A-LEVEL CHEMISTRY • EXERCISE SHEET

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

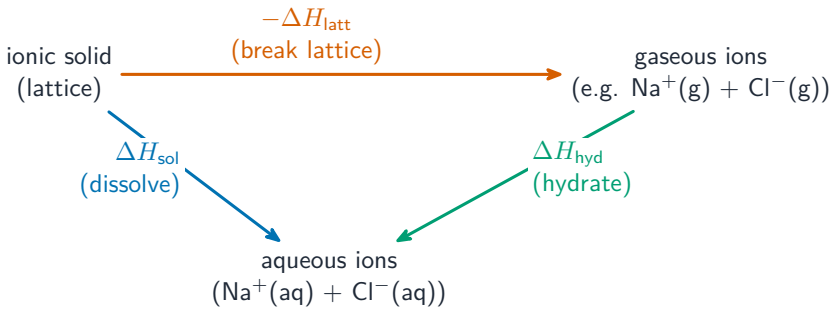
Build the skills to answer exam questions on **enthalpies of solution and hydration** — $\Delta H_{\mathbf{sol}}$ 溶解焓 and $\Delta H_{\mathbf{hyd}}$ 水合焓, the energy cycle, and the effect of ionic charge and radius.

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

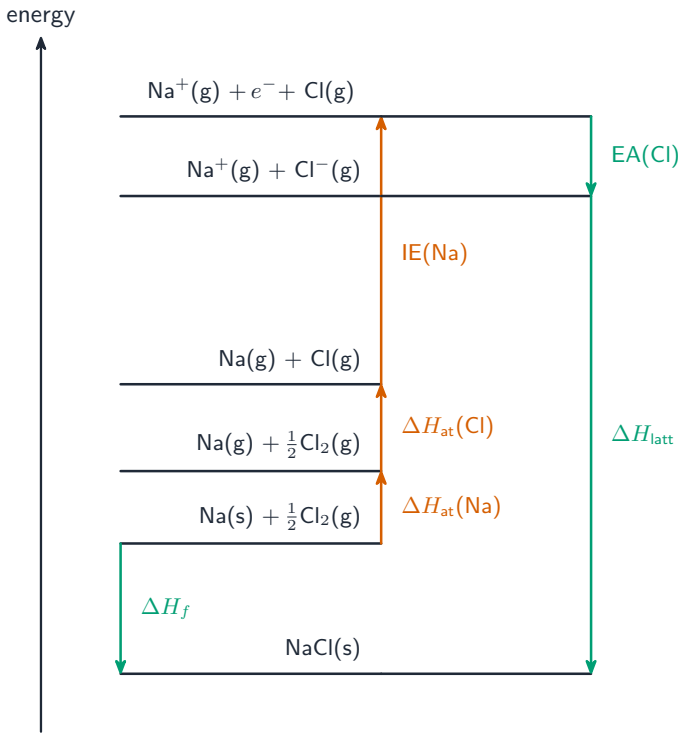
1 Worked examples

The dissolving cycle



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

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



Lattice energy from a Born-Haber cycle

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

- $\Delta H_{\mathbf{hyd}}$ — change when 1 mol of gaseous ions is dissolved in water (always −, water attracts the ion).
- $\Delta H_{\mathbf{sol}}$ — 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 B
C D

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 B
C D

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

3.3 For a salt, $\Delta H_{latt} = -780$ 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.

	$\Delta H^{\ominus}_{\mathbf{LE}}$ / **kJ mol**$^{-1}$	$\Delta H^{\ominus}_{\mathbf{sol}}$ / **kJ mol**$^{-1}$
$MgCl_2$	-2526	-155
$CaCl_2$	-2258	-83

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

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

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