# PHYS1001B College Physics IB

Thermodynamics III — The First Law of Thermodynamics (Ch. 19)

# Introduction

Thermodynamics processes

# Outline

▸ 19.1 Thermodynamic Systems

▸ 19.2 Work Done During Volume Changes

▸ 19.3 Paths Between Thermodynamic States

▸ 19.4 Internal Energy and the First Law of Thermodynamics

▸ 19.5 Kinds of Thermodynamic Processes

▸ 19.6 Internal Energy of an Ideal Gas

▸ 19.7 Heat Capacities of an Ideal Gas

▸ 19.8 Adiabatic Processes for an Ideal Gas
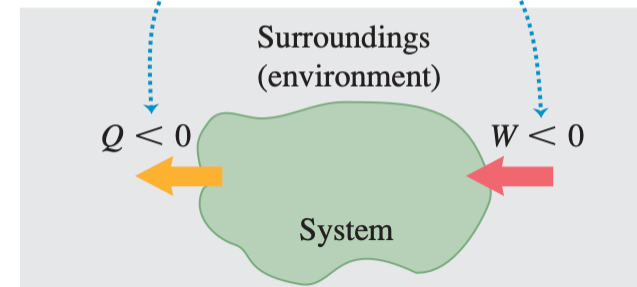
# 19.1 Thermodynamic Systems

**19.3** A thermodynamic system may exchange energy with its surroundings (environment) by means of heat, work, or both. Note the sign conventions for $Q$ and $W$.

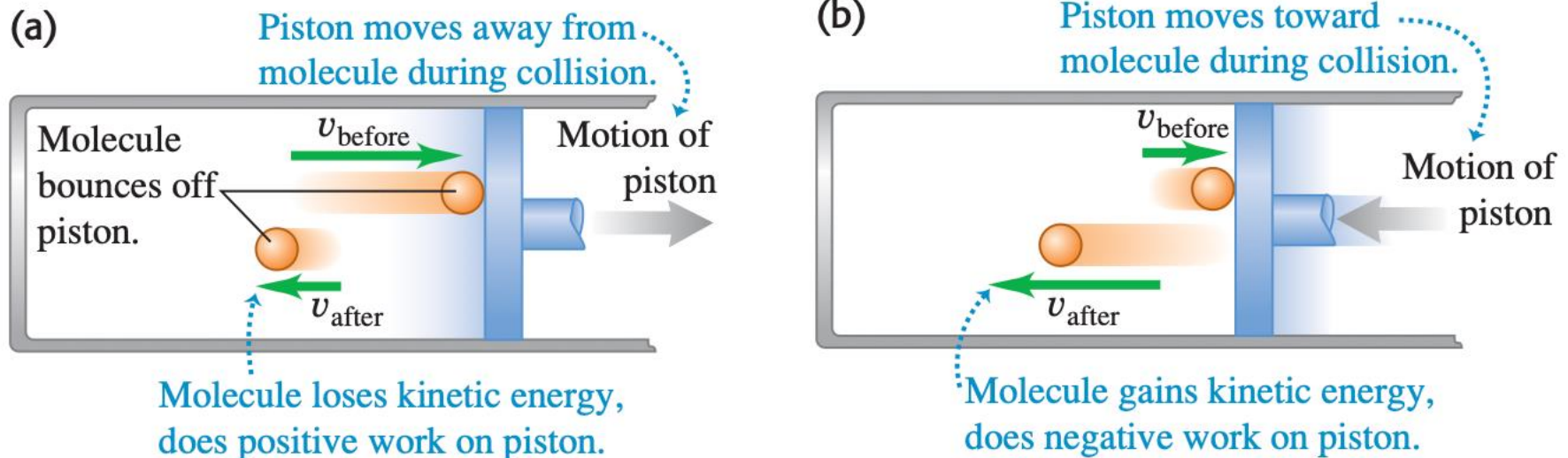Surroundings (environment)

$Q > 0$    System    $W > 0$
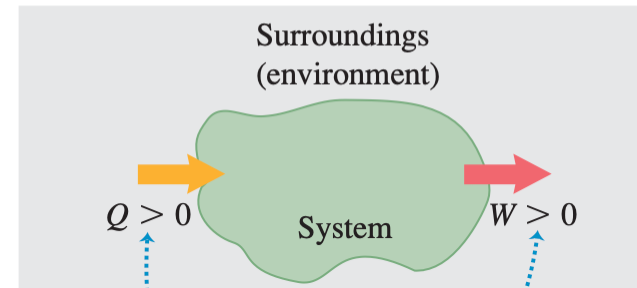
Heat is positive when it *enters* the system, negative when it *leaves* the system.

Work is positive when it is done *by* the system, negative when it is done *on* the system.

Surroundings (environment)

$Q < 0$    $W < 0$

System

# 19.2 Work Done During Volume Changes



(a) Piston moves away from molecule during collision.
Molecule bounces off piston.
$v_{before}$
Motion of piston
$v_{after}$
Molecule loses kinetic energy, does positive work on piston.

(b) Piston moves toward molecule during collision.
$v_{before}$
Motion of piston
$v_{after}$
Molecule gains kinetic energy, does negative work on piston.

# 19.2 Work Done During Volume Changes

**19.5** The infinitesimal work done by the system during the small expansion $dx$ is $dW = pA\ dx$.

System

$dx$

$pA$

$A$

Force that system exerts on piston
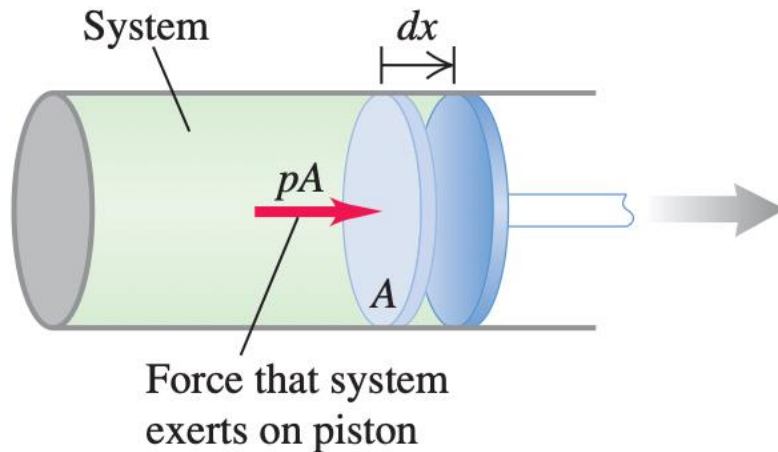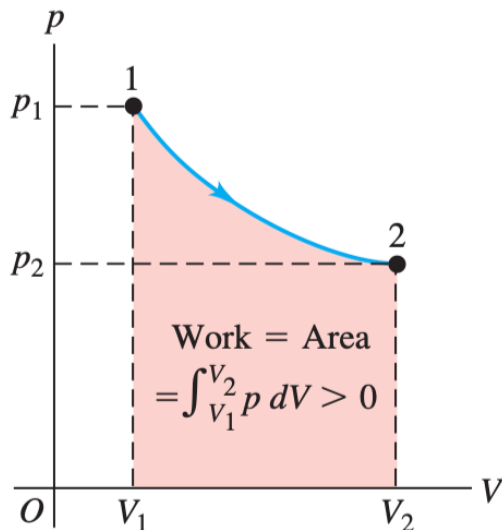
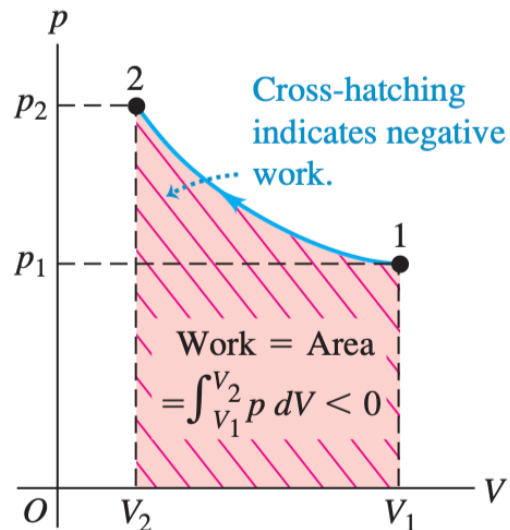$$dW = F\ dx = pA\ dx$$

$$A\ dx = dV$$

$$dW = p\ dV$$

$$W = \int_{V_1}^{V_2} p\ dV \qquad \text{(work done in a volume change)}$$
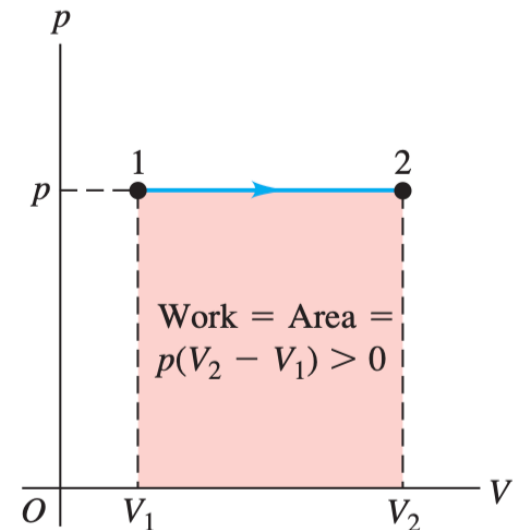
# 19.2 Work Done During Volume Changes

**(a)** $pV$-diagram for a system undergoing an expansion with varying pressure

**(b)** $pV$-diagram for a system undergoing a compression with varying pressure

**(c)** $pV$-diagram for a system undergoing an expansion with constant pressure

Cross-hatching indicates negative work.

Work = Area $= \int_{V_1}^{V_2} p\, dV > 0$

Work = Area $= \int_{V_1}^{V_2} p\, dV < 0$

Work = Area = $p(V_2 - V_1) > 0$

# Sample Problem

| Example 19.1 | Isothermal expansion of an ideal gas |
|---|---|

As an ideal gas undergoes an *isothermal* (constant-temperature) expansion at temperature $T$, its volume changes from $V_1$ to $V_2$. How much work does the gas do?

**EXECUTE:** From Eq. (18.3),

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

We substitute this into the integral of Eq. (19.2), take the constant factor $nRT$ outside, and evaluate the integral:

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

$$= nRT \int_{V_1}^{V_2} \frac{dV}{V} = nRT\ \ln\frac{V_2}{V_1} \quad \text{(ideal gas, isothermal process)}$$

We can rewrite this expression for $W$ in terms of $p_1$ and $p_2$. Because $pV = nRT$ is constant,
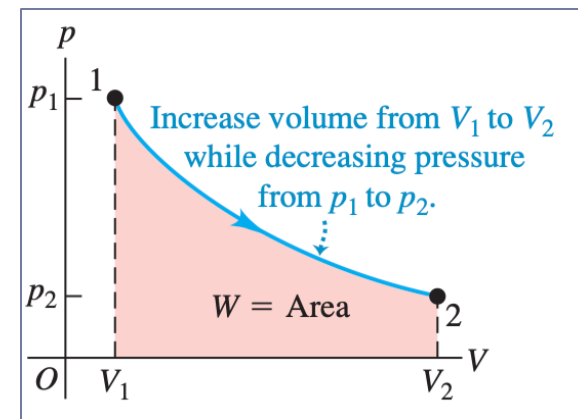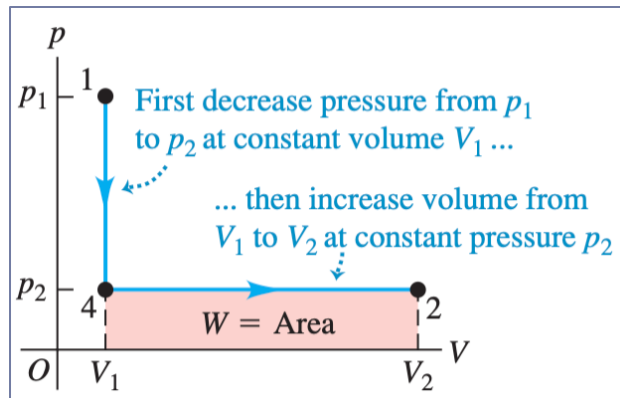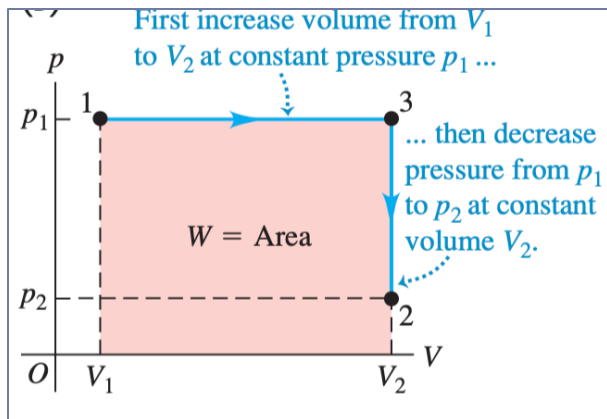
$$p_1 V_1 = p_2 V_2 \quad \text{or} \quad \frac{V_2}{V_1} = \frac{p_1}{p_2}$$

so

$$W = nRT\ \ln\frac{p_1}{p_2} \quad \text{(ideal gas, isothermal process)}$$

# 19.3 Paths Between Thermodynamic States

Work Done in a Thermodynamic Process



These paths give three options for getting from state 1 to state 2.

First increase volume from $V_1$ to $V_2$ at constant pressure $p_1$ ...

... then decrease pressure from $p_1$ to $p_2$ at constant volume $V_2$.

$W = \text{Area}$

First decrease pressure from $p_1$ to $p_2$ at constant volume $V_1$ ...

... then increase volume from $V_1$ to $V_2$ at constant pressure $p_2$.

$W = \text{Area}$

Increase volume from $V_1$ to $V_2$ while decreasing pressure from $p_1$ to $p_2$.

$W = \text{Area}$
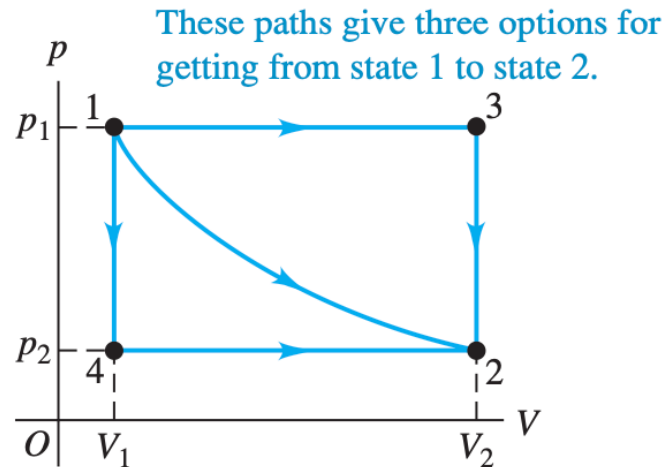
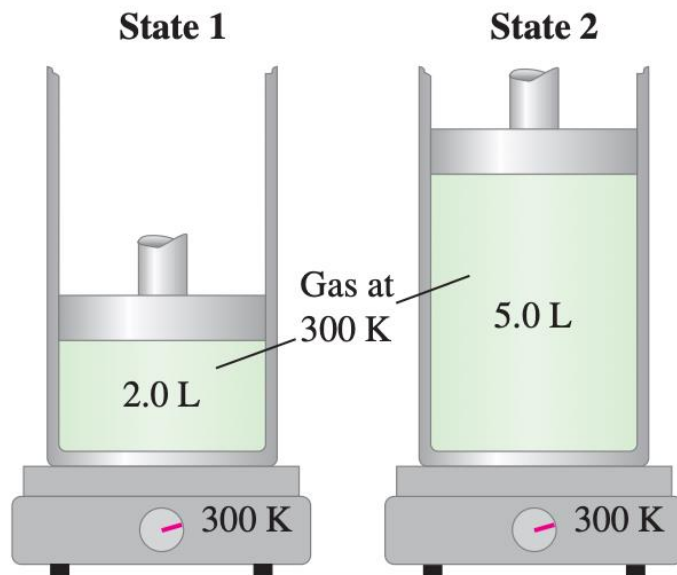# 19.3 Paths Between Thermodynamic States

## Heat Added in a Thermodynamic Process



(a) System does work on piston; hot plate adds heat to system $(W > 0 \text{ and } Q > 0)$.

State 1    State 2

Gas at 300 K

5.0 L

2.0 L

300 K    300 K

(b) System does no work; no heat enters or leaves system $(W = 0 \text{ and } Q = 0)$.

State 1    Insulation    State 2

Vacuum

5.0 L

2.0 L

Breakable partition    Gas at 300 K
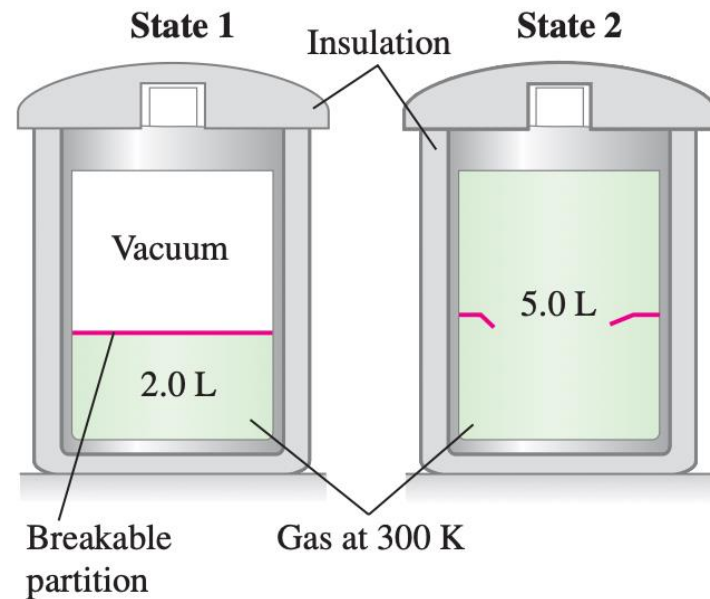
*heat depends not only on the initial and final states but also on the path.*

# 19.4 Internal Energy and the First Law of Thermodynamics
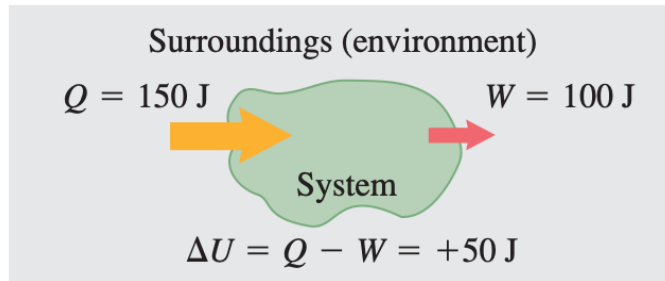
First Law of Thermodynamics

$$U_2 - U_1 = \Delta U = Q - W$$

$U$ for internal energy

**(a)** More heat is added to system than system does work: Internal energy of system increases.

Surroundings (environment)

$Q = 150$ J  →  $W = 100$ J

System

$\Delta U = Q - W = +50$ J

**(b)** More heat flows out of system than work is done: Internal energy of system decreases.

Surroundings (environment)

$Q = -150$ J  ←  $W = -100$ J

System

$\Delta U = Q - W = -50$ J
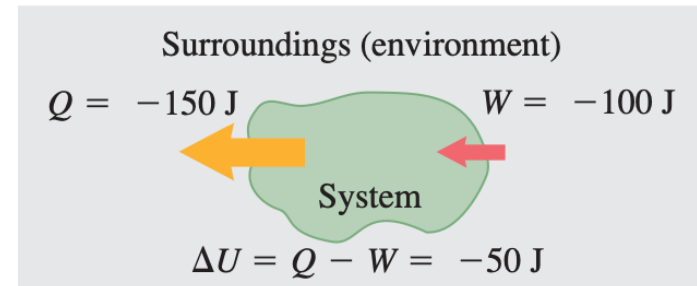
**(c)** Heat added to system equals work done by system: Internal energy of system unchanged.

Surroundings (environment)

$Q = 150$ J  →  $W = 150$ J

System

$\Delta U = Q - W = 0$
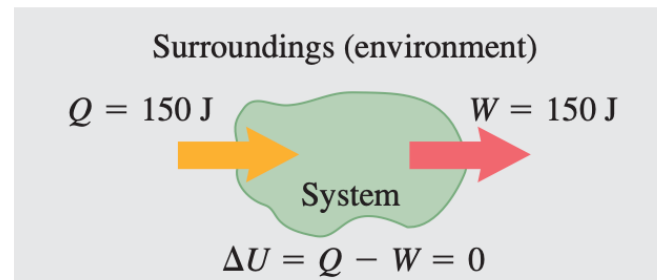
# 19.4 Internal Energy and the First Law of Thermodynamics

**19.10** The internal energy of a cup of coffee depends on just its thermodynamic state—how much water and ground coffee it contains, and what its temperature is. It does not depend on the history of how the coffee was prepared—that is, the thermodynamic path that led to its current state.
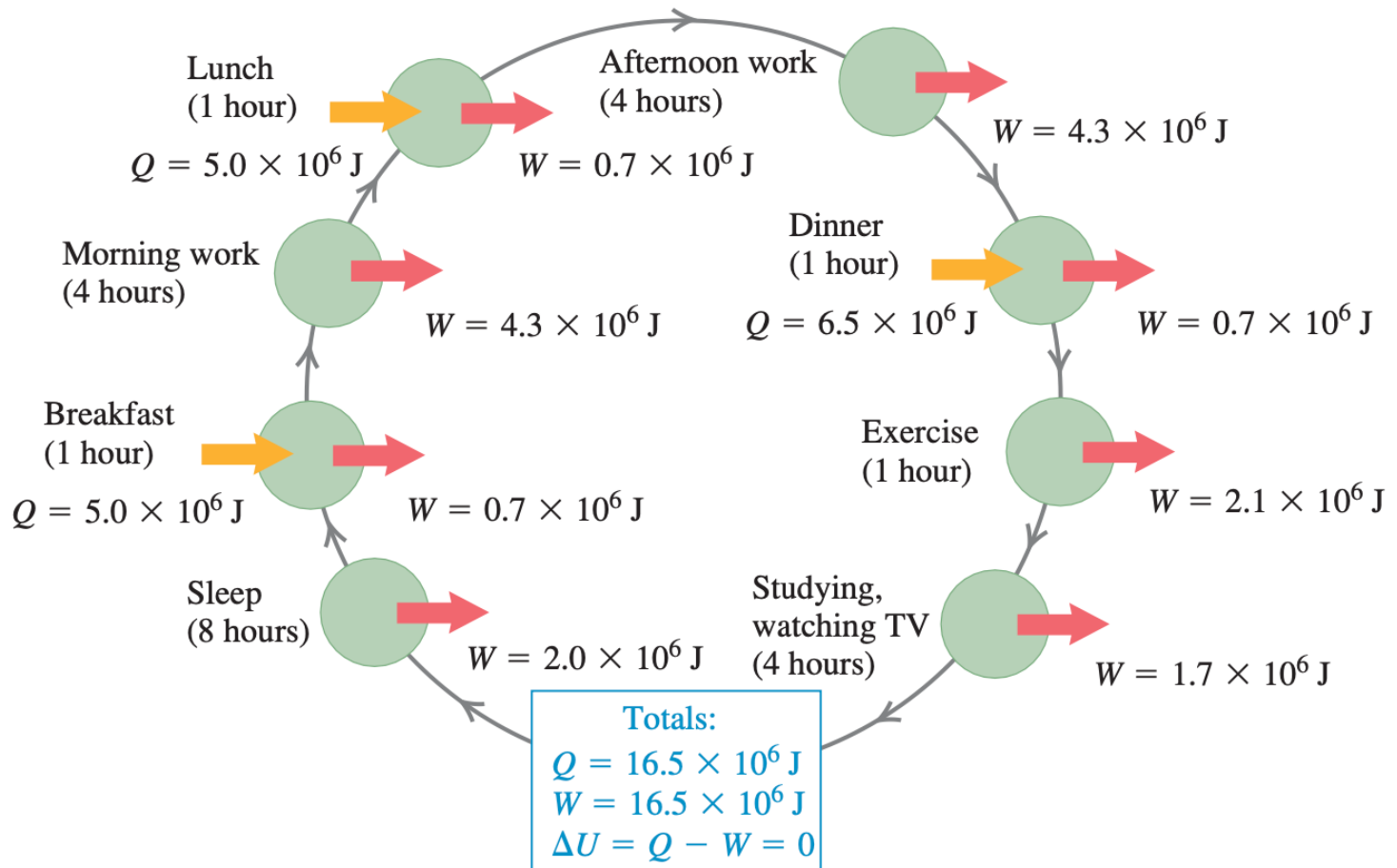
# 19.4 Internal Energy and the First Law of Thermodynamics

Cyclic Processes



Lunch
(1 hour)

$Q = 5.0 \times 10^6$ J

Afternoon work
(4 hours)

$W = 0.7 \times 10^6$ J

$W = 4.3 \times 10^6$ J

Morning work
(4 hours)

$W = 4.3 \times 10^6$ J

Dinner
(1 hour)

$Q = 6.5 \times 10^6$ J

$W = 0.7 \times 10^6$ J

Breakfast
(1 hour)

$Q = 5.0 \times 10^6$ J

$W = 0.7 \times 10^6$ J

Exercise
(1 hour)

$W = 2.1 \times 10^6$ J

Sleep
(8 hours)

$W = 2.0 \times 10^6$ J

Studying,
watching TV
(4 hours)

$W = 1.7 \times 10^6$ J

Totals:
$Q = 16.5 \times 10^6$ J
$W = 16.5 \times 10^6$ J
$\Delta U = Q - W = 0$

# 19.4 Internal Energy and the First Law of Thermodynamics

Isolated system

Another special case occurs in an *isolated system,* one that does no work on its surroundings and has no heat flow to or from its surroundings. For any process taking place in an isolated system,

$$W = Q = 0$$

and therefore

$$U_2 = U_1 = \Delta U = 0$$

In other words, *the internal energy of an isolated system is constant.*

# 19.4 Internal Energy and the First Law of Thermodynamics

## Infinitesimal Changes of State

In the preceding examples the initial and final states differ by a finite amount. Later we will consider *infinitesimal* changes of state in which a small amount of heat $dQ$ is added to the system, the system does a small amount of work $dW$, and its internal energy changes by an amount $dU$. For such a process we state the first law in differential form as

$$dU = dQ - dW \qquad \text{(first law of thermodynamics, infinitesimal process)} \qquad (19.6)$$

For the systems we will discuss, the work $dW$ is given by $dW = p\,dV$, so we can also state the first law as

$$dU = dQ - p\,dV \qquad (19.7)$$

# Sample Problem

**Example 19.2** **Working off your dessert**

You propose to climb several flights of stairs to work off the energy you took in by eating a 900-calorie hot fudge sundae. How high must you climb? Assume that your mass is 60.0 kg.

**EXECUTE:** From the first law of thermodynamics, $\Delta U = 0 = Q - W$, so $W = mgh = Q$. Hence you must climb to height $h = Q/mg$. First convert units: $Q = (900 \text{ kcal})(4186 \text{ J}/1 \text{ kcal}) = 3.77 \times 10^6$ J. Then

$$h = \frac{Q}{mg} = \frac{3.77 \times 10^6 \text{ J}}{(60.0 \text{ kg})(9.80 \text{ m/s}^2)} = 6410 \text{ m}$$

# 19.4 Internal Energy and the First Law of Thermodynamics

Isolated system

Another special case occurs in an *isolated system,* one that does no work on its surroundings and has no heat flow to or from its surroundings. For any process taking place in an isolated system,

$$W = Q = 0$$

and therefore

$$U_2 = U_1 = \Delta U = 0$$

In other words, *the internal energy of an isolated system is constant.*

# 19.5 Kinds of Thermodynamic Processes

## Adiabatic Process

An **adiabatic process** (pronounced "ay-dee-ah-*bat*-ic") is defined as one with no heat transfer into or out of a system; $Q = 0$. We can prevent heat flow either by surrounding the system with thermally insulating material or by carrying out the process so quickly that there is not enough time for appreciable heat flow. From the first law we find that for every adiabatic process,

$$U_2 - U_1 = \Delta U = -W \qquad \text{(adiabatic process)} \qquad (19.8)$$



Open a champagne

# 19.5 Kinds of Thermodynamic Processes

## Isochoric Process

An **isochoric process** (pronounced "eye-so-*kor*-ic") is a *constant-volume* process. When the volume of a thermodynamic system is constant, it does no work on its surroundings. Then $W = 0$ and

$$U_2 - U_1 = \Delta U = Q \qquad \text{(isochoric process)} \qquad (19.9)$$



Pressure cooker

# 19.5 Kinds of Thermodynamic Processes

## Isobaric Process

An **isobaric process** (pronounced "eye-so-*bear*-ic") is a *constant-pressure* process. In general, none of the three quantities $\Delta U$, $Q$, and $W$ is zero in an isobaric process, but calculating $W$ is easy nonetheless. From Eq. (19.3),

$$W = p(V_2 - V_1) \qquad \text{(isobaric process)} \qquad (19.10)$$
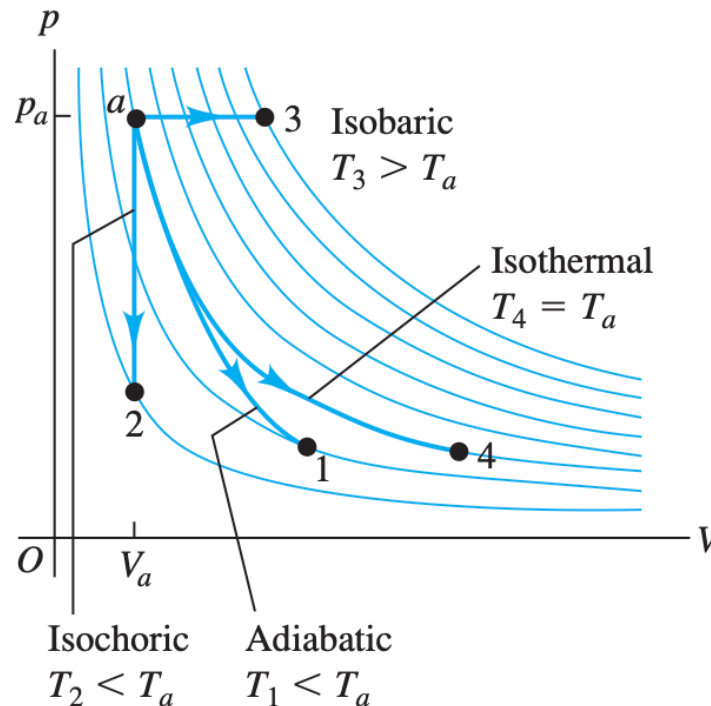
Cooking

# 19.5 Kinds of Thermodynamic Processes

## Isothermal Process

An **isothermal process** is a *constant-temperature* process. For a process to be isothermal, any heat flow into or out of the system must occur slowly enough that thermal equilibrium is maintained. In general, none of the quantities $\Delta U$, $Q$, or $W$ is zero in an isothermal process.
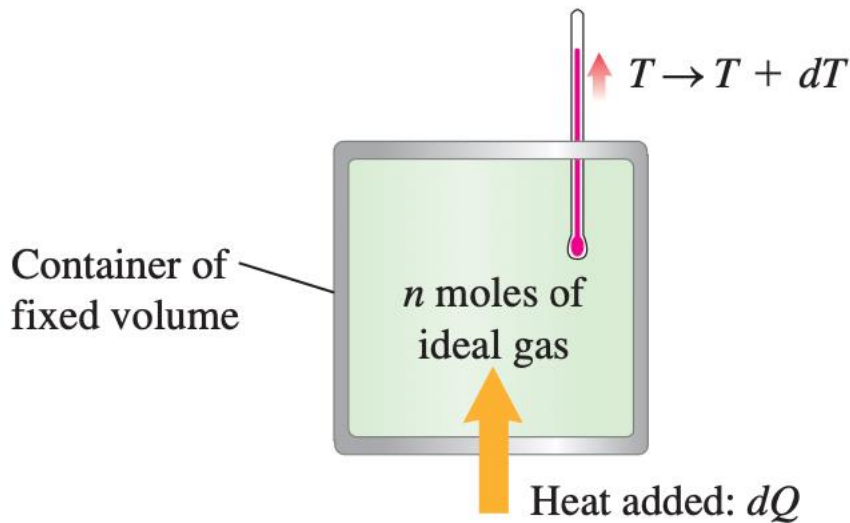
# 19.6 Internal Energy of an Ideal Gas
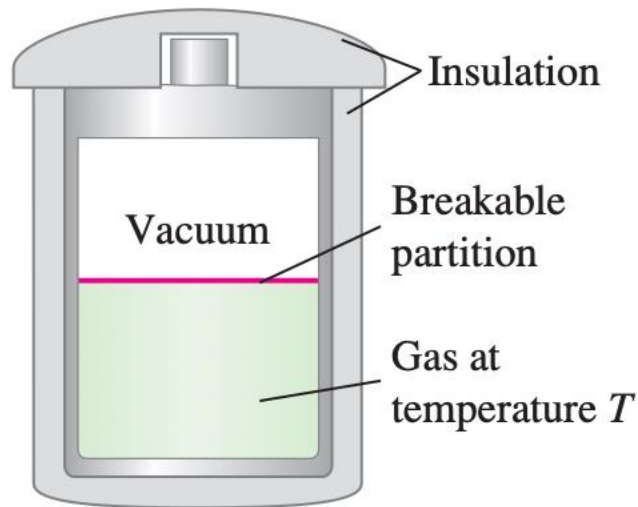


Many experiments have shown that when a low-density gas undergoes a free expansion, its temperature *does not* change.
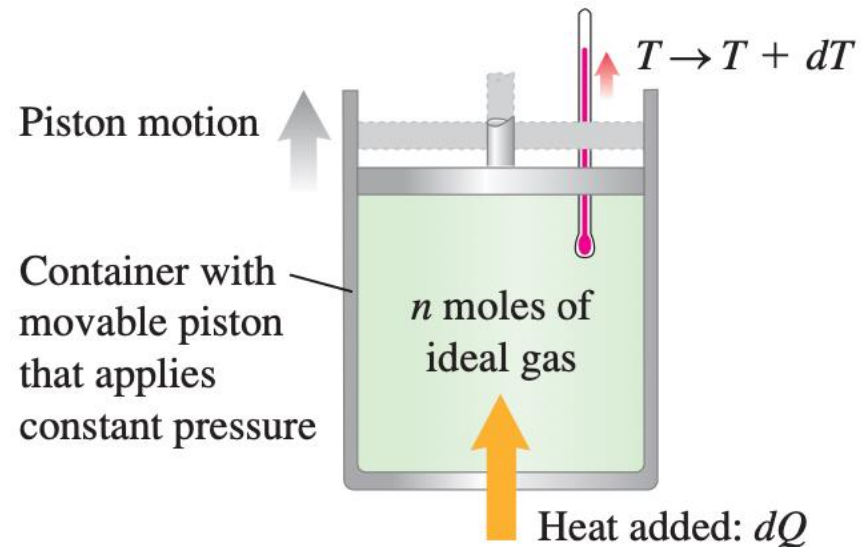
**The internal energy of an ideal gas depends only on its temperature, not on its pressure or volume.**

# 19.7 Heat Capacities of an Ideal Gas

**(a)** Constant volume: $dQ = nC_V \, dT$

**(b)** Constant pressure: $dQ = nC_p \, dT$

Container of fixed volume

$T \rightarrow T + dT$

$n$ moles of ideal gas

Heat added: $dQ$

Piston motion

Container with movable piston that applies constant pressure

$T \rightarrow T + dT$

$n$ moles of ideal gas

Heat added: $dQ$

Measuring the molar heat capacity of an ideal gas (a) at constant volume and (b) at constant pressure.

# 19.7 Heat Capacities of an Ideal Gas

## The Ratio of Heat Capacities

The last column of Table 19.1 lists the values of the dimensionless **ratio of heat capacities,** $C_p/C_V$, denoted by $\gamma$ (the Greek letter gamma):

$$\gamma = \frac{C_p}{C_V} \qquad \text{(ratio of heat capacities)} \qquad \text{(19.18)}$$

(This is sometimes called the "ratio of specific heats.") For gases, $C_p$ is always greater than $C_V$ and $\gamma$ is always greater than unity. This quantity plays an important role in *adiabatic* processes for an ideal gas, which we will study in the next section.

$$C_p = C_V + R \qquad \text{(molar heat capacities of an ideal gas)}$$

# Sample Problem

Example 19.6 **Cooling your room**

A typical dorm room or bedroom contains about 2500 moles of air. Find the change in the internal energy of this much air when it is cooled from 35.0°C to 26.0°C at a constant pressure of 1.00 atm. Treat the air as an ideal gas with $\gamma = 1.400$.
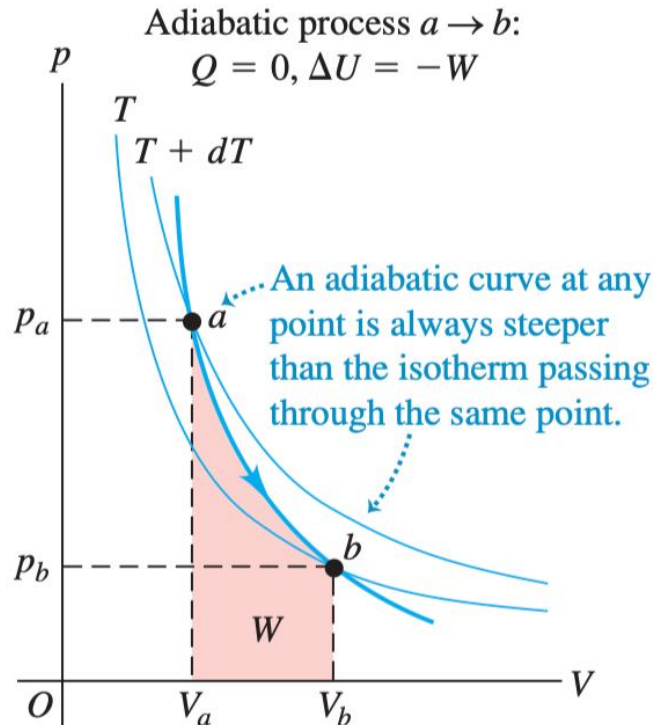
**EXECUTE:** From Eqs. (19.17) and (19.18),

$$\gamma = \frac{C_p}{C_V} = \frac{C_V + R}{C_V} = 1 + \frac{R}{C_V}$$

$$C_V = \frac{R}{\gamma - 1} = \frac{8.314 \text{ J/mol} \cdot \text{K}}{1.400 - 1} = 20.79 \text{ J/mol} \cdot \text{K}$$

Then from Eq. (19.13),

$$\Delta U = nC_V \, \Delta T$$

$$= (2500 \text{ mol})(20.79 \text{ J/mol} \cdot \text{K})(26.0°C - 35.0°C)$$

$$= -4.68 \times 10^5 \text{ J}$$

# 19.8 Adiabatic Processes for an Ideal Gas

Adiabatic process $a \rightarrow b$:
$$Q = 0, \Delta U = -W$$

$p$

$T$

$T + dT$

An adiabatic curve at any point is always steeper than the isotherm passing through the same point.

$p_a$ — $a$

$p_b$ — $b$

$W$

$O$   $V_a$   $V_b$   $V$

Idea gas

$$p_1 V_1^{\gamma} = p_2 V_2^{\gamma}$$

$$T_1 V_1^{\gamma-1} = T_2 V_2^{\gamma-1}$$

**CAUTION**   **"Heating" and "cooling" without heat**  Keep in mind that when we talk about "adiabatic heating" and "adiabatic cooling," we really mean "raising the temperature" and "lowering the temperature," respectively. In an adiabatic process, the temperature change is due to work done by or on the system; there is *no* heat flow at all. ▌

# 19.8 Adiabatic Processes for an Ideal Gas

We can also calculate the *work* done by an ideal gas during an adiabatic process. We know that $Q = 0$ and $W = -\Delta U$ for *any* adiabatic process. For an ideal gas, $\Delta U = nC_V(T_2 - T_1)$. If the number of moles $n$ and the initial and final temperatures $T_1$ and $T_2$ are known, we have simply

$$W = nC_V(T_1 - T_2) \qquad \text{(adiabatic process, ideal gas)} \qquad (19.25)$$

We may also use $pV = nRT$ in this equation to obtain

$$W = \frac{C_V}{R}(p_1V_1 - p_2V_2) = \frac{1}{\gamma - 1}(p_1V_1 - p_2V_2) \qquad \text{(adiabatic process, ideal gas)} \qquad (19.26)$$

Expansion: temperature drop, work is positive
Compression: temperature increases, work is negative

# Sample Problem

| **Example 19.7** | **Adiabatic compression in a diesel engine** |

The compression ratio of a diesel engine is 15.0 to 1; that is, air in a cylinder is compressed to $\frac{1}{(15.0)}$ of its initial volume (Fig. 19.21). (a) If the initial pressure is $1.01 \times 10^5$ Pa and the initial temperature is 27°C (300 K), find the final pressure and the temperature after adiabatic compression. (b) How much work does the gas do during the compression if the initial volume of the cylinder is 1.00 L $= 1.00 \times 10^{-3}$ m³? Use the values $C_V = 20.8$ J/mol·K and $\gamma = 1.400$ for air.

Initial volume

Maximum compression

$V_1$

$V_2 = \dfrac{1}{(15.0)} V_1$
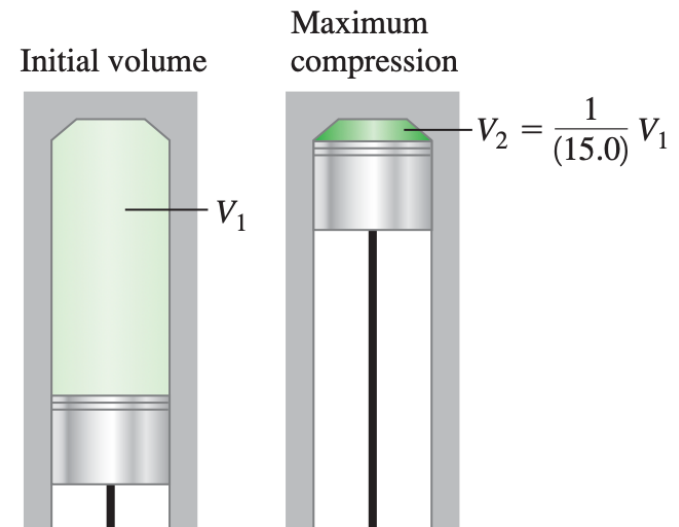
# Sample Problem

**EXECUTE:** (a) From Eqs. (19.22) and (19.24),

$$T_2 = T_1\left(\frac{V_1}{V_2}\right)^{\gamma-1} = (300\text{ K})(15.0)^{0.40} = 886\text{ K} = 613°C$$

$$p_2 = p_1\left(\frac{V_1}{V_2}\right)^{\gamma} = (1.01 \times 10^5\text{ Pa})(15.0)^{1.40}$$

$$= 44.8 \times 10^5\text{ Pa} = 44\text{ atm}$$

(b) From Eq. (19.26), the work done is

$$W = \frac{1}{\gamma - 1}(p_1 V_1 - p_2 V_2)$$

Using $V_1/V_2 = 15.0$, this becomes

$$W = \frac{1}{1.400 - 1}\left[\begin{array}{l}(1.01 \times 10^5\text{ Pa})(1.00 \times 10^{-3}\text{ m}^3) \\ -(44.8 \times 10^5\text{ Pa})\left(\dfrac{1.00 \times 10^{-3}\text{ m}^3}{15.0}\right)\end{array}\right]$$

$$= -494\text{ J}$$

Initial volume    Maximum compression



$-V_2 = \dfrac{1}{(15.0)}V_1$

$-V_1$