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AP Chemistry Unit 6 Study Guide

Thermochemistry (Topics 6.1 to 6.9)

Show all work with units and correct significant figures. In explanations, say what happens to the particles and to the energy, not just whether a value goes up or down.

Constants	c_{water} = 4.184 J/(g·°C)	c_{ice} = 2.09 J/(g·°C)	ΔH_{fus} (H_2O) = 6.01 kJ/mol	ΔH_{vap} (H_2O) = 40.7 kJ/mol
Equations	$q = mc\Delta T$	$q_{system} = -q_{surroundings}$	$q = n\Delta H$	1 kJ = 1000 J
	$\Delta H°_{rxn} = \Sigma\Delta H°_f$(products) − $\Sigma\Delta H°_f$(reactants)	$\Delta H \approx \Sigma$(bonds broken) − Σ(bonds formed)	$\Delta T = T_{final} - T_{initial}$	K = °C + 273

Part 1. Endothermic and Exothermic Processes (Topic 6.1)

1. Complete the table. Treat the substance or reaction named as the system.

Process	**Endothermic or exothermic?**	**Sign of q for the system**	**Temperature of the surroundings (rises or falls)**
Ice melting on a kitchen counter			
Water vapor condensing on a cold glass			
Propane burning in a gas grill			
NH_4NO_3 dissolving in a cold pack (it gets cold)			
Iron rusting in a hand warmer (it gets warm)			

2. When solid calcium chloride, $CaCl_2$, dissolves in water, the temperature of the water rises. Is the dissolving endothermic or exothermic? Explain in terms of the attractions that are broken and formed.

3. Decide whether each statement is true or false. If it is false, rewrite it so it is true.

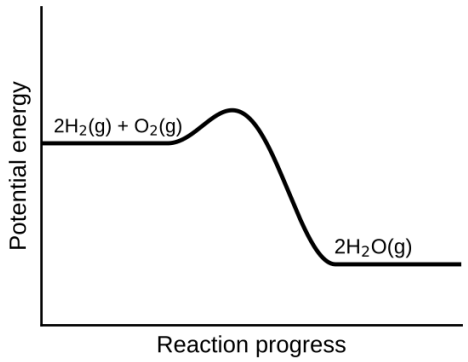
(a) Breaking a chemical bond releases energy.

(b) In an exothermic reaction, the energy lost by the system is gained by the surroundings.

(c) A cold pack feels cold because cold flows out of the pack and into your hand.

Part 2. Energy Diagrams (Topic 6.2)

4. The energy diagram represents the reaction $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$.

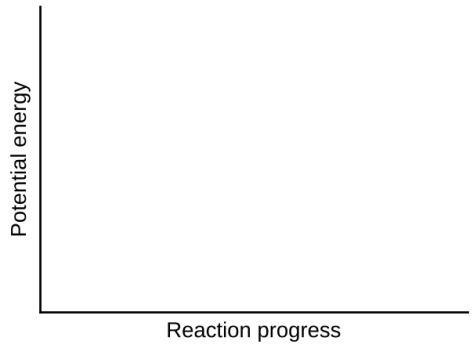


(a) Is the reaction endothermic or exothermic? Use the diagram to justify your answer.

(b) Draw an arrow on the diagram that represents ΔH for the reaction.

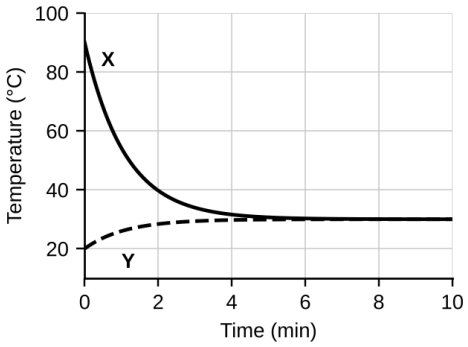
(c) Which have the stronger bonds overall, the reactants or the products? Explain.

5. The reaction $N_2(g) + O_2(g) \rightarrow 2NO(g)$ has $\Delta H = +181$ kJ/mol. On the axes, sketch an energy diagram for this reaction. Label the reactants, the products, and ΔH.



Part 3. Heat Transfer and Thermal Equilibrium (Topic 6.3)

6. A hot piece of metal is placed in a cup of cool water inside an insulated container. The graph shows the temperature of each over time.



(a) Which curve, X or Y, represents the metal? How do you know?

(b) Explain at the particle level how thermal energy moves from the metal to the water.

(c) Once the curves meet, which quantity is the same for the metal and the water? Circle one: **total thermal energy** / **temperature** / **heat capacity**. What is true about the particles at that point?

(d) The metal's temperature changes by 60 °C but the water's changes by only 10 °C, even though the energy lost equals the energy gained. Explain.

Part 4. Heat Capacity and Calorimetry (Topic 6.4)

7. How much heat, in kJ, is needed to warm 250.0 g of water from 22.0 °C to 85.0 °C?

8. The molar heat capacity of liquid water is 75.3 J/(mol·°C). Use $q = nC_m\Delta T$ to find how much heat it takes to raise the temperature of 2.00 mol of water by 15.0 °C.

9. Equal 100. g samples of aluminum (c = 0.897 J/(g·°C)) and iron (c = 0.449 J/(g·°C)) each absorb 500. J. Circle the sample with the larger temperature increase: **aluminum** / **iron** / **same**. Explain.

10. A 45.0 g sample of an unknown metal at 98.5 °C is dropped into 75.0 g of water at 21.0 °C in a coffee-cup calorimeter. The final temperature is 25.0 °C. Assume no heat is lost to the surroundings.

(a) Calculate the heat gained by the water.

(b) Calculate the specific heat of the metal.

(c) Use the table to identify the metal.

Metal	**Aluminum**	**Copper**	**Iron**	**Lead**
c (J/(g·°C))	0.897	0.385	0.449	0.128

11. A student dissolves 5.35 g of NH_4Cl (molar mass 53.49 g/mol) in 100.0 g of water in a coffee-cup calorimeter. The temperature falls from 24.0 °C to 20.6 °C. Assume the solution has a specific heat of 4.184 J/(g·°C).

(a) Is the dissolving endothermic or exothermic? Cite your evidence.

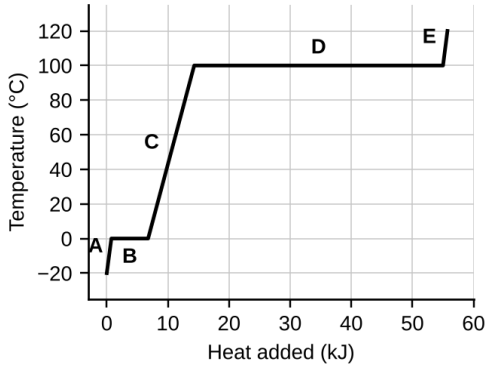
(b) Calculate q for the solution.

(c) Calculate ΔH_{soln} for NH_4Cl in kJ/mol.

(d) The calorimeter was not well insulated, so some heat flowed in from the room during the experiment. Would the calculated ΔH_{soln} be too large or too small? Explain.

Part 5. Energy of Phase Changes (Topic 6.5)

12. The heating curve shows 1 mol of water heated at a steady rate from −20 °C to 120 °C.



(a) In which segments does the temperature stay constant? What is the added energy doing during those segments?

(b) Why is segment D so much longer than segment B?

(c) Is energy absorbed or released when steam condenses? What is $\Delta H_{condensation}$ for water?

13. Calculate the total heat needed to turn 36.0 g of ice at −10.0 °C into liquid water at 25.0 °C. (H_2O = 18.02 g/mol)

14. Steam at 100 °C condenses on a cook's hand.

(a) How much heat is released when 27.0 g of steam condenses to liquid water at 100 °C?

(b) Explain why a burn from steam at 100 °C is usually worse than a burn from the same mass of liquid water at 100 °C.

Part 6. Introduction to Enthalpy of Reaction (Topic 6.6)

15. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -572$ kJ (for the reaction as written)
How much heat is released when 10.0 g of H_2 burns? (H_2 = 2.016 g/mol)

16. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = -890.$ kJ (for the reaction as written)
What mass of methane must burn to release 1.00×10^3 kJ of heat? (CH_4 = 16.04 g/mol)

17. When aqueous solutions of an acid and a base are mixed, the temperature of the mixture rises. Explain at the particle level where the energy that warms the solution comes from.

Part 7. Bond Enthalpies (Topic 6.7)

18. Use the average bond enthalpies to estimate ΔH for the combustion of methane: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Bond	kJ/mol	Bond	kJ/mol	Bond	kJ/mol	Bond	kJ/mol
H–H	436	C–H	413	O–H	467	C–O	358
C–C	347	C=C	614	O=O	495	C=O (in CO_2)	799

(a) Calculate the energy needed to break all the bonds in the reactants.

(b) Calculate the energy released when all the bonds in the products form.

(c) Estimate ΔH. Is the reaction endothermic or exothermic?

19. Estimate ΔH for $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$ using the same table. Hint: draw Lewis structures and count only the bonds that change.

20. Using standard enthalpies of formation, ΔH° for the methane combustion in Question 18 is −802 kJ/mol. Why is the bond enthalpy estimate different?

Part 8. Enthalpy of Formation (Topic 6.8)

21. Which equation represents the standard enthalpy of formation of $NH_3(g)$?

Choice	Equation	Choice	Equation
(A)	$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$	(B)	$\frac{1}{2}N_2(g) + (3/2)H_2(g) \rightarrow NH_3(g)$
(C)	$N(g) + 3H(g) \rightarrow NH_3(g)$	(D)	$NH_3(g) \rightarrow \frac{1}{2}N_2(g) + (3/2)H_2(g)$

(a) Circle the correct equation and explain what is wrong with one of the others.

(b) Why is $\Delta H°_f$ for $O_2(g)$ equal to zero?

22. Propane burns according to $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$. Use the table.

Substance	$C_3H_8(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(l)$
$\Delta H°_f$ (kJ/mol)	−103.8	0	−393.5	−285.8

(a) Calculate ΔH° for the reaction.

(b) How much heat is released when 22.0 g of propane burns? (C_3H_8 = 44.10 g/mol)

23. Ethanol burns according to $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$, ΔH° = −1367 kJ/mol. Use the table in Question 22 to calculate $\Delta H°_f$ for $C_2H_5OH(l)$.

Part 9. Hess's Law (Topic 6.9)

24. For the reaction A + 2B → C, ΔH = −200. kJ. Complete the table.

Equation	ΔH (kJ)	Rule used
C → A + 2B		
2A + 4B → 2C		
½C → ½A + B		

25. Use the equations to calculate ΔH° for $C(s) + 2H_2(g) \rightarrow CH_4(g)$. Show how you combine them.

Equation	ΔH° (kJ/mol)
(1) $C(s) + O_2(g) \rightarrow CO_2(g)$	−393.5
(2) $H_2(g) + ½O_2(g) \rightarrow H_2O(l)$	−285.8
(3) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	−890.3

26. Given $N_2(g) + O_2(g) \rightarrow 2NO(g)$, ΔH° = +180.5 kJ, and $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$, ΔH° = −114.1 kJ, calculate ΔH° for $½N_2(g) + O_2(g) \rightarrow NO_2(g)$. Then explain why ΔH for an overall reaction equals the sum of ΔH for its steps.