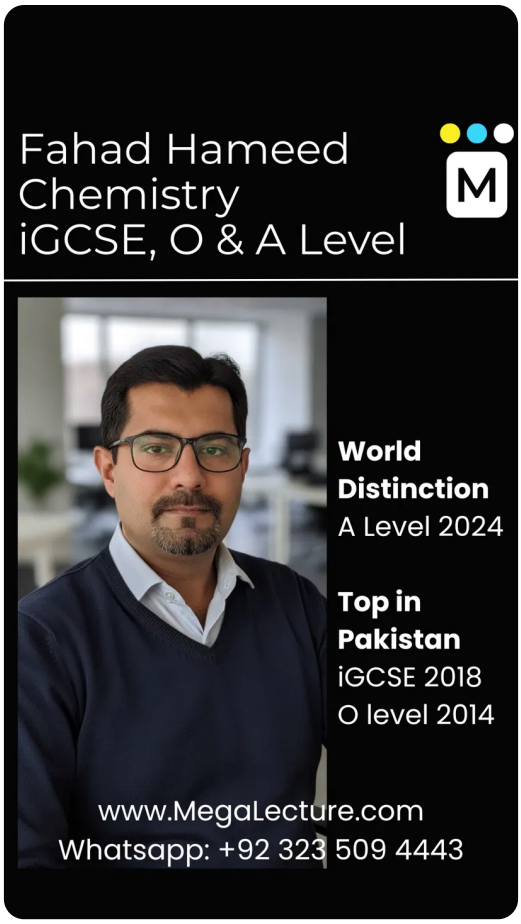
MEGA LECTURE · A LEVEL CHEMISTRY

# Chapter 10: Enthalpy Changes

OCR A Level Chemistry A (H432) — Structured Practice Worksheet




| SPECIFICATION | TOTAL MARKS | SUGGESTED TIME |
|---|---|---|
| 3.1.4 Enthalpy: definitions, calorimetry, Hess's law, bond enthalpies, enthalpy cycles | 86 | 2 hours |

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**Instructions to candidates**

- Answer **all** questions in the spaces provided. Show all working in calculations.
- Take the specific heat capacity of aqueous solutions as 4.18 J $g^{-1}$ $K^{-1}$ and the density of aqueous solutions as 1.00 g $cm^{-3}$, unless told otherwise.
- Marks are shown in brackets [ ] at the end of each part.

## Question 1 — Definitions and enthalpy diagrams 3.1.4

a. Define *standard enthalpy change of reaction*, $\Delta H_r^{\ominus}$. **[2]**

b. State the standard conditions used when measuring standard enthalpy changes. **[2]**

c. Define *standard enthalpy change of formation*, $\Delta H_f^{\ominus}$. **[2]**

d. Define *standard enthalpy change of combustion*, $\Delta H_c^{\ominus}$. **[2]**

e. Sketch a labelled enthalpy level diagram for an exothermic reaction, showing the reactants, products, ΔH and the activation energy. **[4]**

*Sketch the enthalpy level diagram*

**Question 1 total: 12 marks**

## Question 2 — Exothermic and endothermic changes

3.1.4

a. State the sign of ΔH for an exothermic reaction and explain what this means in terms of energy transfer to the surroundings. **[2]**

b. State the sign of ΔH for an endothermic reaction and explain what this means in terms of energy transfer. **[2]**

c. Explain, in terms of bond breaking and bond making, why combustion reactions are exothermic overall. **[3]**

**Question 2 total: 7 marks**

## Question 3 — Calorimetry: combustion

3.1.4 / PAG

A student burns 0.68 g of ethanol, $C_2H_5OH$, and uses the heat released to raise the temperature of 100 g of water from 20.5 °C to 45.8 °C.

a. Calculate the heat energy released, in J, using $q = mc\Delta T$. **[3]**

b. Calculate the amount, in mol, of ethanol burned, and hence calculate the experimental enthalpy of combustion of ethanol in kJ $mol^{-1}$. **[3]**

c. The data book value for the enthalpy of combustion of ethanol is $-1367$ kJ $mol^{-1}$. Compare this with your answer to (b). **[1]**

d.

Suggest two reasons why the experimental value is less exothermic than the data book value, and suggest one improvement to the experiment for each. **[4]**

**Question 3 total: 11 marks**

## Question 4 — Calorimetry: neutralisation

3.1.4 / PAG

50.0 $cm^3$ of 1.00 mol $dm^{-3}$ hydrochloric acid is mixed with 50.0 $cm^3$ of 1.00 mol $dm^{-3}$ sodium hydroxide solution in a polystyrene cup. The temperature rises from 19.2 °C to 26.0 °C.

a. Calculate the heat energy released, assuming the total mass of solution is 100 g. **[3]**

b. Calculate the amount, in mol, of water formed, and hence calculate the enthalpy change of neutralisation in kJ $mol^{-1}$. **[3]**

c. Explain why a polystyrene cup, rather than a glass beaker, is used for this experiment. **[2]**

d. State two ways this experiment could be modified to reduce heat loss to the surroundings. **[2]**

**Question 4 total: 10 marks**

## Question 5 — Hess's law and enthalpy cycles

3.1.4

a. State Hess's law. **[2]**

b. Calculate the enthalpy change for the reaction $C(s) + 2H_2(g) \rightarrow CH_4(g)$, given: $\Delta H_c^{\circ}[C(s)] = -394$ kJ mol$^{-1}$, $\Delta H_c^{\circ}[H_2(g)] = -286$ kJ mol$^{-1}$, $\Delta H_c^{\circ}[CH_4(g)] = -890$ kJ mol$^{-1}$. Draw a labelled Hess's law energy cycle to support your calculation. **[6]**

*Draw the Hess's law cycle here, then complete the calculation below*

c. Calculate the enthalpy change for the reaction $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$, given: $\Delta H_f^{\circ}[CaCO_3(s)] = -1207$ kJ mol$^{-1}$, $\Delta H_f^{\circ}[CaO(s)] = -635$ kJ mol$^{-1}$, $\Delta H_f^{\circ}[CO_2(g)] = -394$ kJ mol$^{-1}$. **[5]**

**Question 5 total: 13 marks**

## Question 6 — Bond enthalpies 3.1.4

a. Define *mean bond enthalpy*. **[2]**

b. Explain why bond enthalpy calculations using mean bond enthalpies from data books usually give only an approximate value for ΔH, rather than an exact one. **[2]**

c. Use the mean bond enthalpies given (in kJ mol$^{-1}$: C–H = 412, C=C = 612, H–H = 436, C–C = 348) to calculate the enthalpy change for the hydrogenation of ethene: $CH_2{=}CH_2(g) + H_2(g) \rightarrow CH_3CH_3(g)$. **[6]**

d. State whether bond breaking is exothermic or endothermic, and whether bond making is exothermic or endothermic. Explain your answer. **[3]**

**Question 6 total: 13 marks**

## Question 7 — Enthalpy cycles using combustion data

3.1.4

Calculate the enthalpy change for the reaction: C(s) + $O_2$(g) → $CO_2$(g), using an alternative route via CO(g), given:
$\Delta H_1$: C(s) + ½$O_2$(g) → CO(g), $\Delta H_1 = -111$ kJ mol$^{-1}$
$\Delta H_2$: CO(g) + ½$O_2$(g) → $CO_2$(g), $\Delta H_2 = -283$ kJ mol$^{-1}$

a. Draw a Hess's law cycle linking C(s) + $O_2$(g), CO(g) + ½$O_2$(g), and $CO_2$(g), and use it to calculate the enthalpy change for the direct combustion of carbon to $CO_2$. **[4]**

*Draw the cycle and show your calculation*

b. Explain why it is not possible to measure $\Delta H_1$ directly by experiment. **[2]**

**Question 7 total: 6 marks**

## Question 8 — Practical evaluation

3.1.4 / PAG

a. A student measures the enthalpy of combustion of three alcohols (methanol, ethanol, propan-1-ol) using simple calorimetry. State and explain the trend you would expect in the experimental values as the number of carbon atoms increases. **[2]**

b. **[3]**

Explain why simple "spirit burner" calorimetry experiments generally give values that are significantly less exothermic than data book values, referring to at least two sources of error.

c. Suggest how a bomb calorimeter gives more accurate results than a simple spirit burner experiment. **[3]**

**Question 8 total: 8 marks**

## Question 9 — Extended response

Synoptic

**A student wants to determine the enthalpy change of formation of magnesium oxide, MgO(s), from its elements. Explain why this cannot be measured directly by experiment, and describe how the value could be found indirectly using Hess's law, referring to the enthalpy changes of combustion of magnesium and of the reaction needed.** **[6]**

*This question is marked using a level of response.*

**Question 9 total: 6 marks**

**PAPER TOTAL: 86 MARKS**

# MARK SCHEME — Chapter 10: Enthalpy Changes

*Award marks independently unless stated. ECF applies to calculations.*

### Q1 (a) [2]

- The enthalpy change when the molar quantities of reactants react, as specified in the equation... (1)
- ...under standard conditions, with all reactants and products in their standard states. (1)

### Q1 (b) [2]

- 100 kPa pressure. (1)
- A stated temperature, usually 298 K (25 °C) — concentration 1 mol $dm^{-3}$ for solutions. (1)

### Q1 (c) [2]

- The enthalpy change when 1 mole of a compound is formed from its constituent elements... (1)
- ...with all substances in their standard states under standard conditions. (1)

### Q1 (d) [2]

- The enthalpy change when 1 mole of a substance is completely burned/reacts completely with oxygen... (1)
- ...under standard conditions, with all substances in their standard states. (1)

### Q1 (e) [4]

- Reactants drawn at a higher energy level than products. (1)
- ΔH arrow drawn pointing downward from reactants to products, labelled negative. (1)
- A "hump" (activation energy pathway) drawn between reactants and products. (1)
- Activation energy, $E_a$, correctly labelled from reactants to the top of the hump. (1)

### Q2 (a) [2]

- ΔH is negative. (1)
- Energy is transferred from the system to the surroundings (surroundings get hotter). (1)

### Q2 (b) [2]

- ΔH is positive. (1)
- Energy is transferred from the surroundings to the system (surroundings get colder). (1)

### Q2 (c) [3]

- Breaking bonds in the reactants requires energy (endothermic). (1)
- Making bonds in the products ($CO_2$ and $H_2O$) releases energy (exothermic). (1)
- In combustion, more energy is released making new bonds than is required to break the original bonds, so the overall reaction is exothermic. (1)

### Q3 (a) [3]

- $q = mc\Delta T = 100 \times 4.18 \times (45.8 - 20.5)$ (1)
- $= 100 \times 4.18 \times 25.3$ (1)
- = 10 575.4 J (accept 10.6 kJ). (1)

### Q3 (b) [3]

- $M(C_2H_5OH) = 46.0$ g $mol^{-1}$; $n = 0.68 \div 46.0 = 0.0148$ mol (1)
- $\Delta H_c = -10\,575.4 \div 0.0148$ (1)
- $= -715$ kJ $mol^{-1}$ (accept −700 to −720). (1)

### Q3 (c) [1]

- The experimental value is considerably less exothermic (smaller in magnitude) than the data book value. (1)

### Q3 (d) [4]

- Heat loss to the surroundings (not all heat released reaches the water) — improvement: use a lid/draught excluder or insulate the calorimeter. (2)
- Incomplete combustion of the ethanol (some carbon forms soot rather than fully burning to $CO_2$) — improvement: ensure adequate air supply / use a calorimeter that improves combustion efficiency. (2)
- *Also accept: heat lost through evaporation of ethanol before burning, or heat capacity of the container not accounted for.*

**Q4 (a) [3]**

- $q = mc\Delta T = 100 \times 4.18 \times (26.0 - 19.2)$ (1)
- $= 100 \times 4.18 \times 6.8$ (1)
- $= 2842.4$ J (2.84 kJ). (1)

**Q4 (b) [3]**

- $n(HCl) = 0.0500 \times 1.00 = 0.0500$ mol $= n(H_2O)$ formed (1:1 ratio). (1)
- $\Delta H_{neut} = -2842.4 \div 0.0500$ (1)
- $= -56.8$ kJ mol$^{-1}$. (1)

**Q4 (c) [2]**

- Polystyrene is a good insulator (poor conductor of heat). (1)
- This minimises heat loss to the surroundings, giving a more accurate temperature change/result than glass (which conducts heat away). (1)

**Q4 (d) [2]**

- Use a lid on the cup to reduce heat loss by evaporation/convection. (1)
- Use a insulating jacket / place the cup inside a larger beaker with insulation, or use a plastic (not metal) stirrer to reduce further heat loss. (1)

**Q5 (a) [2]**

- If a reaction can occur by more than one route... (1)
- ...the total enthalpy change is the same regardless of the route taken (provided initial and final conditions are the same). (1)

**Q5 (b) [6]**

- Correct Hess's cycle drawn: C(s) + 2$H_2$(g) at top, $CH_4$(g) at bottom, both routes via $CO_2$(g) + 2$H_2O$(l) (combustion products) at the base. (2)
- $\Delta H = \Delta H_c$(reactants) $- \Delta H_c$(products) $= [(-394) + 2(-286)] - (-890)$ (2)
- $= (-966) - (-890)$ (1)
- $= -76$ kJ mol$^{-1}$. (1)

**Q5 (c) [5]**

- $\Delta H = \Sigma\Delta H_f$(products) $- \Sigma\Delta H_f$(reactants) (1)
- $\Delta H_f$(products) $= (-635) + (-394) = -1029$ kJ mol$^{-1}$ (1)
- $\Delta H_f$(reactants) $= -1207$ kJ mol$^{-1}$ (1)
- $\Delta H = -1029 - (-1207)$ (1)
- $= +178$ kJ mol$^{-1}$. (1)

**Q6 (a) [2]**

- The average energy required to break one mole of a given type of bond... (1)
- ...in the gaseous state, averaged over many different molecules/compounds. (1)

**Q6 (b) [2]**

- Mean bond enthalpies are averaged over many different compounds/environments. (1)
- The actual bond enthalpy for a specific bond in a specific molecule may differ from this average value. (1)

**Q6 (c) [6]**

- Bonds broken: one C=C and four C–H remain but one C–H pair reforms... correctly identify bonds broken = 1 × C=C + 1 × H–H = 612 + 436 = 1048 kJ mol$^{-1}$. (2)
- Bonds formed: 1 × C–C + 2 × C–H (the two new C–H bonds formed on each carbon) = 348 + 2(412) = 1172 kJ mol$^{-1}$. (2)
- $\Delta H$ = bonds broken − bonds formed = 1048 − 1172 (1)
- $= -124$ kJ mol$^{-1}$. (1)

**Q6 (d) [3]**

- Bond breaking is endothermic (energy must be put in to separate atoms). (1)
- Bond making is exothermic (energy is released as new bonds form). (1)
- Explanation: breaking bonds requires overcoming attractive forces between atoms (energy absorbed); forming bonds creates new attractive forces (energy released). (1)

**Q7 (a) [4]**

- Cycle correctly drawn with C(s) + $O_2$(g) at top, $CO_2$(g) at bottom (direct route, ΔH to find), and CO(g) + ½$O_2$(g) as the alternative route via $\Delta H_1$ then $\Delta H_2$. (2)
- $\Delta H = \Delta H_1 + \Delta H_2 = (-111) + (-283)$ (1)
- $= -394$ kJ $mol^{-1}$. (1)

**Q7 (b) [2]**

- Burning carbon in a limited/controlled supply of oxygen always produces some $CO_2$ as well as CO — it is impossible to obtain pure CO as the only product. (1)
- This means $\Delta H_1$ cannot be measured directly and must be found indirectly (via Hess's law). (1)

**Q8 (a) [2]**

- The experimental (and theoretical) enthalpy of combustion becomes more exothermic (larger negative value) as the number of carbon atoms increases. (1)
- More bonds are present per mole, so more energy is released overall when the substance burns completely. (1)

**Q8 (b) [3]**

- Heat loss to the surroundings (radiation, convection) rather than all being transferred to the water. (1)
- Incomplete combustion, producing soot/carbon monoxide instead of fully oxidising to $CO_2$ (releasing less energy than expected). (1)
- Evaporation of the alcohol before/during burning also loses fuel without releasing usable heat. (1)

**Q8 (c) [3]**

- A bomb calorimeter is a sealed, insulated container, so almost all heat released is transferred to the surrounding water with minimal loss. (1)
- Combustion occurs in an excess/atmosphere of pure oxygen under pressure, ensuring complete combustion. (1)
- This gives a much more accurate (closer to data book) value for the enthalpy of combustion. (1)

**Q9 [6] — Level of response**

- **Level 3 (5–6):** Clearly explains why direct formation cannot be measured and correctly describes an indirect Hess's law route using combustion data, with a correctly structured cycle description.
- **Level 2 (3–4):** Explains why direct measurement is not possible and gives a broadly correct but incomplete description of the indirect method.
- **Level 1 (1–2):** Some relevant, isolated points about Hess's law or combustion with limited application to this example.

***Indicative content:*** *Burning magnesium in air/oxygen is extremely difficult to control precisely enough (rapid, very exothermic reaction; magnesium may also react with nitrogen in air) to measure $\Delta H_f$ directly and accurately by simple calorimetry. Instead, $\Delta H_f$(MgO) can be found using an indirect Hess's law cycle: measure the (easily obtainable) standard enthalpy of combustion of magnesium, $\Delta H_c^{\circ}$[Mg(s)] (Mg(s) + ½$O_2$(g) → MgO(s)), which is in fact numerically identical to $\Delta H_f$(MgO) in this case since the combustion product is the oxide itself — so the enthalpy of combustion of magnesium directly gives the enthalpy of formation via Hess's law reasoning (same reactants and products, so same ΔH regardless of the term used).*

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**END OF MARK SCHEME · 86 marks**