10.01 Group 2: physical properties

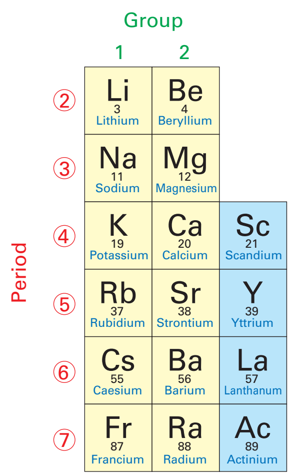
OBJECTIVES

already from AS level, you

- can explain the trend atomic radius of the period 3 elements
- understand how ionization energies give evidence for energy levels
- understand that metallic bonding involves a lattice of positive ions surrounded by delocalized electrons

and after this spread you should

- understand the trends in atomic radius, first ionization energy, and melting point of the elements in group 2



Group 2 is in the s block of the Periodic Table. It contains the elements Be to Ra.

Shielding

The number of protons increases down group 2. You might expect the attraction for the electrons in the highest occupied energy level to increase, leading to a decrease in the atomic radius. The electrons in the lower occupied energy levels become closer to the nucleus, but they shield the electrons in the highest occupied energy level from the attraction of the nucleus. So, the main factor in determining atomic redius becomes the number of occupied energy levels.

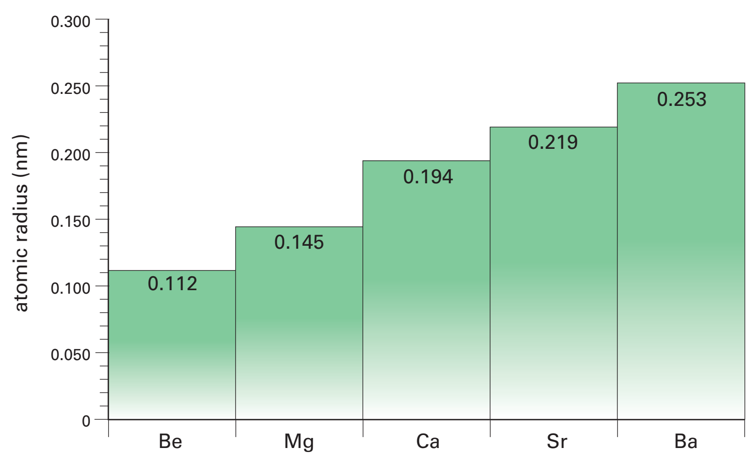
The elements of group 2 are also called the **alkaline earth metals**. This is because they form basic oxides. The group 2 elements have different uses, related to their physical and chemical properties.

- Beryllium alloys and magnesium alloys are light and strong. They are used in aircraft, spacecraft and missiles.
- Calcium carbonate is used in the manufacture of cement, concrete, and steel.
- Strontium compounds burn with a bright crimson flame. They are used in tracer bullets, fireworks, and signal flares.
- Many barium compounds are poisonous. Barium carbonate is used as rat poison. Barium sulfate is safe to use in medical X-ray photographs because it is insoluble and not absorbed.

The elements of group 2. From left to right: beryllium, magnesium, calcium, strontium, and barium. Radium is very radioactive and has little everyday practical use.

Trend in atomic radius

The atomic radius increases as you go down group 2. This is because there are more filled energy levels between the nucleus and the electrons in the highest occupied energy level.

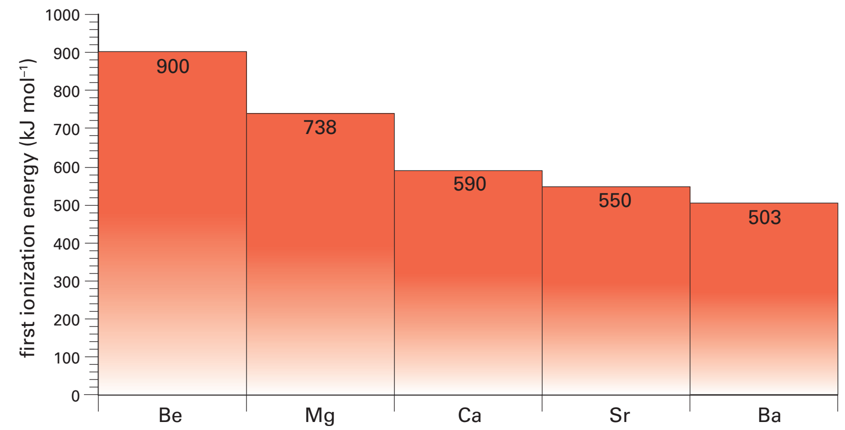


Atomic radius increases down group 2.

Trend in first ionization energy

The first ionization energy decreases as you go down group 2. In each case, an electron from the highest occupied a sub-level is being removed.

As you go down group 2, each successive element has more occupied energy levels. The distance between the outer electrons and the nucleus increases. Electrons in higher energy levels are less strongly attracted to the nucleus than electrons closer to it. So, even though the nuclear charge increases down the group, less energy is needed to remove an outer electron.



First ionization energy decreases down group 2.

First ionization energy

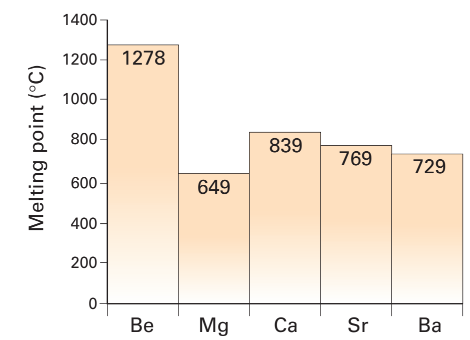
The first ionization energy is the energy needed to remove one electron from a gaseous atom. The general equation for this process is:

$M(g) \rightarrow M^+(g) + e^-$

where M stands for the chemical symbol of the element.

Trend in melting point

The melting point generally decreases as you go down group 2. The radius of the metal ions increases going down group 2 so their charge density decreases. The force of attraction between the metal ions and the delocalized electrons decreases, reducing the strength of the metallic bonding.



Melting point T_m generally decreases down group 2.

Magnesium

You may have noticed that magnesium has a lower melting point than might be expected. One possible explanation involves the structures of the solid metals. Beryllium and magnesium have a different structure from the other elements in the group. They have a hexagonal close-packed structure, but calcium and strontium have a face-centred cubic structure, and barium has a body-centred cubic structure. But magnesium also has a lower boiling point than expected. Boiling involves the state change from liquid to gas, so the different solid structures do not seem a likely explanation.

There should be two moles of delocalized electrons per mole of metal atoms in the group 2 elements. If there were fewer than two moles of delocalized electrons per mole of magnesium atoms, the metallic bonding would be weaker than expected, leading to a lower melting point. Is there any evidence for fewer delocalized electrons in magnesium? Metals conduct electricity because they have delocalized electrons, and magnesium has a lower electrical conductivity than beryllium or calcium.

Check your understanding

1. Describe and explain the trends down group 2 in
   a atomic radius
   b first ionization energy
   c melting point

10.02 Group 2: reactions with water

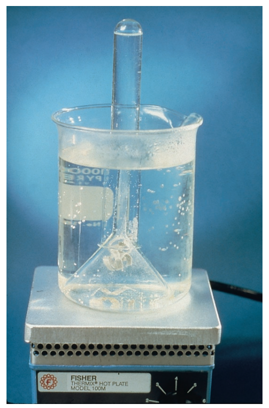
OBJECTIVES

already from GCSE, you know that

- the group 1 elements react with water to form alkaline hydroxides and hydrogen gas
- the group 1 elements become more reactive as you go down the group

and after this spread you should know

- the reactions of the group 2 elements magnesium to barium with water, and recognize the trend



Magnesium reacts very slowly with water. Hydrogen is collected over several days by upward displacement through an inverted filter funnel.

Testing for hydrogen

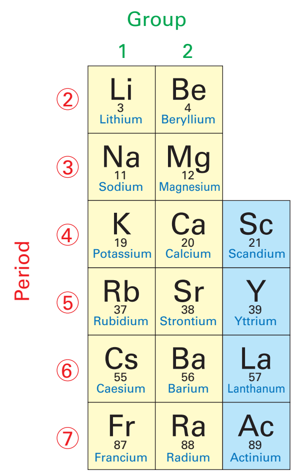
There is a simple laboratory test for the presence of hydrogen. A lighted wooden splint will ignite a test tube of hydrogen gas with a popping sound.

Calcium reacts steadily with water.

Beryllium at the top of group 2 does not react with water or steam, even if it is heated until it glows. But the other group 2 elements react with water to produce the metal hydroxide and hydrogen:

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

The reactions become more vigorous as you go down the group.



Group 2 is in the s block of the Periodic Table. It contains the elements Be to Ra.

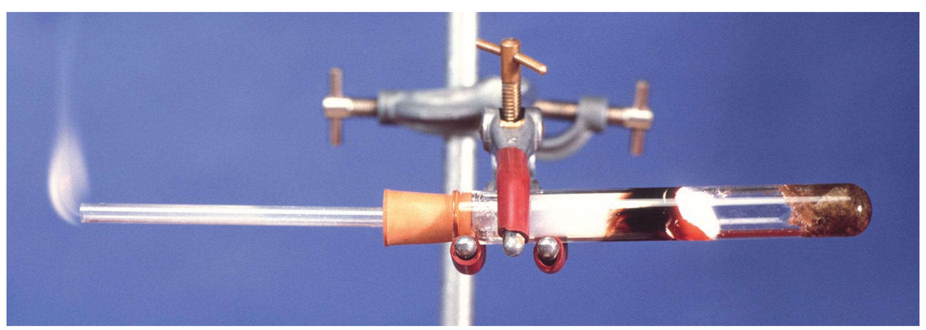
Magnesium

Magnesium burns vigorously with a bright white flame, so people often expect it to react vigorously with water, too. Instead, it only reacts very slowly with water. It may take several days to collect enough hydrogen to test. The magnesium hydroxide produced is sparingly soluble. Just enough dissolves to produce an alkaline solution.

Magnesium and steam

Magnesium burns with a bright white light when it is heated in steam. The hydrogen gas can be led out through a tube and, with care, it can be ignited. A white solid, magnesium oxide, is produced in this reaction instead of magnesium hydroxide.

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



Magnesium reacts vigorously with steam.

Calcium

Calcium reacts steadily with cold water. The reaction is exothermic, and the rate of reaction increases as the reaction mixture heats up. Calcium

hydroxide is slightly soluble and produces an alkaline solution. This usually looks cloudy white because undissolved calcium hydroxide forms a suspension.

Strontium and barium

Strontium and barium react with water more quickly than magnesium and calcium do. Barium, near the bottom of group 2, reacts immediately with water and produces hydrogen gas very quickly. Strontium and barium are stored in oil, just like sodium and other elements in group 1.

Strontium is stored in oil to keep air and water away.

Barium reacts quickly with water.

Explaining the trend

The calculated standard enthalpy changes for the reactions between the group 2 elements and water are very similar to each other. This is also true of beryllium, even though this element does not react with water. So another factor must be responsible for the different reactivities.

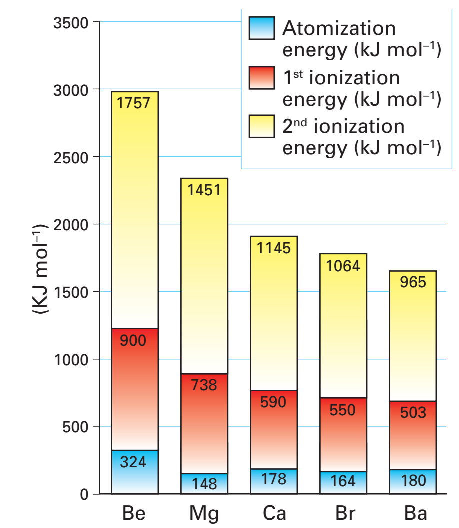
group 2 element	$\Delta H^{\ominus}$ (kJ mol^{-1})
Be	–331
Mg	–353
Ca	–415
Sr	–387
Ba	–373

Standard enthalpy changes for the reactions of the Group 2 elements with water.

The group 2 hydroxides are ionic compounds. They contain the metal ion M^{2+} and the hydroxide ion OH^-. The metal atoms become metal ions in the reaction with water. You can calculate the energy needed to do this. It involves three processes:

- **atomization energy (the energy needed to break up the metal lattice)**
- **first ionization energy (for the process $M(g) \rightarrow M^+(g) + e^-$)**
- **second ionization energy (for the process $M^+(g) \rightarrow M^{2+}(g) + e^-$)**

The reactions between the group 2 metals and water are exothermic. These three processes are all endothermic. They represent an activation energy that must be overcome for the reactions to happen. The graph shows the standard enthalpy changes for the formation of the M^{2+} ions. Notice that the energy needed decreases down the group. The reactions happen faster when the activation energy is lower.



The standard enthalpy changes for forming M^{2+} ions from group 2 atoms.

Check your understanding

1. What is the trend in the reactivity of group 2 elements towards water?
2. Describe, with the help of relevant equations, the following reactions:
   a the reaction between calcium and water
   b the reaction between magnesium and steam
3. a Use the bar chart to estimate the standard enthalpy change for forming Ra^{2+} ions from radium atoms.
   b Predict what you would observe if a piece of radium were added to water, and explain how you answer to part **a** helps you to do this.

10.03 Group 2: hydroxides

OBJECTIVES

already from AS Level, you know

- the reactions of the group 2 elements magnesium to barium with water, and can recognize the trend

and after this spread you should know

- the relative solubilities of the hydroxides of the elements magnesium to barium
- the use of magnesium hydroxide in medicine
- the use of calcium hydroxide in agriculture

Magnesium hydroxide is used as a medicine to treat indigestion, heartburn, and constipation.

Solubility

The solubility of a substance is a measure of how much of it will dissolve in a certain solvent. It is the maximum mass of the substance that will dissolve under stated conditions. It is not how quickly a solute dissolves.

There are several factors that can affect the solubility of an individual substance in a solvent, including

- the volume of the solvent
- the temperature of the solvent

More solute will dissolve in a larger volume of solvent, and the solubility of most solids increases as the temperature increases. To take these factors into account, solubility is given for a specified temperature and volume of solvent.

IUPAC recommend that solubility *s* should be measured in $mol\,m^{-3}$. But you will usually see solubility measured in g solute per $100\,cm^3$ solvent. Note that if the solvent is water, the unit may instead be g solute per 100 g H_2O. This is because the density of water is approximately $1\,g\,cm^{-3}$, so $1\,cm^3$ has a mass of 1 g.

group 2 hydroxide	solubility (g per $100\,cm^3$ of water)
$Mg(OH)_2$	0.0012
$Ca(OH)_2$	0.113
$Sr(OH)_2$	0.410
$Ba(OH)_2$	2.57

The solubility of Group 2 hydroxides increases down the group.

Group 2 hydroxides

The solubility of the group 2 hydroxides increases down the group. Magnesium hydroxide is **sparingly soluble**: very little of it dissolves. Barium hydroxide is much more soluble, and it is possible to prepare a $0.1\,mol\,dm^{-3}$ solution of it.

Magnesium hydroxide

Magnesium hydroxide is used as an **antacid**. It forms a white suspension, usually called milk of magnesia. Stomach acid is hydrochloric acid. Indigestion is caused by excess stomach acid, and heartburn by stomach acid being pushed up into the oesophagus (gullet). Both conditions are painful. Magnesium hydroxide and other antacids relieve the symptoms by neutralizing the acid. Magnesium hydroxide reacts with hydrochloric acid to produce magnesium chloride and water:

$$Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$

Calcium hydroxide

Limestone is a common rock made from calcium carbonate. This decomposes on heating to form calcium oxide:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Calcium oxide reacts with water to form calcium hydroxide:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

Farmers 'lime' their fields to control the acidity in the soil. They add powdered limestone, calcium oxide, or calcium hydroxide. This is important because different crops grow best in soils with different degrees of acidity. For example, potatoes grow best at pH 5.5 to 6.5, oilseed rape at pH 6.0 to 7.5, and barley at pH 6.5 to 7.5.

Farmers use a soil test kit to find the pH of the soil, then add sufficient lime to bring the pH of the soil into the required range. This needs to be repeated regularly because liming is not a permanent answer to the problem of soil acidity. Acids from acid rain, plants, and fertilizers react with the lime. This reduces its effectiveness over time.

Soil acidity is controlled by spreading calcium hydroxide on fields.

Limewater

Calcium hydroxide solution is also called limewater. This is used to detect the presence of carbon dioxide. When carbon dioxide is bubbled through limewater, a cloudy white suspension of calcium carbonate is formed:

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

This gradually clears when excess carbon dioxide is bubbled through:

$$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$$

Laxative effect

Magnesium hydroxide also acts as a laxative. It relieves constipation and makes it easier to pass faeces. The magnesium ions are responsible for this action, rather than the hydroxide ions. Magnesium ions are not absorbed by the intestines very well, causing water to be drawn into the intestines. This makes the faeces softer and easier to pass. But it also means that indigestion sufferers need to make sure they do not swallow too much milk of magnesia.

Limewater is commonly used to test for carbon dioxide in the breath.

Check your understanding

1. **a** What is the trend in the solubility of the group 2 hydroxides?
   b Predict the relative solubility of radium hydroxide, and give a reason for your answer.
2. Explain the use of
   a magnesium hydroxide in medicine
   b calcium hydroxide in agriculture

10.04 Group 2: sulfates

OBJECTIVES

already from AS level, you know

- the relative solubilities of the hydroxides of the elements magnesium to barium

and after this spread you should

- know the relative solubilities of the sulfates of the elements magnesium to barium
- know the use of barium sulfate in medicine
- understand why acidified barium chloride solution is used as a reagent to test for sulfate ions

The solubility of the group 2 sulfates decreases as you go down the group. Note that this is the opposite trend to the one shown by the group 2 hydroxides. Magnesium hydroxide is sparingly soluble but magnesium sulfate is soluble.

group 2 sulfate	solubility (g per 100 cm^3 of water)
$MgSO_4$	25.5
$CaSO_4$	0.24
$SrSO_4$	0.013
$BaSO_4$	0.00022

The solubility of group 2 sulfates decreases down the group.

Other sulfates

Magnesium sulfate-7-water, $MgSO_4.7H_2O$, is also called Epsom salts. It gets its name from mineral water at Epsom in Surrey, which naturally contains magnesium sulfate. It is used in bath salts, fireproofing, and artificial snow for film sets. It is a laxative, just like magnesium hydroxide.

Plaster of Paris is $2CaSO_4.H_2O$. When it is mixed with water, it hydrates to form calcium sulfate-2-water, $CaSO_4.2H_2O$. This expands slightly and sets hard. It is used to make plaster casts to keep broken bones still, and by forensic scientists to make plaster casts of shoeprints at crime scenes.

Plaster of Paris is used to make plasterboard, and for plastering walls and ceilings.

Barium sulfate

Barium compounds are toxic. Soluble barium compounds are particularly hazardous as they could be absorbed through the intestines if swallowed. Barium sulfate is very insoluble in water. It is also opaque to X-rays, which means that they cannot pass through it. These features make barium sulfate useful as a *contrast medium* in medical X-rays of the digestive system. The lower part of the digestive system shows up in X-rays if the patient is given an enema of barium sulfate suspension. The upper part shows up in X-rays if the patient swallows barium sulfate suspension. In both cases, the barium sulfate eventually passes harmlessly out of the body.

Barium sulfate shows up in medical X-rays, like this one of the large intestine.

Rat poison

Barium carbonate is used as rat poison. It is not soluble, but it reacts with stomach acid to produce barium chloride, which is soluble and also toxic:

$BaCO_3(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

Testing for sulfates

The insolubility of barium sulfate is the basis of a simple laboratory test for the presence of sulfate ions in solution. A white precipitate of barium sulfate forms when barium chloride solution is added to a solution containing sulfate ions. For example, when barium chloride solution is added to sodium sulfate solution the following reaction occurs.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

This can be simplified to an ionic equation, since the Na^+ and Cl^- ions remain in solution as spectator ions:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The test works just as well if barium nitrate is used instead. But in both cases it is important to first acidify the test sample using hydrochloric acid or nitric acid. This is needed to prevent the false detection of sulfite ions SO_3^{2-} in the test. Barium sulfite $BaSO_3$ is insoluble. It would also form a white precipitate in the test. The acids react with the sulfite ion to form sulfur dioxide:

$$2H^+(aq) + SO_3^{2-}(aq) \rightarrow SO_2(g) + H_2O(l)$$

Acidified barium chloride solution produces a white precipitate if sulfate ions are present in a sample.

Check your understanding

1. **a** What is the trend in the solubility of the group 2 sulfates?
   b In what way is this trend different from the trend in the solubility of the group 2 hydroxides?
2. Explain the use of barium sulfate in medicine, even though barium compounds are toxic.
3. **a** Describe the laboratory test for the presence of aqueous sulfate ions.
   b Explain why the barium chloride used in the test can be acidified with hydrochloric acid, but not with sulfuric acid, H_2SO_4.

10.05 Testing for anions

OBJECTIVES

already from AS level, you understand

- why acidified silver nitrate solution is used as a reagent to identify and distinguish between F^-, Cl^-, Br^-, and I^- ions
- why acidified barium chloride solution is used as a reagent to test for sulfate ions

and after this spread you should know

- how to carry out anion tests competently

One of the recommended Practical Skills Assessment tasks is to carry out some inorganic tests. You may be asked to test some unknown solutions to find out which anions they contain. For example, they may contain chloride, bromide, iodide, or sulfate ions. You will be assessed on your ability to

- work safely and carefully
- use sensible volumes of the reagents
- obtain the correct observations for each test solution

Since you will be following a set of instructions given to you, it is possible that you might have to carry out tests for other anions. For example, you might also test for the presence of carbonate or nitrate ions.

Substance	Test	Observations	Inference
sodium chloride solution	A few drops of dil. HCl added, then a few drops of $AgNO_3(aq)$	White precipitate forms	Cl^- ions present
A			
B			
C			

You may wish to draw a table like this one. It has space to record what you did, your observations, and what you think your observations mean.

Results

It is always wise to draw a blank results table before you start work. That way, you can be certain that you have completed all the tasks before you tidy away.
If you are uncertain of an observation and repeat a certain test, do not erase your original record. Just put a line through your writing in such a way that you can still see your original observations. You may have been correct the first time around.

Watch the volumes

A common mistake in test tube reactions is to add too much reagent. When you do this, it is very difficult to mix the contents effectively. It is best to limit the volume of your test substance to about 1 cm^3, unless you are told otherwise.
If you are asked to add a reagent dropwise, this is exactly what it means.

- Add a drop using a teat pipette.
- Hold the tube at the top.
- Mix the contents by shaking the bottom of the tube from side to side.

Anion tests – a reminder

Halide ions

You can test for the presence of Cl^-, Br^-, and I^- ions using silver nitrate solution.

- Add about 1 cm^3 of your test substance to a clean test tube.
- Add a few drops of dilute nitric acid and shake.
- Add a few drops of silver nitrate solution.

observations	inference
No precipitate	Does not contain Cl^-, Br^-, or I^- ions
White precipitate	Cl^- ions present
Cream precipitate	Br^- ions present
Yellow precipitate	I^- ions present

Typical observations and inferences for halide ion tests. Note that the absence of a precipitate may also mean that the concentration of halide ions is too low to detect.

You may need to carry out a confirmatory test on each precipitate using aqueous ammonia. Remember that ammonia is corrosive and has a sharp, irritating smell. Concentrated ammonia should be used in a fume cupboard. You may need to add an excess of ammonia.

- White precipitate of AgCl redissolves in dilute aqueous ammonia.
- Cream precipitate of AgBr redissolves in concentrated ammonia.
- Yellow precipitate of AgI does not redissolve in dilute or concentrated ammonia.

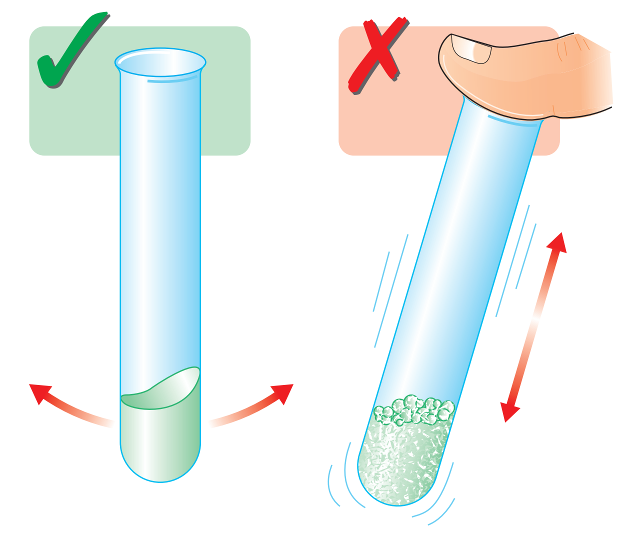
Sulfate ions

You can test for the presence of SO_4^{2-} ions using aqueous barium chloride or barium nitrate solution.

- Add about 1 cm^3 of your test substance to a clean test tube.
- Add a few drops of dilute hydrochloric acid and shake.
- Add a few drops of barium chloride or barium nitrate solution.

observations	inference
No precipitate	Does not contain SO_4^{2-} ions
White precipitate	SO_4^{2-} ions present

Typical observations and inferences for sulfate ion tests. Note that the absence of a precipitate may also mean that the concentration of sulfate ions is too low to detect.



Do not put your thumb over the top of the test tube and shake it up and down.

Check your understanding

1. Suggest why the reagents should be made up in distilled or de-ionized water, rather than tap water.
2. Explain why you should be careful to
   - a use the correct acids in these tests
   - b use clean test tubes
   - c record your observations from one test before carrying out the next one