

Exercise walk-through

6.9

Hess's Law

AP Chemistry

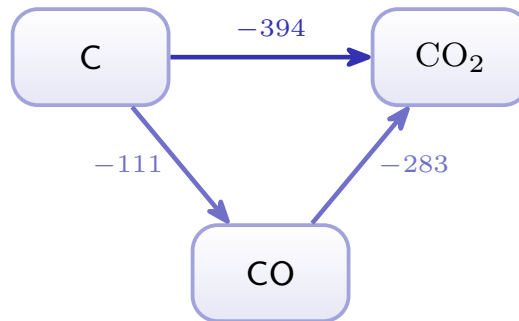
Enthalpy of formation and Hess's law

From a table of ΔH_f°

$$\Delta H_{rxn}^\circ = \sum \Delta H_f^\circ(\text{prod}) - \sum \Delta H_f^\circ(\text{react})$$

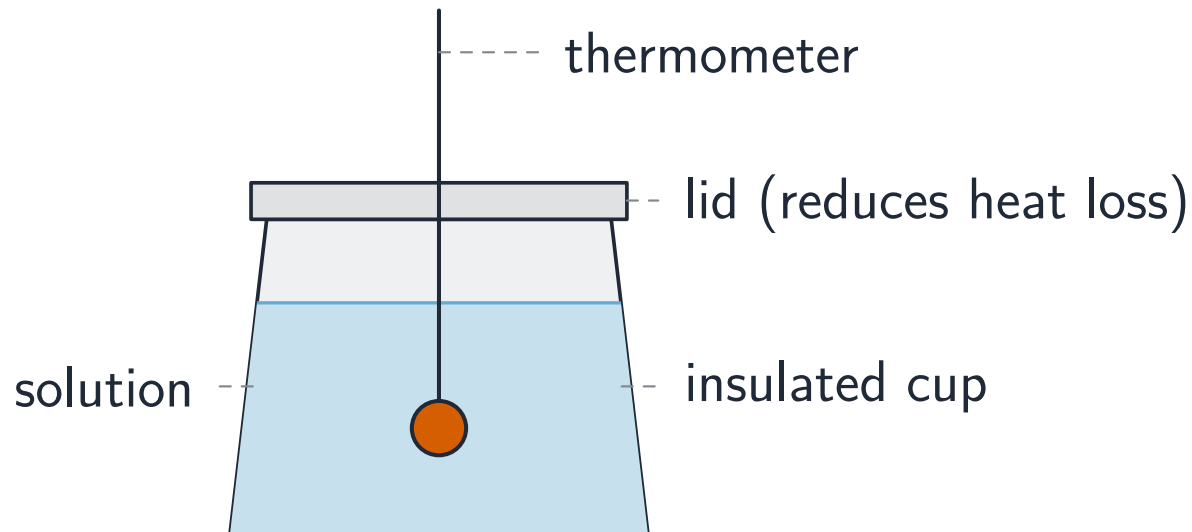
An element in its standard state has $\Delta H_f^\circ = 0$.

Hess's law : ΔH is a **state function** – add steps. **Reverse** flips the sign; **scale** multiplies it.



Same start
and end $\Rightarrow$ same ΔH .

Heat Capacity and Calorimetry



Calorimetry: measure the temperature change of a known mass of solution

Heat, equilibrium, and the enthalpies

heat

flows from **hot to cold** until objects reach **thermal equilibrium**

equilibrium means

equal temperature – not equal energy, and not no motion

enthalpy of reaction

ΔH_{rxn} – the heat released or absorbed at constant pressure

the sign

negative for exothermic, positive for endothermic

standard enthalpy of formation

ΔH_f° – one mole of a compound from its elements in their standard states

why it is useful

$$\Delta H_{\text{rxn}} = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$$

An element in its standard state has $\Delta H_f^\circ = 0$ **by definition**. Forgetting that is what breaks the subtraction.

Enthalpy of Formation

formation ΔH_f :

elements → 1 mol of the compound

combustion ΔH_c :

1 mol substance + O₂ → CO₂ + H₂O

Formation makes a compound from its elements; combustion burns it in oxygen

Endothermic and Exothermic Processes



temperature
rises

exothermic

heat given out, so
the surroundings
get hotter
($\Delta H < 0$)



temperature
falls

endothermic

heat taken in, so
the surroundings
get colder
($\Delta H > 0$)

An exothermic reaction warms the surroundings; an endothermic one cools them

Exercises 2.1-2.3

2.1 State Hess's law in one sentence.

[1]

2.2 If reversing a reaction, what happens to ΔH ?

[1]

2.3 A reaction has $\Delta H = -100$ kJ. What is ΔH if you multiply the reaction by 2? [1]

Exercises 2.1-2.3

2.1 State Hess's law in one sentence.

[1]

2.1

- The enthalpy change of a reaction is the same regardless of the path taken

2.2 If reversing a reaction, what happens to ΔH ?

[1]

2.2

- Its sign changes

2.3 A reaction has $\Delta H = -100$ kJ. What is ΔH if you multiply the reaction by 2? [1]

2.3

- -200 kJ

Exercise 3.1

3.1 Hess's law works because ΔH is a

[1]

☐

A state function (path-independent)

B rate

C catalyst

D temperature

Exercise 3.1

3.1 Hess's law works because ΔH is a

[1]

☐

A state function (path-independent)

B rate

C catalyst

D temperature

3.1 A

- ΔH is a state function

Exercise 3.2 (a)

3.2 Find ΔH for $\text{N}_2 + 2\text{O}_2 \rightarrow 2\text{NO}_2$ using: (1) $\text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}$, $\Delta H = +180 \text{ kJ}$; (2) $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$, $\Delta H = -112 \text{ kJ}$.

(a) State how to combine (1) and (2).

[1]

Exercise 3.2 (a)

3.2 Find ΔH for $\text{N}_2 + 2\text{O}_2 \rightarrow 2\text{NO}_2$ using: (1) $\text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}$, $\Delta H = +180 \text{ kJ}$; (2) $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$, $\Delta H = -112 \text{ kJ}$.

(a) State how to combine (1) and (2).

[1]

(a)

- Add (1) and (2) directly – the 2NO cancels

Exercise 3.2 (b)

3.2 Find ΔH for $\text{N}_2 + 2\text{O}_2 \rightarrow 2\text{NO}_2$ using: (1) $\text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}$, $\Delta H = +180 \text{ kJ}$; (2) $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$, $\Delta H = -112 \text{ kJ}$.

(b) Find the target ΔH .

[2]

Exercise 3.2 (b)

3.2 Find ΔH for $\text{N}_2 + 2\text{O}_2 \rightarrow 2\text{NO}_2$ using: (1) $\text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}$, $\Delta H = +180 \text{ kJ}$; (2) $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$, $\Delta H = -112 \text{ kJ}$.

(b) Find the target ΔH .

[2]

(b)

- $\Delta H = 180 + (-112) = +68 \text{ kJ}$

Exercise 3.3

3.3 Find ΔH for $\text{C} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}$ using $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ (-394 kJ) and $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$ (-283 kJ). Show your reasoning. [4]

Exercise 3.3

3.3 Find ΔH for $\text{C} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}$ using $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ (-394 kJ) and $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$ (-283 kJ). Show your reasoning. [4]

3.3

- Keep reaction 1 as is (-394); reverse reaction 2 ($\text{CO}_2 \rightarrow \text{CO} + \frac{1}{2}\text{O}_2$, $+283$); add – CO_2 cancels, leaving $\text{C} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}$; $\Delta H = -394 + 283 = -111 \text{ kJ}$