

12.1 Calorimetry and chemical energy

Name: _____ Class: _____ Date: _____

Total: 10 marks

KEY WORDS enthalpy change • exothermic

Objective

Calorimetry and chemical energy. Supported HL concept preparation; original classroom practice, not an official IB paper or assessed learner work.

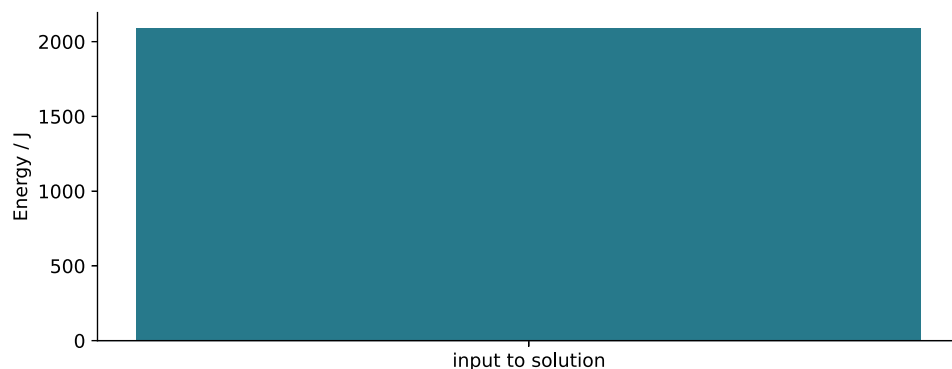
1 Key knowledge and worked examples

■ What would explain this observation?

- A cup warms when two solutions react. The temperature rise measures energy transferred to the surroundings; it does not directly equal the **enthalpy change**.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ Build the model

- **Exothermic** reactions transfer energy to surroundings. Endothermic reactions take energy from surroundings. Bond breaking requires energy; bond formation releases energy.
- exothermic: Transferring energy to the surroundings; enthalpy change: Heat change at constant pressure for a stated process.



■ Choose evidence that can test it

- Use energy transferred = mass $\times$ specific heat capacity $\times$ temperature change. Convert joules to kilojoules before dividing by reaction amount. An exothermic molar enthalpy change has a negative sign.
- Use insulation and a lid, measure starting temperatures consistently, stir, and record a temperature-time series. Estimate the reaction temperature from an appropriate extrapolation rather than ignoring cooling during measurement.

■ Work from known quantities

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- Known: 100 g solution rises by 5.0 °C; specific heat capacity is 4.18 J per gram per degree. $q = mc\Delta T$. $q = 100 \times 4.18 \times 5.0 = 2,090 \text{ J} = 2.09 \text{ kJ}$. If 0.050 mol reacts, $\Delta H = -q/n = -2.09/0.050 = -41.8 \text{ kJ per mole}$.

Example:

50 g water rises 4 °C. Use $c = 4.2 \text{ J per gram per degree}$ to find q . Use the same sequence: known quantities $\rightarrow$ model $\rightarrow$ relation $\rightarrow$ substitution $\rightarrow$ unit and interpretation.

■ Check the conclusion and its limits

- Heat loss usually lowers the observed temperature rise. The solution gaining heat and the reaction losing heat have opposite signs.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Bond breaking releases energy. This claim is false: Heat loss usually lowers the observed temperature rise. The solution gaining heat and the reaction losing heat have opposite signs.

Key:

Calorimetry and chemical energy: Use energy transferred = mass $\times$ specific heat capacity $\times$ temperature change. Convert joules to kilojoules before dividing by reaction amount. An exothermic molar enthalpy change has a negative sign.

2 Practice

2.1 Which process requires energy? Choose one answer.

- **A** Breaking a covalent bond
- **B** Forming every chemical bond
- **C** Cooling a hot solution

[1]

2.2 50 g water rises 4 °C. Use $c = 4.2 \text{ J per gram per degree}$ to find q . Show the relation, substitution and units.

[2]

2.3 Decide whether this claim is true and justify: Bond breaking releases energy.

[2]

2.4 Transferring energy to the surroundings: write the technical term.

[1]

3 Exam-style questions

3.1 Fictional solution gains 1.5 kJ when 0.030 mol reacts. Find molar reaction enthalpy using the stated heat-transfer model, and explain the sign and effect of uncorrected heat loss.

[4]