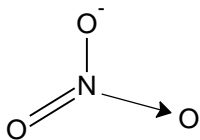
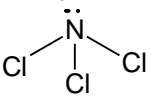
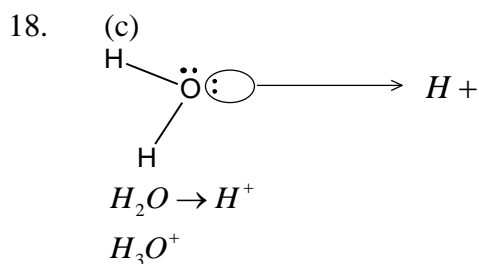
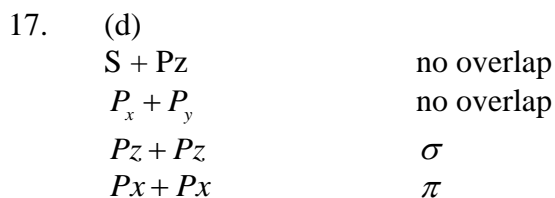
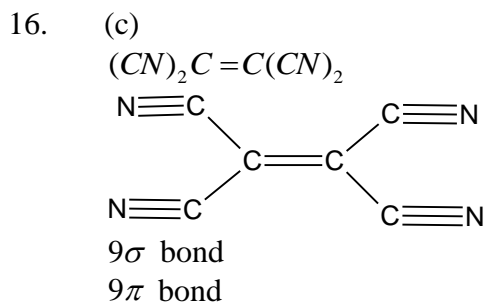
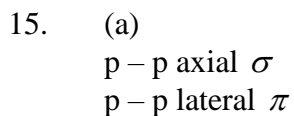
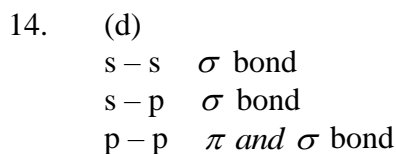
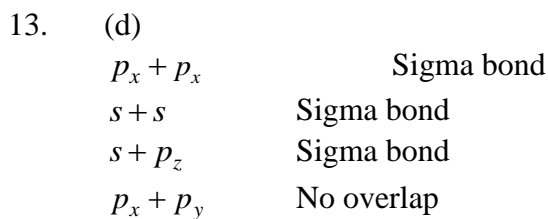
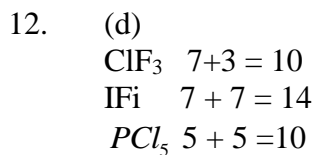
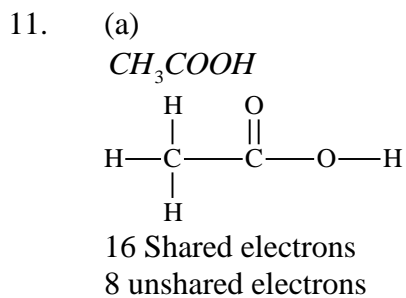


**EXERCISE - 1 [A]**

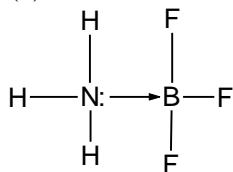
- (b)  
Lower the ionisation energy, cation is easily formed. Similarly higher the electron affinity, anion is easily formed.
- (c)  
 $K^+I^-$
- (b)  
 $H-Cl$
- (b)  
$$L.E. \propto \frac{q^{\oplus}q^{\ominus}}{r^{\oplus} + r^{\ominus}}$$

Charge	↑	L.E.	↑
Size	↓	L.E.	↑
- (b)  
Covalent bond (directional)
- (d)  
In  $ClF_3$ , Cl has 10 valence electrons.
- (d)  


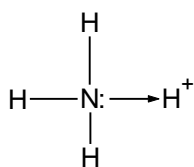
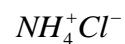
4 bond pair electrons & 0 lone pair electrons.
- (d)  
 $NCl_3$   

- (c)  
 $AgNO_3 + CCl_4 \rightarrow$  No reaction
- (c)  
 $K^+CN^-$   
 $^-C \equiv N$



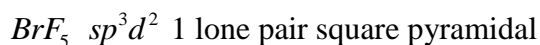
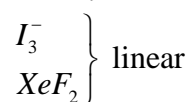
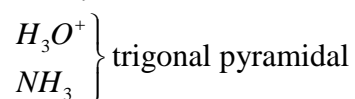
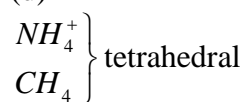
19. (c)



20. (b)



21. (d)

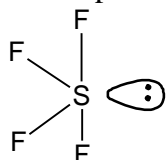


22. (b)



4 $\sigma$  bond

1 lone pair



$sp^3d$

G: trigonal bipyramidal

S : See – saw

23. (d)



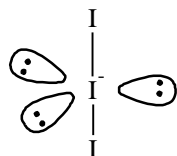
2 $\sigma$  bond

3 lone pair

$sp^3d$

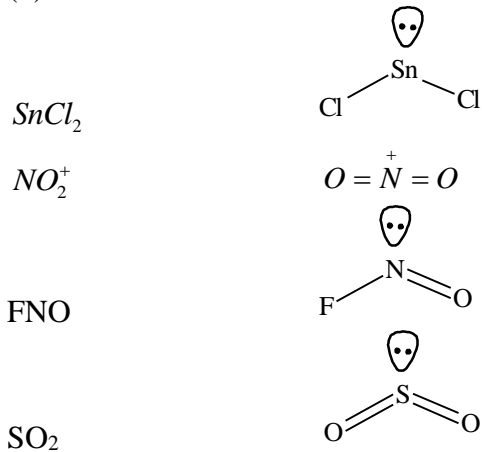
G: trigonal bipyramidal

S : linear



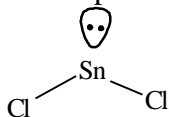
24. (a)  
 $PCl_6^-$  6 $\sigma$  bond and 0 lone pair  $sp^3d^2$  octahedral  
 $SF_4$   $sp^3d$  (see saw)  
 $BO_3^{3-}$   $sp^2$  (trigonal planar)  
 $BF_4^-$   $sp^3$  (tetrahedral)

25. (b)

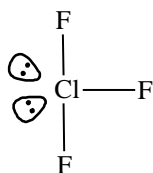


26. (c)  
 2 $\sigma$  and 2 $\pi$  bond 1 lone pair }  $sp^2$  trigonal planar.

27. (b)  
 $MCl_2$   
 Atomic no. = 50 (Sn)  
 $SnCl_2$   
 1 lone pair and 2  $\sigma$ -bond



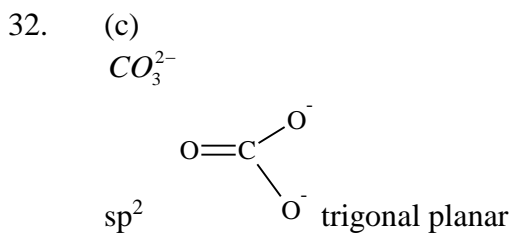
28. (c)  
 $ClF_3$   
 2 lone pair }  $sp^3d$   
 3 $\sigma$  bond }



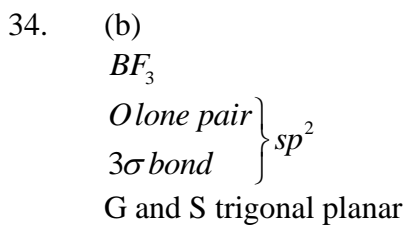
G : trigonal bipyramidal  
 S : T-shape

29. (d)  
 ${}^1C H_2 = {}^2C H - {}^3C \equiv {}^4C H$   
 $\uparrow \quad \uparrow$   
 $sp^2 \quad sp$

30. (b)  
Size  $sp < sp^2 < sp^3$
31. (a)  
 $CO_2$   
0 lone pair, 2  $\sigma$  bond, sp (linear)  $O = C = O$

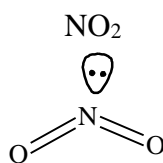
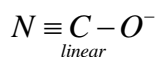
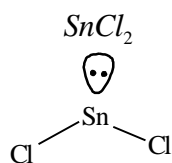


33. (b)  
 $s$  and  $p$   
 $sp$  hybridisation angle =  $180^\circ$



35. (a)  
Iso electronic  $\left\{ \begin{array}{l} NO_3^- \quad sp^2 \\ CO_3^{2-} \quad sp^2 \end{array} \right\}$  trigonal planar  
 $SO_3$   $sp^2$  trigonal planar  
 $ClO_3^-$   $sp^3$  trigonal pyramidal

36. (c)  
 $CO_2$   
 $O = C = O$  linear



37. (b)  
 $C_2H_2$   
 $H - C \equiv C - H$   
0 lone pair, 2  $\sigma$  bond, sp linear

38. (b)  
In  $F_2$ ,  $P_z$  &  $P_z$  orbitals overlap to form  $\sigma$  bond.

39. (b)

$$sp^2 \quad \text{and} \quad sp^3d$$

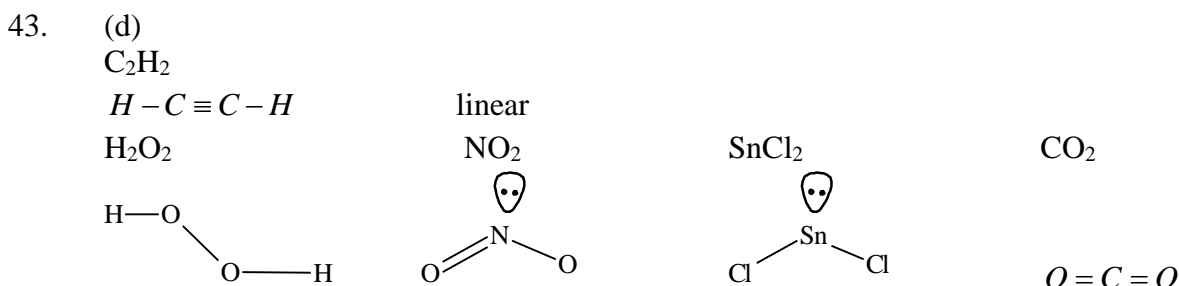
$$\%s = \frac{1}{3} \times 100 \quad \frac{1}{5} \times 100$$

$$= 33.33\% \quad = 20\%$$

40. (b)  
Electronegativity  $sp > sp^2 > sp^3 > sp^3d$

41. (b)  
 $SO_3$   $sp^2$   
 $PCl_5$   $sp^3d$   
 $SO_2$   $sp^2$   
 $PCl_3$   $sp^3$

42. (c)

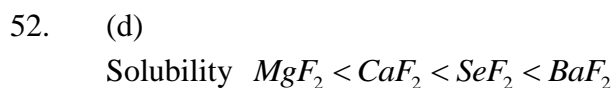
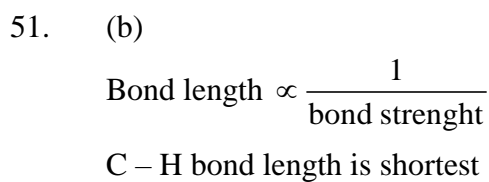
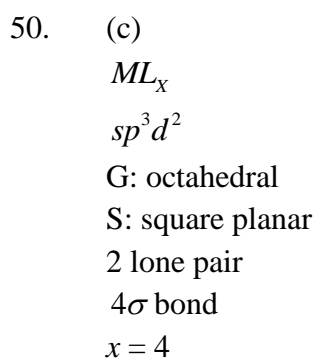
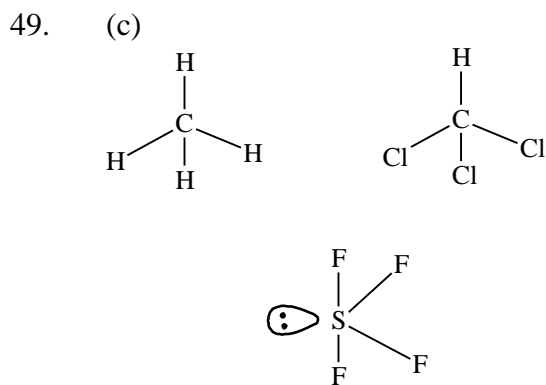
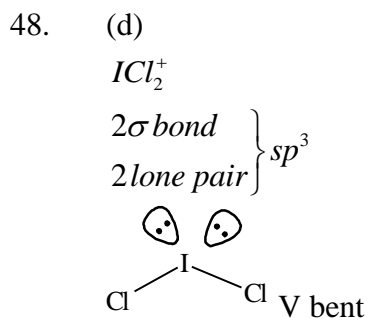
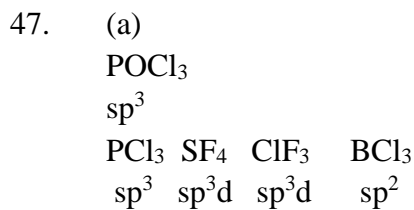
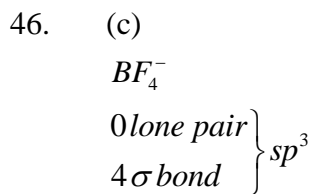
$$OF_2 \left. \begin{array}{l} 2\sigma \text{ bond} \\ 2 \text{ lone pair} \end{array} \right\} sp^3$$


44. (b)

$BF_3$   $sp^2$  (trigonal planar)  
 $NH_3$   $sp^3$  (tetrahedral)  
 $AlF_3$   $sp^2$  (trigonal planar)  
 $BeF_2$   $sp$  (linear)  
 $H_2O$   $sp^3$  (tetrahedral)  
 $BCl_3$   $sp^2$  (trigonal planar)  
 $PCl_3$   $sp^3$  (tetrahedral)

45. (a)

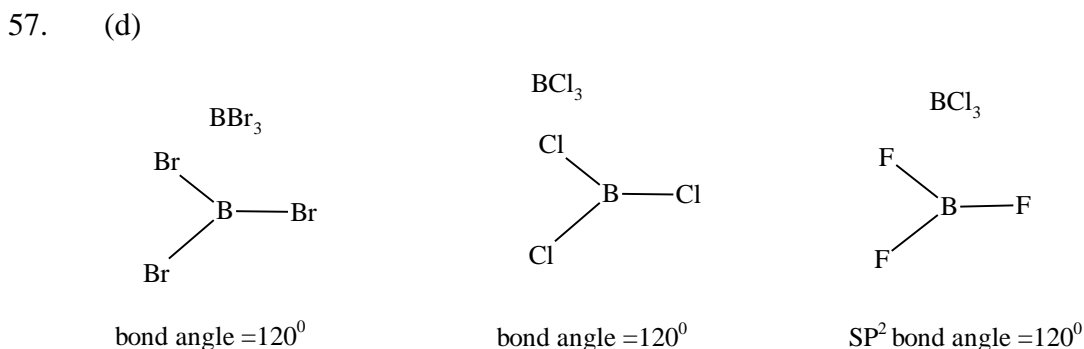
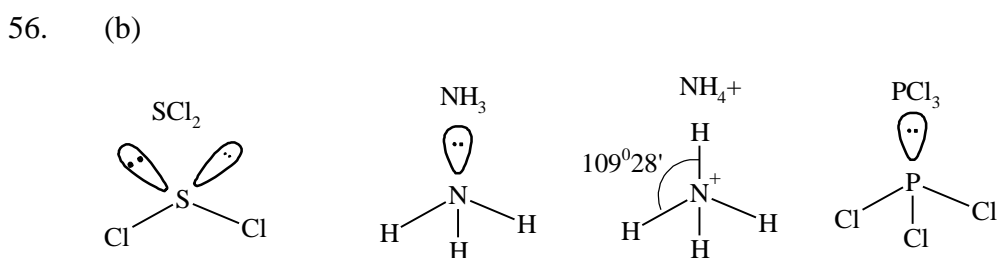
$PCl_3$   $sp^3$  (trigonal pyramidal)  
 $NH_3$   $sp^3$  (trigonal pyramidal)  
 $CF_4$   $sp^3$  (tetrahedral)  
 $SF_4$   $sp^3d$  (tetrahedral)  
 $PbCl_2$   $sp^2$  (V-bent)  
 $CO_2$   $sp$  (Linear)  
 $PF_5$   $sp^3d$  (trigonal bipyramidal)  
 $IF_5$   $sp^3d^2$  (square pyramidal)



53. (b)  
 Covalent correct  
 $\text{LiCl} > \text{RbCl}$   
 $\text{LiCl} < \text{BeCl}_2$   
 $\text{BeCl}_2 > \text{MgCl}_2$

54. (b)  
 $\text{AgI}$   
 $\text{I}^-$  Large anion  
 Greatest covalent character least soluble

55. (a)  
 $\text{AgCl}$   
 Greater covalent  $\text{Ag}^+$  (less soluble)  
 PIGC (pseudo inert gas centi)



58. (d)

	$\text{CO}$	$\text{CO}_3^{2-}$	$\text{CO}_2$
B.O.	3	$4/3$	2

$b.l \propto \frac{1}{b.o}$   
 $b.l : \text{CO}_3^{2-} > \text{CO}_2 > \text{CO}$

59. (b)  
 Least covalent character highest melting point.

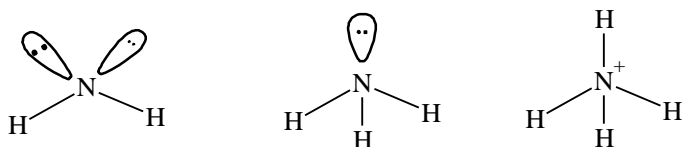
60. (c)  
 Covalent  $\text{LiCl} > \text{NaCl}$   
 Solubility in water :  $\text{LiCl} < \text{NaCl}$   
 Organic solvent :  $\text{LiCl} > \text{NaCl}$



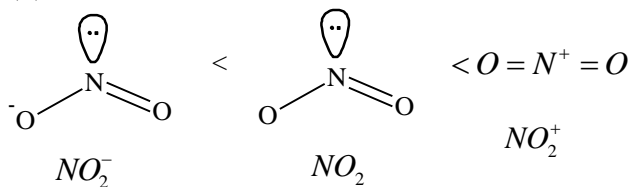
61. (c)  
Covalent  $BeCl_2 > MgCl_2 > CaCl_2 > SrCl_2 > BaCl_2$

62. (c)  
Charge  $\uparrow$  Polarisation  $\uparrow$  covalent characters  $\uparrow$   
 $Na_3N$   
 $F^- < O^{2-} < N^{3-}$

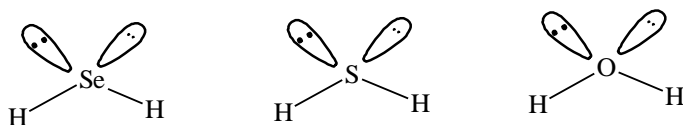
63. (b)  
 $NH_2^- < NH_3 < NH_4^+$



64. (d)  
 $NO_2^- < NO_2 < NO_2^+$



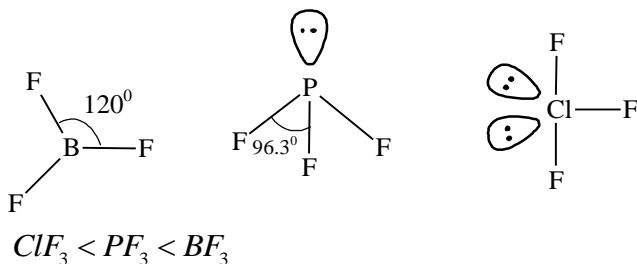
65. (c)  
 $H_2Se < H_2S < H_2O$



66. (d)  
 $BF_3$   $BCl_3$   $BBr_3$   
 $sp^2$   $sp^2$   $sp^2$  trigonal planar bond angle =  $120^\circ$

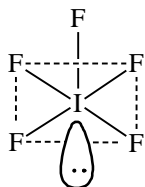
67. (b)  
 $NH_3$   $CH_4$   $H_2O$   $H_2S$   
 $107^\circ$   $109^\circ 28'$   $104.5^\circ$   $92^\circ$

68. (c)  
 $BF_3$   $PF_3$   $ClF_3$



69. (b)  
Covalent Character  $KCl < CaCl_2 < AlCl_3 < SiCl_4$

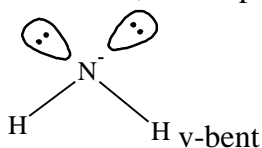
70. (c)  
 $IF_5$   
1 lone pair  
5  $\sigma$  bond  
 $sp^3d^2$  octahedral  
S: square pyramidal  
 $90^\circ$  and  $180^\circ$



71. (d)  
Size  $Na^+ > Mg^{2+}$   
 $S^{2-} > Cl^-$   
Covalent  $MgS > NaCl$   
Solubility in water :  $MgS < NaCl$

72. (d)  
 $I^-$  largest anion (greater polarisation)  
Covalent  $AlI_3 > AlBr_3 > AlCl_3 > AlF_3$

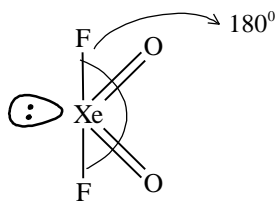
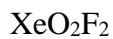
73. (c)  
 $NH_2^-$   
2  $\sigma$  bond, 2 lone pair }  $sp^3$



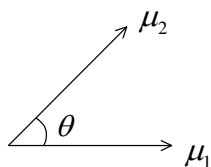
74. (c)

Bond  $H_2Se < NH_4^+ < NOCl < SO_3$  angle

75. (c)



76. (a)

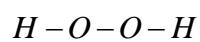


$$\mu_{net} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos \theta}$$

$$0 < \theta < 90^\circ \quad \theta \uparrow \mu_{net} \uparrow$$

$$90^\circ < \theta < 180^\circ \quad \theta \uparrow \mu_{net} \downarrow$$

77. (c)



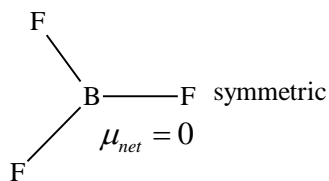
Non polar bond

78. (c)

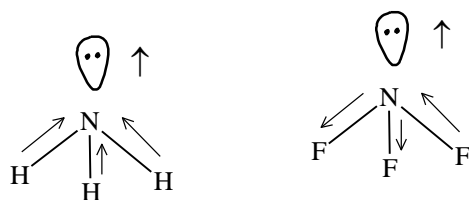
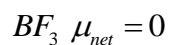


$\Delta(E.N) \uparrow$  Polar character  $\uparrow$

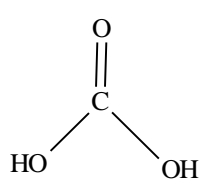
79. (b)



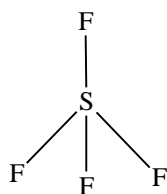
80. (d)



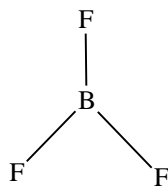
81. (a)



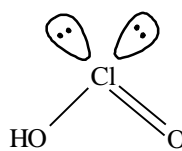
$sp^2$   
Polar



$sp^3$   
Non polar

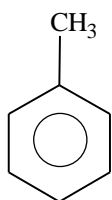


$sp^2$   
Non polar

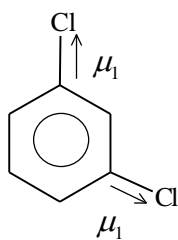


$sp^3$   
Polar

82. (b)



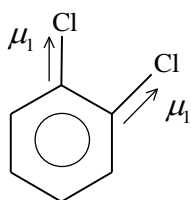
I



II

$$\theta = 120^\circ$$

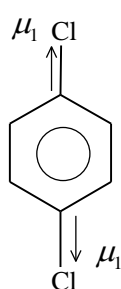
$$\mu_{net} = \mu_1$$



III

$$\theta = 60^\circ$$

$$\mu_{net} = \sqrt{3}\mu_1$$



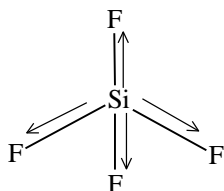
IV

$$\theta = 180^\circ$$

$$\mu_{net} = 0$$

$$IV < I < II < III$$

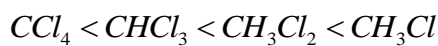
83. (c)



$$\mu_{net} = 0$$

84. (a)

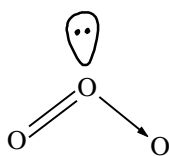
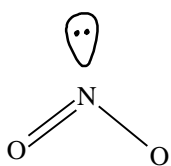
Dipole moment



85. (b)

$$\mu_{net} \text{ of } \text{SiF}_4 \text{ and } \text{CO}_2 = 0$$

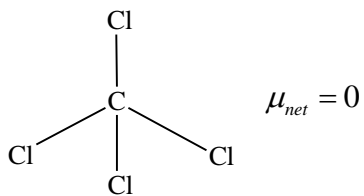
$$\mu_{net} \text{ of } \text{NO}_2 \text{ and } \text{O}_3 \neq 0$$



86. (c)

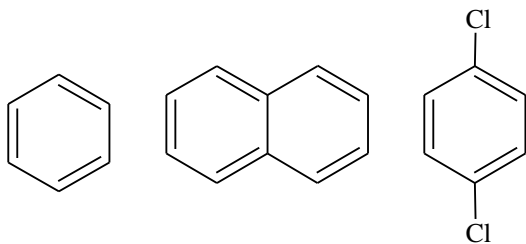
Covalent (directional)

87. (d)



88. (b)  
C – H  
 $\Delta(E.N)$  is very less

89. (d)



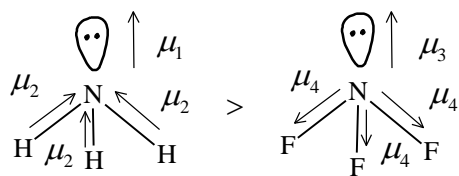
90. (a)
- 
- $\mu_{net} \neq 0$
- Other  $\mu_{net} = 0$   
( $CO_2, BF_3, CCl_4$ )

91. (d)
- 
- $\mu_{net} = 0$

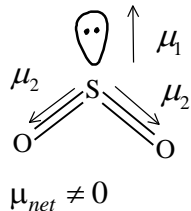
92. (c)
- 
- $\mu_{net} \neq 0$
- Others  $\mu_{net} = 0$

93. (b)

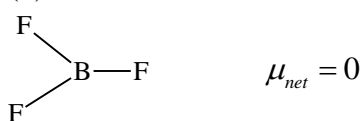
$\mu_{net}$  of  $BF_3$  and  $B_2H_6 = 0$



94. (c)



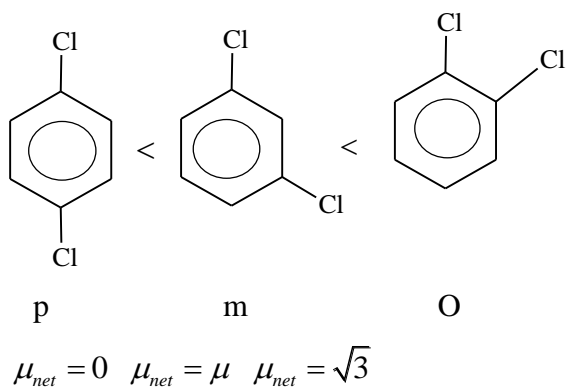
95. (b)



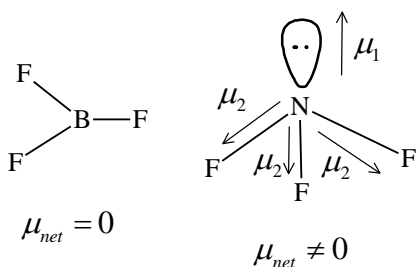
96. (a)

$\mu_{net}(NF_3) < \mu_{net}(NH_3)$   
 $\mu_{net}$  of others = 0

97. (d)



98. (c)



99. (a)

$\Delta(E.N) \uparrow$  ionic  $\uparrow$   
 $HI < HBr < HCl < HF$

100. (c)

	$O_2$	$H_2O_2(O_2^{2-})$	$O_3$
B.O	2	1	1.5

*b.l.*:  $O_2 < O_3 < H_2O_2$

101. (a)

	$N_2$	$N_2^+$	$O_2$	$O_2^+$
B.O	$\frac{6-0}{2} = 3$	$\frac{5-0}{2} = 2.5$	$\frac{6-2}{2} = 2$	$\frac{6-1}{2} = 2.5$

*b.l.*:  $N_2 < N_2^+$

$O_2^+ < O_2$

102. (d)

	$C_2$	$CN$	$N_2$	$O_2$
B.O	$\frac{4-0}{2} = 2$	$\frac{5-0}{2} = 2.5$	$\frac{6-0}{2} = 3$	$\frac{6-2}{2} = 2$
B.O.	$\frac{3-0}{2} = 1.5$	$\frac{4-0}{2} = 2$	$\frac{5-0}{2} = 2.5$	$\frac{6-1}{2} = 2.5$

103. (c)

Energy  $\sigma_{1s}^* < \sigma_{2s}$

$n = 1 \quad n = 2$

104. (c)

Only same energy atomic orbitals combine to give molecular orbital.

105. (d)

$$O_2^{2-} \quad B.O. = \frac{6-4}{2} = \frac{2}{2} = 1$$

106. (a)

	$O_2^+$	$N_2^-$	$CN^-$	$NO^+$	CO
B.O	$\frac{6-1}{2} = 2.5$	$\frac{6-1}{2} = 2.5$	$\frac{6-0}{2} = 3$	$\frac{6-0}{2} = 3$	$\frac{6-0}{2} = 3$

107. (d)

	$H_2^+$	$H_2^-$	$H_2$
B.O	$\frac{1-0}{2} = 0.5$	$\frac{2-1}{2} = 0.5$	$\frac{2-0}{2} = 1$

*b.l.*:  $H_2 < H_2^+ < H_2^-$

108. (b)

	$O_2$	$O_2^+$	$O_2^-$	$O_2^{2-}$
B.O.	$\frac{6-2}{2} = 2$	$\frac{6-1}{2} = 2.5$	$\frac{6-3}{2} = 1.5$	$\frac{6-4}{2} = 1$

Bond strength  $O_2^{2-} < O_2^- < O_2 < O_2^+$

109. (a)

	$NO$	$NO^+$	$NO^-$
B.O.	$\frac{6-1}{2} = 2.5$	$\frac{6-0}{2} = 3$	$\frac{6-2}{2} = 2$

B.E.  $NO^- < NO < NO^+$

110. (b)

$O_2$	$O_2^{2-}$	$O_2^-$	$O_2^+$
2	4	3	1

111. (a)

$O_2$  has unpaired electrons in antibonding molecular orbitals.

112. (a)

$O_2$  is paramagnetic with 2 unpaired electrons.

113. (b)

	Unpaired electrons
$B_2$	2
$O_2$	2
$NO$	1

114. (b)

$$\text{B.O.} = \frac{2-0}{2} = 1$$

115. (a)

$C_2$

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

116. (c)

$O_2$  Paramagnetic

$O_2^-$  Paramagnetic

$C_2$  Diamagnetic

$N_2^+$  Paramagnetic

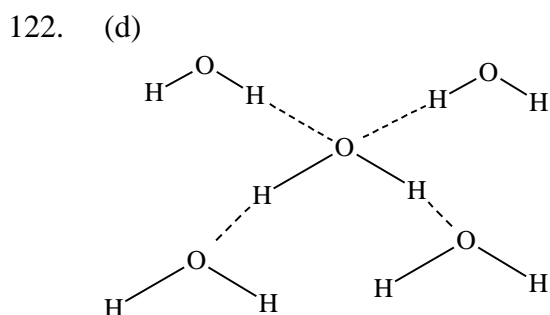
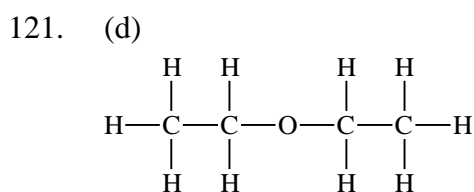
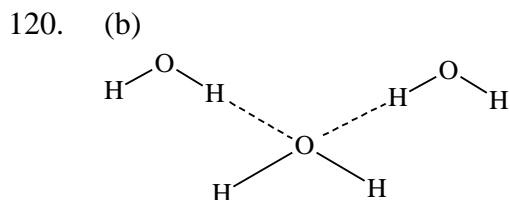
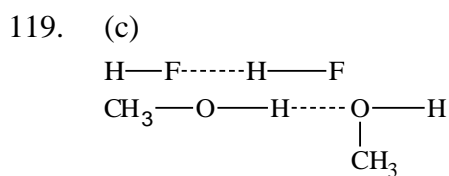
117. (c)

$CaC_2$   $C_2^{2-}$

$$b.o. = \frac{6-0}{2} = 3$$



118. (b)  
For antibonding molecular orbitals,  
Energy increases on decreasing distance between atoms.



123. (b)  
Volatility  $\uparrow$  if intermolecular hydrogen  $\downarrow$  and  
B.P.  $\downarrow$       bonding  $\downarrow$   
Molar mass  $\uparrow$  B.P.  $\uparrow$  volatility  $\downarrow$   
B.P.  $\text{HCl} < \text{HBr} < \text{HI} < \text{HF}$   
Volatility :  $\text{HCl} > \text{HBr} > \text{HI} > \text{HF}$

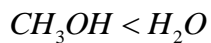
124. (a)  
B.P.  $\text{H}_2\text{S} < \text{H}_2\text{Se} < \text{H}_2\text{Te} < \text{H}_2\text{O}$   $\uparrow$  due to intermolecular H-bonding

125. (d)  
HCl (no H – bonding)

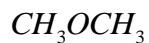
126. (a)  
H-bonding Strength  
 $\text{H} \cdots \cdots \text{F} > \text{H} \cdots \cdots \text{O} > \text{H} \cdots \cdots \text{N}$

127. (b)  
Bond length of H – bond  $>$  bond length of covalent bond

128. (b)

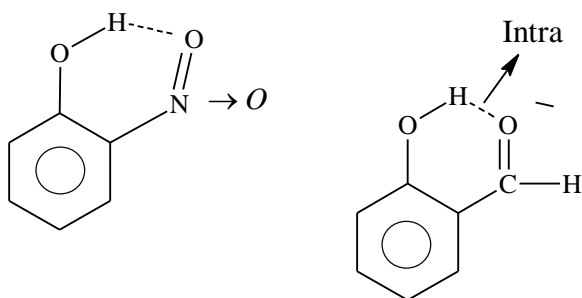


Greater strength of H-bonding in water

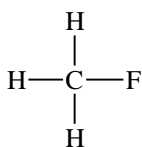


No H-bonding

129. (d)



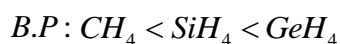
130. (d)



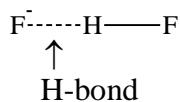
131. (d)

Hydrogen bond is weaker than other bonds

132. (d)

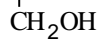
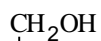
Molar mass  $\uparrow$  *B.P.*  $\uparrow$ 

133. (c)

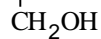
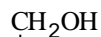


134. (a)

Inter molecular

H-bonding  $\uparrow$  viscosity  $\uparrow$ 

Glycerol

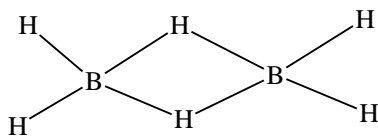


Glycol

135. (b)

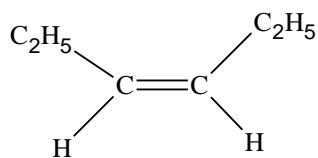
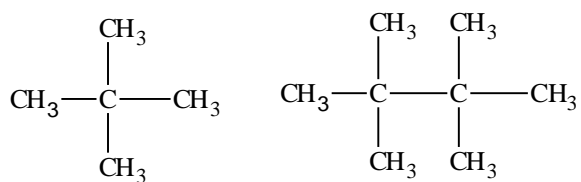
Covalent bond is directional.

136. (d)



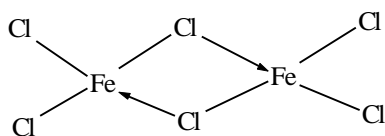
(2c 2e bond or banana bond)

137. (c)



max. dipole moment

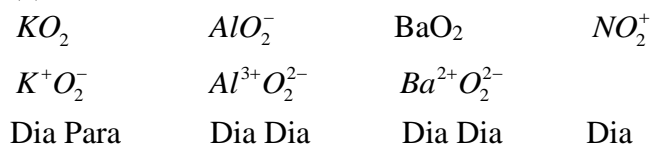
138. (a)



139. (b)

B.P. HF > HBr > HCl

140. (c)



141. (a)

% ionic character

$$= 16\Delta(E.N) + 3.5\Delta(E.N)^2$$

$$= 16(2) + 3.5(2)^2$$

$$= 32 + 14$$

$$= 46$$

142. (d)

Identical bonding is not necessary

143. (d)

Odd electrons molecules are paramagnetic  $ClO_2$

144. (B)

Atomic no = 10 (Ne)

Monoatomic

B.O. of  $Ne_2$  = 0

145. (c)

$CO_2$   $H_2O$   $SiO_2$   $He$   
gas Liquid Solid gas

146. (a)

	$O_2$	$Na_2O_2O_2^-$	$KO_2$	$O_3$
B.O	$\frac{6-2}{2} = 2$	$\frac{6-4}{2} = 1$	$\frac{6-3}{2} = 1.5$	1.5

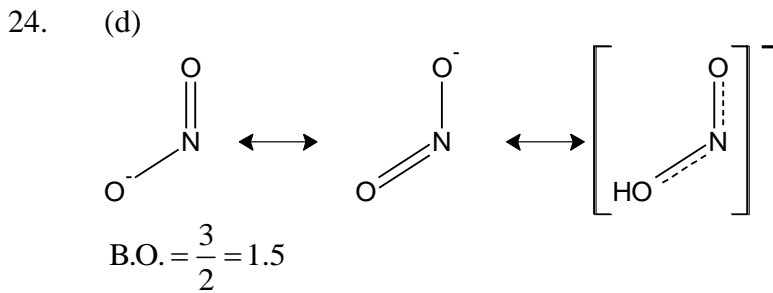
## EXERCISE - 1 [B]

1. (b)  
 $Li^{2+} : 1s^1$   
 $Be^{2+} : 1s^2 = [He]$   
 $B^{2+} : 1s^2 2s^1$   
 $C^{2+} : 1s^2 2s^2$
  
2. (d)  
 $Cl^- = 1s^2 2s^2 2p^6 3s^2 3p^6$   
 $= [Ar]$
  
3. (d)  
 $Ti^{4+} : [Ar]$   
 $Kr : [Ar] 3d^{10} 4s^2 4p^6$   
 $Cl^- : [Ne] 3s^2 3p^6$
  
4. (a)  
 $Ca^+ : [Ar] 4s^1$
  
5. (d)  
 $L.E. \propto \frac{q^+ q^-}{r^+ + r^-}$
  
6. (a)  
 $AB(s) \longrightarrow A^+(g) + B^-(g)$   
 $\Delta H = L.E.$
  
7. (d)  
 $L.E. \propto \frac{q^+ q^-}{r^+ + r^-}$
  
8. (d)  
 $[Kr] 4d^{10}$  is ion of p-block elements or 11<sup>th</sup> or 12<sup>th</sup> group.  
 $Sn = [Kr] 5s^2 4d^{10} 5p^2$   
 $Sn^{4+} = [Kr] 5s^0 4d^{10}$   
 $Ag = [Kr] 5s^1 4d^{10}$   
 $Ag^+ = [Kr] 5s^0 4d^{10}$
  
9. (c)  
 $Fe = [Ar] 4s^2 3d^6$   
 $Fe^{2+} = [Ar] 4s^0 3d^6$

10. (c)  
Bond formed between non-metals is covalent.
11. (b)  
Bond order  $\uparrow$ , Bond length  $\downarrow$
12. (b)  
E.N. of Halogens is higher.
13. (a)  
E.N. of Alkali metals is lower.
14. (d)  
Left to right, E.N.  $\uparrow$
15. (d)  
Dipole moment : HF > HCl > HBr > HI
16. (b)  
As  $\Delta$ E.N.  $\uparrow$  polarity  $\uparrow$   
It is least for *p* & *s*.
17. (b)  
 $\Delta$ E.N.  $\uparrow$  polarity  $\uparrow$   
It is least for O & F & maximum for Mg & F.
18. (a)  
 $\Delta$ E.N.(O-F) = 4 - 3.5 = 0.5  
 $\Delta$ E.N.(Cl-F) = 4 - 3 = 1
19. (b)  
 $\mu = \delta \times d$   
$$\delta = \frac{\mu}{d} = \frac{0.44 \times 3.34 \times 10^{-30}}{1.61 \times 10^{-10}}$$
  
$$\frac{\delta}{e} = \frac{0.44 \times 3.34 \times 10^{-30}}{1.61 \times 10^{-10} \times 1.6 \times 10^{-19}}$$
  
$$= 0.057$$
20. (c)  
H— $\ddot{\text{N}} \equiv \ddot{\text{N}}$ —H
21. (d)  
4 + 3 + 4 + 2 + 7 = 20
22. (b)  

$$S = \begin{array}{c} 3s \qquad \qquad 3p \\ \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow} \quad \boxed{\uparrow} \end{array}$$

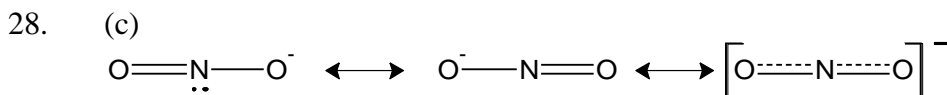
23. (a)
- |                             |    |                 |                               |
|-----------------------------|----|-----------------|-------------------------------|
|                             | CO | CO <sub>2</sub> | CO <sub>3</sub> <sup>2-</sup> |
| Bond order :                | 3  | 2               | 1.33                          |
| Bond order ↑, Bond length ↑ |    |                 |                               |



25. (a)  
In resonance, delocalisation of electrons take place.

26. (c)  
 $x + 3(-1) = 0$   
 $x = +3$

27. (c)  
In resonance, delocalisation of electrons take place.



29. (d)  
In SO<sub>4</sub><sup>2-</sup>, electrons are delocalised on 4 bond pair.

30. (b)  
Due to resonance, bonds are identical in all.

31. (c)  
O can not expand its octet.

32. (c)  
CO has fully filled molecular orbitals of 2p.

33. (d)  
Sr = [Kr] 5s<sup>2</sup>  
Sr = [Kr] 5s<sup>1</sup>

34. (d)  
SeF<sub>4</sub>  
6 + 4 = 10 valence electrons in Se.

35. (c)  
CF<sub>4</sub>  
4 + 4 = 8  
Valence Electrons.

36. (c)  
BF<sub>3</sub>  
3 + 3 = 6  
Valence Electrons.

37. (a)  
ClF<sub>3</sub>  
7 + 3 = 10  
Valence Electrons.  
In SO<sub>3</sub>, coordinate bond can be formed.

38. (c)  
ICl<sub>5</sub>  
7 + 5 = 12  
Valence Electrons.

39. (b)  
BeH<sub>2</sub>  
2 + 2 = 4  
Valence Electrons.

40. (a)

	O <sub>2</sub> <sup>2+</sup>	O <sub>2</sub> <sup>+</sup>	O <sub>2</sub> <sup>-</sup>	O <sub>2</sub> <sup>2-</sup>
Bond order	$\frac{6-0}{2} = 3$	$\frac{6-1}{2} = 2.5$	$\frac{6-3}{2} = 1.5$	$\frac{6-4}{2} = 1$

B. 1 O<sub>2</sub><sup>2-</sup> > O<sub>2</sub><sup>-</sup> > O<sub>2</sub><sup>+</sup> > O<sub>2</sub><sup>2+</sup>

41. (a)  
Bond enthalpy is energy required to break 1 mole of covalent bond.

42. (c)

NO <sup>+</sup>	C <sub>2</sub> <sup>2-</sup>	CN <sup>-</sup>	N <sub>2</sub>
14	14	14	14

*isoelectronic*

43. (d)  
Bond order increasing.  
Bond strength increasing.  
Bond length decreasing.

44. (d)  
Bond order increasing.  
Bond strength increasing.  
Bond length decreasing.

45. (d)  
Bond order increasing.  
Bond strength increasing.  
Bond length decreasing.



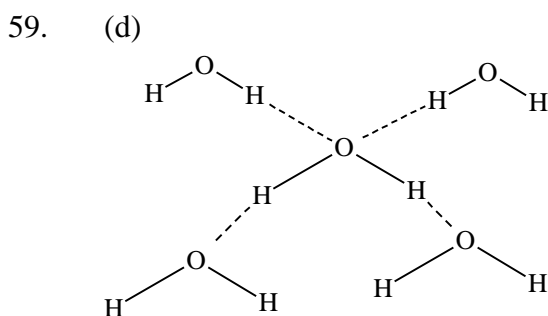
46. (a)  
Solid  $\text{PCl}_5$  exists as  
 $[\text{PCl}_4]^+ [\text{PCl}_6]^-$   
 $sp^3 \quad sp^3d$
47. (b)
- |                |                   |                |              |
|----------------|-------------------|----------------|--------------|
| $\text{O}_2^-$ | $\text{O}_2^{2-}$ | $\text{O}_2^+$ | $\text{O}_2$ |
| Paramagnetic   | Diamagnetic       | Paramagnetic   | Paramagnetic |
48. (b)
- $$\text{H}-\underset{\text{O}}{\underset{\parallel}{\text{C}}}-\text{O}-\text{H}\cdots\text{C}-\underset{\text{O}}{\underset{\parallel}{\text{C}}}-\text{CH}_3$$
49. (d)  
 $\text{PCl}_5$   
 $5\sigma$  &  $0 \ell.p.$   
 $sp^3d$
50. (d)
- |                |                |                |                |
|----------------|----------------|----------------|----------------|
| $\text{BBr}_3$ | $\text{XeF}_4$ | $\text{BCl}_3$ | $\text{XeO}_3$ |
| $sp^2$         | $sp^3d^2$      | $sp^2$         | $sp^3$         |
51. (c)
- |                |               |                  |
|----------------|---------------|------------------|
| $\text{I}_3^-$ | $\text{CO}_2$ | $\text{ClO}_2^-$ |
| Linear         | Linear        | V-bent           |
52. (c)  
 $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$   
 $[\text{Cu}(\text{H}_2\text{O})_4]\text{SO}_4 \cdot \text{H}_2\text{O}$   
Covalent bond in  $\text{H}_2\text{O}$  & in  $\text{SO}_4$   
Ionic bond between  $[\text{Cu}(\text{H}_2\text{O})_4]^{2+}$  &  $\text{SO}_4^{2-}$ ,  
Coordinate bond between  $\text{Cu}^{2+}$  &  $\text{H}_2\text{O}$ .
53. (c)
- |               |               |                |                |
|---------------|---------------|----------------|----------------|
| $\text{IF}_7$ | $\text{SF}_6$ | $\text{PCl}_3$ | $\text{PCl}_5$ |
| Non polar     | Non polar     | Polar          | Non polar      |
54. (d)
- |                      |              |               |               |
|----------------------|--------------|---------------|---------------|
| $\text{H}_2\text{O}$ | $\text{O}_3$ | $\text{SO}_2$ | $\text{CO}_2$ |
| Non polar            | Non polar    | Polar         | Non polar     |
55. (d)
- |                 |               |               |                 |
|-----------------|---------------|---------------|-----------------|
| $\text{BeCl}_2$ | $\text{CS}_2$ | $\text{CO}_2$ | $\text{SO}_3$   |
| Linear          | Linear        | Linear        | Trigonal planer |

56. (b)  
Polarisation  $\uparrow$  m.p.  $\downarrow$   
Polarisation  $\propto$  size of anion

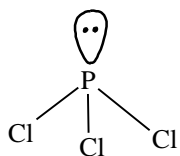
57. (d)  
Polarisation  $\propto$  charge  
 $\propto$  size of anion  
 $\propto \frac{1}{\text{size of cation}}$

58. (d)

$\text{O}_2^-$	$\text{O}_2$	$\text{O}_2^{2-}$
$2 + 2 + 3$	$2 + 2 + 2$	$2 + 2 + 4$
$= 7$	$= 6$	$= 8$



60. (c)  
 $\text{PCl}_3$   
 $3\sigma \text{ bond}$   
 $1 \text{ lone pair}$   $\left. \vphantom{\begin{matrix} 3\sigma \text{ bond} \\ 1 \text{ lone pair} \end{matrix}} \right] sp^3$



Trigonal pyramidal

61. (b)  
Charge  $\uparrow$  and size  $\downarrow$   $\text{AlCl}_3$

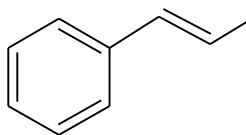
62. (c)  
S character  $\uparrow$  electronegativity  $\uparrow$   
 $sp > sp^2 > sp^3 > sp^3d > \dots$

63. (a)

	$\text{N}_2$	$\text{O}_2$	$\text{F}_2$	$\text{Cl}_2$
B.O.	3	2	1	1

$b.o \uparrow$   $b.l \downarrow$   
Smallest  $b.l$   $\text{N}_2$

64. (c)



4 $\pi$  bond 9 $\sigma$  bond

65. (c)



2 $\sigma$  bond

3 $lp$

$sp^3d$

G: trigonal bipyramidal

S: linear



2 $\sigma$  bond

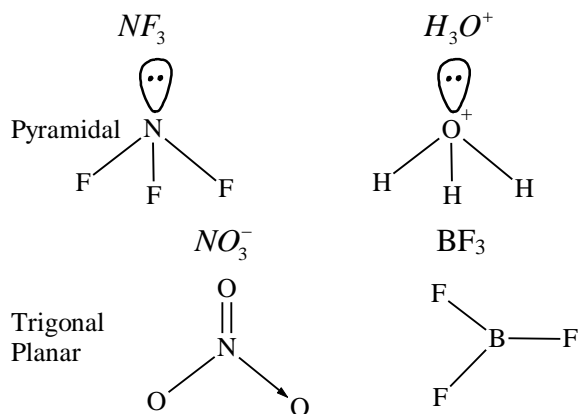
0 lone pair

$sp$

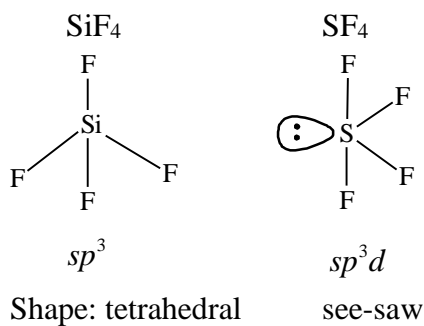
linear

linear

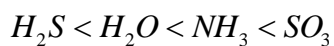
66. (c)



67. (d)



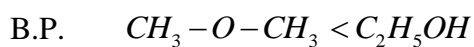
68. (c)



90 $^\circ$

120 $^\circ$

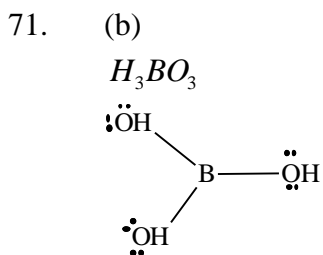
69. (d)



Volatility  $CH_3-O-CH_3 > C_2H_5OH$

Due to hydrogen bonding in  $C_2H_5OH$

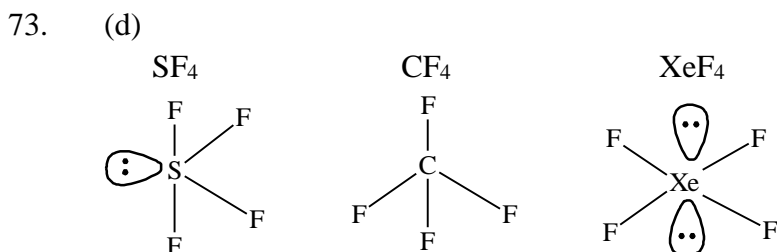
70. (b)
- |      |     |                 |
|------|-----|-----------------|
|      | NO  | NO <sup>+</sup> |
| b.o. | 2.5 | 3               |
| b.o. | ↑   | b.l. ↓          |
- b.l.*: NO > NO<sup>+</sup>



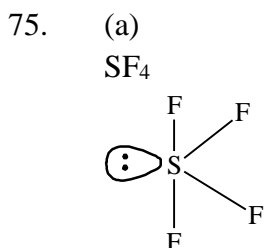
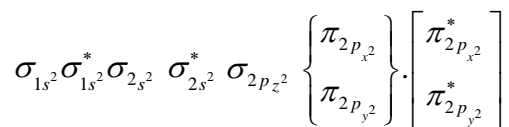
For B : 3  $\sigma$  bond ; 0 l.p ( $sp^2$ )

For O : 2  $\sigma$  ; 2 l.p ( $sp^3$ )

72. (b)
- m.p. NaCl > KCl > RbCl > LiCl



74. (a)
- O<sub>2</sub><sup>2-</sup> Diamagnetic



*B.l.*(S-F)<sub>equ</sub> ≠ *B.l.*(S-F)<sub>axial</sub>

76. (b)
- O<sub>2</sub><sup>2-</sup>
- Diamagnetic

77. (d)  
 $K^+ Ca^{2+} Mg^{2+} Be^{2+}$

Size of cation ↓

And

Charge ↑

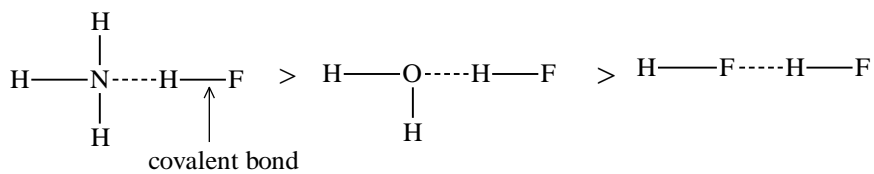
Polarisation ↑

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

78. (c)

$O_2 \rightarrow O_2^+$   
b.o.  $\frac{6-2}{2} = 2$        $\frac{6-1}{2} = 2.5$   
para                      Para

79. (c)



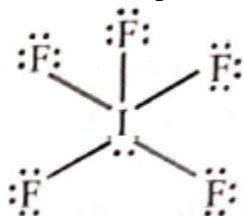
## EXERCISE - 1 [C]

1. (1)  
9 $\sigma$  and 9 $\pi$  bond are present.
2. (6)  
$${}^4C_2 = \frac{4!}{2!2!} = 6$$
3. (3)  
SF<sub>4</sub>, BF<sub>2</sub>Cl, PF<sub>3</sub>Cl<sub>2</sub>
4. (5)  
I<sub>3</sub><sup>+</sup>, XeF<sub>4</sub>, H<sub>2</sub>O, NH<sub>2</sub><sup>-</sup>, H<sub>2</sub>S
5. (6)  
CH<sub>4</sub>, CCl<sub>4</sub>, CHCl<sub>3</sub>, XeF<sub>6</sub>, POCl<sub>3</sub>, XeF<sub>4</sub>
6. (6)  
Excited state  

SO <sub>3</sub> → 2	CH <sub>4</sub> → 1
OF <sub>2</sub> → 0	SF <sub>2</sub> → 0
XeF <sub>4</sub> → 2	POCl <sub>3</sub> → 1
7. (8)  
[XeF<sub>4</sub>]<sup>2+</sup>, [ClF<sub>4</sub>]<sup>+</sup>, [SF<sub>5</sub>]<sup>+</sup>, [IF<sub>2</sub>]<sup>-</sup>, [SiF<sub>5</sub>]<sup>-</sup>  
XeF<sub>2</sub>, [SF<sub>2</sub>]<sup>2-</sup>, [PF<sub>2</sub>]<sup>3-</sup> have steric number = 5
8. (5)  
H<sub>3</sub>PO<sub>4</sub>, H<sub>2</sub>SO<sub>4</sub>, H<sub>3</sub>PO<sub>3</sub>, H<sub>3</sub>PO<sub>2</sub>, HClO<sub>4</sub>
9. (4)  
BF<sub>3</sub>, CCl<sub>4</sub>, SF<sub>6</sub>, SO<sub>3</sub>
10. (3)  
H<sub>2</sub>O, XeF<sub>5</sub><sup>-</sup>, ClF<sub>3</sub>
11. (4)  
(a) BF<sub>3</sub>; (b) BCl<sub>3</sub>; (c) CO<sub>3</sub><sup>2-</sup>; (d) SO<sub>3</sub>
12. (6)  
Cl<sub>2</sub>O<sub>6</sub>, Cl<sub>2</sub>O<sub>7</sub>, Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>, S<sub>2</sub>O<sub>7</sub><sup>2-</sup>, SO<sub>2</sub>Cl<sub>2</sub>, SOCl<sub>2</sub>

13. (4)  
CCl<sub>4</sub>, SO<sub>3</sub>, SF<sub>6</sub>, PCl<sub>5</sub>
14. (4)  
BF<sub>3</sub>, SF<sub>6</sub>, B<sub>3</sub>N<sub>3</sub>H<sub>6</sub>, PCl<sub>3</sub>F<sub>2</sub>
15. (9)  
I<sub>3</sub><sup>+</sup>, XeF<sub>4</sub>, C<sub>2</sub>F<sub>4</sub>, BrF<sub>4</sub><sup>-</sup>, SO<sub>3</sub>, NOCl, ClF<sub>3</sub>, F<sub>2</sub>CO, XeF<sub>5</sub><sup>-</sup>
16. (6)  
O<sub>2</sub><sup>2+</sup>, O<sub>2</sub><sup>2-</sup>, B<sub>2</sub><sup>2-</sup>, C<sub>2</sub><sup>2-</sup>, C<sub>2</sub>, N<sub>2</sub>
17. (3)  
KO<sub>2</sub>, NO<sub>2</sub>, NO
18. (3)  
B.L. : N<sub>2</sub><sup>-</sup> > N<sub>2</sub>; NO > NO<sup>+</sup>; H<sub>2</sub><sup>+</sup> < H<sub>2</sub><sup>-</sup>
19. (2)  
All axial bonds.
20. (1)

1. (b)  
The geometry of  $\text{IF}_5$  is square pyramid with an unsymmetric charge distribution, therefore this molecule is polar.



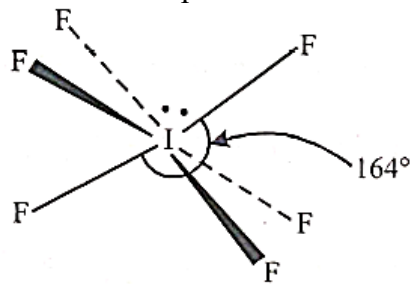
2. (d)  
Given  $e = 1.60 \times 10^{-19} \text{ C}$   
 $d = 9.17 \times 10^{-11} \text{ m}$

From  $\mu = e \times d$

$$\mu = 1.60 \times 10^{-19} \times 9.17 \times 10^{-11} = 14.672 \times 10^{-30}$$

$$\begin{aligned} \% \text{ ionic character} &= \frac{\text{Observed dipole moment}}{\text{Dipole moment for 100\% ionic bond}} \\ &= \frac{6.104 \times 10^{-30}}{14.672 \times 10^{-30}} \times 100 = 41.5\% \end{aligned}$$

3. (a)  
The structure of  $\text{IF}_6^-$  is distorted octahedral.  
This is due to presence of a “weak” lone pair.




4. (b)  
According to Fajans rules, small and highly charged cation has greatest covalent character while large cation with small charge has greatest ionic character.

5. (d)  
 $\text{H} \overset{\sigma}{-} \text{C} \begin{matrix} \pi, \sigma \\ \equiv \\ \pi \end{matrix} \text{N}$

Therefore, HCN has  $2\pi$  and  $2\sigma$  bonds.

6. (b)  
Hydrogen bond is a type of strong electrostatic dipole-dipole interaction and dependent on the inverse cube of distance between the molecules.



7. (b)  
 Dipole moment ( $\mu$ ) =  $q \times d$   
 $\Rightarrow 1\text{D} \approx 10^{-18} \text{ esu cm}$   
 $0.38 \times 10^{-18} \text{ esu cm} = q \times (1.617 \times 10^{-8} \text{ cm})$   
 $q = 2.35 \times 10^{-11} \text{ esu}$   
 So, fractional charge =  $\frac{\text{Partial charge}}{\text{Total charge}} = \frac{q}{Q}$   
 $= \frac{2.35 \times 10^{-11} \text{ esu}}{4.802 \times 10^{-10} \text{ esu}} = 0.049 \approx 0.05$
8. (c)  
 Assertion is correct but reason is incorrect. Bonding MO shows constructive interference of the combining electron waves.
9. (c)  
 Hybridisation ( $H$ ) = [No. of valence electrons of central atom + No. of monovalent atoms attached to it + (-ve charge if any) - (+ve charge in any)]  
 $\text{NO}_2^+$   $sp$  hybridization  
 $\text{NO}_2^-$   $sp^2$  hybridization  
 $\text{NO}_3^-$   $sp^2$  hybridization  
 The Lewis structures of  $\text{NO}_2$  shows a bent molecular geometry with trigonal planar electron pair geometry, hence the hybridization will be  $sp^2$ .
10. (c)  
  
 $\therefore$  Total number of lone pair of electrons is 9.
11. (c)  
 $\text{KCl}$  is an ionic compound while others ( $\text{PH}_3$ ,  $\text{O}_2$ ,  $\text{B}_2\text{H}_6$ , and  $\text{H}_2\text{SO}_4$ ) are covalent compounds.
12. (c)  
 Electronic configuration of  $\text{O}_2$  is  
 $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2$   
 $\pi^* 2p_x^1 = \pi^* 2p_y^1$   
 When an electron is added in  $\text{O}_2$  to form  $\text{O}_2^-$ , the incoming electron goes to  $\pi^* 2p_x$  or  $\pi^* 2p_y$  orbital.

13. (a)

Species	Hybridisation
$\text{ICl}_2^-$	$sp^3d$
$\text{ICl}_4^-$	$sp^3d^2$
$\text{BrF}_2^-$	$sp^3d$
$\text{IF}_6^-$	$sp^3d^3$

14. (b)

The molecules with no unpaired electrons are diamagnetic.

Molecule	Bond order
NO	1
CO	Zero
$\text{O}_2$	2
$\text{B}_2$	2

Since CO has no unpaired electron. Hence CO is diamagnetic.

15. (d)

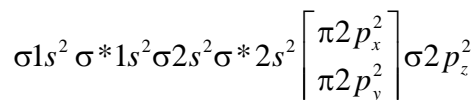
$\mu_{\text{CCl}_4} = \mu_{\text{CH}_4} = 0$  due to symmetrical structure but  $\mu_{\text{CHCl}_3} \neq 0$ .

So dipole moment order is :  $\text{CHCl}_3 > \text{CH}_4 = \text{CCl}_4$

16. (b)

Total number of electron in  $\text{CN}^- = 6 + 7 + 1 = 14$

$\therefore$  Molecular orbital distribution



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

$\text{CN}^-$  is diamagnetic because all electrons are paired.

17. (d)

(I) Ion-ion interaction energy  $\propto \left(\frac{1}{r}\right)$ .

(II) Dipole-dipole interaction energy  $\propto \left(\frac{1}{r^3}\right)$ .

(III) London dispersion  $\propto \left(\frac{1}{r^6}\right)$ .

18. (d)

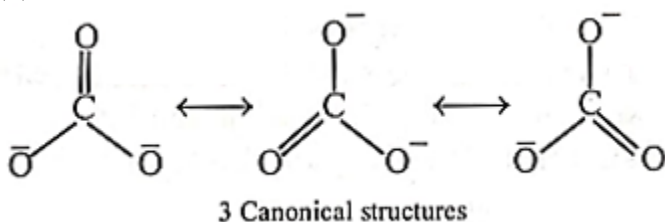
Among given intermolecular forces, ionic interaction are strongest force. Thus, correct order is ion-ion > ion-dipole > dipole-dipole

19. (c)

For London dispersion forces,

$$E \propto r^{-6}; \text{So, } x = -6$$

20. (d)



21. (b)

The chemical species having capability to donate electron pair is called Lewis base.  $\text{PCl}_5$  cannot function as a Lewis base as the central atom P does not have lone pair of electrons.

22. (c)

Species	Hybridisation	Bond length
$\text{XeF}_4$	$sp^3d^2$ (sq. planar)	All bond lengths equal
$\text{BF}_4^-$	$sp^3$ (Tetrahedral)	All bond lengths equal
$\text{SF}_4$	$sp^3d$ (See-saw)	Axial B.L. > Equatorial B.L.
$\text{SiF}_4$	$sp^3$ (Tetrahedral)	All bond lengths equal

23. (a)

According to the electronic configuration of  $\text{O}_2^-$  ion is

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

$$\text{Bond order} = \frac{10 - 7}{2} = 1.5$$

It also exhibits paramagnetic character due to the presence of unpaired electron.

24. (d)

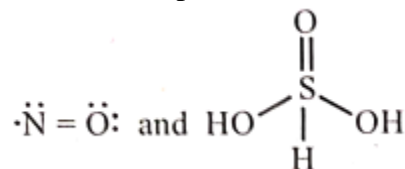
(A)-(III); (B)-(IV); (C)-(I); (D)-(II)

Molecule	Bond order
$\text{Ne}_2$	0
$\text{N}_2$	3
$\text{F}_2$	1
$\text{O}_2$	2

25. (b)

$\text{NO} \rightarrow$  Odd Electron molecule

$\text{H}_2\text{SO}_4 \rightarrow$  Expanded octet of S

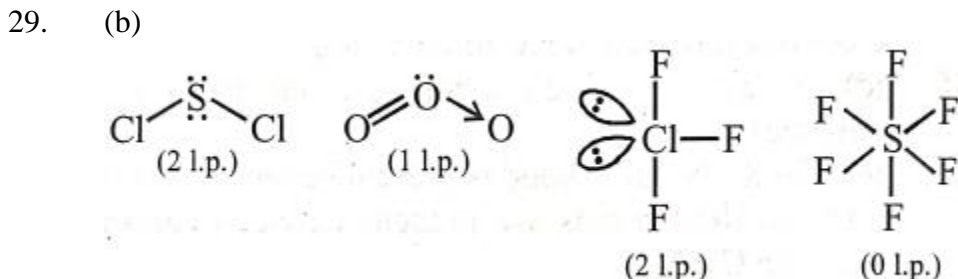


26. (b)

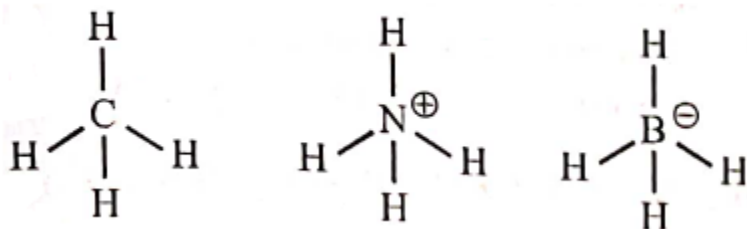
According to Fajan's rule, Covalent character  $\propto$  size of Anion. As the cation is the same in all the options. Therefore, covalent character will depend upon the size of anion. Covalent character:  $\text{A} < \text{B} < \text{C} < \text{D}$ ; Size of anion:  $\text{F}^- < \text{Cl}^- < \text{Br}^- < \text{I}^-$

27. (c)  
According to Fajan's rule, small size cation and larger size anion will have high covalent character.  
So the order is :  $\text{LiCl} > \text{NaCl} > \text{KCl} > \text{CsCl}$

28. (c)  
Electron deficient species have less than 8 electrons (or two electrons for H) in their valence (incomplete octet).  $\text{B}_2\text{H}_6$ ,  $\text{BCl}_3$  have incomplete octet.

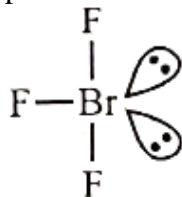


30. (b)  
Isoelectronic species have same number of electrons valence



All have steric number is 4 so  
All are tetrahedral and each have 10 electrons.

31. (c)  
As Fluorine is more electronegative it occupies the axial position and lone pairs occupy equatorial position



Steric no. = 5 ( $sp^3d$ ), lone pair = 2  
Bent T shape.

32. (c)  
The  $sp^3d$  hybridization in  $\text{PCl}_5$  is a combination of [ $sp^2 + pd$ ] mixing.  
 $sp^2$  - hybrid orbital form equatorial bonds and  $pd$  hybrid orbital forms axial bonds.  
Hence, axial bonds are longer and weaker than equatorial bonds.

33. (b)  
 $\text{PCl}_5$  forms five bonds by using the d-orbitals to "expand the octet". But  $\text{NCl}_5$  does not exist because there are no d-orbitals in the valence shell ( $2^{\text{nd}}$  shell).  
Therefore, there is no way to expand the octet.

34. (a)  
In case of tetrahedral, rectangular planar and square planar geometries, there are only bond pair of electrons by four bonds.  
Square pyramidal geometry is possible with four bond pairs and one lone pair of  $e^-$ . Hence, it will be polar due to dipole moment of these  $e^-$ .

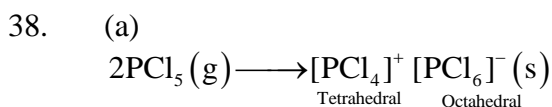
35. (c)  
The strength of H-bonding depends upon two factors:  
(1) Electronegativity difference between the hydrogen and Halogen element.  
(2) Number of H-bonding interaction.  
So, the order of H-bonding is  $CH_4 < HCN < NH_3$ .

36. (a)

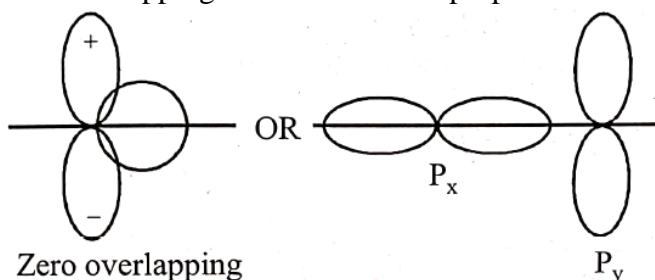
ion / molecule	Number of $e^-$ in BMO	Number of $e^-$ in ABMO	Bond order
$O_2^+$	10	5	2.5
$O_2$	10	6	2
$O_2^-$	10	7	1.5
$O_2^{2-}$	10	8	1

Bond order :  $O_2^{2-} < O_2^- < O_2 < O_2^+$

37. (c)  
Bond strength  $\propto$  Bond order  
Removal of electron from antibonding MO increases B.O.  
NO and  $O_2$  have valence  $e^-$  in  $\pi^*$  orbital.

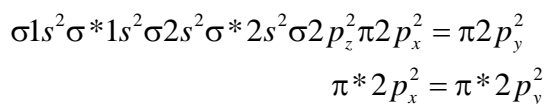


39. (a)  
Zero overlapping is occur due to improper orientation or orbitals and it is an out of phase overlapping.



40. (c)  
**Statement I:**  $O_2$ ,  $Cu^{2+}$  and  $Fe^{3+}$  all are paramagnetic species. Hence, they will be attracted by magnetic field and will be magnetized in the same direction as magnetic field.  
**Statement II:** Due to no unpaired  $e^-$  in  $Na^+$ ,  $Cl^-$ ,  $H^+$  and  $O^{2-}$ , they will not be magnetized in a magnetic field.

41. (0)  
Molecular orbital configuration of  $O_2^{2-}$  is :



Total unpaired electron in  $O_2^{2-}$  is zero.

42. (0)  
Bond order of CO = 3;  
Due to isoelectronic species, bond order of  $NO^+$  = 3  
Difference =  $0 = \frac{x}{2} \Rightarrow x = 0$

43. (15)  
Bond order =  $\frac{\text{No. of } e^- \text{ in BMO} - \text{No. of } e^- \text{ in ABMO}}{2} = 2.5$

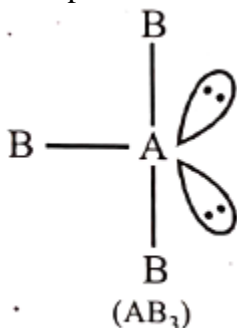
$$\Rightarrow \text{No. of } e^- \text{ in BMO} - \text{No. of } e^- \text{ in AMBO} = 5$$

As AX is diatomic molecule (neutral), the only possible case is NO.

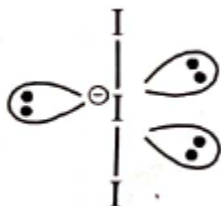
Total number of electrons = 15

**Note :** Total number of electrons equal to 13 will also have the 2.5 bond order. But in this case neutral diatomic molecule will not be possible.

44. (2)  
T-shaped molecule means central atom has 3 sigma bond and 2 lone pairs of electron.



45. (3)  
 $I_3^-$  : It is  $sp^3d$  hybridised.



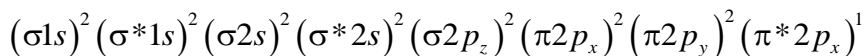
The number of lone pairs of electron on the central atom is 3.

46. (2)

Species	Structure	Lone Pair
SF <sub>4</sub>		1
BF <sub>4</sub> <sup>-</sup>		0
ClF <sub>3</sub>		2
AsF <sub>3</sub>		1
PCl <sub>5</sub>		0
BrF <sub>5</sub>		1
XeF <sub>4</sub>		2
SF <sub>6</sub>		0

47. (1)  
 S.N. = [5 B.P. + 0 L.P.)  
 PF<sub>5</sub> ⇒ sp<sup>3</sup>d hybridization

48. (4)  
 The molecular orbital configuration of B<sub>2</sub> is : (σ1s)<sup>2</sup>(σ\*1s)<sup>2</sup>(σ2s)<sup>2</sup>(σ\*2s)<sup>2</sup>(π2p<sub>x</sub>)<sup>1</sup>(π2p<sub>y</sub>)<sup>1</sup>  
 The molecular orbital configuration of C<sub>2</sub><sup>+</sup> is : (σ1s)<sup>2</sup>(σ\*1s)<sup>2</sup>(σ2s)<sup>2</sup>(σ\*2s)<sup>2</sup>(π2p<sub>x</sub>)<sup>2</sup>(π2p<sub>y</sub>)<sup>2</sup>(σ2p<sub>z</sub>)<sup>1</sup>  
 The molecular orbital configuration O<sub>2</sub><sup>+</sup> is :



The molecular orbital configuration of  $\text{He}_2^+$  is :  $(\sigma 1s)^2 (\sigma^* 1s)^1$

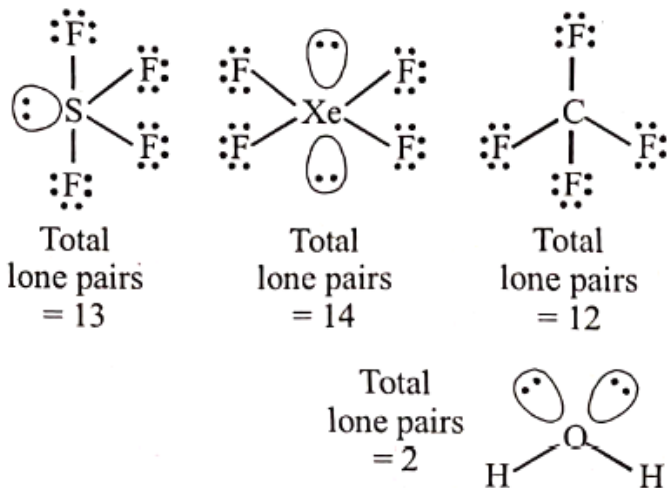
Hence,  $\text{B}_2$ ,  $\text{C}_2^-$ ,  $\text{O}_2^+$  and  $\text{He}_2^+$  are paramagnetic species.

49. (3)

$\text{CN}^-$ ,  $\text{NO}^+$  and  $\text{O}_2^{2+}$  all have bond order 3.

50. (1)

The electron pairs which do not take part in bonding and resides on atom are called lone pair.



51. (3)

Symmetrical molecules have zero dipole moment like  $\text{BF}_3$ ,  $\text{BeF}_2$ ,  $\text{SiF}_4$ ,  $\text{CCl}_4$  and  $\text{PF}_5$ .

Unsymmetrical molecules have net dipole moment like  $\text{NF}_3$ ,  $\text{CHCl}_3$  and  $\text{H}_2\text{S}$

