Physics 1501

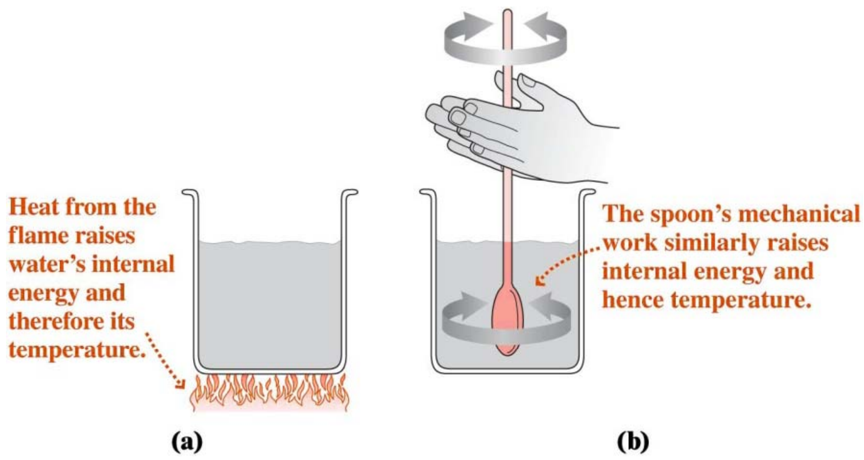
Fall 2008

Mechanics, Thermodynamics, Waves, Fluids

Lecture 32: Heat and Work II

Recap: the first law of thermodynamics

- Two ways to raise temperature:
  - Thermally: flow of heat
    - Energy flow resulting from a temperature difference
  - Mechanically: doing work
  - End result is the same for the same energy input:
    - An increase ΔU in the system's **internal energy** U



- First law of thermodynamics
    - Change in internal energy of a system is the heat added *to* the system minus the work done *by* the system:

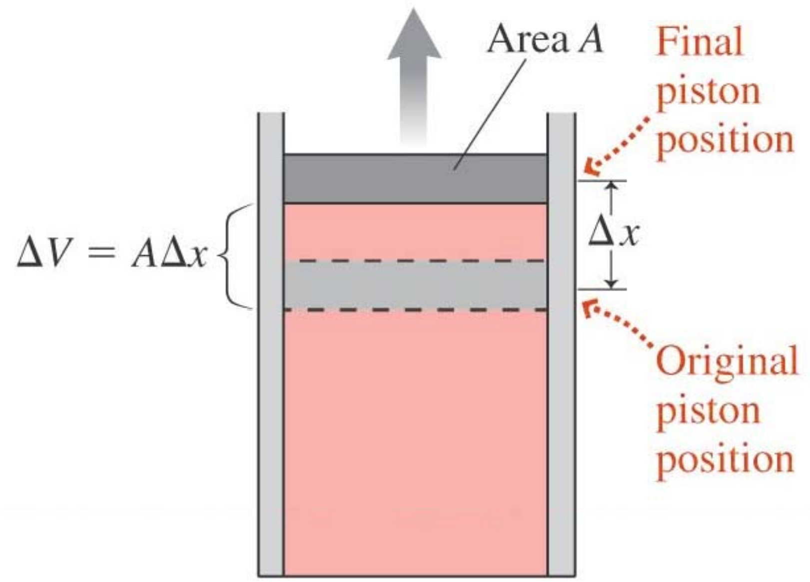
$$\Delta U = Q - W$$

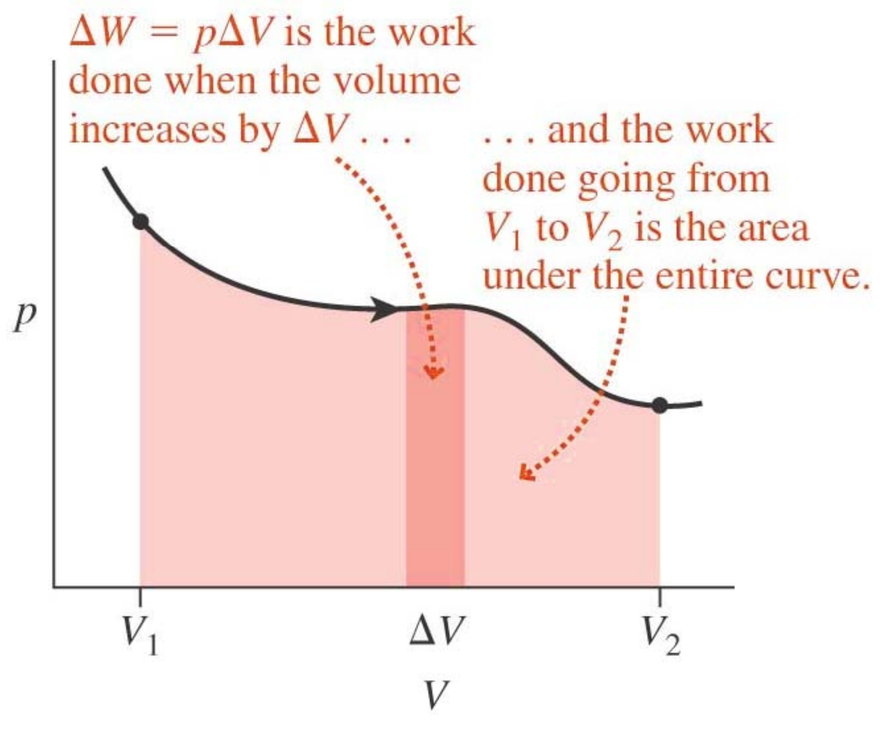
    - Extends conservation of energy to include thermal processes

Recap: Doing work

- A piston-cylinder system is a useful device for describing the thermodynamic behavior of a gas.
  - The piston seals the cylinder, allowing the gas volume to change without any gas escaping.
  - Work can be done on or by the gas as the piston moves.
  - If the bottom is uninsulated, heat can flow in or out.
- The work done on or by the gas is the area under the pV curve:

$$W = \int_{V_1}^{V_2} p\, dV$$

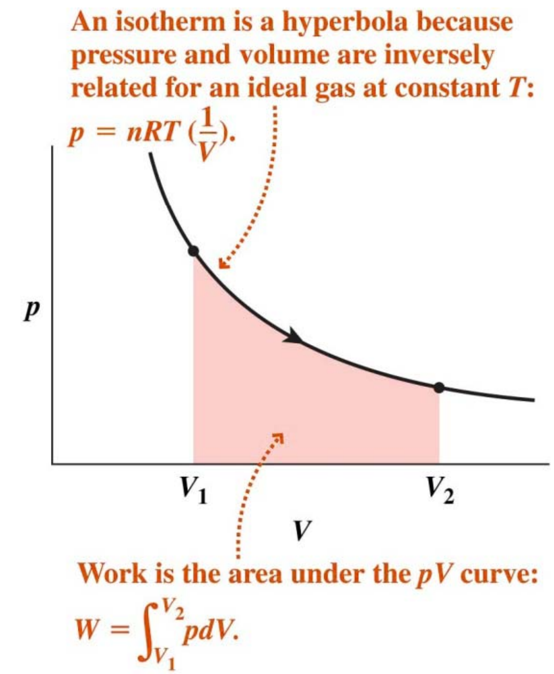




Recap: isothermal processes

- An **isothermal process** takes place at constant temperature.
  - One way to achieve this is to keep the system in thermal contact with a heat reservoir—a much larger system held at constant temperature.
  - The ideal gas law gives $P = nRT/V$.
  - Then, with constant T, the work done is

$$W = \int_{V_1}^{V_2} p\,dV = nRT\int_{V_1}^{V_2} \frac{dV}{V} = nRT\ln\left(V_2/V_1\right)$$



  - Since the temperature doesn't change, neither does the internal energy of an ideal gas. Therefore the first law gives

$$Q = W = nRT\ln\left(V_2/V_1\right)$$

Recap: constant-volume processes

- In a constant-volume process, the heat added to the gas and the resulting temperature change are related by

$$Q = nC_V\Delta T$$

  - Here C_V is the **molar specific heat at constant volume.**
    - Its units are J/K·mol.
- No work is done in a constant-volume process, so $W = 0$ and so the first law reads $\Delta U = Q$.
  - Therefore $\Delta U = nC_V\Delta T$.
  - For an ideal gas, internal energy depends on temperature alone, so this relationship $\Delta U = nC_V\Delta T$ between ΔU and ΔT holds for *any* process.

Recap: isobaric processes

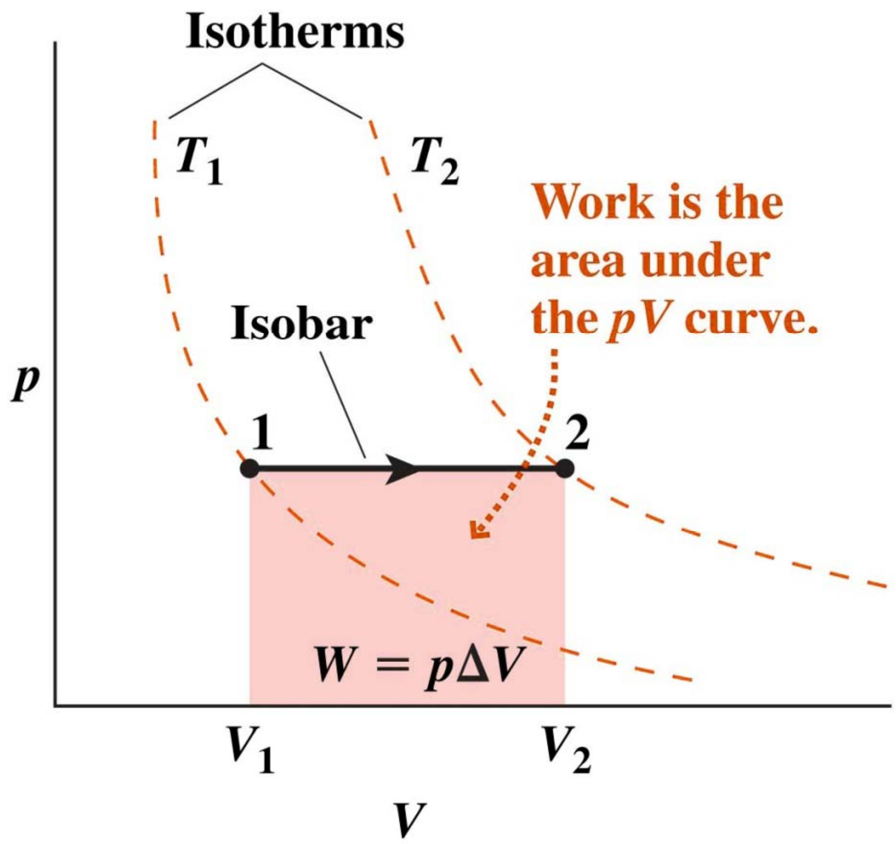
- An **isobaric process** takes place at constant pressure.
- Then the work done is $W = p\,\Delta V$.
- Adding heat to an ideal gas results both in a temperature change and in work being done.
  - Therefore it takes *more* heat to effect a given temperature change:

    $$Q = nC_V\Delta T + p\,\Delta V$$

  - The **molar specific heat at constant pressure**, C_p, expresses this extra work:

    $$nC_p\Delta T = nC_V\Delta T + p\,\Delta V$$

  - Here $C_p = C_V + R$.

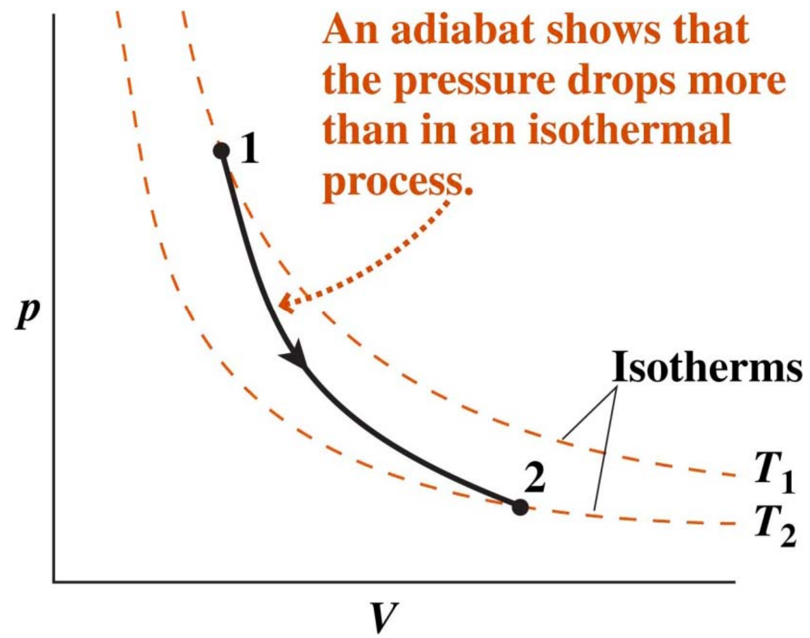


Recap: adiabatic processes

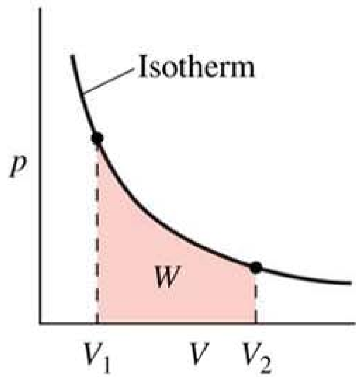
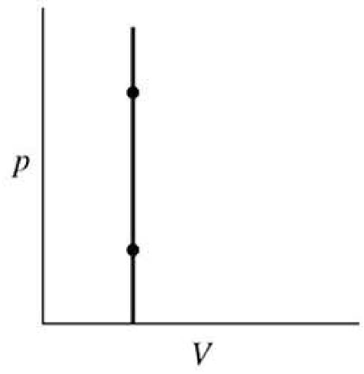
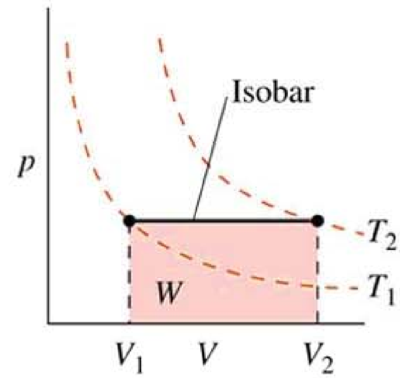
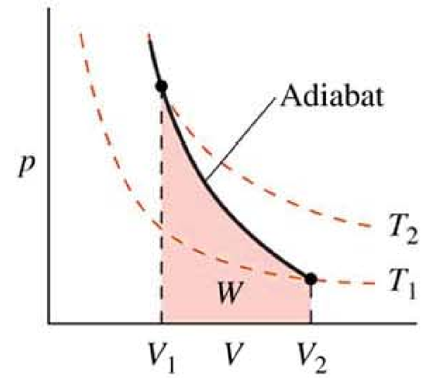
- In an **adiabatic process**, no heat flows into or out of the system.
  - Therefore $Q = 0$ and the first law reads $\Delta U = -W$.
  - Analysis of the adiabatic process for an ideal gas shows that

$$pV^{\gamma} = \text{constant}$$

    where $\gamma = C_p/C_V$ is the ratio of specific heats of the gas.
  - An adiabatic curve or **adiabat** is steeper than an isotherm because the gas does work.
    - It therefore loses internal energy and its temperature drops.



Ideal-gas processes: a comparison

	ISOTHERMAL	CONSTANT-VOLUME	ISOBARIC	ADIABATIC
pV diagram				
Defining characteristic	T = constant	V = constant	p = constant	$Q = 0$
First law	$Q = W$	$Q = \Delta U$	$Q = \Delta U + W$	$\Delta U = -W$
Work done by gas	$W = nRT \ln\left(\frac{V_2}{V_1}\right)$	$W = 0$	$W = p(V_2 - V_1)$	$W = \frac{p_1V_1 - p_2V_2}{\gamma - 1}$
Other relationships	pV = constant	$Q = nC_V\Delta T$	$Q = nC_p\Delta T$ $C_p = C_V + R$	pV^γ = constant $TV^{\gamma-1}$ = constant

question

- The ideal-gas law states that $pV = nRT$ but for an adiabatic process, we are taught that $pV^{\gamma} =$ constant. How can we reconcile this apparent contradiction?

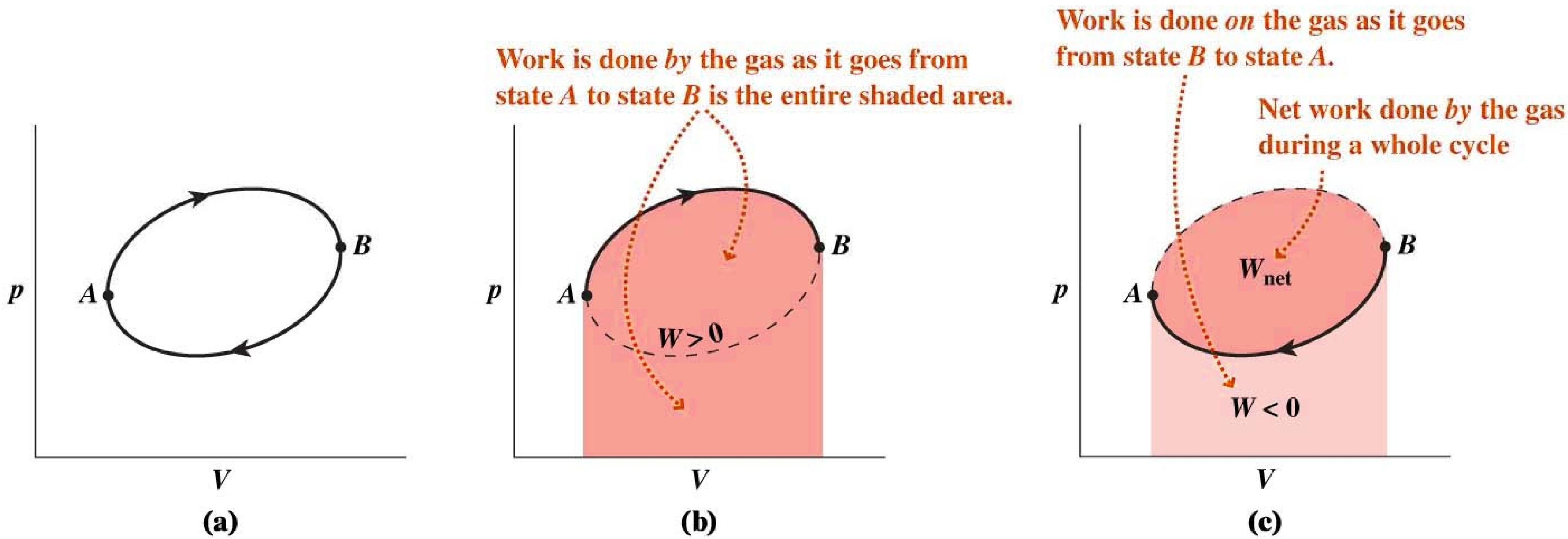
A. The ideal gas law does not apply to adiabatic processes.

B. The term pV in the ideal-gas law need not remain constant.

C. In an adiabatic process, γ changes such that pV^{γ} remains constant.

Cyclic processes

- Cyclic processes combine the basic processes of other thermodynamic processes to take a system around a complete cycle and back to its starting state.
  - Cyclic processes are important in technological systems like engines.
  - They also occur in natural systems from sound waves to oscillating stars.
  - The work done in or by the gas in a cyclic process is the area enclosed by the process curve in the pV diagram.

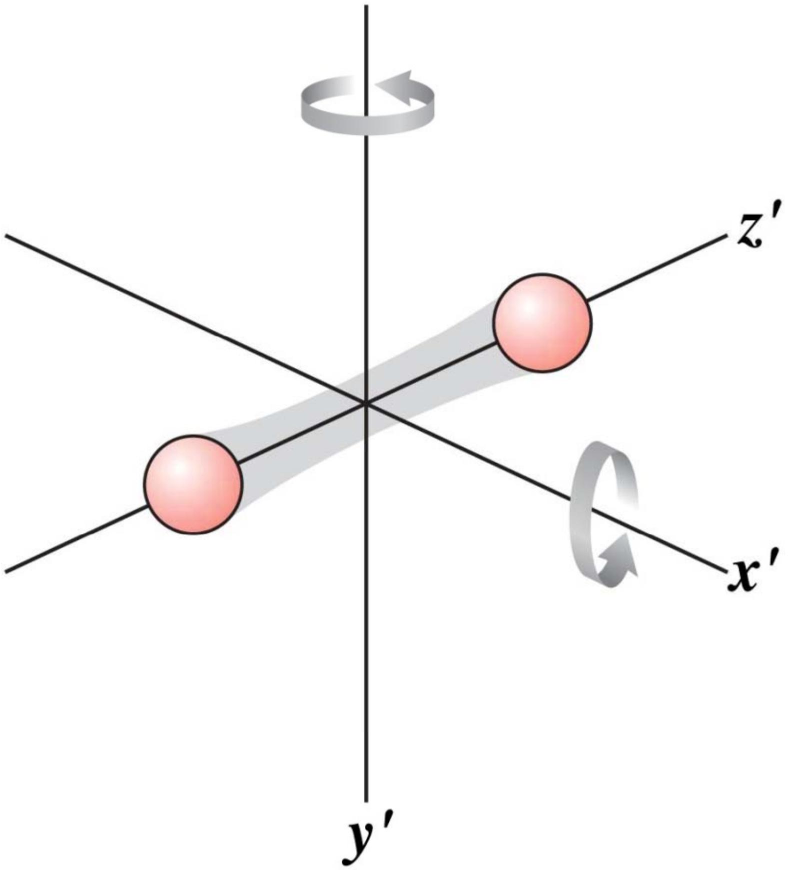


Specific heats of an ideal gas

- The specific heat of an ideal gas depends on its molecular structure.
  - More complex molecules have more **degrees of freedom**, or ways they can absorb energy.
  - Monatomic gases have 3 degrees of freedom per molecule (the three directions of translational motion), and have volume specific heat

    $$C_V = \tfrac{3}{2}R$$

    and adiabatic exponent $\gamma = 1.67$.
  - Diatomic gases have 5 degrees of freedom per molecule (three translational and 2 rotational), and have volume specific heat $C_V = \tfrac{5}{2}R$ and adiabatic exponent $\gamma = 1.4$.



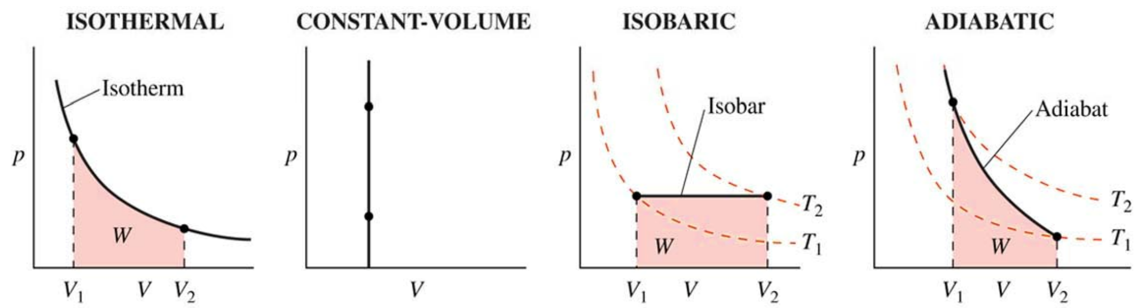
A diatomic molecule has rotational as well as translational energy.

Summary

- The first law of thermodynamics states that the change in a system's internal energy is equal to the heat added to the system minus the work done by the system:

$$\Delta U = Q - W$$

- Thermodynamic processes take systems between states in their pV diagrams.
  - Reversible processes remain in equilibrium and follow a definite path in the pV diagram.
  - Irreversible processes result in temporary loss of equilibrium and don't follow a definite path.
- Important ideal gas processes include
  - Isothermal (constant temperature)
  - Constant volume
  - Isobaric (constant pressure)
  - Adiabatic (no heat flow)



  - Cyclic processes that combine these or other processes to take a system through a complete cycle
- The specific heats of an ideal gas depend on the structure of the gas molecules.
  - More complex molecules have more degrees of freedom for absorbing energy.