

Journal Prompt: The Limits of Your Calorimeter

Every calorimetry experiment contains built-in error. Understanding where error comes from separates a skilled chemist from someone who simply inserts numbers into a formula.

Suppose you completed a calorimetry experiment to find the molar enthalpy of a combustion reaction. Your calculated ΔH is more negative than the accepted value. Write a journal entry. Explain at least two specific sources of heat loss or measurement error that could cause this difference. Describe how each source would affect your calculated $Q = mc\Delta T$ value ([Outcome](#), [Outcome](#)).

Where to start:

- What assumption does $Q = mc\Delta T$ make about where the heat released actually goes?
- Where might thermal energy escape before your thermometer or probe registers it?
- Does your calorimeter design (a polystyrene cup versus an open container) change how much heat is lost to the surroundings?
- How would incomplete combustion affect the amount of energy actually released, apart from measurement error?

Stretch:

Use $Q = mc\Delta T$ to calculate the heat released in the combustion of a known mass of a substance. Explain using $Q = mc\Delta T$ you would expect the reported molar enthalpy to move closer to the accepted value if your change succeeded ([Outcome](#)).