The Ultimate AP® Physics 2 Unit 9 Cheat Sheet

9.1 Kinetic Theory: Pressure

Gas Model Assumptions

Ideal gas: many tiny particles in random motion, no forces except during collisions; collisions are elastic (KE conserved); particle volume ≈ 0 vs. the container.
Real gas ≈ ideal at low P, high T (particles far apart, fast).

Where Pressure Comes From

Pressure: avg force per unit area from molecules hitting the wall. P = F/A, measured in pascals (1 Pa = 1 N/m²).
One collision: $\Delta p = 2mv_x$ (elastic bounce reverses v_x). Wall pushes molecule; molecule pushes wall (N3L).
Force on wall: $F_{avg} = \Delta p/\Delta t$ summed over all hits. More hits per second or harder hits → P ↑. Fewer, softer → P ↓.

Change	Hits/sec	Δp per hit	P
T ↑ (V const)	↑	↑	↑
V ↓ (T const)	↑	same	↑
N ↑ (V,T const)	↑	same	↑
m ↑ (same T)	↓	↑	same

Micro ↔ Macro

Microscopic: v, m, KE of single molecules. **Macroscopic**: P, V, T of whole gas = averages over huge N.
Pressure ≠ force: same F on smaller A → bigger P.
Units: 1 Pa = 1 N/m². 1 atm ≈ 1.0×10^5 Pa. Gauge P = absolute P − atmospheric P (a tire gauge reads gauge P).
Kinetic link: $P = (1/3)(N/V)m\,v_{rms}^2 \rightarrow P \propto (N/V)\cdot K_{avg}$.
Sealed can heated: V fixed, T ↑ → P ↑ → can may burst.
Piston at rest: FBD → $P_{gas}A = P_{atm}A + mg$.
Same P on all walls: random motion, no preferred direction.

⚠ Molecules hitting each other don't change wall pressure; only wall collisions push on the wall.

✓ FRQ: explain P ↑ with BOTH factors: more frequent collisions AND bigger Δp per collision (when T ↑).

9.1 Temperature & Molecular Speeds

Temperature = Avg KE

Temperature: a measure of the average translational KE per molecule. $K_{avg} = (3/2)k_BT$, w/ T in kelvin.
Kelvin only: T(K) = T(°C) + 273. Doubling °C ≠ doubling KE.
$k_B = 1.38\times10^{-23}$ J/K (on equation sheet). $K_{avg} \propto T$, not √T.

Speed Distribution

Maxwell-Boltzmann: molecules have a RANGE of speeds; the curve peaks, then a long tail to high v.
T ↑ → peak shifts right & flattens; area (total N) stays same.
$v_{rms} = \sqrt{(3k_BT/m)}$: ∝ √T, ∝ 1/√m. 4× T → 2× v_{rms}.

Same T: gas	Avg KE	rms speed
He (light)	same	faster
O_2 (heavier)	same	slower
Mix at equil.	equal	lighter faster

Interpreting Graphs

Histogram of speeds: a wider spread + a peak at higher v → a higher T (for the same gas, same N).
Two gases, same curve: if they're at the same T, the curve shifted right belongs to the lighter gas (smaller m).
Single molecule: its speed is random & changes every collision; T describes the average, not each molecule.
Ex: 300 K → 1200 K: K_{avg} 4×, v_{rms} 2×.

Ideal gas energy: only KE (no PE b/w molecules) → internal energy U is set by T alone; ΔT = 0 → ΔU = 0.
Mixed gases: collisions swap energy until K_{avg} equal.

0 K: lowest possible T; gas-law T never negative.

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

✓ Ratio trick for rms speed: $v_2/v_1 = \sqrt{(T_2/T_1)}$. Cancel constants; never plug in Boltzmann's constant.

9.2 The Ideal Gas Law

Two Forms

PV = nRT (n = moles, R = 8.31 J/mol·K) or **$PV = Nk_BT$** (N = # of molecules). Link: $N = nN_A$ (Avogadro's #).

Units: P in Pa, V in m³ (1 L = 10^{-3} m³), T in K. Wrong units are the top error on gas-law problems; convert first.

Ratio Method (Fixed Gas)

Closed container (n const): $P_1V_1/T_1 = P_2V_2/T_2$. Cross out whatever stays constant, then solve for the unknown.

Held const	Law	Relation
T (isotherm)	Boyle	P ∝ 1/V
P (isobar)	Charles	V ∝ T
V (isochor)	Gay-Lussac	P ∝ T
P, T	Avogadro	V ∝ n

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

Reading PV Graphs

PV ∝ T: a point farther from the origin (bigger PV) → hotter gas. Same PV product → same T (same isotherm).
Isotherm: hyperbola P = const/V; higher curve = higher T.
Find T at a point: T = PV/(nR). Compare T at 2 points w/ $T_2/T_1 = P_2V_2/P_1V_1$ (n fixed); no R needed for a ratio.

Lab: linearize. Plot P vs 1/V (const T) → slope = nRT. Plot P vs T (const V) → slope = nR/V. Use the best-fit line.
Ex: 2.0 mol, 300 K, 0.050 m³ → P = (2)(8.31)(300)/0.050 ≈ 1.0×10^5 Pa, about 1 atmosphere of pressure.
Density: n/V = P/RT → hot gas is less dense at the same P, so a hot-air balloon rises in the cooler air around it.
Link constants: $R = N_Ak_B$ (per mole vs. per molecule).

⚠ Celsius in PV = nRT gives nonsense (0°C ≠ zero energy). Always convert to kelvin first (+273).

✓ Gas leaks out? n changes, so use PV = nRT at each state separately, not the combined ratio law.

9.2–9.4 PV Diagrams & Work

Work = Area

Work on gas: W = −PΔV (const P). In general: W = −(area under the PV curve) for any path taken.
Expansion (V ↑): gas does work, so W on gas < 0.
Compression (V ↓): W on the gas > 0.
No ΔV → W = 0, even if P changes hugely.

Four Processes

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

Adiabatic: steeper than isotherm on PV graph. Fast compression → T ↑ (e.g. bike pump warms).
Isothermal expansion: all Q in → work out (Q = −W).

Cycles

Closed cycle: back to start → ΔU = 0, so $Q_{net} = -W_{net}$.

Net work = enclosed area. **Clockwise** → gas does net work (heat engine). **Counterclockwise** → net work is done on the gas (refrigerator, heat pump). Bigger loop → more W.
Path matters: W & Q depend on the path; ΔU depends only on the endpoints (a state function like T).
Ex: isobaric at 3×10^5 Pa, V 0.02 → 0.05 m³ → W = −9000 J.
Compare paths: same endpoints, the higher-P path → more area under it → more work done by the gas.
Free expansion (into a vacuum): W = 0, Q = 0 → ΔU = 0, so T stays constant (for an ideal gas only).
Isotherm vs adiabat: same start & same final V, the adiabatic expansion ends at lower P & lower T (gas cools).

⚠ AP sign: W = work done ON gas. Gas expanding → W negative. Check the equation sheet's convention.

✓ Curved path: count grid squares under the curve × (P per square × V per square) = |W|; then add the sign.

9.3 Thermal Energy Transfer & Equilibrium

Heat vs Temperature

Heat (Q): energy transferred b/c of a ΔT. It's not stored; objects store **internal energy** instead (U).
Direction: spontaneously hot → cold only, until T equal.
Microscopic: fast molecules of hot obj collide w/ slow ones at boundary → KE passes to cold side.

Three Mechanisms

Mode	How	Example
Conduction	collisions	metal spoon
Convection	fluid moves	boiling pot
Radiation	EM waves	sunlight

Radiation: needs no medium; works through a vacuum. A hotter object emits more & at shorter λ (e.g., Sun).
Convection: warm fluid is less dense → rises; cool fluid sinks. Needs a fluid + gravity to make a current.

Thermal Equilibrium

Equilibrium: same T → no net energy transfer (molecules still collide, but the exchange balances both ways).
Zeroth law: A in equilibrium w/ C & B in equilibrium w/ C → A in equilibrium w/ B. Why thermometers work.
Isolated system: $Q_{lost,hot} = Q_{gained,cold}$. ΣQ = 0.

Final T always lands between the 2 starting temps, closer to the object w/ the bigger mc (more "thermal mass").
Rate ∝ ΔT: big temp gap → fast transfer; the rate slows as the two temps converge toward equilibrium.
Ex: 1 kg water at 80°C + 1 kg at 20°C → T_f = 50°C.

Define system first; energy out of one = energy into other.

⚠ Metal feels colder than wood at the same T b/c it conducts energy from your hand faster, not b/c it's colder.

✓ Energy bar charts: show Q leaving hot system and entering cold one with equal size bars.

9.4 First Law of Thermodynamics

The Law

ΔU = Q + W: energy conservation for a gas. Q = heat INTO the gas, W = work done ON the gas by surroundings.
Internal energy (ideal monatomic): U = (3/2)nRT = $(3/2)Nk_BT$. ΔU ∝ ΔT only, not the path taken.

Sign	Q	W	ΔU
+	heat in	compressed	T ↑
−	heat out	expands	T ↓
0	insulated	V const	T const

Solving Workflow

1) Name the process (what's const?). **2)** Get W from graph area & sign. **3)** Get ΔU from ΔT (or ΔU = (3/2)Δ(PV)). **4)** Q = ΔU − W. **5)** Check every sign against the graph.

Ex: gas expands at 2×10^5 Pa by 0.01 m³ → W = −2000 J. ΔU = +3000 J → Q = 3000 − (−2000) = +5000 J.

Quick Reasoning

Isothermal: ΔU = 0 → Q = −W. **Isochoric**: W = 0 → ΔU = Q.
Adiabatic: Q = 0 → ΔU = W. Expansion cools the gas; compression heats it (e.g., a bike pump warms).
Isobaric expansion: T ↑ AND gas does work → Q > ΔU.
Same ΔT, diff path: ΔU same; Q & W split differently.
Fast process (no time for heat to flow) ≈ adiabatic; slow, in a big water bath ≈ isothermal (T held by the bath).
Bar chart: $U_i + Q + W = U_f$; bars must balance.

Monatomic: AP uses U = (3/2)nRT unless told otherwise.
Ex: rigid tank gets 600 J of heat → W = 0 (no ΔV) → ΔU = +600 J, so the gas's T rises; all Q stays as U.
Expands w/ no heat (adiabatic): U pays for the work → T ↓.

⚠ Q ≠ ΔU. Adding heat need not raise T if the gas does an equal amount of work (isothermal process).

✓ Make a Q / W / ΔU sign table per step before any math. FRQs award points for signs + justification.

9.5 Specific Heat & Thermal Conductivity

Specific Heat

$Q = mc\Delta T$: c = specific heat (J/kg·K) = energy needed to raise 1 kg of a material by 1 K. Water's c is very high.
High c (e.g., water, 4186 J/kg·K) → slow to heat & slow to cool; that's why it moderates coastal climates.
Mixing: $m_1c_1(T_f - T_1) + m_2c_2(T_f - T_2) = 0$. Solve for T_f.

Phase change: T stays constant while Q breaks bonds; Q = mcΔT doesn't apply during melting or boiling.

Conduction Rate

$Q/\Delta t = kA\Delta T/L$: k = thermal conductivity, A = cross-section area, L = thickness (length of the heat path).

Change	Rate Q/Δt
A doubled	2×
L doubled (thicker)	½×
ΔT doubled	2×
Higher k (metal)	↑

Insulators: low k (air, foam, wood, fiberglass). **Conductors**: high k (metals like Cu, Al). Trapped air = good insulator.
Rate vs total: Q/Δt = power (W). Total Q = rate × time.

Graph Skills

T vs Q graph: slope = 1/(mc). Steeper → smaller mc. A flat part = a phase change (T const while Q goes in).
Ex: 0.5 kg Al (c ≈ 900 J/kg·K) warmed by 20 K → Q = mcΔT = 0.5 · 900 · 20 = 9000 J of heat added.
Steady state: T at each point const; energy in = energy out.
Layers: in steady state, the same rate flows through each layer; biggest ΔT across the worst conductor.
Equal mass, same Q: lower c → bigger ΔT (sand heats up faster than water at the beach on a sunny day).

⚠ c is per kg. Same Q into 2× mass of same stuff → ½ ΔT.

✓ Lab FRQ: plot ΔT vs Q (or vs heater time); slope → c. State what's held constant & why it matters.

9.6 Entropy & Second Law

Second Law

Second law: the total entropy of an isolated system never decreases; $\Delta S_{total} \geq 0$ for every real process.

Entropy (S): measure of # of microscopic arrangements; more spread-out energy → higher entropy S.
Spontaneous: heat flows hot → cold b/c that ↑ S_{total}. The reverse never happens on its own (no work).

Entropy Changes

Event	S of system
Gas expands freely	↑
Heat added (T ↑)	↑
Melting / boiling	↑
Gas compressed, cooled	↓ (env ↑ more)

Local ↓ allowed: a fridge lowers the inside S but dumps more S into the room; total S still ↑ (needs work input).
Irreversible: friction, mixing & free expansion. Real processes always ↑ S_{total}; none run backward on their own.

Engines & FRQ Moves

Heat engine: takes in Q_H from a hot reservoir, does W, dumps Q_C to a cold one. e = W/Q_H < 1 always.

Claim → evidence → reasoning: name the law, cite graph/data, link micro (collisions) to macro (P, T).
Translate reps: PV graph ↔ energy bar chart ↔ equation; all three must agree on every sign (Q, W, ΔU).
Reversible (ideal limit): $\Delta S_{total} = 0$; never fully real.

Microstates: all molecules crowded on one side of a box is far less likely than spread out evenly in it.

⚠ "Entropy always increases" is wrong for a part of a system. Only the TOTAL (isolated) can't decrease.

✓ No engine is 100% efficient: some Q must always flow out to the cold reservoir ($Q_C > 0$), even ideally.