InkChemistry

ANSWER BOOK

Enthalpies of solution and hydration

Thermodynamics · Unit C08

AQA 3.1.8.1 · OCR A 5.2.1

17 questions · 57 marks

NAVY EDITION

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SECTION A · Short, 1 to 2 marks

A1 ESH-Q3

B1 $Mg^{2+}(g) \rightarrow Mg^{2+}(aq)$, starting from the gaseous ion. **B1** $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$, with two chloride ions. [2]

SECTION B · Standard, 3 to 5 marks

B1 ESH-Q1

B1 enthalpy of hydration: the enthalpy change when one mole of gaseous ions. **B1** dissolves in water to form one mole of aqueous ions; reject 'one mole of compound' for the hydration enthalpy. **B1** enthalpy of solution: the enthalpy change when one mole of an ionic solid dissolves in enough water that the ions are well separated and do not interact. **B1** hydration starts from separate gaseous ions, whereas solution starts from the solid lattice. [4]

B2 ESH-Q4

B1 around Na^+ the water molecules point their δ- oxygen atoms towards the ion. **B1** around Cl^- the water molecules point a δ+ hydrogen atom towards the ion. **B1** the gaseous ion attracts the oppositely charged end of each polar water molecule, an ion–dipole attraction; reject 'covalent bonds form between the ion and water'. **B1** attractions form and none are broken, so energy is released and the enthalpy change is negative. [4]

B3 ESH-Q9

M1 $\Delta sol H = \Delta LEH(\text{dissociation}) + \Delta hyd H(Mg^{2+}) + 2 \times \Delta hyd H(Cl^-)$. **M1** the chloride value is doubled: $2 \times (-364) = -728$ kJ mol⁻¹. **M1** the hydration enthalpies sum to $-1920 - 728 = -2648$ kJ mol⁻¹; error carried forward from M2 if the chloride was not doubled. **A1** $\Delta sol H = +2525 - 2648 = -123$ kJ mol⁻¹, three significant figures, negative sign required. [4]

B4 ESH-Q15

B1 the claim is contradicted by observation: sodium chloride dissolves readily in water at room temperature. **B1** the enthalpy change alone does not decide whether a change is feasible. **B1** dissolving increases entropy, because ions spread through the water are more disordered than ions in a lattice. **B1** the $T\Delta S$ term is larger than the small positive ΔH, so $\Delta G = \Delta H - T\Delta S$ is negative and dissolving is feasible; the claim is wrong. [4]

B5 ESH-Q2

M1 $\Delta sol H = \Delta LEH(\text{dissociation}) + \Delta hyd H(Na^+) + \Delta hyd H(F^-)$, so $\Delta hyd H(F^-) = \Delta sol H - \Delta LEH(\text{dissociation}) - \Delta hyd H(Na^+)$. **M1** $\Delta hyd H(F^-) = +1 - 913 - (-406)$. **A1** $\Delta hyd H(F^-) = -506$ kJ mol⁻¹, three significant figures, negative sign required. [3]

B6 ESH-Q5

B1 Mg^{2+} has a larger charge than Na^+. **B1** Mg^{2+} is also smaller than Na^+, so it has a much higher charge density. **B1** it attracts the δ- oxygen atoms of the water molecules more strongly, so more energy is released when the ion is hydrated. [3]

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B7 ESH-Q6

B1 the ions all have a charge of +1 but the ionic radius increases down the group. **B1** the water molecules cannot get as close to the centre of the charge, so the charge density falls. **B1** the attraction between the ion and the δ- oxygen of the water is weaker, so less energy is released. [3]

B8 ESH-Q8

M1 $\Delta_{sol}H = \Delta LEH(\text{dissociation}) + \Delta_{hyd}H(K^+) + \Delta_{hyd}H(Br^-)$. **M1** the hydration enthalpies sum to -322 + (-348) = -670 kJ mol^{-1}. **A1** $\Delta_{sol}H$ = +689 - 670 = +19 kJ mol^{-1}, two significant figures; accept 19 with 'endothermic' in place of the + sign. [3]

B9 ESH-Q10

M1 $\Delta LEH(\text{dissociation}) = \Delta_{sol}H - \Delta_{hyd}H(Li^+) - \Delta_{hyd}H(Cl^-)$, from rearranging the cycle. **M1** $\Delta LEH(\text{dissociation})$ = -37 - (-519) - (-364). **A1** $\Delta LEH(\text{dissociation})$ = +846 kJ mol^{-1}, three significant figures, positive sign shown. [3]

B10 ESH-Q12

B1 the hydroxide ion is small, so the distance between the ions depends strongly on the size of the cation. **B1** as the cation gets larger down the group, the lattice enthalpy falls quickly, and the hydration enthalpy of the cation also falls. **B1** the lattice enthalpy falls faster than the hydration enthalpy, so the enthalpy of solution becomes less endothermic and the hydroxides more soluble. [3]

B11 ESH-Q14

B1 both are exothermic, because each ion attracts polar water molecules by ion–dipole attractions. **B1** the water molecules face Na^+ with their δ- oxygen atoms, but face Cl^- with a δ+ hydrogen atom. **B1** Na^+ is the smaller ion with the higher charge density, so its hydration enthalpy is the more exothermic, -406 kJ mol^{-1} against -364 kJ mol^{-1}. [3]

B12 ESH-Q17

B1 a measured volume or mass of water is placed in a polystyrene cup and its starting temperature recorded. **B1** a weighed mass of the salt is added and stirred until it dissolves, and the highest or lowest temperature is recorded. **B1** $q = mc\Delta T$ is calculated and divided by the amount of salt in moles, with a positive sign if the temperature fell and a negative sign if it rose. [3]

B13 ESH-Q19

B1 breaking up the lattice and spreading the ions through the water increases the disorder of the ions. **B1** Mg^{2+} is small and doubly charged, with a high charge density, so it holds a shell of water molecules very tightly. **B1** those water molecules lose much of their freedom to move, and this increase in order outweighs the disorder gained, so the overall entropy change is negative. [3]

B14 ESH-Q20

B1 in the hydrate the ions are already surrounded by water molecules within the crystal. **B1** much of the hydration enthalpy was released when the hydrate formed, so far less is released when it dissolves. **B1** the energy needed to break up the hydrated crystal is then greater than the energy released, so the enthalpy of solution is positive. [3]

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B15 ESH-Q21

B1 water is polar, and its δ- oxygen and δ+ hydrogen atoms are attracted to the cations and anions, releasing hydration enthalpy close in size to the lattice enthalpy. **B1** hexane is non-polar, so it has no partial charges to attract the ions and almost no energy is released when ions are surrounded by it. **B1** in hexane nothing repays the large lattice enthalpy of dissociation, so the lattice does not break up. [3]

SECTION C · Long, 6 marks or more

C1 ESH-Q16

Indicative content [6]

B1 the enthalpy of solution is the lattice enthalpy of dissociation plus the hydration enthalpies of the two ions. **B1** the sulfate ion is large, so the distance between the ions is dominated by the sulfate and changes little from Mg^{2+} to Ba^{2+}. **B1** the lattice enthalpy of dissociation therefore falls only slightly from $MgSO_4$ to $BaSO_4$. **B1** Ba^{2+} is much larger than Mg^{2+}, so its charge density is lower and its hydration enthalpy much less exothermic, -1360 against -1920 kJ mol^{-1}. **B1** the hydration enthalpy falls by much more than the lattice enthalpy, so the enthalpy of solution is much more endothermic for $BaSO_4$. **B1** the more endothermic the enthalpy of solution, the less soluble the salt, so $MgSO_4$ dissolves and $BaSO_4$ does not.

Levels

Level 3 (5–6 marks) the enthalpy of solution is set out as the sum of the two terms for both salts, the lattice enthalpy is shown to change little because the large sulfate ion dominates the spacing, the hydration enthalpy is shown to fall much more because of the lower charge density of Ba^{2+}, and the more endothermic enthalpy of solution is linked to insolubility.

Level 2 (3–4 marks) both terms are used and Ba^{2+} is said to hydrate less exothermically because of its size; why the lattice enthalpy hardly changes, or how the enthalpy of solution decides solubility, is not explained.

Level 1 (1–2 marks) lattice or hydration enthalpy is mentioned on its own; barium sulfate is often said to be insoluble because its lattice is the stronger one.

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