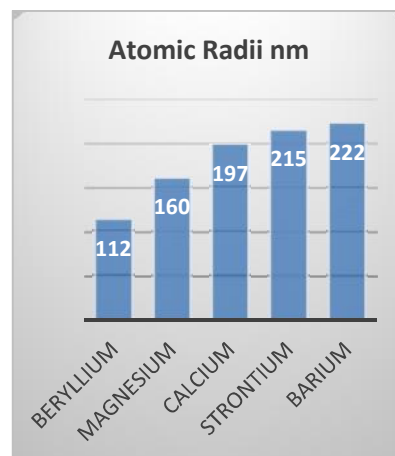


## Group 2: The Alkaline Earth Metals

### Atomic radius

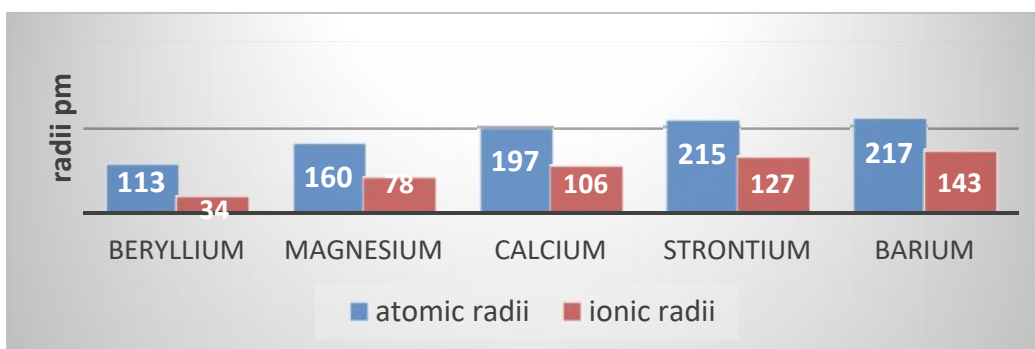
Atomic **radius increases** down Group 2.

As one goes down the group, the atoms have **more shells of electrons** making the atom bigger and there is a decrease in effective nuclear charge with successive elements because of increased screening, and so the electrons are less firmly held to the nucleus.



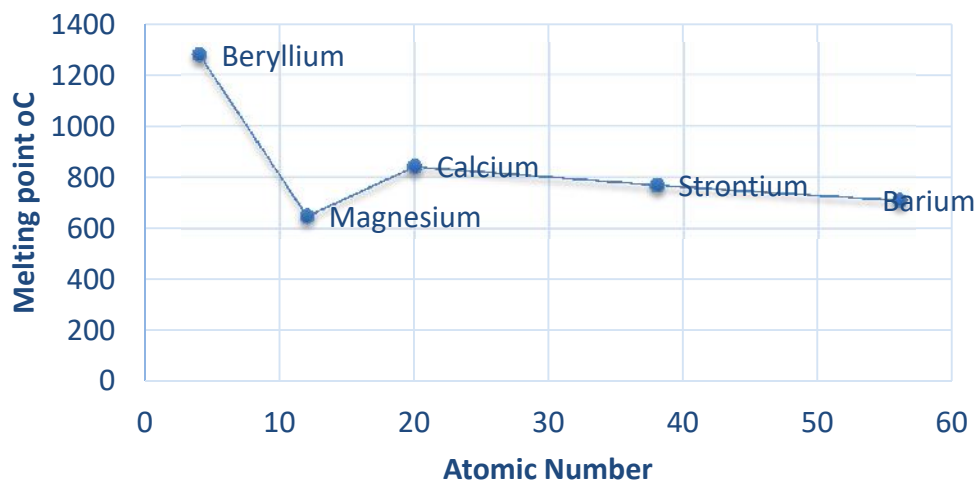
### Ionic Radius

The ionic radius of a group 2 metal will be smaller than the corresponding atomic radius, because it has lost one shell of electrons and the remaining protons hold onto the remaining electrons more strongly and pulls them in more tightly.



### Melting points

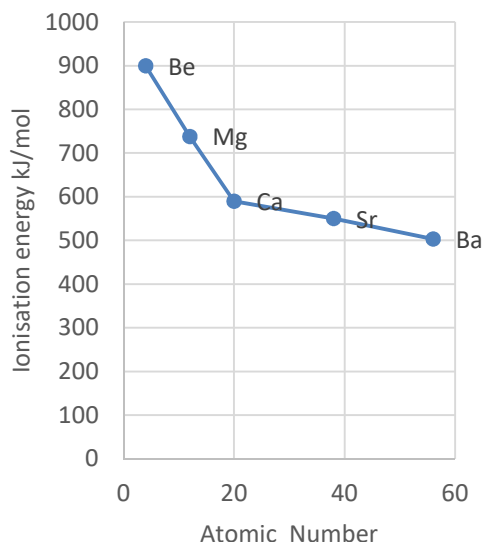
The melting points decrease going down the group. The metallic bonding weakens as the atomic size increases. The distance between the positive ions and delocalized electrons increases. Therefore the electrostatic attractive forces between the positive ions and the delocalized electrons weaken.



### 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

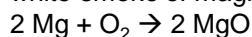
Reactivity of group 2 metals **increases** down the group

The reactivity increases down the group as the atomic radii increase there is more shielding. The nuclear attraction decreases and it is easier to remove outer electrons and so cations form more easily.

### Reactions with oxygen.

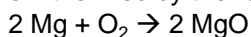
The group 2 metals will burn in oxygen.

Magnesium burns with a bright white flame and produces a white smoke of magnesium oxide.



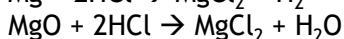
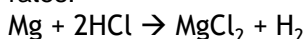
MgO is a white solid with a high melting point due to its ionic bonding.

Magnesium will also react slowly with 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.



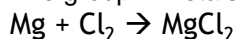
The magnesium oxide needs to be cleaned off by emery paper before doing reactions with magnesium ribbon.

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



### Reactions with halogens

The group 2 metals will react with halogens to produce ionic halide solids.

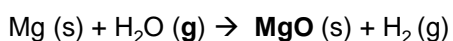


All group 2 halides (except beryllium) are white, ionically bonded, solids.

## Reactions with water.

### Reaction of magnesium with steam

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

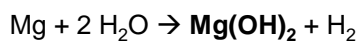


In the picture on the right, the hydrogen produced is set on fire at the end of the tube in the bung.



### Reaction of Magnesium with Water

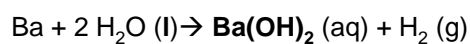
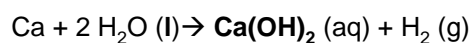
Mg will also react with warm water, giving **magnesium hydroxide** as the product. In the photo on the right magnesium powder was used and there was a slow fizzing.



This is a much slower reaction than the reaction of magnesium 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

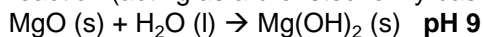
One would observe:

- 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)

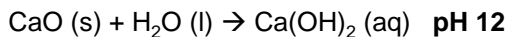
## Reactions of the oxides of group 2 elements with water

**Group 2 ionic oxides** react with water to form hydroxides

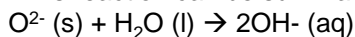
The ionic oxides are basic as the oxide ions accept protons to become hydroxide ions in this reaction (acting as a bronsted lowry base)



$\text{Mg(OH)}_2$  is only slightly soluble in water so fewer free  $\text{OH}^-$  ions are produced and so lower pH

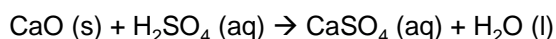
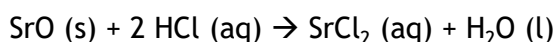
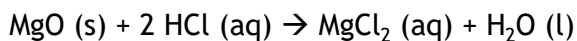


This reaction can be summarised ionically as

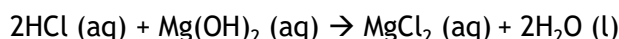
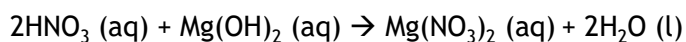


The pH of an aqueous solution of a hydroxide depends on the extent to which the metal hydroxide can be split to produce independent ions. The greater the ease of producing  $\text{OH}^-$  ions, the more alkaline is the solution formed

## Reactions of the oxides of group 2 elements with acids



## Reactions of the hydroxides of group 2 elements with acids



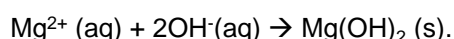
### Solubility of hydroxides

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

All insoluble group II hydroxides are white solids.

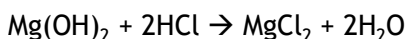
Magnesium hydroxide is classed as insoluble in water.

**Simplest Ionic Equation** for formation of  $\text{Mg(OH)}_2$  (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.

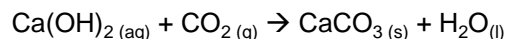
Magnesium hydroxide is used in medicine (in suspension as milk of magnesia) to neutralise excess acid in the stomach and to treat constipation.



It is safe to use because it so weakly alkaline.

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

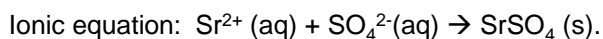


## Solubility of sulfates

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

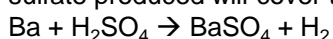
BaSO<sub>4</sub> is the least soluble.

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



Barium sulfate (BaSO<sub>4</sub>) 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, barium sulfate is safe to use here because of its low solubility so does not get absorbed into the body.

If barium metal is reacted with sulfuric acid, the barium will only react slowly because 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.

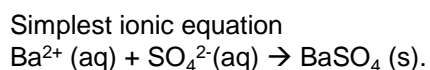
The same effect does not happen with other acids like hydrochloric or nitric as they form soluble group 2 salts.

## Testing for presence of a sulfate

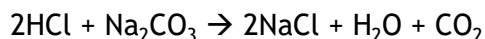
**Acidified BaCl<sub>2</sub> solution** is used as a reagent to test for sulphate ions.

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

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



The 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 CO<sub>2</sub> would be observed if a carbonate was present.



**white precipitate** of barium sulfate

## Explaining trends in solubility

### Hydroxides

Group 2 hydroxides become more soluble down the group.

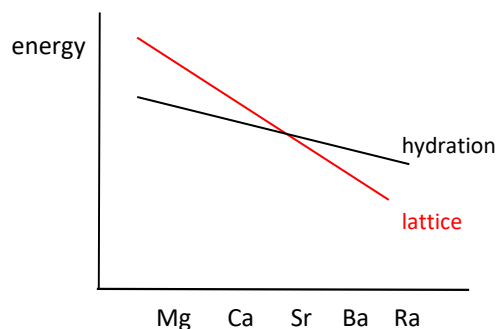
This can be explained by changes in the lattice enthalpy and hydration enthalpy.

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.

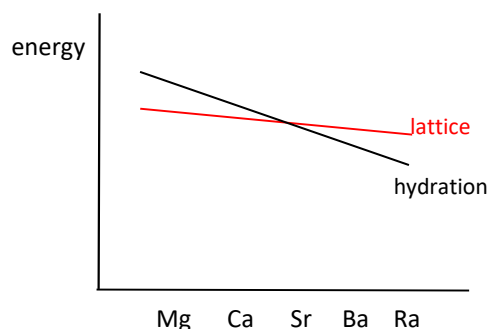
This can be explained by same changes in the lattice enthalpy and hydration enthalpy as the above case.

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

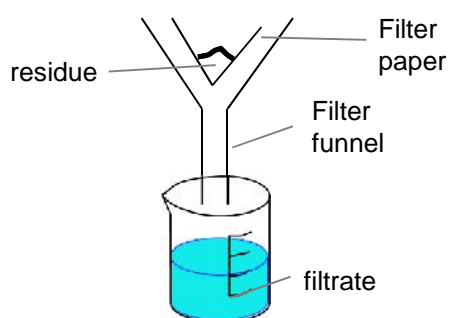
Insoluble salts can be made by mixing appropriate solutions of ions so that a **precipitate** is formed  
 barium nitrate (aq) + sodium sulfate (aq) → **barium sulfate (s)** + sodium nitrate (aq)  
 These are called **precipitation** reactions. A **precipitate is a solid**

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

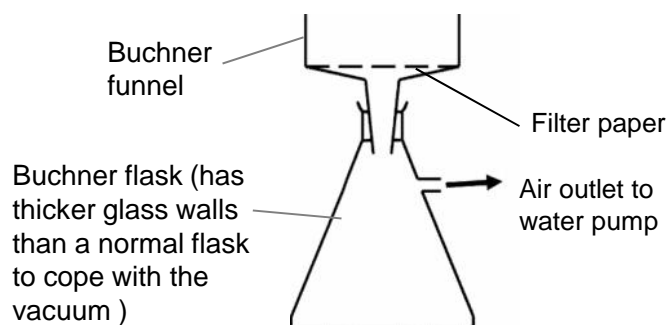
Soluble salts	Insoluble salts
All sodium, potassium and ammonium salts	
All nitrates	
Most chlorides, bromides, iodides	Silver, lead chlorides, bromides iodides
Most sulfates	Lead, strontium and barium sulfates
Sodium, potassium and ammonium carbonates	Most other carbonates
Sodium, potassium and ammonium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

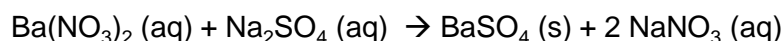
### Writing ionic equations for precipitation reactions

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

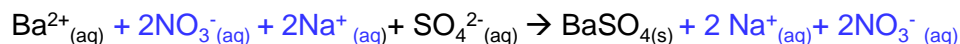
Spectator ions are ions that are:

- Not changing state
- Not changing oxidation number

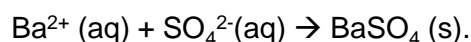
Take full equation



Separate (aq) solutions into ions



Cancel out spectator ions leaving the simplest ionic equation



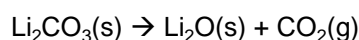
## Thermal decomposition of group 2 carbonates

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



Group 2 carbonates become more thermally stable going 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 don't decompose with the exception of lithium. Group 1 ions only have +1 charges so they do not 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.

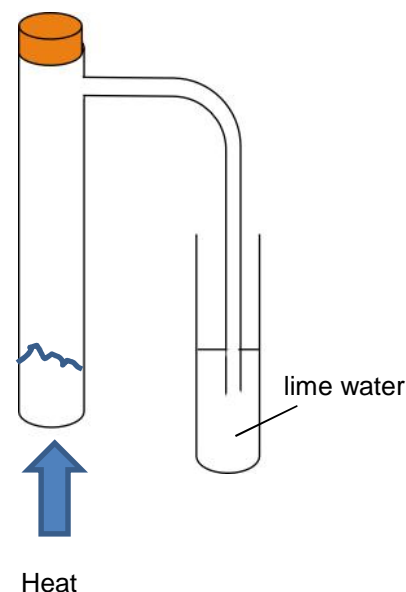


There are a number of experiments that can be done to investigate the ease of decomposition.

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

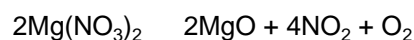
The ease of thermal decomposition **decreases** down the group



## Thermal decomposition of group 2 nitrates

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

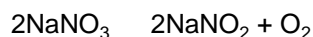
You would observe brown gas evolving ( $\text{NO}_2$ ) and the white nitrate solid is seen to melt to a colourless solution and then re-solidify.



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

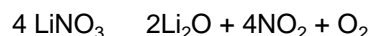
The explanation for change in thermal stability is the same as for carbonates. Magnesium nitrate decomposes the easiest because the  $\text{Mg}^{2+}$  ion is smallest and has the greater charge density. It causes more polarisation of the nitrate anion and weakens the N-O bond

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





## Flame tests

The elements of group 1 and some of group 2 emit brightly coloured flames when they are heated. These coloured flames can be used to identify the metal ion present in an analysis

### Method for Flame test

Use a **nichrome wire** ( nichrome is an unreactive metal and will not give out any flame colour)

Clean the wire by dipping **in concentrated hydrochloric acid** and then heating in Bunsen flame

If the sample is not powdered then grind it up.

Dip wire in solid and put in Bunsen flame and observe flame

Lithium : Scarlet red

Sodium : Yellow

Potassium : lilac

Rubidium : red

Caesium: blue

Magnesium: no flame colour (energy emitted of a wavelength outside visible spectrum)

Calcium: brick red

Strontium: red

Barium: apple green

### Explanation for occurrence of flame

In a flame test the heat causes the electron to move to a higher energy level.

The electron is unstable at the higher energy level and so drops back down. As it drops **back down from the higher to a lower energy level**, energy is emitted in the form of visible light energy with the wavelength of the observed light

## Group 2 Questions- physical trends, reactions and acid base properties

- 1 a) State and explain the trend in melting point of group 2 metals down the group  
b) State and explain the trend in atomic radius of group 2 metals down the group  
c) State and explain the trend in the first ionisation energy of group 2 metals down the group
- 2 a) State and explain the trend in reactivity with water of the Group 2 metals from Mg to Ba  
b) Write equations for the reaction of barium with water and calcium with water.
- 3 a) Write an equation for the reaction of magnesium with steam  
b) Write an equation for the reaction of magnesium with water  
c) State two differences in observations between these two reactions.
- 4 Group 2 oxides and hydroxides are bases. Write equations for the neutralisation reactions of
  - a)  $\text{MgO} + \text{HCl}$
  - b)  $\text{SrO} + \text{H}_2\text{SO}_4$
  - c)  $\text{Ba}(\text{OH})_2 + \text{HNO}_3$
- 5 a) The water produced by an industrial plants is acidic due to the presence of hydrochloric acid. Lime,  $\text{Ca}(\text{OH})_2$ , is added to neutralise this acid. Write an equation for this reaction.  
b) Lime is used because it is relatively inexpensive and available in large quantities. Lime can also be used to remove  $\text{SO}_2$  from power station emissions. Write an equation for the reaction of Lime and  $\text{SO}_2$ . Why is it important to remove the  $\text{SO}_2$  from the emissions? Give a use for the product of this reaction.
- 6 a) When acid spillages occur, calcium carbonate is often used to neutralise the acid. Apart from cost, explain why the use of calcium carbonate is more appropriate than using sodium hydroxide.  
b) Write an equation for the reaction of calcium carbonate with hydrochloric acid.
- 7 The magnesium used in a laboratory experiment was supplied as a ribbon. The ribbon was stored in an open plastic bag exposed to the air. Explain why it is important to clean the surface of this magnesium ribbon when investigating the rate of its reaction with hydrochloric acid.

## Group 2 Questions- Solubility

- 1(a) State the trends in solubility of the hydroxides and of the sulfates of the Group II elements Mg–Ba.  
(b) Give the formula of the least soluble hydroxide of the Group 2 metals from Mg to Ba
- 2 Magnesium hydroxide can be used to neutralise excess stomach acid .
  - (a) Explain why it would be preferable to use magnesium hydroxide rather than barium hydroxide for this use.
  - (b) What advantage does magnesium hydroxide have over calcium carbonate in neutralising stomach acid?
  - (c) Explain the reason why sugars are often added to antacid tablets.
- 3 State how barium sulfate is used in medicine. Explain why this use is possible, given that solutions containing barium ions are poisonous.
- 4 Acidified barium chloride solution is used as a reagent to test for sulfate ions.
  - (a) Generally hydrochloric acid is used to acidify the solution. Why is it used?
  - (b) State why sulfuric acid should not be used to acidify the barium chloride.
  - (c) Write the simplest ionic equation for the reaction that occurs when acidified barium chloride solution is added to a solution containing sulfate ions.

**5** The following pair of compounds **BaCl<sub>2</sub>(aq)** and **MgCl<sub>2</sub>(aq)** can be distinguished by observing what happens in test-tube reactions.

Give a suitable aqueous reagent that could be added separately to each compound.

Describe what you would observe in each case.

**6** An aqueous solution of sodium chloride may be distinguished from an aqueous solution of sodium sulfate using a simple chemical test.

**(a)** Identify a reagent for this test.

**(b)** State the observations you would expect to make if the reagent identified in part a is added to a separate sample of each solution. Write an equation for any reaction which occurs.

**7** This question concerns the chemistry of the Group II metals Mg to Ba.

An aqueous solution of a Group II metal chloride, **XCl<sub>2</sub>**, forms a white precipitate when dilute aqueous sodium hydroxide is added. A separate sample of the solution of **XCl<sub>2</sub>** does **not** form a precipitate when dilute aqueous sodium sulfate is added.

An aqueous solution of a different Group II metal chloride, **YCl<sub>2</sub>**, does **not** form a precipitate when dilute aqueous sodium hydroxide is added. A separate sample of the solution of **YCl<sub>2</sub>** forms a white precipitate when dilute aqueous sodium sulfate is added.

Suggest identities for the Group II metals **X** and **Y**. Write ionic equations, including state symbols, for the reactions

**8** Barium metal react vigorously with hydrochloric acid but with sulfuric acid the reaction is slow.

Write equations for the reactions of Barium with both acids

Explain why Barium reacts more slowly with sulfuric acid

**9** Pure magnesium reacts completely with an excess of dilute sulfuric acid.

The reaction of pure calcium with an excess of dilute sulfuric acid is very rapid to start with and then the reaction slows down and stops before all of the calcium has been used up.

Use your knowledge of the solubility of Group 2 sulfates to explain why these reactions of magnesium and calcium with dilute sulfuric acid are different.

**10**

Both strontium carbonate and strontium sulfate are white solids which are insoluble in water.

Strontium carbonate reacts with hydrochloric acid to produce a solution of strontium chloride. Strontium sulfate does not react with hydrochloric acid.

Describe how you would obtain strontium sulfate from a mixture of strontium carbonate and strontium sulfate.