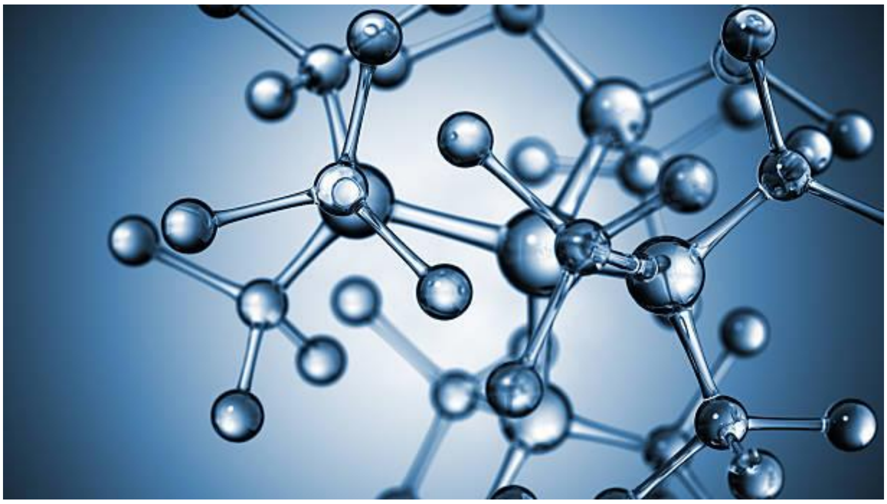
Master

Kirkcaldy High School



Higher Chemistry

Unit 1 - part 2

Chemical changes and structure

Name: ______________________

Class: ______________________

Teacher: ______________________

Assessment Page

Homework

Homework title	Date	Mark/Total Mark
N5 Formulae + H Periodicity		/
N5 Covalent Bonding + H Periodicity	.	/

Notes/comments

Check tests

Test title	Date	Mark/Total Mark
Unit 1: Key area 1bi		/

Notes/comments

Date: ________________

The Bonding Continuum

Overarching question(s) for this topic

- How does the differences in electronegativity between bonding atoms effect the type of bonding in the compound?

Covalent and Ionic Bonds

Covalent Bonds

- **Definition:** A chemical bond where atoms **share pairs** of electrons.

- **Formation:** Two positive **nuclei** are held together by their common attraction for the **shared pair** of **electrons**.

Ionic Bonds

- **Definition:** A chemical bond formed through the **electrostatic attraction** between **positive** and **negative ions.**

- **Formation:** Electrons are **transferred** from one atom to another, resulting in the formation of **ions**.

Electronegativity

- **Electronegativity:** The measure of an atom's ability to attract electrons in a chemical bond.

Trends

- **Across a Period: Increases** from left to right.

- **Down a Group: Decreases** from top to bottom.

Bonding Continuum

The bonding continuum describes the range of bond types from **pure** covalent (non-polar covalent) to **ionic**, with **polar covalent** lying in between.

This range is the result of the **electronegativity difference** between the bonding atoms.

Pure covalent (non-polar): A covalent bond where there is equal sharing of the bonding electrons.

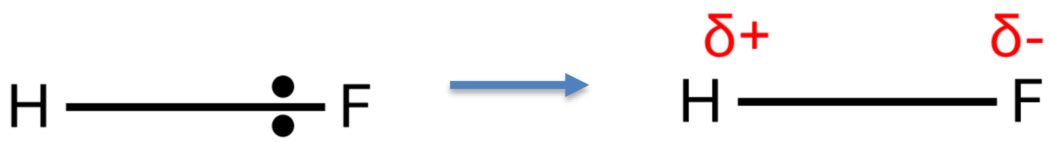
This requires a small or no difference in electronegativity. E.g. H_2
(Approx. ≤ 0.4).

● = electron

Polar covalent: A covalent bond where there is **unequal** sharing of the bonding electrons, resulting in **slight** positive and negative charges.

Slight charges are denoted by a Greek **delta** symbol (δ). The more electronegative atom becomes δ- and the less electronegative atom becomes δ+. These are known as **permanent dipoles**.

A significant difference in electronegativity is required for this. E.g HF
(Approx. > 0.4 ≤ 1.8)



Ionic: A bond where there is such an **unequal** share of the electrons that the **more** electronegative atom holds **both** of the electrons in the bond resulting in the formation of a **full positive** and **negative** charge.

This is the result of **large** electronegativity difference. E.g CsF
(Approx. >1.6 and generally metal to non-metal)

Note: The ranges for each bonding type can be amiguous around the boundaries and more information is sometimes required but the ranges listed are good enough for most cases.

Using page 11 of your data booklet to find the electronegativity differences, place the following bonds within the correct type of bonding below with the correct notation for the charges if there are any.

C – H	P – H	O – H	H – Cl	Na – Cl
S – H	Ca – O	Si – O	Li – O	C – F
Cl – Cl	Br – I	N – H	Al – Cl	K – Br

Increasing electronegativity difference

More covalent character *More ionic character*

Type of bonding	Pure covalent (non-polar covalent)	Polar covalent	Ionic
Approx E.N. difference	**≤ 0.4**	**> 0.4 ≤ 1.8**	**> 1.6**
Type of elements bonding	**Non-metal to non-metal**	**typically non-metal to non-metal**	**typically metal to non-metal**
Examples	H – H	$H^{\delta+}$ – $F^{\delta-}$	$Cs^{+}F^{-}$

Get a laptop and [Google 'Phet Molecule Polarity'] to help you understand how dipoles form. Only use the 'Two Atoms' mode for now.

For more practice, use the QR code or follow this [link]:	

Date: _______________

Intermolecular forces

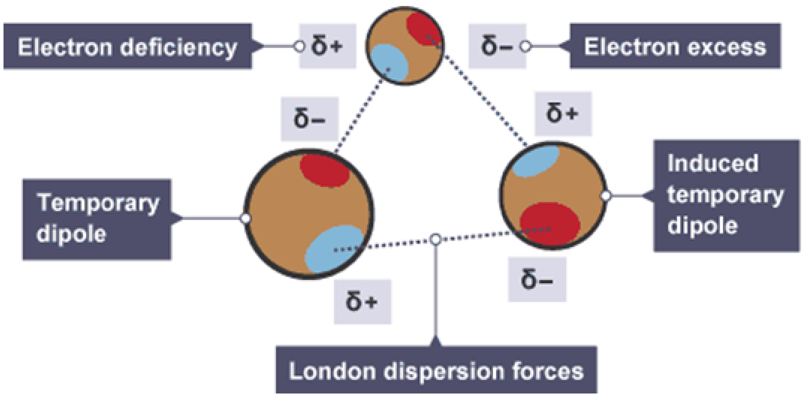
Overarching question(s) for this topic

- How does the type of bonding effect the properties of a substance?
- How does the differences in electronegativity between atoms effect the type of intermolecular forces present in a covalent molecular substance?
- How does the shape of a molecule effect the polarity of the molecule?

Types of intermolecular forces

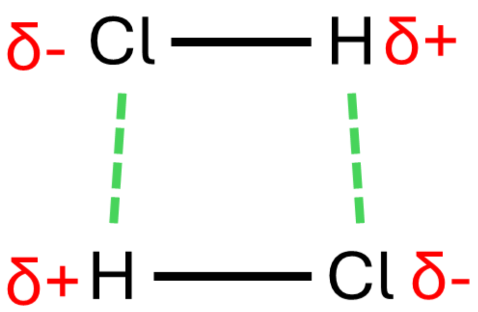
London Dispersion Forces (LDF) – see (Unit 1 part 1)

- **Definition**: The weakest intermolecular forces and arise due to **temporary** dipoles **induced** in atoms or molecules.
- **Formation**: Occur when the electrons in two adjacent atoms occupy positions that make the atoms form **temporary** dipoles.
- **Characteristics**: Present in all atoms and molecules, whether they are polar or non-polar. The strength of London dispersion forces increases with the number of electrons in the molecule. **Non-polar** substances **only** contain LDF.
- **Example**: Noble gases (e.g., Argon, Ar), and non-polar molecules like methane (CH_4).



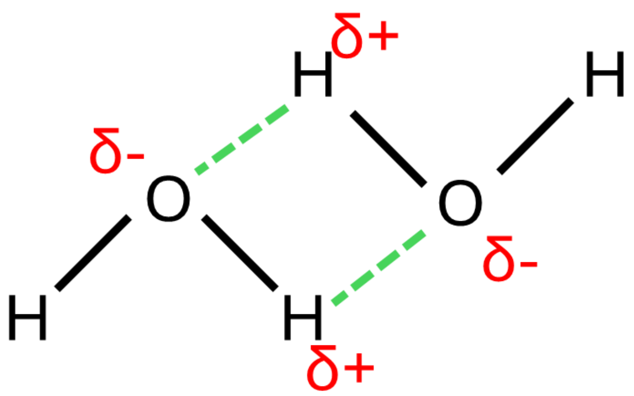
Permanent Dipole-Permanent Dipole Interactions (PD-PDI)

- **Definition**: Attractive forces between the **positive** end of one **polar** molecule and the **negative** end of another **polar** molecule.
- **Formation**: Occur in molecules that have **permanent** dipoles, i.e., molecules with a significant difference in electronegativity between bonded atoms, so have **δ-** and **δ+** part of the molecule.
- **Characteristics**: **Stronger** than London dispersion forces but **weaker** than hydrogen bonds. The **larger** the **difference** in electronegativity, the **stronger** the PD-PDI.
- **Example**: Hydrogen chloride (HCl), sulfur dioxide (SO_2).



Hydrogen Bonding

- **Definition**: A special type of Permanent Dipole-Permanent Dipole Interaction that occurs when **hydrogen** is covalently bonded to a **highly** electronegative atom (**fluorine**, **oxygen**, or **nitrogen**).
- **Formation**: The same as PD-PDI but forms a stronger bond.
- **Characteristics**: The **strongest** intermolecular force. Hydrogen bonding greatly influences the physical properties of compounds, such as **increasing** boiling points and **solubility** in water.
- **Example**: Water (H_2O), ammonia (NH_3), hydrogen fluoride (HF).



Summary Table

Type of Force	Relative strength	Example	Key Features
London Dispersion Forces	**Weakest**	**Argon (Ar), Methane (CH_4), Chlorine (Cl_2)**	**Present in <u>all</u> atoms/molecules** **Non-polar molecules <u>only</u> have LDF**
PD-PDI Interactions	**Intermediate**	**Hydrogen chloride (HCl), Sulfur dioxide (SO_2)**	**Occur in polar molecules**
Hydrogen Bonds	**Strongest**	**Water (H_2O), Ammonia (NH_3), Hydrogen Fluoride (HF)**	**H bonded to F, O, or N**

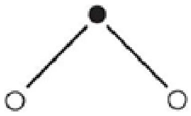
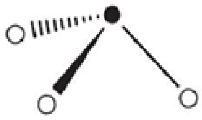
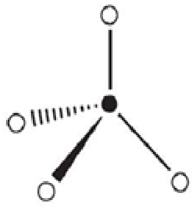
Questions (2 atom bonds only)

Fill in the table below to categorise the type of intermolecular force(s) present in the following molecules. You must determine first if they are polar (determine their electronegativity differences). We will only focus on 2 atoms for now.

Bond	Electronegativity Difference	Polar or Non-Polar	Type of Intermolecular Force(s) Present
H – Cl	1.0	Polar	LDF and PD-PDI
O = O			
S – H			
C – H			
C – O			
Br – Br			
H – F			
C – O			
C – Cl			
Cl – Cl			
C = C			
P – O			
H – Br			
N ≡ N			
H – O			
C – N			
N – H			

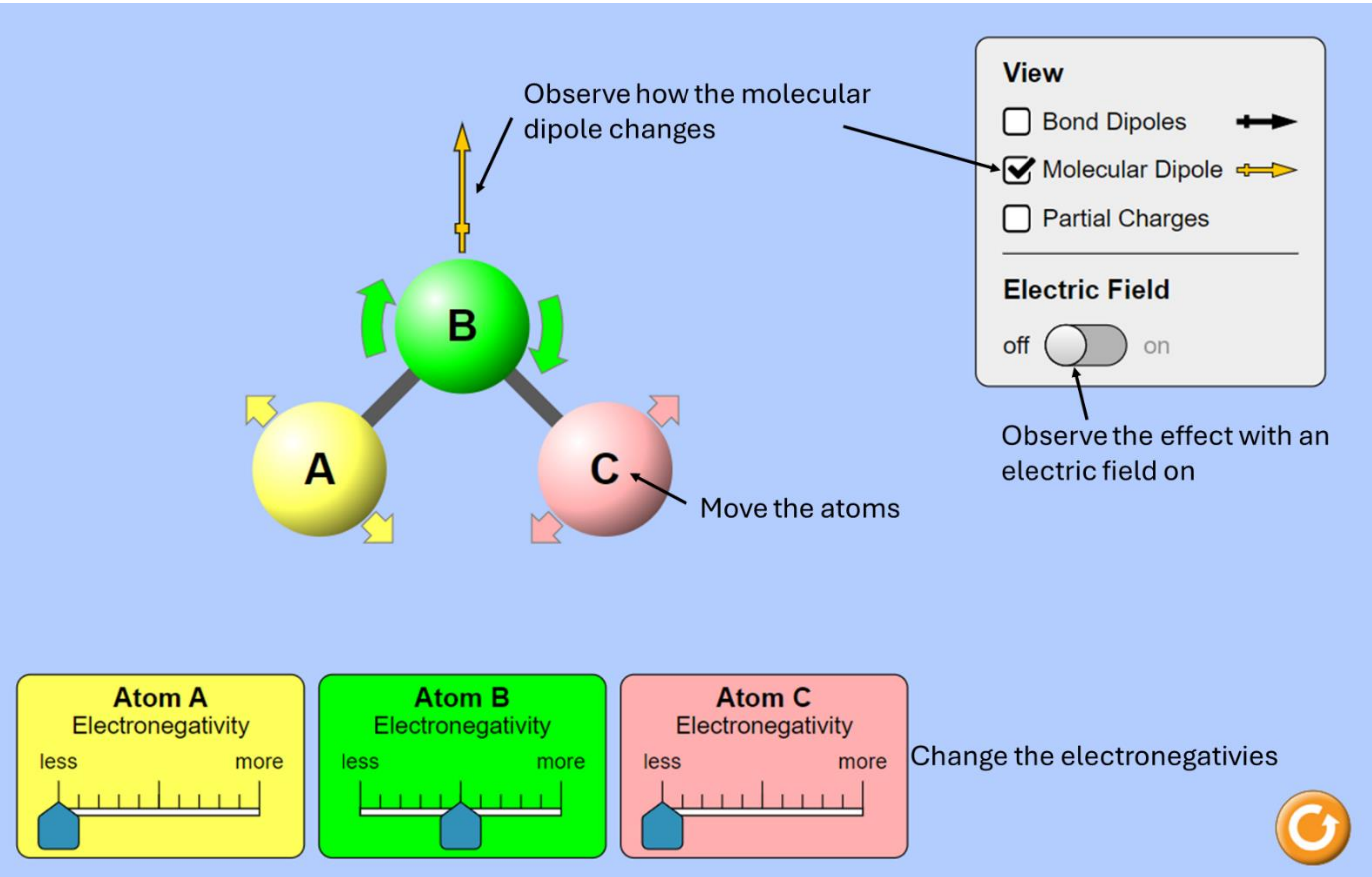
Polar and Non-polar molecules

It is possible for a molecule to contain polar bonds but overall the molecule becomes non-polar. This is due to the symmetry of the molecule causing the electronegativity differences to cancel out. For this, we must revisit the shapes of molecules we learned about in National 5.

Name of shape	Linear	Angular	Trigonal Pyramidal	Tetrahedral
Shape				

To understand how a molecule with polar bonds can become non-polar:

Get a laptop and [Google 'Phet Molecule Polarity'] to help you understand how dipoles form. Only use the 'Three Atoms' mode.



1. Fill the following examples table with the drawing of the following molecules

   HCl, NH_3, SO_2, CO_2, SH_2, PH_3, CH_4 methane, CCl_4 (tetrachloromethane), CCl_3H (trichloromethane), CCl_2H_2 (dichloromethane), $CClH_3$ (chloromethane)

2. Assign the δ- and δ+ parts of the molecules where necessary.
3. Assign whether it is a:
   a. Non-polar molecule with non-polar bonds (NPM-NPB)
   b. Polar-molecule with polar bonds (PM-PB)
   c. Non-polar molecule with polar bonds (NPM-PB)

Linear	**Angular**	**Trigonal Pyramidal**	**Tetrahedral**

Only linear and tetrahedral have the potential to cancel out due to the symmetry of the molecule if the bonds are polar.

For more practice, use the QR code or follow this link:	

Date: _______________

Properties of polar/non-polar molecules

Overarching question(s) for this topic

- How does the polarity and intermolecular forces within a molecule effect it's properties?

Types of properties

Melting Point

The melting point of a substance is the temperature at which it changes from a solid to a liquid.

Stronger intermolecular forces require **more** energy to overcome, leading to **higher** melting points.

- **Example:** Substances with hydrogen bonding (and LDF), like water (H_2O), have significantly higher melting points than those with only London dispersion forces, like methane (CH_4).

Boiling Point

The boiling point is the temperature at which a substance changes from a liquid to a gas.

As intermolecular forces **increase**, more energy is needed to **separate** the molecules, resulting in **higher** boiling points.

- **Example:** Hydrogen fluoride (HF), which exhibits hydrogen bonding (and LDF), has a higher boiling point than hydrogen chloride (HCl), which only exhibits PD-PDI interactions (and LDF).

Summary: As intermolecular force strength increases, both melting and boiling points increase because more energy is required to overcome the attractions between molecules.

Solubility

General Principle: "Like dissolves like."

Polar substances tend to dissolve well in **polar solvents** due to the **strong** intermolecular forces (e.g., hydrogen bonds, PD-PDI) between solute and solvent.

Non-polar substances dissolve better in **non-polar solvents** where London dispersion forces are the dominant interaction.

Polar and non-polar substances **repel** each other, this is why oil (**non-polar** molecules) do not mix with water (**polar** molecules).

With Increasing Intermolecular Forces:

- **Polar Solvents:** As the polarity and intermolecular forces of a solute **increase**, its solubility in **polar** solvents (e.g., water) generally **increases**.

- **Non-Polar Solvents:** Strong intermolecular forces in a **polar** solute may **reduce** its solubility in non-polar solvents because the intermolecular forces between the solute and solvent are **weaker**.

Examples:

- Hydrogen Fluoride (HF) dissolves well in water (H_2O) as they are both polar and have strong hydrogen bonding between them.

- Hydrogen chloride (HCl) dissolves well in water (H_2O) as they are both polar and have relatively strong PD-PDI between them.

- Oil (a non-polar substance) does not dissolve in water but will dissolve in hexane (a C_6H_{14} - non-polar solvent).

Additional Note:

- Ionic substances, such as NaCl, dissolve very well in water due to very **strong ion-dipole** interactions.

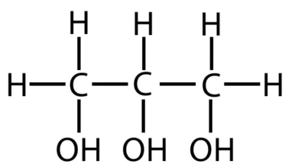
Viscosity

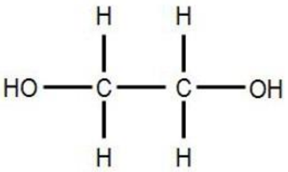
- **Viscosity** is a measure of a liquid's resistance to flow (how thick or runny it is). Substances with **strong** intermolecular forces tend to have **higher** viscosities because the molecules are more **strongly** attracted to each other, making it **harder** for them to **move past** one another.

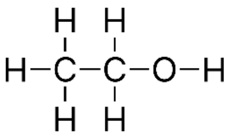
With Increasing Intermolecular Forces:

- As intermolecular force strength increases, the viscosity of a substance **increases**.

- **Example**: Glycerol (propan-1,2,3-triol), which has multiple hydroxyl groups capable of hydrogen bonding, has a much higher viscosity than ethanol. Ethan-1,2-diol has an intermediate viscosity between these.


Glycerol (propan-1,2,3-triol)


ethan-1,2-diol


ethanol

Volatility

- Volatility refers to how readily a liquid substance **vaporises** (turns into a gas).

- Substances with **weaker** intermolecular forces are more volatile because their molecules can escape into the gas phase more easily.

- Nice smelling volatile substances are added to candles and cosmetics, so they produce a strong pleasant smell. The molecules must be physically in your nose to smell them.

With Increasing Intermolecular Forces:

- As intermolecular forces **strengthen**, the volatility of a substance **decreases**.

- **Example**: Ethanol (which has hydrogen bonding) is less volatile than diethyl ether (which primarily has PD-PDI interactions).

Summary of Changes with Increasing Intermolecular Forces

- **Melting and Boiling Points**: Increase with **stronger** intermolecular forces due to the **higher** energy required to overcome these forces.

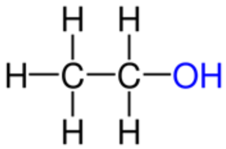
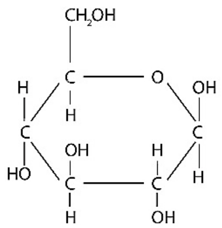
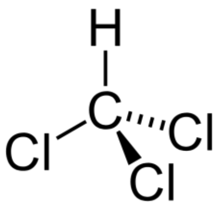
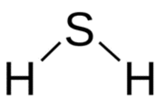
- **Solubility**: Solubility increases in solvents with **similar** intermolecular forces (e.g., polar solutes in polar solvents). Polar substances with strong intermolecular forces dissolve poorly in non-polar solvents.

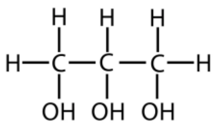
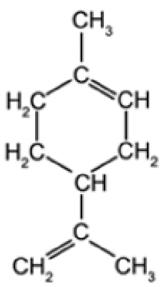
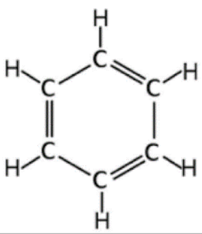
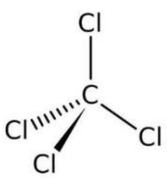
- **Viscosity**: Increases as intermolecular forces **strengthen**, making the liquid flow more **slowly**.

- **Volatility**: Decreases with **stronger** intermolecular forces because more energy is required for the molecules to escape into the gas phase.

Summary Diagram/Flowchart

Questions

Molecule	Soluble in polar solvent - e.g. Water (Yes/No)	Soluble in non-polar solvents – e.g. hexane (Yes/No)	Estimated Melting Point/Boiling Point (low/High)	Look up: State at Room Temperature (solid/Liquid/Gas)	If Liquid: Viscosity (low/high)	If Liquid: Volatility (low/High)
Ethanol: 						
Glucose: 						
Propane: **C_3H_8**						
Chloroform (trichloromethane): 						
Hydrogen Chloride: **HCl**						
Ammonia: **NH_3**						
Hydrogen Sulfide: 						

Molecule	Soluble in polar solvent - e.g. Water (Yes/No)	Soluble in non-polar solvents – e.g. hexane (Yes/No)	Estimated Melting Point/Boiling Point (low/High)	Look up: State at Room Temperature (solid/Liquid/Gas)	If Liquid: Viscosity (low/high)	If Liquid: Volatility (low/High)
Glycerol (propan-1,2,3-triol) 						
Bromine: **Br_2**						
Salt: **NaCl**						
Limonene: 						
Benzene: 						
Tetrachloromethane: 						

Understanding Molecular Properties Through Intermolecular Forces

Objective: You are tasked with analysing the differences in properties among several molecules. Using the information provided about each molecule, your goal is to determine how the type and strength of intermolecular forces affect their physical properties, such as boiling point, solubility, and viscosity.

Molecules to Analyse:

Molecule	Structure/Forces	Boiling Point	Solubility	Viscosity
Water (H_2O)	Polar, hydrogen bonding	100°C	Highly soluble in polar solvents (e.g., water)	Low viscosity
Methane (CH_4)	Non-polar, London dispersion forces	-161.5°C	Insoluble in water, soluble in non-polar solvents	Very low viscosity
Ethanol (C_2H_5OH)	Polar, hydrogen bonding	78.37°C	Soluble in both polar and non-polar solvents	Low viscosity, slightly higher than water
Hexane (C_6H_{14})	Non-polar, London dispersion forces	68.73°C	Insoluble in water, soluble in non-polar solvents	Low viscosity
Chloroform ($CHCl_3$)	Polar, permanent dipole-permanent dipole interactions (PD-PDI)	61.2°C	Soluble in non-polar solvents, slightly soluble in water	Low viscosity

Instructions:

Explain Property Differences:

1. Water vs. Methane: Explain why water has a much higher boiling point than methane, despite both being small molecules.
2. Ethanol vs. Water: Discuss why ethanol, which also has hydrogen bonding, has a lower boiling point than water but is still higher than hexane.
3. Hexane vs. Ethanol: Consider why hexane has a much lower boiling point than ethanol and different solubility characteristics.
4. Chloroform (PD-PDI): Analyse how the permanent dipole-permanent dipole interactions in chloroform affect its boiling point and solubility compared to the other molecules.
5. Viscosity Comparison: Compare the viscosities of these substances, explaining why water, ethanol, and chloroform have different viscosities despite all being polar molecules with different types of intermolecular forces.

Conclusion:

6. Summarise how the type and strength of intermolecular forces present in each molecule lead to the observed differences in boiling points, solubility, and viscosity. Reflecting on the importance of understanding these forces when predicting and explaining the physical properties of various substances.

Questions

1. Explain fully why butan-1-ol is more viscous than ethanol. (3 marks)

2. The boiling point of ammonia, NH_3, is much higher than the boiling point of nitrogen, N_2. Explain fully why the boiling point of ammonia is much higher than the boiling point of nitrogen. (3 marks)

3. Liquid carbon dioxide has a boiling point of –78.5°C. Explain fully why carbon dioxide is a gas at room temperature. (3 marks)

4. Explain fully why pentan-1-ol is less volatile than pentane. (3 marks)

5. The boiling point of chloromethane, $CClH_3$, is much higher than the boiling point of methane, CH_4.

   Explain fully why the boiling point of chloromethane (-24.2 ^oC) is much higher than the boiling point of methane (-161.5 ^oC). (3 marks)

6. Liquid chlorine has a boiling point of –34.6°C. Explain fully why chlorine is a gas at room temperature. (3 marks)

7. Explain fully why glycerol (propan-1,2,3-triol) is more viscous than ethanol. (3 marks)

8. The boiling point of hydrogen peroxide, H_2O, is much higher than the boiling point of oxygen, O_2. Explain fully why the boiling point of hydrogen peroxide is much higher than the boiling point of oxygen. (3 marks)

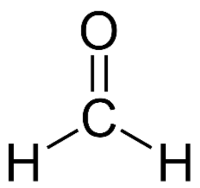
9. Liquid ethane has a boiling point of –88.6°C. Explain fully why ethane is a gas at room temperature. (3 marks)

10. Explain fully why pentane is more viscous than propane. (3 marks)

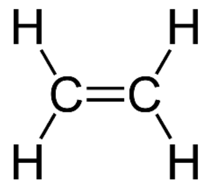
Challenge Questions

1. **Explain fully** why iodine monochloride (ICl) and bromine (Br_2) make a good comparison to determine the relative strength of intermolecular forces. (3 marks)

2. **Explain fully** why comparing formaldehyde (CH_2O) and ethene (C_2H_4) is useful for understanding the relative strength of intermolecular forces. (3 marks)



formaldehyde



ethene

Flash Cards:

[Structure and Bonding – The bonding Continuum]

Extension questions:

Hodder and Gibson

Blue Chemcord book

SCHOLAR

Past Papers

S & B 1st 20	2015	2016	2017	2018	2019	2021	2022	2023
MC	1,4,5	1,2,3	1,3,4,5	5	1,2,16	3, 5	1, 3	1,2,9
S2	1	3a, 12ai	5cii	2b-c, 9b	4d.i, 6a.i, 12 b.i	1.c, 7c	1c, 6bii,	2a, 2biii, 7ci,9a

<u>TEAMS: Check Test – Unit 1: Key Area 1bi</u>

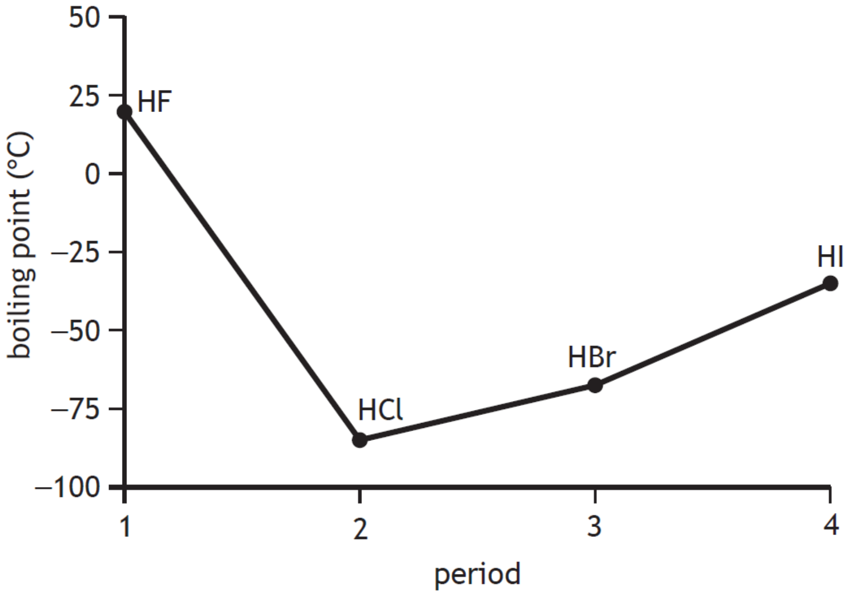
Past Paper Questions (in your jotter)

1. Fluorine has a greater attraction for bonding electrons than hydrogen.

   State the term used to describe the type of **covalent** bond in hydrogen fluoride. (1 mark)

2. Hydrogen halides are diatomic molecules formed between hydrogen and the elements fluorine, chlorine, bromine and iodine.

   The boiling points of the hydrogen halides are shown in the graph below.



Hydrogen fluoride, HF, has the highest boiling point of the hydrogen halides.

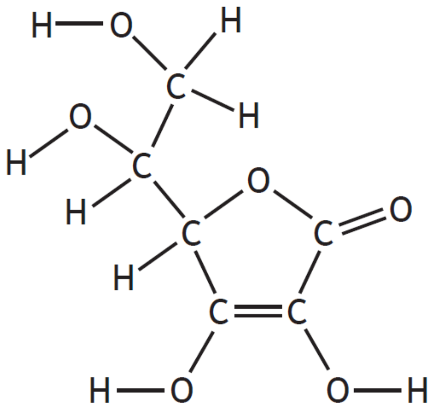
a. **State** the name of the strongest type of intermolecular force found between hydrogen fluoride molecules and **explain** how this type of intermolecular force arises. (2 marks)

The table shows the boiling points of hydrogen chloride, bromide and hydrogen iodide

Hydrogen halide	Boiling point (°C)
Hydrogen chloride	–85
Hydrogen bromide	–66
Hydrogen iodide	–35

b. **Explain fully** why the boiling point increases from hydrogen chloride to hydrogen iodide. (2 marks)

3. The structure of Vitamin C is shown

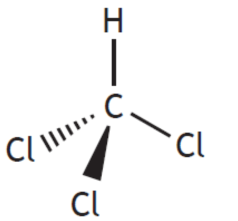


Explain fully why Vitamin C is soluble in water. (2 marks)

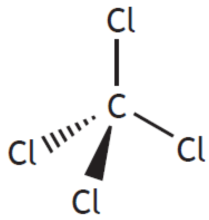
4. Chlorine is added to tap water to make it safe to drink.

The chlorine can react with substances in water to produce trichloromethane, $CHCl_3$.

Trichloromethane is more soluble in water than tetrachloromethane due to the polarities of the molecules.



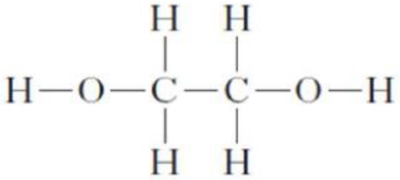
trichloromethane



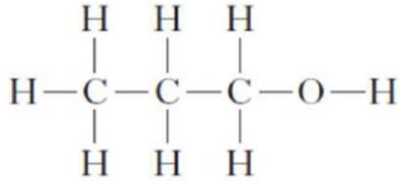
tetrachloromethane

Explain the difference in polarities of trichloromethane and tetrachloromethane molecules. (2 marks)

5. **Explain fully** why ethane-1,2-diol is more viscous than propan-1-ol (2 marks)



ethane-1,2-diol



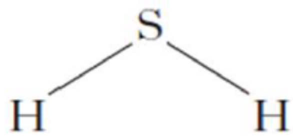
propan-1-ol

6. The boiling pont of hydrogen fluoride, HF, is much higher than the boiling point of F_2.

H — F	F — F
boiling point: 19·5 °C	boiling point: −188 °C

Explain fully why the boiling point of hydrogen fluoride is much higher than the boiling point of fluoride. (3 marks)

7. Liquid hydrogen sulfie has a boiling point of –60 °C.



Explain fully why hydrogen sulfide is a gas at room temperature. (3 marks)