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Topic 6.6: Introduction to Enthalpy of Reaction

Worksheet · AP Chemistry

UNIT 6 · THERMOCHEMISTRY

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How to Use This Worksheet

Work in order. Section A checks the definitions and rules: what enthalpy is, what a state function means, what a thermochemical equation must carry, and the standard conditions behind ΔH°. Section B practises the two working rules, reversing and scaling, on six different reactions, including the per-mole pricing that every heat question reduces to. Section C is exam style: a multi-part peroxide problem, six claims to judge, and a Haber process calculation that ends with heat per gram of product.

Reference data. Thermochemical equations used throughout: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -572$ kJ. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, $\Delta H = -92$ kJ. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, $\Delta H = -890$ kJ. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$, $\Delta H = +178$ kJ. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$, $\Delta H = -566$ kJ. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$, $\Delta H = -184$ kJ. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$, $\Delta H = -198$ kJ. $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$, $\Delta H = -196$ kJ. Molar masses: H_2O_2 34.02 g mol^{-1}, NH_3 17.03 g mol^{-1}, N_2 28.02 g mol^{-1}. Signs are taken from the system's point of view: heat released by the system is negative, heat absorbed is positive.

Three habits protect every answer. Read the ΔH as pricing the equation exactly as written, never as a per-mole value, until you have divided by a coefficient yourself. When you rewrite an equation, carry the ΔH with it: reverse it and flip the sign, scale it and scale the value by the same factor. Finish by checking the sign against the equation's own ΔH: formation and combustion reactions here release heat, so their answers must come out negative; the limestone decomposition absorbs heat, so its answer comes out positive, but the peroxide

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decomposition in C1 releases heat, so its answer comes out negative. Never guess the sign from the reaction type alone.

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

Section A: Recall and Concepts

A1. (a) Define enthalpy and write the equation that defines it, naming each symbol. (b) At constant pressure, the heat exchanged by a reaction equals which quantity? (c) Use that fact to explain why enthalpy is more useful than internal energy for reactions run in open flasks. (d) State the sign of ΔH for an exothermic reaction and for an endothermic reaction, and state what each sign says about the direction of heat flow.

A2. (a) State what it means to call enthalpy a state function. (b) The same reactants reach the same products by two different routes, one using a catalyst and one not. Compare the ΔH of the two routes and justify the comparison. (c) Write ΔH in terms of the enthalpies of the reactants and the products. (d) A reaction runs in five small steps instead of one. State, with a reason, whether its overall ΔH changes.

A3. (a) List the three things a thermochemical equation must include. (b) Explain why the physical states belong in it. (c) For $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -572$ kJ, state exactly which amounts of which substances the −572 kJ refers to. (d) A card shows only "$H_2 + O_2 \rightarrow H_2O$". Name three faults in it as a thermochemical equation.

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A4. (a) Write the symbol for a standard enthalpy change and state what the symbol's extra mark means. (b) State the standard pressure for a gas. (c) State the temperature at which thermochemical data are usually tabulated. (d) State the standard concentration for a substance in solution. (e) State the standard enthalpy of formation of $O_2(g)$, and explain why it has that value.

A5. Start from $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -572$ kJ. Write the ΔH for each rewritten version: (a) the equation halved; (b) the equation tripled; (c) the equation reversed; (d) the amount of H_2 named in the original equation replaced by 1 mol of H_2 reacting, coefficients adjusted to match; (e) the heat released per mole of O_2 consumed.

Section B: Core Practice

B1. Ammonia synthesis: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, $\Delta H = -92$ kJ. (a) Calculate the heat released per mole of NH_3 formed. (b) Write the thermochemical equation for the formation of 1 mol of NH_3, fractional coefficients allowed, and state its ΔH. (c) Calculate the heat released when 6.0 mol of H_2 reacts completely. (d) Calculate the heat released when 0.25 mol of N_2 reacts completely.

B2. Methane combustion: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, $\Delta H = -890$ kJ. (a) Calculate the heat released when 2.5 mol of CH_4 burns completely. (b) Calculate the heat released per mole of H_2O formed. (c) Write the thermochemical equation for the combustion of 2 mol of CH_4 and state its ΔH. (d) State whether the ΔH in part (c) describes a more or less exothermic reaction than the original, and explain the answer in one sentence.

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B3. Limestone decomposition: $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$, $\Delta H = +178$ kJ. (a) Calculate the heat absorbed when 0.50 mol of $CaCO_3$ decomposes. (b) Write the thermochemical equation for the reverse process, CaO reacting with CO_2, and state its ΔH. (c) Calculate the heat released when 2.0 mol of CO_2 is captured by CaO in that reverse process. (d) State, with a reason, whether the decomposition is endothermic or exothermic.

B4. Carbon monoxide combustion: $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$, $\Delta H = -566$ kJ. (a) Calculate the heat released per mole of CO burned. (b) Calculate the heat released per mole of CO_2 formed, and comment on how it compares with part (a). (c) Calculate the heat released when 0.40 mol of CO burns. (d) Write the thermochemical equation rewritten per mole of CO, and state its ΔH.

B5. Hydrogen chloride formation: $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$, $\Delta H = -184$ kJ. (a) Calculate the heat released when 5.0 mol of HCl is formed. (b) State the ΔH for $2H_2(g) + 2Cl_2(g) \rightarrow 4HCl(g)$. (c) Calculate the heat absorbed when 1.0 mol of HCl is decomposed back into its elements. (d) Calculate the heat released per mole of Cl_2 consumed.

B6. Sulfur trioxide formation: $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$, $\Delta H = -198$ kJ. (a) Calculate the heat released per mole of SO_3 formed. (b) Calculate the heat released when 10 mol of SO_3 is formed. (c) This ΔH was measured with the reaction open to the air. State which quantity the measured heat equals, and the condition that makes the equality hold. (d) A student claims a negative ΔH means this reaction must be fast. Correct the claim, using the state function idea.

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Section C: Exam Style

C1. Hydrogen peroxide decomposes: $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$, $\Delta H = -196$ kJ. (a) Calculate the heat released per mole of H_2O_2 decomposed. (b) A first-aid solution decomposes completely, and the peroxide in it amounts to 3.4 g. Calculate the heat released. (c) Write the thermochemical equation for the reverse process and state its ΔH. (d) A manganese(IV) oxide catalyst is added and the same sample decomposes in seconds instead of days. State the ΔH of the catalysed decomposition and justify the answer. (e) Calculate the heat released when 0.50 mol of H_2O_2 decomposes.

C2. Judge each claim, supporting each verdict with the rule it depends on. (a) "Doubling every coefficient in a thermochemical equation leaves ΔH unchanged, because it is still the same reaction." (b) "The reverse reaction has the same ΔH as the forward reaction, sign included." (c) "For a reaction run in an open beaker, the heat exchanged equals ΔH." (d) "$\Delta H = -572$ kJ for $2H_2 + O_2 \rightarrow 2H_2O$ means each mole of H_2 releases 572 kJ." (e) "The standard enthalpy of formation of $O_2(g)$ is zero by definition." (f) "A positive ΔH means the reaction gives out heat to the surroundings."

C3. The Haber process runs $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, $\Delta H = -92$ kJ. (a) A batch uses 28.02 g of N_2 with excess H_2 and converts it completely. Calculate the number of moles of N_2 and the heat released. (b) Calculate the mass of NH_3 the batch produces. (c) Calculate the heat released per gram of NH_3 produced. (d) A second batch uses half as much of each reactant. State the heat it releases and the mass of NH_3 it produces. (e) The plant then doubles its hourly throughput at the same recipe. State what happens to (i) the total heat released per hour and (ii) the ΔH of the equation as written, and explain why the two answers differ.

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Answer Key

A1. (a) Enthalpy is the internal energy of a system plus the product of its pressure and volume: H = E + PV, where E is the internal energy, P the pressure and V the volume. (b) The enthalpy change, ΔH: at constant pressure, q_p = ΔH. (c) Open flasks sit at the constant pressure of the atmosphere, so the heat a bench reaction exchanges is ΔH directly. Internal energy misses the expansion work done against the atmosphere when gases are produced, and enthalpy includes it through the PV term. (d) Exothermic: ΔH is negative, because the system loses enthalpy as heat flows out to the surroundings. Endothermic: ΔH is positive, because the system gains enthalpy as heat flows in from the surroundings.

A2. (a) Its value depends only on the current state of the system, the substances, amounts, states, temperature and pressure, and not on the path by which that state was reached. (b) They are equal. ΔH is fixed by the initial and final states alone, and both routes share the same two states; a catalyst changes the path and the speed, neither of which a state function records. (c) ΔH = H(products) − H(reactants). (d) It does not change. The overall change is the difference between the same two endpoint states, however many steps connect them.

A3. (a) A balanced chemical equation; the physical state of every reactant and product; and the ΔH value for the equation as written. (b) Different physical states are different states with different enthalpies: forming liquid water and forming steam from the same reactants give different ΔH values, so an equation without states does not fix which change its ΔH prices. (c) It refers to 2 mol of H_2 reacting with 1 mol of O_2 to form 2 mol of liquid water, releasing 572 kJ in total for those amounts. (d) It is not balanced (one O atom on the right against two on the left); it shows no physical states; and it carries no ΔH value, so no heat is priced at all.

A4. (a) ΔH°. The small circle marks standard conditions: the value applies when every substance is in its standard state. (b) 1 atm (1 bar in the newer convention). (c) 298 K. (d) 1 mol L^{-1}. (e) Zero. By convention the standard enthalpy of formation of an element in its standard state is defined as zero, making the elements the reference from which all formation values are measured; $O_2(g)$ is oxygen's standard state, so it sits on the reference line itself.

A5. (a) Halving scales ΔH by one half: −572 ÷ 2 = −286 kJ. (b) Tripling scales it by three: 3 × (−572) = −1,716 kJ. (c) Reversing keeps the size and flips the sign: +572 kJ. (d) One mole of H_2 is half the equation's 2 mol, so the release is 572 ÷ 2 = 286 kJ, and ΔH = −286 kJ for $H_2(g) + ½O_2(g) \rightarrow H_2O(l)$. (e) The equation consumes 1 mol of O_2 for the full 572 kJ, so the heat released per mole of O_2 consumed is 572 kJ.

B1. (a) The equation forms 2 mol of NH_3 per 92 kJ released, so per mole: 92 ÷ 2 = 46 kJ released. (b) $½N_2(g) + ³⁄₂H_2(g) \rightarrow NH_3(g)$, ΔH = −46 kJ: halving the equation halves ΔH. (c) The equation consumes 3 mol of H_2 per 92 kJ released. 6.0 mol of H_2 is two equation-loads, so the heat released is 2 × 92 =

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184 kJ. (d) The equation consumes 1 mol of N_2 per 92 kJ released, so 0.25 mol releases 0.25 × 92 = 23 kJ.

B2. (a) The equation burns 1 mol of CH_4 per 890 kJ released, so 2.5 mol releases 2.5 × 890 = 2,225 kJ. (b) The equation forms 2 mol of H_2O per 890 kJ released, so per mole of water: 890 ÷ 2 = 445 kJ released. (c) $2CH_4(g) + 4O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$, $\Delta H = 2 \times (-890) = -1{,}780$ kJ. (d) Neither more nor less exothermic per mole of methane: doubling the equation doubles the material burned, so the heat doubles with it, while the release per mole of CH_4 is unchanged at 890 kJ.

B3. (a) The equation decomposes 1 mol of $CaCO_3$ per 178 kJ absorbed, so 0.50 mol absorbs 0.50 × 178 = 89 kJ. (b) $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$, $\Delta H = -178$ kJ: reversal keeps the magnitude and flips the sign. (c) The reverse equation captures 1 mol of CO_2 per 178 kJ released, so 2.0 mol of CO_2 releases 2.0 × 178 = 356 kJ. (d) Endothermic: its ΔH is positive, meaning the system absorbs heat from the surroundings, which a kiln must supply continuously to keep limestone decomposing.

B4. (a) The equation burns 2 mol of CO per 566 kJ released, so per mole: 566 ÷ 2 = 283 kJ released. (b) The equation also forms 2 mol of CO_2 per 566 kJ, so per mole of CO_2: 566 ÷ 2 = 283 kJ released, exactly the same as part (a), because CO and CO_2 carry the same coefficient in this equation. (c) 0.40 mol of CO at 283 kJ per mole releases 0.40 × 283 = 113 kJ. (d) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$, $\Delta H = -283$ kJ: halving the equation halves ΔH.

B5. (a) The equation forms 2 mol of HCl per 184 kJ released, so per mole: 92 kJ. Forming 5.0 mol releases 5.0 × 92 = 460 kJ. (b) Doubling the equation doubles ΔH: $2 \times (-184) = -368$ kJ. (c) Decomposition is the reverse process: forming 1 mol of HCl releases 92 kJ, so decomposing 1.0 mol of HCl absorbs 92 kJ. (d) The equation consumes 1 mol of Cl_2 per 184 kJ released, so the heat released per mole of Cl_2 consumed is 184 kJ.

B6. (a) The equation forms 2 mol of SO_3 per 198 kJ released, so per mole: 198 ÷ 2 = 99 kJ released. (b) 10 mol × 99 kJ = 990 kJ released. (c) The measured heat equals ΔH, and the equality holds because the reaction runs at constant pressure, open to the air: $q_p = \Delta H$. (d) The claim confuses thermodynamics with kinetics. ΔH compares only the enthalpies of the reactants and products, the two endpoint states; it records nothing about the path between them or how fast the reaction travels it, so a very exothermic reaction can still be very slow.

C1. (a) The equation decomposes 2 mol of H_2O_2 per 196 kJ released, so per mole: 196 ÷ 2 = 98 kJ released. (b) n = 3.4 ÷ 34.02 = 0.10 mol of H_2O_2. At 98 kJ released per mole, the heat released is 0.10 × 98 = 9.8 kJ. (c) $2H_2O(l) + O_2(g) \rightarrow 2H_2O_2(l)$, $\Delta H = +196$ kJ: the reverse process keeps the magnitude and flips the sign. (d) ΔH is unchanged at −196 kJ for the equation as written. Enthalpy is a state function, so ΔH depends only on the peroxide at the start and the water and oxygen at the end; the catalyst changes the route and the speed, never the endpoints. (e) 0.50 mol at 98 kJ released per mole gives 0.50 × 98 = 49 kJ released.

C2. (a) False. ΔH is extensive: it prices the amounts in the equation, so doubling the amounts doubles the heat and ΔH doubles with the coefficients. (b) False. Reversing swaps reactants and

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products in ΔH = H(products) − H(reactants), so the value keeps its magnitude and flips its sign. (c) True. An open beaker sits at the constant pressure of the atmosphere, and at constant pressure the heat exchanged is exactly ΔH. (d) False. The 572 kJ prices the whole equation as written, in which 2 mol of H_2 reacts; per mole of H_2 the release is 572 ÷ 2 = 286 kJ. (e) True. An element in its standard state is the reference of the formation scale, and its standard enthalpy of formation is defined as zero. (f) False. Positive ΔH means the system absorbs heat from the surroundings: the process is endothermic. Heat given out carries a negative ΔH.

C3. (a) n = 28.02 ÷ 28.02 = 1.00 mol of N_2. The equation consumes 1 mol of N_2 per 92 kJ released, so the batch releases 92 kJ. (b) The equation forms 2 mol of NH_3 per mole of N_2, so the batch forms 2.00 mol of NH_3, a mass of 2.00 × 17.03 = 34.06 g, about 34.1 g. (c) Heat per gram of NH_3: 92 ÷ 34.06 = 2.70 kJ per gram. (d) Half the material follows the same equation at half the scale: the heat released is 46 kJ, and the NH_3 produced is 1.00 mol, a mass of 17.03 g. (e) (i) The total heat released per hour doubles, because twice as much reaction happens per hour and each equation-load of reaction releases its 92 kJ. (ii) ΔH for the equation as written is unchanged at −92 kJ. The two answers differ because ΔH is a price per equation-load, a fixed property of the reaction's endpoints, while the total heat is that price multiplied by how many loads run: throughput scales the total, never the price.

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