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Topic 6.3: Heat Transfer and Thermal Equilibrium

Revision Notes · AP Chemistry

UNIT 6 · THERMOCHEMISTRY

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Heat Flows From Hot to Cold

Heat is energy transferred between objects because their temperatures differ. The flow has one fixed direction: from higher temperature to lower temperature.

Temperature alone sets the direction. Mass does not, density does not, and stored internal energy does not. A bath at 40 °C holds far more internal energy than a nail at 600 °C, yet heat flows from the nail into the water.

At particle level, fast particles in the hot object collide with slower ones in the cold object and hand kinetic energy across.

Heat Is Not Temperature

Heat and temperature are different quantities. Heat is energy in transit, measured in joules: it exists only while energy is crossing from one object to another. Temperature measures the average kinetic energy of the particles, in degrees Celsius or kelvin.

A firework spark sits above 1,000 °C but carries almost no energy, because so few particles are hot; a bath of water at 40 °C stores vastly more internal energy. Temperature says how hot; heat says how much energy moved.

Thermal Equilibrium

Two objects in contact exchange energy until their temperatures are equal. That state is thermal equilibrium: temperatures match, so there is no net flow of heat.

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Net zero does not mean the particles stop exchanging energy. Collisions still pass energy both ways at the boundary, but the flows are equal, so neither temperature changes. A thermometer works by this rule: it exchanges heat with a sample until it shares the sample's temperature. Equilibrium is a temperature condition, not an energy condition.

The Heat Equation: q = mcΔT

While a substance stays in a single phase, the heat it absorbs or releases is

$q = mc\Delta T$

m is the mass in grams, c is the specific heat capacity in J g^{-1} $°C^{-1}$, and ΔT = T(final) − T(initial). Temperature intervals have the same size in degrees Celsius and in kelvin, so either scale works inside ΔT.

Signs are taken from the system's point of view, as in Topic 6.1. Warming gives positive ΔT and positive q: heat absorbed. Cooling gives negative ΔT and negative q: heat released.

When 200 g of water cools from 60 °C to 25 °C, ΔT = 25 − 60 = −35 °C, and q = 200 × 4.18 × (−35) = −29,260 J: the system released 29.3 kJ.

The equation stops at a phase boundary: melting and boiling transfer heat at constant temperature, where ΔT is zero but q is not; that energy is priced in Topic 6.5.

Specific Heat Capacity: Why Water Is Stubborn

Specific heat capacity, c, is the heat needed to raise one gram of a substance by one degree Celsius, in J g^{-1} $°C^{-1}$: a property of the substance, not of the sample.

Liquid water's value, 4.18 J g^{-1} $°C^{-1}$, is unusually high; common metals sit far lower: aluminium 0.897, iron 0.450, copper 0.385 J g^{-1} $°C^{-1}$. Energy fed into water is shared among the strong attractions between its molecules, so its temperature creeps; fed into copper, the temperature jumps.

Two consequences drive most questions. For fixed q and m, ΔT is inversely proportional to c: halve c and ΔT doubles. And mc, the sample's own heat capacity, states the heat that sample needs per degree; mc, not mass alone, decides how a mixture settles.

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Mixing: Heat Lost Equals Heat Gained

Place hot and cold samples in an insulated container and they trade heat until one temperature serves both. No heat escapes, so the trade balances exactly: heat lost by the hot sample equals heat gained by the cold sample.

$m_1c_1(T_1 - T(\text{final})) = m_2c_2(T(\text{final}) - T_2)$

The final temperature is pulled toward the start of the sample with the larger mc product, because it needs more heat per degree of movement. Only when the two mc products are equal does it sit halfway; equal masses of one substance are that case, landing on the average of the two temperatures.

How the Transfer Happens

Conduction transfers heat through direct contact: particles collide and pass kinetic energy along, with no bulk movement of material. It is the route across the boundary between any two touching objects in this topic. Convection carries energy with a moving fluid, and radiation transfers it as electromagnetic waves, needing no contact at all.

Worked Examples

Example 1: Warming water. 50 g of water warms from 20 °C to 30 °C. ΔT = 30 − 20 = +10 °C, so q = mcΔT = 50 × 4.18 × 10 = +2,090 J: the water absorbed 2.09 kJ, and the positive sign says so.

Example 2: Finding an unknown specific heat. A 25 g metal sample absorbs 160 J and warms from 20 °C to 36 °C, so ΔT = +16 °C. Rearranged, c = q ÷ (mΔT) = 160 ÷ (25 × 16) = 0.40 J g^{-1} $°C^{-1}$: down among the metals, far below water's 4.18.

Example 3: The same heat into water and into iron. 418 J is fed into 20 g of water and, separately, into 20 g of iron. Water: ΔT = 418 ÷ (20 × 4.18) = 5.0 °C. Iron: ΔT = 418 ÷ (20 × 0.450) = 46.4 °C. Equal masses and equal heat, yet the iron warms over nine times as much, because its c is about a ninth of water's.

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Example 4: A hot block of iron in water. 100 g of iron at 90 °C is placed in 100 g of water at 20 °C, in an insulated cup. The mc products are 45.0 J $°C^{-1}$ (iron: 100 × 0.450) and 418 J $°C^{-1}$ (water: 100 × 4.18). Heat lost = heat gained: 45.0 × (90 − T(final)) = 418 × (T(final) − 20), giving T(final) = 26.8 °C, just 6.8 °C above the water's start. The heat exchanged is 45.0 × 63.2 = 2,844 J each way.

Traps That Cost Marks in Topic 6.3

"Heat flows from the object with the most heat." Objects do not contain heat; they contain internal energy, and heat is the transfer itself. Direction is set by temperature alone, so heat leaves the small hot nail for the large warm bath.

"At equilibrium the two objects hold the same internal energy." Equilibrium means equal temperatures and zero net flow. A kilogram and a gram of water at one temperature are in equilibrium, yet the kilogram holds a thousand times the internal energy.

"ΔT is the final temperature." ΔT is final minus initial. Using the final temperature in q = mcΔT breaks cooling calculations, where ΔT and q must both come out negative.

"A bigger rise means more heat was absorbed." ΔT depends on m and c as well as q. The 385 J that raises 50 g of copper by 20.0 °C raises 50 g of water by only 1.84 °C: same heat, different rises.

"The final temperature is the average of the two starts." Averaging works only when the two mc products are equal, as with equal masses of one substance. The average of 90 and 20 is 55 °C; Example 4's equilibrium is 26.8 °C.

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Key Relationships to Remember

- Heat flows spontaneously from higher to lower temperature until the two are equal; mass, density and internal energy never set the direction
- Heat (q) is energy in transit, in joules; temperature measures average particle kinetic energy, in °C or K
- Thermal equilibrium: equal temperatures, zero net flow, with energy still exchanged equally both ways at particle level
- $q = mc\Delta T$, with $\Delta T = T(\text{final}) - T(\text{initial})$, holds within a single phase; q is positive when the system absorbs heat, negative when it releases it
- c is the heat needed to raise 1 g by 1 °C, in J g^{-1} $°C^{-1}$: water 4.18, aluminium 0.897, iron 0.450, copper 0.385
- For fixed q and m, ΔT is inversely proportional to c; for fixed m and c, doubling ΔT doubles q
- In an insulated mixture, heat lost by the hot sample equals heat gained by the cold; the final temperature sits closest to the start of the sample with the larger mc
- Conduction transfers heat by direct contact, convection by a moving fluid, radiation by waves with no contact needed

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