[AQA A2 Specimen Paper 1 2015]

A 5.00 g sample of potassium chloride was added to 50.0 g of water initially at 20.0 °C. The mixture was stirred and as the potassium chloride dissolved, the temperature of the solution decreased. The temperature of the water decreased to 14.6 °C.

a) Calculate a value, in $kJmol^{-1}$, for the enthalpy of solution of potassium chloride.
(You should assume that only the 50.0 g of water changes in temperature and that the specific heat capacity of water is $4.18\ JK^{-1}g^{-1}$.)

① Calculate the temperature change:

$20.0 - 14.6 = 5.40$

② Calculate the energy taken in:

$q = 50 \times 4.18 \times 5.4$

$= +1128.6\ J$

$q = mc\Delta T$

③ Calculate the moles of KCl that reacts:

$\text{moles} = \dfrac{5.00}{(39.1 + 35.5)}$

$= 0.0670\ \text{moles}$

$\text{moles} = \dfrac{\text{mass}}{M_r}$

④ Use these values to find the enthalpy change per mole:

$\Delta H_{soln.} = \dfrac{+1128.6}{0.0670}$

$= +16839\ (Jmol^{-1})$

$\Rightarrow \underline{+16.8\ kJmol^{-1}}$

$\Delta H = \dfrac{q}{\text{moles}}$

the reaction is endothermic so ΔH is a positive value.

The enthalpy of solution of calcium chloride is –82.9 kJmol^{-1}. The enthalpies of hydration for calcium ions and chloride ions are –1650 and –364 kJmol^{-1}, respectively.

b) Use these values to calculate a value for the lattice enthalpy of dissociation of calcium chloride.

① Set up a Hess's Law diagram:

$$\begin{tikzcd}
CaCl_2 \arrow[rr, "-82.9"] \arrow[dr, "\Delta H_{lat.}"'] & & Ca^{2+}_{(aq)} + 2Cl^-_{(aq)} \\
& Ca^{2+}_{(g)} + 2Cl^-_{(g)} \arrow[ur, "\Delta H_{hyd.}"'] &
\end{tikzcd}$$

② Calculate the total enthalpy of hydration:

$$\Delta H_{hyd.} = -1650 + 2(-364)$$
$$= -2378\ kJmol^{-1}$$

③ Use this to find lattice dissociation enthalpy:

$$-82.9 = \Delta H_{lat.} + -2378$$

$$\Rightarrow \Delta H_{lat.} = -82.9 + 2378$$
$$= +2295.1\ldots$$
$$\Rightarrow \underline{+2300\ kJmol^{-1}}$$

Remember to treat the arrows as vectors.