



Cambridge (CIE) A Level Chemistry



Your notes

Similarities & Trends in the Properties of the Group 2 Metals, Magnesium to Barium, & Their Compounds

Contents

- * Similarities, Trends & Compounds of Magnesium to Barium



Ionic Radius & Thermal Stability of Group 2 Nitrates & Carbonates

- The Group 2 nitrates and carbonates become **more thermally stable** going down the group
- The **charge density of the cation** (Group 2 metal ion) and the **polarisation of the anion** (the nitrate and carbonate ion) attribute towards this increased stability

Trends in thermal stability going down the group

- All Group 2 metals form 2+ ions as they lose two electrons from their valence shells
- The metal cations at the top of the group are **smaller in size than those at the bottom**
 - For example, the atomic radius of beryllium (the first element in Group 2) is 112 pm whereas the atomic radius of calcium (further down the group) is 197 pm
- The metal cations at the top of Group 2, therefore, have the **greatest charge density** as the same charge (2+) is packed into a smaller volume
- As a result, smaller Group 2 ions have a **greater polarising effect** on neighbouring negative ions
- When a carbonate or nitrate ion approaches the cation, it becomes polarised
 - This is because the metal cation draws the electrons in the carbonate or nitrate ion towards itself
- The **more polarised the anion** is, the **less heat is required** to thermally decompose them
- Therefore, the **thermal stability increases down the group**
 - As down the group, the cation becomes larger
 - Thus has a smaller charge density
 - And a smaller polarising effect on the carbonate or nitrate anion
 - So the anion is less polarised
 - Therefore, more heat is required to thermally decompose them

Solubility & Enthalpy Change of Solution of Group 2 Hydroxides & Sulfates

- The **solubility** of Group 2 hydroxides **increases** down the group



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- In contrast, the Group 2 sulfates show a **decrease in solubility** going down the group
- Compounds that have very **low** solubility are said to be **sparingly soluble**
 - For example, calcium sulfate (CaSO_4) has low solubility as only 0.21 g of CaSO_4 dissolves in 100 g of water
- Most of the sulfates are soluble in warm water with the exception of **barium sulfate** which is **insoluble**

Enthalpy change of hydration and lattice energy

- The **standard enthalpy of solution** ($\Delta H_{\text{sol}}^{\ddagger}$) is the energy **absorbed** or **released** when 1 mole of ionic solid dissolves in enough water to form a dilute solution (under standard conditions)
 - The $\Delta H_{\text{sol}}^{\ddagger}$ can be either **exothermic** or **endothermic**
- For example, the $\Delta H_{\text{sol}}^{\ddagger}$ of sodium chloride (NaCl) is +3.9 kJ mol⁻¹



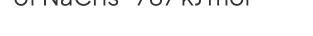
OR



- This means, that 3.9 kJ mol⁻¹ of energy is **absorbed** when one mole of NaCl is dissolved in enough water to form a dilute solution

$$\Delta H_{\text{sol}}^{\ddagger} = \Delta H_{\text{hyd}}^{\ddagger} - \Delta H_{\text{latt}}^{\ddagger}$$

- The **lattice (formation) energy** is the energy **released** when **gaseous ions** combine to form **one mole** of an ionic compound under (standard conditions)
 - Since energy is released when an ionic compound is formed, the $\Delta H_{\text{latt}}^{\ddagger}$ is always **exothermic**
- For example, the $\Delta H_{\text{latt}}^{\ddagger}$ of NaCl is -787 kJ mol⁻¹



- This means, that 787 kJ mol⁻¹ of energy is released when NaCl is formed from its gaseous ions
- The **standard enthalpy of hydration** is the **energy released when gaseous ions dissolve in enough water to form a dilute solution** (under standard conditions)
 - Since energy is released when gaseous ions become hydrated, the $\Delta H_{\text{hyd}}^{\ddagger}$ is always **exothermic**
- For example, the $\Delta H_{\text{hyd}}^{\ddagger}$ of the sodium (Na^+) ion is -406 kJ mol⁻¹



- This means, that 406 kJ mol⁻¹ of energy is released when Na^+ ions become hydrated

Trends of enthalpy change of solution

- Going down the group, the $\Delta H_{\text{latt}}^{\ddagger}$ of the ionic compounds **decreases**



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- The positively charged **cations** become larger, so there is more space between the negatively and positively charged ions in the ionic compound
- So there are weaker **attractive forces** between the ions
- Therefore, there is less energy released upon formation of the ionic compound from its gaseous ions
 - Therefore, the ΔH_{latt}^∞ becomes **less exothermic**
- Going down the group, the ΔH_{hyd}^∞ also **decreases**
 - The positively charged ions become larger going down the group
 - The **ion-dipole bonds** between the cations and water molecules get **weaker**
 - This means that **less energy** is released when the gaseous Group 2 ions become hydrated
 - The ΔH_{hyd}^∞ , therefore, becomes **less exothermic**
- For **Group 2 hydroxides**:
 - Hydroxide ions are relatively small ions
 - The ΔH_{latt}^∞ falls faster than the ΔH_{hyd}^∞
 - The **enthalpy change of solution** is, therefore, more **exothermic** going down the group
- For **Group 2 sulfates**:
 - Sulfate ions are relatively large ions
 - The ΔH_{latt}^∞ falls slower than the ΔH_{hyd}^∞ enthalpy
 - The ΔH_{sol}^∞ will become more **endothermic** going down the group
- The more **exothermic** the ΔH_{sol}^∞ the **more soluble** the compound
 - This is why the **sulfates** become **less** soluble going down the group and the **hydroxides more** soluble