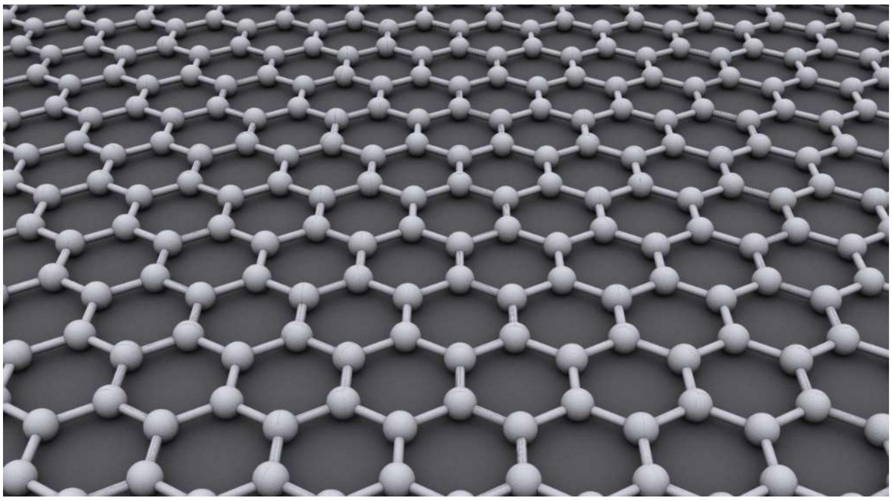
Kirkcaldy High School



Higher Chemistry

Unit 1

Chemical changes and structure

Name:______________________

Class:______________________

Teacher:______________________

Expectations and Outcomes Learner Evaluation

Topic: Chemical changes and structure

Experience and Outcomes	Date Completed (dd/mm/yy)	Evaluation How happy are you with it? (☺ ? ☹)

1. Periodicity Bonding in Elements 1-20

The Periodic Table

The periodic table was invented by Dimitri Mendeleev (1869). He arranged elements in order of increasing atomic number, and noted that their properties e.g. Melting point, boiling point and density were periodic in nature (repeating patterns existed). The periodic table allows chemists to make accurate predictions of physical properties and chemical behaviour for any element, based on its position. Features of the table are:

- **Groups:** vertical columns within the table contain elements with similar chemical properties resulting from a common number of electrons in the outer shell.
- **Periods:** rows of elements arranged with increasing atomic number, demonstrating an increasing number of outer electrons and a move from metallic to non-metallic characteristics.

Bonding and Structure in Elements 1-20

Chemical Bonding describes how atoms join together. Chemical bonds are electrostatic forces (attractions between positive and negative charges) which hold atoms together. Atoms form bonds to become more stable - by losing, gaining or sharing electrons.

The type of bond formed in a substance depends on the elements involved and their position in the periodic table.

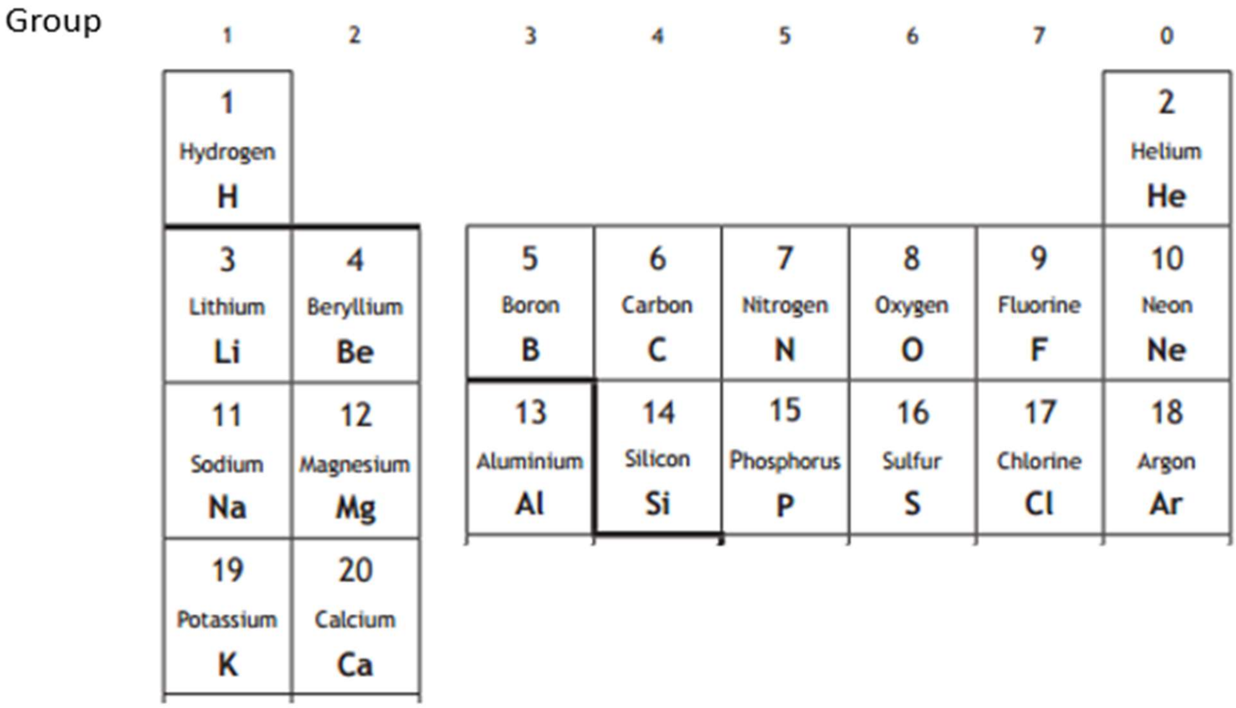
Structure describes how the atoms in the element/compound are arranged.

Types of Chemical bond (Intramolecular bonding)

Type of Bond	Description of bonding	Structure	Properties	Example
Metallic	Positive metal ions surrounded by a sea of delocalised electrons.		Usually high melting and boiling points. Delocalised electrons allows conduction of electricity.	
Covalent Molecular	Shared pair of electrons between the positive nuclei of two atoms.		Low melting point and boiling point. Does not conduct electricity.	
Covalent Network	Shared pair of electrons between the positive nuclei of two atoms. Many thousands of atoms are joined together by covalent bonds.		High melting point and boiling point. Does not conduct electricity (except graphite)	
Ionic	The attraction between positive and negative ions form a structured lattice.		High melting and boiling points. Free ions allow conduction of electricity only when molten or in solution.	

Elements 1-20 Types of bonding

Colour in the Periodic table and key to show the different types of bonding in the elements 1-20.

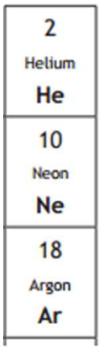


Group	1	2	3	4	5	6	7	0
	1 Hydrogen H							2 Helium He
	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulfur S	17 Chlorine Cl	18 Argon Ar
	19 Potassium K	20 Calcium Ca						

Key

Colour	Bonding type
	Metallic bonding
	Covalent network
	Discrete covalent molecules
	Discrete covalent molecules (diatomic elements)
	Monatomic elements

Monatomic Elements



Group = ___ atoms.

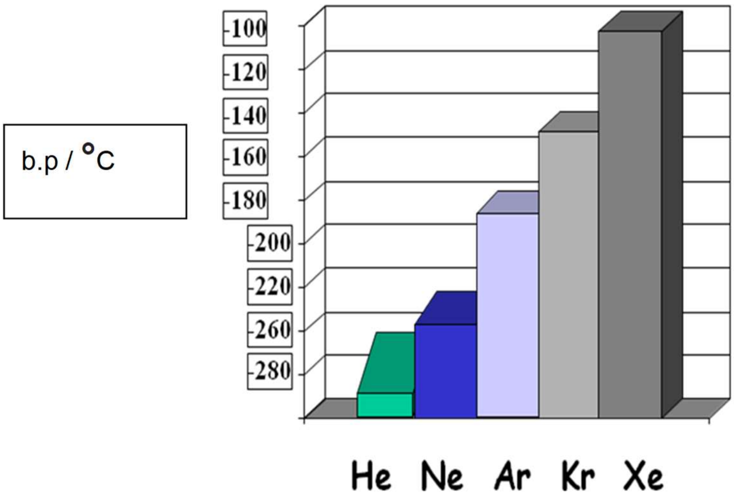
Monoatomic = Single atoms

Elements have a full outer shell of electrons, as a result forms no bonds to other atoms.

Properties

Monatomic elements do not conduct electricity because there are no free electrons.

Monatomic elements are gases with low melting and boiling points.



Using the graph above, describe the trend in Boiling point as you go down the group 8 elements. (Noble gases).

Boiling point indicates the strength of the intermolecular forces between the molecules.

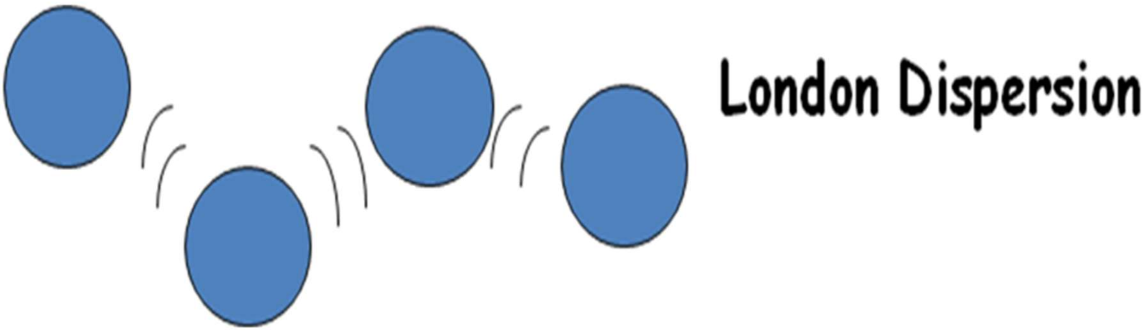
Describe the trend in strength of intermolecular forces between the molecules as you go down the group 8 elements.

Intermolecular forces will be explained in detail later. However, a brief introduction on London dispersion forces is needed as this is the intermolecular force present in the monatomic elements.

London Dispersion Forces (LDF)

London dispersion forces are a type of Van der Waals force that exists between molecules.

Complete the diagram to show how Van der Waals forces arise.

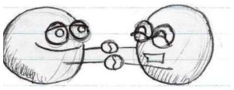


The strength of London dispersion forces (LDF) is related to the **number of electrons** within an atom or molecule – the more electrons the stronger the LDFs.

Example:

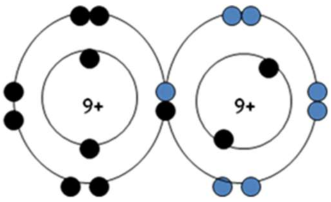
Explain why Xenon has a higher boiling point than Neon.

Covalent bonding

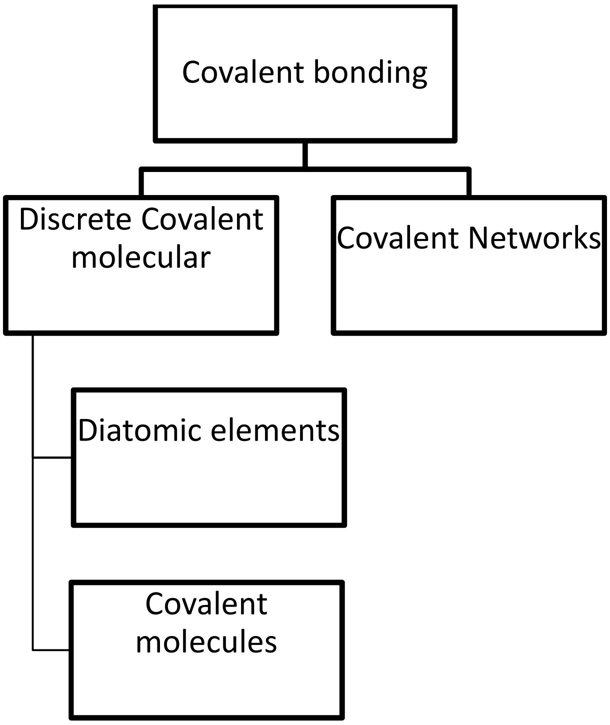


Covalent bonding involves an electrostatic attraction, where the atoms are held together by the attraction of the atoms two positive nuclei to the shared pair of electrons. Molecules which involve covalent bonds can be either discrete covalent molecules or covalent networks. Covalent bonds are strong and require a lot of energy to break them apart.

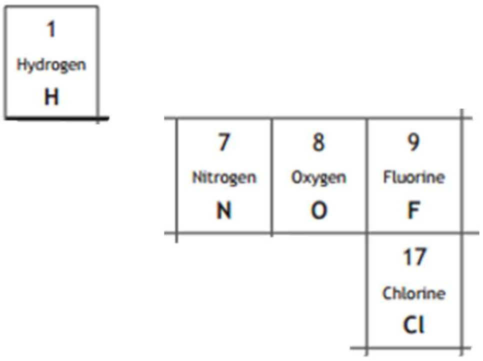
Example: Fluorine, F_2



Covalent bonds are found in covalent molecules and covalent networks. The bonding and structure impact the physical properties of the substance.

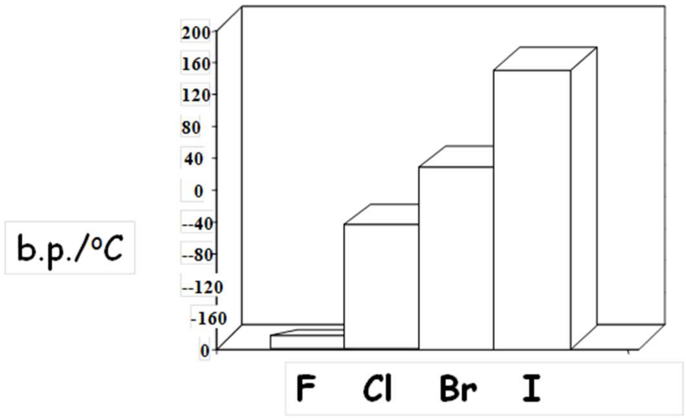


Discrete Covalent Molecules



Diatomic Elements (I_2, H_2, N_2, Br_2, O_2, Cl_2, F_2)
Two atoms bonded to each other.

Properties
A covalent molecular substance has a low melting/boiling point.



Using the graph above, describe the trend in Boiling point as you go down the group.

When fluorine melts or boils the strong covalent bonds remain intact, whilst the weak intermolecular forces are broken. In fluorine the intermolecular forces are London dispersion forces (LDF). Draw a diagram to show this:

Explain why Iodine has a higher boiling point than Fluorine.

Discrete Covalent Molecules (Sulfur and Phosphorus, Fullerene)

Phosphorus and Sulfur exist as **discrete** (separate) molecules but are NOT diatomic.

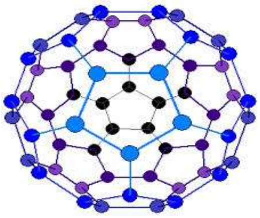
	Phosphorous = P_4	Sulfur = S_8
Structure		
Melting point (°C)		
Boiling point (°C)		
Conducts electricity?		

Use your data booklet to find the melting points and boiling points for Sulfur and Phosphorus.

Explain the difference in the melting points of Phosphorus and Sulfur.

Sulfur has a higher boiling point than Phosphorus. This is due to Sulfur having stronger London dispersion forces than Phosphorus. The strength of LDF is related to the number of electrons present. Sulfer exists as S8 molecule (this has ____electrons) whilst Phosphorus exists as a P4 molecule (this has ____ electrons).

Fullerenes



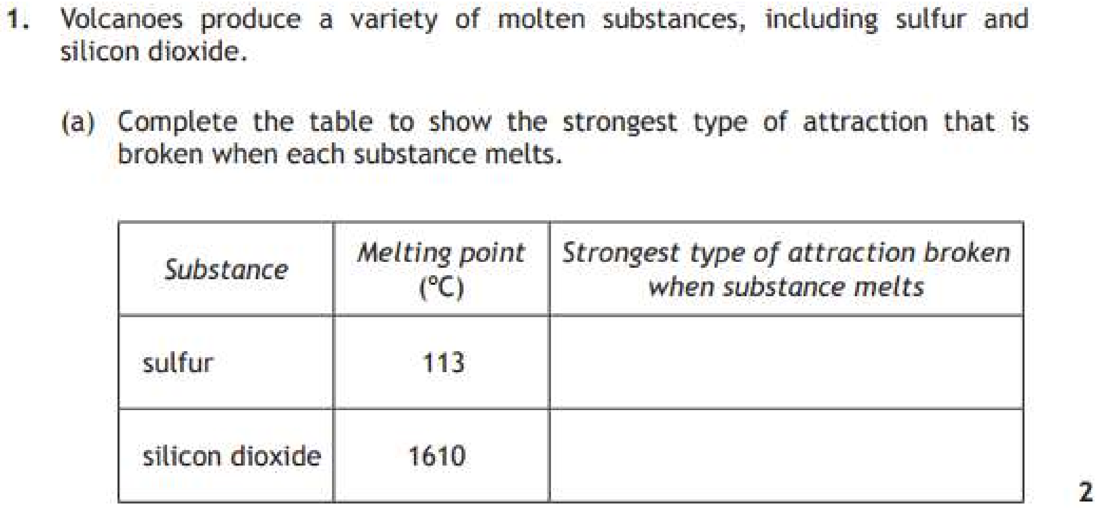
Fullerene (C60) **Carbon Nanotubes**

- Fullerene (C_{60}) is ***one*** of the **<u>three</u>** forms of carbon.
- It does **not conduct** electricity- as there are no delocalised electrons.
- Due to being a large molecule and having more electrons, fullerenes have stronger London dispersion forces than smaller molecules.
- It does **not conduct** electricity- as there are no delocalised electrons.

Covalent Molecular Summary

- Discrete <u>molecular elements</u> the intramolecular forces (within the molecules) are all strong covalent bonds. The intermolecular force (between the molecules) are weak London dispersion forces.
- It is the weak forces that are broken during boiling and melting. The larger the molecule is, the greater the number of electrons, the stronger the London dispersion forces.

Past Paper Questions



1. Volcanoes produce a variety of molten substances, including sulfur and silicon dioxide.

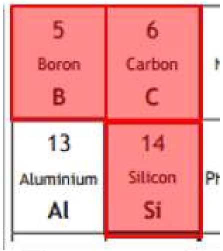
(a) Complete the table to show the strongest type of attraction that is broken when each substance melts.

Substance	*Melting point (°C)*	*Strongest type of attraction broken when substance melts*
sulfur	113	
silicon dioxide	1610	

2

Covalent Networks

Every atom is linked to another by **strong** covalent bonds.

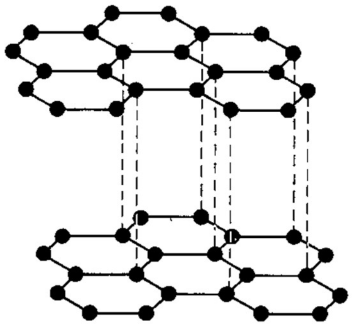


5 Boron B	6 Carbon C
13 Aluminium Al	14 Silicon Si

Covalent network structures contain many thousands of strong covalent bonds. These are the bonds that are broken when melting/boiling occurs.

Carbon has two covalent network structures:

- Diamond and Graphite



Graphite

3 bonds per carbon atom – layered structure with

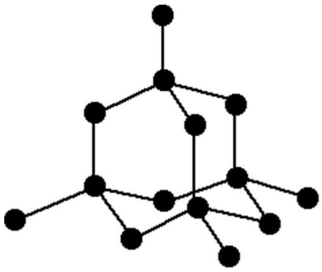
Delocalised electrons in between layers.

London dispersion forces between the layers

Conductor of electricity due to delocalised electrons between the layers – used in electrodes.

Very soft material.

High melting point.



Diamond

4 bonds per carbon atom – tetrahedral structure.

Non-conductor of electricity as no free electrons.

The hardest natural substance.

Very high melting point

Metallic Bonding

Metallic bonding involves the attraction of negatively charged electrons for the positive metal ions (**electrostatic attraction).**

1 Hydrogen H	
3 Lithium Li	4 Beryllium Be
11 Sodium Na	12 Magnesium Mg
19 Potassium K	20 Calcium Ca

Draw a diagram to represent metallic bonding

Properties

Good conductors as delocalised electrons are mobile and can carry

Most are solids (except Hg) with high melting and boiling points, due to the closely packed lattice structure with lots of strong bonds to break.

Questions

1 Which type of structure can be found in a fullerene?

A Ionic lattice

B Metallic lattice

C Covalent network

D Covalent molecular

2 Which of the following has a covalent molecular structure?

A Radium chloride

B A noble gas

C Silicon dioxide

D A fullerene

3 Which of the following elements exists as discrete molecules?

A Boron

B Carbon (diamond)

C Silicon

D Sulphur

4 Particles with the same electron arrangement are said to be isoelectronic.
Which of the following compounds contains ions which are isoelectronic?

A Na_2S

B $MgCl_2$

C KBr

D $CaCl_2$

5 The elements nitrogen, oxygen, fluorine and neon

A can form negative ions

B are made up of diatomic molecules

C have single bonds between the atoms

D are gases at room temperature.

6 Which type of structure is found in phosphorus?

A Covalent network

B Covalent molecular

C Monatomic

D Metallic lattice

7 Which element is a solid at room temperature and consists of discrete molecules?

A Carbon

B Silicon

C Sulphur

D Boron

8 Graphite, a form of carbon, conducts electricity because it has

A metallic bonding

B van der Waals bonding

C delocalised electrons

D pure covalent bonding.

Complete Chemcord Questions Page 7 in your jotter.

Complete Hodder and Gibson Textbook questions page

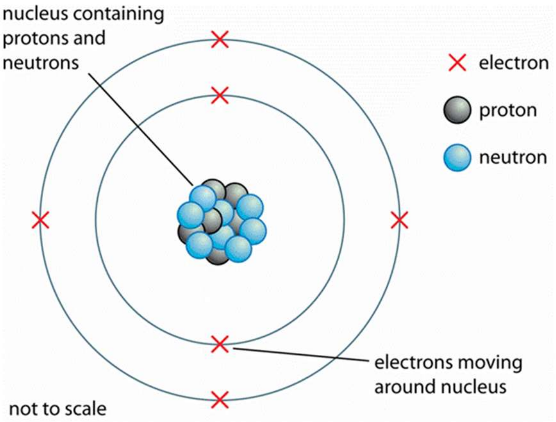
Periodicity

In this section you will learn about the trends in:

- Covalent Radii
- Ionisation Energy
- Electronegativity

N5 Recap – The Structure of the Atom.

Annotate the diagram with the Charges of protons, electrons and neutrons.



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To explain the trends in covalent radius, ionisation energy and electronegativity, we need first learn about nuclear charge and the shielding effect.

Nuclear Charge

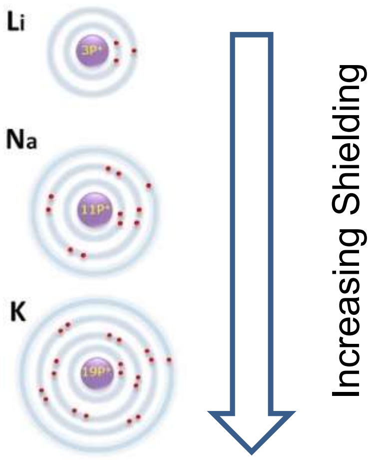
Nuclear charge is the positive **charge** on the **nucleus** of an atom. The greater the number of protons in the nucleus, the greater the nuclear charge.

Complete the blanks

The effects of nuclear charge on atomic size.

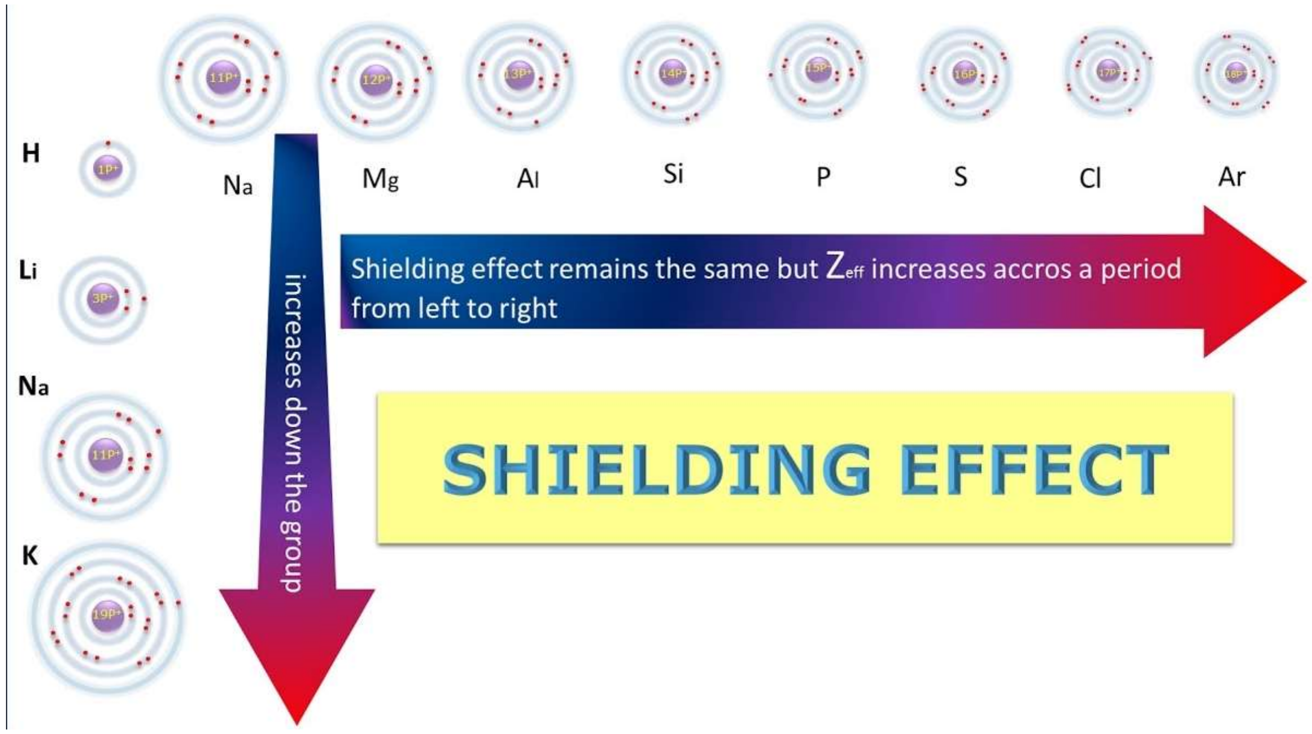
Going across a period, the atomic number _____________. Therefore, the number of protons ____________, resulting in the nuclear charge ___________. Electrons are ___________ charged, increasing atomic number results in the electrons being more strongly attracted to the nucleus which means that the size of the atom gets smaller.

Shielding Effect



Inner (core) electron shells will shield the outer electrons from the nucleus. Therefore, the full nuclear charge cannot be felt by the outer electron and will decrease to an **effective nuclear charge**.

The greater the number of electron shells the greater the shielding effect.



Covalent Radius

- The size of an atom is indicated by its covalent radius.
- The Covalent radius is half the distance between the nuclei of two covalently bonded atoms
- Draw a diagram to show what the term covalent radius means.

Trends in Covalent radii across a period.

Use your data book to complete the table.

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
Atomic Number	3	4	5				
Nuclear charge	3+	4+	5+				
Electron Arrangement	2,1	2,2					
Covalent Radius							

State the trend in covalent radius as you go across a period.

Explain the trend in covalent radius as you go across a period.

Covalent radii down a group

Element	**Li**	Na	K	Rb	Cs
Atomic Number					
Nuclear charge					
Electron Arrangement					
Covalent Radius					

State the trend in covalent radius as you go down a group.

Explain the trend in covalent radius as you down a group.

Covalent Radius Summary

Across the period the covalent radius **decreases** because the ***nuclear charge increases,*** the electrons are more strongly attracted to the nucleus.

Down the group the covalent radius **increases** because the number of **filled electron shells increase. The extra shells of electrons 'shield' the outer electrons from the nuclear charge meaning they are less strongly attracted to the nucleus.**

Questions

1. State the **two factors** contribute to the size of an atom?

2. Explain why are fluorine atoms smaller than carbon atoms?

3. Explain why are iodine atoms bigger than chlorine atoms?

4. Describe and explain the difference in size between a potassium and oxygen atom?

5. **Explain fully** why the atomic size increases down a group?

6. **Explain fully** why the atomic size decreases across a period?

Ionisation Energy

- **Definition of ionisation energies.**
- **Calculating ionisation energy for a given reaction.**
- **Trends in Ionisation.**

First Ionisation energy is the energy required to remove one mole of outer electrons from one mole of gaseous atoms.

This can be represented by an equation written in this format:

$E_{(g)} \longrightarrow E^{+}_{(g)} + e^{-}$ First Ionisation Energy

Ionisation energy data can be found in the databook (p).

Examples

1. Write the equation for the first ionisation of sodium.
2. Write the equation for the first energy of magnesium.
3. Write the equation for the first energy of aluminium.

- 2nd Ionisation energy: the amount of energy required to remove one mole of electrons from one mole of gaseous 1^{+} ions.

$E^{+}_{(g)} \longrightarrow E^{+}_{(g)} + e-$ [Second Ionisation Energy]

1. Write the equation for the second ionisation of sodium.

2. Write the equation for the second energy of magnesium.

3. Write the equation for the second energy of aluminium.

Calculating energy changes for reactions using ionisation energy

- Some reactions involve more than 1 mole of electrons being removed from an atom or ion. The energy change for these reactions can be calculated.

Example 1. Calculate the enthalpy change for the following reaction

$Al_{(g)} \longrightarrow Al^{3+}_{(g)} + 3e^-$

Questions

Calculate the energy required for the following changes:

1. $Ca_{(g)} \longrightarrow Ca^{2+}_{(g)} + 2e^-$

2. $K^+_{(g)} \longrightarrow K^{3+}_{(g)} + 2e^-$

3. $Li^+_{(g)} \longrightarrow Li^{3+}_{(g)} + 2e^-$

4. $Be_{(g)} \longrightarrow Be^{3+}_{(g)} + 3e^-$

5. $Na_{(g)} \longrightarrow Na^{2+}_{(g)} + 2e^-$

Trends in Ionisation Energy - across a period

Use your data booklet to describe the trend in ionisation energy as you go across a period.

Explain the trend in ionisation energy as your go across a period.

Across the period the first ionisation energy **increases** because the **nuclear charge increases,** the electrons are more strongly attracted to the nucleus resulting in more energy required to remove a mole of electron.

Trends in Ionisation Energy - down a period

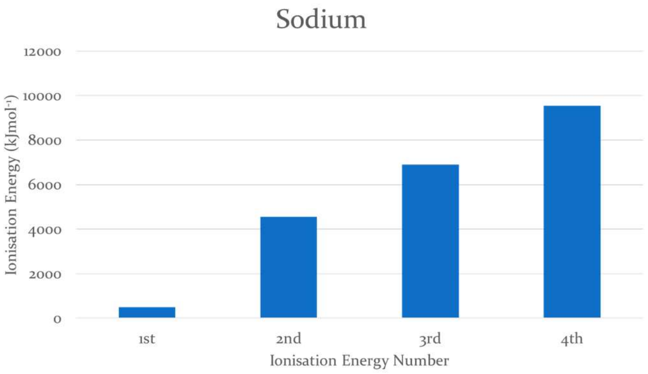
Use your databook to describe the trend in ionisation energy as you go down a group.

Explain the trend

Comparing Ionisation Energies

– 1st Versus 2nd

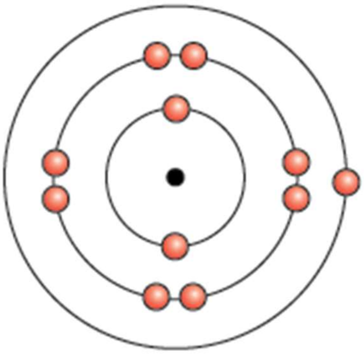
Why is the 2nd ionisation energy always greater than the 1st? Consider Sodium.

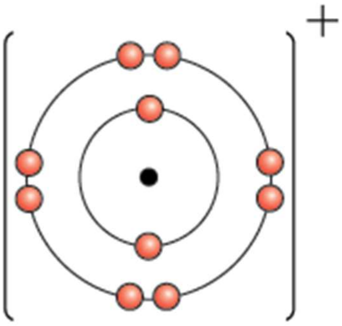


Once an electron is removed you are left with a positive ion that has more protons than electrons. The attraction of the outer electrons to the nucleus is stronger in the Sodium ion than it is in the sodium atom.

For sodium, the second ionisation energy is significantly higher than the first ionisation energy. Removing the second mole of electrons involves:

- a full, stable outer shell.
- an outer electron shell that is closer to the nucleus.
- an electron shell is less shielded (less inner electron shells).
- attraction of nucleus to the outer electron shell is greater.





Electron arrangement: Na: 2,8, 1 Na^+: 2, 8

Ionisation equation:

Ionisation Energy:

Questions

1. Describe what is meant by the first ionisation energy of an element?

2. Describe the trend in first ionisation energy of atoms of the elements across a period

3. Describe the trend in ionisation energy going down a group.

4. Explain why lithium atom has a larger first ionisation energy than a potassium Atom.

5. Explain the large difference between the second and third ionisation energies of calcium.

6. (a) Write an equation corresponding to:

   i) the first ionisation energy of rubidium,

   ii) the second ionisation energy of strontium,

   iii) the third ionisation energy of aluminium.

   (b) Refer to the Data Booklet and alongside each of the above equations write in the value for the energy change for this reaction.

7. Calculate the energy required to bring about the following changes.

   (a) $Al\ (g) \rightarrow Al^{3+}\ (g) + 3e^-$

   (b) $K^+\ (g) \rightarrow K^{3+}\ (g) + 2e^-$

Past Paper Questions

1 Which of the following equations represents the first ionisation energy of fluorine?

- A $F^-(g) \rightarrow F(g) + e^-$
- B $F^-(g) \rightarrow \frac{1}{2}F_2(g) + e^-$
- C $F(g) \rightarrow F^+(g) + e^-$
- D $\frac{1}{2}F_2(g) \rightarrow F^+(g) + e^-$

2 The table shows the first three ionisation energies of aluminium.

Ionisation energy (kJ mol^{-1})		
First	*Second*	*Third*
578	1817	2745

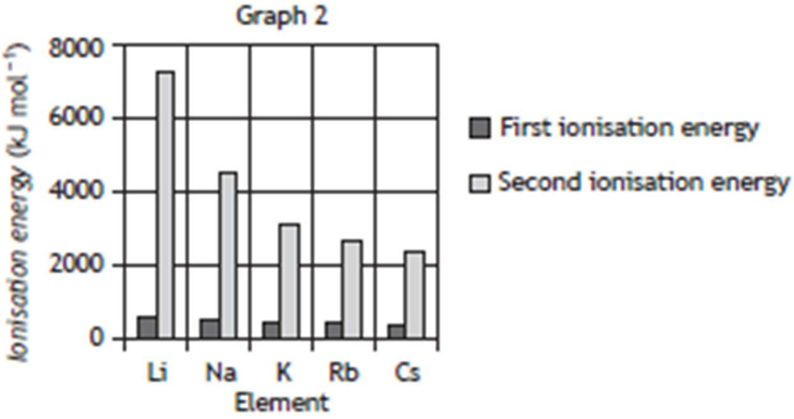
Using this information, what is the enthalpy change, in kJ mol^{-1}, for the following reaction?

$$Al^+(g) \rightarrow Al^{3+}(g) + 2e^-$$

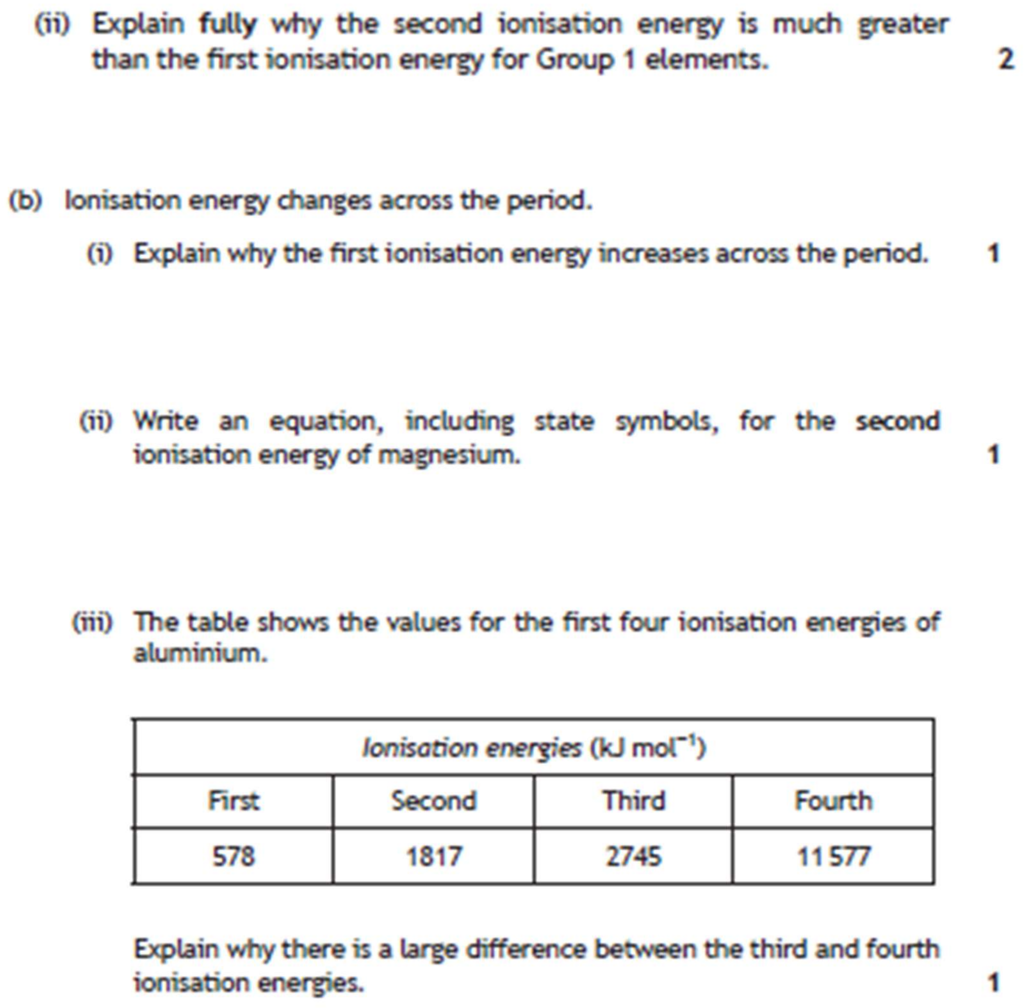
- A 1817
- B 2395
- C 4562
- D 5140

2. (continued)

(b) Graph 2 shows the first and second ionisation energies of elements in Group 1 of the Periodic Table.



(i) Explain why the first ionisation energy decreases going down Group 1. **1**



(ii) Explain **fully** why the second ionisation energy is much greater than the first ionisation energy for Group 1 elements. **2**

(b) Ionisation energy changes across the period.

(i) Explain why the first ionisation energy increases across the period. **1**

(ii) Write an equation, including state symbols, for the second ionisation energy of magnesium. **1**

(iii) The table shows the values for the first four ionisation energies of aluminium.

Ionisation energies (kJ mol^{-1})			
First	Second	Third	Fourth
578	1817	2745	11 577

Explain why there is a large difference between the third and fourth ionisation energies. **1**

Electronegativity

Electronegativity is a measure of the attraction an atom involved in a bond has for the electrons in the bond.

Atoms of different elements have different attractions for bonding electrons (page 12 db). High values indicate that electrons are attracted very strongly to atom.

Trends in Electronegativity across a period

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
Electronegativity							

Describe the trend

Explain the trend

Trends in Electronegativity down a group

Element	**Li**	Na	K	Rb	Cs
Electronegativity					

Describe the trend

Explain the trend

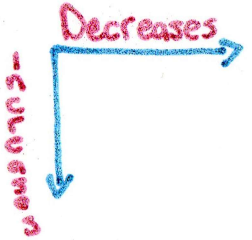
Electronegativity Summary

Across the period the electronegativity **increases** because the ***nuclear charge increases,*** the atom attracts bonding electrons more strongly.

Down the group the electronegativity **decreases** because the atomic size increases as there are more shells of electrons**.** The increase in the number of inner shells of electrons ‘shield’ or ‘screen’ the bonded electrons from the nuclear charge meaning they are less strongly attracted to the atoms resulting in a decrease in electronegativity.

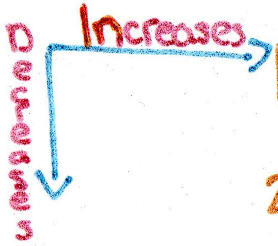
Trends in the Periodic Table

- **Atomic size**



- Across a period the nuclear charge and no. of electrons increase. The electrons are attracted to the nucleus, therefore atomic size decreases.
- Down a group, an extra electron shell is added. This shields the pull of electrons to the nucleus, therefore atomic number increases.

- **Ionisation Energy**



1st: "The amount of energy required to remove 1 mole of electrons from 1 mole of atoms in the gaseous state". $M_{(g)} \longrightarrow M^{+}_{(g)} + e^{-}$ (1st)

2nd: "The amount of energy required to remove 1 mole of electrons from 1 mole of gaseous 1+ ions". $M^{+}_{(g)} \longrightarrow M^{2+}_{(g)} + e^{-}$ (2nd)

- Across a period the nuclear charge increases. The electrons are more strongly attracted to the nucleus => more energy must be supplied to remove an electron, therefore ionisiation energy increases.
- Down a group, electrons are further away from the nucleus. This screening effect shields the attraction to the nucleus, therefore less energy must be supplied to remove an electron. Ionisation energy decreases.

- **Electronegativity**

"Attraction for bonding electrons".

- Across a period the nuclear charge increases, attracting electrons more strongly to the nucleus, therefore electronegativity increases.
- Down a group, the number of electron shells increase. As a result, this shielding increases the distance of outer shell from nucleus, therefore electronegativity decreases.

Structure & Bonding in Compounds

Elements in compounds and molecules are held together in different ways. The properties of chemical compounds are determined by the bonding between atoms and the attractive intermolecular forces between molecules. We will study the following in more detail:

- Pure covalent bonds
- Polar covalent bonds
- Ionic bonds

Pure Covalent Bonds (Non-polar covalent bonds)

- Occurs when electrons are equally shared when there is little or no-difference between the electronegativity of elements. Both atoms have an equal attraction for the shared electrons, they have the same electronegativity.

Examples include diatomic elements such as Hydrogen or fluorine.

H —— H

Polar Covalent Bonds

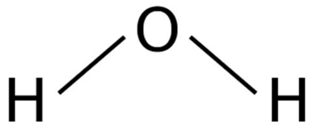
- Polar covalent bonding is formed in compounds between different non-metal elements which have different electronegativities.
- In polar covalent bonds the electrons are not equally shared between the atoms.
- The atom with the highest electronegativity attracts electrons more strongly than the other atom.
- One atom will have a slight negative charge and the other will have a slight positive charge.

For example:

Cl —— H

H-Cl (the electronegativity of Hydrogen is ___, and the electronegativity of Chlorine is ____. This means that the electrons will be pulled closer to _____________.

Partial charges (δ+/δ-) can be added.



Questions

For the following atoms look up the electronegativity values in the databook. Use this to draw the atoms and bond and add in the partial charges (δ+/δ-)

1. H and C

2. Cl and H

3. Cl and F

4. N and C

5. H and N

6. S and N

7. F and Br

Order by polarity using the electronegativity difference

1. HCl
2. HF
3. HBr
4. BrI

Explain the difference between a polar covalent bond and a non-polar covalent bond.

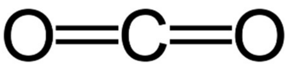
Polar bonds and permanent dipoles

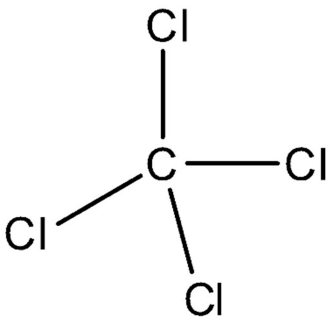
- Polar bonds are formed in molecules when there is a difference in Electronegativity between the atoms.
- When a molecule has an uneven distribution of electrons the molecule has a permanent dipole.

BUT Not all polar covalent bonds create a polar molecule.

- If there is a symmetrical arrangement of polar bonds, the polarity cancels out over the molecule as a whole and is therefore the molecule is non-polar.
- If the molecule has a linear shape with the same element bonded to the central element or has a tetrahedral shape with the same element bonded to the central element. The molecule will be **non-polar.**

Annotate the diagram below to show the partial charge δ+/δ-





Questions

Identify which molecules are polar (draw the structure to help you)

1. CH_3Cl
2. CF_4
3. H_2S
4. CO_2
5. CH_2Cl_2
6. $SiCl_4$

Which element has the greatest attraction for bonding electrons?

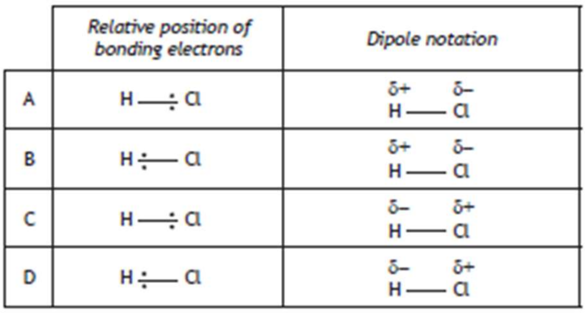
A Lithium

B Chlorine

C Sodium

D Bromine

Which line in the table is correct for the polar covalent bond in hydrogen chloride?



	Relative position of bonding electrons	*Dipole notation*
A	H —÷ Cl	$\overset{\delta+}{H}—\overset{\delta-}{Cl}$
B	H ÷— Cl	$\overset{\delta+}{H}—\overset{\delta-}{Cl}$
C	H —÷ Cl	$\overset{\delta-}{H}—\overset{\delta+}{Cl}$
D	H ÷— Cl	$\overset{\delta-}{H}—\overset{\delta+}{Cl}$

1. Hydrogen will form a non-polar covalent bond with an element that has an electronegativity value of

A 0·9

B 1·5

C 2·2

D 2·5.

2. Which of the following is a polar molecule?

A CCl_4

B NH_3

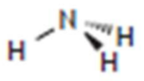
C CO_2

D CH_4

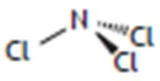
Past Paper Question

Chloramines are less soluble in water than ammonia due to the polarities of the molecules, and so readily escape into the atmosphere, causing irritation to the eyes.

(i) Explain the difference in polarities of ammonia and trichloramine molecules.



ammonia

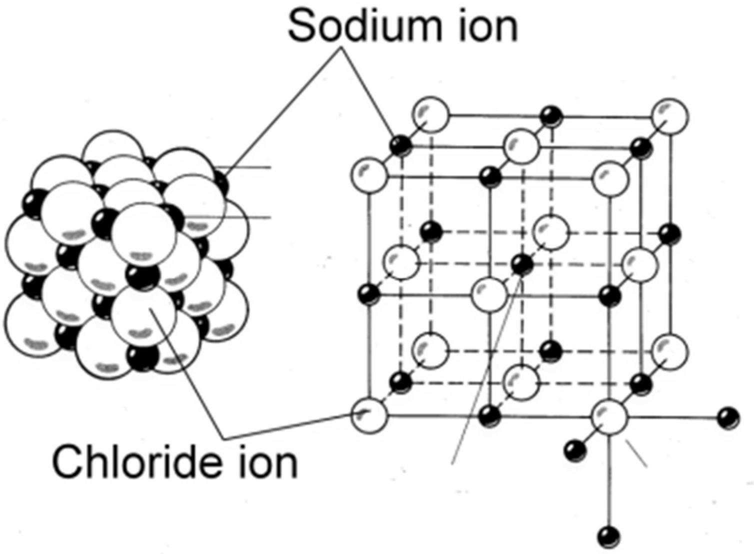


trichloramine

2

Ionic Bonding

- Formed between metal and non-metal elements with a large difference in Electronegativity.
- Electrostatic force of attraction between positively and negatively charged ions.
- Ionic bonding involves the transfer of electrons.
- The non-metal element with the high Electronegativity GAINS electrons to form a negative (-) ion.
- The element with the low Electronegativity LOSES electrons to form a positive (+) ion.
- Forms a regular structure-Lattice.
- Chemical formula gives ratio of atoms.



Properties

Solids at room temperature.

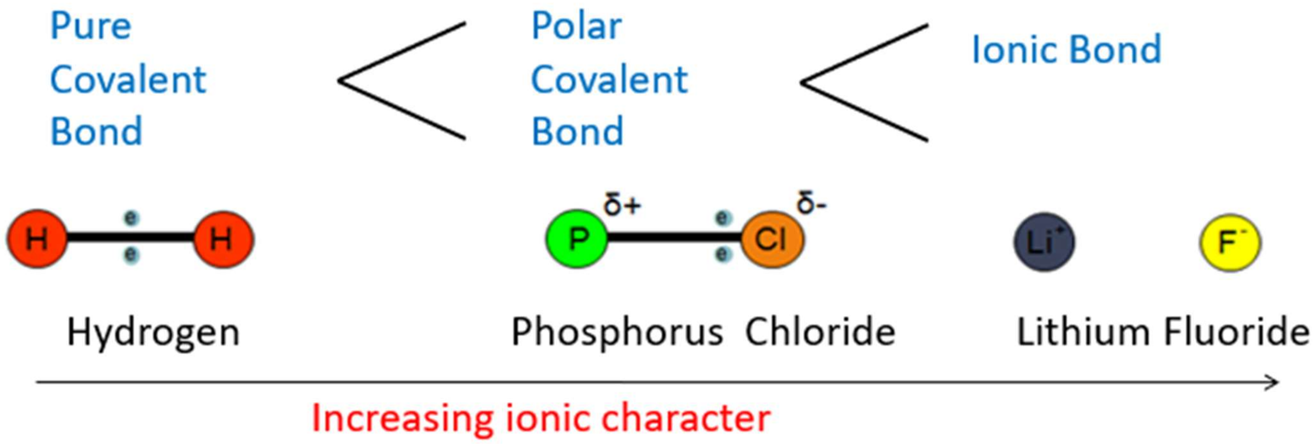
High melting and boiling points (as lots of energy needed to break strong ionic bonds).

Conduct electricity when molten or in solution.

The Bonding Continuum

Labelling a compound as ionic, polar covalent or non-polar covalent is convenient as it helps us understand how elements bond when forming compounds. In reality these categories are quite rigid and do not reflect the wide variety of bonding found in compounds. The concept of a bonding continuum can be used to help us appreciate the differences in bonding.

The bonding continuum places ionic bonding and pure covalent bonding at opposite ends with polar covalent bonding in the middle. The type of bonding present in a compound is determined by the differences in electronegativity between the elements involved. Whilst this provides a fair estimation of the type of bonding, this can only be conclusively determined by examination of the properties of the compound.



Questions

1. Calculate the difference in electronegativity for the elements in the following compounds and indicate which compound will be the most ionic and which will be the least ionic.
   - KBr
   - NaI
   - CsCl

2. The electronegativity of Potassium and Hydrogen are 0.8 and 2.2 respectively. Suggest what kind of bonding Potassium hydride might have.

Intermolecular forces and physical properties

Introduction

Intermolecular forces are the forces that **hold molecules together**. They are crucial in determining the **physical** and **chemical** properties of substances. In this section, we will explore three main types of intermolecular forces**:**

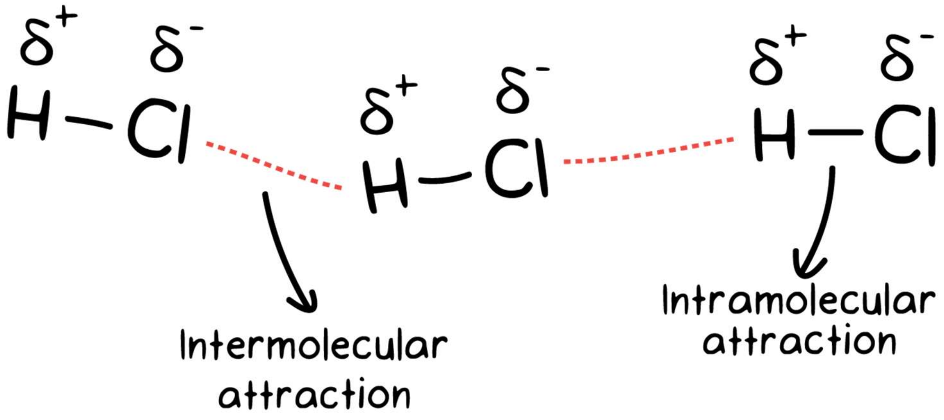
- London Dispersion Forces (LDF),
- Permanent Dipole-Permanent Dipole Interactions (PDPDI)
- Hydrogen Bonding.

These are much weaker forces than intermolecular bonds (covalent, ionic, metallic).

***Intra*molecular bond**	**Strength (kJ mol $^{-1}$)**
Metallic	80 to 600
Ionic	100 to 500
Covalent	100 to 500

***Inter*molecular bond**	**Strength (kJ mol $^{-1}$)**
London's forces	1 to 20
Dipole-Dipole	30
Hydrogen	40

These forces play an important role in determining the physical properties of substances, such as their **boiling** and **melting** points, **volatility**, **viscosity**, and **solubility**.



London Dispersion Forces (LDF)

London Dispersion Forces,are the weakest type of intermolecular force. They occur between all types of molecules and atoms.

These forces arise due to the **temporary uneven distribution** of **electrons** around a molecule, creating an **instantaneous dipole**. This temporary dipole induces a dipole in a neighbouring molecule, and an attractive force develops between them.

Remember:
Molecules with stronger intermolecular forces have higher melting points, boiling points and viscosity. They also have lower volatility

LDF's affect on physical properties

The strength of LDFs increases with the size of the molecules and the **number of electrons present**. The **more electrons** a molecule has, the **stronger** the **London Dispersion Forces**.

Draw a diagram of LDF

Questions

1. What are London Dispersion Forces (LDF), and how do they arise in different types of molecules?

2. State what causes an increase in the strength of LDF?

3. In what way do London Dispersion Forces affect the boiling and melting points of various substances?

4. Compare the London Dispersion Forces in diatomic molecules like nitrogen (N_2) and oxygen (O_2). Which molecule has stronger LDFs and why?

5. How do London Dispersion Forces influence the viscosity of different liquids?

6. What factors determine the strength of LDFs in monoatomic noble gases such as helium (He), neon (Ne), and argon (Ar)?

7. Explain how London Dispersion Forces contribute to the volatility of different substances.

8. How do LDFs in covalent molecular solids like sulfur (S_8) and phosphorus (P_4) compare to those in diatomic and monoatomic molecules?

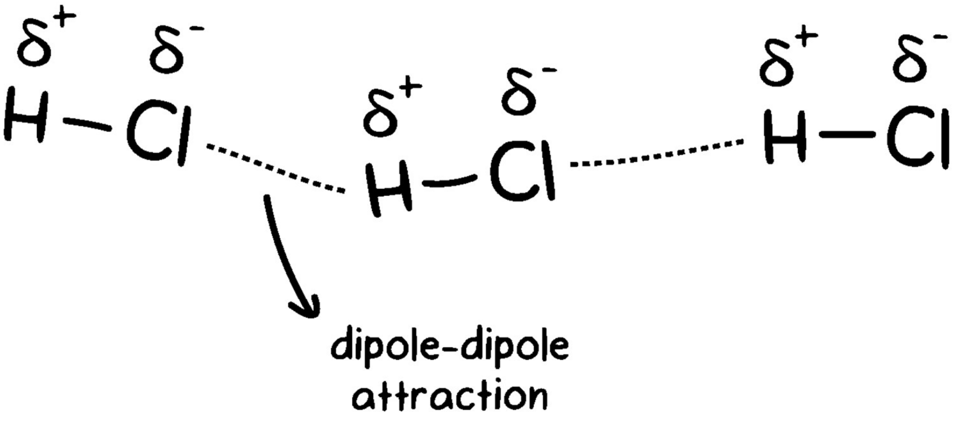
Permanent Dipoles

Permanent Dipole-Permanent Dipole Interactions (PDPDI)

Permanent Dipole-Permanent Dipole Interactions occur between **polar** molecules, which have a permanent uneven distribution of electron density, leading to distinct **positive** and **negative** ends called **dipoles**.

These interactions happen when the positive end of one polar molecule is attracted to the negative end of another polar molecule. The **strength** of **PDPDIs depends** on the **difference in electronegativity** and the distance between the interacting molecules.

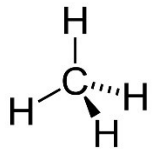
Remember:
Molecules with bonds that have a difference in electronegativity of > 0.4 are polar



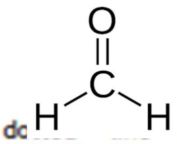
PDPDIs are generally **stronger** than **London Dispersion Forces** but **weaker** than **hydrogen bonding**.

Questions

1. Do these atoms/molecules have Pdp-Pdp or LDF?



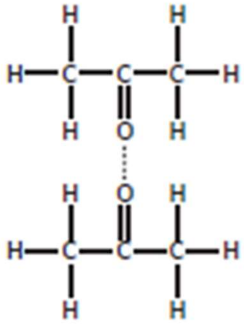
$Cl - H$



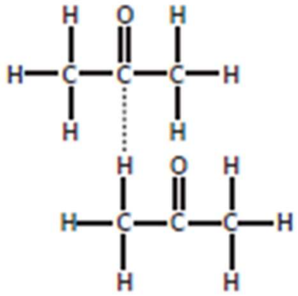
Argon atoms

In he following diagrams does the dc represent a permanent dipole-permanent dipole interaction between propanone molecules?

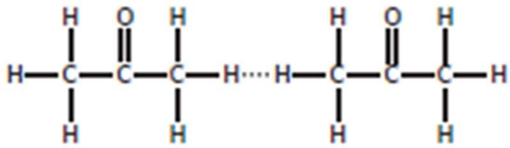
A



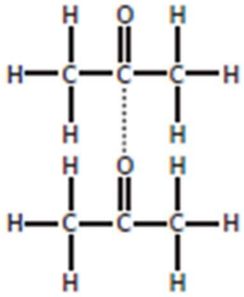
B



C



D



2. Draw the structures of the following molecules and show polarities where appropriate:

   Phosphorus hydride,

   Hydrogen bromide

   Iodine monochloride

Hydrogen Bonding

Hydrogen bonding is a special type of intermolecular force that occurs between molecules containing a hydrogen atom bonded to a highly electronegative atom, such as nitrogen (N), oxygen (O), or fluorine (F). This type of bonding is stronger than both LDF and PDPDI. In a hydrogen bond, the hydrogen atom's positive charge is attracted to the negative charge of the electronegative atom in a neighbouring molecule. Hydrogen bonding plays a crucial role in determining the properties of many substances, including water. It is responsible for water's high boiling point, surface tension, and its unique ability to expand upon freezing.

Draw a diagram to show the hydrogen bonds present in water.

Questions

Identify the molecule which have Hydrogen bonding as an intermolecular force:

H_2S
H_2O
CH_3OH
CH_2O
HF
HCl
NH_3

Intermoleculer forces Summary

Covalent molecules are held together by forces of attraction known as **van der Waals' forces,** these are *inter*molecular forces. There are three main types of van der Waals' forces:

1) *London dispersion forces:*
    - Between all molecules
    - The **only** force in **non-polar** molecules and monatomic elements.
    - The weakest force
2) *Permanent dipole-permanent dipole:*
    - Between **polar** molecules
    - Stronger than London dispersion forces
3) *Hydrogen Bonding:*
    - Between **polar** molecules where an H atom is directly bonded to an N, O or F.
    - Strongest intermolecular force.

Questions

Hodder and Gibson Textbook page 60

Physical Properties

Definitions

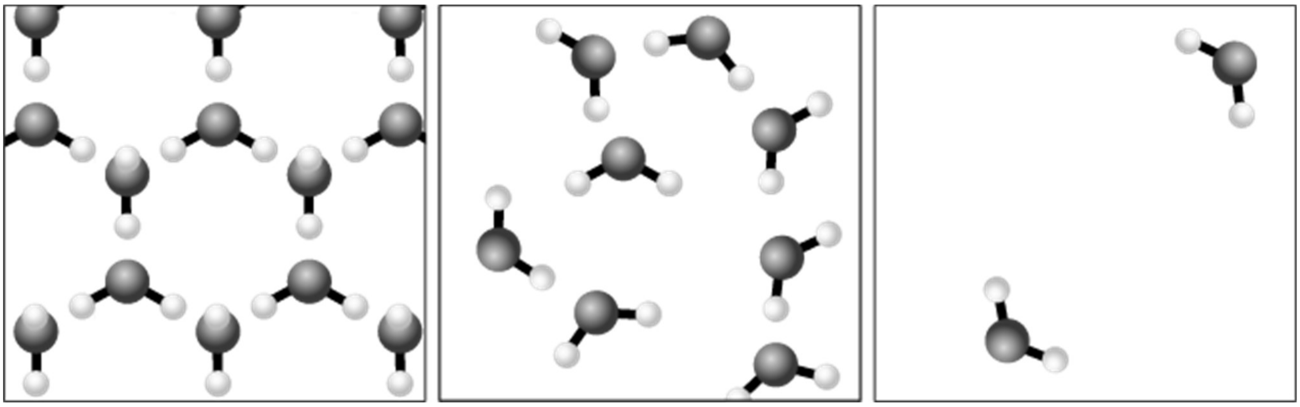
Viscosity: __ Viscosity is a measure of a fluid's resistance to flow.

Volatility: ____ refers to the tendency of a substance to evaporate Viscosity is a measure of a fluid's resistance to flow __________

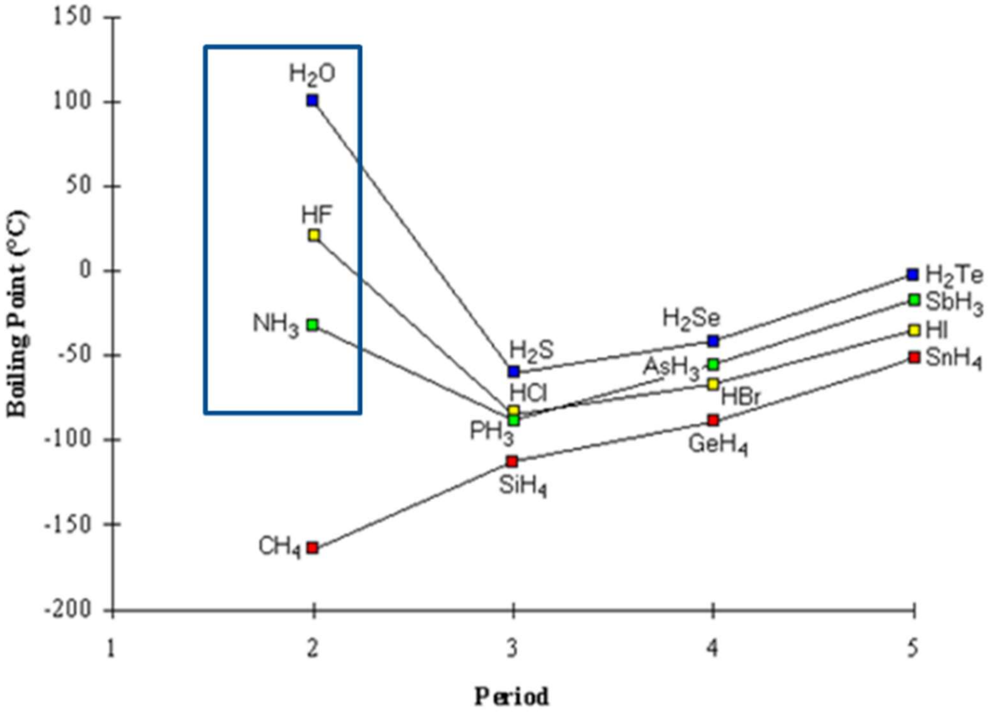
Solubility: _____ Solubility is the ability of a substance to dissolve in a solvent__

Boiling and Melting Points

Boiling and **melting** points are the temperatures at which a substance **transitions** from **one phase** to **another** (liquid to gas or solid to liquid, respectively). The strength of intermolecular forces plays a crucial role in determining these temperatures. Substances with **stronger intermolecular forces require more energy to overcome the forces holding the molecules together**, resulting in higher boiling and melting points. Conversely, substances with weaker intermolecular forces have lower boiling and melting points.



Water has a much higher boiling point than other hydrogen containing compound. This is due to Hydrogen bonding (strongest intermolecular force). The anomalous boiling points of ammonia, water and hydrogen fluoride are a result of hydrogen bonding.

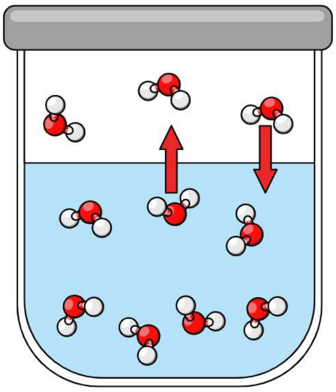


Viscosity

Viscosity is a measure of a fluid's resistance to flow (how thick the liquid is). Substances with **stronger intermolecular forces** tend to **have higher viscosity** because the forces between the molecules create more resistance to flow. Conversely, substances with weaker intermolecular forces exhibit lower viscosity, as the forces between molecules offer less resistance, allowing the fluid to flow more easily.

Volatility

Substances with **weak intermolecular** forces are **more volatile**, as it is **easier** for their molecules to **escape** the **liquid phase** and enter the **gas phase**. On the other hand, substances with **stronger intermolecular** forces are **less volatile** because more energy is needed to overcome the forces holding the molecules together.

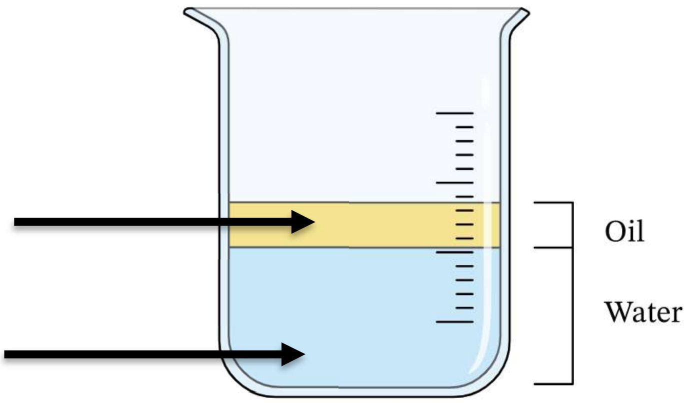


Solubility

The solubility of a substance is influenced by the intermolecular forces between the solute and solvent molecules. In general, substances with **similar types** of **intermolecular forces** are **more likely** to be **soluble** in each other.

Solubility rule: “like dissolves like”

For example, polar solvents like water, which exhibit strong hydrogen bonding, can dissolve polar solutes that also have strong intermolecular forces, while non-polar solvents like hexane are better suited for dissolving nonpolar solutes with weaker LDFs. Polar and non-polar substances **repel each other**.


Oil
Water

Viscosity Investigation

Aim:____

____

Method:

Results: *(graph paper on the following page)*

Conclusion:

____

____

____

____

Evaluation:

____

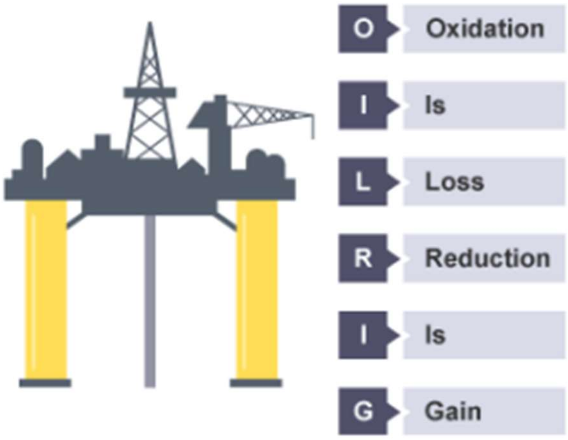
____

Draw a graph using your results.

Oxidising and Reducing Agents

Redox reactions involve a simultaneous oxidation and reduction reaction. This involves the transfer of electrons from one atom/ion to another atom/ion.

N5 Recap



<u>Reduction</u> involves the gain of electrons (RIG):

$Cu^{2+} + 2e^- \rightarrow Cu$

$Ag^+ + e- \rightarrow Ag$

The ion-electron equation for reduction is taken directly from data book (P12)

<u>Oxidation</u> involves the loss of electrons (OIL):

For example:

$Fe \rightarrow Fe^{2+} + 2e^-$

$Mg \rightarrow Mg^2 + 2e^-$

The ion-electron equation for oxidation must be written in reverse.

Writing Oxidation and Reduction Reactions.

Write the ion electron equation for each of the following reactions:

1. oxidation of Magnesium atoms to ions.
2. oxidation of Iron(II) ions
3. oxidation of iodide ions
4. reduction of aluminium ions
5. reduction of chlorine molecules

REDOX reactions

In REDOX reactions, an oxidation and reduction reaction occur at the same time. The combined reaction is called redox.

$\mathrm{Mg(s)} \longrightarrow \mathrm{Mg^{2+}(aq)} + 2e^-$ Oxidation

$\mathrm{Cu^{2+}(aq)} + 2e^- \longrightarrow \mathrm{Cu(s)}$ Reduction

$\mathrm{Mg(s)} + \mathrm{Cu^{2+}(aq)} \longrightarrow \mathrm{Mg^{2+}(aq)} + \mathrm{Cu(s)}$ Redox

The electrons for the oxidation and reductions reactions must be equal before they can be added together to form the REDOX equation. REDOX equations do not show any electrons.

Questions – Writing REDOX equations page 105 Chemcord book

Oxidising and Reducing Agents

A reducing agent causes the reduction of another atom/ion. In doing so it donates electrons (and is itself oxidised).

An oxidising agent causes the oxidation of another atom/ion. In doing so accepts electrons (and is itself reduced).

Identifying an oxidising or reducing agent first requires identification of the reduction and oxidation reactions within a redox equation.

Example 1: slide 21

$\mathrm{Mg\ (s)} + \mathrm{Zn^{2+}\ (aq)} \longrightarrow \mathrm{Mg^{2+}\ (aq)} + \mathrm{Zn\ (s)}$ REDOX

$\mathrm{Mg} \longrightarrow \mathrm{Mg^{2+}}$ Oxidation

$\mathrm{Zn^{2+}} \longrightarrow \mathrm{Zn}$ Reduction

Example 2

$\mathrm{Cu(s)} + 2\mathrm{Ag^{+}(aq)} + 2\mathrm{NO_3^{-}(aq)} \rightarrow 2\mathrm{Ag(s)} + \mathrm{Cu^{2+}(aq)} + 2\mathrm{NO_3^{-}(aq)}$

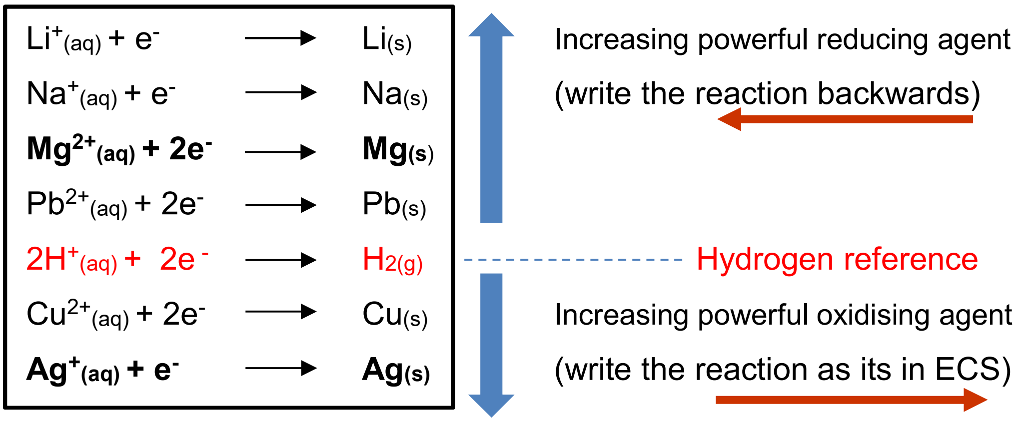
1. Identify any spectator ions
2. Identify oxidation reaction
3. Identify reduction reaction

Questions:

Electrochemical Series

- The electrochemical series (ECS) can be used to indicate the ability of an element or compound to act as an oxidising and reducing agents.
- The electrochemical series is a series of reduction reactions. (data booklet page 12).

- Elements and compounds at the top right of the ECS are strong reducing agents (donates electrons) – Label your ECS in your data book
- Elements and compounds at the bottom left of the ECS are strong oxidising agents (accepts electrons)- Label your ECS in your data book



$Li^{+}_{(aq)} + e^{-} \longrightarrow Li_{(s)}$

$Na^{+}_{(aq)} + e^{-} \longrightarrow Na_{(s)}$

$\mathbf{Mg^{2+}_{(aq)} + 2e^{-} \longrightarrow Mg_{(s)}}$

$Pb^{2+}_{(aq)} + 2e^{-} \longrightarrow Pb_{(s)}$

$2H^{+}_{(aq)} + 2e^{-} \longrightarrow H_{2(g)}$

$Cu^{2+}_{(aq)} + 2e^{-} \longrightarrow Cu_{(s)}$

$\mathbf{Ag^{+}_{(aq)} + e^{-} \longrightarrow Ag_{(s)}}$

Consider the two ion-equations,

1) $Mg^{2+}_{(aq)} + 2e^{-} \rightarrow Mg_{(s)}$ 2) $Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$

Magnesium is higher up the electrochemical series, and act as the **reducing agent.** (i.e. Magnesium is oxidised itself, the ion-electron equation would be written backwards).
Whilst Silver is at a lower position in the electrochemical series and acts as the **oxidising agent** (i.e. Silver ions are reduced themselves, the ion electron equation would be written as it is in the ECS).

Questions

1.Decide if the elements/ions below would be strong oxidising or reducing agents by looking in the ECS

K		Cl_2	
MnO_4^-		Li	

2) Which would be a stronger reducing agent:

a) Cs or Na b) Mg or Fe

3) Which would be a stronger oxidising agent:

a) Br_2 or Cl_2 b) $Cr_2O_7^{2-}$ or MnO_4^-

Electronegativity & Oxidising/Reducing Agents

Elements with a high electronegativity are likely to act as _______agents as they tend to form ions by_______ electrons. These tend to be non-metals.

Example: Fluorine when it reacts will ________ electrons to become a fluoride ion (F^-)

Elements with low electronegativity tend to form ions by _______________ electrons so act as _____________ agents. These tend to be metals.

The strongest reducing agents are found in group ___ (the ________ __________) whilst the strongest oxidising agents come from group ____ (the ________).

- Non-metals tend to be **oxidising** agents as they have **high** electronegativities and **gain** electrons and form **negatively-charged** ions.
- Metals tend to be **reducing** agents as they have low electronegativities and **lose** electrons and form **positively-charged** ions.

Compounds as Oxidising and reducing Agents.

Compounds, group ions and molecules can also act as oxidising and reducing agents

Compounds as oxidising agents

Dichromate ($Cr_2O_7^{2-}$) – in acidic solutions (H^+)

Permanganate (MnO_4^-) – in acidic solutions (H^+)

Hydrogen peroxide (H_2O_2)

Reducing agents

Carbon monoxide (CO) is a gas which is used in the blast furnace and reduces Iron (III) ions to Iron atoms. $Fe^{3+} + 3e^{-} \rightarrow Fe$

Everyday Uses for Strong Oxidising agents

- Oxidising agents are used because of the effectiveness with which they can kill fungi and bacteria, and can inactivate viruses.
- Potassium permanganate uses to treat the fungal infection athletes' foot and is used to prevent disease in fish in ponds.
- Hydrogen peroxide can break down coloured compounds, ideal for use as 'bleach' for clothes and hair.

Complex Ion Equations

Example 1

Write the ion-electron equation for $\mathbf{SO_3^{2-} \rightarrow SO_4^{2-}}$

Step 1: Balance main element, $\mathbf{SO_3^{2-} \rightarrow SO_4^{2-}}$

Step 2: Balance oxygen using H_2O, $\mathbf{SO_3^{2-} + H_2O \rightarrow SO_4^{2-}}$

Step 3: Balance hydrogen using H^+ ions, $\mathbf{SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+}$

Step 4: Balance charge using electrons on same side as H^+ $\mathbf{SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-}$

Example 2 Write the ion-electron equation for $\mathbf{MnO_4^- \rightarrow Mn^{2+}}$

Questions

Write the balanced ion electron equation for each of the following reactions and state if its an oxidation or reduction.

1. $IO_3 \longrightarrow I_2$

2. $PbO_2 \longrightarrow Pb^{2+}$

3. $XeO_3 \longrightarrow Xe$

4. $ClO_3^- \longrightarrow Cl^-$

2. $FeO_4^{2-} \longrightarrow Fe^{3+(}aq)$

3. $Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$

Combining two complex ion-equations

Example 1:

Acidified potassium permanganate solution reacting with iron (II) sulfate solution.

K^+ and SO_4^{2-} are spectator ions

1. Write the ion-electron equation for $MnO_4^- \rightarrow Mn^{2+}$
2. Write the combined redox equation using your equation calculated above and: $Fe^{2+} \rightarrow Fe^{3+} + e^-$

Keyword	**Meaning**
Bonding continuum	A concept applied to bonding. Ionic and pure covalent bonds lie at opposite ends of the boding continuum with polar covalent bonding in between.
Covalent bond	A shared pair of electrons attracted to two positive nuclei.
Covalent molecule	Small number of atoms covalently bonded to each other.
Covalent network	Large number of atoms covalently bonded in a large structure (examples include diamond (C), SiO_2.
Covalent radius	A measure of the size of an atom.
Delocalised electrons	Electrons that are mobile and free to move away from the atom that they came from.
Electronegativity	Is a measure of the attraction an atom in a bond has for electrons in a bond.
Groups	Vertical column of elements in the periodic table. Elements in the same group have the same number of outer electrons and have similar chemical properties.
Hydrogen bonding	An intermolecular force that occurs in polar molecules which contain N-H, F-H and O-H. The strongest of the 3 intermolecular forces.
Intermolecular force	Attractive forces that occur between molecules also called Van der Waals forces (3 different types; LDF, Hydrogen bonding, permanent dipole-permanent dipole interaction).
Ionic bond	An electrostatic attraction between a positive ion and a negative ion.
Ionisation energy	The energy required to remove one mole of electrons from one mole of gaseous atoms.
Lattice	The structure formed in ionic substances.
London dispersion forces	The weakest type of Van der Waals force. Occurs in all atoms
Metallic bonding	Positive metal ions surrounded by a ‘sea’ of delocalised electrons.
Miscible	Liquids that form a homogeneous mixture.
Non- polar covalent bond (also called pure covalent bond)	A covalent bond where the atoms in the bond are shared equally between the atoms due to the atoms having the same electronegativity.

Keyword	Meaning
Nuclear charge	The charge on the nucleus of an atom that is dependent on the number of protons.
Oxidation reaction	A reaction where electrons are lost.
Oxidising agent	A substance that causes the oxidation of another atom/molecule.
Periods	Rows of elements arranged with increasing atomic number.
Permanent dipole	Formed in a molecule where there is a large difference in electronegativity, resulting in electrons in the bond being unequally shared.
Permanent dipole-permanent dipole	An intermolecular force between polar molecules.
Polar covalent bond	A covalent bond where the electrons are **not** equally shared between the atoms. The bond has δ+ and δ- side as a result.
Pure covalent bond (also called non-polar covalent bond)	A covalent bond where the atoms in the bond are shared equally between the atoms due to the atoms having the same electronegativity.
Polar molecule	A molecule that has a permanent dipole (uneven distribution of electrons).
Reducing agent	A substance that causes the reduction of another atom/molecule.
Reduction reaction	A reaction that involves the gain of electrons.
Shielding effect	Inner (core) electron shells will **shield** the outer electrons from the nucleus. Therefore, the full nuclear charge (Z) cannot be felt by the outer electron and will decrease to an effective nuclear charge.
Temporary dipole	A temporary uneven distribution of electrons, induces dipole in
Unsaturated	Molecule containing a double or triple carbon to carbon bond.
Viscosity	The thickness of a liquid.