(a) LiCl, NaCl, KCl, RbCl, CsCl

Size of cation is the deciding factor (charge of each ion and size of Cl is same)

$$\xrightarrow[\text{size of cation}]{Li^+ < Na^+ < K^+ < Rb^+ < Cs^+}$$

$$\xleftarrow[\text{polarisation}]{}$$

$$\xleftarrow[\text{covalent}]{LiCl > NaCl > KCl > RbCl > CsCl}$$

$$\xrightarrow[\text{mp/bp}]{}$$

$$\xrightarrow[\text{conductivity}]{}$$

$$\xrightarrow[\text{solubility in } H_2O \text{ (polar solvents)}]{}$$

$$\xleftarrow[\text{solubility in non polar solvent}]{}$$

(b) NaF, NaCl, NaBr, NaI

Deciding factor is the size of the anion. Larger size of the anion, greater the polarization hence greater the covalent nature.

$$\xrightarrow[\text{size of anion}]{F^- < Cl^- < Br^- < I^-}$$

$$\xrightarrow[\text{polarisation}]{F^- < Cl^- < Br^- < I^-}$$

$$\xrightarrow[\text{covalent nature}]{NaF < NaCl < NaBr < NaI}$$

(c) $NaCl, MgCl_2, AlCl_3, SiCl_4, PCl_5$

$$\xleftarrow[\text{size of cation}]{Na^+ > Mg^{2+} > Al^{3+} > Si^{4+} > P^{5+}}$$

$$\xrightarrow[\text{charge on cation}]{}$$

> ***Note :*** Greater the charge, smaller the size then greater the polarisation
>
> $$\xrightarrow[\text{polarisation}]{Na^+ < Mg^{2+} < Al^{3+} < Si^{4+} < P^{5+}}$$
>
> $$\xrightarrow[\text{covalent nature}]{NaCl < MgCl_2 < AlCl_3 < SiCl_4 < PCl_5}$$

(d) NaF, Na_2O, Na_3N

$\xrightarrow{F^- < O^{2-} < N^{3-}}$ size of anion

$\xrightarrow{F^- < O^{2-} < N^{3-}}$ charge on anion

$\xrightarrow{F^- < O^{2-} < N^{3-}}$ polarisation

$\xrightarrow{NaF < Na_2O < Na_3N}$ covalent nature

(e) CuCl and NaCl

$[Cu^+]$	$[Ar]\ 3d^{10}$
$[Na^+]$	$[Ne]$

Cations with 18– electron shells have greater polarising power than the 8-electron shell ions with the same charge and size. This is due to the increased electronegativity of the 18-electron shell ions as the inner electrons have poor shielding effect on the nucleus. Thus CuCl is covalent and NaCl is ionic.

SHAPES OF MOLECULES (VSEPR THEORY) Section - 5

The shapes or geometry of a molecule is quite accurately predicted by VSEPR (Valence Shell Electron Pair Repulsion) theory. According to this theory *all valence shell electron pairs surrounding the central atom arrange themselves in such a manner as to be as far away from each other as possible.* By separating the electrons from each other, the electrostatic repulsion, (that is the cause of higher energy), is minimised. As a result each molecule tends to acquire a state of lowest energy.

The basic ideas can be summarized as follows :

Valence shell pairs of electrons are arranged about the central atoms so that repulsion among them is minimised, or so that there is maximum separation among the regions of high electron density (bond pairs) about the atom. For instance, two regions of high electron density would be most stable on opposite side of the central atom (*Linear arrangement*), while three regions would be most stable when they are arranged at the corners of an equilateral triangle *Trigonal planar arrangement)*. The resulting arrangement of these regions is referred to as the electronic geometry of the central atom.

Number of Bond Pairs	Electronic Geometry	Bond Angles
2	Linear	180°
3	Trigonal planar	120°
4	Tetrahedral	$109^{\circ}\,28'$
5	Trigonal Bi-pyramidal	90°, 120° , 180°
6	Octahedral	90°, 180°

As mentioned earlier, there are two distinct types of electron pairs around the central atom : Bonding electron pairs (bp) and Nonbonding electron pairs or lone pairs (–*l*p). The strength of repulsion between the electron pairs varies as :

–*l*p-–*l*p > –*l*p-bp > bp-bp

This is very significant while deciding the final geometry of a molecule. We will divide our study of shapes in two categories :

- *The molecules containing only bond–pairs of electrons and*
- *The molecules containing both bond–pairs and lone (non-bonding) pairs.*

Note : Before we go into the details of shapes, first let us discuss the *Theory of Covalent Bonding*.

THEORIES OF COVALENT BONDING Section - 6

Orbital theory of Covalent Bonding :

As discussed earlier that covalent bonding is the result of sharing of electrons between atoms. The main requirement for sharing of electrons is: *only a single electron must be present in an orbital of the outer shell.* Such an orbital (i.e. containing an odd electron) will combine with another orbital (containing an odd electron *but of opposite spin*) of same kind or of another kind of atom to form covalent bond. This combination of orbitals is known as *overlapping of orbitals.* So according to atomic orbital theory, *the covalent bond is formed by the overlapping of atomic orbitals having electrons with opposite spins, belonging to the external shells of two atoms of same or of different kinds*. After overlapping the electron–pair now belongs to both the orbitals of two atoms.