KNOWLEDGE

Chemistry Topics 1 & 2
Atomic Structure and Periodic Table

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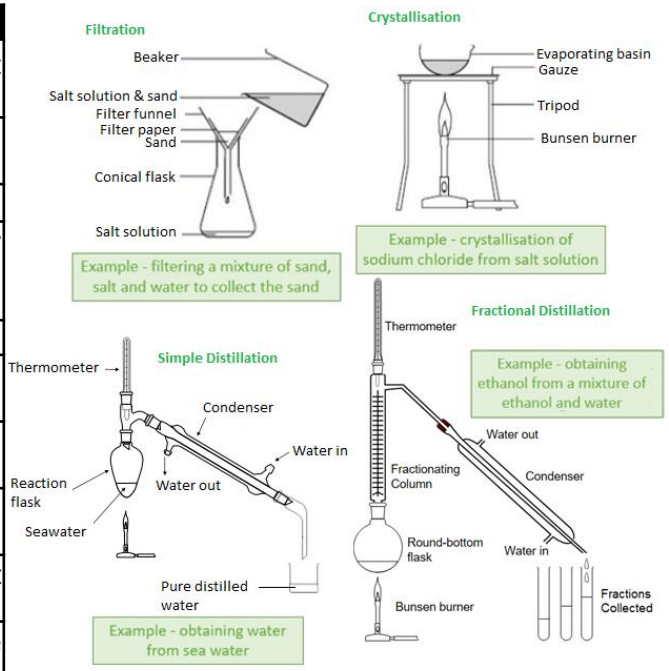
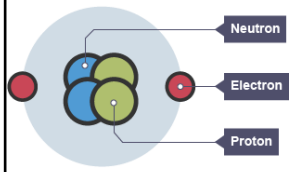
Section 1: Key Terms	
Atom	The **smallest part of an element** that can still be recognised as that element. **No** overall **electrical charge**. **Very small**, radius of 0.1nm.
Element	An element **contains only one type of atom**. Found on the Periodic Table. There are about 100 elements.
Compound	**Two or more elements chemically bonded** with each other.
Mixture	**Contains two or more elements or compounds not chemically bonded**. Can be separated using physical methods e.g. by filtration, crystallisation, distillation and chromatography.
Filtration	A process that **separates** mixtures of **insoluble solids and liquids**.
Crystallisation	A process that **separates a soluble solid from a solvent** by **evaporating** the liquid to leave crystals.
Distillation	A process that **separates a mixture of liquids** based on their **boiling points**.
Chromatography	A process that **separates mixtures** by **how quickly they move through a stationary phase** (e.g. paper chromatography)
Isotope	An atom of the **same element with same number of protons** but **different numbers of neutrons**.
Relative atomic mass	An **average value of mass** that takes account of the **abundance of the isotopes** of the element.



Section 2: Development of Atomic Model	
Plum Pudding 	Thompson's plum pudding model shows that the atom is a **ball of positive charge** with **negative electrons embedded** in it. Was **incorrect**.
Nuclear Model 	**Rutherford's** alpha particle scattering experiment found a central area of positive charge. The nuclear model has a **positive nucleus** and **electrons in shells**. **Chadwick** later discovered **neutrons**. **Bohr** discovered the arrangement of **electrons in shells**.

Mass number – the total number of **protons** and **neutrons**

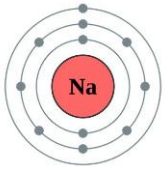
Atomic number – the **number of protons** (the number of electrons is the same in an atom)

Electron configuration– Electrons fill the first energy level (shell) first.
Maximum electrons: **2 electrons in first shell, 8 in the 2nd, 8 in the 3rd.**

Section 3: Properties of Sub-Atomic Particles

Sub-atomic particle	Mass	Charge	Position in Atom
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus
Electron	Very small	-1	Orbiting in shells

$^{23}_{11}Na$



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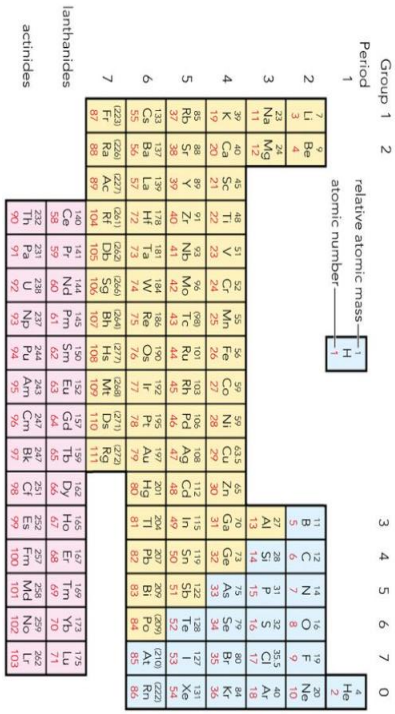
Chemistry Topics 1 & 2
Atomic Structure and Periodic Table

Section 4: Periodic Table	
Group	Elements in the **same vertical column** are in the same group. Elements in the same group have the **same number of electrons in their outer shell**, and therefore **similar properties**.
Period	Elements in the **same horizontal row**. The atomic number increases by one moving across the period from left to right.
Metal	Elements that react to form positive ions (except Hydrogen). Left and centre of periodic table
Non-Metal	Elements that react to form negative ions. Right hand side of periodic table.
Mendeleev	Was able to make a relatively accurate periodic table by **leaving gaps for undiscovered elements** and **re-arranging some elements** (Mendeleev could only measure relative atomic mass, not atomic number). Hence he arranged the elements in **order of mass** number and **predicted the properties** of the elements in the gaps

Section 5: Groups of the Periodic Table			
Sub-atomic particle	**Properties**	**Trends**	**Reactions**
Group 0 (Noble Gases)	**Unreactive** and **do not form diatomic molecules**.	**Boiling point increases** going **down the group**.	Very unreactive because they **have full outer shells**.
Group 1 (Alkali Metals)	**Reactive** because they can easily lose their one outermost electron. Always form ionic compounds Low density	**Reactivity increases** going **down the group**. **Melting points and boiling point decrease** going **down the group**.	**With water:** Metal + water → Metal hydroxide + hydrogen **With oxygen:** Metal + oxygen → Metal oxide **With chlorine:** Metal + chlorine → Metal chloride
Group 7 (Halogens)	Low melting points and boiling points. Poor conductors of heat and electricity. Form diatomic molecules	**Reactivity decreases** going **down the group**. **Boiling point** and **melting point increase** going **down the group.**	A **more reactive halogen** can **displace** a **less reactive halogen** from a solution of its salt. Chlorine + sodium bromide → sodium chloride + bromine

Elements in the modern periodic table are **arranged by atomic (proton) number**.



Group – Vertical column

Period – Horizontal Row

Metals are on the left, non-metals on the right.

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Chemistry Topic 4
Chemical calculations

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Section 1: Chemical calculations Key Terms	
Law of conservation of mass	**No atoms** are **destroyed** or **created** during a **chemical reaction**. The total mass of the products is the same as the total mass of the reactants. Some reactions **appear to give a change in mass**, but this is because a **gas may have escaped** from the reaction container.
Relative atomic mass (A_r)	The **average mass** of an **atom of an element** compared to Carbon-12.
Relative formula mass (M_r)	The **sum** of **all the atomic masses** of the atoms in a **formula** of a substance (e.g. CO_2).
Uncertainty	The **interval** within which the **true value** can be **expected to lie**. E.g. 25°C $\pm$ 2°C – the true value lies between 23°C and 27°C.
Mole (HT)	A measurement for the amount of a chemical. It is the amount of substance in the relative atomic or formula mass of a substance in grams. The **mass** (in grams) of **6.02×10^{23}** (the Avogadro constant) **atoms of an element**. Symbol: mol.
Balanced equation (HT)	Balanced symbol equations show **the number of moles that react**. e.g. $Ca + 2HCl \rightarrow CaCl_2 + H_2$ Shows one mole of Calcium reacting with two moles of hydrochloric acid to form one mole of Calcium chloride and one mole of hydrogen.
Limiting reactant (HT)	The **reactant** that **gets used up first** in a chemical reaction. It **limits the amount of product** formed.
Excess reactant (HT)	The reactant that is **not completely used up** in a chemical reaction. There is some reactant left at the end.
Concentration	A measure of the **number of particles** of a chemical in a **volume**. Can be measured in **g/dm^3**.
Decimetre^3 (dm^3)	A **measurement of volume**. Contains 1**000cm^3**.

Section 2: Calculating relative formula mass (M_r)	
Add up all the atomic masses in a formula.	e.g. CO_2 Mass of C = 12. Mass of oxygen = 16. 12 + (2x16) = 44

Section 3: Calculating moles and masses (HT)	
Number of moles = $\frac{\text{mass (g)}}{M_r}$	1) How many moles are there in 9.8g of sulfuric acid H_2SO_4? **Number of moles = $\frac{9.8}{98}$ = 0.1 moles** 2) What is the mass of 2.5 moles of Carbon dioxide? **Mass = 2.5 x 44 = 88g**

Section 4: Equations and calculations (HT)	
Number of moles = $\frac{\text{mass (g)}}{M_r}$	1) What masses of reactants and products are involved in the balanced symbol equation $H_2 + Cl_2 \rightarrow 2HCl$ **Reactants: (2x1) + (2x35.5) = 73** **Products: 2 x 36.5 = 73** 2) What mass of oxygen will react with 72.0g of magnesium? $2Mg + O_2 \rightarrow 2MgO$ **Moles Mg = 72/12 = 3 moles** **Molar ratio Mg:O_2 is 2:1** **Moles O_2 = 3/2 = 1.5 moles** **Mass O_2 = 1.5 x 32 = 48g**

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Chemistry Topic 4
Chemical calculations

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Section 5: From masses to balanced equations (HT)

Number of moles = $\frac{\text{mass (g)}}{M_r}$

1) 8.08g of Potassium nitrate KNO_3 was decomposed on heating to form 6.8g of potassium nitrite KNO_2 and 1.28g of oxygen.

a) Calculate the number of moles of KNO_3, KNO_2 and O_2 and hence

Moles KNO_3 = 8.08/101 = 0.08
Moles KNO_2 = 6.8/85 = 0.08
Moles O_2 = 1.28/32 = 0.04

b) Use your answers to a) to work out the simplest whole number ratio of these values and use this to write a balanced equation for the reaction.

Moles KNO_3: KNO_2 : O_2
0.08 : 0.08 : 0.04
2 : 2 : 1
Hence equation is
$2KNO_3 \rightarrow 2KNO_2 + O_2$

Section 7: Expressing concentrations (in g/dm^3)

If you are working in decimetres cubed (dm^3)

Concentration (g/dm^3) = $\frac{\text{mass of solute (g)}}{\text{volume (dm}^3\text{)}}$

If you are working in centimetres cubed (cm^3)

Concentration (g/dm^3) = $\frac{\text{mass of solute (g) x 1000}}{\text{volume (cm}^3\text{)}}$

1) Calculate the concentration in g/dm^3 of 6g of magnesium chloride dissolved in 1.5 dm^{3}of solution
Concentration = 6/1.5 = 4 g/dm^3

2) Calculate the concentration in g/dm^3 of 40g of sodium hydroxide dissolved in 500 cm^{3}of solution
Concentration = 40/500 x 1000 = 80 g/dm^3

Section 6: Limiting reactants (HT)

Number of moles = $\frac{\text{mass (g)}}{M_r}$

Remember:

A **limiting reactant** is the **reactant** that **gets used up first** in a chemical reaction. It **limits the amount of product** formed.

Excess reactant is the **reactant** that is **not completely used up** in a chemical reaction. There is some reactant left at the end.

1) If you have 7.2g of magnesium reacting with 10.95g of dilute hydrochloric acid, which reactant is in excess?

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

Moles Mg = 7.2/24 = 0.3 mol
Moles HCl = 10.95/36.5 = 0.3 mol
From the balanced equation you see that 1 mole of Mg reacts with 2 moles of HCl.
Hence 0.3 mol of Mg requires 0.6 mol of HCl to react completely. We only have 0.3 mol of HCl so dilute hydrochloric acid is the limiting reactant.

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Chemistry Topic 5
Chemical changes

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Section 1: Key Terms

Displacement reaction	A **more reactive metal** will **displace** a **less reactive metal** from a **compound**. e.g. Iron is more reactive than copper hence will displace copper from solution. $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$
Oxidation	Two definitions: Chemicals are oxidised if they **gain oxygen** in a reaction. Chemicals are oxidised if they **lose electrons** in a reaction. (HT)
Reduction	Two definitions: Chemicals are oxidised if they **lose oxygen** in a reaction. Chemicals are oxidised if they **gain electrons** in a reaction. (HT)
Acid	A chemical that **dissolves in water** to produce **H^+ ions**. Acids are proton donors
Base	A chemical that **reacts with acids** and **neutralise** them. E.g. **metal oxides**, **metal hydroxides**, **metal carbonate**
Alkali	A soluble **base** that produces **OH- ions** in solution.
Neutralisation	When a **neutral solution** is formed from reacting an **acid** and **alkali**. Ionic equation: $\mathbf{H^+ + OH^- \rightarrow H_2O}$
pH	A scale to **measure acidity/ alkalinity**. A **decrease of one pH** unit causes a **10x increase in concentration of H^+ ions**. (HT)
Strong acid (HT)	Strong acids **completely ionise** in solution. E.g. **hydrochloric**, **nitric** and **sulfuric** acids.
Weak acid (HT)	A weak acid is **only partially ionised** in solution. E.g. **ethanoic**, **citric** and **carbonic** acids.

Section 2: The Reactivity Series

Metals can be placed in order of reactivity by their reactions with water and dilute acid. Hydrogen gas is given off when metals react with acid or water. The gas gives a squeaky pop with a lighted spill.

Element	Reaction with water	Reaction with acid	Reactivity
Potassium	Potassium **melts**, floats & moves around very quickly. It **sets on fire** with a **lilac flame**. Alkaline solution forms.	Explodes	↑
Sodium	Sodium **melts** to form a ball that moves around on the surface. It **fizzes rapidly**. Alkaline solution forms.	Explodes	
Lithium	Lithium floats. It **fizzes steadily** and becomes smaller. Alkaline solution formed.	Explodes	
Calcium	It fizzes steadily leaving an alkaline solution.	**Fizzes quickly** with dilute **acid**.	
Magnesium	Very slow reaction	**Fizzes quickly** with dilute **acid**.	
(Carbon)			
Zinc	Very slow reaction	**Bubbles slowly** with dilute **acid**.	
Iron	Very slow reaction	**Very slow reaction** with dilute **acid**.	
(Hydrogen)			
Copper	No reaction	No reaction	
Silver	No reaction	No reaction	
Gold	No reaction	No reaction	

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Chemistry Topic 5
Chemical changes

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Section 3: Extracting Metals

Very unreactive metals e.g. Silver and gold	Found **naturally** in the ground. Extracted using **mining**.
Metals **less reactive than carbon** e.g. Zinc, Iron & Lead	Metals less reactive than carbon can be extracted from their ores by **reduction** using carbon, coke or charcoal. $2PbO(s) + C(s) \rightarrow 2Pb(s) + CO_2(g)$ Carbon has displaced lead from its oxide because carbon is more reactive than lead. This extraction takes place in a **blast furnace** at high temperature.
Metals **less reactive than hydrogen** e.g. Tungsten	Metals less reactive than hydrogen can be extracted from their ores by **reduction** using hydrogen. Tungsten is obtained from its oxide by reduction using hydrogen. $WO_3(s) + 3H_2(g) \rightarrow W(s) + 3H_2O(g)$
Metals more reactive than carbon e.g. Aluminium	Extracted by **electrolysis**.

Section 4a: Salts from metals (neutralisation reactions)

With metal	Acid + Metal → Salt + Hydrogen $2HCl(aq) + Fe(s) \rightarrow FeCl_2(aq) + H_2(g)$
With alkali	Acid + Metal Hydroxide → Salt + Water $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
With metal oxide	Acid + Metal Oxide → Salt + Water $2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$
With carbonate	Acid + Metal Carbonate → Salt + Water + Carbon Dioxide $2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Section 4b: Making a Soluble Salt

A salt is a compound formed when the hydrogen in an acid is wholly, or partially, replaced by metal or ammonium ions.
Salts are made when a suitable metal, metal carbonate, metal oxide or metal hydroxide is reacted with acid.

Crystallisation	Evaporation
Pure dry crystals can be obtained from solution by: • **Add solid** metal, metal carbonate, metal oxide or metal hydroxide **to an acid**. • Add solid **until no more reacts** (saturated solution). • **Filter** off excess solid. • **Evaporate** to remove some of the water. • Leave to **crystallise**. • Filter the crystals • Leave to dry **in air**/in a **desiccator/oven**.	When you react an acid with an alkali, you need to be able to tell when the acid and alkali **have completely reacted**. Then you can collect pure dry crystals of the salt. • Carry out an **acid/alkali titration** using an indicator to see how much acid **reacts completely** with alkali • **Run that volume of acid again** into solution of alkali but **without indicator**. • Pour solution into evaporating basin • Heat • **Leave to crystallise** / boil off water

Section 5: Strong and weak acids

Aqueous solutions of **weak acids have higher pH** than solutions **of strong acids with the same concentration**. Strong acids **completely ionise** in solution to produce hydrogen ions. e.g. $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
Weak acids **only partially ionise** in solution. The reaction is **reversible** (unlike the ionisation of strong acids.) So as the molecules of the weak acid split up to form its ions, the ions recombine to form the original molecule.
e.g. Ethanoic acid: $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
A position of **equilibrium** is reached in which both the original molecule (majority) and its ions (minority) are present.

Measuring acidity or alkalinity

Indicators are substances that change colour when you add an acid or an alkali. Litmus is an indicator that turns red in acid and blue in alkali. You can also use a pH meter which gives a digital reading of pH.

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Chemistry Topic 6 Electrolysis

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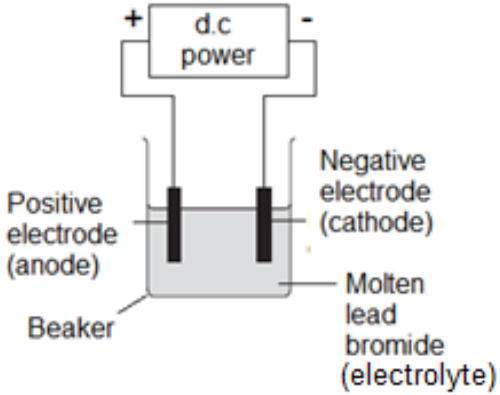
Section 1 Electrolysis key terms	
Electrolysis	The process of **splitting an ionic compound** by passing **electricity** through it.
Electrolyte	An **ionic compound** that is **molten** (melted) or **dissolved in water**. The electrolyte is broken down by electricity enabling its **ions** to and hence carry a charge. **move freely**
Electrode	An **electrical conductor** that is placed in the **electrolyte** and connected to the **power supply**.
Cathode	The **negative electrode**. The electrode attached to the negative terminal of the power supply.
Anode	The **positive electrode**. The electrode attached to the positive terminal of the power supply.
Oxidation	Loss of electrons
Reduction	Gain of electrons



Positive
Anode
Negative
Is
Cathode

Section 2a: Changes at the electrodes – Pure ionic compounds		
Electrolyte	**Cathode**	**Anode**
Molten Compound	Metal	Non-metal produced.
Molten lead bromide (diagram above)	**Lead metal** is produced $Pb^{2+} + 2e^- \rightarrow Pb$	**Bromine** is produced $2Br^- \rightarrow Br_2 + 2e^-$

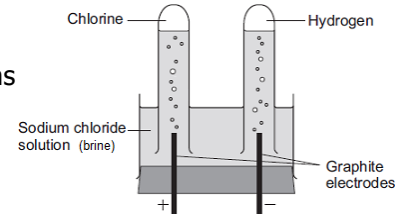
Section 2b: Changes at the electrodes – Aqueous solutions		
Electrolyte	**Cathode**	**Anode**
Dissolved compound (aqueous solution)	The **metal** if the metal is **less reactive than hydrogen**. **Hydrogen** is produced if the **metal is more reactive than hydrogen**.	**Oxygen** is produced **unless the solution contains halide ions** (chloride, bromide, iodide) when the **halogen** (chlorine, bromine, iodine) is produced.

Electrolyte	Cathode	Anode
$CuBr_{2(aq)}$	Copper	Bromine
$NaCl_{(aq)}$	Hydrogen	Chlorine
$KI_{(aq)}$	Hydrogen	Iodine
$Na_2SO_{4(aq)}$	Hydrogen	Oxygen

Electrolysis of Brine (concentrated sodium chloride solution)

In the electrolysis of brine, **three products** are formed, **hydrogen, chlorine** and **sodium hydroxide.**

Sodium chloride solution → hydrogen gas + chlorine gas + sodium hydroxide solution



At the **cathode hydrogen** gas forms
$2H^+ + 2e^- \rightarrow H_2$ (**reduction**)

At the **anode**, **chlorine** gas forms
$2Cl^- \rightarrow Cl_2 + 2e^-$ (**Oxidation**)

Sodium ions stay in solution (as sodium is more reactive than hydrogen) and **combine with hydroxide ions** to form sodium hydroxide.
$Na^+ + OH^- \rightarrow NaOH$

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Chemistry Topic 6 Electrolysis

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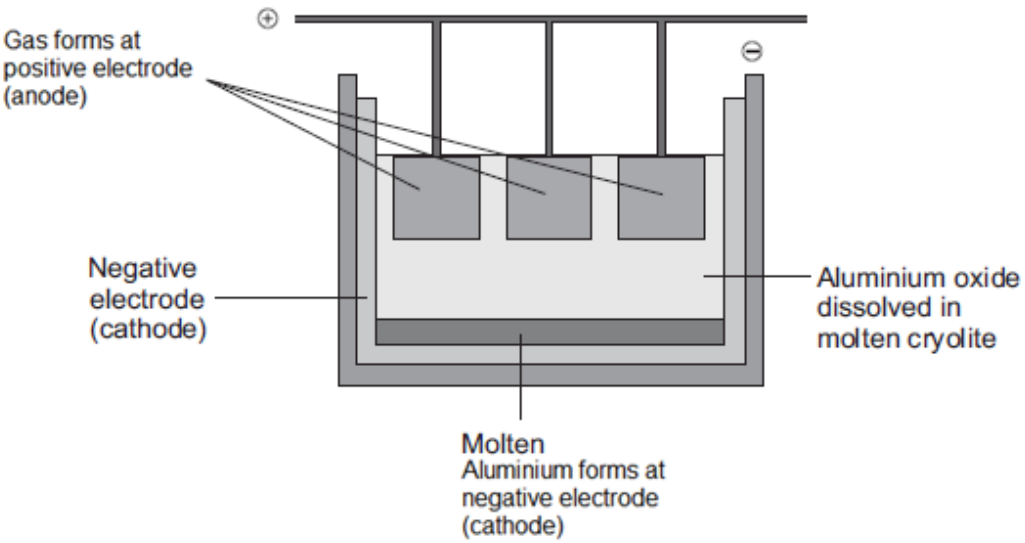
Section 3a: The extraction of Aluminium by electrolysis	
Bauxite	You get aluminium oxide from the ore called **Bauxite**, the ore is mined by **open cast mining**.
Cryolite	**Aluminium oxide** is **dissolved in cryolite** to **lower its melting point**. This **saves money on energy costs**.
Graphite	The **electrodes** are made from **graphite** (carbon) as graphite can conduct electricity (due to it having delocalised electrons between it's layers.)
Cathode	Positive **Al^{3+} ions move to the cathode**. Aluminium is produced (reduction). **$Al^{3+} + 3e^- \rightarrow Al$**
Anode	Negative **O^{2-} ions move to the anode**. Oxygen is made (oxidation). **$2O^{2-} \rightarrow O_2 + 4e^-$** The anode **wears away** gradually as the **carbon graphite anode reacts with oxygen to form carbon dioxide**.



Section 3b: Uses of Aluminium

Aluminium is a very important metal, the uses of its metal or alloys include:

- Pans
- Overhead power cables
- Aeroplanes
- Cooking foil
- Drink cans
- Window and patio door frames
- Bicycle frames and car bodies

KNOWLEDGE

Chemistry Topic 7
Energy changes

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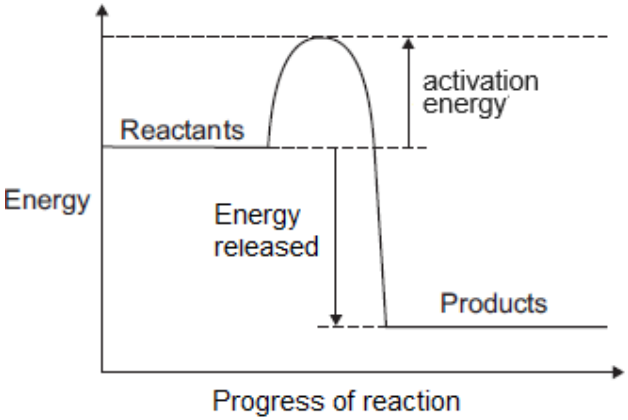
Section 1 Energy Changes Key Terms

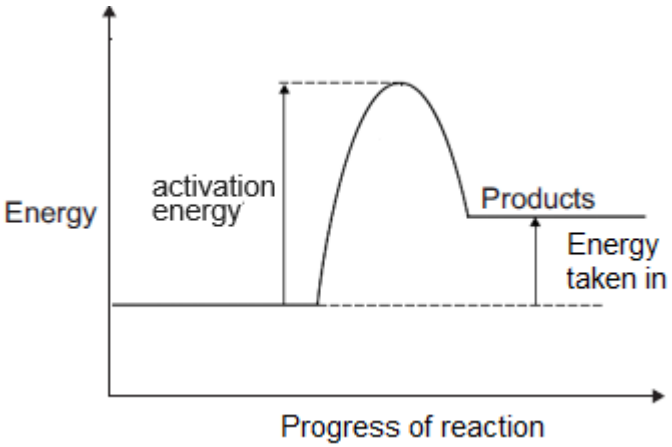
Conservation of energy	Energy is **neither created or destroyed**, only **transferred from one store to another**
ΔH	**Change in energy** of a system in a reaction, its units are KJ/mol
Exothermic	A reaction that **transfers energy to the surroundings** so the **temperature of the surroundings increases**, e.g. **combustion** and **neutralisation** reactions. Used in **self-heating cans** and **hand warmers**. Has a negative value of ΔH
Endothermic	A reaction that **takes in energy from the surroundings** so the **temperature of the surroundings decreases**, e.g. **thermal decomposition**. Used in **sports injury packs. Has a positive value of** ΔH
Activation energy	The **energy needed** for **particles to successfully react**.

Section 2a Reaction profiles – Exothermic reaction

The products are at a **lower energy** than the reactants. This means that energy has been **transferred to the surroundings**. Hence the surroundings gets **hotter** and the temperature **rises**.



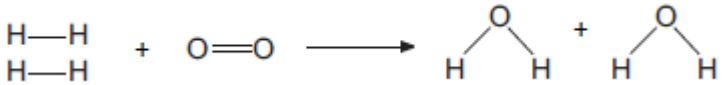
Section 2b Reaction profiles – Endothermic reaction



The products are at a **higher energy** than the reactants. This means that energy has been **transferred from the surroundings**. Hence the surroundings gets **colder** and the temperature **decreases**.

Section 3 Bond breaking and making (HT)

Breaking bonds	**Energy is needed** to break bonds (Endothermic).
Forming bonds	**Energy is released** when bonds are formed (Exothermic).



Hydrogen and oxygen react together to make water. The bonds between hydrogen and oxygen have to be broken so that new bonds can form between hydrogen and oxygen.

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Chemistry Topic 7
Energy changes

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Section 4 Bond energy calculations (HT)

You can calculate the **overall energy change** in a chemical reaction **using bond energies**. *Bond energy data will be given to you in an exam, hence you don't need to revise the data.*

Equation:
Bond energy = energy required to break bonds in the reactants – energy required to make bonds in the products

1) Use bond energies to estimate the overall energy change for the reaction:

$$H_2 + Cl_2 \rightarrow 2HCl$$

H–H + Cl–Cl → H–Cl + H-Cl
Bonds broken - Bonds made
1 H – H 436 2H – Cl 2x432
1 Cl – Cl 243
+ 436 + 243 – (2 x 432)
679 – 864
ΔH = -185 KJ/mol

Figure 1

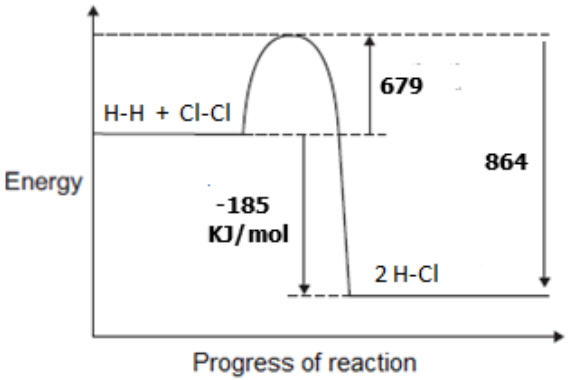

Figure 1 shows the **energy profile diagram** for the reaction between hydrogen and chlorine.

679 KJ/mol of energy was **taken in** when the **reactants bonds** where **broken**, 864 KJ/mol of energy was **released** when the **products bonds** where **formed**, hence the **overall energy of the reaction was -185 KJ/mol.**

Because the energy change ΔH is **negative**, energy was **transferred to the surroundings** in an **exothermic** reaction.

Common Bond energies KJ/mol	
C-C	347
C-O	358
C-H	413
C-N	286
C-Cl	346
Cl-Cl	243
H-Cl	432
H-O	464
H-N	391
H-H	436
O=O	498
N≡N	945