

2.2 Group 2

Atomic radius

Atomic **radius increases** down the Group.
As one goes down the group, the atoms have **more shells of electrons** making the atom bigger.

Melting points

Down the group the melting **points decrease**. The **metallic bonding** weakens as the **atomic size increases**. The distance between the positive ions and delocalized electrons increases. Therefore the electrostatic attractive forces between the **positive ions** and the **delocalized electrons** weaken.

1st ionisation energy

The outermost electrons are held more weakly because they are **successively further from the nucleus** in additional shells.

In addition, the outer shell electrons become **more shielded** from the attraction of the nucleus by the repulsive force of inner shell electrons

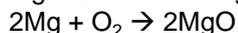
Group 2 reactions

Reactivity of group 2 metals **increases** down the group

Reactions with oxygen.

The group 2 metals will burn in oxygen.

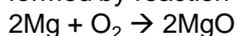
Mg burns with a bright white flame.



MgO is a white solid with a high melting point due to its ionic bonding.

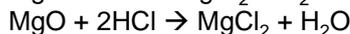
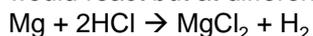
Mg will also react slowly with oxygen without a flame.

Mg ribbon will often have a thin layer of magnesium oxide on it formed by reaction with oxygen.



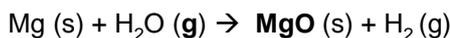
This needs to be cleaned off by emery paper before doing reactions with Mg ribbon.

If testing for reaction rates with Mg and acid, an un-cleaned Mg ribbon would give a false result because both the Mg and MgO would react but at different rates.

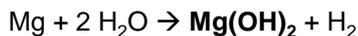


Reactions with water.

Magnesium burns in **steam** to produce **magnesium oxide** and hydrogen. The Mg would burn with a bright white flame.



Mg will also react with **warm water**, giving a different **magnesium hydroxide** product.



This is a much slower reaction than the reaction with steam and there is no flame.

Make sure you learn the difference between the reaction of Magnesium with steam and that of warm water

The hydroxides produced make the water alkaline (if they are soluble in water)

The other group 2 metals will react with **cold water** with increasing vigour down the group to form **hydroxides**.



One would observe:

- fizzing, (more vigorous down group)
- the metal dissolving, (faster down group)
- the solution heating up (more down group)
- and with calcium a white precipitate appearing (less precipitate forms down group)

Using Magnesium to Extract titanium

Titanium is a very useful metal because it is abundant, has a low density and is corrosion resistant – it is used for making strong, light alloys for use in aircraft for example.

Titanium is extracted by reaction with a more reactive metal (e.g. Mg, Na).

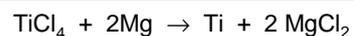
Steps in extracting Titanium

1. TiO_2 (solid) is converted to TiCl_4 (liquid) at 900°C :
2. The TiCl_4 is purified by fractional distillation in an Ar atmosphere.
3. The Ti is extracted by Mg in an Ar atmosphere at 500°C

Titanium is expensive because

1. The expensive cost of the Mg
2. This is a batch process which makes it expensive because the process is slower (having to fill up and empty reactors takes time) and requires more labour and the energy is lost when the reactor is cooled down after stopping
3. The process is also expensive due to the Ar, and the need to remove moisture (as TiCl_4 is susceptible to hydrolysis).
4. High temperatures required in both steps

Titanium cannot be extracted with carbon because titanium carbide (TiC) it is formed rather than titanium (similar reactions take place for vanadium, tungsten and molybdenum). Titanium cannot be extracted by electrolysis because it has to be very pure.



TiO_2 is converted to TiCl_4 as it can be purified by fractional distillation, TiCl_4 being molecular (liquid at room temperature) rather than ionic like TiO_2 (solid at room temperature).

This all makes titanium expensive even though it is a relatively abundant metal. It is only therefore used to a limited amount even though it has useful properties

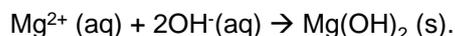
Solubility of Hydroxides

Group II hydroxides become **more soluble** down the group.

All Group II hydroxides when not soluble appear as white precipitates.

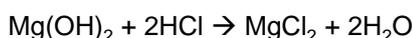
Magnesium hydroxide is classed as insoluble in water.

Simplest Ionic Equation for formation of Mg(OH)_2 (s)



A suspension of magnesium hydroxide in water will appear slightly alkaline (pH 9) so some hydroxide ions must therefore have been produced by a very slight dissolving.

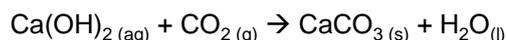
Magnesium hydroxide is used in medicine (in suspension as milk of magnesia) to neutralise excess acid in the stomach and to treat constipation.



It is safe to use because it is so weakly alkaline. It is preferable to using calcium carbonate as it will not produce carbon dioxide gas.

Calcium hydroxide is reasonably soluble in water. It is used in agriculture to neutralise acidic soils.

An aqueous solution of calcium hydroxide is called lime water and can be used a test for carbon dioxide. The limewater turns cloudy as white calcium carbonate is produced.



Barium hydroxide would easily dissolve in water. The hydroxide ions present would make the solution strongly alkaline.

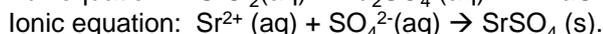


Solubility of Sulphates

Group II sulphates become **less soluble** down the group.

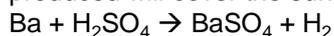
BaSO_4 is the least soluble.

An equation for the formation of the precipitate can be written as a full equation or simplest ionic equation



BaSO_4 is used in medicine as a 'Barium meal' given to patients who need x-rays of their intestines. The Barium absorbs the x-rays and so the gut shows up on the x-ray image. Even though Barium compounds are toxic it is safe to use here because of its low solubility.

If Barium metal is reacted with sulphuric acid it will only react slowly as the insoluble Barium sulphate produced will cover the surface of the metal and act as a barrier to further attack.



The same effect will happen to a lesser extent with metals going up the group as the solubility increases.

The same effect does not happen with other acids like hydrochloric or nitric as they form soluble group 2 salts.

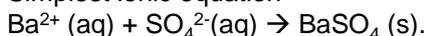
Testing for Presence of a Sulphate ion

BaCl_2 solution acidified with hydrochloric acid is used as a reagent to test for sulphate ions.

If acidified **Barium Chloride** is added to a solution that contains sulphate ions a **white precipitate** of Barium Sulphate forms.

Other anions should give a negative result which is no precipitate forming.

Simplest ionic equation



The hydrochloric acid is needed to react with carbonate impurities that are often found in salts which would form a white Barium carbonate precipitate and so give a false result. You could not use sulphuric acid because it contains sulphate ions and so would give a false positive result.



Fizzing due to CO_2 would be observed if a carbonate was present.

More on Insoluble salts and Precipitation reactions

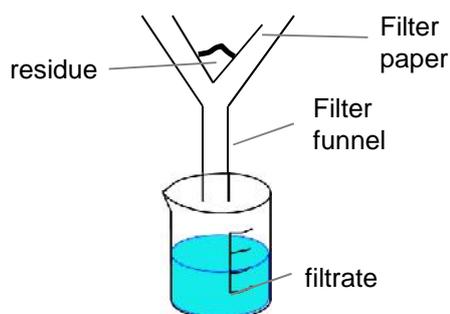
Insoluble salts can be made by mixing appropriate solutions of ions so that a **precipitate** is formed
 Barium nitrate (aq) + sodium sulphate (aq) → **Barium Sulphate (s)** + sodium nitrate (aq)
 These are called **precipitation** reactions. A **precipitate is a solid**

There are some common rules for solubility of salts. No syllabus requires these to be learnt but a good chemist does know them.

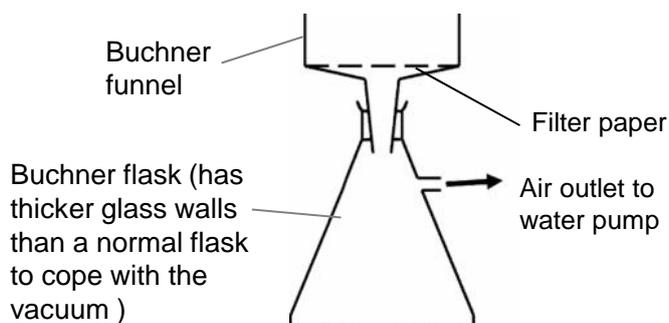
Soluble salts	Insoluble salts
All sodium, potassium and ammonium salts	
All nitrates	
Most chlorides, bromides, iodides	Silver, lead chlorides, bromides iodides
Most sulphates	Lead sulphate strontium and barium sulphate
Sodium, potassium and ammonium carbonates	Most other carbonates
Sodium, potassium and ammonium hydroxides	Most other hydroxides

When making an insoluble salt, normally the salt would be removed by **filtration**, washed with **distilled water to remove soluble impurities** and then **dried on filter paper**

Filtration



This is gravitational filtration.
 Use if small amounts of solid are formed.



This is vacuum filtration. The apparatus is connected to a water pump which will produce a vacuum. Use if larger amounts of solid are formed.

For both types of filtration apparatus AQA expect filter paper to be drawn on the diagram

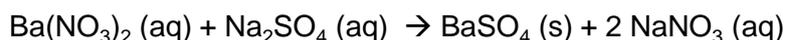
Writing Ionic equations for precipitation reactions

We usually write ionic equations to show precipitation reactions. Ionic equations only show the ions that are reacting and leave out spectator ions.

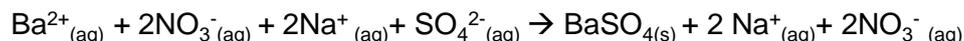
Spectator ions are ions that are

- Not changing state
- Not changing oxidation number

Take full equation



Separate (aq) solutions into ions



Cancel out spectator ions leaving the simplest ionic equation

