

## 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.

	<i>Be</i>	<i>Mg</i>	<i>Ca</i>	<i>Sr</i>	<i>Ba</i>
Atomic Number	4	12	20	38	56
Electronic configuration	$1s^2 2s^2$	$[\text{Ne}] 3s^2$	$[\text{Ar}] 4s^2$	$[\text{Kr}] 5s^2$	$[\text{Xe}] 6s^2$

### PHYSICAL PROPERTIES

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

	<i>Be</i>	<i>Mg</i>	<i>Ca</i>	<i>Sr</i>	<i>Ba</i>
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.

	$\text{Be}^{2+}$	$\text{Mg}^{2+}$	$\text{Ca}^{2+}$	$\text{Sr}^{2+}$	$\text{Ba}^{2+}$
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.

	<i>Be</i>	<i>Mg</i>	<i>Ca</i>	<i>Sr</i>	<i>Ba</i>
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.

	<i>Be</i>	<i>Mg</i>	<i>Ca</i>	<i>Sr</i>	<i>Ba</i>
1st I.E. / $\text{kJ mol}^{-1}$	899	738	590	550	500
2nd I.E. / $\text{kJ mol}^{-1}$	1800	1500	1100	1100	1000
3rd I.E. / $\text{kJ mol}^{-1}$	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 <sup>2+</sup>	Mg <sup>2+</sup>	Ca <sup>2+</sup>	Sr <sup>2+</sup>	Ba <sup>2+</sup>
Hydration Enthalpy / kJ mol <sup>-1</sup>		-1920	-1650	-1480	-1360

## CHEMICAL PROPERTIES

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

Mg burns readily with a bright white flame



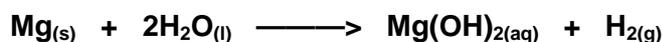
Ca, Sr, Ba Similar reaction

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

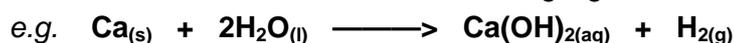
Mg reacts very slowly with cold water



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 ( $M^{2+}$ ) / $\text{kJ mol}^{-1}$	Solubility moles/100g
$\text{MgSO}_4$	0.064	-1891	$3600 \times 10^{-4}$
$\text{CaSO}_4$	0.094	-1562	$11 \times 10^{-4}$
$\text{SrSO}_4$	0.110	-1413	$0.62 \times 10^{-4}$
$\text{BaSO}_4$	0.134	-1273	$0.009 \times 10^{-4}$

- 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)<sub>2</sub> *insoluble*

Mg(OH)<sub>2</sub> *sparingly soluble*

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

Sr(OH)<sub>2</sub> *quite soluble*

Ba(OH)<sub>2</sub> *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<sup>-</sup> ions and larger unipositive ions
- the ions will split away from each other more easily
- there will be a greater concentration of OH<sup>-</sup> ions in water



*'The greater the concentration of OH<sup>-</sup> 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 (BeCl<sub>2</sub>) 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

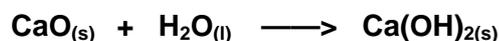
<b>Compounds</b>	limestone	calcium carbonate	$\text{CaCO}_3(\text{s})$	<i>making cement</i>
	quicklime	calcium oxide	$\text{CaO}(\text{s})$	<i>iron purification</i>
	slaked lime	solid calcium hydroxide	$\text{Ca}(\text{OH})_2(\text{s})$	<i>soil treatment</i>
	lime water	aqueous calcium hydroxide	$\text{Ca}(\text{OH})_2(\text{aq})$	<i>testing for <math>\text{CO}_2</math></i>

### Review of Reactions

limestone decomposes on strong heating giving



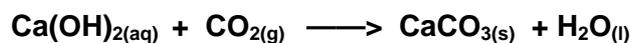
quicklime reacts with water to produce **slaked lime** (calcium hydroxide)



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



lime water reacts with carbon dioxide



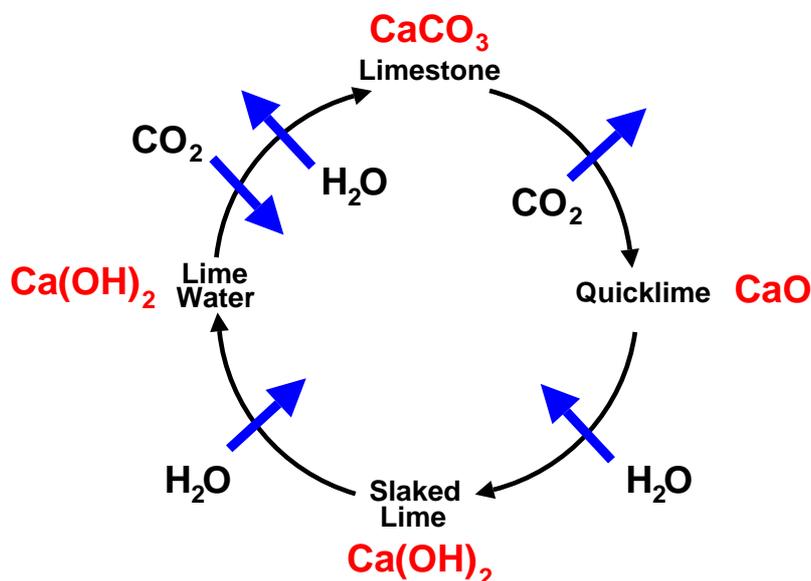
excess carbon dioxide produces calcium hydrogencarbonate



heating reverses the process

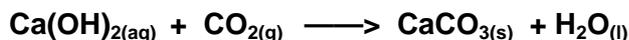


### THE CALCIUM CYCLE



*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 -  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$
- insoluble calcium and magnesium salts do not cause hardness
- formed by the action of rain and carbon dioxide on limestone



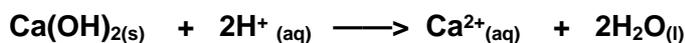
- calcium hydrogencarbonate produces temporary hardness
- temporary hardness is removed by heating the water - cause of 'fur' in kettles



- 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

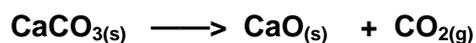


- much safer to use than sodium hydroxide which is caustic

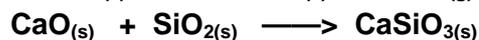
*Limestone*

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

it thermally decomposes to give quicklime



calcium oxide reacts with silica



calcium silicate (slag) is formed

*Quicklime*

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

