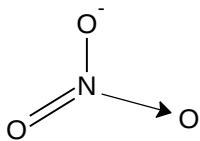
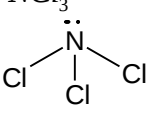


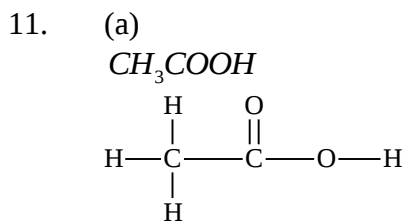
EXERCISE - 1 [A]

1. (b)
Lower the ionisation energy, cation is easily formed. Similarly higher the electron affinity, anion is easily formed.
2. (c)
 K^+I^-
3. (b)
 $H-Cl$
4. (b)

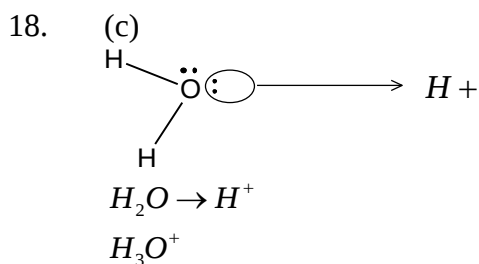
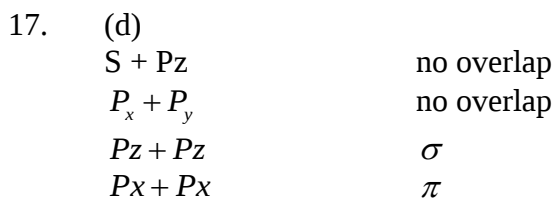
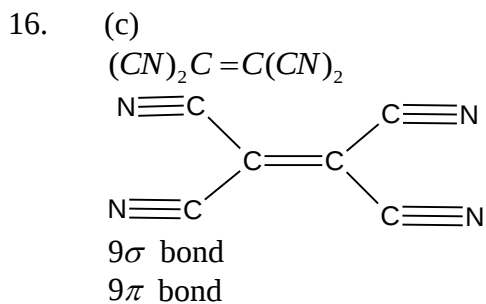
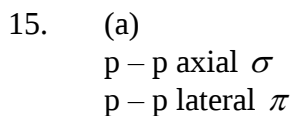
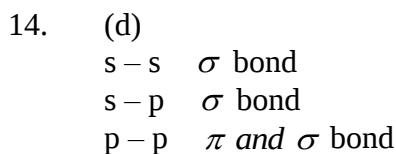
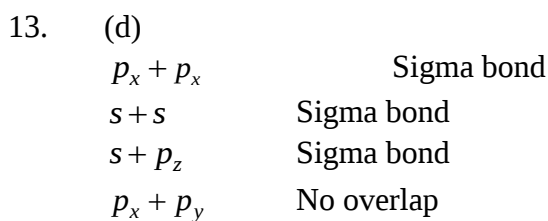
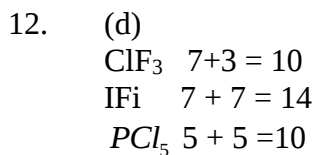
$$L.E. \propto \frac{q^+ q^-}{r^+ + r^-}$$

Charge $\uparrow$	L.E. $\uparrow$
Size $\downarrow$	L.E. $\uparrow$
5. (b)
Covalent bond (directional)
6. (d)
In ClF_3 , Cl has 10 valence electrons.
7. (d)


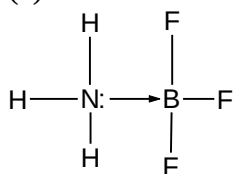
4 bond pair electrons & 0 lone pair electrons.
8. (d)
 NCl_3

9. (c)
 $AgNO_3 + CCl_4 \rightarrow \text{No reaction}$
10. (c)
 K^+CN^-
 $^-C \equiv N$



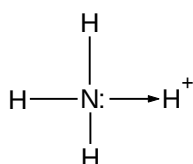
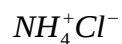
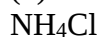
16 Shared electrons
8 unshared electrons



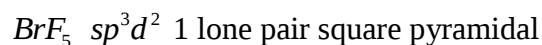
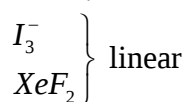
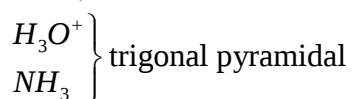
19. (c)



20. (b)



21. (d)

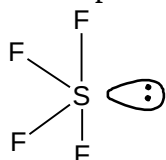


22. (b)



4 σ bond

1 lone pair



G: trigonal bipyramidal

S : See – saw

23. (d)



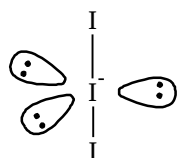
2 σ bond

3 lone pair



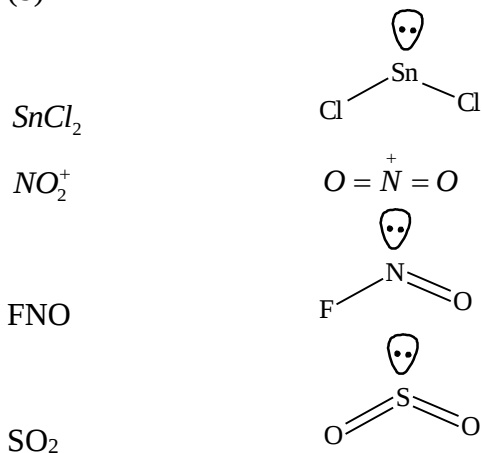
G: trigonal bipyramidal

S : linear



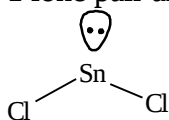
24. (a)
 PCl_6^- 6 σ bond and 0 lone pair sp^3d^2 octahedral
 SF_4 sp^3d (see saw)
 BO_3^{3-} sp^2 (trigonal planar)
 BF_4^- sp^3 (tetrahedral)

25. (b)

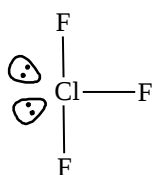


26. (c)
 2 σ and 2 π bond 1 lone pair } sp^2 trigonal planar.

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

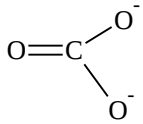


28. (c)
 ClF_3
 2 lone pair } sp^3d
 3 σ bond }



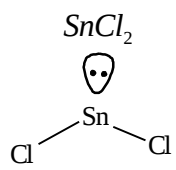
G : trigonal bipyramidal
 S : T-shape

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

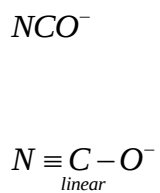
30. (b)
Size $sp < sp^2 < sp^3$
31. (a)
 CO_2
0 lone pair, 2 σ bond, sp (linear) $O = C = O$
32. (c)
 CO_3^{2-}
- 

sp^2 trigonal planar
33. (b)
s and p
sp hybridisation angle $= 180^\circ$
34. (b)
 BF_3

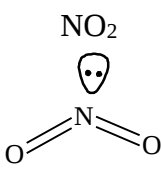
$$\left. \begin{array}{l} \text{O lone pair} \\ 3\sigma \text{ bond} \end{array} \right\} sp^2$$
 trigonal planar
35. (a)
Iso electronic $\left\{ \begin{array}{l} NO_3^- \quad sp^2 \\ CO_3^{2-} \quad sp^2 \end{array} \right\}$ trigonal planar
 SO_3 sp^2 trigonal planar
 ClO_3^- sp^3 trigonal pyramidal
36. (c)
 CO_2
 $O = C = O$ linear
- $SnCl_2$

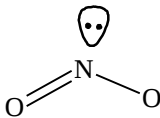
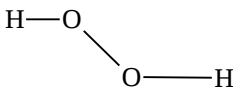
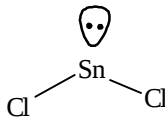


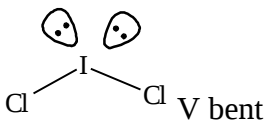
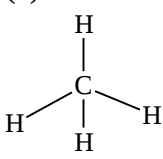
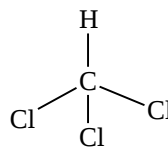
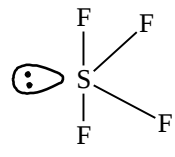
NCO^-



NO_2


37. (b)
 C_2H_2
 $H - C \equiv C - H$
0 lone pair, 2 σ bond, sp linear
38. (b)
In F_2 , P_z & P_z orbitals overlap to form σ bond.

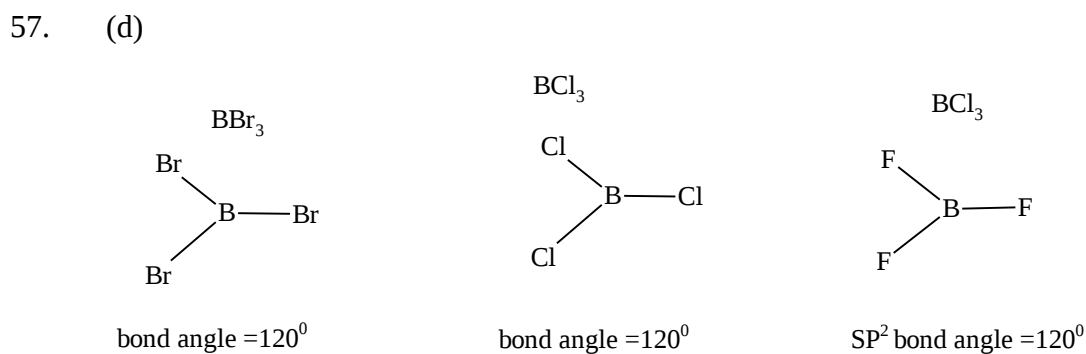
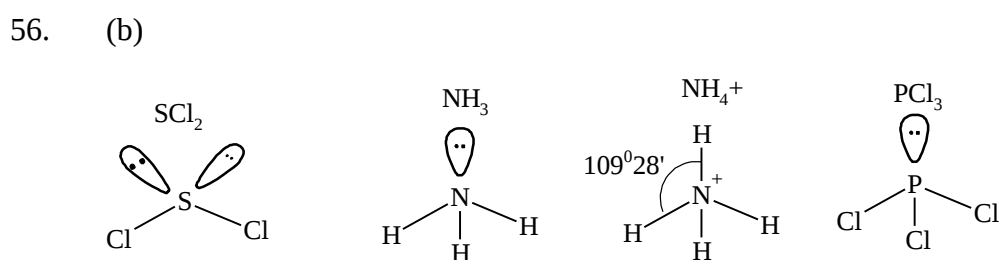
39. (b)
- $$\begin{array}{cc} sp^2 & \text{and} & sp^3d \\ \%s = \frac{1}{3} \times 100 & & \frac{1}{5} \times 100 \\ = 33.33\% & & = 20\% \end{array}$$
40. (b)
- Electronegativity $sp > sp^2 > sp^3 > sp^3d$
41. (b)
- SO_3 sp^2
- PCl_5 sp^3d
- SO_2 sp^2
- PCl_3 sp^3
42. (c)
- $$OF_2 \left. \begin{array}{l} 2\sigma \text{ bond} \\ 2 \text{ lone pair} \end{array} \right\} sp^3$$
43. (d)
- C_2H_2 linear
- $H-C \equiv C-H$ NO_2
- H_2O_2 
-   CO_2
- $O=C=O$
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)

46. (c)
 BF_4^-
 $\left. \begin{array}{l} 0 \text{ lone pair} \\ 4 \sigma \text{ bond} \end{array} \right\} sp^3$
47. (a)
 $POCl_3$
 sp^3
 PCl_3 SF_4 ClF_3 BCl_3
 sp^3 sp^3d sp^3d sp^2
48. (d)
 ICl_2^+
 $\left. \begin{array}{l} 2 \sigma \text{ bond} \\ 2 \text{ lone pair} \end{array} \right\} sp^3$

V bent
49. (c)



50. (c)
 ML_x
 sp^3d^2
G: octahedral
S: square planar
2 lone pair
4 σ bond
 $x = 4$
51. (b)
Bond length $\propto \frac{1}{\text{bond strength}}$
C – H bond length is shortest
52. (d)
Solubility $MgF_2 < CaF_2 < SeF_2 < BaF_2$

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

54. (b)
 AgI
 I^- Large anion
Greatest covalent character least soluble

55. (a)
 AgCl
Greater covalent Ag^+ (less soluble)
PIGC (pseudo inert gas centi)



58. (d)

	CO	CO_3^{2-}	CO_2
B.O.	3	4/3	2

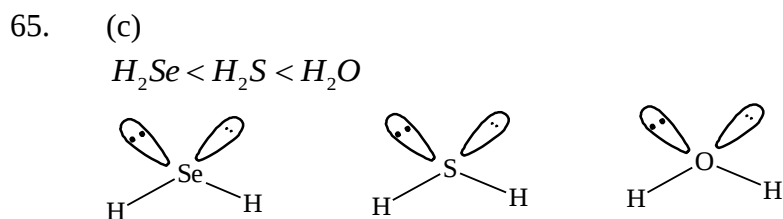
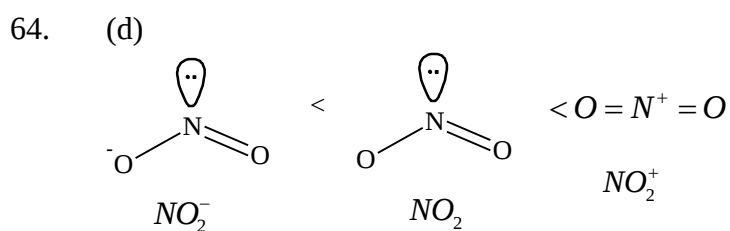
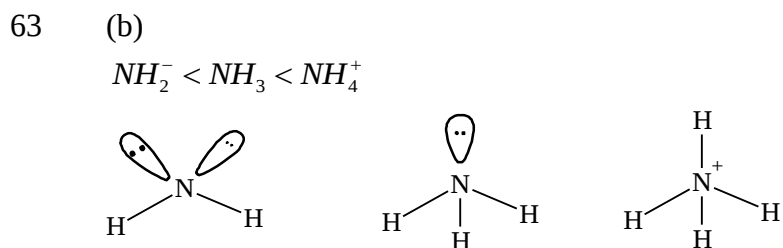
$b.l \propto \frac{1}{b.o}$
 $b.l : \text{CO}_3^{2-} > \text{CO}_2 > \text{CO}$

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 $BeCl_2 > MgCl_2 > CaCl_2 > SrCl_2 > BaCl_2$

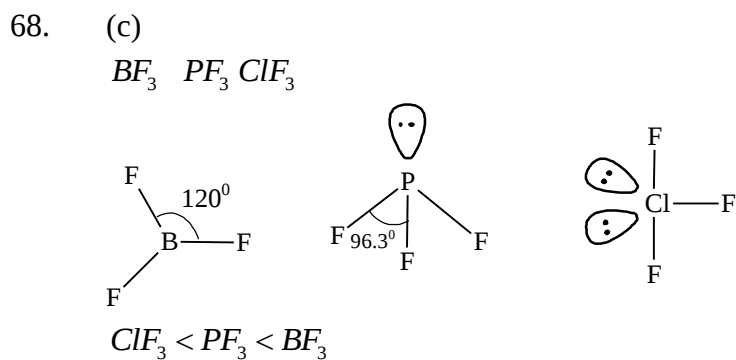
62. (c)
Charge $\uparrow$ Polarisation $\uparrow$ covalent characters $\uparrow$
 Na_3N
 $F^- < O^{2-} < N^{3-}$



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

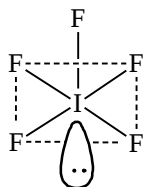
67. (b)

NH_3	CH_4	H_2O	H_2S
107°	$109^\circ 28'$	104.5°	92°



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

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



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

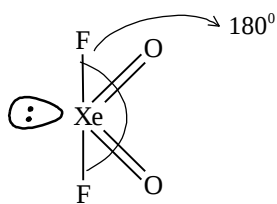
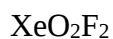
72. (d)
 I^- largest anion (greater polarisation)
Covalent $AlI_3 > AlBr_3 > AlCl_3 > AlF_3$

73. (c)
 NH_2^-
2 σ bond, 2 lone pair } sp^3
 v-bent

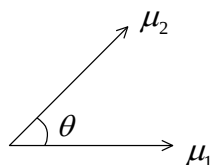
74. (c)

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

75. (c)



76. (a)

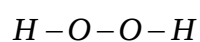


$$\mu_{net} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos \theta}$$

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

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

77. (c)



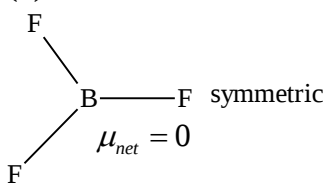
Non polar bond

78. (c)

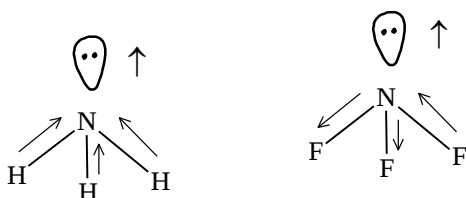
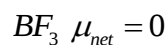


$\Delta(E.N) \uparrow$ Polar character $\uparrow$

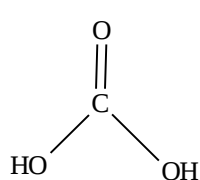
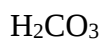
79. (b)



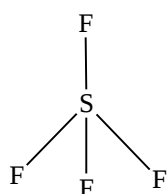
80. (d)



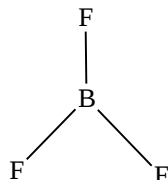
81. (a)



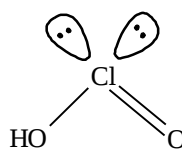
sp^2
Polar



sp^3
Non polar

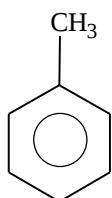


sp^2
Non polar

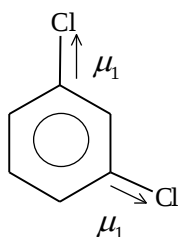


sp^3
Polar

82. (b)



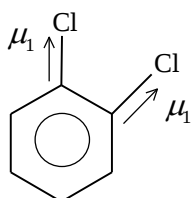
I



II

$$\theta = 120^\circ$$

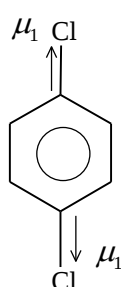
$$\mu_{\text{net}} = \mu_1$$



III

$$\theta = 60^\circ$$

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



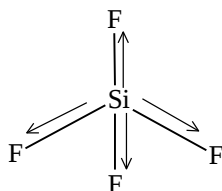
IV

$$\theta = 180^\circ$$

$$\mu_{\text{net}} = 0$$

$$IV < I < II < III$$

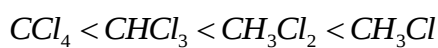
83. (c)



$$\mu_{\text{net}} = 0$$

84. (a)

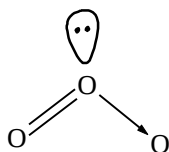
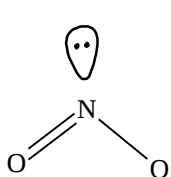
Dipole moment



85. (b)

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

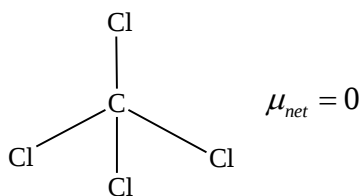
$$\mu_{\text{net}} \text{ of } \text{NO}_2 \text{ and } \text{O}_3 \neq 0$$



86. (c)

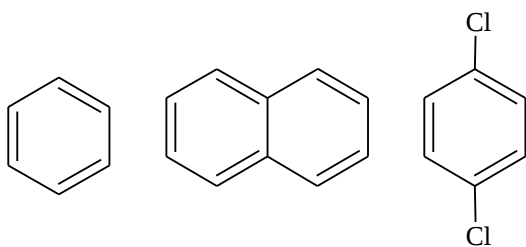
Covalent (directional)

87. (d)



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

89. (d)

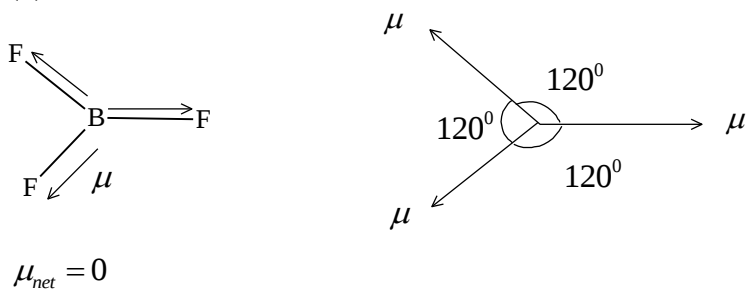


$\mu_{net} = 0$

90. (a)

Other $\mu_{net} = 0$
(CO_2, BF_3, CCl_4)

91. (d)

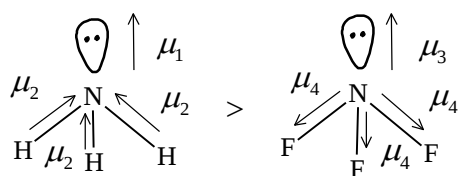


92. (c)

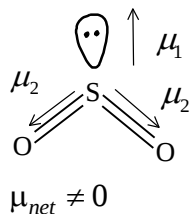
 $\mu_{net} \neq 0$
Others $\mu_{net} = 0$

93. (b)

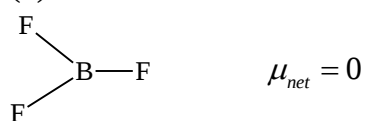
μ_{net} of BF_3 and $B_2H_6 = 0$



94. (c)



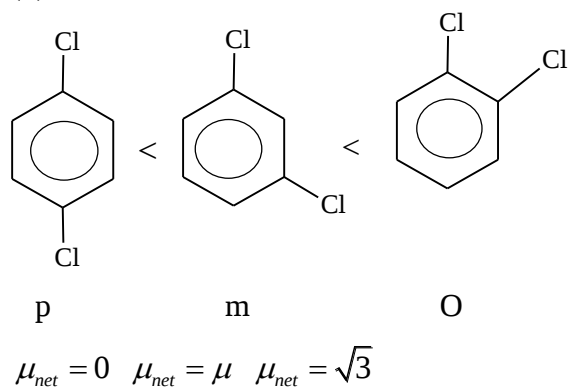
95. (b)



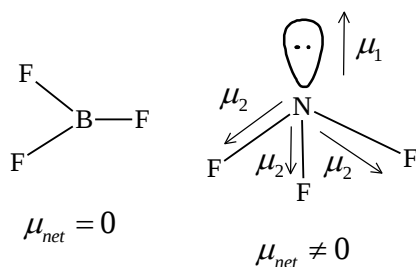
96. (a)

$\mu_{net}(NF_3) < \mu_{net}(NH_3)$
 μ_{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 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)

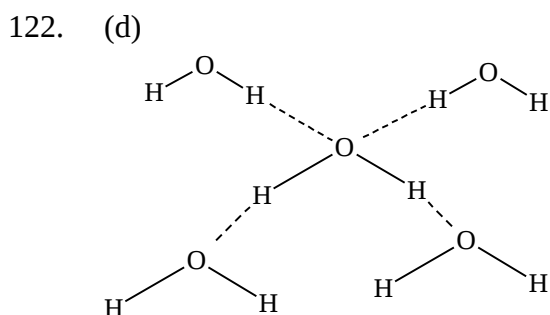
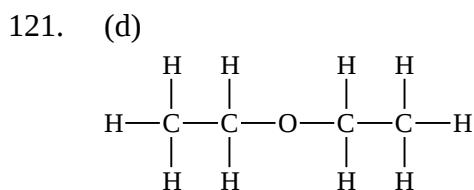
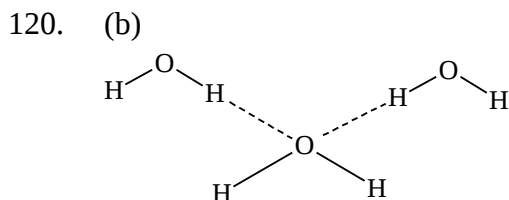
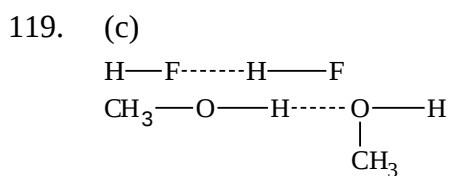
O_2	Paramagnetic
O_2^-	Paramagnetic
C_2	Diamagnetic
N_2^+	Paramagnetic

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 : $\text{HCl} > \text{HBr} > \text{HI} > \text{HF}$

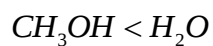
124. (a)
B.P. $\text{H}_2\text{S} < \text{H}_2\text{Se} < \text{H}_2\text{Te} < \text{H}_2\text{O}$ $\uparrow$ due to intermolecular H-bonding

125. (d)
HCl (no H – bonding)

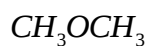
126. (a)
H-bonding Strength
 $\text{H} \cdots \cdots \text{F} > \text{H} \cdots \cdots \text{O} > \text{H} \cdots \cdots \text{N}$

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

128. (b)

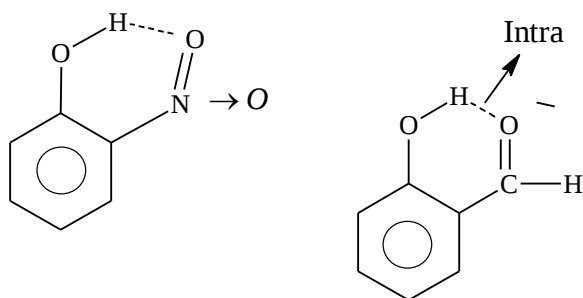


Greater strength of H-bonding in water

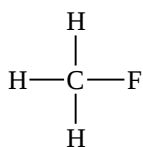


No H-bonding

129. (d)



130. (d)

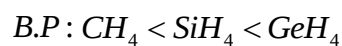


131. (d)

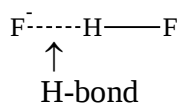
Hydrogen bond is weaker than other bonds

132. (d)

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



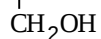
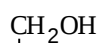
133. (c)



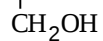
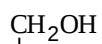
134. (a)

Inter molecular

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



Glycerol

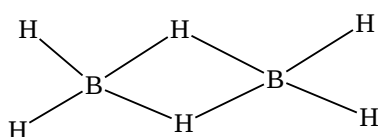


Glycol

135. (b)

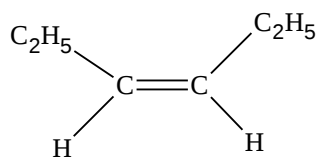
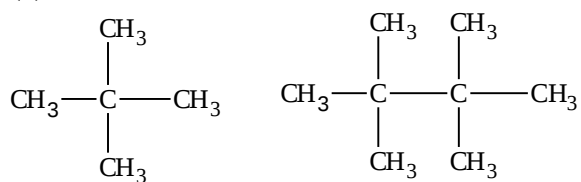
Covalent bond is directional.

136. (d)



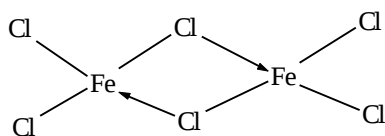
(2c 2e bond or banana bond)

137. (c)



max. dipole moment

138. (a)



139. (b)

B.P. $\text{HF} > \text{HBr} > \text{HCl}$

140. (c)

KO_2	AlO_2^-	BaO_2	NO_2^+
$\text{K}^+ \text{O}_2^-$	$\text{Al}^{3+} \text{O}_2^{2-}$	$\text{Ba}^{2+} \text{O}_2^{2-}$	
Dia Para	Dia Dia	Dia Dia	Dia

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 $\text{Ne}_2 = 0$

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)

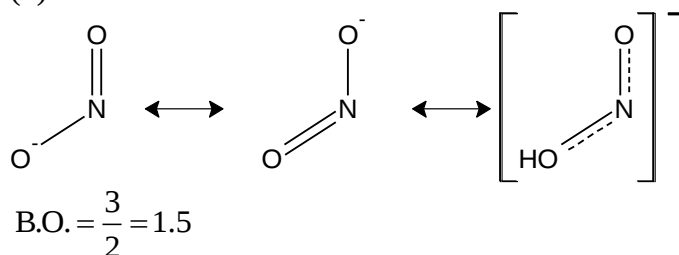
$$L.E. \propto \frac{q^+ q^-}{r^+ + r^-}$$
8. (d)
 $[Kr] 4d^{10}$ is ion of p-block elements or 11th or 12th group.
 $Sn = [Kr] 5s^2 4d^{10} 5p^2$
 $Sn^{4+} = [Kr] 5s^0 4d^{10}$
 $Ag = [Kr] 5s^1 4d^{10}$
 $Ag^+ = [Kr] 5s^0 4d^{10}$
9. (c)
 $Fe = [Ar] 4s^2 3d^6$
 $Fe^{2+} = [Ar] 4s^0 3d^6$

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 : $\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$
16. (b)
As $\Delta\text{E.N.} \uparrow$ polarity $\uparrow$
It is least for p & s .
17. (b)
 $\Delta\text{E.N.} \uparrow$ polarity $\uparrow$
It is least for O & F & maximum for Mg & F.
18. (a)
 $\Delta\text{E.N.}(\text{O}-\text{F}) = 4 - 3.5 = 0.5$
 $\Delta\text{E.N.}(\text{Cl}-\text{F}) = 4 - 3 = 1$
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)
 $\text{H} - \ddot{\text{N}} = \ddot{\text{N}} - \text{H}$
21. (d)
 $4 + 3 + 4 + 2 + 7 = 20$
22. (b)
$$S = \begin{array}{c} 3s \qquad 3p \\ \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow} \quad \boxed{\uparrow} \end{array}$$

23. (a)

	CO	CO ₂	CO ₃ ²⁻
Bond order :	3	2	1.33
Bond order ↑, Bond length ↑			

24. (d)



25. (a)

In resonance, delocalisation of electrons take place.

26. (c)

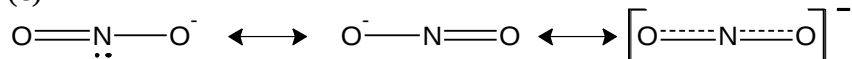
$$x + 3(-1) = 0$$

$$x = +3$$

27. (c)

In resonance, delocalisation of electrons take place.

28. (c)



29. (d)

In SO₄²⁻, electrons are delocalised on 4 bond pair.

30. (b)

Due to resonance, bonds are identical in all.

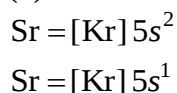
31. (c)

O can not expand its octet.

32. (c)

CO has fully filled molecular orbitals of 2p.

33. (d)



34. (d)

SeF₄
6 + 4 = 10 valence electrons in Se.

35. (c)

CF₄
4 + 4 = 8
Valence Electrons.

36. (c)
 BF_3
 $3 + 3 = 6$
 Valence Electrons.
37. (a)
 ClF_3
 $7 + 3 = 10$
 Valence Electrons.
 In SO_3 , coordinate bond can be formed.
38. (c)
 ICl_5
 $7 + 5 = 12$
 Valence Electrons.
39. (b)
 BeH_2
 $2 + 2 = 4$
 Valence Electrons.
40. (a)
- | | | | | |
|------------|---------------------|-----------------------|-----------------------|---------------------|
| | O_2^{2+} | O_2^+ | O_2^- | O_2^{2-} |
| 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 $\text{O}_2^{2-} > \text{O}_2^- > \text{O}_2^+ > \text{O}_2^{2+}$
41. (a)
 Bond enthalpy is energy required to break 1 mole of covalent bond.
42. (c)
- | | | | |
|---------------|-------------------|---------------|--------------|
| NO^+ | C_2^{2-} | CN^- | N_2 |
| 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.

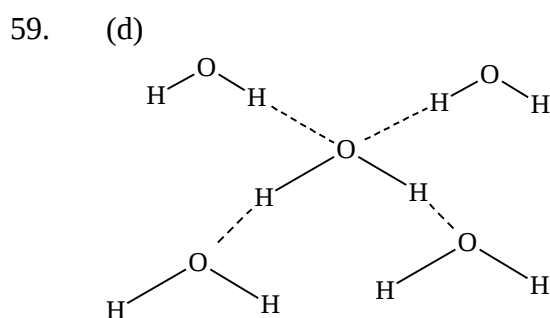
46. (a)
Solid PCl_5 exists as
 $[\text{PCl}_4]^+ [\text{PCl}_6]^-$
 $sp^3 \quad sp^3d$
47. (b)
- | | | | |
|----------------|-------------------|----------------|--------------|
| O_2^- | O_2^{2-} | O_2^+ | O_2 |
| Paramagnetic | Diamagnetic | Paramagnetic | Paramagnetic |
48. (b)
- $$\text{H}-\text{C}(=\text{O})-\text{O}-\text{H} \cdots \text{C}(\text{H})-\text{C}(=\text{O})-\text{CH}_3$$
49. (d)
 PCl_5
 $5\sigma \text{ \& } 0 \ell.p.$
 sp^3d
50. (d)
- | | | | |
|----------------|----------------|----------------|----------------|
| BBr_3 | XeF_4 | BCl_3 | XeO_3 |
| sp^2 | sp^3d^2 | sp^2 | sp^3 |
51. (c)
- | | | |
|----------------|---------------|------------------|
| I_3^- | CO_2 | 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 H_2O & in SO_4
Ionic bond between $[\text{Cu}(\text{H}_2\text{O})_4]^{2+}$ & SO_4^{2-} ,
Coordinate bond between Cu^{2+} & H_2O .
53. (c)
- | | | | |
|---------------|---------------|----------------|----------------|
| IF_7 | SF_6 | PCl_3 | PCl_5 |
| Non polar | Non polar | Polar | Non polar |
54. (d)
- | | | | |
|----------------------|--------------|---------------|---------------|
| H_2O | O_3 | SO_2 | CO_2 |
| Non polar | Non polar | Polar | Non polar |
55. (d)
- | | | | |
|-----------------|---------------|---------------|-----------------|
| BeCl_2 | CS_2 | CO_2 | 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)

O_2^-	O_2	O_2^{2-}
$2 + 2 + 3$	$2 + 2 + 2$	$2 + 2 + 4$
$= 7$	$= 6$	$= 8$



60. (c)
 PCl_3
 $\left. \begin{array}{l} 3\sigma \text{ bond} \\ 1 \text{ lone pair} \end{array} \right\} sp^3$

Trigonal pyramidal

61. (b)
Charge $\uparrow$ and size $\downarrow$ AlCl_3

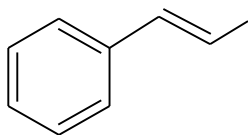
62. (c)
S character $\uparrow$ electronegativity $\uparrow$
 $sp > sp^2 > sp^3 > sp^3d > \dots$

63. (a)

	N_2	O_2	F_2	Cl_2
B.O.	3	2	1	1

$b.o \uparrow$ $b.l \downarrow$
Smallest $b.l$ N_2

64. (c)



4 π bond 9 σ bond

65. (c)



2 σ bond

3 ℓp

sp^3d

G: trigonal bipyramidal

S: linear



2 σ bond

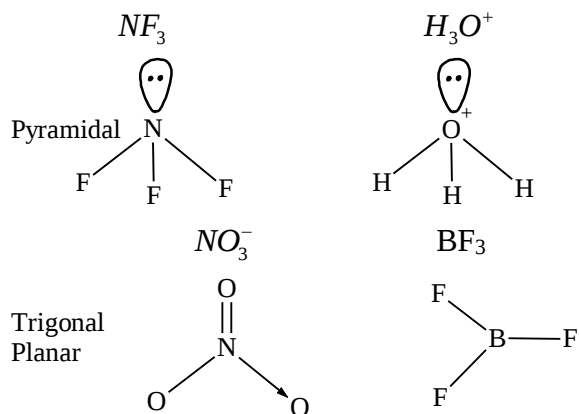
0 lone pair

sp

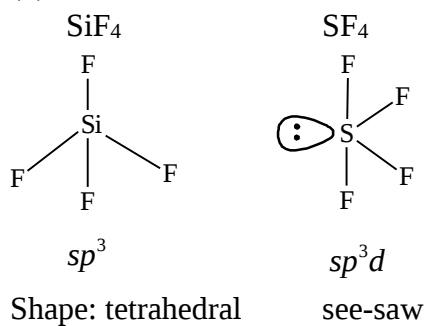
linear

linear

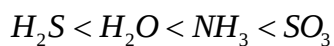
66. (c)



67. (d)



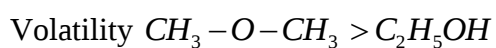
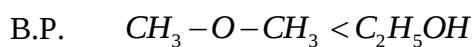
68. (c)



90°

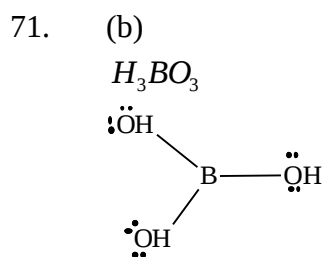
120°

69. (d)



Due to hydrogen bonding in $\text{C}_2\text{H}_5\text{OH}$

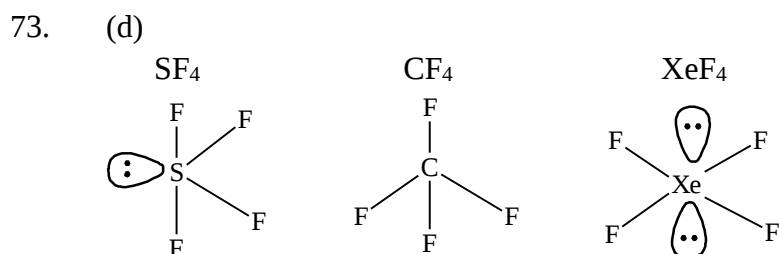
70. (b)
- | | | |
|------|-----|-----------------|
| | NO | NO ⁺ |
| b.o. | 2.5 | 3 |
| b.o. | ↑ | b.ℓ.↓ |
- b.ℓ.:* NO > NO⁺



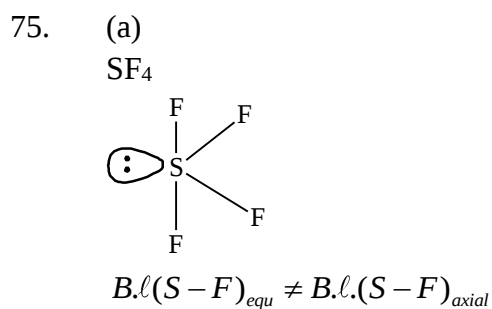
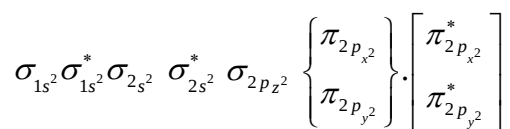
For B : 3 σ bond ; 0 l.p (sp^2)

For O: 2 σ ; 2ℓp (sp^3)

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



74. (a)
- O_2^{2-} Diamagnetic



76. (b)
- O_2^{2-}
- Diamagnetic

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

Size of cation ↓

And

Charge ↑

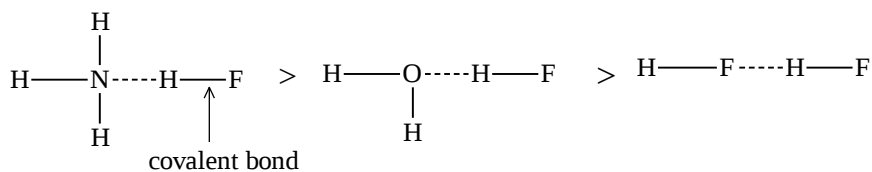
Polarisation ↑

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

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]

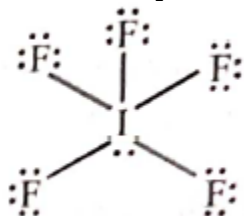
1. (1)
9σ and 9π bond are present.
2. (6)
$${}^4C_2 = \frac{4!}{2!2!} = 6$$
3. (3)
SF₄, BF₂Cl, PF₃Cl₂
4. (5)
I₃⁺, XeF₄, H₂O, NH₂⁻, H₂S
5. (6)
CH₄, CCl₄, CHCl₃, XeF₆, POCl₃, XeF₄
6. (6)
Excited state

SO ₃ → 2	CH ₄ → 1
OF ₂ → 0	SF ₂ → 0
XeF ₄ → 2	POCl ₃ → 1
7. (8)
[XeF₄]²⁺, [ClF₄]⁺, [SF₅]⁺, [IF₂]⁻, [SiF₅]⁻
XeF₂, [SF₂]²⁻, [PF₂]³⁻ have steric number = 5
8. (5)
H₃PO₄, H₂SO₄, H₃PO₃, H₃PO₂, HClO₄
9. (4)
BF₃, CCl₄, SF₆, SO₃
10. (3)
H₂O, XeF₅⁻, ClF₃
11. (4)
(a) BF₃; (b) BCl₃; (c) CO₃²⁻; (d) SO₃
12. (6)
Cl₂O₆, Cl₂O₇, Cr₂O₇²⁻, S₂O₇²⁻, SO₂Cl₂, SOCl₂

13. (4)
 $\text{CCl}_4, \text{SO}_3, \text{SF}_6, \text{PCl}_5$
14. (4)
 $\text{BF}_3, \text{SF}_6, \text{B}_3\text{N}_3\text{H}_6, \text{PCl}_3\text{F}_2$
15. (9)
 $\text{I}_3^+, \text{XeF}_4, \text{C}_2\text{F}_4, \text{BrF}_4^-, \text{SO}_3, \text{NOCl}, \text{ClF}_3, \text{F}_2\text{CO}, \text{XeF}_5^-$
16. (6)
 $\text{O}_2^{2+}, \text{O}_2^{2-}, \text{B}_2^{2-}, \text{C}_2^{2-}, \text{C}_2, \text{N}_2$
17. (3)
 $\text{KO}_2, \text{NO}_2, \text{NO}$
18. (3)
B.L. : $\text{N}_2^- > \text{N}_2; \text{NO} > \text{NO}^+; \text{H}_2^+ < \text{H}_2^-$
19. (2)
All axial bonds.
20. (1)

SECTION-I

1. (b)
The geometry of IF_5 is square pyramid with an unsymmetric charge distribution, therefore this molecule is polar.



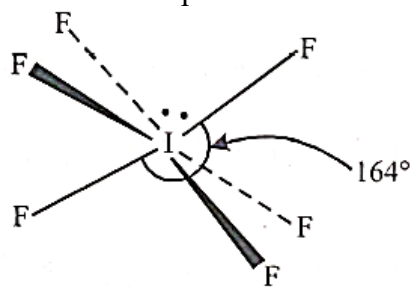
2. (d)
Given $e = 1.60 \times 10^{-19} \text{ C}$
 $d = 9.17 \times 10^{-11} \text{ m}$

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 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)
 $\text{H} \xrightarrow{\sigma} \text{C} \begin{matrix} \xrightarrow{\pi, \sigma} \\ \xleftarrow{\pi} \end{matrix} \text{N}$

Therefore, HCN has 2π and 2σ 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 (μ) = $q \times d$
 $\Rightarrow 1\text{D} \approx 10^{-18} \text{ esu cm}$
 $0.38 \times 10^{-18} \text{ esu cm} = q \times (1.617 \times 10^{-8} \text{ cm})$
 $q = 2.35 \times 10^{-11} \text{ esu}$
 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 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)
KCl is an ionic compound while others (PH_3 , O_2 , B_2H_6 , and H_2SO_4) are covalent compounds.
12. (c)
Electronic configuration of 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 O_2 to form O_2^- , the incoming electron goes to $\pi^* 2p_x$ or $\pi^* 2p_y$ orbital.

13. (a)

Species	Hybridisation
ICl_2^-	sp^3d
ICl_4^-	sp^3d^2
BrF_2^-	sp^3d
IF_6^-	sp^3d^3

14. (b)

The molecules with no unpaired electrons are diamagnetic.

Molecule	Bond order
NO	1
CO	Zero
O_2	2
B_2	2

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

15. (d)

$\mu_{CCl_4} = \mu_{CH_4} = 0$ due to symmetrical structure but $\mu_{CHCl_3} \neq 0$.

So dipole moment order is : $CHCl_3 > CH_4 = CCl_4$

16. (b)

Total number of electron in $CN^- = 6 + 7 + 1 = 14$

$\therefore$ Molecular orbital distribution

$$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \left[\begin{matrix} \pi 2p_x^2 \\ \pi 2p_y^2 \end{matrix} \right] \sigma 2p_z^2$$

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

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 $\propto \left(\frac{1}{r^3}\right)$.

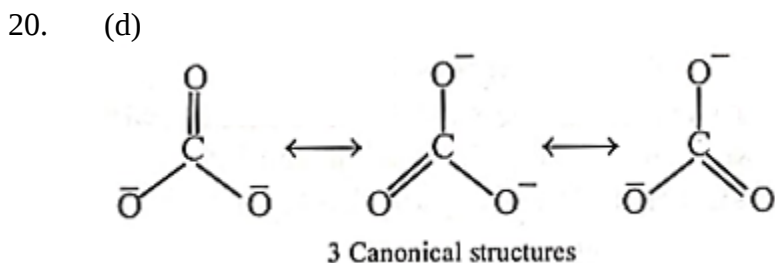
(III) London dispersion $\propto \left(\frac{1}{r^6}\right)$.

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



21. (b)
The chemical species having capability to donate electron pair is called lewis base. PCl_5 cannot function as a Lewis base as the central atom P does not have lone pair of electrons.

22. (c)

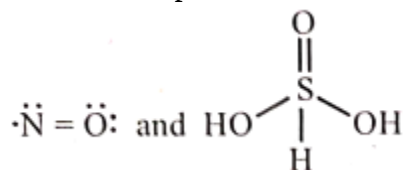
Species	Hybridisation	Bond length
XeF_4	sp^3d^2 (sq. planar)	All bond lengths equal
BF_4^-	sp^3 (Tetrahedral)	All bond lengths equal
SF_4	sp^3d (See-saw)	Axial B.L. > Equatorial B.L.
SiF_4	sp^3 (Tetrahedral)	All bond lengths equal

23. (a)
According to the electronic configuration of O_2^- ion is
 $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^1 = \pi^* 2p_y^1$
 Bond order $= \frac{10 - 7}{2} = 1.5$
 It also exhibit paramagnetic character due to the presence of unpaired electron.

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

Molecule	Bond order
Ne_2	0
N_2	3
F_2	1
O_2	2

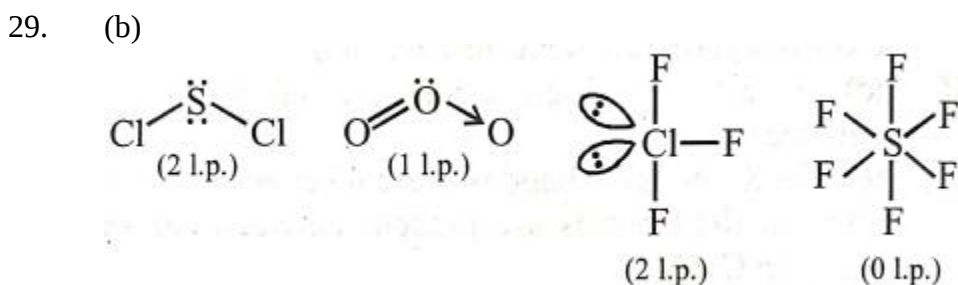
25. (b)
 $\text{NO} \rightarrow$ Odd Electron molecule
 $\text{H}_2\text{SO}_4 \rightarrow$ Expanded octet of S



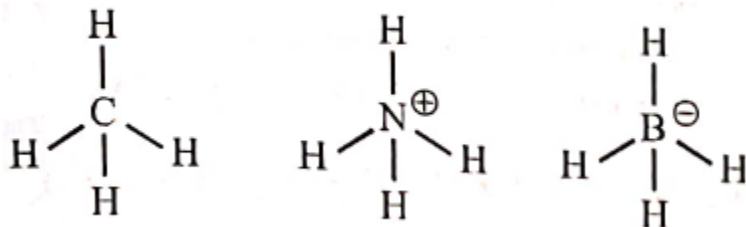
26. (b)
According to Fajan's rule, Covalent character $\propto$ size of Anion As the cation is same in all the option. Therefore covalent character will depend upon the size of anion. Covalent character : $A < B < C < D$;
Size of anion : $F^- < Cl^- < Br^- < I^-$

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

28. (c)
Electron deficient species have less than 8 electrons (or two electrons for H) in their valence (incomplete octet). B_2H_6 , BCl_3 have incomplete octet.

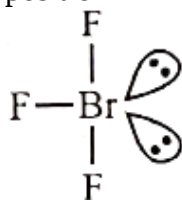


30. (b)
Isoelectronic species have same number of electrons valence



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 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)
 PCl_5 forms five bonds by using the d-orbitals to “expand the octet”. But NCl_5 does not exist because there are no d-orbitals in the valence shell (2^{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 $\text{CH}_4 < \text{HCN} < \text{NH}_3$.

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 |

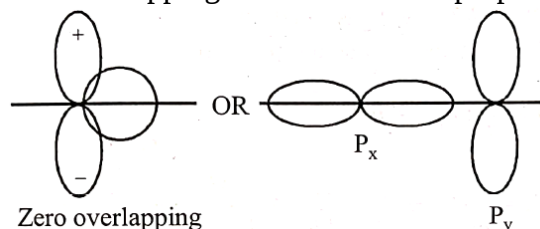
Bond order: $\text{O}_2^{2-} < \text{O}_2^- < \text{O}_2 < \text{O}_2^+$

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

38. (a)

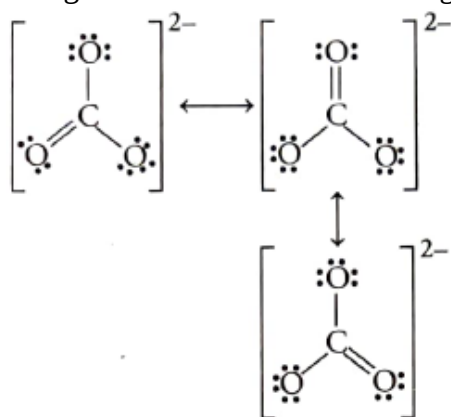
$$2\text{PCl}_5(\text{g}) \longrightarrow \underset{\text{Tetrahedral}}{[\text{PCl}_4]^+} \underset{\text{Octahedral}}{[\text{PCl}_6]^-}(\text{s})$$

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.

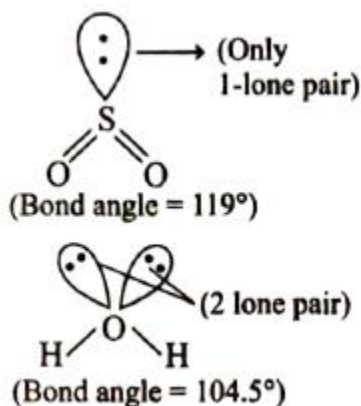
41. (d)
 CO_3^{2-} has a single structure i.e., resonance hybrid of CO_3^{2-} gives three canonical structures.
 The given structures are resonating structures of carbonate ion.



42. (d)
 Greater the difference of electronegativity between the cation and anion, greater will be the ionic character and lesser will be the covalent character.
 Hence, correct order is $\text{KI} > \text{KF}$; $\text{LiF} > \text{KF}$; $\text{CuCl} > \text{NaCl}$
 Smaller the size of the cation, higher will be its polarising power. Since, Sn^{4+} is smaller in size than Sn^{2+} hence it will have more covalent character in its compounds. So, $\text{SnCl}_4 > \text{SnCl}_2$.

43. (c)
 Assertion (A) is correct but Reason (R) is incorrect.
 Dipole moment is a vector quantity i.e., it has magnitude as well as definite directions.
 In chemistry, the dipole moment is represented by cross arrow on the positive center and arrowhead on the negative center. This arrow symbolises the shift of electron density in the molecules.

44. (b)
 Shape of SO_2 and H_2O are given as



Hence, both has V-shape, but bond angle in SO_2 is greater than H_2O , due to repulsion between lone pairs at oxygen.
 So, statement I is correct but statement II is incorrect.

45. (a)
 The correct match is A-II, B-I, C-IV, D-III.

46. (d)

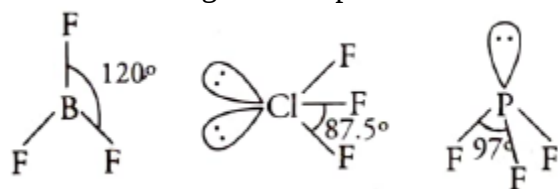
The bond order of acetylide ion C_2^{2+} ($^-\text{C} \equiv \text{C}^+$) is 3 and is diamagnetic in nature.

- Bond order of $\text{O}_2^- = \frac{10-7}{2} = 1.5$ and paramagnetic
- Bond order of $\text{N}_2^- = \frac{10-5}{2} = 2.5$ and paramagnetic
- Bond order of $\text{NO}^+ = \frac{10-4}{2} = 3$ and paramagnetic
- Bond order of $\text{O}_2^+ = \frac{10-5}{2} = 2.5$ and paramagnetic

Thus, bond order and magnetic property of acetylide ion are same as that of NO^+ .

47. (b)

The structure of given compounds are as follows.



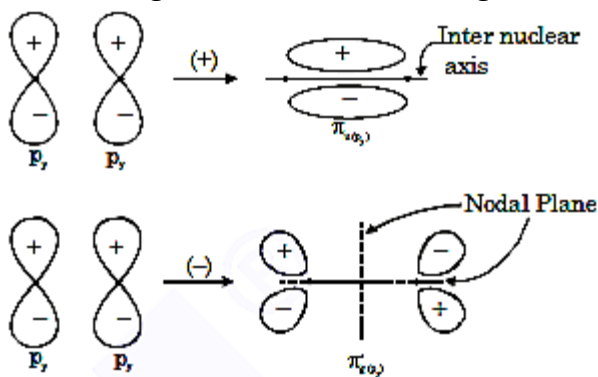
So, the correct order of bond angle is $\text{ClF}_3 < \text{PF}_3 < \text{BF}_3$.

48. (b)

Among the given molecules CH_4 and PF_5 are non-polar. Hence they have net zero dipole momentum while among NH_3 and NF_3 , the dipole moment of NH_3 is greater than NF_3 . It is because in NH_3 vector addition of bond momentum and lone pair takes place while in NF_3 vector subtraction of bond moment and lone pair takes place.

49. (c)

A π bonding molecular orbital has higher electron density above and below inter nuclear axis



50. (a)

The ionic character in a molecule depend upon electronegativity difference between the constituting atoms.

Greater the electronegativity difference, greater will be the ionic character of molecules.

Therefore, the correct sequence of increasing order of ionic character.

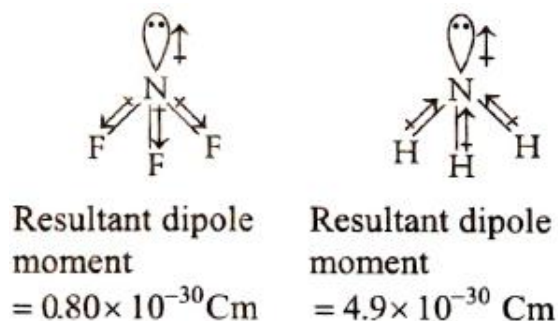


Hence, option (a) is correct.

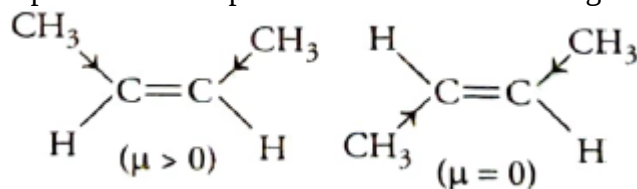
51. (c)
According to Fajan's rule, Ag and Cl bond will be covalent in nature not ionic.
The ionic character of the given compounds is in the order,
 $\text{AgCl} < \text{CoCl}_2 < \text{BaCl}_2 < \text{KCl}$
Hence, AgCl is least ionic.

52. (d)
Polar molecules are those molecules in which there exists an electronegativity difference between the bonded atom. They have non-zero dipole moment. Hence, CHCl_3 is a polar molecule. Non-polar molecules are those molecules in which electrons are shared equally. They have zero dipole moment, $\text{CH}_2 = \text{CH}_2$, CCl_4 , CO_2 are non-polar molecules.
Hence, option (d) is correct.

53. (a)
Both A and R are correct and R is the correction explanation of A.
The bond structure of NF_3 and NH_3 with dipole moment direction are;

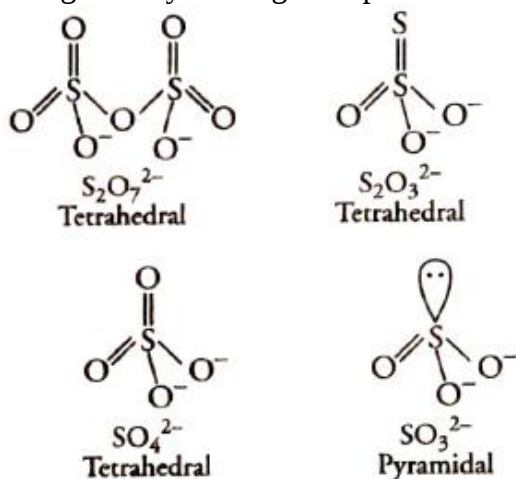


54. (a)
Both A and R are correct and R is the correct explanation of A.
Dipole moment is a vector quantity and for compound net dipole moment is the vector sum of all dipoles. Hence dipole moment of *cis*-form is greater than *trans*-form.



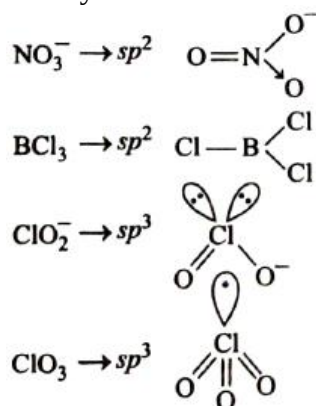
55. (c)
The hybridisation of central atom of given species in different options are
 BF_3 , NO_2^- , $\text{NO}_2 \Rightarrow sp^2$ -hybridisation
 H_2O and $\text{NH}_2^- \Rightarrow sp^3$ -hybridisation
So, the pair of central atoms that exhibit sp^2 -hybridisation is BF_3 and NO_2^- .

56. (c)
The geometry of the given species are as follows



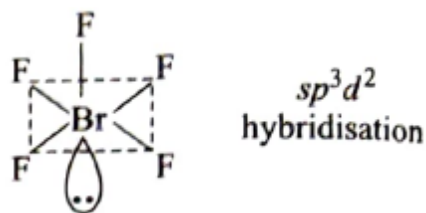
Thus, among the given species only SO_3^{2-} has pyramidal geometry.

57. (b)
The hybridisation of central atom in the given molecules/ ions are as follows

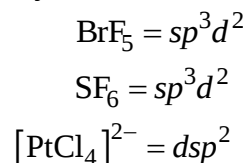


Thus, two molecules / ions i.e. ClO_2^- and ClO_3 has sp^3 hybridisation.

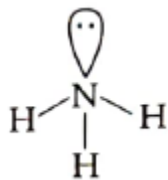
58. (a)
Among the given options, BrF_5 has square pyramidal shape.



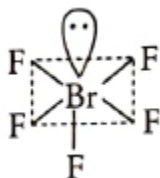
59. (b)
Among the given options, only $[\text{Co}(\text{NH}_3)_6]^{3+}$ shows d^2sp^3 hybridisation.
Hybridisation of rest of the members are as follows



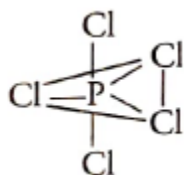
60. (c)
The correct match is A-II, B-I, C-IV, D-II
 NH_3 = Trigonal pyramidal



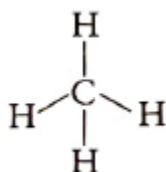
BrF_5 = Square pyramidal



PCl_5 = Trigonal bipyramidal



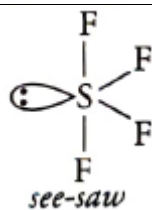
CH_4 = Tetrahedral



61. (a)
- | Molecules | Hybridisation | Shape/geometry |
|-----------|---------------|----------------|
|-----------|---------------|----------------|

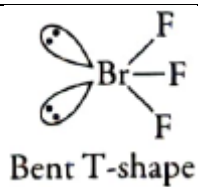
SF_4

sp^3d



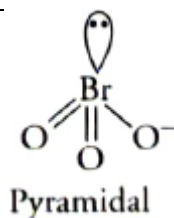
BrF_3

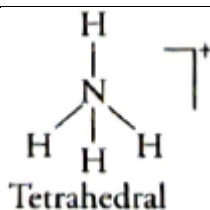
sp^3d



BrO_3^-

sp^3



 sp^3 

Hence, correct match is A-III, B-IV, C-II, D-I.

62. (c)

The correct match is A-III, B-I, C-II, D-IV.

- SO_2Cl_2 has sp^3 -hybridisation and has tetrahedral geometry.
- NO is paramagnetic in nature.
- NO_2^- is diamagnetic in nature.
- I_3^- has sp^3d -hybridisation and has linear geometry.

63. (c)

The correct match is A-III, B-IV, C-I, D-II.

Hybridisation Orientation

$sp^3 \rightarrow$ Tetrahedral

$dsp^2 \rightarrow$ Square planar

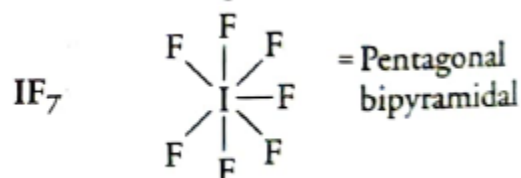
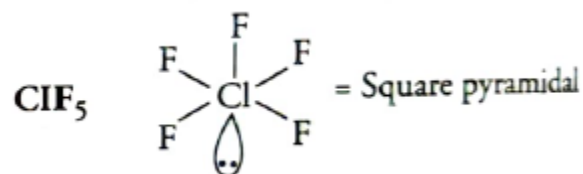
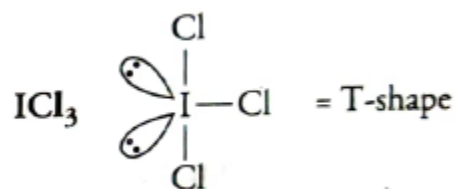
$sp^3d \rightarrow$ Trigonal bipyramidal

$sp^3d^2 \rightarrow$ Octahedral

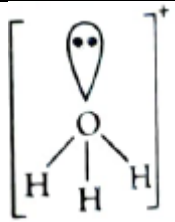
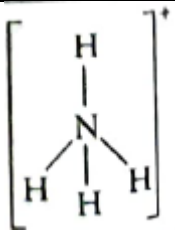
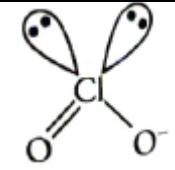
64. (a)

The correct match is A-IV, B-I, C-II, D-III.

ICl = Linear



65. (d)
The correct match is A-III, B-II, C-I, D-IV.

Molecules	Hybridisation	Shape	Structure
A. H_3O^+	sp^3	Pyramidal (III)	
B. Acetylide anion	sp	Linear (II)	$\ominus\text{C}\equiv\text{C}^\ominus$
C. NH_4^+	sp^3	Tetrahedral (I)	
D. ClO_2^-	sp^3	Bent (IV)	

66. (a)
Antibonding molecular orbital (ABMO) are formed by destructive interference of wave function.
(AMBO) $\sigma^* = \Psi_A - \Psi_B$

67. (a)
Statements given in A, D and E are correct while Statement given in B and C are incorrect. Their correct form are as follows
- Intramolecular hydrogen bonding is present in *o*-nitrophenol.
 - Intermolecular hydrogen bonding is present in HF.

SECTION-II

1. (0)
Molecular orbital configuration of O_2^{2-} is :

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

$$\pi^* 2p_x^2 = \pi^* 2p_y^2$$

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

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

$$\text{Difference} = 0 = \frac{x}{2} \Rightarrow x = 0$$

3. (15)

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

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

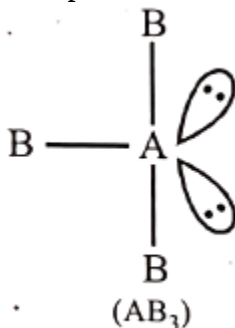
As AX is diatomic molecule (neutral), the only possible case is NO.

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.

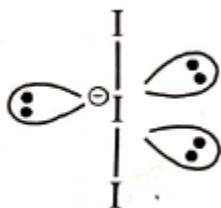
4. (2)

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



5. (3)

I₃⁻ : It is sp³d hybridised.



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

6. (2)

Species	Structure	Lone Pair
SF ₄		1
BF ₄ ⁻		0
ClF ₃		2

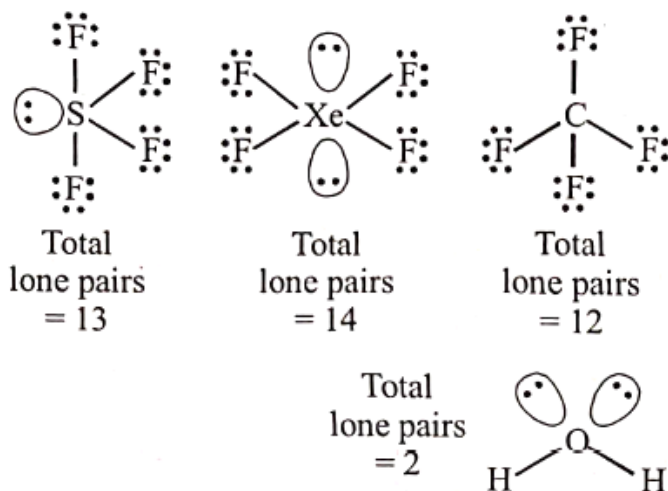
AsF ₃	$\begin{array}{c} \text{F} - \overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{As}}} - \text{F} \\ \\ \text{F} \end{array}$	1
PCl ₅	$\begin{array}{c} \text{Cl} \\ \\ \text{Cl} - \text{P} - \text{Cl} \\ \quad \diagup \quad \diagdown \\ \text{Cl} \quad \text{Cl} \end{array}$	0
BrF ₅	$\begin{array}{c} \text{F} \quad \quad \text{F} \\ \diagdown \quad \diagup \\ \text{Br} \\ \diagup \quad \diagdown \\ \text{F} \quad \quad \text{F} \\ \\ \text{F} \end{array}$	1
XeF ₄	$\begin{array}{c} \text{F} \quad \quad \text{F} \\ \diagdown \quad \diagup \\ \text{Xe} \\ \diagup \quad \diagdown \\ \text{F} \quad \quad \text{F} \end{array}$	2
SF ₆	$\begin{array}{c} \text{F} \\ \\ \text{F} \quad \text{S} \quad \text{F} \\ \diagdown \quad \diagup \\ \text{F} \quad \text{S} \quad \text{F} \\ \\ \text{F} \end{array}$	0

7. (1)
 S.N. = [5 B.P. + 0 L.P.)
 $\text{PF}_5 \Rightarrow \text{sp}^3\text{d}$ hybridization

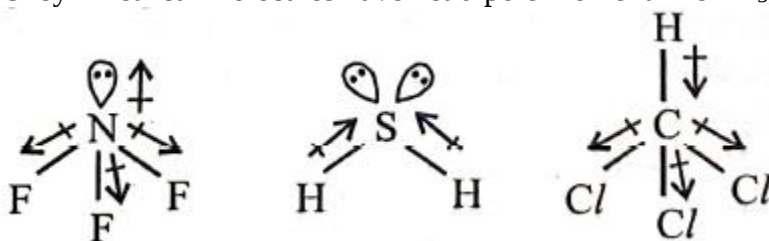
8. (4)
 The molecular orbital configuration of B₂ is : $(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\pi 2p_x)^1 (\pi 2p_y)^1$
 The molecular orbital configuration of C₂⁺ is : $(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\pi 2p_x)^2 (\pi 2p_y)^2 (\sigma 2p_z)^1$
 The molecular orbital configuration O₂⁺ is :
 $(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x)^2 (\pi 2p_y)^2 (\pi^* 2p_x)^1$
 The molecular orbital configuration of He₂⁺ is : $(\sigma 1s)^2 (\sigma^* 1s)^1$
 Hence, B₂, C₂⁻, O₂⁺ and He₂⁺ are paramagnetic species.

9. (3)
 CN⁻, NO⁺ and O₂²⁺ all have bond order 3.

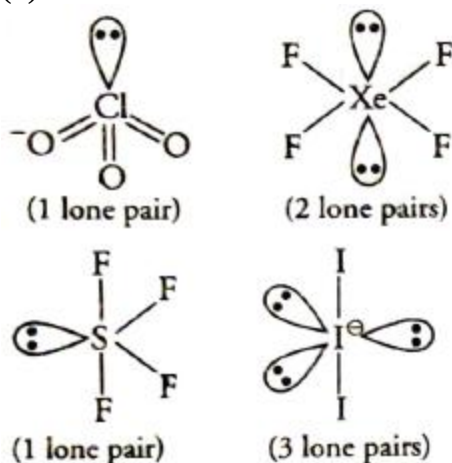
10. (1)
 The electron pairs which do not take part in bonding and resides on atom are called lone pair.



11. (3)
 Symmetrical molecules have zero dipole moment like BF_3 , BeF_2 , SiF_4 , CCl_4 and PF_5 .
 Unsymmetrical molecules have net dipole moment like NF_3 , CHCl_3 and H_2S



12. (3)

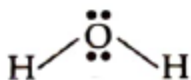


Among the given species I_3^- contains maximum number of unpaired electrons on its central atom i.e. 3.

13. (4)

In water (H_2O) molecule, oxygen is bonded to two hydrogen atoms. Total number of valence electrons present in oxygen atom is 6. It shares one electron with each hydrogen atom to form two single bonds.

Since, four valence electrons do not participate in any bond formation, there are two lone pairs in water molecule.



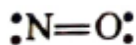
In CO, each carbon and oxygen atom contains one lone pair of electrons.



The nitrogen atoms are bonded with triple bonds. So, each of them have a lone pair of electron. Hence, total number of lone pair is 2.

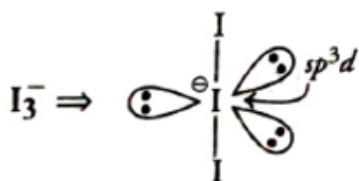


In nitric oxide (NO). Total number of lone pair is 2.

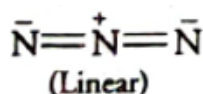


14. (3)

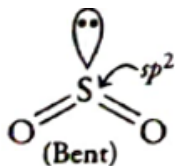
N_3^- and I_3^- are linear whereas NO_2^- , O_3 and SO_2 are bent



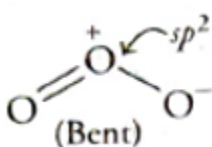
Also, N_3^- structure is linear and has sp -hybridisation.



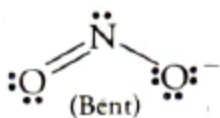
SO_2 is bent in shape.



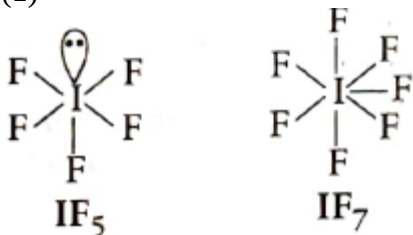
O_3 is bent in shape.



NO_2^- , that is nitrite ion, the N-atom have sp^2 -hybridisation, thus it adopts bent geometry.



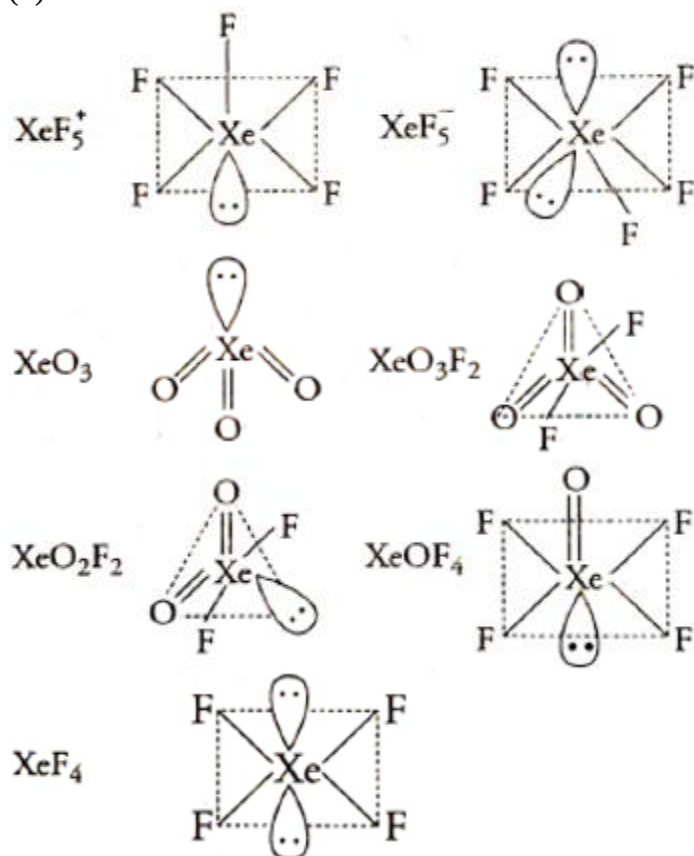
15. (1)



Lone pairs present on the central atom of interhalogen, IF_5 is one and IF_7 is zero.

So, sum of lone pairs present on the central atom of the interhalogen IF_5 and IF_7 is $(1 + 0)$ equal to one.

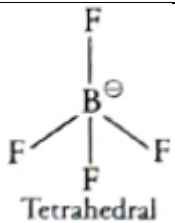
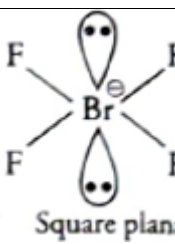
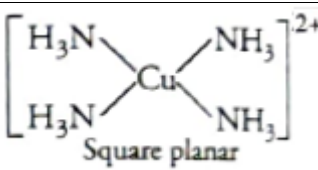
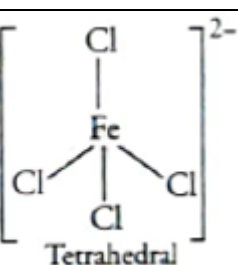
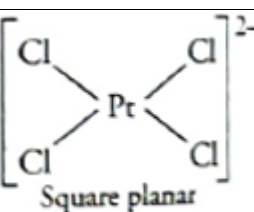
16. (4)



Thus, XeF_5^+ , XeO_3 , XeO_2F_2 and XeOF_4 have single lone pair on central atom.

17. (4)

Species	Hybridisation	Structure
XeF_4	sp^3d^2	Square planar
SF_4	sp^3d	see-saw
SiF_4	sp^3	Tetrahedral

BF_4^-	sp^3	 Tetrahedral
BrF_4^-	dsp^2	 Square planar
$[\text{Cu}(\text{NH}_3)_4]^{2+}$	dsp^2	 Square planar
$[\text{FeCl}_4]^{2-}$	sp^3	 Tetrahedral
$[\text{PtCl}_4]^{2-}$	dsp^2	 Square planar

So, total 4 species have square planar shape.

18. (3)

PF_5	sp^3d	Trigonal pyramidal
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BrF_4^-	sp^3d^2	Square planar
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IF_5	sp^3d^2	Square pyramidal
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BrF_5	sp^3d^2	Square pyramidal
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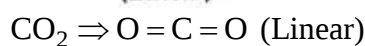
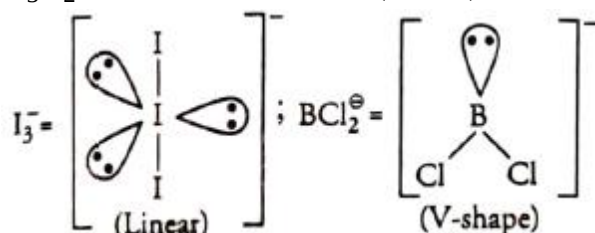
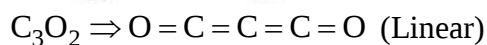
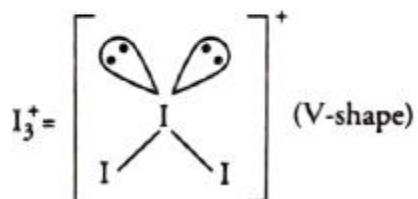
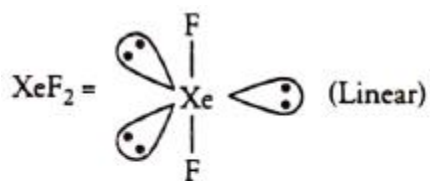
XeOF_4	sp^3d^2	Square pyramidal
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ICl_4^-	sp^3d^2	Square planar
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In above molecules, 3 molecules (IF_5 , BrF_5 , XeOF_4) have square pyramidal structure.

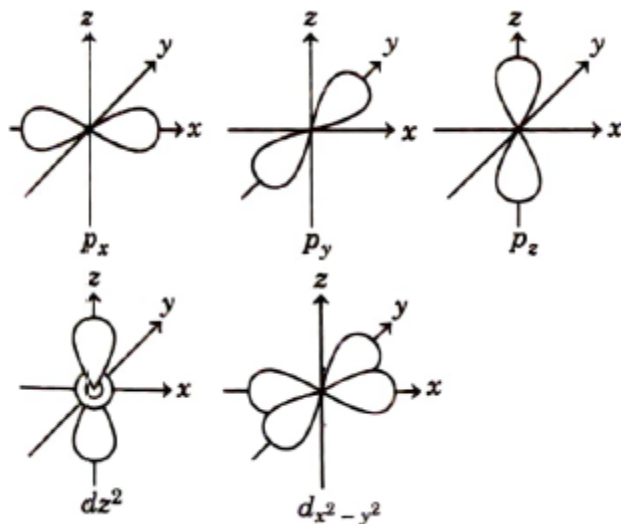
19. (5)

Among the given molecule only five i.e. XeF_2 , C_3O_2 , I_3^- , CO_2 , BeCl_2 have linear structure.



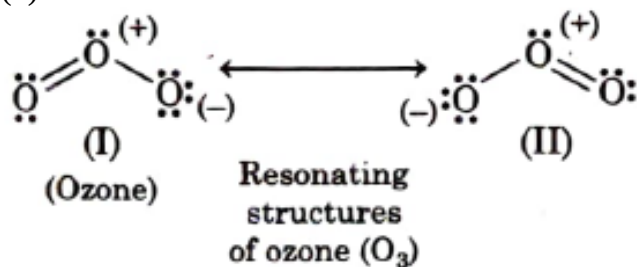
20. (5)

p_x , p_y , p_z , p_{z^2} and $d_{x^2-y^2}$ orbitals have electron density along the axes.



d_{xy} and d_{yz} have electron density between the axes (on the plane).

21. (6)



Ozone has 6 lone pair of electrons.

22. (6)

NO_2 , H_2SO_4 , BF_3 , ClO_2 , PCl_5 , BeF_2 . These are exception of octet rule.

Hence, among the given molecules total 6 molecule are exception of octet rule.

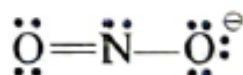
23. (6)

Among the given, H_2 , CO_2 , BF_3 , CH_4 , SiF_4 , BeF_2 are symmetrical molecules.

Hence, total 6 molecules will show zero dipole moment.

24. (8)

The Lewis dot structure of NO_2^- is



Number of valence e^- around N-atom = 8.

25. (5)

Polar molecules have non-zero dipole moments while non-polar molecules have zero dipole moments. Thus, among the given species/compounds NF_3 , H_2O , H_2S , HBr and HCl will have non-zero dipole moment.

26. (3)

Molecule with net zero dipole moment are CO_2 , CH_4 and BF_3 i.e. 3

Remaining all molecules have non-zero dipole moment.

27. (4)

The number of non-polar molecules are 4. These are CO_2 , H_2 , CH_4 and BF_3 .

Remaining all compound are polar molecules.

HF , H_2O , SO_2 , NH_3 , $CHCl_3$, HCl .

28. (6)

Among the given species, total 6 species have sp^2 hybrid orbital for bonding with central atoms.

These are SO_2 , C_2H_4 , BCl_3 , $HCHO$, C_6H_6 , BF_3

29. (3)

Molecules / Ions	Shapes
PF_5 , PCl_5 , $Fe(CO)_5$	Trigonal bipyramidal
BrF_5	Square pyramidal

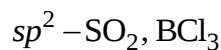
$[\text{PrCl}_4]^{2-}$	Square planar
BF_3	Trigonal planar

Hence, total 3 molecules have trigonal bipyramidal shape.

30. (4)

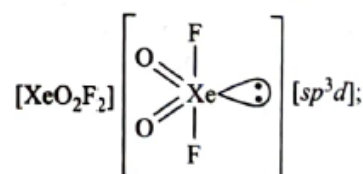
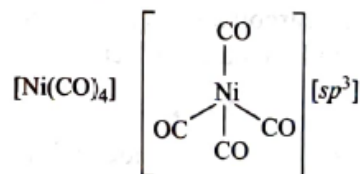
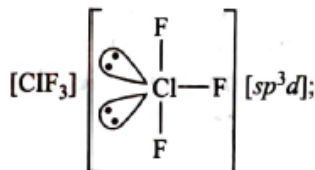
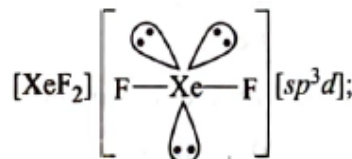
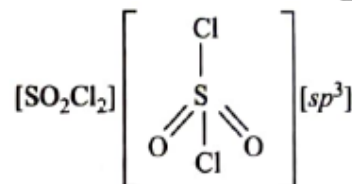
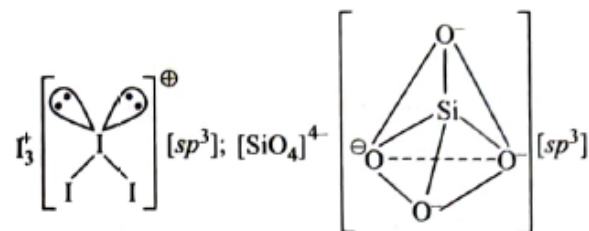
The number of species that uses sp^3 hybrid orbitals in bonding are 4 i.e., H_2O , NH_3 , SiO_2 , CH_4 .

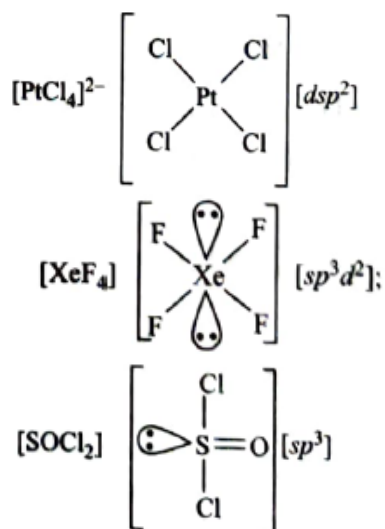
While other species uses sp^2 and sp -hybrid orbitals which are as follows.



31. (5)

The structures of given species among with their hybridisation is given below





Thus, I_3^+ , $[\text{SiO}_4]^{4-}$, SO_2Cl_2 , $\text{Ni}(\text{CO})_4$ and SOCl_2 have sp^3 -hybridised central atom.

32. (6)

The electronic configuration of $\text{O}_2(16e^-)$ is

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

$$\pi 2p_x^2 = \pi 2p_y^2, \pi^* 2p_x^1 = \pi^* 2p_y^1$$

So, number of e^- in π^* of $\text{O}_2 = 2$

Number of e^- in π^* of $\text{O}_2^+ = 1$

Number of e^- in π^* of $\text{O}_2^- = 3$

So, total number of e^- in $\pi^* = 2 + 1 + 3 = 6$

33. (4)

Species	Unpaired electrons
N_2	0
O_2	2
C_2^-	1
O_2^-	1
O_2^{2-}	0
H_2^+	1
CN^-	0
He_2^+	1

So, one unpaired electron is present in C_2^- , O_2^- , H_2^+ and He_2^+ .

34. (2)

$$\text{Formula for Bond order (B.O)} = \frac{N_B - N_A}{2}$$

Where, N_B = Number of electron in bonding orbital.

N_A = Number of electron in antibonding orbital.

For C_2

$$\text{B.O.} = \frac{8-4}{2} = 2$$

For O_2

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

For Be_2

$$\text{B.O.} = \frac{4-4}{2} = 0$$

For Li_2

$$\text{B.O.} = \frac{4-2}{2} = 1$$

For Ne_2

$$\text{B.O.} = \frac{4-2}{2} = 1$$

For N_2

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

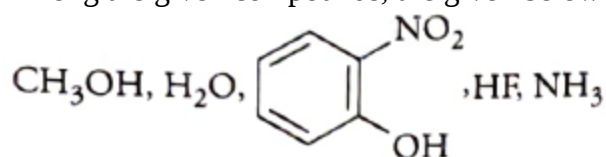
For He_2

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

Hence, total 2 molecules have bond order of 2.

35. (5)

Among the given compounds, the given below compounds exhibit hydrogen bonding,



Hence, total 5 compounds are there.

36. (2)

According to molecular orbital theory, number of unpaired electrons in the given molecules / species as follows

O_2 = Number of unpaired electrons is 2

O_2^{-1} = Number of unpaired electron is 1

NO = Number of unpaired electron is 1

CN^{-1} = Number of unpaired electron is zero

O_2^{2-} = Number of unpaired electron is zero

37. (8)
Molecular orbital formed by 2s atomic orbital = 2($\sigma 2s$ and $\sigma^* 2s$)
Molecular orbital formed by 2p atomic orbital
= 6($\sigma 2p_z$ and $\sigma^* 2p_z$, $\pi 2p_x$ $\pi 2p_y$ and $\pi^* 2p_x$, $\pi^* 2p_y$)
So, total molecular orbitals = 2 + 6 = 8
38. (4)
Antibonding molecular orbital from 2s = 1
Antibonding molecular orbital from 2p = 3
Therefore, total number of antibonding molecular orbitals are = 1 + 3 = 4

39. (1)

Ion/molecule	Magnetic behaviour	Bond order
H ₂	Diamagnetic	1
He ₂ ⁺	Paramagnetic	0.5
N ₂ ²⁻	Paramagnetic	2
O ₂ ²⁻	Diamagnetic	1
O ₂ ⁺	Paramagnetic	2.5
F ₂	Diamagnetic	1
Ne ₂ ⁺	Paramagnetic	0.5
B ₂	Paramagnetic	1

Only B₂ has paramagnetic magnetic behaviour and bond order of 1.

40. (6)
NO⁺ and CO both are isoelectronic species and each of them is having bond order 3.
 $\therefore$ Sum of bond order of NO⁺ and CO = Bond order (NO⁺) + Bond order (CO) = 3 + 3 = 6