Reactivity 1.2 · Energy cycles in reactions · supported focus

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

Supported SL core focus. First assessment 2025; current brief acquired; full Chemistry guide not acquired. Remaining guide, assessment and practical requirements retain their recorded holds.

Prerequisites: enthalpy change and calorimetry from Reactivity 1.1, balancing equations, and arithmetic with signed quantities. Each lesson develops its own method before independent transfer.

These are original or explicitly fictional teaching examples, not actual measurements or completed assessed learner investigations.

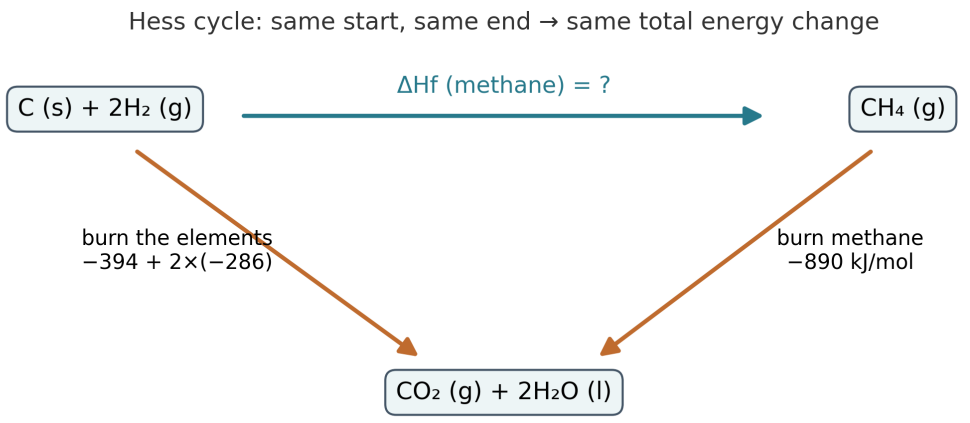
Energy cycles in reactions

■ What would explain this observation?

- Some enthalpy changes cannot be measured directly: carbon and hydrogen will not simply combine into methane in a calorimeter. How can we still know the enthalpy change of that reaction?
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ Build the model

- **Hess's law**: the total enthalpy change of a reaction is the same whatever route is taken, because energy is conserved. An **energy cycle** links the direct route to an indirect route through shared substances, so the unknown change is found from measurable ones.
- With **standard enthalpy of formation** data the rule is $\Delta H = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants})$. With **standard enthalpy of combustion** data the cycle runs through the combustion products: $\Delta H = \Sigma\Delta H_c(\text{reactants}) - \Sigma\Delta H_c(\text{products})$. Elements in their standard states have $\Delta H_f = 0$.
- energy cycle: A diagram linking two routes between the same start and end points; standard enthalpy of formation: The enthalpy change when one mole of a compound forms from its elements under standard conditions; standard enthalpy of combustion: The enthalpy change when one mole burns completely under standard conditions.



■ Choose evidence that can test it

- Measure the enthalpies of combustion of carbon, hydrogen and methane by calorimetry; each is directly measurable. The cycle then predicts the formation enthalpy of methane, and the same method applied to a reaction that CAN be measured directly (such as neutralisation) confirms that route and direct measurement agree within experimental uncertainty.
- State the states of every substance and the standard conditions; an enthalpy value without states is incomplete.

■ 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: $\Delta Hc(C) = -394$ kJ/mol, $\Delta Hc(H_2) = -286$ kJ/mol, $\Delta Hc(CH_4) = -890$ kJ/mol. The target is ΔHf for C (s) + $2H_2$ (g) → CH_4 (g). By Hess's law along the combustion route: $\Delta Hf = [\Delta Hc(C) + 2\Delta Hc(H_2)] - \Delta Hc(CH_4)$. Substitution: $\Delta Hf = [-394 + 2 \times (-286)] - (-890) = -966 + 890 = -76$ kJ/mol. The negative sign means the formation of methane from its elements is exothermic.

Example:

Ethanol burns: C_2H_5OH (l) + $3O_2$ (g) → $2CO_2$ (g) + $3H_2O$ (l). Using $\Delta Hf(CO_2) = -394$, $\Delta Hf(H_2O) = -286$ and $\Delta Hf(C_2H_5OH) = -278$ kJ/mol, find ΔHc of ethanol with the same sequence: known quantities → model → relation → substitution → unit and interpretation. (Answer: $\Delta H = [2 \times (-394) + 3 \times (-286)] - [-278 + 0] = -1646 + 278 = -1368$ kJ/mol; oxygen is an element in its standard state, so its ΔHf is 0.)

Check the conclusion and its limits

- Hess's law assumes the routes share the same start and end states and that the tabulated values apply at the stated conditions. Real calorimetry loses heat, so measured values carry uncertainty; the law itself is exact.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

The enthalpy change of a reaction depends on the route by which the reaction happens. This claim is false: Hess's law follows from the conservation of energy – the total change depends only on the start and end states, so an unmeasurable change can be computed from a measurable indirect route.

Key:

Energy cycles in reactions: ΔH is route-independent. Use $\Delta H = \Sigma\Delta Hf(\text{products}) - \Sigma\Delta Hf(\text{reactants})$ from formation data, or $\Delta H = \Sigma\Delta Hc(\text{reactants}) - \Sigma\Delta Hc(\text{products})$ from combustion data, keeping signs and states explicit.