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Topic 6.6: Introduction to Enthalpy of Reaction

Revision Notes · AP Chemistry

UNIT 6 · THERMOCHEMISTRY

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What Enthalpy Is

Enthalpy, H, is defined as

$H = E + PV$:

the internal energy of a system plus the product of its pressure and its volume.

The PV term exists for one practical reason. When a reaction produces gas at constant pressure, some of its energy is spent pushing back the atmosphere. Enthalpy change is the internal energy change with that expansion work already accounted for, so the heat that actually flows equals it: at constant pressure,

$q_p = \Delta H$

Bench chemistry runs in open beakers and flasks, at the constant pressure of the room. That is why enthalpy, not internal energy, is the working measure of reaction heat.

A State Function: Only the Endpoints Matter

Enthalpy is a state function. Its value depends only on the state of the system: which substances are present, their amounts, their physical states, the temperature and the pressure.

It does not depend on the path. Reaching the same products from the same reactants gives the same ΔH whether the reaction runs in one step or several, quickly or slowly, with a catalyst or without one.

So ΔH is fixed by two states only:

$\Delta H = H(\text{products}) - H(\text{reactants})$

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Every rule in this topic is bookkeeping on that single difference, with signs from the system's point of view.

The Thermochemical Equation

A thermochemical equation is a balanced chemical equation with its ΔH written beside it, physical states included:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -572$ kJ

The ΔH belongs to the equation exactly as written. It prices the amounts the coefficients name: 572 kJ is released when 2 mol of H_2 and 1 mol of O_2 react to form 2 mol of liquid water.

The states matter because they are part of the state. Forming liquid water and forming steam from the same reactants are different final states.

The sign reads straight off the equation. Negative ΔH: the products sit lower in enthalpy than the reactants, the difference leaves as heat, and the reaction is exothermic. Positive ΔH: the products sit higher, heat is absorbed from the surroundings, and the reaction is endothermic. In equation form, absorbed heat behaves like a reactant and released heat like a product.

Rule One: Reversing the Reaction Flips the Sign

Reversing an equation swaps the roles of reactants and products, so it swaps the two states in

$\Delta H = H(\text{products}) - H(\text{reactants})$,

and the difference changes sign while keeping its size.

Decomposing water is the exact reverse of forming it:

$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$, $\Delta H = +572$ kJ

The 572 kJ that formation releases, decomposition must absorb.

This is the sign-flip rule of Topic 6.5 applied to reactions: a reverse change always costs exactly what the forward change pays out.

Rule Two: Scaling the Equation Scales ΔH

Enthalpy is an extensive quantity: it scales with the amount of material. Multiply every coefficient in a thermochemical equation by a factor and ΔH is multiplied by the same factor, sign unchanged.

Halve the water equation and ΔH halves:

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$H_2(g) + ½O_2(g) \rightarrow H_2O(l)$, $\Delta H = -286$ kJ

Triple it, and $\Delta H = 3 \times (-572) = -1{,}716$ kJ for 6 mol of water formed.

Fractional coefficients are legitimate in thermochemical equations precisely because of this rule: they are how an equation gets rewritten per mole of one chosen species.

Heat for Any Amount of Any Species

Any heat question is answered by anchoring to the equation as written, then scaling. First find the price per mole of the species named: divide ΔH by that species' coefficient. Then multiply by the number of moles in the question.

For ammonia synthesis,

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, $\Delta H = -92$ kJ

Per mole of NH_3 the price is $-92 \div 2 = -46$ kJ: each mole of ammonia formed releases 46 kJ. Per mole of N_2 consumed it is $-92 \div 1 = -92$ kJ.

The same division gives the ΔH of the equation rewritten per mole of product:

$½N_2(g) + {}^{3}⁄_{2}H_2(g) \rightarrow NH_3(g)$, $\Delta H = -46$ kJ

Standard Enthalpy of Reaction, ΔH°

The symbol ΔH° is the standard enthalpy of reaction: the ΔH when reactants and products are all in their standard states.

Standard conditions: a gas at a pressure of 1 atm (1 bar in the newer convention), a dissolved substance at 1 mol L^{-1}, and a temperature of usually 298 K. Each substance is in its standard state, its most stable form under those conditions.

One convention anchors the whole scale: the standard enthalpy of formation of an element in its standard state is defined as zero. $O_2(g)$, $N_2(g)$ and graphite carry $\Delta H°f = 0$, not because their enthalpy is nothing, but because every other value is measured from the elements as the reference. Topic 6.8 builds on those values.

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Worked Examples

Example 1: Per mole, then an amount. For $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -572$ kJ. Per mole of water formed: $-572 \div 2 = -286$ kJ mol^{-1}. Forming 3.0 mol of water therefore releases $3.0 \times 286 = 858$ kJ.

Example 2: The reverse reaction. Electrolysis decomposes water: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$. Reversal flips the sign: $\Delta H = +572$ kJ per 2 mol decomposed, or +286 kJ per mole. Decomposing 0.50 mol of water absorbs $0.50 \times 286 = 143$ kJ.

Example 3: Scaling to a reactant amount. For $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, $\Delta H = -92$ kJ, the equation prices 1 mol of N_2 at 92 kJ released. When 0.25 mol of N_2 reacts completely, the heat released is $0.25 \times 92 = 23$ kJ. Written for 6 mol of ammonia, the equation triples: $3N_2(g) + 9H_2(g) \rightarrow 6NH_3(g)$, $\Delta H = 3 \times (-92) = -276$ kJ.

Example 4: Two species, two prices. For $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$, $\Delta H = -566$ kJ. Per mole of CO burned: $-566 \div 2 = -283$ kJ. Per mole of CO_2 formed: also $-566 \div 2 = -283$ kJ, because the two coefficients are equal here. Forming 0.80 mol of CO_2 releases $0.80 \times 283 = 226$ kJ.

Traps That Cost Marks in Topic 6.6

"$\Delta H = -572$ kJ means 572 kJ per mole of H_2." It means 572 kJ for the equation as written, which burns 2 mol of H_2. Per mole of H_2 the release is 286 kJ. Divide by the coefficient before calling anything per mole.

"The reverse reaction has the same ΔH." Same size, opposite sign. Forgetting the flip turns an energy input into an energy output.

"Halving the coefficients leaves ΔH alone." ΔH is extensive. Half the material, half the enthalpy change; the ΔH must be scaled with the equation, every time.

"A catalyst changes ΔH." It changes the path and the speed, never the endpoints. Enthalpy is a state function, so ΔH is identical with or without a catalyst.

"Positive ΔH means heat is given out." Positive means absorbed: the system gains enthalpy, the surroundings cool, and the process is endothermic. Released heat carries the negative sign.

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"Standard temperature is 273 K, and standard pressure for a solution is 1 atm." Thermochemical tables use 298 K. The 1 atm (or 1 bar) condition belongs to gases; solutions are standard at 1 mol L^{-1}.

Key Relationships to Remember

- Enthalpy is defined as H = E + PV; at constant pressure the heat exchanged equals the enthalpy change, q_p = ΔH
- Enthalpy is a state function: ΔH = H(products) − H(reactants), independent of path, steps, speed and catalyst
- A thermochemical equation is a balanced equation with physical states and its ΔH; the ΔH prices exactly the amounts the coefficients name
- Negative ΔH is exothermic (heat released), positive ΔH is endothermic (heat absorbed), from the system's point of view
- Reversing a reaction keeps the magnitude of ΔH and flips its sign
- Scaling every coefficient by a factor scales ΔH by the same factor; enthalpy is extensive
- Heat for any amount: divide ΔH by the named species' coefficient to get its per-mole price, then multiply by the moles asked for
- ΔH° is measured with everything in standard states: gases at 1 atm (1 bar), solutions at 1 mol L^{-1}, usually 298 K; an element in its standard state has ΔH°f defined as zero

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