

Journal Prompt: Trusting the Numbers

A calorimetry experiment always loses some heat to the surroundings, yet chemists still trust the data enough to calculate enthalpy changes.

Prompt: Explain how you would use calorimetry data to determine the enthalpy change of a reaction, and discuss why the calculated value might differ from an accepted value found using standard enthalpies of formation ([Outcome](#), [Outcome](#)).

Where to start:

- What does $Q = mc\Delta T$ represent in a calorimetry setup, and what does each variable represent ([Outcome](#))?
- How is enthalpy, ΔH , for a chemical reaction ([Outcome](#))?
- What assumptions does a simple calorimeter make about heat loss to the surroundings?
- How would you use $Q = mc\Delta T$ to find your result ([Outcome](#))?

Stretch:

Explain how Hess' law could let you check your experimental value against a calculated value built from other known reactions, without repeating the calorimetry ([Outcome](#)).