Solutions

Marking: **M1** a correct method, **A1** the correct answer, **B1** an independent mark.

Each answer shows the steps, then the **result**.

2.1

- the enthalpy change for a reaction is independent of the route taken
- provided the initial and final conditions are the same

2.2

- oxygen is an element in its standard state
- forming it from itself involves no change, so no enthalpy change

2.3

- $\Delta H_\mathrm{r} = \Delta H_1 + \Delta H_2$
- $= \mathbf{-393\ kJ\,mol^{-1}}$

3.1

- $\Delta H_\mathrm{f}(\mathrm{H_2}) = 0$
- $\Delta H = -85 - (+52)$
- $= \mathbf{-137\ kJ\,mol^{-1}}$

3.2

(a) burning carbon in limited oxygen always gives some $\mathrm{CO_2}$ too

(b) $\mathrm{C} + \frac{1}{2}\mathrm{O_2} \rightarrow \mathrm{CO}$; cycle to $\mathrm{CO_2}$; $\Delta H_\mathrm{f} = -394 - (-283) = \mathbf{-111\ kJ\,mol^{-1}}$

3.3

(a) products: $-1131 - 286 - 394 = -1811$; reactants: $2 \times (-951) = -1902$; $\Delta H = -1811 - (-1902) = \mathbf{+91\ kJ\,mol^{-1}}$

(b) **endothermic**: energy must be supplied all the time to break bonds faster than forming them releases energy

3.4

- cycle down to $2\mathrm{CO_2} + 3\mathrm{H_2O}$
- reactants: $2 \times (-394) + 3 \times (-286) = -1646$
- $\Delta H_\mathrm{f} = -1646 - (-1367)$
- $= \mathbf{-279\ kJ\,mol^{-1}}$

3.5

(a) $\mathrm{CH_3OH} + \frac{3}{2}\mathrm{O_2} \rightarrow \mathrm{CO_2} + 2\mathrm{H_2O}$; cycle with arrows up from $\mathrm{C} + 2\mathrm{H_2} + \frac{1}{2}\mathrm{O_2}$ (and $\frac{3}{2}\mathrm{O_2}$) to both sides; products $-394 + 2 \times (-286) = -966$; $\Delta H_\mathrm{c} = -966 - (-239) = \mathbf{-727\ kJ\,mol^{-1}}$

(b) it is an element in its standard state, so its value is zero

(c) **less** exothermic; heat lost to the surroundings / incomplete combustion

(d) Hess's law; energy is conserved, so the enthalpy change does not depend on the route

(e) bond energies are averages from many compounds; formation values are measured for these exact substances in their standard states