ANSWER KEY Answers and explanations appear after each question, marked by a bar at the left. Answers written in table cells are in *italics*.

ZornScience
zornscience.org

AP Chemistry Unit 6 Study Guide | ANSWER KEY

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	*Endothermic*	+	*Falls*
Water vapor condensing on a cold glass	*Exothermic*	−	*Rises*
Propane burning in a gas grill	*Exothermic*	−	*Rises*
NH_4NO_3 dissolving in a cold pack (it gets cold)	*Endothermic*	+	*Falls*
Iron rusting in a hand warmer (it gets warm)	*Exothermic*	−	*Rises*

> Energy is conserved. Whatever energy the system gains (+q, endothermic) comes from the surroundings, so they cool. Whatever energy the system loses (−q, exothermic) goes to the surroundings, so they 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.

> **Exothermic.** Energy is absorbed to pull the Ca^{2+} and Cl^- ions apart (ion-ion attractions) and to push some water molecules apart (hydrogen bonds). Energy is released when new ion-dipole attractions form between the ions and water. More energy is released forming ion-dipole attractions than is absorbed breaking the old attractions, so the net energy goes into the water and its temperature rises.

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.

> **False.** Breaking a bond always requires (absorbs) energy. Energy is released when bonds form.

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

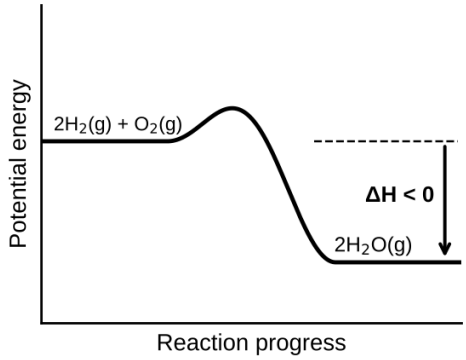
> **True.** Energy is conserved, so energy that leaves the system ends up in the surroundings.

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

> **False.** Thermal energy flows from your warmer hand into the pack, where the endothermic dissolving absorbs it. Your hand loses energy, so it feels cold. "Cold" is not something that flows.

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.

> **Exothermic.** The products are lower in potential energy than the reactants, so the difference in energy is released to the surroundings.

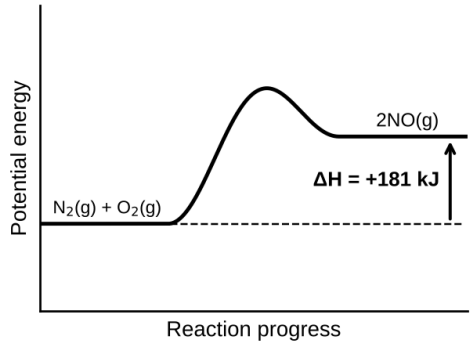
(b) Draw an arrow on the diagram that represents ΔH for the reaction.

> The arrow starts at the reactant level and points down to the product level. ΔH is negative.

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

> **The products.** Lower potential energy means more energy is released forming the bonds in $2H_2O$ than is absorbed breaking the bonds in $2H_2$ and O_2.

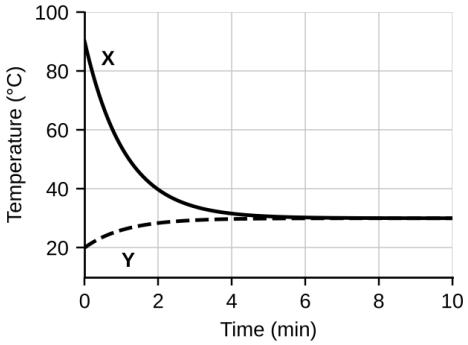
5. The reaction $N_2(g) + O_2(g) \rightarrow 2NO(g)$ has ΔH = +181 kJ/mol. On the axes, sketch an energy diagram for this reaction. Label the reactants, the products, and ΔH.



> The products (2NO) are drawn higher than the reactants ($N_2 + O_2$). The ΔH arrow points up from the reactant level to the product level and is labeled +181 kJ. The hump between them is the activation energy, which you studied in Unit 5; it is fine to include it.

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?

> **Curve X.** It starts at the higher temperature and falls as the metal loses energy to the water.

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

> The metal atoms have a greater average kinetic energy than the water molecules. When metal atoms collide with water molecules at the surface, energy is transferred to the water molecules. The metal particles slow down and the water molecules speed up.

(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?

> **Temperature.** At thermal equilibrium the particles in both have the same average kinetic energy. They keep colliding, but there is no longer any net transfer of energy.

(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.

> The water has a much larger heat capacity (mass × specific heat) than the metal, so the same amount of energy produces a smaller temperature change in the water.

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?

> $q = mc\Delta T = (250.0\text{ g})(4.184\text{ J/(g·°C)})(85.0\text{ °C} - 22.0\text{ °C}) = 6.59 \times 10^4\text{ J} =$ **65.9 kJ**

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.

> $q = nC_m\Delta T = (2.00\text{ mol})(75.3\text{ J/(mol·°C)})(15.0\text{ °C}) = 2.26 \times 10^3\text{ J} =$ **2.26 kJ**

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.

> **Iron.** ΔT = q/(mc). Iron's specific heat is about half of aluminum's, so the same energy raises its temperature about twice as much.
> Iron: ΔT = 500. J ÷ (100. g × 0.449 J/(g·°C)) = 11.1 °C. Aluminum: ΔT = 5.57 °C.

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.

> $q_{water} = (75.0\text{ g})(4.184\text{ J/(g·°C)})(25.0\text{ °C} - 21.0\text{ °C}) = 1255\text{ J}$, or **$1.3 \times 10^3$ J** (ΔT = 4.0 °C has 2 significant figures).

(b) Calculate the specific heat of the metal.

> $q_{metal} = -q_{water} = -1255\text{ J}$
> $c = q/(m\Delta T) = -1255\text{ J} \div [(45.0\text{ g})(25.0\text{ °C} - 98.5\text{ °C})] =$ **0.38 J/(g·°C)**
> (Carry the unrounded 1255 J; using 1.3×10^3 J gives 0.39 J/(g·°C), also acceptable.)

(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

> **Copper.** Its specific heat, 0.385 J/(g·°C), is closest to the measured 0.38 J/(g·°C).

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.

> **Endothermic.** The temperature of the solution fell, so the dissolving absorbed energy from the water.

(b) Calculate q for the solution.

> m = 100.0 g + 5.35 g = 105.35 g. q_{soln} = (105.35 g)(4.184 J/(g·°C))(20.6 °C − 24.0 °C) = **-1.5×10^3 J**

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

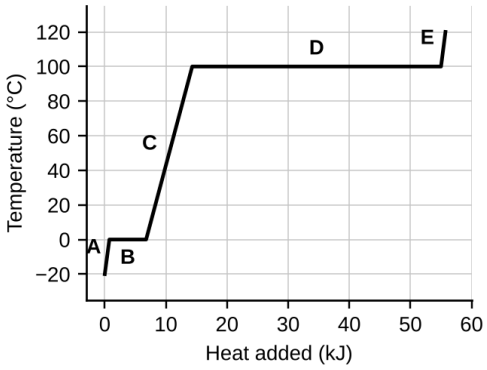
> n = 5.35 g ÷ 53.49 g/mol = 0.1000 mol. $q_{dissolving} = -q_{soln} = +1.5 \times 10^3$ J
> ΔH_{soln} = +1.5 kJ ÷ 0.1000 mol = **+15 kJ/mol**

(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.

> **Too small.** Heat from the room replaced some of the energy the dissolving absorbed, so the measured temperature drop was smaller than it should be. A smaller ΔT gives a smaller q and a ΔH_{soln} that is smaller than the true value.

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 and D.** The added energy overcomes the attractions between water molecules (it raises their potential energy) instead of making them move faster, so the average kinetic energy and the temperature stay the same.

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

> Boiling (D) separates the molecules completely, so nearly all of the hydrogen bonds must be overcome. Melting (B) only loosens them. That is why ΔH_{vap} (40.7 kJ/mol) is much larger than ΔH_{fus} (6.01 kJ/mol).

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

> **Released.** Condensing is the reverse of vaporizing, so $\Delta H_{condensation} = -\Delta H_{vap}$ = **−40.7 kJ/mol**.

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)

> Warm the ice: q_1 = (36.0 g)(2.09 J/(g·°C))(10.0 °C) = 752 J = 0.752 kJ
> Melt the ice: q_2 = (36.0 g ÷ 18.02 g/mol)(6.01 kJ/mol) = (1.998 mol)(6.01 kJ/mol) = 12.0 kJ
> Warm the water: q_3 = (36.0 g)(4.184 J/(g·°C))(25.0 °C) = 3766 J = 3.77 kJ
> Total: q = 0.752 + 12.0 + 3.77 = **16.5 kJ**. Most of the energy goes into melting.

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?

> n = 27.0 g ÷ 18.02 g/mol = 1.498 mol. q = (1.498 mol)(40.7 kJ/mol) = **61.0 kJ released** (ΔH = −61.0 kJ).

(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.

> The steam first releases its large enthalpy of vaporization as it condenses on the skin, and then the hot liquid releases more energy as it cools. Liquid water at 100 °C only releases the energy from cooling, so much less energy reaches the skin.

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)

> mol H_2 = 10.0 g ÷ 2.016 g/mol = 4.96 mol. The reaction releases 572 kJ for every 2 mol of H_2.
> q = 4.96 mol H_2 × (572 kJ / 2 mol H_2) = **1.42×10^3 kJ released**

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)

> n = 1.00×10^3 kJ × (1 mol CH_4 / 890. kJ) = 1.12 mol CH_4. Mass = (1.12 mol)(16.04 g/mol) = **18.0 g**

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.

> The bonds formed in the products are stronger than the bonds broken in the reactants, so the products have lower chemical potential energy. That potential energy difference becomes kinetic energy of the particles in the solution. Faster-moving particles mean a higher temperature.

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.

> Bonds broken: 4 C–H + 2 O=O = 4(413) + 2(495) = **2642 kJ**

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

> Bonds formed: 2 C=O + 4 O–H = 2(799) + 4(467) = **3466 kJ**

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

> $\Delta H \approx 2642$ kJ − 3466 kJ = **−824 kJ/mol**. Exothermic: more energy is released forming bonds than is absorbed breaking them.

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.

> Broken: one C=C and one H–H = 614 + 436 = 1050 kJ. Formed: one C–C and two new C–H = 347 + 2(413) = 1173 kJ.
> $\Delta H \approx 1050 - 1173$ = **−123 kJ/mol**. (Counting every bond also works; the four C–H bonds in C_2H_4 cancel out.)

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

> Bond enthalpies are averages taken over many different molecules. The actual strength of a bond depends on the molecule it is in, so a bond enthalpy calculation is only an estimate.

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).** A formation reaction makes exactly 1 mol of the compound from its elements in their standard states. (A) makes 2 mol, (C) starts from atoms instead of N_2 and H_2, and (D) is the reverse reaction.

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

> $O_2(g)$ is already an element in its standard state, so forming it from itself involves no change.

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 $\Delta H°$ for the reaction.

> $\Delta H° = [3(-393.5) + 4(-285.8)] - [(-103.8) + 5(0)] = (-2323.7) - (-103.8) =$ **−2219.9 kJ/mol**

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

> n = 22.0 g ÷ 44.10 g/mol = 0.499 mol. q = (0.499 mol)(2219.9 kJ/mol) = **1.11×10^3 kJ released**

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

> $-1367 = [2(-393.5) + 3(-285.8)] - [\Delta H°_f(C_2H_5OH) + 3(0)] = -1644.4 - \Delta H°_f$, so $\Delta H°_f(C_2H_5OH) =$ **−277.4 kJ/mol**

Part 9. Hess's Law (Topic 6.9)

24. For the reaction $A + 2B \rightarrow C$, $\Delta H = -200.$ kJ. Complete the table.

Equation	ΔH (kJ)	Rule used
$C \rightarrow A + 2B$	*+200.*	*Reversed: change the sign*
$2A + 4B \rightarrow 2C$	*−400.*	*Multiplied by 2: multiply ΔH by 2*
$\frac{1}{2}C \rightarrow \frac{1}{2}A + B$	*+100.*	*Reversed and multiplied by ½: change the sign and halve ΔH*

25. Use the equations to calculate $\Delta H°$ 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) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	−285.8
(3) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	−890.3

> Keep (1): −393.5 kJ. Multiply (2) by 2: −571.6 kJ. Reverse (3): +890.3 kJ. Adding them cancels CO_2, H_2O, and O_2, leaving $C(s) + 2H_2(g) \rightarrow CH_4(g)$.
> $\Delta H° = -393.5 - 571.6 + 890.3 =$ **−74.8 kJ/mol**

26. Given $N_2(g) + O_2(g) \rightarrow 2NO(g)$, $\Delta H° = +180.5$ kJ, and $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$, $\Delta H° = -114.1$ kJ, calculate $\Delta H°$ for $\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g)$. Then explain why ΔH for an overall reaction equals the sum of ΔH for its steps.

> Add the equations (2NO cancels): $N_2 + 2O_2 \rightarrow 2NO_2$, $\Delta H° = +180.5 - 114.1 = +66.4$ kJ. Halve it: $\Delta H° =$ **+33.2 kJ**.
> Energy is conserved, so the net energy transferred by the sequence is the sum of the energy transferred in each step.