Chemistry Knowledge Organisers

Unit 1: Atomic Structure and the Periodic table

Unit 2: Bonding, Structure and the properties of matter

Unit 3: Quantitative Chemistry

Unit 4: Chemical Changes

Unit 5: Energy Changes

Unit 6: The rate and extent of chemical changes

Unit 7: Organic Chemistry

Unit 8: Chemical Analysis

Unit 9: Chemistry of the atmosphere

Unit 10: Using Resources

Chemistry Exam 1: Units 1-5
Chemistry Exam 2: Unit 6-10

Revision technique: Read, cover, write, check, repeat!

Read your notes.

Cover your notes up and write down as much as you can remember.

Check how you did. Did you miss any information out?

Repeat the whole process.

TIPS: Only try and do a few of the squares at a time.
Don't keep doing the ones you know well.
Keep repeating the ones you struggle to remember.

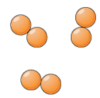
Elements, Mixtures and Compounds

Rule 1 – If two identical elements combine then the name doesn't change

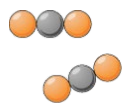
Rule 2 – When two elements join the end is usually ________ide.

Rule 3 – When three or more elements combine and one of them is oxygen the ending is ______ate

An element is just a pure substance, for example oxygen (O_2)

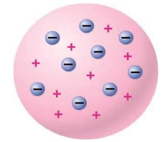


A compound is a material that is made up of more than one type of atom chemically bonded together, for example Carbon Dioxide (CO_2)



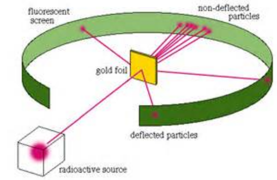
A mixture contains two or more different types of compounds or elements that are not chemically bonded together

Atomic Structure



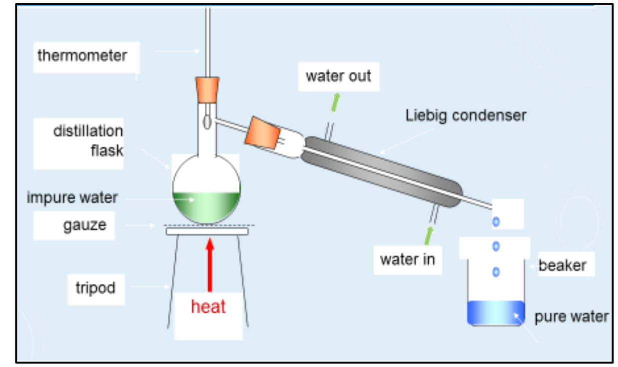
1. In 1901 JJ Thompson suggested the **plum pudding model – this was an atom** that the atom is a ball of positive charge with negative electrons embedded in it.

2. In 1909 Rutherford changed the accepted model using his alpha scattering experiment. The results from the alpha particle scattering experiment led to the conclusion that the mass of an atom was concentrated at the centre (nucleus) and that the nucleus was charged. This nuclear model replaced the plum pudding model.



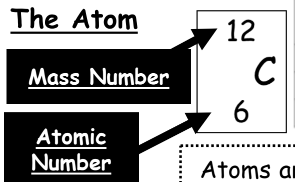
3. Niels Bohr adapted the nuclear model by suggesting that electrons orbit the nucleus at specific distances.
4. 20 years later, James Chadwick provided the evidence to show the existence of neutrons within the nucleus.

Distillation



Distillation can be used to separate liquids from a mixture, if they have different boiling points. Distillation is the process in which evaporation of a liquid is followed by condensation

The Atom



Name of particle	Relative charge	Relative mass
Proton	+1	1
Neutron	0	1
Electron	–1	Very small

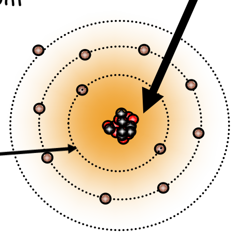
Atoms are very small, having a radius of about 0.1 nm (1×10^{-10} m).

The radius of a nucleus is less than 1/10 000 of that of the atom (about 1×10^{-14} m).

The mass number tells us the number of protons + neutrons.

The number of protons in an atom is known as its atomic number, this is also the number of electrons

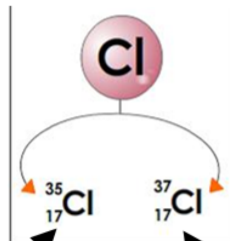
The Nucleus
a dense core of protons and neutrons containing nearly all the mass of the atom



'Shells' of electrons
electrons are really very very tiny so the atom is mostly empty space.

Relative Atomic Mass

RAM is the average mass of all the stable isotopes of that element and includes the relative abundance.



Chlorine - 35 Chlorine - 37

Element	Relative mass of isotope	Relative abundance
Chlorine	35	3
	37	1

$$R.A.M. = \frac{(35 \times 3) + (37 \times 1)}{3 + 1} = 35.5$$

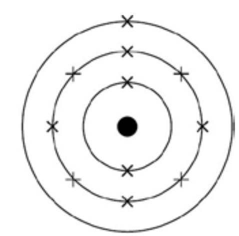
Electronic Structure

The electrons in an atom occupy the lowest available energy levels (innermost available shells).

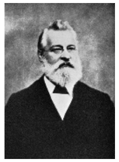
The electronic structure of an atom can be represented by numbers or by a diagram.

Up to two electrons can occupy the lowest energy level, up to eight in the second energy level and up to eight in the third energy level.

For example, the electronic structure of sodium is 2,8,1.



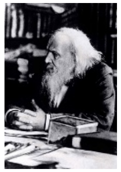
Development of the Periodic Table



Newlands

Before the discovery of protons, neutrons and electrons, scientists attempted to classify the elements by arranging them in order of their atomic weights.

The early periodic tables were incomplete and some elements were placed in inappropriate groups if the strict order of atomic weights was followed.



Mendeleev

Mendeleev overcame some of the problems by leaving gaps for elements that he thought had not been discovered and in some places changed the order based on atomic weights.

Elements with properties predicted by Mendeleev were discovered and filled the gaps. Knowledge of isotopes made it possible to explain why the order based on atomic weights was not always correct.

Transition Metals (Triple Only)

The transition elements are metals with similar properties. Their properties are different from those found in Group 1. Lots of transition metals are used as catalysts.

Properties of transition metals;

- High melting + boiling point
- Form positive ions
- Good electrical conductors
- High thermal conductivity
- Malleable
- Form colored compounds

Copper Good conductor of heat and electricity	**Iron** Alloys are very strong	**Manganese** Resistant to corrosion
Cobalt Strong when alloyed with other metals	**Chromium** Can speed up reactions (Catalyst)	**Nickel** Alloys are resistant to corrosion

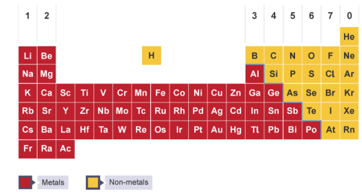
Metals and non-metals

Elements that react to form positive ions are metals.
Elements that do not form positive ions are non-metals.

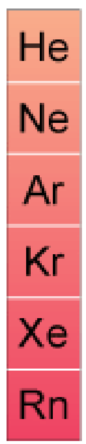
The formation of ions can be worked out using the Periodic Table:

- Group 1 elements form 1+ ions, group 2 elements form 2+ ions and group 3 elements form 3+ ions.
- Group 5 elements form 3- ions, group 6 elements form 2- ions and group 7 elements form 1- ions.
- Group 0 do not form ions due to having a stable structure/full outer shell.

The majority of elements are metals. Metals are found to the left and towards the bottom of the periodic table. Non-metals are found towards the right and top of the periodic table.



Group 0

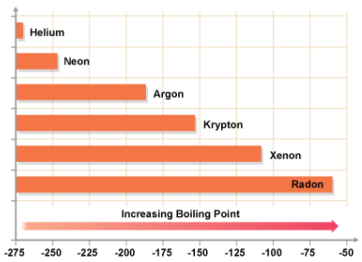


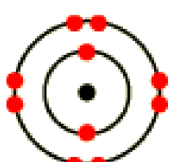
The elements in Group 0 of the periodic table are called the noble gases.

They are unreactive and do not easily form molecules because their atoms have stable arrangements of electrons.

The noble gases have eight electrons in their outer shell, except for helium, which has only two electrons.

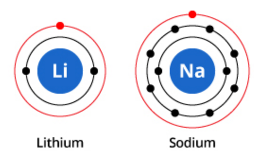
The boiling points of the noble gases increase with increasing relative atomic mass (down the group).

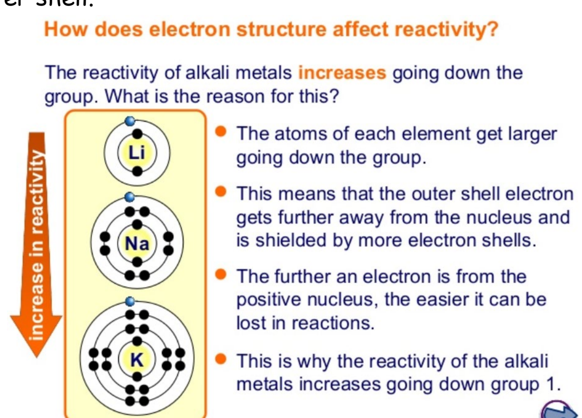




Group 1

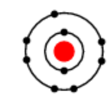
The elements in Group 1 of the periodic table are known as the alkali metals and have characteristic properties because of the single electron in their outer shell.





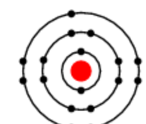
Group 7

The elements in Group 7 of the periodic table are known as the halogens and have similar reactions because they all have seven electrons in their outer shell.



F

The halogens are non-metals and consist of molecules made of pairs of atoms.



Cl

In Group 7, the further down the group an element is the higher its relative molecular mass, melting point and boiling point.

In Group 7, the reactivity of the elements decreases going down the group.

A more reactive halogen can displace a less reactive halogen from an aqueous solution of its salt.

Displaced is just a chemist's word for pushed out.

chlorine + sodium bromide → sodium chloride + bromine

$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$

Chemical bonds

There are three types of strong chemical bonds: ionic, covalent and metallic.

IONIC

For ionic bonding the particles are oppositely charged ions.

Ionic bonding occurs between metals and non-metals.

COVALENT

For covalent bonding the particles are atoms which share pairs of electrons.

Covalent bonding occurs in non-metals.

METALLIC

For metallic bonding the particles are atoms which share delocalised electrons.

Metallic bonding occurs in metallic elements and alloys.

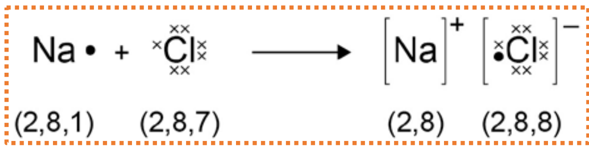
Ionic bonding

When a metal atom reacts with a non-metal atom electrons in the outer shell of the metal atom are transferred.

Metal atoms lose electrons to become positively charged ions.

Non-metal atoms gain electrons to become negatively charged ions.

The ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 have the electronic structure of a noble gas (Group 0).



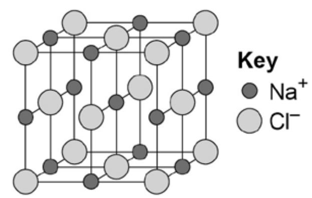
Ionic compounds

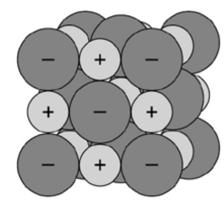
Ionic compounds are held together by strong electrostatic forces of attraction between oppositely charged ions.

These forces act in all directions in the lattice and this is called ionic bonding.

They have high melting and boiling points due to the strong forces of attraction in all directions holding the ions together.

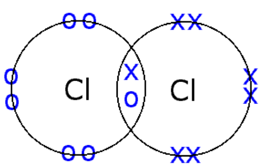
They also conduct electricity when molten or dissolved in water as the ions are free to move so the charge can flow.



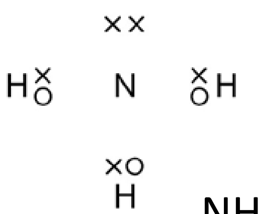


Covalent Bonding

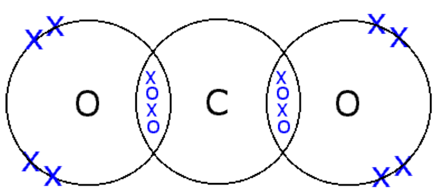
Covalent bonds happen where two or more non-metal elements make a bond. These bonds share electrons, therefore completing the outer shell for all the atoms.



Cl_2

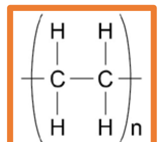


NH_3



CO_2

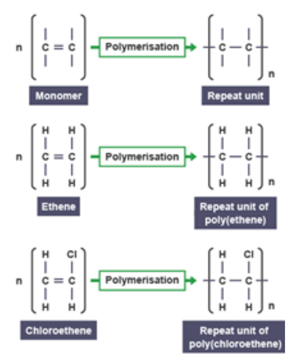
Polymers



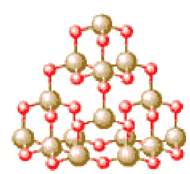
poly(ethene)

Polymers have very large molecules, with the atoms in polymer molecules linked to other atoms by strong covalent bonds. They are ,made by joining thousands of small identical molecules together (monomers). The monomers are often alkene molecules (e.g. ethene). The intermolecular forces between polymer molecules are relatively strong so they are solid at room temperature

By changing the monomer used, we can change the properties of the polymer formed.



Giant Covalent Structures 1

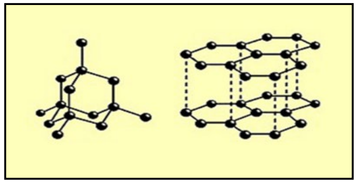


Silicon dioxide (silica)

Each silicon atom is covalently bonded to four oxygen atoms. Each oxygen atom is covalently bonded to two silicon atoms. This means that, overall, the ratio is two oxygen atoms to each silicon atom, giving the formula SiO_2.

Silicon dioxide is very **hard**. It has a very **high melting point** (1,610 °C) and **boiling point** (2,230 °C), is insoluble in water, and does not conduct electricity. These properties result from the very strong covalent bonds that hold the silicon and oxygen atoms in the giant covalent structure.

Giant Covalent Structures 2



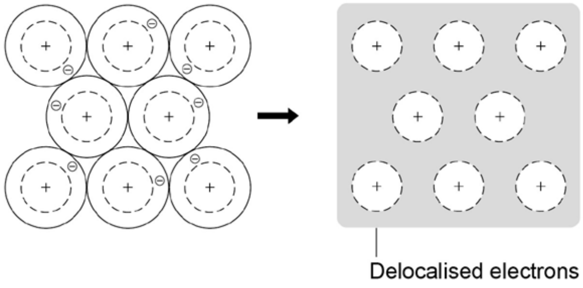
Diamond and graphite are both allotropes of carbon.

In Diamond each carbon atom is bonded to four other carbon atoms by very strong covalent bonds and therefore has no free electrons. The four strong covalent bonds give diamond a very high melting point.

In Graphite each carbon is bonded to 3 carbon atoms with weak intermolecular forces between the layers, which allows the layers to easily slide over each other. They also have a delocalised electron which allows graphite to conduct electricity. Graphite is used in lubricants as the layers can slide.

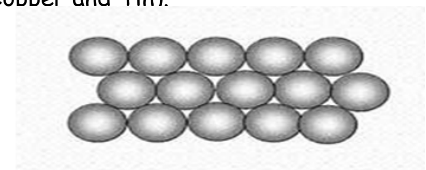
Metallic Bonding

Metals consist of giant structures of atoms arranged in a regular pattern. The electrons in the outer shell of metal atoms are delocalised and so are free to move through the whole structure. The sharing of delocalised electrons gives rise to strong metallic bonds.



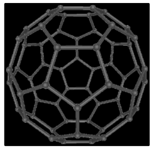
Giant Metallic Structures

Metals also form alloys. In alloys they contain at least two different types of atom which distorts the rigid regimented structure of the metal. As the layers are unable to slide over each other this causes metal alloys to be much stronger than the pure metals. Examples of alloys include Bronze (Copper and tin), Steel (Iron and Carbon) and Bronze (Copper and Tin).



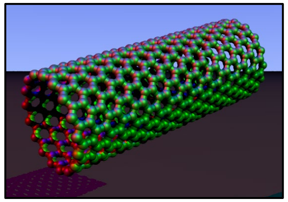
Giant Covalent Structures 3

Fullerenes are molecules of carbon atoms with hollow shapes, based on hexagonal rings of carbon atoms but they may also contain five or seven carbon atoms. Buckminsterfullerene C_{60} was the first to be discovered.



Carbon nanotubes are cylindrical fullerenes with high length to diameter ratios, this makes them useful for nanotechnology, electronics and materials.

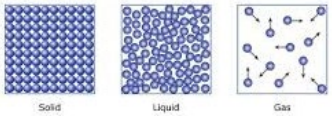
Fullerenes are often good lubricants due to their ability to roll.



Particle Model

The particles in a solid are tightly packed together and can only vibrate. They cannot be pushed any closer together.

The particles in a liquid are in contact with each other, but are arranged randomly. They can roll over each other, that is why a liquid can be poured.



The particles in a gas can move around freely. There are large spaces between the particles, so they can be pushed closer. This is why a gas can be compressed

In melting and boiling the strength of the forces between particles becomes less due to the increased kinetic energy, resulting in more space between the particles and more random arrangement

Nanoparticles (TRIPLE)

Unit name	Unit symbol	Meaning
gigametre	Gm	one billion metres
megametre	Mm	one million metres
kilometre	km	one thousand metres
metre	m	one metre
millimetre	mm	one thousandth of a metre
micrometre	μm	one millionth of a metre
nanometre	nm	one billionth of a metre

Nanoscience refers to structures that are 1^{-100} nm in size, of the order of a few hundred atoms.

Nanoparticles, are smaller than fine particles (PM2.5), which have diameters between 100 and 2500 nm (1×10^{-7} m and 2.5×10^{-6} m).

Coarse particles (PM10) have diameters between 1×10^{-5} m and 2.5×10^{-6} m. Coarse particles are often referred to as dust.

Nanoparticles may have properties different from those for the same materials in bulk because of their high surface area to volume ratio. It may also mean that smaller quantities are needed to be effective than for materials with normal particle sizes.

As the side of cube decreases by a factor of 10 the surface area to volume ratio increases by a factor of 10.

Conservation of mass

Mass is never lost or gained in chemical reactions. We say that mass is always **conserved**. In other words, the total mass of products at the end of the reaction is equal to the total mass of the reactants at the beginning.

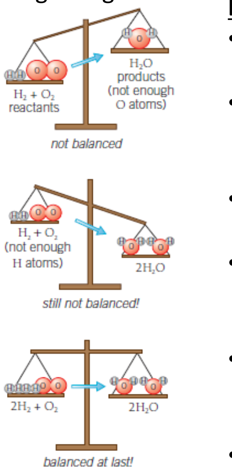


Figure 1 *Balancing an equation*

Balancing equations rules

- Never change the chemical formula
- Total number of reactants must equal total number of products
- Never put a small number yourself
- The big number in front applies to all the atoms in the compound/element
- The small number behind an element applies to that element only
- Use big numbers only and start with 2

Relative formula mass M_r

Mass number = number of protons + number of neutrons
Atomic number = number of protons
Neutron number = mass number – atomic number

The mass of a molecule is called the relative formula mass, M_r. This is calculated by adding up the relative atomic masses of all the atoms in the molecule.

What is the M_r (Relative Formula Mass) of carbon dioxide? **Formula for carbon dioxide: CO_2**

Element	Number of atoms in compound		Mass Number (A_r)		Relative atomic mass of atom(s) in compound
C	1	x	12	=	12
O	2	x	16	=	+ 32
Relative Formula Mass (M_r) of carbon dioxide (CO_2) is....					44

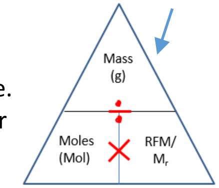
Examples of M_r below:

H_2SO_4 → M_r = (1x2=2) + 32 + (16x4=64) = 98
$Ca(OH)_2$ → M_r = 40 + (16 x 2=32) + (1 x 2=2) = 74
$Mg(HCO_3)_2$ → M_r = 24 + (1x2=2) + (12x2=24) + (16x6=96) = 146
$Al_2(SO_4)_3$ → M_r = (27x2=54) + (32x3=96) + (16x12=192) = 342

Moles and Reacting Masses

One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance. The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant which is 6.02×10^{23} per mole.

Formula to calculate moles



The rules for working out **reacting masses** & example:

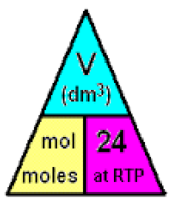
If 28 g of iron reacts with copper sulphate solution, what mass of copper will be made?

- **Step 1.** Write down the balanced symbol equation.
  Fe (28g) + $CuSO_4$ → Cu (?) + $FeSO_4$
- **Step 2.** Write down the relative atomic/formula masses.
  Fe = 56 Cu = 64
- **Step 3.** Write down the ratio of reactants and products.
  Fe : Cu = 1 : 1
- **Step 4.** Convert to ratio of reacting masses.
  Fe : Cu = 1 : 1 = 56 g : 64 g
- **Step 5.** Calculate the scale factor and apply this to the ratio of reacting masses.
  scale factor = 28 g / 56 g = 0.5
  mass of Cu made = 64 g x 0.5 = 32 g

Limiting Reactant (LR)
Is the reactant that gets used up first in a reaction. This is the reactant that is NOT in excess. Therefore, the amounts of product formed in a chemical reaction are determined by the LR

Volume of Gases

One mole of any gas has a volume of 24 dm³ or 24,000 cm³ at rtp (room temperature (20°C) and pressure (1 atmosphere)). This volume is called the molar volume of a gas.



Concentrations

The **concentration** of a solution is usually expressed as the amount of **solute (mol)** dissolved in a given **volume (dm³)** of solution.

Figure 1 *The orange squash is getting less concentrated going left to right (the darker colour indicates more squash is in the same volume of its solution)*

Figure 2 *Volumetric flasks are used to make up solutions. They have a graduation mark around their narrow necks. Water is added to the solute until the bottom of its meniscus (the curve at the surface of the solution when viewed from the side) is level with the mark*

Concentration continued…

The equations to calculate concentration:

$$\text{concentration (g/dm}^3) = \frac{\text{amount of solute (g)}}{\text{volume of solution (dm}^3)}$$

If you are working in centimetres cubed (cm³), convert the volume to dm³ by dividing it by 1000, and use the equation above. Alternatively, substitute your data in cm³ into the following equation:

$$\text{concentration (g/dm}^3) = \frac{\text{amount of solute (g)}}{\text{volume of solution (cm}^3)} \times 1000$$

*** to convert cm³→ dm³, divide by 1000 (0.001 dm³)**
*** to convert dm³→ cm³, multiply by 1000 (1000 cm³)**

You can increase the concentration of an aqueous solution by:

- adding more solute and dissolving it in the same volume of its solution
- evaporating off some of the water from the solution so you have the same mass of solute in a smaller volume of solution.

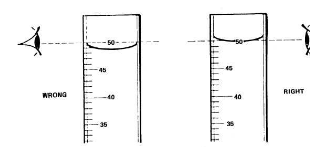
Titrations (TRIPLE ONLY)

Measuring the EXACT volumes of acid and alkali that are needed to react together. **What is this reaction called?** **NEUTRALISATION**

$$H^+ + OH^- \rightarrow H_2O$$

You can measure the exact volumes of acid and alkali needed to react with each other using a technique called **titration**. The point at which the acid and alkali have reacted completely is called the **end point** of the reaction. You judge when the end point has been reacted using an acid/base indicator, such as **Phenolphthalein Indicator**. It turns colourless in an neutral solution and pink in an alkaline solution.

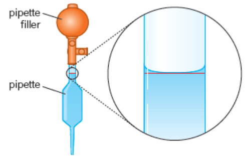
Measuring to the meniscus

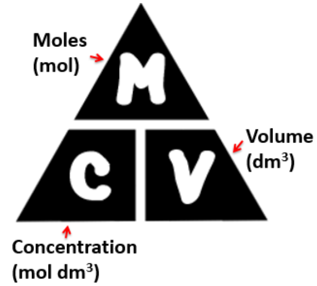


Titrations continued…Carrying out a titration

1. First wash the pipette with distilled water, then with some alkali. Empty alkali into a conical flask.
2. Add a few drops of indicator to the conical flask. Swirl
3. Rinse a **burette** with distilled water and then with some acid. Acid added to burette, starting volume of acid is read accurately.
4. Record the reading on the burette. Open tap to release a bit of acid into flask, swirl.
5. Repeat step 4 until acid in burette has almost run in, then add one drop at a time. Neutralisation occurs. The volume of acid recorded.
6. Repeat 3 times. Discard anomalous results. Repeat the titrations until two results are within of 0.1 cm^3 each other. These precise results are called **concordant**. Calculate a mean.
7. Calculate the concentration of the acid or alkali.

- A volumetric pipette is used to accurately measure a volume of an alkali.
- A pipette filler is used to draw solution into the pipette safely.
- Neutralisation is a change in colour when acid and alkali have been mixed = titration is complete.
- Titre is the volume recorded from a burette





Percentage yield and Atom economy (TRIPLE)

$$\% \text{ yield} = \frac{\text{mass of product obtained}}{\text{maximum theoretical mass of product}} \times 100$$

- The reaction may be reversible – as products form they react to re-form the reactants again. You show reversible reactions using this symbol ⇌ instead of the normal arrow between reactants and products. Chemists can manipulate reversible reactions by the conditions they choose in the reaction vessels in chemical plants.
- Some reactants may react to give unexpected or unwanted products in alternative reactions.
- Some of the product may be lost in handling or left in the apparatus.
- The reactants may not be pure (as in the case of the lime kiln).
- Some of the desired product may be lost during its separation from the reaction mixture.

$$\text{Atom economy} = \frac{\text{mass of wanted product from equation}}{\text{total mass of products from equation}} \times 100$$

Yield Industrial processes –
Industrial processes need as high a percentage yield as possible, because this:
1) Reduces the waste of reactants
2) Reduces the cost of the process

Atom Industrial processes –
Industrial processes need as high an atom economy as possible, because this:
1) Reduces the production of unwanted products
2) Makes the process more *sustainable*
3) Conserve the Earth's resources and minimise pollution

Extraction of Metals + Metal Oxides

Metals react with oxygen to form metal oxides

Chromium + Oxygen → Chromium oxide
Iron + Oxygen → Iron oxide
Copper + Oxygen → Copper oxide

Many metals are found in the ground as metal compounds. The metal needs to be extracted. For metals that are below carbon in the reactivity series this can be done by heating the metal compound with carbon The carbon removes the oxygen from the metal oxide.

K	Potassium	Most reactive
Na	Sodium	
Ca	Calcium	
Mg	Magnesium	
Al	Aluminium	
C	Carbon	
Zn	Zinc	
Fe	Iron	
Sn	Tin	
Pb	Lead	
H	Hydrogen	
Cu	Copper	Least reactive
Ag	Silver	
Au	Gold	
Pt	Platinum	

C H *added for comparison*

Reactivity Series of Metals

1. Copper oxide + Carbon → Carbon dioxide + Copper
2. Lead oxide + Carbon → Carbon dioxide + Lead
3. Iron oxide + Carbon → Carbon dioxide + Iron

Oxidation and Reduction

Oxidation is the gain of oxygen and the loss of electrons, reduction is the loss of oxygen and gain of electrons.

A chemical reaction where both oxidation and reduction occur is called a redox reaction.

The equation below shows a word equation, a balanced symbol equation, ionic and half equations which show the movement of electrons.

Zinc + copper sulphate → zinc sulphate + copper
$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
Zn - oxidised Cu^{2+} - reduced
$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
$Zn - 2e^- \rightarrow Zn^{2+}$ or $Zn \rightarrow Zn^{2+} + 2e^-$
$Cu^{2+} + 2e^- \rightarrow Cu$

Metals + Acids and Metal Carbonates + Acid

Metal + Acid → Metal salt + Hydrogen

Calcium + Hydrochloric acid → $Ca + 2HCl$	Calcium chloride + Hydrogen $CaCl_2 + H_2$
Zinc + Hydrochloric acid → $Zn + 2HCl$	Zinc chloride + Hydrogen $ZnCl_2 + H_2$

Metal Carbonate + Acid → Metal salt + Carbon Dioxide + Water

Calcium Carbonate + Hydrochloric acid → $CaCO_3 + 2HCl$	Calcium chloride + Carbon Dioxide + Water $CaCl_2 + CO_2 + H_2O$
Potassium Carbonate + Nitric acid → $K_2CO_3 + 2HNO_3$	Potassium nitrate + Carbon Dioxide + Water $2KNO_3 + CO_2 + H_2O$

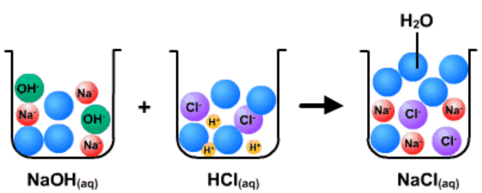
Acid + Alkali → Metal salt + Water

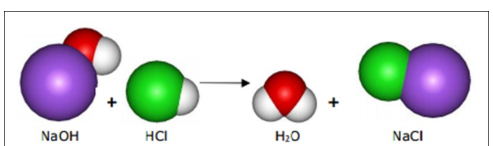
Hydrochloric acid + Sodium hydroxide → $2HCl + NaOH$	Sodium chloride + Water $NaCl_2 + H_2O$
Sulphuric acid + Potassium hydroxide → $H_2SO_4 + 2KOH$	Potassium Sulphate + Water $K_2SO_4 + 2H_2O$

Neutralisation

The acid used will determine the salt produced in a neutralisation reaction:

- hydrochloric acid produces chlorides
- nitric acid produces nitrates
- sulfuric acid produces sulfates





Soluble salts (Required practical)

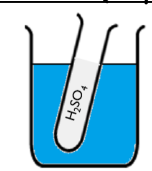
Soluble salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates.

The solid is added to the acid until no more reacts and the excess solid is filtered off to produce a solution of the salt.

Salt solutions can be crystallised to produce solid salts.

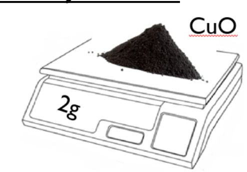
Soluble salts (Required practical): Method

1



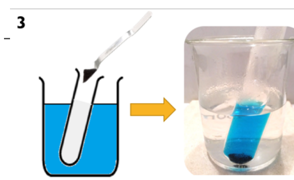
Sulfuric acid is warmed in a water bath

2



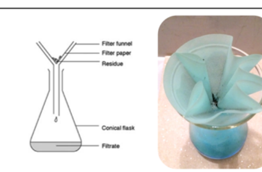
Weigh 2g of black copper oxide powder

3



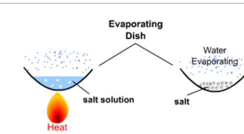
Add copper oxide to the sulphuric acid until a blue solution is formed and excess copper oxide sinks to the bottom of the tube.

4



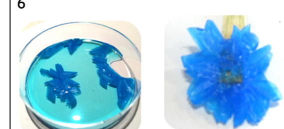
Filter the unreacted copper oxide from the solution and collect the filtrate

5



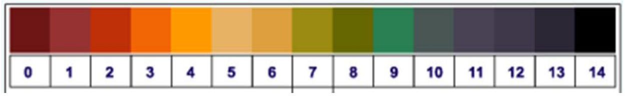
Transfer the solution to an evaporating dish and heat gently

6



Leave to cool, copper sulfate crystals will form. Remove and dry crystals.

pH and Acids + Alkalis



Acids produce H^+ (as H_3O^+) ions in water and they taste sour. They also corrode metals and have a pH of less than 7. They also turns blue litmus paper to red.

Alkalis produce OH^- ions in water and they taste bitter with a pH greater than 7. Alkalis turns red litmus paper to blue.

A solution is defined as an acid if the concentration of H^+ ions is greater than the concentration of OH^- ions. $[H^+] > [OH^-]$

A solution is defined as alkali/base if the concentration of hydrogen ions is less than the concentration of hydroxide ions. $[H^+] < [OH^-]$

Strong and weak acids

A strong acid is completely ionised in aqueous solution.

$$HCl + H_2O \longrightarrow H^+ + Cl^-$$

Examples of strong acids are hydrochloric, nitric and sulfuric acids.

A weak acid is only partially ionised in aqueous solution.

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H^+$$

Examples of weak acids are ethanoic, citric and carbonic acids.

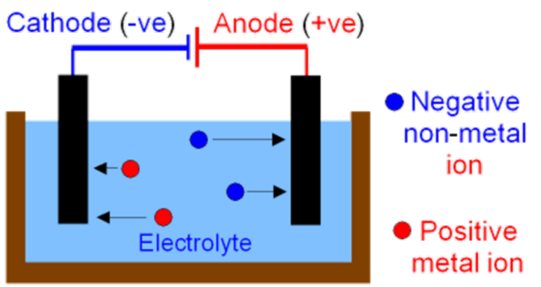
For a given concentration of aqueous solutions, the stronger an acid, the lower the pH.

As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.

	$[H^+]$	pH	Example
	1×10^{0}	0	HCl
	1×10^{-1}	1	Stomach acid
	1×10^{-2}	2	Lemon juice
Acids	1×10^{-3}	3	Vinegar
	1×10^{-4}	4	Soda
	1×10^{-5}	5	Rainwater
	1×10^{-6}	6	Milk
Neutral	1×10^{-7}	7	Pure water
	1×10^{-8}	8	Egg whites
	1×10^{-9}	9	Baking soda
	1×10^{-10}	10	Tums® antacid
Bases	1×10^{-11}	11	Ammonia
	1×10^{-12}	12	Mineral lime - $Ca(OH)_2$
	1×10^{-13}	13	Drano®
	1×10^{-14}	14	NaOH

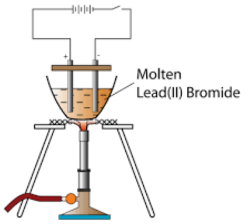
Electrolysis

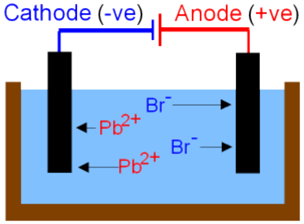
When an ionic compound is melted or dissolved in water, the **ions** are free to move about within the liquid or solution. These liquids and solutions are able to conduct electricity and are called electrolytes. Passing an electric current through electrolytes causes the ions to move to the electrodes. Positively charged ions move to the negative electrode (the cathode), and negatively charged ions move to the positive electrode (the anode).



Electrolysis of molten ionic compounds

When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes, the metal (lead) is produced at the cathode and the non-metal (bromine) is produced at the anode.





Cathode (-ve electrode)

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

Anode (+ve electrode)

$$2Br^- - 2e^- \rightarrow Br_2$$

OR $$2Br^- \rightarrow Br_2 + 2e^-$$

Electrolysis Extended

At the negative electrode, hydrogen is produced if the metal is more reactive than hydrogen.
At the positive electrode oxygen is produced unless the solution contains halide ions when the halogen is produced.

This is due to water molecules breaking down in aqueous solution to form hydrogen and hydroxide ions.

At the cathode positively charged ions gain electrons, whereas as the negatively charged ions lose electrons at the anode. These are both examples of oxidation and reduction. These can be represented as half equations.

$2H^+ + 2e^- \rightarrow H_2$

$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

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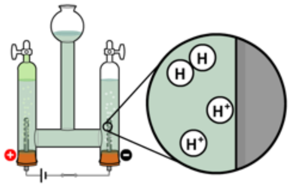


At the cathode

Whether hydrogen or a metal is produced at the cathode depends on the position of the metal in the metal **reactivity series**:

- the metal is produced at the cathode if it is less **reactive** than hydrogen
- hydrogen is produced at the cathode if the metal is more reactive than hydrogen

Rules for determining products

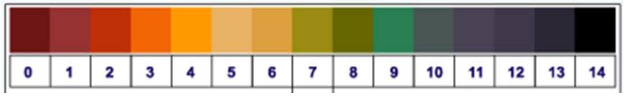
At the anode

Oxygen is produced (from hydroxide ions), unless **halide** ions (chloride, bromide or iodide ions) are present. In that case, the negatively charged halide ions lose electrons and form the corresponding **halogen** (chlorine, bromine or iodine).

The table summarises the product formed at the anode during the electrolysis of different **electrolytes** in solution.

Negative ion	Element given off at anode
Chloride, Cl^-	Chlorine, Cl_2
Bromide, Br^-	Bromine, Br_2
Iodide, I^-	Iodine, I_2
Sulfate, SO_4^{2-}	Oxygen, O_2
Nitrate, NO_3^-	Oxygen, O_2

pH and Acids + Alkalis



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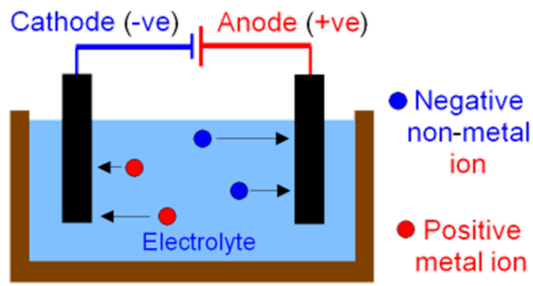

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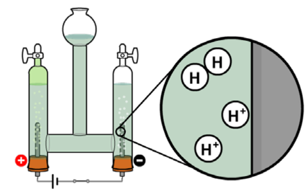
$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

$4OH^- - 4e^- \rightarrow O_2 + 2H_2O$

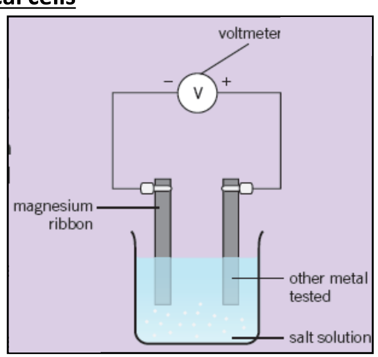


Cells and batteries continued…

- Metals lose electrons and form positive ions.
- When 2 metals are dipped in a salt solution and joined by a wire, the more reactive metal will donate electrons to the less reactive metal. This forms a simple electrical cell.
- The greater the difference in reactivity between the 2 metals, the higher the voltage produced by the cell.

Investigating chemical cells

This apparatus is used to investigate the voltage produced by different metals paired with magnesium ribbon. You can compare magnesium against zinc, iron, copper & tin in your electrical cells.



Fuel Cells

Scientists are developing hydrogen as a fuel.

Hydrogen + Oxygen → Water

$2H_2 + O_2 \rightarrow 2H_2O$

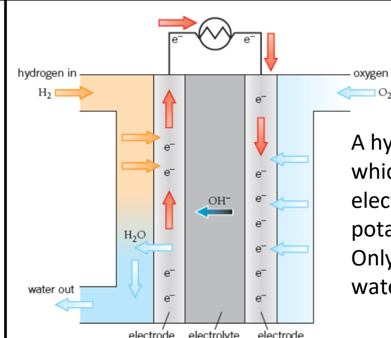
- The world relies on fossil fuels. However, they are non-renewable and they cause pollution.
- Hydrogen is one alternative fuel. It can be burned in combustion engines or used in fuel cells to power vehicles.
- Hydrogen gas is oxidised and provides a source of electrons in the hydrogen fuel cell, in which the only waste product is water.

Hydrogen gas is supplied as a fuel to the negative electrode. It diffuses through the graphite electrode and reacts with hydroxide ions to form water and provides a source of electrons to an external circuit.

$$2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$$

Oxygen is supplied to the positive electrode. It diffuses through the graphite and reacts to form hydroxide ions, accepting electrons from the external circuit.

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$



A hydrogen fuel cell which has an alkaline electrolyte, such as potassium hydroxide. Only waste product is water.

Advantages of hydrogen fuel cells –
1) Do not need to be electrically recharged
2) No pollutants are produced
3) Can be a range of sizes for different uses

Disadvantages of hydrogen fuel cells–
1) Hydrogen is highly flammable
2) Hydrogen is sometimes produced for the cell by non-renewable sources
3) Hydrogen is difficult to store

Exothermic and endothermic reactions

Exothermic reactions **release** thermal energy (heat) into their surroundings. They can occur spontaneously and some are explosive. Most chemical reactions are exothermic. Temperature **increases**.

What are some examples?

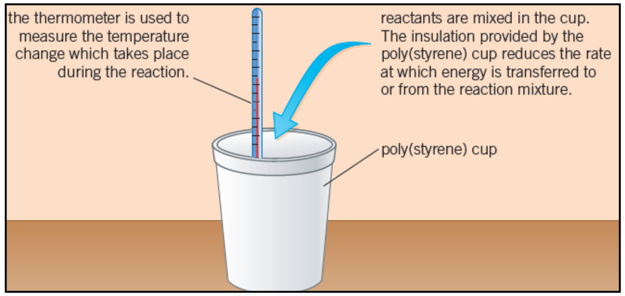
- combustion
- respiration
- neutralization of acids with alkalis
- reactions of metals with acids
- Mg (s) + HCl (aq) → $MgCl_2$ (aq) + H_2 (g)
- the Thermite Process.
- Endothermic reactions absorb thermal energy, and so cause a **decrease** in temperature.

What are some examples?

- thermal decomposition, e.g. calcium carbonate in a blast furnace
- photosynthesis
- some types of electrolysis
- Sherbet
- NH_4NO_3 (s) + H_2O (l)→NH_4^+ (aq) + NO_3^- (aq)

Investigating temperature changes

Record the initial temperatures of any solutions, and the maximum and minimum temperatures reached in the course of the reaction.



Using energy transfers from reactions

- Exothermic changes can be used in hand warmers and self heating cans. Crystallisation of the supersaturated solution is used in reusable warmers. However, disposable, one-off hand warmers heat up the surrounding for longer.
- Endothermic changes can be used in instant cold packs for sports injuries.

Reaction profiles and Activation energy

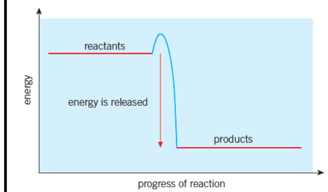


Figure 1 *The reaction profile for an exothermic reaction*

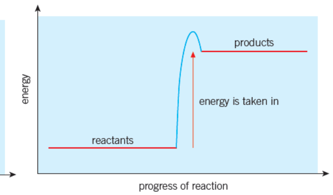


Figure 2 *The reaction profile for an endothermic reaction*

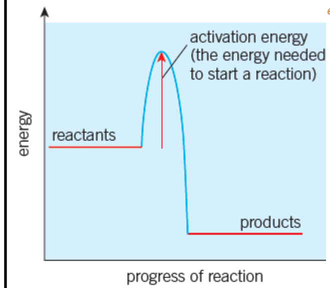


Figure 3 *This reaction profile shows the activation energy for an exothermic reaction*

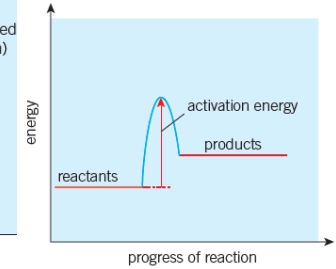


Figure 4 *This reaction profile shows the activation energy for an endothermic reaction*

Bond breaking is endothermic whereas bond making is exothermic.

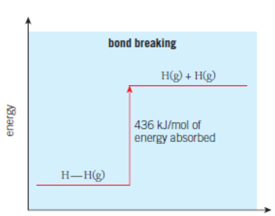
Bond energy calculations

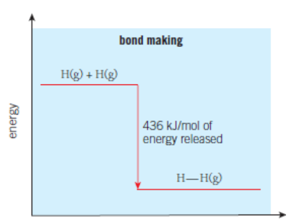
The energy needed to break a bond between 2 atoms is called the **bond energy** for that bond. They are measured in KJ/mol.

Table 1 *Common bond energies*

Bond	Bond energy in kJ/mol
C—C	347
C—O	358
C—H	413
C—N	286
C—Cl	346
Cl—Cl	243

Bond	Bond energy in kJ/mol
H—Cl	432
H—O	464
H—N	391
H—H	436
O═O	498
N≡N	945



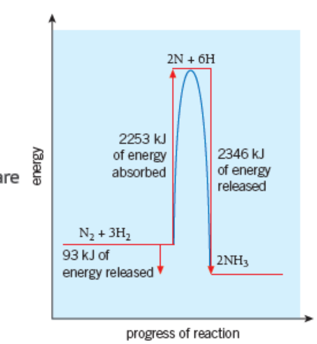


Breaking and making a particular bond always involves the same amount of energy

The formation of ammonia. The energy released, 93KJ, is from the formation of 2 moles of ammonia (see balanced equation below). So if you wanted to know the energy change for the reaction per mole of ammonia formed, it would release exactly half this, i.e. 46.5kJ/mol.

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

- In chemical reactions, energy must be supplied to break the bonds between atoms in the reactants.
- When new bonds are formed between atoms in a chemical reaction, energy is released.
- In an exothermic reaction, the energy released when new bonds are formed is greater than the energy absorbed when bonds are broken.
- In an endothermic reaction, the energy released when new bonds are formed is less than the energy absorbed when bonds are broken.
- You can calculate the overall energy change in a chemical reaction using bond energies.



Cells and batteries

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

The sulfate ions do not change in the displacement reaction above. They are spectator ions.

So you can leave them out of the equation and write an ionic equation:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

You can think of this redox reaction as two half equations.

One will represent reduction:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

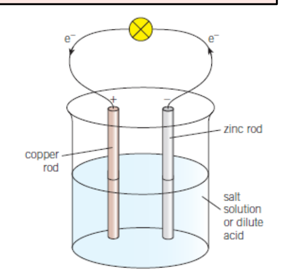
The Cu^{2+} ions are reduced to Cu.

The other will be an oxidation reaction:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

The Zn atoms are oxidised to Zn^{2+} ions.

An electrical cell made from zinc and copper → The electrons flow from the more reactive metal (zinc) to the less reactive metal (copper). So zinc acts as the negative terminal of the cell, providing electrons to the external circuit.

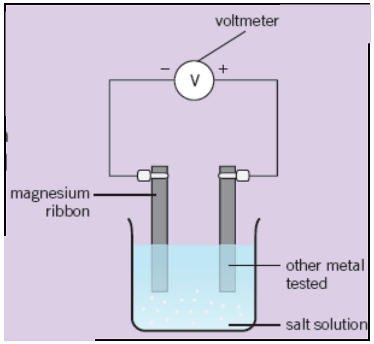


Cells and batteries continued…

- Metals lose electrons and form positive ions.
- When 2 metals are dipped in a salt solution and joined by a wire, the more reactive metal will donate electrons to the less reactive metal. This forms a simple electrical cell.
- The greater the difference in reactivity between the 2 metals, the higher the voltage produced by the cell.

Investigating chemical cells

This apparatus is used to investigate the voltage produced by different metals paired with magnesium ribbon. You can compare magnesium against zinc, iron, copper & tin in your electrical cells.



Fuel Cells

Scientists are developing hydrogen as a fuel.

Hydrogen + Oxygen → Water

$$2H_2 + O_2 \rightarrow 2H_2O$$

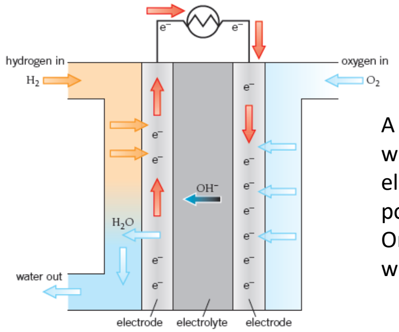
- The world relies on fossil fuels. However, they are non-renewable and they cause pollution.
- Hydrogen is one alternative fuel. It can be burned in combustion engines or used in fuel cells to power vehicles.
- Hydrogen gas is oxidised and provides a source of electrons in the hydrogen fuel cell, in which the only waste product is water.

Hydrogen gas is supplied as a fuel to the negative electrode. It diffuses through the graphite electrode and reacts with hydroxide ions to form water and provides a source of electrons to an external circuit.

$$2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$$

Oxygen is supplied to the positive electrode. It diffuses through the graphite and reacts to form hydroxide ions, accepting electrons from the external circuit.

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$



A hydrogen fuel cell which has an alkaline electrolyte, such as potassium hydroxide. Only waste product is water.

Advantages of hydrogen fuel cells –

1) Do not need to be electrically recharged
2) No pollutants are produced
3) Can be a range of sizes for different uses

Disadvantages of hydrogen fuel cells–

1) Hydrogen is highly flammable
2) Hydrogen is sometimes produced for the cell by non-renewable sources
3) Hydrogen is difficult to store

The rate of a reaction can be measured by the rate at which a reactant is used up, or the rate at which a product is formed.

- We can **measure the rate** of a reaction by looking at:
- how fast solid **reactants are used** up,
- how quickly **gas is produced** or
- how **quickly light is blocked** (the disappearing cross)

The quantity of reactant or product can be measured by:

- mass in grams or volume in cm^3. The units are: **g/s or cm^3/s.**
- HT: quantity of reactants in terms of moles and units for rate of reaction in **mol/s.**

$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time}} \text{ or } \frac{\text{quantity of product formed}}{\text{time}}$$

There are **3 different methods** that can be used to measure the rate of a reaction. Measuring the;

1. Decreasing mass of a reaction mixture (e.g. marble chips (calcium carbonate) & HCl
2. Increasing volume of a gas given off
3. Decreasing light passing through a solution (i.e. disappearing X)

Reactions, particles and collisions

Reactions take place when particles collide with a certain amount of energy.

The minimum amount of energy needed for the particles to react is called the **activation energy**, and is different for each reaction.

The rate of a reaction depends on two things:

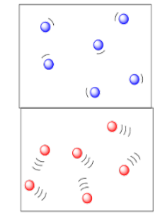
- the **frequency** of collisions between particles
- the **energy** with which particles collide.

If particles collide with less energy than the activation energy, they will not react. The particles will just bounce off each other.

What factors affect the rate of reactions?

- increased **temperature**
- increased **concentration** of dissolved reactants
- increased **pressure** of gaseous reactants
- increased **surface area** of solid reactants
- use of a **catalyst**

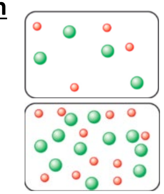
Effect of temperature on rate of reaction



At a higher temperature, particles have more energy. This means they move faster and are more likely to collide with other particles.

When the particles collide, they do so with more energy, and so the number of successful collisions increases.

Effect of concentration on rate of reaction



At a higher concentration, there are more particles in the same amount of space. This means that the particles are more likely to collide and therefore more likely to react.

Effect of pressure on rate of reaction

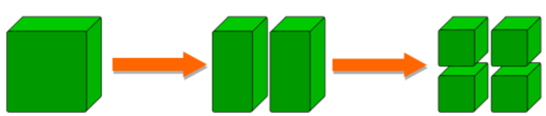
As the pressure increases, the space in which the gas particles are moving becomes smaller.

The gas particles become closer together, increasing the frequency of collisions. This means that the particles are more likely to react.

Effect of surface area on rate of reaction

This means that there is an increased area for the reactant particles to collide with.

The smaller the pieces, the larger the surface area. This means more collisions and a greater chance of reaction.



low surface area — high surface area

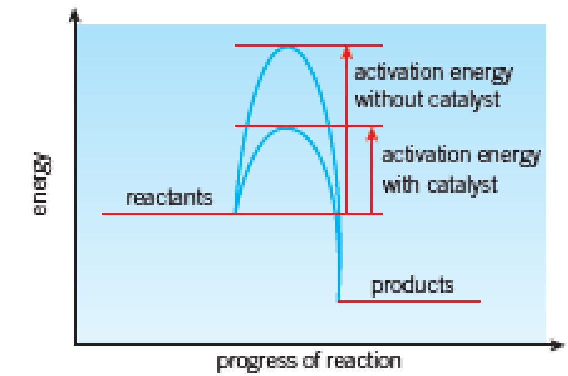
Effect of catalysts on rate of reaction

Catalysts are substances that change the rate of a reaction without being used up in the reaction.

Catalysts never produce more product – they just produce the same amount more quickly.

Different catalysts work in different ways, but most lower the reaction's activation energy (E_a).

- **Nickel** is a catalyst in the production of margarine (hydrogenation of vegetable oils).
- **Iron** is a catalyst in the production of ammonia from nitrogen and hydrogen (the Haber process).
- **Platinum** is a catalyst in the catalytic converters of car exhausts. It catalyses the conversion of carbon monoxide and nitrogen oxide into the less polluting carbon dioxide and nitrogen.

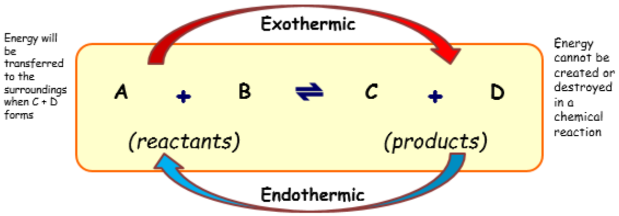


Why are catalysts so important for industry?

- Products can be made more quickly, saving time and money.
- Catalysts reduce the need for high temperatures, saving fuel and reducing pollution.
- Catalysts often come in the form of **powders, pellets** or **fine gauzes,** this provides the largest possible surface area for them to work.

Reversible reactions

Reversible reactions occur when the backwards reaction (reactants → products) takes place relatively easily under certain conditions. The products turn back into the reactants.



Litmus is a complex molecule. This can be represented as HLit (where H is hydrogen). HLit is red. If you add alkali, HLit turns into the Lit− ion by losing an H+ ion. Lit− is blue. If you then add more acid, blue Lit− changes back to red HLit, and so on.

$$HLit(aq) \rightleftharpoons H^+(aq) + Lit^+(aq)$$

red litmus — blue litmus

Reversible reactions can be endothermic and exothermic. The energy transferred **from** the surroundings by the endothermic reaction is **equal to** the energy transferred **to** the surroundings during the exothermic reaction. E.g. thermal decomposition of hydrated copper sulfate.

HT: Le Chatelier's Principle

If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change. The effects of changing conditions on a system at equilibrium can be predicted using Le Chatelier's Principle.

1. Temperature…

If the temperature of a system at equilibrium is increased:

- the relative amount of products at equilibrium increases for an endothermic reaction
- the relative amount of products at equilibrium decreases for an exothermic reaction.

Equilibrium

When reversible reactions **reach equilibrium** the forward and reverse reactions are still happening but at the same rate, so the concentrations of reactants and products do not **change**. The balance point can be affected by temperature, and also by pressure for gasses in **equilibrium**

When you heat ammonium chloride, a reversible reaction takes place. Ammonium chloride breaks down on heating. It forms ammonium chloride and hydrogen gases (colourless gases). This is an example of a DECOMPPOSITION REACTION.

ammonium chloride	⇌	ammonia	+	hydrogen chloride
NH_4Cl (s)	⇌	NH_3 (g)	+	HCl (g)

NH_4Cl decomposes back into NH_3 and HCl gases, when heated. White solid NH_4Cl reforms in the cooler part of the test tube.

1. Temperature continued…

If the temperature of a system at equilibrium is decreased: • the relative amount of products at equilibrium decreases for an endothermic reaction •the relative amount of products at equilibrium increases for an exothermic reaction.

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

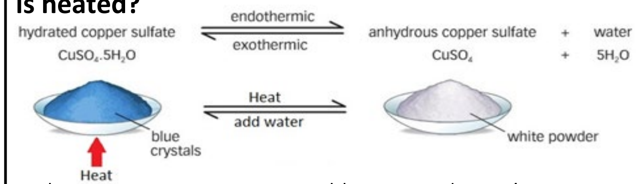
Here the forward reaction is exothermic – a decrease in temperature moves equilibrium to the right (more NH_3).

2. Pressure…

For gaseous reactions at equilibrium:

- an increase in pressure causes the equilibrium position to shift towards the side with the smaller number of molecules as shown by the symbol equation for that reaction
- a decrease in pressure causes the equilibrium position to shift towards the side with the larger number of molecules as shown by the symbol equation for that reaction.

What happens when hydrated copper (II) sulfate is heated?



When a reaction is at equilibrium it doesn't mean the amounts of reactants and products are equal.

- If the equilibrium lies to the right, the concentration of products is greater than that of the reactants.
- If the equilibrium lies to the left, the concentration of reactants is greater than that of the products.

The position of equilibrium depends on the following conditions:

1. Temperature
2. Pressure (this only affects equilibria of gases)
3. Concentration of the reactants and products

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

There are 4 moles to the left (1 of N_2 and 3 of H_2) but only 2 on the right. So if you increase the pressure, the equilibrium shifts to the right (more NH_3).

3. Concentration…

If the concentration of one of the reactants or products is changed, the system is no longer at equilibrium and the concentrations of all the substances will change until equilibrium is reached again.

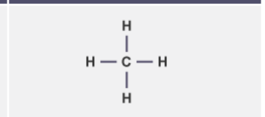
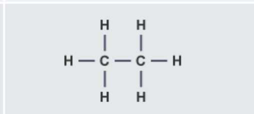
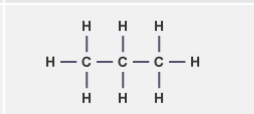
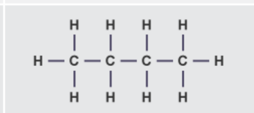
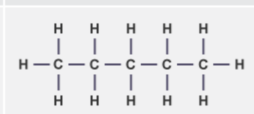
- If the concentration of a reactant is increased, more products will be formed until equilibrium is reached again.
- If the concentration of a product is decreased, more reactants will react until equilibrium is reached again.

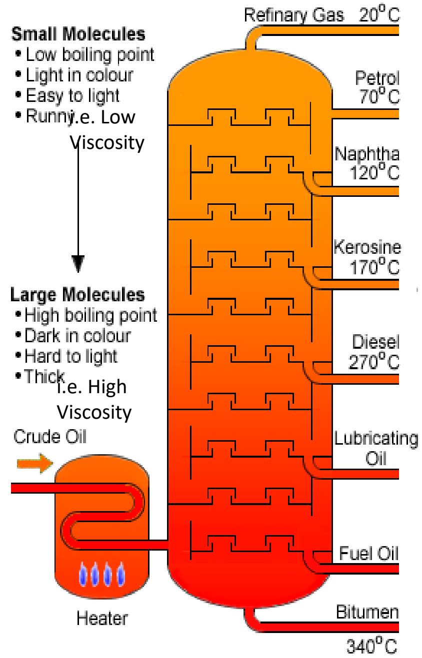
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

If more N_2 or H_2 is added, the forward reaction increases to produce more NH_3.

<u>Hydrocarbons:</u> Are fuels that are made of just <u>hydrogen and carbon</u> atoms only, joined together by <u>single</u> chemical bonds called <u>covalent bonds</u>.

<u>Alkanes:</u> <u>Are saturated hydrocarbons</u> – This means that their carbon atoms are joined to each other by **single C-C bonds** and that they can have as many hydrogen atoms as possible. **Alkanes have the formula: $C_nH_{(2n+2)}$**

Alkane	Molecular formula	Structural formula
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	
Butane	C_4H_{10}	
Pentane	C_5H_{12}	

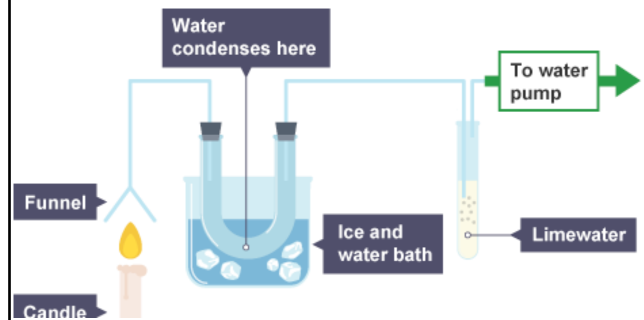


Fractional Distillation

What factors affect the properties of hydrocarbons?

Small molecules means that they have short carbon chains.
Large molecules means that they have long carbon chains.

<u>Combustion (burning)</u>
When a hydrocarbon is burned with sufficient oxygen supply, the products are always <u>carbon dioxide</u> and <u>water vapour</u>.
Hydrocarbon + oxygen → carbon dioxide + water (+ energy)
E.g. Butane + Oxygen → Carbon Dioxide + water
$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$
During combustion both carbon and hydrogen from the hydrocarbon are <u>oxidised</u>. Hydrocarbons are used as <u>fuels</u> due to the amount of energy released when they combust completely.



<u>Crude Oil</u>
Formed from the <u>buried remains of plants and animals (mainly plankton)</u>. Over millions of years with high temperature and pressure, the remains turn into <u>crude oil</u>. Fossil fuels such as coal, oil and gas are non-renewable. Crude oil Is a <u>mixture</u> of lots of different hydrocarbons (mostly alkanes). Crude oil can be split up into separate fractions by <u>fractional distillation.</u>
<u>Fractional Distillation</u> (see diagram in the middle box above)
1) Crude oil enters the **bottom** of a fractional distillation column and is **heated** to about **350°C** until most of it has turned to gas
2) The temperature is controlled
3) Most of the substances in the crude oil **<u>evaporate</u>**. The mixture of vapours then passes up the tower and **<u>condense</u>**
4) Hydrocarbons with **high** boiling points (long chains) **condense** first, **low down** in the tower
5) Some hydrocarbons have **very low** boiling points and so they are gases. They don't condense but are collected as 'fuel gases'.

<u>Crude Oil uses are important in the modern world</u>
Oil provides the fuel for most modern transport – cars, trains, planes etc. E.g. diesel, kerosene, heavy fuel oil etc. come from crude oil..
The **petrochemicals industry** uses some of the hydrocarbons from crude oil as **feedstock** (raw material to supply or fuel a machine or industrial process) to make new compounds for use in things such as *solvents, lubricants, polymers, detergents etc.*

<u>Cracking (is a thermal decomposition reaction)</u>
This mean splitting up long-chain hydrocarbons using heat. Short-chain hydrocarbons are flammable so make good fuels and are high in demand. Long-chain hydrocarbons are thick, gooey liquids = not that useful. **Cracking** is the breakdown of large , long-chain hydrocarbon **alkanes** into smaller, more useful **alkanes** and **alkenes**. This process requires high temperatures and high pressure.
Alkenes are used as a **starting material** when making lots f other compounds and be used to make **polymers**.

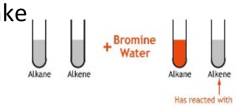
<u>Methods for Cracking:</u> Catalytic and steam cracking.
1) Heat long-chain hydrocarbons to **<u>vaporise</u>** them (turn them into gas)
2) **<u>Vapour</u>** is passed over **<u>hot</u>** powdered aluminum oxide **<u>catalyst</u>**
3) Long-chain molecules split apart on the surface of the speck of the **<u>catalyst</u>** = **<u>catalytic cracking</u>**

1. **<u>Vaporise</u>** hydrocarbons and mix them with **<u>steam</u>**
2. **<u>Heat</u>** to very high temperature = **<u>steam cracking</u>**

$$\underset{\text{decane}}{C_{10}H_{22}} \xrightarrow{500\,^{\circ}C\ +\ \text{catalyst}} \underset{\text{pentane}}{C_5H_{12}} + \underset{\text{propene}}{C_3H_6} + \underset{\text{ethene}}{C_2H_4}$$

This cracking reaction is an example of thermal decomposition.

<u>**Alkenes**</u>: Are unsaturated hydrocarbons. This means they have 2 fewer H atoms and are joined by **double C=C bonds**.
<u>Testing to see if it's an alkane or an alkene</u>
Add <u>orange</u> bromine water and shake
Alkane = stays orange
Alkene = colourless



Purity

In chemistry a pure substance contains only an **element** or a **compound**. It's not mixed with anything else. But in everyday language, a pure substance can mean a substance that has had nothing added to it, so it is in its natural state, e.g. pure milk.

The melting point (MP) or boiling point (BP) tells you how pure a substance is

- Pure elements and compounds **melt** and **boil** at **specific** temperatures
- You can test the purity of a sample by measuring its BP and MP, and then compare it to the BP and MP of pure substances (find from a data book)
- The closer your measured value is to the actual BP or MP, the **purer** your sample is. i.e. the purer the compound the narrower the range.

Impurities in your sample;

- **Lower** the MP and **increase** the melting range of your substance
- **Increase** the boiling point and may result in your sample boiling at a **range** of temperatures

Formulations

Is a mixture that has been designed to produce a useful product with a precise purpose, that are made by following a 'formula' (a recipe).

E.g. of formulations: paint, medicinal drugs, fragrance additives, fuels, fertilisers, pesticides, alloys, cosmetics & food products.

Paints are formulations, they contain:

- A **pigment**, to provide colour
- A **binder (resin)**, to help the paint attach itself to an object and to form a protective film when dry
- A **solvent**, to help the pigment and binder spread well (dissolve) during painting by thinning them out (alter the viscosity)
- An **additive**, to further change the physical and chemical properties of the paint.

Washing up liquids are formulations, they contain:

- A **surfactant**, the actual detergent that removes the grease.

Continued…

- **Water**, to thin out the mixture so it can squirt more easily from the bottle.
- **Colouring and fragrance additives**, to improve the appeal of the product to customers.
- **Rinse agent**, to help water drain off crockery

Formulations in the industry

Are very important. E.g. pharmaceutical industry

Medicines are formulations:

Alter formulations of a pill, to ensure it delivers the drug to the correct part of the body; At the right concentration; To make sure it can be consumed; It has a long enough shelf life etc.

E.g. products have info about composition on the packaging;

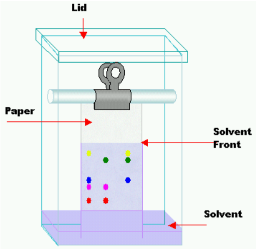
Ratio/percentage of each component

Choose the right composition for your particular use

Chromatography

Can be used to separate mixtures and give information to help identify substances.

Chromatography always involves a **mobile phase** and a **stationary phase**.



- The mobile phase, where the molecules **can** move. Always liquid or gas.
- The stationary phase, where the molecules **cannot** move. Solid or really thick liquid.
- During chromatography, the substance constantly move between mobile (M) and a stationary (S) phase = Equilibrium formed
- The mobile phase, moves through the stationary phase, and anything dissolved in the mobile phase moves with it.
- How quick a chemical moves depends on 'distribution' between phases, i.e. how much more time it spends in M or S phase.
- More time in M phase = move further
- Components in a mixture normally separate through S phase

- **Pure substance** = **one spot only**, one substance, in any solvent
- If the unknown sample is a mixture of compounds, there is usually more than one spot formed on the chromatogram.
- A substance with a stronger force of attraction between itself and the mobile phase is carried further
- than a substance with a stronger force of attraction between itself and the stationary phase.
- In paper chromatography the **mobile** phase is the **solvent** (e.g. water or ethanol)
- The **stationary** phase is the **paper**.

How long molecules spend in each phase depend:

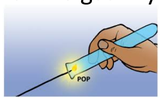
1) how soluble they are in the solvent
2) how attached they are to paper

- Molecules with **higher solubility** and **less attracted** to paper = spend more time in M phase

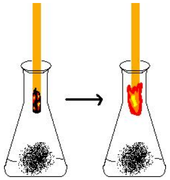
$$Rf = \frac{\text{distance moved by the substance}}{\text{distance moved by the solvent}}$$

Test for gases

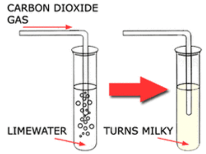
Hydrogen: Use a burning splint held at the open end of a test tube of the gas. Hydrogen burns rapidly with a pop sound.



Oxygen: Use a glowing splint inserted into a test tube of the gas. The splint relights in oxygen.



Carbon dioxide: Use an aqueous solution of calcium hydroxide (lime water). When carbon dioxide is shaken with or bubbled through limewater, the limewater turns milky (cloudy).



Chlorine: Use litmus paper. When damp litmus paper is put into chlorine gas the litmus paper is bleached and turns white.

Early and Current Atmosphere

During the first billion years of the Earth's existence there was intense volcanic activity that released gases that formed the early atmosphere and water vapour which condensed to form the oceans. Similar to the atmospheres of Mars and Venus today, consisting of mainly carbon dioxide with little or no oxygen gas.

Volcanoes also produced nitrogen which gradually built up in the atmosphere along small proportions of methane and ammonia. The carbon dioxide dissolved in the formed oceans and carbonates were precipitated producing sediments, reducing the amount of carbon dioxide

Present Atmosphere
~80% Nitrogen
~20% Oxygen
Trace amounts of CO2, Water Vapour and noble gases

Changes from the early atmosphere

Algae first produced oxygen about 2.7 billion years ago and soon after this oxygen appeared in the atmosphere. Over the next billion years plants evolved and the percentage of oxygen gradually increased to a level that enabled animals to evolve.

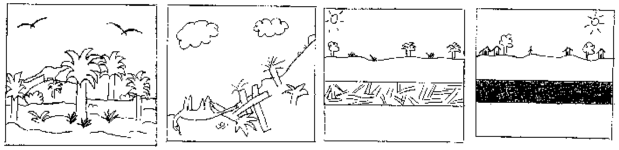
Photosynthesis

$$6CO_2 + 6H_2O \longrightarrow C_6H_{12}O_6 + 6O_2$$

carbon dioxide + water ⟶ glucose + oxygen

Algae also decreased the amount of Carbon dioxide in the atmosphere via photosynthesis, along with carbon dioxide forming sedimentary rocks and fossil fuels

Formation of Coal, Gas, Crude Oil



Coal is formed from trees in swamps millions of years ago. When these trees and animals die they get buried in mud. Layers form over them and the pressure and heat over time results in the formation of coal which is then mined. Oil and Natural gas are also formed in this process except they are formed by marine organisms in the sea.

Limestone is also produced from dead living organisms. The creatures themselves have decayed but their skeletons and shells undergo compaction form Limestone (Calcium Carbonate) $CaCO_3$

Global Warming

Scientists believe that greenhouse gases, such as Methane and Carbon Dioxide, are causing the planets temperature to increase, resulting in global climate change.

The burning of fossil fuels is one way in which we are increasing the amount of Carbon Dioxide in our atmosphere. The increase in the amount of cattle also results in more Methane which equally increases the temperature.

Global Warming can effect;

- Agriculture due to desertification
- Extreme weather conditions
- Increase in sea levels due to glaciers melting
- Changing of natural wildlife habitats

These will also have social effects on businesses who rely on the income generated from agriculture in the effected regions, furthermore homes will also be destroyed due to increased sea levels.

Atmospheric Pollutants

When fuels undergo combustion the gases released;

- Carbon Dioxide
- Carbon Monoxide
- Sulfur Dioxide
- Nitrogen Oxides
- Particulates

Fuels undergo either complete or incomplete combustion

fuel + oxygen ⟶ carbon dioxide + water

$$C_mH_n + O_2 \longrightarrow CO_2 + H_2O$$

fuel + oxygen (not enough) ⟶ carbon + water

$$C_mH_n + O_2 \longrightarrow C + H_2O$$

carbon monoxide CO

carbon dioxide CO_2

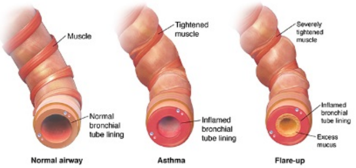
other carbon oxides C_xO_y

Atmospheric Pollutants

Carbon Monoxide is a toxic gas (the silent killer) as it is colorless, odorless and not easily detectable.

Sulphur Dioxide and Nitrogen oxides cause acid by dissolving into water droplets in clouds, this makes the rain more acidic which can damage buildings and wildlife.

Particulates are unburnt carbon particles. These are absorbed into the clouds and cause more water droplets to form in clouds. Theyr also make clouds better at reflecting sunlight, which causes global dimming.



Sulfur dioxide, Nitrogen Oxides and particulates also cause respiratory health problems for humans

Earths Resources + Recycling

Humans use the Earth's resources to provide warmth, shelter, food and transport.

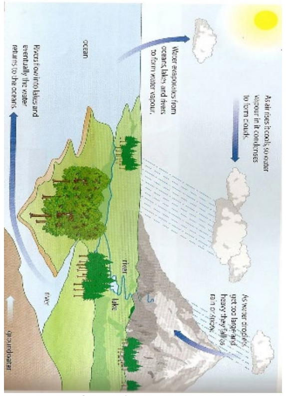
Natural resources, supplemented by agriculture, provide food, timber, clothing and fuels. For example Wood, Cotton and Leather.

Finite resources from the Earth, oceans and atmosphere are processed to provide energy and materials. For example, Stone, Gold and Crude oil.

Obtaining raw materials from the Earth by quarrying and mining causes environmental impacts. Some products, such as glass bottles, can be reused, other products cannot be reused. Metals can be recycled by melting and recasting or reforming into different products. Different materials require different levels of separation before recycling. For example, some scrap steel can be added to iron from a blast furnace to reduce the amount of iron that needs to be extracted from iron ore

Potable Water

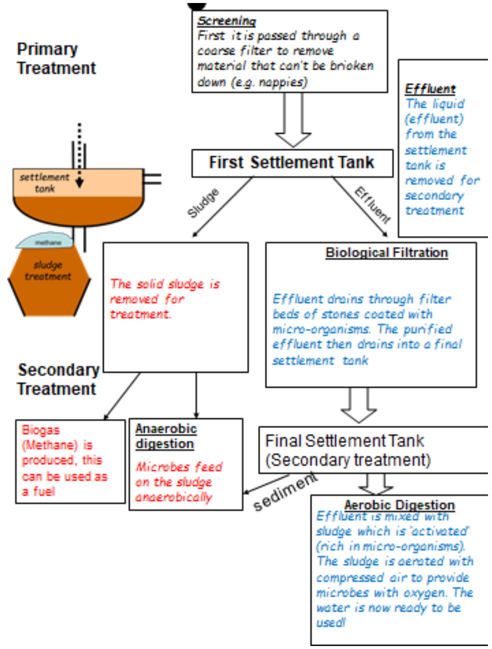
Water that is safe to drink is called potable water. Potable water is not pure water in the chemical sense because it contains dissolved substances.



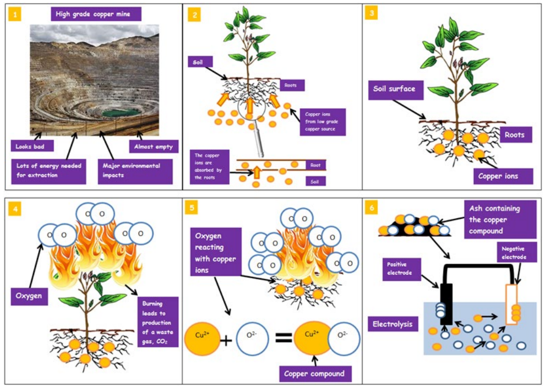
In the UK rain provides low levels of fresh water that collects in; the ground, lakes and rivers. It is then passed through filter beds to remove larger contaminants it is then sterilized by using either chlorine, ozone or ultraviolet light to kill bacteria.

It is not always possible to access fresh water. Desalination of sea water can be done by distillation or by processes that use membranes such as reverse osmosis. These processes require large amounts of energy.

Waste Water Treatment



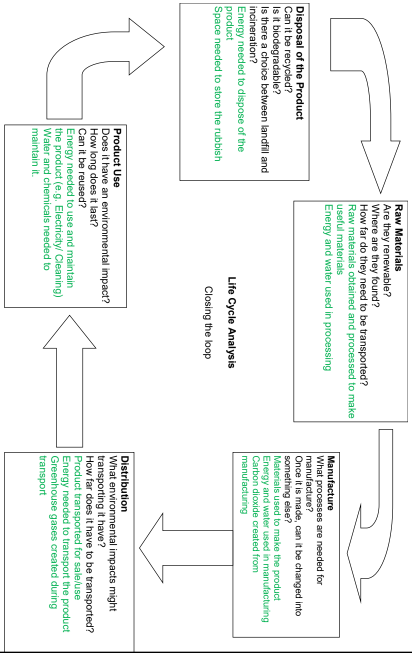
Phytomining



Bioleaching

In Bioleaching, bacteria are mixed with low grade copper ore. The bacteria produce leachate solutions that contain metal compounds from the ore. These metal compounds can then be separated by displacement or electrolysis

Life Cycle Assessment

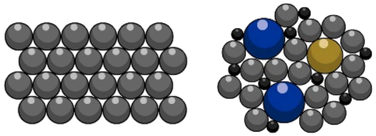


LCA'S assess the environmental impact of products in each of these stages

- Extracting and processing raw materials
- Manufacturing and packaging
- Use and operation during its life time
- Disposal at the end of its useful life

Alloys (Triple Only)

Most metals in everyday use are alloys.



Bronze is an alloy of copper and tin and used in coins. Brass is an alloy of copper and zinc and used in piping.

Gold used as jewelry is usually an alloy with silver, copper and zinc. The proportion of gold in the alloy is measured in carats. 24 carat being 100 % (pure gold), and 18 carat being 75 % gold.

Steels are alloys of iron that contain specific amounts of carbon and other metals. High carbon steel is strong but brittle. Low carbon steel is softer and more easily shaped. Steels containing chromium and nickel (stainless steels) are hard and resistant.

AQA Science: Using resources

Ceramics, Composites and Glass (Triple Only)

Glass
Most of the glass we use is soda-lime glass, made by heating a mixture of sand, sodium carbonate and limestone. Borosilicate glass, made from sand and boron trioxide, melts at higher temperatures than soda-lime glass.

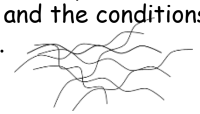
Ceramics
Pottery and brick ceramics are made by shaping wet clay and then heating in a furnace

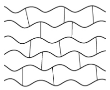
Composites
Most composites are made of two materials, a matrix or binder surround and binding together fragments of the different material called the reinforcement

Uses
- Fiberglass
- Dental Fillings
- MDF

Polymers

Polymers depend on what monomers they are made from and the conditions under which they are made.

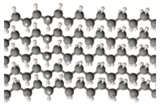


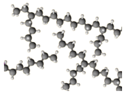


Thermosoftening
Individual tangles polymers that melt when heated

Thermosetting
Polymer chains with cross links between them, they do not melt when heated

LDPE AND HDPE
Ethene can be polymerized in slightly different ways to produce low density polyethene and high density polyethene. LDPE has side branches that stop the polymer molecules lining up properly, it is not crystalline, therefore it has weaker bonds and has a lower melting point. HDPE has a crystalline structure Therefore has a higher melting point due to the force of attractions





Corrosion Prevention (Triple Only)

Rusting is an example of corrosion. Both air and water are necessary for iron to rust.

There are 2 main methods to help prevent rusting

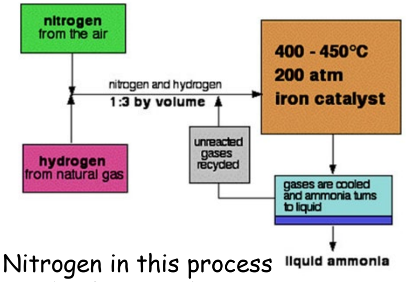
1) Putting a barrier between the metal and the air and moisture e.g. grease, paint, plastic, unreactive metal

2) Sacrificial protection -The iron is covered or connected to a more reactive metal so it will 'corrode' ahead of the iron.

Corrosion = Metals corrode when they react with oxygen to form oxides in the presence of water/moisture

Copper + Oxygen $\xrightarrow{\text{Water}}$ Copper oxide

Haber Process (Triple Only)



The Nitrogen in this process is from the fractional distillation of air and the hydrogen is obtained from the natural gas (CH4) or electrolysis of water

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The reaction is reversible. TO separate the ammonia the, it is cooled and removed, the remaining nitrogen is recycled

NPK Fertilisers (Triple Only)

Compounds such as Nitrogen, phosphorus and potassium (NPK) are used in agriculture.

Ammonia can be used to manufacture ammonium salts and nitric acid to give the nitrogen (N). Potassium (K) comes from mining of potassium chloride and potassium sulphate. Phosphorous (P) comes from mining phosphate rock.

Acid	Alkali	Fertiliser
Nitric Acid	Ammonia Solution	Ammonium Nitrate
Phosphoric Acid	Ammonia Solution	Ammonium Phosphate
Sulfuric Acid	Ammonia Solution	Ammonium Sulphate