

GROUP II ELEMENTS

Beryllium to Barium

Introduction

Elements in Group I (*alkali metals*) and Group II (*alkaline earths*) are known as **s-block elements** because their valence (bonding) electrons are in s orbitals.

	Be	Mg	Ca	Sr	Ba
Atomic Number	4	12	20	38	56
Electronic configuration	1s ² 2s ²	[Ne] 3s ²	[Ar] 4s ²	[Kr] 5s ²	[Xe] 6s ²

PHYSICAL PROPERTIES

Atomic Radius

Increases down each group electrons are in shells further from the nucleus

	Be	Mg	Ca	Sr	Ba
Atomic radius / nm	0.106	0.140	0.174	0.191	0.198

Ionic Size

Increases down the group

The size of positive ions is less than the original atom because the nuclear charge exceeds the electronic charge.

	Be ²⁺	Mg ²⁺	Ca ²⁺	Sr ²⁺	Ba ²⁺
Ionic radius / nm	0.030	0.064	0.094	0.110	0.134

Melting Points

Decrease down each group metallic bonding gets weaker due to increased size. Each atom contributes two electrons to the delocalised cloud. Melting points tend not to give a decent trend as different crystalline structures affect the melting point.

	Be	Mg	Ca	Sr	Ba
Melting point / °C	1283	650	850	770	710

Ionisation Energy

Decreases down the group atomic size increases

Values for Group I are low because the electron has just gone into a new level and is shielded by filled inner levels. This makes them reactive. Group II elements have higher values than their Group I equivalents due to the increased nuclear charge.

	Be	Mg	Ca	Sr	Ba
1st I.E. / kJ mol ⁻¹	899	738	590	550	500
2nd I.E. / kJ mol ⁻¹	1800	1500	1100	1100	1000
3rd I.E. / kJ mol ⁻¹	14849	7733	4912	4120	3390

There is a **large increase for the 3rd I.E.** as the electron is now being removed from a **shell nearer the nucleus** and there is **less shielding**.

Electronegativity **Decreases down the group**

Increased shielding makes the shared pair less strongly attracted to the nucleus

	Be	Mg	Ca	Sr	Ba
Electronegativity (Pauling)	1.5	1.2	1.0	0.95	0.89

Hydration

Enthalpy

This is a measure of an ion's attraction for water

Decreases (gets less negative) down each group

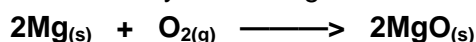
Charge density of the ions decreases thus reducing the attraction for water

	Be ²⁺	Mg ²⁺	Ca ²⁺	Sr ²⁺	Ba ²⁺
Hydration Enthalpy / kJ mol ⁻¹		-1920	-1650	-1480	-1360

CHEMICAL PROPERTIES

Oxygen • **react with increasing vigour down the group**

Mg burns readily with a bright white flame



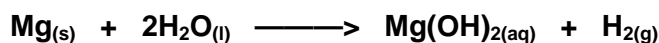
Ca, Sr, Ba Similar reaction

Ca brick red flame
Sr crimson flame
Ba apple green flame

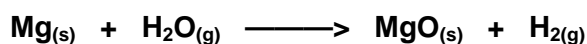
Water • **react with increasing vigour down the group**

Be does not react with water or steam

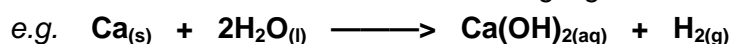
Mg reacts very slowly with cold water



but reacts quickly with steam



Ca, Sr, Ba react with cold water with increasing vigour



Reason(s) for the difference in reactivity...

COMPOUNDS

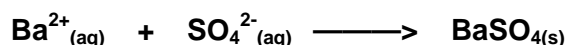
- Sulphates*
- white crystalline solids
 - solubility in water decreases down the Group

Salt	Ionic radius (M^{2+}) / nm	Hydration Enthalpy (M^{2+}) / kJ mol^{-1}	Solubility moles/100g
MgSO_4	0.064	-1891	3600×10^{-4}
CaSO_4	0.094	-1562	11×10^{-4}
SrSO_4	0.110	-1413	0.62×10^{-4}
BaSO_4	0.134	-1273	0.009×10^{-4}

- reasons for solubility decreasing down the group ...
 - there is little change in the lattice enthalpy BUT
 - as the **cation gets larger** the **hydration enthalpy gets much smaller**
 - a **larger cation** has a **lower charge density** and so is **less attracted to water**

Testing for sulphates

- barium sulphate's insolubility is used as a test for sulphates
- Method
- make up a solution of the compound to be tested
 - acidify it with dilute hydrochloric (or nitric) acid *
 - add a few drops of barium chloride solution
 - white precipitate of barium sulphate confirms the presence of a sulphate



* adding acid prevents the precipitation of other insoluble ions such as carbonate

- Hydroxides*
- white crystalline solids
 - solubility in water increases down the Group

Be(OH)₂ *insoluble*

Mg(OH)₂ *sparingly soluble*

Ca(OH)₂ *slightly soluble* - an aqueous solution is known as 'lime water'

Sr(OH)₂ *quite soluble*

Ba(OH)₂ *very soluble*

- **basic strength also increases down group**
- the **metal ions get larger** so charge density decreases
- there is a lower attraction between the OH⁻ ions and larger unipositive ions
- the ions will split away from each other more easily
- there will be a greater concentration of OH⁻ ions in water



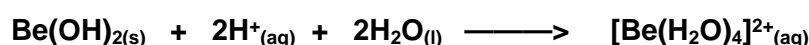
'The greater the concentration of OH⁻ ions in water the greater the alkalinity'

THE ATYPICAL NATURE OF BERYLLIUM

Theory Beryllium differs from the other Group II elements; it has properties closer to that of aluminium - THE DIAGONAL RELATIONSHIP. Being the **head element** of a Group...

- it has*
- a much **smaller ionic size** (a **greater charge/size ratio - highly polarising**)
 - a much **larger ionisation energies** than those elements below it

- so*
- is less likely to form ions
 - compounds (BeCl₂) show covalent character
 - often soluble in organic solvents
 - have lower melting points
 - often hydrolysed by water
 - maximum co-ordination number of 4
 - due to small size
 - beryllium hydroxide is AMPHOTERIC
 - dissolves in both acids and bases



CALCIUM AND ITS COMPOUNDS

Source Most calcium is found as calcium carbonate in limestone

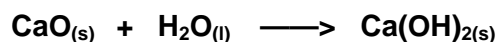
Compounds	limestone	calcium carbonate	$\text{CaCO}_{3(s)}$	<i>making cement</i>
	quicklime	calcium oxide	$\text{CaO}_{(s)}$	<i>iron purification</i>
	slaked lime	solid calcium hydroxide	$\text{Ca(OH)}_{2(s)}$	<i>soil treatment</i>
	lime water	aqueous calcium hydroxide	$\text{Ca(OH)}_{2(aq)}$	<i>testing for CO_2</i>

Review of Reactions

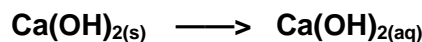
limestone decomposes on strong heating giving



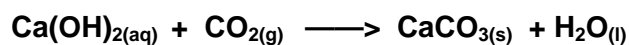
quicklime reacts with water to produce **slaked lime** (calcium hydroxide)



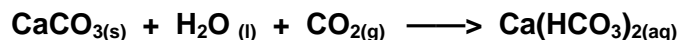
calcium hydroxide is sparingly soluble - an aqueous solution is called **lime water** (pH = 9-10)



lime water reacts with carbon dioxide



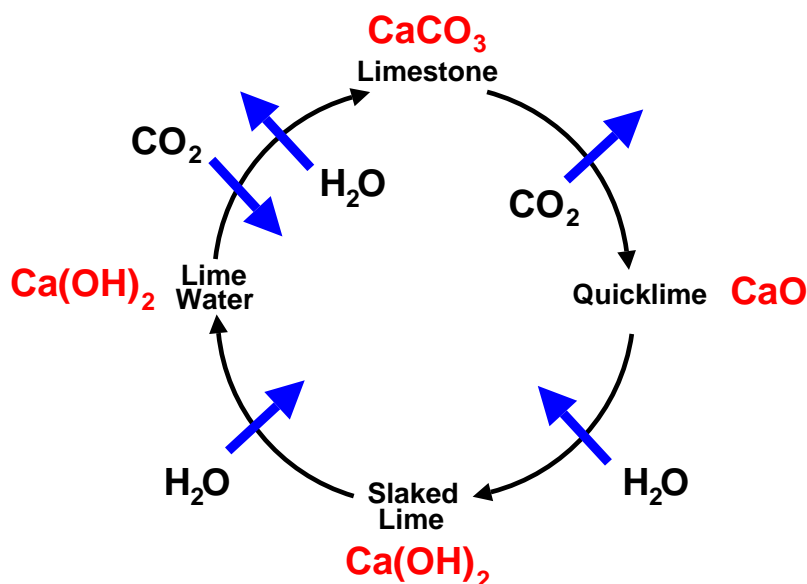
excess carbon dioxide produces calcium hydrogencarbonate



heating reverses the process

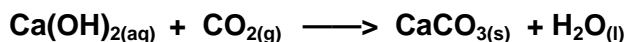


THE CALCIUM CYCLE

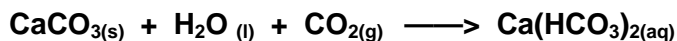


Lime water

- an aqueous solution of calcium hydroxide - **pH = 9-10**
- calcium hydroxide is only sparingly soluble in water
- used as a laboratory test for carbon dioxide
- passing carbon dioxide through lime water produces a white precipitate



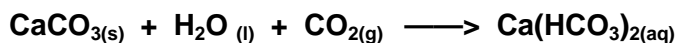
- the white precipitate re-dissolves with EXCESS carbon dioxide



- warming the solution reverses the process and precipitates calcium carbonate

*Hard water*

- water that doesn't form a lather easily with soap
- contains **soluble** calcium and magnesium ions - Ca^{2+} and Mg^{2+}
- insoluble calcium and magnesium salts do not cause hardness
- formed by the action of rain and carbon dioxide on limestone



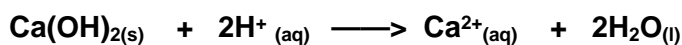
- calcium hydrogencarbonate produces temporary hardness
- temporary hardness is removed by heating the water - cause of 'fur' in kettles



- other ways of removing hardness include...
 - ion exchange
 - distillation
 - using excess soap
 - using washing soda

Slaked lime

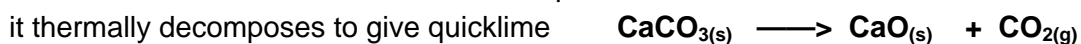
- solid calcium hydroxide
- used to reduce the acidity of soil to get better crop yields



- much safer to use than sodium hydroxide which is caustic

Limestone

- solid calcium carbonate
- used in the extraction of iron to remove impurities...



calcium silicate (slag) is formed

Quicklime

- solid calcium oxide
- removes impurities in iron and steel making - see above

MAJOR USES OF GROUP II COMPOUNDS - A SHORT REVIEW

<i>General</i>	Group II elements and compounds are used widely in industry.
Magnesium	<i>Property</i> burns with a bright white light
	<i>Use</i> flares, incendiary bombs, tracer bullets
	<i>Property</i> high in reactivity series
	<i>Use</i> sacrificial metal to protect steel from corroding on ships' hulls extraction of titanium

Magnesium hydroxide	<i>Property</i> weak alkali
	<i>Use</i> in indigestion tablets and toothpaste to remove acidity $\text{Mg(OH)}_{2(s)} + 2\text{H}^+_{(aq)} \longrightarrow \text{Mg}^{2+}_{(aq)} + 2\text{H}_2\text{O}_{(l)}$

Calcium hydroxide	<i>Property</i> weak alkali
	<i>Use</i> sold is added to soil to reduce acidity $\text{Ca(OH)}_{2(s)} + 2\text{H}^+_{(aq)} \longrightarrow \text{Ca}^{2+}_{(aq)} + 2\text{H}_2\text{O}_{(l)}$

Calcium hydroxide and magnesium hydroxide are preferable for neutralising acid than any Group I hydroxide such as NaOH. Sodium hydroxide is caustic and would burn skin or damage plants if it came into contact with them.

Calcium oxide	<i>Property</i> basic oxide
	<i>Use</i> removal of impurities in the blast furnace $\text{CaO}_{(s)} + \text{SiO}_{2(s)} \longrightarrow \text{CaSiO}_{3(s)}$ <i>base acid slag'</i>

Calcium carbonate	<i>Property</i> basic oxide
	<i>Use</i> removal of impurities in the blast furnace manufacture of cement by heating it with clay