

Group 2

A2

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Read first: [Group 2, Equilibria](#)

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What you need to know

By the end of this topic you should be able to do the following for the Group 2 metals magnesium to barium:

1. **Thermal stability.** Describe how the thermal stability of the nitrates and the carbonates changes down the group, and explain it in words: as the cation gets bigger, it distorts (polarises) the large nitrate or carbonate ion less, so the compound is harder to break up.
2. **Solubility.** Describe how the solubility of the hydroxides and of the sulfates changes down the group, and how their enthalpy change of solution, ΔH_{sol} , changes. Explain both by comparing how much the lattice energy and the enthalpy change of hydration change.

Both explanations are qualitative: you are not asked to calculate a trend. But the Paper 4 questions on this topic always come with the AS Group 2 work (the equations for the decompositions and for the oxides with water and acids) and often with a pH or energy-cycle calculation from other A Level topics. This note shows those parts too, so you can answer the whole question.

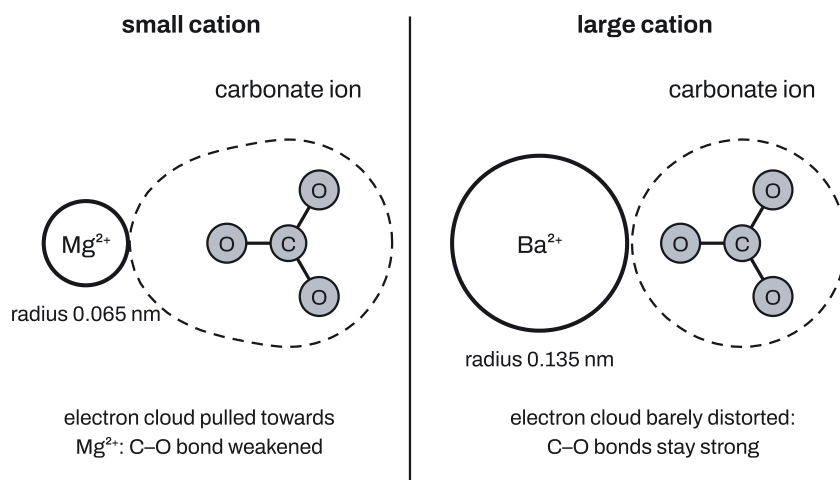
Understand it

What "polarising the anion" means

A Group 2 ion, M^{2+} , has lost its two outer electrons, so it is small and carries a 2+ charge. How strongly it pulls on nearby electrons depends on its **charge density**: the same 2+ charge squeezed into a smaller ion means a stronger pull. The ionic radii in the Data Booklet grow down the group:

Ion	Mg^{2+}	Ca^{2+}	Sr^{2+}	Ba^{2+}
ionic radius / nm	0.065	0.099	0.113	0.135

The nitrate ion, NO_3^- , and the carbonate ion, CO_3^{2-} , are **large** anions, and their electron clouds are spread over several atoms, so they are easy to distort. A small cation with a high charge density pulls the anion's electron cloud towards itself. We say the cation **polarises** the anion.



Pulling electron density towards the cation drains it from the bonds inside the anion, so a C–O bond (or an N–O bond in a nitrate) is **weakened**. A weaker bond breaks at a lower temperature.

Why thermal stability increases down the group

On heating, the carbonate ion splits into a carbon dioxide molecule and an oxide ion, and the nitrate ion breaks up in a similar way:



Put together the argument, one link at a time:

1. Down the group the cation's **ionic radius increases** (one more shell each time) with the same 2+ charge, so its **charge density decreases**.
2. So the cation **polarises (distorts) the anion less**.
3. So the C–O or N–O bonds in the anion are **weakened less**.
4. So a **higher temperature** is needed to decompose the compound: thermal stability **increases** down the group.

Magnesium carbonate, with the smallest cation, decomposes at the lowest temperature; barium carbonate needs the highest.

The same idea works for any compound with a large, easily distorted anion. Group 2 **hydroxides** ($\text{M}(\text{OH})_2 \rightarrow \text{MO} + \text{H}_2\text{O}$), **hydrogencarbonates** and **ethanedioates** ($\text{C}_2\text{O}_4^{2-}$) all get more stable down the group. It also works for other metals: a Cu^{2+} , Zn^{2+} or Pb^{2+} nitrate that "decomposes like a Group 2 nitrate" is compared with the Group 2 compounds by the **ionic radius** of its cation, and the Group 1 nitrates and hydrogencarbonates get more stable down Group 1 for the same reason.

An energy view of the same trend. You can also look at the enthalpy change of decomposition. Both lattice energies, of MCO_3 and of MO , become less exothermic down the group as the cation gets bigger. But the oxide ion is small, so in MO the cation's size makes up a large part of the distance between ion centres, and $\Delta H_{\text{latt}}(\text{MO})$ falls **faster** than $\Delta H_{\text{latt}}(\text{MCO}_3)$. Forming the oxide gets less and less of an energy gain, so the enthalpy change of decomposition becomes **more positive** (more endothermic) down the group.

The energy terms behind solubility

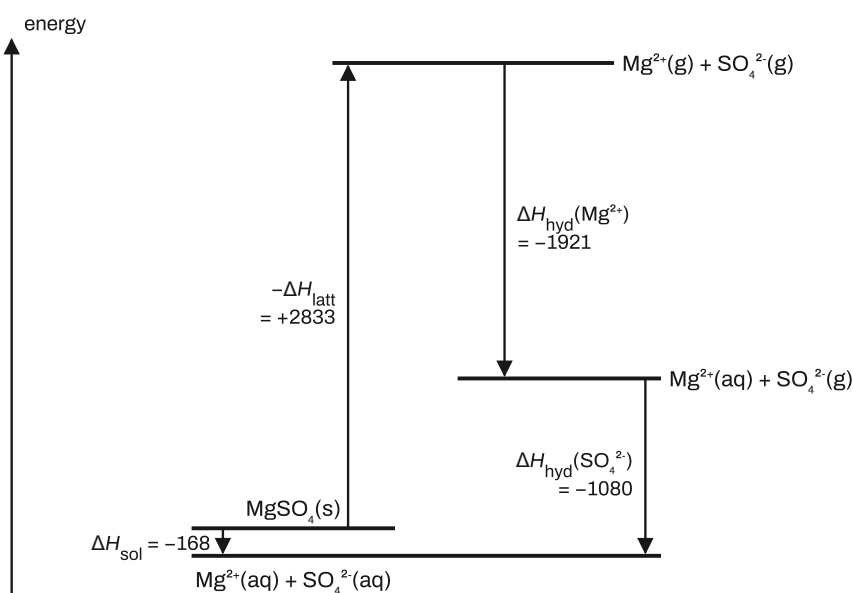
Dissolving an ionic solid can be split into two imaginary steps:

1. Pull the solid apart into gaseous ions. This takes in energy equal to the lattice energy with its sign reversed, $-\Delta H_{\text{latt}}$. The **lattice energy** is the enthalpy change when one mole of an ionic solid forms from its gaseous ions, so it is always negative (exothermic).
2. Put the gaseous ions into water. The **enthalpy change of hydration**, ΔH_{hyd} , is the enthalpy change when one mole of gaseous ions dissolves in water to give a solution. It is exothermic, because the ions attract the polar water molecules: a cation attracts the δ^- oxygen atoms, an anion the δ^+ hydrogen atoms.

By Hess's law,

$$\Delta H_{\text{sol}} = \Delta H_{\text{hyd}}(\text{cation}) + \Delta H_{\text{hyd}}(\text{anions}) - \Delta H_{\text{latt}}.$$

Count every ion: for $\text{M}(\text{OH})_2$ there are **two** hydroxide ions.



For magnesium sulfate, with the approximate literature values in the table below (these are not in the Data Booklet; a question gives you any values it wants you to use):

$$\Delta H_{\text{sol}} = (-1921) + (-1080) - (-2833) = -168 \text{ kJ mol}^{-1}.$$

Both terms are huge and ΔH_{sol} is the small difference between them, so a small change in either one decides which way ΔH_{sol} moves. The more exothermic (or less endothermic) ΔH_{sol} is, the more soluble the compound tends to be. (Entropy matters too, but the syllabus explains the trends with enthalpy only.)

What sets the size of each term. Both depend on the **charge** and the **radius** of the ions:

- A higher charge or a smaller radius means stronger attraction, so a **more exothermic** lattice energy and a **more exothermic** hydration enthalpy.
- Down Group 2 the cation gets bigger with the same charge, so **both** ΔH_{latt} and $\Delta H_{\text{hyd}}(\text{M}^{2+})$ become **less exothermic**.

The question is which one falls **more**. That depends on the size of the anion.

Sulfates: large anion, solubility decreases

The lattice energy depends on the attraction between cation and anion, which falls as the distance between their centres, $r_+ + r_-$, grows. The sulfate ion is large, so r_- is the bigger part of that distance. Making the cation bigger changes the total distance only a little, so the **lattice energy falls only a little** from MgSO_4 to BaSO_4 .

The hydration enthalpy of the cation depends on the cation alone, so it **falls a lot** from Mg^{2+} to Ba^{2+} .

The hydration term falls more, so ΔH_{sol} becomes **less exothermic** (more endothermic) and the sulfates become **less soluble** down the group.

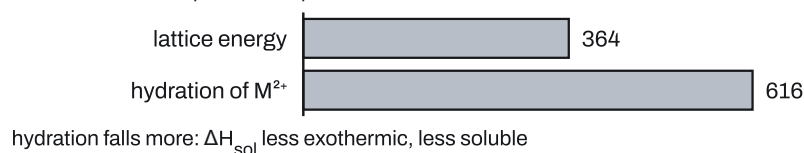
Hydroxides: small anion, solubility increases

The hydroxide ion is small, so now the cation's radius is a large part of $r_+ + r_-$. Making the cation bigger increases the distance a lot, so the **lattice energy falls more** than the hydration enthalpy. ΔH_{sol} becomes **more exothermic** (less endothermic), and the hydroxides become **more soluble** down the group. The fluoride ion, F^- , is also small, so the Group 2 fluorides follow the hydroxide pattern.

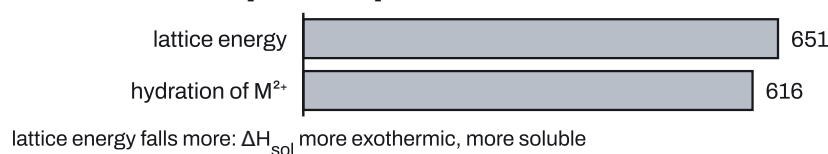
Approximate literature values / kJ mol^{-1}	Mg	Ba	Change from Mg to Ba
$\Delta H_{\text{hyd}}(\text{M}^{2+})$	−1921	−1305	616 less exothermic
$\Delta H_{\text{latt}}(\text{MSO}_4)$	−2833	−2469	364 less exothermic
$\Delta H_{\text{latt}}(\text{M}(\text{OH})_2)$	−2993	−2342	651 less exothermic

Other hydration values used here: SO_4^{2-} −1080, OH^- −460. These are approximate: different sources quote values that differ by up to about 100 kJ mol^{-1} , so use them only to see the pattern, which is the same in every source.

Sulfates, $\text{MgSO}_4 \rightarrow \text{BaSO}_4$



Hydroxides, $\text{Mg}(\text{OH})_2 \rightarrow \text{Ba}(\text{OH})_2$



Bar length: how much less exothermic each term becomes from Mg to Ba, in kJ mol^{-1}

With these values, ΔH_{sol} goes from −168 (MgSO_4) to $(-1305) + (-1080) - (-2469) = +84 \text{ kJ mol}^{-1}$ (BaSO_4). Worked example 4 does the same for the hydroxides.

What you see. A more soluble hydroxide gives a higher concentration of OH^- in a saturated solution, so the **pH of a saturated solution rises** down the group. A less soluble sulfate precipitates sooner: a barium ion solution needs only a little sulfuric acid before white barium sulfate appears.

The reactions you use in these questions

From the AS topic, written with M for any Group 2 metal:

- oxide with water: $\text{MO} + \text{H}_2\text{O} \rightarrow \text{M}(\text{OH})_2$
- oxide with acid: $\text{MO} + \text{H}_2\text{SO}_4 \rightarrow \text{MSO}_4 + \text{H}_2\text{O}$, or with 2HCl to give $\text{MCl}_2 + \text{H}_2\text{O}$
- metal with acid: $\text{M} + 2\text{HCl} \rightarrow \text{MCl}_2 + \text{H}_2$ (fizzing)
- hydrogencarbonate on heating: $\text{M}(\text{HCO}_3)_2 \rightarrow \text{MCO}_3 + \text{CO}_2 + \text{H}_2\text{O}$

Group 2 hydroxides are strong bases

A dissolved Group 2 hydroxide is fully ionised, and each formula unit gives **two** hydroxide ions: $\text{M}(\text{OH})_2 \rightarrow \text{M}^{2+} + 2\text{OH}^-$. So $[\text{OH}^-] = 2 \times [\text{M}(\text{OH})_2]$. Then, at 298 K,

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]}, \quad \text{pH} = -\log[\text{H}^+], \quad \text{or} \quad \text{pOH} = -\log[\text{OH}^-], \quad \text{pH} = 14.00 - \text{pOH}.$$

Key facts and formulas

"Given" means it is in the 9701 Data Booklet (or given in the question). "Learn it" means you must remember it.

Formula	Status
Ionic radii: Mg^{2+} 0.065, Ca^{2+} 0.099, Sr^{2+} 0.113, Ba^{2+} 0.135 nm	Given
Relative atomic masses (Mg 24.3, Ca 40.1, Sr 87.6, Ba 137.3, O 16.0, H 1.0)	Given
$K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ at 298 K	Given
Thermal stability of nitrates and carbonates increases down the group (radius up, charge density down, anion polarised less)	Learn it
$\text{MCO}_3 \rightarrow \text{MO} + \text{CO}_2$	Learn it
$\text{M}(\text{NO}_3)_2 \rightarrow \text{MO} + 2\text{NO}_2 + \frac{1}{2}\text{O}_2$	Learn it
$\text{M}(\text{HCO}_3)_2 \rightarrow \text{MCO}_3 + \text{CO}_2 + \text{H}_2\text{O}$	Learn it
$\Delta H_{\text{sol}} = \sum \Delta H_{\text{hyd}}(\text{ions}) - \Delta H_{\text{latt}}$	Learn it

Formula	Status
Down the group: ΔH_{latt} and ΔH_{hyd} both become less exothermic	Learn it
Sulfates (large anion): ΔH_{hyd} falls more, ΔH_{sol} less exothermic, solubility decreases	Learn it
Hydroxides (small anion): ΔH_{latt} falls more, ΔH_{sol} more exothermic, solubility increases	Learn it
$[\text{OH}^-] = 2 \times [\text{M}(\text{OH})_2]$; $\text{pH} = -\log[\text{H}^+]$; $\text{pH} + \text{pOH} = 14.00$ at 298 K	Learn it

How to do it

In the Paper 4 questions we have tagged for this topic (37 questions, 2021–2025, of which 27 are different and the rest are the same question on another variant), the types came up this often. 30 of the 37 questions mix more than one type, so the counts add up to 77.

Question type	Questions
Explain or compare thermal stability	26
Write the equation for a thermal decomposition	20
Explain a solubility trend with energy terms	19
Products or observations for metals and oxides with water or acids	6
pH or amount for a Group 2 hydroxide solution	6

1. Explain or compare thermal stability (26 questions)

- State the trend:** thermal stability **increases** down the group (the decomposition temperature rises).
- Name the cause:** the **cation's** ionic radius increases (or its charge density decreases). Say which ion: " Ba^{2+} is larger than Ca^{2+} ", not just "the radius increases".
- Give the effect on the anion:** the larger cation **polarises (distorts) the anion** (nitrate or carbonate) **less**, so the N–O or C–O bonds are **weakened less**.

Which of two compounds decomposes at a lower temperature? Name the one with the **smaller** cation, then give points 2 and 3 for that pair. For example, $\text{Ca}(\text{NO}_3)_2$ decomposes at a lower temperature than $\text{Sr}(\text{NO}_3)_2$, because Ca^{2+} is smaller than Sr^{2+} and so distorts the nitrate ion more.

Variations you will meet:

- Another anion "like the nitrates":** hydroxides, hydrogencarbonates or ethanedioates. Use exactly the same three steps; only the anion's name changes.

- **A cation from outside Group 2** (a Cu^{2+} , Zn^{2+} or Pb^{2+} salt): compare its **ionic** radius (from a table or the Data Booklet) with the Group 2 ions. Smaller cation, less stable compound. If the question lists radii, find every Group 2 ion that is smaller than the new one: only those compounds are less stable.
- **Group 1 salts**: the trend and the reason are the same down Group 1.
- **Estimate a temperature**: if one nitrate starts to decompose at a given temperature, a compound above it in the group gets a lower value and one below it a higher value.
- **Energy version**: the enthalpy change of decomposition becomes **more positive** down the group, because the small oxide ion makes $\Delta H_{\text{latt}}(\text{MO})$ fall faster than $\Delta H_{\text{latt}}(\text{MCO}_3)$.

2. Write the equation for a thermal decomposition (20 questions)

1. **Carbonate**: $\text{MCO}_3 \rightarrow \text{MO} + \text{CO}_2$.
2. **Nitrate**: $\text{M}(\text{NO}_3)_2 \rightarrow \text{MO} + 2\text{NO}_2 + \frac{1}{2}\text{O}_2$, or double it: $2\text{M}(\text{NO}_3)_2 \rightarrow 2\text{MO} + 4\text{NO}_2 + \text{O}_2$.
3. **Hydrogencarbonate**: $\text{M}(\text{HCO}_3)_2 \rightarrow \text{MCO}_3 + \text{CO}_2 + \text{H}_2\text{O}$; as an ionic equation, $2\text{HCO}_3^- \rightarrow \text{CO}_3^{2-} + \text{CO}_2 + \text{H}_2\text{O}$.
4. **Check the atoms balance** and that the formulas are right (the oxide of a 2+ ion is MO; of a 1+ ion, M_2O).

Variations you will meet:

- **"Decomposes in a similar way to the Group 2 compounds"**: use the same pattern for CuCO_3 , $\text{Cu}(\text{NO}_3)_2$ or $\text{Zn}(\text{NO}_3)_2$, giving the metal oxide. For lithium nitrate the oxide is Li_2O : $2\text{LiNO}_3 \rightarrow \text{Li}_2\text{O} + 2\text{NO}_2 + \frac{1}{2}\text{O}_2$.
- **Ethanedioates**, MC_2O_4 : the clues give the two gases. If one turns limewater milky and the other is carbon monoxide, $\text{MC}_2\text{O}_4 \rightarrow \text{MO} + \text{CO} + \text{CO}_2$. If the solid left fizzes with acid, it is the carbonate instead: $\text{MC}_2\text{O}_4 \rightarrow \text{MCO}_3 + \text{CO}$.
- **Curly arrows for the carbonate ion**: one arrow from a lone pair on a negative oxygen to the C–O bond next to it (making a C=O), and one from a different C–O bond to its oxygen, which leaves as O^{2-} . That gives $\text{CO}_2 + \text{O}^{2-}$.
- **Observations on heating a hydrated nitrate**: steam (condensation) first, then brown fumes of NO_2 , and a glowing splint relights in the oxygen; the metal oxide is left (white for Group 2; other metals' oxides can be coloured, such as black).

3. Explain a solubility trend with energy terms (19 questions)

For a 3–4 mark "explain", give four points in this order:

1. **The trend**: sulfates **decrease** in solubility down the group; hydroxides **increase**. If asked, write the order, for example most soluble $\text{Sr}(\text{OH})_2 > \text{Ca}(\text{OH})_2 > \text{Mg}(\text{OH})_2$ least soluble.
2. **Both terms fall**: ΔH_{latt} and ΔH_{hyd} **both** become less exothermic down the group (the cation gets bigger).
3. **Which falls more** (the mark most often lost):
 - sulfates: ΔH_{hyd} falls more (it is the dominant factor), because the large anion means ΔH_{latt} changes only a little;

- hydroxides: ΔH_{latt} falls more, because the anion is small.

4. **So ΔH_{sol} :** becomes less exothermic (more endothermic) for sulfates; more exothermic for hydroxides.

Variations you will meet:

- **An anion told to behave like one you know:** treat a large anion (such as chromate, CrO_4^{2-} , or selenate) like sulfate, and a small anion (fluoride) like hydroxide. Then give the four points.
- **Compare two named compounds**, such as "why is CaSO_4 more soluble than SrSO_4 ": write the points as a comparison, " CaSO_4 has the more exothermic ΔH_{latt} and ΔH_{hyd} ; the difference in ΔH_{hyd} is greater; so ΔH_{sol} of CaSO_4 is more exothermic".
- **A practical setting:** equal amounts of two oxides shaken with water (the one with the **less** soluble hydroxide needs **more** water to dissolve), or sulfuric acid added to two solutions until a precipitate appears (the one whose sulfate is **more** soluble needs **more** acid). Then give the energy explanation.
- **pH of saturated solutions:** pH increases down the group, because the hydroxide is more soluble, so $[\text{OH}^-]$ is higher. Say "higher concentration of OH^- ", not "more OH^- ".
- **A definition or a calculation from the energetics topic** in the same question: ΔH_{hyd} is the enthalpy change when one mole of gaseous ions dissolves in water; it is more exothermic for smaller or more highly charged ions, because they attract water molecules more strongly. For ΔH_{sol} use $\Delta H_{\text{sol}} = \sum \Delta H_{\text{hyd}} - \Delta H_{\text{latt}}$, with a $\times 2$ for two anions.

4. Products or observations for metals and oxides with water or acids (6 questions)

1. **Oxide + water** gives the hydroxide only: $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$. No hydrogen.
2. **Oxide + acid** gives a salt and water: $\text{MgO} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}$.
3. **Metal + acid** gives a salt and hydrogen, for example with ethanoic acid $\text{Ca} + 2\text{CH}_3\text{COOH} \rightarrow (\text{CH}_3\text{COO})_2\text{Ca} + \text{H}_2$.
4. **Observations:** fizzing (effervescence) for any metal with acid. With **dilute sulfuric acid**, magnesium dissolves completely, but barium fizzes and a **white solid** forms, because barium sulfate is insoluble; it coats the metal and the fizzing stops.

These parts are usually followed by "explain why A is more (or less) soluble than B", which is type 3.

5. pH or amount for a Group 2 hydroxide solution (6 questions)

From a mass to a pH:

1. Moles of the oxide or hydroxide: $n = \frac{m}{M_r}$. One mole of oxide gives one mole of hydroxide.
2. Moles of $\text{OH}^- = 2n$.
3. $[\text{OH}^-] = \frac{\text{moles}}{\text{volume in dm}^3}$.
4. $[\text{H}^+] = \frac{1.00 \times 10^{-14}}{[\text{OH}^-]}$, then $\text{pH} = -\log[\text{H}^+]$ (or use $\text{pH} = 14.00 - \text{pOH}$). Give the number of decimal places asked.

From a pH back to a mass: reverse it. $[\text{H}^+] = 10^{-\text{pH}}$, $[\text{OH}^-] = \frac{K_w}{[\text{H}^+]}$, $[\text{M}(\text{OH})_2] = \frac{1}{2}[\text{OH}^-]$, then moles in the volume and mass = $n \times M_r$.

Variations you will meet:

- **Identify the hydroxide:** find $[\text{OH}^-]$ from the pH, halve it, find the moles in the volume, then $M_r = \frac{\text{mass}}{\text{moles}}$. Subtract 2×17.0 for the two OH groups to get A_r of the metal, and match it to the Periodic Table.
- **A given concentration:** $0.050 \text{ mol dm}^{-3} \text{ Ba}(\text{OH})_2$ has $[\text{OH}^-] = 0.100$, so $\text{pH} = 13.00$.
- **A mass or pH for a given solution:** use the steps above forwards or backwards; one mole of oxide gives one mole of hydroxide.
- **Solubility product** (from the Equilibria topic): for $\text{M}(\text{OH})_2$ with solubility s , $K_{\text{sp}} = [\text{M}^{2+}][\text{OH}^-]^2 = s(2s)^2 = 4s^3$, units $\text{mol}^3 \text{ dm}^{-9}$. Don't write $[2\text{OH}^-]$.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
 - **A1** accuracy mark: for a correct answer; it needs the M mark before it.
 - **B1** mark for a correct result on its own, with no method needed.
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Example 1

Magnesium carbonate, strontium carbonate and barium carbonate each break down to the metal oxide and carbon dioxide on strong heating.

(a) Write an equation for the reaction when magnesium carbonate is heated. **[1]**

(b) List the three carbonates in order of increasing thermal stability. **[1]**

(c) Strontium carbonate starts to break down at about 1100°C . Suggest a starting temperature for barium carbonate, and account for the order in (b) in terms of the ions present. **[3]**

(a) $\text{MgCO}_3 \rightarrow \text{MgO} + \text{CO}_2$ **B1**

(b) $\text{MgCO}_3 < \text{SrCO}_3 < \text{BaCO}_3$ **B1**

(c) Any temperature above 1100°C , for example about 1300°C . **B1**

The cation's ionic radius increases from Mg^{2+} (0.065 nm) to Sr^{2+} (0.113 nm) to Ba^{2+} (0.135 nm), so its charge density decreases. **B1**

A larger cation polarises (distorts) the carbonate ion less, so its C–O bonds are weakened less and a higher temperature is needed to break it up. **B1**

Example 2

The Group 2 selenates, MSeO_4 , have a large anion and show the same solubility trend as the Group 2 sulfates.

(a) Put magnesium selenate, calcium selenate and barium selenate in order of solubility, most soluble first. **[1]**

(b) Explain the trend in solubility. Refer to the relevant energy terms. **[3]**

(c) Suggest how the solubility trend of the Group 2 hydroxides differs, and give the reason for the difference. **[2]**

(a) $\text{MgSeO}_4 > \text{CaSeO}_4 > \text{BaSeO}_4$ **B1**

(b) Down the group the cation gets larger, so the lattice energy and the hydration enthalpy of the cation **both** become less exothermic. **M1**

The selenate ion is large, so the lattice energy changes only a little; the hydration enthalpy changes more (it is the dominant factor). **M1**

So ΔH_{sol} becomes less exothermic (more endothermic) down the group, and the selenates become less soluble. **A1**

(c) The hydroxides become **more** soluble down the group. **B1** The hydroxide ion is small, so the lattice energy falls more than the hydration enthalpy, and ΔH_{sol} becomes more exothermic. **B1**

Example 3

A small piece of a Group 2 metal, X, reacts with cold water. The alkaline solution that forms is made up with water to give a solution of the hydroxide, $\text{X}(\text{OH})_2$.

(a) Write a general equation for the reaction of X with water. **[1]**

(b) A separate solution is made by dissolving 0.912 g of pure $\text{X}(\text{OH})_2$ and making it up to 200.0 cm^3 . Its pH at 298 K is 12.88. Use these data to identify X. **[4]**

(c) Suggest how the pH of a saturated solution of $\text{X}(\text{OH})_2$ compares with that of a saturated solution of magnesium hydroxide, and explain why. **[2]**

(a) $\text{X} + 2\text{H}_2\text{O} \rightarrow \text{X}(\text{OH})_2 + \text{H}_2$ **B1**

(b) $[\text{H}^+] = 10^{-12.88} = 1.318 \times 10^{-13}$, so $[\text{OH}^-] = \frac{1.00 \times 10^{-14}}{1.318 \times 10^{-13}} = 0.07586 \text{ mol dm}^{-3}$ **M1**

Each $\text{X}(\text{OH})_2$ gives two OH^- , so $[\text{X}(\text{OH})_2] = 0.03793 \text{ mol dm}^{-3}$, and in 200.0 cm^3 : $0.03793 \times 0.2000 = 7.586 \times 10^{-3} \text{ mol}$ **M1**

$$M_r = \frac{0.912}{7.586 \times 10^{-3}} = 120.2 \text{ M1}$$

$$A_r(\text{X}) = 120.2 - 2 \times 17.0 = 86.2, \text{ closest to strontium (87.6): X is } \mathbf{\text{strontium. A1}}$$

(c) Higher. **B1** Strontium hydroxide is more soluble than magnesium hydroxide (the hydroxides get more soluble down the group), so its saturated solution has a higher concentration of OH^- . **B1**

Example 4

The table gives approximate values of some energy changes, in kJ mol^{-1} .

	value
lattice energy of $\text{Mg}(\text{OH})_2$	−2993
lattice energy of $\text{Ba}(\text{OH})_2$	−2342
enthalpy change of hydration of Mg^{2+}	−1921
enthalpy change of hydration of Ba^{2+}	−1305
enthalpy change of hydration of OH^-	−460

(a) Calculate the enthalpy change of solution of $\text{Mg}(\text{OH})_2$ and of $\text{Ba}(\text{OH})_2$. [3]

(b) Use your answers and the data to explain why barium hydroxide is much more soluble than magnesium hydroxide. [3]

(a) $\Delta H_{\text{sol}} = \Delta H_{\text{hyd}}(\text{M}^{2+}) + 2\Delta H_{\text{hyd}}(\text{OH}^-) - \Delta H_{\text{latt}}$ **M1** (the $\times 2$ for two hydroxide ions)

$$\text{Mg}(\text{OH})_2: (-1921) + 2(-460) - (-2993) = +152 \text{ kJ mol}^{-1} \text{ A1}$$

$$\text{Ba}(\text{OH})_2: (-1305) + 2(-460) - (-2342) = +117 \text{ kJ mol}^{-1} \text{ A1}$$

(b) From Mg to Ba both terms become less exothermic: the lattice energy by $2993 - 2342 = 651$ and the hydration enthalpy of the cation by $1921 - 1305 = 616 \text{ kJ mol}^{-1}$. **M1**

The lattice energy changes more, because the hydroxide ion is small, so the cation's size has a big effect on the distance between the ions. **M1**

So ΔH_{sol} becomes less endothermic (from +152 to +117), and barium hydroxide is more soluble. **A1**

Common mistakes

- **Polarising the wrong ion.** Examiners see "the cation is polarised" or "the anion polarises the cation" every year. It is always the **cation** that polarises the **anion**.
- **Leaving out the anion step.** Many answers give "the cation radius increases" and stop. The mark for "so the nitrate (or carbonate) ion is polarised less" is the one most often missing.

- **Talking about attraction between the ions** instead of distortion of the anion when explaining thermal stability. Stability here is about the bonds inside the anion.
- **Ambiguous "radius increases"**. Say which ion: " Ca^{2+} has a smaller ionic radius than Ba^{2+} ". Using **atomic** radii instead of ionic radii for a metal such as lead is a common error.
- **Answering about the wrong property or compound**. Some candidates explain solubility when asked about thermal stability, or explain the hydroxides when the question was about the sulfates (or compare with sodium hydroxide).
- **Missing "which term changes more"**. Saying both ΔH_{latt} and ΔH_{hyd} decrease is only one mark; the comparison is what explains the trend.
- **Mixing up "ionic radius" and "charge density"** as if they were two separate factors. Charge density comes from the charge and the radius; give radius and charge as the factors, then say what they do to the attraction.
- **Unbalanced nitrate equations**, or products such as the metal, NO or M_2O for a $2+$ ion. The pattern is 2 : 2 : 4 : 1.
- **pH calculations**: forgetting the $\times 2$ for two OH^- ions, or the volume conversion from cm^3 to dm^3 . Label every number you write (" $[\text{OH}^-] = 0.0158$ "): examiners cannot credit an unlabelled number.

How to get the marks

- **"Describe the trend"** needs the direction and "down the group" (or "from Mg to Ba"). **"Explain"** then needs the chain of reasons, each step a mark.
- **Thermal stability** is usually 3 marks: trend; cation radius increases / charge density decreases; anion polarised (distorted) less, or the N–O / C–O bonds weakened less.
- **Solubility** is usually 3–4 marks: order or trend; both ΔH_{latt} and ΔH_{hyd} less exothermic; which one changes more; so ΔH_{sol} less (or more) exothermic. The marks are often awarded independently, so write all four points even if you are unsure of one.
- **"State which ... decomposes at a lower temperature. Explain"**: the name alone is not enough; the name and its reason (the smaller cation) often share a mark.
- **"Suggest"** means apply the same idea to an unfamiliar compound (ethanedioates, chromates, fluorides, a transition-metal nitrate). Follow the rule for its anion's size, or compare its cation's ionic radius.
- **Equations**: $\frac{1}{2}\text{O}_2$ is accepted; the doubled equation earns the same mark. State symbols only when asked.
- **Calculations**: show every step with labels, and give the precision asked ("two decimal places" for a pH is strict). Wrong earlier values are carried through (error carried forward), so keep going.

Quick check

1. Write equations for the thermal decomposition of (a) barium nitrate and (b) magnesium hydrogencarbonate.
2. State which of strontium carbonate and barium carbonate needs the higher temperature to decompose, and explain why.

3. Put $\text{Mg}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$ and $\text{Ba}(\text{OH})_2$ in order of solubility, most soluble first, and give the energy term whose change decides the order.
4. Calculate the pH of $0.0350 \text{ mol dm}^{-3}$ strontium hydroxide solution at 298 K, to two decimal places.
5. (a) Name the products when strontium oxide reacts with dilute nitric acid. (b) Describe what you see when a small piece of strontium is added to an excess of dilute sulfuric acid.

Answers

1. (a) $2\text{Ba}(\text{NO}_3)_2 \rightarrow 2\text{BaO} + 4\text{NO}_2 + \text{O}_2$. (b) $\text{Mg}(\text{HCO}_3)_2 \rightarrow \text{MgCO}_3 + \text{CO}_2 + \text{H}_2\text{O}$.
2. Barium carbonate. Ba^{2+} has a larger ionic radius (0.135 nm, against 0.113 nm for Sr^{2+}), so a lower charge density; it polarises the carbonate ion less, so the C–O bonds are weakened less.
3. $\text{Ba}(\text{OH})_2 > \text{Sr}(\text{OH})_2 > \text{Mg}(\text{OH})_2$. The lattice energy becomes less exothermic down the group by more than the hydration enthalpy does, so ΔH_{sol} becomes more exothermic.
4. $[\text{OH}^-] = 2 \times 0.0350 = 0.0700 \text{ mol dm}^{-3}$. $[\text{H}^+] = \frac{1.00 \times 10^{-14}}{0.0700} = 1.429 \times 10^{-13}$, so $\text{pH} = -\log(1.429 \times 10^{-13}) = 12.85$.
5. (a) Strontium nitrate and water: $\text{SrO} + 2\text{HNO}_3 \rightarrow \text{Sr}(\text{NO}_3)_2 + \text{H}_2\text{O}$. (b) Fizzing (bubbles of hydrogen) at first, and a white solid (strontium sulfate, which is sparingly soluble) forms on the metal; the fizzing slows and stops before all the metal has reacted.

Past questions to try

- **9701/42/M/J/23 Q1**: curly arrows for the carbonate ion, the stability trend, and its energy explanation with lattice energies.
- **9701/42/M/J/21 Q2**: nitrate stability, lead(II) nitrate compared by ionic radius, an ethanedioate equation and a pH.
- **9701/43/O/N/21 Q4**: a practical setting for thermal stability and for sulfate solubility.
- **9701/42/M/J/24 Q1**: the sulfate solubility order and its energy explanation.