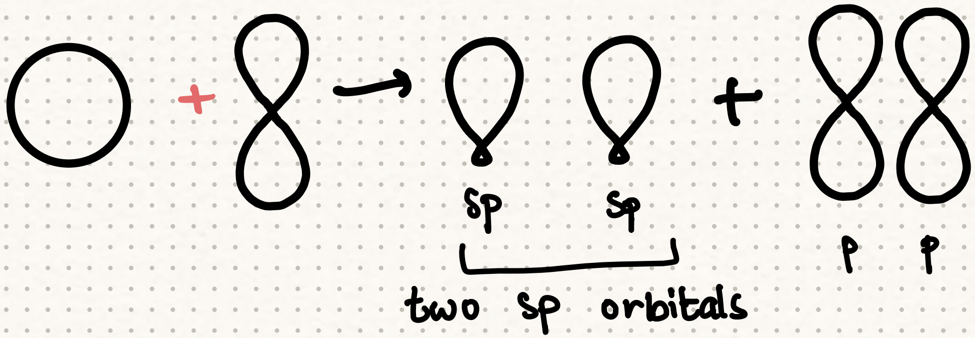
CHEMICAL BONDING

Sp - hybridization



LP + BP	Hybridisation	Geometry
2	s - p	180° linear
3	$s - p^2$	120° trigonal Planar
4	sp^3	109.5° Tetrahedral
5	sp^3d	120° 90° triagonal bi pyramidal
6	sp^3d^2	Square bi-pyramidal (octahedral)
7	sp^3d^3	72° Pentagonal bi-pyramidal <decahedral>

VSEPR theory

two types of e^- pairs

- Lone Pair (LP)
- Bond pair (BP)

Repulsion :

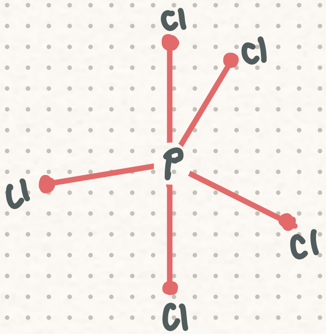
LP : LP (most) > BP : LP > BP : BP (least)

← increases

→ VSEPR theory talks about the behavior of the valence electrons and this helps us to predict the hybridisation state of the atom

PCl_5

LP = 0   BP = 5   **T = 5**



hybridisation state: **Sp^3d**

Geometry: Triangular bi-pyramidal

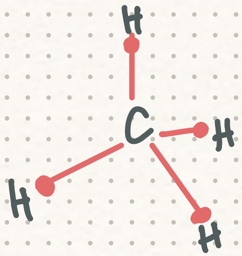
Structure ↑

CH_4

(4)

BP → 4   LP → 0 = **4**

hybridisation State: **Sp^3**



Geometry: tetrahedral

↰ Structure

$AlCl_3$

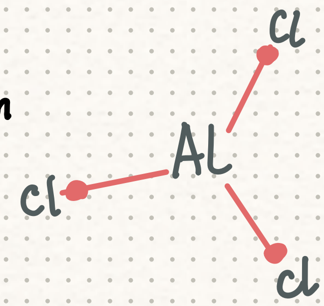
(3)

BP = 3   LP = 0

T = 3

hybridisation State: **Sp**

Geometry: trigonal planar



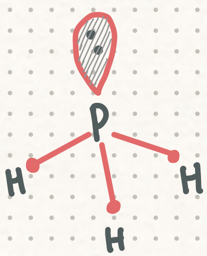
PH_3

(5)

BP = 3   LP = 1

T = **4**

Structure



hybridisation State: **Sp^3**

Geometry: **tetrahedral**

Shape: Pyramidal △

H_2S

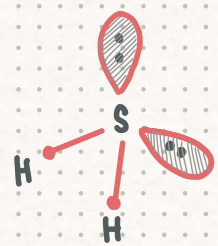
(6)

BP = 2   LP = 2

T = **4**

hybridisation State: **Sp^3**

Structure:



Geometry: tetrahedral

Shape: 109.5° bent angle//

ClF_3

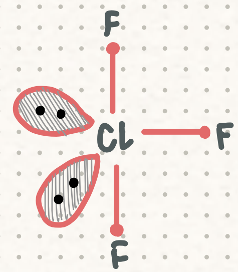
(7)

BP = 3   LP = 2

T = **5**

hybridisation State: **Sp^3d**

Structure ↓



Shape: **T-shaped**

Geometry: trigonal bi-pyramidal//

What about IONS ???

Example:

$IBr_6^{\oplus}$

{ Total Valence $e^{\ominus}$ = 7 - 1 }

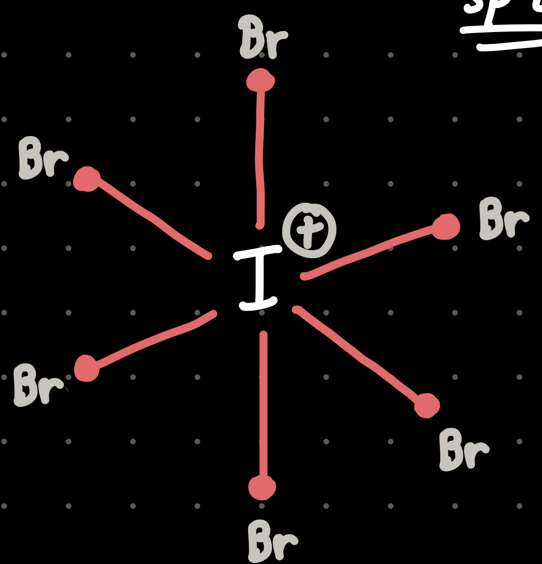
(6) → BP = 6, LP = 0

T = 6

hybridisation state: Sp^3d^2

Geometry: octahedral

Structure:



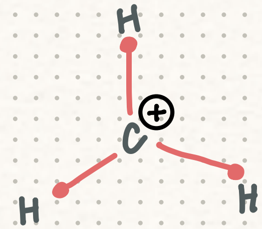
$CH_3^{\oplus}$

(3) → BP = 3, LP = 0

T = 0

hybridisation State: Sp^2

Structure:



Geometry: triangular planar

$CH_3^{\ominus}$

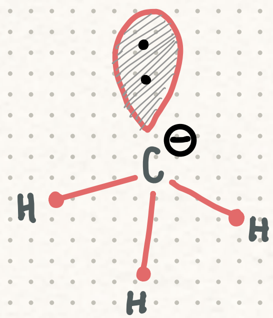
Structure

(5) → BP = 3, LP = 1

T = 4

hybridisation State: Sp^3

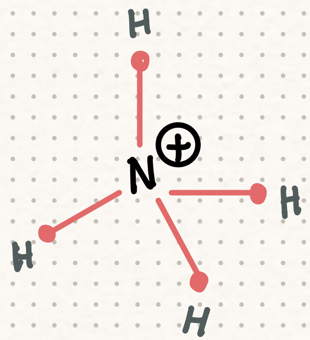
Geometry: tetrahedral



Shape: triangular pyramidal

NH_4^{+}

Structure:



4 → BP = 4, LP = 0

T = 4

hybridisation State:

sp^3

Geometry: tetrahedral

CO_2

Structure:

O = C = O

4 → BP = 2, LP = 0

T = 2

Geometry:

Linear

hybridisation state:

Sp

sp^3 hybridisation

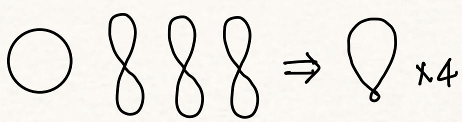
C (6) $1s^2 2s^2 2p^2$

2s: [⇅] 2p: [↑][↑][ ]

2s: [↑] 2p: [↑][↑][↑]

⇒ Involves intermixing of 1 orbital of s-subshell and three orbitals of p-subshell

i.e.



S: 25%; P_z, P_y, P_x: 75% ⇒ sp^3 hybrids <×4>

Polarity in Covalent bonds:

- Electronegativity can be defined as the ability of an atom, to attract the bonding or shared pair of electrons towards itself in a molecule.

Eg: X: Y

X > Y

electronegativity

Covalent Bond

non-polar

- Covalent bond b/w two same atoms.
- e^- are not displaced towards a specific atom
- like H-H, Cl-Cl, N≡N, O=O

Polar

- Covalent bond formed b/w atoms of different elements.
- Shared e^- pair is more inclined towards the element of high electronegativity.
- Greater the difference in the electronegativity, more will be the polarity.
- Eg. $\overset{\delta+}{H} - \overset{\delta-}{Cl}$, $\overset{\delta-}{O}$ with $\overset{\delta+}{H}$ / \ $\overset{\delta+}{H}$

Such Polar covalent bonds develop a partial ionic character.
<as a result of difference in the electronegativity>

Dipole Moment

Defined as the product of the magnitude of charges on any one of the atoms and the distance between the atoms.

Represented by ⟶ μ

expressed as : $\mu = e \times d$ [where e = charge on any one atoms, d = distance b/w the charges]

SI unit = C·m < coulomb × metre >