

23.2 Enthalpies of solution and hydration

Name: _____ Class: _____ Date: _____

Total: 11 marks

Objective

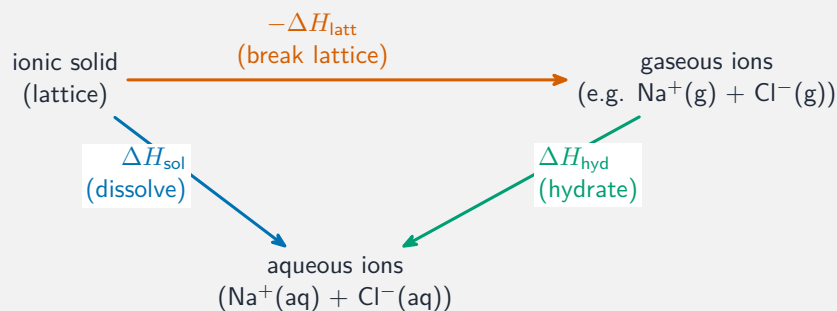
Build the skills to answer exam questions on **enthalpies of solution and hydration** — ΔH_{sol} 溶解焓 and ΔH_{hyd} 水合焓, the energy cycle, and the effect of ionic charge and radius.

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

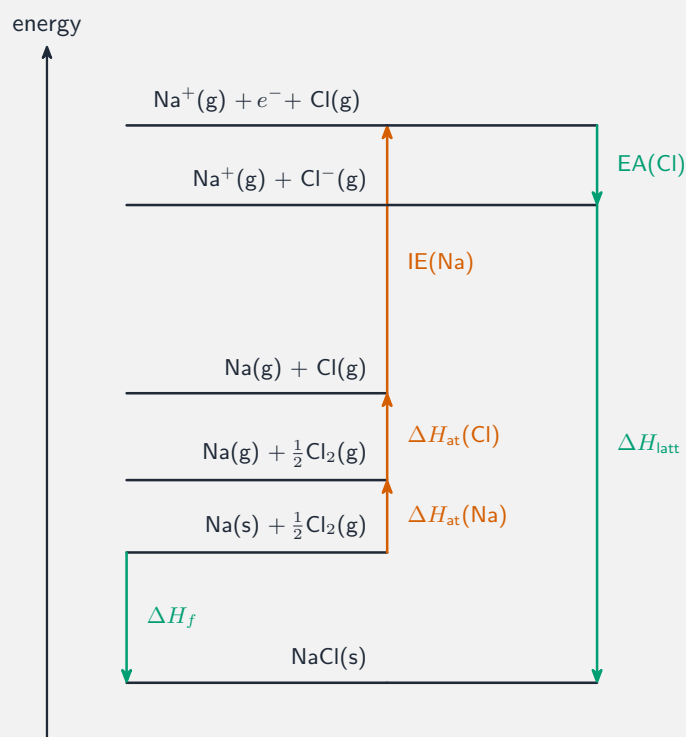
1 Worked examples

■ The dissolving cycle



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

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]

4 Go further

You are now ready for the real exam questions on this subtopic. Open the **23.2 Enthalpies of solution and hydration** past-paper sheet in the Library, or try these in **Practice mode**:

- 9701/41 N25 —Q1 (enthalpies of solution and hydration)
 - 9701/42 N25 —Q1 (enthalpies of solution and hydration)
-

- 9701/44 N25 —Q2 (enthalpies of solution and hydration)

Solutions

2.1 the enthalpy change when 1 mol of gaseous ions is dissolved in (a large amount of) water.

2.2 a higher ionic charge; a smaller ionic radius.

3.1 B. A small, highly-charged ion has the highest charge density, so the most exothermic hydration.

3.2 B. $\Delta H_{\text{sol}} = -\Delta H_{\text{latt}} + \Delta H_{\text{hyd}}$.

3.3 (a) $\Delta H_{\text{sol}} = -(-780) + (-820) = +780 - 820 = -40 \text{ kJ mol}^{-1}$.
(b) exothermic.

3.4 Mg^{2+} has a higher charge and a smaller radius than Na^{+} , so a higher charge density that attracts the polar water molecules more strongly —more exothermic hydration.