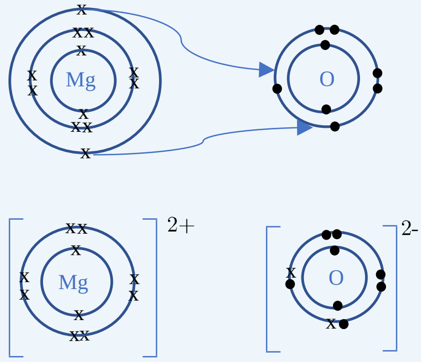
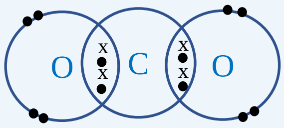
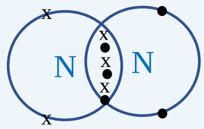
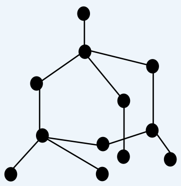
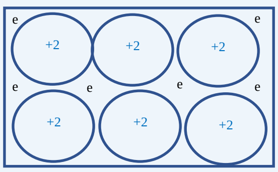
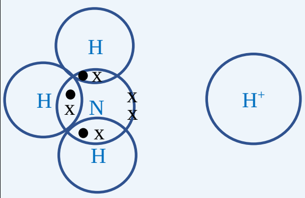
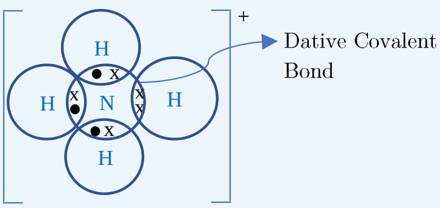
Bonding	Examples	Structure	Properties
<u>**Ionic Bonding**</u> **Electrostatic attraction** between **oppositely charged ions** Occurs between metals & non metals All about the transfer of electrons to gain a full outer shell of electrons		Giant ionic lattice structure.	High melting and boiling point due to **strong electrostatic attractions.** Conduct electricity only in the **molten & aqueous** forms due to **mobile ions** in these states. Brittle
Covalent Bonding (simple covalent) Occurs between non-metals & non-metals.	<u>**E.g.1**</u>  Carbon has 4 electrons in its outer shell. It needs 4 more electrons from 2 O atoms. Hence, each O gives 2 electrons <u>**E.g 2**</u>  Nitrogen atom has 5 electrons in its outer shell. It needs 3 more electrons from another N atom. Hence, each N provide 3 electrons to be shared.	Simple Molecular	Low melting and boiling point due to **weak intermolecular** forces. Examples of Intermolecular forces include: ▪ Van der Waal (London forces) ▪ Permanent dipole - dipole ▪ Hydrogen Bonding Do not conduct electricity as no mobile ions or electrons.
Covalent Bonding (Giant Macromolecular)	**E.g. Diamond** 	Giant Covalent Macromolecular	**Diamond:** Each C atom covalently bonded to four other C atoms. High Melting and Boiling points due to strong covalent bonds. Does not conduct electricity. --- **Graphite:** Each Carbon atom covalently bonded to 3 other C atoms. Conducts electricity due to mobile electrons Layers can slide over each other as weak intermolecular forces b/w layers.
Metallic Bonding **Electrostatic attraction** b/w positive **metal ions** and **sea of delocalised electrons.** Occurs in metals.	E.g. Magnesium 	Giant Metallic Lattice	High melting and boiling points due to **strong electrostatic attraction**. Conducts electricity due **to sea of delocalised electrons.** **Malleable & Ductile**
<u>**Dative Covalent Bonding**</u> One atom provides **both** electrons to be shared to form a bond. Also, known as **coordinate bonding.**	 		

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Electronegativity

Electronegativity is the ability of an atom to **attract the electron pair in a covalent bond towards itself**. The electronegativity is affected by:

- **Atomic Radius**: the smaller the atomic radius, the greater the electronegativity.
- **Nuclear Charge:** greater the nuclear charge, given shielding remains constant, the greater the electronegativity.
- **Shielding:** lower the electron shell shielding (no. of shells), greater the electronegativity.

Trends in Electronegativity

The Pauling scale is used to show the electronegativity of different atoms. The table below shows the electronegativities of few elements in period 2 and Group 7. Fluorine is the most electronegative species.

Electronegativity Increases: Atomic Radius decreases, Nuclear charge increases, shielding remains constant →				
C: 2.5	N: 3.0	O: 3.5	F: 4.0	↑ **Electronegativity increases** Atomic Radius decreases Shielding decreases
			Cl: 3.0	
			Br: 2.8	
			I: 2.6	

Polarity

Polarity is defined as the **uneven distribution of electrons in a covalent bond.**

Cl —ₓ•— Cl

Both Chlorine atoms have the same electronegativity. This results in an even distribution of electrons, thus Cl_2 is referred to as a **non-polar** molecule.

$^{\delta+}$H —ₓ•— Cl$^{\delta-}$

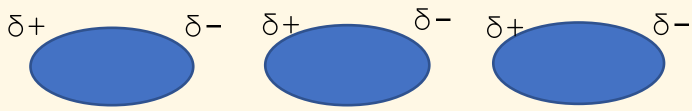
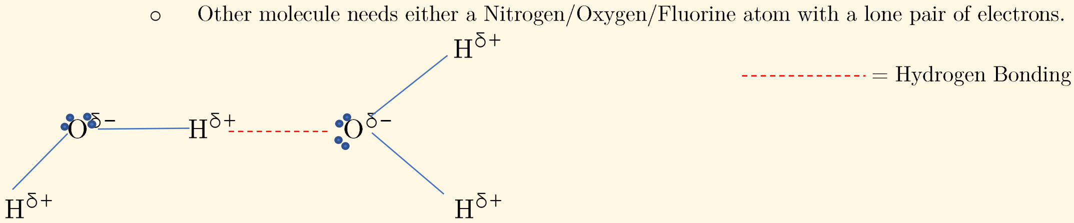
Chlorine is more electronegative than H, therefore has a greater ability to attract the electrons within the covalent bond towards itself. This results in an uneven distribution of electrons, thus HCl is referred to as a **polar** molecule.

δ^+ / δ^- = are known as **dipoles**. Represent partial positive and partial negative charges.

Intermolecular Forces

As the name suggest, intermolecular forces occur **between** molecules. Note, we tend to refer to intermolecular forces when discussing interactions between **simple covalent molecules.** The three types of intermolecular forces are discussed below:

Strength Increases ↓

Van Der Waal Forces	▪ Occurs in all molecules and in noble gases which exist as atoms. ▪ At any given instant, the electrons distribution is towards one side of the atom/molecule, hence **forming temporary dipoles**. These instantaneous/temporary dipoles **induce** dipoles on neighbouring atoms/molecules. Attraction between these temporary dipoles is known as Van der Waal forces.  ▪ Note that Van der Waal forces are **continuously created and destroyed**, as the temporary dipoles are created and destroyed. ▪ **Greater** the number of electrons, the **greater** the Van der Waal forces.
Permanent Dipole Dipole	▪ Permanent dipole dipole interactions occur between polar molecules, as they have permanent dipoles which arise from the electronegativity difference between atoms. $^{\delta+}$H—Cl$^{\delta-}$ $^{\delta+}$H—Cl$^{\delta-}$ $^{\delta+}$H—Cl$^{\delta-}$
Hydrogen Bonding	▪ For Hydrogen Bonding to occur, the below two conditions must be satisfied: ○ One molecule needs an $^{\delta+}$H attached to either a Nitrogen/Oxygen/Fluorine atom ○ Other molecule needs either a Nitrogen/Oxygen/Fluorine atom with a lone pair of electrons.  Note: In Ice the molecules are **less closely packed together** than in liquid water, hence **ice is less dense than water** i.e. floats on water. In ice the hydrogen bonds hold the water molecules in a fixed structure, where in liquid water they are continuously created and destroyed as the molecules move.

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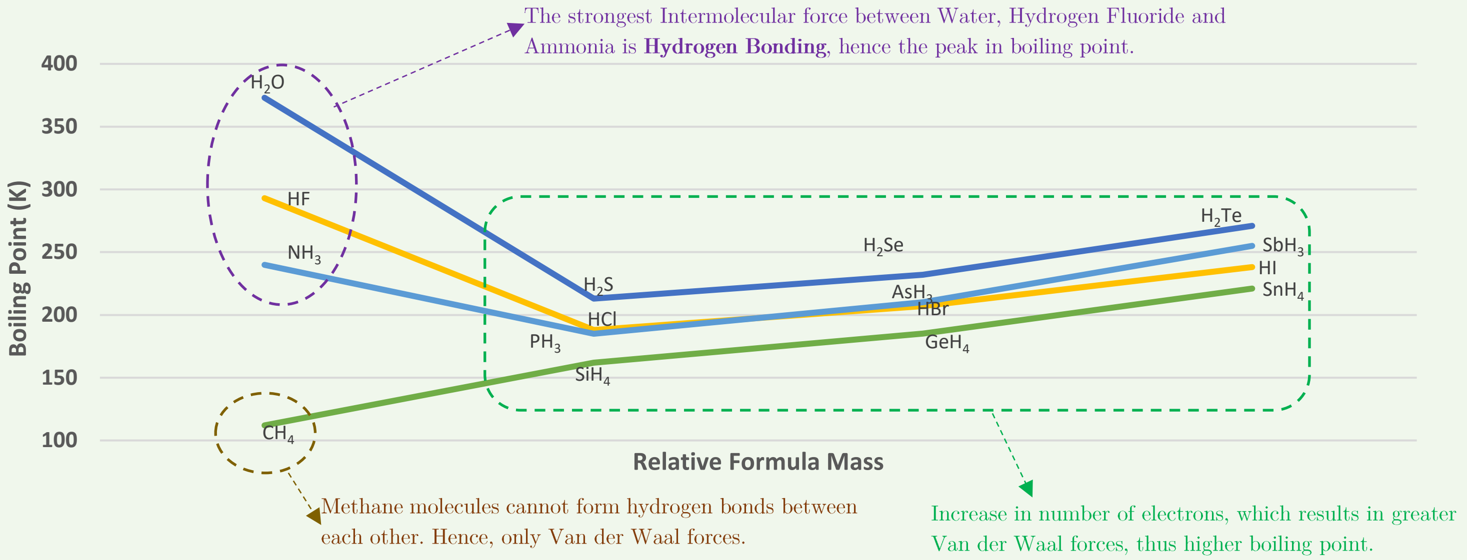
Example 1

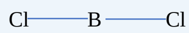
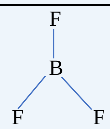
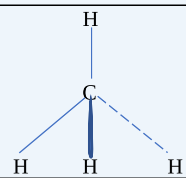
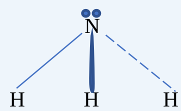
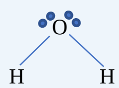
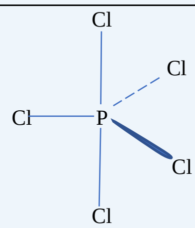
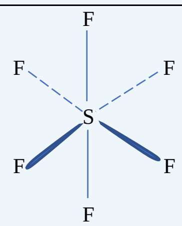
Below is a graph highlighting the boiling points of the Group 4, 5, 6 and 7 hydrides as you move down the respective groups.



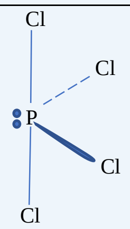
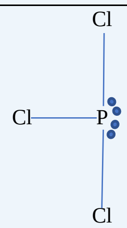
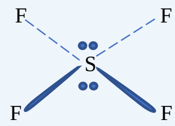
Shapes of Molecules

The shape of a molecule depends on the **number of electron pairs on the central atom**. Bonds arrange themselves in such a way to **minimize repulsion.** Note: Treat double/triple bonds as one electron pair. Also note, the below VSEPR (Valence Shell Electron Repulsion Theory):

Bonding Pair vs Bonding Pair Repulsion < Bonding Pair vs Lone Pair Repulsion < Lone Pair vs Lone pair Repulsion

Shape Name	No. of Bonding Pairs	No. of Lone Pairs	Shape	Bond Angle
Linear	2	0		180°
Trigonal Planar	3	0		120°
Tetrahedral	4	0		109.5°
Trigonal Pyramidal	3	1		107°
Bent/Non-linear	2	2		104.5°
Trigonal Bipyramidal	5	0		120°, 90°
Octahedral	6	0		90°

More Complex Shapes:

Shape Name	No. of Bonding Pairs	No. of Lone Pairs	Shape	Bond Angle
See-saw	4	1		<90° <120°
T-shaped	3	2		<90°
Square Planar	4	2		90°

To identify shape of molecule:

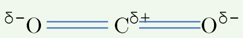
4) Draw a dot and cross diagram of the molecule.
5) Identify no of bonding pairs and lone pairs.
6) Recall shape and bond angle from the above table.

To explain shape of molecule:

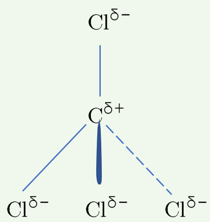
1) Mention no of electron pairs on central atom (bonding & lone pairs)
2) Explain electrons within bonds repel, thus arrange themselves to minimise repulsion.
3) If lone pairs present, refer to lone pair repulsion being greater than bonding pair repulsion.

Example 2

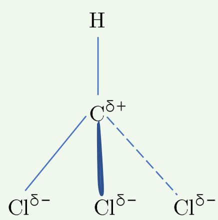
Some molecules may have polar bonds but overall be non-polar, as dipoles cancel out. This occurs when there is symmetry in the distribution of charge within the molecule.



Symmetry in distribution of charge, hence dipoles cancel out. CO_2 is a non-polar molecule.



Symmetry in distribution of charge, hence dipoles cancel out. CCl_4 is a non-polar molecule. Commonly used as non-polar solvent.



No symmetry in distribution of charge, as Hydrogen has no dipoles. Therefore, $CHCl_3$ is a polar molecule. Commonly used as polar solvent.

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