

10. 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

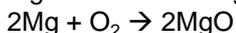
Group 2 reactions

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

Reactions with oxygen.

The group 2 metals will burn in oxygen.

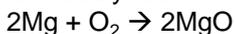
Mg burns with a bright white flame.



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

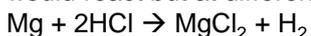
Mg will also react slowly with oxygen without a flame.

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



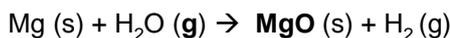
This needs to be cleaned off by emery paper before doing reactions with Mg ribbon.

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.

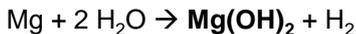


Reactions with water.

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**.

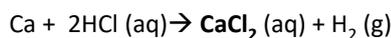


One would observe:

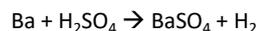
- fizzing, (more vigorous down group)
- the metal dissolving, (faster down group)
- the solution heating up (more down group)
- and with calcium a white precipitate appearing (less precipitate forms down group)

Reactions with acid

The group 2 metals will react with acids with increasing vigour down the group to form a salt and hydrogen



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.



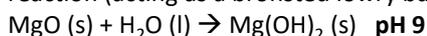
The same effect will happen to a lesser extent with metals going up the group as the solubility increases.

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

Reactions of the Oxides of Group 2 elements with water

Group 2 ionic oxides react with water to form hydroxides

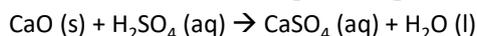
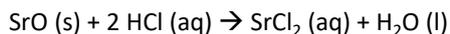
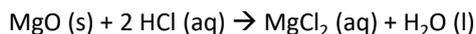
The ionic oxides are basic as the oxide ions accept protons to become hydroxide ions in this reaction (acting as a bronsted lowry base)



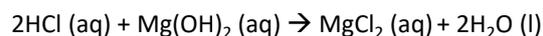
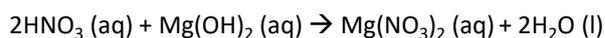
Mg(OH)₂ is only slightly soluble in water so fewer free OH⁻ ions are produced and so lower pH



Reactions of the Oxides of Group 2 elements with Acids



Reactions of the hydroxides of Group 2 elements with Acids



Reactions of the carbonates of Group 2 elements with Acids



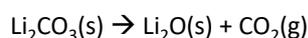
Thermal decomposition of Group 2 carbonates

Group 2 carbonates decompose on heating to produce group 2 oxides and carbon dioxide gas.



Group 2 carbonates are more thermally stable as you go down the group. As the cations get bigger they have less of a polarising effect and distort the carbonate ion less. The C-O bond is weakened less so it less easily breaks down

Group 1 carbonates do not decompose with the exception of Lithium. As they only have +1 charges they don't have a big enough charge density to polarise the carbonate ion. Lithium is the exception because its ion is small enough to have a polarising effect

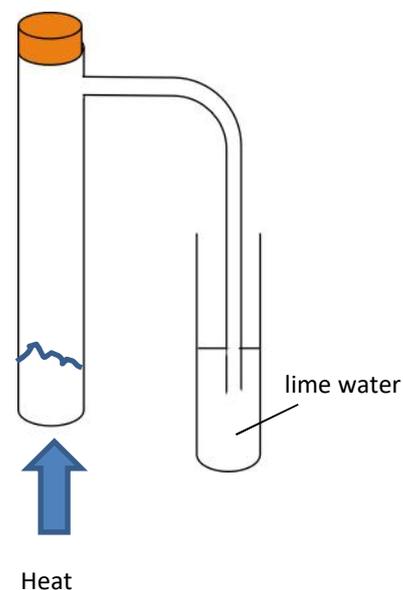


There are a number of experiments that can be done to investigate the ease of decomposition.

One is to heat a known mass of carbonate in a side arm boiling tube and pass the gas produced through lime water. Time for the first permanent cloudiness to appear in the limewater. Repeat for different carbonates using the same moles of carbonate/same volume of limewater/same Bunsen flame and height of tube above flame.

Thermal decomposition is defined as the use of heat to break down a reactant into more than one product

The ease of thermal decomposition **decreases** down the group



Thermal decomposition of Group 2 Nitrates

Group 2 nitrates decompose on heating to produce group 2 oxides, oxygen and nitrogen dioxide gas.

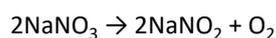
You would observe brown gas evolving (NO_2) and the white nitrate solid is seen to melt to a colourless solution and then resolidify



The ease of thermal decomposition **decreases** down the group

The explanation for change in thermal stability is the same as for carbonates. Magnesium nitrate decomposes the easiest because the Mg^{2+} ion is smallest and has the greater charge density. It causes more polarisation of the nitrate anion and weakens the N—O bond

Group 1 nitrates do not decompose in the same way as group 2 with the exception of Lithium nitrate. They decompose to give a Nitrate (III) salt and oxygen.



Sodium Nitrate(V) Sodium Nitrate(III)

Lithium nitrate decomposes in the same way as group 2 nitrates



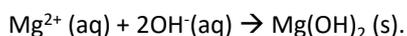
Solubility of Hydroxides

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

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

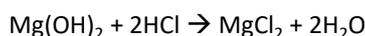
Magnesium hydroxide is classed as insoluble in water.

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.

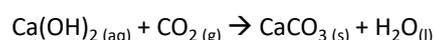


It is safe to use because it is so weakly alkaline. It is preferable to using calcium carbonate as it will not produce carbon dioxide gas.

Calcium hydroxide is 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 as 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.

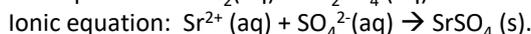
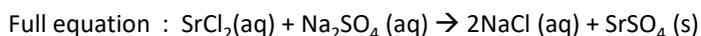


Solubility of Sulphates

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

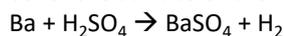
BaSO_4 is the least soluble.

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.

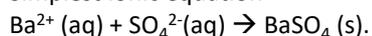
Testing for Presence of a Sulphate ion

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

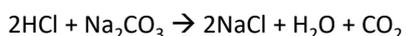
If acidified **Barium Chloride** is added to a solution that contains sulphate ions a **white precipitate** of Barium Sulphate forms.

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.

Explaining trends in solubility

Hydroxides

Group 2 hydroxides become more soluble down the group.

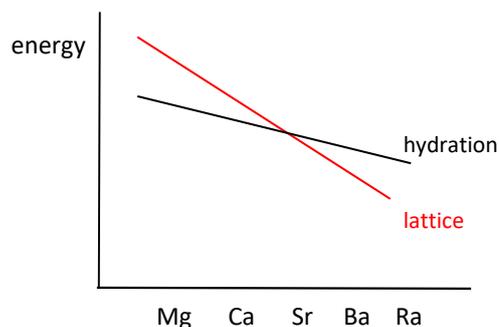
This can be explained by changes in the lattice enthalpy and hydration enthalpy

Lattice enthalpy drops down the group because the cations get bigger leading to a weaker ionic attraction

Water molecules are more strongly attracted to smaller ions with a larger charge.

Hydration enthalpy drops because the cations becomes bigger and less polarising so they attract the water molecules less.

The hydroxides become more soluble because the lattice enthalpy drops more than the hydration enthalpy



For magnesium hydroxide the lattice enthalpy is larger than the hydration enthalpy leading to an endothermic enthalpy of solution which makes it less likely to dissolve

For barium hydroxide the lattice enthalpy is smaller than the hydration enthalpy leading to an exothermic enthalpy of solution which makes it more likely to dissolve

Sulfates

Group 2 sulfates become **less** soluble down the group.

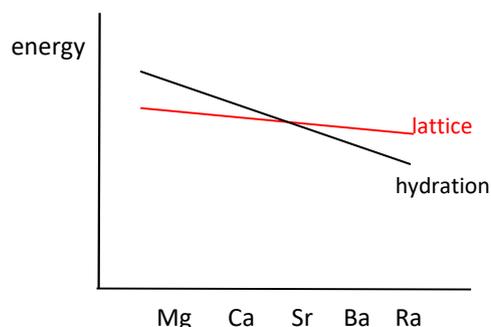
This can be explained by same changes in the lattice enthalpy and hydration enthalpy as the above case

Lattice enthalpy drops down the group because the cations get bigger leading to a weaker ionic attraction. The key difference between sulfates and hydroxides is the sulfate ion is much larger than the metal ions (and hydroxide ion). The decreasing size of the cation has less of an effect in this case and the lattice enthalpy falls less going down the group

Water molecules are more strongly attracted to smaller ions with a larger charge.

Hydration enthalpy drops because the cations becomes bigger and less polarising so they attract the water molecules less.

The hydroxides become less soluble because the lattice enthalpy drops less than the hydration enthalpy



For magnesium sulfate the lattice enthalpy is smaller than the hydration enthalpy leading to an exothermic enthalpy of solution which makes it more likely to dissolve

For barium sulfate the lattice enthalpy is larger than the hydration enthalpy leading to an endothermic enthalpy of solution which makes it less likely to dissolve

More on Insoluble salts and Precipitation reactions

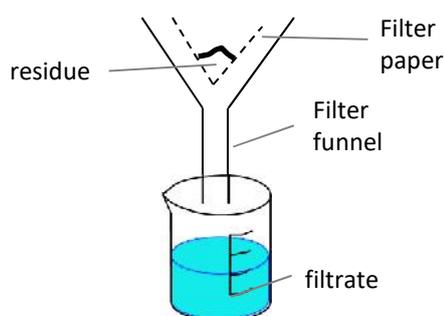
Insoluble salts can be made by mixing appropriate solutions of ions so that a **precipitate** is formed
 Barium nitrate (aq) + sodium sulphate (aq) → **Barium Sulphate (s)** + sodium nitrate (aq)
 These are called **precipitation** reactions. A **precipitate is a solid**

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

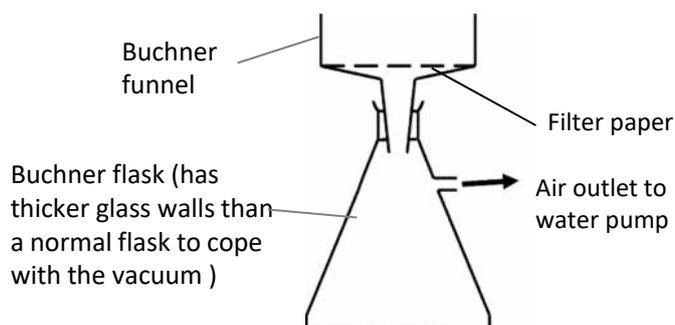
Soluble salts	Insoluble salts
All sodium, potassium and ammonium salts	
All nitrates	
Most chlorides, bromides, iodides	Silver, lead chlorides, bromides iodides
Most sulphates	Lead sulphate strontium and barium sulphate
Sodium, potassium and ammonium carbonates	Most other carbonates
Sodium, potassium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

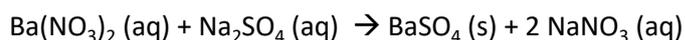
Writing Ionic equations for precipitation reactions

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

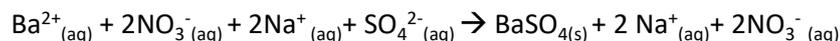
Spectator ions are ions that are

- Not changing state
- Not changing oxidation number

Take full equation



Separate (aq) solutions into ions



Cancel out spectator ions leaving the simplest ionic equation

