

Unit 9: Thermodynamics

Unit 9 covers the thermal behavior of matter. It starts with the kinetic theory of gases, which explains pressure and temperature as the result of molecular motion, then builds the ideal gas law, the first law of thermodynamics with the AP sign convention, PV diagrams and the four standard processes, heat engines and efficiency, the second law and entropy, and the three modes of heat transfer.

AP Physics 2

Thermodynamics

About 15 minutes to read

How to use this guide

Read it in order the first time because the topics build on each other. Kinetic theory gives physical meaning to temperature, and the ideal gas law and the first law both depend on it. PV diagrams turn the first law into geometry, and heat engines combine the first and second laws. Exam questions usually describe a scenario, such as a gas expanding in a piston or a cycle on a PV diagram, and ask you to assign signs to Q , W , and ΔU .

After the first read, use the trap boxes and the tables to review the distinctions the exam tests most often. The sign convention table and the process table are the two highest-value references in the guide. Finish with the practice questions, then complete the recall check on the last page out loud and note any items you cannot explain yet.

What this unit is worth. Thermodynamics is about 15 to 18 percent of the AP Physics 2 exam. On top of that, the sign discipline this unit demands carries into every problem that mixes heat and work, so a shaky convention here costs points across the whole exam.

9.1 Kinetic Theory: Temperature as Molecular Motion

The **kinetic theory of gases** explains macroscopic behavior from microscopic motion. A gas is a huge number of molecules moving randomly and colliding elastically with each other and with the walls of the container. Each collision with a wall transfers momentum to the wall, and **pressure** is the result of countless such collisions: the average force per unit area from molecules striking the walls. More molecules, faster molecules, or more frequent collisions all raise the pressure.

Temperature, in this picture, measures the average translational kinetic energy of one molecule: $K_{\text{avg}} = (3/2) k_B T$, where $k_B = 1.38 \times 10^{-23} \text{ J/K}$ is the Boltzmann constant and T must be in kelvin. This relation defines absolute temperature for an ideal gas. Doubling the kelvin temperature doubles the average molecular kinetic energy. It does not double each molecule's speed, because speed grows as the square root of energy.

Worked example: at room temperature, $T = 300 \text{ K}$, the average translational kinetic energy per molecule is $K_{\text{avg}} = (3/2)(1.38 \times 10^{-23})(300)$. First multiply $(1.38 \times 10^{-23})(300) = 4.14 \times 10^{-21}$, then multiply by $3/2$ to get $6.21 \times 10^{-21} \text{ J}$. A tiny number per molecule, but a mole of gas holds about 6.02×10^{23} of them, which is how microscopic motion becomes macroscopic thermal energy.

Molecular speeds

The **root-mean-square speed** follows from the same relation: $v_{\text{rms}} = \sqrt{(3k_B T / m)} = \sqrt{(3RT / M)}$, where m is the mass of one molecule, M is the molar mass in kg/mol, and $R = 8.31 \text{ J/(mol}\cdot\text{K)}$. For nitrogen gas, N_2 , $M = 0.028 \text{ kg/mol}$. At 300 K , $v_{\text{rms}} = \sqrt{(3 \times 8.31 \times 300 / 0.028)} = \sqrt{(267,107)} \approx 517 \text{ m/s}$. Heavier molecules move more slowly at the same temperature and lighter ones move faster, which is why the same formula gives a much higher speed for hydrogen. Note that v_{rms} is not the plain average speed; it is the square root of the average of the squared speeds, and it is the speed that connects directly to kinetic energy.

Trap. Temperature measures average kinetic energy *per molecule*, not total thermal energy. A bathtub of lukewarm water holds far more thermal energy than a cup of boiling water, because thermal energy scales with the number of molecules. Questions that contrast a small hot object with a large warm one are testing exactly this.

Two related definitions worth keeping straight: **thermal energy** is the total internal kinetic and potential energy of all the particles in a sample, and **heat** is energy in transit because of a temperature difference. Two objects are in **thermal equilibrium** when they are at the same temperature and no net heat flows between them. Heat flows spontaneously from hot to cold, never the reverse without outside work.

9.2 The Ideal Gas Law

The **ideal gas law** relates the four state variables of a gas: $PV = nRT = Nk_B T$. Here P is pressure in pascals, V is volume in cubic meters, T is temperature in kelvin, n is the amount of gas in moles, and N is the number of molecules. The two forms are equivalent because $R = N_A k_B$, where $N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$ is Avogadro's number. An ideal gas is a model in which molecules are point particles with no forces between them except elastic collisions; real gases behave this way at ordinary temperatures and low pressures.

Worked example: a sealed rigid container holds gas at $P = 2.0 \times 10^5 \text{ Pa}$, $V = 0.010 \text{ m}^3$, and $T = 300 \text{ K}$. Solve for n : $n = PV / RT = (2.0 \times 10^5)(0.010) / (8.31 \times 300)$. The numerator is 2000 J and the denominator is 2493 J/mol, so $n = 2000 / 2493 \approx 0.80 \text{ mol}$. That is about $0.80 \times 6.02 \times 10^{23} \approx 4.8 \times 10^{23}$ molecules.

When the amount of gas is fixed, PV/T is constant, which gives the **combined gas law**: $P_1 V_1 / T_1 = P_2 V_2 / T_2$. This is the fastest tool for "before and after" problems. A rigid tank has constant V , so pressure scales directly with kelvin temperature. Heating gas from 27°C to 177°C means $T_1 = 300 \text{ K}$ and $T_2 = 450 \text{ K}$, so $P_2 = P_1 \times 450/300 = 1.5 P_1$. A 150 K rise from a 300 K starting point raises the pressure by 50 percent.

Trap. The gas law needs absolute temperature, so convert Celsius to kelvin (add 273) before any ratio or substitution. In the example above, using $177/27 \approx 6.6$ instead of $450/300 = 1.5$ is a classic wrong answer, and it appears as a distractor. Volume must be in m^3 and pressure in Pa if you want n in moles out of $PV = nRT$.

One more relation the first law will need: for a monatomic ideal gas, the internal energy is $U = (3/2)nRT$, so $\Delta U = (3/2)nR\Delta T$. Internal energy depends only on temperature. Whenever the temperature of an ideal gas does not change, $\Delta U = 0$, and the sign of ΔU always matches the sign of ΔT .

9.3 The First Law of Thermodynamics

The **first law of thermodynamics** is conservation of energy written for a thermodynamic system: $\Delta U = Q + W$. The change in the system's internal energy equals the heat added to it plus the work done on it.

The sign convention is the part that decides points, so fix it before touching any numbers. This is the convention on the AP Physics 2 equation sheet:

SYMBOL	POSITIVE WHEN	NEGATIVE WHEN
ΔU	Internal energy rises. For an ideal gas the temperature rises.	Internal energy falls. For an ideal gas the temperature falls.
Q	Heat flows into the system.	Heat flows out of the system.
W	Work is done on the system, as in compression.	The system expands and does work on its surroundings.

For a constant-pressure change, the work done **on** the system is $W = -P\Delta V$. The minus sign carries the convention: when the gas expands ($\Delta V > 0$), W is negative, because the system is doing the work rather than receiving it. When the gas is compressed ($\Delta V < 0$), W is positive.

Worked example: a gas absorbs 450 J of heat, so $Q = +450$ J. It expands and does 300 J of work on its surroundings, so the work done *on* the system is $W = -300$ J. Then $\Delta U = Q + W = 450 + (-300) = +150$ J. The internal energy rises by 150 J, so the gas gets hotter. The energy accounting reads plainly: 450 J came in as heat, 300 J left as work, and 150 J stayed behind as internal energy.

Trap. The number-one error in this unit is plugging in the work done *by* the gas where W belongs. On the AP equation sheet W means work done *on* the system, and $W = -P\Delta V$. If the gas expands, W is negative. Before you substitute, ask whether the gas expanded or was compressed, and set the sign of W from the answer. A problem that says "the gas does 300 J of work" hands you $W = -300$ J, not +300 J.

9.4 PV Diagrams and Thermodynamic Processes

A **PV diagram** plots pressure against volume. Each point on it is an equilibrium state of the gas, and each path between two points is a process. The magnitude of the work in a process equals the **area under the PV curve** between the initial and final volumes. In the AP convention, moving right (expansion) means the system does work on the surroundings, so the work done *on* the system is negative; moving left (compression) makes it positive. Over a complete cycle the gas returns to its starting point, so $\Delta U = 0$ and the net work equals the area enclosed by the loop.

Worked example: a gas expands at constant pressure $P = 3.0 \times 10^5$ Pa from $V = 0.020$ m³ to $V = 0.040$ m³. On the PV diagram this is a horizontal line, and the area under it is a rectangle: work done *by* the gas = $P\Delta V = (3.0 \times 10^5)(0.020) = 6000$ J. The work done *on* the system is therefore $W = -6000$ J. If 8000 J of heat flows into the gas during the expansion, the first law gives $\Delta U = 8000 + (-6000) = +2000$ J, so the gas warms up.

There are four standard processes. Each one holds something fixed, and that constraint zeroes out one term of the first law, which is what makes them fast to analyze.

PROCESS	HELD CONSTANT	Q	W (WORK ON SYSTEM)	ΔU
Isochoric	Volume ($\Delta V = 0$); vertical line	Equals ΔU	0, since $\Delta V = 0$	= Q. Heating at constant V raises T and P.
Isobaric	Pressure; horizontal line	$\Delta U - W$	$-P\Delta V$: negative in expansion, positive in compression	= Q + W
Isothermal	Temperature; hyperbola	= $-W$: positive in expansion (heat must enter), negative in compression (heat must leave)	Negative in expansion, positive in compression	0, since $\Delta T = 0$
Adiabatic	No heat flow (insulated or very fast)	0	= ΔU : negative in expansion, positive in compression	= W. Expansion cools the gas; compression heats it.

The isothermal and adiabatic rows deserve a closer look because the exam loves to pair them. On a PV diagram through the same starting point, the **adiabatic curve is steeper** than the isothermal curve. The reason is physical: in an isothermal expansion, heat flows in to hold the temperature fixed, so the pressure falls only because the volume grows. In an adiabatic expansion, no heat enters, so the work done comes out of the gas's own internal energy, the gas cools, and the pressure falls for two reasons at once. Expanding to the same final volume, the adiabatic path therefore does less work (smaller area) and ends at a lower temperature.

Trap. Isothermal and adiabatic are easy to swap under time pressure. Isothermal means constant temperature, so $\Delta U = 0$ and heat must flow to sustain the temperature. Adiabatic means no heat flow, so $Q = 0$ and the gas temperature changes: expansion cools it, compression heats it. If a question asks which process cools the gas on expansion, the answer is the adiabatic one, and its PV curve is the steeper of the two.

A useful habit for any PV-diagram question: first read the direction of the path, then assign W , then use the process constraint to find Q and ΔU . Rightward means expansion, so W (on the system) is negative. Leftward means compression, so W is positive. Upward at constant volume means heating with $W = 0$. Once two of the three quantities are signed, the first law gives the third.

Cycles. A clockwise cycle on a PV diagram is a heat engine: the expansion happens at higher pressure than the compression, so the net work done *by* the system is positive and equals the enclosed area. A counterclockwise cycle is a refrigerator or heat pump: net work is done *on* the system. Either way, $\Delta U = 0$ over the full loop.

9.5 Heat Engines and Efficiency

A **heat engine** runs a working substance through a repeating cycle. Each cycle it absorbs heat Q_h from a hot reservoir, converts part of that heat into net work W_{net} , and rejects the leftover heat Q_c to a cold reservoir. Because the cycle returns to its starting point, $\Delta U = 0$ over one cycle, and the first law gives W_{net} (done *by* the engine) = $Q_h - Q_c$.

Thermal efficiency measures what fraction of the absorbed heat becomes useful work: $e = W_{\text{net}} / Q_h = 1 - Q_c / Q_h$. Efficiency is always less than 1, because Q_c can never be zero. Some heat must always be dumped to the cold reservoir.

Worked example: an engine absorbs $Q_h = 4000 \text{ J}$ from the hot reservoir each cycle and rejects $Q_c = 3000 \text{ J}$ to the cold reservoir. The net work per cycle is $W_{\text{net}} = 4000 - 3000 = 1000 \text{ J}$. The efficiency is $e = 1000 / 4000 = 0.25$, or 25%. Check with the second form: $1 - 3000/4000 = 1 - 0.75 = 0.25$. The remaining 75 percent is not a mistake or a loss to fix; it is heat the engine is required to reject.

Trap. Efficiency divides by Q_h , the heat taken from the *hot* reservoir. The ratio $Q_c/Q_h = 3000/4000 = 0.75$ is the fraction of heat thrown away, not the efficiency. When a problem gives you two heat numbers, label each one Q_h or Q_c before you divide. The efficiency formula also tells you that W_{net}/Q_c and Q_h/Q_c are not efficiencies, even though they look similar.

9.6 The Second Law and Entropy

The **second law of thermodynamics** has two equivalent statements. Heat flows spontaneously from a hotter object to a colder one, never from cold to hot without work being done, which is why refrigerators need a power supply. And no engine operating in a cycle can convert heat from a single reservoir entirely into work; some heat must always be rejected, which is why efficiency stays below 100 percent.

Entropy is a measure of disorder, or more precisely the number of microscopic arrangements consistent with a macroscopic state. The second law can be restated in terms of it: the entropy of an **isolated system never decreases**. It increases in every real, irreversible process and stays constant only in an idealized reversible process. When heat flows from hot to cold, the cold object's entropy rises by more than the hot object's entropy falls, so the total rises.

Trap. The entropy of *one part* of a system can decrease. Water freezing in a freezer becomes more ordered and its entropy falls, but the freezer dumps a larger amount of entropy into the kitchen as waste heat, so the total entropy of the isolated whole still rises. Whenever a question asks about entropy, check whether it means one object or the entire isolated system, because the rule "never decreases" applies only to the whole.

9.7 Heat Transfer: Conduction, Convection, Radiation

Heat moves by three mechanisms. Exam questions usually describe a scenario and ask which mode dominates, so learn the telltale conditions for each.

MODE	NEEDS A MEDIUM?	HOW THE ENERGY MOVES	EXAMPLE
Conduction	Yes, matter in contact	Molecule-to-molecule collisions pass kinetic energy along a material. Metals conduct well because their free electrons carry energy quickly.	A metal spoon handle warming up in hot soup.
Convection	Yes, a fluid that can flow	Bulk motion of the fluid carries warm material into cooler regions and cool material back to be warmed.	Warm air rising above a heater and circulating around a room.
Radiation	No	Electromagnetic waves, mostly infrared, carry energy across empty space. Every object with a temperature above absolute zero radiates.	Sunlight warming your skin across 150 million kilometers of vacuum.

The quick sort: contact between solids or still fluids points to conduction, moving fluids point to convection, and energy crossing a vacuum or arriving as light points to radiation. A campfire warms you mostly by radiation; the air above it rises by convection; the metal poker handle heats your hand by conduction.

Confusions That Cost Points

PAIR	HOW TO KEEP THEM STRAIGHT
Work on vs work by	AP uses work done <i>on</i> the system: $W = -P\Delta V$. Expansion makes W negative, compression makes it positive. If a problem states the work done <i>by</i> the gas, flip the sign before using $\Delta U = Q + W$.
Isothermal vs adiabatic	Isothermal: T constant, so $\Delta U = 0$ and heat must flow to sustain the temperature. Adiabatic: $Q = 0$, so the gas cools on expansion and heats on compression. The adiabatic curve is the steeper one on a PV diagram.
Q_h vs Q_c in efficiency	Efficiency = W_{net}/Q_h . Q_h is the heat absorbed from the hot reservoir. Q_c/Q_h is the waste fraction, not the efficiency.
Celsius vs kelvin	The gas law and kinetic theory need absolute temperature. Add 273 before any ratio or substitution. A rise from 27 °C to 177 °C is 300 K to 450 K, a factor of 1.5, not 177/27.
Heat vs temperature vs thermal energy	Temperature is the average kinetic energy per molecule. Thermal energy is the total over all molecules. Heat is energy in transit because of a temperature difference.
Entropy of a part vs the whole	One object's entropy can fall, as in freezing water. The entropy of the entire isolated system never decreases.
Conduction vs convection vs radiation	Conduction needs contact, convection needs a flowing fluid, radiation needs nothing. Match the scenario to the condition, not to the word "heat."

Practice Questions

Original questions written for this guide in the style of the AP exam. Answers and explanations are on the next page, so complete the questions before checking them.

1. A gas absorbs 200 J of heat from its surroundings and expands, doing 150 J of work on its surroundings. What is the change in internal energy of the gas?

1. -350 J
2. -50 J
3. +50 J
4. +350 J

2. A gas is compressed at a constant pressure of 2.0×10^5 Pa from a volume of 0.030 m^3 to a volume of 0.010 m^3 . What is the work done *on* the gas?

1. -4000 J
2. 0 J
3. +8000 J
4. +4000 J

3. An ideal gas expands isothermally. Which of the following is true?

1. The internal energy of the gas increases.
2. The gas absorbs heat equal in magnitude to the work it does.
3. No heat is transferred between the gas and its surroundings.
4. The temperature of the gas rises.

4. A rigid sealed container holds a gas at 27°C . The gas is heated until its temperature reaches 177°C . By what factor does the pressure change?

1. 0.67
2. 1.5
3. 6.6
4. It does not change

5. An ideal gas expands from the same initial state to the same larger volume in two separate experiments: once isothermally and once adiabatically. Which statement is correct?

1. The adiabatic expansion does more work, because no heat escapes.
2. Both processes end at the same temperature.
3. Both processes do zero work because $\Delta U = 0$.
4. The isothermal expansion does more work, because its PV curve stays above the adiabatic curve.

6. A heat engine absorbs 4000 J of heat from a hot reservoir each cycle and rejects 3000 J of heat to a cold reservoir. What is the thermal efficiency of the engine?

1. 25%
2. 75%
3. 33%
4. 133%

7. A tray of ice melts in a warm room. Which statement about entropy is correct?

1. The total entropy stays the same because the heat lost by the room equals the heat gained by the ice.
2. The entropy of the room decreases by more than the entropy of the ice increases.
3. The total entropy of the ice and the room increases.
4. The entropy of the ice decreases as it melts.

8. One end of a metal rod is held in a flame. The other end gradually becomes too hot to hold, even though the metal itself does not move. The dominant mode of heat transfer through the rod is

1. conduction
2. convection
3. radiation
4. evaporation

Answer Key

- 1. C.** $Q = +200 \text{ J}$ (heat in) and $W = -150 \text{ J}$ (work done *by* the gas means work done *on* it is negative). $\Delta U = 200 + (-150) = +50 \text{ J}$. D adds the work with the wrong sign, the most common first-law error. A flips both signs. B flips the sign of Q , treating absorbed heat as negative.
- 2. D.** $W = -P\Delta V = -(2.0 \times 10^5)(0.010 - 0.030) = -(2.0 \times 10^5)(-0.020) = +4000 \text{ J}$. Compression makes ΔV negative, and the two negatives give positive work done on the gas. A reports the work done *by* the gas instead. B may come from confusing ΔV with zero or from skipping the calculation. C doubles the volume change or uses the final volume instead of the change.
- 3. B.** Isothermal means $\Delta T = 0$, so for an ideal gas $\Delta U = 0$. The first law then gives $0 = Q + W$, so $Q = -W$: the heat absorbed equals the magnitude of the (negative) work done on the system, which is the work done by the gas. A contradicts $\Delta U = 0$ for constant temperature. C describes an adiabatic process, not an isothermal one. D contradicts the definition of isothermal.
- 4. B.** At constant volume, P scales with kelvin temperature: $T_1 = 27 + 273 = 300 \text{ K}$ and $T_2 = 177 + 273 = 450 \text{ K}$, so $P_2/P_1 = 450/300 = 1.5$. C uses the Celsius ratio $177/27 \approx 6.6$, the classic kelvin trap. D ignores the temperature change entirely. A inverts the ratio, giving the factor for cooling instead of heating.
- 5. D.** The adiabatic curve drops more steeply than the isothermal curve, so over the same volume change the isothermal curve encloses the larger area and the gas does more work. A reverses the steepness and the area comparison. B is wrong because the adiabatic expansion cools the gas while the isothermal one holds its temperature. C is wrong twice over: work is the nonzero area under each curve, and $\Delta U = 0$ holds only for the isothermal case.
- 6. A.** $W_{\text{net}} = Q_h - Q_c = 4000 - 3000 = 1000 \text{ J}$, and $e = W_{\text{net}}/Q_h = 1000/4000 = 0.25 = 25\%$. B computes $Q_c/Q_h = 75\%$, the waste fraction, which is the Q_h/Q_c mix-up. C computes $W_{\text{net}}/Q_c = 1000/3000 \approx 33\%$, dividing by the wrong heat. D divides Q_h by W_{net} , inverting the efficiency.
- 7. C.** Melting is irreversible, so the total entropy of the isolated system (ice plus room) increases. The room's entropy falls by less than the ice's entropy rises because the heat leaves at a higher temperature than it arrives at. D gets the ice's entropy change backwards; melting increases disorder. A confuses energy conservation (heat in equals heat out) with entropy, which is not conserved. B reverses the magnitudes: the room's entropy drop is the smaller term.
- 8. A.** Energy travels through the solid metal by molecule-to-molecule collisions with no bulk motion of the material, which is conduction. B requires a flowing fluid, and the rod is solid and stationary. C is a real effect near the flame, but it is not what carries heat along the length of the rod. D is not a heat transfer mode at all.

When you check your answers, note which distinction each miss came from. Make a flashcard for that distinction and drill it spaced out over the next few days instead of rereading the whole section. If you missed one of these questions, the same distinction is worth practicing again in Rycal, where the Thermodynamics deck has flashcards for it and more practice questions use the same kinds of traps.

One-Page Recall Check

Say each answer out loud before you look back, and mark the ones you cannot finish. Anything you cannot say out loud yet belongs in your flashcard deck. In Rycal, add those items to the Thermodynamics deck and let spaced review bring them back over the next few days.

- Explain pressure in terms of molecular collisions with the container walls.
- Write $K_{\text{avg}} = (3/2)k_B T$ and explain what each symbol means, including the required unit of T.
- Explain why doubling the kelvin temperature does not double the molecular speed.
- Write the ideal gas law in both forms and list the SI units of each quantity.
- Convert 27 °C to kelvin and explain why the gas law needs absolute temperature.
- State the first law with the AP sign convention, defining the sign of each term.
- For a gas expanding at constant pressure, write W (work done on the system) in terms of P and ΔV and give its sign.
- Explain how to read work off a PV diagram and how the sign differs for expansion and compression.
- For each of the four processes, state what is held fixed and what becomes zero, then write the first law for it.
- Explain why the adiabatic curve is steeper than the isothermal curve on a PV diagram.
- State the thermal efficiency formula and identify Q_h and Q_c in words.
- State the second law in terms of heat flow and in terms of entropy.
- Explain why the entropy of an isolated system cannot decrease, using the melting ice example.
- Name the three heat transfer modes and state the medium each one requires.

Where to go next. Turn every missed item above into flashcards and drill them spaced out over several days rather than in one sitting. In Rycal, open the Thermodynamics deck under AP Physics 2. The deck covers the terms in this guide, and its practice questions target the same traps named here. If you have a test date, add it in the Test Planner. You can also start your next review with a Brain Dump, then check what you missed against this guide.

Key terms for this unit

Kinetic theory of gases, Pressure, Temperature, Thermal energy, Heat, Thermal equilibrium, Average molecular kinetic energy, Boltzmann constant, Root-mean-square speed, Ideal gas law, Combined gas law, Internal energy, First law of thermodynamics, Sign convention, Work done on the system, PV diagram, Isochoric process, Isobaric process, Isothermal process, Adiabatic process, Heat engine, Hot reservoir, Cold reservoir, Thermal efficiency, Second law of thermodynamics, Entropy, Conduction, Convection, Radiation.

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