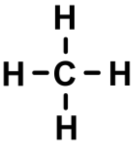
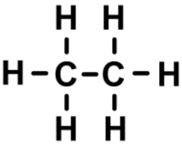
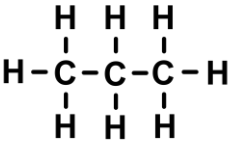
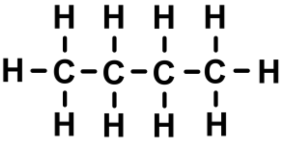
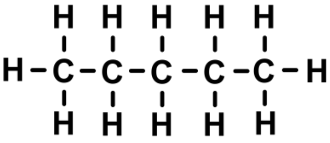
Contents

Organic Chemistry

Alkanes

These are hydrocarbons with **single covalent** bonds between the carbon atoms. They are referred to as **saturated hydrocarbons** for this reason.

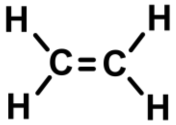
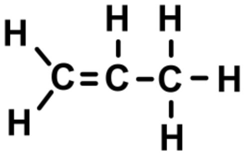
The general formula for an alkane – $C_nH_{(2n+2)}$

Name	Molecular Formula	Structural formula	Skeletal structure
Methane	CH_4	 $H-\underset{H}{\overset{H}{C}}-H$	
Ethane	C_2H_6	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	
Propane	C_3H_8	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	
Butane	C_4H_{10}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	
Pentane	C_5H_{12}	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$	

Alkenes

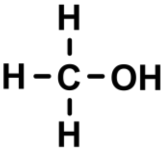
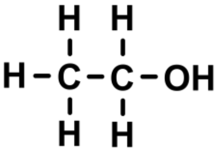
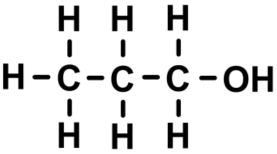
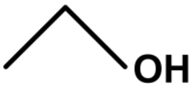
An alkene is an organic compound that contains a double bond between two carbon atoms. For this reason an alkene is referred to as an **unsaturated hydrocarbon**.

The general formula for an alkene – C_nH_{2n}

Name	**Molecular Formula**	**Structural formula**	**Skeletal structure**
Ethene	C_2H_4		
Propene	C_3H_6		

Alcohols

The general formula for a simple alcohol – $C_nH_{(2n+1)}OH$

Name	**Molecular Formula**	**Structural formula**	**Skeletal structure**
Methanol	CH_3OH		
Ethanol	C_2H_5OH		
Propanol	C_3H_7OH		

Functional group

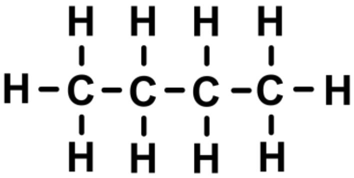
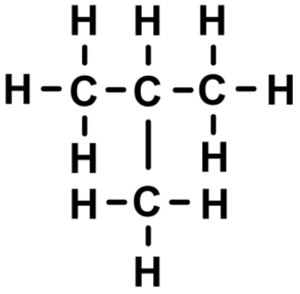
Functional groups are groups of atoms within molecules that are responsible for the characteristic chemical reactions of those molecules. The same functional group will undergo the same or similar chemical reaction(s) regardless of the size of the molecule it is a part of.
The functional group for alkenes is the double bond between the carbon atoms C=C.
The functional group for alcohols is the –OH group.

Isomers

In organic chemistry, isomers are molecules with the same molecular formula (i.e. the same number of atoms of each element), but different structural or spatial arrangements of the atoms within the molecule.

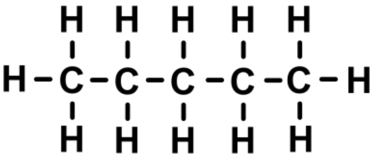
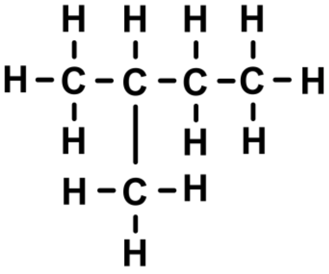
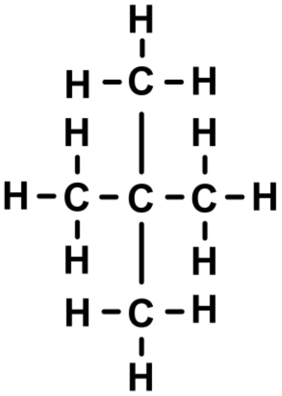
Isomer – has the same molecular formula but has a different structure.

Structural isomers of Butane

butane	methylpropane / 2-methylpropane
	

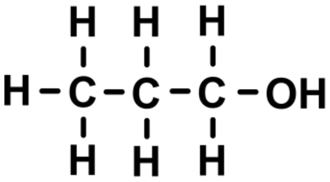
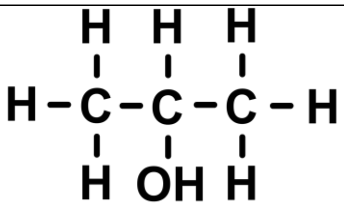
Structural isomers of Pentane

Note: You will not be required to remember the names of the isomers

pentane	2-methylbutane	2,2-di-methylpropane
		

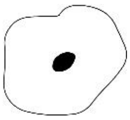
Positional isomers of Propanol

Positional isomers occur when the ‘functional group’ of the molecule changes position.

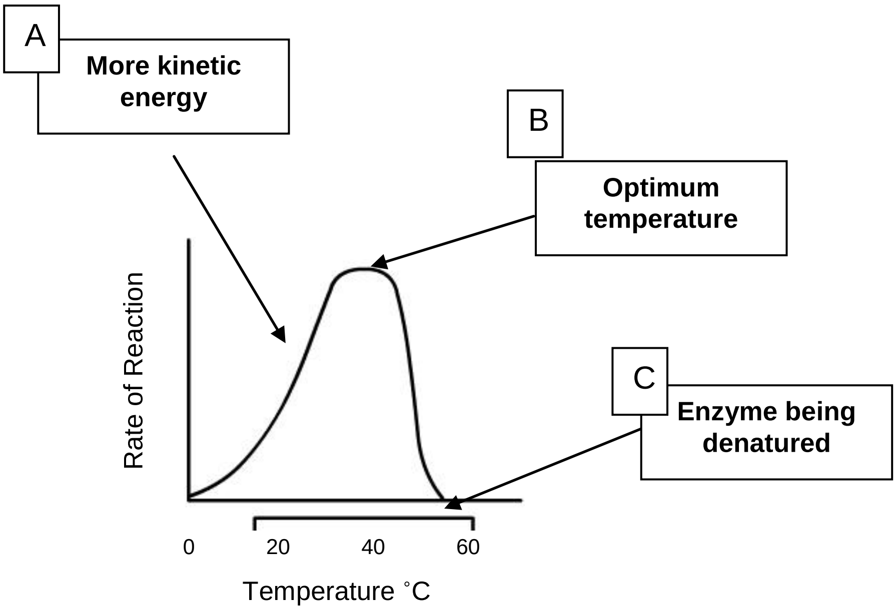
propan-1-ol	Propan-2-ol
	

Enzymes

- Enzymes are catalysts formed from living cells



- A catalyst is a substance that speeds up a reaction (without getting used up in the process) and remains unchanged at the end
- Enzymes are globular proteins
- Temperature affects the reaction rates catalysed by enzymes.



A Reacting molecules have more and more kinetic energy as temperature increases. This increases the chances of a successful collision therefore the rate increases.

B The temperature at which the enzyme works best is referred to as the optimum temperature. The optimal temperature for human body enzymes is 37.5°C

C Too much heat destroys the activity of the enzyme as the structure begins to break down (denature). They are denatured at temperatures over 60°C.

Enzymes are used in **biotechnology** and in the **baking**, **brewing** and **milk** industries.

Fermentation

Ethanol is produced from the fermentation of glucose by yeast.

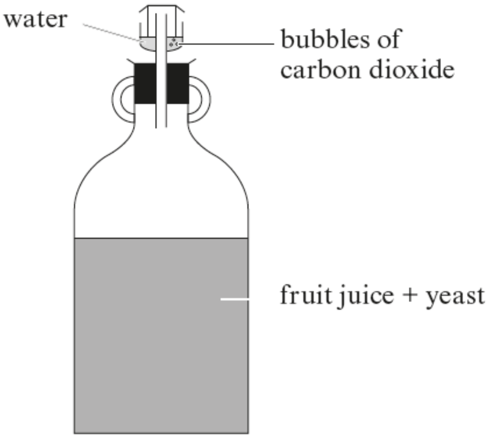
Yeast is a living, single – cell microorganism that belongs to the fungi kingdom.

Yeast contains enzymes that catalyse the breakdown of glucose to ethanol and carbon dioxide.

Beer and **wine** are produced by **fermenting glucose** with **yeast.**

$$\text{glucose} \xrightarrow{\text{yeast enzymes}} \text{ethanol} + \text{carbon dioxide}$$

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$



Conditions necessary for fermentation to occur

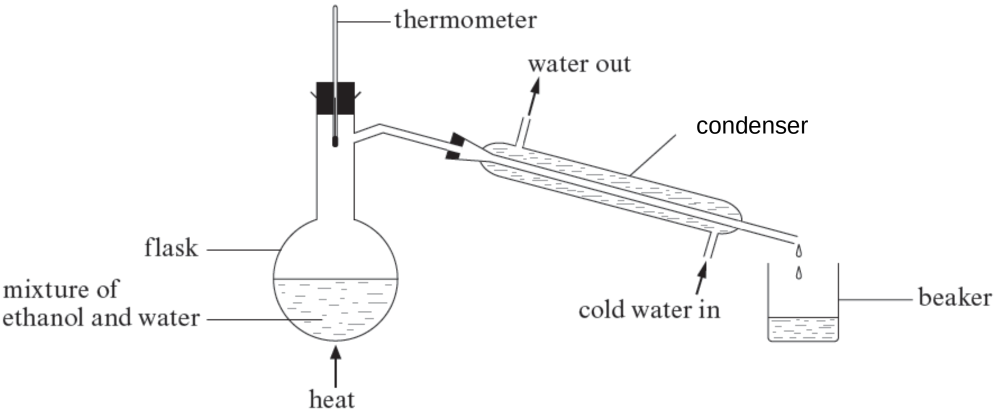
For the yeast enzymes to work they need

- a glucose solution (glucose and water)
- a temperature in the range of 20-40°C
- absence of oxygen
- a pH in the range of 4 to 7

To obtain ethanol from the mixture, yeast is removed by **filtering**.

Then the ethanol and water (and some sugar) mixture is **distilled**.

Ethanol has a lower boiling point (79°C) than water (100°C) or water glucose mixture (100-110°C). This physical property allows ethanol to evaporate quicker in the flask. A cold condenser will allow this vapour to change state into liquid ethanol leaving the water/glucose solution behind.



Ethanol - Health, Social and Economic impacts

Health problems with excessive use of alcohol over many years can include

- liver or kidney damage
- vitamin deficiency
- heart disease
- memory loss
- depression
- stomach disorders
- cancer
- brain damage
- high blood pressure

Drinking excessively can lead to a number of harmful effects such as **alcohol poisoning** and **cirrhosis of the liver**.

Alcohol poisoning*

Alcohol poisoning occurs when excessive amounts of alcohol start to interfere with the body's automatic functions such as:
- breathing
- heart rate
- gag reflex, which prevents you choking

Alcohol poisoning can cause a person to fall into a coma and could lead to their death.

Cirrhosis of the liver*

Cirrhosis is scarring of the liver as a result of continuous, long-term liver damage. Scar tissue replaces healthy tissue in the liver and prevents the liver from working properly.
The damage caused by cirrhosis can't be reversed and eventually can become so extensive your liver stops functioning. This is called **liver failure.**

Social and economic impact of alcoholic drinks

Excessive use of alcohol can result in anti-social behaviour

- Aggressive behaviour
- Domestic violence
- Road accidents due to drink driving
- Wasting emergency services time
- Increases the cost of emergency services
- Tax on alcoholic drinks
- The tax raised from sale of alcoholic drinks generates significant revenue for the government.

*NHS Direct

Ethanol as a fuel

Ethanol can be produced from the fermentation of plants such as sugar cane. These are referred to as fuel crops and after distillation results in the production of bioethanol.

Bioethanol produces only carbon dioxide and water as waste products.

Bioethanol is **carbon neutral** because the carbon dioxide released during fermentation and combustion is equivalent to the amount removed from the atmosphere while the crop is growing.

Bioethanol is also **renewable**.

To decrease on the dependence on fossil fuels and increase the energy from renewable sources, European governments have agreed to add some bioethanol to some petrol blends.

(Less sulfur dioxide will be formed which prevents acid rain forming.)

Some critics warn of deforestation, and land being grabbed from food crops. This will increase food poverty as food prices are forced up.

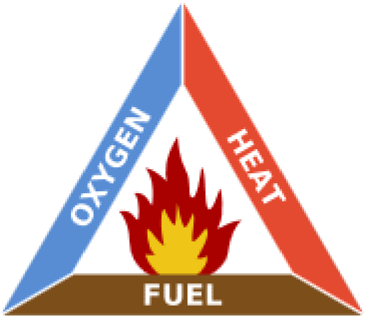
Ethanol as a solvent

Not all substances dissolve in water. Ethanol is used extensively as a solvent:-

- in the manufacture of varnishes and perfumes;
- as a preservative for biological specimens;
- in the preparation of essences and flavorings;
- in many medicines and drugs;
- as a disinfectant

The Fire Triangle

The fire triangle is a simple way of understanding the factors essential for fire. Each side of the triangle represents one of the three factors required for the creation and maintenance of any fire; **oxygen, heat and fuel**. Remove any one of these, the triangle is broken and the fire is stopped.



Removing Heat

Heat can be removed by the addition of something to reduce it. Water is used to put housefires and bonfires out.

Removing Oxygen

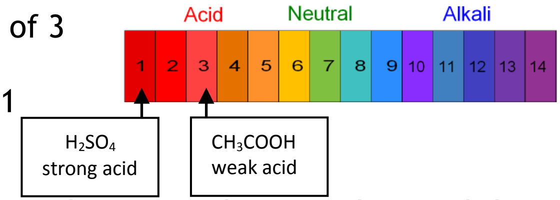
Cover things that are burning with foam, carbon dioxide or a fire blanket to remove the air supply.

- A fire blanket is used to extinguish a chip pan fire or a person on fire.
- Carbon dioxide powder is used to put out indoor fires, chemical and electrical fires.
- Foam is used to extinguish aircraft fire.

Removing Fuel

Without fuel a fire will stop.
Switch off the electrical or gas supply,
Fire-breaks are used to put forest fires out. This is when a section of trees is cleared deliberately to remove the fuel.

Comparing the reactions of Ethanoic Acid and Sulfuric Acid



Ethanoic Acid CH_3COOH is a weak acid, it has a pH of 3

Sulfuric Acid H_2SO_4 is a strong acid, it has a pH of 1

Ethanoic acid and sulfuric acid react similarly but as ethanoic acid is a weaker acid the **reactions are slower.**

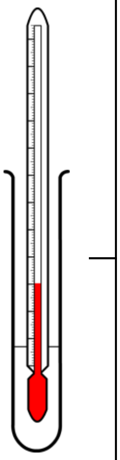
Ethanoic acid forms **ethanoate** salts. Sulfuric acid forms **sulfate** salts

ACID + ALKALI → SALT + WATER

sulfuric acid + sodium hydroxide → sodium sulfate + water

ethanoic acid + sodium hydroxide → sodium ethanoate + water

Observations: temperature rise in both reactions



ACID + BASE → SALT + WATER

sulfuric acid + copper oxide → copper sulfate + water

ethanoic acid + copper oxide → copper ethanoate + water

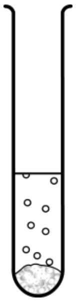
Observations: temperature rise in both reactions, base dissolves slower in ethanoic acid

ACID + CARBONATE → SALT + WATER + CARBON DIOXIDE

sulfuric acid + sodium carbonate → sodium sulfate + water + carbon dioxide

ethanoic acid + sodium carbonate → sodium ethanoate + water + carbon dioxide

Observations: Bubbles of CO_2 given off in both, more bubbles given off in the sulfuric acid reaction

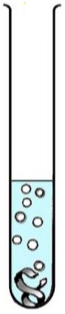


METAL + ACID → SALT + HYDROGEN

magnesium + sulfuric acid → magnesium sulfate + hydrogen

magnesium + ethanoic acid → magnesium ethanoate + hydrogen

Observations: fizzing due to hydrogen gas in both reactions, more fizzing in the sulfuric acid reaction, Mg reacts faster in sulfuric acid



Conclusion: Both acids react similarly but the rate of reaction for ethanoic acid is slower as it is a weaker acid.

Industrial Processes

Creating Ammonia – The Haber process

Hydrogen and atmospheric nitrogen need to react to form ammonia. This reaction is different to most because it is **reversible.** This means the reaction can go forwards or backwards depending on the conditions.

The theoretical conditions needed for a high yield, forward reaction to occur would be *low temperature* with *high pressure*.

The word and symbol equation for this are:-

$$\text{Nitrogen} + \text{Hydrogen} \rightleftharpoons \text{Ammonia}$$

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

The industrial process of making ammonia is called the **Haber process.**

The conditions used in the manufacture of ammonia are:-

350-450˚C	(relatively high temperature)
150-200 atmospheres	(relatively low pressure)
Iron catalyst	

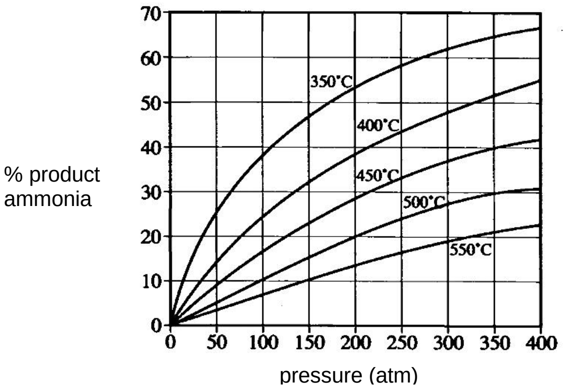
This is an exothermic reaction that creates liquid ammonia on condensing.

From an industrial point of view, a lower temperature would cause the process to be too slow. This explains the moderately high temperature chosen.

A pressure of 150-200 atmospheres is used as creating equipment to maintain a higher pressure is too expensive.

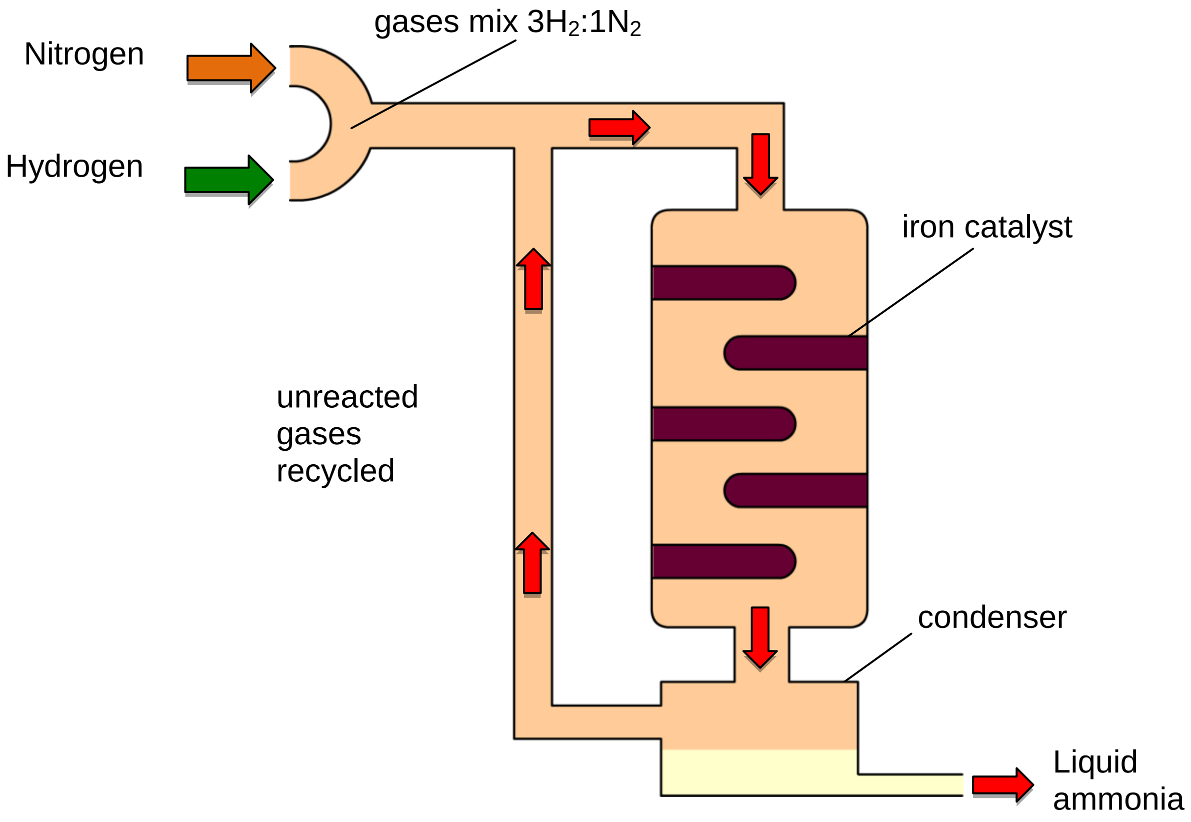
The yield is compromised to 15-40%. The unreacted nitrogen and hydrogen are recycled. This way the greatest amount of ammonia per day/week/month is achieved.

The following graph illustrates the effect of temperature and pressure on the yield of ammonia formed



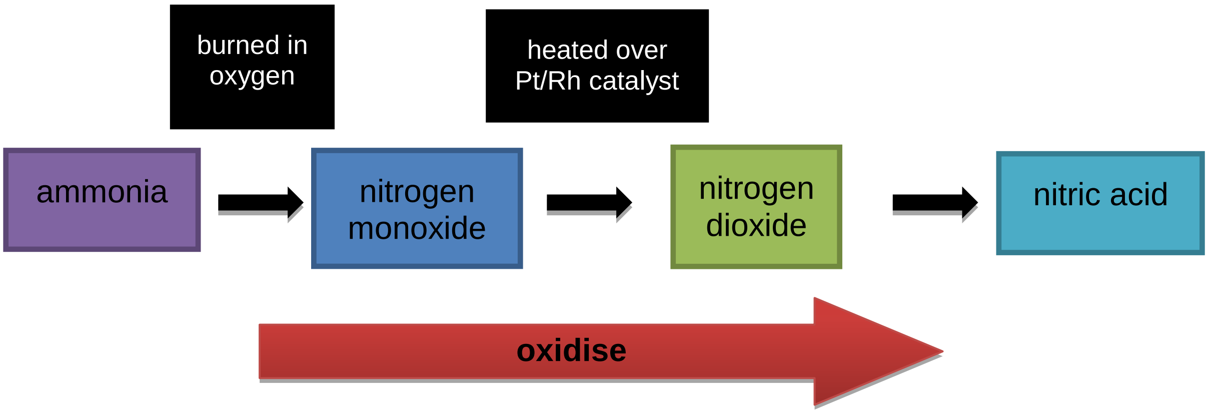
Although the iron catalyst speeds up the reaction, over time this will become poisoned and will reduced the speed at which ammonia is created.

Here is a diagram to show how ammonia is formed and the unreacted gases recycled.



Formation of nitric acid

Some ammonia is converted into nitric acid by oxidation. A **platinum/rhodium** catalyst is used to speed up the reaction.



Formation of sulfuric acid – The Contact Process

The industrial process of making sulfuric acid is called the **Contact process.**

There are three stages to the process.

Stage 1 – Production of Sulfur dioxide

Sulfur dioxide is obtained by burning sulphur in air;

$$\text{sulfur + oxygen} \longrightarrow \text{sulfur dioxide}$$

$$S\,(s) + O_2\,(g) \longrightarrow SO_2\,(g)$$

Stage 2 – Production of Sulfur trioxide (Important step)

The sulfur dioxide is then oxidised to sulfur trioxide by reacting with **excess air** over a **vanadium oxide (V_2O_5)** catalyst (a transition metal compound) at **420°C** at pressures not much above **atmospheric**. A yield of about **95 %** is obtained.

$$\text{sulfur dioxide + oxygen} \xrightleftharpoons[420°C]{\text{vanadium oxide catalyst}} \text{sulfur trioxide}$$

$$2SO_2\,(g) + O_2\,(g) \xrightleftharpoons[420°C]{V_2O_5} 2SO_3\,(g)$$

As the reaction is exothermic a low temperature favours a forward reaction. The speed of sulfur trioxide formation at low temperature would be too slow, therefore a temperature of 420°C is chosen.

Stage 3 – Conversion to Sulfuric Acid

Sulfur trioxide is **absorbed in sulfuric acid** (98% H_2SO_4, 2 % water). Sulfur trioxide cannot be absorbed directly in 100% water, as the reaction is violent and produces a mist of the acid. The SO_3 reacts with the small quantity of water.

$$\text{sulfur trioxide + water} \longrightarrow \text{sulfuric acid}$$

$$SO_3\,(g) + H_2O\,(l) \longrightarrow H_2SO_4\,(l)$$

Fertilizers

Most of the ammonia formed is used to make fertilizers. These are nitrogen rich compounds which are spread on farmlands for better plant growth.

When these dissolve in rainwater nitrogen is released to the soil.

Healthy plants need nitrogen to make protein.

Ammonia is an alkali and can be **neutralised** with acids to form ammonium salts.

Formation of two fertilizers

ammonium sulfate

ammonia + sulphuric acid → ammonium sulfate

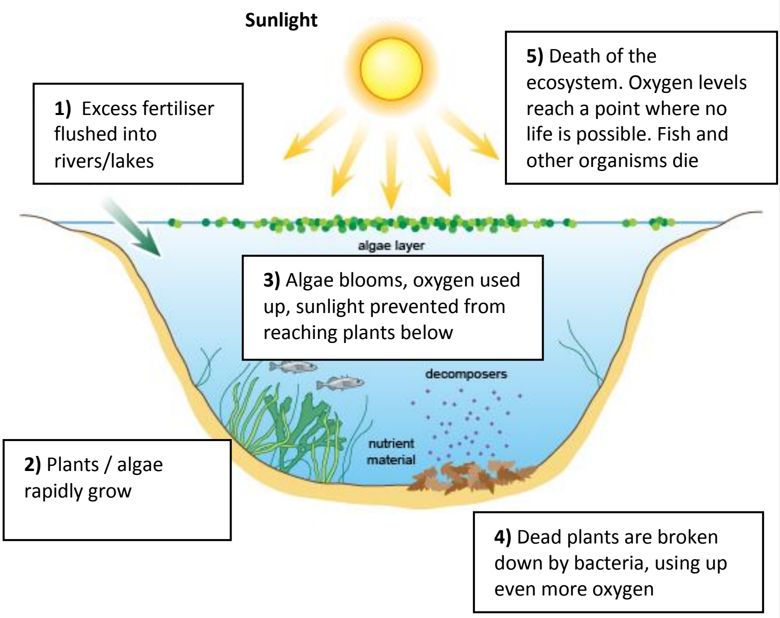
ammonium nitrate

ammonia + nitric acid → ammonium nitrate

Advantages and disadvantages of fertilizers

Advantages	**Disadvantages**
Increases crop yields	Eutrophication
Healthier plants	Could enter water supply Blue baby syndrome
Relatively cheap	
Improves poor quality land	

Eutrophication: If large amounts of fertilisers, especially nitrates, are washed out of the soil into rivers/lakes they can seriously unbalance the equilibria of the natural food chains and life cycles. Fertilisers are used up by water plants which rapidly cover the water. Underwater plants die and decompose as they do not get enough sunlight, as a result creatures such as fish die as the oxygen has been used up by decomposing bacteria. The result is an overgrowth of algae



1) Excess fertiliser flushed into rivers/lakes

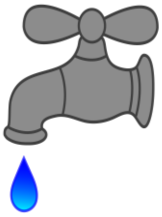
2) Plants / algae rapidly grow

3) Algae blooms, oxygen used up, sunlight prevented from reaching plants below

4) Dead plants are broken down by bacteria, using up even more oxygen

5) Death of the ecosystem. Oxygen levels reach a point where no life is possible. Fish and other organisms die

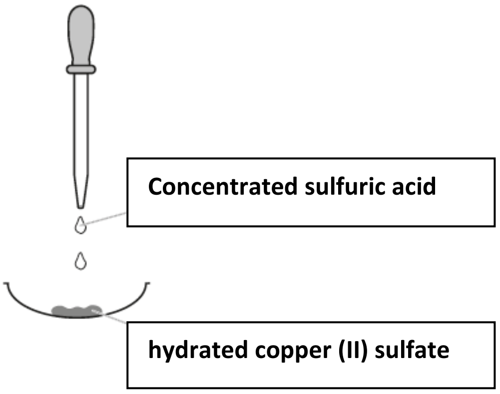
Nitrates in the water supply: High concentration of nitrates from nitrogenous fertilisers can cause cancer in humans. Toxic chemicals are formed due to the presence of nitrates. Nitrates can interfere with the transport of oxygen in the blood in particular in infants where **blue baby syndrome** can occur.



Other uses of sulfuric acid

Sulfuric acid can be used as a **dehydrating agent**. This means that the elements of water can be removed from a substance.

Dehydration of copper sulfate.



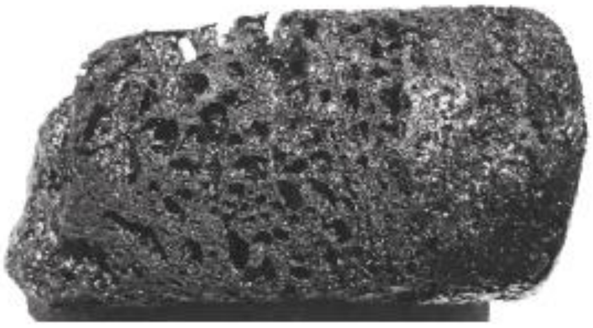
The blue hydrated copper(II) sulfate turns white.

The crystalline hydrated copper (II) sulfate turns powdery and crumbly losing its crystalline appearance.

Dehydration of sugar

When concentrated sulfuric acid is added to sugar a black solid is formed.

Acid dehydrates the sugar removing the elements of water namely hydrogen and oxygen. The black mass left is due to the carbon.



Limestone

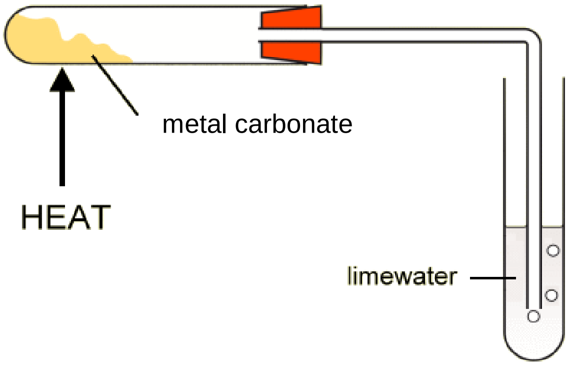
Metal carbonates such as calcium carbonate or copper carbonate break down when heated strongly. When a compound splits into two or more by heat it is referred to as **thermal decomposition**.

e.g.

calcium carbonate → calcium oxide + carbon dioxide

$CaCO_3 (s) \rightarrow CaO (s) + CO_2 (s)$

The rate at which a metal carbonates decompose can be investigated using the following apparatus.



The speed of decomposition can be measured by timing the colour change of limestone. It turns from clear to milky in the presence of carbon dioxide

Metal Carbonate	**Sodium carbonate**	**Calcium carbonate**	**Copper carbonate**
Colour before heating	white	white	green
Colour after heating	white	white	black
Gas evolved	none	carbon dioxide	carbon dioxide
Ease of decomposition	very difficult	fairly easy	easy

No reaction is observed with sodium carbonate as the more reactive metal makes the carbonate more stable.

Calcium is less reactive than sodium therefore heat is able to decompose calcium carbonate fairly easily.

calcium carbonate → calcium oxide + carbon dioxide

$CaCO_3 \rightarrow CaO + CO_2$

Copper is the least reactive metal and decomposes rapidly.

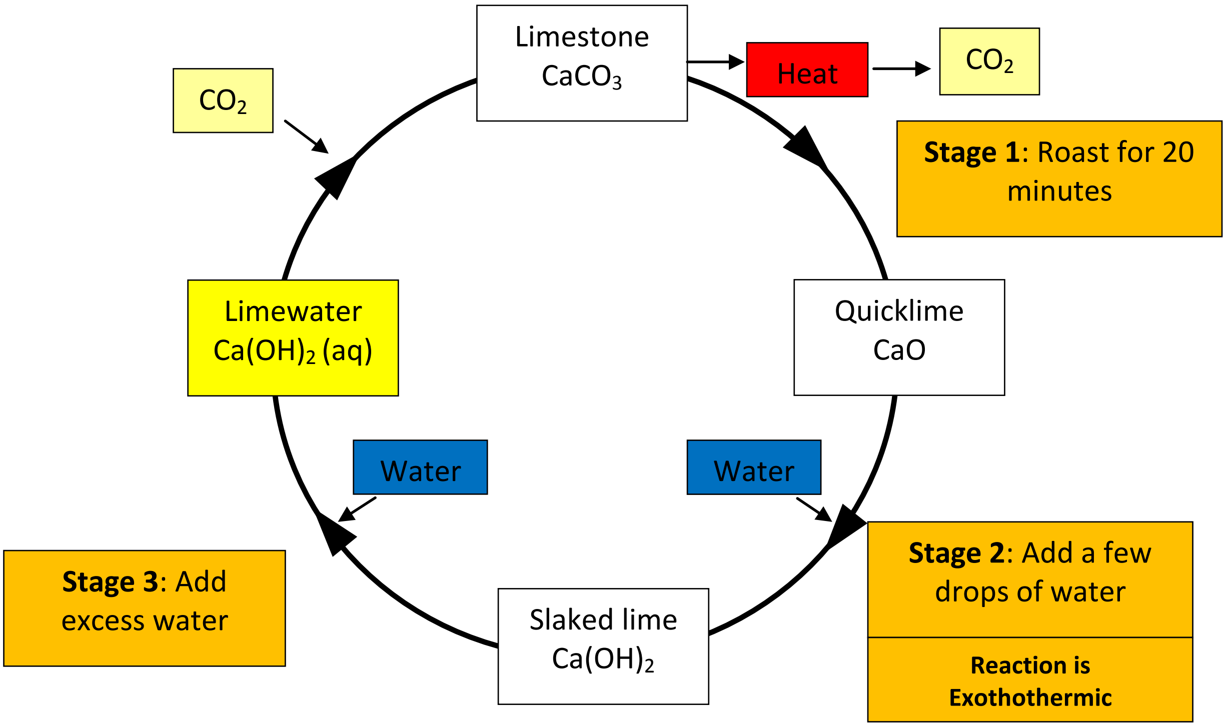
copper carbonate → copper oxide + carbon dioxide

$CuCO_3 \rightarrow CuO + CO_2$

The more reactive the metal the more stable the carbonate

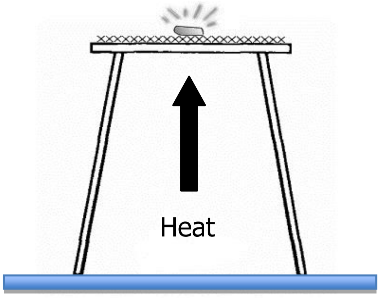
The Limestone Cycle

Limestone can be used to create limewater. There are 3 stages to its formation.



Stage 1: Roast for 20 minutes

Calcium carbonate (limestone) is heated for 20 minutes. Limestone glows and becomes crumbly. This decomposes it to calcium oxide (quicklime)



calcium carbonate → calcium oxide + carbon dioxide
(limstone) (quicklime)

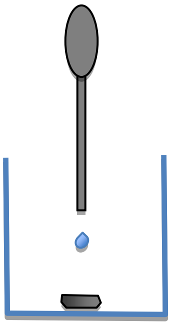
$CaCO_3 \rightarrow CaO + CO_2$

Stage 2: Add a few drops of water

A few drops of water are added to the calcium oxide (quicklime). This causes the compound to sizzle and release steam. This forms calcium hydroxide (slaked lime). The reaction is exothermic.

calcium oxide + water → calcium hydroxide
(quicklime) (slaked lime)

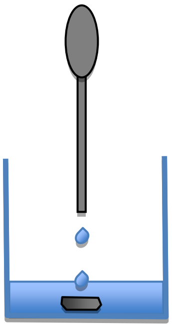
$CaO + H_2O \rightarrow Ca(OH)_2$



calcium hydroxide
$Ca(OH)_2$
(slaked lime)

Stage 3: Add excess water

Calcium hydroxide (slaked lime) dissolves a little in water. Excess water is added to form an alkaline solution called limewater.



calcium hydroxide
$Ca(OH)_2$ (aq)
(limewater)

The reaction of limewater with carbon dioxide

Carbon dioxide passed into limewater gives a milky solution, this is due to the insoluble calcium carbonate formed

calcium hydroxide + carbon dioxide → calcium carbonate + water

$Ca(OH)_2 (aq) + CO_2 (g) \rightarrow CaCO_3 (s) + H_2O (l)$

Clear limewater turning milky is the test for carbon dioxide gas

Uses of Limestone

Limestone is an important raw material obtained by **quarrying**. There are many advantages of using limestone but there are also disadvantages regarding the effects of quarrying for it.

Raw material uses.

- It is used in the production of steel and iron. It is added to iron ore and coke in the blast furnace to remove impurities.
- It is used as a road stone when mixed with bitumen as it has strong physical properties.
- It is used to make cement when it is mixed with clay or sandstone. This can then be mixed with aggregates (mixture of building rocks) to form concrete.
- Limestone is used to neutralise and raise the pH of acidic soils (which are usually less than pH 6.5).
- Quicklime and slaked lime are added to acidic lakes to improve the diversity of aquatic life.
- Limestone is used to absorb acidic waste gases like sulfur dioxide in power station chimneys. These are referred to as ‘sulfur scrubbers’.
- It is used to make glass
- Used in the manufacture of medicinal antacids
- It is used in toothpastes as abrasive

Effects of quarrying limestone.

There are also disadvantages associated with limestone quarrying.

The main disadvantages include :-

- the formation of dust
- noise due to rock blasting
- increased traffic with heavy lorries
- landscape becomes unsightly
- destroys wildlife

The mole

The mole is a term that describes a specific number - like the word ‘dozen’ represents the number 12. The mole however is a much larger number 6.02×10^{23} atoms. (6 followed by 23 zeros).

A mole is defined as the number of atoms in exactly 12 grams of ^{12}C. 12 is the mass number of carbon, so one mole of carbon atoms has a mass of 12 grams.

Mole of atoms
The mass of one mole of the atoms of any element is its ‘relative atomic mass’ in grams

$^{7}_{3}Li$

Relative Atomic Mass (RAM)

One mole of Lithium has a mass of 7 g

Element	R.A.M (Ar)	Mass of 1 mole
H	1	1g
C	12	12g
Ne	20	20g
Mg	24	24g
Ca	40	40g

Moles of molecules
Example: Hydrogen gas exists as H_2 molecules. Because the molecule contains two atoms of hydrogen it has a Relative Molecular Mass (Mr) of 2 (1 x 2). Therefore one mole of hydrogen gas has a mass of 2 g

Molecule	R.M.M (Mr)	Mass of 1 mole
H_2	1 x 2 = 2	2g
O_2	16 x 2 = 32	32g
N_2	14 x 2 = 28	28g
Cl_2	35.5 x 2 = 71	71g

Moles of compounds
To find the mass of one mole of a compound add up the RAMs of the elements of the compound taking account of the formula. The number you get is the Relative Formula Mass which is the Mr of the compound.
For example CO_2

1 x Carbon atom =	1 x 12 =	12
2 x oxygen atoms =	2 x 16 =	32
TOTAL		44

Compound	R.M.M (Mr)	Mass of 1 mole (Molar Mass)
NaCl	23 + 35.5	58.5g
$MgCl_2$	24 + (35.5x2)	95g
$CaCO_3$	40 + 12 + (16x3)	100g

The molar mass of a compound is its relative molecular (formula) mass expressed in grams

Converting Mass into Moles

Moles of atoms = $\frac{\text{mass}}{\text{Mr}}$

Moles of atoms = amount of substance
Mass = mass in grams
Mr = Molecular mass

How many moles of atoms are there in;
4.8 g of carbon ?

moles = $\frac{4.8}{12}$ = **0.4 moles**

Ar (C) = 12

How many moles are there in; 640 g of oxygen (O_2) ?

moles = $\frac{640}{32}$ = **0.4 moles**

Ar (O) = 16, Mr(O_2) = 16 x 2 = 32

How many moles are there in; 10 g of $CaCO_3$?

moles = $\frac{10}{100}$ = **0.1 moles**

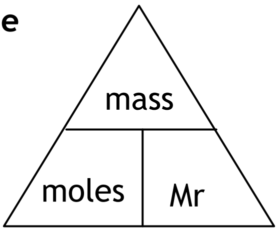
Ar (Ca) = 40, (C) = 12, (O) = 16
Mr $CaCO_3$ = 100

Converting Moles into mass

You can rearrange the equation to form

mass = moles x Mr

If you find rearranging difficult you can use the triangle



What is the mass of 0.05 moles of carbon atoms?

mass = 0.05 x 12 = **0.6 g**

Ar (C) = 12

What is the mass of 0.6 moles of chlorine molecules (Cl_2)?

mass = 0.6 x 71 = **42.6 g**

Ar (Cl) = 35.5, Mr(Cl_2) = 35.5 x 2 = 71

What is the mass of 0.1 moles of calcium carbonate?

mass = 0.1 x 100 = **10.0 g**

Ar (Ca) = 40, (C) = 12, (O) = 16
Mr $CaCO_3$ = 100

It is also possible to calculate the molar mass Mr when mass and the number of moles are known

Mr = $\frac{\text{mass}}{\text{moles}}$

0.5 moles of a compound weighs 80g, calculate its Mr

Mr = $\frac{80}{0.5}$ = 160

Moles and Concentration

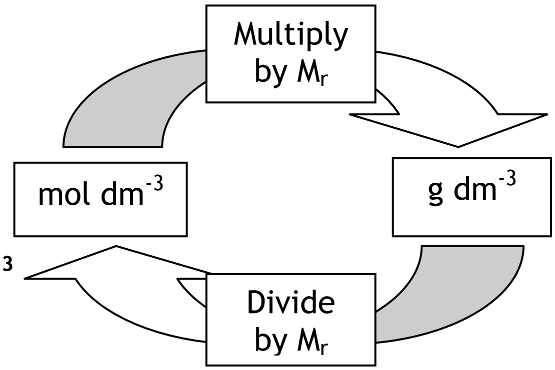
All bottles of solutions in a laboratory must be labelled with its concentration

Concentrations are normally stated in mol dm^{-3}

A 1 mol dm^{-3} solution contains 1 mole of the substance dissolved in 1 dm^3 (or 1000 cm^3)

Preparing a 1 mol dm^{-3} solution is easy.
You weigh out its molecular mass and dissolve in 1000 cm^3 water

Compound	**R.M.M**	**Mass in 1 mol dm^{-3} solution**
NaCl	23 + 35.5	58.5g

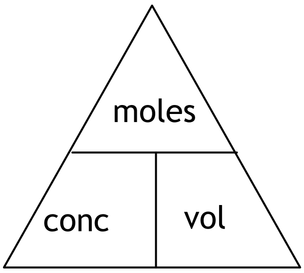


Concentration can sometimes be written as g dm^{-3}
To convert a value from mol dm^{-3} to g dm^{-3}
Multiply by Mr
e.g. 0.01 mol dm^{-3} of NaCl = 0.01 x 58 = **0.58 g dm^{-3}**

To convert a value from g dm^{-3} to mol dm^{-3}
Divide by Mr. e.g. 0.58 g dm^{-3} of NaCl = 0.58 / 58 = **0.01 mol dm^{-3}**

If we know the concentration and volume we can calculate the number of moles in any solution

$$\text{moles} = \text{concentration} \times \frac{\text{volume (in } cm^3\text{)}}{1000}$$



How many moles of sodium chloride are there in 200 cm^3 of a 2.0 mol dm^{-3} solution?

$$\text{moles} = 2.0 \times \frac{200}{1000} = \mathbf{0.4\ moles}$$

What is the concentration of a 100 cm^3 0.05 mole solution?

First we need to rearrange the original equation:

$$\text{concentration} = \frac{\text{moles}}{\text{volume}/1000}$$

$$\frac{0.05}{0.1} = \mathbf{0.1\ mol\ dm^{-3}}$$

A solution of concentration 1 mol dm^{-3} contained 0.25 mole. What was the volume of the solution?

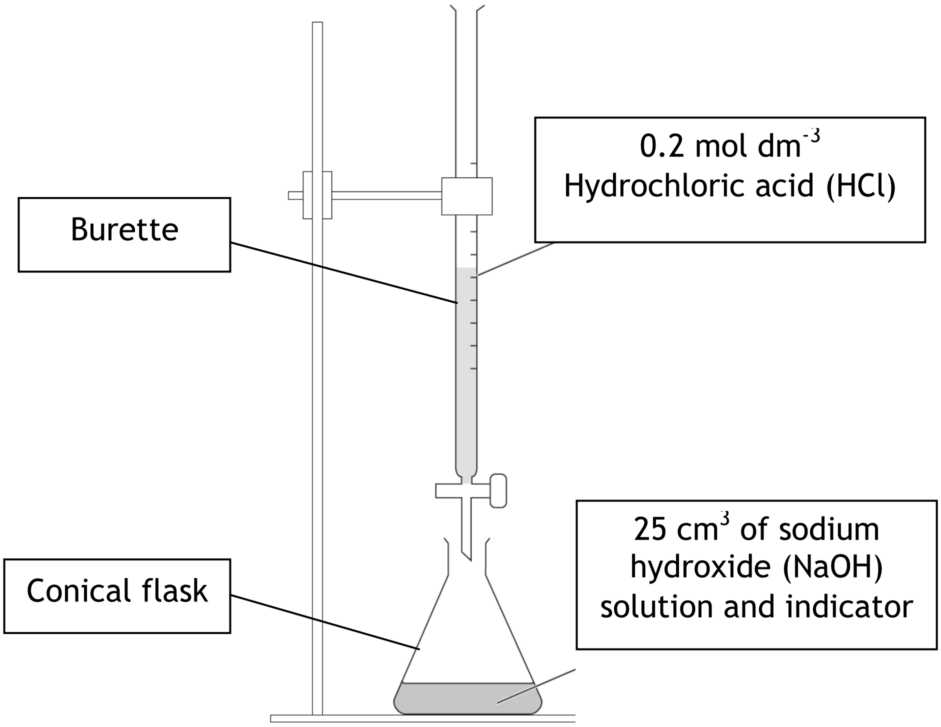
$$\text{volume} = \frac{\text{moles}}{\text{concentration}} = \frac{0.25}{1} = 0.25\ dm^3 \times 1000 = \mathbf{250\ cm^3}$$

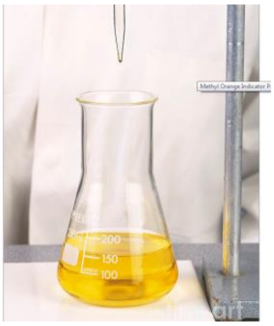
(x1000 to get the answer in cm^3)

Titrations

You can find out the number of moles that react together in solutions using a technique called **titration**.
In titration experiments, you use a burette to add one solution to another.
You need a way to decide when the reaction is complete - The **end point** of reaction.
Indicators are used to determine end points in acid/alkali titrations.

Example: A student was asked to carry out a titration to check the concentration of the sodium hydroxide using the apparatus below



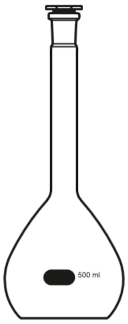


Methyl orange in alkali

End point - neutral

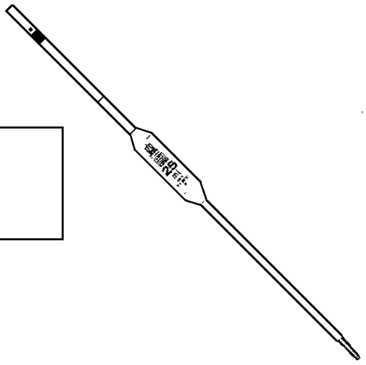
Some practical points:

- When solution are made up – they are made up using a **volumetric flask**
- A **volumetric pipette** is used to accurately measure the solutions that go into the conical flask
- When you fill the burette make sure that you remove the funnel afterwards – it could affect the level of solution in the burette and your results.
- A white tile is placed under the flask so that the colour change is easier to see.



Volumetric flask

Volumetric pipette



- In the example the acid is added slowly by means of a burette, the volume of acid needed to change the indicator colour is recorded.
- It is easy to overshoot the end point the first time (turning the mixture acidic rather than neutral); the first titration is normally a practice run.

Results:

	Titration				
	1	**2**	**3**	**4**	**Average**
Volume of Hydrochloric acid added (cm^3)	**23.50 ***	**20.00**	**20.05**	**19.95**	**20.00**

Calculating the average= Titration $\frac{2 + 3 + 4}{3} = \frac{20.00 + 20.05 + 19.95}{3} = 20.00\ cm^3$

* Titration 1 not used in average as it is not reliable (practice run)

Calculating the concentration of the sodium hydroxide solution

Step 1: Write the balanced equation for the reaction

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

Step 2: Gather the information

	NaOH	HCl
concentration	?	0.2
volume	25 / 1000	20 / 1000
moles	0.004 (step 4)	0.004 (step 3)

Step 3: Calculate the number of moles of HCl used

moles = concentration x $\frac{\text{volume (in } cm^3)}{1000}$

Moles (HCl) $= 0.2 \times \frac{20}{1000} = 0.004$ moles

Step 4: Check the mole ratio

One mole of HCl reacts with one mole of NaOH 1:1

As 1: 1 ratio; 0.004 mole of HCl reacts with 0.004 mole NaOH

Step 5: Calculate the concentration of NaOH

Rearranging gives;

conc = $\frac{\text{moles}}{\text{volume}/1000}$

Concentration = $\frac{0.004}{25/1000} = 0.16\ mol\ dm^{-3}$

Example 2:

25.0 cm^3 of sodium hydroxide (NaOH) solution of unknown concentration was titrated with dilute sulfuric acid (H_2SO_4) of concentration 0.050 mol dm^{-3}. 20.0 cm^3 of the acid was required to neutralise the alkali. Find the concentration of the sodium hydroxide solution in mol dm^{-3}

Step 1: Write the balanced equation for the reaction

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$$

Step 2: Gather the information

	2NaOH	H_2SO_4
concentration	?	0.05
volume	25 / 1000	20 / 1000
moles	0.002	0.001

Step 3: Calculate the number of moles of H_2SO_4 used

$$\text{Moles } (H_2SO_4) = 0.05 \text{ x } \frac{20}{1000} = 0.001 \text{ moles}$$

Step 4: Check the mole ratio

One mole of H_2SO_4 reacts with two moles of NaOH **1** H_2SO_4 : **2** NaOH

Therefore 0.001 mole H_2SO_4 reacts with 0.002 (0.001 x 2) mole NaOH

Step 5: Calculate the concentration of NaOH

$$\text{Concentration} = \frac{0.002}{25/1000} = 0.08 \text{ mol dm}^{-3}$$

Brain teasers:

1. A student reacted 25 cm^3 of 1 mol dm^{-3} HCl with 25 cm^3 of 1 mol dm^{-3} NaOH, is the solution alkali, acidic or neutral ?
2. A student reacted 25 cm^3 of acid (HCl) with 20 cm^3 alkali (NaOH) to make a neutral solution, which had the highest concentration, the acid or the alkali?
3. A student reacted 10 cm^3 of acid (HCl) with 40 cm^3 alkali (NaOH) to make a neutral solution, which had the weakest concentration, the acid or the alkali?

Answers; 1 neutral, 2 alkali, 3 alkali

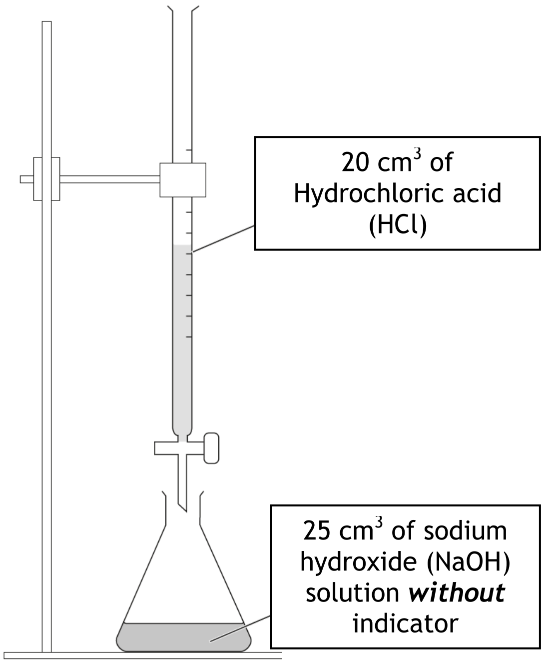
Using titration to prepare pure solutions

Once you have established the end point of a titration you can repeat the experiment without an indicator to obtain a pure sample of the compound.

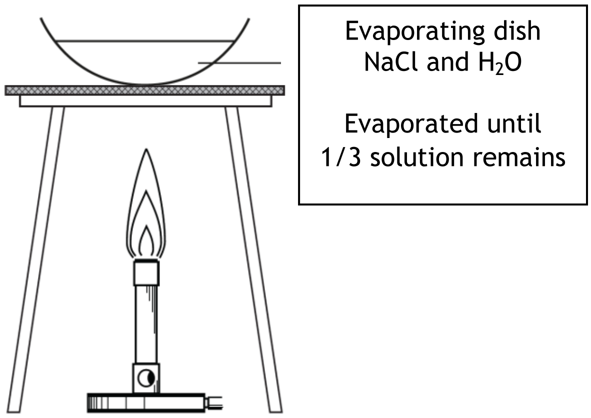
$$NaOH + HCl \longrightarrow NaCl + H_2O$$

In the reaction on page 23, using an indicator it was found that exactly 20.00 cm^3 of the hydrochloric acid was required to neutralise 25.00 cm^3 of the sodium hydroxide alkaline solution. The experiment could be repeated without an indicator to make a pure solution which could be further evaporated to give pure sodium chloride salt

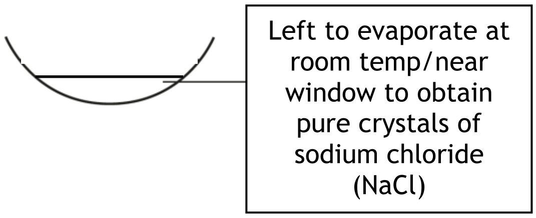
Step 1: Neutralisation



Step 2: Evaporation

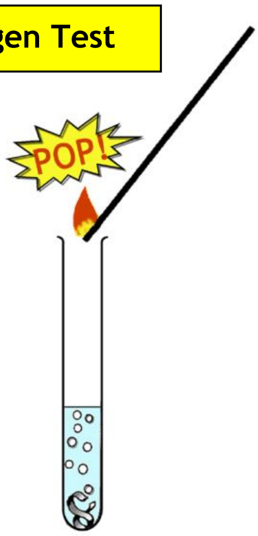


Step 3: Evaporate slowly to dryness



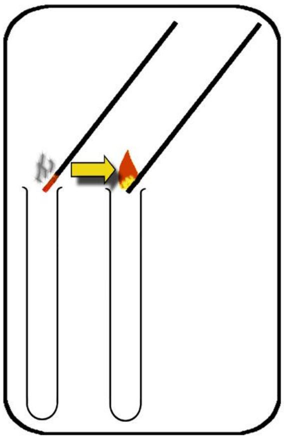
Chemical Analysis - Chemical tests for gases

Hydrogen Test



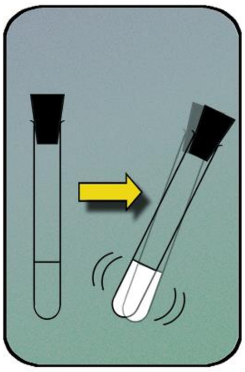
If a lit splint is placed in hydrogen it will create a squeaky 'pop' sound.

Oxygen Test



Oxygen re-lights a glowing splint

Carbon dioxide test

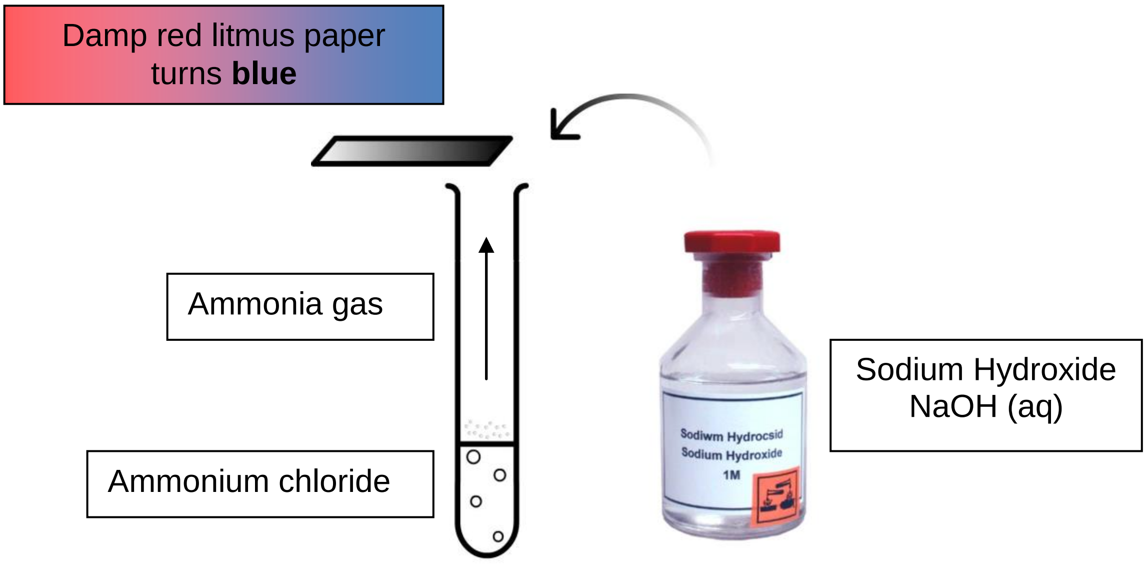


Carbon dioxide turns clear limewater milky.

Ammonia - identifying ammonium salt

If a salt containing ammonium reacts with sodium hydroxide it forms ammonia gas.

Ammonia gas will change damp red litmus paper blue.

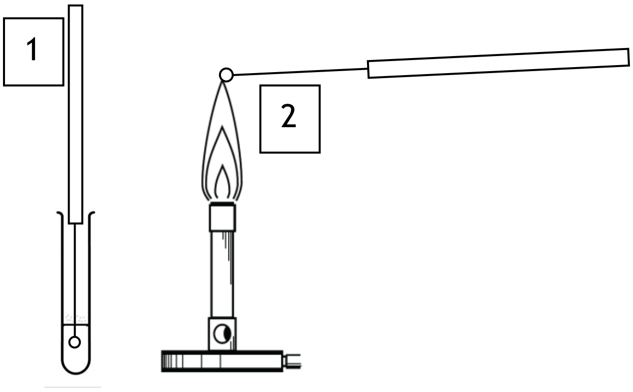


ammonium chloride + sodium hydroxide → sodium chloride + **ammonia** + water

Chemical Analysis – Flame tests

Method

1. Dip a clean wire loop in the sample solution
2. Hold the flame test wire loop at the edge of a Bunsen flame
3. Observe the changed colour of the flame, and decide which metal it indicates
4. Clean the loop in acid, rinse with water and repeat procedure with another sample



Metal	Ion	Flame test
Sodium	Na^+	Yellow-orange
Potassium	K^+	Lilac
Calcium	Ca^{2+}	Brick red
Copper	Cu^{2+}	Green


Sodium


Potassium


Calcium


Copper

Atomic Spectroscopy

This method is used to identify and show the amount (concentration) of specific atoms/ions present in the sample.

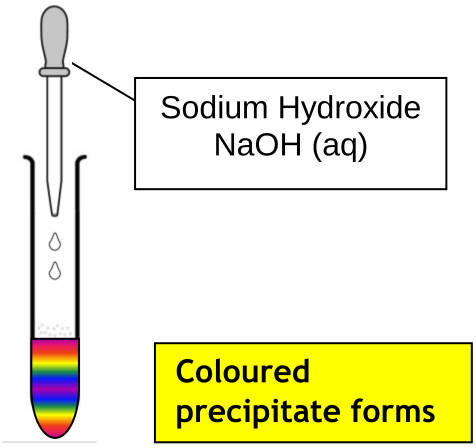
The colour of the light emitted during a flame test corresponds to a specific frequency

The intensity of the emission is measured – this corresponds to the amount of the metal present.

Chemical Analysis – Precipitation Reaction

- Transition metal hydroxides are insoluble in water.
- If a solution of any soluble transition metal compound is mixed with sodium hydroxide then we get a displacement reaction. The sodium is the more reactive metal, and displaces the transition metal from its compound.
- The transition metal hydroxide is formed as a result. As this is insoluble it appears as a solid in the liquid - this is called a **precipitate**

Add Sodium Hydroxide



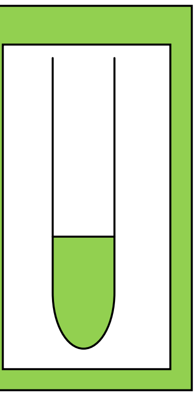
Iron (II) / Fe^{2+}

When we add sodium hydroxide to an iron (II) salt a **dirty green precipitate** is formed

iron (II) sulfate + sodium hydroxide ⟶ iron (II) hydroxide + sodium sulfate

$$FeSO_4 + 2NaOH \longrightarrow Fe(OH)_2 + Na_2SO_4$$

$$Fe^{2+} (aq) + 2OH^- (aq) \longrightarrow Fe(OH)_2 (s)$$



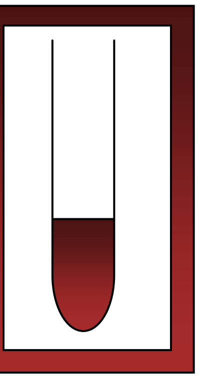
Iron (III) / Fe^{3+}

When we add sodium hydroxide to an iron (III) salt a **red brown precipitate** is formed

iron (III) nitrate + sodium hydroxide ⟶ iron (III) hydroxide + sodium nitrate

$$Fe(NO_3)_3 + 3NaOH \longrightarrow Fe(OH)_3 + 3NaNO_3$$

$$Fe^{2+} (aq) + 3OH^- (aq) \longrightarrow Fe(OH)_3 (s)$$



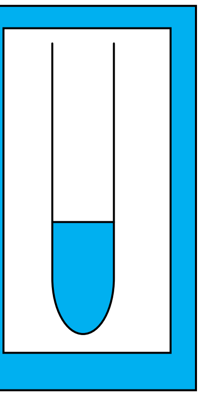
Copper (II) / Cu^{2+}

When we add sodium hydroxide to copper (II) salts a **blue precipitate** is formed

copper (II) sulfate + sodium hydroxide ⟶ copper (II) hydroxide + sodium sulfate

$$CuSO_4 + 2NaOH \longrightarrow Cu(OH)_2 + Na_2SO_4$$

$$Cu^{2+} (aq) + 2OH^- (aq) \longrightarrow Cu(OH)_2 (s)$$



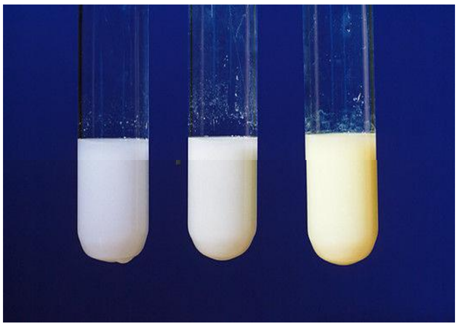
Chemical Analysis – negative ions

Testing for the halide ions

Add dilute nitric acid followed by silver nitrate

Nitric acid is added first to remove impurities. When silver nitrate is then added to a compound that contains a halide ion a precipitate is formed, the colour corresponds to the halide.

Non-metal	Ion	Silver Nitrate test
Chloride	Cl^-	white
Bromide	Br^-	cream
Iodide	I^-	yellow



Example equation

silver nitrate + sodium chloride → silver chloride + sodium nitrate

$AgNO_3\,(aq) + NaCl\,(aq) \longrightarrow AgCl\,(s) + NaNO_3\,(aq)$

ionic equation: $Ag^+\,(aq) + Cl^-\,(aq) \longrightarrow AgCl\,(s)$

Testing for a carbonate ion CO_3^{2-}

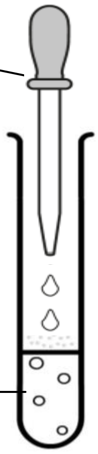
Add acid

When acid reacts with a carbonate **fizzing** is observed. Bubbles are of CO_2 gas which turns limewater milky

Example equations

sodium carbonate + hydrochloric acid → sodium chloride + water + carbon dioxide

$Na_2CO_3\,(aq) + 2HCl\,(aq) \longrightarrow 2NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$



bubbles / fizz

Testing for a sulfate ion SO_4^{2-}

Add dilute hydrochloric acid followed by barium chloride

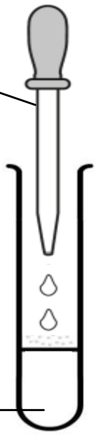
When barium chloride is added to a compound that contains a sulfate ion a white precipitate is seen - the white precipitate is barium sulfate

Example equations

barium chloride + sodium sulfate → sodium chloride + barium sulfate

$BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow 2NaCl\,(aq) + BaSO_4\,(s)$

$Ba^{2+}\,(aq) + SO_4^{2-}\,(aq) \longrightarrow BaSO_4\,(s)$

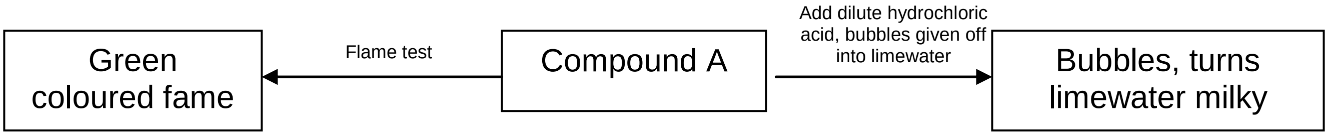


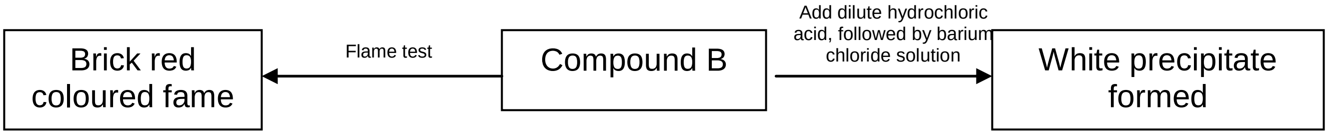
white precipitate forms

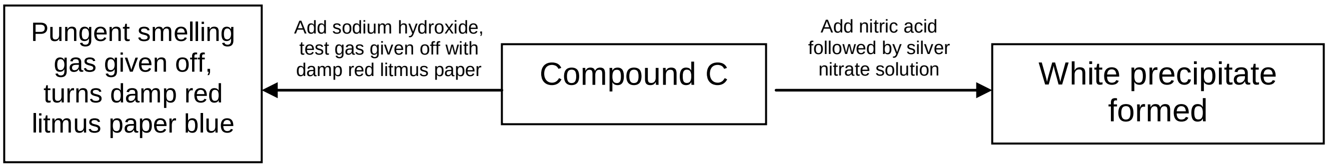
Chemical Analysis – Problem solving

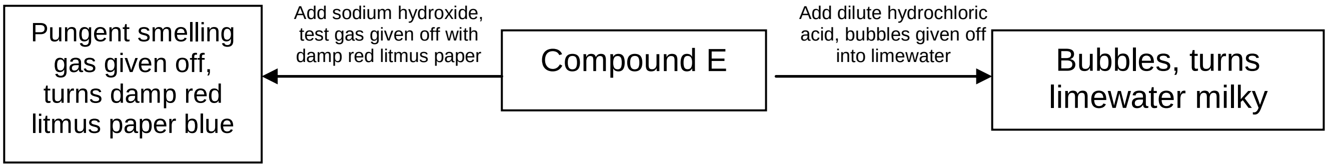
All the tests described so far can be used to identify unknown solutions. The flow charts below show tests carried out on compounds A, B and C and the results of those tests.

NH_4^+ Ca^{2+} Cu^{2+} K^+ Na^+ Cl^- Br^- I^- SO_4^{2-} CO_3^{2-}







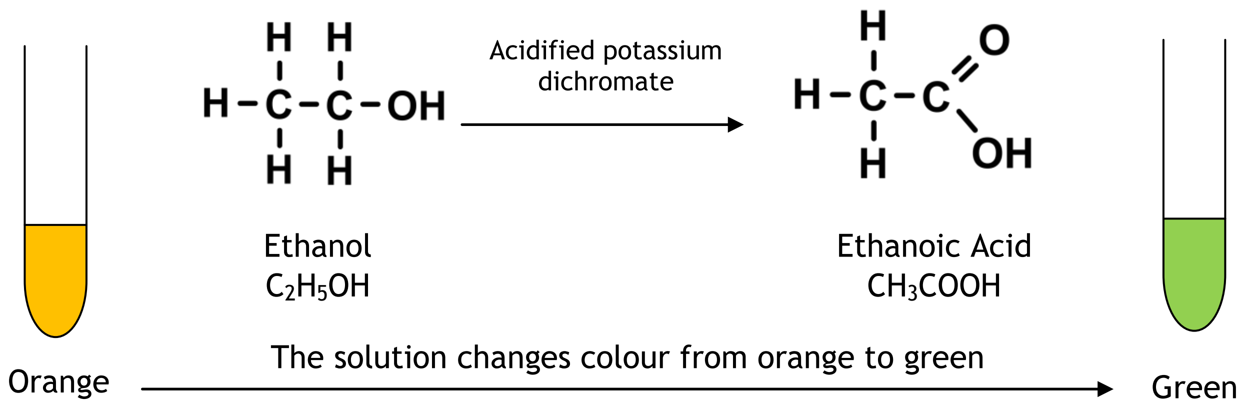


Answers: A is Copper carbonate, B is Calcium sulfate, C is Ammonium chloride, D is Sodium iodide, E is Ammonium carbonate

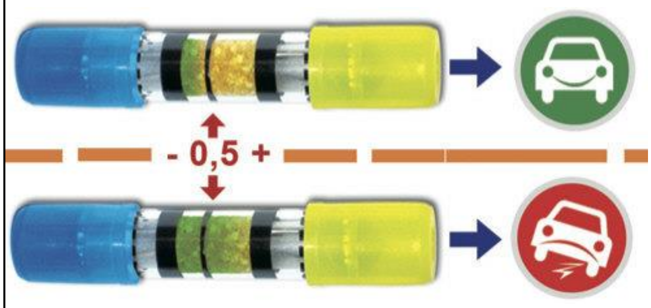
Chemical Tests for Organic Compounds

Oxidation of alcohols

Alcohols are oxidised to carboxylic acids with acidified potassium dichromate



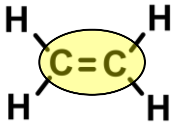
The early breathalyser



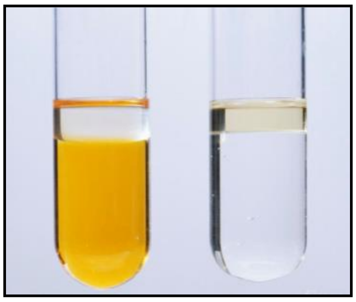
The oxidation reaction above was the basis of the early breathalyser. Tubes were used that contained orange dichromate crystals, the driver blew through the tube, if the driver had been drinking alcohol the crystals would turn green, the amount of crystals that were changed to green corresponded to the amount of alcohol in the drivers breath. More accurate techniques such as infrared are used in police stations today.

Testing for an alkene

Add bromine water



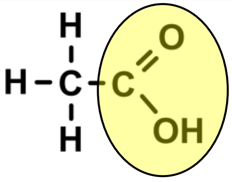
If a you think a compound has a C=C bond use the bromine water test



Orange bromine water turns colourless

Testing for a carboxylic acid

Add sodium carbonate (Na_2CO_3)



If a you think a compound has a -COOH functional group, add sodium carbonate, it will fizz and turn limewater milky

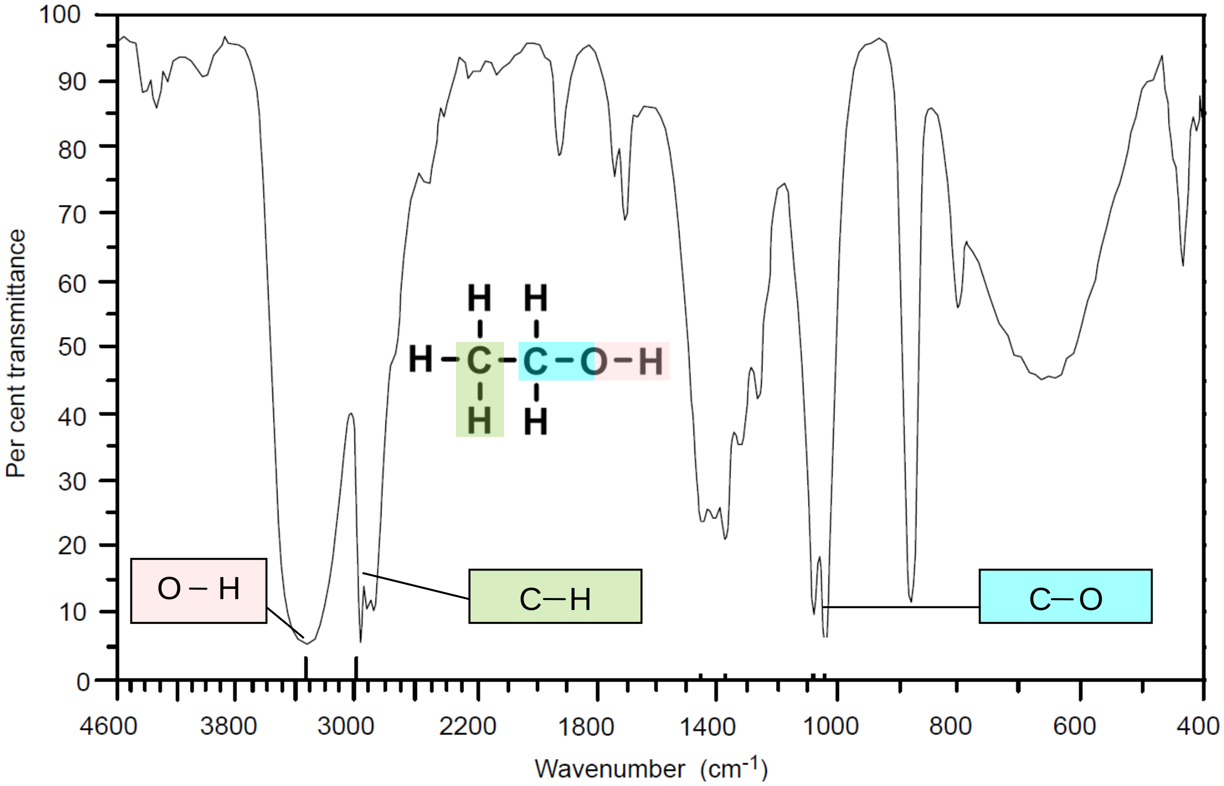
Effervescence - Bubbles of CO_2 produced which turns limewater milky

Infrared Spectroscopy

Infrared spectroscopy is used to identify the presence of certain bonds in organic molecules. All bonds in a molecule vibrate; different bonds will vibrate at different frequencies. The absorption values will be given in the exam, examples are shown below;

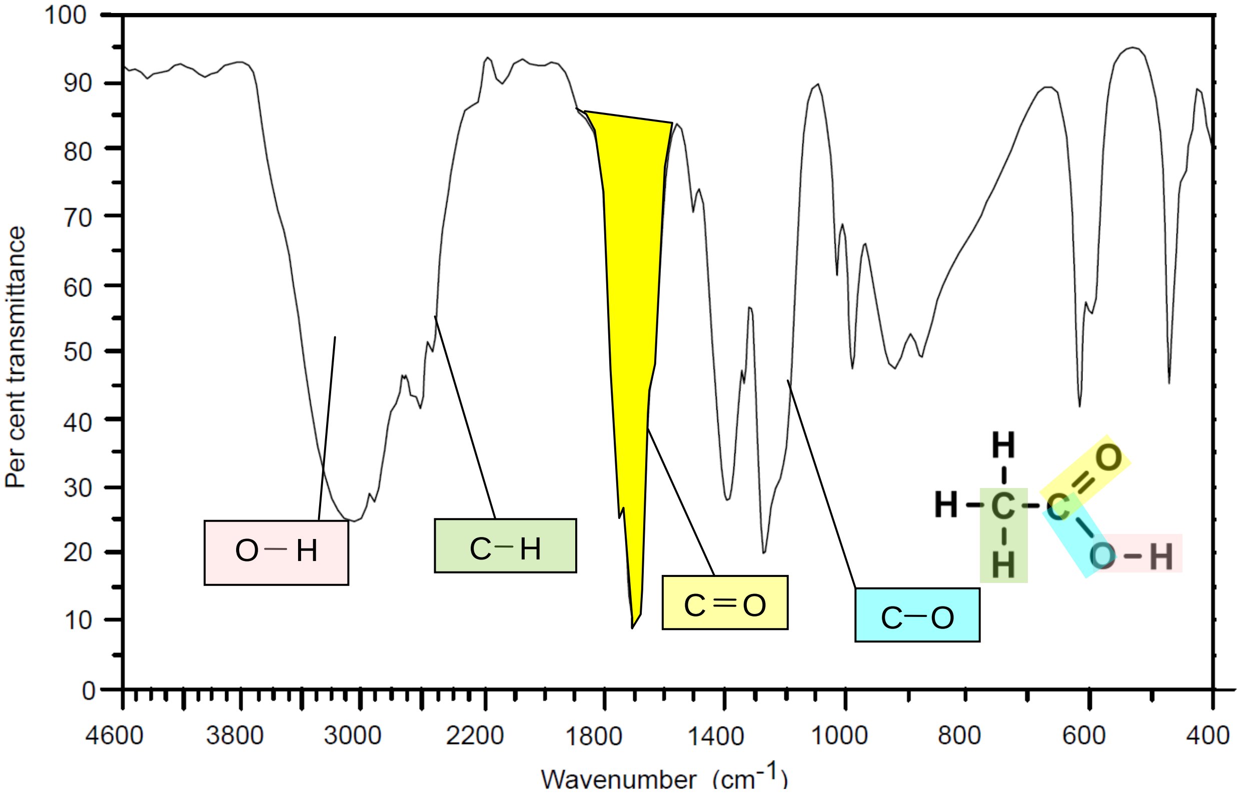
Infrared spectroscopy characteristic absorption values	
Bond	**Wavenumber / cm^{-1}**
C—O	1000 - 1300
C=C	1620 – 1670
C=O	1650 – 1750
C—H	2800 – 3100
O—H	2500 - 3550

The infrared spectrum below shows the characteristic frequencies for Ethanol



The infrared spectrum of ethanol

Infrared spectroscopy can be used as a tool to identify important functional groups in a molecule. It can also be used to check if reactions have been successful. Consider the reaction of ethanol with acidified potassium dichromate, in this reaction ethanol is oxidised to ethanoic acid. The infrared spectrum of the ethanoic acid produced is shown on the next page.



The infrared spectrum of ethanoic acid

The appearance of the peak at 1700 cm^{-1} (C=O) proves that ethanol has been oxidised. This was not present in the infrared spectrum of ethanol. This spectrum is typical of a carboxylic acid which has the C=O and -OH groups (-COOH).

Uses of Infrared Spectroscopy (IR)

As spectroscopic techniques are simple, fast and accurate they have replaced the need for chemical test in large laboratories, spectroscopy is an important tool in drug development for medicine and the development of new products in industry.

POSITIVE IONS		NEGATIVE IONS	
Name	**Formula**	**Name**	**Formula**
Aluminium	Al^{3+}	Bromide	Br^{-}
Ammonium	NH_4^{+}	Carbonate	CO_3^{2-}
Barium	Ba^{2+}	Chloride	Cl^{-}
Calcium	Ca^{2+}	Fluoride	F^{-}
Copper(II)	Cu^{2+}	Hydroxide	OH^{-}
Hydrogen	H^{+}	Iodide	I^{-}
Iron(II)	Fe^{2+}	Nitrate	NO_3^{-}
Iron(III)	Fe^{3+}	Oxide	O^{2-}
Lithium	Li^{+}	Sulfate	SO_4^{2-}
Magnesium	Mg^{2+}		
Nickel	Ni^{2+}		
Potassium	K^{+}		
Silver	Ag^{+}		
Sodium	Na^{+}		
Zinc	Zn^{2+}		

PERIODIC TABLE OF ELEMENTS

1	2						Group					3	4	5	6	7	0
								$^{1}_{1}\mathrm{H}$ Hydrogen									$^{4}_{2}\mathrm{He}$ Helium
$^{7}_{3}\mathrm{Li}$ Lithium	$^{9}_{4}\mathrm{Be}$ Beryllium											$^{11}_{5}\mathrm{B}$ Boron	$^{12}_{6}\mathrm{C}$ Carbon	$^{14}_{7}\mathrm{N}$ Nitrogen	$^{16}_{8}\mathrm{O}$ Oxygen	$^{19}_{9}\mathrm{F}$ Fluorine	$^{20}_{10}\mathrm{Ne}$ Neon
$^{23}_{11}\mathrm{Na}$ Sodium	$^{24}_{12}\mathrm{Mg}$ Magnesium											$^{27}_{13}\mathrm{Al}$ Aluminium	$^{28}_{14}\mathrm{Si}$ Silicon	$^{31}_{15}\mathrm{P}$ Phosphorus	$^{32}_{16}\mathrm{S}$ Sulfur	$^{35}_{17}\mathrm{Cl}$ Chlorine	$^{40}_{18}\mathrm{Ar}$ Argon
$^{39}_{19}\mathrm{K}$ Potassium	$^{40}_{20}\mathrm{Ca}$ Calcium	$^{45}_{21}\mathrm{Sc}$ Scandium	$^{48}_{22}\mathrm{Ti}$ Titanium	$^{51}_{23}\mathrm{V}$ Vanadium	$^{52}_{24}\mathrm{Cr}$ Chromium	$^{55}_{25}\mathrm{Mn}$ Manganese	$^{56}_{26}\mathrm{Fe}$ Iron	$^{59}_{27}\mathrm{Co}$ Cobalt	$^{59}_{28}\mathrm{Ni}$ Nickel	$^{64}_{29}\mathrm{Cu}$ Copper	$^{65}_{30}\mathrm{Zn}$ Zinc	$^{70}_{31}\mathrm{Ga}$ Gallium	$^{73}_{32}\mathrm{Ge}$ Germanium	$^{75}_{33}\mathrm{As}$ Arsenic	$^{79}_{34}\mathrm{Se}$ Selenium	$^{80}_{35}\mathrm{Br}$ Bromine	$^{84}_{36}\mathrm{Kr}$ Krypton
$^{86}_{37}\mathrm{Rb}$ Rubidium	$^{88}_{38}\mathrm{Sr}$ Strontium	$^{89}_{39}\mathrm{Y}$ Yttrium	$^{91}_{40}\mathrm{Zr}$ Zirconium	$^{93}_{41}\mathrm{Nb}$ Niobium	$^{96}_{42}\mathrm{Mo}$ Molybdenum	$^{99}_{43}\mathrm{Tc}$ Technetium	$^{101}_{44}\mathrm{Ru}$ Ruthenium	$^{103}_{45}\mathrm{Rh}$ Rhodium	$^{106}_{46}\mathrm{Pd}$ Palladium	$^{108}_{47}\mathrm{Ag}$ Silver	$^{112}_{48}\mathrm{Cd}$ Cadmium	$^{115}_{49}\mathrm{In}$ Indium	$^{119}_{50}\mathrm{Sn}$ Tin	$^{122}_{51}\mathrm{Sb}$ Antimony	$^{128}_{52}\mathrm{Te}$ Tellurium	$^{127}_{53}\mathrm{I}$ Iodine	$^{131}_{54}\mathrm{Xe}$ Xenon
$^{133}_{55}\mathrm{Cs}$ Caesium	$^{137}_{56}\mathrm{Ba}$ Barium	$^{139}_{57}\mathrm{La}$ Lanthanum	$^{179}_{72}\mathrm{Hf}$ Hafnium	$^{181}_{73}\mathrm{Ta}$ Tantalum	$^{184}_{74}\mathrm{W}$ Tungsten	$^{186}_{75}\mathrm{Re}$ Rhenium	$^{190}_{76}\mathrm{Os}$ Osmium	$^{192}_{77}\mathrm{Ir}$ Iridium	$^{195}_{78}\mathrm{Pt}$ Platinum	$^{197}_{79}\mathrm{Au}$ Gold	$^{201}_{80}\mathrm{Hg}$ Mercury	$^{204}_{81}\mathrm{Tl}$ Thallium	$^{207}_{82}\mathrm{Pb}$ Lead	$^{209}_{83}\mathrm{Bi}$ Bismuth	$^{210}_{84}\mathrm{Po}$ Polonium	$^{210}_{85}\mathrm{At}$ Astatine	$^{222}_{86}\mathrm{Rn}$ Radon
$^{223}_{87}\mathrm{Fr}$ Francium	$^{226}_{88}\mathrm{Ra}$ Radium	$^{227}_{89}\mathrm{Ac}$ Actinium															

Key: