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— TUITION —

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CHEMISTRY

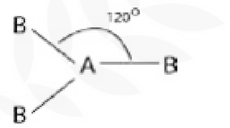
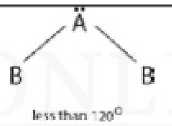
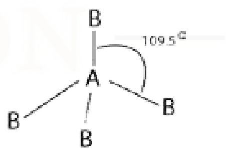
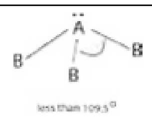
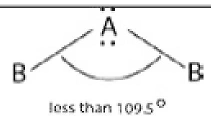
REVISION NOTES

CHEMICAL BONDING -2

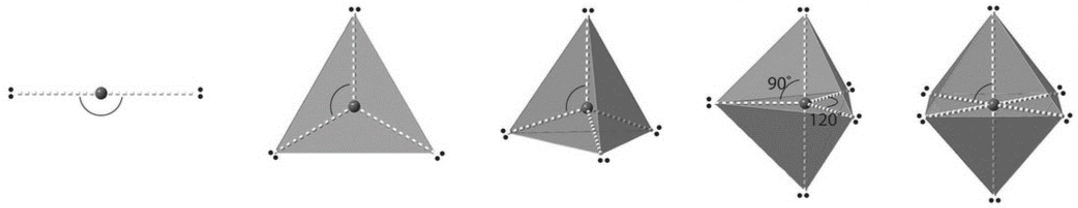
Shape of Molecules & Intermolecular Forces

- The **valence shell electron pair repulsion theory** (VSEPR) explains molecules' shape and bond angles.
- Electrons repel other electrons when close to each other.
- In a molecule, the **bonding pairs of electrons** will repel other electrons around the **central atom,** causing the molecule to adopt a shape in which these **repulsive forces** are minimized.
- Following VSEPR rules should be considered when determining the **shape** and **bond** angles of a molecule:
  - Valence shell electrons repel each other as they have the same charge.
  - Lone pair electrons repel each other more than bonded pairs.
  - Repulsion between multiple and single bonds is treated the same as repulsion between single bonds.
  - The repulsion between pairs of double bonds is greater.
  - The most stable shape is adopted to minimize the repulsion forces.

- Different types of electron pairs have different repulsive forces.
  - Lone pairs of electrons have a more concentrated electron charge cloud than bonding pairs of electrons.
  - The cloud charges are wider and closer to the central atom's nucleus.
  - Therefore, the order of repulsion is lone pair – lone pair > lone pair-bond pair > bond pair-bond pair.

Type	Electron Pairs			Arrangement of pairs	Molecular geometry	Shape	Example
	Total	Bonding	Lone				
AB_2	2	2	0	Linear	Linear	B-A-B	$BeCl_2$ $HgCl_2$
AB_3	3	3	0	Trigonal planar	Trignol planar		BH_3, BF_3 $AlCl_3$
		2	1		Bent (or angular)		$SnCl_2$, SO_2
AB_4	4	4	0	Tetrahedral	Tetrahedral		CH_4, $SiCl_4$, CCl_4, BF_4, NH_4^+, SO_4^{2-}
		3	1		Trignol pyramidal		NH_3, NF_3, PH_3
		2	2		Bent (or angular		H_2O, H_2S

Electron Groups	2	3	4	5	6
Geometry	Linear	Trigonal planar	Tetrahedral	Trigonal bipyramidal	Octahedral
Predicted Bond Angles	180°	120°	109.5°	90°, 120°	90°



Exam Question

The diagram shows bond angles in ammonia and water.

Explain why the bond angle in water is less than the bond angle in ammonia.

Draw a diagram of the ammonia molecule, clearly showing its shape.
Include any lone pairs of electrons and the value of the bond angle.

Draw a diagram of a hydrogen bond between two water molecules in ice.

Show the value of the H–O–H angle within a molecule and the value of the O–H–O angle between the two molecules.

ELECTRONEGATIVITY

Electronegativity is the relative tendency of an **atom in a covalent bond** in a molecule **to attract electrons** in a covalent bond **to itself**.

Electronegativity is measured on the **Pauling scale** (ranges from 0 to 4)

17^{th} group elements	Electro negativity
^{9}F	4.0
^{17}Cl	3.2
^{35}Br	3.0
^{53}I	2.7

Electronegativity decreasing in a group

Factors affecting electronegativity

Several factors impact electronegativity.

Across the period:
When moving across a period, the number of protons in the nucleus increases while the atomic radius decreases. This results in a greater attraction between the nucleus and the electrons within the same shell, making it more challenging to remove them.

Down the group:
When moving down a group, the distance between the nucleus and outer electrons increases, and the inner shell electrons' shielding effect also increases. This makes it easier to remove the outermost electrons, ultimately reducing the atom's electronegativity.

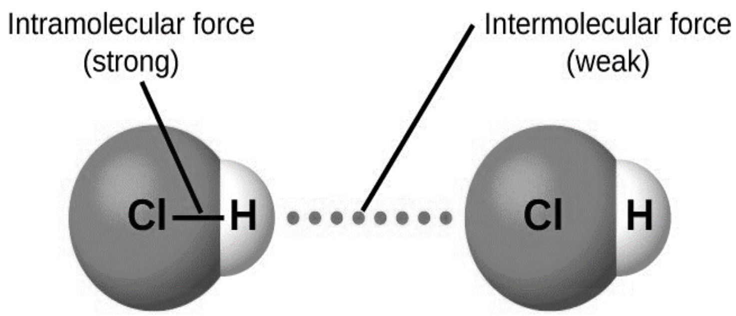
increasing electronegativity →

decreasing electronegativity ↓

group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H 2.2																	He —
Li 1.0	Be 1.6											B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne —
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar —
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	La–Lu 1.1–1.3	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.3	Au 2.5	Hg 2.0	Tl 1.6	Pb 2.3	Bi 2.0	Po 2.0	At 2.2	Rn —
Fr 0.7	Ra 0.9	Ac–Lr 1.1–1.4	Rf —	Db —	Sg —	Bh —	Hs —	Mt —	Ds —	Rg —	Cn —	Nh —	Fl —	Mc —	Lv —	Ts —	Og —

Intramolecular Forces vs. Intermolecular Forces:

Intramolecular forces: Forces within a molecule among atoms and are usually covalent bonds.
Intermolecular forces: Molecules also contain weaker intermolecular forces, which are forces between the molecules.



Types of intermolecular forces:

- **Induced dipole–dipole forces** are also called van der Waals or London dispersion forces.

- **Permanent dipole–dipole forces** are the attractive forces between two neighboring molecules with a permanent dipole.

    - **Hydrogen Bonding** is a special type of permanent dipole - permanent dipole forces.

Polar bonds & Non-Polar bonds :

- The bond is nonpolar if two atoms in a covalent bond have equal electronegativity. For example, two chlorine atoms share bonding electrons equally.

- If two atoms are involved in a covalent bond and have different electronegativities, the bond is considered polar. In such a case, the electrons involved in the bond will be drawn towards the more electronegative atom. For example, Cl has a greater electronegativity than H, and this causes the electrons to be more strongly attracted to the Cl atom. As a result, the Cl atom becomes delta-negative, and the H atom becomes delta-positive.
  As a result of this:
    - Electron distribution is asymmetric.
    - The less electronegative atom gets a partial charge of δ+ (delta positive)
    - The more electronegative atom gets a partial charge of δ- (delta negative)

The bond becomes more polar as the difference in electronegativity increases.

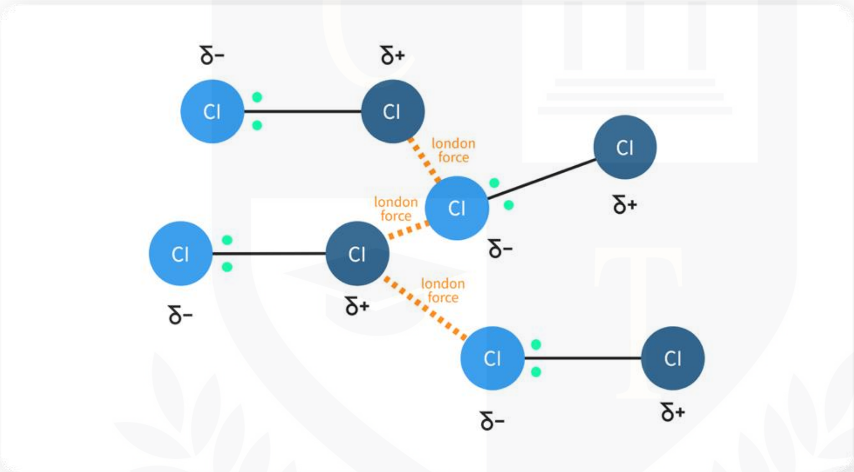
Permanent dipole - dipole forces:

- Polar molecules have a permanent dipole with a negatively and positively charged end, resulting in permanent dipole-dipole forces between two such molecules.
- The δ+ end of one molecule's dipole is attracted to the δ- end of a neighboring molecule's dipole.

Dipole–dipole interaction between a δ+ atom of one molecule and a δ– atom of another molecule.

Induced Dipole-induced Dipole Forces / Van Dar Waal forces:

- The electron charge cloud in non-polar molecules or atoms is in constant motion. As it moves, the electron charge cloud can be more concentrated on one side of the atom or molecule than the other, creating a temporary dipole on neighboring molecules.
- The δ+ end of one molecule's dipole is attracted to the δ- end of a neighboring molecule's dipole.
- Exist between all atoms or molecules.



Exam Questions

The intermolecular attractions between halogen molecules are London forces.

(i) Describe how London forces form between halogen molecules.

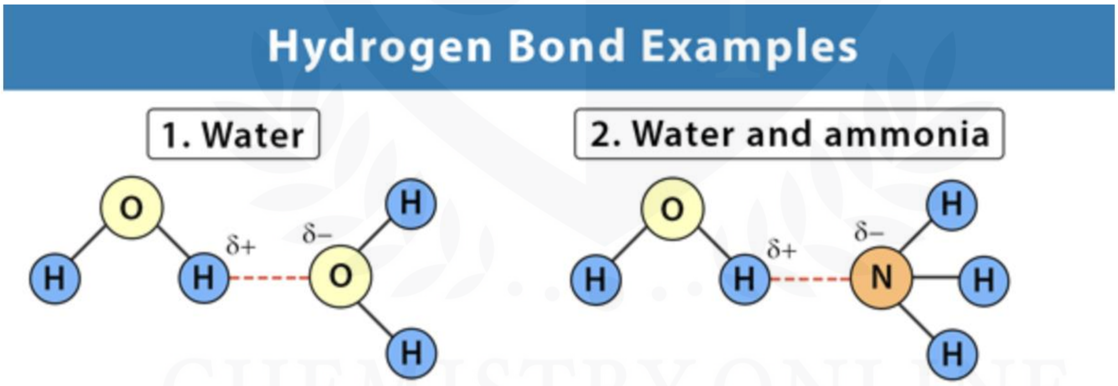
The boiling temperatures of chlorine and bromine are shown in the table.

Halogen	Boiling temperature / °C
chlorine	–34
bromine	59

Explain why bromine has a higher boiling temperature than chlorine.

<u>Hydrogen bonding</u>:

- It is the most potent form of intermolecular bonding.
- Hydrogen bonding is a type of permanent dipole – dipole bonding.
    - For hydrogen bonding to occur, a species must have an electronegative atom (O, N, or F) bonded to hydrogen.
- The bond becomes highly polarized when hydrogen is covalently bonded to oxygen, nitrogen, or fluorine. Consequently, the hydrogen atom acquires a positive charge (δ+), making it capable of forming a bond with the lone pair of an oxygen, nitrogen, or fluorine atom in a different molecule.
    - For example, Water can form two hydrogen bonds because the O has two lone pairs.

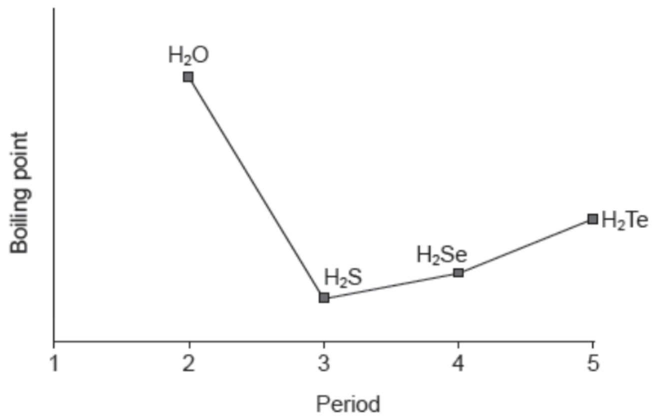


<u>Properties of water</u>:

Hydrogen bonding in water causes it to have anomalous properties such as high melting and boiling points, high surface tension, and a higher density in the liquid than the solid.

<u>High melting & boiling points</u>:

Water has high melting and boiling points due to the strong intermolecular forces of hydrogen bonding between the molecules in its solid (ice) and liquid forms. These forces make it difficult to separate the water molecules and require a lot of energy to melt or boil water.

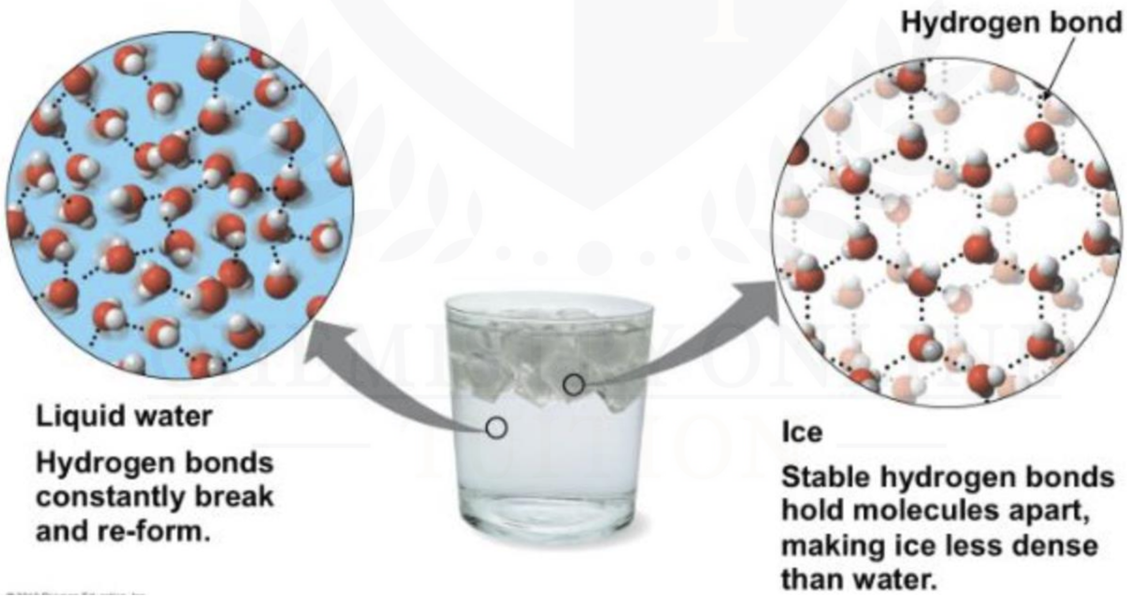


High surface tension:

- **Surface tension** is the ability of a **liquid surface** to resist any **external forces.**
- The water molecules at the **surface** of the liquid are bonded to other water molecules through **hydrogen bonds**. These molecules **pull the surface molecules downwards,** causing their surface to become compressed and more tightly together at the surface, so water has a **high surface tension.**

Density:

- **Solids** are **denser** than their **liquids.**
- The water molecules form an open lattice and are slightly further apart due to the long hydrogen bonds.



Exam Question

Water might be expected to have a lower boiling temperature than hydrogen sulfide but it actually has a higher boiling temperature.

Comment on this statement by referring to the intermolecular forces in both these substances.

A detailed description of how the intermolecular forces arise is not required.

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* Methanol, CH_3OH, is miscible with water in all proportions.
Sodium chloride is much less soluble in methanol than in water.

Explain these statements using your knowledge of the interactions between solutes and solvents.

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