Solutions

Each answer shows the steps, then the **result**.

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)

- $\Delta H_{\text{sol}} = -(-780) + (-820) = +780 - 820 = -40$ 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

3.5

(a)

- The cycle goes: solid $\rightarrow$ gaseous ions ($-\Delta H_{\text{LE}}$) $\rightarrow$ aqueous ions ($\sum \Delta H_{\text{hyd}}$), and directly solid $\rightarrow$ aqueous ions (ΔH_{sol}). So $\Delta H_{\text{sol}} = -\Delta H_{\text{LE}} + \Delta H_{\text{hyd}}(Mg^{2+}) + 2\Delta H_{\text{hyd}}(Cl^{-})$; $-155 = 2526 + \Delta H_{\text{hyd}}(Mg^{2+}) + 2(-364)$; $\Delta H_{\text{hyd}}(Mg^{2+}) = -155 - 2526 + 728 = -1953$ kJ mol^{-1}

(b)

- Mg^{2+} is the smaller ion but carries the same 2+ charge, so it has the greater charge density; it attracts the $\delta-$ oxygen of the water molecules more strongly, releasing more energy per mole

(c)

- ΔH_{sol} is the small **difference** between two very large opposing terms — the energy needed to pull the lattice apart and the energy released on hydration; they nearly cancel, so the net value is small

(d)

- The purely ionic model assumes perfect spheres with no covalent character; a small, highly charged cation polarises a large anion, giving some covalent character and a larger experimental lattice enthalpy than calculated