Edexcel
IGCSE

Chemistry

CODE: (0620)

Unit 2

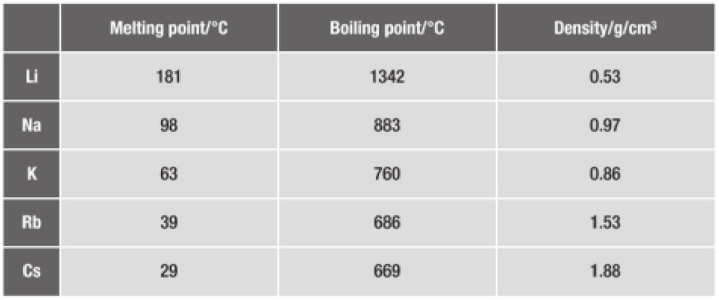
Inorganic Chemistry

2.1 The alkali metals

Group 1 of the Periodic Table contains **alkali metals**, including **radioactive** element francium. Its isotopes are produced during uranium-235 decay, with only 20-30g present in Earth's crust. Understanding the rest of Group 1 allows for predictions about francium's characteristics.

Physical properties

Metals have low melting and boiling points, increasing in densities as they move down the group. They are soft, easily cut, and shiny when freshly cut, but tarnish quickly in air exposure.



	Melting point/°C	Boiling point/°C	Density/g/cm³
Li	181	1342	0.53
Na	98	883	0.97
K	63	760	0.86
Rb	39	686	1.53
Cs	29	669	1.88

Storage and handling

Group 1 metals are highly reactive, forming oxides and alkaline solutions with oxygen and water. To prevent reactions, lithium, sodium, and potassium are stored under oil, while rubidium and caesium are stored in sealed glass tubes. It's crucial not to touch these metals with bare fingers, as sweat can cause reactions and **corrosive** metal hydroxides. It's essential to avoid contact with these metals to prevent potential health risks.

A family of elements

The chemical properties depend on the number of electrons in the outer shell. The Group 1 elements react in very similar ways because they all have the same number of electrons in the outer shell (one), so reason 2 is just a consequence of reason 1.

Reactions with water

All these metals react in the same way with water to produce a metal hydroxide and hydrogen:

alkali metal + water → alkali metal hydroxide + hydrogen

$$2M + 2H_2O \rightarrow 2MOH + H_2$$

The main difference between the reactions is how quickly they happen.

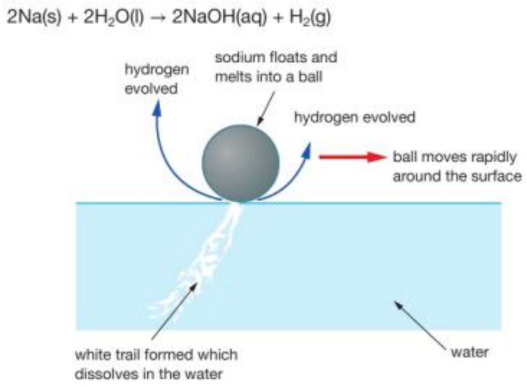
As you go down the group, the metals become more reactive, and the reactions occur more rapidly.
The reaction between sodium and water is typical.

WITH SODIUM

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

The main observations you can make when this reaction occurs are:

■The sodium floats because it is less dense than water.
■The sodium melts into a ball because its melting point is low, and a lot of heat is produced by the reaction.
■There is fizzing because hydrogen gas is produced.
■The sodium moves around on the surface of the water.
■The piece of sodium gets smaller and eventually disappears. The sodium is used up in the reaction.
■If you test the solution that is formed with universal indicator solution, you will see that the universal indicator goes blue, indicating an alkaline solution has been formed.



▲ Figure 11.5 Sodium reacting with water. The white trail is the sodium hydroxide, which dissolves in water to form a strongly alkaline solution.

LITHIUM

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

The reaction is very similar to sodium's reaction, except that it is slower. Lithium's melting point is higher, and the heat isn't produced so quickly, so the lithium doesn't melt.

POTASSIUM

$2K(s) + 2H_2O(l) \rightarrow 2KOH\ (aq) + H_2(g)$

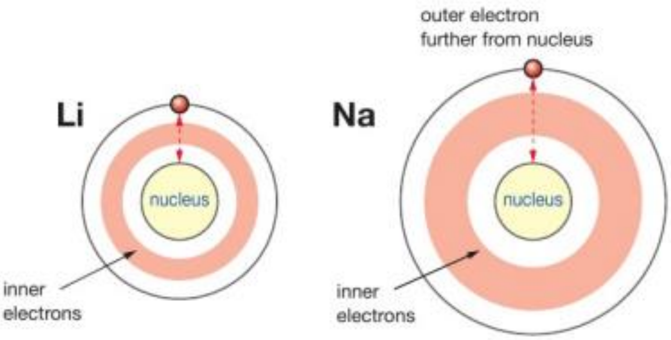
Potassium's reaction is faster than sodium's. Enough heat is produced to ignite the hydrogen, which burns with a lilac flame. The reaction often ends with the potassium spitting around and exploding.

RUBIDIUM AND CAESIUM

These react even more violently than potassium, and the reaction can be explosive. Rubidium hydroxide and caesium hydroxide are formed.

Explaining the increase in reactivity

The differences in reactions are based on the loss of the outer electron of a metal, which is attracted to the nucleus in the original atom. As atoms grow larger, the outer electron is further from the nucleus, making it less strongly attracted and more easily lost. This is because the nucleus attracts negative electrons, and the atom's size increases.



▲ Figure 11.7 Electrons of lithium and sodium

Reactions of the alkali metals with air

Lithium, sodium, and potassium are stored in oil due to their reactions with air. When exposed to oxygen, sodium tarnishes rapidly, while lithium tarnishes slower due to its slower reaction. Potassium tarnishes more rapidly than sodium, indicating its higher reactivity. Each metal reacts with oxygen in the air to form an oxide with the formula M2O. When heated using a Bunsen burner, the reactions become more vigorous, making it difficult to determine the most reactive metal. Lithium forms lithium oxide with a red flame, sodium forms sodium oxide with a yellow flame, and potassium forms potassium oxide with a lilac flame. The equation for these reactions is,

$4M(s) + O_2(g) \rightarrow 2M_2O(s)$.

Compounds of the alkali metals

Group 1 metal ions are colorless, unless combined with a negative ion. Their compounds are typically ionic solids, mostly soluble in water.

Summarizing the main features of the group 1 elements

Group 1 elements:

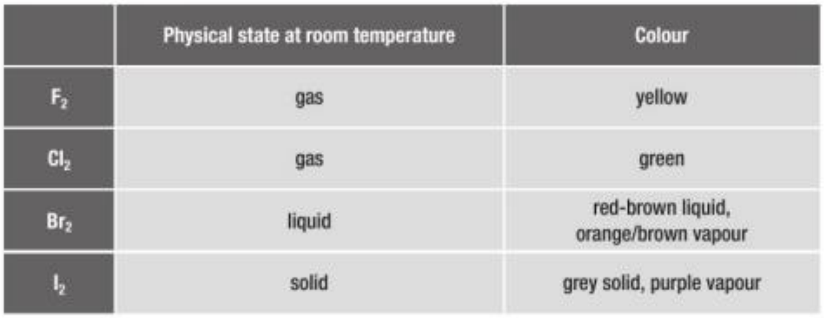
- Are metals
- Are soft with melting points and densities which are very low for metals
- Must be stored out of contact with air or water
- React rapidly with air to form coatings of the metal oxide
- React with water to produce an alkaline solution of the metal hydroxide and hydrogen gas
- Increase in reactivity as you go down the group
- Form compounds in which the metal has a 1^+ ion
- Have mainly white/colourless compounds which dissolve to produce colourless solutions.

Predicting the properties of francium

We can predict that francium:
- Is very soft
- Will have a melting point around room temperature
- Has density which is probably just over 2 g/cm3
- Will be a silvery metal, but will tarnish almost instantly in air
- Will react violently with water to give francium hydroxide and hydrogen
- Will be more reactive than caesium
- Will have a hydroxide, francium hydroxide, with the formula FrOH, which will be soluble in water and form a strongly alkaline solution
- Will form compounds that are white/colourless and dissolve in water to give colourless solutions.

2.2 The halogens

Table 12.1 The melting and boiling points of the halogens increase and the colour becomes darker down the group



	Physical state at room temperature	Colour
F_2	gas	yellow
Cl_2	gas	green
Br_2	liquid	red-brown liquid, orange/brown vapour
I_2	solid	grey solid, purple vapour

The halogens

Halogens, such as fluorine, chlorine, bromine, iodine, and astatine, are non-metallic elements with diatomic molecules. They are part of Group 7 in the Periodic Table, known for their salt-producing properties. These elements react similarly with metals, producing various salts containing halogen ions.

Safety

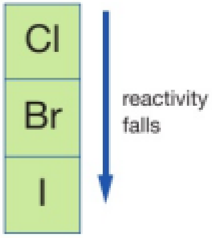
Fluorine and chlorine react in school labs, causing dangerously reactive vapors. Halogens have poisonous **vapours,** necessitating fume cupboard handling. Liquid bromine is corrosive, requiring careful handling.

Reactions of the halogens

The halogens react with hydrogen to form **hydrogen halides.**

Displacement reactions involving the halogens.

Displacement reactions involve the reaction of a halogen solution with a solution containing halide ions, demonstrating that chlorine is more reactive than bromine and iodine.



▲ Figure 12.5 Reactivity decreases down the group.

Reacting chlorine with potassium bromide or potassium Iodide solutions

If you add chlorine solution ('chlorine water') to colourless potassium bromide solution, the solution becomes orange as bromine is formed:

$$2KBr(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + Br_2(aq)$$

potassium bromide + chlorine → potassium chloride + bromine

Similarly, adding chlorine solution to potassium iodide solution gives a brown (orange if it is very dilute) solution of iodine.

$$2KI(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + I_2(aq)$$

In the same way, the more reactive bromine displaces the less reactive iodine from potassium iodide solution. Adding bromine solution ('bromine water') to colourless potassium iodide solution gives a brown solution of iodine:

$$2KI(aq) + Br_2(aq) \rightarrow 2KBr(aq) + I_2(aq)$$

Ionic equation for displacement reactions

The **spectator ion**, a metal ion, does not affect reactions like chlorine with sodium or magnesium bromide, unlike the metal ion in the chlorine-potassium iodide reaction.

Displacement reactions as redox reactions

If we look at the reaction between chlorine and either bromide ions or iodide ions, we can see why they are described as redox reactions.

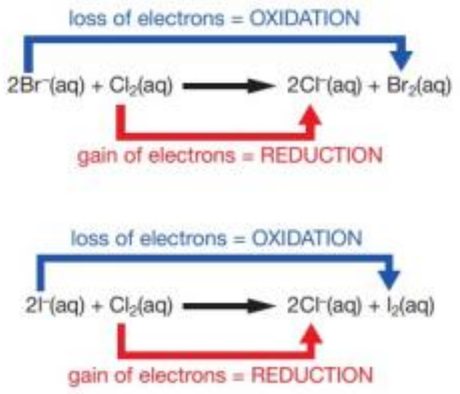


Figure 12.8 The reaction of chlorine with potassium bromide (top) or potassium iodide (below).

In the reaction between bromine and iodide ions, the bromine is the oxidising agent:

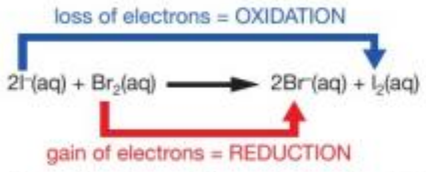
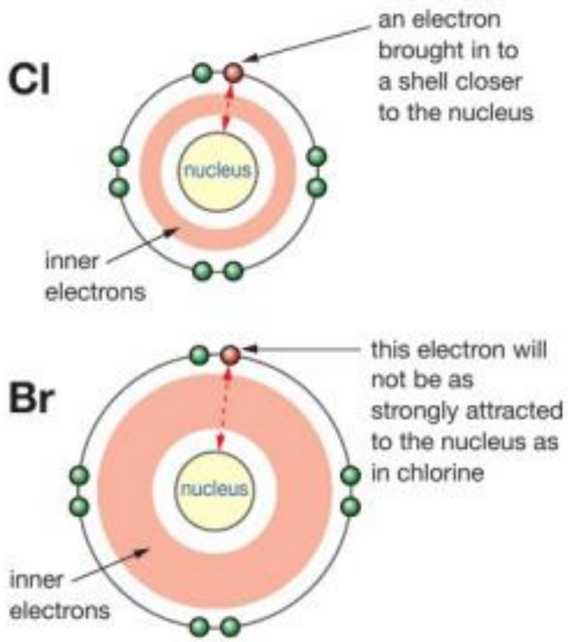


Figure 12.9 The reaction between bromine solution and potassium iodide solution.

Explaining the trend in the reactivity of the halogens

Halogens' reactivity is attributed to their ability to form negative ions. When reacting, they form 1-ions by taking electrons away from another atom. Chlorine, being smaller than bromine, has a stronger tendency to form 1-ions due to its proximity to the nucleus. On the other hand, bromine is larger and has a weaker attraction to the nucleus, resulting in a stronger tendency to form 1-ions.



▲ Figure 12.10 Chlorine has a stronger attraction for an electron than bromine.

2.3 Gases in the atmosphere

The composition of the air

The approximate percentages (by volume) of the four most abundant gases present in unpolluted, dry air are shown in Table 13.1.

Gas	Amount in air (%)	Amount in air (fraction)
nitrogen	78.1	about 4/5
oxygen	21.0	about 1/5
argon	0.9	
carbon dioxide	0.04	

The combustion of elements of oxygen

Burning magnesium

Magnesium burns in oxygen with an extremely bright white flame to give a white, powdery ash of magnesium oxide:

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

The white powder formed is not very soluble in water, but a very small amount does dissolve to form an alkaline solution:

$$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$$

Buring sulfur

Sulfur burns in oxygen with a blue flame. Poisonous, colourless sulfur dioxide gas is produced.

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

The sulfur dioxide dissolves in water to form an acidic solution of sulfurous acid:

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

Burning hydrogen

Hydrogen burns in oxygen with a pale blue flame. The product is water:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

The properties of oxides

We can make some generalisations about the properties of oxides formed when elements burn in oxygen.

Meatal oxides

- Metal oxides are ionic compounds containing O2- ions.
- Metal oxides are usually basic oxides, which means that they react with acids to form salts.
- Metal oxides are usually insoluble in water. Those metal oxides that are soluble in water react with it to form alkaline solutions containing hydroxide (OH-) ions.

Nonmetal oxides

- Non-metal oxides are covalent compounds.
- Non-metal oxides are usually acidic oxides, which react with alkalis/bases to form salts.
- Non-metal oxides are often soluble in water and react with it to form acidic solutions containing hydrogen (H+) ions.

Carbon dioxide

Carbon dioxide is a colourless gas that is most easily made in the laboratory by the reaction between dilute hydrochloric acid and calcium carbonate in the form of marble chips:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Carbon dioxide can also be obtained when metal carbonates are heated strongly.

Most carbonates split to give the metal oxide and carbon dioxide when you heat them. This is an example of **thermal decomposition,** breaking up something by heating it.

carbon dioxide and global warming: the greenhouse effect

Carbon dioxide is produced when fossil fuels (coal, oil, and gas) burn. For instance, when coal, which is mostly carbon, burns in excess oxygen:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

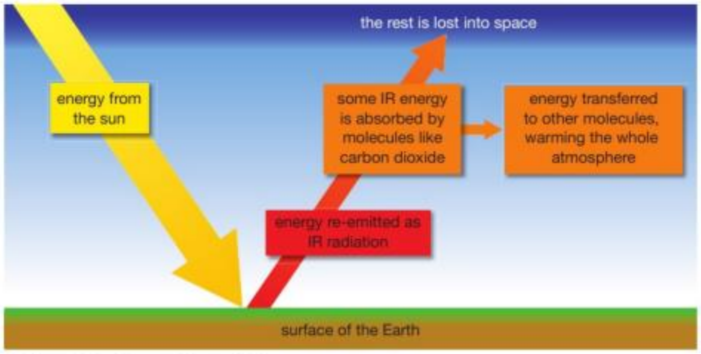
Petrol is a mixture containing many different hydrocarbons. An example of a reaction that occurs when petrol burns is:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$$

The greenhouse effect occurs when high-energy **UV** and visible light from the sun warm up Earth's surface, absorbing infrared (IR) radiation and releasing it in all directions. Carbon dioxide, a **greenhouse gas**, has been increasing in the atmosphere for around 200 years due to the industrial revolution, burning fossil fuels, and deforestation. This increase in carbon dioxide levels is believed to contribute to climate change, causing the Earth's surface to radiate infrared radiation.

The exact nature of this climate change is difficult to know, but some of the things that scientists believe could happen are:

- Polar ice caps could melt sea levels could rise
- There could be more extreme weather (such as floods, droughts, and heat waves).



▲ Figure 13.8 The greenhouse effect

2.4 Reactivity series

Displacement reactions involving metal oxides.

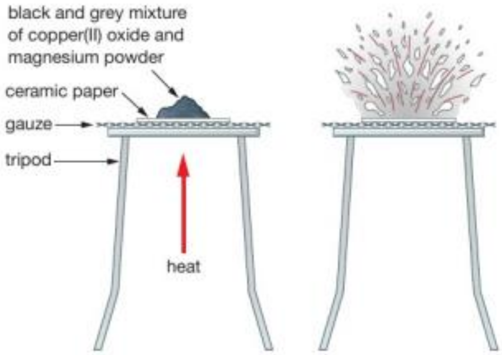
The reaction between magnesium and copper oxide

Magnesium powder and copper (II) oxide are mixed and heated very strongly. At the end, traces of white magnesium oxide are left on the ceramic paper.

magnesium + copper(II) oxide → magnesium oxide + copper

$$Mg(s) + CuO(s) \rightarrow MgO(s) + Cu(s)$$

If you heated copper with magnesium oxide, nothing would happen because copper is less reactive than magnesium. Copper isn't capable of displacing magnesium from magnesium oxide.



▲ Figure 14.3 Magnesium reacts with copper(II) oxide.

The reaction between magnesium and zinc oxide

Heating magnesium with zinc oxide produces zinc metal:

magnesium + zinc oxide → magnesium oxide + zinc

$$Mg(s) + ZnO(s) \rightarrow MgO(s) + Zn(s)$$

The reaction between carbon and copper oxide

A black mixture of carbon and copper (II) oxide is heated in a test-tube. The mixture glows red hot because of the heat given out during the reaction, and you are left with pink-brown copper in the tube.

$C(s)+2CuO(s)\ CO_2(g) +2Cu(s)$

Carbon is above copper in the reactivity series and displaces the copper from copper (II) oxide.

Oxidation and reduction

A substance has been oxidised if it gains oxygen. **Oxidation** is gain of oxygen.
A substance has been reduced if it loses oxygen**. Reduction** is loss of oxygen.

Consider the reaction between magnesium and copper (II) oxide again. Figure 14.5 shows what happens in terms of oxidation and reduction.

▲ Figure 14.5 Magnesium reacting with copper(II) oxide

A **redox** reaction is one in which both reduction and oxidation are occurring. Oxidation and reduction always occur together because if something loses oxygen, something else must gain it.

■**A reducing agent** is a substance that reduces something else. In this case, the magnesium is the reducing agent because it takes the oxygen away from the copper oxide, reducing it.
■**An oxidising** agent is a substance that oxidises something else. The copper (II) oxide is the oxidising agent in this reaction because it gives oxygen to the magnesium, oxidising it.

Electron transfer

We are now going to look very closely at what happens in the reaction between magnesium and copper (II) oxide in terms of the various particles involved. Here is the equation again:

$$Mg(s) + CuO(s) \rightarrow MgO(s) + Cu(s)$$

The magnesium and the copper are metals, and are made of metal atoms, but the copper(II) oxide and the magnesium oxide are both ionic compounds.

The copper (II) oxide contains Cu^{2+} and O^{2-}ions, and the magnesium oxide contains Mg^{2+} and O+ions. Writing these into the equation gives:

This reaction involves magnesium atoms forming magnesium ions, which lose electrons, and copper ions gaining these electrons to form copper atoms. The oxygen is not the important part, and a more fundamental definition of oxidation and reduction is needed to better understand this process.

Oxidation is loss of electrons.
Reduction is gain of electrons.

Displacement reactions involving solutions of salt.

The reaction between zinc and copper (II) sulfate solution

The copper is displaced by the more reactive zinc. The blue colour of the copper (II) sulfate solution fades as colourless zinc sulfate solution is formed:

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

The zinc and the copper are metals consisting simply of atoms, but the copper(II) sulfate and the zinc sulfate are metal compounds and so are ionic. The equation can be rewritten showing the ions:

$$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

We can see that the sulfate ions are the same on both sides of the equation - they have not changed at all in the reaction. We call the sulfate ions here spectator ions. Removing the spectator ions (because they aren't changed during the reaction) leaves:

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

The Zn atoms are oxidised to Zn2+ ions because they lose electrons. The Cu2+ ions gain electrons and are reduced to Cu atoms. We can split up the ionic equation to show the individual oxidation and reduction processes:

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

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

These are called ionic half-equations as they each only show one of the processes (either oxidation or reduction) occurring in the reaction.

oxidizing and reducing agents in terms of electrons

■A couple more definitions (this time in terms of electrons): An oxidising agent is something that oxidises something else by taking electrons away from it. Oxidising agents accept electrons and therefore are reduced in a reaction.

■A reducing agent is something that reduces something else by giving electrons to it. Reducing agents give away electrons and therefore are oxidised in a reaction.

The reaction between copper and silver nitrate solution

Silver is below copper in the reactivity series, so a coil of copper wire in silver nitrate solution will produce metallic silver. Figure 14.7 shows the silver being produced as a mixture of grey 'fur' and delicate crystals. Notice the solution becoming blue, as copper(II) nitrate is produced:

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

This time the nitrate ions are spectator ions, and the final version of the ionic equation looks like this:

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

This is another redox reaction.

The Cu is the reducing agent (it is oxidised) because it gives electrons to the Ag+ ions to reduce them to Ag. The Ag+ ion is the oxidising agent (it is reduced) because it takes electrons away from the Cu atoms.

Reactions of metal with water

Reactions of potassium, sodium, or lithium with cold water

These reactions are described in detail on pages 125-126. They are very vigorous reactions but become less violent in the following order: potassium > sodium > lithium. The equations all look like this:

$$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$$

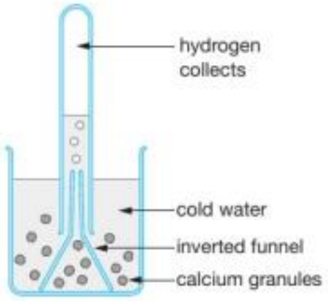
Replace M by K, Na or Li, depending on which metal you want.

Reaction of calcium with cold water

Calcium reacts gently with cold water. The grey granules sink but are carried back to the surface again as bubbles of hydrogen are formed around them. The mixture becomes warm as heat is produced.

Calcium hydroxide is formed. This isn't very soluble in water. Some of it dissolves to give a colourless solution, but most of it is left as a white, insoluble solid.

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq\ or\ s) + H_2(g)$$



▲ Figure 14.9 Calcium reacting with cold water

Reaction of magnesium with cold water

There is almost no reaction. If the magnesium is very clean, a few bubbles of hydrogen form on it, but the reaction soon stops again. This is because the magnesium becomes coated with insoluble magnesium hydroxide, which prevents any more water encountering the magnesium.

Reaction of magnesium with steam

Magnesium ribbon can be heated in steam using the apparatus shown in Figure 14.10.

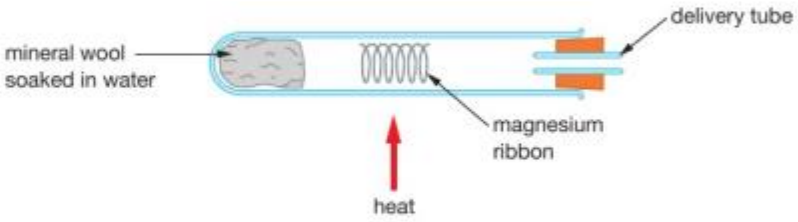
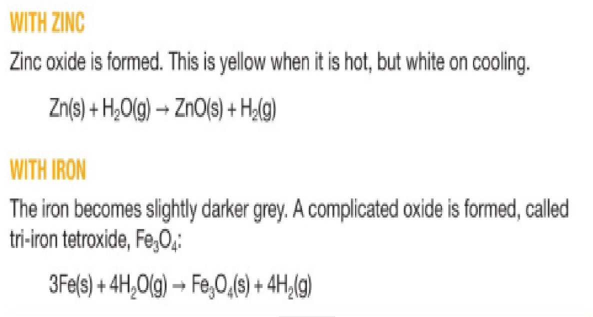


Figure 14.10 Magnesium reacting with steam

The mineral wool isn't heated directly. Enough heat moves back along the test-tube to turn the water to steam.
The magnesium burns with a bright white flame in the steam, producing hydrogen, which can be ignited at the end of the delivery tube. White magnesium oxide is formed:

$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$

Reaction of zinc or iron with stem



WITH ZINC

Zinc oxide is formed. This is yellow when it is hot, but white on cooling.

$$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$$

WITH IRON

The iron becomes slightly darker grey. A complicated oxide is formed, called tri-iron tetroxide, Fe_3O_4:

$$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$$

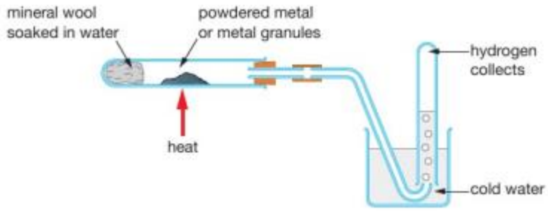


Figure 14.11 Zinc or iron reacting with steam

With both zinc and iron, the hydrogen comes off slowly enough to be collected. Neither metal burns.

Reactions of metal with dilute acids

Metals above hydrogen in the reactivity series

Metals above hydrogen react with acids to form a salt (e.g. magnesium sulfate or zinc chloride) and hydrogen. The higher the metal in the series, the more violent the reaction.

metal + acid salt + hydrogen
metal + dilute sulfuric acid → metal sulfate + hydrogen
metal + dilute hydrochloric acid → metal chloride + hydrogen

metals below hydrogen in the reactivity series

Metals such as copper, silver and gold do not react with simple dilute acids such as sulfuric or hydrochloric acid.

Potassium, sodium, lithium, and calcium with dilute acids

These are too reactive to add safely to acids, the reaction is too violent. Calcium can be used if the acid is very dilute.

Metals from magnesium to iron in the reactivity series

Magnesium reacts vigorously with cold dilute acids, and the mixture becomes hot. A colourless solution of magnesium sulfate or chloride is formed. With dilute sulfuric acid:

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

Aluminium is slow to start reacting, but after warming it reacts very vigorously. There is a very thin, but very strong, layer of aluminium oxide on the surface of the aluminium, which stops the acid from getting to it. On heating, the acid removes this layer, and the aluminium can show its true reactivity. With dilute hydrochloric acid:

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

Zinc and iron react slowly in the cold, but more rapidly on heating. The vigour of the reactions is less than that of aluminium. The zinc forms zinc sulfate or zinc chloride and hydrogen. The iron forms iron (II) sulfate or iron(II) chloride and hydrogen.

Finding approximate position of a metal in the reactivity series using water and dilute acids

Add a very small piece of metal to some cold water. If there is any rapid reaction, then the metal must be above magnesium in the reactivity series.
If there isn't any reaction, add a small amount of metal to some dilute hydrochloric acid (or dilute sulfuric acid). If there isn't any reaction in the cold acid, warm it carefully.
If there is still no reaction, the metal is probably below hydrogen in the reactivity series. If there is a reaction, then it is somewhere between magnesium and hydrogen.

Making predictions using the reactivity series

You can make predictions about the reactions of unfamiliar metals if you know their position in the reactivity series.

Suppose you have a question as follows.

Manganese (Mn) lies between aluminium and zinc in the reactivity series and forms a 2+ ion. Solutions of manganese (II) salts are very pale pink (almost colourless).

a. Use the reactivity series to predict whether manganese will react with copper (II) sulfate solution. If it will react, describe what you would see, name the products, and write an equation for the reaction.

b. Explain why you would expect manganese to react with steam. Name the products of the reaction and write the equation.

a. The reaction between manganese and copper (II) sulfate solution
Manganese is above copper in the reactivity series and so will displace it from the copper (II) sulfate:

A pink-brown deposit of copper will be formed. The colour of the solution will fade from blue and leave a very pale pink (virtually colourless) solution of manganese (II) sulfate.

$Mn(s) + CuSO_4(aq) \rightarrow MnSO_4(aq) + Cu(s)$

b. The reaction between manganese and steam
Manganese is above hydrogen in the reactivity series and so reacts with steam to give hydrogen and the metal oxide, in this case manganese (II) oxide.
You couldn't predict the colour of the manganese (II) oxide, and the question doesn't ask you to do this.

$Mn(s) + H_2O(g) \rightarrow MnO(s) + H_2(g)$

Rusting for iron

Iron **rusts** in the presence of oxygen and water. Rusting occurs with iron and the most common alloy of iron, mild steel (see below).

▲ Figure 14.17 Rusting is accelerated by salty water.

Preventing rusting by using barriers

The most obvious way of preventing rusting is to keep water and oxygen away from the iron. You can do this by painting it, coating it in oil or grease, or covering it in plastic. Coating the iron with a metal below it in the reactivity series (e.g. coating steel with tin for tin cans) is also a **barrier method.**

Preventing rusting by galvanizing

Galvanised iron is iron that is coated with a layer of zinc. If the zinc layer is unscratched, it serves as a barrier to air and water.

Preventing rusting by using sacrificial protection

Sacrificial anodes, made of zinc, magnesium, or aluminium, are used to protect iron/steel from rusting on metal hulls or keels of ships. These more reactive metals react more readily in oxygen/water, preventing iron from rusting. They are replaced periodically when more reactive metal has been oxidized. Galvanising combines barrier methods and **sacrificial protection** for underground pipelines.

2.5 Extraction and uses of metals

Extracting metals from their ores

Minerals and ores
Most metals are found in the Earth's crust combined with other elements. The individual compounds are called **minerals.**

An **ore** is a sample of rock that contains enough of a mineral for it to be worthwhile to extract the metal. Most metals are extracted from ores found in the Earth's crust.

A few very unreactive metals, such as gold, are found **native.**

Extracting the metals

Many ores contain either oxides or compounds that are easily converted to oxides. Sulfides such as sphalerite (zinc blende), ZnS, can be easily converted into an oxide by heating in air, a process known as roasting.

$2ZnS(s) + 30_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$

Methods of extraction and the reactivity series

A manufacturer obviously wants to use the cheapest possible method of extracting a metal from an ore. There are two main economic factors to consider:

- The cost of energy
- The cost of the reducing agent.

potassium
sodium
calcium
magnesium
aluminium
(carbon)
zinc
iron
copper

Figure 15.3 A part of the reactivity series

Metals below carbon in the reactivity series

For a metal below carbon in the reactivity series, the cheapest method of reducing the ore is often to heat it with carbon.

One of the main ores of iron contains a high percentage of iron (III) oxide. The iron can be extracted from this by heating with carbon:

$Fe2O3(s) + 3C(s) \rightarrow 2Fe(l) + 3CO(g)$

Carbon is higher in the reactivity series than iron and will take the oxygen away from the iron oxide. This is a redox reaction; the Fe_2O3 is reduced to Fe in the reaction and the C is oxidised to CO. In this reaction the carbon is the reducing agent, it reduces the iron (III) oxide.

The extraction of iron is carried out in a blast furnace. The process is a bit more complicated than we have shown here. Other reactions also occur, and the main reducing agent is carbon monoxide:

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$

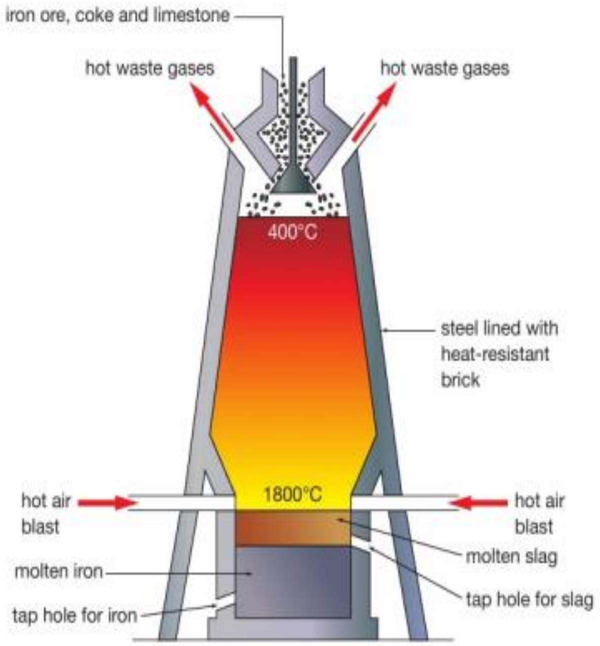


Figure 15.4 Iron is extracted from its ore in a blast furnace.

Metals above carbon in the reactivity series

Metals higher in reactivity series than zinc cannot be reduced using carbon due to their higher reactivity and are typically produced through electrolysis.

Aluminium is extracted by the electrolysis of aluminium oxide (Al_2O_3) dissolved in a molten salt called **cryolite.**

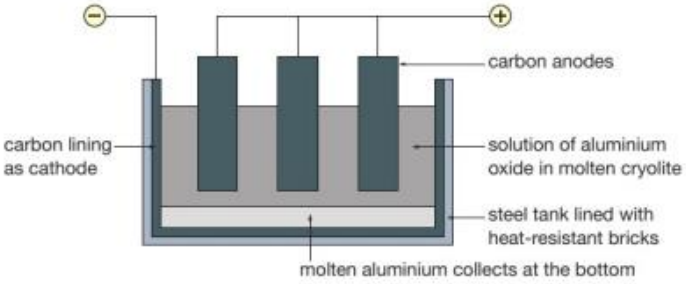


Figure 15.5 Aluminium is produced by electrolysis.

At the cathode the reaction is $Al^{3+} + 3e^- \rightarrow Al$ reduction

At the anode the reaction is $2O^{2-} \rightarrow O_2 + 4e^-$ oxidation

Electrolysis is an expensive process for extracting metals like aluminium, iron, and titanium. Despite being more abundant, it is more expensive than reduction with carbon. Other metals like titanium can be extracted by heating compounds with reactive metals, which is also expensive due to the initial extraction process.

Alloys

An alloy is a mixture of a metal with, usually, other metals or carbon.

Alloys are harder than the individual pure metals from which they are made. In an alloy, the different metals/elements have slightly differently sized atoms.

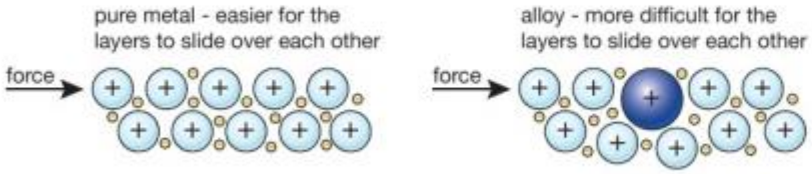


Figure 15.8 The layers cannot slide over each other as easily in an alloy as in a pure metal.

Other common alloys include bronze (a mixture of copper and tin), stainless steel (an alloy of iron with chromium and usually nickel), and the mixture of copper and nickel (cupronickel) which is usually used to make 'silver' coins.

Properties and uses of some metals

Uses of aluminum

Aluminium alloys, made with elements like silicon, copper, or magnesium, are commonly used due to its low density, strength, ability to conduct electricity and heat, and resistance to corrosion. Its uses include planes, electricity cables, and pots and pans.

Properties and uses of different kinds of steel

Mild steel

There are lots of different alloys of iron that contain various proportions of carbon and sometimes other metals. An alloy of iron and carbon is called **steel.**

Mild steel is an iron alloy with 0.25% carbon, enhancing its hardness and strength. It's easily hammered into shapes and wires, used in nails, car bodies, ship building, girders, and bridges.

A disadvantage of mild steel is that it rusts when exposed to oxygen and water. It is also about three times denser than aluminium. Some car bodies are made from aluminium. This has the advantage that the car body will not rust and, because the car is lighter, less fuel will have to be used.

Figure 15.10 Mild steel is used for car bodies . . .

Figure 15.11 Some cars have a body made of aluminium. In this photo the car body has been painted. The paint is mainly for decoration; it is not required to prevent rusting.

High – carbon steel

Figure 15.12 High-carbon steel is used to make cutting tools.

High-carbon steel is iron containing about 0.6-1.2% carbon (these amounts are variable). High-carbon steel is harder and more resistant to wear than mild steel but more brittle (not as malleable and ductile). It is used for cutting tools. High-carbon steel also usually contains small amounts of manganese.

stainless steel

Stainless steel is an alloy of iron, chromium, and nickel, known for its resistance to corrosion. It is used in various industries, including kitchen sinks, saucepans, knives, forks, and gardening tools. It is also used in brewing, dairy, and chemical industries for corrosion-resistant vessels. Stainless steel is more expensive than mild steel.

Figure 15.13 Some cars have a stainless steel body. They do not need to be painted.

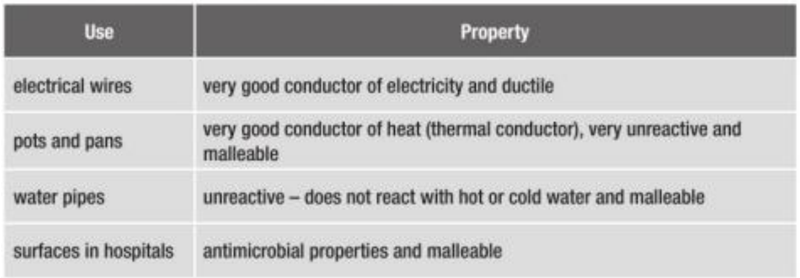
Table 15.1 A summary of types of steel

Type of steel	Iron mixed with	Some uses
mild steel	up to 0.25% carbon	nails, car bodies, ship building, girders
high-carbon steel	0.6–1.2% carbon	cutting tools, masonry nails
stainless steel	chromium (and nickel)	cutlery, cooking utensils, kitchen sinks

Copper

Some properties and uses of copper and its alloys are summarised in Table 15.2.

Table 15.2 Copper and its alloys have a variety of uses



Use	Property
electrical wires	very good conductor of electricity and ductile
pots and pans	very good conductor of heat (thermal conductor), very unreactive and malleable
water pipes	unreactive – does not react with hot or cold water and malleable
surfaces in hospitals	antimicrobial properties and malleable

Figure 15.14 Copper pots and pans

2.6 Acids, alkalis, and titrations

pH and indicators

The pH scales.

The pH scale ranges from about 0 to about 14 and tells you how acidic or how alkaline a solution is.

We can classify substances as strongly acidic/alkaline or weakly acidic/alkaline based on their pH. Remember, when writing pH, it is a small p and a big H (the symbol for hydrogen).

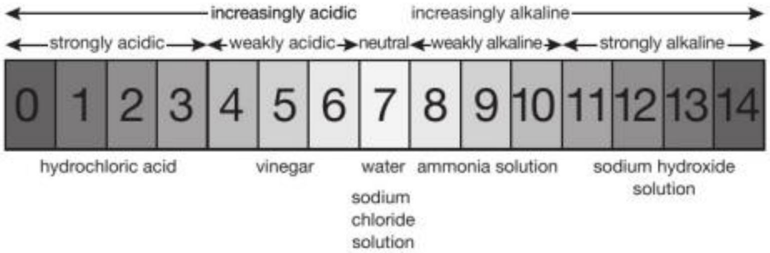
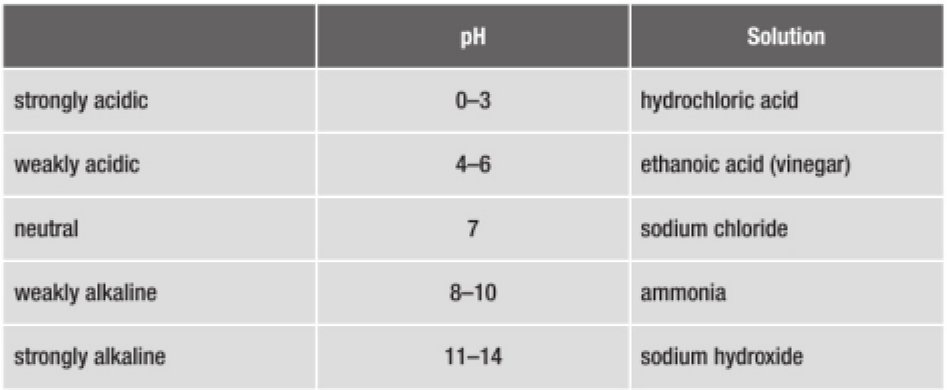


Figure 16.3 The pH scale. Vinegar is a solution of ethanoic acid.

Table 16.1 We can classify solutions according to their pH



	pH	Solution
strongly acidic	0–3	hydrochloric acid
weakly acidic	4–6	ethanoic acid (vinegar)
neutral	7	sodium chloride
weakly alkaline	8–10	ammonia
strongly alkaline	11–14	sodium hydroxide

Measuring pH

Using universal indicator

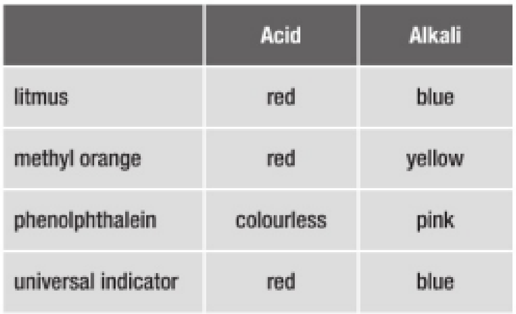
Universal indicator is a mixture of indicators that gradually changes color over pH values, used as a solution or paper. The most common form is full-range universal indicator, which changes colors from pH 1 to 14, but isn't accurate. It's checked against a chart to determine pH.

▲ Figure 16.4 Using universal indicator solution to measure the pH of various solutions.

Acid alkali indicators

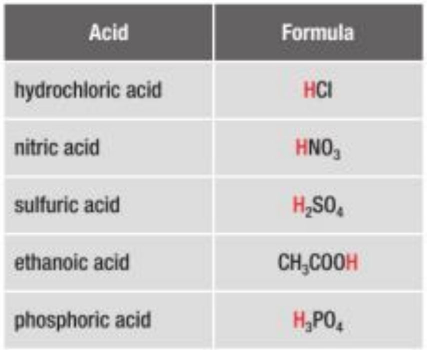
Any substance that has different colours depending on the pH can be used as an indicator. Some common indicators and their colours in acidic and alkaline solutions are shown in Table 16.2.

Table 16.2 The colours of various indicators in acidic and alkaline solutions



	Acid	Alkali
litmus	red	blue
methyl orange	red	yellow
phenolphthalein	colourless	pink
universal indicator	red	blue

Table 16.3 Some acids showing the replaceable hydrogen



Acid	Formula
hydrochloric acid	HCl
nitric acid	HNO_3
sulfuric acid	H_2SO_4
ethanoic acid	CH_3COOH
phosphoric acid	H_3PO_4

Litmus is red in acidic solutions and blue in alkaline ones. In neutral solutions, the colour is purple, which is an equal mixture of the red and blue forms.
Universal indicator is green in neutral solutions.

Acids

The formulae of some acids are given in Table 16.3.
All acids contain hydrogen and when acids react the hydrogen shown in red in Table 16.3 is replaced by something else; all acids have replaceable.

Not all the hydrogens in acids are replaceable.

Bases

Bases are substances that neutralise acids by combining with the hydrogen ions in them. When we are referring to a base at this level, we usually mean a metal oxide, a metal hydroxide or ammonia (NH_3)

Alkalis

Some bases dissolve in water to form solutions containing hydroxide ions. These are alkalis. Alkalis are a source of hydroxide (OH) ions in solution.

Examples of alkalis are sodium hydroxide and potassium hydroxide (all the Group 1 hydroxides). When sodium hydroxide is in water it breaks apart to form sodium and hydroxide ions.

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

The other alkali you will meet is a solution of ammonia (NH3). The ammonia reacts with the water to form ammonium ions and hydroxide ions;

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

These alkalis all have a pH greater than 7.

Other alkaline solutions

There are some other substances, such as soluble metal carbonates, that react with water to form hydroxide ions

$$Na_2CO_3(aq) + H_2O(l) \rightleftharpoons NaOH(aq) + NaHCO_3(aq)$$

Soluble carbonates like sodium and potassium carbonate are alkalis with a pH above 7, due to OH- ions in the solution, while only some carbonate ions react with water.

Reacting acids with bases and alkalis

Acids react with bases or alkalis in a neutralisation reaction.
Metal oxides, such as copper (II) oxide and magnesium oxide, are bases. Copper (II) oxide reacts with hot dilute sulfuric acid, in a neutralisation reaction, to produce a solution of copper (II) sulfate and water:

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

Copper (II) oxide is an ionic compound containing the O2-ion and what has happened in this reaction is that the H+ ions from the acid have combined with the O2-ions from the base to form water (H_2O).

Reacting with alkalis

Sodium hydroxide solution (an alkali) reacts with dilute hydrochloric acid to form sodium chloride and water:

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

This is a neutralisation reaction. Sodium hydroxide and sodium chloride are both ionic compounds and so will be present as ions in solution. HCl(aq) is an acid and so will dissociate into H+(aq) and Cl(aq). We can therefore re-write this equation showing all the ions.

$$\boxed{Na^+(aq)} + OH^-(aq) + H^+(aq) + \boxed{Cl^-(aq)} \rightarrow \boxed{Na^+(aq)} + \boxed{Cl^-(aq)} + H_2O(l)$$

The Na+(aq) and Cl-(aq) ions are the same on both sides of the equation and so are spectator ions. We can leave the spectator ions out of this equation to just show the things that have changed:

$$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$$

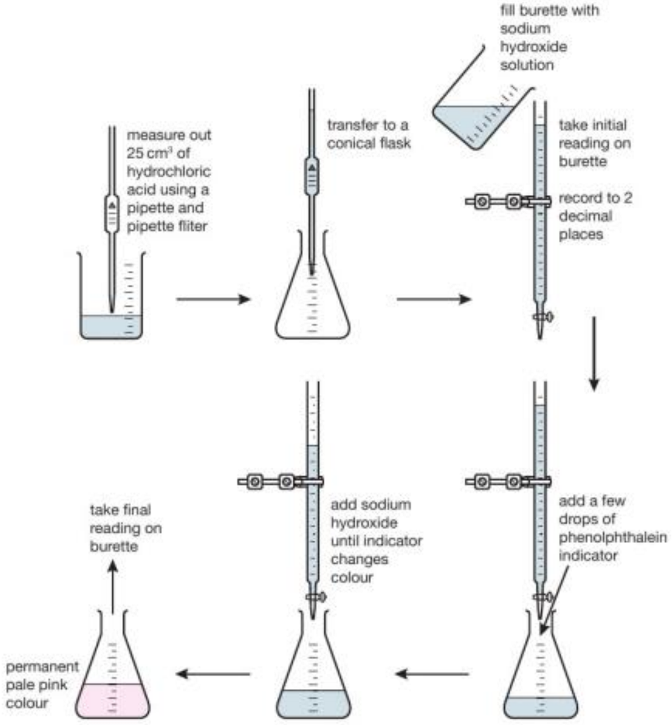
This is an ionic equation. All neutralisation reactions for an acid reacting with an alkali will have the same ionic equation. This is because they all involve the OH- ions from the alkali reacting with the H+ ions from the acid to form water.

Titration

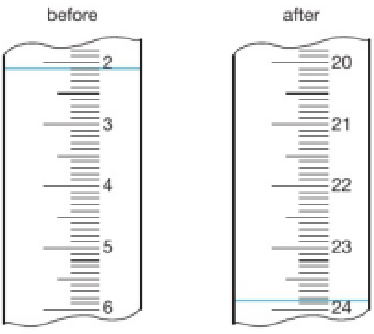
We can use titration to follow the course of a neutralisation reaction between an acid and alkali. Titration can be used to find out how much of the acid/alkali reacts with a certain volume of the alkali/acid.

The process involves adding alkali to an acid from a burette until the indicator changes color. A rough titration is conducted to determine the required amount of alkali. The experiment is repeated, this time adding 20 cm3 of alkali quickly and swirling the conical flask. The alkali is added slowly until one drop causes the indicator to change from acid to alkali color. The amount needed for neutralisation is then known within one drop. Readings from a burette should be recorded to the nearest 0.05 cm3, with the second decimal place representing 'O' if the level of solution is on the line and '5' if it is between the lines.

The technique for carrying out a titration using phenolphthalein as indicator is shown in Figure 16.7.

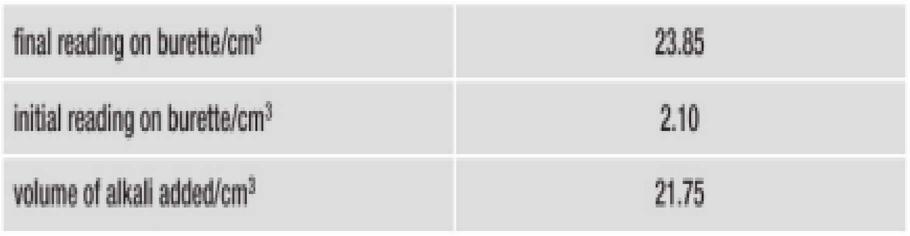


▲ Figure 16.7 How to carry out an acid–alkali titration



▲ Figure 16.8 Reading a burette

Table 16.4 Results table from a titration experiment



final reading on burette/cm^3	23.85
initial reading on burette/cm^3	2.10
volume of alkali added/cm^3	21.75

The volume of alkali added is obtained by subtracting the initial reading from the final reading:

volume of alkali = $23.85-2.1021.75cm^3$

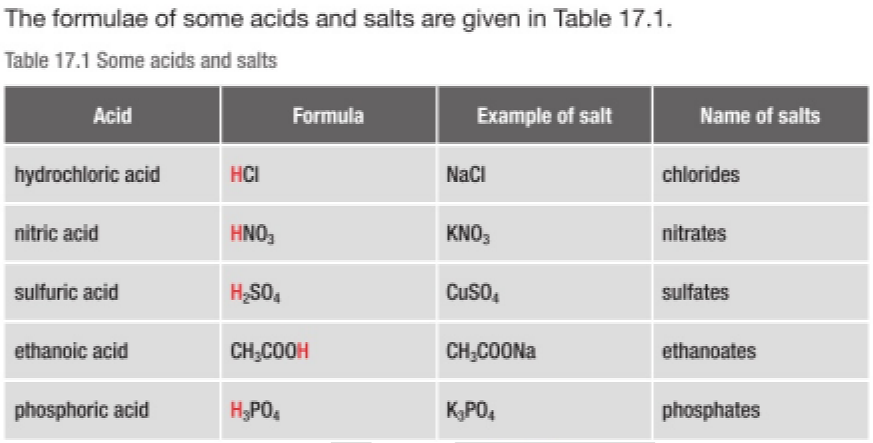
Various other indicators can also be used, such as methyl orange, but not universal indicator. We do not use universal indicator because it has a range of colours and changes gradually between them; this means we would not be able to see a clear endpoint for the titration.

2.7 Acids bases and salt preparations

salts

All acids contain hydrogen. When that hydrogen is replaced by a metal, the compound formed is called a salt. Magnesium sulfate is a **salt**, and so is zinc chloride, and so is potassium nitrate.

Sulfuric acid is the parent acid of all sulfates, which are salts formed from it. However, direct hydrogen replacement is not possible, as copper (II) sulfate cannot be made from copper and dilute sulfuric acid. Alternative methods exist, such as copper (II) sulfate being a salt. Ammonium salts, such as $NHCI$ and $(NH4)2SO4$, are formed when hydrogen is replaced with $NH4$.



The formulae of some acids and salts are given in Table 17.1.

Table 17.1 Some acids and salts

Acid	Formula	Example of salt	Name of salts
hydrochloric acid	HCl	$NaCl$	chlorides
nitric acid	HNO_3	KNO_3	nitrates
sulfuric acid	H_2SO_4	$CuSO_4$	sulfates
ethanoic acid	CH_3COOH	CH_3COONa	ethanoates
phosphoric acid	H_3PO_4	K_3PO_4	phosphates

Reactions of acids

Before we look at some practical methods to make salts, we need to study the reactions of acids.

Reacting acids with metals

Simple dilute acids react with metals depending on their positions in the reactivity series.
- Metals below hydrogen in the series don't react with dilute acids.
- Metals above hydrogen in the series react to produce hydrogen gas.
- The higher the metal is in the reactivity series, the more vigorous the reaction. You would never mix metals such as sodium or potassium with acids because their reactions are too violent.

A summary equation for metals above hydrogen in the reactivity series: metal + acid → salt + hydrogen

Reactions involving magnesium and acids.

With dilute sulfuric acids

There is rapid fizzing and a colourless gas is evolved, which pops with a lighted splint (the test for hydrogen). The reaction mixture becomes very warm as heat is produced. The magnesium gradually disappears to leave a colourless solution of magnesium sulfate.

$Mg(s) + H2SO4(aq) \rightarrow MgSO4(aq) + H2(g)$

This is a displacement reaction. The more reactive magnesium has displaced the less reactive hydrogen.

With dilute sulfuric acids

Reactions involving zinc and acids.

Again, the reactions between zinc and the two acids look the same. The reactions are slower because zinc is lower down the reactivity series than magnesium.
The equations are:

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Bases

We saw in Chapter 16 that bases are substances that neutralise acids by combining with the hydrogen ions in them to produce water. When we are referring to a base at this level, we usually mean a metal oxide, a metal hydroxide or ammonia.

Reactions acids with bases

Metal oxides, such as copper (II) oxide and magnesium oxide, are bases.

Reacting dilute sulfuric acid with copper (II) oxide

The copper (II) oxide (black powder) reacts with hot dilute sulfuric acid to produce a blue solution of copper (II) sulfate.

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

All the metal oxide and acid combinations that you will meet at International GCSE behave in exactly the same way as the reaction between copper (II) oxide and dilute sulfuric acid, that is, they produce a salt and water. Most need to be heated for the reaction to start.

Copper (II) oxide is an ionic compound containing the O^{2-}-ion. What has
happened in this reaction is that the H^+ ions from the acid have combined with the O^{2-}-ions to form water (H_2O).
The general equation for the reaction of a metal oxide (base) with an acid is;

metal oxide + acid → salt + water

This is a neutralisation reaction - the base neutralises the acid (water is formed).

Bases and alkalis

Some metal oxides are soluble in water and react with it to form solutions of metal hydroxides, for example:
$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

All the Group 1 oxides do this reaction, so for potassium oxide we would get:
$K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$

Most other metal oxides are not soluble in water.
Calcium oxide dissolves slightly to form calcium hydroxide:
$CaO(s) + H_2O(l) \rightarrow Ca\ (OH)_2(aq)$

Another alkali you will meet is a solution of ammonia (NH3). The ammonia reacts with the water to form ammonium ions and hydroxide ions:

$NH_3(aq) + H_2O\ (1)\ NH_2+(aq) + OH-(aq)$

All these solutions contain hydroxide ions. We learned in the last chapter that we can define alkalis as solutions that are a source of hydroxide (OH-) ions, so all these solutions are alkalis.
All metal hydroxides react with acids in a neutralisation reaction:

metal hydroxide + acid → salt + water

reacting acids with metal hydroxides

All metal hydroxides react with acids in a neutralisation reaction:

metal hydroxide + acid ⟶ salt + water

reacting dilute hydrochloric acid with sodium hydroxide solution

Mixing sodium hydroxide solution and dilute hydrochloric acid produces a colourless solution so not much seems to have happened. But if you repeat the reaction with a thermometer in the beaker, the temperature rises several degrees, showing that there has been a chemical change. Sodium chloride solution has been formed:

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H2O(l)$

We saw in Chapter 16 that the ionic equation for this reaction is:

$OH\ (aq) + H^+(aq) \rightarrow H_2O(l)$

All neutralisation reactions for an acid reacting with an alkali have the same ionic equation: they all involve the hydroxide ions from the alkali reacting with the H+ ions from the acid to form water.

Reactions acids with carbohydrates

Carbonates react with cold dilute acids to produce carbon dioxide gas. A summary equation for acids and carbonates:

carbonate + acid → salt + carbon dioxide + water.

The reactions between copper (II) carbonate and dilute acids

Green copper (II) carbonate reacts with the common dilute acids to give a blue or blue-green solution of copper(II) sulfate, copper(II) nitrate or copper(II) chloride. Carbon dioxide gas is given off. You can recognise this because it turns limewater milky.
All the equations have the same form:

$CuCO_3(s) + H2SO_4(aq) \rightarrow CuSO_4(aq) + CO_2(g) + H_2O(l)$
$CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu\ (NO3)_2(aq) + CO_2(g) + H_2O(l)$
$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + CO2(g) + H_2O(l)$

The reaction between sodium carbonate and dilute acid

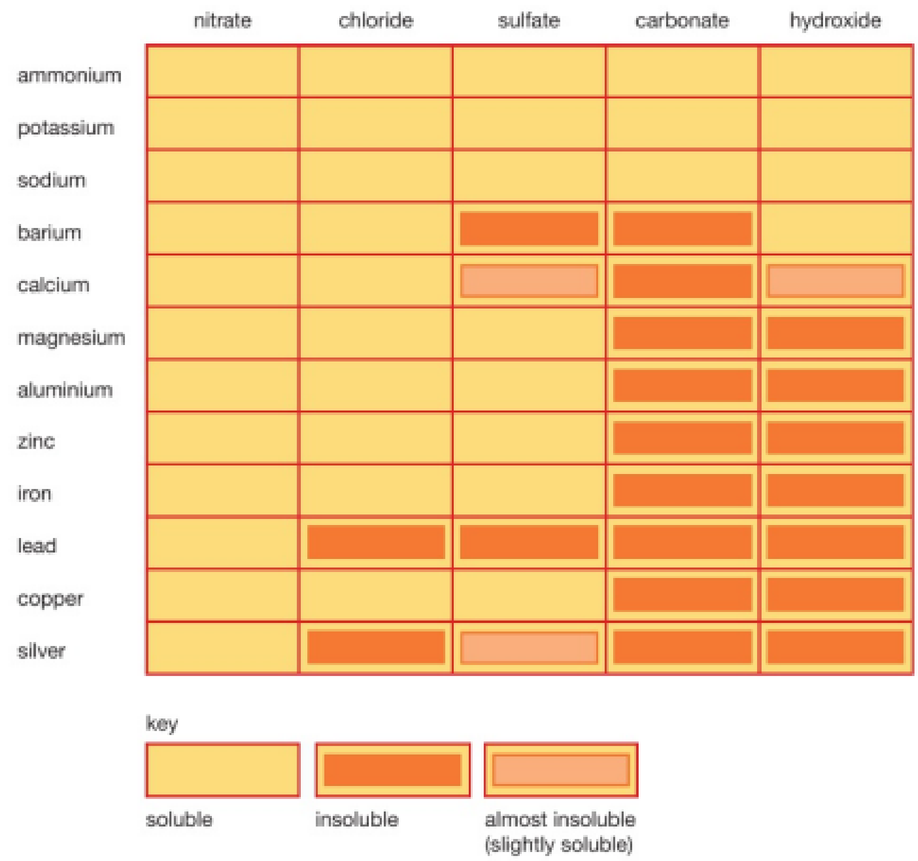
Sodium carbonate is soluble in water. The equation for the reaction of sodium carbonate with hydrochloric acid is:

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

Salt preparations

In the next sections we will learn about some practical methods to make salts. Before we can do this, we need to know which substances are soluble and which are insoluble in water because the method we use depends on this.

The solubility of ironic compound in the water



	nitrate	chloride	sulfate	carbonate	hydroxide
ammonium	🟨	🟨	🟨	🟨	🟨
potassium	🟨	🟨	🟨	🟨	🟨
sodium	🟨	🟨	🟨	🟨	🟨
barium	🟨	🟨	🟧	🟧	🟨
calcium	🟨	🟨	▨	🟧	▨
magnesium	🟨	🟨	🟨	🟧	🟧
aluminium	🟨	🟨	🟨	🟧	🟧
zinc	🟨	🟨	🟨	🟧	🟧
iron	🟨	🟨	🟨	🟧	🟧
lead	🟨	🟧	🟧	🟧	🟧
copper	🟨	🟨	🟨	🟧	🟧
silver	🟨	🟧	▨	🟧	🟧

key

🟨 soluble

🟧 insoluble

▨ almost insoluble (slightly soluble)

▲ Figure 17.4 **Solubility patterns**

Making soluble salts (accept sodium, potassium, and ammonium salts)

These all involve reacting a solid with an acid. You can use any of the following mixtures:

■Acid + metal (but only for the moderately reactive metals from magnesium to iron in the reactivity series)
■Acid + metal oxide or hydroxide
■Acid + carbonate.

Whatever mixture you use, the method is basically the same.

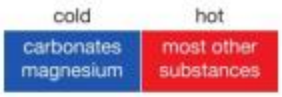
Making magnesium sulphate crystals

When we add the magnesium the reaction mixture will fizz (hydrogen is given off). We keep adding magnesium until the fizzing stops and there is magnesium left in the beaker. This means that all the acid has reacted. The equation for the formation of the solution of magnesium sulfate is:

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
Again, the solution is concentrated by heating it and allowed to crystallise. The crystallisation reaction is:

$MgSO_4(aq) + 7H_2O(l) \rightarrow MgSO4\ 7H_2O(s)$

How do you know whether you need to heat the mixture?
Carbonates react with dilute acids in the cold, and so does magnesium. Most other things that you are likely to come across need to be heated.



cold	hot
carbonates magnesium	most other substances

Figure 17.8 Do you need to heat the mixture?

Making sodium potassium and ammonium salts

The need for a different method
The method of adding an excess of a solid to an acid and filtering off the unreacted solid is not effective for making sodium, potassium, and ammonium salts. These compounds are soluble in water, and any excess solid would dissolve in the water, leaving no visible excess to filter off. There is no simple way to determine when just enough solid has been added to neutralize the acid.

Solving the problem by doing a titrations
Salts can be made from sodium, potassium hydroxide, ammonia solution, or carbonates. These solutions are alkaline, and a neutral solution can be determined using an indicator. Titration is a method used to determine the amount of acid and alkali needed for neutralization. After determining the required amount, a pure solution of salt can be created by mixing the same volumes again without the indicator.

Making sodium sulphate crystals

- 25 cm^3 of sodium hydroxide solution is transferred to a conical flask using a pipette, and a few drops of methyl orange are added as the indicator.

- Dilute sulfuric acid is run in from the burette until the indicator just turns from yellow to orange.

- The volume of acid needed is noted, and the same volumes of acid and alkali are mixed in a clean flask without any indicator.

The equation for the formation of the solution is:
$2NaOH(aq) + H2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O\ (1)$

- The solution is heated to evaporate off some of the water until a saturated solution is formed. It is then left to cool so that crystals form.
- The crystals are finally separated from any remaining solution by filtration (or the remaining solution can be poured away).
- The crystals are dried by patting them dry with a paper towel or by leaving them in a warm place.

The equation for the crystallisation process is:
$Na_2SO_4(aq) + 10H_2O(l) \rightarrow Na_2SO\ 10H_2O(s)$

Making sodium chloride crystals

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
You would need to do the titration using dilute hydrochloric acid rather than dilute sulfuric acid. You would then follow the same procedure as for sodium sulfate.

Making ammonium sulphate crystals

$2NH_3(aq) + H_2SO_4(aq) \rightarrow (NH_4)2SO_4(aq)$

Using ammonia solution rather than sodium hydroxide solution makes no difference to the method.

Making insoluble salts

Precipitation reactions

The basic procedure is to mix solutions of two soluble salts to form an insoluble salt and a solution of a soluble one. For instance, to make insoluble silver chloride we would mix solutions of silver nitrate and sodium chloride.

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

This is called a precipitation reaction. A precipitate is a solid that is formed by a chemical reaction involving liquids or gases.

Explaining what's happens?

Silver nitrate and sodium chloride solutions are attracted to each other but not strong enough to stick together. When mixed, the ions form a solid when silver ions meet chloride ions, while sodium and nitrate ions remain in solution due to their insufficient attraction. The attraction between the ions is not enough to stick together.

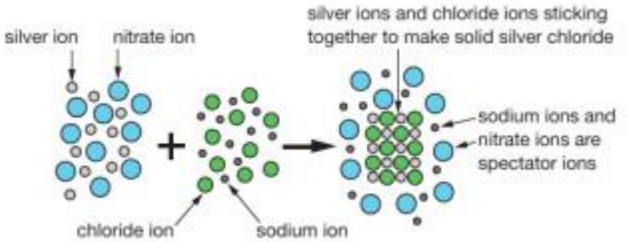


Figure 17.11 Precipitation of silver chloride

What do we mix to make insoluble salts?

Our procedure for making insoluble salts is to mix two solutions containing soluble salts, but how do we know what to mix? To determine this, two useful facts are:

■All nitrates are soluble
■All sodium and potassium salts are soluble.

So, all we must do is mix the nitrate of the metal part of our insoluble salt with the sodium or potassium salt of the non-metal part.

For instance, if we must make lead (II) iodide:

■The metal part is lead, so we use a solution of lead (II) nitrate.
■The non-metal part is iodide, so we use a solution of sodium iodide.

▲ Figure 17.12 A precipitate of lead(II) iodide

The equation for the reaction is:

$Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

Figure 17.12 shows the yellow precipitate of lead (II) iodide.

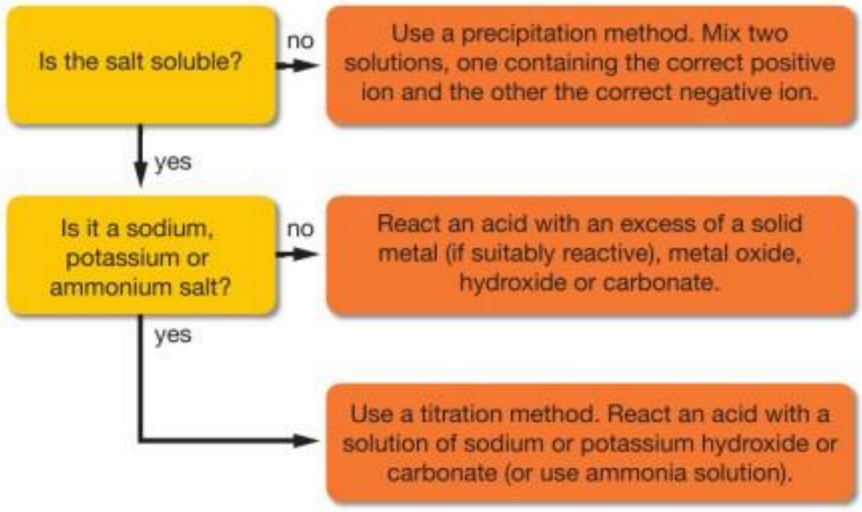
To make barium sulphate

We can use barium nitrate solution and potassium sulfate solution:

$$Ba(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KNO_3(aq)$$

The barium sulfate is a white precipitate.
There are other ways we could make barium sulfate. We could also use barium chloride, which is soluble in water. The sulfate part doesn't necessarily have to come from a salt; dilute sulfuric acid contains sulfate ions.

Summarize the methods making salts.



▲ Figure 17.13 Methods of making salts

Theories of acids and bases

The idea that an acid is something that produces H+(aq) ions in solution and an alkali is something that produces OH-(aq) ions in
solution. This is called the Arrhenius theory of acids and bases. For example, when hydrogen chloride gas is dissolved in water to form hydrochloric acid the molecules dissociate to form H+ ions:

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

A neutralisation reaction can be shown as H+ from an acid reacting with OH- from an alkali to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

The problem with this definition of acids and bases is that it only applies to reactions that occur in aqueous solution and is not more widely applicable.

There is a more general theory of acids and bases that can be used in more situations:

■An acid is a proton (hydrogen ion) donor.
■A base is a proton (hydrogen ion) acceptor. This is called the **Brønsted-Lowry theory.**

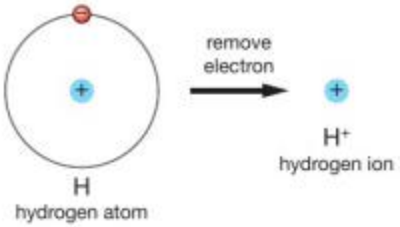


Figure 17.14 A hydrogen ion (H^+) is a proton. If an electron is removed from a hydrogen atom, only a proton is left.

The H_2O^+(aq) ion is called **the hydroxonium ion**. This is the ion that we usually write simply as H+(aq). You can think of it as a hydrogen ion joined to a water molecule. Cl- has a negative charge because the H+ left an electron behind when it transferred to the H_2O.

THE REACTION BETWEEN NH, AND HCl

Dilute hydrochloric acid reacts with ammonia solution in a neutralisation reaction to form a salt, ammonium chloride:

$$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$$

Ammonium chloride is an ionic salt so we can show the ions that are present:

$$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

Now we can see that, to form ammonium chloride, the HCl must donate a proton (H+) to the NH3. The HCl is an acid because it donates a proton and the NH3 is a base because it accepts the proton.

Figure 17.16 HCl is an acid and donates a proton to NH_3 (a base).

More acid- base reactions

Acids in solution are acidic because they donate a proton (H+) to water to form the hydroxonium ion. If we take a general acid (HA), this is what happens when it is put into water:

$$HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$$

HA donates a proton to H2O, which acts as a base. For nitric acid this would be;

$$\underset{\text{acid}}{HNO_3(aq)} + \underset{\text{base}}{H_2O(l)} \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

When bases, such as metal oxides (e.g. Na2O), react with water they accept a proton from water:

$$(Na^+)_2O^{2-}(s) + H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq)$$

The O2-ion accepts a proton from the water. In this reaction the water acts as an acid because it donates a proton.

When sulfuric acid reacts with copper (II) oxide (CuO) we have:

$$Cu^{2+}O^{2-}(s) + H_2SO_4(aq) \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq) + H_2O(l)$$

H2SO, is an acid, it donates protons to CuO, the base. Ammonia is also a base.

$$\underset{\text{base}}{NH_3(aq)} + \underset{\text{acid}}{H_2O(l)} \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

the ammonia accepts a proton (H+) from the water.

For reactions such as:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

What happens is that the H_3O^+ ion formed when the acid reacts with water donates a proton to the base, OH^-

$$\underset{\text{acid}}{H_3O^+(aq)} + \underset{\text{base}}{OH^-(aq)} \rightarrow 2H_2O(l)$$

But for International GCSE purposes, we almost always use the simplified version:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

2.8 Chemical tests

Testing for gases

Hydrogen H_2

THE TEST FOR HYDROGEN GAS

A lighted splint is held to the mouth of the tube. The hydrogen explodes with a squeaky pop. The hydrogen combines explosively with oxygen in the air to make water.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Oxygen O_2

THE TEST FOR OXYGEN GAS

A glowing splint is put into the tube containing the gas. Oxygen relights a glowing splint,

Carbon dioxide ($C0_2$)

THE TEST FOR CARBON DIOXIDE GAS

The carbon dioxide is bubbled through limewater. Carbon dioxide turns limewater milky/chalky/cloudy. Limewater is calcium hydroxide solution. Carbon dioxide reacts with it to form a white precipitate of calcium carbonate.

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

Chlorine, cl₂

The test for chlorine gas

A damp litmus paper or universal indicator paper is placed in a test tube, and chlorine, a green gas, bleaches it. If blue litmus paper is used, it turns red first due to the chlorine's dissolution in water, then white.

Figure 18.2 Chlorine bleaches damp litmus paper.

Ammonia NH₃

THE TEST FOR AMMONIA GAS

Hold a piece of damp universal indicator paper or red litmus paper at the mouth of the test-tube. Ammonia turns the universal indicator paper/litmus paper blue.

Testing for water

Using anhydrous copper sulphate

Water turns white anhydrous copper (II) sulfate blue. Anhydrous copper (II) sulfate lacks water of crystallisation and is white. Dropping water onto it replaces the water of crystallisation and turns it blue.

anhydrous copper sulfate + water → hydrated copper sulfate

$$\underset{\text{white}}{CuSO_4(s)} + 5H_2O(l) \rightarrow \underset{\text{blue}}{CuSO_4{\cdot}5H_2O(s)}$$

Testing for ions

All salts contain at least one cation (positive ion) and anion (negative ion). In this section we will learn how to test for these anions and cations. We do the tests for these anions and cations separately.

Flame tests.

A flame test is a method to detect the presence of metal ions in a compound by dipping a platinum or nichrome wire in concentrated hydrochloric acid and salt, then holding the wire and salt within a non-luminous Bunsen burner flame.

TESTING FOR CATIONS (POSITIVE IONS) USING SODIUM HYDROXIDE SOLUTION

Testing for positive ions using sodium hydroxide solution involves dissolved salt in distilled water and adding a volume of dilute sodium hydroxide solution. Most metal hydroxides are insoluble, so adding sodium hydroxide solution to a solution containing metal ions results in a precipitate of metal hydroxide. The goal is to observe the formation of a cloudy solid form.

Figure 18.9 The blue precipitate of copper(II) hydroxide.

The test for copper (II) irons

Formation of a blue precipitate shows the presence of copper (II) ions. The precipitate is copper (II) hydroxide.

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

Any copper (II) salt in solution will react with sodium hydroxide solution in this way. For example, with copper (II) sulfate solution, the full equation is:

$$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

Test for ions (III) irons.

Formation of a blue precipitate shows the presence of copper (II) ions. The precipitate is copper (II) hydroxide.

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

Any copper (II) salt in solution will react with sodium hydroxide solution in this way. For example, with copper (II) sulfate solution, the full equation is:

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$$

Test for ions (II) irons.

Formation of an orange-brown precipitate shows the presence of iron (III) ions (Fe^{3+}). The precipitate is iron (III) hydroxide.

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

Any iron (III) compound in solution will give this precipitate. An example of a full equation might be:

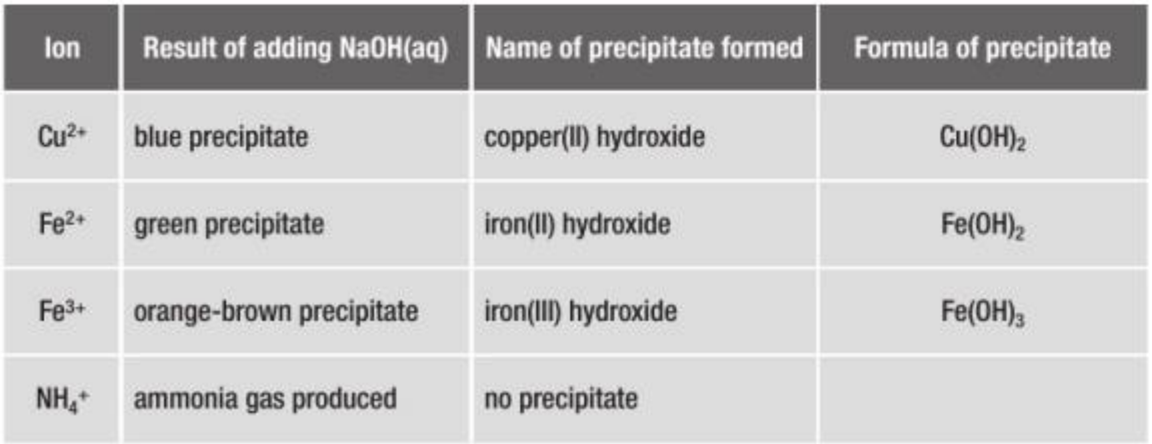
$$FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$$

Notice how much more complicated the full equations for these reactions are. They also hide what is going on. Use ionic equations for precipitation reactions wherever possible.

THE test for ammonium iorns

This test indicates the presence of ammonium salts by reacting sodium hydroxide solution with ammonium salts, producing ammonia gas. The gas can be detected by smelling it in cold conditions. When heated, the gas can be tested using damp red litmus paper or universal indicator paper, turning it blue.

Table 18.1 A summary of the reactions of positive ions with sodium hydroxide solution



Ion	Result of adding NaOH(aq)	Name of precipitate formed	Formula of precipitate
Cu^{2+}	blue precipitate	copper(II) hydroxide	$Cu(OH)_2$
Fe^{2+}	green precipitate	iron(II) hydroxide	$Fe(OH)_2$
Fe^{3+}	orange-brown precipitate	iron(III) hydroxide	$Fe(OH)_3$
NH_4^{+}	ammonia gas produced	no precipitate	

Testing for carbonates $C0_3^{2-}$

Add dilute hydrochloric acid to a salt sample and observe fizzing gas bubbles. Test for carbon dioxide by bubbling it through limewater. Most carbonates are insoluble in water, so this test works on solid samples like sodium, potassium, and ammonium carbonates. Soluble carbonates can also be tested.

Testing for sulphate (SO_4^{2-})

Acid is added to a solution to react with and remove anions, such as carbonate ions, to prevent white precipitates when adding barium chloride solution. It must be added before the solution, as carbonate ions are white and insoluble. It is important to avoid acidifying the solution with sulfuric acid, as it contains sulfate ions, which can result in a precipitate of barium sulfate.

Testing for chlorides (cl^-) bromides (Br^-) and iodides (I^-)

Make a solution of your suspected chloride, bromide or iodide and add enough dilute nitric acid to make it acidic. Then add some silver nitrate solution.

A white precipitate (of silver chloride) shows the presence of chloride ions:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

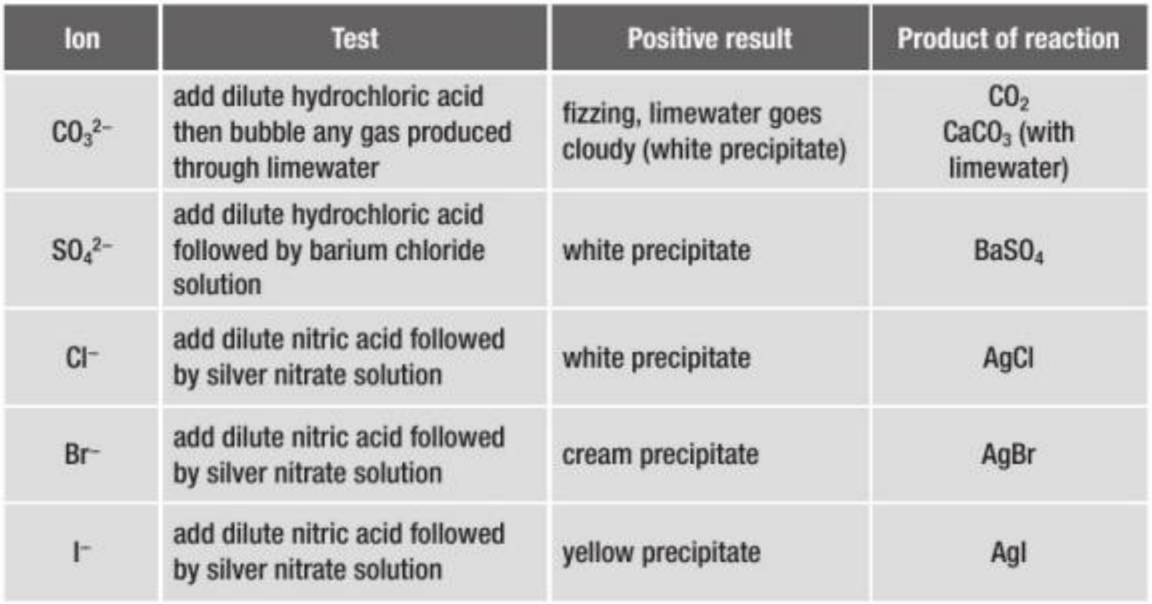
A cream precipitate (of silver bromide) shows the presence of bromide ions:

$$Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$$

A yellow precipitate (of silver iodide) shows the presence of iodide ions:

$$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

Table 18.2 A summary of the tests for negative ions (anions)



Ion	Test	Positive result	Product of reaction
CO_3^{2-}	add dilute hydrochloric acid then bubble any gas produced through limewater	fizzing, limewater goes cloudy (white precipitate)	CO_2 $CaCO_3$ (with limewater)
SO_4^{2-}	add dilute hydrochloric acid followed by barium chloride solution	white precipitate	$BaSO_4$
Cl^-	add dilute nitric acid followed by silver nitrate solution	white precipitate	AgCl
Br^-	add dilute nitric acid followed by silver nitrate solution	cream precipitate	AgBr
I^-	add dilute nitric acid followed by silver nitrate solution	yellow precipitate	AgI

Revision questions

1) Lithium, potassium, and caesium are three metals in Group 1 of the Periodic Table.

(a) A small piece of each metal is placed on water in separate large troughs. Complete the table by giving the correct metal, lithium, potassium, or caesium, for each description.

Description of reaction	Metal
explodes on contact with water	
fizzes gently	
reacts violently and forms a lilac flame	

(b) (i) Give the name and formula of the gas formed when potassium reacts with water

(ii) Give the name and formula of the compound formed when lithium reacts with water

(iii) Describe how you could show that an alkaline solution is formed when caesium reacts with water

2) Solutions of lead (II) nitrate and sodium chloride react together to form a precipitate of lead (II) chloride. The equation for the reaction is.

$$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$

A student carries out a series of experiments to find how much precipitate is formed when different volumes of lead (II) nitrate are added. She uses this method.

- place 15 cm3 of sodium chloride solution into a boiling tube
- add 2.0 cm3 of lead (II) nitrate solution
- allow the precipitate to settle
- measure the height of the precipitate
- repeat the experiment using different volumes of lead (II) nitrate solution The table shows the student’s results.

Volume in cm^3 of lead(II) nitrate added	Height of precipitate in cm
2.0	0.6
4.0	1.2
6.0	1.8
8.0	2.1
10.0	2.5
12.0	2.1
14.0	2.1

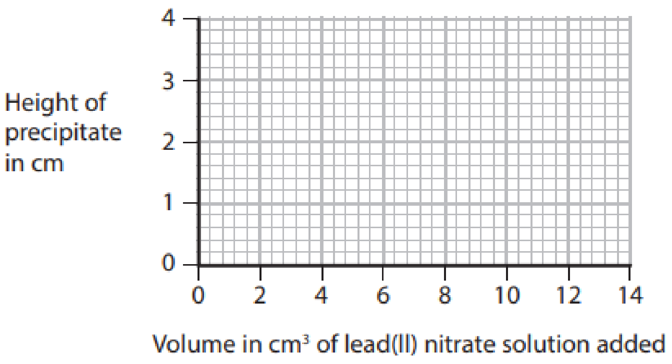
(a) Suggest why the height of the precipitate eventually stops increasing as more lead (II) nitrate solution is added.

(b) (i) Plot the student’s results on the grid. Draw a straight line of best fit through the origin and the first three points, and another straight line of best fit through the last four points. Make sure that the two lines cross.

(ii) Draw a circle on the grid around the point that represents the anomalous result

(iii) Why should the graph line pass through the origin?

(iv) Use your graph to estimate the volume of lead (II) nitrate solution that would be required to react completely with 15 cm^3 of the sodium chloride solution.



3) Three aqueous solutions are sodium chloride, sodium iodide and silver nitrate. They are in containers labelled X, Y and Z. It is not known which solution is in each container. The solutions are mixed together as shown in the table, and the observations recorded.

Experiment	Observation
solution X added to solution Y	yellow precipitate formed
solution X added to solution Z	no change
solution Y added to solution Z	white precipitate formed

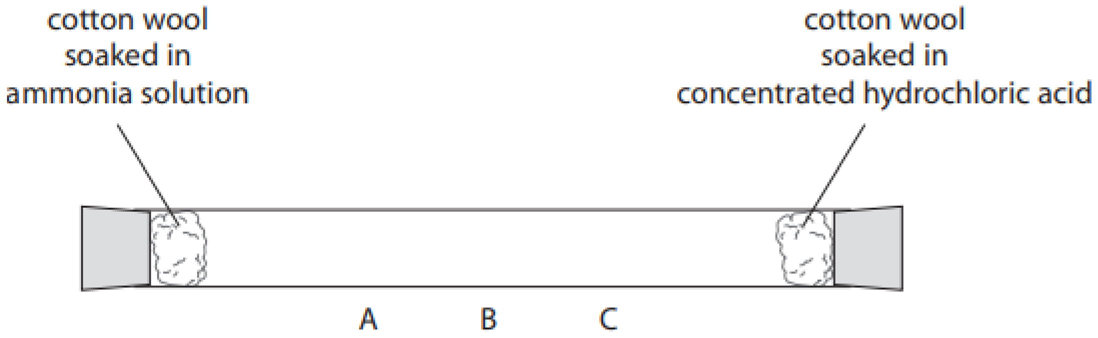
(a) Explain how the results show that Y is aqueous silver nitrate
(b) Explain how the results can be used to identify both X and Z
(c) Aqueous chlorine is added to separate aqueous solutions of sodium chloride and sodium iodide. Explain how the observations made can be used to distinguish between sodium chloride and sodium iodide.

4)A teacher added some of the Group 1 elements to separate samples of water.
(a) State two observations that could be made when a small piece of sodium is added to a large trough containing water
(b) In another experiment she added a small piece of a different Group 1 element and noticed that the reaction was less vigorous. Which element did she add in this experiment?
(c) In another experiment she added a small piece of potassium to a large trough containing water. This time she observed a lilac flame.
(i) Identify the gas that burned.
(ii) Give the formula of the ion that caused the flame to be lilac.
(d) When the Group 1 elements react with water, each of their atoms loses an electron from its outer shell. For sodium and potassium, these processes can be represented by the equations.

- $Na \rightarrow Na^+ + e^-$
- $K \rightarrow K^+ + e^-$

Explain, by referring to the electronic configurations of sodium and potassium, why potassium is more reactive than sodium.

5) In the second experiment, she soaks two pieces of cotton wool in different liquids and places them at opposite ends of a glass tube. She immediately seals the tube with bungs. The diagram shows the apparatus at the start of the experiment.



During the experiment a white ring appears in the tube.

(i) State whether the white ring appears at A, B or C.
(ii) Explain your choice.

b) A sample of a chlorofluorocarbon (CFC) contains 0.24 g of carbon, 0.38 g of fluorine and 1.42 g of chlorine.
(a) (i) Show, by calculation, that the empirical formula of the CFC is $CFCl_2$
(ii) The relative formula mass of the CFC is 204. Deduce the molecular formula of the CFC.

6) This question is about elements in Group 7 of the Periodic Table. (a) Complete the table to show the physical state at room temperature of fluorine and astatine, and the colour of liquid bromine

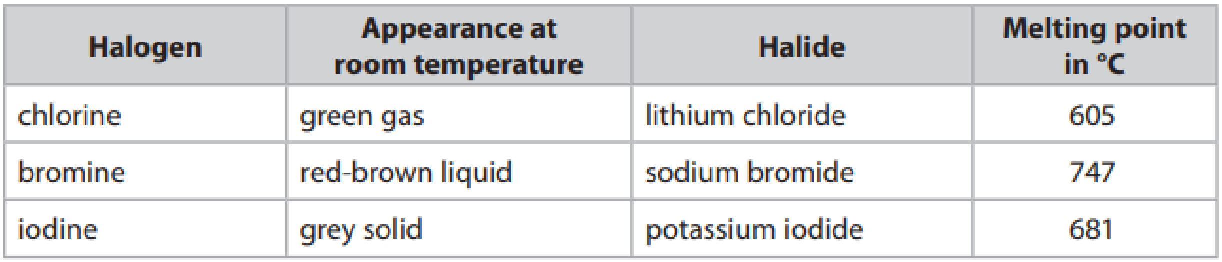
Element	Colour	Physical state at room temperature
fluorine	pale yellow	
chlorine	pale green	gas
bromine		liquid
iodine	dark grey	solid
astatine	black	

(b) Chlorine reacts with hydrogen to form hydrogen chloride. A piece of magnesium ribbon is added to hydrogen chloride in three separate experiments under different conditions. The table below shows the observations made under these different conditions.

Experiment	Conditions	Observations
1	Hydrogen chloride gas	No visible change
2	Hydrogen chloride dissolved in water	The magnesium ribbon gets smaller and bubbles are seen
3	Hydrogen chloride dissolved in methylbenzene	No visible change

(i) Write the formulae of two ions formed in the solution produced in experiment 2
Positive ion............................... Negative ion...............................
(ii) Identify the gas formed in experiment 2 and give a test for it.
gas.. test ....
(iii) Silver nitrate solution and dilute nitric acid are added to the solution produced in experiment 2. State what is observed and name the substance responsible for this observation. Explain why dilute nitric acid is added
(iv) Explain why there is no reaction in experiment 3.

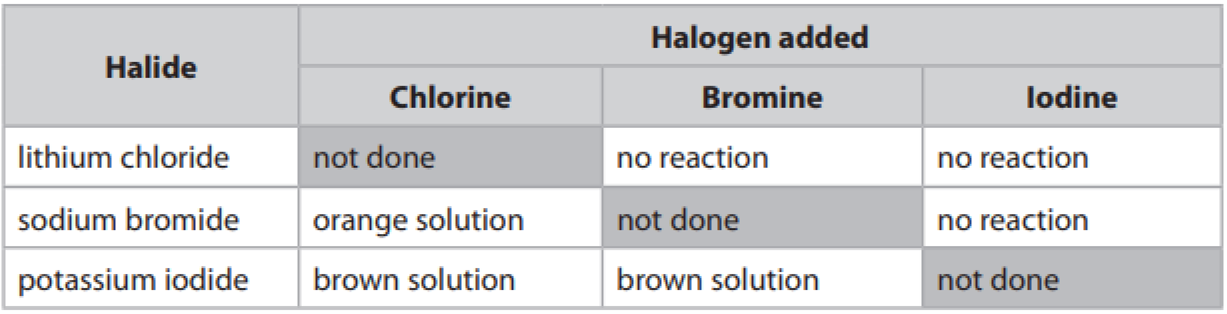
6) The halogens are elements in Group 7 of the Periodic Table. The halogens react with metals to form compounds called halides. Table 1 shows information about some halogens and their halides.



Halogen	Appearance at room temperature	Halide	Melting point in °C
chlorine	green gas	lithium chloride	605
bromine	red-brown liquid	sodium bromide	747
iodine	grey solid	potassium iodide	681

Table 1

(a) (i) Predict the physical state of fluorine at room temperature.
(ii) Predict how the colour of astatine at room temperature compares with the colour of iodine
(b) Each of the halides in table 1 was dissolved in water to form a solution. A sample of each of the halogens was then added to some of the halide solutions. Table 2 shows the results.

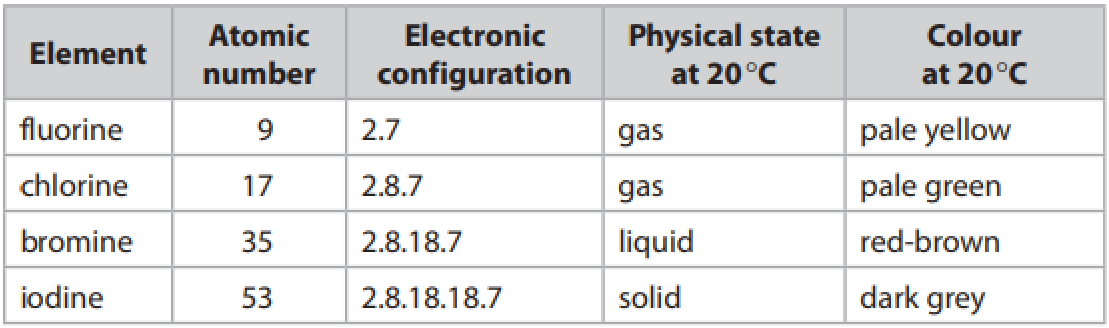


Halide	Halogen added		
	Chlorine	Bromine	Iodine
lithium chloride	not done	no reaction	no reaction
sodium bromide	orange solution	not done	no reaction
potassium iodide	brown solution	brown solution	not done

Table 2

(i) Suggest why there is no reason to add chlorine to lithium chloride solution.
(ii) Why was there no reaction when iodine was added to sodium bromide solution?
(iii) Name the substance with the brown colour that formed when chlorine was added to potassium iodide solution.

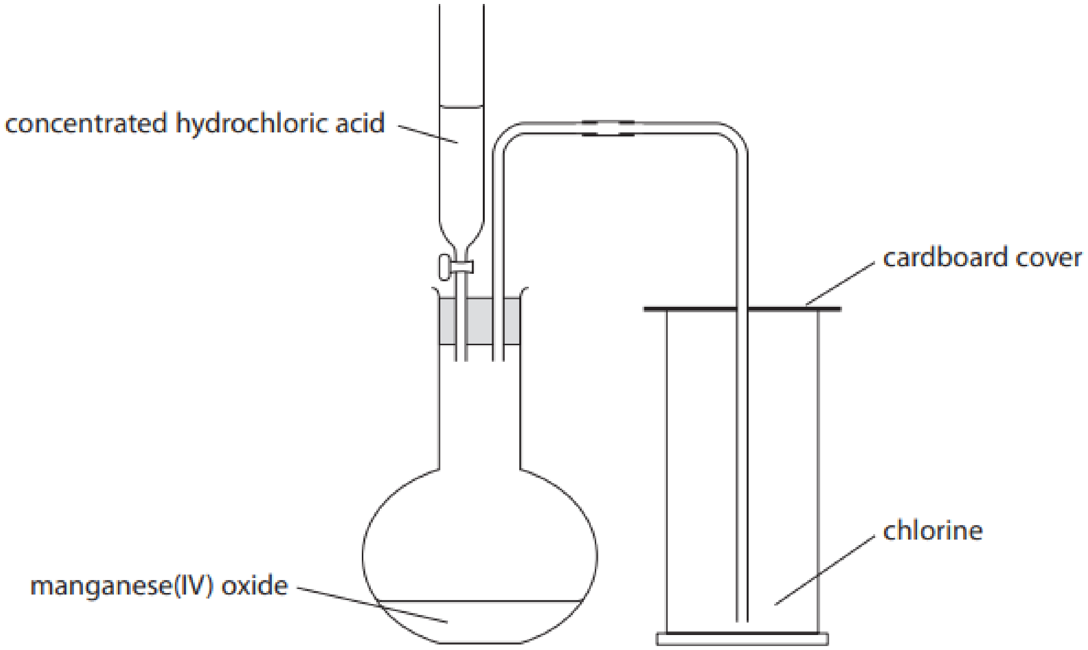
7) The table gives information about the first four elements in Group 7 of the Periodic Table



Element	Atomic number	Electronic configuration	Physical state at 20 °C	Colour at 20 °C
fluorine	9	2.7	gas	pale yellow
chlorine	17	2.8.7	gas	pale green
bromine	35	2.8.18.7	liquid	red-brown
iodine	53	2.8.18.18.7	solid	dark grey

a) Astatine (At) has an atomic number of 85 and is the fifth element in Group 7. It is possible to make predictions about astatine by comparison with the other elements in Group 7.
(i) How many electrons does an atom of astatine have in its outer shell?
(ii) What physical state and colour would you expect for astatine at 20°C?
(iii) Predict the formula of the compound formed between astatine and hydrogen. Suggest a name for this compound
(iv) Suggest how the reactivity of astatine compares to that of iodine. Explain your answer

8) a)Chlorine gas can be prepared by heating a mixture of concentrated hydrochloric acid and manganese(IV) oxide using this apparatus.



(i) Balance the equation for the reaction.

.... HCl(aq) + ... MnO_2(s) → .... $MnCl_2$(aq) + ... H_2O(l) + ... Cl_2(g)

(ii) State what you would observe when a piece of damp litmus paper is placed into the gas jar containing chlorine

b) Bromine is an element in Group 7 of the Periodic Table. (
(i) State the number of outer electrons in an atom of bromine.
(ii) Identify an element in Group 7 that is a solid at room temperature.
(iii) Identify an element in Group 7 that is more reactive than bromine.

9) A student carries out some experiments to investigate displacement reactions. She adds some halogen solutions to halide solutions and observes whether a reaction occurs. The table shows her results.

	Halogen solution added		
Halide solution	**bromine**	**chlorine**	**iodine**
lithium chloride	no reaction	(not done)	no reaction
sodium bromide	(not done)	reaction occurs	no reaction
potassium iodide	reaction occurs	reaction occurs	(not done)

(i) The table shows that she did not do three experiments. Suggest why she did not do these experiments.
(ii) The table shows that there was no reaction in three experiments. Why was there no reaction in these experiments?
(iii) The student writes this word equation for one of the experiments in which a reaction occurs

bromine + potassium iodide → potassium bromine + iodine

The name of one of the substances is incorrect. Write the correct name of this substance.

(iv) A reaction occurs when the student adds chlorine solution to potassium iodide solution. Complete the chemical equation for this reaction.

Cl_2 + ...KI → .... + ...

(v) All displacement reactions are examples of redox reactions. State the meaning of the term redox.