Midetrm 1 Study Guide

Mondai, Physics 122, Winter 2026, Cal Poly

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Efficiency

Concept Summary

Efficiency measures how much useful output energy (or work) you get compared to the energy input:

$$e = \frac{E_{\text{useful}}}{E_{\text{input}}}$$

Efficiency is often expressed as a percentage.

Common pitfalls:

- Forgetting that efficiency is always less than 1 (or 100%).

- Mixing up input and useful output energy.

Example Problem

An engine takes in $500\,\text{J}$ of energy and produces $150\,\text{J}$ of useful work.

1. What is the efficiency?

2. How much energy is wasted?

Solution

Concept: Efficiency compares useful output energy to total input energy.

Diagram: [DIAGRAM PLACEHOLDER: Energy input arrow splitting into useful work and wasted energy]

Algebra:

$$e = \frac{E_{\text{useful}}}{E_{\text{input}}}$$

Numbers:

$$e = \frac{150}{500} = 0.30 = 30\%$$

$$E_{\text{wasted}} = 500 - 150 = 350\,\text{J}$$

Energy Use in the Human Body

Concept Summary

The human body converts chemical energy from food into useful work and thermal energy. Like all systems, the body is not 100% efficient.

Common pitfalls:

- Assuming all food energy becomes useful work.

- Forgetting that most energy becomes thermal energy.

Example Problem

A cyclist uses $2000\,\mathrm{J}$ of chemical energy while pedaling and produces $400\,\mathrm{J}$ of mechanical work.

1. What is the efficiency of the cyclist?

2. How much energy becomes thermal energy?

Solution

Concept: The body converts chemical energy into work, with the remainder becoming thermal energy.

Diagram: [DIAGRAM PLACEHOLDER: Chemical energy in body converting to work and thermal energy]

Algebra:
$$e = \frac{W_{\text{out}}}{E_{\text{chem}}}$$

Numbers:
$$e = \frac{400}{2000} = 0.20 = 20\%$$
$$E_{\text{thermal}} = 2000 - 400 = 1600\,\mathrm{J}$$

Assess: A 20% efficiency is reasonable for human muscles. Energy is conserved.

The First Law of Thermodynamics

Concept Summary

The First Law of Thermodynamics relates changes in a system's internal energy to energy transfers:

$$\Delta E_{\text{int}} = Q + W_{env}$$

where Q is heat added to the system and W is work done by the system.

Common pitfalls:

- Using the wrong sign for work.

- Forgetting that work done *by* the system is positive if the dispalcement and F_{gas} are in the same direction.

Example Problem

A gas absorbs $300\,\text{J}$ of heat and does $100\,\text{J}$ of work on its surroundings.

1. What is the change in internal energy of the gas?

Solution

Concept: Internal energy changes based on heat added and work done by the system.

Diagram: [DIAGRAM PLACEHOLDER: Gas in cylinder with heat entering and piston moving]

Algebra:

$$\Delta E_{\text{int}} = Q + W_{env}$$

Numbers:

$$\Delta E_{\text{int}} = 300 - 100 = 200\,\text{J}$$

Assess: The internal energy increases, which makes sense because more heat was added than work done.

Example: Heat and Work from a PV Diagram

Problem

A gas undergoes the process shown on the PV diagram below. The pressure remains constant at $P = 200\,\text{kPa}$ while the volume increases from $V_1 = 1.0\,\text{m}^3$ to $V_2 = 3.0\,\text{m}^3$.

During this process, the internal energy of the gas increases by $200\,\text{J}$.

1. What is the work done by the environment on the gas?

2. How much heat is transferred to the gas?

PV Diagram
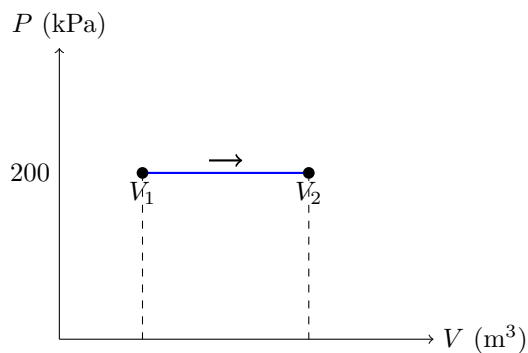
Solution

Concept: On a PV diagram, the work magnitude equals the area under the curve. For expansion, the environment does negative work on the gas. Heat is found using the First Law, $\Delta E = Q + W_{\text{env}}$.

Algebra:
$$W_{\text{env}} = -P\,\Delta V$$
$$\Delta E = Q + W_{\text{env}}$$

Numbers:
$$\Delta V = 3.0 - 1.0 = 2.0\,\text{m}^3$$
$$W_{\text{env}} = -(200\,\text{kPa})(2.0\,\text{m}^3) = -(2.0 \times 10^5)(2.0) = -4.0 \times 10^5\,\text{J}$$
$$Q = \Delta E - W_{\text{env}} = 200 - (-4.0 \times 10^5) = 4.002 \times 10^5\,\text{J}$$

Assess: The negative work confirms expansion. Most of the added heat goes into work done on the surroundings, which is consistent with the large area under the curve.

Example: Entropy (Conceptual)

Concept Summary

Entropy is a measure of how energy is spread out or how many microscopic arrangements (microstates) are possible. Natural processes tend to move toward higher entropy.

Example Problem

A hot metal block is placed in contact with a cold metal block inside an insulated box. Energy flows from the hot block to the cold block until they reach the same temperature.

1. Does the entropy of the hot block increase or decrease?

2. Does the entropy of the cold block increase or decrease?

3. Does the total entropy of the two-block system increase, decrease, or stay the same?

Solution

Concept: Entropy increases when energy spreads out. Heat flowing from hot to cold increases the number of accessible microstates.

Diagram: [DIAGRAM PLACEHOLDER: Hot block and cold block with heat flowing from hot to cold]

Reasoning:

- The hot block loses energy, so its entropy decreases.

- The cold block gains energy, so its entropy increases.

- The increase in entropy of the cold block is greater than the decrease in entropy of the hot block.

Conclusion: The total entropy of the two-block system increases, which is why heat flows spontaneously from hot to cold.

Assess: This process is irreversible and consistent with the Second Law of Thermodynamics: entropy of an isolated system increases.

Chapter 12: Temperature and Heat

Temperature, Thermal Energy, and Molecular Speed

Concept Summary

- **Temperature** measures the *average kinetic energy* of the particles in a substance.

- **Thermal energy** is the *total* kinetic energy of all particles and depends on both temperature and the amount of substance.

- The **root-mean-square speed** v_{rms} describes the typical speed of particles in a gas and depends only on temperature and particle mass.

Key ideas:

- Higher temperature $\Rightarrow$ faster average particle motion.

- Larger objects can have more thermal energy even at the same temperature.

- At the same temperature, lighter molecules move faster than heavier ones.

Common pitfalls:

- Confusing temperature with thermal energy.

- Thinking individual particles all move at the same speed.

- Assuming heavier molecules move faster at the same temperature.

Example Problem

Two sealed containers, A and B, hold ideal gases at the same temperature. Container A holds helium atoms, while container B holds oxygen molecules.

1. Which container has gas particles with the greater average kinetic energy?

2. Which gas has the greater v_{rms}?

3. If container B has more gas particles than container A, which container has more thermal energy?

Solution

Concept: Temperature determines average kinetic energy. Thermal energy depends on both temperature and the number of particles. The rms speed depends on temperature and particle mass.

Diagram: [DIAGRAM PLACEHOLDER: Two containers at same temperature; light helium particles moving faster than heavier oxygen molecules]

Reasoning:

- Since the temperatures are the same, particles in both containers have the same average kinetic energy.

- Helium atoms are much lighter than oxygen molecules, so helium has the larger v_{rms}.

- Container B has more particles, so it has greater total thermal energy.

Assess: These results are consistent with molecular theory: temperature sets average kinetic energy, mass affects speed, and thermal energy scales with the number of particles.

The Ideal Gas Law

Concept Summary

An ideal gas is a simplified model in which gas particles:

- Have negligible volume compared to the container.

- Do not exert forces on each other except during collisions.

For an ideal gas, pressure, volume, temperature, and the amount of gas are related by:
$$PV = nRT$$

where:

- P is pressure,

- V is volume,

- n is the number of moles,

- $R = 8.31 \, \text{J/(mol·K)}$,

- T is temperature in kelvins.

Key ideas:

- Temperature must always be in kelvins.

- Increasing temperature increases molecular speed and pressure (if volume is fixed).

- The ideal gas law connects macroscopic quantities to microscopic motion.

Common pitfalls:

- Using Celsius instead of kelvins.

- Forgetting that pressure depends on molecular collisions with container walls.

Example Problem

A container holds $1.0 \, \text{mol}$ of an ideal gas at a temperature of $300 \, \text{K}$. The gas occupies a volume of $0.25 \, \text{m}^3$.

1. What is the pressure of the gas?

2. If the temperature increases while volume stays constant, what happens to the pressure?

Solution

Concept: The ideal gas law relates pressure to temperature, volume, and the number of moles. At fixed volume, pressure is directly proportional to temperature.

Diagram: [DIAGRAM PLACEHOLDER: Gas particles colliding with the walls of a rigid container]

Algebra:

$$P = \frac{nRT}{V}$$

Numbers:

$$P = \frac{(1.0)(8.31)(300)}{0.25}$$

$$P = \frac{2493}{0.25} \approx 1.0 \times 10^4 \, \text{Pa}$$

Reasoning (Part 2): If the temperature increases while the volume stays constant, gas particles move faster and collide more often with the walls, so the pressure increases.

Assess: The pressure is reasonable for a low-density gas. The direct relationship between pressure and temperature is consistent with kinetic theory.

Constant-Pressure (Isobaric) Process

Example: One mole of an ideal gas expands at constant pressure $P = 2.0 \times 10^5$ Pa. The volume increases from $V_i = 0.010$ m^3 to $V_f = 0.030$ m^3.

Concept: At constant pressure, the PV path is a horizontal line. The work done by the environment equals the negative of the area under this line on a PV diagram.
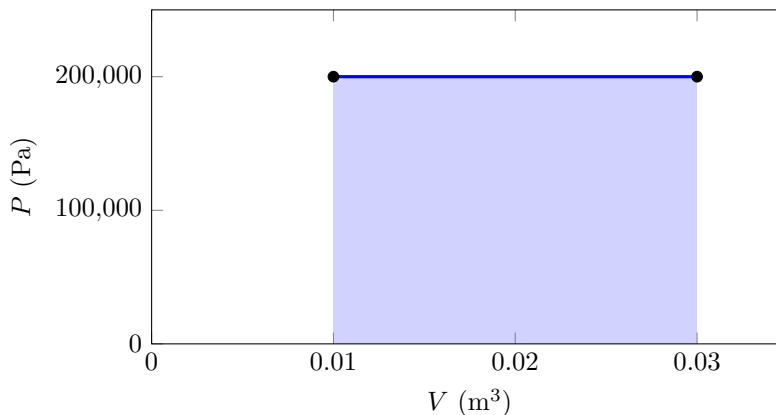
Algebra:

$$W_{\text{env}} = -P\,\Delta V$$

Numbers:

$$\Delta V = 0.030 - 0.010 = 0.020 \text{ m}^3$$

$$W_{\text{env}} = -(2.0 \times 10^5)(0.020) = -4.0 \times 10^3 \text{ J}$$

Assess: The negative sign is correct because the gas expands. The magnitude is reasonable for this pressure and volume change.

Constant-Volume (Isochoric) Process

Example: One mole of an ideal gas is heated at constant volume $V = 0.020$ m^3. The pressure increases from $P_i = 1.0 \times 10^5$ Pa to $P_f = 3.0 \times 10^5$ Pa. **Find the heat Q added to the gas.**

    Concept: In a constant-volume process the PV path is vertical, so no work is done by or on the gas. Any change in internal energy must come entirely from heat transfer.
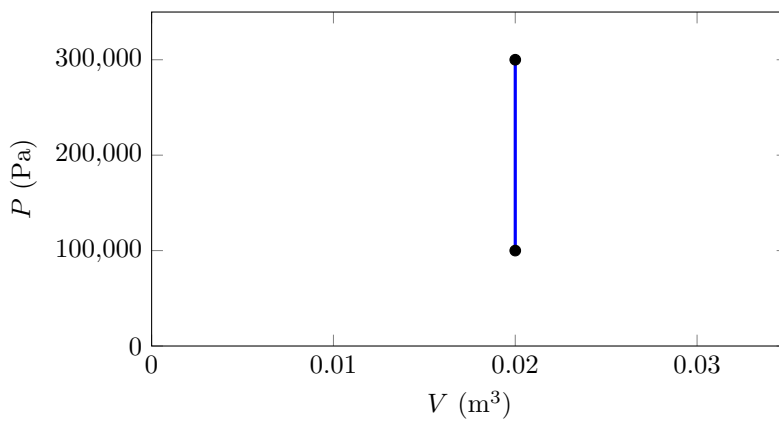
    Algebra: First Law of Thermodynamics:

$$\Delta E = Q + W_{\text{env}}$$

For a constant-volume process:

$$W_{\text{env}} = 0$$

For an ideal gas:

$$\Delta E = \frac{3}{2}nR\Delta T$$

Numbers: Use the ideal gas law to find the temperature change:

$$T = \frac{PV}{nR}$$

$$T_i = \frac{(1.0 \times 10^5)(0.020)}{(1)(8.31)} \approx 240 \text{ K}$$

$$T_f = \frac{(3.0 \times 10^5)(0.020)}{(1)(8.31)} \approx 720 \text{ K}$$

$$\Delta T = 720 - 240 = 480 \text{ K}$$

$$\Delta E = \frac{3}{2}(1)(8.31)(480) \approx 6.0 \times 10^3 \text{ J}$$

Since $W_{\text{env}} = 0$:

$$Q = \Delta E \approx 6.0 \times 10^3 \text{ J}$$

Assess: The heat is positive, as expected for heating the gas. The magnitude is reasonable for a large temperature increase at fixed volume.

Constant-Temperature (Isothermal) Process

Example: One mole of an ideal gas undergoes an isothermal expansion at temperature $T = 300$ K. The volume increases from $V_i = 0.010$ m^3 to $V_f = 0.040$ m^3.

Concept: In an isothermal process the temperature remains constant, so the gas obeys $PV = $ constant. The PV diagram is a downward-curving hyperbola, and the work done by the environment equals the signed area under this curve.
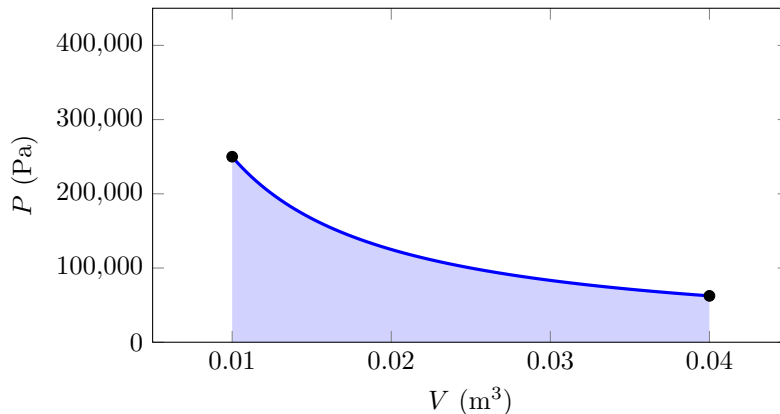
Adiabatic Process (Conceptual)

Example: An ideal gas expands adiabatically inside an insulated cylinder, so no heat enters or leaves the gas during the process.

Concept: In an adiabatic process there is no heat transfer ($Q = 0$), so any work done by the gas comes at the expense of its internal energy. As the gas expands, its temperature decreases. On a PV diagram, an adiabatic curve slopes downward more steeply than an isothermal curve passing through the same initial point.
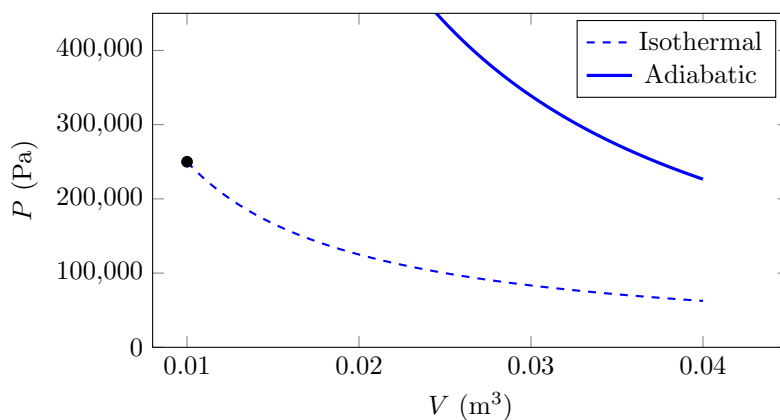
Assess: The steeper adiabatic curve reflects the fact that pressure drops more rapidly with volume than in an isothermal process because the gas cools as it does work with no heat input.

1D Thermal Expansion (Linear)

Example: A steel rod has an initial length $L_0 = 2.0$ m at 20°C. The temperature increases to 120°C. The coefficient of linear expansion for steel is $\alpha = 1.2 \times 10^{-5}$ °C^{-1}. Find the change in length of the rod.

Concept: In one-dimensional thermal expansion, the change in length is proportional to the original length and the temperature change.

Algebra: The linear expansion formula is $\Delta L = \alpha L_0 \Delta T$.

Numbers: The temperature change is $\Delta T = 120 - 20 = 100$°C. Substituting gives $\Delta L = (1.2 \times 10^{-5})(2.0)(100) = 2.4 \times 10^{-3}$ m.

Assess: The rod length increases by 2.4 mm, which is reasonable for steel over a 100°C temperature change.

3D Thermal Expansion (Volume)

Example: A solid aluminum block has an initial volume $V_0 = 1.0 \times 10^{-2}$ m^3 at 20°C. The temperature increases to 70°C. The coefficient of linear expansion for aluminum is $\alpha = 2.3 \times 10^{-5}$ °C^{-1}. Find the change in volume of the block.

Concept: In three dimensions the object expands independently along three perpendicular directions. For small temperature changes, the volume expansion coefficient is approximately 3α, extending the 2D result where the area expansion coefficient is 2α.

Algebra: The volume expansion formula is $\Delta V = 3\alpha V_0 \Delta T$.

Numbers: The temperature change is $\Delta T = 70 - 20 = 50$°C. Substituting gives $\Delta V = 3(2.3 \times 10^{-5})(1.0 \times 10^{-2})(50) = 3.45 \times 10^{-5}$ m^3.

Assess: The volume increase is much smaller than the original volume, consistent with the assumption of small thermal expansion and validating the linear approximation.