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Topic 6.1: Endothermic and Exothermic Processes

Worksheet · AP Chemistry

UNIT 6 · THERMOCHEMISTRY

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How to Use This Worksheet

Work in order. Section A checks the definitions and the system and surroundings labels. Section B practises the skills: reading enthalpy diagrams, running an energy ledger between system and surroundings, and classifying phase changes with numbers. Section C is exam style: judging claims and two full scenarios.

Reference facts. ΔH is always reported from the system's point of view: positive for endothermic (system absorbs heat), negative for exothermic (system releases heat). Whatever the system gains, the surroundings lose, and the reverse: the two changes are equal in size and opposite in sign. Separating particles against their attractions (melting, evaporating, subliming) is endothermic; forming attractions (freezing, condensing, depositing) is exothermic. Useful values: melting 1 mol of ice absorbs 6.01 kJ; evaporating 1 mol of water absorbs 40.7 kJ; burning methane has $\Delta H = -890$ kJ/mol; neutralisation of a strong acid with a strong base has $\Delta H \approx -57$ kJ/mol.

Three habits protect every answer. Name the system before you assign any sign. Check every energy ledger balances: gain on one side, equal loss on the other. And read what a hand or thermometer feels as a report about the surroundings, not the system.

Do not look at the answer key until you have attempted every question.

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Section A: Recall and Concepts

A1. (a) Define the system in a chemistry experiment. (b) Define the surroundings. (c) In a coffee-cup calorimetry experiment, a student mixes two solutions and records the temperature change of the mixture. State what the system is, and state two things that belong to the surroundings. (d) State why the choice of boundary matters when a sign is assigned to ΔH.

A2. (a) State what heat is, and the condition that makes heat flow. (b) State the direction of heat flow in an endothermic process. (c) State the direction of heat flow in an exothermic process. (d) State what happens to the temperature of the surroundings in each case.

A3. (a) State the sign of ΔH for an endothermic process, from the system's point of view, and explain the sign. (b) State the sign of ΔH for an exothermic process, from the system's point of view, and explain the sign. (c) A process has ΔH = +45 kJ/mol. State whether it is endothermic or exothermic, and state how much heat is transferred, and in which direction, when 1 mol reacts. (d) A reaction gives off 120 kJ of heat per mole. Write its ΔH value with its sign and units.

A4. Classify each process as endothermic or exothermic, giving a one-line reason for each. (a) Ice melting in a drink. (b) Water freezing in a freezer tray. (c) Alcohol evaporating from skin. (d) Methane burning in oxygen: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. (e) Steam condensing on a bathroom mirror. (f) Dry ice subliming at room temperature.

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A5. (a) State what temperature measures at the particle level. (b) State whether heat and temperature are the same quantity, and justify your answer in one sentence. (c) The same amount of heat is released into 50 g of water and into 500 g of water. State which water shows the larger temperature rise, and explain why in terms of particles. (d) A thermometer in a calorimeter shows a temperature rise. State which side of the boundary the thermometer is reporting, and state what that rise tells you about the reaction.

Section B: Core Practice

B1. An enthalpy diagram shows reactants at 95 kJ/mol and products at 140 kJ/mol. (a) Calculate ΔH, showing the subtraction you use. (b) Classify the reaction as endothermic or exothermic, using the sign and the relative levels. (c) State how much heat is transferred when 1 mol reacts, and in which direction. (d) Calculate ΔH for the reverse reaction and classify it. (e) State what the surroundings gain or lose when 1 mol of the forward reaction occurs.

B2. Methane burns with $\Delta H = -890$ kJ/mol, from the system's point of view. (a) State how much heat is released to the surroundings when 1 mol of methane burns. (b) Calculate the heat released when 0.50 mol of methane burns. (c) Write the energy ledger for that 0.50 mol burn: the change for the system and the change for the surroundings. (d) State what happens to the average kinetic energy of the particles in the surroundings, and what a thermometer placed there shows. (e) Explain, at the level of bonds, why this reaction is exothermic.

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B3. A student dissolves 5.35 g of ammonium chloride (0.100 mol) in water and the temperature of the solution falls. The dissolving absorbs 1.5 kJ of heat. (a) Calculate ΔH for the dissolving, per mole, with its sign. (b) State the system and the surroundings in this experiment. (c) Explain why the beaker feels cold, in terms of energy crossing the boundary. (d) State what happens to the average kinetic energy of the water particles. (e) A classmate says the cold solution "contains more cold than before". Correct the statement.

B4. Use the reference values: melting ice, ΔH = +6.01 kJ/mol; evaporating water, ΔH = +40.7 kJ/mol, both from the system's point of view. (a) Calculate the heat absorbed when 0.25 mol of water evaporates. (b) Write ΔH for condensing steam, per mole, with its sign, and calculate the heat transferred when 2.0 mol of steam condenses, stating the direction. (c) Write ΔH for freezing water, per mole, with its sign, and calculate the heat transferred when 5.0 mol of water freezes, stating the direction. (d) Explain why evaporation cools the skin while condensation warms a mirror, using the same particle-level idea for both.

B5. A student mixes HCl solution with NaOH solution in a foam cup. The temperature of the mixture rises by several degrees: $HCl + NaOH \rightarrow NaCl + H_2O$. (a) State the system and the surroundings. (b) State the direction of heat flow across the boundary. (c) Classify the neutralisation as endothermic or exothermic. (d) State the sign of ΔH, from the system's point of view, and justify it from the temperature reading. (e) If 0.20 mol of water forms and ΔH = −57 kJ/mol, calculate the heat gained by the surroundings.

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B6. Photosynthesis and respiration are reverse processes: photosynthesis, $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$; respiration, $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$. (a) Photosynthesis is endothermic. State the sign of its ΔH and state where the absorbed energy ends up. (b) State the classification and the sign of ΔH for respiration. (c) The combustion of glucose has $\Delta H = -2{,}808$ kJ/mol. State the value of ΔH for photosynthesis, per mole of glucose formed. (d) Explain why the two ΔH values must be equal in size and opposite in sign. (e) A leaf in sunlight stays cool while photosynthesis runs. Suggest why, in terms of energy flow.

Section C: Exam Style

C1. Judge each claim, supporting each verdict with the definition or the sign rule it depends on. (a) "The beaker felt cold during the reaction, so the reaction released heat." (b) "$\Delta H = +26$ kJ/mol means the surroundings gained 26 kJ of enthalpy per mole." (c) "Energy is conserved in every reaction, so ΔH is always zero." (d) "Condensation is endothermic, because the gas particles slow down as they become a liquid." (e) "Doubling the amount of reactant doubles the heat transferred, but ΔH in kJ/mol stays the same."

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C2. A cold pack contains ammonium nitrate and water in separate compartments. Squeezing the pack mixes them and the pack turns cold: $NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$. In a test, dissolving 0.100 mol absorbs 2.6 kJ of heat. (a) Calculate ΔH for the dissolving, per mole, with its sign, from the system's point of view. (b) Identify the system and the surroundings when the pack is held against an injured ankle. (c) Explain, in terms of heat flow and particle kinetic energy, why the ankle feels cold. (d) State the energy change of the surroundings for the 0.100 mol test, with its sign, and show that the total energy change of system plus surroundings is zero. (e) The pack is left until it returns to room temperature. State whether the dissolving process was endothermic or exothermic, and state what supplied the energy that the system gained.

C3. An athlete's muscles respire glucose: $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$, $\Delta H = -2{,}808$ kJ/mol from the system's point of view. (a) Calculate the heat released to the surroundings when 0.10 mol of glucose is respired. (b) State where that heat goes in the body, and what happens to the average kinetic energy of the particles it reaches. (c) The athlete sweats, and the sweat evaporates from the skin. Classify evaporation as endothermic or exothermic, state the direction of heat flow, and explain why it cools the athlete. (d) A spectator says: "Respiration releases heat, so the glucose must end up with more enthalpy than it started with." Correct the statement, using the definition of ΔH. (e) State the classification and ΔH sign for the reverse of respiration, and name the process that carries it out.

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Answer Key

A1. (a) The part under study: the reactants and products of the process being investigated. (b) Everything outside the system that can exchange energy with it: solvent, container, thermometer, air, and anything touching the apparatus. (c) The system is the reacting chemicals (the two solutions as they react). The surroundings include the water the reaction sits in, the cup and lid, the thermometer and the stirrer (any two). (d) Because ΔH is defined from the system's side of the boundary: the same heat flow is a loss for one side and a gain for the other, so the sign flips if the labels are swapped.

A2. (a) Heat is energy transferred across the boundary because of a temperature difference between system and surroundings. (b) From the surroundings into the system. (c) From the system into the surroundings. (d) Endothermic: the surroundings lose energy and their temperature falls. Exothermic: the surroundings gain energy and their temperature rises.

A3. (a) Positive. The system absorbs heat and gains enthalpy, so H(products) is greater than H(reactants) and the difference is positive. (b) Negative. The system releases heat and loses enthalpy, so H(products) is less than H(reactants) and the difference is negative. (c) Endothermic: 45 kJ of heat is absorbed by the system from the surroundings for each mole that reacts. (d) ΔH = −120 kJ/mol: the heat is given off, so the system loses enthalpy and the sign is negative.

A4. (a) Endothermic: melting separates particles against their attractions, which costs energy drawn from the drink. (b) Exothermic: freezing forms attractions as the liquid becomes solid, releasing heat to the surroundings. (c) Endothermic: evaporation separates the liquid's particles, drawing heat from the skin. (d) Exothermic: combustion forms strong new bonds in CO_2 and H_2O, releasing more energy than breaking the reactant bonds cost. (e) Exothermic: condensing brings vapour particles together and forms attractions, releasing heat to the mirror. (f) Endothermic: subliming separates the solid's particles completely, absorbing heat from the room.

A5. (a) Temperature measures the average kinetic energy of the particles. (b) No: heat is an amount of energy transferred, while temperature is an average per particle, so the same heat can produce different temperatures in different amounts of material. (c) The 50 g of water: the same energy is shared among fewer particles, so the average kinetic energy per particle, and therefore the temperature, rises more. (d) The surroundings: the thermometer sits outside the reacting system, in the water and cup. A rise there means heat arrived in the surroundings, so the system released it and the reaction is exothermic.

B1. (a) ΔH = H(products) − H(reactants) = 140 − 95 = +45 kJ/mol. (b) Endothermic: ΔH is positive and the products sit higher in enthalpy than the reactants. (c) 45 kJ of heat is absorbed by the system from the surroundings per mole. (d) ΔH(reverse) = −45 kJ/mol, exothermic: the reverse reaction

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runs downhill from 140 to 95 kJ/mol and releases the 45 kJ. (e) The surroundings lose 45 kJ per mole: they supply exactly what the system gains, so their temperature falls.

B2. (a) 890 kJ per mole burned. (b) 0.50 × 890 = 445 kJ released. (c) System: −445 kJ. Surroundings: +445 kJ. The two cancel, as conservation requires. (d) It increases: the transferred energy becomes kinetic energy of the surroundings' particles, and the thermometer shows a temperature rise. (e) The bonds formed in the CO_2 and H_2O products are strong, and forming them releases more energy than was needed to break the bonds in CH_4 and O_2. The surplus leaves the system as heat.

B3. (a) ΔH = 1.5 ÷ 0.100 = +15 kJ/mol. Positive, because the dissolving system absorbs the heat. (b) The system is the dissolving ammonium chloride (the solute as it dissolves). The surroundings are the water, the beaker and the air and hand touching it. (c) Heat flows from the surroundings into the system to supply the dissolving. The beaker is part of the surroundings, so it loses energy and its temperature drops; your hand detects that loss as cold. (d) It decreases: the water particles gave up energy to the system, so their average kinetic energy, and the temperature, fall. (e) There is no such substance as cold to contain. The solution feels cold because it holds less energy than before: heat left the solution and entered the dissolving system.

B4. (a) 0.25 × 40.7 = 10.2 kJ absorbed by the system from the surroundings. (b) Condensing is the reverse of evaporating, so ΔH = −40.7 kJ/mol. For 2.0 mol: 2.0 × 40.7 = 81.4 kJ released by the system to the surroundings. (c) Freezing is the reverse of melting, so ΔH = −6.01 kJ/mol. For 5.0 mol: 5.0 × 6.01 = 30.05 kJ released by the system to the surroundings. (d) Evaporation must pull particles apart against their attractions, and the energy for that separation is taken from the skin, whose particles lose kinetic energy and cool. Condensation brings particles together and forms those attractions, releasing energy that arrives in the mirror and warms it.

B5. (a) The system is the reacting HCl and NaOH (the neutralisation reaction). The surroundings are the solution the reaction sits in, the foam cup, the thermometer and the air. (b) From the system into the surroundings. (c) Exothermic. (d) Negative. The thermometer sits in the surroundings and shows their temperature rising, which means heat arrived there from the system; a system that loses enthalpy has ΔH negative from its own point of view. (e) Heat released by the system = 0.20 × 57 = 11.4 kJ, and all of it (in the ideal calorimetry picture) is gained by the surroundings, so the surroundings gain 11.4 kJ.

B6. (a) Positive: an endothermic process has ΔH greater than zero from the system's point of view. The absorbed energy is stored as chemical potential energy in the glucose and oxygen formed. (b) Respiration is exothermic, with ΔH negative. (c) +2,808 kJ/mol: photosynthesis is the exact reverse of the combustion, so its ΔH has the same size and the opposite sign. (d) Reversing a process swaps the roles of reactants and products in ΔH = H(products) − H(reactants), which flips the sign while keeping the size, because the same two enthalpy levels are involved. (e) The sunlight energy the leaf absorbs is carried into stored chemical energy by photosynthesis instead of raising the leaf particles' kinetic energy.

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C1. (a) False. A cold beaker is losing heat to the process inside it; heat flowing into the system is an endothermic process, not a release. (b) False. ΔH is defined from the system's point of view: +26 kJ/mol means the system gained 26 kJ per mole. The surroundings lost 26 kJ per mole. (c) False. Conservation applies to system and surroundings combined: their changes must cancel, but the system's own ΔH can be any positive or negative value. (d) False. The particles do slow down, but the classification follows the heat flow: condensing forms attractions and releases heat to the surroundings, so it is exothermic. Slowing particles are the result of losing energy, not evidence of absorbing it. (e) True. Heat transferred is an extensive amount and scales with the quantity reacted; ΔH in kJ/mol is defined per mole, so it is unchanged when the amount doubles.

C2. (a) ΔH = 2.6 ÷ 0.100 = +26 kJ/mol. (b) The system is the dissolving ammonium nitrate and the mixture it forms inside the pack. The surroundings are the pack material, the ankle and skin it touches, and the air. (c) The dissolving absorbs heat, so heat flows from the skin into the pack. The skin's particles lose kinetic energy, their average kinetic energy falls, and that fall is what is felt as cold. (d) The system gained 2.6 kJ, so the surroundings changed by −2.6 kJ. Total: +2.6 + (−2.6) = 0, as conservation requires. (e) Endothermic: it absorbed heat throughout. The energy came from the surroundings: the water and pack material and, while held in place, the ankle, which is why they cooled.

C3. (a) 0.10 × 2,808 = 280.8 kJ released to the surroundings. (b) It passes to the body's tissues and blood, the surroundings of the reacting system, whose particles gain kinetic energy, which is part of how the body stays warm. (c) Endothermic. Heat flows from the skin into the evaporating sweat. Separating the water particles against their attractions takes energy, the skin supplies it and loses kinetic energy, so the skin cools. (d) The statement has the enthalpy change backwards. ΔH = H(products) − H(reactants) = −2,808 kJ/mol is negative, so the products (CO_2 and H_2O) have lower enthalpy than the glucose and oxygen started with; the difference is exactly the heat released. (e) Endothermic, with ΔH positive: photosynthesis, carried out by plants, reverses respiration and stores the energy again.

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