

[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 $4.18 \text{ JK}^{-1}\text{g}^{-1}$.)

① Calculate the temperature change:

$$20.0 - 14.6 = 5.40$$

② Calculate the energy taken in:

$$q = 50 \times 4.18 \times 5.4$$

$$= +1128.6 \text{ J}$$

$$q = mc\Delta T$$

③ Calculate the moles of KCl that reacts:

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

$$= 0.0670 \text{ moles}$$

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

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

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

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

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

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

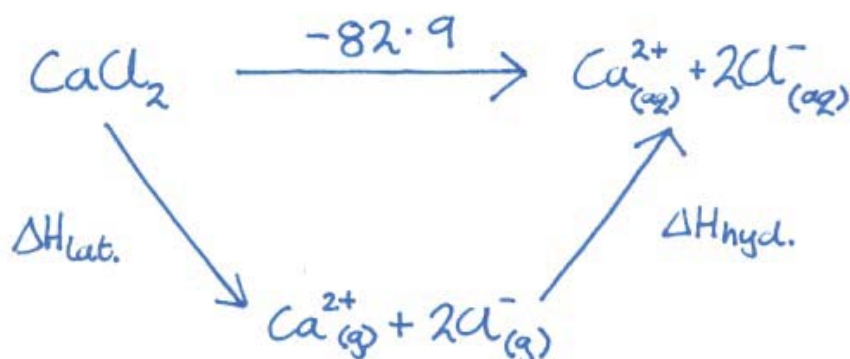
the reaction is endothermic so ΔH is a positive value.



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:



② Calculate the total enthalpy of hydration:

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

③ Use this to find lattice dissociation enthalpy:

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

$$\begin{aligned}
 \Rightarrow \Delta H_{\text{lat.}} &= -82.9 + 2378 \\
 &= +2295.1\dots
 \end{aligned}$$

Remember to
break the arrows
as vectors.

$$\Rightarrow \underline{\underline{+2300 \text{ kJmol}^{-1}}}$$

