CIE Chemistry A Level

10 : Group 2

Notes

Trends in the properties of Group 2 metals, magnesium to barium, and their compounds

Group 2 metals and oxygen

Generally, the Group 2 metals burn in oxygen to form a **metal oxide.**

- **Beryllium** is coated in a thin layer of **beryllium oxide** which **inhibits** the reaction meaning it only reacts in a powder form.
  $2Be + O_2 \rightarrow 2BeO$
- **Magnesium** burns in oxygen with a **bright white flame**.
  $2Mg + O_2 \rightarrow 2MgO$
- **Calcium** burns with a **bright white flame** which is slightly red at the top.
  $2Ca + O_2 \rightarrow 2CaO$
- **Strontium** is reluctant to start burning but burns intensely with a **white flame.**
  $2Sr + O_2 \rightarrow 2SrO$
- **Barium** burns in oxygen with **white flame**.
  $2Ba + O_2 \rightarrow 2BaO$

Group 2 metals and water

The reactions of the Group 2 metals with water or steam can be used to see the **trend in reactivity** down the group.

- **Beryllium** reacts with steam only at **very high temperatures.**
  $Be + H_2O \rightarrow BeO + H_2$
- **Magnesium** has a very **slight reaction** with **cold water.** The reaction stops due to the production of an insoluble coat of magnesium hydroxide.
  $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$
  Magnesium burns in **steam** more readily than cold water:
  $Mg + H_2O \rightarrow MgO + H_2$
- **Calcium**, **strontium** and **barium** all react in cold water to produce their hydroxide and hydrogen gas. The reactions become **increasingly vigorous** down the group.
  E.g. $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

Beryllium only reacts with steam at high temperatures but, going down Group 2, the metals react more readily and rapidly with cold water, with barium reacting the fastest. This shows that **reactivity increases down the group.**

Group 2 metals and dilute acids

Hydrochloric acid

All Group 2 metals react with dilute hydrochloric acid to produce a **metal chloride** and **hydrogen** gas. The reactions get **more vigorous** as you go down the group. The general equation for this reaction is: $X + 2HCl \rightarrow XCl_2 + H_2$ (where X is a Group 2 metal).
E.g. $Ca + 2HCl \rightarrow CaCl_2 + H_2$

Sulfuric acid

Dilute sulfuric acid reacts with Group 2 metals to produce a **metal sulfate** and **hydrogen**. The general equation for this reaction is: $X + H_2SO_4 \rightarrow XSO_4 + H_2$ (where X is a Group 2 metal).
E.g. $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

The reactions with dilute sulfuric acid **do not get more vigorous** down the group due to the **solubility of the sulfates** produced. Beryllium and magnesium produce soluble sulfates so their reactions with sulfuric acid are similar to their reactions with hydrochloric acid.
Calcium produces a **sparingly soluble** sulfate. **Strontium** and **barium** produce **insoluble sulfates.** This means calcium, strontium and barium will only react with sulfuric acid for a short period of time because the **reaction will stop** once the **insoluble sulfate forms on the metal**.

Group 2 oxides

Reactions with water

Apart from **beryllium**, all Group 2 oxides react with water to produce hydroxides.

- **Magnesium oxide** produces a solution that is around **pH 9**. This is because the **magnesium hydroxide** is only **sparingly soluble** so not many OH^- ions are released into the solution.
  $MgO + H_2O \rightarrow Mg(OH)_2$
- **Calcium oxide** (quicklime) undergoes an **exothermic** reaction to produce calcium hydroxide (also known as slaked lime or lime water). **Calcium hydroxide** is **partially soluble** so the resulting solution is **pH 12.**
  $CaO + H_2O \rightarrow Ca(OH)_2$
- **Strontium oxide** and **barium oxide** produce hydroxides which are **increasingly soluble**. They react in the same way as calcium but produce solutions with a **higher pH** as more OH^- ions get released into the solution.

Reactions with dilute acids

All Group 2 oxides react with dilute acids to produce **salt and water.**
The general equations for these reactions (where X is a group 2 metal) are:
$XO + 2HCl \rightarrow XCl_2 + H_2O$
$XO + 2HNO_3 \rightarrow X(NO_3)_2 + H_2O$
$XO + H_2SO_4 \rightarrow XSO_4 + H_2O$

The reactions with **hydrochloric and nitric acid are standard** and reactivity increases down the group.

The reactions with **sulfuric acid are different** due to the different **solubilities** of the products. Magnesium and beryllium oxides react as expected. **Calcium**, **barium** and **strontium** oxides react differently because their sulfates are **increasingly insoluble**. The sulfate formed during the reaction coats the metal oxide, **slowing or stopping** the reaction.

Group 2 hydroxides

Reactions with water

The Group 2 hydroxides **do not react** with water.

Reactions with dilute acids

The Group 2 hydroxides react with dilute acids in the **same way as the metal oxides** (explained on the previous page). The only difference is that two water molecules are produced rather than one.

E.g. $Sr(OH)_2 + 2HCl \rightarrow SrCl_2 + 2H_2O$

Metal carbonates

Reactions with water

The Group 2 metal carbonates are insoluble so they **do not react** with water.

Reactions with dilute acids

Group 2 carbonates react with dilute acids to produce a **salt**, **water** and **carbon dioxide.**

The general equations for these reactions (where X is a Group 2 metal) are:

$XCO_3 + 2HCl \rightarrow XCl_2 + H_2O + CO_2$

$XCO_3 + 2HNO_3 \rightarrow X(NO_3)_2 + H_2O + CO_2$

$XCO_3 + H_2SO_4 \rightarrow XSO_4 + H_2O + CO_2$

The reactions with **hydrochloric and nitric acid are standard** and **reactivity increases** down the group.

The reactions with **sulfuric acid are different** due to the different **solubilities** of the products. Magnesium and beryllium carbonates react as expected. **Calcium**, **barium** and **strontium** carbonates react differently because their sulfates are **increasingly insoluble**. The sulfate formed during the reaction coats the metal carbonate, **slowing or stopping** the reaction.

Thermal decomposition of the nitrates and carbonates

Nitrates

All Group 2 nitrates undergo **thermal decomposition** to produce a **metal oxide**, **oxygen** and **nitrogen dioxide**. The nitrates are **heated more strongly** as you go down the group.

The general equation for this reaction (where X is the Group 2 metal) is:

$2X(NO_3)_2 \rightarrow 2XO + 4NO_2 + O_2$

Observations: The nitrate and the oxide are both **white solids**. Nitrogen dioxide is a **brown gas**.

Carbonates
All Group 2 carbonates undergo **thermal decomposition** to produce a **metal oxide** and **carbon dioxide**. The carbonates are **heated more strongly** as you go down the group.
The general equation for this reaction (where X is the Group 2 metal) is:
$XCO_3 \rightarrow XO + CO_2$
Observations: The carbonate and the oxide are both **white solids**. Carbon dioxide is a **colourless gas**.

Both Group 2 carbonates and nitrates become **more stable to heat** as you go down the group. This means as you go down the group, the nitrates and carbonates have to be **heated more strongly** for thermal decomposition to occur.

Thermal stability of the nitrates and carbonates (A Level only)

Both **carbonates** and **nitrates** become **more thermally stable** as you go down Group 2. The stability of the compounds is influenced by **charge density of the cation** and how **polarised** the **anion** becomes.
Ions formed from elements at the top of group 2 are **smaller** than those at the bottom. A **smaller +2 ion** has a **greater charge density** because the same charge is packed into in a small volume. This means smaller group 2 ions have a **greater polarising effect** on neighbouring negative ions.

When a carbonate or nitrate ion is placed near the metal cation, the **anion becomes polarised** because the **cation draws the electrons** towards itself. Smaller Group 2 ions are more polarising due to their greater charge density. The **more polarised the anion** is, the **less heat is required** for thermal decomposition to occur. This means that **thermal stability increases down the group** because ions further down the group are worse at polarising the anion so more heat energy is required to break bonds for thermal decomposition.

Predicting trends
The reactions of Group 2 metals, metal oxides, metal hydroxides and metal carbonates generally indicate that **reactivity increases as you go down Group 2.**
Reactivity increases down the group because **ionisation energy decreases** due to the **increasing atomic radius** and the **shielding effect** of electrons. This means further down the group, the electrons become easier to remove so reactivity increases.

Some **exceptions** to this trend can be seen when **sulfates and hydroxides are produced**, however this is due to the **insolubility** of some sulfates and hydroxides inhibiting reactions.

Solubility of hydroxides and sulfates

The **solubilities** of the Group 2 metal **hydroxides** and **sulfates** show trends in the group. The trend in the solubility of sulfates is **opposite** to the trend in the hydroxides:

Group 2 element - X	Hydroxide - $X(OH)_2$	Sulfate - XSO_4
Magnesium	Least soluble	Most soluble
Calcium		
Beryllium		
Barium	Most soluble	Least soluble

Compounds with very **low solubilities**, like magnesium hydroxide, are often said to be **sparingly soluble**. Most sulfates are soluble in warm water except barium sulfate which is insoluble.

Solubility in terms of enthalpy change of hydration and lattice energy (A Level only)

The trends in the solubilities of the Group 2 hydroxides and sulfates can be seen above.
As you go **down Group 2**, the **lattice enthalpy** required to break up the compound **decreases** because the **size of the positive ions increases**. Larger cations means there is **more space between ions** in the compound so there are **weaker forces of attraction** between the ions.
As the **cations increase in size** down the group, **the enthalpy change of hydration** (the amount of energy released as the ions bond to water molecules) also **decreases**.

<u>Hydroxides</u>
For hydroxide ions (relatively **small** ions), the **lattice enthalpy falls faster than the enthalpy change of hydration** of the cations. This means the **enthalpy change of solution** will become **more negative** down the group (more exothermic).

<u>Sulfates</u>
For sulfate ions (relatively **large** ions), the **lattice enthalpy falls slower than the hydration enthalpy** of the cations. This means the **enthalpy of solution** will become **more positive** down the group (more endothermic).

The **more exothermic** the **enthalpy of solution** is, the **more soluble** a compound is. Therefore sulfates become less soluble down the group and hydroxides become more soluble down the group.

Uses of Group 2 compounds

Calcium hydroxide and **calcium carbonate** are both compounds used in **agriculture**. Calcium carbonate is **powdered limestone**. Calcium hydroxide is formed when calcium oxide is added to water. Calcium oxide and calcium hydroxide are often referred to as **lime** and **slaked lime**, respectively.

Crops grow best in soil is around **pH 6**. If soil becomes **too acidic**, calcium carbonate or calcium hydroxide can be added to raise the pH. This is because both compounds **react with and neutralise acids**.

Calcium carbonate **reacts more slowly** than calcium hydroxide since it is **not water soluble**, however it is used more often as it is **cheaper** and **easier to handle.**