Name: _____

CHY 113 | Guided Practice: Thermochemistry

Thermochemical Equations • Calorimetry • Hess's Law • Standard Enthalpies of Formation

Friday group activity | Discuss the reasoning before calculating. Show setups, units, and signs. Follow the prompts in Sections 1–4; Section 5 is independent practice.

Section 1 — Thermochemical Equations

Model: $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$ $\quad \Delta H = -571.6$ kJ

1. For the reaction as written, is energy absorbed or released? How many moles of H_2 correspond to $\Delta H = -571.6$ kJ?

2. If 1.00 mol H_2 reacts, what is ΔH? Explain how you adjusted the thermochemical equation and its enthalpy.

3. Calculate the heat transferred when 9.00 g H_2 (*2.016 g/mol*) reacts completely with excess O_2. First convert grams → moles; then use the mole-to-energy relationship. Report the sign and the magnitude released.

4. If the reaction is reversed to decompose liquid water, what happens to the sign of ΔH? What is ΔH when 1.00 mol $H_2O(l)$ decomposes?

5. $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ $\quad \Delta H = -92.2$ kJ. (a) How many moles of NH_3 correspond to −92.2 kJ? (b) Find ΔH when 0.750 mol NH_3 forms. (c) Is the reaction exothermic or endothermic?

6. Using the ammonia equation above, 6.00 g H_2 reacts with excess N_2. (a) Convert H_2 to moles (M = 2.016 g/mol). (b) Use the balanced equation to determine q. (c) What changes if the equation is reversed?

Section 2 — Calorimetry

Model: $q = mc\Delta T$, $\Delta T = T_{final} - T_{initial}$. Assume $c_{water} = c_{solution} = 4.184\ J/(g\cdot°C)$

7. A reaction warms 125.0 g of aqueous solution from 21.0°C to 27.5°C. Identify the system and surroundings. Predict the signs of qsolution and qreaction before calculating.

8. For Question 7, calculate ΔT and $q_{solution}$ in J. Then calculate $q_{reaction}$ in kJ. Assume the calorimeter itself absorbs negligible heat.

9. When 6.00 g of a solid dissolves in 100.0 g of water, the temperature increases from 22.0°C to 28.0°C. Assuming all 106.0 g of solution has c = 4.184 J/(g·°C), what mass belongs in q = mcΔT? Calculate $q_{solution}$ and $q_{dissolution}$.

10. A 60.0 g metal sample initially at 95.0°C is placed in 150.0 g water at 20.0°C. The final temperature is 24.0°C. Which substance gains heat? Which loses heat? Write the conservation-of-energy equation before substituting.

11. Use Question 10 to calculate the metal's specific heat capacity. Assume no heat is lost to the surroundings.

12. 4.00 g KNO_3 dissolves in 80.0 g water. The temperature drops from 24.0°C to 21.5°C. (a) What is the mass of solution? (b) Predict the signs of $q_{solution}$ and $q_{dissolution}$. (c) Calculate both values.

13. A 40.0 g hot metal sample at 88.0°C is added to 100.0 g water at 19.0°C. The final temperature is 22.0°C. (a) Write the equation for $q_{metal} = -q_{water}$. (b) Determine ΔT for each substance. (c) Calculate the specific heat of the metal; assume no heat loss.

Section 3 — Hess's Law

Model: ΔH depends on initial and final states, not the path. Reverse an equation → change the sign of ΔH. Multiply an equation → multiply ΔH by the same factor. Add equations, cancel identical species on opposite sides, then add ΔH values.

14. Two reaction pathways have the same starting substances and final products. Must they have the same ΔH? Explain using the idea of a state function.

Target: $C(s, graphite) + \frac{1}{2} O_2(g) \rightarrow CO(g)$

I. $C(s, graphite) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ

II. $CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g)$ $\Delta H = -283.0$ kJ

15. Compare the target with equations I and II. Which equation must be reversed? Write the adjusted equation and its new ΔH.

16. Add the adjusted equations, cross out species on both sides, and verify that the result matches the target. Only then calculate ΔH for the target.

Target: $2 NO(g) + O_2(g) \rightarrow 2 NO_2(g)$

I. $N_2(g) + O_2(g) \rightarrow 2 NO(g)$ $\Delta H = +180.6$ kJ

II. $N_2(g) + 2 O_2(g) \rightarrow 2 NO_2(g)$ $\Delta H = +66.4$ kJ

17. Identify what needs to happen to equation I to place NO on the reactant side. Write the adjusted equation and ΔH.

18. Combine the equations and cancel common species. Determine ΔH for the target and classify the reaction.

19. If the target were $NO(g) + \frac{1}{2} O_2(g) \rightarrow NO_2(g)$, how would your ΔH change? Explain the coefficient–enthalpy relationship.

20. Determine ΔH for $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$ from these equations: (I) $H_2(g) + Br_2(l) \rightarrow 2\ HBr(g)$, ΔH = −72.5 kJ; (II) $HBr(g) + \frac{1}{2} Cl_2(g) \rightarrow HCl(g) + \frac{1}{2} Br_2(l)$, ΔH = −36.3 kJ. (a) Which equation must be multiplied? (b) Show cancellations. (c) Calculate ΔH.

21. Target: $2\ C(s, \text{graphite}) + H_2(g) \rightarrow C_2H_2(g)$. Given (I) $2\ C(s) + 2\ O_2(g) \rightarrow 2\ CO_2(g)$, ΔH = −787.0 kJ; (II) $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$, ΔH = −285.8 kJ; (III) $2\ C_2H_2(g) + 5\ O_2(g) \rightarrow 4\ CO_2(g) + 2\ H_2O(l)$, ΔH = −2599.2 kJ. (a) Determine which equation to reverse and scale. (b) Show the final cancellations. (c) Find ΔH.

Section 4 — Standard Enthalpies of Formation

Model: $\Delta H°f$ is the enthalpy change for formation of exactly 1 mol of a compound from its elements in their standard states. Elements in their standard states have $\Delta H°f = 0$.

22. Write balanced standard formation equations, including states, for (a) $MgO(s)$, (b) $NH_3(g)$, and (c) $H_2O(l)$. Why might a coefficient of ½ be necessary?

23. Explain why $\Delta H°f[O_2(g)] = 0$ but $\Delta H°f[H_2O(l)] \neq 0$. Would $\Delta H°f[O(g)]$ equal zero? Explain.

Model: $\Delta H°rxn = \Sigma[n\Delta H°f(products)] - \Sigma[n\Delta H°f(reactants)]$. Include coefficients and states. All values below are kJ/mol.

$2\ H_2S(g) + 3\ O_2(g) \rightarrow 2\ H_2O(l) + 2\ SO_2(g)$

$\Delta H°f$: $H_2S(g) = -20.1$; $O_2(g) = 0$; $H_2O(l) = -285.8$; $SO_2(g) = -296.1$

24. Write the products sum, showing every coefficient and formation enthalpy. Then write the reactants sum, including O_2.

25. Subtract the reactants sum from the products sum. Calculate $\Delta H°rxn$ for the equation as written, including units and sign.

26. If only 0.500 mol H_2S reacts with excess O_2, how much heat is released? Explain how the balanced coefficients connect to the enthalpy value.

27. Write standard formation equations for (a) $CaCl_2(s)$ and (b) $C_2H_6(g)$, using elemental standard states and forming exactly 1 mol of product. Explain the fractional coefficient in the ethane equation.

28. New example: For 2 $NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$, use $\Delta H°f[NO(g)] = +90.3$ kJ/mol and $\Delta H°f[NO_2(g)] = +33.2$ kJ/mol. (a) Set up products minus reactants, including coefficients. (b) Calculate $\Delta H°rxn$. (c) Determine q when 0.250 mol NO reacts completely.

Section 5 — General Practice (Choose Your Own Method)

29. For $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$, $\Delta H = -1203.6$ kJ: calculate the heat released when 12.0 g Mg reacts with excess oxygen. M(Mg) = 24.31 g/mol.

30. A 50.0 g sample of water cools from 65.0°C to 30.0°C. Determine qwater. Is heat absorbed or released by the water? cwater = 4.184 J/(g·°C).

31. A 75.0 g metal sample at 95.0°C is placed in 150.0 g water at 20.0°C. Final temperature is 24.0°C. Determine the metal's specific heat capacity, assuming no heat loss. cwater = 4.184 J/(g·°C).

For Questions 32–33:

A. $2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$ $\Delta H = -197.8$ kJ

B. $S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H = -296.8$ kJ

32. Use Hess's law to calculate ΔH for $2\ S(s) + 3\ O_2(g) \rightarrow 2\ SO_3(g)$. Show all equation manipulations.

33. Calculate ΔH for $SO_3(g) \rightarrow SO_2(g) + \frac{1}{2}\ O_2(g)$. State which equation you manipulated and how.

34. Calculate $\Delta H°_{rxn}$ for $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$. $\Delta H°_f$ (kJ/mol): $CH_4(g)$ −74.8, $CO_2(g)$ −393.5, $H_2O(l)$ −285.8, $O_2(g)$ 0.

Section 5 — General Practice: Limiting Reactants and Energy

35. $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$, ΔH = −571.6 kJ. If 5.00 g H_2 and 20.0 g O_2 react, identify the limiting reactant and calculate the heat released. $M(H_2)$ = 2.016 g/mol; $M(O_2)$ = 32.00 g/mol.

36. $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$, ΔH = −1203.6 kJ. If 10.0 g Mg and 5.00 g O_2 react, identify the limiting reactant and calculate q for the reaction. M(Mg) = 24.31 g/mol; $M(O_2)$ = 32.00 g/mol.

37. $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$, ΔH = −890.3 kJ. A mixture contains 8.00 g CH_4 and 24.0 g O_2. Calculate the heat released and the mass of excess reactant remaining. $M(CH_4)$ = 16.04 g/mol; $M(O_2)$ = 32.00 g/mol.

38. Explain why using the initial mass of the excess reactant to calculate heat can give an incorrect result. What quantity actually determines the reaction extent?

Final check: Did you identify the system? Match ΔH to the balanced equation? Use the correct mass? Track signs? Cancel Hess's law intermediates? Include formation coefficients? Check for a limiting reactant?