Preliminary Chemistry.

Water I.

CHECKPOINT

By the end of the lesson, students should be able to:

Solutions

☐ Compare the state, percentage and distribution of water in the biosphere, lithosphere, hydrosphere and atmosphere

☐ Define the terms solute, solvent and solution

☐ Identify the importance of water as a solvent

☐ Outline the significance of the different states of water on Earth in terms of water as:

- a constituent of cells and its role as both a solvent and a raw material in metabolism
- a habitat in which temperature extremes are less than nearby terrestrial habitats
- an agent of weathering of rocks both as liquid and solid
- a natural resource for humans and other organisms

Intermolecular Forces

☐ Identify the water molecule as a polar molecule

☐ Describe hydrogen bonding between molecules

☐ Describe the attractive forces between polar molecules as dipole-dipole forces

☐ Construct Lewis electron dot structures of water, ammonia and hydrogen sulfide to identify the distribution of electrons

☐ Compare the molecular structure of water, ammonia and hydrogen sulfide, the differences in their molecular shapes and in their melting and boiling points

Practicals and Research

☐ Plan and perform an investigation to identify and describe the effect of anti-freeze or salt on the boiling point of water

WATER

Earth's position in the solar system enables water to exist in solid, liquid and gaseous forms on and around the earth's surface. To begin this new topic, we will look at the distribution of water around the earth and its significance in biological and geological processes. Further, we will look further at the molecular structure of water and introduce intermolecular forces. We will then use this to analyse the different boiling and melting points of substances, before ending with an analysis on the unusual density changes of water.

DISTRIBUTION OF WATER

Water is present in all spheres of the earth. However, its abundance, distribution and state varies.

Sphere	Description of Sphere	States of water	Abundance of water	Distribution
Lithosphere	Crust of the earth	liquid, solid	<10%	Mostly as groundwater, with some as ice on mountaintops
Hydrosphere	All water ON the earth's surface	liquid, solid	>95%	Liquid within oceans, rivers and lakes etc. Solid form present in glaciers
Atmosphere	The air surrounding the earth	gas, liquid, solid	0.5%-5%	Majority as water vapour, with remaining as rain and hail
Biosphere	The part of the earth where living organisms are found	liquid	60-90%	Within living organisms, in the habitat of living organisms

SOLUTIONS

Before we continue, it is first important to define the terms **solute, solvent** and **solution.**

Definitions

A **solvent** is …….

A **solute** is ……

A **solution** is ………………………………………………………………………………………………………

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If a substance dissolves in a solvent, then it is **soluble**. Otherwise, it is **insoluble**. The process by which a solute dissolves into a solvent is called **dissolution**.

Question 1 (1 mark)

A key ingredient in the manufacture of sodium carbonate is brine, which consists of a large amount of sodium chloride dissolved in pure water. Identify the solute, solvent and solution in the above statement.

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THE IMPORTANCE OF WATER

MAINTAINING METABOLISM

Solvent

Water, being a highly effective solvent is able to dissolve many substances. This makes it important for

- **Transport –** nutrients (e.g. salt, oxygen) as well as waste (e.g. carbon dioxide, urea) can be dissolved and transported within, around and out of organisms to maintain proper metabolism. Blood and urine are two substances made mainly out of water which are solutions used for transportation.
- **Reaction medium –** substances dissolved in water can undergo many chemical reactions necessary to support life

Raw Material

Water is involved as a reactant in many reactions necessary for life. For example, it is used in photosynthesis which produces glucose – the basic unit of energy in the biosphere

$$6CO_{2(g)} + 6H_2O_{(l)} \xrightarrow{\text{light}} C_6H_{12}O_{6(aq)} + 6O_{2(g)}$$

TEMPERATURE MODERATION

Water has a high heat capacity (this will be elaborated on later). This means it can absorb larger amounts of heat energy with smaller changes in temperature, resulting in water bodies having a smaller temperature range than the surrounding environment. It is thus a suitable habitat for life forms such as fish and bacteria which can only survive in certain temperature ranges.

WEATHERING

Water in both liquid and ice form breaks down rocks physically and chemically in a process called weathering. Liquid water can detach small particles and dissolve the compounds that constitute rock.

Ice can also cause erosion. Water can seep into cracks and then expand upon freezing, causing shattering of the rock. Eventually large chunks of rock are broken away.

Natural Resource

Water is a versatile substance, and humans use it for many purposes including

- Recreation e.g. fishing, swimming
- Irrigation
- Electricity Generation (hydroelectricity)
- Drinking and food preparation

Question 2 (3 marks)

Explain the importance of water in biological and geographic processes

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POLAR BONDS & MOLECULES

The properties of water, such as its high heat capacity and its efficacy as a solvent are determined by the structure of water molecules and the forces between them. This section will develop an understanding of the forces that hold molecules together and why they are important in determining the properties of a substance

LEWIS DOT STRUCTURES

Lewis dot structures show the distribution of electrons in bonded atoms. This section is a short review of this concept.

Talent Tip:

Question 3 (3 marks)

Draw the Lewis electron dot diagrams for

a) F_2 **1**

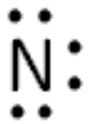
b) N_2O **1**

c) OH^- **1**

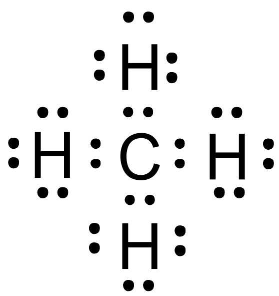
Question 4 (3 marks)

What is wrong with the following Lewis dot diagrams?

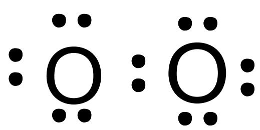
a) **1**



b) **1**



c) **1**



POLAR COVALENT BONDS

We know that a covalent bond forms when two atoms share electrons and are 'tied' together by electrostatic forces. Importantly, we need to analyse **how** those electrons are distributed in the bond, and **what** this means for the two atoms. How the electrons are distributed depends on electronegativity.

Electronegativity is ………………………………………………………………………………………………………..

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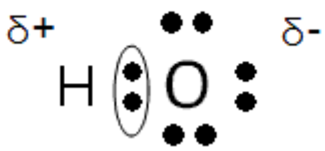
Electronegativity differs between different types of atoms. If one atom in a bond has a higher electronegativity than the other, then the shared electrons would tend to be located closer to the more electronegative atom. This would create a very small net negative charge (δ^-) closer to the more electronegative atom, and a very small net positive charge (δ^+) closer to the less electronegative atom (note that we are dealing with neutral atoms). This creates dipoles – an occurrence of equal positive and negative charges - and forms a **polar covalent bond.**

Covalent bonds in which the electrons are unequally shared are called ……………………………

…………………………………………… These bonds have …………………………………………………………………

For example, the hydrogen – oxygen bonds in water are polar since oxygen is more electronegative then hydrogen. This means that the shared electrons (circled) are grouped more closely to oxygen, creating the net negative and positive charges.



Question 4 (8 marks)

a) Explain the trend in electronegativity in the following cases

i) Down a group **2**

ii) Across a period **2**

b) Using this information along with a periodic table, identify whether these covalent bonds will be polar or non-polar, giving a qualitative estimate on the degree of polarity (e.g. weak, strong)

i) $\mathrm{H-Cl}$ **1**

ii) $\mathrm{I-Br}$ **1**

iii) $\mathrm{O-O}$ **1**

iv) $\mathrm{C-Cl}$ **1**

POLAR MOLECULES

However, simply because the individual bonds in a molecule are polar does not mean the molecule itself is polar (that is, having a net dipole). The placement of the polar bonds can cause dipoles to neutralise each other, making the molecule neutral overall. This happens for **symmetrical** molecular structures – you will only be expected to know that this occurs for **linear** and **tetrahedral** structures.

A polar molecule needs to have ………………………………………………………………………………………….

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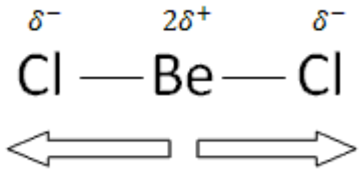
We can look at this concept with some examples, asking ourselves two questions

- Are the individual bonds in the molecule polar?
- Does the structure of the molecule allow for a net dipole?

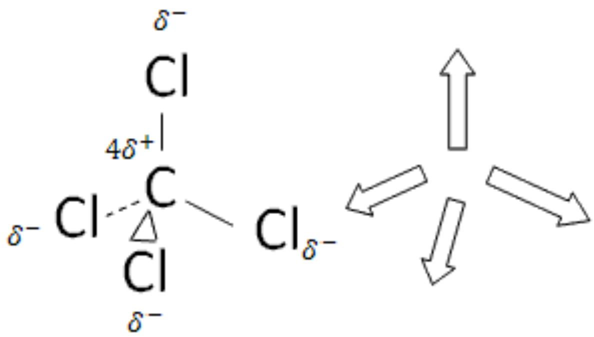
Beryllium Chloride ($BeCl_2$)



- **Polar bonds -** significant difference in electronegativity between Be and Cl.
- **Linear structure –** net dipoles cancel each other out

The molecule is **non-polar**. If we place a charge anywhere around the molecule, the net force experienced would be 0.

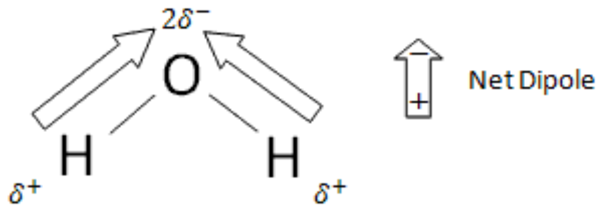
Carbon Tetrachloride (CCl_4)



- **Polar bonds –** significant difference in electronegativity between C and H
- **Tetrahedral structure –** net dipoles cancel each other out

The molecule is **non-polar**

Water (H_2O)



- **Polar bonds –** significant difference in electronegativity between O and H
- **Bent structure –** dipoles **do not** cancel each other out

The molecule is **polar**

Question 5 (4 marks)

Explain whether the following molecules have a net dipole.

a) CH_4 **1**

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b) HF **1**

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c) PCl_3 (pyramidal structure) **1**

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d) H_2S **1**

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INTERMOLECULAR FORCES

DEFINITION OF INTERMOLECULAR FORCES

Intermolecular Forces (van der Waals forces)

Intermolecular forces are ………………………………………………………………………………………………..

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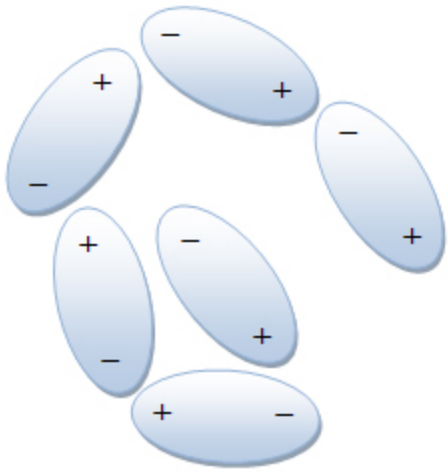
Intermolecular forces are **electrostatic forces**, and they arise as the result of the formation of dipoles. As discussed earlier, polar molecules have negative and positive poles which attract each other. The more polar the molecule, the stronger the attraction and the more tightly the molecules will be held together. This becomes very important when we start talking about physical properties of the substance.



Intermolecular forces can be divided into three types: hydrogen bonds, dipole-dipole interactions and dispersion forces.

TYPES OF INTERMOLECULAR FORCES

Hydrogen Bonds

When atoms in a bond have a large difference in electronegativity, this means that the bond is highly polar. When there are particularly strong polar bonds in a polar molecule, these form strong dipoles and thus strong intermolecular forces. Such polar bonds are formed between H and N, O or F, which have very large differences in electronegativity

A typical hydrogen bond is about $1/10^{th}$ the strength of a typical covalent bond. Nonetheless, molecules with hydrogen bonding are held together quite strongly.

Dipole-dipole interactions

In these molecules, the polar bonds are present but not as strong as those that form hydrogen bonds. Regardless, they still form significant dipoles and the intermolecular forces are significant.

Dispersion forces

Certain molecules do not have polar bonds but are still held together by intermolecular forces. This happens because the movement of electrons around atoms is random, so at one time, simply by chance, electrons may be more concentrated in one part of the atom. This forms a very weak temporary dipole that will induce another temporary dipole in a nearby molecule, forming an intermolecular force known as a dispersion force. Dispersion forces are present in all molecules, whether or not they contain polar bonds. They are more significant in larger molecules due to the large surface area of contact between molecules, and in atoms/functional groups with more electrons.

Summary of Types of Intermolecular Forces

1.

2.

3.

Question 6 (4 marks)

Determine the type of intermolecular force(s) acting between molecules of these substances

a) I_2 **1**

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b) HF **1**

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c) ClF **1**

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d) $CH_3CO - OH$ **1**

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INTRAMOLECULAR FORCES

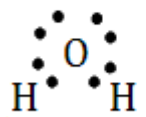
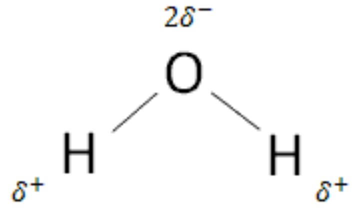
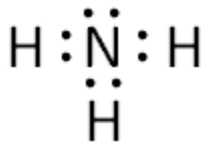
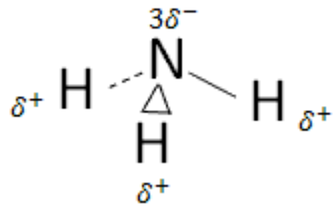
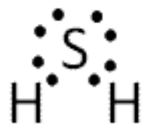
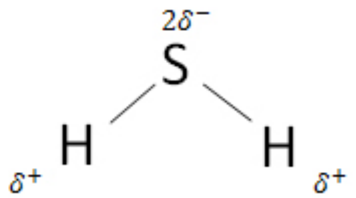
It is important to distinguish between **inter**molecular forces and **intra**molecular forces. Intramolecular forces hold together the atoms within a molecule or compound, and include **covalent bonds, ionic bonds and metallic bonds.**

STRUCTURES OF SIMPLE MOLECULES

Here, we will apply our knowledge of intermolecular forces to understand the reasons why boiling points and melting points vary between substances. This section will specifically compare water, ammonia and hydrogen sulfide.

WATER, AMMONIA AND HYDROGEN SULFIDE

The structures of water, ammonia and hydrogen sulfide are summarised below.

Compound	Molecular Formula	Lewis Dot	VESPR Diagram	Type of Intermolecular Force
Water	H_2O			Hydrogen bonding
Ammonia	NH_3			Hydrogen Bonding
Hydrogen Sulfide	H_2S			Dipole-dipole interactions

The bonds in these molecules are polar, and their structures are not completely symmetrical. Thus, they are all polar molecules.

Melting and Boiling Points

There are two factors affecting melting and boiling points – strength of intermolecular forces and the mass of the molecules.

Question 7 (2 marks)

If we take two substances, one which is polar and another which is non-polar but otherwise the same, which one will have a higher boiling/melting point, and why?

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Question 8 (6 marks)

a) Draw the Lewis dot diagrams for water and oxygen **2**

b) Explain the intermolecular forces present in these compounds **2**

c) Hence, explain why the boiling point of water is higher than the boiling point of oxygen. **2**

Melting/Boiling Points of Water, Ammonia and Hydrogen Sulfide

Compound	Melting Point (^{O}C)	Boiling Point (^{O}C)
Water	0	100
Ammonia	-78	-33
Hydrogen Sulfide	-82	-60

Although all these compounds have polar bonds, there are obvious differences in their melting and boiling points.

- Although both ammonia and water have hydrogen bonds, the pyramidal structure of ammonia means there is a larger degree of dipole cancelling, reducing the polarity of the molecule. This would make the boiling/melting point of ammonia lower than water
- The $\mathrm{H-S}$ bond is less polar than the $\mathrm{H-O}$ bond due to a smaller difference in electronegativity. This cancels out the fact that hydrogen sulfide is heavier than water.

Comparison of BP/MP of Water with other Small Molecules

Molecule	Melting Point ($^\circ\mathrm{C}$)	Boiling Point ($^\circ\mathrm{C}$)
Water ($\mathrm{H_2O}$)	0	100
Methane ($\mathrm{CH_4}$)	-183	-162
Ammonia ($\mathrm{NH_3}$)	-77.7	-33.3
Ethanol ($\mathrm{C_2H_5OH}$)	-114.1	78.3
Ozone ($\mathrm{O_3}$)	-193	-111

Although water has the lowest atomic mass of all the molecules present, it has by far the highest melting and boiling points. This can be attributed to the strength of its intermolecular forces – the hydrogen bonds that keep the water molecules together.