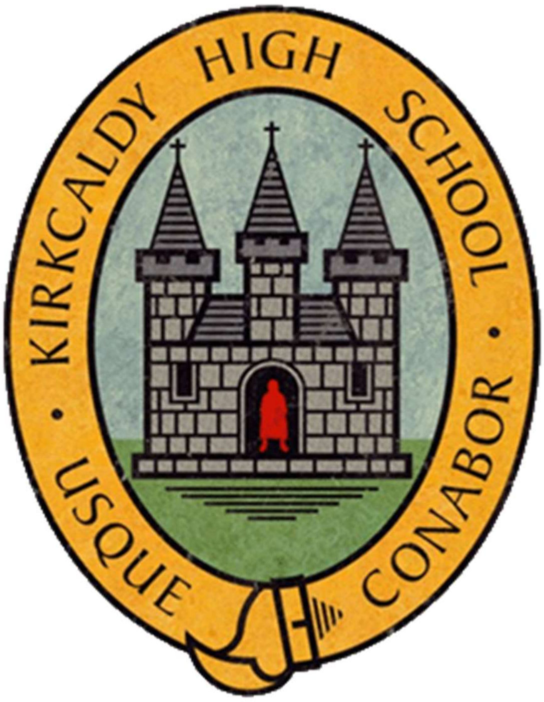
Kirkcaldy High School



Chemistry

National 5

Unit 3 – Chemistry in Society

NOTES

Course Overview

Contents

The National 5 Chemistry Course is split into three units. *Italic* shows the contents of this notes booklet.

Unit 1 – Chemical Changes in Structure

(a) Rates of reaction
(b) Atomic structure and bonding related to properties of materials
(c) Formulae and reacting quantities
(d) Acids and Bases

Unit 2 – Nature's Chemistry

(a) Homologous Series
(b) Everyday consumer products
(c) Energy from Fuels

Unit 3 – Chemistry in Society

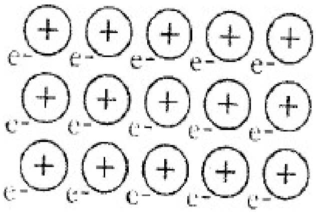
Assessment

- There is a **final exam** for National 5 Chemistry. It lasts for **2 h 30 mins** and contains **100 marks**. There are:
    - 25 marks of multiple-choice questions
    - 75 marks of written answer questions
- There is an **Assignment** (essay on a Chemistry topic) to write for National 5 Chemistry.
    - The Assignment is **externally marked** by an SQA marker (NOT your teacher)
    - The Assignment contains **20 marks,** but these are scaled to 25 marks such that the assignment is worth one quarter of the total marks.
    - You have unlimited **time to research** and gather data for your Assignment.
    - You can **1 h 30 mins to write** the Assignment.

(a) Metals

Metallic Bonding

Metallic bonding is between atoms of **metal elements** (left hand side of periodic table).



- Metallic bonding is an electrostatic force of attraction between positive metal ions and a 'sea' of negative electrons (called delocalised electrons).
- All metals **conduct electricity** because the *electrons are free to move*.
- Electricity is a flow of charged particles.

Reactivity Series

- Metals react with a variety of substances to form compounds, but some metals react more readily than others.
- Those which react very easily are termed **reactive**.
- Those which do not react readily are termed **unreactive or inactive**.
- The reactivity series places the metals in order of reactivity, **most reactive at the top** and **least reactive at the bottom**.
- Reaction of Metals - Not all metals undergo all these reactions – it depends on their reactivity (page 10 data book)

Reaction of metals with oxygen

- For all metals above copper in the electrochemical series (data book p10)

Metal + Oxygen → Metal Oxide

- *e.g.*

Magnesium + Oxygen → Magnesium Oxide

- Chemical formula equation

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

- Ionic equation

$$2Mg(s) + O_2(g) \rightarrow 2Mg^{2+}O^{2-}(s)$$

Reaction of metals with water

- For all metals above magnesium in the electrochemical series (data book p10)

Metal + Water → Metal Hydroxide + Hydrogen (*Magnesium and above*)

- *e.g.*

Lithium + Water → Lithium Hydroxide + Hydrogen

- Chemical formula equation

$$2Li(s) + H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$$

- Ionic equation

$$2Li(s) + H_2O(l) \rightarrow 2Li^{+} (aq) + 2OH^{-} (aq) + H_2(g)$$

Reaction of metals with acid

- For all metals above hydrogen in the electrochemical series (data book p10)

Metal + Acid → Metal Salt + Hydrogen

- *e.g.*

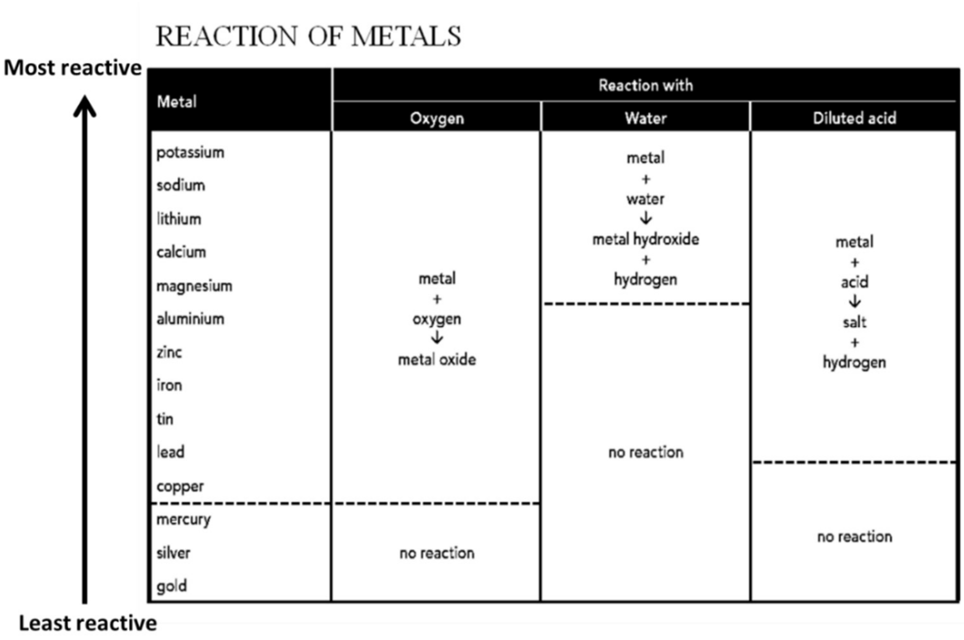
Magnesium + Hydrochloric Acid → Magnesium Chloride + Hydrogen

- Chemical formula equation

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

- Ionic equation

$$Mg(s) + 2H^{+} (aq) + 2Cl^{-} (aq) \rightarrow Mg^{2+} (aq) + 2Cl^{-} (aq) + H_2(g)$$



REACTION OF METALS

Most reactive

Metal	Reaction with Oxygen	Reaction with Water	Reaction with Diluted acid
potassium	metal + oxygen → metal oxide	metal + water → metal hydroxide + hydrogen	metal + acid → salt + hydrogen
sodium			
lithium			
calcium			
magnesium			
aluminium		no reaction	
zinc			
iron			
tin			
lead			
copper			no reaction
mercury	no reaction		
silver			
gold			

Least reactive

Redox

Oxidation and Reduction

- OIL RIG
    - Oxidation is Loss (of electrons)
    - Reduction is Gain (of electrons)
    - **Red*ox*** reactions involve both **reduction** and ***ox*idation** reactions

Writing redox equations

- Ion-electron equations (page 10 data booklet) – written as reduction equations (left to right); flip around and write right to left for oxidation equations
- Metal ions are formed from the loss of electrons [electrons go on the same side as the ions]
- Non-metal ions are formed from the gain of electrons [electrons go on the same side as the atoms]

- *e.g.*

$Li^+ (aq) + e^- \rightarrow Li (s)$ [reduction – gain of electrons]

$2I^- (aq) \rightarrow I_2 (s) + 2e^-$ [oxidation – loss of electrons]

- Redox equations – combine oxidation and reduction ion-electron equations, MUST ensure they have the same number of electrons on the LHS and RHS before combining.
- Redox equations do **NOT** contain electrons or spectator ions

Combining Ion-electron equations by cancelling electrons

- *e.g.*

Oxidation: $Al (s) \rightarrow Al^{3+}(aq) + 3e^-$

Reduction: $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$

- To ensure they both have the same number of electrons multiply the oxidation equation by the number of electrons in the reduction equation and vice versa.

Oxidation:	$2Al (s)$	$\rightarrow$	$2Al^{3+}(aq) + \cancel{6e^-}$	**(x2)**
Reduction:	$6H^+ (aq) + \cancel{6e^-}$	$\rightarrow$	$3H_2 (g)$	(x3)
Redox Equation	$2Al(s) + 6H^+ (aq)$	$\rightarrow$	$2Al^{3+}(aq) + 2H_2 (g)$	

Removing Spectator Ions from Redox Equations

- *e.g.*

Chemical formula equation	**$Zn(s) + NiCl_2\ (aq)$**	→	**$ZnCl_2\ (aq) + Ni(s)$**	
Ionic equation	$Zn(s) + Ni^{2+}(aq) + 2Cl^-\ (aq)$	→	$Zn^{2+}\ (aq) + 2Cl^- + Ni(s)$	
Redox equation (spectator ions removed)	$Zn(s) + Ni^{2+}(aq)$	→	$Zn^{2+}\ (aq) + Ni(s)$	
Oxidation ion-electron equation	$Zn(s)$	→	$Zn^{2+}\ (aq) + 2e^-$	(e^- on same side as metal ions)
Reduction ion-electron equation:	$Ni^{2+}(aq) + 2e^-$	→	$Ni(s)$	(e- on same side as metal ions)

- *e.g.*

Chemical formula equation:	**Cl_2 (g) + 2NaBr (aq)**	→	**2NaCl (aq) + Br_2 (l)**	
Ionic equation:	Cl_2 (g) + $2Na^+$(aq) + $2Br^-$ (aq)	→	$2Na^+$+ $2Cl^-$ (aq) + Br_2 (l)	
Redox equation (spectator ions removed):	Cl_2 (g)+ $2Br^-$ (aq)	→	$2Cl^-$ (aq) + Br_2 (l)	
Oxidation ion-electron equation:	$2Br^-$ (aq)	→	Br_2 (l) + $2e^-$	(e- on same side as non-metal atoms)
Reduction ion-electron equation:	Cl_2 (g)+ $2e^-$	→	$2Cl^-$ (aq)	(e- on same side as non-metal atoms)

displacement reactions

- When a metal reacts with an acid it **DISPLACES** the hydrogen.
- *e.g.*

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

- Displacement reactions are examples of **redox** reactions
- *e.g.*

$$Mg + CuSO_4 \rightarrow MgSO_4 + Cu$$

- The magnesium has displaced (pushed out) the copper
- A more reactive metal (higher in the Electrochemical Series on Data Book p10) can **displace** a less reactive metal (lower in the Electrochemical Series) from a solution.

Example: aluminium will displace zinc from solution, but it will not displace magnesium from solution

Extraction of Metals

- Most metals are found combined with other elements in **metal ores**
- Changing metal ions to metal atoms = reduction – this happens in the **extraction of metals**
- There are 3 methods of extraction:
- **Electrolysis:** Most reactive metals, (the more stable its compound and the harder it is to extract from its ore) – Metals **above** and including **Aluminium** (potassium, sodium, lithium, calcium, magnesium, aluminium)
- Heating with carbon or carbon monoxide: Zinc to Copper (zinc, iron, nickel, lead, copper)
- The carbon is doing the reducing (removing the oxygen) so it is described as the **reducing agent.**

Metal oxide + Carbon → Metal + Carbon dioxide

- **Heating alone:** (decomposition) Least reactive metals - Metals **below copper** (mercury, silver, gold and platinum)

	Metal	**Process by which metal is extracted from ores**
Most Reactive	Potassium	Electrolysis
↑	Sodium	
	Lithium	
	Calcium	
	Magnesium	
	Aluminium	
	Zinc	Heat plus carbon or carbon monoxide
	Iron	Blast Furnace
	Tin	Heat plus carbon or carbon monoxide
	Lead	
	Copper	
	Mercury	Heat alone
	Silver	
Least Reactive	Gold	

Electrolysis

- **Electrolysis** is using **electricity** to separate a compound into its elements.
- Compound is heated, the ionic lattice breaks down and the ions are free to move.

- The positive metal ions are attracted to the negative electrode and gain electrons to form metal atoms - *i.e.* the metal ions are reduced.
- *e.g.*

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

- The negative non-metal ions are attracted to the positive electrode and lose electrons to form non-metal atoms/molecules i.e. the non-metal ions are oxidised.
- *e.g.*

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

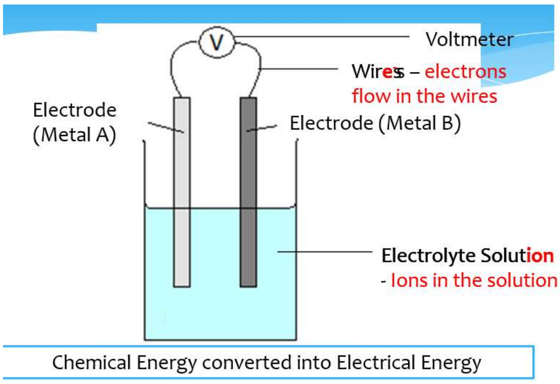
Heating alone

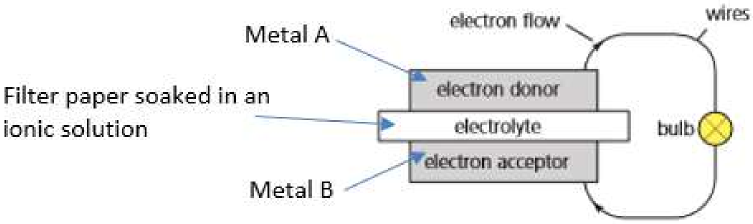
- decomposition reaction

Metal oxide → Metal + Oxygen

Electrochemical Cells

Simple cells:

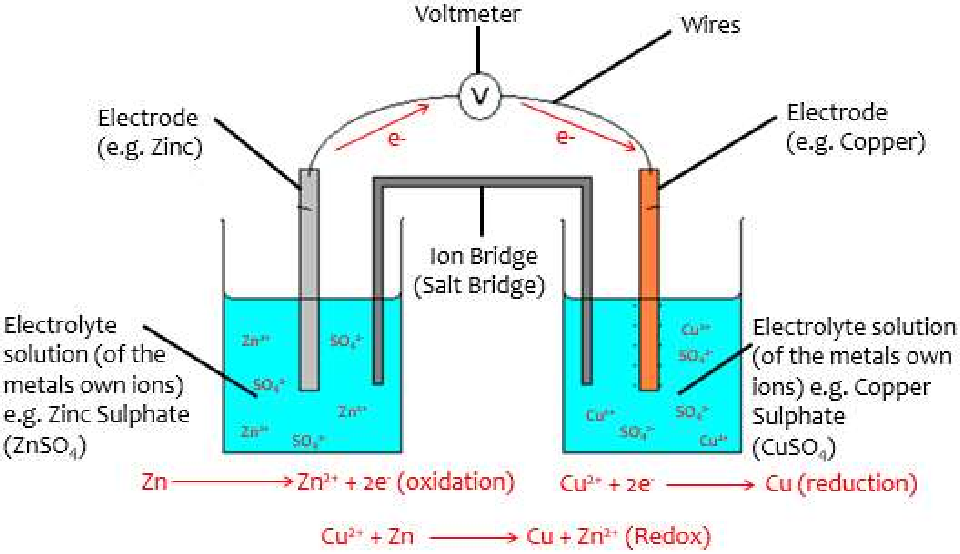




- Made up of:
- **Two electrodes** = two different metals.
- **Wires** = connect the electrodes - **electrons** flow from one electrode to the other electrode through the wires.
- **Electrolyte** solution = conducting solution of ions (salt water, acid, alkali). This **completes** the circuit because it allows the ions to flow.
- Electric current is seen through a voltmeter in the circuit (or a bulb or buzzer)

- ** The further apart the metals are in the electrochemical series, the higher the voltage. **
- **Electrons flow from the metal higher in the series to the metal lower in the series **
- If the two metal electrodes are the same there is no flow = 0 voltage

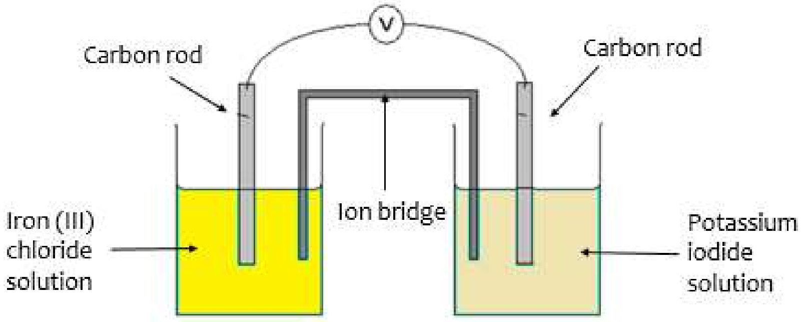
Complex cells – redox cells:



- Electron flow from metal higher in the Electrochemical Series (Data Book p10) to metal lower in the Electrochemical Series – **electrons ONLY** flow through the wires (the 'external circuit')
- The purpose of the **ion bridge** (salt bridge) is to complete the circuit and connect the two solutions – it allows the **ions** to flow
- The direction of the current tells you which metal is oxidised. Electrons **always flow from the metal that is oxidised.**
- If the cell were left for a long time the zinc electrode would decrease in mass and the copper electrode would increase in mass

Non-metal electrodes

- Non-metal electrodes –must **conduct electricity** and be **unreactive** = graphite (carbon) rods
- *e.g.*



- Iodide ions lose electrons to the carbon electrode which flow through the voltmeter to the iron ions in the other beaker, the iron ions gain electrons from the carbon electrode.

Iodide ions are oxidised: $2I^- \rightarrow I_2 + 2e^-$

Iron (III) ions are reduced: $Fe^{3+} + e^- \rightarrow Fe^{2+}$

Redox Equation $Fe^{3+} + 2I^- \rightarrow Fe^{2+} + I_2$

- Using the Electrochemical Series (Data Book p10) the oxidation equation is above the reduction equation

Rechargeable Batteries

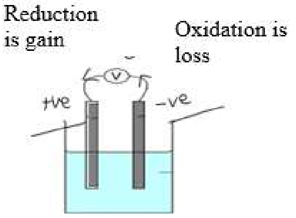
- Non-rechargeable (dry-cell – reactants get used up over time and it stops working)
- Rechargeable (can be regenerated once flat, electrical energy **reverses the redox reaction**, the original reactants are regenerated and can be **re-used**)

When a rechargeable battery is being used to make electricity, it is like any other battery ... an oxidation reaction provides electrons at the negative electrode and a reduction reaction accepts electrons at the positive electrode.

negative electrode: $X \longrightarrow X^+ + e^-$ (oxidation)

cell in use

positive electrode: $Y^+ + e^- \longrightarrow Y$ (reduction)

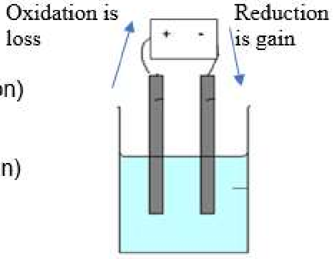


However, when it is being charged, mains electricity is being used to reverse the chemical reactions in order to regenerate the reactants.

negative electrode: $X^+ + e^- \longrightarrow X$ (reduction)

on charge

positive electrode: $Y \longrightarrow Y^+ + e^-$ (oxidation)



(b) Plastics

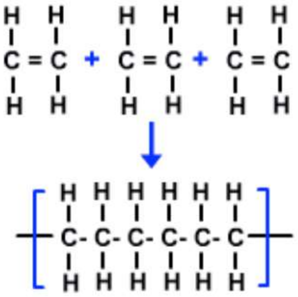
(i) Addition Polymerisation

- Plastics are a group of important materials. They are long-chain (big) molecules called **<u>polymers</u>**.
- **Polymers** are made from small molecules (units) called **<u>monomers</u>** joined.
- They are made in a process called **<u>polymerisation</u>** and their names can be deduced from the name of the monomer.
- *e.g.*

  monomer = ethene; polymer = poly(ethene) – *i.e.* put ‘poly’ in front of the monomer name
- Thermosoftening plastics can be reshaped once heated as they melt (and can be recycled)
- Thermosetting plastics cannot be reshaped once moulded
- Plastics can be made from alkenes in **addition polymerisation** reactions,
- **Addition polymerisation** – a chemical reaction in which several small **unsaturated** molecules (*e.g.* alkenes) join to form a long chain molecule. **No other product is formed**.
- The monomers used all contain a carbon to carbon double bond.

Representation of the structure of monomers and polymers

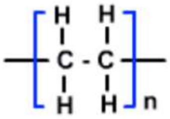
- *e.g.* polyethene (trade name: polythene) made from ethene (alkene) monomers



Monomers

Polymer – MUST include a bond off each end carbon, i.e. make sure each carbon has 4 bonds

Double bonds are all broken, and it has ONLY carbon atoms in its backbone.



Repeating unit = the smallest part of the structure which when repeated again and again gives the structure of the polymer chain

It does not contain a double bond as it is of the polymer

Make sure each end carbon has a bond coming off it (often it is the monomer without the double bond, or two different monomers combined)

(c) Fertilisers

(i) Commercial production of fertilisers

- We need a good supply of healthy crops to meet the demands of the growing population.
- Fertilisers are used to provide plants with the essential elements needed for healthy plant growth.

- The three essential elements are:

Nitrogen (N)
Phosphorous (P)
Potassium (K)

- Remember 'ammonium' contains nitrogen (NH_4)

- Fertilisers must be ***soluble*** in water to be taken up through the roots of plants (data book pg8)

Natural (Organic) Fertilisers

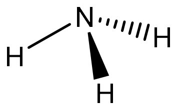
- These are substances that are obtained from plant and animal waste. *Examples:* Compost and Manure

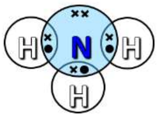
Artificial (Synthetic) Fertilisers

- Made by chemists (manmade) from ***neutralisation*** reactions and can be designed for a specific crop. *e.g.* Ammonium Nitrate (NH_4NO_3)

Ammonia

- Nitrogen trihydride
- Chemical formula: NH_3
- Bonding electron diagram:
- Shape: Trigonal Pyramidal (make sure it looks like a pyramid with N above all 3 H's)





Haber and Ostwald Processes

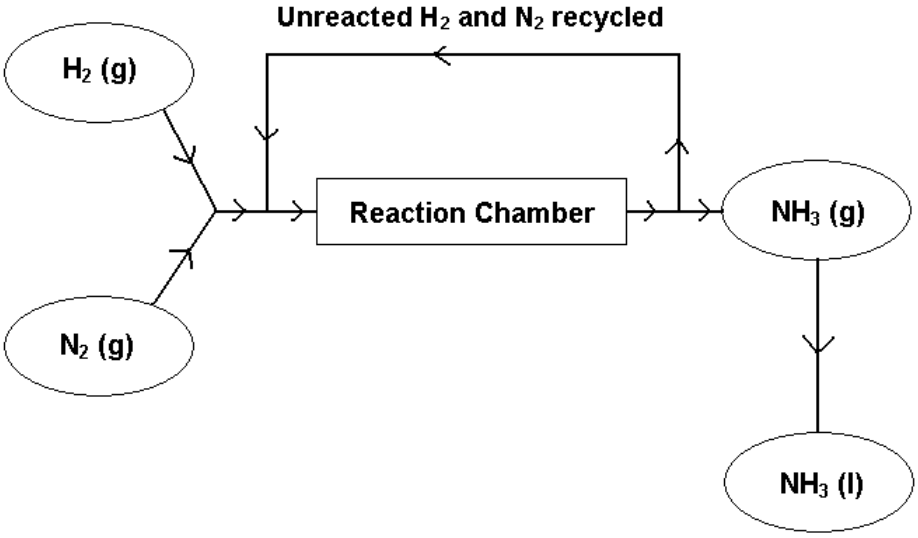
Industrial Manufacture of Ammonia:

- Process = Haber Process

- Made from reaction of nitrogen and hydrogen in a reversible reaction (shown by arrow)

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

- The reaction is carried out in a sealed vessel, at high pressure and an ***iron catalyst*** is used to get a fast reaction at a low temperature.



- This process is '**economical**' because the unreacted reactants are recycled and reused.

Uses of Ammonia

- It is a starting material (feedstock) in the production of **nitric acid (OSTWALD PROCESS).**
- It can be used to make ammonium salts used as fertilisers.

Industrial Production of Nitric Acid

- Nitric acid = HNO_3
- Process = Ostwald Process
- Ammonia is **oxidised** to nitric oxide (NO (g)) by passing a mixture of ammonia and air over a heated **platinum** catalyst.

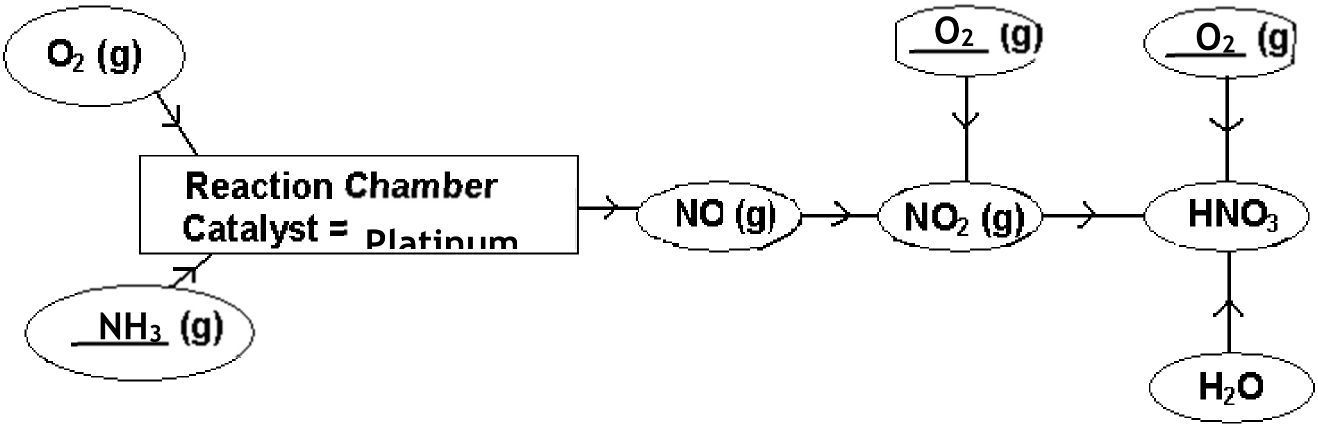
Ammonia + Oxygen → Nitrogen Monoxide + Water

$4NH_3 (g) + 5O_2 (g) \rightarrow 4NO (g) + 6H_2O (l)$

- Nitric oxide (NO (g)) is **oxidised** using oxygen in the air to form brown fumes of nitrogen dioxide (NO_2 (g)).

Nitrogen Monoxide + Oxygen → Nitrogen Dioxide

$2NO (g) + O_2 (g) \rightarrow 2NO_2(g)$



- Nitrogen dioxide (NO_2 (g)) is **dissolved in water** in the presence of air to form nitric acid (HNO_3).

Nitrogen Dioxide + Water + Oxygen → Nitric Acid

$4NO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 4HNO_3 (aq)$

Making Salts as Fertilisers

- Formation of salts are examples of **neutralisation** reactions

Acid + Base → Salt + Water

- Naming a salt: **first** part of the name comes from the **base**, **second** part of the name comes from the **acid**
- Nitric Acid + *Potassium* Hydroxide → *Potassium* Nitrate + Water

Acid Name	Salt name ends in
Phosphoric acid	Phosphate
Hydrochloric acid	Chloride
Sulphuric acid	Sulphate
Nitric acid	Nitrate

- **Ammonia is an alkaline gas** when it reacts with an acid it forms an ammonium salt
- *e.g.*

Ammonia + Nitric Acid → Ammonium Nitrate

$$NH_3\ (g) + HNO_3\ (aq) \rightarrow NH_4NO_3\ (aq)$$

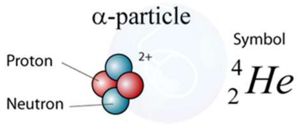
(d) Nuclear Chemistry

(i) Radiation

- Nuclear chemistry is the study of the break-up of unstable nuclei into more stable nuclei, which results in the emission of radiation and the release of energy.
- The stability of a nucleus depends on the proton to neutron ratio.

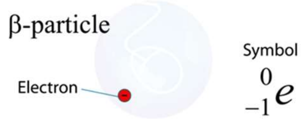
Three types of radiation:

Alpha (α): *Helium nuclei* containing two protons and two neutrons (2+ charge)



Alpha particle is nucleus of helium

- **Beta (β):** *Electrons* formed from a neutron breaking up into a proton and electron (1- charge)



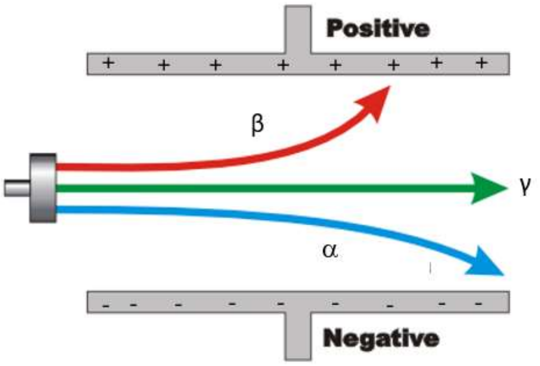
Beta particle is high speed electron

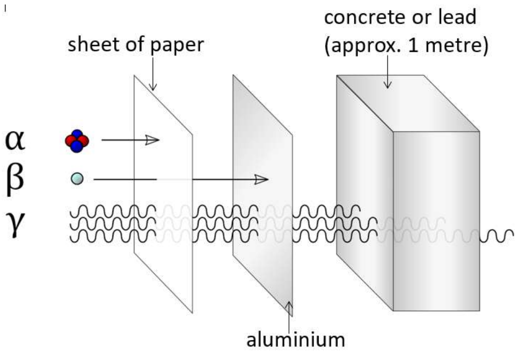
- **Gamma (γ):** *Electromagnetic wave* (no charge)

Gamma ray is electromagnetic wave

Radiation Type	Chemical Symbol	Charge	Relative Mass	Stopped By	Nature
Alpha (α)	$^{4}_{2}He$	2+	4	Paper	Helium nuclei
Beta (β)	$^{0}_{-1}e$	1-	0	Aluminum	Electron
Gamma (γ)	$^{0}_{-0}\gamma$	0	0	Greatly reduced by concrete or lead	Wave

- Alpha, beta and gamma radiation are all **ionising radiations** which can damage living cells
- Ionising radiation can break molecules into ions that can take place in alternative chemical reactions in living cells resulting in living cells dying or becoming cancerous
- Ionising ability (most) Alpha > Beta > Gamma (least)



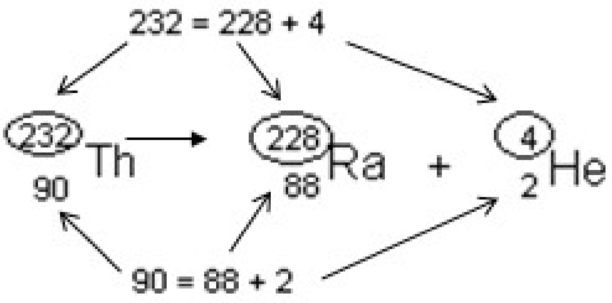


Nuclear equations

- In a balanced nuclear equation, the total mass number on the left hand side of the equation is equal to the total mass number on the RHS. The same is true for the atomic number.
- If the atomic number changes – this means the element symbol also changes

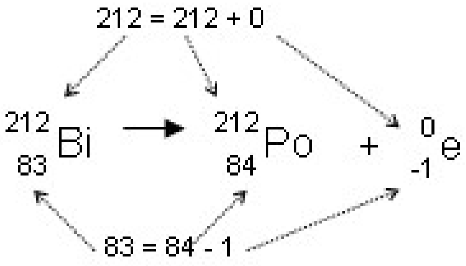
Alpha Decay

- When a nucleus emits an alpha particle its atomic number will decrease by 2 (loss of two protons) and its mass number will decrease by 4 (loss of two protons and two neutrons)
- *e.g.*



Beta Decay

- When a nucleus undergoes beta decay its atomic number will increase by 1 but its mass number will remain the same
- *e.g.*



Gamma Decay

- Gamma rays have no mass and no charge, emission of gamma rays has no effect on the mass number or atomic number of a radioactive atom.

Half life

- Radioactive decay is a random event.
- The rate of decay of a radioisotope is measured by a period of time known as the **half-life**

Half-life= $t_{\frac{1}{2}}$

- The half-life is the time it takes for half of the starting nuclei of a isotope to decay.
- After one half-life has passed, 50% of the starting nuclei have decayed and changed into new isotopes.
- The half-life, for a isotope, is constant so radioactive isotopes can be used to date materials.

- *e.g.* $t_{\frac{1}{2}}$ = 14 days, 10 g mass at the start

Number of half-lives	**0**	**1**	**2**	**3**
Time (days)	0	14	28	42
Mass of original sample (g)	10	5	2.5	1.25
Mass of product (g)	0	5	7.5	8.75
Percentage (%) of original sample	100	50	25	12.5
Fraction of original sample	1	$\frac{1}{2}$	$\frac{1}{4}$	$\frac{1}{8}$

- *e.g.*
- A radioisotope of phosphorus has a half-life of 14 days, a sample of the radioisotope has a mass of 80 g, calculate the remaining mass of the sample after 56 days?

Starting quantity = 80 g
End quantity = ? g
Half life = 14 days
Total time = 56 days
Number of half-lives = ?

$$\mathbf{80\text{ g}}\ \overbrace{\xrightarrow{14\text{ days}}\ \mathbf{40\text{ g}}\ \xrightarrow{14\text{ days}}\ \mathbf{20\text{ g}}\ \xrightarrow{14\text{ days}}\ \mathbf{10\text{ g}}\ \xrightarrow{14\text{ days}}}^{\textbf{These add up to 56 days}}\ \mathbf{5\text{ g}}$$

The remaining mass of the sample would be 5 g

Uses of radioisotopes

Medical Uses

- Cancer cells can be destroyed by γ-radiation from the isotope **cobalt-60**
- Skin cancer is treated by the penetrating β-emitter **phosphorus-23**
- Radioactive **gold** needles are often planted in tumours in the body.
- Sodium chloride solution from radioactive **sodium-24** can be injected into the body and used to investigate abnormal blood circulation.
- Disease of the thyroid gland can be detected using radioactive iodine, **iodine-132** and **iodine-123**

Industrial Uses

- Gamma radiation is used as a radioactive tracer to identify leaks in underground pipes.
- Smoke alarms contain **americium-241**, an α-emitter. Smoke will change the normal steady level of radiation passing across an air gap sufficiently to trigger the alarm (this has a long half-life too)

Energy production

- Radioisotope **uranium-235** is used in nuclear reactors.

Carbon dating

- **Carbon-14** is used to calculate the age of plant and animal remains. It decays by beta emission and has a half-life of 5600 years. Carbon-14 is absorbed by plants during photosynthesis and so the tissues of all living plants and animals contain the radioisotope. When the living organism dies it can no longer absorb carbon-14 and so the level of radioactivity will decrease.
- *e.g.*
- A radioisotope has a half-life of 7,000 years. Calculate the time it will take for 48 g of the radioisotope to decay to leave 6 g.

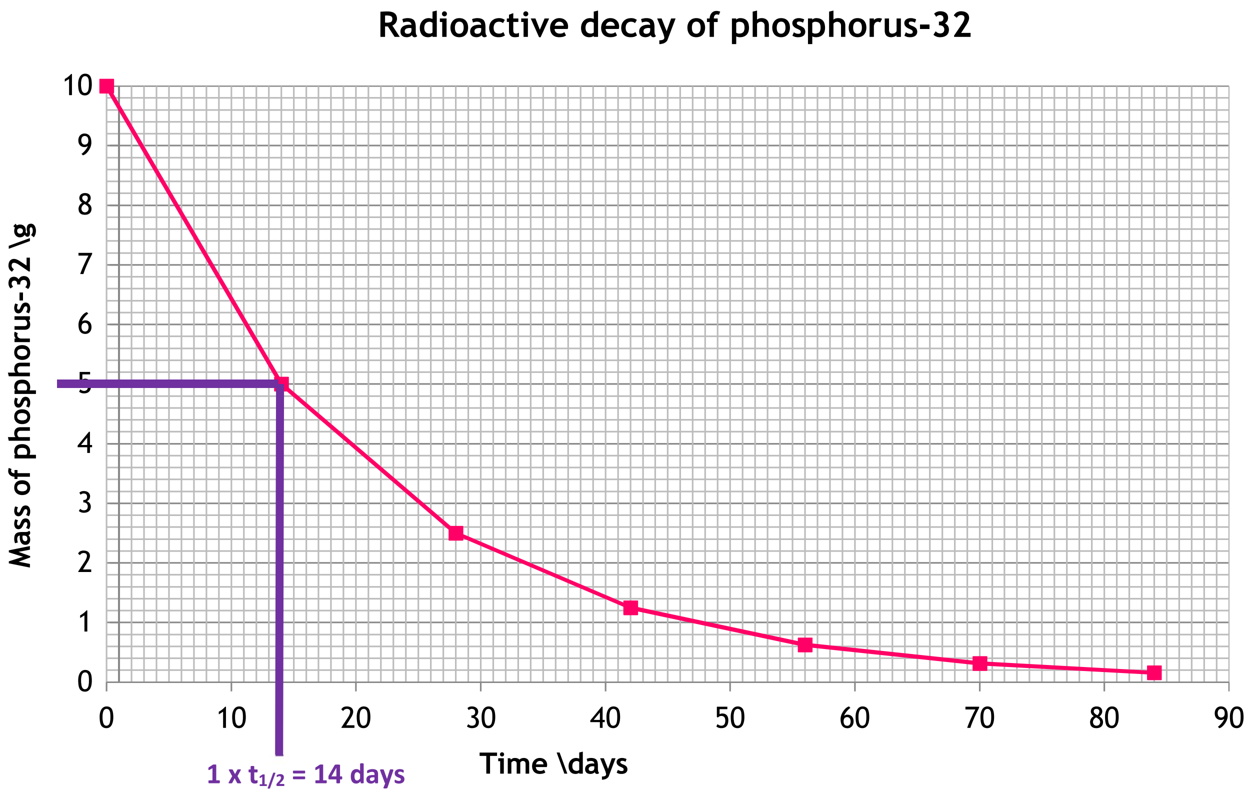
Starting quantity = 48 g
End quantity = 6 g
Half life = 7000 years
Total time = ? years
Number of half-lives = ?

$$48\text{ g} \; \overbrace{\xrightarrow{7000\text{ years}} \; 24\text{ g} \; \xrightarrow{7000\text{ years}} \; 12\text{ g} \; \xrightarrow{7000\text{ years}}}^{\textbf{These add up to 21 000 years}} \; 6\text{ g}$$

The time it will take for 48 g of the radioisotope to decay to leave 6 g is 21 000 years.

Half-life from a graph:

- Look at the starting mass *e.g.* 10 g, half it for half-life *e.g.* 5 g – draw a line to the curve then down to the time



(e) Chemical Analysis

(i) Common Chemical Apparatus

You should be familiar with the following from your in-class experimental work.

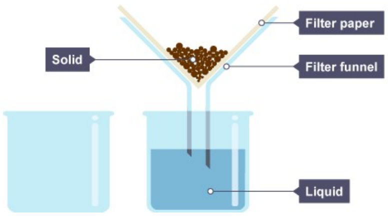
- conical flask
- beaker
- measuring cylinder
- delivery tube
- dropper
- test tubes/boiling tubes
- funnel
- filter paper
- evaporating basin
- pipette with safety filler
- burette
- thermometer

General practical techniques

Filtration and Evaporation

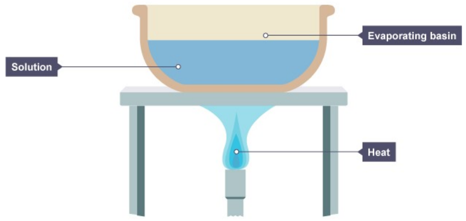
- ***Filtration*** – to separate an **insoluble solid** from a liquid (*e.g.* sand and water OR silver iodide and water)
- ***Evaporation*** – to separate a **soluble solid** from a liquid (*e.g.* sodium chloride (salt) in water)

Filtration



The insoluble solid collects in the filter paperand the liquid in the beaker

Evaporation



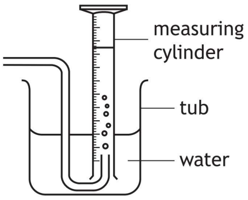
The liquid is evaporated leaving the soluble solid, leave solid to air dry

Using a Balance:

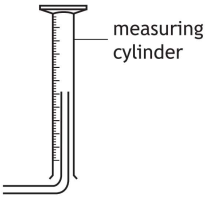
- A balance is used to weigh compounds.
- When weighing a solid the following method is used:
- The weighing boat is placed on the balance
- The balance is set to 0 by pressing the tare button.
- Sample is added to the weighing boat.
- Mass recorded on the digital scale.
- Remove the weighing boat from the balance.
- When weighing a liquid, the following method is used:
- A cork and reaction flask are placed on the balance.
- The balance is set to 0 by pressing the tare button
- The liquid is transferred to the reaction flask.
- Mass recorded on the digital scale.
- Remove the cork ring and reaction flask from the balance.

Collection of Gases

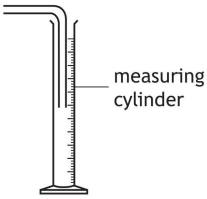
Insoluble Gases – Collection over water



Soluble gases less dense than air – downward displacement of air

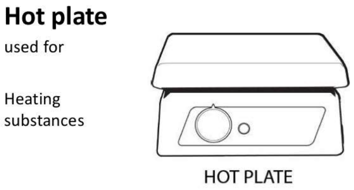


Soluble gases denser than air – upward displacement of air



Methods for heating

- To heat samples a Bunsen burner or a hot plate can be used. A hot plate is used if the sample being heated is flammable as it is safer.



Preparation of soluble salts

- Reactions of reaction of acids with metals, metal oxides, metal hydroxides and metal carbonates often produce soluble salts
- *e.g.*

sodium hydroxide + sulfuric acid → calcium sulfate + water

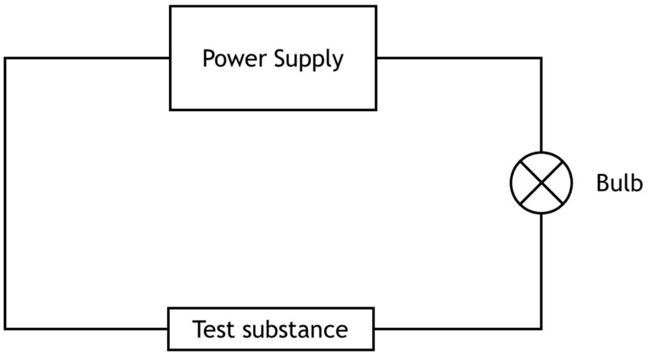
Precipitation Reactions

- The formation of an **insoluble solid** from two **aqueous solutions** (soluble solids in water).
- The insoluble solid is called the precipitate
- *e.g.*

$AgNO_3$ (**aq**) + NaCl (**aq**) → $AgCl$ (**s**) + $NaNO_3$ (aq) [ions have swapped]

- If the state symbols are not given you MUST use table **on page 8** of data booklet to tell you the first two salts ($AgNO_3$ and NaCl) are soluble and AgCl is insoluble *i.e.* the precipitate.

Testing Electrical Conductivity

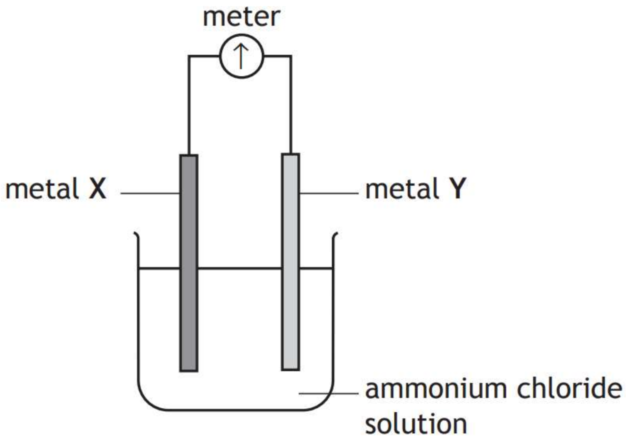


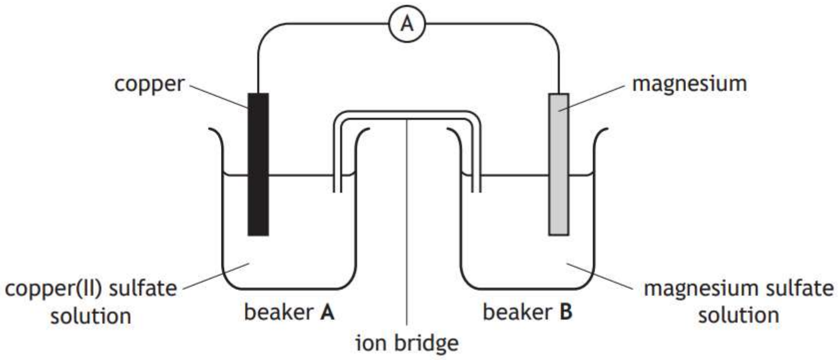
- If the bulb lights then the substance conducts electricity.

Setting up and Electrochemical Cell

Metal Electrodes

- *e.g.*

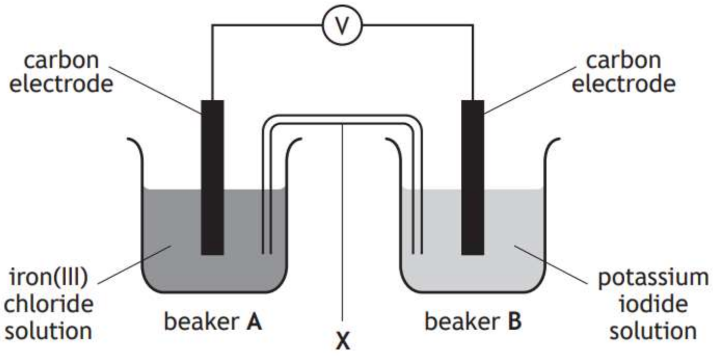




- **Electrons** flow through the **wire** from the metal **higher** (more reactive) to the metal **lower** (less reactive) in the electrochemical series (data book p10).
- **Ions** flow through the **solutions** and the **ion bridge**.
- The **largest voltage** is obtained when the two metals are **furthest apart** in the electrochemical series (data book p10).

Non-metal electrodes

- *e.g.*



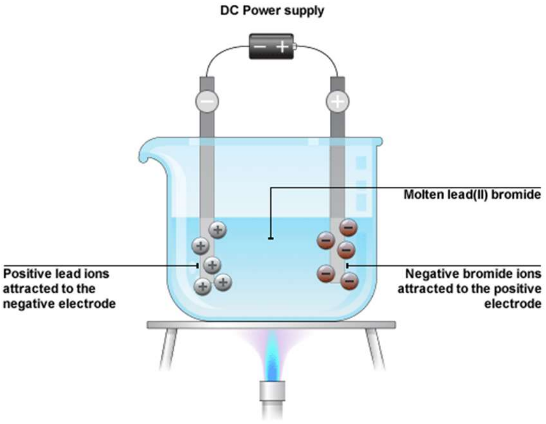
- The reactions taking place are

beaker A $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$

beaker B $2I^-(aq) \rightarrow I_2(l) + 2e^-$

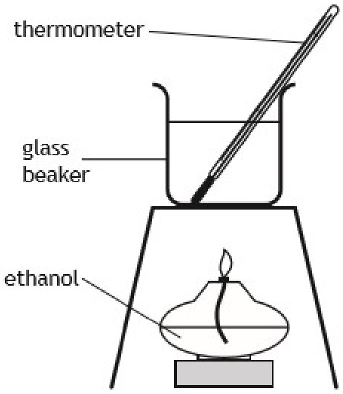
Electrolysis using a d.c. supply

- Electrolysis is the process by which ionic substances are decomposed (broken down) into simpler substances when an **electric current** is passed through them.
- For electrolysis to work, the **ions must be free to move**.
- Ions are free to move when an ionic substance is **dissolved in water** or when **melted.**
- A **direct current (d.c)** is used in electrolysis so that products can be identified.
- The positive ion is attracted to the negative electrode
- The negative ion is attracted to the positive electrode.
- *e.g.*
- electrolysis of lead bromide:



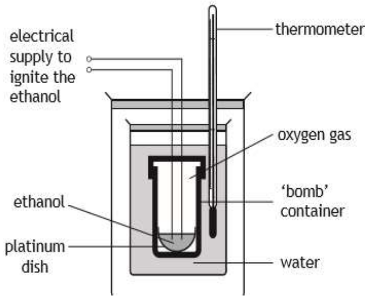
Experimental set up to calculate energy released (E_h)

- Fuel is burned in a spirit burner
- The heat released is used to heat water in a beaker
- Mass of water in the beaker is measured before and the change in temperature is measured in the experiment
- The mass of fuel used to heat the water can be calculated by measuring the mass of the spirit burner at the start and at the end
- Must ensure the distance between the flame and the beaker is kept constant
- Light burned under the beaker to ensure as much energy as possible is transferred to the water



- There are two problems with this experimental set up:
- Heat is lost to the surroundings.
- Incomplete combustion occurs
- (A copper can and draught shield may be used instead of a glass beaker to minimise heat loss)

- The bomb calorimeter – improvement:



- Heat no longer lost to surroundings – water surrounds the container
- Complete combustion: calorimeter is attached to an oxygen cylinder.

$$E_h = cm\Delta T$$

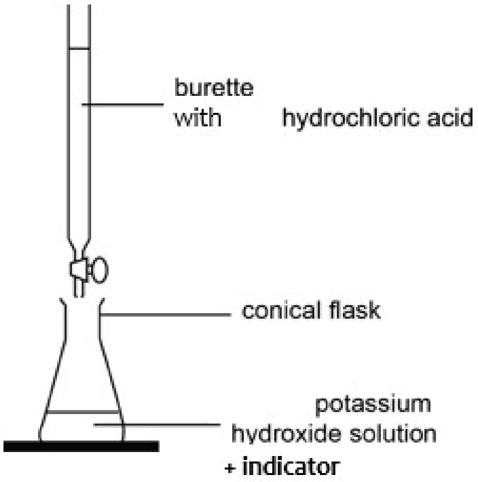
- equation and c given in data booklet page 3
- **E_h** = energy released (kJ)

- **c** = specific heat capacity of water (4.18 kJ kg^{-1} $°C^{-1}$) = energy needed to raise the temperature of 1 kg of a substance by 1 °C
- **m** = mass of **water** (kg) [we assume that 1 litre = 1 kg] (if given volume in cm^3 must divide by 1000)
- **ΔT** = temperature change (°C)

Analytical techniques

Acid/Base Titrations

- A neutralisation takes place
- An **indicator** is added to the conical flask to show the end-point
- The titration is repeated until the results are concordant
- (within 0.2 cm^3) to get more accurate results and an average taken
- A **white tile** is used to identify the colour change easily.
- Solutions of accurately known concentrations are known as **standard solutions**
- Flame Testing
- Flame testing can be used to identify **metal ions** through the colour they produce when burned in a flame
- You MUST use table **on page 6** of data booklet to tell you – write the colour exactly as is stated



Flame Tests

- Different metal ions can be identified from their colours in a flame.
- You can find the colours on p6 of the data book.

Testing Gases:

- The test for hydrogen gas is it burns with a squeaky pop
- The test for oxygen gas is it relights a glowing splint
- The test for carbon dioxide gas is it turns limewater cloudy

Reporting Experimental Work

- Throughout your time in Chemistry you have written experiments with:
    - **Aim** – what you are trying to find out
    - **Method** – what you did (often with a **labelled diagram**)
    - Results – what happened (often a table of data with headings and units)
    - **Conclusion** – what you found out
    - **Evaluation** – how the experiment could be improved
- Experiments that produce **quantitative results** (numbers) are often **repeated** and an **average** (mean) calculated.
- Experiments that produce **quantitative results** (numbers) are often **plotted** as a **bar, line** or **scatter graph** with suitable **scale(s)** and **labels** and a **line of best fit** (straight or curved).