Chemistry Paper 1 (Combined Higher)

CONTENTS

1. Atoms, Elements, Compounds and Mixtures

Keyword	Definition
Atom	smallest part of an element
Element	made up of only one type of atom
Compound	made from at least two elements, chemically combined
Mixture	made of two or more elements or compounds not chemically combined together

Radius of an atom = 0.1nm (1 x 10^{-10}m).

Radius of a nucleus is less than 1/10 000 of that of an atom.
This is 1 x 10^{-14}m.

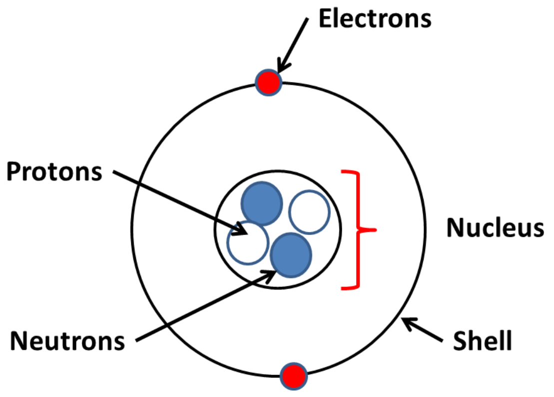
Atoms are **neutral** (no electrical charge) because:
-The number of protons and electrons are the same.
-The charges cancel out

Atomic number = Proton number

Mass number = Number of protons and neutrons

Number of electrons = Number of protons

Structure of the atom (Nuclear model)



Subatomic particle	Relative charge	Relative mass
Proton	+1	1
Neutron	0	1
Electron	-1	1/1840

2. Structure of the Atom

7	**Top number**
Li	
3	**Bottom number**

Proton = bottom number

Electron = bottom number

Neutron = top number – bottom number

Electronic Configuration

Electrons are arranged in shells.

1st shell – maximum of 2 electrons

2nd shell – maximum of 8 electrons

3rd shell – maximum of 8 electrons

Isotopes:

Atoms of the same element that have different numbers of neutrons but the same number of protons and electrons.

They have the same chemical properties but different physical properties.

$^{39}_{18}\mathrm{Ar}$	$^{38}_{18}\mathrm{Ar}$
18 protons 18 electrons 21 neutrons	18 protons 18 electrons 20 neutrons

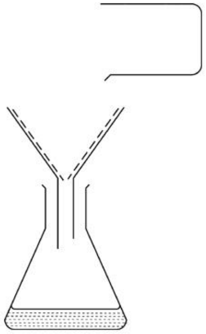
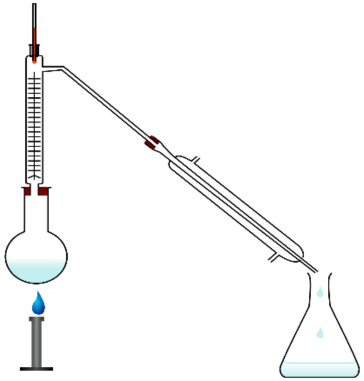
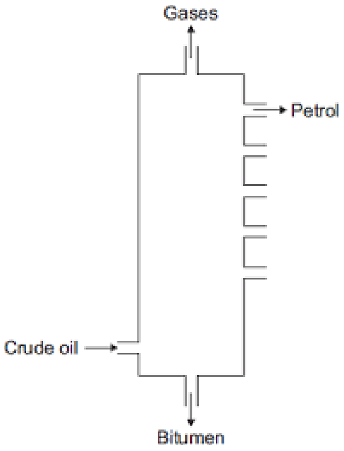
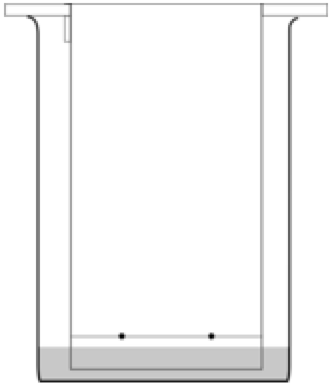
Calculating Relative Isotopic Abundance

Mass number	Abundance (%)
39	93.1
41	6.9

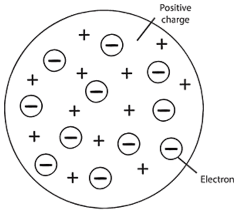
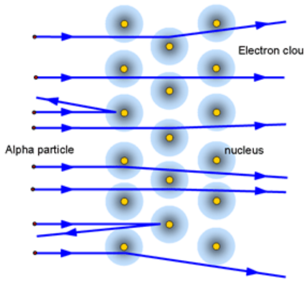
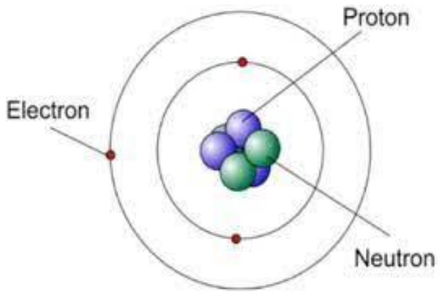
$$= \frac{(39 \times 93.1) + (41 \times 6.9)}{93.1 + 6.9}$$

$$= 39.1$$

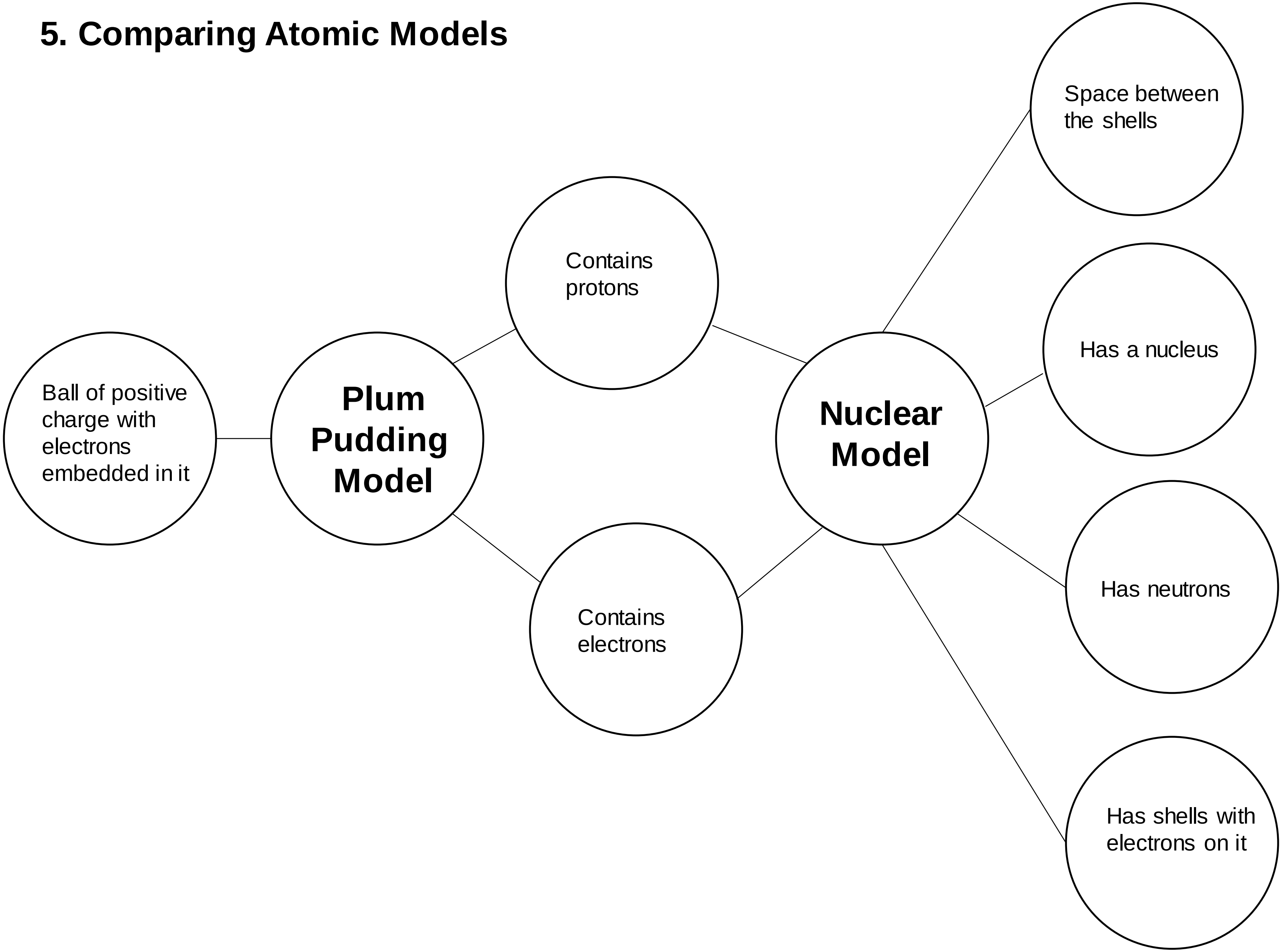
3. Separating Mixtures

Process	**Filtration**	**Distillation**	**Fractional distillation**	**Chromatography**
Diagram				
Physical property	Difference in solubility	Difference in boiling points	Difference in boiling points	Difference in solubility
Example	Sand and salt	Ink and water	Ink, water and oil	Different colours in dyes

4. History of the Atom

Atomic model	**Plum pudding model**		**Nuclear model**		
Diagram	 Plum pudding model				
Discovery	Electron	Positive nucleus in the centre of the atom	Electrons occupy shells Electrons are at specific distances from the nucleus	Neutrons	• Atomic radius: 1×10^{-10} m • Radius of a nucleus is less than 1/10 000 of the radius of an atom. • Most of the mass of an atom is concentrated in the nucleus. • The electrons are arranged at different distances from the nucleus.
Description	The atom is a ball of positive charge with negative electrons embedded in it.	Positively charged alpha particles were fired at thin gold foil. Most alpha particles went straight through the foil. A few were scattered in different directions by the atoms in the foil. It showed that the mass of an atom was in the centre (the nucleus) and the nucleus was positively charged.		Proved the existence of isotopes	
Discovered by	Thompson	Rutherford	Bohr	Chadwick	

5. Comparing Atomic Models



Ball of positive charge with electrons embedded in it

Plum Pudding Model

Contains protons

Contains electrons

Nuclear Model

Space between the shells

Has a nucleus

Has neutrons

Has shells with electrons on it

6. Ionic and Covalent Bonding

Ionic Bonding (metal & non-metal)

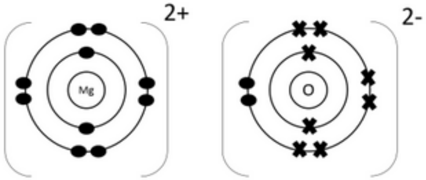
Structure: Giant ionic lattice

Electrons are lost or gained to achieve a full outer shell.

Ionic bond: Electrostatic attraction between oppositely charged ions.

Ions held in a fixed lattice.

Charge of ion: +2 (loses 2 electrons) and -2 (gains 2 electrons)



Describing the formation of an ionic compound

Example 1: NaF

Na atom loses 1 electron to form Na^{1+} ion.

F atom gains 1 electron to form F^{1-} ion

Example 2: Na_2O

Two Na atoms each lose 1 electron to form two Na^{1+} ions.

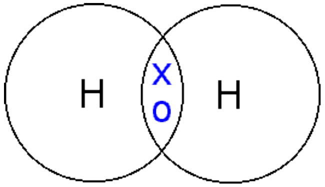
One O atom gains 2 electrons to form O^{2-} ion.

Covalent Bonding (2 x non-metals)

Covalent bond: Pairs of electrons are shared between the atoms.

Sharing one pair of electrons = single bond

Sharing two pairs of electrons = double bond



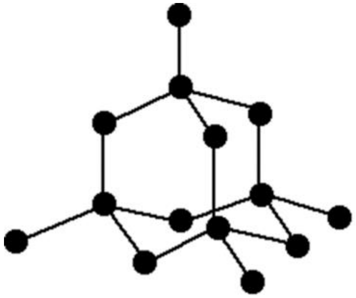
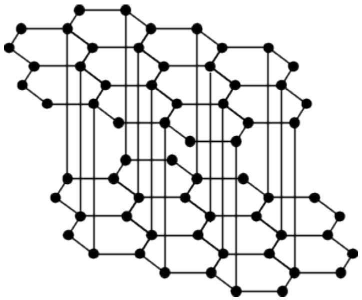
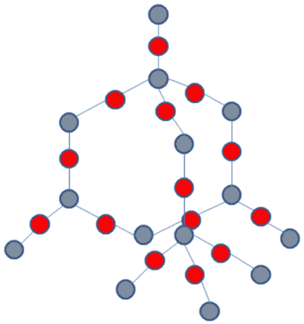
Simple Molecules

(2 x non-metals, covalent bonding)

Simple molecules (small molecules)

e.g. H_2, Cl_2, O_2, N_2, HC*l*, H_2O

7. Giant Covalent Bonding

	Diamond	**Graphite**	**Silicon dioxide**
Bonding	Giant covalent	Giant covalent	Giant covalent
Made of	Carbon	Carbon	Silicon and oxygen
Structure	Each carbon atom forms four C-C covalent bonds.	Each carbon atom forms three covalent bonds with three other carbon atoms, forming layers of hexagonal rings . The 4th electron is delocalised	Each silicon atom forms four covalent bonds with oxygen atoms
Diagram			

8. Metallic Bonding and Alloys

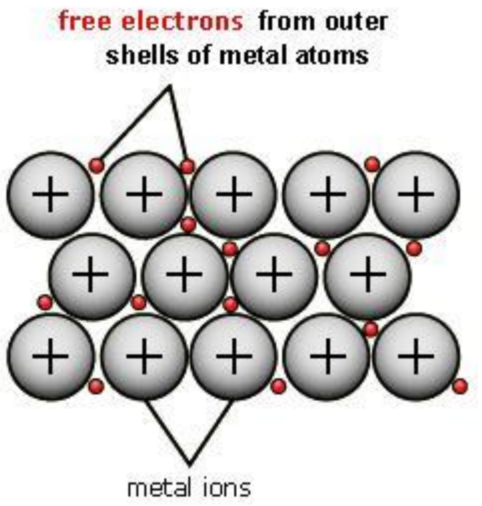
Metallic Bonding

Metallic bond: Attraction

between the positive metal ion and delocalised electrons.

Structure: Layers of metal positive ions surrounded by

delocalised electrons

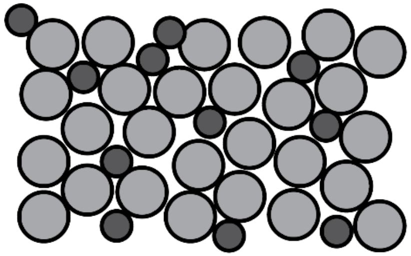


Alloy

Mixtures of metals with metals or a non-metal

e.g. stainless steel is a mixture of iron and carbon

Structure: Irregular layers



9. Quantitative Chemistry

Relative formula mass (RFM or M_r)

This is the mass in grams of 1 mole of the substance.

To calculate M_r (top number) you need to add up the atomic mass (Ar) of all of the atoms in the molecule.

Example1. $NaCl = Na + Cl = 23 + 35.5 = 58.5$

Example 2. $MgF_2 = Mg + (2 x F) = 24 + (2 x 19) = 62$

% Mass of an Element in a compound

$$\text{\% mass of an element} = \frac{\text{Atomic mass of element x number of atoms}}{\text{Relative formula mass of compound}} \text{ X } 100$$

Remember: $\frac{\text{part}}{\text{whole}}$ x 100

Conservation of Mass

During a chemical reaction, no atoms are made, no atoms are destroyed.

Decrease in mass:

$$CaCO_3 (s) \longrightarrow CaO (s) + CO_2 (g)$$

Carbon dioxide is a gas which is a product

Carbon dioxide escapes into the air.

Increase in mass:

$$2Mg(s) + O_2 (g) \longrightarrow 2MgO (s)$$

Mg reacts with oxygen in the air

Oxygen has added to the magnesium

Concentration of a solution

$$\begin{tikzcd} dm^3 \arrow[r, shift left, "\text{X } 1000"] & cm^3 \arrow[l, shift left, "\div 1000"] \end{tikzcd}$$

Concentration (g/dm^3) = mass (g) ÷ volume (dm^3)

10. Acids and Alkalis

Acid	Chemical formula
Sulfuric acid	H_2SO_4
Nitric acid	HNO_3
Hydrochloric acid	HCl

Alkali	Chemical formula
Sodium hydroxide	$NaOH$
Potassium hydroxide	KOH

Acid	Salt name ending
Hydrochloric	-chloride
Nitric acid	-nitrate
Sulfuric	-sulfate

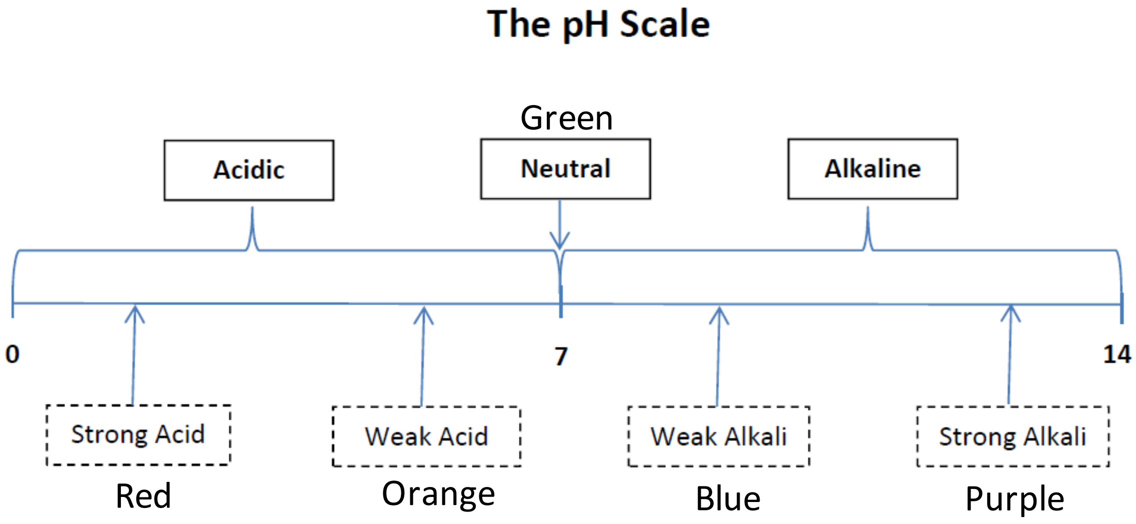
The pH Scale

It can be measured with a pH probe, or universal indicator.

Acid: pH 0-6

Neutral: pH 7

Alkali: pH 8-14



Neutralisation

Acids contain hydrogen ions (H^+)

Alkalis contain hydroxide ions (OH^-)

acid + alkali → water

Ionic equation: $H^+ (aq) + OH^- (aq) \rightarrow H_2O (l)$

11. Reactions of Acids to Make a Salt (Neutralisation)

Reaction 1	**Reactions of Acids with Metals (Neutralisation)**
Rule	acid + metal → salt + hydrogen
Example	hydrochloric acid + magnesium → magnesium chloride + hydrogen

Reaction 2	**Reactions of Acids with Metal Oxide (Neutralisation)**
Rule	acid + metal oxide → salt + water
Example	sulfuric acid + magnesium oxide → magnesium sulfate + water

Reaction 3	**Reactions of Acids with Metal Hydroxide (Neutralisation)**
Rule	acid + metal hydroxide → salt + water
Example	nitric acid + magnesium hydroxide → magnesium nitrate + water

Reaction 4	**Reactions of Acids with Metal Carbonate (Neutralisation)**
Rule	acid + metal carbonate → salt + water + carbon dioxide
Example	nitric acid + magnesium carbonate → magnesium nitrate + water + carbon dioxide

12. Strong and Weak Acids

Strong acid

Completely ionised (breaks down) in aqueous solution.

$HCl \rightarrow H^+ + Cl^-$

***Examples**:* Hydrochloric acid (HCl), nitric acid (HNO_3) and sulfuric acid (H_2SO_4).

Lower pH numbers (pH 1-3)

The stronger the acid, the more it ionises in solution, and the more hydrogen ions there are in the solution.

Concentrated acid

More hydrogen ions (H^+) per volume

Weak acid

Partially ionised (breaks down) in aqueous solution.

$CH_3COOH \rightarrow CH_3COO^- + H^+$

Examples: Ethanoic acid, citric acid and carbonic acid.

Higher pH numbers (pH 4-6)

pH

If the hydrogen ion concentration in a solution increases by a factor of 10, the pH of the solution decreases by 1.

Volume of acid (cm^3)	pH
10	3
1000	5

13. Energy Changes

Exothermic Reaction. Energy is transferred from particles to the surroundings. Temperature increases.

Examples: Combustion, many oxidation reactions, neutralisation.

Every day uses: self-heating cans and hand warmers.

Endothermic reaction. Energy is transferred from the surroundings to the particles. Temperature decreases.

Example: Thermal decomposition and the reaction between citric acid and sodium hydrogencarbonate.

Every day uses: sports injury packs.

Activation energy: minimum amount of energy required for the reaction to start.

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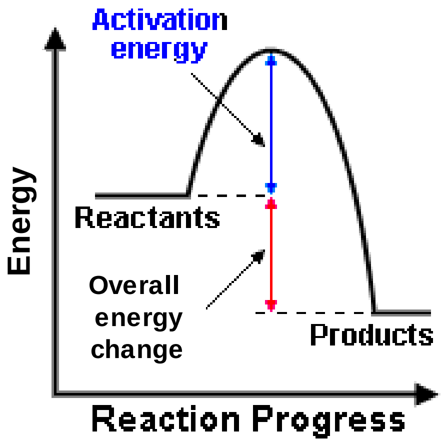
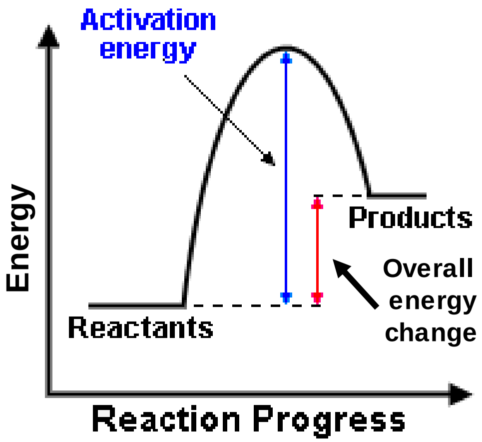


Exothermic reaction



Endothermic reaction

Exothermic energy profile:

Reactants are **higher** in energy than the products.

Energy is released to the surroundings.

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Endothermic energy profile:

Reactants are **lower** in energy than the products.

Energy is absorbed by the surroundings.

.

14. Calculating Bond Enthalpy

Energy IN
Total energy
needed to break
the bonds in
the reactants

—

Energy OUT
Total energy needed
to form the bonds in
the products

=

Energy CHANGE
Overall energy change

Exothermic reaction.

Negative value

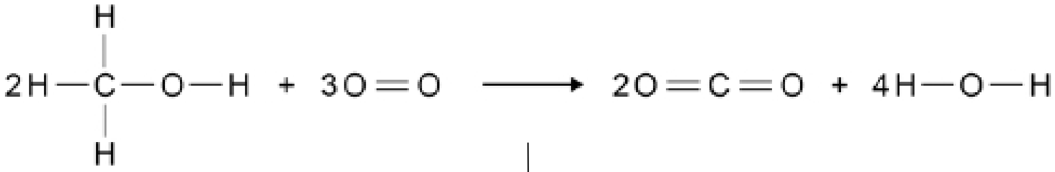
Total energy needed to break the bonds in the reactants ≪ Total energy needed to form the bonds in the products

Endothermic reaction.

Positive value.

Total energy needed to break the bonds in the reactants ≫ Total energy needed to form the bonds in the products

	C–H	**C–O**	**O–H**	**O=O**	**C=O**
Bond energy in kJ / mol	412	360	464	498	805



$$2\,H-\underset{H}{\overset{H}{C}}-O-H + 3\,O=O \longrightarrow 2\,O=C=O + 4\,H-O-H$$

6 x (C-H) = 6 x 412 = 2472
2 x (C-O) = 2 x 360 = 720
2 x (O-H) = 2 x 464 = 928
3 x (O=O) = 3 x 498 = 1494
TOTAL = 5614

4 x (C=O) = 4 x 805 = 3220
8 x (O-H) = 8 x 464 = 3712
TOTAL = 6932

5614 – 6932= -1318

15. The Development of the Periodic Table

Newland's Periodic Table	**Similarities**	**Mendeleev's Periodic Table**
Included only the elements known at the time	Ordered elements by atomic weight	Left gaps for elements he predicted would be discovered later
Maintained a strict order of atomic weights	Missing noble gases	Swapped the order of some elements if that fitted their properties better e.g. Te and I
Every eighth element had similar properties		Elements in groups had similar properties
Was criticised by other scientists for grouping some elements with others when they were obviously very different to each other		Was seen as a curiosity to begin with by other scientists, but then as a useful tool when the predicted elements were discovered later
Mendeleev's version was **accepted** because the newly discovered elements fitted in these gaps. The properties of the elements were predicted correctly.		

Modern Periodic Table

It is called a **Periodic Table** because similar properties occur at regular intervals

Elements arranged in order of atomic number (proton number)

Groups (columns): Elements with similar chemical properties

Group number = number of outer shell electrons = similar chemical properties

Period (row): Elements have the same number of shells

16. Chemical Formulae

Group number	Charge of ion formed
1	+1
2	+2
3	+3
5	-3
6	-2
7	-1

Name of ion	Chemical formula of ion
Sulfate	SO_4^{2-}
Hydroxide	OH^{1-}
Ammonium	NH_4^{1+}
Nitrate	NO_3^{1-}
Carbonate	CO_3^{2-}

Chemical Formulae

NaCl – 1 x Na atom and 1 x Cl atom

H_2O – 2 x H atoms and 1 x O atom

$Mg(OH)_2$ – 1 x Mg atom, 2 x O atoms and 2 x H atoms

$CaCO_3$ – 1 x Ca atom, 1 x C atom and 3 x O atoms

How to deduce chemical formulae

	Mg^{2+}	**Br^{1-}**
Identify the number	2	1
Swap the numbers	1	2
Chemical formula	$MgBr_2$	

	NH_4^{1+}	**SO_4^{2-}**
Identify the number	1	2
Swap the numbers	2	1
Chemical formula	$(NH_4)_2SO_4$	

17. Reactions of Group 1, Group 7 and Group 0

	Group 1	**Group 7**	**Group 0**
Name	Alkali Metals	Halogens (non-metal)	Noble gases
Reactivity	Increases down the group	Decreases down the group	Unreactive (inert). Does not form ions or molecules
Reactivity explanation	The outer electron is further from the nucleus. There is less attraction between the nucleus and the outer electron. The atom loses an electron more easily.	In fluorine, outer shell greater attraction between the nucleus and the outer shell, easier to gain an electron.	Already has a full outer shell of 8 electrons (except helium which has 2). No need to react.
Trend in melting point	Decreases	Increases	Increases
Explanation for trend in melting point		Mass increases. Stronger intermolecular forces. More energy is required to break these forces	Mass increases. More energy is required
Reactions	**Reaction with oxygen:** $4M + O_2 \rightarrow 2M_2O$ Forms the metal oxide (M_2O)	**Displacement:** A more reactive halogen can displace a less reactive halogen from its salt e.g. $2KBr + Cl_2 \rightarrow 2KCl + Br_2$ Chlorine more reactive than bromine. Displacement occurs. $2KBr + I_2 \rightarrow$ no reaction Iodine cannot displace bromine	
	Reaction with chlorine: $2M + Cl_2 \rightarrow 2MCl$ Forms the metal chloride (MCl) Vigorous reaction Na = silver solid; Cl_2 = green gas Reaction = orange flame Product = white solid NaCl produced		
	Reaction with water: $2M + 2H_2O \rightarrow 2MOH + H_2$ Hydroxide ions (OH^-) make solutions alkali. Metal floats and moves. Effervescence.		

18. Reactivity of Metals

Oxidation and Reduction (adding and losing oxygen)

Oxidation: When the metal gains oxygen to become a metal oxide.

Reduction: When the metal oxide loses oxygen to become a metal.

The Reactivity Series

Potassium
Sodium
Calcium
Magnesium
Aluminium

Extracted by electrolysis of a molten ionic compound

Carbon
Zinc
Iron
Tin
Lead
Hydrogen
Copper

Extracted from its oxide by reduction using carbon

Silver
Gold
Platinum – least reactive

Extraction of metals

Metals above carbon in the reactivity series: Extracted by electrolysis

Metals below carbon: Extracted from their oxides by reduction with carbon.

iron oxide + carbon → iron + carbon dioxide

The iron has been reduced – it has lost oxygen. The carbon has been oxidised.

Silver, gold and platinum: Found in the Earth as the metal itself because they are unreactive.

Oxidation and Reduction (adding and losing electrons)

Oxidation: Loss of electrons.

Reduction: Gain of electrons.

Remember **OIL RIG**

For example:

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

The iron ion gains two electrons and becomes an iron atom.

The iron has been reduced – it has gained two electrons.

19. Properties of Ionic Compounds and Simple Molecules

Property of Ionic Compounds	Explanation
High melting point	Giant ionic structure. Lots of energy needed to break strong electrostatic attraction between ions.
Conducts electricity in solution/molten	Ions are mobile and carry charge.
Does not conduct electricity as a solid	Ions are in a fixed lattice. Ions are not mobile so cannot carry a charge

Simple Molecules

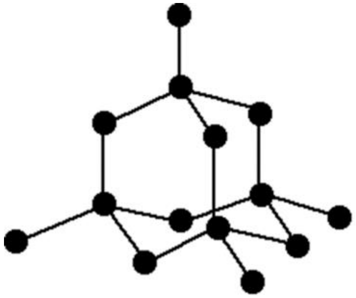
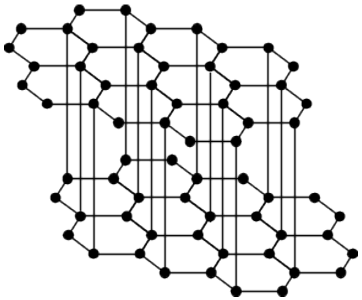
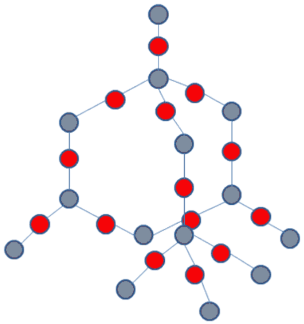
(2 x non-metals, covalent bonding)

Simple molecules (small molecules)

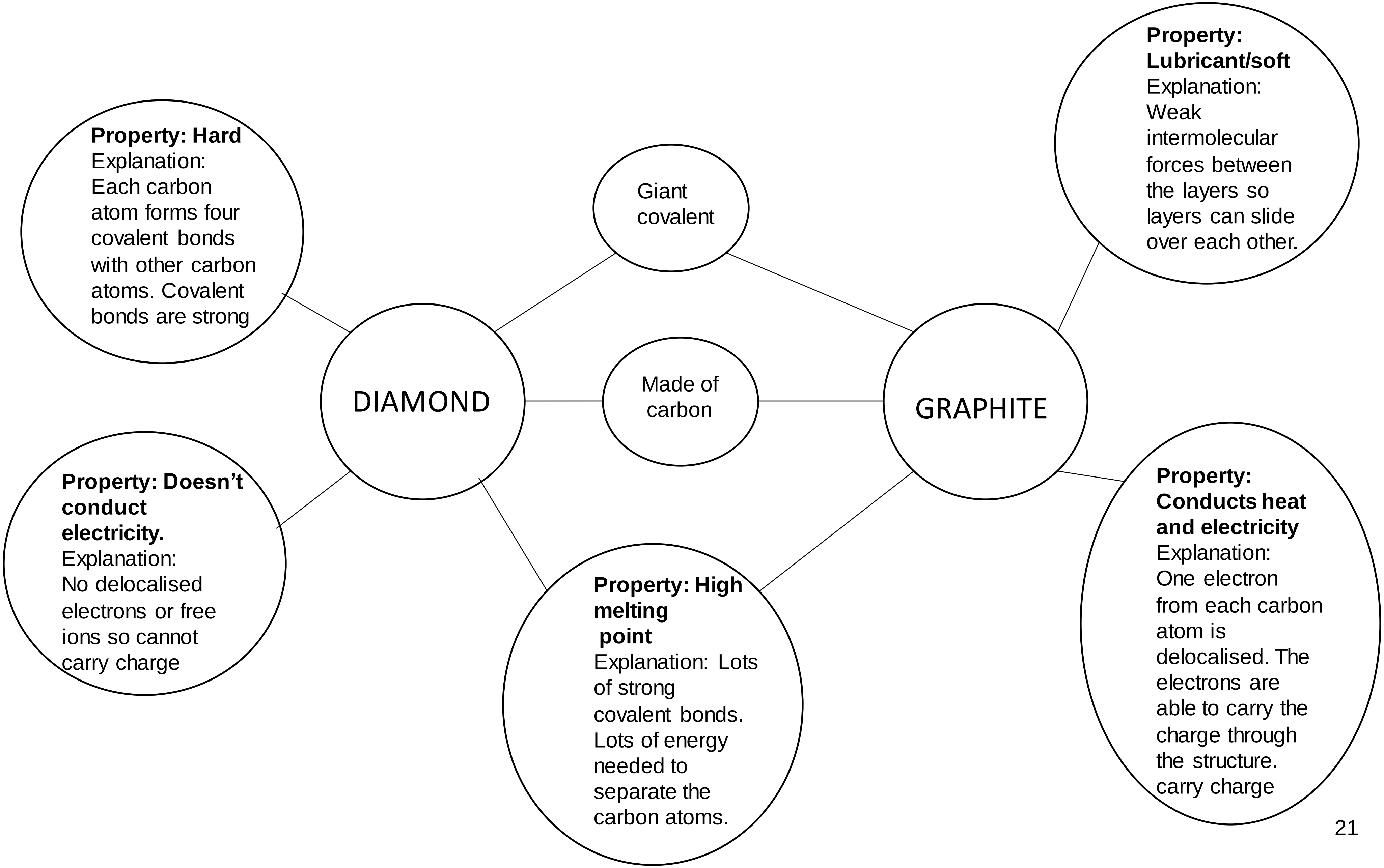
e.g. H_2, F_2, Cl_2, O_2, N_2, HCl, H_2O, CO_2

Property of Simple Molecules	Explanation
Low melting points and boiling points. **(Gas at room temperature)**	-Simple molecule -Weak intermolecular forces between the molecules. -Little energy needed to overcome these forces.
Does not conduct electricity	Molecules do not have any mobile ions or delocalised electrons

20. Structure of Giant Covalent Substances

	Diamond	**Graphite**	**Silicon dioxide**
Bonding	Giant covalent	Giant covalent	Giant covalent
Made of	Carbon	Carbon	Silicon and oxygen
Structure	Each carbon atom forms four C-C covalent bonds.	Each carbon atom forms three covalent bonds with three other carbon atoms, forming layers of hexagonal rings . The 4th electron is delocalised	Each silicon atom forms four covalent bonds with oxygen atoms
Diagram			

21. Comparing Properties of Giant Covalent Substances: Diamond and Graphite



Property: Hard
Explanation:
Each carbon atom forms four covalent bonds with other carbon atoms. Covalent bonds are strong

Giant covalent

Property: Lubricant/soft
Explanation:
Weak intermolecular forces between the layers so layers can slide over each other.

DIAMOND

Made of carbon

GRAPHITE

Property: Doesn't conduct electricity.
Explanation:
No delocalised electrons or free ions so cannot carry charge

Property: High melting point
Explanation: Lots of strong covalent bonds. Lots of energy needed to separate the carbon atoms.

Property: Conducts heat and electricity
Explanation:
One electron from each carbon atom is delocalised. The electrons are able to carry the charge through the structure. carry charge

22. Graphene

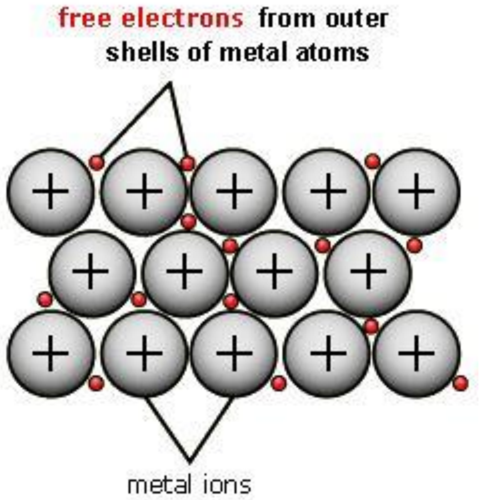
Graphene	
Single layer of graphite. Made of carbon atoms. One atom thick	
Structure	Each carbon atom forms three covalent bonds with three other carbon atoms, forming layers of hexagonal rings . The 4th electron is delocalised and carries the charge through the structure.
Property & Explanation	Conducts heat and electricity **Explanation**: One electron from each carbon atom is delocalised. The electrons are able to carry the charge through the structure.
Property & Explanation	High melting point **Explanation**: Lots of strong covalent bonds. Lots of energy needed to separate the carbon atoms.

23. Structure of Metals and Alloys

Metallic Bonding

Metallic bond: Attraction between the positive metal ion and delocalised electrons.

Structure: Layers of metal positive ions surrounded by delocalised electrons

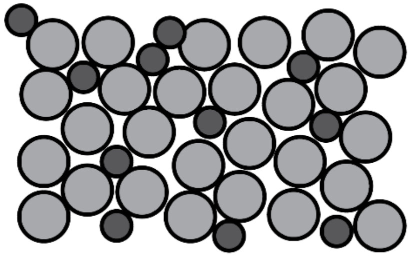


Alloy

Mixtures of metals with metals or a non-metal

e.g. stainless steel is a mixture of iron and carbon

Structure: Irregular layers



24. Properties of Metallic Bonding and Alloys

Property of metals	Explanation
Conduct electricity	Delocalised electrons are free to move and carry the charge through the metal.
Conducts thermal energy	Delocalised electrons move Energy transferred
Strong **High melting point**	Strong attraction between the metal positive ion and the delocalised electrons, so lots of energy needed to overcome attraction
Bent and shaped (malleable)	Layers of atoms are able to slide over each other.

Property of alloys	Explanation
Harder than pure metals	The atoms are different sizes. Layers are distorted and cannot easily slide over each other.

25. Polymers

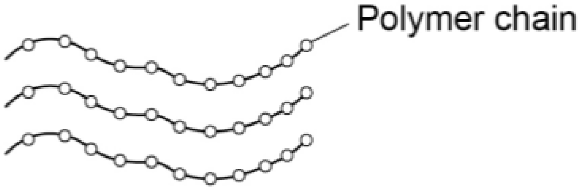
Keyword	Definition
Monomer	Made of a C=C bond. An alkene
Polymers	Large molecules linked to other atoms by strong covalent bonds.
n	Number of monomers/repeating units
Polymerisation	The C=C double bond in the monomer breaks open. Many monomers join together to form a long chain molecule (polymer.

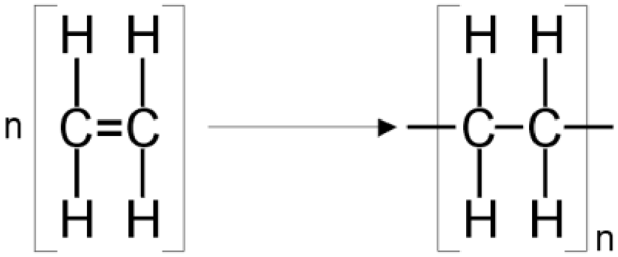
Property of polymers	Explanation of property
Solid at room temperature/ Low melting point	The intermolecular forces between polymer molecules are relatively strong. Lots of energy needed to break bonds.

Structure and bonding in a polymer chain

Strong covalent bonds between the atoms

Weak intermolecular forces between the chains



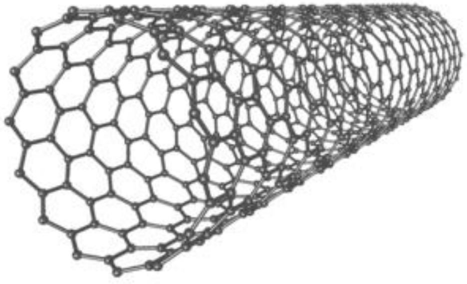


Name of monomer	Name of polymer
Vinyl chloride	Polyvinyl chloride
Styrene	Polystyrene
Ethene	Polyethene

26. Fullerenes

Fullerenes

Fullerenes are molecules of carbon atoms with hollow shapes based on hexagonal rings of carbon atoms.



Properties: High tensile strength, electrical conductivity and conducts heat.

Uses:

Drug delivery into the body as it has a hollow structure.

Lubricants

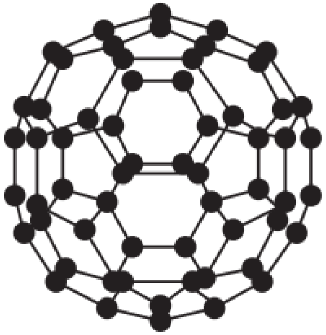
Catalysts.

Buckminster fullerene:

Molecular formula: C_{60}

Spherical shaped

Uses: Lubricant as they can roll over each other



27. Electrolysis

Electrolysis: The splitting of an ionic compound into its elements using electricity.

Electrolyte: A molten ionic compound or an ionic solution e.g. sodium chloride. They conduct electricity.

Reaction condition for electrolysis to occur:

In a solid, ions are not free to move.

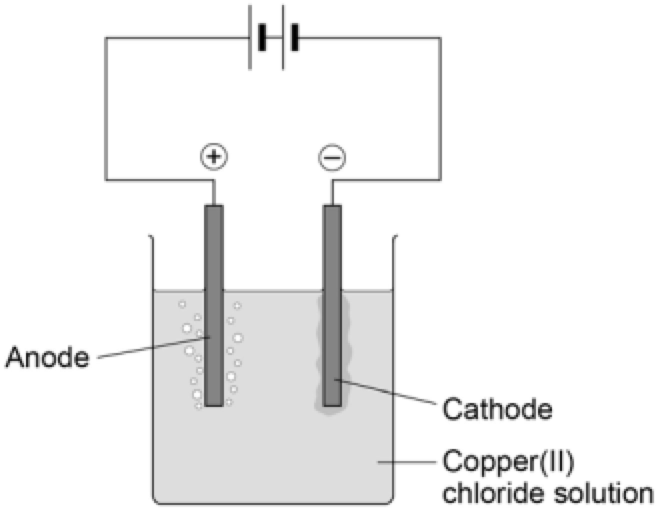
In solution or molten, the ions are free to move and carry the charge.

Electrolysis Apparatus

Remember PANIC (Positive Anode Negative Is Cathode)

Positive ions move to the cathode (negative electrode)

Negative ions move to the anode (positive electrode)



28. Processes Occurring During Electrolysis

Reaction at the Anode

Non-metal ions (anions) move to the anode. Non-metal molecules are produced.

Half Equation:

$2Cl^- - 2e^- \rightarrow Cl_2$ ($2Cl^- \rightarrow Cl_2 + 2e^-$)

Each chloride ions lose 1 electron to form a chlorine molecule. It has been oxidised.

Remember OIL RIG (<u>O</u>xidation <u>I</u>s <u>L</u>oss Reduction <u>I</u>s <u>G</u>ain)

Processes at the anode

If the anion is sulfate (SO_4^{2-}) or a nitrate (NO_3^{1-}) oxygen gas (O_2) is produced

If the non-metal ion is a halide e.g. Br^-, the halogen molecule will be produced (Br_2)

Reaction at the Cathode

Metal ions (cations) move to the cathode. Metal atoms are produced.

Half Equation: $Li^+ + e^- \rightarrow Li$

Lithium ion has gained 1 electron to form lithium atoms.

It has been reduced.

Remember OIL RIG (<u>O</u>xidation <u>I</u>s <u>L</u>oss Reduction <u>I</u>s <u>G</u>ain)

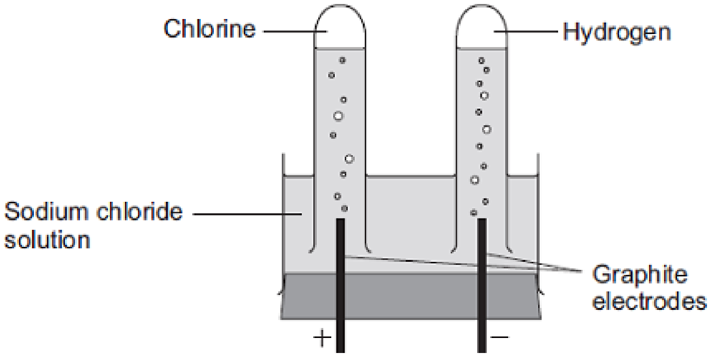
Competition between two positive ions at the cathode

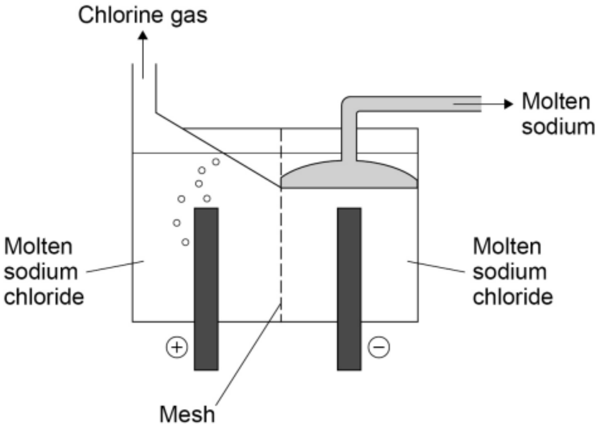
A positive metal ion e.g. K^+, and a positive hydrogen ion, H^+ are both in solution.

At the cathode, hydrogen gas (H_2) is produced if the metal is more reactive than hydrogen e.g. K^+ and H^+ ions are in solution.

Refer to reactivity series on page **18**

29. Electrolysis as an industrial process (sodium chloride)





Half equation for the production of sodium
$Na^+ + e^- \rightarrow Na$

Mesh is used to keep the products of the electrolysis apart so the products do not react

Ions pass through the mesh

Products for the electrolysis of sodium chloride
Hydrogen gas (H_2), chlorine gas (Cl_2) and sodium hydroxide (NaOH)

Ions present in solution
Na^+ and Cl^- (from NaCl),
H^+ and OH^- (from water)

How sodium hydroxide (alkali) solution is produced:
Sodium ions and hydroxide ions are left in solution
Hydrogen ions are released at the negative electrode to form hydrogen gas
$2\ H^+ + 2\ e^- \rightarrow H_2$
Chloride ions are released at the positive electrode to form chlorine gas
$2\ Cl^- \rightarrow Cl_2 + 2\ e^-$

30. Extraction of Aluminium Using Electrolysis

Electrolysis to extract metals

Metals above carbon in the reactivity series – extracted from their ores using electrolysis.

Metals below carbon in the reactivity series – extracted from their ores using carbon. This is called reduction.

Aluminium

Aluminium ore – Bauxite (aluminium oxide, Al_2O_3)

Uses of aluminium: make cars and plane and tin foil

Reaction at the cathode

$Al^{3+} + 3e^- \rightarrow Al$ Reduction

Al^{3+} has gained 3 electrons to form Al atoms.

Expensive - Large amounts of energy are needed to melt the metal compound, and to produce electricity.

Why a molten mixture of aluminium oxide is used:

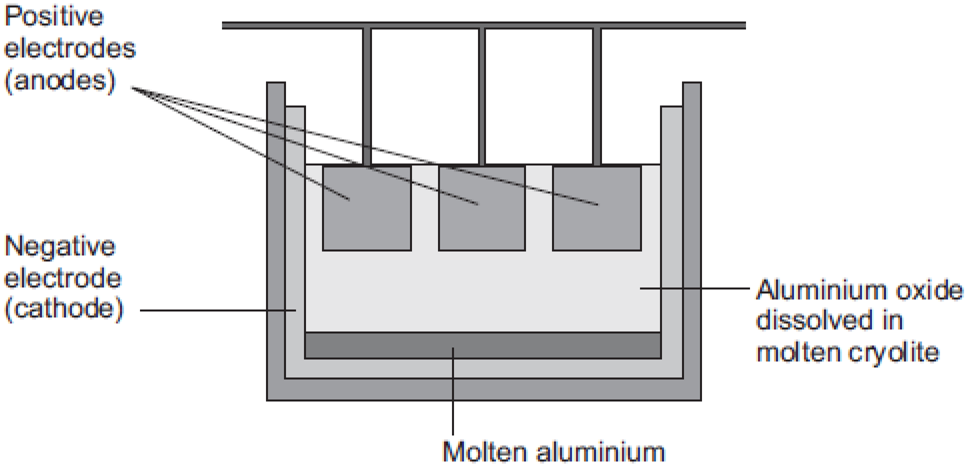
Mixed with cryolite. This lowers the melting point, so less energy is needed.

Carbon anodes replaced because the carbon anode reacts with oxygen produced at the anode. The anode fizzles away as CO_2 is produced.

Reaction at the anode

$2O^{2-} \rightarrow O_2 + 4e-$ Oxidation

Two O^{2-} ions have lost 2 electrons each to form an O_2



31. Required Practicals 1: Making a salt and Electrolysis

Making a soluble salt

1. Add excess copper oxide to sulfuric acid in a beaker
2. Stir using a stirring rod
3. Filter using a funnel and filter paper into a conical flask.

4. Evaporate the water from the copper sulfate solution in an evaporating dish using gentle heat until half the volume is left.

5. Leave on windowsill to form crystals.

6. Pat dry crystals.

Reasoning for the steps

Step 1: Excess metal oxide used so that all the acid reacts.

Step 2: Reaction stirred so all the chemicals react.

Step 3: Removal of excess copper oxide. Excess copper oxide used as it is easier to remove than excess acid

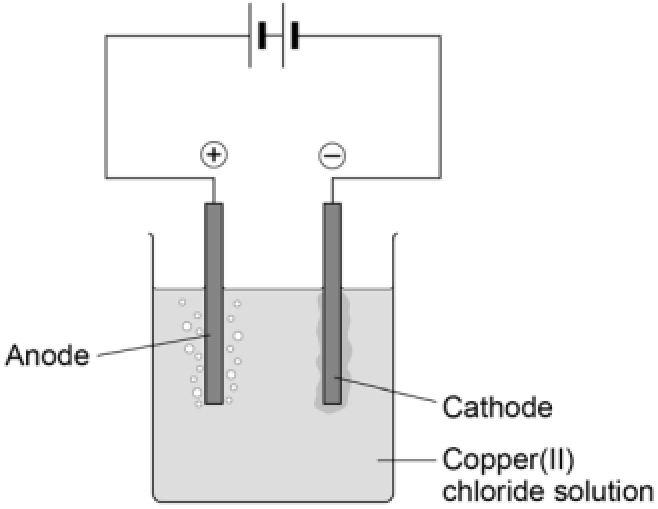
Step 4: Slow this step down by using a water bath

Observations:

Black solid (copper oxide) is left in the filter paper

Colour change

Electrolysis of aqueous solutions



Cathode: Metal attracted. Metal atoms are formed.

If the **metal is more reactive than hydrogen**, the metal ion will stay in solution and hydrogen ions will attract to the cathode, producing hydrogen gas

Anode: If the anion is sulfate (SO_4^{2-}) or a nitrate (NO_3^{1-}) oxygen gas (O_2) is produced

32. Required Practicals 2 – Energy Changes

Reacting two solutions, e.g. acid and alkali

1.Place the polystyrene cup inside the glass beaker

2.Using a measuring cylinder, measure 25 cm^3 of acid

3. Add to polystyrene cup.

4.Record the temperature of the acid using a thermometer.

5.Add 5cm^3 of alkali to the polystyrene cup and record the temperature obtained.

6.Repeat with 5cm^3 of alkali until 40 cm3 of alkali has been added

IV: Volume of alkali

DV: Temperature of reaction mixture

CV: Type of acid and alkali, volume of acid

To improve the accuracy

Use polystyrene cup

Add a lid

Repeat the experiment and calculate the mean ignoring anomalous results

Valid results: Repeat 3 times, identify the anomalous results, calculate the mean

Reacting a solid with a solution, e.g. metal and solution

1.Place the polystyrene cup inside the glass beaker to make it more stable.

2.Using a measuring cylinder, measure 25 cm^3 of copper sulfate solution

3.Place the solution in a polystyrene cup.

4.Record the temperature of the solution using a thermometer.

5.Using a balance, weigh out 1g zinc powder

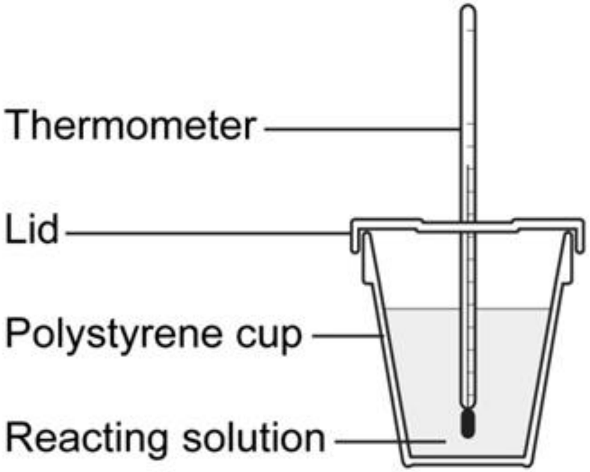
6.Add the zinc powder and record the temperature.

7.Repeat steps 1-6 with different masses of zinc powder

IV: Mass of metal

DV: Temperature of reaction mixture

CV: Concentration and volume of copper sulfate solution



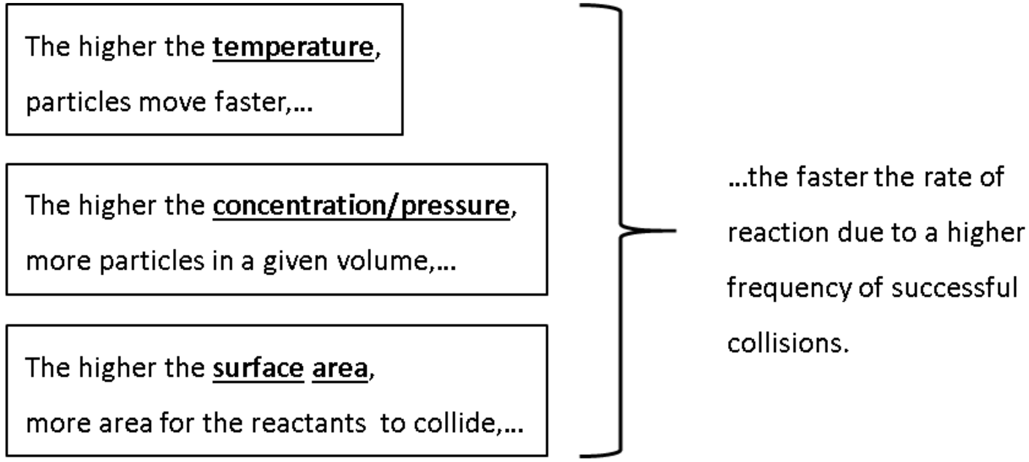
Chemistry Paper 2 (Combined Higher)

CONTENTS

33. Rates of Reaction

Collision theory	Chemical reactions can occur only when reacting particles collide with each other and with sufficient energy.
Activation energy	The minimum amount of energy that particles must have to react energy.
Factors that affect the rate of a reaction	Concentration; Temperature Pressure; Catalyst Surface area

Explaining the rate of reaction in terms of particles



The higher the **temperature**, particles move faster,...

The higher the **concentration/pressure**, more particles in a given volume,...

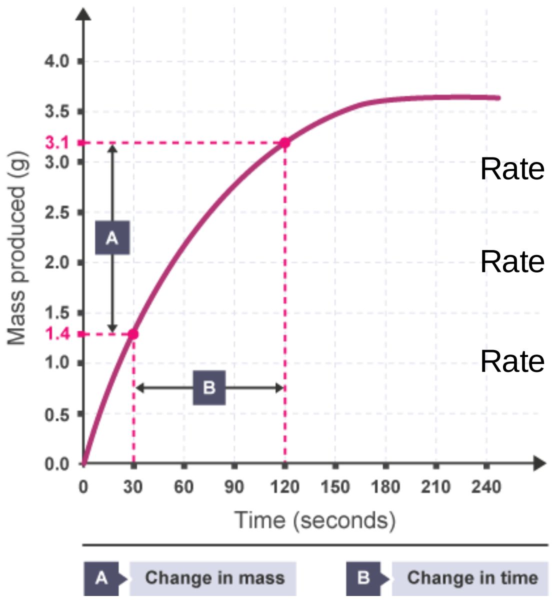
The higher the **surface area**, more area for the reactants to collide,...

...the faster the rate of reaction due to a higher frequency of successful collisions.

Measure the rate of reaction by:	**Equipment needed:** Stop clock Balance or measuring cylinder/gas syringe
	a) Loss of mass of the reactants (use a balance) b) Volume of gas produced (use a gas syringe or upturned measuring cylinder) c) Time taken for the solution to become cloudy (place conical flask on cross and watch it disappear)

34. Rates of Reaction Graphs

Calculating average rate

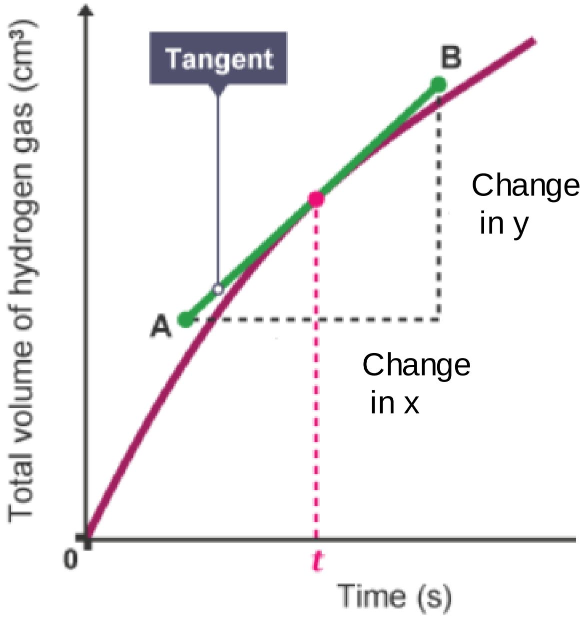


$$\text{Rate} = \frac{\text{change in mass}}{\text{change in time}}$$

$$\text{Rate} = \frac{3.1 - 1.4}{120\text{-}30}$$

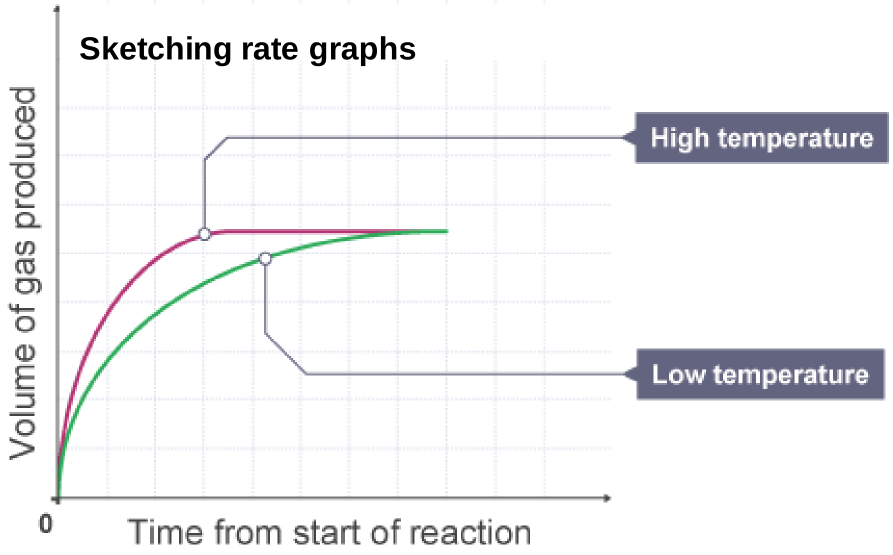
$$\text{Rate} = 0.0189 \text{ /s}$$

Calculating the rate at a specific time



1. Draw a tangent at that point approximately 10 cm long.
2. Draw a triangle
3. Calculate change in y
4. Calculate change in x
5. Gradient $= \frac{\text{change in y}}{\text{change in x}}$

Sketching rate graphs



Steeper the curve	Faster the rate of reaction
Horizontal line on graph	Reaction is finished (reactants used up)

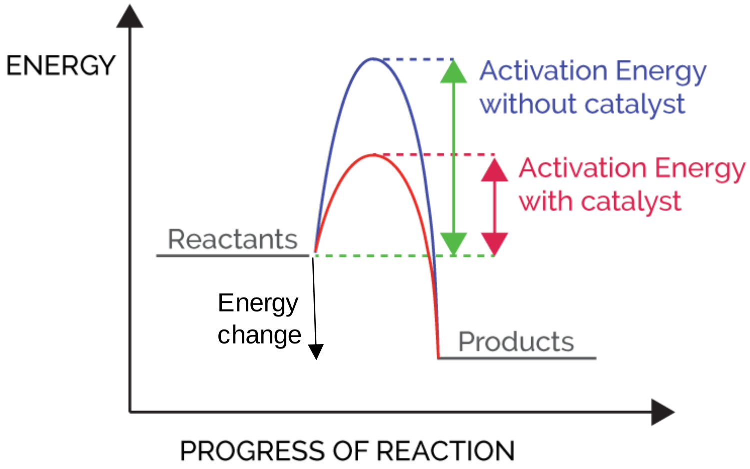
35. Rates of Reaction and Equilibrium

Catalysts increase the rate of reaction by providing a different pathway for the reaction that has a lower activation energy.

They reduce energy costs.

Catalysts are not included in the chemical equation for the reaction.

Biological catalyst: enzyme



Reversible reaction

Anhydrous	+	water	$\rightleftharpoons$	Hydrated
copper sulfate				copper sulfate
(white)				(blue)

Closed system	When reactants or products cannot enter or leave the system
What does it mean by equilibrium?	The rate of the forward and reverse reaction is the same. The concentrations of reactants and products are constant. It is a closed system
Equilibrium and temperature	Increase in temperature – reaction moves in the endothermic direction. Decrease in temperature – reaction moves in the exothermic direction.
Equilibrium and pressure	Increase in pressure – reaction moves to the side of the fewer moles. Decrease in pressure – reaction moves to the side of the most moles.
Equilibrium and concentration	Increase in concentration of a chemical– reaction moves to the opposite side to use up excess chemical. Decrease in concentration of a chemical– moves to this side to create more of this chemical.
Equilibrium and a catalyst	No effect on the position of equilibrium. A catalyst allows the reaction to reach equilibrium faster. Increases the rate of the forward and the reverse reaction by the same amount.

36. Evolution of the Atmosphere

Volcanoes released water vapour (H_2O), carbon dioxide (CO_2), methane (CH_4), ammonia (NH_3).

Volcanoes were a source of nitrogen.

Not certain of exact % of each gas as there was no evidence

⇨

Temperature cooled down. Water vapour condensed to form oceans

⇨

Algae and plants used up carbon dioxide by photosynthesis to produce oxygen.

⇨

Methane reacted with oxygen to form carbon dioxide and water.

Ammonia reacted with oxygen to form nitrogen and water.

Today's atmosphere:

- 78 % Nitrogen (N_2)
- 21 % oxygen (O_2)
- 1 % other gases

Reasons why O_2 levels increased	• Algae and plants began to photosynthesise, producing oxygen • Oxygen levels increased, allowing animals to evolve.
Reasons why CO_2 levels decreased	•Absorbed by oceans. •Locked up as sedimentary rocks and fossil fuels. •Used in photosynthesis to produce oxygen

How coal was formed from carbon dioxide present in the early atmosphere:

- Carbon dioxide was used during photosynthesis by trees
- Trees die and are compressed over millions of years

37. Greenhouse effect

Greenhouse Gases

- Water vapour (H_2O)
- Carbon dioxide (CO_2)
- Methane (CH_4)

Effects of Global Climate Change
Sea level rise, which may cause flooding and increased coastal erosion
More frequent and severe storms
Changes to the distribution of wildlife species

Human Activities Which Increase Greenhouse Gases
Combustion of fossil fuels releasing more carbon dioxide
Deforestation leading to less trees so less photosynthesis occurring
More animal farming (digestion, waste decomposition) so more methane released
Decomposition of rubbish in landfill sites so more methane released)

Short wavelength radiation enters the atmosphere

Absorbed by materials and re-emitted as a longer wavelength radiation

The longer wavelength radiation is trapped by a greenhouse gas which stops radiation escaping from the atmosphere.

38. Polluting our Atmosphere

Pollutant	How it is made	Effect on health/environment
Sulfur dioxide (SO_2)	Sulfur in fossil fuels reactions with oxygen to form sulphur dioxide.	Cause respiratory problems in humans and causes acid rain. Acid rain damages plants and buildings.
Carbon monoxide (CO)	Incomplete combustion of hydrocarbons.	A toxic gas which causes death.
Carbon particulates (unburned hydrocarbons)	Incomplete combustion of hydrocarbons.	Causes global dimming and damages lungs.
Oxides of nitrogen (NO_x)	Made from nitrogen and oxygen in air reacting at a high temperature in a car engine.	Causes respiratory problems in humans and cause acid rain.

Carbon Footprint	The total amount of carbon dioxide and other greenhouse gases emitted over the full life cycle of a product, service or event.
How to Reduce the Carbon Footprint	- Increased use of alternative energy supplies e.g. wind - Use energy efficient appliances - Carbon capture and storage (CCS)
Problems on Reducing the Carbon Footprint	- Lifestyle changes e.g. using public transport - Economic considerations e.g. can countries afford to build more wind turbines?

39. Further Quantitative Chemistry 1: Equations and definitions

Mole	Mole= mass (g) / relative formula mass Mole = mass (g) /relative atomic mass
Avogadro’s Number	6.02×10^{23} The number of particles (atoms, ions or electrons) in one mole of substance.
Volume	$cm^3 \xrightarrow{\div 1000} dm^3$
Concentration	Concentration (g/dm^3) = mass (g) / volume (dm^3)

40. Further Quantitative Chemistry 2: Limiting reactants and Theoretical Yield

Limiting Reactants

The reactant that is completely used up is called the limiting reactant because it limits the amount of products.

For example: 3 g of <u>Mg</u> react with 7 g of $\underline{O_2}$. Which is the limiting reagent?

	2$\underline{Mg}$	+	$\underline{O_2}$	→ 2MgO
Mass	3 g			
Mr	24			
(don’t include big 2)				
Moles	3/24=0.125		0.125/2 = 0.0625	
Ratio		**2**Mg :**1**O_2		

O_2 is the limiting reactant as there is only 0.0625 moles. Once the oxygen has reacted, the reaction is over.

Theoretical Yield Calculation

128 grams of <u>hydrogen peroxide</u> break down into water and oxygen. What mass of <u>oxygen</u> is produced?

	$\underline{2H_2O_2}$	→	$2H_2O$ +	$\underline{O_2}$
Mass	128 g			60.2g
Mr	34			32
(don’t include big 2)				
Moles	128/34=3.76			3.76/2 = 1.88
Ratio		**2**H_2O_2 :**1**O_2		

a) Underline the 2 substances from the question in the equation.
b) Add the information from the question under mass, Mr and moles.
c) Use ratios (the big numbers to calculate the new moles).
d) Follow the U-arrow to calculate new mass

41. Further Quantitative Chemistry 3: Balancing equations

Question: Work out the balanced equation when 12 grams of magnesium reacts completely with 38.5g of HCl, to make 49.5 grams of $MgCl_2$ and 1 gram of H_2

:

	Mg +	HCl →	$MgCl_2$ +	H_2
Step 1: work out the moles of each reactant and product.	12 g/ 24 = 0.5	38.5 g/ 38.5 = 1	49.5/99 = 0.5	1/2 = 0.5
Step 2: divide through by the smallest number	0.5/0.5=1	1/0.5 = 2	1/0.5=1	0.5/0.5=1
Step 3: write the balanced equation	$Mg + 2HCl \rightarrow MgCl_2 + H_2$			

42. Alkanes

Hydrocarbon	Made of **only** hydrogen and carbon
Alkane	A hydrocarbon made of C-C single bonds.
Alkane General Formula	$C_n H_{2n+2}$
Functional group of an alkane	C-C single bond Alkanes are saturated as all the C bonds are used up.
Homologous series	A family of hydrocarbons with similar chemical properties who share the same general formula

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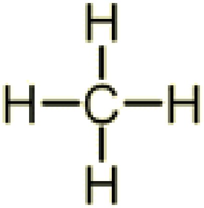
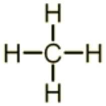
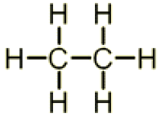
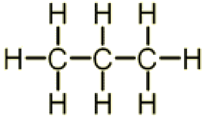
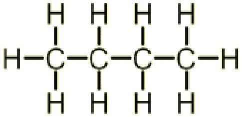


Methane
A compound
A hydrocarbon
Covalent bonds between the C-H atoms
Homologous series: Alkanes

Alkane	**Molecular Formula**	**Displayed formula**
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	
Butane	C_4H_{10}	

43. Fractional Distillation of Crude Oil

Keyword	Definition
Boiling point	The temperature at which a liquid turns into a gas
Combustion	Burning in oxygen
Flammability	How easily a substance ignites (catches on fire)
Fossil fuels	(non-renewable/finite fuels) Coal, oil, natural gas
Fraction	Molecules with a similar number of carbon atoms
Viscosity	The runniness of a liquid Higher the viscosity of the liquid, the longer it will take for the liquid to flow
Volatility	How easily a liquid changes into a gas

Physical property:
Fractional distillation relies of mixtures having different **boiling points** to enable the mixture to be separated

How coal is made: Trees die and are compressed over millions of years.

How crude oil is made: Made by the decomposition of plankton buried in mud over millions of years

Coal has more carbon than oil and natural gas

Fractional distillation of crude oil

- Crude oil is heated and evaporated.
- Fractions in crude oil separate depending on their boiling point and size of fraction.
- At the top of the column, short fractions with low boiling point condense
- At the bottom of the column, long fractions with high boiling point condense

Properties of fractions as you go down the column

Boiling point - increase with increasing molecular size

Viscosity - increase with increasing molecular size

Flammability - decreases with increasing molecular size

44. Combustion and Cracking

	Complete combustion (FO COW)	**Incomplete combustion**
Reaction conditions	Lots of oxygen	Little oxygen
Reactants	Fuel and oxygen	Fuel and oxygen
Products	Carbon dioxide and water	Carbon monoxide and water
	Test for carbon dioxide: Bubble through limewater **Result**: Turns cloudy	Carbon monoxide is toxic

Cracking vs Distillation

Cracking Requires a catalyst

Distillation Does not require a catalyst

Cracking - Hydrocarbons can be broken down (cracked) to produce smaller, more useful molecules. Also known as thermal decomposition.

Thermal decomposition – breaking down a compound using heat.

Example:

$$C_{30}H_{62} \rightarrow C_{20}H_{42} + C_{10}H_{20}$$

Long alkane	more useful shorter alkane	alkene (make plastics)

Reason for cracking: Turns long hydrocarbon chains into more useful shorter hydrocarbon chains.

Short alkanes are useful as they are flammable

Alkenes are used to make plastics via polymerisation (see page 24)

Catalytic Cracking

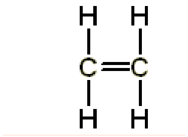
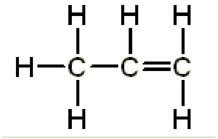
Reaction conditions: High temperature and a catalyst

Steam Cracking

Reaction conditions: High temperature

45. Alkenes

Alkene	A hydrocarbon made of C=C double bonds.
Alkane General Formula	$C_n H_{2n}$
Functional group of an alkane	C=C double bond Alkanes are unsaturated
Chemical test for alkene	Add bromine water Alkene = Orange to colourless Alkane = stays orange

Alkene	**Molecular Formula**	**Displayed formula**
Ethene	C_2H_4	
Propene	C_3H_6	

46. Mixtures, Test for Gases and Test for Water

Keyword	Definition
Boiling point	The temperature at which a liquid turns into a gas. Water has a boiling point of 100 ^oC
Formulation	A mixture that has been designed as a useful product e.g. shampoo Formulations include fuels, cleaning products, medicines, paints, alloys, fertilisers and foods.
Melting point	The temperature at which a solid turns into a liquid. Ice has a melting point of 0 ^oC
Pure substance	A single element or compound

Gas	Chemical test	Result
Hydrogen (H_2)	Lit splint	Pop sound
Oxygen (O_2)	Glowing splint	Splint relights in oxygen
Carbon Dioxide (CO_2)	Bubble through limewater	Turns milky/cloudy
Chlorine (Cl_2)	Damp litmus paper	Paper is bleached (white)

	Test	Result
Pure water	Boil it	Boils at exactly 100 ^oC
Water	Add anhydrous copper sulfate	Turns from white to blue

47. Chromatography

Chromatography can be used to separate mixtures and identify substances.

Relies on the difference in **solubility** (physical property) of the mixture

Mobile phase – the solvent e.g. water running up the chromatogram.

Stationary phase – the paper.

Evidence that the dye is a mixture

- More than 1 spot
- In a vertical column

Substances move between the phases. If a substance is more attracted to the mobile phase, it will move further up.

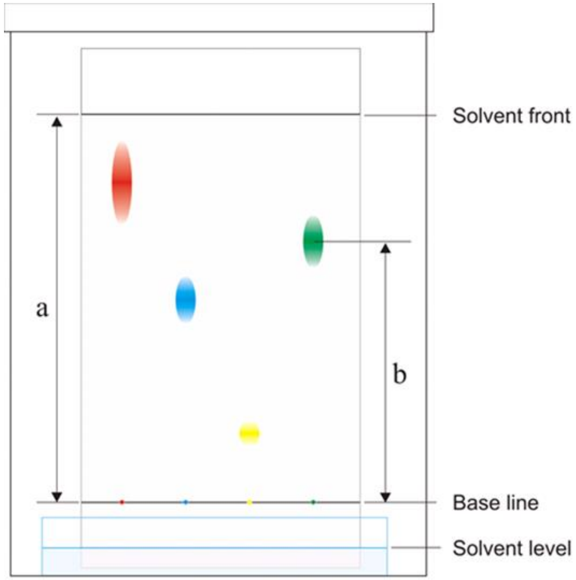
The R_f value tells you how far the substance has moved, relative to the solvent.

$$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$$

The R_f value can be used to identify the substance.

The R_f values would be compared to the known substance.

Rf value will always be **less than 1**



$R_f = b \div a$

48. Potable Water

Finite resource (non-renewable):

A source from the Earth that is running out e.g. coal

Renewable source:

A source that isn't running out e.g. wood

Potable water.

Safe to drink. Contains **low** levels of dissolved salts and microbes. Not pure.

	Potable water from salty water using distillation	**Potable water from rainwater/groundwater**	**Potable water from the sea (desalination)**	**Potable water from waste water (sewage)**
Method	1. Heat salty water. 2. Water evaporates. 3. Cool the water vapour 4. The vapour condenses to form potable water	1. Rainwater collected in reservoirs. 2. Passing the water through filter beds to remove any solids. 3. Sterilise to kill microbes. **Sterilising agents:** chlorine, ozone or ultra-violet light.	Distillation or by processes that use membranes such as reverse osmosis.	1. Removal of organic matter and harmful chemicals 2. Screening and grit removal 3. Sedimentation to produce sewage sludge and effluent 4. Anaerobic digestion of sewage sludge 5. Aerobic biological treatment of effluent.
Issues		Reliant on rainfall	These processes require large amounts of energy.	Expensive: Needs filtering and sterilising to remove harmful bacteria. Lots of steps

49. Saving Resources

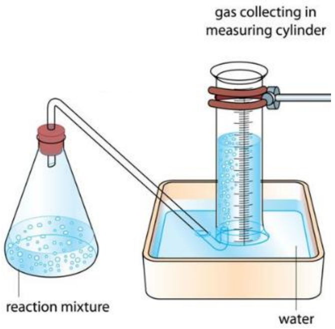
Reduces	Limits the use of raw materials, energy consumption, waste and environmental impacts (quarrying and mining for raw materials).
Reuse	Use the item for another purpose e.g. a glass bottle is refilled.
Recycle	Turn the item into something else e.g. plastic bottles recycled to make fleeces, scrap steel is added to iron from a blast furnace. **Benefits**: conserves metal ores; uses less energy; reduces waste
Sustainable development	Development that meets the needs of current generations without compromising the resources for future generations.
Life Cycle Assessments (LCAs)	To assess the environmental impact (of the stages in the life of a product). • Extracting the raw material • Processing the raw material • Manufacturing • Disposal at the end of its useful life

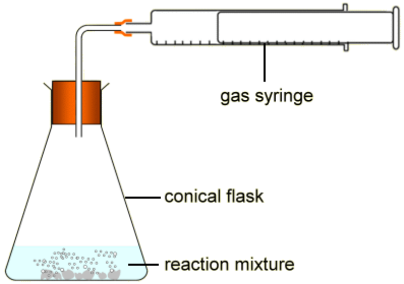
50. Copper Extraction

Keyword	Definition
High grade copper ore	Rock that contains enough copper that makes it economically viable to extract it.
Low grade copper ore	Extract using phytomining or bioloeaching.
Bioleaching	Uses bacteria to produce leachate solutions that contain metal compounds. **Advantages**: Used to clean up toxic metals from industrial sites. Extracting copper from low grade ores. **Disadvantages**: Requires lots of energy in smelting and electrolysis process.
Phytomining	Grow plants on land containing copper ores. Plants are burnt to produce ash. Ash dissolved in acid to produce a solution of a copper compound. Electrolysis of solution containing the copper compound. **Advantages:** reduces the need to obtain new ore by mining. Conserves limited supplies of more valuable ores with higher metal content **Disadvantage**: Takes a long time. Large area of land required.

51. Required practicals 4: Rates of Reaction

Measuring the rate of reaction by collecting a gas





Method

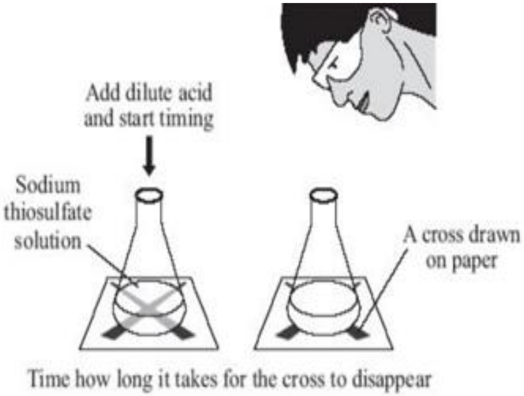
1. Set up equipment as shown in diagram.
2. Add 5 cm magnesium strip and 30 cm^3 of a highly concentrated acid.
3. Collect gas for 1 minute.
4. Repeat steps 1-3 with different concentrations of acid

IV: concentration of acid

DV: volume of gas collected in 1 minute

CV: volume and type of acid, length of magnesium strip, time period of gas collection.

Measuring the rate of reaction by the formation of a precipitate



Method

1. Place conical flask on a black cross
2. Add sodium thiosulfate and hydrochloric acid to the flask.
3. Time how long it take for the cross to disappear.
4. Repeat steps 1-3 with different concentrations of sodium thiosulfate.

IV: concentration of acid

DV: time taken for cross to disappear

CV: volume and type of acid

Why there is mass loss:

- Sulfur dioxide gas is made
- Escapes into the air

Why the solution goes cloudy:

Solid sulfur is made

52. Required practicals 5: Chromatography and Potable Water

Chromatography

Method:

1. Draw pencil start line on chromatography paper and place spot of dye on start line.
2. Place solvent in beaker and place chromatography paper in beaker so the paper is in solvent but solvent is below start line.
3. Wait for solvent to travel up the paper and mark solvent front.
4. Dry the paper

Measurements to take:

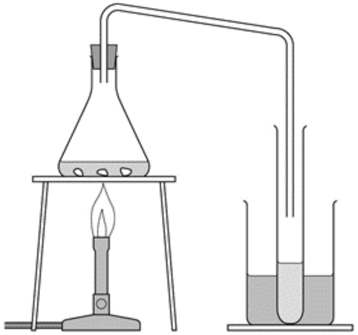
Measure distance between start line and centre of spot.

Measure distance between start line and solvent front.

Use of measurements to determine Rf value

Use of pencil – pencil is insoluble. Does not interfere with ink.

Line is above solvent level – so ink travels up the paper with the rising solvent



Method:

1. Heat seawater in conical flask.
2. Water evaporates
3. Water vapour condenses in delivery tube
4. Condenses in test tube

Chemical test	**Test for seawater in conical flask**	**Test for pure water in test tube**
Flame test to test for Na^+ ions. Dip wooden splint in each type of water and heat in blue Bunsen flame	Orange flame.	No change in colour
Test for Cl^- ions. Add silver nitrate	White precipitate	No change in colour

53. Maths in Science 1

Anomalous result	A number that does not fit the pattern
Mean	Adding up a list of numbers and dividing by how many numbers are in the list. Exclude the anomalous result.
Median	The middle value when a list of numbers is put in order from smallest to largest
Mode	The most common value in a list of numbers. If two values are tied then there are two modes. If more than two values are tied then there is no mode.
Range	The largest number take away the smallest value in a set of data or written as X-Y.
Uncertainty	range ÷ 2
Surface area of a cube	(area of 1 side) x 6 sides
Volume of a cube	Width x height x depth
Area of a circle	∏ x (radius)^2

Prefixes

1 kJ = 1 x 10^3 J = 1000 J

1 pm = 1 x 10^{-12} m

1 mm = 1 x 10^{-3} m= 0.001 m

kilo	10 ^{3}
centi	10 $^{-2}$
milli	10 $^{-3}$
micro	10 $^{-6}$
nano	10 $^{-9}$
pico	10 $^{-12}$

5607.376

Standard form: 5.607 x 10^3

2 decimal places: 5607.38

3 significant figures: 5610

0.03581

Standard form: 3.581 x 10^{-2}

2 decimal places: 0.04

3 significant figures: 0.0358

54. Maths in Science 2

Calculating percentage: (part ÷ whole) x 100

e.g. Out of 90 insects, 40 of them were ladybirds. What is the % of ladybirds?

(40 ÷ 90) x 100 = 44 %

Calculating percentage change:

(difference ÷ starting value) x 100

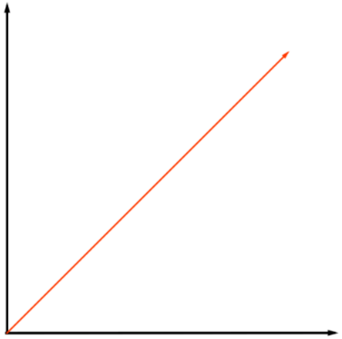
(0.59 ÷ 2.22) x 100 = 26.6 %

Conc of Sucrose (M)	Mass of potato at start (g)	Mass of potato at end (g)	Change in mass (g)
0	2.22	2.81	0.59

Graphs

Proportional (α)

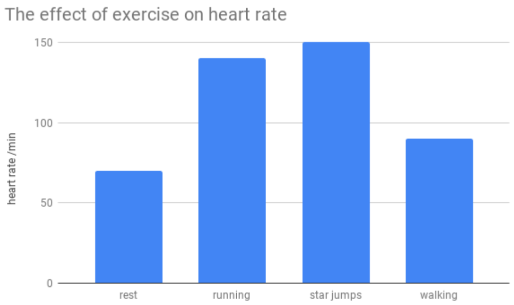
When the line passes through the origin



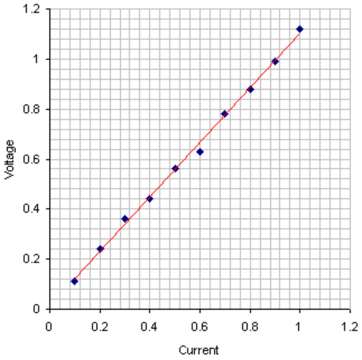
x axis = independent variable = left hand column of results table

y axis = dependent variable = right hand column of results table

Categoric data: data put into groups e.g. colour of eyes
Draw a bar chart

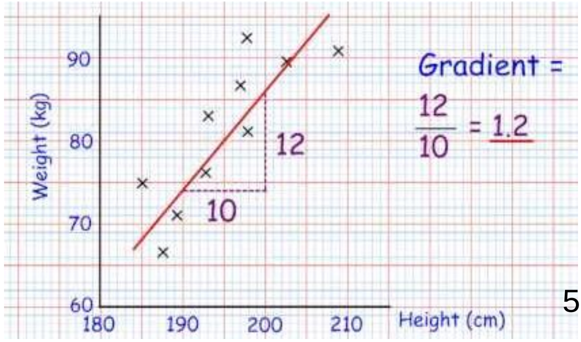


Continuous data: data that can take any value e.g. current
Draw a line graph



Gradient and Graphs

Gradient = $\frac{\text{Change in y}}{\text{Change in x}}$



Key

relative atomic mass
atomic symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
							1 **H** hydrogen 1										4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4											11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chlorine 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	63.5 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	[98] **Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La*** lanthanum 57	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89	[261] **Rf** rutherfordium 104	[262] **Db** dubnium 105	[266] **Sg** seaborgium 106	[264] **Bh** bohrium 107	[277] **Hs** hassium 108	[268] **Mt** meitnerium 109	[271] **Ds** darmstadtium 110	[272] **Rg** roentgenium 111	[285] **Cn** copernicium 112	[286] **Uut** ununtrium 113	[289] **Fl** flerovium 114	[289] **Uup** ununpentium 115	[293] **Lv** livermorium 116	[294] **Uus** ununseptium 117	[294] **Uuo** ununoctium 118

* The Lanthanides (atomic numbers 58 – 71) and the Actinides (atomic numbers 90 – 103) have been omitted.

Relative atomic masses for **Cu** and **Cl** have not been rounded to the nearest whole number.