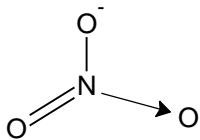
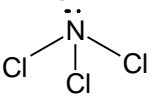
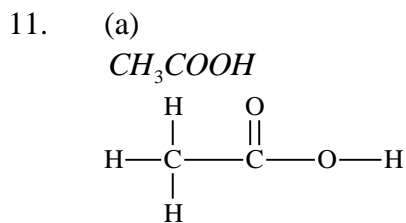


EXERCISE - 1 [A]

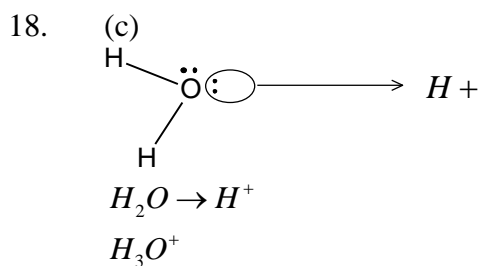
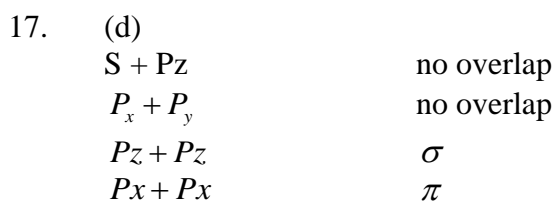
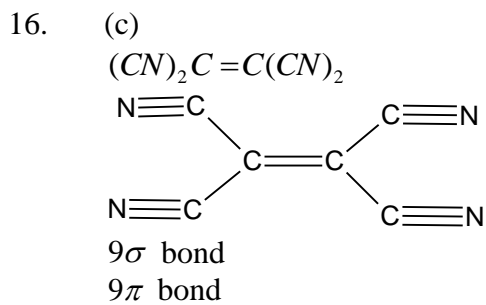
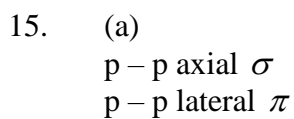
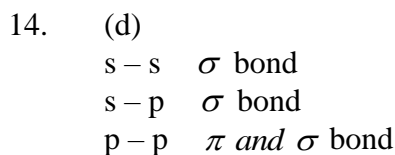
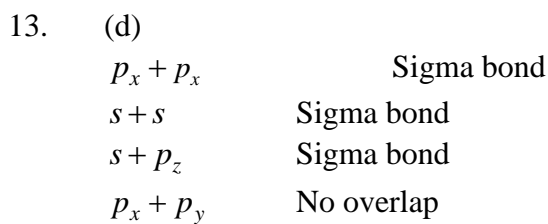
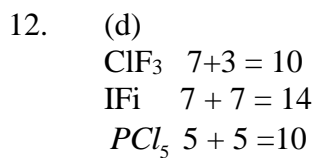
- (b)
Lower the ionisation energy, cation is easily formed. Similarly higher the electron affinity, anion is easily formed.
- (c)
 K^+I^-
- (b)
 $H-Cl$
- (b)
$$L.E. \propto \frac{q^{\oplus} q^{\ominus}}{r^{\oplus} + r^{\ominus}}$$

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


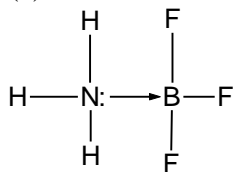
4 bond pair electrons & 0 lone pair electrons.
- (d)
 NCl_3

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



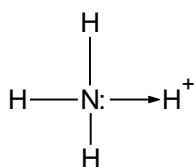
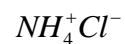
16 Shared electrons
8 unshared electrons



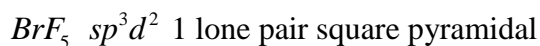
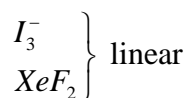
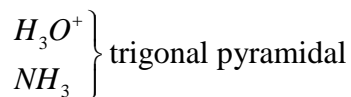
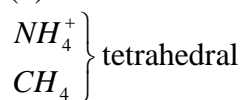
19. (c)



20. (b)



21. (d)

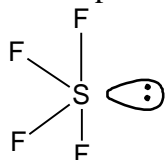


22. (b)



4 σ bond

1 lone pair



G: trigonal bipyramidal

S : See – saw

23. (d)



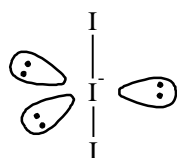
2 σ bond

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.

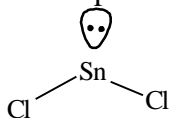
27. (b)



Atomic no. = 50 (Sn)



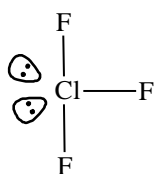
1 lone pair and 2 σ -bond



28. (c)



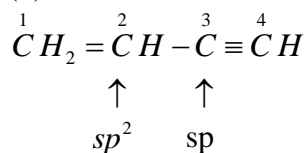
2 lone pair } sp^3d
 3 σ bond }



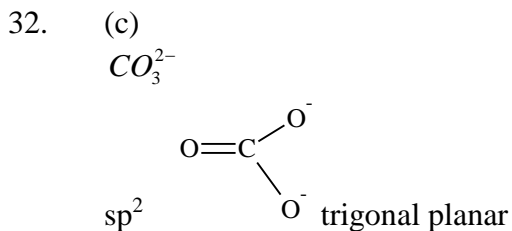
G : trigonal bipyramidal

S : T-shape

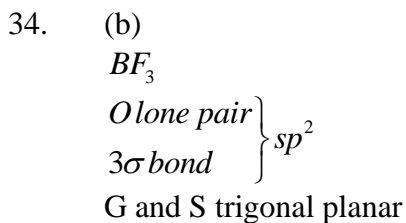
29. (d)



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

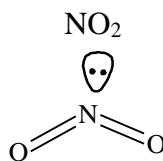
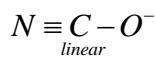
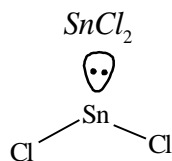


33. (b)
 s and p
 sp hybridisation angle = 180°



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 σ bond, sp linear

38. (b)
In F_2 , P_z & P_z orbitals overlap to form σ bond.

39. (b)

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

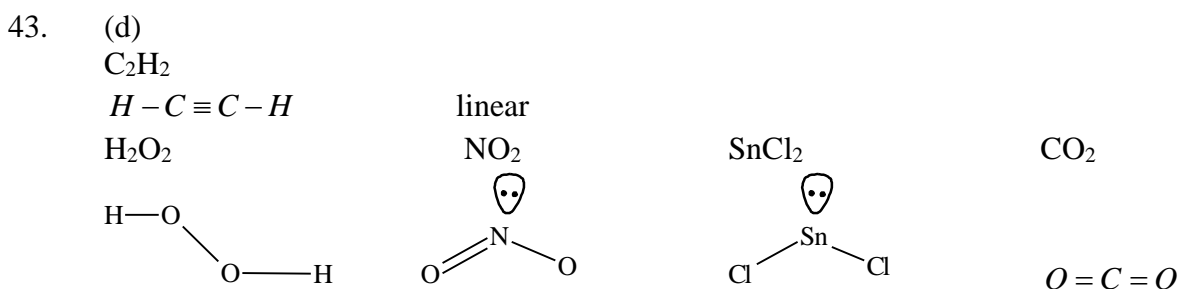
$$\%s = \frac{1}{3} \times 100 \quad \frac{1}{5} \times 100$$

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

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)

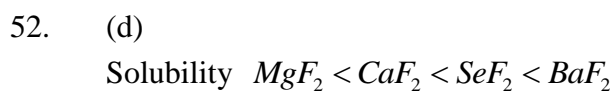
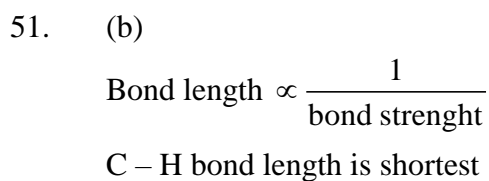
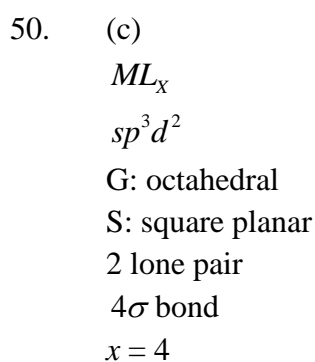
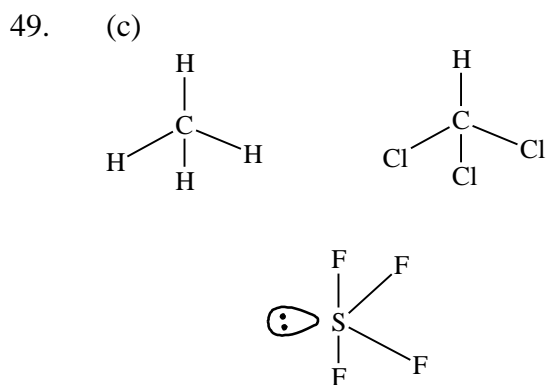
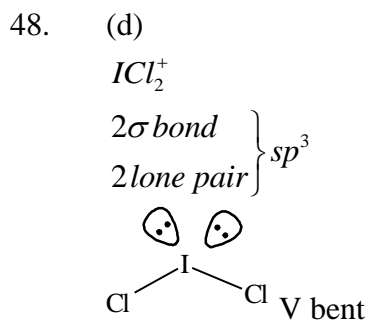
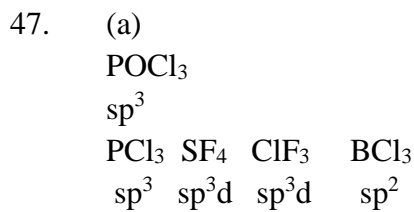
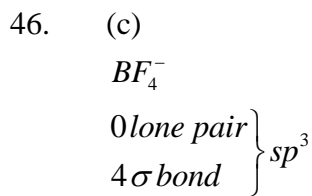
$$OF_2 \left. \begin{array}{l} 2\sigma \text{ bond} \\ 2 \text{ lone pair} \end{array} \right\} sp^3$$


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)

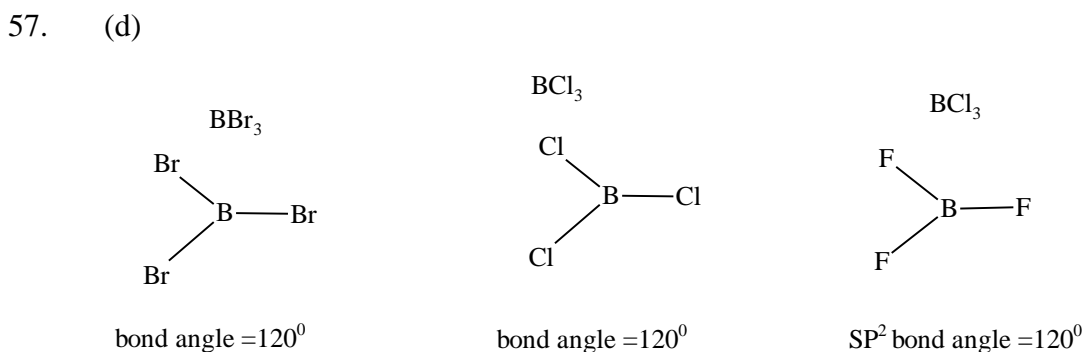
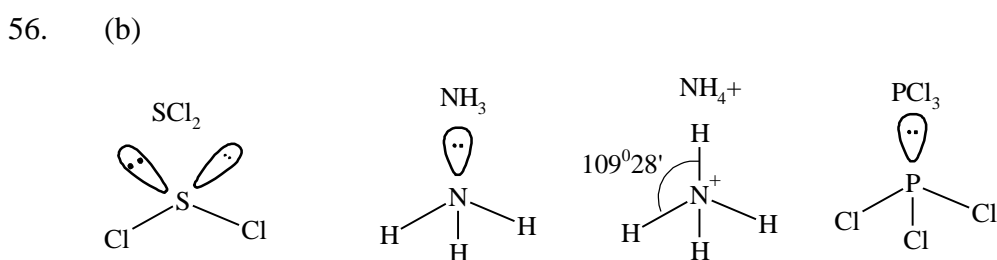
PCl_3 sp^3 (trigonal pyramidal)
 NH_3 sp^3 (trigonal pyramidal)
 CF_4 sp^3 (tetrahedral)
 SF_4 sp^3d (tetrahedral)
 $PbCl_2$ sp^2 (V-bent)
 CO_2 sp (Linear)
 PF_5 sp^3d (trigonal bipyramidal)
 IF_5 sp^3d^2 (square pyramidal)



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

54. (b)
 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

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

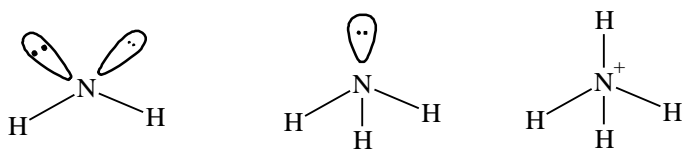
59. (b)
Least covalent character highest melting point.

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

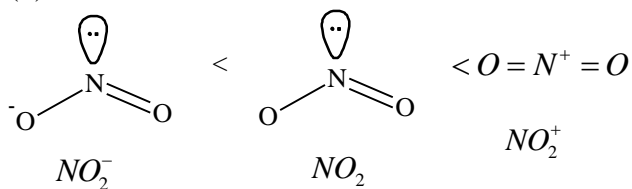
61. (c)
Covalent $BeCl_2 > MgCl_2 > CaCl_2 > SrCl_2 > BaCl_2$

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

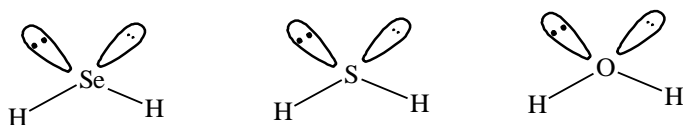
63. (b)
 $NH_2^- < NH_3 < NH_4^+$



64. (d)
 $NO_2^- < NO_2 < O=N^+=O$
 NO_2^- NO_2 NO_2^+



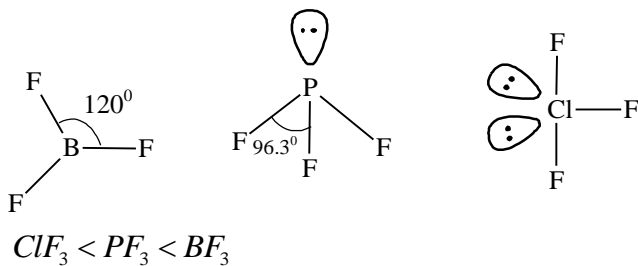
65. (c)
 $H_2Se < H_2S < H_2O$



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

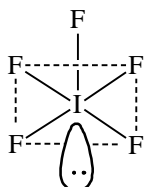
67. (b)
 NH_3 CH_4 H_2O H_2S
 107° $109^\circ 28'$ 104.5° 92°

68. (c)
 BF_3 PF_3 ClF_3



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

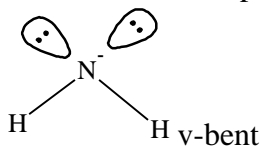
70. (c)
 IF_5
1 lone pair
5 σ bond
 sp^3d^2 octahedral
S: square pyramidal
 90° and 180°



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 $AlI_3 > AlBr_3 > AlCl_3 > AlF_3$

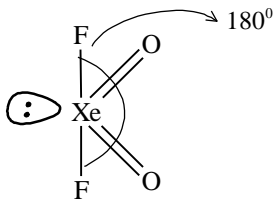
73. (c)
 NH_2^-
2 σ bond , 2 lone pair } sp^3



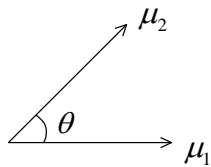
74. (c)

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

75. (c)
XeO₂F₂



76. (a)

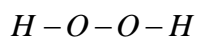


$$\mu_{net} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos \theta}$$

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

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

77. (c)



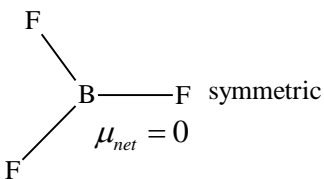
Non polar bond

78. (c)

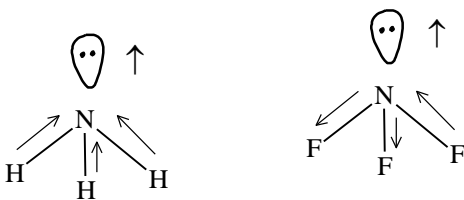
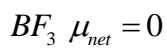


$\Delta(E.N) \uparrow$ Polar character \uparrow

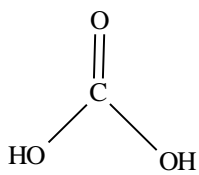
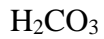
79. (b)



80. (d)

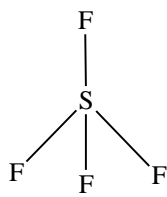


81. (a)



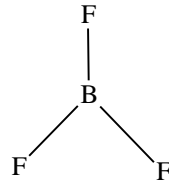
sp^2

Polar



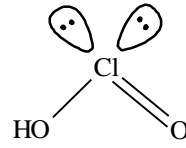
sp^3

Non polar



sp^2

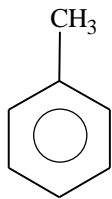
Non polar



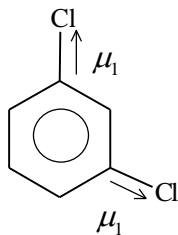
sp^3

Polar

82. (b)



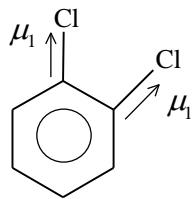
I



II

$$\theta = 120^\circ$$

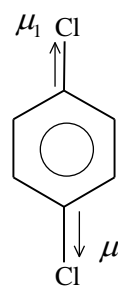
$$\mu_{net} = \mu_1$$



III

$$\theta = 60^\circ$$

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



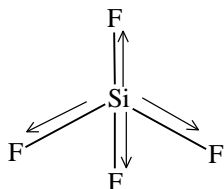
IV

$$\theta = 180^\circ$$

$$\mu_{net} = 0$$

$$IV < I < II < III$$

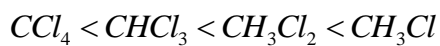
83. (c)



$$\mu_{net} = 0$$

84. (a)

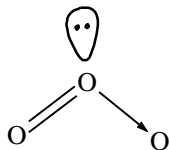
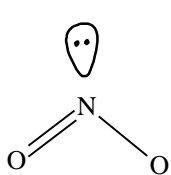
Dipole moment



85. (b)

μ_{net} of SiF_4 and $\text{CO}_2 = 0$

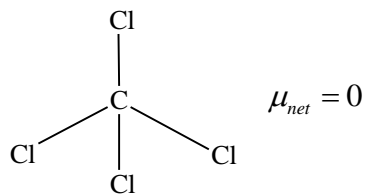
μ_{net} of NO_2 and $\text{O}_3 \neq 0$



86. (c)

Covalent (directional)

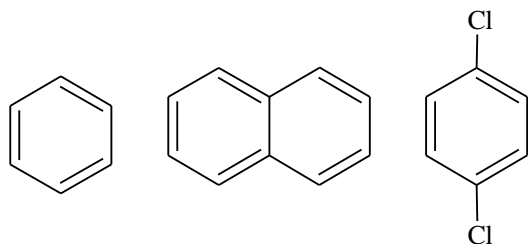
87. (d)



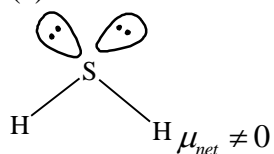
88. (b)

C - H
 $\Delta(E.N)$ is very less

89. (d)

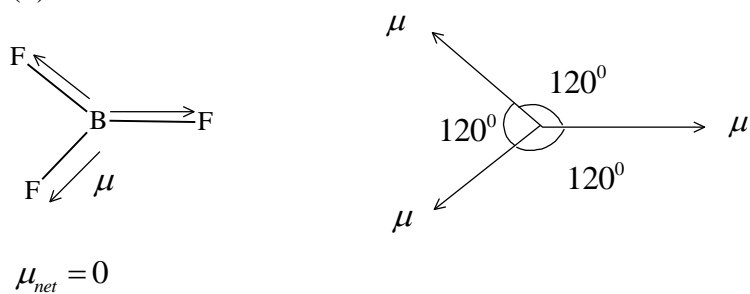


90. (a)

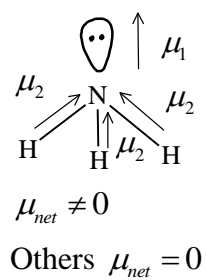


Other $\mu_{net} = 0$
(CO_2, BF_3, CCl_4)

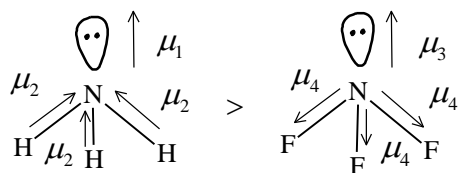
91. (d)



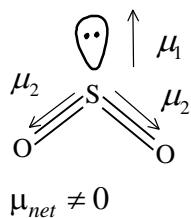
92. (c)



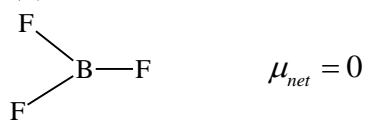
93. (b)
 μ_{net} of BF_3 and $B_2H_6 = 0$



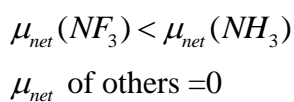
94. (c)



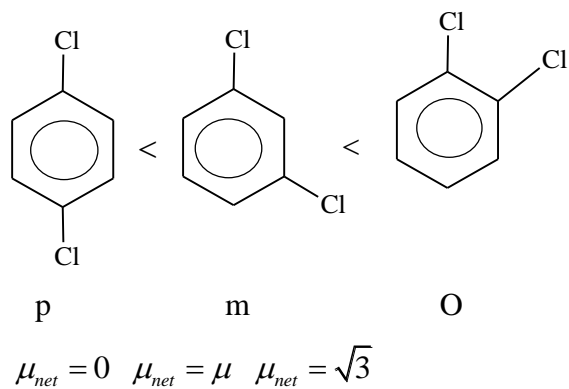
95. (b)



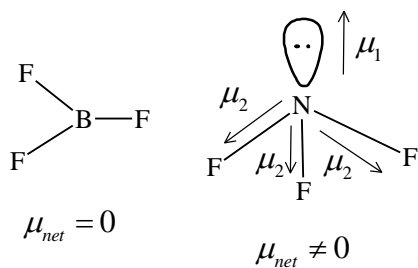
96. (a)



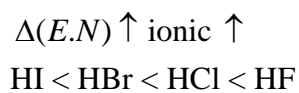
97. (d)



98. (c)



99. (a)



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

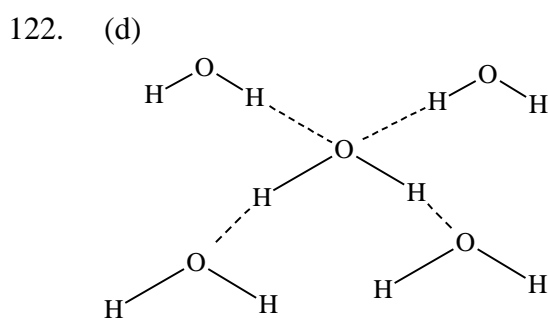
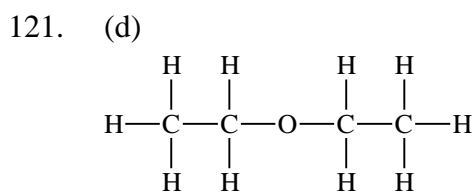
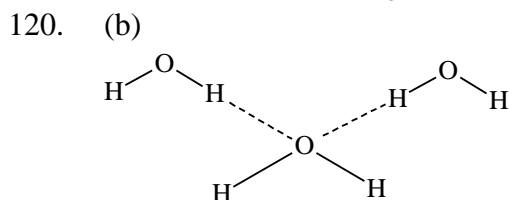
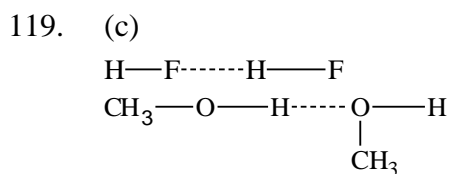
N_2^+ Paramagnetic

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

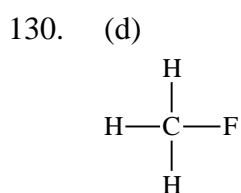
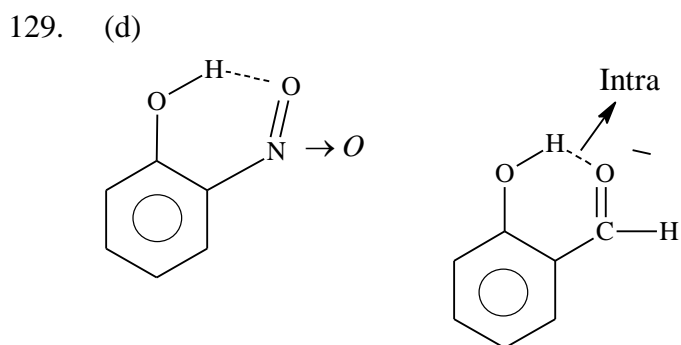
124. (a)
B.P. $\text{H}_2\text{S} < \text{H}_2\text{Se} < \text{H}_2\text{Te} < \text{H}_2\text{O}$ \uparrow due to intermolecular H-bonding

125. (d)
HCl (no H – bonding)

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

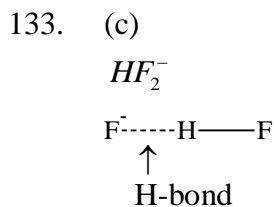
127. (b)
Bond length of H – bond $>$ bond length of covalent bond

128. (b) $CH_3OH < H_2O$ CH_3OCH_3
 Greater strength of H-bonding in water No H-bonding



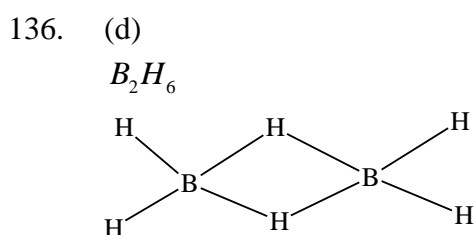
131. (d)
 Hydrogen bond is weaker than other bonds

132. (d)
 Molar mass \uparrow B.P. \uparrow
 B.P. : $CH_4 < SiH_4 < GeH_4$



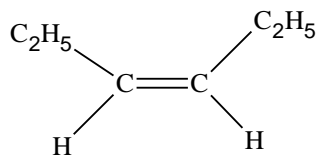
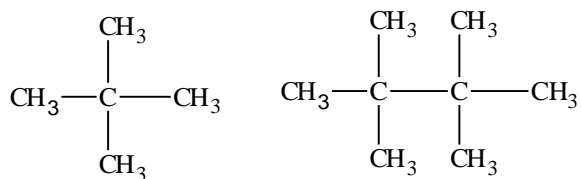
134. (a)
- | | |
|---|---|
| Inter molecular
$\begin{array}{c} \text{CH}_2\text{OH} \\ \\ \text{CHOH} \\ \\ \text{CH}_2\text{OH} \end{array}$
Glycerol | H-bonding \uparrow viscosity \uparrow
$\begin{array}{c} \text{CH}_2\text{OH} \\ \\ \text{CH}_2\text{OH} \end{array}$
Glycol |
|---|---|

135. (b)
 Covalent bond is directional.



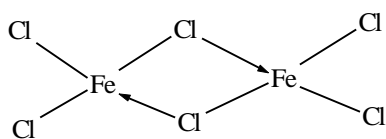
(2c 2e bond or banana bond)

137. (c)



max. dipole moment

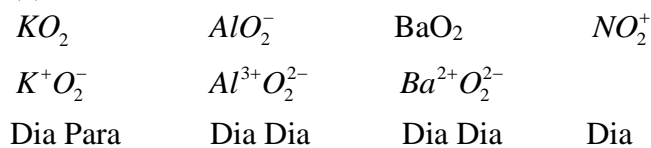
138. (a)



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 11th or 12th group.
 $Sn = [Kr] 5s^2 4d^{10} 5p^2$
 $Sn^{4+} = [Kr] 5s^0 4d^{10}$
 $Ag = [Kr] 5s^1 4d^{10}$
 $Ag^+ = [Kr] 5s^0 4d^{10}$

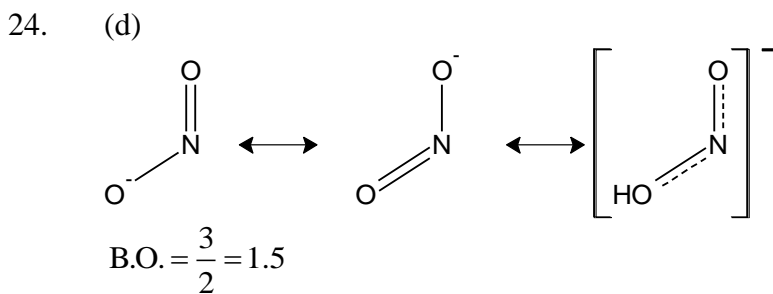
9. (c)
 $Fe = [Ar] 4s^2 3d^6$
 $Fe^{2+} = [Ar] 4s^0 3d^6$

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

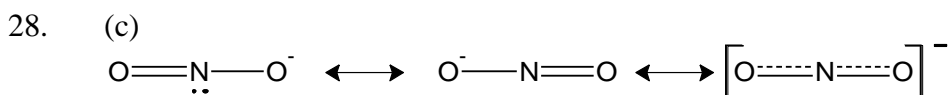
23. (a)
- | | | | |
|-----------------------------|----|-----------------|-------------------------------|
| | CO | CO ₂ | CO ₃ ²⁻ |
| Bond order : | 3 | 2 | 1.33 |
| Bond order ↑, Bond length ↑ | | | |



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 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)
Sr = [Kr] 5s²
Sr = [Kr] 5s¹

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
$\underbrace{\hspace{10em}}_{\text{isoelectronic}}$			

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.

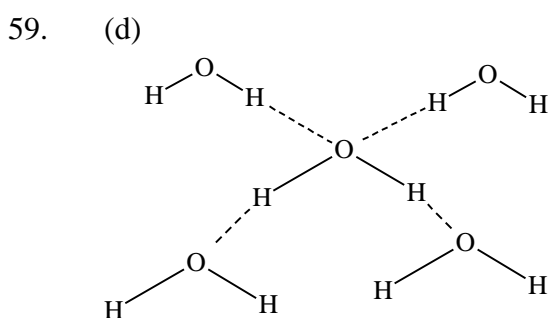
46. (a)
Solid PCl_5 exists as
 $[\text{PCl}_4]^+ [\text{PCl}_6]^-$
 $sp^3 \quad sp^3d$
47. (b)
- | | | | |
|----------------|-------------------|----------------|--------------|
| O_2^- | O_2^{2-} | O_2^+ | O_2 |
| Paramagnetic | Diamagnetic | Paramagnetic | Paramagnetic |
48. (b)
- $$\text{H}-\underset{\text{O}}{\underset{\parallel}{\text{C}}}-\text{O}-\text{H}\cdots\text{C}-\underset{\text{O}}{\underset{\parallel}{\text{C}}}-\text{CH}_3$$
49. (d)
 PCl_5
 5σ & $0 \ell.p.$
 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

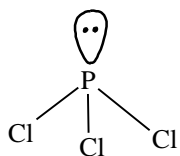
57. (d)
Polarisation \propto charge
 \propto size of anion
 $\propto \frac{1}{\text{size of cation}}$

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
 $3\sigma \text{ bond}$
 1 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 AlCl_3

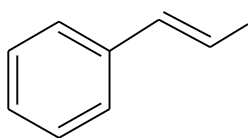
62. (c)
S character \uparrow electronegativity \uparrow
 $sp > sp^2 > sp^3 > sp^3d > \dots$

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

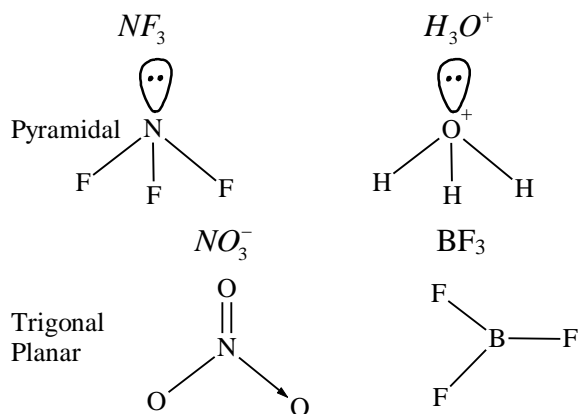
0 lone pair

sp

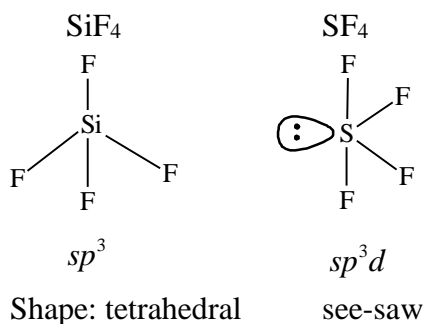
linear

linear

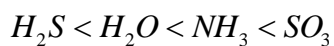
66. (c)



67. (d)



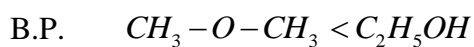
68. (c)



90°

120°

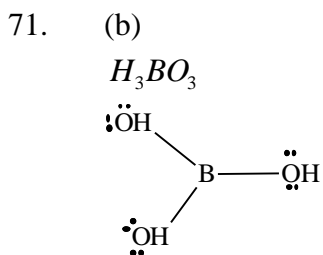
69. (d)



Volatility $CH_3-O-CH_3 > C_2H_5OH$

Due to hydrogen bonding in C_2H_5OH

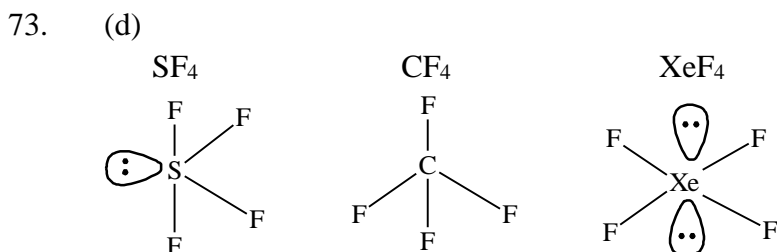
70. (b)
- | | | |
|------|-----|-----------------|
| | NO | NO ⁺ |
| b.o. | 2.5 | 3 |
| b.o. | ↑ | b.l. ↓ |
- b.l.*: NO > NO⁺



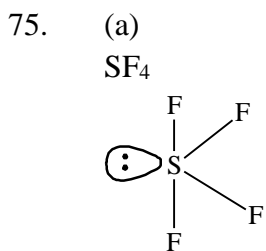
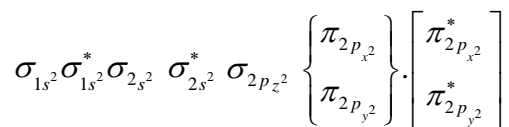
For B : 3 σ bond ; 0 l.p (sp^2)

For O: 2 σ ; 2 l.p (sp^3)

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



74. (a)
- O₂²⁻ Diamagnetic



B.l.(S-F)_{equ} ≠ *B.l.*(S-F)_{axial}

76. (b)
- O₂²⁻
- Diamagnetic

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

Size of cation ↓

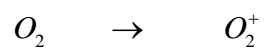
And

Charge ↑

Polarisation ↑

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

78. (c)

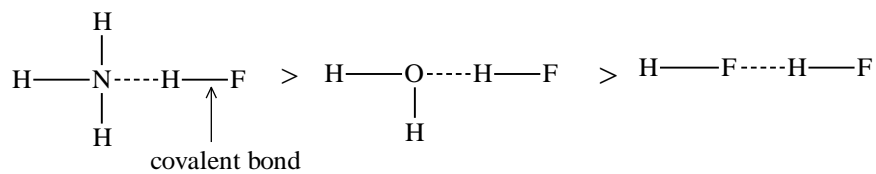


$$\text{b.o. } \frac{6-2}{2} = 2 \quad \frac{6-1}{2} = 2.5$$

para

Para

79. (c)



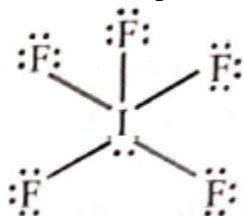
EXERCISE - 1 [C]

1. (1)
9 σ and 9 π 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)
CCl₄, SO₃, SF₆, PCl₅
14. (4)
BF₃, SF₆, B₃N₃H₆, PCl₃F₂
15. (9)
I₃⁺, XeF₄, C₂F₄, BrF₄⁻, SO₃, NOCl, ClF₃, F₂CO, XeF₅⁻
16. (6)
O₂²⁺, O₂²⁻, B₂²⁻, C₂²⁻, C₂, N₂
17. (3)
KO₂, NO₂, NO
18. (3)
B.L. : N₂⁻ > N₂; NO > NO⁺; H₂⁺ < H₂⁻
19. (2)
All axial bonds.
20. (1)

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



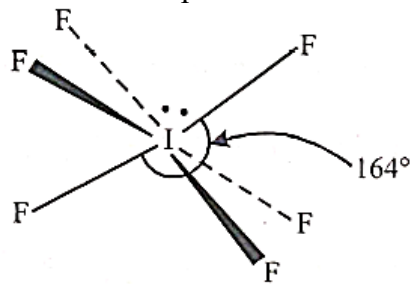
2. (d)
Given $e = 1.60 \times 10^{-19} \text{ C}$
 $d = 9.17 \times 10^{-11} \text{ m}$

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

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)]
 NO_2^+ sp hybridization
 NO_2^- sp^2 hybridization
 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)

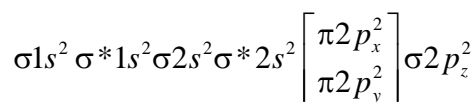
$\mu_{\text{CCl}_4} = \mu_{\text{CH}_4} = 0$ due to symmetrical structure but $\mu_{\text{CHCl}_3} \neq 0$.

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



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

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.

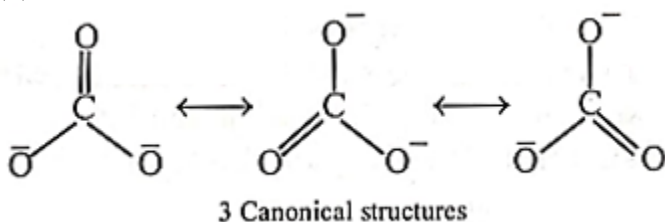
Thus, correct order is ion-ion > ion-dipole > dipole-dipole

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. 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

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

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

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

24. (d)

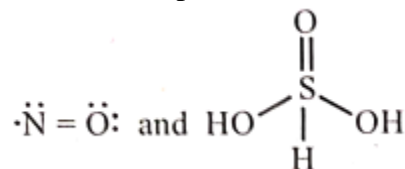
(A)-(III); (B)-(IV); (C)-(I); (D)-(II)

Molecule	Bond order
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

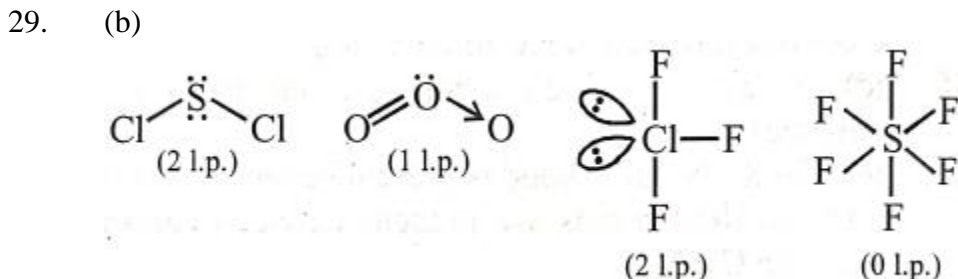


26. (b)

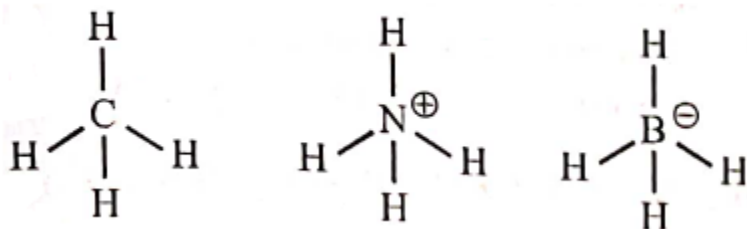
According to Fajan's rule, Covalent character \propto size of Anion. As the cation is the same in all the options. Therefore, covalent character will depend upon the size of anion. Covalent character: $\text{A} < \text{B} < \text{C} < \text{D}$; Size of anion: $\text{F}^- < \text{Cl}^- < \text{Br}^-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

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). B_2H_6 , BCl_3 have incomplete octet.

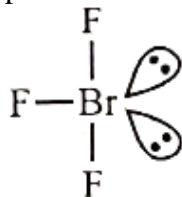


30. (b)
Isoelectronic species have same number of electrons valence



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^- .

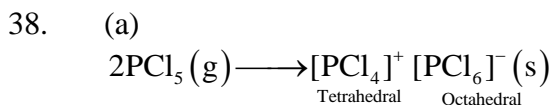
35. (c)
The strength of H-bonding depends upon two factors:
(1) Electronegativity difference between the hydrogen and Halogen element.
(2) Number of H-bonding interaction.
So, the order of H-bonding is $CH_4 < HCN < NH_3$.

36. (a)

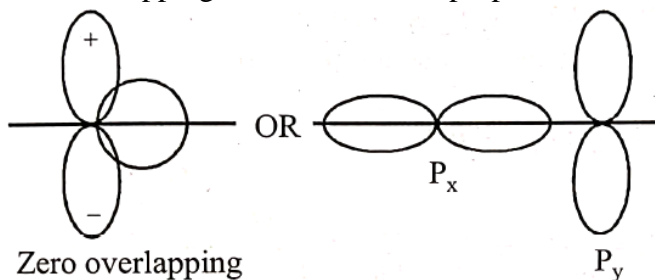
ion / molecule	Number of e^- in BMO	Number of e^- in ABMO	Bond order
O_2^+	10	5	2.5
O_2	10	6	2
O_2^-	10	7	1.5
O_2^{2-}	10	8	1

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 π^* orbital.

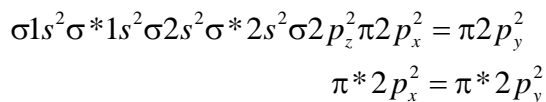


39. (a)
Zero overlapping is occur due to improper orientation or orbitals and it is an out of phase overlapping.



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

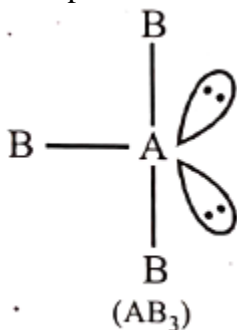
$$\Rightarrow \text{No. of } e^- \text{ in BMO} - \text{No. of } e^- \text{ in AMBO} = 5$$

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

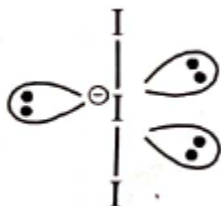
Total number of electrons = 15

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.



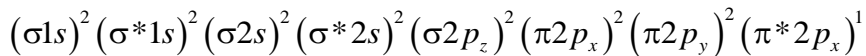
The number of lone pairs of electron on the central atom is 3.

46. (2)

Species	Structure	Lone Pair
SF ₄		1
BF ₄ ⁻		0
ClF ₃		2
AsF ₃		1
PCl ₅		0
BrF ₅		1
XeF ₄		2
SF ₆		0

47. (1)
 S.N. = [5 B.P. + 0 L.P.]
 PF₅ ⇒ sp³d hybridization

48. (4)
 The molecular orbital configuration of B₂ is : (σ1s)²(σ*1s)²(σ2s)²(σ*2s)²(π2p_x)¹(π2p_y)¹
 The molecular orbital configuration of C₂⁺ is : (σ1s)²(σ*1s)²(σ2s)²(σ*2s)²(π2p_x)²(π2p_y)²(σ2p_z)¹
 The molecular orbital configuration O₂⁺ is :



The molecular orbital configuration of He_2^+ is : $(\sigma 1s)^2 (\sigma^* 1s)^1$

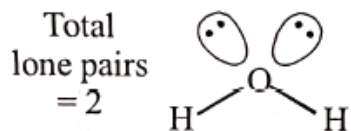
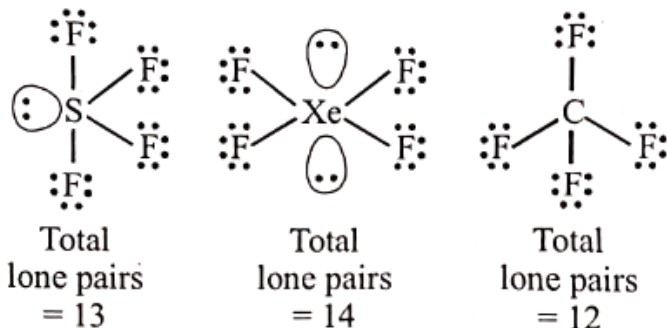
Hence, B_2 , C_2^- , O_2^+ and He_2^+ are paramagnetic species.

49. (3)

CN^- , NO^+ and O_2^{2+} all have bond order 3.

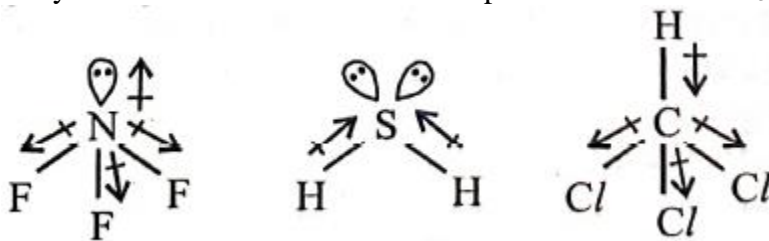
50. (1)

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 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



EXERCISE - 2 [A]

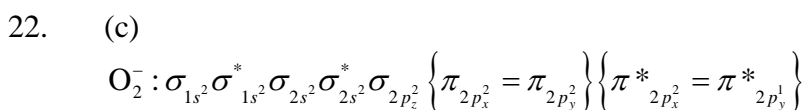
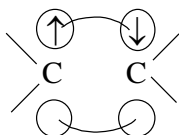
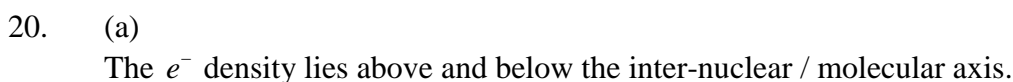
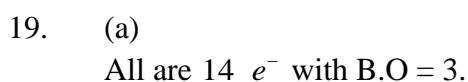
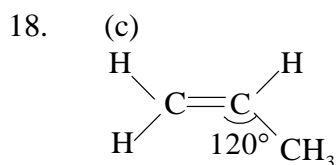
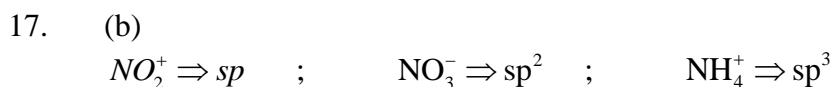
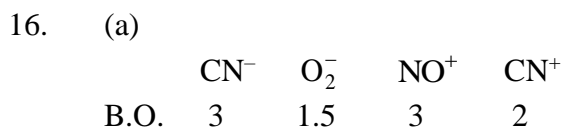
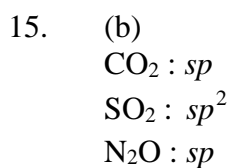
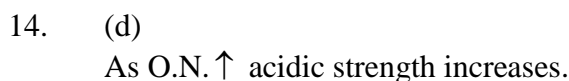
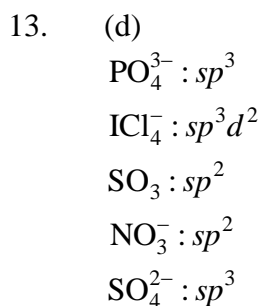
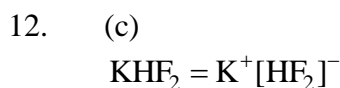
1. (b)
4 B.P and Zero L.P is SiF_4 .
2. (d)
 ICl_2^+ has 2 B.P and 2 L.P.
 sp^3 hybridization and bent shape.
3. (c)
 sp^2 hybridization and trigonal planar geometry.
4. (d)
 $\text{I}_3^- : 3, \text{XeF}_4 : 2, \text{SF}_4 : 1, \text{ClO}_3^- : 1$
5. (c)
 $\text{ClO}_4^- : 4 \text{ B.P} + \text{Zero L.P } sp^3$ hybridization.
6. (d)
 n of $\sigma_{2s} > n$ of σ_{1s}^*
7. (a)

$$\begin{array}{c} \text{Cl} - \text{H}^{\delta+} - \text{O}^{\delta-} \\ | \quad \quad \quad \diagdown \\ \text{Cl} - \text{C} - \text{C} - \text{H} \\ | \quad \quad \quad \diagup \\ \text{Cl} - \text{H}^{\delta+} - \text{O}^{\delta-} \end{array}$$
8. (b)
Bond order of $\text{O}_2^{2-} = 1$; $\text{O}_2 = 2$
 $\text{O}_2^- = 1.5$; $\text{O}_2^+ = 2.5$
Hence, Stability $\text{O}_2^+ > \text{O}_2 > \text{O}_2^- > \text{O}_2^{2-}$
9. (b)
 CH_3 free radical is formed which is sp^2 hybridized.
10. (b)
Both BN & O_2^{2-} are diamagnetic
11. (c)

	N_2	N_2^+	O_2	O_2^+	NO	NO^+
B.O.	3	2.5	2	2.5	2.5	3

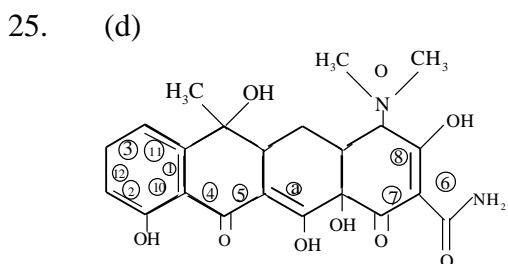
Bond dissociation energy : $\text{N}_2 > \text{N}_2^+$
 $\text{O}_2^+ > \text{O}_2$

Bond length : $\text{N}_2^+ > \text{N}_2$
 $\text{NO} > \text{NO}^+$

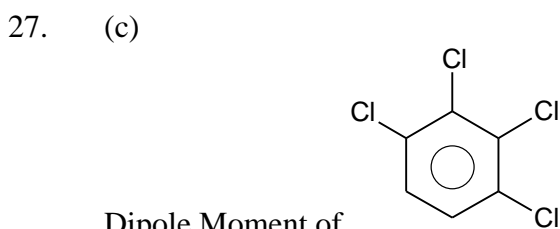
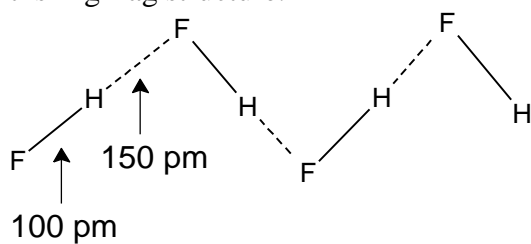


23. (b)
More the number of H – atoms, more is the H – Bonding hence, Due to higher intermolecular forces of attraction, higher is the melting point.

24. (a)
Lead Oxide PbO_2 .



26. (c)
It is Zig Zag structure.



Dipole Moment of

$$= \sqrt{3} \times 1.5 \text{ D} \quad \left(\sqrt{p_1^2 + p_2^2 + 2p_1p_2 \cos 60^\circ} \right)$$

$$= 2.6 \text{ D} \quad P_1 = P_2.$$

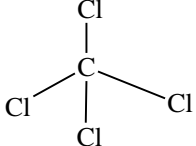
28. (d)
O – H bond is more polar than N – H bond. & SO + on H is more in O – H bond.
Also N is better donor than O.

29. (c)
 MX_3
 $\mu_{net} = 0$
Trigonal planar
 sp^2

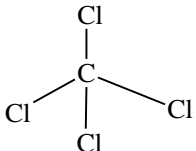
30. (a)
CO CN^-
 $14e^-$ $14e^-$

31. (a)
 CO_2
 $\text{O} = \text{C} = \text{O}$ (linear sp)

32. (b)
 CCl_4
 $\mu_{net} = 0$



33. (b)
 $\mu_{net} = 0$



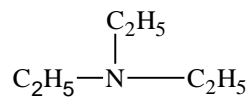
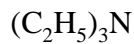
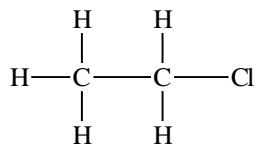
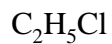
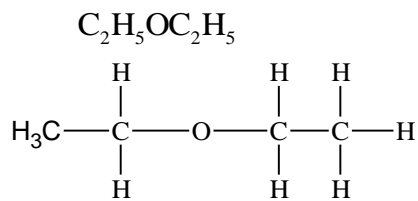
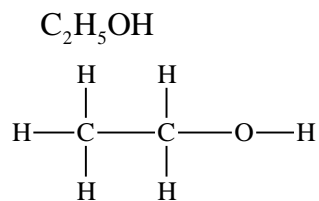
34. (d)
 $\text{H} - \text{Cl}$
 In F, N, O, H-bond are formed.

35. (a)
 $\text{NO} : \sigma_{1s^2} \sigma_{1s^2}^* \sigma_{2s^2} \sigma_{2s^2}^*$
 $\sigma_{2p_z^2} \left\{ \begin{array}{l} \pi_{2p_x^2} \\ \pi_{2p_y^2} \end{array} \right\} \left[\begin{array}{l} \pi_{2p_x^1}^* \\ \pi_{2p_y^0}^* \end{array} \right]$

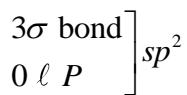
36. (c)
 For strongest H-bond, H should be attached to F with covalent bond.

37. (c)
 SO_2
 $\left. \begin{array}{l} 1l.p \\ 2\sigma \text{ bond} \end{array} \right\} sp^2$

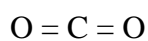
38. (a)



39. (c)

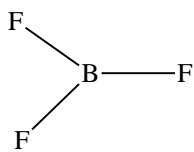


40. (a)



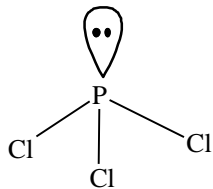
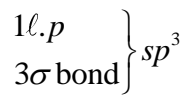
sp

41. (b)

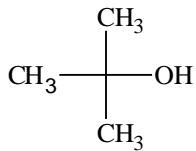


$$\mu_{net} = 0$$

42. (a)



43. (c)
 $(\text{CH}_3)_3\text{COH}$



4 σ bond

0 *l.p*

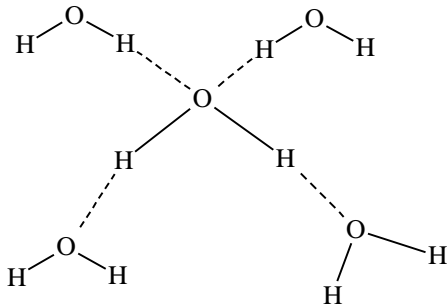
sp^3

44. (a)



2 σ bond }
 2 *l.p* } sp^3

45. (b)



4 hydrogen bond per molecule

46. (c)

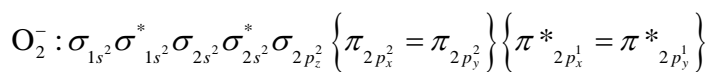


Oxide of non-metal (covalent)

Ionic $\text{MnO} > \text{Mn}_2\text{O}_7$

+2 +7

47. (d)



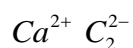
No. of paired O = 14

48. (c)



S: trigonal pyramidal S: trigonal planar

49. (b)



B.O. = 3

$\text{C} \equiv \text{C}$ and 1σ and 2π

50. (c)
- | | | | |
|--------------------------|---------------------------------|---------------------------------|-----------------|
| KO_2 | AlO_2^- | BaO_2 | NO_2^+ |
| K^+O_2^- | $\text{Al}^{3+}\text{O}_2^{4-}$ | $\text{Ba}^{2+}\text{O}_2^{2-}$ | Dia |
| Dia para | Dia dia | Dia Dia | |
| Para | Dia | dia | |

51. (a)
- H_2CO_3
-
- sp^2
- $\mu_{net} \neq 0$

52. (c)
- $\text{H}-\text{O}-\text{O}-\text{H}$
- ↑ ↑
- Non polar
- Polar

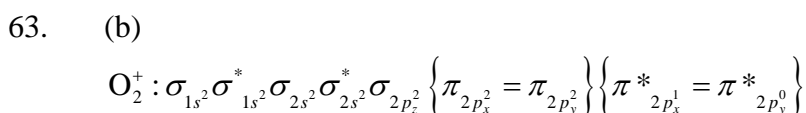
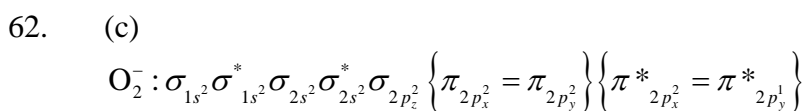
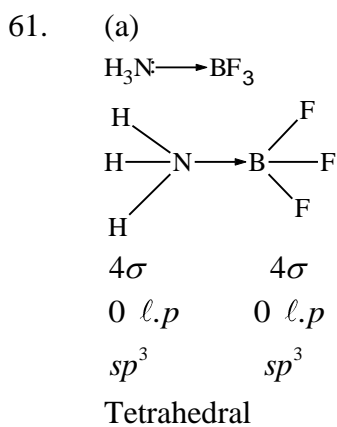
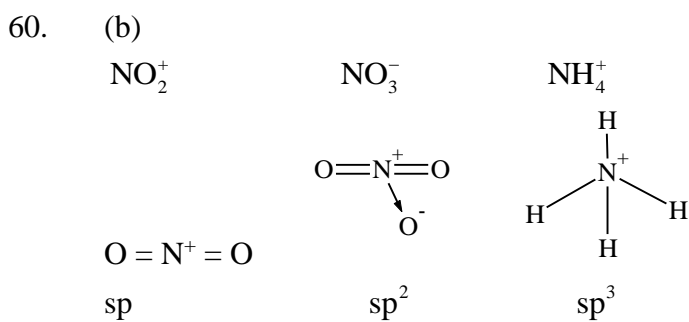
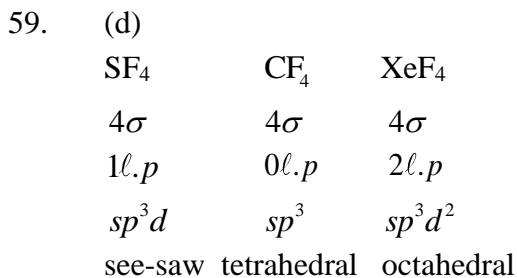
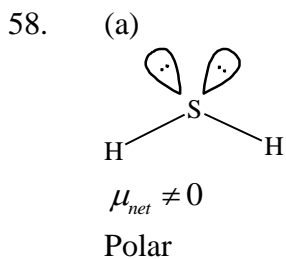
53. (d)
- | | |
|--------------|----------------------|
| O_2 | H_2O |
| Non-polar | polar |

54. (b)
- SO_2
- 1 lone pair
2 σ bond $\left. \vphantom{\begin{array}{l} 1 \text{ lone pair} \\ 2 \sigma \text{ bond} \end{array}} \right\} \text{sp}^2$

55. (a)
-
- Or

56. (c)
- LiCl
- Li^+ small size polarisation \uparrow Covalent character \uparrow

57. (d)
- | | | | |
|-------------|--|----------------------|---------------|
| | CO | CO_3^{2-} | CO_2 |
| Bond order | 3 | $\frac{4}{3} = 1.33$ | 2 |
| Bond length | $\text{CO}_3^{2-} > \text{CO}_2 > \text{CO}$ | | |



$$\text{Bond order} = \frac{6-1}{2} = 2.5$$

Paramagnetic.

EXERCISE - 2 [B]

One or More Than One Option Correct

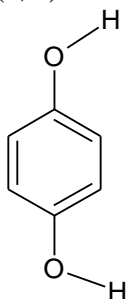
1. (a, b, c)

O_2^{2-} (Na_2O_2)	:	Diamagnetic
O_3	:	Diamagnetic
N_2O	:	Diamagnetic
O_2^- (KO_2)	:	Paramagnetic

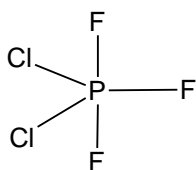
2. (b, c, d)

	CO	NO^-	NO^+	CN^-	N_2
B.O.	3	2	3	3	3

3. (a, c)

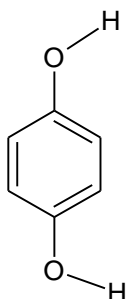


Polar

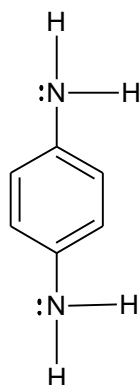


Polar

4. (b, c)



Polar

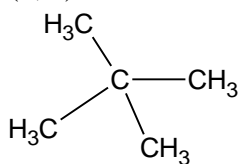


Polar

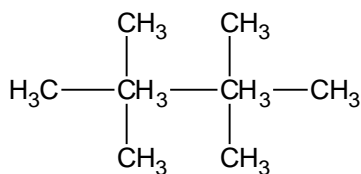
5. (a, b, c)

ICl_2^-	ClF_3	XeO_2F_2	XeF_4
sp^3d	sp^3d	sp^3d	sp^3d^2

6. (b, c)



Non-Polar



Non-Polar

7. (b, c)

	N_2	N_2^+	N_2^-	N_2^{2-}
Bond order :	3	2.5	2.5	2

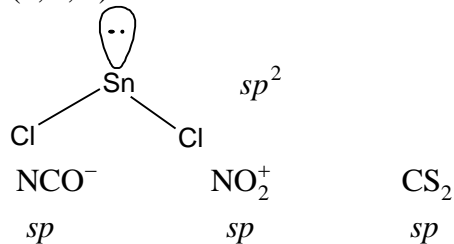
8. (a, b)

O_2	N_2^{2-}	N_2	O_2^{2-}
Paramagnetic	Paramagnetic	Diamagnetic	Diamagnetic

9. (a, b, c, d)

All have lone pair of electrons. Therefore geometry is distorted.

10. (b, c, d)



11. (a, d)

sp^3 tetrahedral

sp^3d trigonal bipyramidal

12. (a, c)

Bond strength : $\sigma > \pi$
 $\equiv > =$

13. (a, c, d)

$HgCl_2$	ClF_3	ICl_4^-
Linear	T-shaped	square planar

14. (a, b)

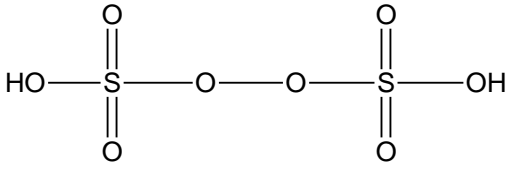
PCl_5	PCl_4^+	PCl_6^-
sp^3d	sp^3	sp^3d^2
AlH_3	$Li^+[AlH_4]^-$	
sp^2	sp^3	
NH_3	NH_4^+	
sp^3	sp^3	
H_3PO_2	H_3PO_3	PH_3
sp^3	sp^3	sp^3

15. (a, c)

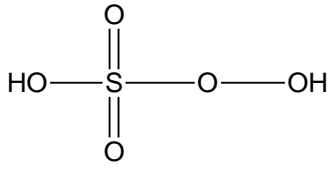
$Cl - Hg - Cl$; $H - C \equiv C - H$

16. (b, c, d)

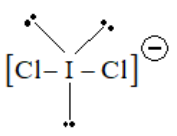
$S = C = S$; $O = \overset{+}{N} = O$

17. (a, c)
 CN^- and NO^+ has B.O = 3.
18. (a, c, d)
 B_2H_6 & $\text{Al}_2(\text{CH}_3)_6$ have 3 carbon & 2 electron bond.
 Al_2Cl_6 has 3 carbon & 4 electron bond
 Acidic strength : $\text{BCl}_3 > \text{AlCl}_3$
19. (b, d)
 H_3O^+ , NH_3 and CH_3^\ominus have Sp^3 hybridization.
20. (b)
 BF_3
 sp^2 (trigonal planar)
21. (a, c)
- 

$\text{H}_2\text{S}_2\text{O}_8$

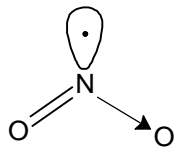


H_2SO_5
22. (b, d)
- | | | | |
|---------------|--------------|-------------------|--------------|
| CN^- | NO | O_2^{2-} | O_2 |
| Diamagnetic | Paramagnetic | Diamagnetic | Paramagnetic |
23. (a, b, c, d)
 More electronegative element in each case, i.e. O and F occupy axial position in I & II respectively.
24. (b, c, d)
 Ionic bonds are non directional bonds.
 Ionic compounds have high melting point. They have high solubility in polar solvent.
25. (a, b, d)
 NH_4^+ has one coordinate bond.
 Bond strength : $\sigma > \text{H}$
 Polarity : $\text{HF} > \text{HCl}$
26. (a, b, c)
 HF and H_2O have abnormally high B.pt due to H – Bonds
27. (a, c, d)
- $\text{Cl} - \text{Hg} - \text{Cl}$;

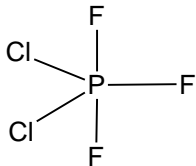


; $\text{S} = \text{C} = \text{S}$

28. (b, d)



Polar

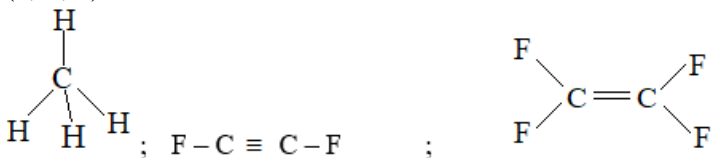


Polar

29. (d)

Hydrogen can form 1 bond with itself or with other elements to give 1 σ bond.

30. (a, b, c)



31. (a, b, c, d)

In all intermolecular hydrogen bonding take place.

32. (b, c)



33. (a, c)

SF_6 is octahedral in shape.

It has 12 electrons in valence shell of S but Ar has 8 electrons.

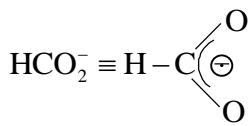
34. (a, c, d)

Only O_2 and B_2 are paramagnetic.

Bond length : $\text{NO} > \text{NO}^+$

Bond length : $\text{CO} > \text{CO}^+$

35. (c, d)

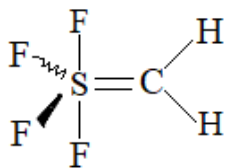


B.O of HCO_2^- (1.5) $>$ B.O of CO_3^{2-} (1.33)

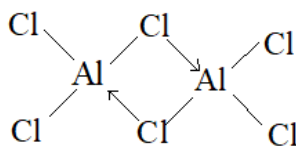
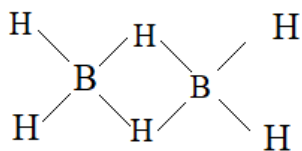
[formate ion] [Carbonate ion]

\therefore Bond length of $\text{HCO}_2^- <$ Bond length of CO_3^{2-}

36. (a, c)



37. (a, b)



Comprehension Type

Passage-I

1. (d)

Both F – atoms are present in axial positions.

2. (a)

ClF_3 has Sp^3d hybridization ; 3 B. P + 2 L.P

3. (d)

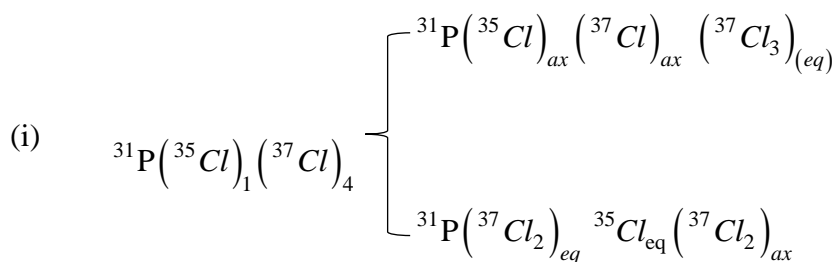
PI_5 and $[\text{PBr}_6]^-$ do not exist due to steric factors. In PH_5 , Because of its low electronegative H-atom cannot effect the contraction of dz^2 orbital to form Sp^3 hybridized P-atom.

4. (b)

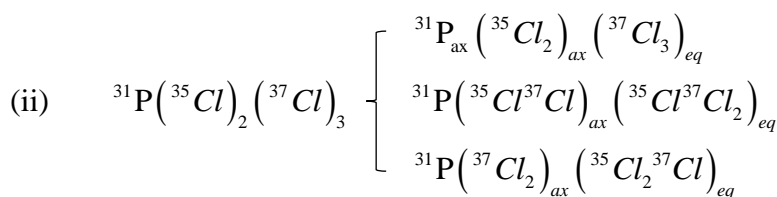
$[\text{PBr}_6]^-$ does not exist since P cannot accommodate 6 Br – atoms around it due to steric hindrance.

5. (d)

$${}^{30}\text{P}({}^{35}\text{Cl})_5 + {}^{30}\text{P}({}^{37}\text{Cl})_5 + {}^{31}\text{P}({}^{35}\text{Cl})_5 + {}^{31}\text{P}({}^{37}\text{Cl})_5 = 4$$



= 2 forms



= 3 forms

(iii) ${}^{31}\text{P}({}^{35}\text{Cl})_3({}^{37}\text{Cl})_2 = 3$ forms ; same as (ii)

(iv) ${}^{31}\text{P}({}^{35}\text{Cl})_4({}^{37}\text{Cl}) = 2$ forms ; same as (ii)

Similarly with ${}^{30}\text{P}$ – atom as central Atom, 10 forms will be there. Total = 24.

6. (c)

SF_4 has 4 G.P + 1 L. P ; Hence (C).

7. (c)

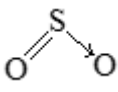
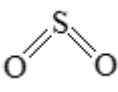
IF_2^- (sp^3d) has 2B.P + 3 L.P

8. (c)
In Sp^3d hybridized, dz^2 orbital is utilized.

Passage-II

1. (a)

i) $I_3^+ : 2B.P + 2L.P, Sp^3$ hybridization

ii) SO_2 can be represented as  and . It is considered that S being in 3rd period can utilize its d-orbitals to form $p\pi - d\pi$ bonds with O atoms.

iii) $XeF_2 : 3 L.P$ and $2 B.P \Rightarrow Sp^3d$ hybridization.

$CO_2 : Sp$ hybridization $\Rightarrow 2 B.P$

iv) $SF_4 : Sp^3d$ hybridization $\Rightarrow 4 B.P + 1 L.P$

$ICl_3 : Sp^3d$ hybridization $\Rightarrow 3 B.P + 2 L.P$

2. (c)

$I_3^- : L.P = 3; \quad B.P = 2 \quad \text{Ratio} \Rightarrow 1.5$

$XeF_4 : L.P = 2; \quad B.P = 4 \quad \text{Ratio} \Rightarrow 0.5$

3. (a)

$NO_2^+ \Rightarrow \text{linear} \Rightarrow 180^\circ$

NO_2 (odd e^- system) $\Rightarrow \text{bent} \Rightarrow 134^\circ$

NO_2 (1 L.P) $\Rightarrow \text{bent} \Rightarrow 115^\circ$

4. (c)

Bond angle $PF_3 < PCl_3 < PBr_3 < PI_3$.

Due to $d\pi - p\pi$ bonds in PF_3 , double bond character develops. Increase in B.P - B.P repulsions lead to increase in Bond Angle.

5. (d)

B.O of $O_2^+ = 2.5, O_2 = 1.5, O_2 = 2, O_2^{2-} = 3, O_2^{2-} = 1$

Bond length $O_2^{2-} > O_2^- > O_2 > O_2^+ > O_2^{2+}$

6. (d)

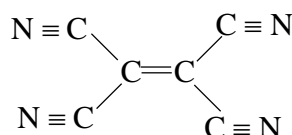
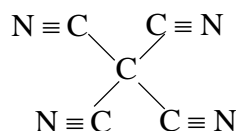
B.O of $N_2^+ = 2.5$

B.O of $N_2^- = 2.5; e^-s$ occupy non-bonding Molecular orbital's.

\therefore bond strength of $N_2^- <$ Bond strength of N_2^+

7. (c)

i)



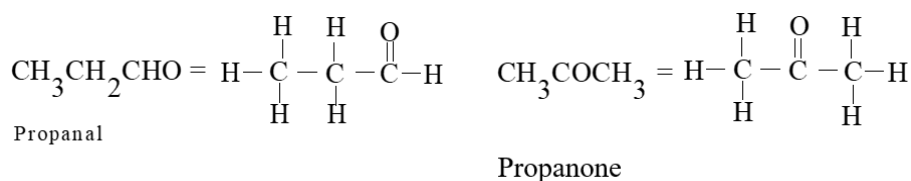
$$\sigma \text{ bonds} = 8$$

$$\pi \text{ bonds} = 8$$

$$\sigma \text{ bonds} = 9$$

$$\pi \text{ bonds} = 9$$

ii)



iii) XeF_4 is square planar in shape.

SF_4 has See – saw Shape.

iv) Due to smaller orbital size, effective overlapping take place in Cl – Cl bond.

Passage-III

1. (d)
Hydrogen bond is very weak compared to covalent bond.

2. (a)
In HF_2^- H bonding exists between HF and F^- .

3. (b)
H attached to more E.N. element with covalent bond forms stronger H-bond.

Passage-IV

1. (a)
 BeCl_2 has maximum covalent character because of smallest size of cation (Be^{+2})

2. (d)
 I^- is the largest Anion.
It has maximum polar ability.

3. (c)
Highest positive charge and smallest size of cation (Al^{+3}), hence maximum polarization and covalent character.

4. (b)
 LiCl has maximum covalent character, hence it will be most soluble in non – polar solvent ether.

5. (d)
 CaI_2 has maximum covalent character hence least M.pt...

Passage-V

1. (a, b)
These are planar molecules. XeF_4 has square planar shape.
 XeF_4 has trigonal planar shape.

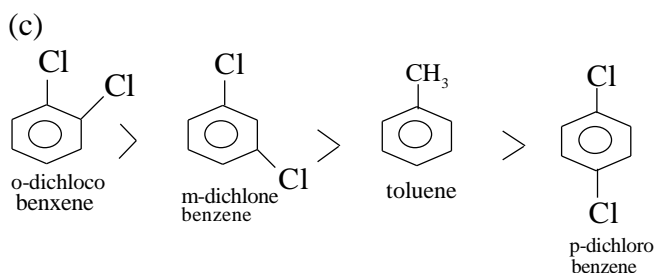
2. (c)
 $p = q \times l$

$$q = \frac{p}{l} = \frac{1.2 \times 10^{-18} \text{ esu cm}}{10^{-8} \text{ cm}}$$

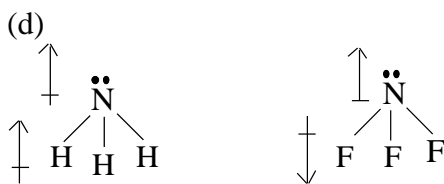
$$q = 1.2 \times 10^{-10} \text{ esu}$$

$$\text{Fraction of charge} = \frac{1.2 \times 10^{-10}}{4.8 \times 10^{-10}} = \frac{1}{4} = 0.25.$$

3.



4.



5.

(b)

$$\% \text{ Ionic character} = \frac{1.5 \times 10^{-29} \text{ C.m}}{1.6 \times 10^{-19} \text{ C} \times 150 \times 10^{-12} \text{ m}} \times 100 = 62.5\%$$

Passage-VI

1. (b)

$$\text{Bond order} = \frac{6}{4} = 1.5$$

2. (a)

	O_2	O_2^-	O_2^{2-}	O_3
Bond order :	2	1.5	1	1.5
Bond strength :	$\text{O}_2 > \text{O} = \text{O}_3 > \text{O}_2^{2-}$			

3. (a)

	$\text{CH}_3 - \text{CH}_3$	Benzene
Bond order :	1	1.5
Bond length of C-C :	$\text{CH}_3 - \text{CH}_3 > \text{benzene}$	

Fill in the Blanks

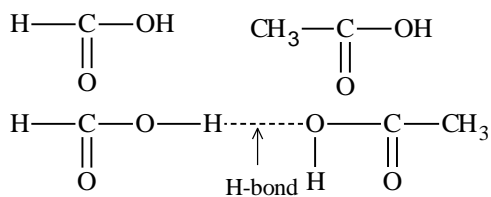
1. CO_2

Bond angle



$$180^\circ \quad 109^\circ 28' \quad 104.5^\circ$$

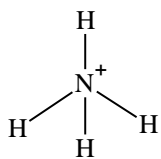
2. HCOOH and CH₃COOH



3. 2

N \equiv N 1 σ and 2 π bond

4. sp³



4 σ and 0 *l.p*

sp³

5. planar

CH₃⁺

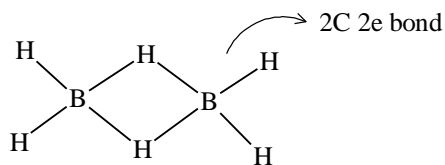
0 lone pair
3 σ bond } 3 hybrid orbitals

sp²

G and S : trigonal planar

6. three centred two electrons bond or banana bond

B₂H₆



banana bond (3c 2e bond)

7. Increases, decreases

	N_2	N_2^+	O_2	O_2^+
Bond order	$\frac{6-0}{2} = 3$	$\frac{5-0}{2} = 2.5$	$\frac{6-2}{2} = 2$	$\frac{6-1}{2} = 2.5$

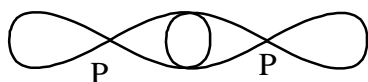
B.I. $\text{N}_2^+ > \text{N}_2$ and $\text{O}_2 > \text{O}_2^+$

8. N_2O , I_3^-
 N_2O SO_2 I_3^+ I_3^-
 $N \equiv N \rightarrow O$
 2σ bond $2l.p$ $3l.p.$

$0l.p.$ $1l.p$ 2σ bond 2σ bond
 sp sp^2 sp^3 sp^3d
linear V. bent linear

True / False

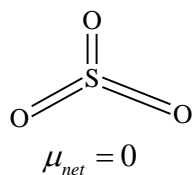
1. True



σ bond

2. False
 CO_2 and SO_3

$O = C = O$
 $\mu_{net} = 0$



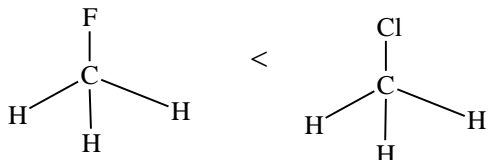
3. True
 $SnCl_2$
 2σ bond } sp^2
 $1l.p$ }
S : v- bent

4. False
 C_6H_6
 sp^2

5. False
 sp^3
% s character = $\frac{1}{4} \times 100 = 25$
% p character = $\frac{3}{4} \times 100 = 75$

6. False
 CO_2 and SO_3
 $\mu_{net} = 0$ $\mu_{net} = 0$

7. False



$$\mu = q \times d$$

Matrix Match

1. A \rightarrow q, r ; B \rightarrow q, s ; C \rightarrow p, r ; D \rightarrow q, r
 B_2^{2-} C_2^{2-} O_2 N_2^{2+}
 Bond order : 2 3 2 2
 Diamagnetic Diamagnetic Paramagnetic Diamagnetic

2. A \rightarrow s ; B \rightarrow q ; C \rightarrow p, s ; D \rightarrow p, r, s
 XeF_2 BH_4^- $XeOF_4$ PF_3Cl_2
 180° $109^\circ 28'$ $90^\circ \text{ \& } 180^\circ$ $90^\circ, 120^\circ, 180^\circ$

3. A \rightarrow p, r ; B \rightarrow q, s ; C \rightarrow q, r ; D \rightarrow p, s
 XeO_2F_2 $XeOF_4$ XeF_4 PCl_5
 sp^3d sp^3d^2 sp^3d^2 sp^3d
 (4 σ & 1 l.p.) (5 σ & 1 l.p.) (4 σ & 2 l.p.) (5 σ & 0 l.p.)

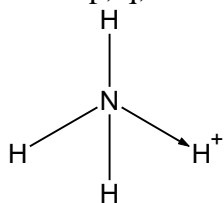
4. A \rightarrow r ; B \rightarrow p, q, s ; C \rightarrow q, s ; D \rightarrow p, q, s
 CO_3^{2-} SO_4^{2-}

 $3\sigma \text{ \& } 1\pi$ $4\sigma \text{ \& } 2\pi$
 $C_2O_4^{2-}$ MnO_4^{2-}

 $2\pi \text{ \& } 5\sigma$ $4\sigma \text{ \& } 2\pi$

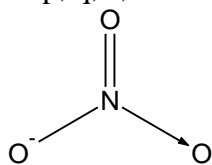
5. A \rightarrow p, q, r, s ; B \rightarrow p, r, s ; C \rightarrow q ; D \rightarrow p
 CS_2 XeF_2 C_2H_2 NCO^-
 Linear Linear Linear Linear
 sp sp^3d sp sp

6. A \rightarrow p, q; B \rightarrow p, q, s; C \rightarrow p, r, s; D \rightarrow r



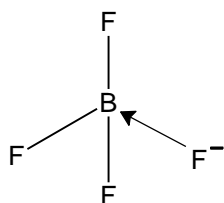
sp^3

tetrahedral



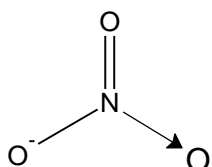
sp^2

planar



sp^3

tetrahedral



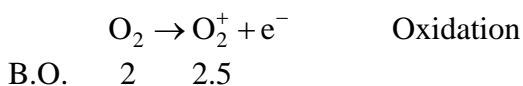
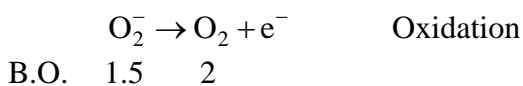
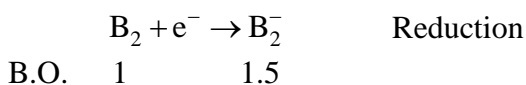
sp^2

planar

7. A \rightarrow p, r, t; B \rightarrow s, t; C \rightarrow p, q; D \rightarrow p, q, s

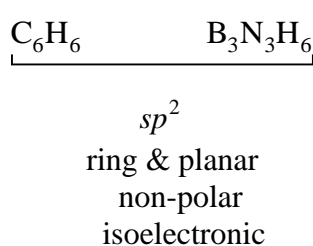
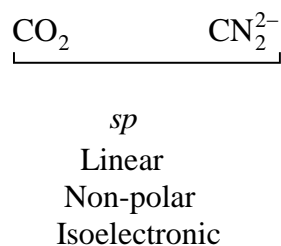
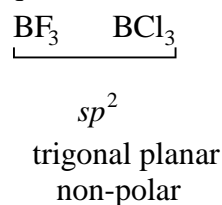
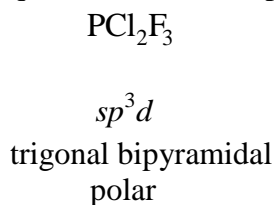
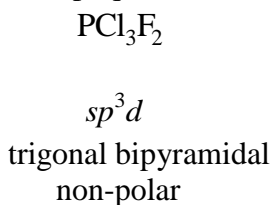
	B_2	N_2	O_2^-	O_2
Bond order :	1	3	1.5	2
	Paramagnetic	Diamagnetic	Paramagnetic	Paramagnetic

In B_2 & N_2 , mixing of s & p orbitals take place.



As bond order \uparrow stability \uparrow

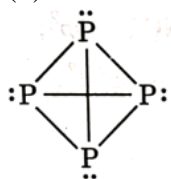
8. A \rightarrow p, q; B \rightarrow p, q, r; C \rightarrow p, q, r, s; D \rightarrow p, q, r, s;



9. A - r - w; B - q - u, C - p - v
- sp^3d : d_{z^2} & trigonal bipyramidal
- sp^3d^2 : $d_{x^2-y^2}, d_{z^2}$ & octahedral
- sp^3d^3 : $d_{xy}, d_{x^2-y^2}, d_{z^2}$ & pentagonal bipyramidal

EXERCISE - 2 [C]

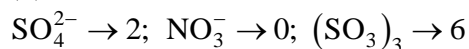
1. (4)



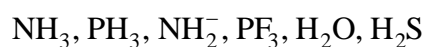
2. (2)



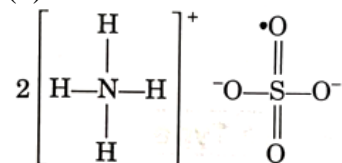
3. (8)



4. (6)

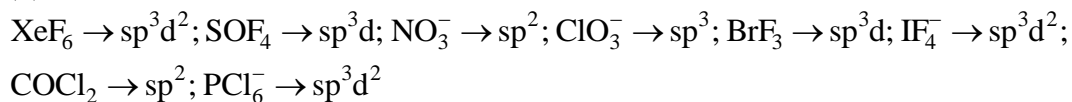


5. (2)

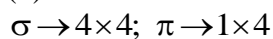


$$p\pi-d\pi = 2$$

6. (5)



7. (4)



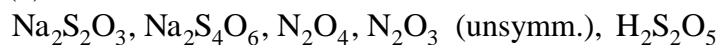
8. (3)



9. (0)

None have two $p\pi-p\pi$ bonds.

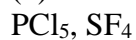
10. (5)



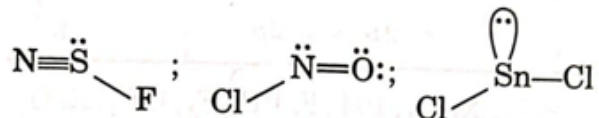
11. (2)

$s-p$ mixing is not possible in : $\text{O}_2^{2+}, \text{O}_2^+$

12. (2)



13. (9)
 $:\text{CF}_2, :\text{SiCl}_2, \text{SnCl}_2, \text{NOCl}, \text{NO}_2^-, \text{O}_3, \text{SO}_2, \text{NSF}, \text{OF}_2.$



14. (9)
 $\text{NO}_3^-, \text{CO}_3^{2-}, \text{F}_2, \text{Cl}_2, \text{Br}_2, \text{O}_2^{2-}, \text{Li}_2^+, \text{He}_2^+, \text{O}_2^-$ have B.O. < 2

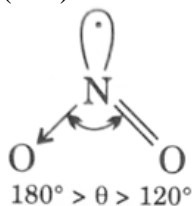
15. (3524)

3	5	2	4
P	Q	R	S

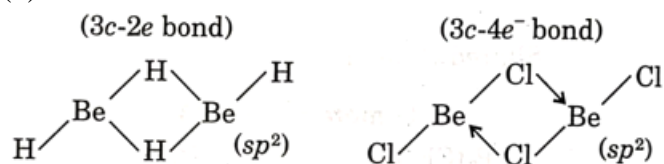
16. (3)
 $\text{B}_2, \text{N}_2, \text{C}_2$

17. (8)
 $K = 5, 3, 0$ (possible values under given condition)

18. (134)



19. (3)



20. (3)
 $\text{H}_2\text{S}, \text{H}_2\text{Se}, \text{PH}_3$

21. (1)
 BF_3 only

22. (1)
 BF_3 only

23. (1)
 $\text{N}(\text{SiH}_3)_3$ only

24. (4)
 $\text{Si}_2\text{H}_6, \text{C}_2\text{H}_6, \text{Si}_2\text{Cl}_6, \text{Al}_2\text{Cl}_6$

Only One Option Correct

1. (a)
Molecular electronic configuration of
CO: $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_y^2 = \pi 2p_z^2, \sigma 2p_x^2$
Therefore, bond order $= \frac{N_b - N_a}{2} = \frac{10 - 4}{2} = 3$
NO⁻: $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_y^2 = \pi 2p_x^2, \{\pi^* 2p_y^1 = \pi^* 2p_x^1\}$
Bond order $= \frac{10 - 6}{2} = 2$
 \therefore NO⁻ has different bond order from that in CO.
2. (d)
(a) In Na₂O₂, we have O₂²⁻ ion. Number of valence electrons of the two oxygen in O₂²⁻ ion
 $= 8 \times 2 + 2 = 18$
Which are present as follows :
 $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_y^2 = \pi 2p_x^2, \pi^* 2p_y^2 = \pi^* 2p_x^2$
 \therefore Number of unpaired electrons = 0, hence, O₂²⁻ is diamagnetic.
(b) No. of valence electrons of all atoms in O₃ = $6 \times 3 = 18$.
Thus, it also does not have any unpaired electron,
Hence, it is diatomic.
(c) No. of valence electron of all atom in N₂O = $2 \times 5 + 6 = 16$.
Hence, here also all electrons are paired.
So, it is diatomic.
(d) In KO₂, were O₂⁻ no. of valence electron of all atoms in O₂⁻ = $2 \times 6 + 1 = 13$,
It has one unpaired electron, hence it is paramagnetic.
3. (a)
If Hund's rule is violated then pairing of e⁻ in the same molecular orbital (of same energy) can take place. Complying with Hund's rule: M.O. configuration: $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_x^1, \pi 2p_y^1$
Molecular orbital configuration of B₂ (10) violating Hund's rule: $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_y^2$
Bond order of B₂ = $\frac{6 - 4}{2} = 1$, B₂ will be diamagnetic.
4. (c)
C₂ = $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^1 \pi 2p_y^1$
When there is no mixing of 2s and 2p atomic orbitals, the energy of $\sigma 2p_z$ molecular orbital will be low. Thus, only C₂ will be paramagnetic.

One or More than One Option Correct

5. (a, c)

(a) The molecular orbital energy configuration of C_2^{2-} is $\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \pi_{2p_x}^2 = \pi_{2p_y}^2, \sigma_{2p_z}^2$

In the MO of C_2^{2-} , there is no unpaired electron hence, it is diamagnetic.

(b) Bond order of O_2^{2+} is 3 and O_2 is 2 therefore bond length of O_2 is greater than O_2^{2+} .

(c) The molecular orbital energy configuration of N_2^+ is $\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \pi_{2p_x}^2 = \pi_{2p_y}^2, \sigma_{2p_z}^1$

$$\text{Bond order of } N_2^+ = \frac{1}{2}(9 - 4) = 2.5$$

The molecular orbital energy configuration of N_2^- is $\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \pi_{2p_x}^2 = \pi_{2p_y}^2, \sigma_{2p_z}^2, \pi_{2p_x}^{*1}$

$$\text{Bond order of } N_2^- = \frac{1}{2}(10 - 5) = 2.5$$

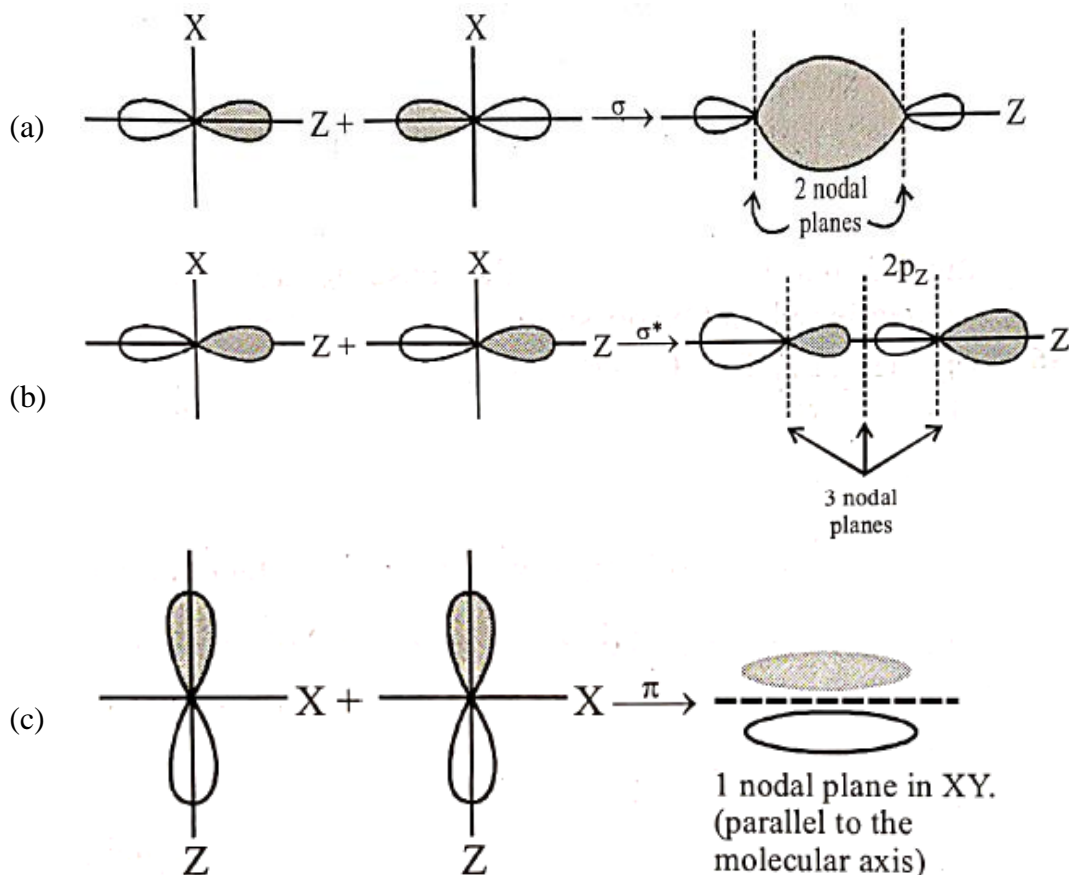
(d) He_2^+ has less energy in comparison to two isolated He atoms because some energy is released during the formation of He_2^+ from 2 He atoms.

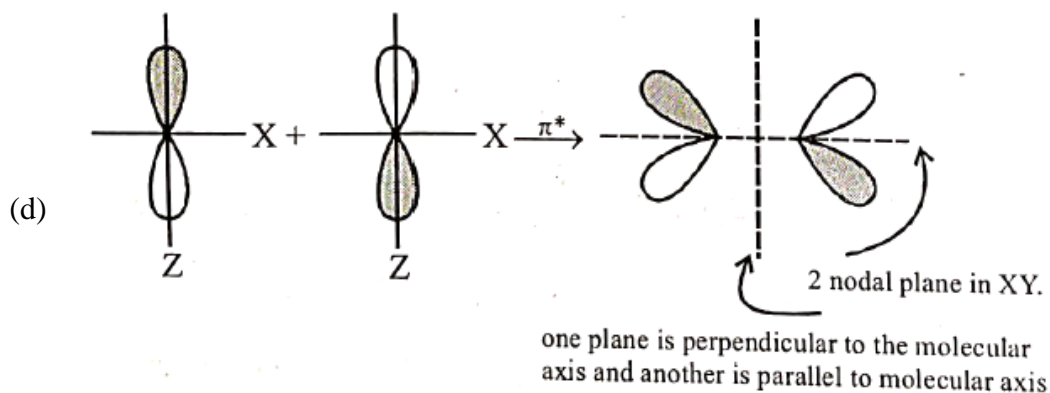
6. (b, c)

Dipole moment (μ) value of $BF_3, SF_6, BeCl_2, CO_2, BCl_3$ is zero.

7. (a, d)

Overlap of two $2p_z$ orbitals.



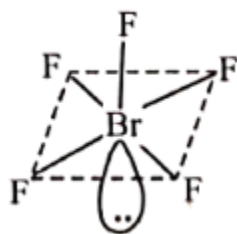


Integer Value Answer

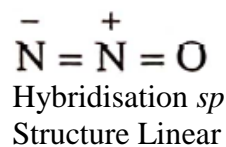
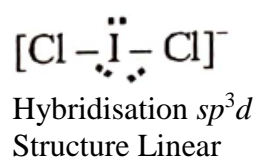
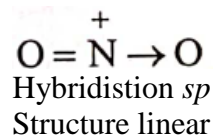
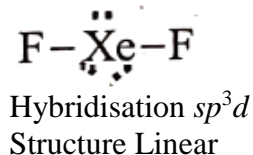
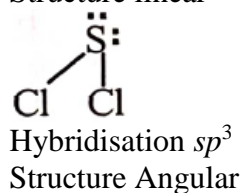
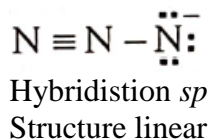
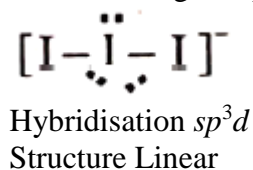
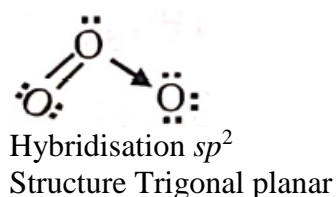
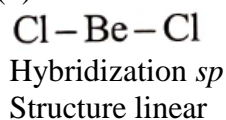
8. (4)

According to VSEPR theory, number of electron pairs around central atom (Br) are 6. Five are bond pairs and one is lone pair.

Its geometry is octahedral but due to lone pair-bond pair repulsion, the four fluorine atoms at equatorial positions are forced towards the axial fluorine atoms, thus reducing F-Br-F angle from 90° to 84.8° . The $F_{\text{eq}} - \text{Br} - F_{\text{eq}}$ angle will remain 90° .



9. (4)



Only BeCl_2 , N_3^- , N_2O and NO_2^+ are linear with sp -hybridisation.

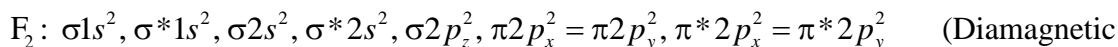
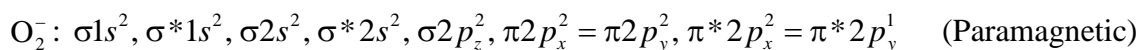
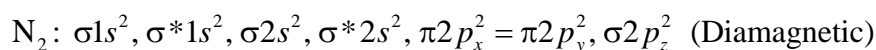
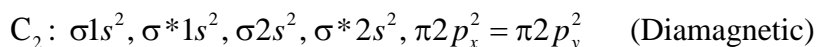
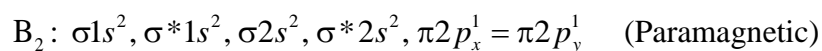
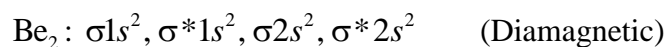
10. (6)

(H_2 , Li_2 , Be_2 , C_2 , N_2 , F_2)

H_2 : $\sigma 1s^2$ (Diamagnetic)

He_2^+ : $\sigma 1s^2, \sigma^* 1s^1$ (Paramagnetic)

Li_2 : $\sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2$ (Diamagnetic)

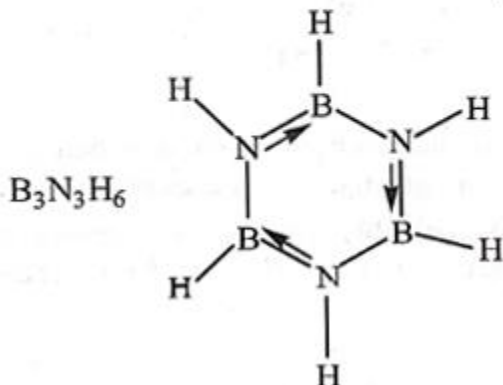
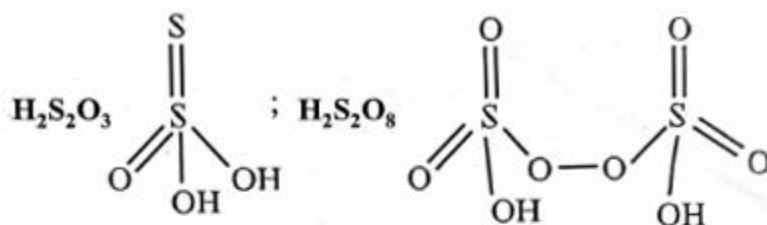
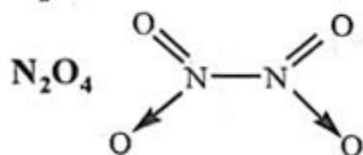
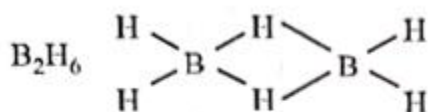


11. (6)

Species	No. of valence e ⁻ with central atom	No. of σ -bonds	No. of lone pairs
$[\text{TeBr}_6]^{2-}$	8	6	1
$[\text{BrF}_2]^+$	6	2	2
SNF_3	6	4	0
$[\text{XeF}_3]^-$	8 + 1	3	3

Sum of number of lone pairs = 1 + 2 + 0 + 3 = 6

12. (4)



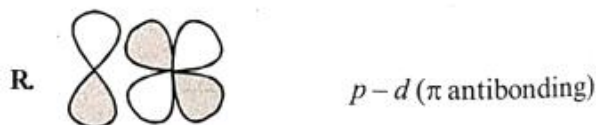
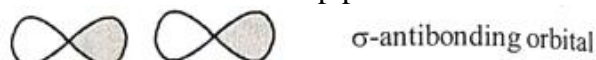
Total no. of molecules containing covalent bond between two atoms of the same kind are 4.

13. (6)
Polar molecules will be attracted/deflected near charged comb.
Polar molecules : HF, H₂O, NH₃, H₂O₂, CHCl₃, C₆H₅Cl (6-polar molecules)
Nonpolar molecules : O₂, CCl₄, C₆H₆

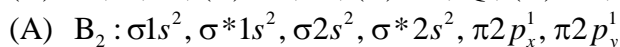
14. (6)
 $2\text{AgNO}_3(s) \xrightarrow{\Delta} 2\text{Ag}(s) + 2\text{NO}_2(g) + \text{O}_2(g)$
NO₂ has one unpaired e⁻ and O₂ has two unpaired e⁻.
Molecular orbital configuration of O₂ : $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi 2p_x^1 \pi 2p_y^1$
Total number of e⁻ present in antibonding molecular orbital = 6.

Matrix-Match

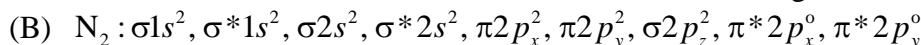
15. (c)
When two same phase overlap, it forms bonding molecular orbital otherwise antibonding molecular orbital. Also axial overlap produces σ -bond and sideways overlap produces π -bond. For example :



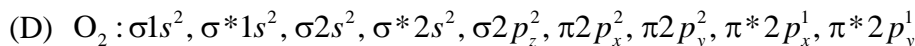
16. (A) – P, R, T ; (B) – S, T ; (C) – P, Q ; (D) – P, Q, S



Addition of e⁻ will increase the B.O., hence it will undergo reduction.



Addition or removal of e⁻ will decrease the B.O., hence it will not undergo oxidation or reduction.



Removal of e⁻ will increase the B.O., hence it will undergo oxidation easily.

(C) Similarly, removal of e⁻ will increase the B.O. for O₂⁻ also. Hence it will undergo oxidation easily.