

Entropy, Gibbs energy and spontaneity

AP Chemistry · Lab answers

The sample data give ΔH close to the accepted value of +25.7 kJ/mol. Your own values will differ with your cup, thermometer and mixing. Check that every calculation uses consistent units and keep three significant figures.

Level: AP Chemistry · **Curriculum:** AP Chem Unit 9

Question 1 Temperature change

Initial 21.9 °C, lowest 16.2 °C, so $\Delta T = 5.7$ °C (a fall).

Question 2 Heat absorbed

$q = 108.0 \times 4.18 \times 5.7 = 2573$ J (2.57 kJ).

Question 3 Enthalpy change per mole

$n = 8.00 / 80.04 = 0.1000$ mol. $\Delta H = +2.57 / 0.1000 = +25.7$ kJ/mol. The sign is positive because heat is absorbed.

Question 4 Why spontaneous

ΔG must be negative. With ΔH positive, $-T\Delta S$ must be negative and larger in size than ΔH , so **ΔS is positive**: ordered ions in the crystal become dispersed in solution.

Question 5 Minimum entropy change

A negative ΔG needs $T\Delta S$ above ΔH , so ΔS must exceed $25\,700 / 295 = 87$ J/(mol·K).

Question 6 Heat gain error

Heat entering from the room makes the lowest temperature higher, so ΔT is too small and ΔH is **underestimated**. Improve it by extrapolating the cooling graph back to 90 s and adding a lid.

Question 7 Larger sample

The amount of solid doubles, so q doubles to 5146 J. Mass is now 116.0 g, so $\Delta T = 5146 / (116.0 \times 4.18) = 10.6$ °C.