Hybridisation

101. (a)Bond angle of hydrides decreases down the group.

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

103. (d) NH_4^+ and SO_4^{2-} both show sp^3 –hybridization and tetrahedral structure.

104. (a) It is shows sp^3d^3 –hybridization. Hence the bond angle is about 72^o.

105. (b) ClO_2

Explanation (Word-friendly format):

- CO_2 → Linear shape (O=C=O), hybridization of C is sp.
- I_3^- → Linear shape, central iodine is sp^3d hybridized (3 lone pairs cause linear geometry).
- ClO_2 → Bent (angular) shape because chlorine has one lone pair and two bonding pairs.

  Hence, **ClO_2 is not linear.**

106. (a) SCl_4

Explanation (Word-friendly format):

- SO_4^{2-} → Tetrahedral (central S is sp^3 hybridized).
- $Ni(CO)_4$ → Tetrahedral (Ni uses sp^3 hybridization).
- $NiCl_4^{2-}$ → Tetrahedral geometry due to sp^3 hybridization in high-spin complex.
- SCl_4 → Has one lone pair on sulfur → actual shape is **see-saw**, not tetrahedral.

Therefore, **SCl_4 is not tetrahedral.**

107. (a) *s*-character increases with increase in bond angle.

Hybridization	s%	Angle
sp	50	180^o
sp^2	33.3	120^o
sp^3	25	109.28^o
sp^3d^1	20	90^o and 120^o

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

109. (a) Spherically symmetrical

Explanation (Word-friendly format):
When all five d orbitals are completely filled (d^{10} configuration), the electron density is uniformly distributed around the nucleus. This results in **spherical symmetry**, similar to s orbitals.
Therefore, a **d^{10} configuration is spherically symmetrical**

110. (a) PCl_3 contain three bonding and one lone pair electron. Hence shows sp^3 –hybridization.

111. (a)Ammonia and $(BF_4)^{-1}$ shows sp^3 –hybridization.

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

113. (b) All carbon atoms of benzene consist of alternate single and double bond and show sp^2 hybridization.

114. (a) **427 kJ·mol^{-1}**

Working (Word-friendly):

Consider the reaction:

$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

Energy required to break bonds = D(H–H) + D(Cl–Cl) = 430 + 242 = 672 kJ.

Energy released on forming 2 H–Cl bonds = 2 × D(H–Cl).

The reaction enthalpy ΔH(reaction) = [energy to break] − [energy released forming bonds]

So: ΔH = 672 − 2·D(H–Cl)

Given ΔH_t for HCl = −91 kJ·mol^{-1} (since 91 kJ is released per mole of HCl formed, the reaction $H_2 + Cl_2 \rightarrow 2HCl$ releases 182 kJ), the reaction enthalpy for 1 mol H_2 + 1 mol $Cl_2 \rightarrow 2$ HCl is:

ΔH = 2 × (−91) = −182 kJ

Plug in: −182 = 672 − 2·D(H–Cl)

Rearrange: 2·D(H–Cl) = 672 + 182 = 854

D(H–Cl) = 854 / 2 = **427 kJ·mol^{-1}**

115. (d) **$PtCl_4^{2-}$**

Explanation (Word-friendly):

- **dsp² hybridization** corresponds to a **square-planar** geometry (one d, one s and two p orbitals mix).
- **$PtCl_4^{2-}$** (Pt(II), d^8) is typically **square planar** → dsp^2.
- **$NiCl_4^{2-}$** is usually **tetrahedral** (sp^3) for high-spin Ni(II).
- **SCl_4** is see-saw (derived from trigonal bipyramidal, sp^3d).
- **NH_4^+** is tetrahedral (sp^3).

116. (c) BCl_3 molecule show sp^2–hybridization and planar structure.

117. (c) BCl_3 Boron trichloride molecule show sp^2–hybridization and trigonal planar structure.

118. (b) SO_2 molecule shows sp^2–hybridization and bent structure.

119. (c) Due to multiple bonding in N_2 molecule.

120. (a) % of *s*-character in

$$\underset{(sp^3)}{CH_4} = \frac{100}{4} = 25\ , \quad \underset{(sp^2)}{C_2H_4} = \frac{100}{3} = 33\ ,$$

$$\underset{(sp)}{C_2H_2} = \frac{100}{2} = 50$$

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