

## 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

Melting **points decrease** down the group. 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

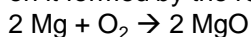
Reactivity of group 2 metals **increases** down the group.

#### Reactions with oxygen.

The group 2 metals will burn in oxygen.  
Mg burns with a bright **white flame**. The MgO appears as a **white powder**.  
 $2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO}$

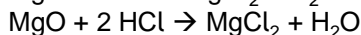
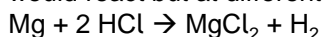
MgO is a white solid with a high melting point due to its ionic bonding.

Magnesium will also react slowly in oxygen without a flame. Magnesium ribbon will often have a thin layer of magnesium oxide on it formed by the reaction with oxygen in the air.



The magnesium oxide needs to be removed by emery paper before doing reactions with magnesium ribbon.

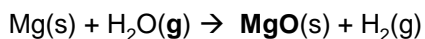
If testing for reaction rates with Mg and acid, an un-cleaned Mg ribbon would give a false result because both the Mg and MgO would react but at different rates.



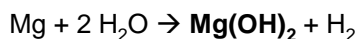
#### Reactions with water.

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

Magnesium reacts **in steam** to produce **magnesium oxide** and hydrogen. The Mg would burn with a bright **white flame**. The MgO appears as a **white powder**.

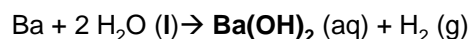
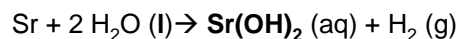
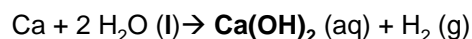


Mg will also react with **warm water**, giving a different **magnesium hydroxide** product.



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

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



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

One would observe:

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

## Using magnesium to extract titanium

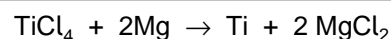
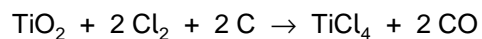
Titanium is a very useful metal because it is abundant, has a low density and is corrosion resistant – it is used for making strong, light alloys for use in aircraft.

Titanium is extracted by reaction with a more reactive metal (e.g. Magnesium).

Titanium cannot be extracted with carbon because titanium carbide (TiC) is formed rather than titanium.  
Titanium cannot be extracted by electrolysis because it has to be very pure.

### Steps in extracting titanium

1.  $\text{TiO}_2$  (solid) is converted to  $\text{TiCl}_4$  (liquid) at  $900^\circ\text{C}$ :
2. The  $\text{TiCl}_4$  is purified by fractional distillation in an argon atmosphere.
3. The Ti is extracted by Mg in an argon atmosphere at  $500^\circ\text{C}$



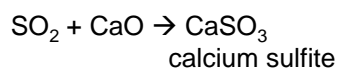
Titanium is expensive because:

1. The expensive cost of the magnesium
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 argon, and the need to remove moisture (because  $\text{TiCl}_4$  is susceptible to hydrolysis).
4. High temperatures required in both steps

$\text{TiO}_2$  is converted to  $\text{TiCl}_4$  as it can be purified by fractional distillation,  $\text{TiCl}_4$  being molecular (liquid at room temperature) rather than ionic like  $\text{TiO}_2$  (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.

Calcium oxide can be used to remove  $\text{SO}_2$  from the waste gases from furnaces (e.g. coal fired power stations) by flue gas desulfurisation. The gases pass through a scrubber containing **basic** calcium oxide which reacts with the **acidic** sulfur dioxide in a **neutralisation** reaction.



The calcium sulfite which is formed can be used to make calcium sulfate for plasterboard.

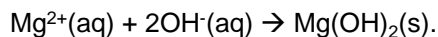
## Solubility of Hydroxides

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

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

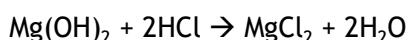
**Magnesium hydroxide** is classed as insoluble in water.

**Simplest ionic equation** for formation of  $\text{Mg}(\text{OH})_2(\text{s})$



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.

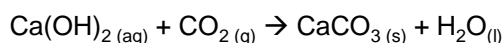


It is safe to use because it is so weakly alkaline. It is preferable to using calcium carbonate as it will not produce carbon dioxide gas.

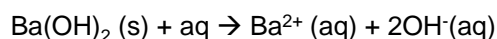
**Calcium hydroxide** is classed as partially soluble in water and will appear as a **white precipitate**. It is used in agriculture to neutralise acidic soils.

A suspension of calcium hydroxide in water will appear more alkaline (pH 11) than magnesium hydroxide as it is more soluble so there will be more hydroxide ions present in solution.

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.

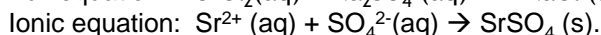
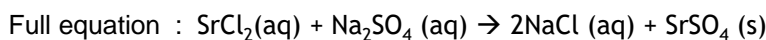


## Solubility of sulfates

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

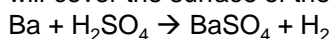
$\text{BaSO}_4$  is the least soluble.

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



$\text{BaSO}_4$  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 barium sulfate's low solubility means it is not absorbed into the blood.

If barium metal is reacted with sulfuric acid it will only react slowly, as the insoluble barium sulfate produced will cover the surface of the metal and act as a barrier to further attack.



The same effect happens to a lesser extent with metals going up the group as the solubility of the sulfates increases.

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

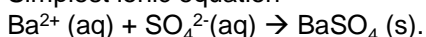
## Testing for presence of a sulfate ion

**$\text{BaCl}_2$  solution acidified with hydrochloric acid** is used as a reagent to test for sulfate ions.

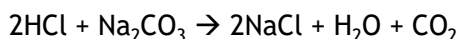
If acidified **barium chloride** is added to a solution that contains sulfate ions a **white precipitate** of barium sulfate forms.

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 sulfuric acid because it contains sulfate ions and so would give a false positive result.



Fizzing due to  $\text{CO}_2$  would be observed if a carbonate was present.

## More on Insoluble salts and Precipitation reactions

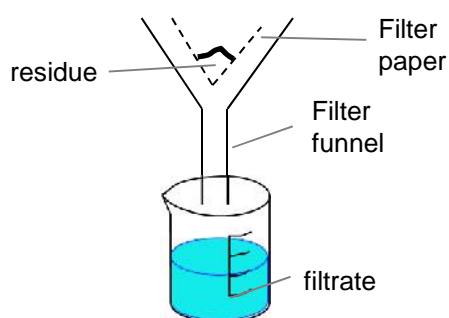
Insoluble salts can be made by mixing appropriate solutions of ions so that a **precipitate** is formed  
barium nitrate (aq) + sodium sulfate (aq) → **barium sulfate (s)** + sodium nitrate (aq)  
These are called **precipitation** reactions. A **precipitate is a solid**.

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

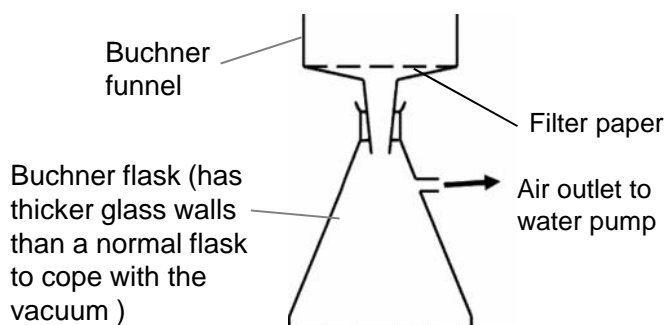
Soluble salts	Insoluble salts
All sodium, potassium and ammonium salts	
All nitrates	
Most chlorides, bromides, iodides	Silver, lead chlorides, bromides iodides
Most sulfates	Lead, strontium and barium sulfates
Sodium, potassium and ammonium carbonates	Most other carbonates
Sodium, potassium and ammonium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

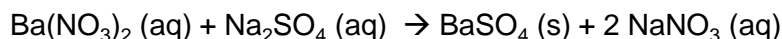
### Writing ionic equations for precipitation reactions

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

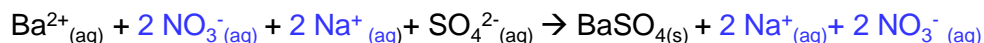
Spectator ions are ions that are

- not changing state
- not changing oxidation number

Take full equation



Separate aqueous solutions into ions.



Cancel out spectator ions leaving the simplest ionic equation.

