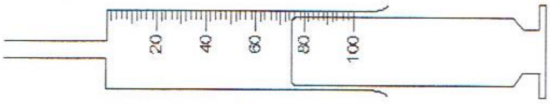
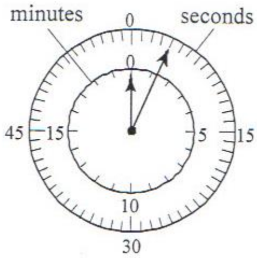
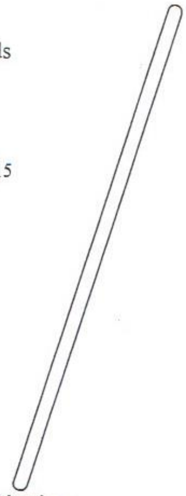
<u>Chemistry IGCSE Paper 6 revision guide</u>



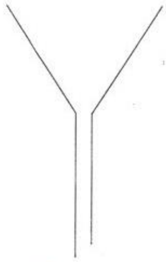
Gas syringe



Stop watch



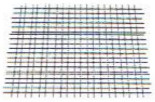
Stirring rod



Funnel

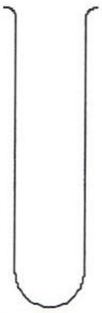
Top pan balance

Stopper cork

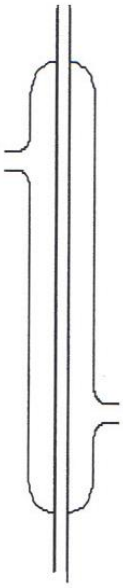


Gauze

Tripod

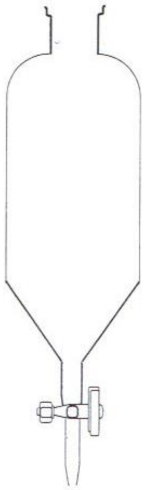


Test tube

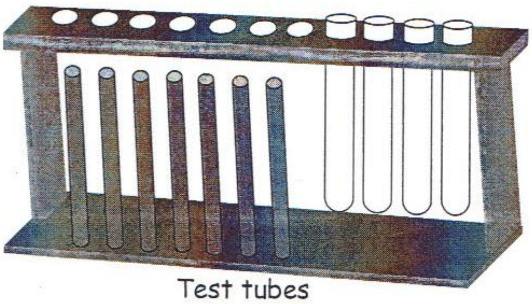


Condenser

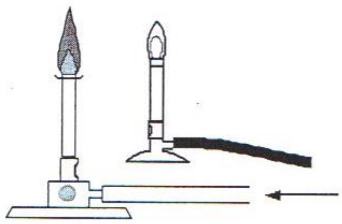
Thistle funnel



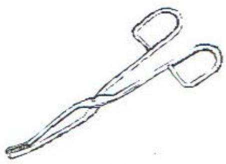
Separating funnel



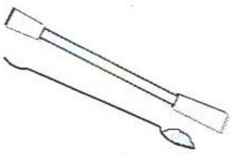
Test tubes **rack**



Bunsen burner



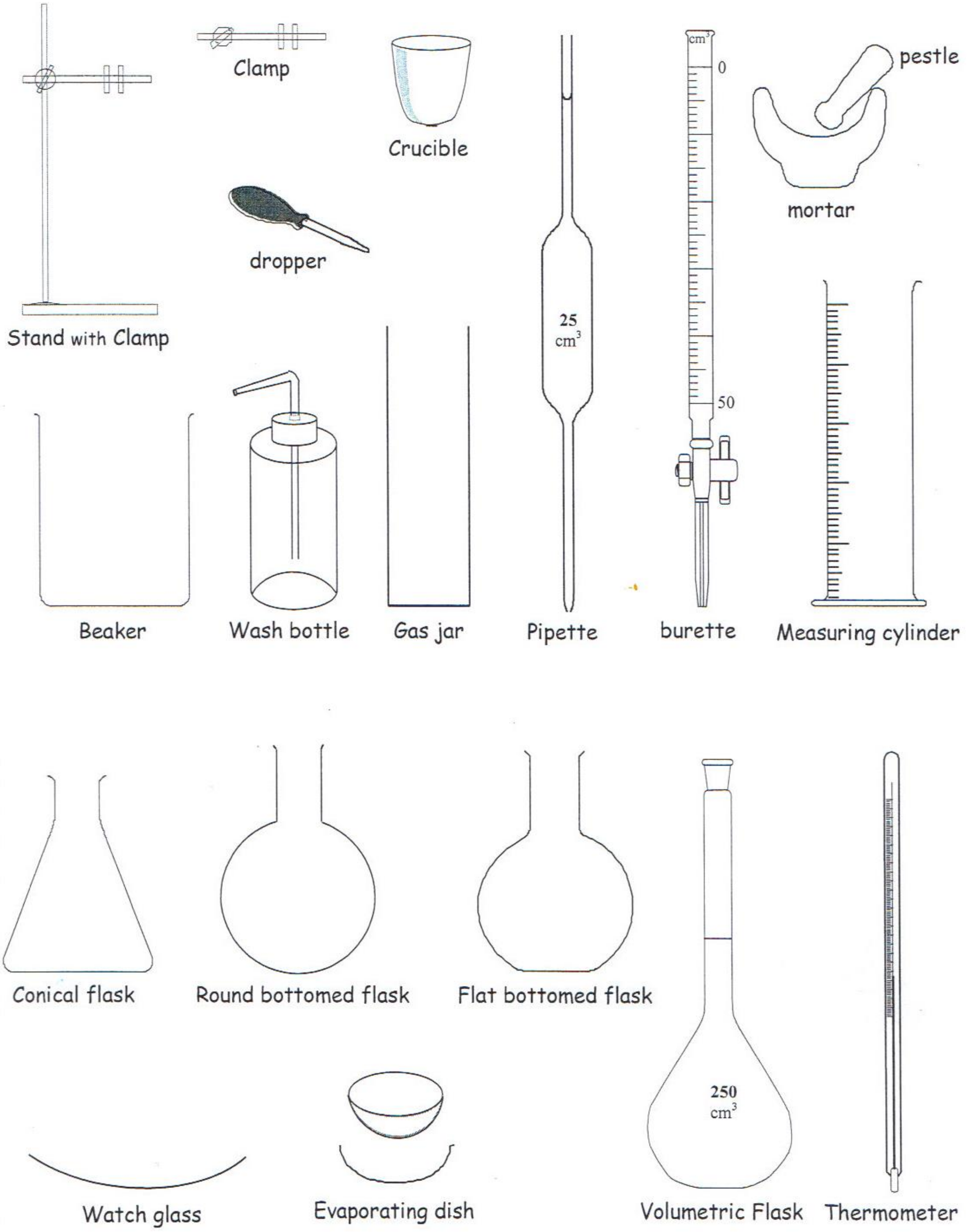
Metallic tong



Spatula

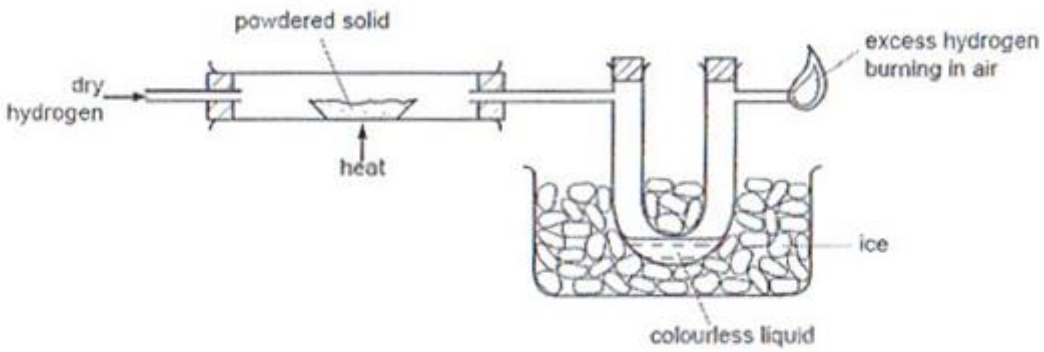
Knife

Spirit burner

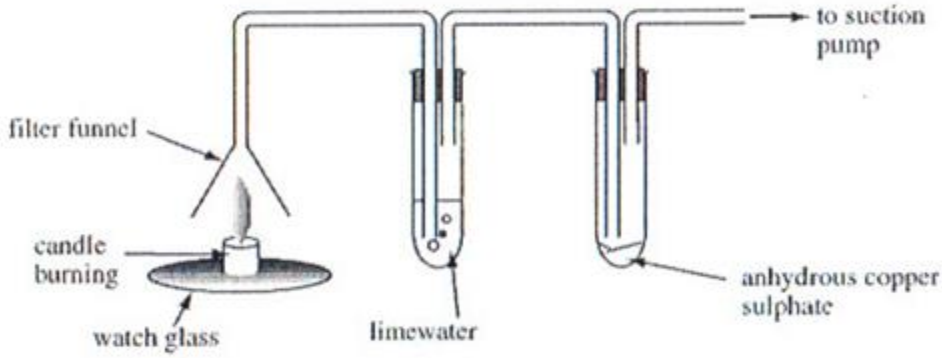


Clamp
Crucible
pestle
mortar
dropper
Stand with Clamp
cm³
0
25
cm³
50
Beaker
Wash bottle
Gas jar
Pipette
burette
Measuring cylinder
Conical flask
Round bottomed flask
Flat bottomed flask
250
cm³
Watch glass
Evaporating dish
Volumetric Flask
Thermometer

Reducing copper (II) oxide to copper



Testing the products of combustion

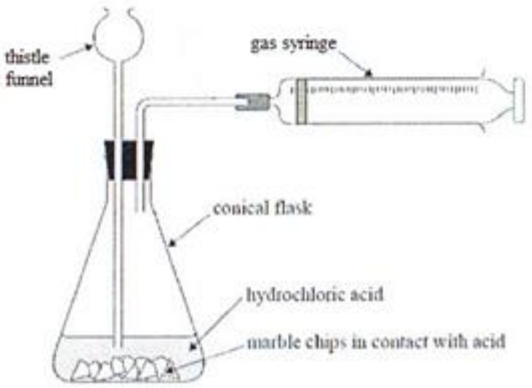


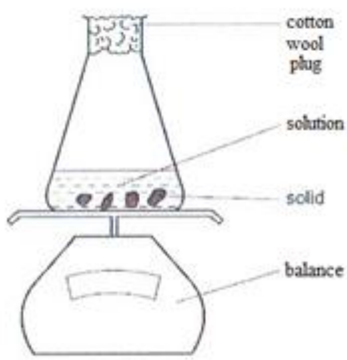
Testing the factors that affect the rate of reaction:

-different temperature acid

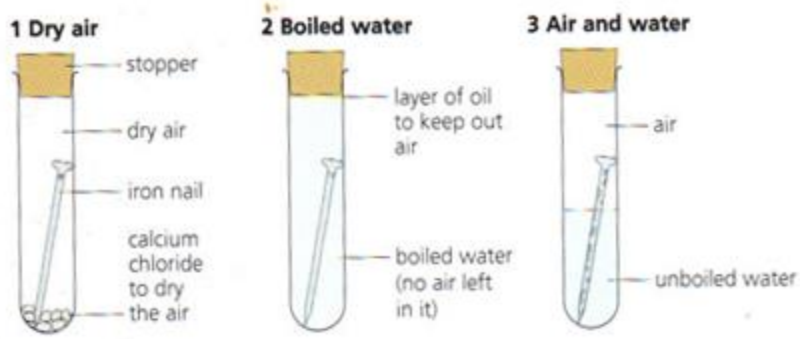
-different size of particles

-concentration of acid





Showing that iron needs both water and oxygen to rust:

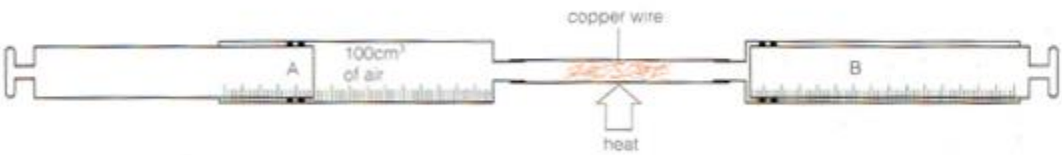


Nails 1 and 2 do not rust. Nail 3 does. The iron is oxidised, like this:

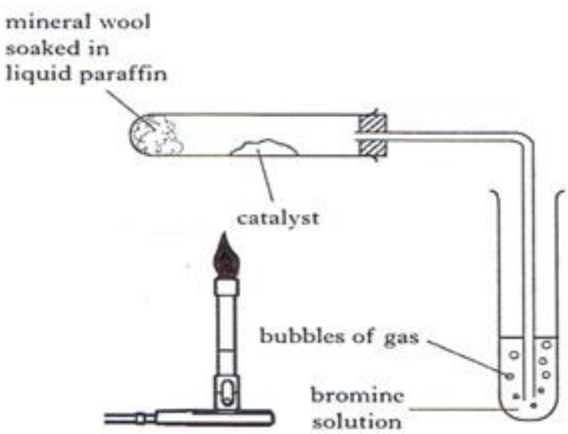
$$4Fe\,(s) + 3O_2\,(g) + 2H_2O\,(l) \longrightarrow 2Fe_2O_3.H_2O\,(s)$$

iron + oxygen + water ⟶ hydrated iron(III) oxide (or rust)

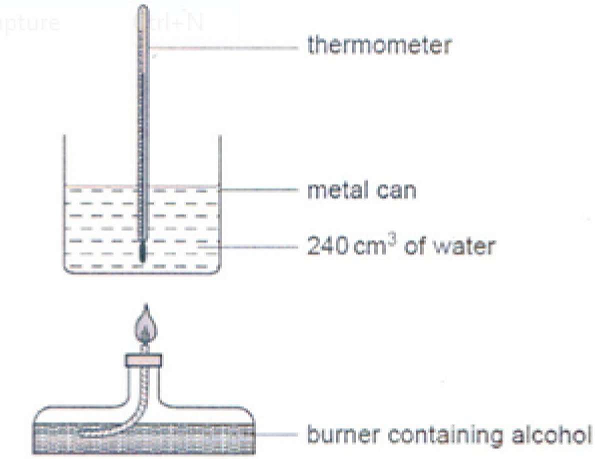
Showing that air is 21% oxygen:



Cracking:



Finding the amount of energy given when an alcohol is burnt:



You need to know:

1. Mass of water
2. Change in mass of burner containing alcohol
3. Specific heat capacity of water
4. Temperature change of water
5. The molecular mass of the alcohol

Then: change in mass / molecular mass = number of moles burnt

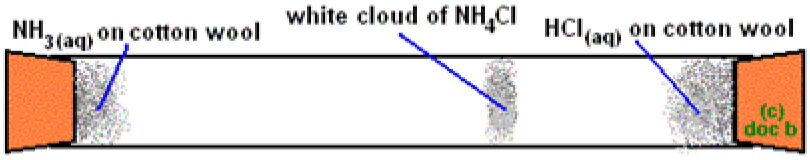
Change in temperature × mass of water × specific heat capacity of water = energy

Energy / moles burnt = amount of energy per mole (J/mol)

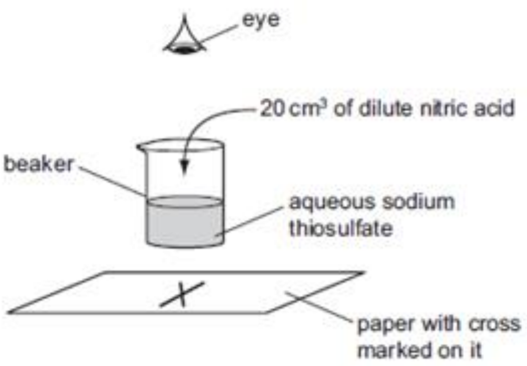
Sources of error: heat escapes into the air

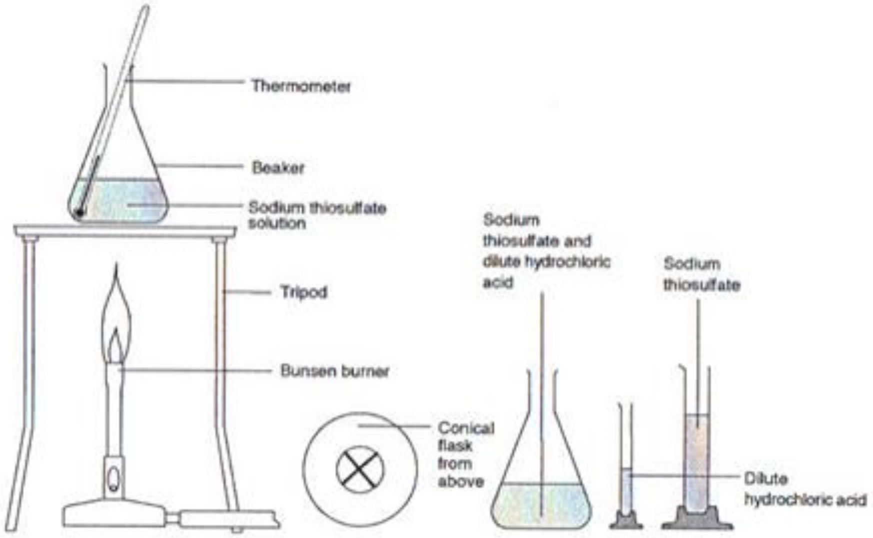
Demonstrating diffusion: potassium manganate(VII) in water & bromine with air

Demonstrating that a higher mass slows rate of diffusion:



Rates of reaction:

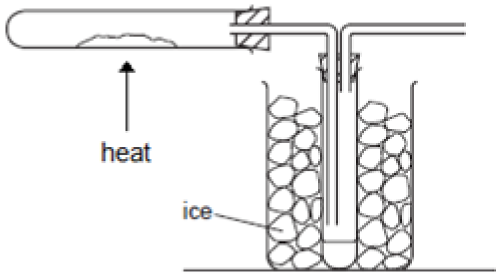




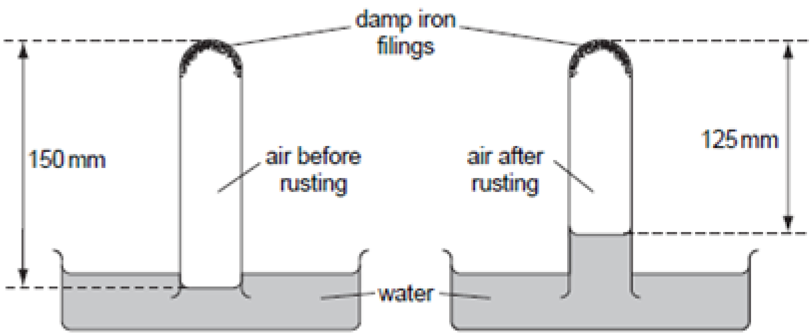
You can change the temperature and concentration used (not both at the same time though)

You need to keep the diameter of the conical flask the same, if it is thinner then the cross will disappear sooner, the cross also has to be the same, volumes too

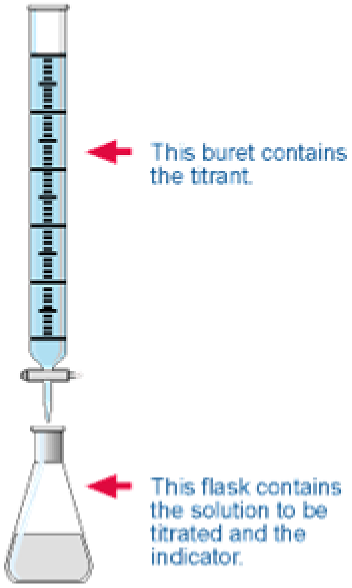
<u>Dehydrating copper (II) sulfate pentahydrate</u> (blue to colorless):



Percentage of oxygen in air v2.0? (from paper 6):



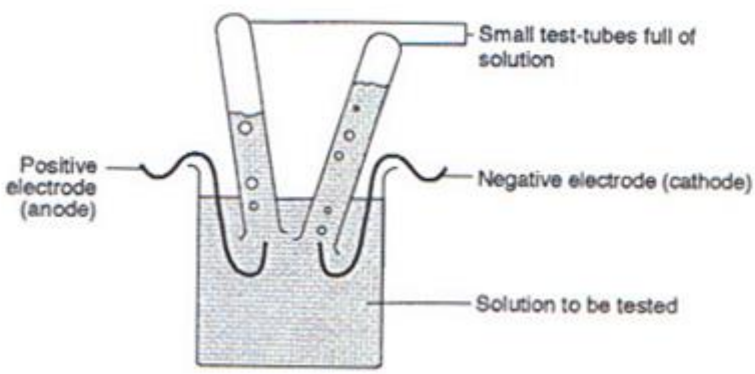
Acid & base titration to find the concentration of a solution:



To explain how it works, I'm going to say that it is an acid in the conical flask and a base in the burette:

1. Measure the volume of acid, then pour it into the conical flask
2. Add an indicator to it so that you will know when the mixture has become neutral
3. Record the initial volume of base in the burette
4. Slowly add the base (which you know the concentration of) from the burette, bit by bit, stirring each time
5. When the indicator shows it is neutral, record the final volume of base
6. Find the amount of base used by doing: final – initial
7. Do: volume × concentration, to find the number of moles of base used
8. Used the balanced equation to find out how many moles of the acid are needed to neutralize the number of moles of base used
9. Do: number of moles of acid needed / volume of acid used (measured in step 1)

Electrolysis of solutions:



(This set up is mostly if you want to test the gases, otherwise you can use the standard set up)

The limestone cycle:

1. Burn the limestone chip/ limestone powder in a test tube in a Bunsen flame → calcium oxide

2. Put calcium oxide in water → calcium hydroxide solution + left over limestone
3. Filter the leftover limestone
4. Blow bubbles using a straw into the calcium hydroxide solution → limestone

Flame colors: (I copied a table from the internet because it isn't in the book, the ones in red I'm pretty sure you have to know, and the ones in the yellow are the ones I think are good to know since they are used often).

Symbol	Element	Color
As	Arsenic	Blue
B	Boron	Bright green
Ba	Barium	Pale/Yellowish Green
Ca	Calcium	Orange to red
Cs	Cesium	Blue
Cu(I)	Copper(I)	Blue
Cu(II)	Copper(II) non-halide	Green
Cu(II)	Copper(II) halide	Blue-green
Fe	Iron	Gold
In	Indium	Blue
K	Potassium	Lilac to red
Li	Lithium	Magenta to carmine
Mg	Magnesium	Bright white
Mn(II)	Manganese(II)	Yellowish green
Mo	Molybdenum	Yellowish green
Na	Sodium	Intense yellow
P	Phosphorus	Pale bluish green
Pb	Lead	Blue
Rb	Rubidium	Red to purple-red
Sb	Antimony	Pale green
Se	Selenium	Azure blue
Sr	Strontium	Crimson
Te	Tellurium	Pale green
Tl	Thallium	Pure green
Zn	Zinc	Bluish green to whitish green

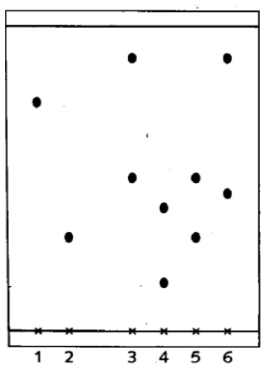
Sources of error for the flame test:

- The test cannot detect low concentrations of most ions.

- The brightness of the signal varies from one sample to another. For example, the yellow emission from sodium is much brighter than the red emission from the same amount of lithium.
- Impurities or contaminants affect the test results. Sodium, in particular, is present in most compounds and will color the flame. Sometimes a blue glass is used to filter out the yellow of sodium.
- The test cannot differentiate between all elements. Several metals produce the same flame color. Some compounds do not change the color of the flame at all.

Separation methods:

-**Paper chromatography**: (To separate substances) a drop of the substance is placed at the centre of a piece of filter paper and allowed to dry. Three or four more drops are added to it. Water is dripped on, drip by drip, so the ink spreads creating different coloured circles. Paper + rings = **chromatogram**. Rings are created because different substances travel at different rates.



(To identify substances) Spots of substances placed onto a pencilled line (as ink would separate) and labelled. Paper goes in solvent, and solvent travels up paper, then paper is taken out. There are spots which have travelled different distances.

-**Interpreting** simple chromatograms:

1. Number of rings/dots = number of substances
2. If two dots travel the same distance up the paper they are the same substance.
3. You can calculate the **R_f value** to identify a substance, given by the formula:

R_f value = distance moved by substance / distance moved by solvent

To make colourless substances you use a **locating agent**: 1. Dry paper in oven 2. Spray it with locating agent 3. Heat it for 10 minutes in oven.

The **stationary phase** is the material on which the separation takes place (e.g. the paper).

The **mobile phase** consists of the mixture you want to separate, dissolved in a solvent.

- Pure substances have a definite, sharp meting/boiling point; a substance + impurity has **lower melting point** and **higher boiling point**, at a range of temperatures; more impurity means bigger change.

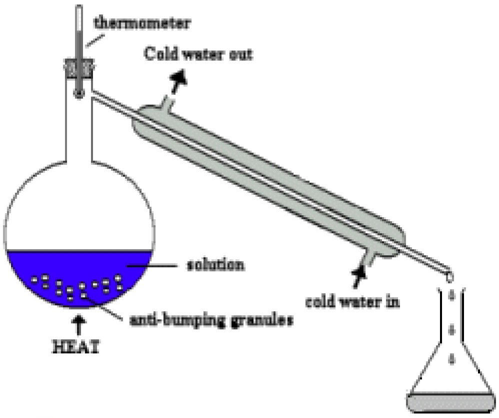
-**Filtration**: Mixture goes in a funnel with filter paper, into a flask. Residue is insoluble and stays at top. Filtrate goes through.

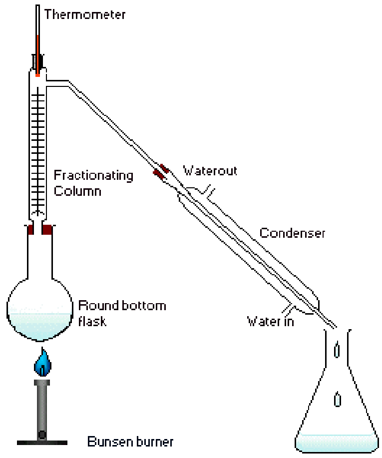
Distillation:

Simple distillation (left picture below) **evaporates** a solvent from a solution.

Fractional distillation (right picture below) removes a liquid from a mixture of liquids, because the liquids have different boiling points. Used to separate substances in crude oil and get ethanol from the products of fermentation.

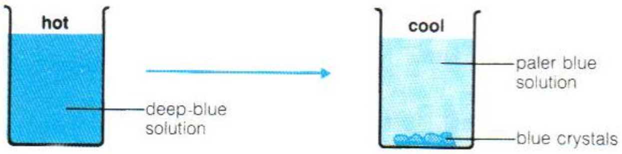
1. mixture is heated to evaporate the substance with the lowest boiling point
2. some of the other liquid(s) will evaporate too. A mixture of gases condense on the beads in the fractional column. So the beads are heated to the boiling point of the lowest substance in this case, so that the substance being removed cannot condense on the beads. The other substances continue to condense and will drip back into the flask.





4 By crystallising

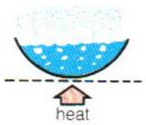
You can obtain many solids from their solutions by allowing them to form crystals. Copper(II) sulphate is an example:



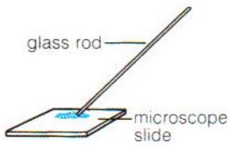
This is a saturated solution of copper(II) sulphate in water at 70 °C. If it is cooled to 20 °C . . .

. . . crystals begin to appear, because the compound is *less soluble* at 20 °C than at 70 °C.

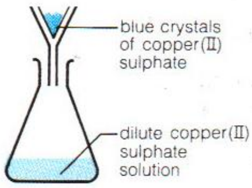
The process is called **crystallisation**. It is carried out like this:



1 A solution of copper(II) sulphate is heated, to get rid of some water. As the water evaporates, the solution becomes more concentrated.



2 The solution can be checked to see if it is ready by placing one drop on a microscope slide. Crystals should form quickly on the cool glass.



3 Then the solution is left to cool and crystallize. The crystals are removed by filtering, rinsed with water and dried with filter paper.

Using a suitable solvent:

Solvent	It dissolves
water	see “soluble salts” in section 8.3 and sugar
white spirit	gloss paint
propanone	grease, nail polish
ethanol	glues, printing inks, scented substances in perfumes and aftershaves

Choosing a suitable separation method:

Method of separation	Used to separate
filter	a solid from a liquid
evaporate	a solid from a solution
crystallise	a solid from a solution
distil	a solvent from a solution
fractional distillation	liquids from each other
chromatography	different substances from a solution

Percent yield = amount of wanted substance / total amount of substance
More colors:
-potassium iodide will go from colourless to red-brown, it is a reducing agent, it therefore is oxidised to produce I_2
$H_2O_2 + 2KI + H_2SO_4 \rightarrow I_2 + K_2SO_4 + 2H_2O$
$2I^-$ (colourless) $\rightarrow I_2$ (red-brown)
-Potassium mangenate, an oxidising agent, will go from purple to colourless. (Recall of equations involving KMnO4 is **not** required.)
-potassium dichromate, an oxidising agent: $Cr_2O_7^{2-}$ (orange) $\rightarrow 2Cr^{3+}$ (green)
<u>Preparing salts in the lab</u> (all on pages 124 and 125)
Starting with a metal:

1. Add excess metal to an acid
2. When bubbling (hydrogen) stops the reaction is done
3. Filter off excess metal (because a metal is a solid, except mercury)

Starting with an insoluble base:

1. Add insoluble base to acid and heat gently, it will dissolve
2. Keep adding until no more dissolves (reaction is done)
3. Filter out the insoluble (excess) base

(**Titration**): Starting with an **alkali** (**soluble base**):

1. Put a certain amount (e.g. $25cm^3$) alkali in a flask
2. Add phenolphthalein (pink in alkaline, colourless in acid or neutral) you could use a different indicator but this is quite simple
3. Add acid from a burette, slowly while stirring, until it goes colourless
4. Find out how much acid you used (using the scale on the burette.
5. Repeat, this time add the same amount of base, but you know exactly how much acid to add to get a neutral solution, don't add indicator though (you don't need it anymore, and it would make it impure)
6. Evaporate the water from the neutral solution using a Bunsen flame and an evaporating dish

Add acid to base in a beaker
-if the salt is <u>insoluble</u> then filter it
-if is <u>soluble</u> then evaporate using an evaporating dish sat on top of a beaker of water on top of a Bunsen flame. When crystals start to form around the edge stop heating and leave to cool for a few days to form crystals, then filter crystals.
If they don't tell you if the salt is soluble or not then you need to know this:
<u>Soluble salts are:</u>
-all potassium, sodium and ammonium salts
-all nitrates
-all halides except silver and lead
-all sulphates except calcium, lead and barium
<u>Insoluble salts are:</u>
-carbonates except potassium, sodium and ammonium
-silver and lead halides
-calcium, lead and barium sulphates

Aqueous cations:

Test	Ion	Result
Add a few drops of dilute sodium hydroxide solution. A precipitate will form.	Cu^{2+}	Pale blue precipitate
	Fe^{2+}	Green precipitate
	Fe^{3+}	Red-Brown precipitate
	Al^{3+}, Zn^{2+}, Ca^{2+}	White precipitate
Divide the solutions into two equal volumes. To one, add double the volume of sodium hydroxide solution. To the other, add double the volume of ammonium hydroxide.	Al^{3+}	The precipitate dissolves again in sodium hydroxide solution giving a colourless solution.
	Zn^{2+}	Precipitate dissolves in both solutions, giving a colourless solution.
	Ca^{2+}	Dissolves in neither.
Take a small amount of the solid or solution. Add a little dilute sodium hydroxide solution and heat gently.	NH_4^+	Ammonia gas given off (it has a strong sharp smell and turns red litmus blue)

Anions:

Test	Ion	Result
Take a small amount of the solid/solution. Add a little dilute hydrochloric acid.	Carbonate	Carbon dioxide gas is produced so there are bubbles and limewater goes milky
Take a small amount of the solution. Add an equal volume of dilute nitric acid. Then add silver nitrate solution. Silver halides are insoluble, so a precipitate forms.	Chloride	White precipitate
	Iodide	Yellow
Take a small amount of the solution. Add a little sodium hydroxide solution. Add aluminium foil and heat gently.	Nitrate	Ammonia gas given off
Take a small amount of the solution. Add an equal volume of dilute hydrochloric acid. Then add barium nitrate solution. Barium sulphate is insoluble so...	Sulphate	White precipitate (barium sulphate)

Gases:

Test	Gas	Result
Smell gas, use indicator paper.	Ammonia	Recognizable odour, indicator paper turns blue
It is a weak acid, so reacts	Carbon Dioxide	white precipitate, solution

with calcium hydroxide (lime water) to form insoluble calcium carbonate		goes milky
Green poisonous gas which bleaches dyes. Hold damp indicator paper in the gas (in a fume cupboard).	Chlorine	Indicator paper turns white.
Collect in a tube then put a lighted splint to it.	Hydrogen	Squeaky pop
Collect in a tube then put a glowing splint to it.	Oxygen	Splint immediately bursts into a flame

Other tests:

Test	Substance	Result
Add some drops of the substance to white anhydrous copper (II) sulphate	water	White crystals go blue
Add to blue cobalt chloride paper		Goes pink
Test for saturation: add bromine water	Alkene	Goes from orange to colorless
	alkane	Stays orange
Blue litmus paper	acid	Goes red
Add a metal carbonates		Effervescence (CO_2)
Red litmus paper	base	Goes blue

Observable properties

For the following properties, I put the ones you might observe in the lab (i.e. brittle green crystal but not diatomic molecule since you can't see that).

Alkali metals properties that might be on paper 6

-React violently with chlorine
-Burst into flames when heated with oxygen:
    -a red flame for lithium
    -a yellow flame for sodium
    -a lilac flame for potassium
-Produce soluble white compounds.
-React with cold water
-Conducts electricity

Halogens, group 7 – physical properties:
-form coloured gases
-are poisonous
-are brittle and crumbly in solid form
-do not conduct electricity

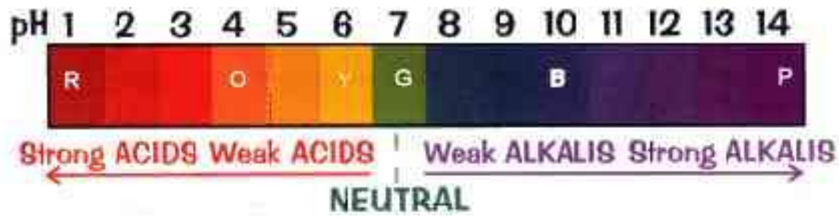
Transition metals
-hard, tough and strong
-high melting points (except mercury)
-malleable (can be hammered into different shapes) and ductile (can be drawn out into wires)
-good conductors of heat and electricity (silver is the best)

-high density (greater than water's density)
-Much less reactive than group one metals, except for iron which rusts easily
-Have no trend in reactivity
-Mostly form coloured compounds
-Can form several differently charged ions: have variable valency, therefore.....
-they can form more than one compound with another element
-can be used as catalysts
-can form complex ions

Indicators

Indicator	Colour in acid	Colour in alkali
Litmus	Red	Blue
phenolphthalein	Colourless	Pink
Methyl orange	red	yellow

Universal indicator can be used as a solution or paper strip:

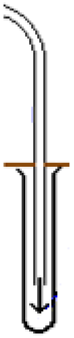
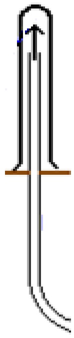
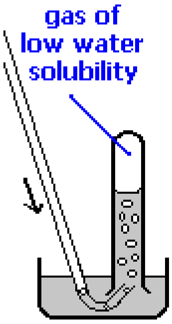
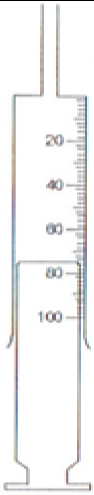


Preparing gases in the lab:

To make....	Place in flask:	Add....	Reaction
CO_2	$CaCO_3$ (marble chips)	Dilute HCl	$CaCO_3(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
Cl_2	Manganese (IV) oxide (as an oxidising agent)	Conc. HCl	$2HCL(aq) + [O] \rightarrow H_2O(l) + Cl_2(g)$
H_2	Pieces of zinc	Dilute HCl	$Zn(s) + SHCL(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
O_2	Manganese (IV) oxide (as a catalyst)	Hydrogen peroxide	$2\ H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

Collecting gases:

Method	Upward displacement of air	Downward displacement of air	Over water	Gas syringe
Use when...	Gas more dense than air	Gas less dense than air	Gas is sparingly soluble in water	To measure the volume

Apparatus				
Examples	Carbon dioxide, chlorine, sulphur dioxide, hydrogen chloride	Ammonia, hydrogen	Carbon dioxide, hydrogen, oxygen	Any gas