The Simple AP® Physics 2 Unit 9 Cheat Sheet

9.1 Kinetic Theory: Pressure

Gas Model

Ideal gas: many tiny particles in random motion; no forces except during collisions; collisions are elastic (no KE lost); particle volume ≈ 0.

Where Pressure Comes From

Pressure (P = F/A): avg force per area from molecules hitting the container wall.
One collision: molecule bounces back → $\Delta p = 2mv_x$. Wall pushes it; it pushes wall (N3L).

More hits/s or harder hits → bigger force → P ↑.

Change	Hits/s	P
T ↑ (V const)	↑ & harder	↑
V ↓ (T const)	↑	↑
More gas (N ↑)	↑	↑

Micro ↔ Macro

Microscopic: speed, mass & KE of single molecules. **Macroscopic**: P, V, T of the whole gas (averages over a huge number N).
Sealed can heated: V fixed, T ↑ → faster, harder hits → P ↑ → the sealed can may burst.
Units: 1 Pa = 1 N/m². 1 atm ≈ 1.0×10^5 Pa.
Pressure ≠ force: the same force spread over a smaller area → a bigger pressure P on it.
Real gas ≈ ideal at low P, high T.

⚠ Only molecules hitting the WALL push on it. Molecules colliding w/ each other don't change the pressure on the wall at all, ever.

✓ FRQ: when T ↑, P ↑ for 2 reasons: more frequent hits AND bigger Δp per hit.

9.1 Temperature & Molecular Speeds

Temperature = Avg KE

Temperature: measures the avg translational KE per molecule. $K_{avg} = (3/2)k_BT$ (T in K).

Kelvin only: T(K) = T(°C) + 273. $K_{avg} \propto T$ in kelvin.

Speed Distribution

Maxwell-Boltzmann curve: molecules have a RANGE of speeds; peak + long tail at high speed.
T ↑ → peak moves right & spreads out; area under the curve (# molecules) stays the same.
rms speed: $v_{rms} = \sqrt{(3k_BT/m)}$. 4× T → 2× speed.

Same T	Avg KE	Speed
He (light)	same	faster
O_2 (heavy)	same	slower

Reading Graphs

Same gas, 2 curves: the one shifted right is hotter.
Ex: 300 K → 1200 K: K_{avg} 4×, v_{rms} 2×.

Single molecule: its speed is random; T describes the average only, not each molecule.
Mixed gases: collisions swap energy until K_{avg} is equal for both gases (same T reached).
Ideal gas energy: all KE (no PE b/w molecules) → internal energy U is set by T alone.
0 K (absolute zero): lowest possible T; kelvin temperatures in gas laws are never negative.

⚠ Same T means same AVG KE, not same speed. Heavier molecules move slower.

✓ Use ratios: $v_2/v_1 = \sqrt{(T_2/T_1)}$. Constants cancel.

9.2 The Ideal Gas Law

Two Forms

PV = nRT (n = moles, R = 8.31 J/mol·K) or **PV = Nk_BT** (N = total # of gas molecules).

Units: P in Pa, V in m³ (1 L = 10^{-3} m³), T in K.

Fixed Amount of Gas

Closed container (n const): $P_1V_1/T_1 = P_2V_2/T_2$.
Cross out whatever stays constant.

Held const	Relation
T (isothermal)	P ∝ 1/V
P (isobaric)	V ∝ T
V (isochoric)	P ∝ T

Ex: V halves, T doubles → P = 2 × 2 = 4× original.

Reading PV Graphs

PV ∝ T: bigger P × V product → hotter gas.
Isotherm (same-T curve): P = const/V; a higher curve = a higher gas temperature.
Lab: plot P vs 1/V (const T) → straight line through the origin, slope = nRT → solve for n or T.
Ex: 2.0 mol, 300 K, 0.050 m³ → P = (2)(8.31)(300)/0.050 ≈ 1.0×10^5 Pa (≈ 1 atm).
Density: hot gas is less dense at the same P, so a hot-air balloon rises in cooler air.
Find T at a point: T = PV/(nR), in kelvin.
Link constants: R = N_Ak_B (per mole).

⚠ Never use °C in PV = nRT. Convert every temperature to kelvin first (add 273).

✓ If gas leaks in or out, n changes: use PV = nRT for each state separately, not the ratio.

9.2–9.4 PV Diagrams & Work

Work = Area

Work on gas: W = −PΔV at const P; in general, W = −(area under the PV curve) for any path.
Expansion (V ↑): gas does work → W on gas < 0.
Compression (V ↓): W on gas > 0.
No volume change → W = 0, any P.

Four Processes

Process	Const	Key fact
Isobaric	P	W = −PΔV
Isochoric	V	W = 0
Isothermal	T	ΔU = 0
Adiabatic	Q = 0	ΔU = W

Cycles

Closed cycle: ends where it began → ΔU = 0.
Net work = area inside the loop. **Clockwise** → gas does net work (heat engine cycle).
Ex: P = 3×10^5 Pa const, V 0.02 → 0.05 m³ → W = −PΔV = −9000 J (the gas expands).
Counterclockwise loop → net work is done ON the gas by its surroundings (refrigerator).
Path matters: W & Q depend on the path; ΔU depends only on the start & end states.
Adiabatic curve is steeper than an isotherm.
Free expansion (into a vacuum): W = 0, Q = 0 → ΔU = 0 → T stays constant for an ideal gas.

⚠ On the AP exam, W = work done ON the gas. Gas expands → W is negative.

✓ Curved path? Count grid squares under the curve × the area of 1 square (P × V).

9.3 Thermal Energy Transfer & Equilibrium

Heat vs Temperature

Heat (Q): energy moving b/c of a temp difference. Objects store **internal energy**, not heat.
Direction: on its own, energy always flows hot → cold until the two temps match.
Why: fast molecules hit slow ones at the boundary and pass along KE to the cold side.

Three Ways

Mode	How
Conduction	collisions (touch)
Convection	fluid moves
Radiation	EM waves

Radiation works through a vacuum (e.g. sunlight).

Thermal Equilibrium

Equilibrium: same T → no NET energy transfer.
Isolated system: energy lost by the hot object = energy gained by the cold object (ΣQ = 0).
Final T lands between the 2 starting temps.
Convection: warm fluid is less dense → rises; cool fluid sinks → a circulating current forms.
Rate ∝ ΔT: bigger temp gap → faster transfer; it slows down as the temps get close.
Ex: 1 kg water at 80°C + 1 kg at 20°C → T_f = 50°C.

⚠ Metal feels colder than wood at the same T b/c it pulls energy from your hand faster.

✓ Energy bar charts: energy leaving the hot object = energy entering the cold one.

9.4 First Law of Thermodynamics

The Law

ΔU = Q + W (energy conservation). Q = heat INTO the gas; W = work done ON the gas.
Internal energy (monatomic ideal gas): U = (3/2)nRT. ΔU depends only on ΔT.

Sign	Q	W	ΔU
+	heat in	compressed	T ↑
−	heat out	expands	T ↓

Solving Steps

1) What's const? **2)** W from area + sign. **3)** ΔU from ΔT. **4)** Q = ΔU − W. Check signs.
Ex: gas does 2000 J of work (W = −2000 J), ΔU = +3000 J → Q = 3000 + 2000 = +5000 J.

Shortcuts

Isothermal: ΔU = 0 → Q = −W. **Isochoric**: W = 0 → ΔU = Q. **Isobaric**: W = −PΔV (box area).
Adiabatic (no heat flow): ΔU = W. Expanding gas cools; compressed gas heats up.
Isobaric expansion: T ↑ AND gas does work → Q > ΔU (the heat pays for both jobs).
Fast process (no time for heat to flow) ≈ adiabatic; slow, in a water bath ≈ isothermal.
Same ΔT, different path: same ΔU, but Q & W split differently between the 2 paths.

⚠ Adding heat doesn't always raise T. If the gas does equal work, ΔU = 0 (isothermal).

✓ Write a +/− sign table for Q, W & ΔU at every step before doing any of the math.

9.5 Specific Heat & Thermal Conductivity

Specific Heat

Q = mcΔT: c (specific heat) = energy needed to warm 1 kg of the material by 1 K (J/kg·K).
High c (e.g., water): slow to heat up and slow to cool down; moderates coastal climates.
Mixing: heat lost by the hot object = heat gained by the cold one. Then solve for T_f.

Ex: 0.5 kg Al (c ≈ 900 J/kg·K) warmed by 20 K → Q = 0.5 × 900 × 20 = 9000 J of heat.

Conduction Rate

Q/Δt = kAΔT/L: k = thermal conductivity, A = area, L = thickness of the slab or wall.

Change	Rate
Area 2×	2×
Thickness 2×	½×
ΔT 2×	2×

Insulators (low k): air, foam, wood. **Conductors** (high k): metals like copper & iron.

Graphs

T vs Q graph: slope = 1/(mc); steeper → smaller mc. A flat part = a phase change (T const).
Phase change: T stays constant while the added Q breaks bonds (melting, boiling).
Rate vs total: Q/Δt is power (in watts, J/s). Total Q = rate × time (J = W × s), e.g., 100 W × 60 s.

⚠ c is per kg. The same Q into twice the mass → half the ΔT (for the same material).

✓ Lab: plot ΔT vs Q; use the slope to find c.

9.6 Entropy & Second Law

Second Law

Second law: total entropy of an isolated system never decreases ($\Delta S_{total} \geq 0$), ever.

Entropy (S): how spread out energy is; more possible arrangements → higher S.
Spontaneous: heat flows hot → cold b/c that raises the total entropy S; never the reverse.

Entropy Changes

Event	S of system
Gas expands	↑
Heat added	↑
Melting / boiling	↑

Local decrease OK: a fridge lowers S inside but raises the room's S by even more.

Engines & FRQ Moves

Heat engine: takes in heat, does work, dumps the rest as heat. Efficiency < 100%.
Claim → evidence → reasoning: name the law, cite the data, link molecules to P & T clearly.
Irreversible: friction, mixing & free expansion all raise total S; they never undo alone.
Microstates: all molecules crowded on one side is far less likely than spread out evenly.

⚠ Only TOTAL entropy can't drop. One part of a system can lose entropy if another gains.

✓ PV graph, bar chart & equation must all agree on the signs of Q, W & ΔU every time.