

# Group 2

## A-Level Chemistry 9701 · Topic 10 revision notes

Everything from the 3 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

1. Group 2 metals and their reactions

2. Group 2 compounds and their trends

3. Uses of Group 2 compounds

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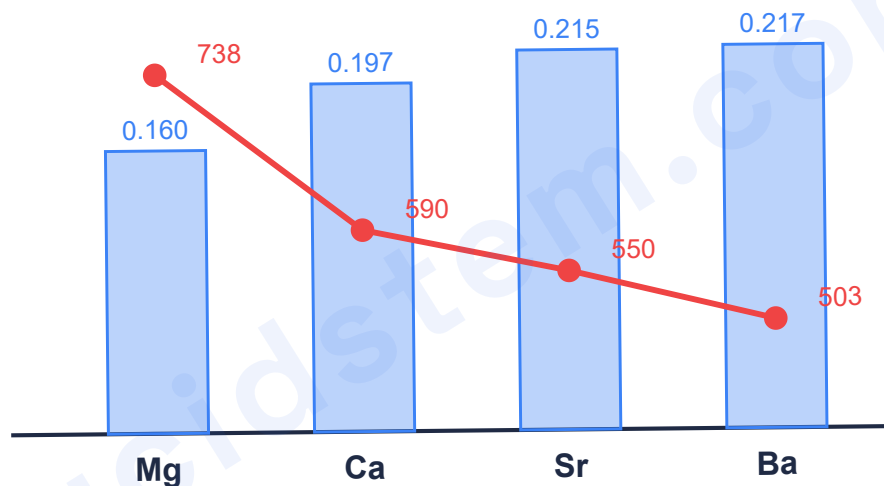
# 1 Group 2 metals and their reactions

## Syllabus 10.1

All Group 2 elements have two outer *s* electrons and form  $M^{2+}$  ions, so their compounds have oxidation number +2. Down the group the atoms gain shells, so **atomic radius increases** and shielding increases; the outer electrons are less strongly held, so the first and second **ionisation energies decrease**. Less energy is needed to form  $M^{2+}$ , so **reactivity increases** from Mg to Ba. The metals burn in oxygen to form  $MO$ , react with water to give hydroxides and hydrogen (Mg only slowly unless it is steam), and react with dilute acids to form salts and hydrogen.

bars: atomic radius / nm

line: 1st IE /  $\text{kJ mol}^{-1}$



bigger atoms lose their two outer electrons more easily

Fig. 1: atomic radius increases and first ionisation energy decreases down Group 2

**Bigger atom, easier to ionise, more reactive.**

- ✓ **Radius** up, **IE** down, **reactivity** up.
- ✓ **Water:**  $M + 2H_2O \rightarrow M(OH)_2 + H_2$ .
- ✓ **Mg + steam**  $\rightarrow MgO + H_2$ .
- ✓ **Oxidation number** always +2.

### Watch out!

Calcium, strontium and barium react with dilute sulfuric acid at first but soon slow down, because their sulfates are insoluble and coat the metal. Do not use this to argue that reactivity decreases down the group. And magnesium with **steam** gives the **oxide**, not the hydroxide.

## 2 Group 2 compounds and their trends

### Syllabus 10.1

Group 2 carbonates decompose on heating to the oxide and carbon dioxide; nitrates decompose to the oxide, nitrogen dioxide and oxygen. **Thermal stability increases down the group**; the  $M^{2+}$  ion gets larger with the same charge, so its **charge density falls** and it **polarises** (distorts) the large carbonate or nitrate anion less. The **hydroxides become more soluble** down the group, so the pH of their saturated solutions rises; the **sulfates become less soluble**, so barium sulfate is insoluble, which is the basis of the test for sulfate ions.



small ion, high charge density:  
anion pulled out of shape

**MgCO<sub>3</sub> decomposes easily**

large ion, low charge density:  
anion hardly distorted

**BaCO<sub>3</sub> is very stable**

thermal stability of carbonates and nitrates increases down the group

Fig. 1: a small  $Mg^{2+}$  ion distorts the carbonate ion much more than a large  $Ba^{2+}$  ion

**It all comes back to the size of  $M^{2+}$ .**

- ✓ **Carbonate**  $\rightarrow$  oxide +  $CO_2$ ; **nitrate**  $\rightarrow$  oxide +  $NO_2$  +  $O_2$ .
- ✓ **Stability** increases down: lower charge density, less polarisation.
- ✓ **Hydroxides** more soluble down; **sulfates** less.
- ✓ **Sulfate test:** HCl then  $BaCl_2$ , white precipitate.

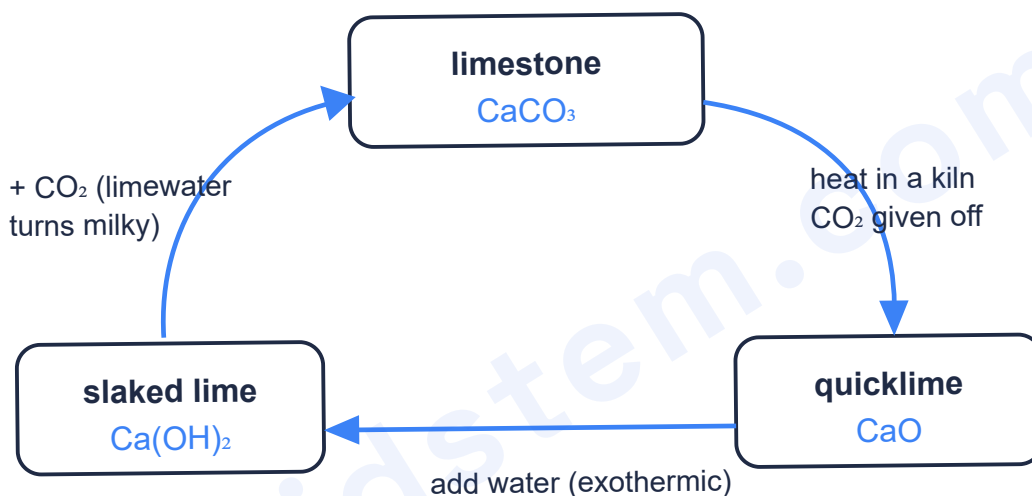
**Watch out!**

The two solubility trends go in **opposite** directions: hydroxides become **more** soluble down the group, sulfates **less**. When explaining thermal stability, talk about the **cation** polarising the **anion**; the anion is the same throughout, so the change must come from the size of  $M^{2+}$ .

### 3 Uses of Group 2 compounds

#### Syllabus 10.2

Heating limestone ( $\text{CaCO}_3$ ) in a lime kiln gives **quicklime** ( $\text{CaO}$ ) and  $\text{CO}_2$ ; adding water gives **slaked lime** ( $\text{Ca(OH)}_2$ ) in an exothermic reaction;  $\text{Ca(OH)}_2$  absorbs  $\text{CO}_2$  to re-form  $\text{CaCO}_3$ , which is why limewater turns milky. Slaked lime and powdered limestone are **bases** used to neutralise acidic soils and lakes. Limestone is used in building and to make cement. Magnesium hydroxide is a mild base in indigestion remedies, and **barium sulfate** is safe to swallow for X-ray "barium meals" because it is **insoluble**.



$\text{CaO}$  and  $\text{Ca(OH)}_2$  neutralise acidic soils;  $\text{CaCO}_3$  is used in building  
1.00 t  $\text{CaCO}_3$  gives 0.560 t  $\text{CaO}$

Fig. 1: the lime cycle

#### Limestone, quicklime, slaked lime, limestone.

- ✓ **Heat:**  $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ .
- ✓ **Slake:**  $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$ , exothermic.
- ✓ **Bases** neutralise acidic soil and stomach acid.
- ✓ **BaSO<sub>4</sub>** is safe because it is insoluble.

#### Watch out!

Barium compounds in general are toxic. The barium meal is safe **only** because  $\text{BaSO}_4$  is insoluble; a soluble salt such as  $\text{BaCl}_2$  must never be used. And slaking quicklime is strongly **exothermic**: the mixture can boil, so water must be added carefully.