Chapter 18

Temperature, Heat, and The First Law Of Thermodynamics

Thermodynamics

Thermodynamics is the study and applications of the thermal energy (or internal energy) of systems.

1. Temperature

Temperature is one of the seven SI base quantities. In physics, we measure temperature on the **kelvin scale**.

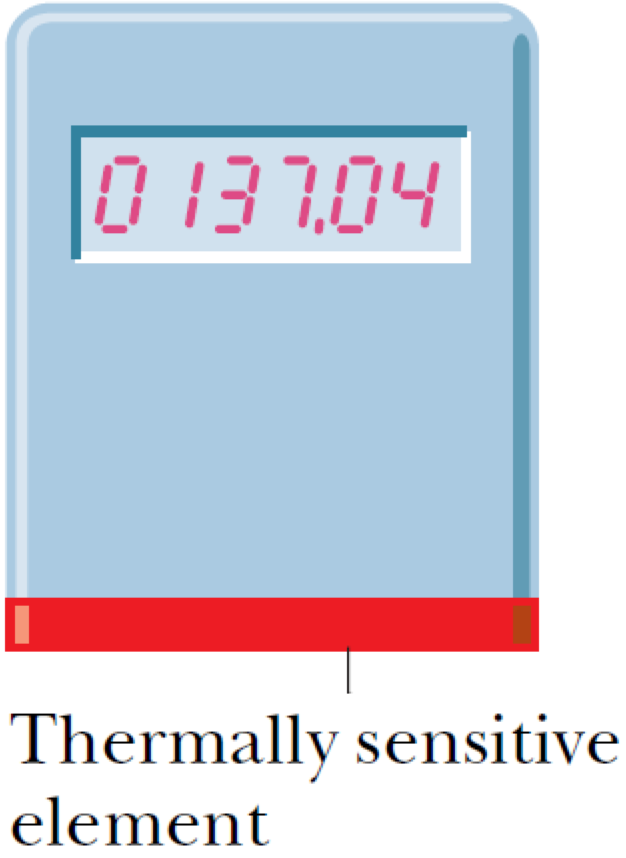
The lower limit of temperature is taken as the zero of the kelvin scale ($0\ \mathrm{K}$). Room temperature is nearly $290\ \mathrm{K}$, above the absolute zero.

2. The Zeroth Law of Thermodynamics

The properties of many bodies changes as their temperature changes, such as volume and resistivity.

We can use any of these properties to build an instrument that measures temperature.

The figure shows such an instrument. The number on it increases with temperature increase and decreases when the temperature decreases. The instrument is not calibrated , and the number on it have no physical meaning yet. This device is a **thermoscope** but not a **thermometer**.

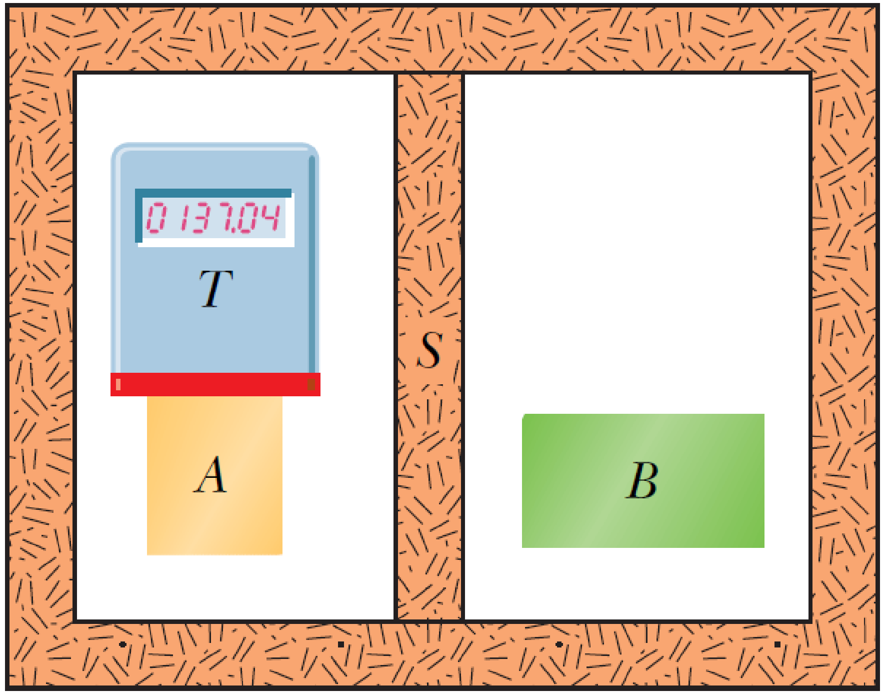


2. The Zeroth Law of Thermodynamics

Suppose that we put the thermoscope (body T) into contact with another body (body A), where the entire system is in an insulating box.

The number on the thermoscope rolls by until it comes to rest (at 137.04, say). The two bodies are now said to be in **thermal equilibrium** with each other.

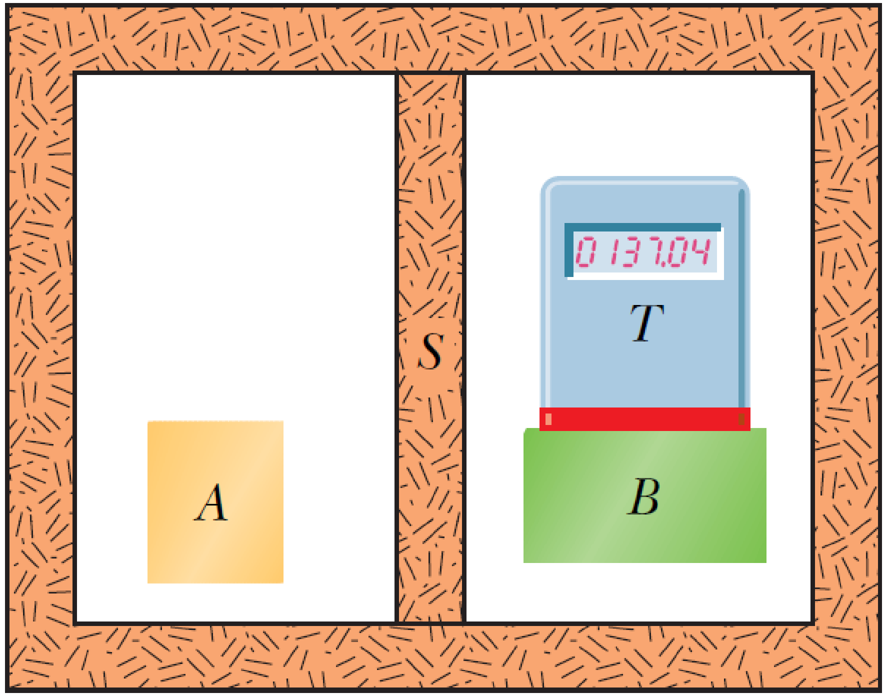
We conclude that bodies T and A must be at the same temperature.



2. The Zeroth Law of Thermodynamics

Suppose that we next put body T into contact with body B, and find that the two bodies come to thermal equilibrium at the same reading (137.04).

We conclude that bodies T and B are at the same temperature.



Are bodies A and B in thermal equilibrium?

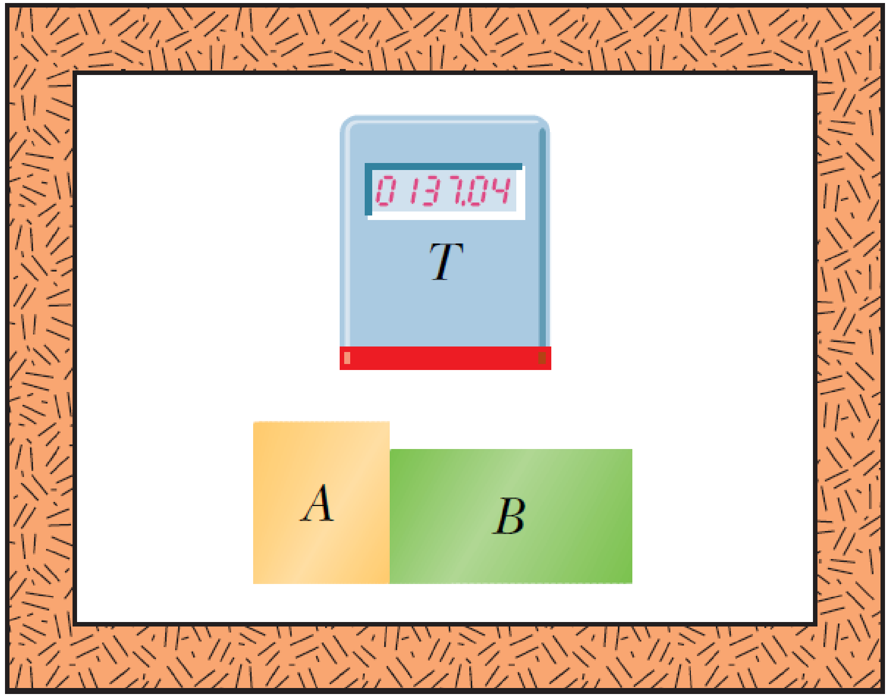
According to experiments, yes.

This experimental fact is summed up in the **zeroth law of thermodynamics**

2. The Zeroth Law of Thermodynamics

“If bodies A and B are each in thermal equilibrium with a third body T, then A and B are in thermal equilibrium with each other.”

In other words, every body has a property called a **temperature**. When two bodies are in thermal equilibrium, their temperature are equal, and vice versa.



3. The Celsius and Fahrenheit Scales

In the Celsius scale (centigrade scale) temperatures are measured in degrees. A Celsius degree has the same size as the kelvin. The zero of the Celsius scale is shifted to a more convenient value. A Celsius temperature T_{C} and a Kelvin temperature T are related by

$$T_{\mathrm{C}} = T - 273.15°.$$

Note that we don’t use the degree symbol ° with Kelvin readings.

The Fahrenheit scale, has a smaller degree that the Celsius scale and a different zero of temperature. The two scales are related by

$$T_{\mathrm{F}} = \frac{9}{5} T_{\mathrm{C}} + 32°.$$

3. The Celsius and Fahrenheit Scales

We use letters C and F with ° on the number to express equivalence of temperature measurements on both scales. For example,

$$0^\circ\mathrm{C} = 32^\circ\mathrm{F},$$

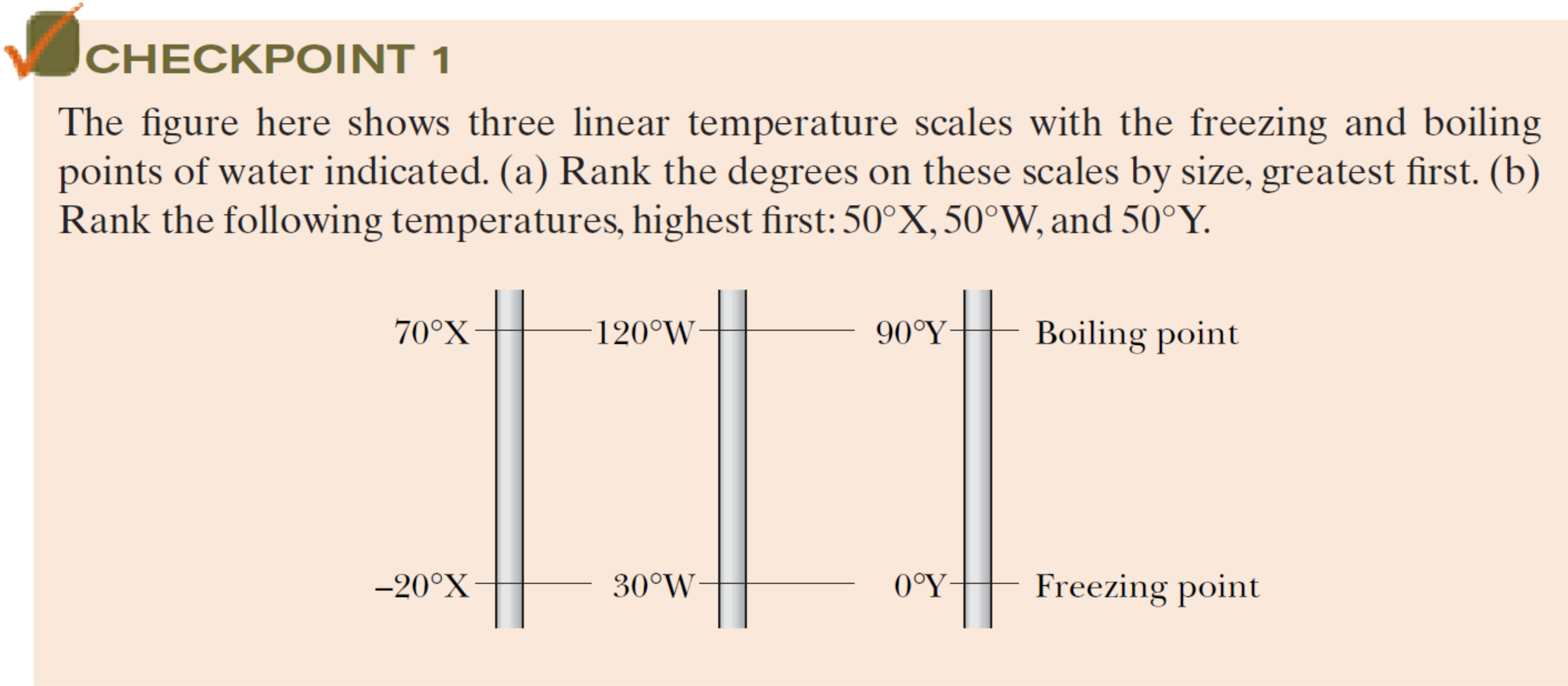
means 0° on the Celsius scale corresponds to 32° on the Fahrenheit scale.

However, we use letters C and F with ° on the letter to express the difference of temperature. For examples,

$$5\ \mathrm{C}^\circ = 9\ \mathrm{F}^\circ,$$

means that a temperature difference of 5 Celsius degrees is equivalent to a temperature difference of 9 Fahrenheit degrees.

3. The Celsius and Fahrenheit Scales



CHECKPOINT 1

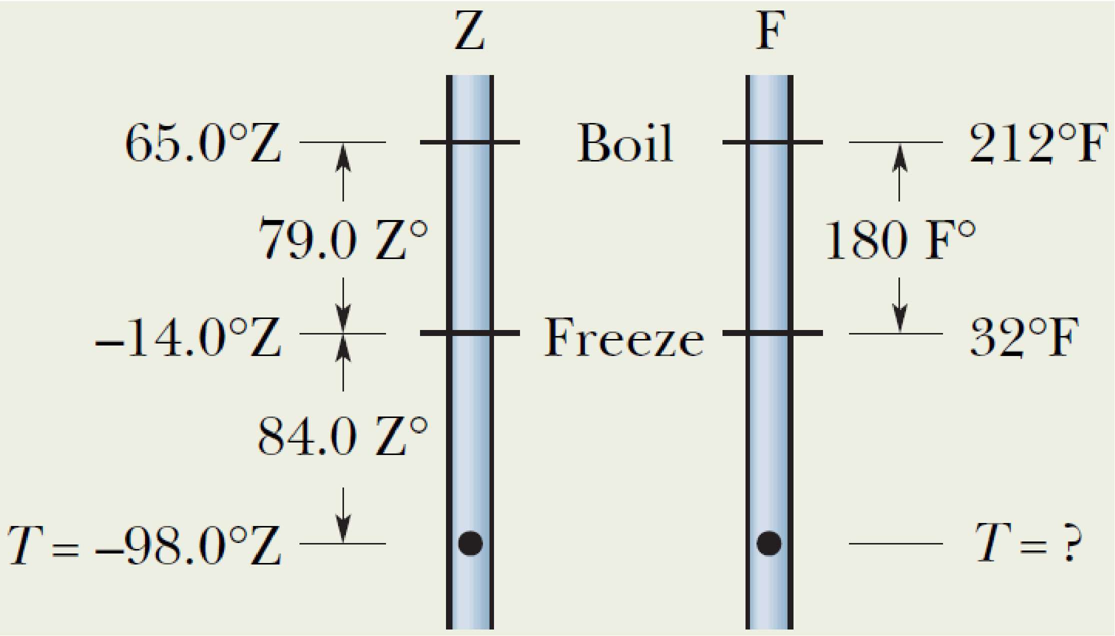
The figure here shows three linear temperature scales with the freezing and boiling points of water indicated. (a) Rank the degrees on these scales by size, greatest first. (b) Rank the following temperatures, highest first: 50°X, 50°W, and 50°Y.

(a) All tie.

(b) 50°X, 50°Y, then 50°W.

3. The Celsius and Fahrenheit Scales

Example 1: Suppose you come across old scientific notes that describe a temperature scale called Z on which the boiling point of water is $65.0°\mathrm{Z}$ and the freezing point is $14°\mathrm{Z}$. To what temperature on the Fahrenheit scale would a temperature of $T = -98.0°\mathrm{Z}$ correspond? Assume that the Z scale is linear; that is, the size of a Z degree is the same everywhere on the Z scale.



3. The Celsius and Fahrenheit Scales

The differences between the boiling and melting points on both scales give that

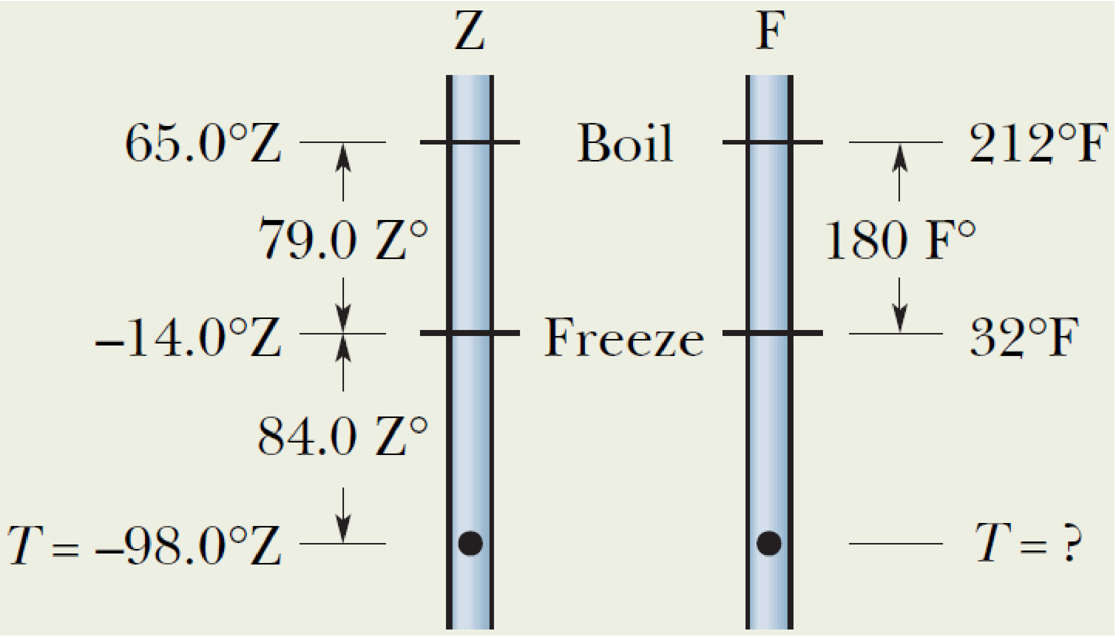
$$79.0\ \mathrm{Z}° = 180.0\ \mathrm{F}°.$$

T is $84.0\ \mathrm{Z}°$ below the freezing point. Therefore, T is below the freezing point of the Fahrenheit scale by

$$\frac{180.0\ \mathrm{F}°}{79.0\ \mathrm{Z}°} 84.0\ \mathrm{Z}° = 191\ \mathrm{F}°.$$

Thus,

$$T = 32.0°\ \mathrm{F} - 191\ \mathrm{F}° = -159°\mathrm{F}.$$



4. Thermal Expansion

Linear Expansion

An increase in the temperature of a metal rod of length L by an amount ΔT results in an increase ΔL in the rod's length, given by

$$\Delta L = L\alpha \, \Delta T,$$

where α is the **coefficient of linear expansion**. Its SI unit is K^{-1} or $\mathrm{C}^{\circ -1}$.

4. Thermal Expansion

Volume Expansion

Because all the dimensions of a solid expand with temperature, the solid's volume expands as well. An increase of ΔT in the temperature of a solid or liquid of volume V results in an increase ΔV in volume given by

$$\Delta V = V\beta\,\Delta T.$$

where β is the **coefficient of volume expansion** of the solid or liquid.

β is related to α by

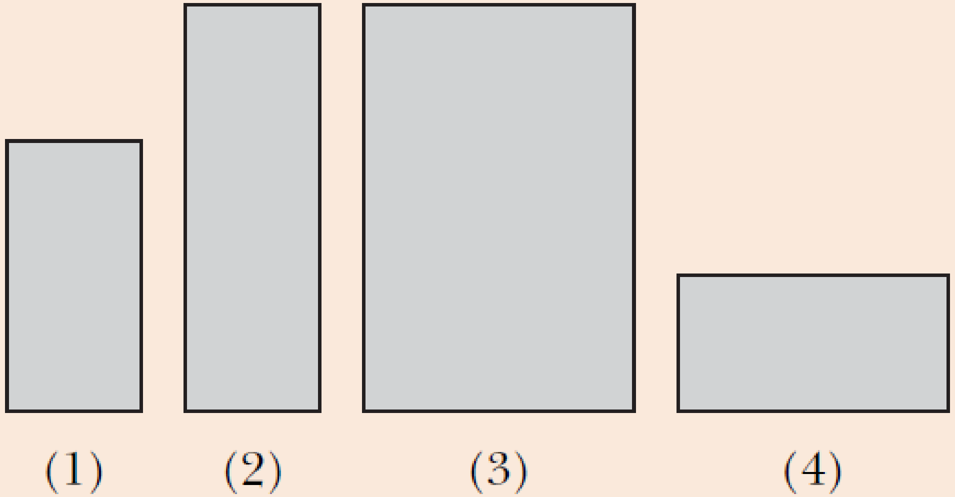
$$\beta = 3\alpha.$$

Water is an exceptional liquid. Between $0°$ and $4°\mathrm{C}$, water expands as the temperature decreases.

4. Thermal Expansion

CHECKPOINT 2

The figure here shows four rectangular metal plates, with sides of L, $2L$, or $3L$. They are all made of the same material, and their temperature is to be increased by the same amount. Rank the plates according to the expected increase in (a) their vertical heights and (b) their areas, greatest first.



(a) 2 & 3, 1, then 4.

(b) 3, 2, then 1 & 4.

$$\Delta L = L\alpha \, \Delta T,$$

4. Thermal Expansion

Example 2: On a hot day in Las Vegas, an oil trucker loaded $37\ 000\ \mathrm{L}$ of diesel fuel. He encountered cold weather on the way to Payson, Utah, where the temperature was $23.0\ \mathrm{K}$ lower than in Las Vegas, and where he delivered his entire load. How many liters did he deliver? The coefficient of volume expansion for diesel fuel is $9.50 \times 10^{-4}/\mathrm{C}°$, and the coefficient of linear expansion for his steel truck tank is $11 \times 10^{-6}/\mathrm{C}°$.

The decrease ΔV in the volume of the fuel is

$$\Delta V = V\beta\,\Delta T = (37\ 000\ \mathrm{L})(9.50\ \times 10^{-4}/\mathrm{C}°)(-23.0\ \mathrm{K}) = -808\ \mathrm{L}.$$

Thus, the amount he delivered is

$$V_{\mathrm{del}} = V + \Delta V = 37\ 000\ \mathrm{L} - 808\ \mathrm{L} = 36\ 190\ \mathrm{L}.$$

5. Temperature and Heat

Consider a system with temperature T_S in an **environment** of temperature T_E. If T_S is not equal to T_E then T_S will change (T_E can change too) until $T_S = T_E$.

The change in temperature is due to a change of the thermal energy of the system because of energy transfer between the system and environment.

The transferred energy is called **heat** Q.

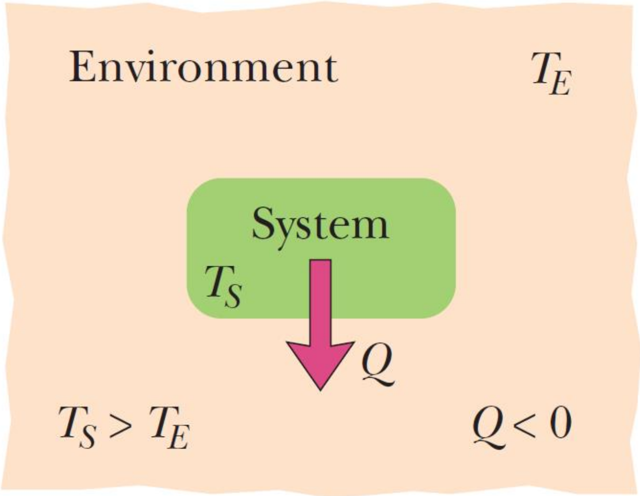
Q is positive when energy is transferred to (absorbed by) a system, and negative when energy is transferred from (released by) a system.

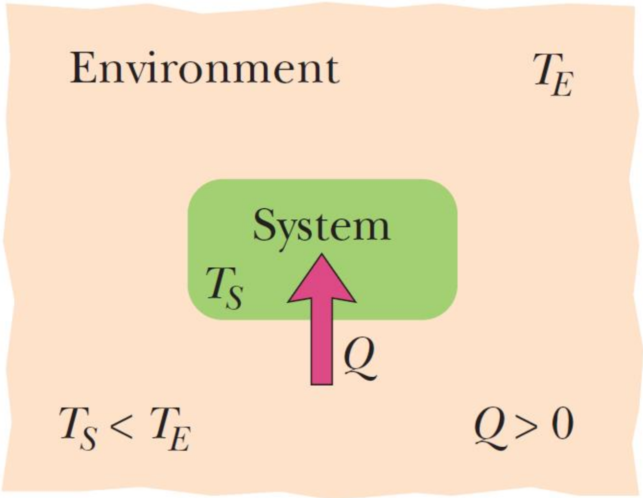
5. Temperature and Heat

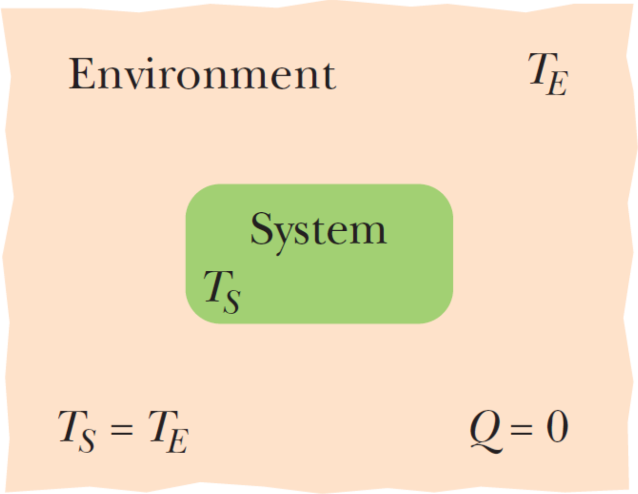
When $T_S > T_E$, energy is transferred from the system to the environment and thus $Q < 0$.

When $T_S < T_E$, energy is transferred to the system from the environment and thus $Q > 0$.

When $T_S = T_E$, there is no energy transfer and $Q = 0$.







5. Temperature and Heat

Heat is defined as the energy transferred between a system and its environment because of a temperature difference that exists between them.

Remember that energy can also be transferred between a system and its environment as a work W.

Heat and work are not intrinsic properties of the system. They only describe the transfer of energy into or out of a system.

The SI unit of heat is the **joule**, since heat is energy.

Other common units of heat are the **British thermal units** (Btu) and the **calorie**, where

$$1\ \mathrm{cal} = 3.968 \times 10^{-3}\mathrm{Btu} = 4.1868\ \mathrm{J}.$$

6. The Absorption of Heat by Solids & Liquids

Heat Capacity

The **heat capacity** C is the proportionality constant between the heat Q absorbed or released and the resulting temperature change ΔT:

$$Q = C\Delta T = C\left(T_f - T_i\right).$$

C has the unit of energy per degree or energy per kelvin.

Warning: The word “capacity” has nothing to do with the amount of energy contained or with an upper limit of energy transfer.

6. The Absorption of Heat by Solids & Liquids

Specific Heat

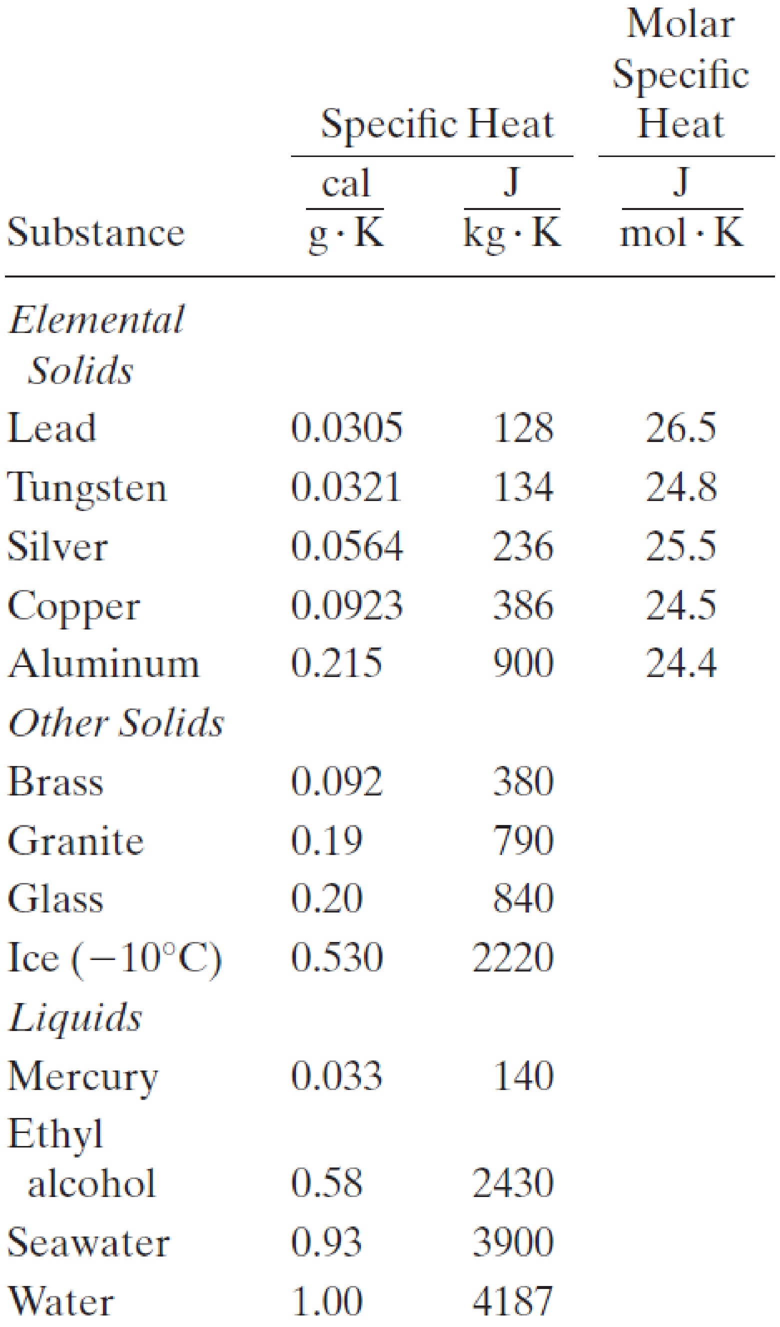
Two objects made of the same material will have heat capacities proportional to their masses. Thus, it is more convenient to define a “heat capacity per unit mass” or **specific heat** c.

In terms of the specific heat c, we write

$$Q = cm\Delta T = cm(T_f - T_i).$$

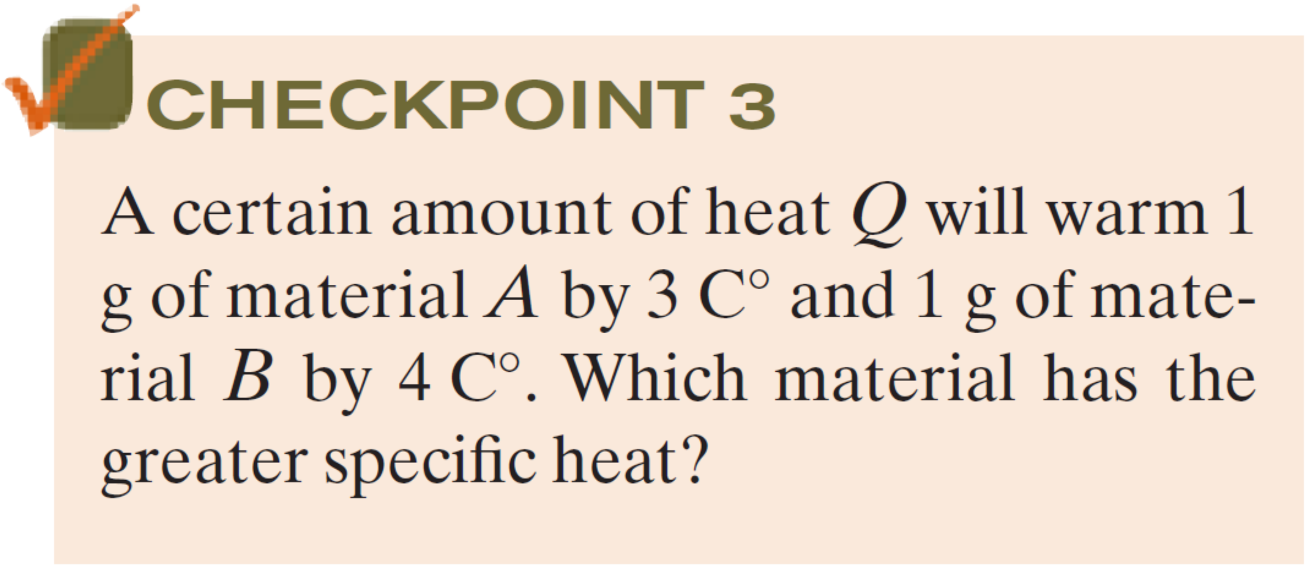
The table shows the specific heats of some substances at room temperature. For example, for water,

$$c = 1.00\,\frac{\text{cal}}{\text{g}\cdot\text{c}^\circ} = 1.00\,\frac{\text{Btu}}{\text{lb}\cdot\text{F}^\circ} = 4186.8\,\frac{\text{J}}{\text{kg}\cdot\text{K}}$$



	Specific Heat		Molar Specific Heat
Substance	$\frac{\text{cal}}{\text{g}\cdot\text{K}}$	$\frac{\text{J}}{\text{kg}\cdot\text{K}}$	$\frac{\text{J}}{\text{mol}\cdot\text{K}}$
Elemental Solids			
Lead	0.0305	128	26.5
Tungsten	0.0321	134	24.8
Silver	0.0564	236	25.5
Copper	0.0923	386	24.5
Aluminum	0.215	900	24.4
Other Solids			
Brass	0.092	380	
Granite	0.19	790	
Glass	0.20	840	
Ice (−10°C)	0.530	2220	
Liquids			
Mercury	0.033	140	
Ethyl alcohol	0.58	2430	
Seawater	0.93	3900	
Water	1.00	4187	

6. The Absorption of Heat by Solids & Liquids



CHECKPOINT 3

A certain amount of heat Q will warm 1 g of material A by 3 C° and 1 g of material B by 4 C°. Which material has the greater specific heat?

Material *A*

$$Q = cm\Delta T$$

$$c = \frac{Q}{m\Delta T}$$

6. The Absorption of Heat by Solids & Liquids

Molar Specific Heat

Sometimes it is more convenient to specify an amount of substance in mole (mol), where

$$1\ \text{mol} = 6.02 \times 10^{23}\ \text{elementary units.}$$

When quantities are expressed in moles, **molar specific heats** must replace specific heats. In terms of the molar specific heat c_{mol}, we write

$$Q = c_{\text{mol}} n \Delta T = c_{\text{mol}} n \left(T_f - T_i\right),$$

where n is the number of moles in the sample.

6. The Absorption of Heat by Solids & Liquids

Heat of Transformation

When heat is absorbed by a solid or liquid sample, the temperature of the sample does not necessarily rise. The sample may change **phase** instead.

To melt a solid into a liquid, energy is required to free the molecules from their rigid structure. The process is reversed in freezing a liquid.

To vaporize a liquid into a gas, like melting, energy is required to free the molecules from their clusters. The process is reversed in condensing a gas.

The amount of heat per unit mass required to completely change the phase of a sample is called the **heat of transformation** L. Thus, when a sample of mass m completely undergoes a phase change, the total energy transferred is

$$Q = Lm.$$

6. The Absorption of Heat by Solids & Liquids

Heat of Transformation

When the phase change is from liquid to gas or from gas to liquid, L is called the heat of vaporization L_V. For water at its normal boiling or condensation temperature,

$$L_V = 539\,\frac{\mathrm{cal}}{\mathrm{g}} = 40.7\,\frac{\mathrm{kJ}}{\mathrm{mol}} = 2256\,\frac{\mathrm{kJ}}{\mathrm{kg}} = 2256\,\frac{\mathrm{J}}{\mathrm{g}}.$$

When the phase change is from solid to liquid or from liquid to solid, L is called the heat of fusion L_F. For water at its normal freezing or melting temperature,

$$L_F = 79.5\,\frac{\mathrm{cal}}{\mathrm{g}} = 6.01\,\frac{\mathrm{kJ}}{\mathrm{mol}} = 333\,\frac{\mathrm{kJ}}{\mathrm{kg}} = 333\,\frac{\mathrm{J}}{\mathrm{g}}.$$

6. The Absorption of Heat by Solids & Liquids

Heat of Transformation

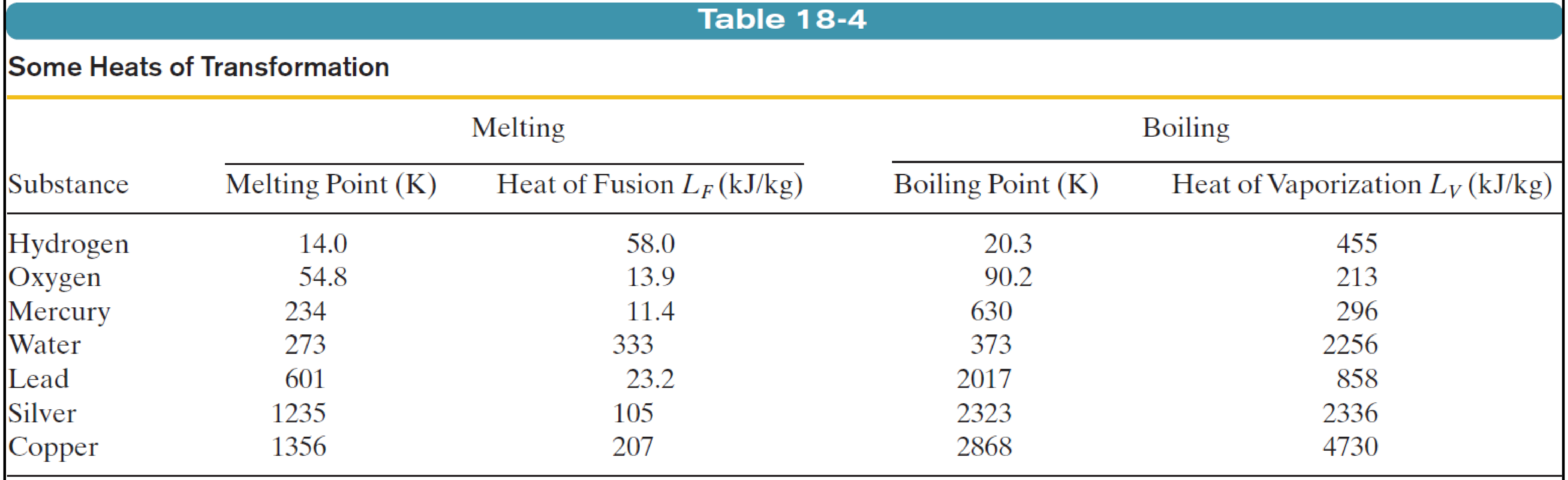
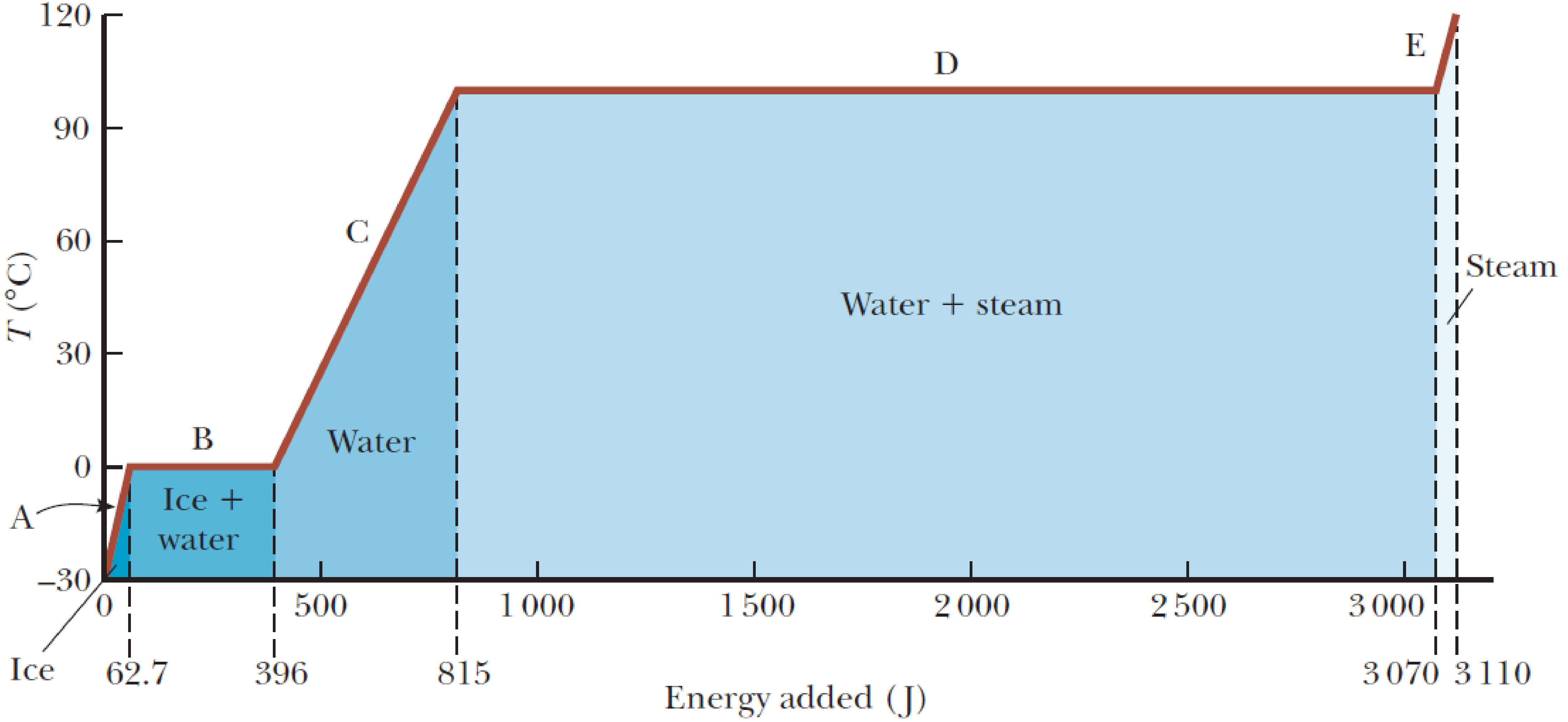


Table 18-4

Some Heats of Transformation

	Melting		Boiling	
Substance	Melting Point (K)	Heat of Fusion L_F (kJ/kg)	Boiling Point (K)	Heat of Vaporization L_V (kJ/kg)
Hydrogen	14.0	58.0	20.3	455
Oxygen	54.8	13.9	90.2	213
Mercury	234	11.4	630	296
Water	273	333	373	2256
Lead	601	23.2	2017	858
Silver	1235	105	2323	2336
Copper	1356	207	2868	4730

6. The Absorption of Heat by Solids & Liquids



6. The Absorption of Heat by Solids & Liquids

Example 3: A copper slug whose mass m_c is 75 g is heated in a laboratory oven to a temperature T of $312°\mathrm{C}$. The slug is then dropped into a glass beaker containing a mass $m_w = 220\ \mathrm{g}$ of water. The heat capacity C_b of the beaker is $45\ \mathrm{cal/K}$. The initial temperature T_i of the water and the beaker is $12°\mathrm{C}$. Assuming that the slug, beaker, and water are an isolated system and the water does not vaporize, find the final temperature T_f of the system at thermal equilibrium.

The heat transfers are related to the temperature changes by

$$Q_w = c_w m\left(T_f - T_i\right),$$

$$Q_b = C_b\left(T_f - T_i\right),$$

$$Q_c = c_c m_c\left(T_f - T_{ic}\right).$$

6. The Absorption of Heat by Solids & Liquids

Because the system is isolated

$$Q_w + Q_b + Q_c = 0.$$

Subsisting we obtain

$$c_w m(T_f - T_i) + C_b(T_f - T_i) + c_c m_c(T_f - T_{ic}) = 0.$$

Solving for T_f gives

$$T_f = \frac{c_c m_c T_{ic} + (C_b + c_w m_w)T_i}{c_w m_w + C_b + c_c m_c}$$

$$= \frac{(0.0923\text{cal/g}\cdot\text{C°})(75\text{ g})(312\text{°C}) + [45\text{cal/K} + (1.00\text{ cal/g}\cdot\text{C°})(220\text{ g})](12\text{°C})}{(1.00\text{ cal/g}\cdot\text{C°})(220\text{ g}) + 45\text{ cal/K} + (0.0923\text{cal/g}\cdot\text{C°})(75\text{ g})}$$

$$= 20\text{°C}.$$

6. The Absorption of Heat by Solids & Liquids

With $T_f = 19.9°\text{C}\ (\approx 20°\text{C})$,

$$Q_w = 1670\ \text{cal},$$
$$Q_b = 342\ \text{cal},$$
$$Q_c = -2020\ \text{cal}.$$

These heats add up to zero, apart form the rounding errors.

6. The Absorption of Heat by Solids & Liquids

Example 4:

(a) How much heat must be absorbed by ice of mass $m = 720\text{ g}$ at $-10°\text{C}$ to take it to the liquid state at $15°\text{C}$?

We need heat Q_1 to bring the ice to its melting temperature:

$$Q_1 = c_{\text{ice}} m\left(T_f - T_i\right) = (2220\text{ J/kg}\cdot\text{K})(0.720\text{ kg})[0 - (-10°\text{C})] = 15.98\text{ kJ}.$$

We need heat Q_2 to melt the ice:

$$Q_2 = L_F m = (333\text{ kJ/kg})(0.720\text{ kg}) = 239.8\text{ kJ}.$$

We need heat Q_3 to heat the water to $15°\text{C}$:

$$Q_3 = c_{\text{w}} m\left(T_f - T_i\right) = (4186.8\text{ J/kg}\cdot\text{K})(0.720\text{ kg})[15°\text{C} - 0] = 45.22\text{ kJ}.$$

The total heat required is therefore

6. The Absorption of Heat by Solids & Liquids

$$Q_1 + Q_2 + Q_3 = 15.98\text{ kJ}+239.8\text{ kJ}+45.22\text{ kJ}=300\text{ kJ}.$$

(b) If we supply the ice with a total energy of only 210 kJ (as heat), what are the final state and temperature of the water?

After heating the ice to the melting point, the remaining heat is

$$Q_{\text{rem}} = 210\text{ kJ} - 15.98\text{ kJ} = 194\text{ kJ}.$$

The mass of the ice melted by the remaining heat is

$$m = \frac{Q_{\text{rem}}}{L_F} = \frac{194\text{ kJ}}{333\text{ kJ/kg}} = 580\text{ g}.$$

The mass of the remaining ice is$720\text{ g} - 580\text{ g} = 140\text{ g}.$

7. A Closer Look at Heat and Work

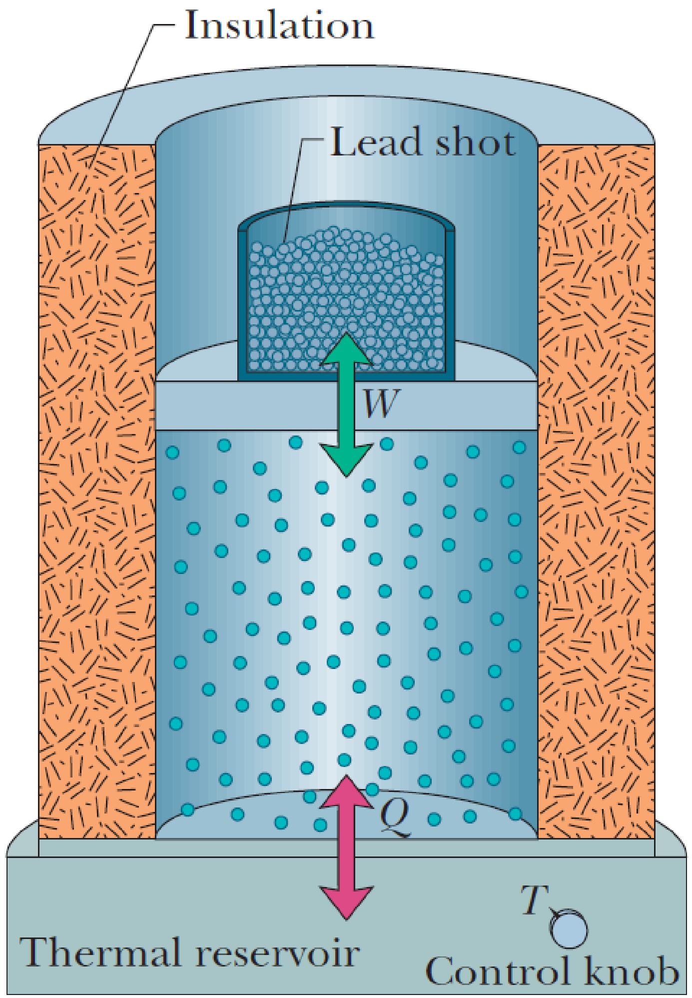
Lets us look in some detail at how energy can be transferred as heat and work between a system and its environment.

Let us consider the system shown in the figure, where a gas is confined to a cylinder with a movable piston.

The pressure multiplied by the piston's area balances the weight of the lead shot.

The walls of the cylinder are insulating.

The bottom of the cylinder is a thermal reservoir of adjustable temperature.

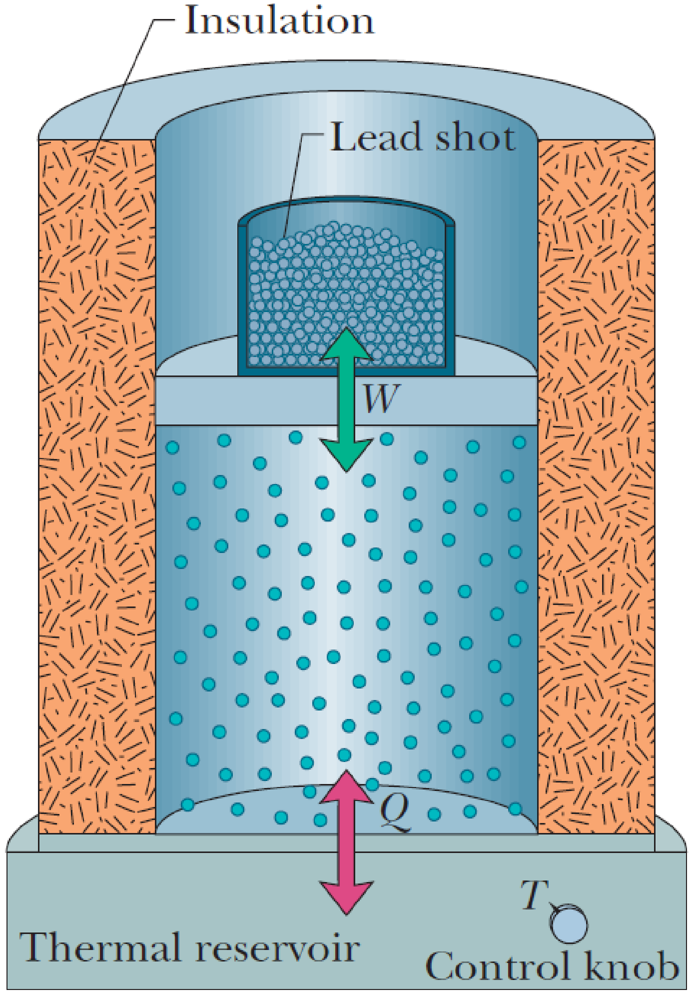


7. A Closer Look at Heat and Work

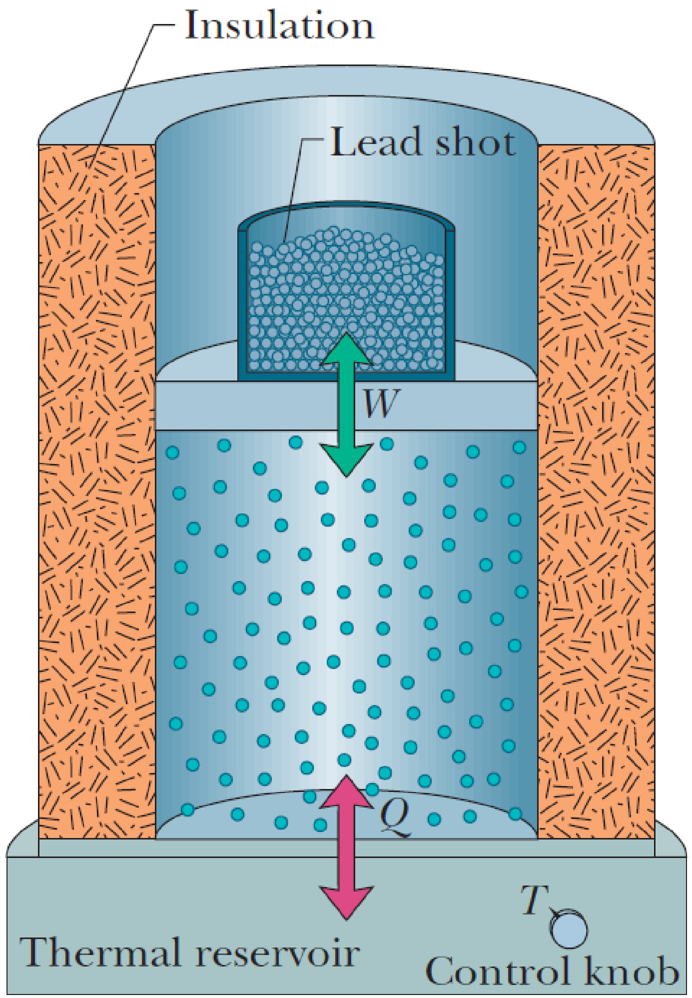
The system (the gas) starts from an **initial state** described by p_i, V_i and T_i.

You want to change the system to a **final state** described by p_f, V_f and T_f.

The process of changing the system from state i to state f is called a **thermodynamic process**.



7. A Closer Look at Heat and Work



During such a process, heat can be transferred to ($Q > 0$) or from ($Q < 0$) the system.

Also, work can be done by the system to raise the piston ($W > 0$) or lower it ($W < 0$).

If you remove a few lead shot, the gas pushes on the piston and move it upward through a differential displacement $d\vec{s}$ with a force $\vec{F} = p\vec{A}$.

The differential work dW done by the gas is then

$$\begin{aligned} dW &= \vec{F} \cdot d\vec{s} = (pA)ds = p(A\,ds) \\ &= p\,dV, \end{aligned}$$

where dV is the differential change in the volume of the gas.

7. A Closer Look at Heat and Work

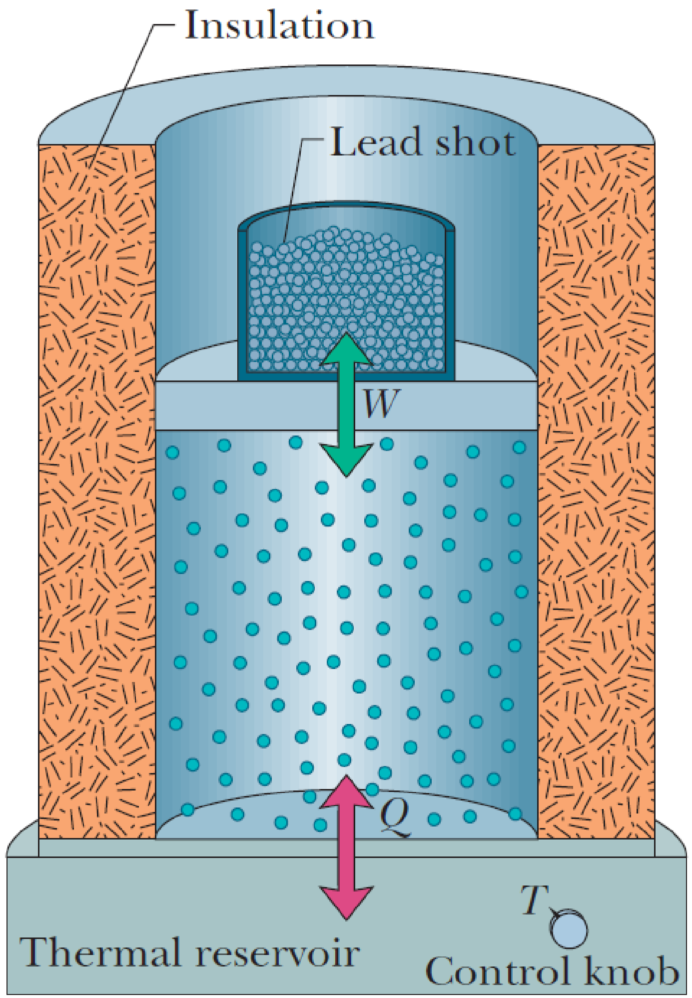
If the gas volume changes from V_i to V_f, the total work done by the gas is

$$W = \int_i^f dW = \int_{V_i}^{V_f} p dV .$$

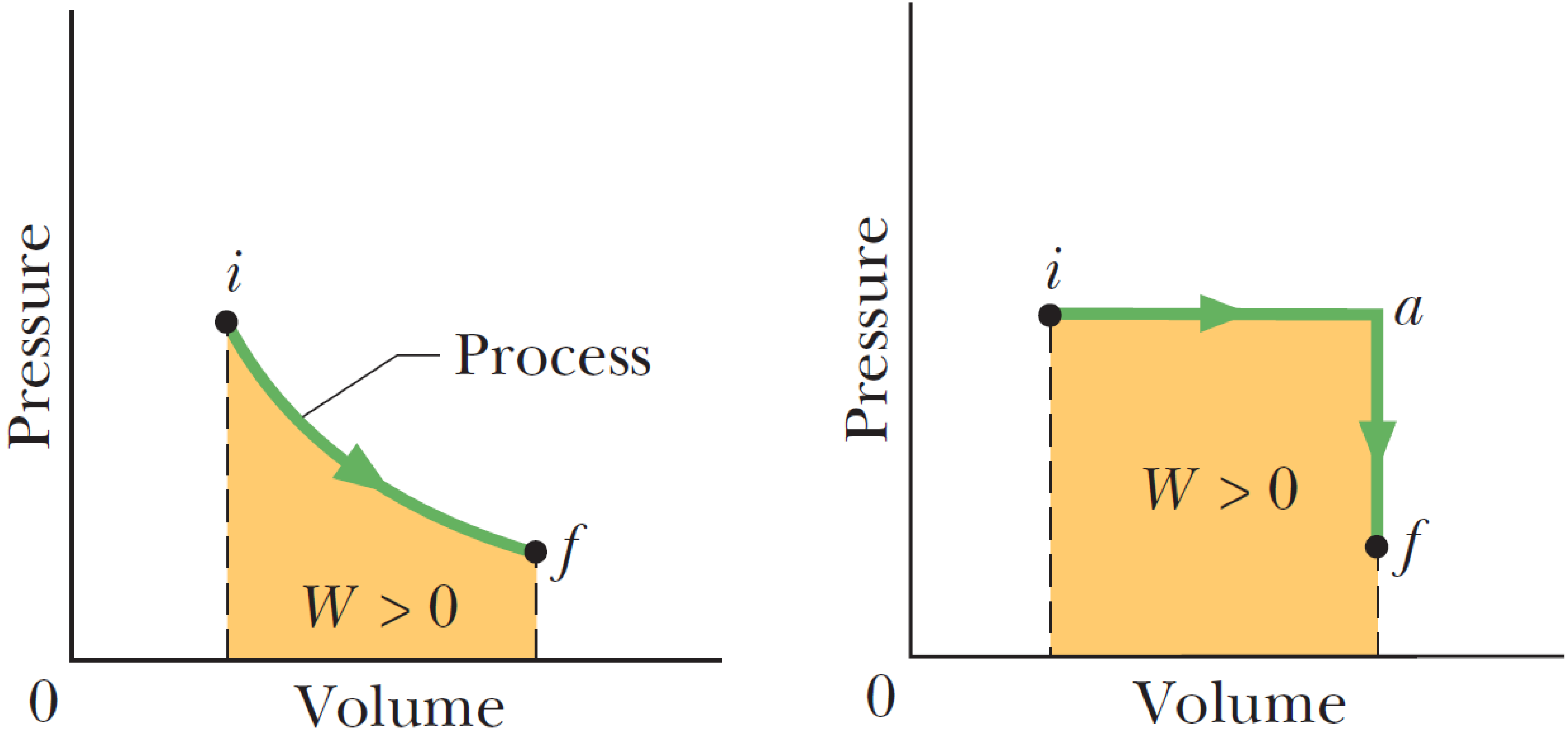
During the volume change, the pressure and temperature may also change.

We need to know how p changes with V (that is $p(V)$) during the process to evaluate the integral.

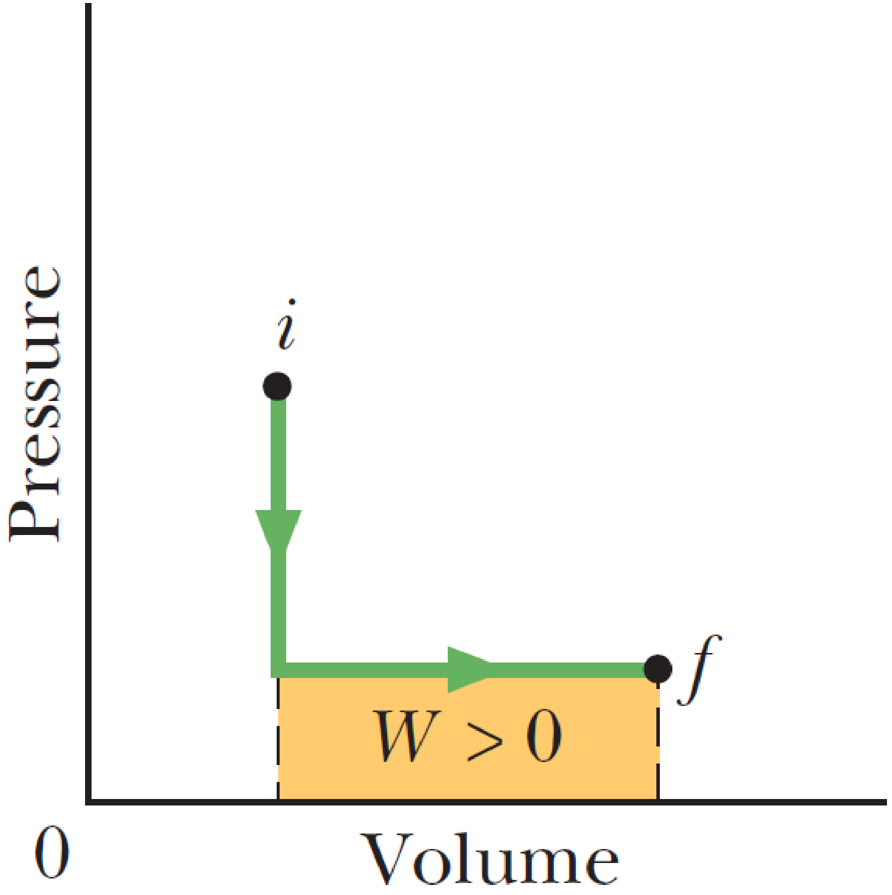
There are many ways to take the system from state i to state f.

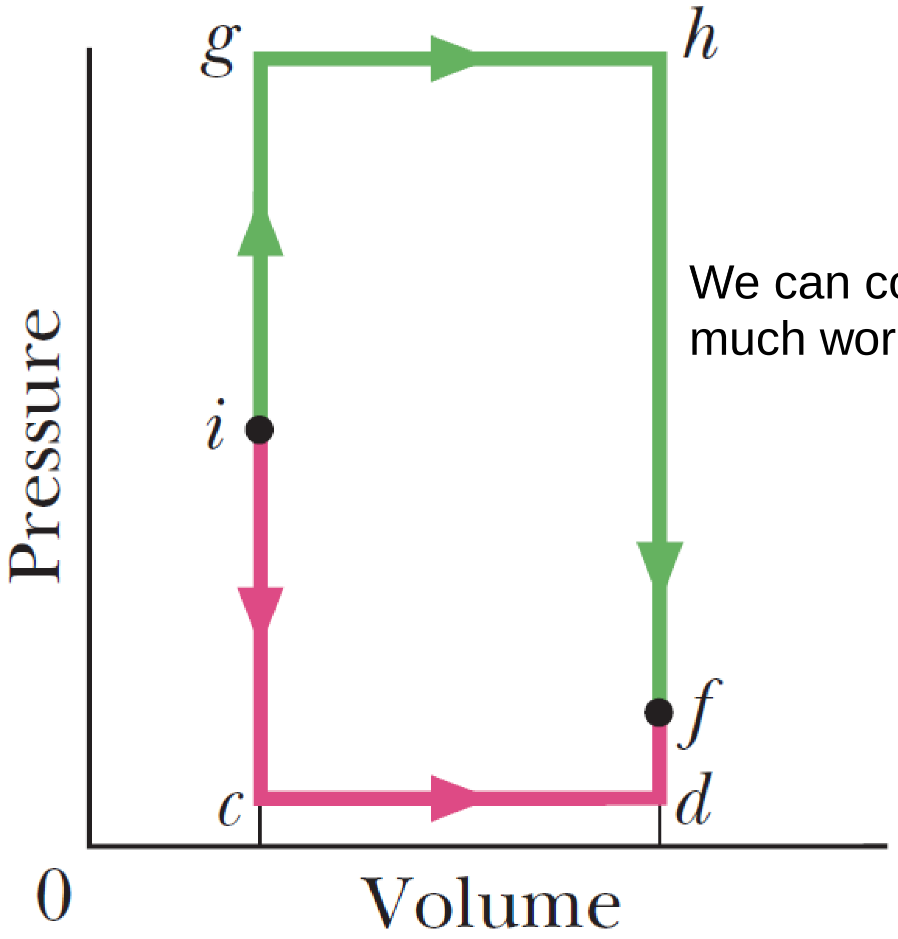


7. A Closer Look at Heat and Work



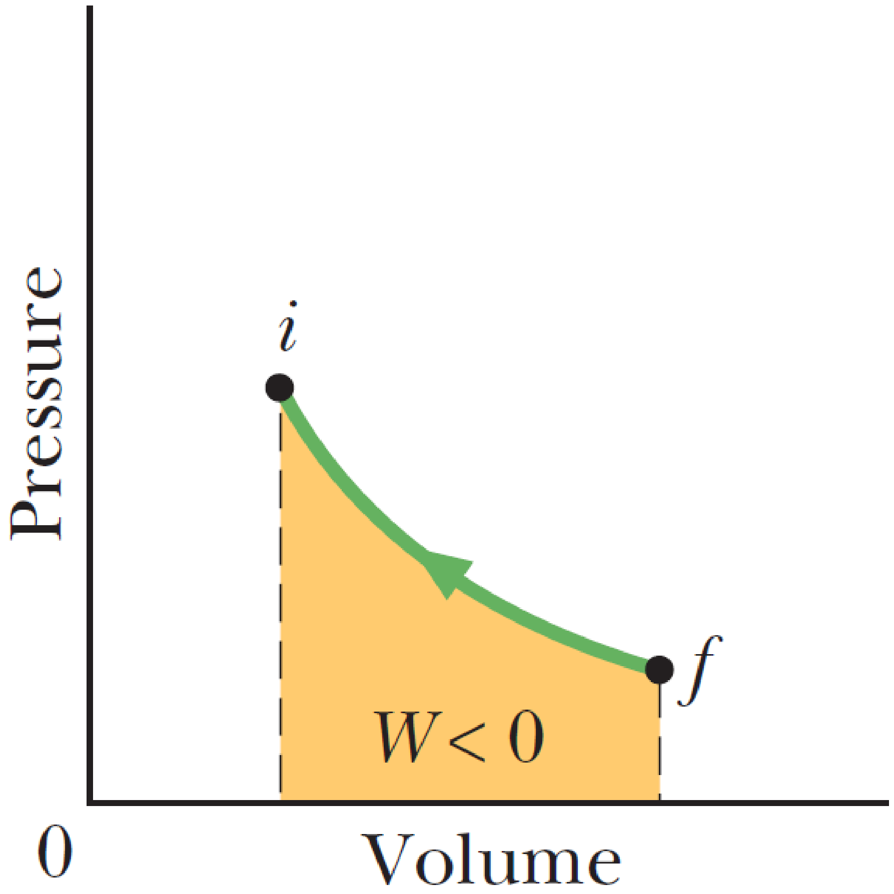
7. A Closer Look at Heat and Work

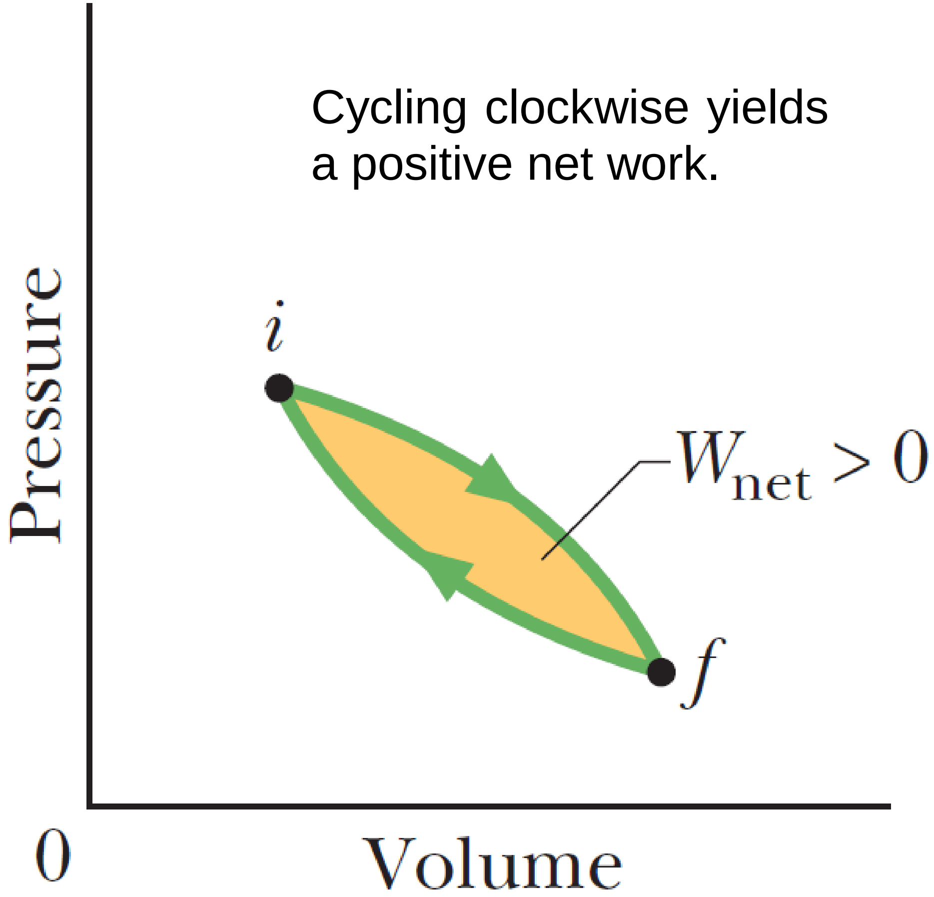




We can control how much work it does.

7. A Closer Look at Heat and Work




Cycling clockwise yields a positive net work.

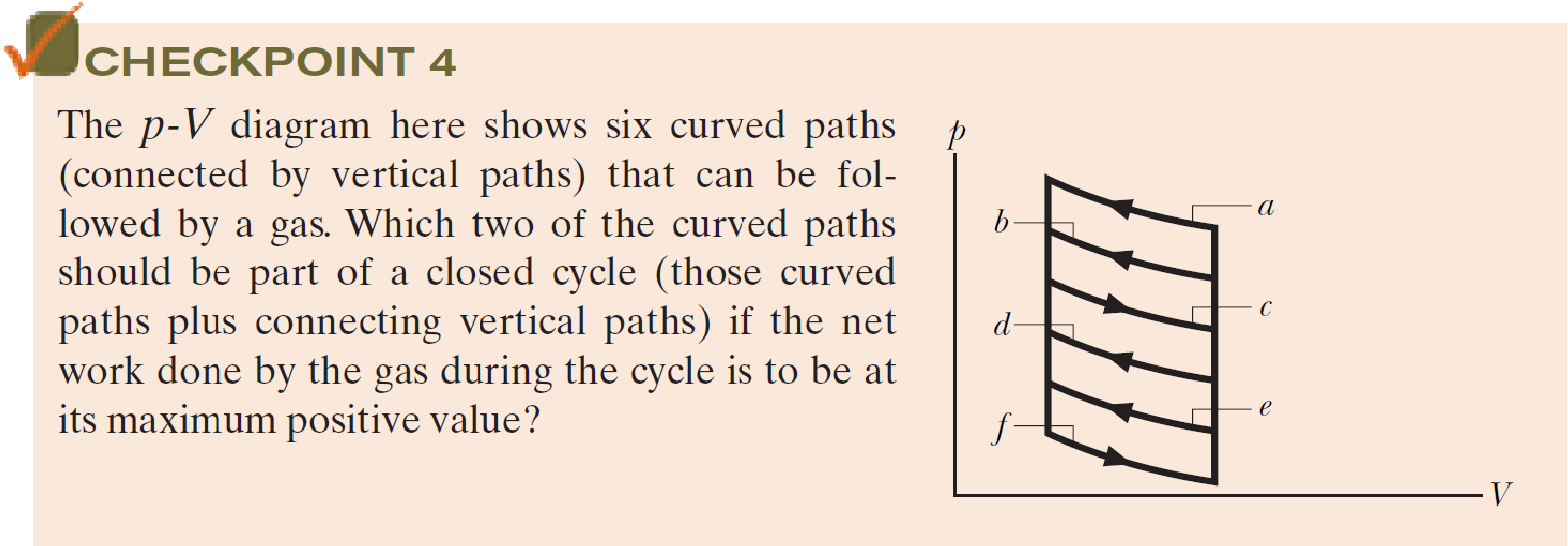
7. A Closer Look at Heat and Work

To sum up, a system can be taken from a given initial state to a given final state by an infinite number of processes.

Heat may or may not be involved.

In general, the work W and and heat Q differ for different processes: W and Q are **path-dependent quantities**.

7. A Closer Look at Heat and Work



CHECKPOINT 4

The p-V diagram here shows six curved paths (connected by vertical paths) that can be followed by a gas. Which two of the curved paths should be part of a closed cycle (those curved paths plus connecting vertical paths) if the net work done by the gas during the cycle is to be at its maximum positive value?

c and e.

8. The First Law of Thermodynamics

We have seen that, when a system changes from state i to state f, both Q and W depend on the nature of the thermodynamic process involved.

However, the quantity $Q - W$ is (experimentally) the same for all processes. It depends only on the initial and final states of the system; it is **path independent**.

The quantity $Q - W$ must represent a change in some intrinsic property of the system. We call this property the internal energy E_{int}. We thus write

$$\Delta E_{\text{int}} = E_{\text{int,f}} - E_{\text{int,i}} = Q - W.$$

This equation is called the **first law of thermodynamics**: "The internal energy E_{int} of a system tends to increase if energy is added as heat Q and tends to decrease if energy is lost as work W done by the system. "

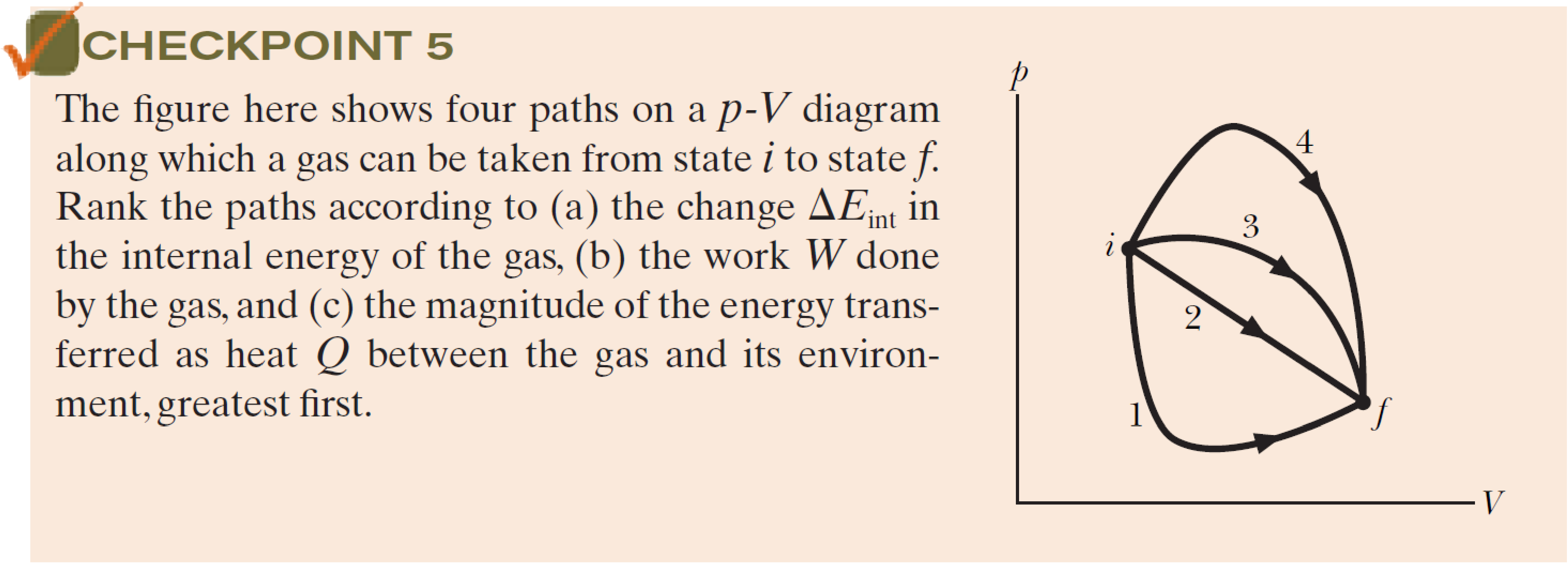
8. The First Law of Thermodynamics

The first law of thermodynamics is an extension of the principle of energy conservation to systems that are not isolated, in which energy can be transferred into or out of the system as either work W or heat Q.

We assume in the first law of thermodynamics that there is no change in the potential energy or kinetic energy of the system $\Delta U = \Delta K = 0$.

Recall that the work W here is the work done by the system.

8. The First Law of Thermodynamics



CHECKPOINT 5

The figure here shows four paths on a p-V diagram along which a gas can be taken from state i to state f. Rank the paths according to (a) the change ΔE_{int} in the internal energy of the gas, (b) the work W done by the gas, and (c) the magnitude of the energy transferred as heat Q between the gas and its environment, greatest first.

(a) All the same

(b) 4, 3, 2, then 1.

(c) 4, 3, 2, then 1.

$$\Delta E_{\text{int}} = Q - W$$

9. Some Special Cases of the First Law of Thermodynamics

Here are four thermodynamics processes:

1. **Adiabatic processes**

An adiabatic process is one that occurs rapidly or in an isolated system so that there is no transfer of energy as heat.

Putting $Q = 0$ in the first law of thermodynamics yields

$$\Delta E_{\text{int}} = -W.$$

The internal energy is decreased by the amount of work done by the system.

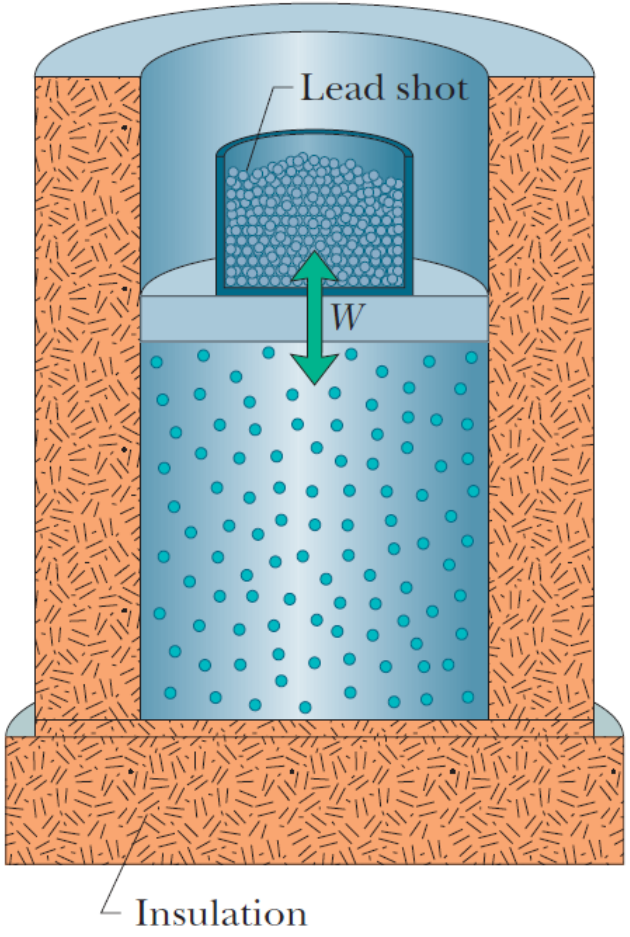
9. Some Special Cases of the First Law of Thermodynamics

1. Adiabatic processes

The figure shows an idealized adiabatic process. The system is isolated; no heat can enter or leave the system.

The only way energy can enter the system is by work. If we remove some lead shot letting the gas expands, positive work is done by the system and the internal energy of the system decreases.

If, instead, we add lead shot and compress the gas, negative work is done by the system and the energy of the system increases.



9. Some Special Cases of the First Law of Thermodynamics

2. Constant-volume (isochoric) processes

If the volume of the system is held constant, $W = 0$.

Putting $W = 0$ in the first law of thermodynamics yields

$$\Delta E_{\mathrm{int}} = Q.$$

The internal energy is increases if it absorbs heat ($Q > 0$), and decreases if the system releases heat ($Q < 0$).

9. Some Special Cases of the First Law of Thermodynamics

3. Cyclical processes

If the system's initial and final states are the same, no intrinsic property of the system is changed and $\Delta E_{\text{int}} = 0$.

The first law of thermodynamics becomes

$$Q = W.$$

The net work done is exactly equal to the net amount of heat transferred.

9. Some Special Cases of the First Law of Thermodynamics

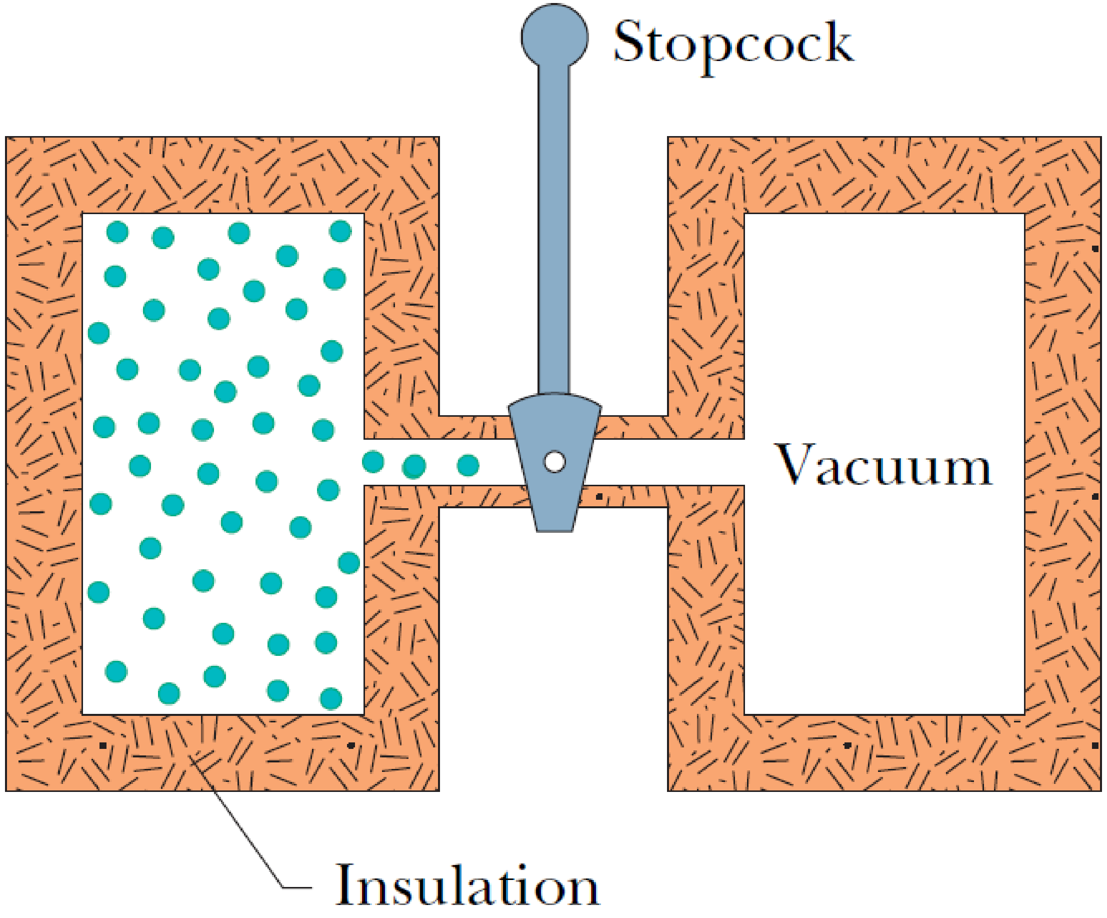
4. Free Expansions

These are processes in which no heat transfer occurs and no work is done on or by the system.

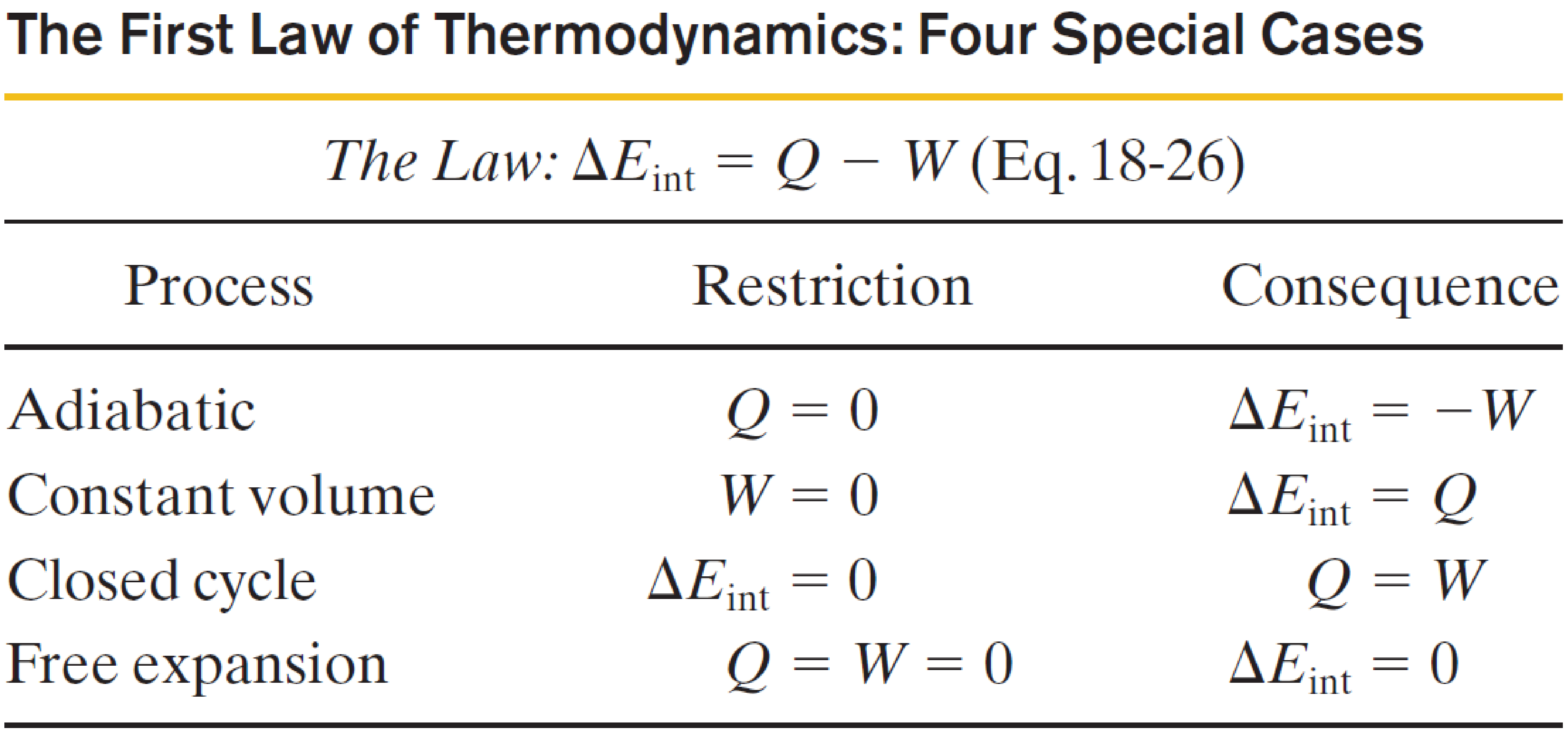
Setting $Q = W = 0$ in the first law of thermodynamics gives

$$\Delta E_{\text{int}} = 0.$$

The figure shows such an expansion.



9. Some Special Cases of the First Law of Thermodynamics



The First Law of Thermodynamics: Four Special Cases

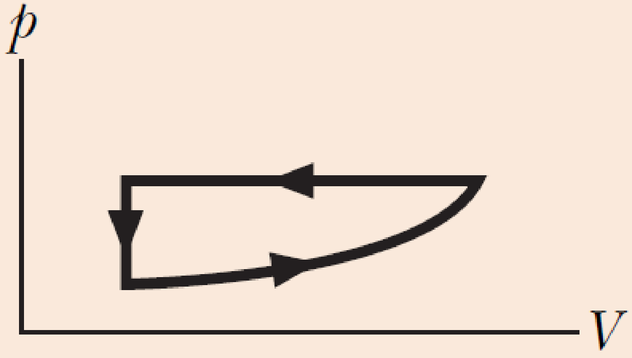
The Law: $\Delta E_{\text{int}} = Q - W$ (Eq. 18-26)

Process	Restriction	Consequence
Adiabatic	$Q = 0$	$\Delta E_{\text{int}} = -W$
Constant volume	$W = 0$	$\Delta E_{\text{int}} = Q$
Closed cycle	$\Delta E_{\text{int}} = 0$	$Q = W$
Free expansion	$Q = W = 0$	$\Delta E_{\text{int}} = 0$

9. Some Special Cases of the First Law of Thermodynamics

CHECKPOINT 6

For one complete cycle as shown in the p-V diagram here, are (a) ΔE_{int} for the gas and (b) the net energy transferred as heat Q positive, negative, or zero?



(a) Zero.

(b) Negative.

$$\Delta E_{\text{int}} = Q - W$$

9. Some Special Cases of the First Law of Thermodynamics

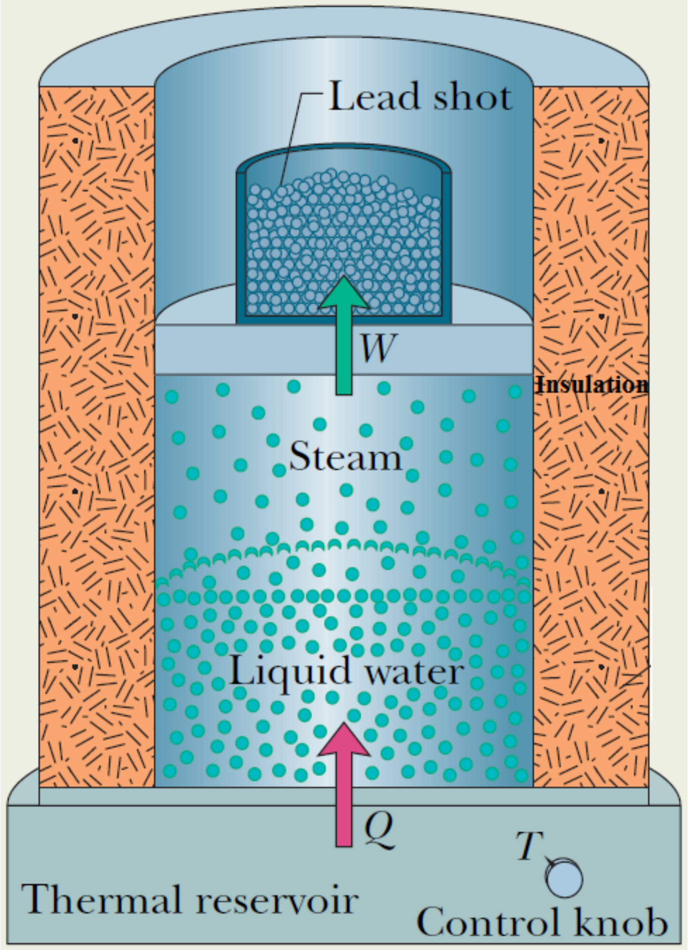
Example 5: Let 1.00 kg of liquid water at 100°C be converted to steam at 100°C by boiling at standard atmospheric pressure (which is 1.00 atm or 1.01×10^5 Pa) in the arrangement of the figure. The volume of that water changes from an initial value of $1.00 \times 10^{-3}\ \mathrm{m}^3$ as a liquid to $1.671\ \mathrm{m}^3$ as steam.

(a) How much work is done by the system during this process?

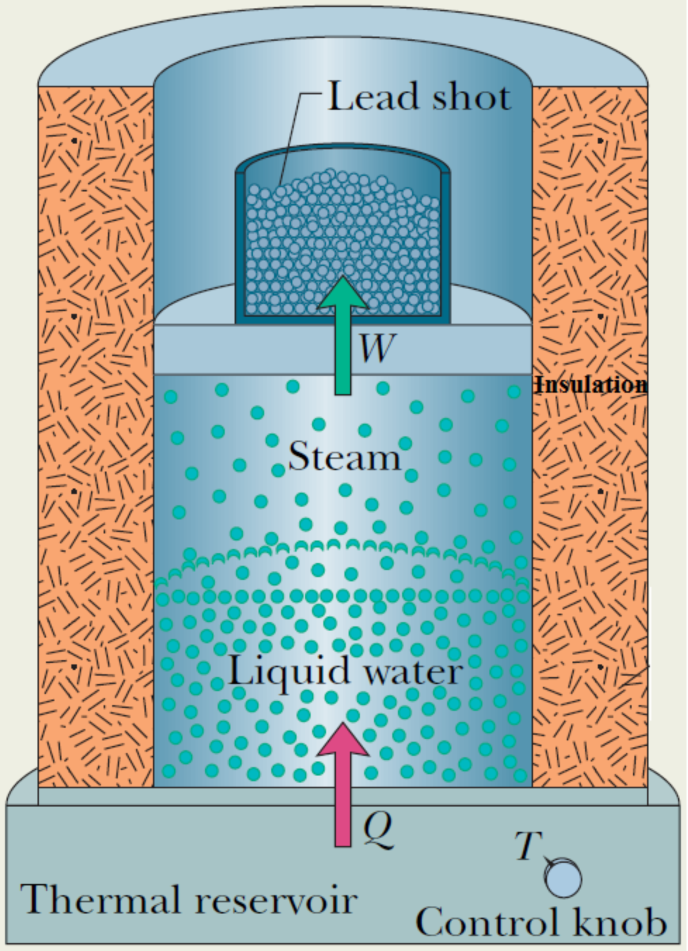
$$W = \int_{V_i}^{V_f} p\,dV = p\int_{V_i}^{V_f} dV = p(V_f - V_i)$$

$$= (1.01 \times 10^5\ \mathrm{Pa})(1.671\ \mathrm{m}^3 - 1.00 \times 10^{-3}\ \mathrm{m}^3)$$

$$= 169\ \mathrm{kJ}.$$



9. Some Special Cases of the First Law of Thermodynamics



(b) How much energy is transferred as heat during the process?

$$Q = mL_v = (1000\text{ g})(2256\text{ J/g}) \approx 2260\text{ kJ}.$$

(c) What is the change in the system's internal energy during the process?

$$\Delta E_{int} = Q - W = 2256\text{ kJ} - 169\text{ kJ} = 2090\text{ kJ}.$$

10. Heat Transfer Mechanisms

Now we describe how heat transfers. There are three main heat transfer mechanism:

(1) Conduction.

(2) Convection

(3) Radiation.

Conduction

If a region in a sample is hotter, the vibration amplitudes of atoms and electrons of that region become relatively higher. These increased vibrational amplitudes are passed via collisions from atom to atom during collisions. This results in an increase in the temperature of the colder regions.

10. Heat Transfer Mechanisms

Consider the situation shown in the figure. Experimentally, the conduction rate P_{cond} is

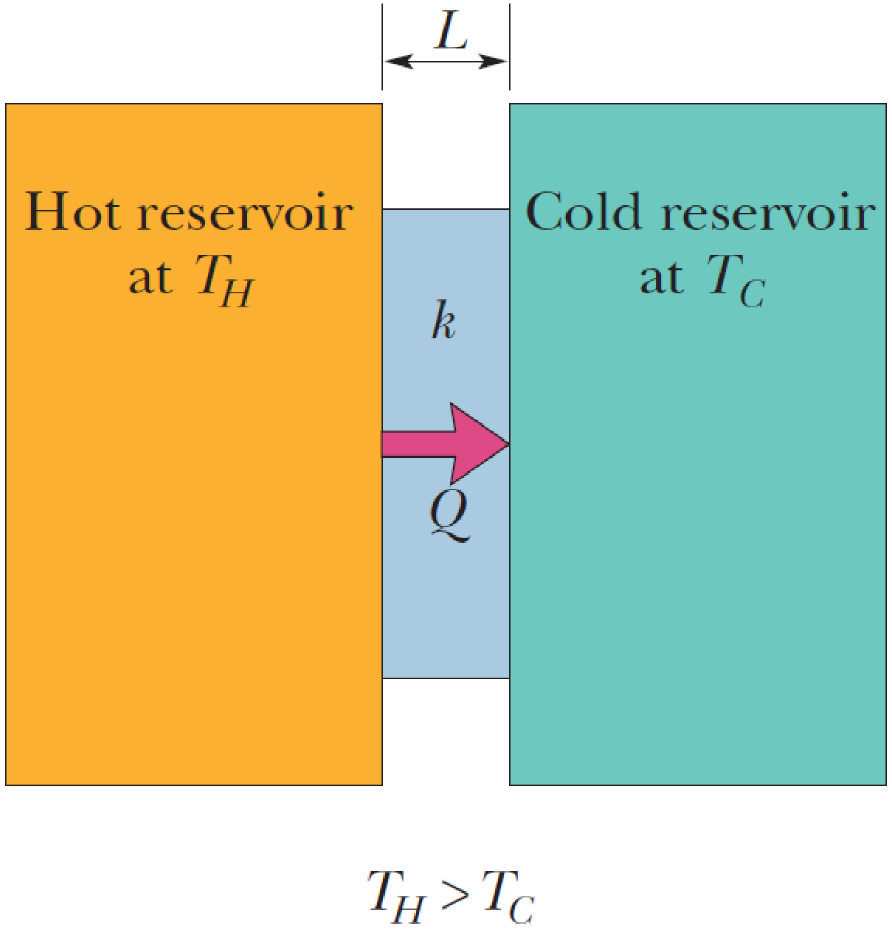
$$P_{\text{cond}} = \frac{Q}{t} = kA\frac{T_H - T_C}{L},$$

where k is the **thermal conductivity**. Its SI unit is $\mathrm{W/(m \cdot K)}$.

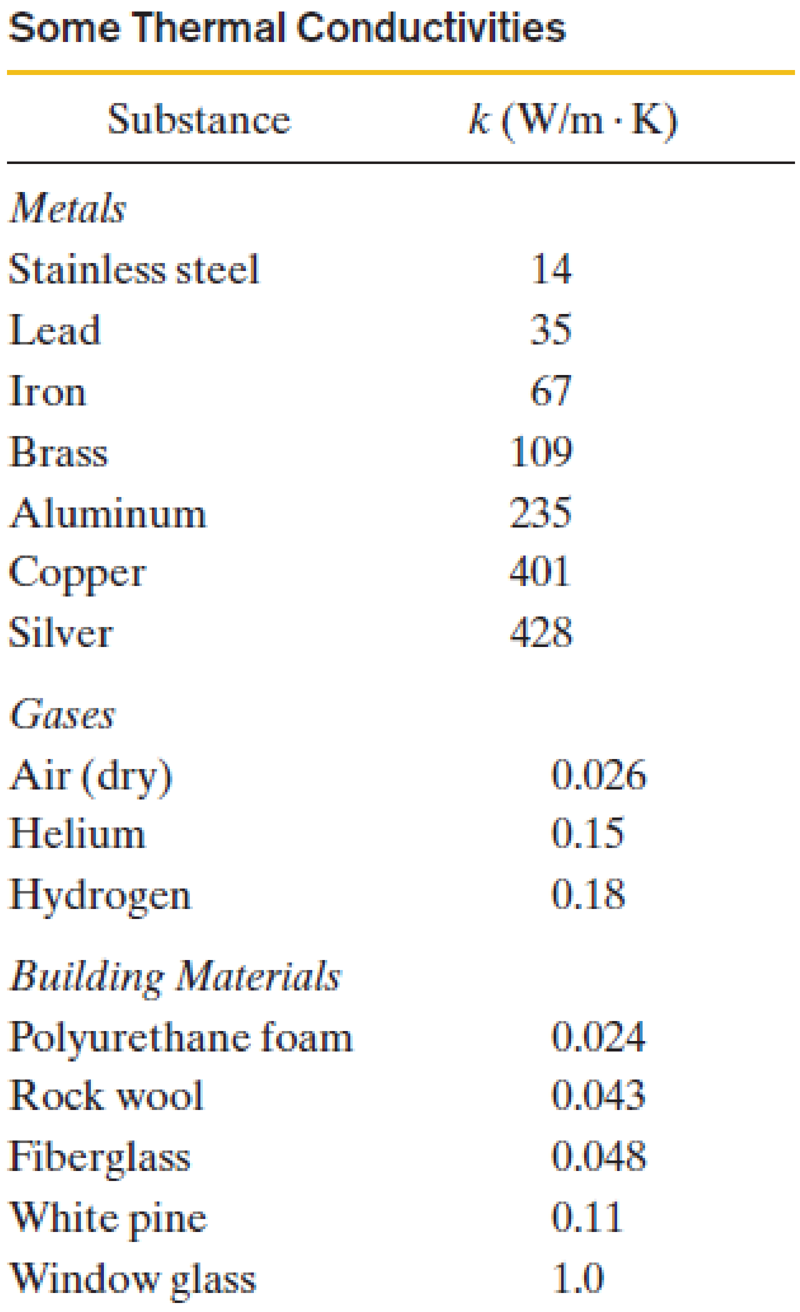
P_{cond} can be recast as

$$P_{\text{cond}} = A\frac{T_H - T_C}{R},$$

where $R \equiv L/k$, is the **thermal resistance** to conduction (R-value).



10. Heat Transfer Mechanisms



Some Thermal Conductivities

Substance	k (W/m · K)
Metals	
Stainless steel	14
Lead	35
Iron	67
Brass	109
Aluminum	235
Copper	401
Silver	428
Gases	
Air (dry)	0.026
Helium	0.15
Hydrogen	0.18
Building Materials	
Polyurethane foam	0.024
Rock wool	0.043
Fiberglass	0.048
White pine	0.11
Window glass	1.0

10. Heat Transfer Mechanisms

Conduction Through a Composite Slap

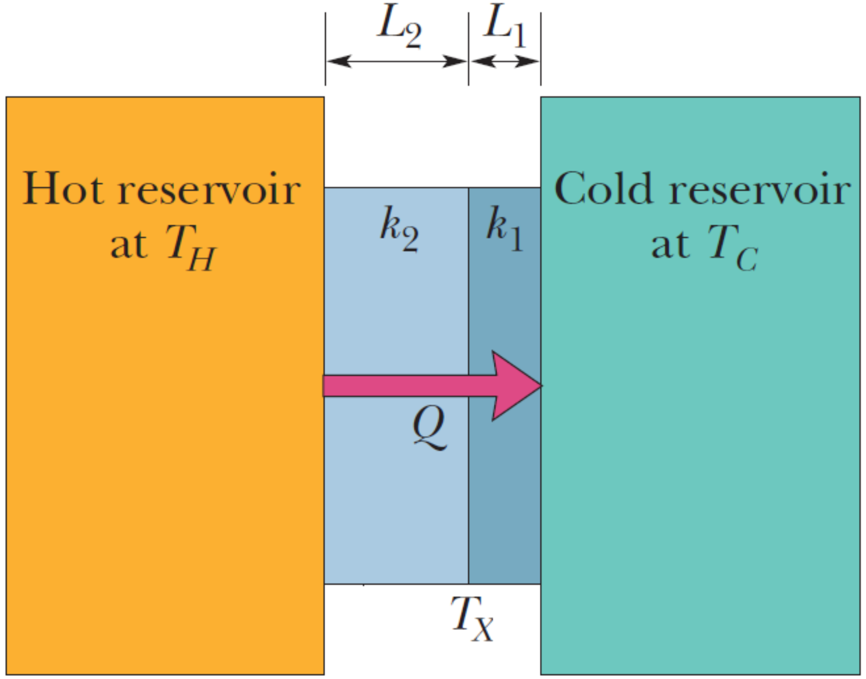
Consider the situation shown in the figure. Let us assume that the transfer is a steady state process (that is, the temperatures everywhere in the slab and the rate of energy transfer do not change with time).

The conduction rates through the two faces of the slap is the same:

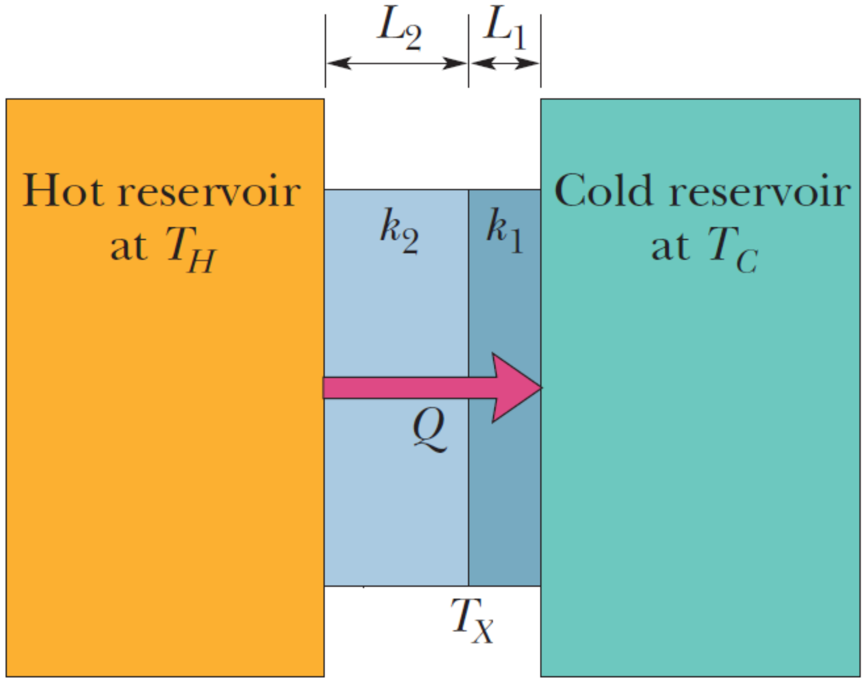
$$P_{\text{cond}} = k_2 A \frac{T_H - T_X}{L_2} = k_1 A \frac{T_X - T_C}{L_1}.$$

Eliminating T_X yields

$$P_{\text{cond}} = A \frac{T_H - T_C}{L_1/k_1 + L_2/k_2} = A \frac{T_H - T_C}{R_1 + R_2}.$$



10. Heat Transfer Mechanisms



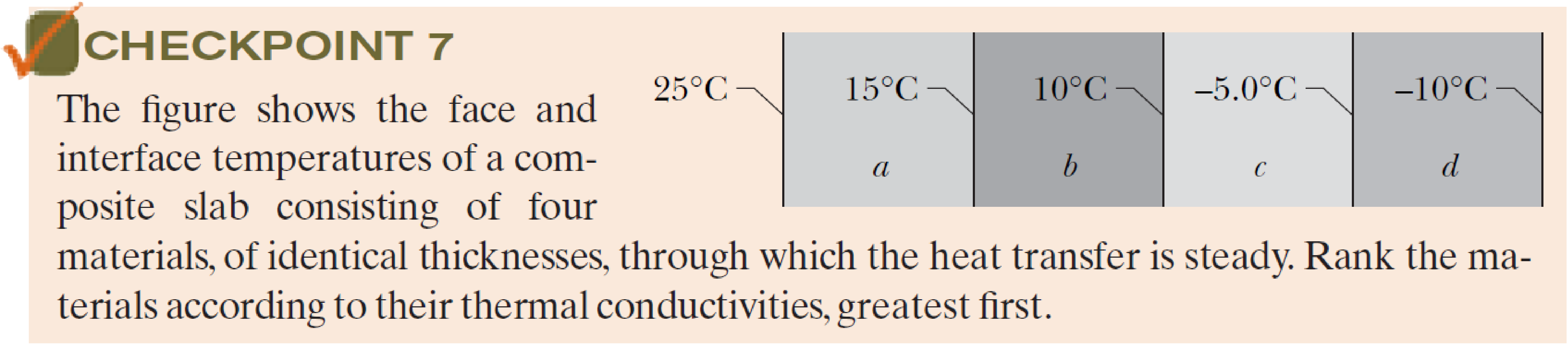
Conduction Through a Composite Slap

For a slab consisting of n materials

$$P_{\text{cond}} = A\frac{T_H - T_C}{\dfrac{L_1}{k_1} + \dfrac{L_2}{k_2} + \dfrac{L_3}{k_3} + \cdots + \dfrac{L_n}{k_n}},$$

$$= A\frac{T_H - T_C}{R_1 + R_2 + R_3 + \cdots + R_n}.$$

10. Heat Transfer Mechanisms



CHECKPOINT 7

The figure shows the face and interface temperatures of a composite slab consisting of four materials, of identical thicknesses, through which the heat transfer is steady. Rank the materials according to their thermal conductivities, greatest first.

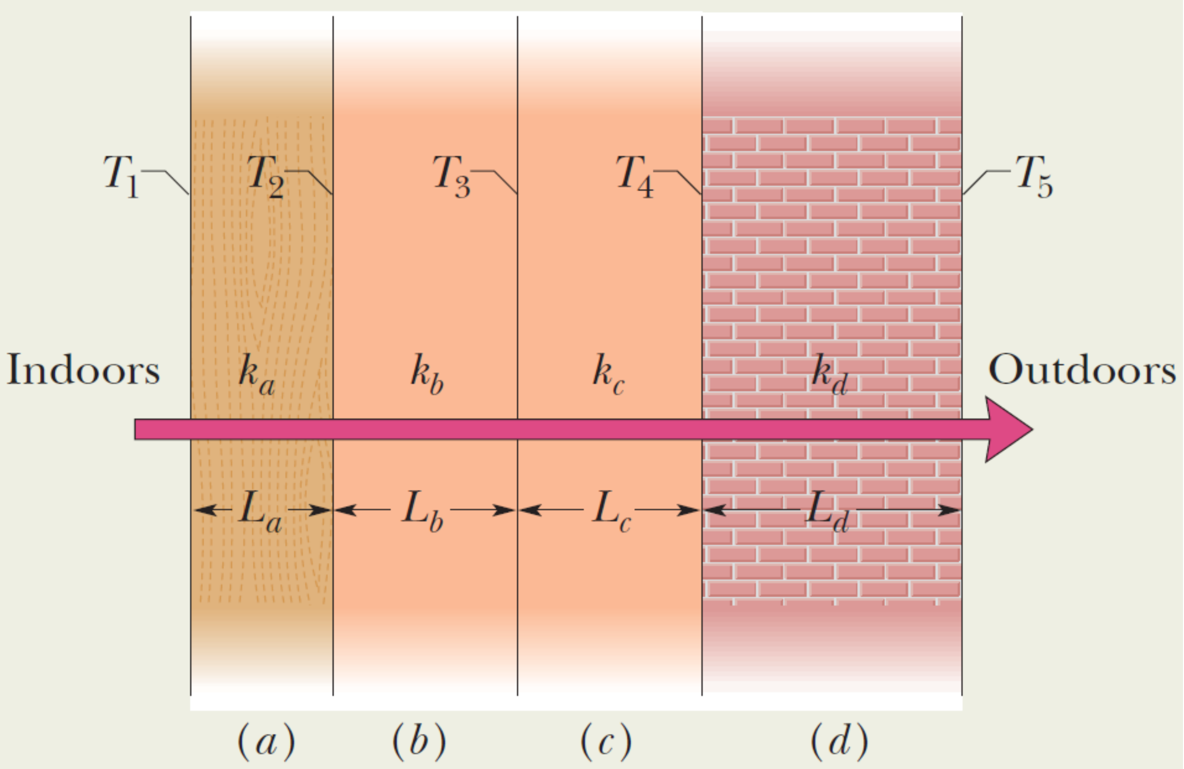
b & d, a then c.

$$P_{\text{cond}} = kA\frac{T_H - T_C}{L}$$

$$k = \frac{LP_{\text{cond}}}{A\,(T_H - T_C)}$$

10. Heat Transfer Mechanisms

Example 6: The figure shows the cross section of a wall made of white pine of thickness L_a and brick of thickness L_d $(= 2.0L_a)$, sandwiching two layers of unknown material with identical thicknesses and thermal conductivities. The thermal conductivity of the pine is k_a and that of the brick is k_d $(= 5.0k_a)$. The face area A of the wall is unknown. Thermal conduction through the wall has reached the steady state; the only known interface temperatures are $T_1 = 25°\text{C}$, $T_2 = 20°\text{C}$, and $T_5 = -10°\text{C}$. What is interface temperature T_4?



10. Heat Transfer Mechanisms

The conduction rates through the pine wood is

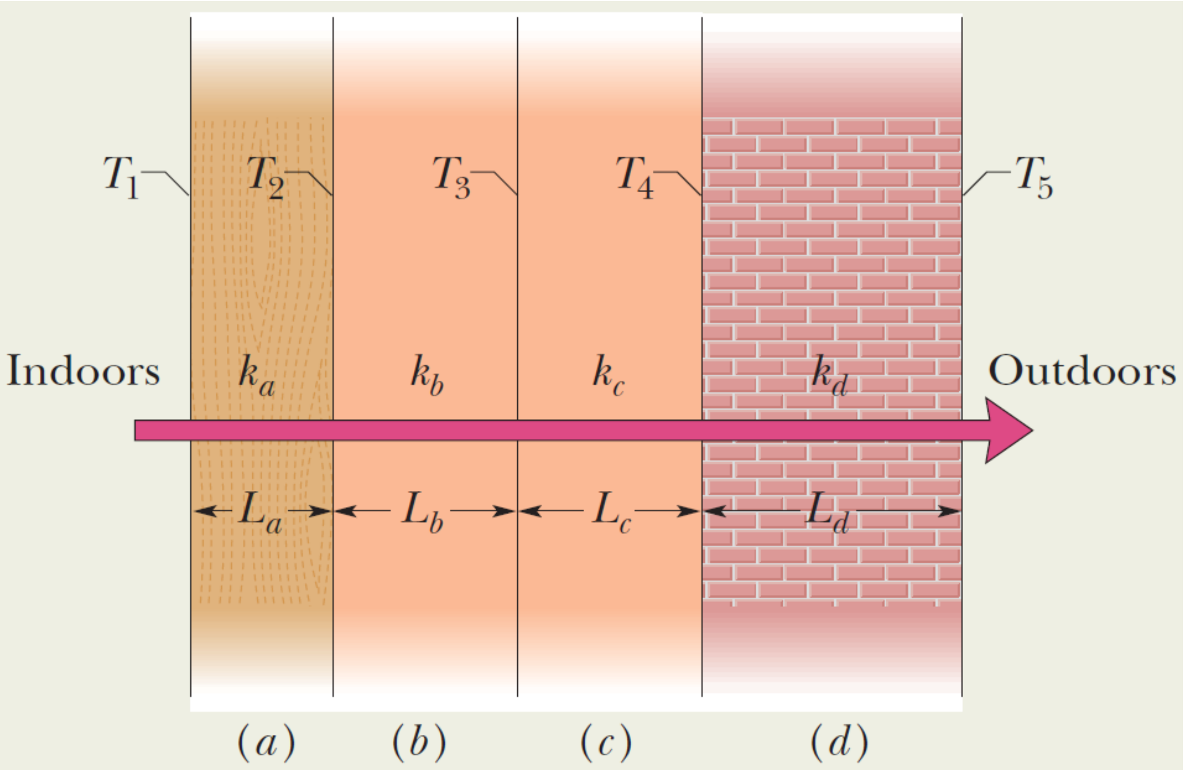
$$P_a = k_a A \frac{T_1 - T_2}{L_a},$$

and through the brick is

$$P_d = k_d A \frac{T_4 - T_5}{L_d}.$$

We know that $P_a = P_d$. This gives

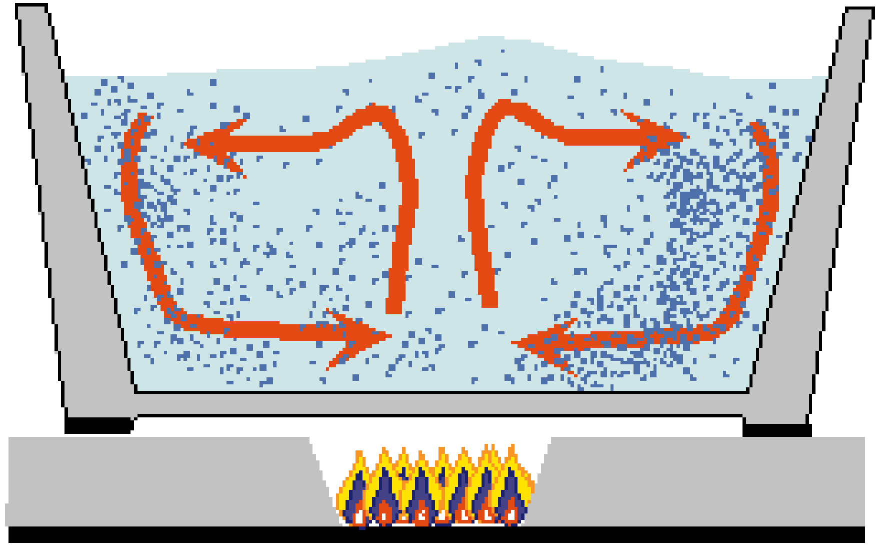
$$\begin{aligned} T_4 &= \frac{k_a L_d}{k_d L_a}(T_1 - T_2) + T_5 \\ &= \frac{(k_a)(2.0L_a)}{(5.0\,k_a)L_a}(25°\text{C} - 20°\text{C}) + (-10°\text{C}) \\ &= -8.0°\text{C}. \end{aligned}$$



10. Heat Transfer Mechanisms

Convection

When a fluid comes in contact with a hotter object, the temperature of the part of the fluid that is in contact with the hot object increases. That part of the fluid expands and thus becomes less dense and lighter than the surrounding cooler fluid. Thus, buoyant forces cause the hotter part to rise and some of the surrounding cooler fluid then flows so as to take the place of the rising warmer fluid, and the process can then continue.



SINGLE CONVECTION CELL

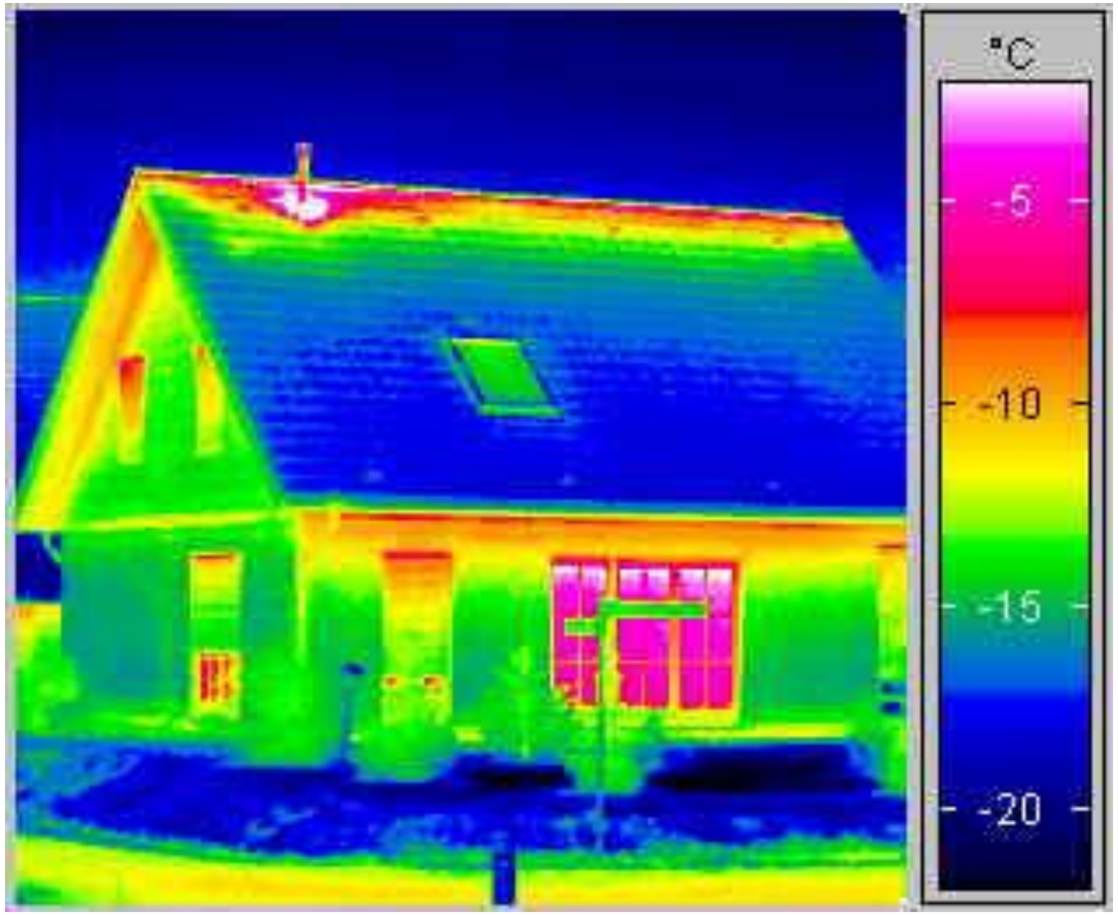
10. Heat Transfer Mechanisms

Radiation

An object of surface area A and temperature T (on the surface) emits electromagnetic radiation (called thermal radiation) at the rate

$$P_{\text{rad}} = \sigma \varepsilon A T^4,$$

where $\sigma = 5.6704 \times 10^{-8}\ \mathrm{W/(m^2 \cdot K^4)}$ is the **Stefan-Boltzmann constant**. ε is called the **emissivity** of the surface, which has a value between 0 and 1.



10. Heat Transfer Mechanisms

Radiation

If an object is in an environment of uniform temperature T_{env}, the object absorbs energy (thermal radiation) at the rate

$$P_{\mathrm{abs}} = \sigma\varepsilon A T_{\mathrm{env}}^4.$$

The object's net rate P_{net} of energy exchange due to thermal radiation is

$$P_{\mathrm{net}} = \sigma\varepsilon A(T_{\mathrm{env}}^4 - T^4).$$

Read sample problem 18.07 on page 538.