

Entropy, Gibbs energy and reaction feasibility

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Lab answers

Model answers follow the sample measurements. Your own results will differ a little because heat is lost to the room, but your ΔH should be positive, your temperature fall should grow with the mass of solid and your explanation should mention entropy.

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Question 1 Moles dissolved

$n = \text{mass} \div 80.0$: Trial 1 = $4.00 \div 80.0 = \mathbf{0.0500 \text{ mol}}$; Trial 2 = $\mathbf{0.100 \text{ mol}}$; Trial 3 = $\mathbf{0.150 \text{ mol}}$.

Question 2 Heat absorbed

$q = 100.0 \times 4.18 \times \Delta T$: Trial 1 = $418 \times 2.8 = \mathbf{1170 \text{ J}}$; Trial 2 = $418 \times 5.5 = \mathbf{2299 \text{ J}}$; Trial 3 = $418 \times 8.4 = \mathbf{3511 \text{ J}}$.

Question 3 Enthalpy of solution

$\Delta H = +q \div n$ (endothermic): Trial 1 = $1.170 \div 0.0500 = 23.4$; Trial 2 = $2.299 \div 0.100 = 23.0$; Trial 3 = $3.511 \div 0.150 = 23.4 \text{ kJ/mol}$. Mean = $\mathbf{+23.3 \text{ kJ/mol}}$ (positive: heat is absorbed).

Question 4 Comparing with the data

$(25.7 - 23.3) \div 25.7 \times 100 = \mathbf{9.3 \%}$ lower than the data-book value.

Question 5 Entropy and spontaneity

$\Delta G = 23.3 - 294 \times 0.1087 = 23.3 - 32.0 = \mathbf{-8.7 \text{ kJ/mol}}$. The entropy increase as the ions spread through the solution gives a $-T\Delta S$ term larger than the positive ΔH , so the process is spontaneous and the solution cools.

Question 6 Minimum entropy change

$\Delta G < 0$ needs $\Delta S > \Delta H \div T = 23\,300 \div 294 = \mathbf{79.3 \text{ J K}^{-1} \text{ mol}^{-1}}$. The actual value of 108.7 is larger, so dissolving is spontaneous.

Question 7 Improving the lab

The cup gains heat from the room during the experiment, so the measured fall is smaller,

giving a **low ΔH** . Improvement: use a lid, a temperature probe and plot temperature against time, extrapolating back to the time of mixing.