AP Chemistry Unit 6 Easier Practice Set

Thermodynamics (Thermochemistry Focus)

Building Confidence with Heat, Enthalpy, and Hess's Law

APChemistryRescue.com

📋 Practice Set Information

- **Total Points:** 70 points
- **Suggested Time:** 60–90 minutes
- **Questions:** 10 (scaffolded from q = mcΔT → calorimetry → Hess's Law → bond energies)
- **Topics:** Heat transfer, calorimetry, enthalpy changes, Hess's Law, standard enthalpies of formation, bond energies
- **Skills:** Heat calculations, sign conventions, energy diagrams, enthalpy manipulations, stoichiometry integration
- **Note:** This set focuses on Unit 6 thermochemistry ONLY. Entropy and Gibbs free energy (Unit 9) are excluded.

🔬 ESSENTIAL THERMOCHEMISTRY FORMULAS & CONSTANTS

HEAT TRANSFER:

$q = mc\Delta T$ (heating/cooling, no phase change)
$q = mL_{fus}$ (melting/freezing at constant T)
$q = mL_{vap}$ (boiling/condensing at constant T)

CALORIMETRY:

Coffee-cup (constant P): $q_{rxn} = -q_{solution} = -mc\Delta T$

Bomb calorimeter (constant V): $q_{rxn} = -C_{cal} \times \Delta T$

ENTHALPY CALCULATIONS:

$\Delta H_{rxn} = \Sigma\ \Delta H°_f(\text{products}) - \Sigma\ \Delta H°_f(\text{reactants})$

$\Delta H_{rxn} = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$

Hess's Law: $\Delta H_{total} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \ldots$

Substance	Specific Heat c (J/g·°C)	L_fus (kJ/mol)	L_vap (kJ/mol)
Water (H_2O)	4.18	6.01	40.7
Ice (H_2O solid)	2.09	—	—
Steam (H_2O gas)	2.01	—	—

🔥 Sign Conventions (CRITICAL!):

- **Exothermic:**

  System releases heat → $q < 0$ or $\Delta H < 0$ (heat flows OUT)
- **Endothermic:**

  System absorbs heat → $q > 0$ or $\Delta H > 0$ (heat flows IN)
- **Calorimetry:**

  $q_{rxn} = -q_{surroundings}$ (conservation of energy)

Question 1: Basic Heat Transfer ($q = mc\Delta T$) (6 points)

A 250.0 g sample of water is heated from 20.0°C to 85.0°C.

$$q = mc\Delta T$$

`c_water = 4.18 J/(g·°C)`

(a) Calculate the **heat absorbed** by the water (in joules). Show all work with units.

Work Space:

(b) Convert your answer from part (a) to **kilojoules**.

Work Space:

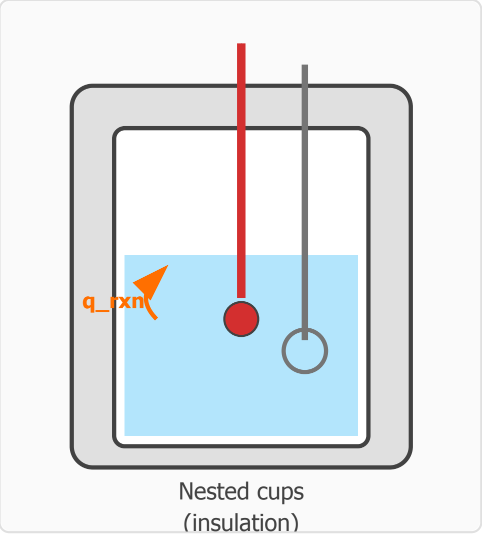
(c) Is this process **endothermic or exothermic** from the water's perspective? What is the **sign of q** for the water?

Work Space:

Question 2: Calorimetry (Coffee-Cup and Bomb) (10 points)

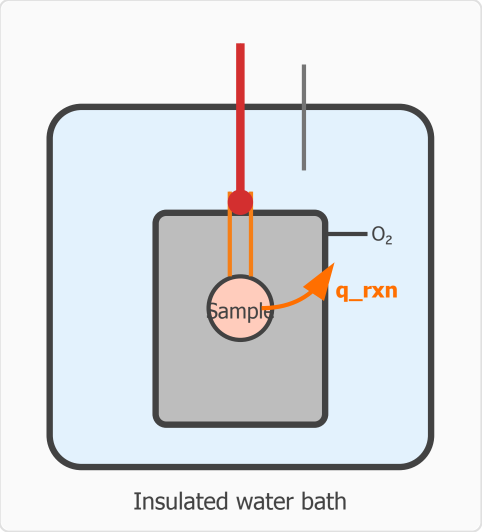
Two types of calorimeters are used to measure heat changes in chemical reactions:

Coffee-Cup Calorimeter (Constant Pressure)



Measures q_p = ΔH
q_rxn = −m_solution c ΔT

Bomb Calorimeter (Constant Volume)



Measures q_v ≈ ΔH
q_rxn = −C_cal × ΔT

Part A: Coffee-Cup Calorimetry

When 5.00 g of solid $CaCl_2$ dissolves in 100.0 g of water in a coffee-cup calorimeter, the temperature rises from 20.0°C to 32.5°C. Assume the specific heat of the solution is 4.18 J/(g·°C).

(a) Calculate the **heat absorbed by the solution** (q_solution). Remember: total mass = mass of water + mass of $CaCl_2$.

Work Space:

(b) What is the **heat released by the reaction** (q_rxn)? Use q_rxn = −q_solution.

Work Space:

(c) Is the dissolution of $CaCl_2$ **exothermic or endothermic**? Explain using the sign of q_rxn and the temperature change.

Work Space:

Part B: Bomb Calorimetry

A 1.50 g sample of glucose ($C_6H_{12}O_6$) is burned in a bomb calorimeter with a heat capacity C_cal = 4.90 kJ/°C. The temperature increases from 22.0°C to 28.4°C.

(d) Calculate the **heat released by the combustion** (q_rxn) using q_rxn = −C_cal × ΔT.

Work Space:

(e) Calculate the **molar enthalpy of combustion** (ΔH_comb in kJ/mol) for glucose. (Molar mass of $C_6H_{12}O_6$ = 180.16 g/mol.)

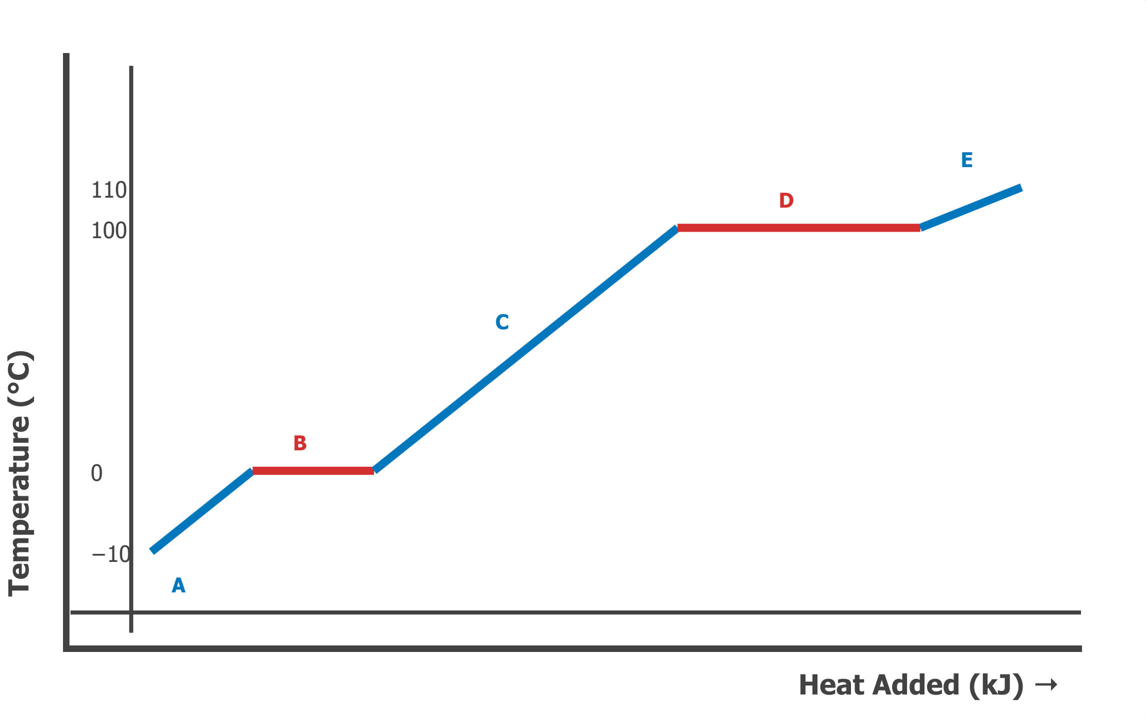
Steps:

1. Find moles of glucose: n = mass/MM
2. ΔH_comb = q_rxn / n

Work Space:

Question 3: Heating Curve and Phase Changes (8 points)

A 50.0 g sample of ice at −10°C is heated until it becomes steam at 110°C. The heating curve shows five distinct segments:



Segment Guide:

A:

Ice warming (−10°C → 0°C), use q = mc_ice ΔT

B:

Melting (0°C, constant T), use q = n × L_fus

C:

Water warming (0°C → 100°C), use q = mc_water ΔT

D:

Boiling (100°C, constant T), use q = n × L_vap

E:

Steam warming (100°C → 110°C), use q = mc_steam ΔT

(a) Calculate the heat required for **Segment A** (warming ice from −10°C to 0°C). Use c_ice = 2.09 J/(g·°C).

Work Space:

(b) Calculate the heat required for **Segment B** (melting ice at 0°C). Use L_fus = 6.01 kJ/mol and molar mass H_2O = 18.02 g/mol.

Work Space:

(c) Why are segments B and D **horizontal** (temperature constant) even though heat is being added?

Work Space:

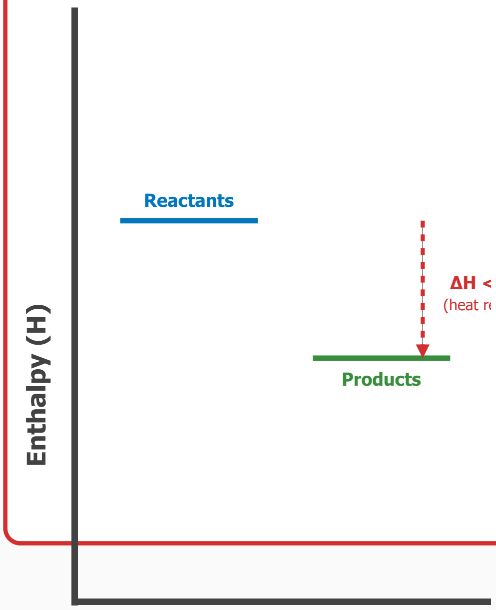
(d) Which segment (C or D) requires **more heat**? Explain using the relative magnitudes of c_water and L_vap.

Work Space:

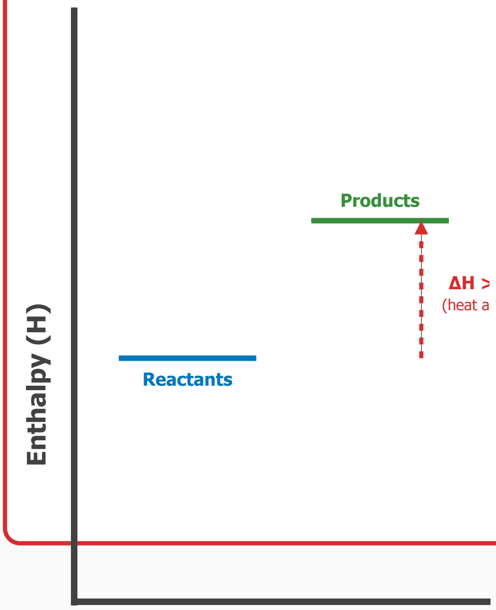
Question 4: Enthalpy Diagrams and Sign Conventions (7 points)

Enthalpy diagrams show the energy change (ΔH) for a reaction. Two examples are shown below:

EXOTHERMIC REACTION (ΔH < 0)



ENDOTHERMIC REACTION (ΔH > 0)



Reaction Progress → Reaction Progress →

(a) For the **exothermic reaction**, which has higher enthalpy: reactants or products? Explain why ΔH is negative.

Work Space:

(b) For the **endothermic reaction**, which has higher enthalpy: reactants or products? Explain why ΔH is positive.

Work Space:

(c) The combustion of methane is represented by: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, $\Delta H = -890$ kJ. Draw a simple enthalpy diagram for this reaction. Label reactants, products, and ΔH.

Enthalpy Diagram for CH_4 Combustion

Question 5: Enthalpy and Stoichiometry (7 points)

The thermochemical equation for the combustion of propane is:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) \quad \Delta H = -2220 \text{ kJ}$$

This means that when **1 mole of C_3H_8** burns completely, **2220 kJ of heat** is released.

(a) How much heat is released when **2.50 moles** of C_3H_8 burn?

Work Space:

(b) How much heat is released when **110.0 g** of C_3H_8 burn? (Molar mass of C_3H_8 = 44.10 g/mol.)

Steps:

1. Convert mass to moles: n = mass/MM
2. Use ΔH per mole: $q = n \times \Delta H$

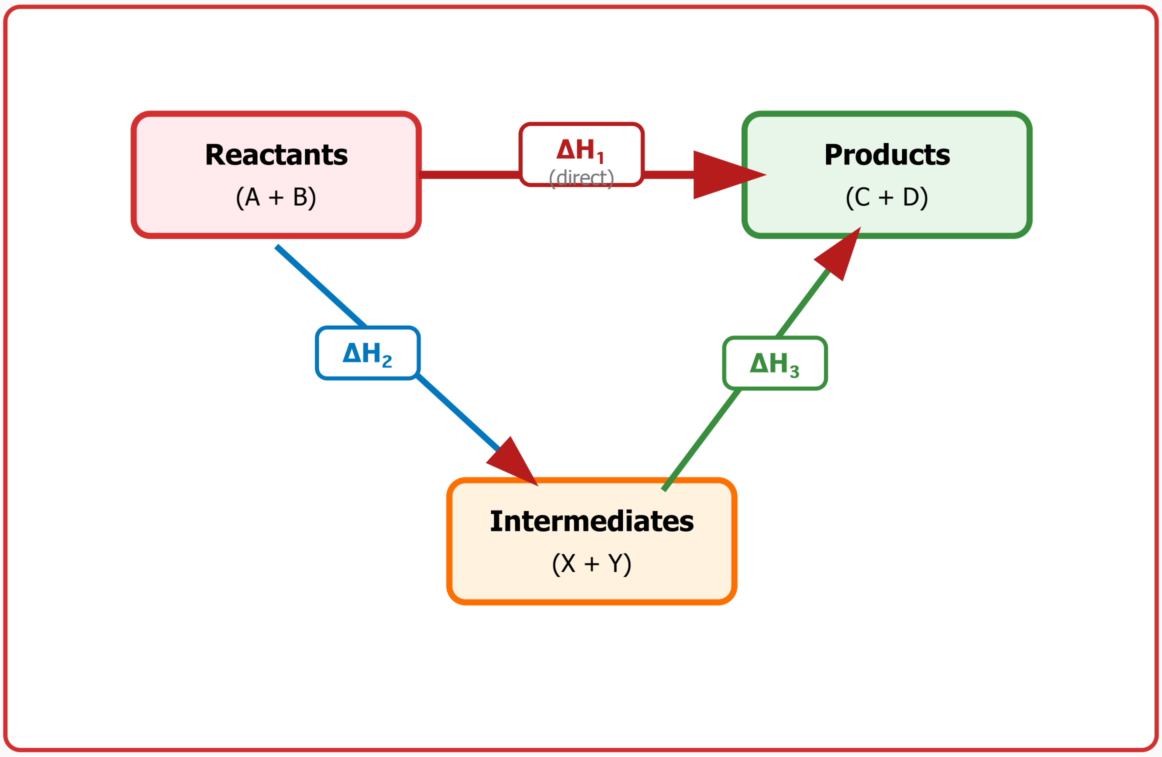
Work Space:

(c) If you wanted to release **1000 kJ of heat**, how many grams of C_3H_8 would you need to burn?

Work Space:

Question 6: Hess's Law and Energy Cycles (9 points)

Hess's Law states that the total enthalpy change for a reaction is the same whether it occurs in one step or multiple steps. The diagram below shows an energy cycle:



Hess's Law:

$\Delta H_1 = \Delta H_2 + \Delta H_3$

(Direct path = Sum of indirect path steps)

Example Problem:

Calculate ΔH for the reaction: $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$

Given the following thermochemical equations:

(1) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_1 = -394$ kJ

(2) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$ $\Delta H_2 = -283$ kJ

(a) Write the **target equation** we want to find ΔH for. This is the formation of CO(g) from elements.

Work Space:

(b) To use Hess's Law, we need to manipulate the given equations. Should equation (2) be used as written, or should it be **reversed**? Explain.

Hint:

CO(g) appears on the left (reactant side) of the target equation, but on which side of equation (2)?

Work Space:

(c) If you reverse equation (2), what happens to the sign of ΔH_2? Write the reversed equation and its ΔH.

Work Space:

(d) Add equation (1) and the reversed equation (2). Show that CO_2 cancels out and you obtain the target equation.

Work Space:

(e) Calculate ΔH for the target equation by adding ΔH_1 and the ΔH of the reversed equation (2).

Work Space:

Question 7: Standard Enthalpies of Formation (8 points)

The **standard enthalpy of formation (ΔH°_f)** is the enthalpy change when 1 mole of a compound is formed from its elements in their standard states (25°C, 1 atm).

Key Rule:

ΔH°_f for any element in its standard state = 0

Examples: ΔH°_f(O_2(g)) = 0, ΔH°_f(C(graphite)) = 0, ΔH°_f(N_2(g)) = 0

Substance	ΔH°_f (kJ/mol)
H_2O(l)	−286

$CO_2(g)$	−394
$NH_3(g)$	−46
$CH_4(g)$	−75
$C_2H_5OH(l)$	−278
$NO(g)$	+90

$$\Delta H^\circ\_rxn = \Sigma\ \Delta H^\circ\_f(\text{products}) - \Sigma\ \Delta H^\circ\_f(\text{reactants})$$

Calculate ΔH°_rxn for the combustion of ethanol:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

(a) Calculate **Σ ΔH°_f(products)**. Remember to multiply each ΔH°_f by its stoichiometric coefficient.

Work Space:

(b) Calculate **Σ ΔH°_f(reactants)**. Remember: ΔH°_f($O_2(g)$) = 0.

Work Space:

(c) Calculate **$\Delta H°\_rxn$** using the formula. Is the combustion exothermic or endothermic?

Work Space:

Question 8: Bond Energies (7 points)

Bond energy is the energy required to **break** one mole of bonds in the gas phase. To estimate ΔH_rxn:

$$\Delta H\_rxn = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$$

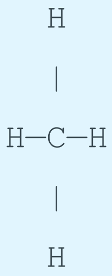
Breaking bonds requires energy (+), forming bonds releases energy (−)

Bond	Bond Energy (kJ/mol)
C–H	413
O=O	498
C=O	799
O–H	463

Estimate ΔH for the combustion of methane:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

REACTANTS



$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

4 C–H bonds

O=O   O=O

2 O=O bonds

PRODUCTS

O=C=O

2 C=O bonds

H—O—H H—O—H

4 O–H bonds

(a) Calculate the total energy required to **break all bonds in the reactants** (4 C–H + 2 O=O).

Work Space:

(b) Calculate the total energy released when **forming all bonds in the products** (2 C=O + 4 O–H).

Work Space:

(c) Calculate **ΔH_rxn** using ΔH = (bonds broken) − (bonds formed). Is the reaction exothermic or endothermic?

Work Space:

Question 9: Conceptual Understanding (4 points)

(a) Explain the difference between **heat (q)** and **enthalpy (H)**. When is $q = \Delta H$?

Work Space:

(b) A student claims: "In an exothermic reaction, the surroundings get colder because the system releases heat." Is this correct? Explain what actually happens to the temperature of the surroundings.

Work Space:

Question 10: Integrated Problem (4 points)

A 100.0 g sample of aluminum at 150°C is dropped into 200.0 g of water at 20.0°C in an insulated container. The final temperature of the mixture is

35.2°C.

Given: c_Al = 0.900 J/(g·°C), c_water = 4.18 J/(g·°C)

(a) Calculate the **heat lost by aluminum** (q_Al). Remember: ΔT = T_final − T_initial, so q_Al will be negative.

Work Space:

(b) Calculate the **heat gained by water** (q_water).

Work Space:

(c) According to the law of conservation of energy, q_Al + q_water should equal zero (in an ideal insulated system). Calculate the sum and comment on any difference.

Work Space:

END OF PRACTICE SET

Total: 70 points | Answer key begins on next page

📘 COMPLETE ANSWER KEY & SCORING GUIDE

Question 1: Basic Heat Transfer (6 points)

(a) Heat absorbed (3 points):

$$q = mc\Delta T$$

$$q = (250.0 \text{ g})(4.18 \text{ J/g·°C})(85.0 - 20.0)°C$$

$$q = (250.0)(4.18)(65.0) = 67,925 \text{ J}$$

(1 pt for formula; 1 pt for substitution; 1 pt for answer with units)

(b) Convert to kJ (1 point):

$$q = 67,925 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 67.9 \text{ kJ}$$

(1 pt)

(c) Endothermic or exothermic (2 points):

The process is **endothermic** from the water's perspective. (1 pt) The water **absorbs heat** (its temperature increases), so **q > 0** (positive). (1 pt)

Scoring: (a) 3 pts; (b) 1 pt; (c) 2 pts

Question 2: Calorimetry (10 points)

Part A: Coffee-Cup Calorimetry

(a) Heat absorbed by solution (2 points):

Total mass = 100.0 g water + 5.00 g $CaCl_2$ = 105.0 g

$$q_{solution} = mc\Delta T = (105.0 \text{ g})(4.18 \text{ J/g·°C})(32.5 - 20.0)°C$$

$$= (105.0)(4.18)(12.5) = 5488 \text{ J} = 5.49 \text{ kJ}$$

(1 pt for including $CaCl_2$ mass; 1 pt for calculation)

(b) Heat released by reaction (2 points):

$$q_{rxn} = -q_{solution} = -5.49 \text{ kJ}$$

(1 pt for negative sign; 1 pt for answer)

(c) Exothermic or endothermic (2 points):

The dissolution is **exothermic**. (1 pt) q_rxn < 0 (negative) means the reaction **releases heat** to the solution, causing the temperature to rise. Heat flows from the dissolving $CaCl_2$ to the water. (1 pt)

Part B: Bomb Calorimetry

(d) Heat released by combustion (2 points):

$$q_{rxn} = -C_{cal} \times \Delta T = -(4.90 \text{ kJ/°C})(28.4 - 22.0) °C$$

$$= -(4.90)(6.4) = -31.4 \text{ kJ}$$

(1 pt for formula and substitution; 1 pt for answer with negative sign)

(e) Molar enthalpy of combustion (2 points):

Moles of glucose:

$$n = \frac{1.50 \text{ g}}{180.16 \text{ g/mol}} = 0.00833 \text{ mol}$$

$$\Delta H_{comb} = \frac{q_{rxn}}{n} = \frac{-31.4 \text{ kJ}}{0.00833 \text{ mol}} = -3770 \text{ kJ/mol}$$

(1 pt for mole calculation; 1 pt for ΔH calculation)

Scoring: Part A: (a) 2 pts, (b) 2 pts, (c) 2 pts; Part B: (d) 2 pts, (e) 2 pts

Question 3: Heating Curve (8 points)

(a) Segment A: Ice warming (2 points):

$$q_A = mc_{ice}\Delta T = (50.0 \text{ g})(2.09 \text{ J/g·°C})(0 - (-10)) °C$$

$$= (50.0)(2.09)(10) = 1045 \text{ J} = 1.05 \text{ kJ}$$

(1 pt for calculation; 1 pt for units)

(b) Segment B: Melting (3 points):

First, convert mass to moles:

$$n = \frac{50.0 \text{ g}}{18.02 \text{ g/mol}} = 2.775 \text{ mol}$$

$$q_B = n \times L_{fus} = (2.775 \text{ mol})(6.01 \text{ kJ/mol}) = 16.68 \text{ kJ}$$

(1 pt for mole conversion; 1 pt for calculation; 1 pt for units)

(c) Why B and D are horizontal (2 points):

During phase changes (melting, boiling), the added energy is used to **overcome intermolecular forces** and change the state of matter, NOT to increase kinetic energy (temperature). (1 pt) The temperature remains constant until the entire sample has completed the phase transition (all ice melted, or all water boiled). (1 pt)

(d) Which requires more heat: C or D? (1 point):

Segment D (boiling) requires more heat. (0.5 pt) L_vap (40.7 kJ/mol) is much larger than the energy needed to warm water 100°C using c_water (4.18 J/g·°C), because breaking H-bonds to vaporize requires far more energy than simply increasing kinetic energy. (0.5 pt)

Scoring: (a) 2 pts; (b) 3 pts; (c) 2 pts; (d) 1 pt

Question 4: Enthalpy Diagrams (7 points)

(a) Exothermic reaction (2 points):

Reactants have higher enthalpy than products. (1 pt) ΔH = H_products − H_reactants. Since products are lower energy, ΔH is negative ($\Delta H < 0$), indicating energy is released. (1 pt)

(b) Endothermic reaction (2 points):

Products have higher enthalpy than reactants. (1 pt) ΔH = H_products − H_reactants. Since products are higher energy, ΔH is positive ($\Delta H > 0$), indicating energy is absorbed from surroundings. (1 pt)

(c) CH_4 combustion diagram (3 points):

Expected diagram features:

- Horizontal line at top labeled "Reactants: $CH_4(g) + 2O_2(g)$" (1 pt)
- Horizontal line at bottom labeled "Products: $CO_2(g) + 2H_2O(l)$" (1 pt)
- Downward arrow between them labeled "$\Delta H = -890$ kJ" (1 pt)

(Products must be LOWER than reactants for exothermic reaction)

Scoring: (a) 2 pts; (b) 2 pts; (c) 3 pts

Question 5: Enthalpy and Stoichiometry (7 points)

(a) Heat from 2.50 mol (2 points):

$$q = n \times \Delta H = (2.50 \text{ mol})(-2220 \text{ kJ/mol}) = -5550 \text{ kJ}$$

(1 pt for calculation; 1 pt for units and sign)

(b) Heat from 110.0 g (3 points):

$$n = \frac{110.0 \text{ g}}{44.10 \text{ g/mol}} = 2.494 \text{ mol}$$

$$q = (2.494 \text{ mol})(-2220 \text{ kJ/mol}) = -5537 \text{ kJ} \approx -5540 \text{ kJ}$$

(1 pt for mole conversion; 1 pt for calculation; 1 pt for answer)

(c) Mass needed for 1000 kJ (2 points):

$$n = \frac{q}{\Delta H} = \frac{-1000 \text{ kJ}}{-2220 \text{ kJ/mol}} = 0.450 \text{ mol}$$

$$\text{mass} = n \times \text{MM} = (0.450 \text{ mol})(44.10 \text{ g/mol}) = 19.8 \text{ g}$$

(1 pt for mole calculation; 1 pt for mass)

Scoring: (a) 2 pts; (b) 3 pts; (c) 2 pts

Question 6: Hess's Law (9 points)

(a) Target equation (1 point):

$$\text{C(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO(g)}$$

(1 pt)

(b) Should equation (2) be reversed? (2 points):

Yes, equation (2) should be reversed. (1 pt) In the target equation, CO(g) is a **product** (right side), but in equation (2) as written, CO(g) is a **reactant** (left side). To match the target, we must reverse (2) so CO appears on the product side. (1 pt)

(c) Reversed equation (2) and ΔH (2 points):

When you reverse a reaction, the sign of ΔH flips. (1 pt)

Reversed: $CO_2(g) \rightarrow CO(g) + \frac{1}{2}O_2(g)$ $\Delta H = +283$ kJ (1 pt)

(d) Add equations (2 points):

Equation (1): $C(s) + O_2(g) \rightarrow CO_2(g)$

Reversed (2): $CO_2(g) \rightarrow CO(g) + \frac{1}{2}O_2(g)$

Sum: $C(s) + O_2(g) + CO_2(g) \rightarrow CO_2(g) + CO(g) + \frac{1}{2}O_2(g)$

Cancel CO_2 (appears on both sides) and simplify O_2 terms:

$$\text{C(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO(g)}$$

(1 pt for adding; 1 pt for canceling)

(e) Calculate ΔH (2 points):

$$\Delta H_{target} = \Delta H_1 + \Delta H_{reversed(2)} = -394 \text{ kJ} + 283 \text{ kJ} = -111 \text{ kJ}$$

(1 pt for addition; 1 pt for answer)

Scoring: (a) 1 pt; (b) 2 pts; (c) 2 pts; (d) 2 pts; (e) 2 pts

Question 7: Standard Enthalpies of Formation (8 points)

(a) Σ ΔH°_f(products) (3 points):

Products: 2 CO_2(g) + 3 H_2O(l)

$$\Sigma \Delta H°_f(\text{products}) = 2(-394) + 3(-286)$$

$$= -788 + (-858) = -1646 \text{ kJ}$$

(1 pt for coefficients; 1 pt for calculation; 1 pt for answer)

(b) Σ ΔH°_f(reactants) (2 points):

Reactants: 1 C_2H_5OH(l) + 3 O_2(g)

$$\Sigma \Delta H°_f(\text{reactants}) = 1(-278) + 3(0)$$

$$= -278 \text{ kJ}$$

(Note: ΔH°_f(O_2(g)) = 0 because it's an element in standard state)

(1 pt for recognizing O_2 = 0; 1 pt for calculation)

(c) ΔH°_rxn and classification (3 points):

$$\Delta H°_{rxn} = \Sigma \Delta H°_f(\text{products}) - \Sigma \Delta H°_f(\text{reactants})$$

$$= -1646 - (-278) = -1646 + 278 = -1368 \text{ kJ}$$

The combustion is **exothermic** ($\Delta H < 0$).

(1 pt for formula; 1 pt for calculation; 1 pt for classification)

Scoring: (a) 3 pts; (b) 2 pts; (c) 3 pts

Question 8: Bond Energies (7 points)

(a) Bonds broken in reactants (2 points):

CH_4: 4 C–H bonds = 4 × 413 = 1652 kJ

2 O_2: 2 O=O bonds = 2 × 498 = 996 kJ

$$\text{Total broken} = 1652 + 996 = 2648 \text{ kJ}$$

(1 pt for calculation; 1 pt for sum)

(b) Bonds formed in products (2 points):

CO_2: 2 C=O bonds = 2 × 799 = 1598 kJ

2 H_2O: 4 O–H bonds = 4 × 463 = 1852 kJ

$$\text{Total formed} = 1598 + 1852 = 3450 \text{ kJ}$$

(1 pt for calculation; 1 pt for sum)

(c) ΔH_rxn and classification (3 points):

$$\Delta H_{rxn} = (\text{bonds broken}) - (\text{bonds formed})$$

$$= 2648 - 3450 = -802 \text{ kJ}$$

The reaction is **exothermic** ($\Delta H < 0$). More energy is released when forming product bonds than is required to break reactant bonds.

(1 pt for formula; 1 pt for calculation; 1 pt for classification)

Scoring: (a) 2 pts; (b) 2 pts; (c) 3 pts

Question 9: Conceptual Understanding (4 points)

(a) Heat vs. enthalpy (2 points):

Heat (q) is the energy transferred between a system and surroundings due to temperature difference. **Enthalpy (H)** is a state function representing the total heat content of a system at constant pressure. (1 pt) $q = \Delta H$ when the process occurs at **constant pressure** with only P-V work (typical for reactions in open containers). (1 pt)

(b) Temperature of surroundings in exothermic reaction (2 points):

Incorrect. In an exothermic reaction, the system releases heat, so the **surroundings get WARMER**, not colder. (1 pt) Heat flows from the system (reaction) to the surroundings, increasing the temperature of the surroundings. For example, in calorimetry, when a reaction is exothermic ($q_{rxn} < 0$), the solution/calorimeter temperature rises ($q_{surroundings} > 0$). (1 pt)

Scoring: (a) 2 pts; (b) 2 pts

Question 10: Integrated Problem (4 points)

(a) Heat lost by aluminum (1.5 points):

$$q_{Al} = mc\Delta T = (100.0 \text{ g})(0.900 \text{ J/g·°C})(35.2 - 150)°C$$

$$= (100.0)(0.900)(-114.8) = -10,332 \text{ J} = -10.3 \text{ kJ}$$

(0.5 pt for ΔT calculation; 1 pt for q with correct negative sign)

(b) Heat gained by water (1.5 points):

$$q_{water} = mc\Delta T = (200.0 \text{ g})(4.18 \text{ J/g·°C})(35.2 - 20.0)°C$$

$$= (200.0)(4.18)(15.2) = 12,707 \text{ J} = 12.7 \text{ kJ}$$

(0.5 pt for ΔT; 1 pt for q with correct positive sign)

(c) Conservation of energy check (1 point):

$$q_{Al} + q_{water} = -10.3 + 12.7 = +2.4 \text{ kJ}$$

Ideally, this sum should be zero. The difference (~2.4 kJ) represents **heat lost to the surroundings** (container walls, air) despite insulation. In a perfectly insulated system, all heat lost by Al would be gained by water. (1 pt for calculation and explanation)

Scoring: (a) 1.5 pts; (b) 1.5 pts; (c) 1 pt