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Chemistry Topic C4
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 reduced if they **lose oxygen** in a reaction. Chemicals are reduced 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 **neutralises** 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: $H^+ + OH^- \rightarrow H_2O$
pH	A scale to **measure acidity/ alkalinity**. A 10x increase in concentration of H^+ ions causes a decrease of one pH unit (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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Section 3: Extracting Metals

Very unreactive metals e.g. Silver and gold	Found **naturally** in the ground. Extracted using **mining**.
Metals l**ess 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

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.**

Evaporation

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.

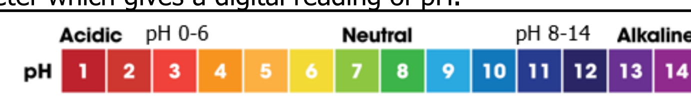
- Carry out an **acid/alkali titration** using an indicator to see how much acid **reacts completely** with alkali
- **Run that volume of acid again** into solution of alkali but **without indicator**.
- Pour solution into evaporating basin
- Heat
- **Leave to crystallise** / boil off water

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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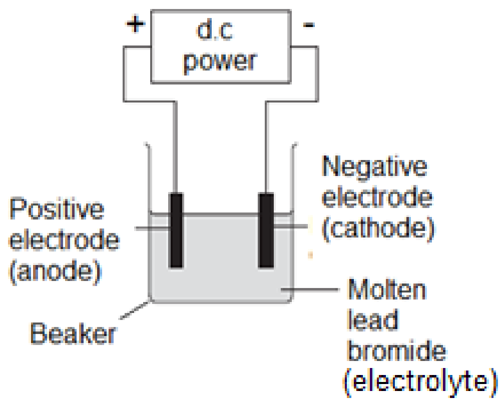
Section 1 Electrolysis key terms	
Electrolysis	The process of **splitting an ionic compound** by passing **electricity** through it.
Electrolyte	An **ionic compound** that is **molten** (melted) or **dissolved in water**. The electrolyte is broken down by electricity enabling its **ions** to and hence 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 $2Br^- \rightarrow Br_2 + 2e^-$

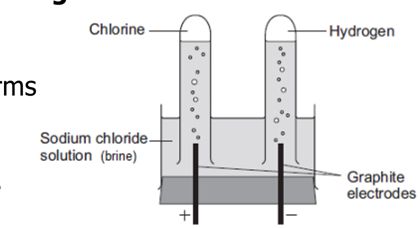
Section 2b: Changes at the electrodes – Aqueous solutions		
Electrolyte	**Cathode**	**Anode**
Dissolved compound (aqueous solution)	The **<u>metal</u>** if the metal is **less reactive than hydrogen**. **<u>Hydrogen</u>** is produced if the **metal is more reactive than hydrogen.**	**<u>Oxygen</u>** is produced **unless the solution contains halide ions** (chloride, bromide, iodide) when the **<u>halogen</u>** (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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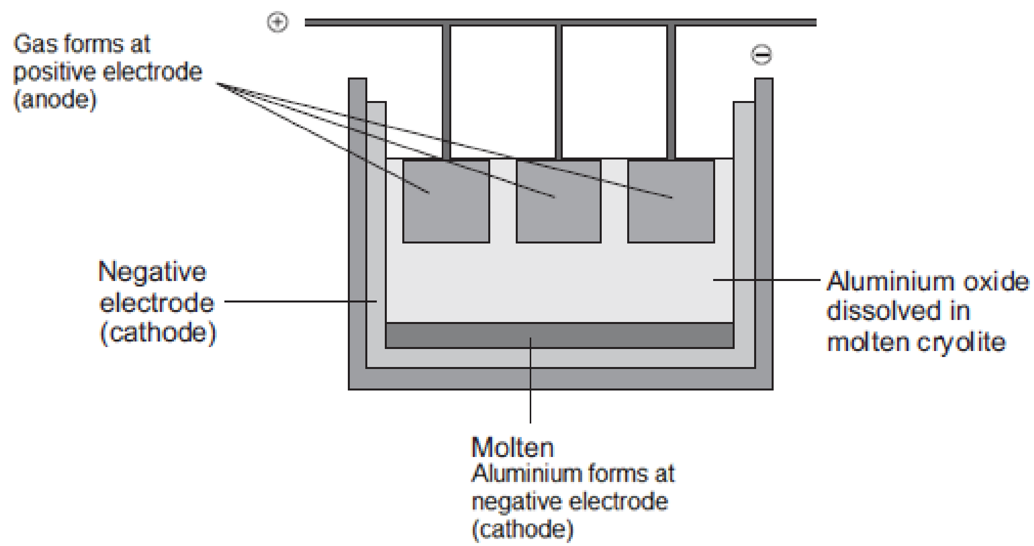
Section 3a: The extraction of Aluminium by electrolysis	
Bauxite	You get aluminium oxide from the ore called **Bauxite**, the ore is mined by **open cast mining**.
Cryolite	**Aluminium oxide** is **dissolved in cryolite** to **lower its melting point**. This **saves money on energy costs**.
Graphite	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