

23.2 Enthalpies of solution and hydration

Name: _____ Class: _____ Date: _____

Total: 23 marks

KEY WORDS enthalpy change of hydration 水合焓变 • enthalpy change of solution 溶解焓变
• enthalpy change 焓变 • lattice energy 晶格能

Objective

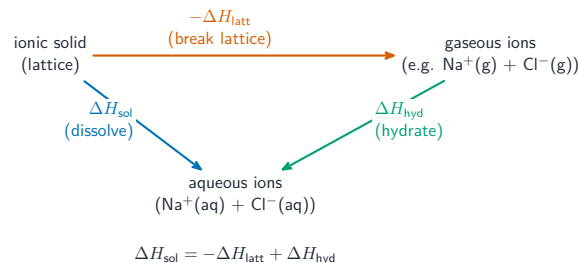
Build the skills to answer exam questions on **enthalpies of solution and hydration** — ΔH_{sol} 溶解焓 and ΔH_{hyd} 水合焓, the energy cycle, and the effect of ionic charge and radius.

You must be able to:

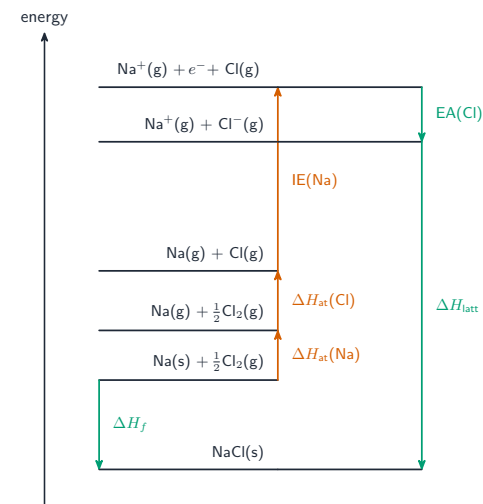
- define enthalpy of solution and enthalpy of hydration
- construct and use an energy cycle linking them with lattice energy
- explain the effect of ionic charge and radius on the magnitude of hydration enthalpy

1 Worked examples

■ The dissolving cycle



Dissolving = breaking the lattice ($-\Delta H_{\text{latt}}$) then hydrating the ions (ΔH_{hyd})



Lattice energy from a Born-Haber cycle

$$\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \sum \Delta H_{\text{hyd}}$$

- ΔH_{hyd} — change when 1 mol of gaseous ions is dissolved in water (always –, water attracts the ion).
- ΔH_{sol} — change when 1 mol of solute dissolves to form a solution.

■ Charge and radius

Hydration is **more exothermic** for ions with a **higher charge** and a **smaller radius** (higher charge density attracts the polar water more strongly).

2 Practice

2.1 Define enthalpy change of hydration.

[2]

2.2 State two factors that make hydration of an ion more exothermic.

[2]

3 Exam-style questions

3.1 Enthalpy of hydration is most exothermic for: [1]

- ☐ A a large, singly-charged ion
☐ B a small, highly-charged ion
☐ C a large, highly-charged ion
☐ D a neutral atom

3.2 Which equation correctly links the enthalpies? [1]

- ☐ A $\Delta H_{\text{sol}} = \Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$
☐ B $\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$
☐ C $\Delta H_{\text{sol}} = \Delta H_{\text{latt}} - \Delta H_{\text{hyd}}$
☐ D $\Delta H_{\text{sol}} = -\Delta H_{\text{hyd}}$

3.3 For a salt, $\Delta H_{\text{latt}} = -780 \text{ kJ mol}^{-1}$ and the sum of the hydration enthalpies of its ions is -820 kJ mol^{-1} .

(a) Calculate the enthalpy change of solution. [2]

(b) State whether the salt dissolving is exothermic or endothermic. [1]

3.4 Explain why the hydration enthalpy of Mg^{2+} is more exothermic than that of Na^+ . [2]

3.5 The table gives data for two Group 2 chlorides.

	$\Delta H_{\text{LE}}^{\ominus} / \text{kJ mol}^{-1}$	$\Delta H_{\text{sol}}^{\ominus} / \text{kJ mol}^{-1}$
MgCl_2	-2526	-155
CaCl_2	-2258	-83

$$\Delta H_{\text{hyd}}^{\ominus}(\text{Cl}^-) = -364 \text{ kJ mol}^{-1}.$$

(a) Draw the energy cycle linking lattice enthalpy, enthalpy of solution and the enthalpies of hydration, and use it to calculate $\Delta H_{\text{hyd}}^{\ominus}(\text{Mg}^{2+})$. [5]

(b) Explain why $\Delta H_{\text{hyd}}^{\ominus}(\text{Mg}^{2+})$ is more exothermic than $\Delta H_{\text{hyd}}^{\ominus}(\text{Ca}^{2+})$. [3]

(c) Both dissolvings are exothermic, yet the magnitudes are small compared with the individual terms. Explain why. [2]

(d) State one reason why a calculated lattice enthalpy from a purely ionic model may differ from the experimental value. [2]