





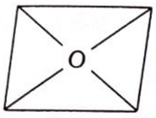




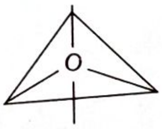
11. (c) As electronegativity of central atom decreases, bond angle decreases.

(Hybridisation and number of lone pair on cental atom are same in all options)

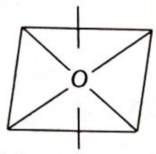
12. (d)



dsp^2 hybridization (Four 90° angles between bond pair and bond pair)



sp^3d hybridization (Six 90° angle between bond pair and bond pair)



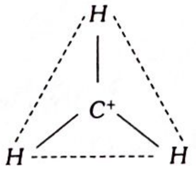
sp^3d^2 hybridization (Twelve 90° angle between bond pair and bond pair)

13. (c) CO_2 has sp – hybridization and is linear.

14. (b) sp-hybridization gives two orbitals at 180° with Linear structure.

15. (a) In PCl_3 molecule, phosphorous is sp^3 – hybridised but due to presence of lone pair of electron, it has pyramidal shape

16. (c) CH_3^+



$H=\frac{1}{2}(4+3-1+0)=3$

$\Rightarrow sp^2$ hybridization

Trick: N, P and Sb are in the same group which are connected to 3 hydrogen atoms in option (a), (b) and (d) so either of them cannot be an answer.

17. (b) Hybridization of N in NH_3 is sp^3 that of Pt in $[PtCl_4]^{2-}$ is dsp^2 that P in PCl_5 is sp^3d and that of B in BCl_3 is sp^2.

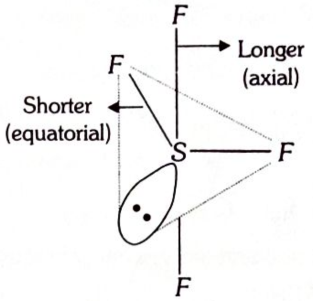
18. (b) IF_7 molecule show sp^3d^3 – hybridization.

19. (c) PF_5 trigonal bipyramidal

BrF_5 square pyramidal (distorted).

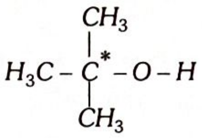
20. (d) $CH_2=\underset{sp^2}{CH}-\underset{\substack{sp^3\\ \text{hybridised}}}{CH_2}-CH_2-C\equiv CH$.

21. (a) The hybridization of sulphur in SF_4 is sp^3d. Equatorial bonds would be shorter than the axial one.



22. (d) An s-orbital being spherical is more close to the nucleus thus more electronegative. So more the s-character more the electronegativity.

23. (c)

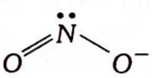


$$H_3C-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{*}{C}}}-O-H$$

All bonds are single covalent bonds so C^* is sp^3 hybridized.

24. (d)

(a) NO_2^-

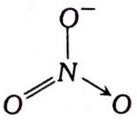


$H=\frac{1}{2}(V+M-C+A)$

$H=\frac{1}{2}(5+0-0+1)=3$

$\Rightarrow sp^2$ hybridization with 1 lone pair and bent shape.

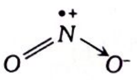
(b) NO_3^-



$H=\frac{1}{2}(5+0-0+1)=3$

$\Rightarrow sp^2$ hybridization and trigonal planar structure.

(c) NO_2



Nitrogen show special type of bonding where it is sp^2 – hybridized and bent shape.

(d) NO_2^+

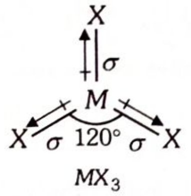
$O=\overset{+}{N}=O$

Linear

$H=\frac{1}{2}(5+0-1+0)=2$

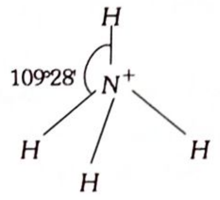
$\Rightarrow sp$ hybridization.

25. (d) In sp^2 hybridization with no lone pair of electrons, the geometry of the molecule MX_3 will be trigonal planar. So the dipole moment will cancel each other. The net dipole moment will be zero.



26. (a) NH_4^+ :

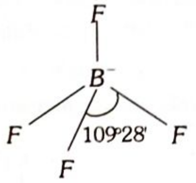
$H=\frac{1}{2}(5+4-1+0)=4$



⇒ sp^3 hybridization with no lone pair and tetrahedral geometry. Bond angles are 109°28'

BF_4^- :

$H=\frac{1}{2}(3+4-0+1)=4$



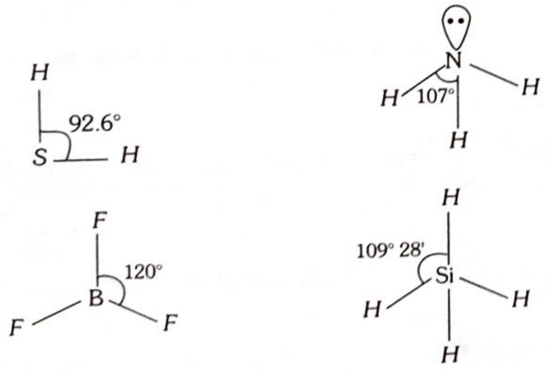
⇒ sp^3 hybridization with no lone pair and tetrahedral geometry. Bond angles are 109°28'.

27. (a) In H_2CO_3 and BF_3 central atom are in sp^2 hybridization but in H_2CO_3 due to the ionic character of $O-H$ bond it will be polar (High electronegativity of oxygen).

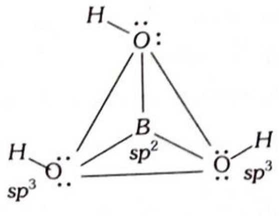
28. (a) The correct order of bond angle (Smallest first) is

$H_2S < NH_3 < SiH_4 < BF_3$

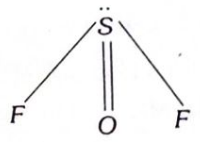
$92.6° < 107° < 109°28' < 120°$



29. (b)



30. (d)



Pyramidal shape

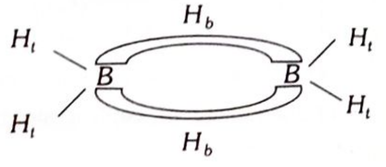
31. (a) Chlorine atom in ClO_2^- is sp^3 hybridised but its shape is angular.

32. (b) For square planar geometry hybridization is dsp^2 involving s, p_x, p_y and $d_{x^2-y^2}$ orbital.

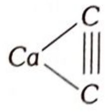
33. (a) The diborane molecule has two types of $B-H$ bond :

(i) $B-H_t$ – It is a normal covalent bond.

(ii) $B-H_b$ – It is a three centred bond.

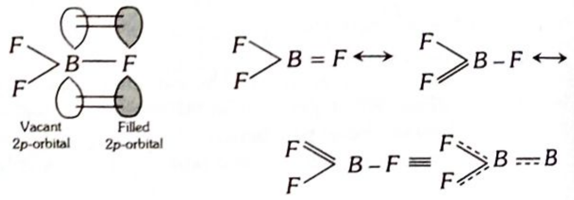


34. (b)



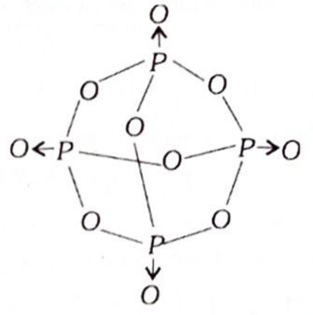
1σ and 2π

35. (c)



Decrease in $B-F$ bond length which results in the higher bond dissociation energy of $B-F$ in BF_3 is due to delocalised $p\pi-p\pi$ bonding between filled p-orbital of F-atom and vacant p-orbital of B atom.

36. (d) Structure of P_4O_{10} is

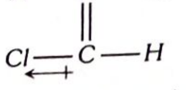


Each phosphorus is attached to 4 oxygen atoms.

37. (b) CCl_4 has no net dipole moment because of its regular tetrahedral structure.

38. (a) H_2S has angular geometry and have some value of dipole moment.

39. (d) Trans –1, 2– dichloroethene will have a net zero dipole moment because of the cancellation of two dipole moments between $C-Cl$ bonds.

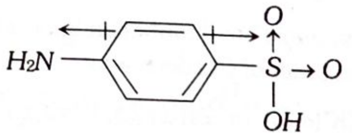


trans –1, 2– dichloroethene

40. (b) Ion–dipole interaction $\propto \frac{1}{r^3}$

41. (b) The correct order of increasing dipole moment is *p*-dichlorobenzene < Toluene < *m*-dichlorobenzene < *o*-dichlorobenzene.

42. (b) Structure of sulphanilic acid is



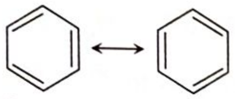
Due to bipolar structure its melting point is high and it is insoluble in organic solvents (like dissolves like).

43. (a) CH_3Cl have one Cl atom which is more electronegative so it will have highest dipole moment.

44. (c) Higher the charge/size ratio, more is the polarizing power.

45. (d) Due to high electronegativity of oxygen, it has partial negative charge.

46. (c) Bond order $= \dfrac{1+2}{2} = 1.5.$



Resonance in C_6H_6

47. (b) $H_2C = \overset{+}{N} = \overset{-}{N}$ (I) $\qquad H_2\overset{+}{C} - N = \overset{-}{N}$ (II)

octet complete — octet incomplete

–ve charge on nitrogen — –ve charge on nitrogen

$H_2\overset{-}{C} - \overset{+}{N} \equiv N$ (III) $\qquad H_2\overset{-}{C} - N = \overset{+}{N}$ (IV)

octet complete — octet incomplete

–ve charge on carbon — –ve charge on carbon.

48. (c) Helium molecule does not exist as bond order of $He_2 = 0$.

49. (c) Configuration of O_2 molecule is

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

No. of pair are 7 so total no. of paired electrons are 14.

50. (b) $O_2 : \sigma1s^2, \sigma^*1s^2, \sigma2s^2, \sigma^*2s^2, \sigma2p_x^{\ 2}\begin{cases}\pi2p_y^{\ 2} \\ \pi2p_z^{\ 2}\end{cases}\begin{cases}\pi^*2p_y^{\ 1} \\ \pi^*2p_z^{\ 1}\end{cases}$

Bond order $= \dfrac{10-6}{2} = 2.0$

(Two unpaired electrons in antibonding molecular orbital)

$O_2^+ : \sigma1s^2, \sigma^*1s^2, \sigma2s^2, \sigma^*2s^2, \sigma2p_x^{\ 2}\begin{cases}\pi2py^2 \\ \pi2pz^2\end{cases}\begin{cases}\pi^*2py^1 \\ \pi^*2pz^0\end{cases}$

Bond order $= \dfrac{10-5}{2} = 2.5$

(One unpaired electron in antibonding molecular orbital so it is paramagnetic).

51. (b) In $C_2 - C_2^+$ electron is removed from bonding molecular orbital so bond order decreases. In $NO \longrightarrow NO^+$, electron is removed from anti bonding molecular orbital so bond order increases and nature changes from paramagnetic to diamagnetic.

52. (a) $CO : \sigma_{1s}^2\sigma_{1s}^{*2}\,\sigma_{2s}^2\sigma_{2s}^{*2}\begin{cases}\pi_{2p_y}^2 \\ \pi_{2p_x}^2\end{cases}\sigma_{2p_x}^2$;

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

$NO^+ : \sigma_{1s}^2\dot{\sigma}_{1s}^{*2}\sigma_{2s}^2\sigma_{2s}^{*2}\sigma_{2p_x}^2\begin{cases}\pi_{2p_y}^2 \\ \pi_{2p_z}^2\end{cases}$;

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

$CN^- : \sigma_{1s}^2\sigma_{1s}^{*2}\sigma_{2s}^2\sigma_{2s}^{*2}\begin{cases}\pi_{2p_y}^2 \\ \pi_{2p_z}^2\end{cases}\sigma_{2p_x}^2$;

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

$N_2 : \sigma_{1s}^2\sigma_{1s}^{*2}\sigma_{2s}^2\sigma_{2s}^{*2}\begin{cases}\pi_{2p_y}^2 \\ \pi_{2p_z}^2\end{cases}\sigma_{2p_x}^2$;

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

$NO^- : \sigma_{1s}^2\,\sigma_{1s}^{*2}\sigma_{2s}^2\sigma_{2s}^{*2}\sigma_{2p_x}^2\begin{cases}\pi_{2p_y}^2 \\ \pi_{2p_z}^2\end{cases}\begin{cases}\pi_{2p_y}^{*1} \\ \pi_{2p_z}^{*1}\end{cases}$;

Bond order $= \dfrac{10-6}{2} = 2$.

53. (b) Higher the bond order, shorter will be the bond length, thus NO^+ having the higher bond order that is 3 as compared to NO having bond order 2 so NO^+ has shorter bond length.

54. (a) B_2 (total number of electrons $=10$)

$$\sigma1s^2\sigma^*1s^2\sigma2s^2\sigma^*2s^2\pi2p_x^{\ 2} = \pi2p_y^{\ 0}\sigma2p_z^{\ 0}$$

So, bond order $= \dfrac{6-4}{2} = 1$ and molecule will be dimagnetic.

55. (b) $Li_2 \ \sigma1s^2 \ \sigma^*1s^2 \ \sigma2s^2$ Bond order $=1$

$Li_2^+ \ \sigma1s^2 \ \sigma^*1s^2 \ \sigma2s^1$ Bond order $=0.5$

$Li_2^- \ \sigma1s^2 \ \sigma^*1s^2 \ \sigma2s^2 \ \sigma^*2s^1$ Bond order $=0.5$

Stability order $Li_2 > Li_2^+ > Li_2^-$

56. (b)

	He_2^+	H_2	H_2^+	H_2^-
$\sigma(1s)$	↑↓	↑↓	↑	↑↓
$\overset{*}{\sigma}(1s)$	↑	_	_	↑
B.O.	$\frac{1}{2}$	1	$\frac{1}{2}$	$\frac{1}{2}$
Magnetic nature	P	D	P	P

(P = Paramagnetic, D = Diamagnetic).

57. (d) $O_2 = (\sigma 1s)^2(\sigma^* 1s)^2(\sigma 2s)^2(\sigma^* 2s)^2(\sigma 2p_z)^2(\pi 2p_x)^2$

$(\pi 2p_y)^2\underbrace{(\pi^* 2p_x)^1(\pi^* 2p_y)^1}_{\text{unpaired electrons}}.$

58. (a) Molecular electronic configuration for given species are

$O_2 = \sigma 1s^2\ \sigma * 1s^2\ \sigma 2s^2\ \sigma * 2s^2\ \sigma 2p_x^2\ \pi 2p_y^2$

$\pi 2p_z^2\ \pi * 2p_y^1\ \pi * 2p_z^1$ (two unpaired electrons)

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

$\pi 2p_y^2\ \pi 2p_z^2\ \pi * 2p_y^2\ \pi * 2p_z^2$ (no unpaired electrons)

$N_2^+ = \sigma 1s^2\ \sigma * 1s^2\ \sigma 2s^2\ \sigma * 2s^2\ \pi 2p_y^2\ \pi 2p_z^2\ \sigma 2p_x^1$

(one unpaired electron)

$B_2 = \sigma 1s^2\ \sigma * 1s^2\ \sigma 2s^2\ \sigma * 2s^2\ \pi 2p_y^1\ \pi 2p_z^1$

(two unpaired electrons)

59. (a) NO is paramagnetic in gaseous state.

60. (d) KO_2 carries superoxide ion, O_2^- whose molecular orbital electronic configuration is

$\sigma_{1s}^2\ \overset{*}{\sigma}{}_{1s}^2\ \sigma_{2s}^2\ \overset{*}{\sigma}{}_{2s}^2\ \sigma_{2p_x}^2\ \pi_{2p_y}^2\ \pi_{2p_z}^2\ \overset{*}{\pi}{}_{2p_y}^2\ \overset{*}{\pi}{}_{2p_z}^1$

Thus it has one unpaired electron and hence paramagnetic in nature with magnetic moment = $\sqrt{n(n+2)} = \sqrt{1(1+2)} = 1.73$ B.M.

61. (ab)

(a) $C_2 = KK'(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p_x)^2(\pi 2p_y)^2$

Hence Diamagnetic

(b) $N_2 = KK'(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p_x)^2(\pi 2p_y)^2(\sigma 2p_z)^2$

Hence, Diamagnetic

(c) $O_2 = KK'(\sigma 2s)^2(\sigma^* 2s)^2(\sigma 2p_z)^2(\pi 2p_x)^2$

$(\pi 2p_y)^2(\pi^* 2p_x)^1(\pi^* 2p_y)^1$

Hence, Paramagnetic

(d) ${}_{16}S = 1s^2, 2s^2 2p^6, 3s^2 3p^4$

S_2 molecule has MO configuration similar to O_2 molecule as :

$S_2 = KK'LL'(\sigma 3s)^2(\sigma^* 3s)^2(\sigma 3p_z)^2(\pi 3p_x)^2(\pi 3p_y)^2$

$(\pi^* 3p_x)^1(\pi^* 3p_y)^1.$

Hence, Paramagnetic.

62. (b) $\underset{1.20Å}{C_2H_2} < \underset{1.34Å}{C_2H_4} < \underset{1.39Å}{C_6H_6} < \underset{1.54Å}{C_2H_6}$.

63. (d) B.O. in CO *i.e.*, $:\overset{-}{C} = \overset{+}{O}:$ is 3, that of $O = C = O$ is 2 while that of CO_3^{2-} ion is 1.33. Since the bond length increases as the bond order decreases i.e. $CO < CO_2 < CO_3^{2-}$. Thus option (d) is correct.

64. (a) Due to synergic bond formation between metal and CO, the bond order of CO decreases.

65. (b) $N_2 = KK'(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p_x)^2(\pi 2p_y)^2(\sigma 2p_z)^2$

$N_b = 8\ ,\ N_a = 2$

For $N_2^+\ :\ N_b = 7\ ,\ N_a = 2$

So when N_2 goes to N_2^+ : Bond order decreases or bond length increases.

$O_2 = KK'(\sigma 2s)^2(\sigma^* 2s)^2(\sigma 2p_z)^2(\pi 2p_x)^2(\pi 2p_y)^2$

$(\pi^* 2p_x)^1(\pi^* 2p_y)^1$

$N_b = 8\ ,\ N_a = 4$

For $O_2^+\ :\ N_b = 8,\quad N_a = 3$

So when O_2 goes to O_2^+ : Bond order increases or bond length decreases.

66. (a) CO

No of electron in $CO = 6 + 8 = 14$

(a) $CO \rightarrow \sigma 1S^2, \sigma^* 1S^2, \sigma 2S^2, \sigma^* 2S^2,$

$\left[\pi 2P_x^2 = \pi 2P_y^2\right]\sigma 2P_z^2$

All electrons are paired so diamagnetic

(b) $O_2 \rightarrow \sigma 1S^2, \sigma^* 1S^2, \sigma 2S^2, \sigma^* 2S^2, \sigma 2P_z^2$

$\left[\pi 2P_x^2 = \pi 2P_y^2\right], \left[\pi^* 2P_x^1 = \pi^* 2P_y^1\right]$

Unpaired electron = 2 (Paramagnetic)

(c) $B_2 \rightarrow \sigma 1S^2, \sigma^* 1S^2, \sigma 2S^2, \sigma^* 2S^2\left[\pi 2P_x^1 = \pi 2P_y^1\right]$

(Paramagnetic)

(d) $NO \rightarrow \sigma 1S^2, \sigma^* 1S^2, \sigma 2S^2, \sigma^* 2S^2, \sigma 2P_z^2$

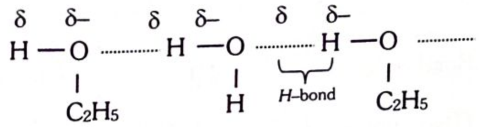
$\left[\pi 2P_x^2 = \pi 2P_y^2\right], \left[\pi^* 2P_x^1 = \pi^* 2P_y^0\right]$

(Paramagnetic)

67. (b) ${H_2}^{2-}$ have bond order zero ∴ do not exist

${H_2}^{2-} \rightarrow \sigma 1s^2, \sigma^* 1s^2, B.O. = \frac{2-2}{2} = 0$.

68. (d)

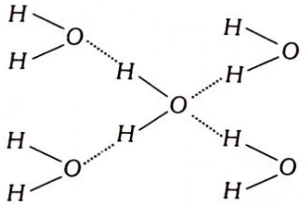


69. (c) The strongest hydrogen bond is in hydrogen fluoride because the power of hydrogen bond ∝ electronegativity of atom and

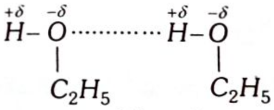
electronegativity $\propto \frac{1}{\text{atomic size}}$

70. (d) Because of low electronegativity of Cl atom, it does not form hydrogen bonding.

71. (d) Each H_2O molecule can participate in 4 hydrogen bonds or 5 molecules are attached together.



72. (a) Due to high electronegative O atom in ethanol, hydrogen bonding is maximum. Diethyl ether, ethyl chloride and triethyl amine does not have an H atom which is bonded to high electronegative element.



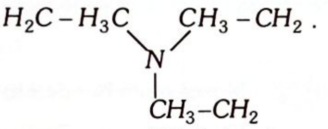
Ethanol

$H_2C-H_3C-\ddot{\underset{\bullet\bullet}{O}}-CH_3-CH_2$

Diethyl ether

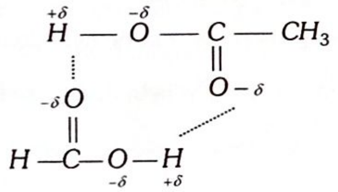
H_2C-CH_3-Cl ,

Ethyl chloride



Triethyl amine

73. (c)



14. NEET/ AIPMT/ CBSE-PMT

1. (b) A metal with low ionization potential will loose electrons easily and a non-metal with high electron affinity can gain electrons easily. So these atoms can form an ionic bond.

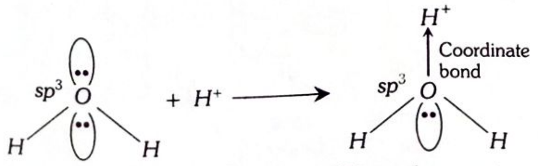
2. (d) $C-F$ bond has the most polar character due to difference of their electronegativity.

3. (c) $BeCl_2$ has covalent bonds, so maximum covalent characters will be shown by $BeCl_2$.

$LiCl$ and $NaCl$ has ionic bonds. The polarizing power of Li^+ is more than Na^+ due to small size of Li^+.

4. (d) Bond dissociation energy decreases with increase in size. So D is smallest.

5. (a)



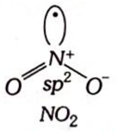
6. (a) H_2S does not have a linear. It is a bent shape and sp^3 hybridisation.

7. (a) BCl_3 has no lone pair of electrons but NCl_3 has a lone pair of electrons and both have different hybridization.

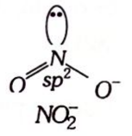
8. (d) $O=\overset{+}{N}=O$

sp

NO_2^+



NO_2



NO_2^-

9. (c) SiF_4 has symmetrical tetrahedral shape which is due to sp^3 hybridization of the central sulphur atom in its excited state configuration. SF_4 has distorted tetrahedral or Sea- Saw geometry which arise due to sp^3d hybridization of central sulphur atom and due to the presence of lone pair of electron in one of the equatorial hybrid orbital.

10. (bc)

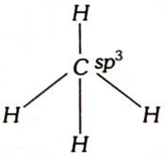
(b) In SO_3^{2-}, ClO_3^- No. of electrons $= 42$

Shape : Pyramidal

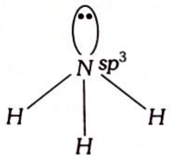
(c) In CO_3^{2-}, NO_3^- No. of electrons $= 32$

Shape : Trigonal planar

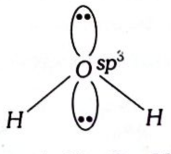
11. (b)



Tetrahedral | $H-C-H = 109°28'$



Trigonal pyramidal | $H-N-H = 107°$



Bent shape | $H-O-H = 104°28'$

In all the three cases, central atom is sp^3 hybridized but due to lone pair of electrons ($lp-lp$ and $lp-bp$ repulsions) bond angle is less in NH_3 and H_2O than regular tetrahedron structure.

12. (a) According to VSEPR theory the repulsion forces between electrons follow this decreasing order :
$lp-lp > lp-bp > bp-bp$.

13. (a) AB_5 molecule can take the structure of trigonal bipyramidal or square pyramidal.

14. (d) BrO_3^-

Electron pair over the central atom $= 3 + \frac{1}{2}[7 - 6 + 1] \Rightarrow 4$

No. of atom bonded to the central atom = 3

Geometry will be pyramidal.

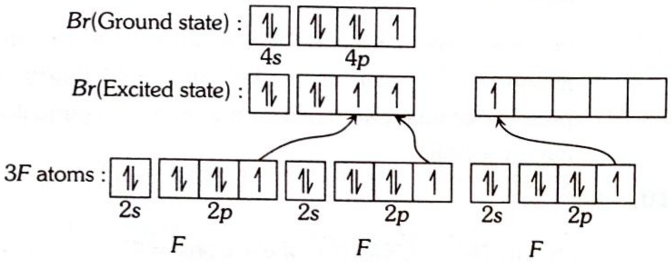
XeO_3

Electron pair $= 3 + \frac{1}{2}[8 - 6] \Rightarrow 4$

No. of atom bonded to the central atom = 3

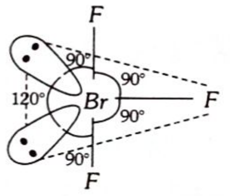
Geometry will be pyramidal.

15. (b)



$3(sp^3d - p)\sigma$ bonds and 2 lone pairs

$\Rightarrow$ T – shape molecule.

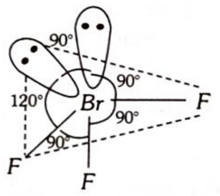


Lone pairs in
equatorial position
Repulsions at 90°

$lp - lp = 0$

$lp - bp = 4$

$bp - bp = 2$



One lone pair in eq. and
one in axial position.
Repulsions at 90°

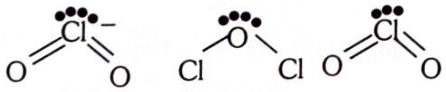
$lp - lp = 1$

$lp - bp = 3$

$bp - bp = 2$

So lone pairs occupy equatorial positions to minimize $lp - lp$ repulsions only.

16. (a) As the number of lone pairs of electrons increases, bond angle decreases due to repulsion between lp-lp. Moreover, as the electronegativity of central atom decreases, bond angle decreases. Hence, the order of bond angle is



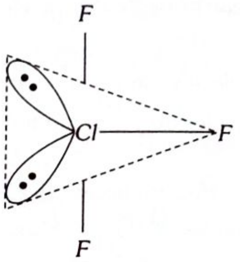
(two lone pairs) (two lone pairs) (one lone pair)

(Cl is less electronegative as compared to O.)

17. (b)

	Column I		Column II
(A)	XX' (e.g. ClF)	(iii)	Linear
(B)	XX_3' (e.g. ClF_3)	(i)	T-shape
(C)	XX_5' (e.g. IF_5)	(iv)	Square-pyramidal
(D)	XX_7' (e.g. IF_7)	(ii)	Pentagonal bipyramidal

18. (b)



Lone pairs $\Rightarrow 2$

Bond pairs $\Rightarrow 3$.

19. (d) In BF_3

$$H = \frac{1}{2}(V + M - C + A)$$

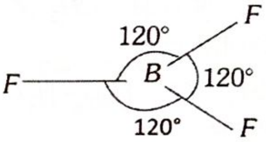
$$H = \frac{1}{2}(3 + 3 - 0 + 0) = 3 \Rightarrow sp^2 \text{hybridization.}$$

In NO_2^-

$$H = \frac{1}{2}(5 + 0 - 0 + 1) = 3 \Rightarrow sp^2 \text{ hybridization.}$$

20. (d) By VSEPR theory

21. (d)



No. of atom bonded to the central atom = 3

In case of 3, geometry is Trigonal planar.

22. (b) In BCl_3, boron is sp^2 hybridized.

23. (b) $NO_2^{(-)}$ and $NO_3^{(-)}$ both have same hybridization.

24. (d) In $SbCl_5^{2-}, Sb$ is sp^3d^2 hybridised, and rest have sp^3d.

25. (b) $SF_4 = sp^3d$.

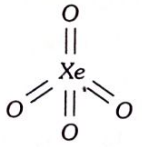
26. (a) $NO_3^{\ominus} = sp^2$

$H_3O^+ = sp^3$.

27. (d) BF_4^- hybridization sp^3, tetrahedral structure.

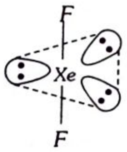
NH_4^+ hybridization sp^3, tetrahedral structure.

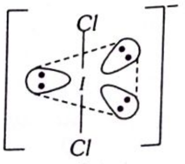
28. (a) Due to sp^3 hybridization, it is tetrahedral.



29. (b) Applying VSEPR theory, both NF_3 and H_2O are sp^3 hybridized.

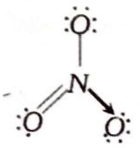
30. (c)





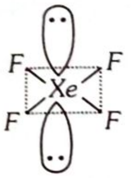
Hybridization $\Rightarrow$	sp^3d	sp^3d
Shape $\Rightarrow$	Linear	Linear
Geometry $\Rightarrow$	Trigonal bipyramidal	Trigonal bipyramidal

31. (d)



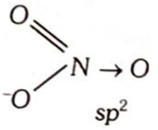
N is sp^2 hybrid and no lone pair.

32. (a)

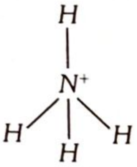


Geometry → Square planar hybridization → sp^3d^2

33. (d) $O=\overset{\oplus}{\underset{sp}{N}}=O$



sp^2



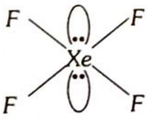
sp^3

34. (c) Due to $lp-lp$ repulsions, bond angle in H_2O is lower (104.5°) than that in NH_3 (107°) and $CH_4(109°28')$. BeF_2 on the other hand, has sp-hybridization and hence has a bond angle of 180°.

35. (c) Compound is carbontetrachloride because CCl_4 has sp^3 –hybridization 4 orbitals giving regular tetrahedron geometry. In others the geometry is little distorted inspite of sp^3 hybridization due to different atoms on the vertices of tetrahedron.

36. (c) In compound $\overset{3}{C}H_2=\overset{2}{C}=\overset{1}{C}H_2$ the second carbon is sp-hybridised.

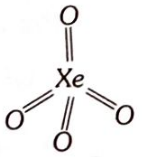
37. (d) XeF_4 :



$$H=\frac{1}{2}(8+4-0+0)=6$$

⇒ sp^3d^2 hybridization and square planar with 2 lone pairs.

XeO_4 :



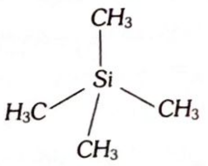
$$H=\frac{1}{2}(8+0-0+0)=4$$

⇒ sp^3 hybridization and tetrahedral.

Trick : In XeF_4 and XeO_4, central atom is same but due to different type of hybridizations, these will have different structures.

38. (c) $\underset{sp^3}{H_3C}-\underset{sp^2}{CH}=\underset{sp}{C}=\underset{sp^2}{CH_2}$.

39. (d) Hybridization on silicon atom

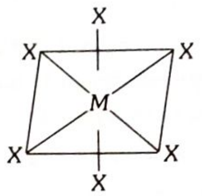


$$H=\frac{1}{2}(V+M-C+A)$$

$$H=\frac{1}{2}(4+4-0+0)=4$$

⇒ sp^3 hybridization and tetrahedral structure.

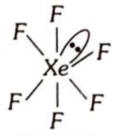
40. (c) In octahedral molecule six hybrid orbitals are directed towards the corner of a regular octahedron with a bond angle of 90°.



according to this geometry, the number of $X-M-X$ bond at 180° must be three.

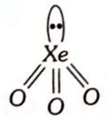
41. (a)

XeF_6



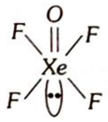
distorted octahedral

XeO_3



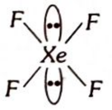
Pyramidal

$XeOF_4$



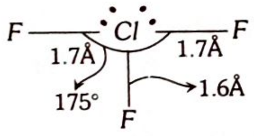
Square pyramidal

XeF_4



Square planar

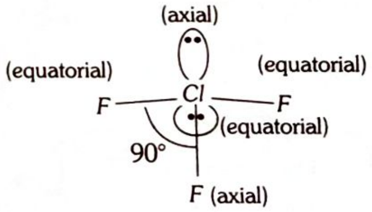
42. (c) ClF_3



$$H=\frac{1}{2}(7+3-0+0)=5$$

⇒ sp^3d hybridization with 2 lone pairs and T-shaped bond with one lone pair at equatorial position and other lone pair of electrons at axial positions.

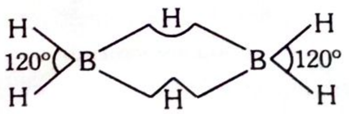
So the correct geometry of T-shaped molecule (actually trigonal bipyramidal) is as drawn in this figure :



So the bond lengths are different as

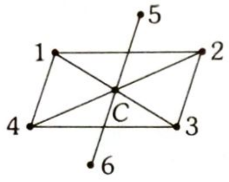
$F-Cl_{\text{axial}} > F-Cl_{\text{equitorial}}$.

43. (c) Each 'B' atom in diborane (B_2H_6) is sp^3 –hybridized of the four hybrid orbitals, three have one electron each, while the fourth is empty. Two orbitals of each B from σ – bonds with two H–atoms, while one of the remaining hybrid orbitals (either filled or empty), 1s orbital of H–atom and one of the hybrid orbitals of other B atom overlap to form three centred two electron bond. So there exists two such type of three centred bond.

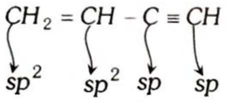


44. (d) Carbon is sp^2 hybridized in graphite so the pair of electrons are concentrated above and below the layers (planar structure). These electrons are free to move and contributes to conduction.

45. (c) In an octahedral structure 4 orbitals lie on x-y plane and 2 orbitals are perpendicular to the plane (axis positions). The four orbitals involved in forming planar structure are $s, p_x, p_y, d_{x^2-y^2}$. Two more perpendicular orbitals will be formed in 2-direction by mixing of p_z and d_{z^2} orbitals.



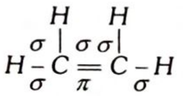
46. (b)



$CH_2 = CH - C \equiv CH$

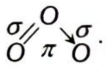
$sp^2 \quad sp^2 \quad sp \quad sp$

47. (c)

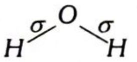


$H \overset{\sigma}{-} C \underset{\pi}{=} C \overset{\sigma}{-} H \Rightarrow 5\sigma$ and 1π bonds.

48. (b)

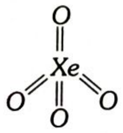


49. (c)



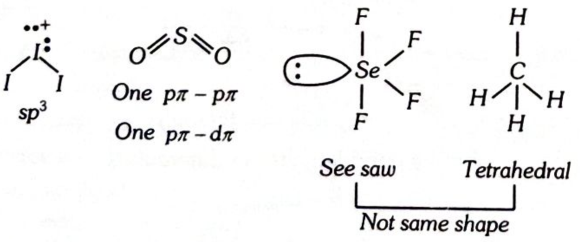
2σ bonds

50. (a) XeO_4

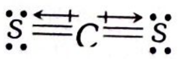


$4\sigma, 4\pi$ bonds are present.

51. (d)

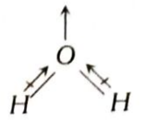


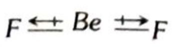
52. (c) The structure of CS_2 is similar to CO_2 which has zero dipole moment.



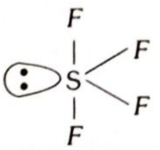
53. (b) CH_4 have regular tetrahedron so its dipole moment is zero.

54. (b) The overall value of the dipole moment of a polar molecule depends on its geometry and shape *i.e.*, vectorial addition of dipole moment of the constituent bonds water has angular structure with bond angle 105° as it has dipole moment. However BeF_2 is a linear molecule since dipole moment summation of all the bonds present in the molecule cancel each other.

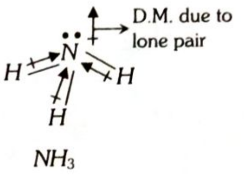




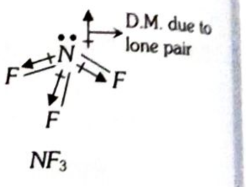
55. (c) Due to distorted tetrahedral geometry SF_4 has permanent dipole moment



56. (d)



In NH_3, the dipole moment due to lone pair is also in the same direction.



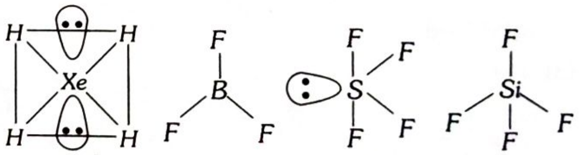
In NF_3, the dipole moment due to lone pair is in opposite direction.

57. (d) Dipole - induced dipole interaction occurs between polar & Non-polar molecule.

HCl	He
Polar	Non-polar
$\mu \neq 0$	$\mu = 0$

58. (a) CO_2 and CH_4 are non-polar molecules. Dipole moment of NH_3 is greater than NF_3.

59. (c) Unsymmetrical distribution of e^- cloud leads to the formation of polar molecule.



$\mu = 0$ $\quad \mu = 0$ $\quad \mu \neq 0$ $\quad \mu = 0$

60. (a) $O_2 \xrightarrow{+1e-} O_2^-$

Molecular orbital configuration of O_2

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

$\pi^* 2p_x^{\ 1} = \pi^* 2p_y^{\ 1}$

therefore incoming e^- would enter into $\pi^* 2p_x$ molecular orbital.

61. (c) Bond order in Be_2 is zero.

62. (b) Bond order $N_2 = 3$

$N_2^- = 2.5$

$N_2^{2-} = 2.0$.

63. (d) The electronic configuration of O_2^-

$O_2^- \to (\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2$

$(\pi 2p_x)^2 (\pi 2p_y)^2 (\pi^* 2p_x)^2 (\pi^* 2p_y)^1$

Bond order = $1/2\,[N_b - N_a]$

$= 1/2[10-7] = 3/2 = 1.5$.

64. (a) Both O_2^{2-} and B_2 had bond order equal to 1.

65. (b) $He_2^+ = 3e^-$; $B.O. = 0.5$

$NO = 15e^-$; $B.O. = 2.5$

$O_2^- = 17e^-$; $B.O. = 1.5$

$C_2^{2-} = 14e^-$; $B.O. = 3$.

66. (c) In O_2^- and NO both have unpaired electrons in molecular orbitals so they are paramagnetic.

67. (c) Due to unpaired e^- ClO_2 is paramagnetic.

68. (d) Correct sequence of $B.O.$ in increasing order is

$$\underset{(B.O. = 2)}{O_2} > \underset{(B.O. = 1.5)}{O_2^-}$$

$$\underset{(B.O. = 3)}{N_2} > \underset{(B.O. = 2.5)}{N_2^-}$$

Since $B.O.$ is fractional so cannot be diamagnetic order of bond length is $O_2^- > O_2$.

69. (d) BCl_3 is sp^2 hybridized and central atom does not have any lone pair of electrons, so all bond angles are 120°.

70. (c) Both CN^- and CO have equal number of electrons, So their bond order will be same.

71. (b) $CN^- \Rightarrow$ bond order is 3

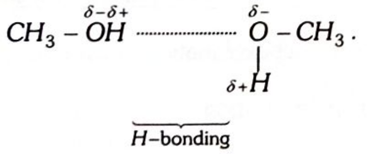
$NO \Rightarrow$ bond order is 2.5

$CN^+ \Rightarrow$ bond order is 2

$CN \Rightarrow$ bond order is 2.5

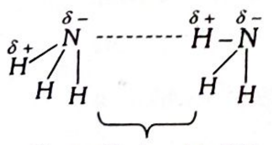
Hence, CN^- has higher bond order.

72. (a)



$$\underbrace{CH_3 - \overset{\delta-\delta+}{OH} \cdots\cdots \overset{\delta-}{\underset{\overset{|}{\delta+H}}{O}} - CH_3}_{H\text{–bonding}}.$$

73. (d)



Strong H – bond in NH_3

74. (a) Hydrogen bonding in ice creates a cage like structure of water molecules having vacant spaces in between which reduces the density of ice.

75. (a) Cellulose is a polysaccharide with the formula $[C_6H_7O_2(OH)_3]_n$ where the oxygen atom and the hydrogen atom connected to other oxygen atom (in $-H$ group) can form intramolecular hydrogen bond.

76. (d) Due to delocalized and mobile electrons present in metallic bond, they are non-directional bonds.

15. AIIMS

1. (c) BCl_3 is electron deficient compound because it has only '6' electrons after forming bond.

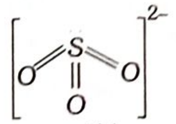
2. (c) When sulphur reacts with chlorine in 1 : 2 ratio, then SCl_4 is obtained which on hydrolysis gives sulphurous acid (H_2SO_3).

$$\underset{1}{S} + \underset{2}{2Cl_2} \rightarrow \underset{(X)}{SCl_4}$$

$$SCl_4 + 4H_2O \rightarrow S(OH)_4 + 4HCl$$

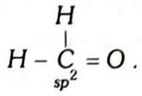
$$S(OH)_4 \rightarrow \underset{\substack{\text{sulphurous acid}\\(Y)}}{H_2SO_3} + H_2O$$

The anion of (Y) is SO_3^{2-}



pyramidal
(sp^3 hybridisation)

3. (b) In $HCHO$, carbon is sp^2 hybridized

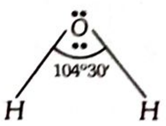


$$H - \underset{sp^2}{\overset{\overset{H}{|}}{C}} = O.$$

4. (b) Bond angle increases with change in hybridisation in following order $sp^3 < sp^2 < sp$.

5. (d) $H = \frac{1}{2}(V + M - C + A)$

$H = \frac{1}{2}(6 + 2 - 0 + 0) = 4$



So it will show sp^3 hybridization with 2 lone pairs and bent structure. Bond angle is 104.5°.

6. (d) dsp^3 hybrid orbitals have bond angles $120°, 90°$.

7. (a) It is shows sp^3d^3–hybridization. Hence the bond angle is about $72°$.

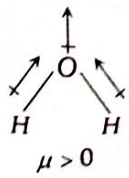
8. (d) C in CO_3^{2-} and B in BCl_3 are sp^2 hybridised thus acquire planar geometry. Although N in $N(SiH_3)_3$ is sp^3 hybridised but the steric hindrance caused by three bulky groups make it planar for higher stability.

9. (b) CO_2 is a symmetrical molecule so its dipole moment is zero.

$$\underset{\longleftrightarrow\ \ \longleftrightarrow}{O = C = O}$$

$\mu = 0$

(a)



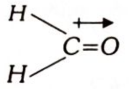
(c) $\underset{\longleftrightarrow}{H - F}$ $\mu > 0$

(d) $\underset{\longleftrightarrow}{H - Br}$ $\mu > 0$

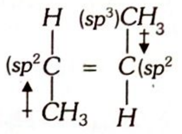
10. (a) The dipole moment of two dipoles inclined at an angle θ is given by the equation $\mu = \sqrt{X^2 + Y^2 + 2XY\cos\theta}$

$\cos 90° = 0$. Since the angle increases from 90 to 180, the value of $\cos\theta$ becomes more and more $-ve$ and hence resultant decreases. Thus, dipole moment is maximum when $\theta = 90°$.

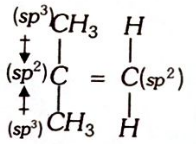
11. (a)



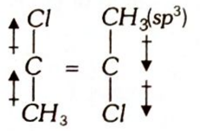
High D.M. due to high electronegativity of oxygen atom



$\mu = 0$



$\mu = 0$

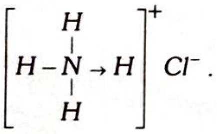


$\mu = 0$

12. (d) Ammonia molecule is more basic than nitrogen trifluoride and Boron trifluoride.

13. (b) When electronegativity difference is more between two joined atoms then covalent bond becomes polar and electron pair forming a bond don't remain in the centre.

14. (b) NH_4Cl has both types of bonds polar and non-polar



$$\left[\begin{array}{c} H \\ | \\ H-N\rightarrow H \\ | \\ H \end{array}\right]^+ Cl^-.$$

15. (a) We know that greater the difference in electronegativity of two atoms forming a covalent bond. More is its polar nature. In HF there is a much difference in the electronegatives of hydrogen and fluorine. Therefore (HF) is a most polar compound.

16. (c) Oxygen is paramagnetic due to the presence of two unpaired electron:

$$O_2 = \sigma(1s)^2\sigma^*(1s)^2\sigma(2s)^2\sigma^*(2s)^2$$

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

17. (a) No. of electrons in $S^{2-} = 18$

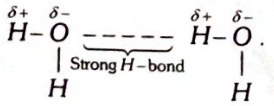
$$S^{2-} = KK'(\sigma 2s)^2(\sigma^* 2s)^2(\sigma 2p_z)^2(\pi 2p_x)^2(\pi 2p_y)^2$$

$$(\pi^* 2p_x)^2(\pi^* 2p_y)^2$$

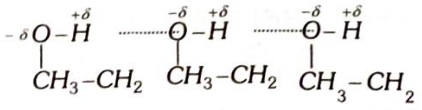
No. unpaired electrons ⇒ Dimagnetic.

18. (c) Nitrogen molecule has highest bond energy due to presence of triple bond.

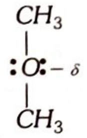
19. (d) Hydrogen bonding increases the boiling point of compound.



20. (a)

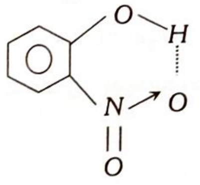


Hydrogen bonding in ethanol

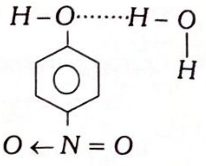


Hydrogen is not available to form hydrogen bonding in dimethyl ether.

21. (c)



o- Nitrophenol



p-Nitrophenol

p-nitrophenol forms intermolecular hydrogen bond so it has high boiling point and less volatile.

22. (a) Generally, zero group elements are linked by the vander Waals force. Hence these show weakest intermolecular forces.

16. Assertion & Reason

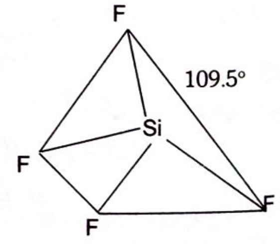
1. (a) Solubility in water depends on hydration energy and lattice energy.

2. (a) Polarity in covalent bond is developed due to shifting of electrons towards one of the bonded atoms.

3. (a) Boron has only 3 valence electrons so it cannot form expected number of covalent bonds to complete its octet.

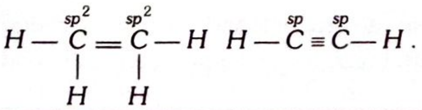
4. (a) Reason is the correct explanation of assertion.

5. (c) SiF_4 have sp^3 hybridization & shape of regular tetrahedral where the bond angle of $F-Si-F$ are found $109.5°$ which is greater than $90°$ but less than $180°$.



Repulsion sequence are $lp-lp > lp-bp > bp-bp$ so assertion are true but the reason is false.

6. (c) The correct reason is the carbon atom is sp^2 hybridised in ethene and sp hybridised in ethyne.



$$\begin{array}{c} H-\overset{sp^2}{C}=\overset{sp^2}{C}-H \\ \;\;\,|\quad\;\; | \\ \;\;H\quad\; H \end{array} \quad H-\overset{sp}{C}\equiv\overset{sp}{C}-H.$$

7. (a) It is correct that during formation of Ice from water there are vacant spaces between hydrogen bonded molecules of Ice. Ice has a cage like structure. Due to this reason Ice is less dense than liquid water. hence both assertion and reason are true and reason is the correct explanation of assertion.

8. (a) o and p-nitrophenols can be separated by steam distillation because o-nitrophenol is steam volatile. Here, both assertion and reason are correct and reason is correct explanation of assertion.

9. (d) In B_2, total number of electrons = 10

$B_2 \rightarrow \sigma(1s)^2\ \sigma^*(1s^2)\ \sigma(2s)^2\ \sigma^*(2s)^2\ \sigma(2p_x)^1\ \pi(2p_y)^1$

Presence of unpaired electron shows the paramagnetic nature.

The highest occupied molecular orbital is of π-type.

10. (c) S–atom in S_8 molecule are sp^3–hybridised and contain two lone pair of electrons on each and exist as staggered eight atom rings.

11. (a) MO electronic configuration of F_2 molecule is

$F_2 = \sigma 1s^2, \sigma * 1s^2, \sigma 2s^2, \sigma * 2s^2, \sigma 2p_x^2,$

$\pi 2p_y^2 \approx \pi 2p_z^2, \pi * 2p_y^2 \approx \pi * 2p_z^2$

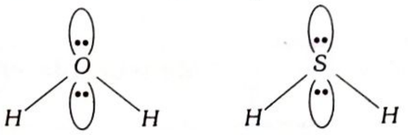
Bond order $= \frac{N_b - N_a}{2}$

Where, N_b =Number of electrons in bonding molecular orbitals

N_a = Number of electrons in antibonding molecular orbitals $= \frac{8-6}{2} = 1$.

12. (c) Assertion : $H-O-H$ bond angle in $H_2O = 104.5°$

$H-S-H$ bond angle in $H_2O \approx 90° (> 90°)$

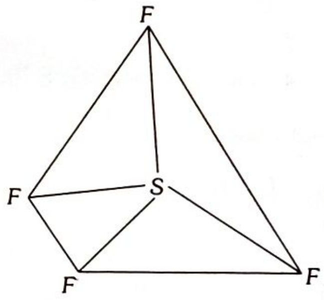


Reason : The lone pairs gets bigger and more diffuse as we go down in a group which means that $lp-lp$ repulsion will decrease rather than increase.

$(lp-lp)_{H_2O} > (lp-lp)_{H_2S}$

Reduction in bond angle has to do with the bond pair – bond pair repulsion. The $O-H$ bond is more polar because of the high electronegativity of oxygen (than sulphur). This means that a lot of electron density of the bonds is pulled towards the oxygen atom. Thus there is a large $bp-bp$ repulsion. This effect will decrease in H_2S.

13. (c) SF_4 has sp^3 hybridization and shape of regular tetrahedral where the bond angle of $F-S-F$ are found 109.5° which is greater than 90° but less than 180°.



Repulsion sequence is $lp-lp > lp-bp > bp-bp$ so assertion is true but the reason is false.