

Part 3 : Thermodynamics

Chapter 19

Temperature, Heat, and the First Law of Thermodynamics

Key contents:

Temperature scales

Thermal expansion

Temperature and heat, specific heat

Heat and work, the 1st law of thermodynamics

Heat transfer mechanisms

19.1 Temperature

What is 'temperature' ?

Feeling of hot and cold

To be defined,
measured with other
properties of a system

A state variable of a
system

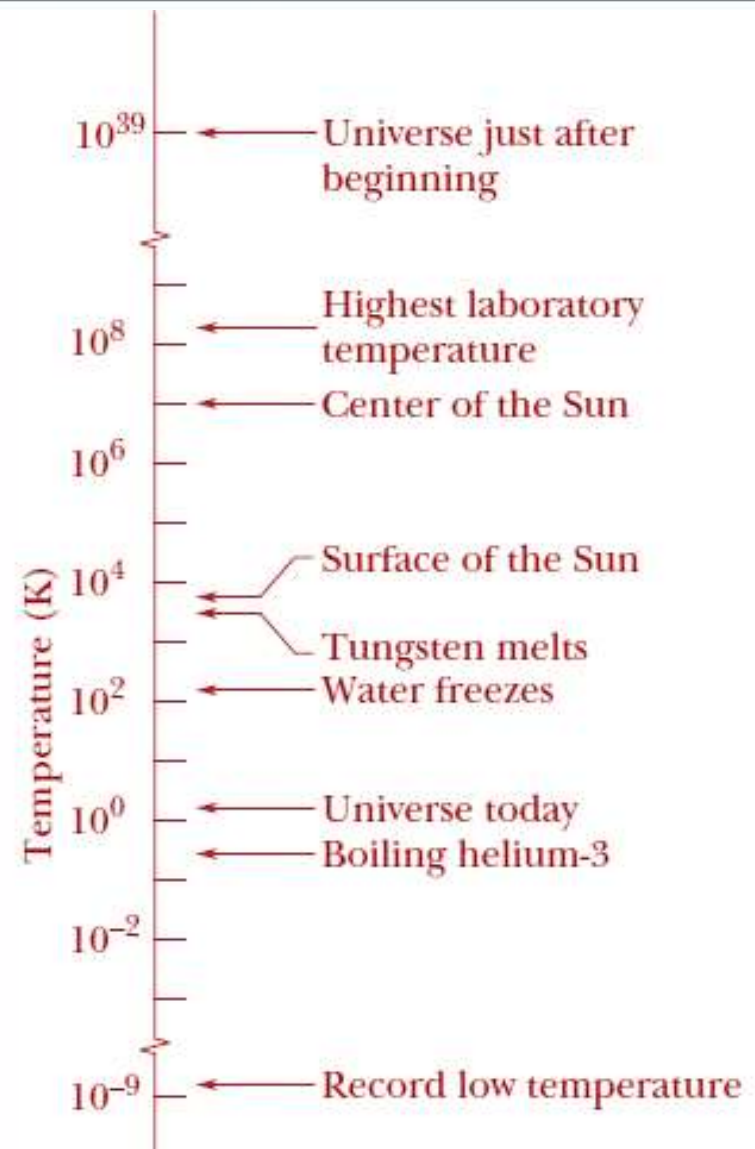


Fig. 18-1 Some temperatures on the Kelvin scale. Temperature $T = 0$ corresponds to $10^{-\infty}$ and cannot be plotted on this logarithmic scale.

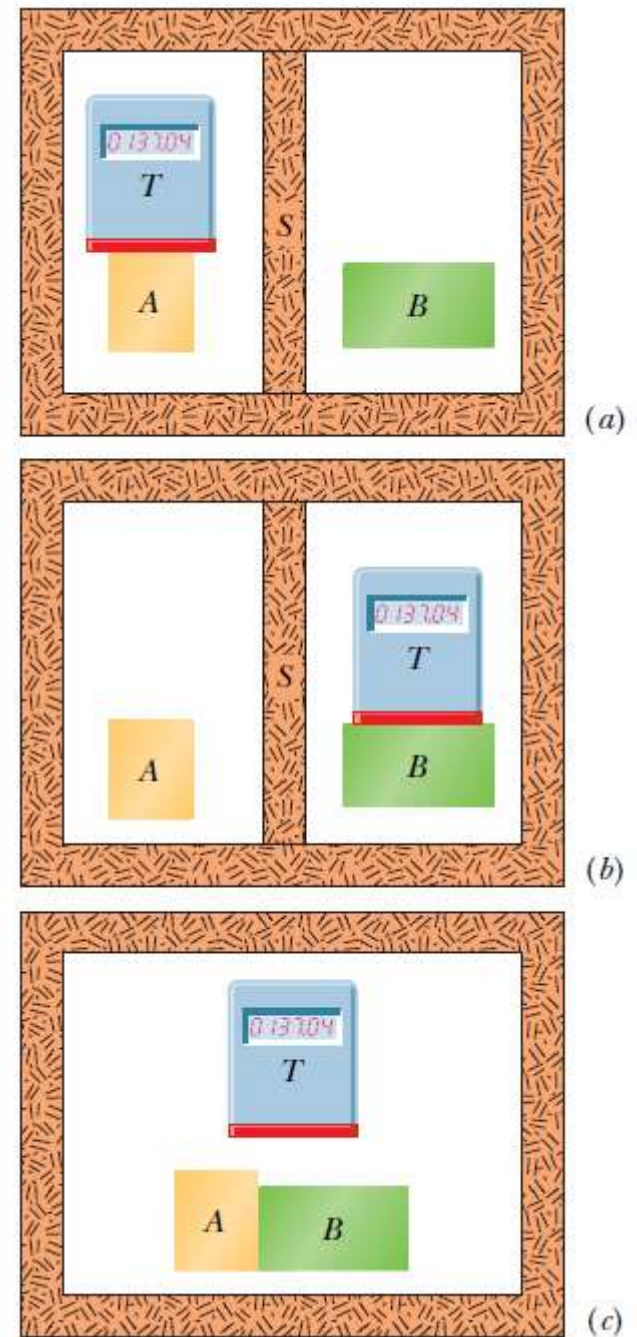
19.2: The Zeroth Law of Thermodynamics

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.

Thermal equilibrium:

The temperature is constant in time.

Fig. 18-3 (a) Body T (a thermoscope) and body A are in thermal equilibrium. (Body S is a thermally insulating screen.) (b) Body T and body B are also in thermal equilibrium, at the same reading of the thermoscope. (c) If (a) and (b) are true, the zeroth law of thermodynamics states that body A and body B are also in thermal equilibrium.



19.3 Measuring Temperature

$$T_3 = 273.16 \text{ K} \quad (\text{triple-point temperature}),$$

The Kelvin scale,
the absolute temperature scale

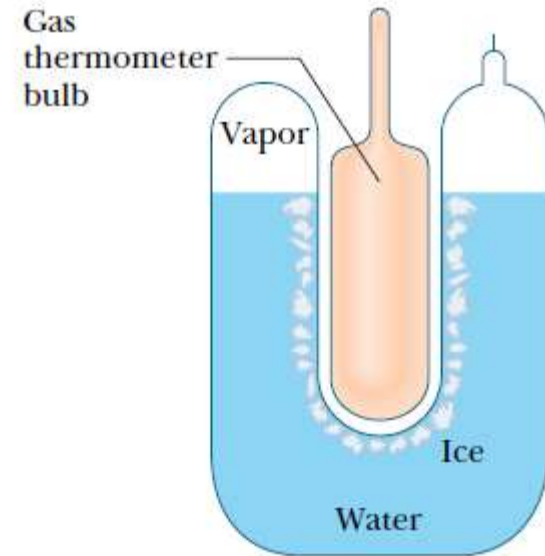


Fig. 18-4 A triple-point cell, in which solid ice, liquid water, and water vapor co-exist in thermal equilibrium. By international agreement, the temperature of this mixture has been defined to be 273.16 K. The bulb of a constant-volume gas thermometer is shown inserted into the well of the cell.

19.4 The Celsius and Fahrenheit Scales

$$T_C = T - 273.15^\circ$$

$$T_F = \frac{9}{5}T_C + 32^\circ,$$

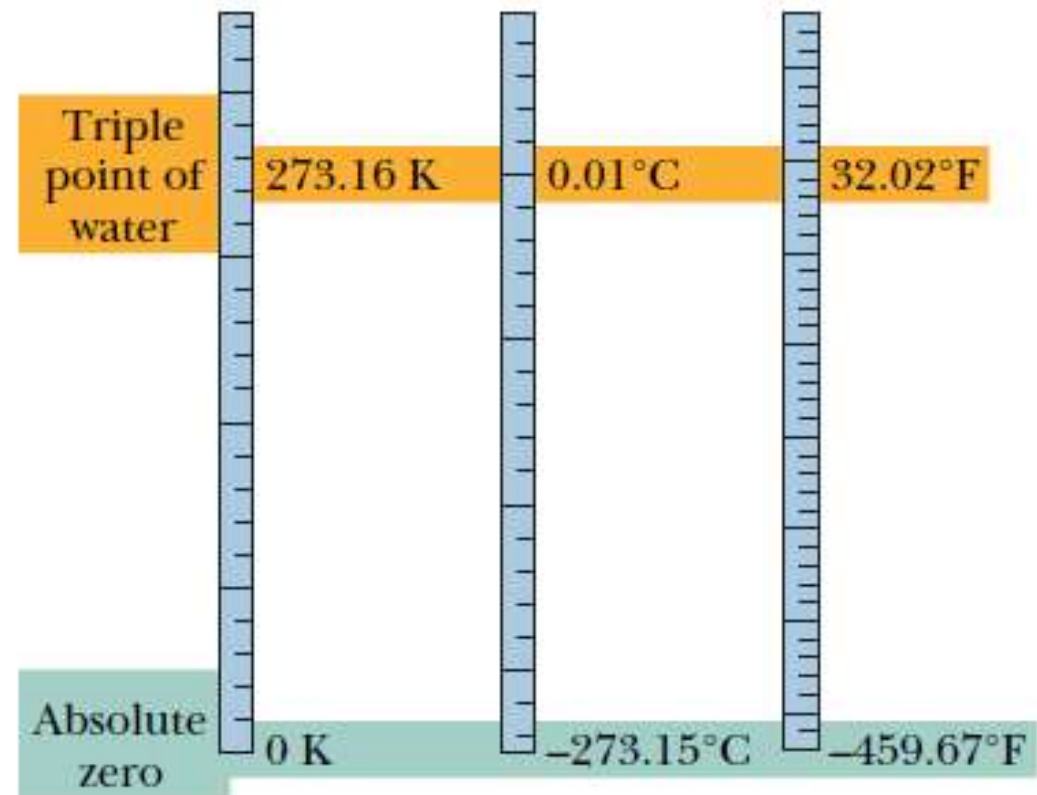


Fig. 18-7 The Kelvin, Celsius, and Fahrenheit temperature scales compared.

19.4 Measuring Temperature, The Constant Volume Gas Thermometer

The temperature of a body can be defined as $T = Cp$, where p is the pressure in the bulb. Assuming at the triple point, we also have $T_3 = Cp_3$, with the same constant C .

Therefore,

$$T = T_3 \left(\frac{p}{p_3} \right) = (273.16 \text{ K}) \left(\frac{p}{p_3} \right)$$

But only when the gas is of a very small amount, this measurement gives consistent results among different materials used.

$$T = (273.16 \text{ K}) \left(\lim_{p \rightarrow 0} \frac{p}{p_3} \right)$$

This is called the

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‘ideal gas temperature’.

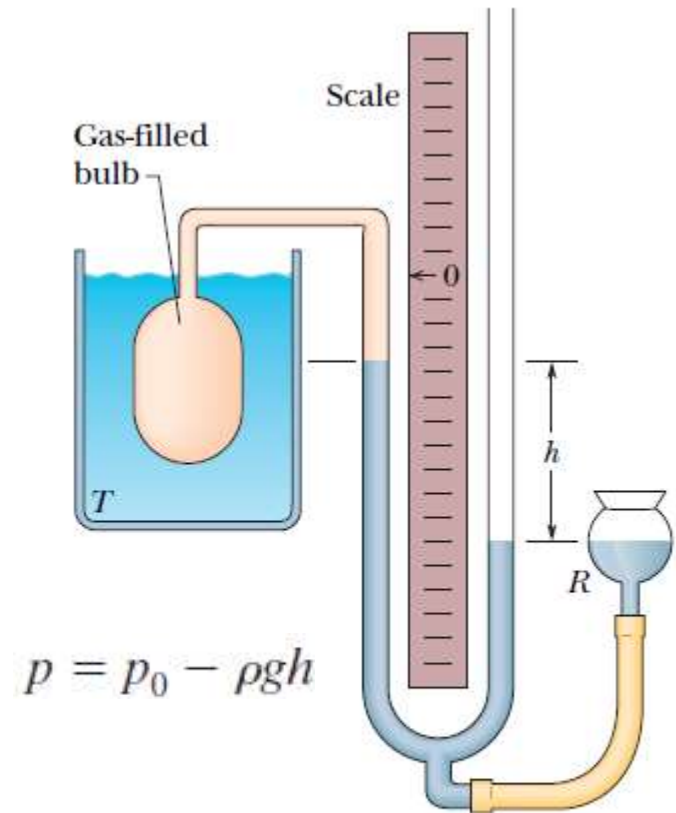


Fig. 18-5 A constant-volume gas thermometer, its bulb immersed in a liquid whose temperature T is to be measured.

19.5 The Celsius and Fahrenheit Scales

Table 18-1

Some Corresponding Temperatures

Temperature	$^{\circ}\text{C}$	$^{\circ}\text{F}$
Boiling point of water ^a	100	212
Normal body temperature	37.0	98.6
Accepted comfort level	20	68
Freezing point of water ^a	0	32
Zero of Fahrenheit scale	≈ -18	0
Scales coincide	-40	-40

^aStrictly, the boiling point of water on the Celsius scale is 99.975°C , and the freezing point is 0.00°C . Thus, there is slightly less than 100°C between those two points.

Example, Conversion Between Temperature Scales

Suppose you come across old scientific notes that describe a temperature scale called Z on which the boiling point of water is 65.0°Z and the freezing point is -14.0°Z . To what temperature on the Fahrenheit scale would a temperature of $T = -98.0^{\circ}\text{Z}$ correspond? Assume that the Z scale is linear; that is, the size of a Z degree is the same everywhere on the Z scale.

KEY IDEA

A conversion factor between two (linear) temperature scales can be calculated by using two known (benchmark) temperatures, such as the boiling and freezing points of water. The number of degrees between the known temperatures on one scale is equivalent to the number of degrees between them on the other scale.

Calculations: We begin by relating the given temperature T to *either* known temperature on the Z scale. Since $T = -98.0^{\circ}\text{Z}$ is closer to the freezing point (-14.0°Z) than to the boiling point (65.0°Z), we use the freezing point. Then we note that the T we seek is *below this point* by $-14.0^{\circ}\text{Z} - (-98.0^{\circ}\text{Z}) = 84.0^{\circ}\text{Z}$ (Fig. 18-8). (Read this difference as “84.0 Z degrees.”)

Next, we set up a conversion factor between the Z and Fahrenheit scales to convert this difference. To do so, we use *both* known temperatures on the Z scale and the corre-

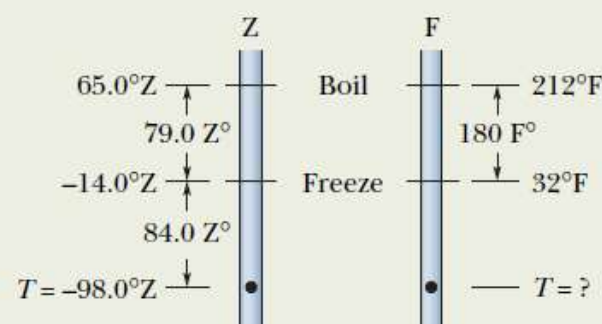


Fig. 18-8 An unknown temperature scale compared with the Fahrenheit temperature scale.

sponding temperatures on the Fahrenheit scale. On the Z scale, the difference between the boiling and freezing points is $65.0^{\circ}\text{Z} - (-14.0^{\circ}\text{Z}) = 79.0^{\circ}\text{Z}$. On the Fahrenheit scale, it is $212^{\circ}\text{F} - 32.0^{\circ}\text{F} = 180^{\circ}\text{F}$. Thus, a temperature difference of 79.0°Z is equivalent to a temperature difference of 180°F (Fig. 18-8), and we can use the ratio $(180^{\circ}\text{F})/(79.0^{\circ}\text{Z})$ as our conversion factor.

Now, since T is below the freezing point by 84.0°Z , it must also be below the freezing point by

$$(84.0^{\circ}\text{Z}) \frac{180^{\circ}\text{F}}{79.0^{\circ}\text{Z}} = 191^{\circ}\text{F}.$$

Because the freezing point is at 32.0°F , this means that

$$T = 32.0^{\circ}\text{F} - 191^{\circ}\text{F} = -159^{\circ}\text{F}. \quad (\text{Answer})$$

19.6: Thermal Expansion

When the temperature of an object is raised, the body usually exhibits “thermal expansion”.

With the added thermal energy, the atoms can move a bit farther from one another than usual, against the spring-like interatomic forces that hold every solid together.

The atoms in the metal move farther apart than those in the glass, which makes a metal object expand more than a glass object.



Fig. 18-9 When a Concorde flew faster than the speed of sound, thermal expansion due to the rubbing by passing air increased the aircraft's length by about 12.5 cm. (The temperature increased to about 128°C at the aircraft nose and about 90°C at the tail, and cabin windows were noticeably warm to the touch.) (*Hugh Thomas/BWP Media/Getty Images News and Sport Services*)

19.6: Thermal Expansion, Linear Expansion

If the temperature of a metal rod of length L is raised by an amount ΔT , its length is found to increase by an amount

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

in which α is a constant called the **coefficient of linear expansion**.

Table 18-2

Some Coefficients of Linear Expansion^a

Substance	α ($10^{-6}/\text{C}^\circ$)	Substance	α ($10^{-6}/\text{C}^\circ$)
Ice (at 0°C)	51	Steel	11
Lead	29	Glass (ordinary)	9
Aluminum	23	Glass (Pyrex)	3.2
Brass	19	Diamond	1.2
Copper	17	Invar ^b	0.7
Concrete	12	Fused quartz	0.5

^aRoom temperature values except for the listing for ice.

^bThis alloy was designed to have a low coefficient of expansion. The word is a shortened form of “invariable.”

19.6: Thermal Expansion, Volume Expansion

$$\Delta V = V\beta \Delta T$$

β is the **coefficient of volume expansion**. The coefficients of volume expansion and linear expansion are related by

$$\beta = 3\alpha.$$

$$\# \quad V \propto L^3$$

$$1 + b = (1 + a)^3 \gg 1 + 3a$$

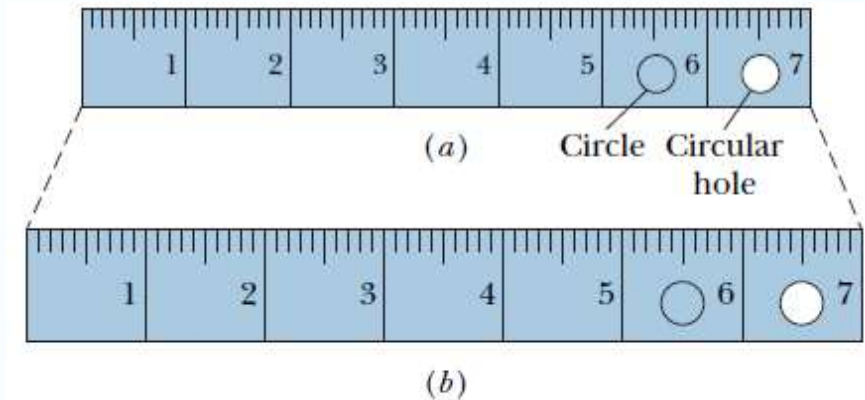


Fig. 18-11 The same steel ruler at two different temperatures. When it expands, the scale, the numbers, the thickness, and the diameters of the circle and circular hole are all increased by the same factor. (The expansion has been exaggerated for clarity.)

19.6: Thermal Expansion, Anomalous Expansion of Water

The most common liquid, water, does not behave like other liquids. Above about 4°C , water expands as the temperature rises, as we would expect.

Between 0 and about 4°C , however, water *contracts with increasing temperature*. Thus, at about 4°C , the density of water passes through a maximum.

At all other temperatures, the density of water is less than this maximum value.

Thus the surface of a pond freezes while the lower water is still liquid.

Example, Thermal Expansion of Volume:

On a hot day in Las Vegas, an oil trucker loaded 37 000 L of diesel fuel. He encountered cold weather on the way to Payson, Utah, where the temperature was 23.0 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 $9.50 \times 10^{-4}/\text{C}^\circ$, and the coefficient of linear expansion for his steel truck tank is $11 \times 10^{-6}/\text{C}^\circ$.

KEY IDEA

The volume of the diesel fuel depends directly on the temperature. Thus, because the temperature decreased, the

volume of the fuel did also, as given by Eq. 18-10 ($\Delta V = V\beta\Delta T$).

Calculations: We find

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

Thus, the amount delivered was

$$\begin{aligned} V_{\text{del}} &= V + \Delta V = 37\,000\text{ L} - 808\text{ L} \\ &= 36\,190\text{ L}. \end{aligned} \quad (\text{Answer})$$

Note that the thermal expansion of the steel tank has nothing to do with the problem. Question: Who paid for the “missing” diesel fuel?

19.7: Temperature and Heat

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

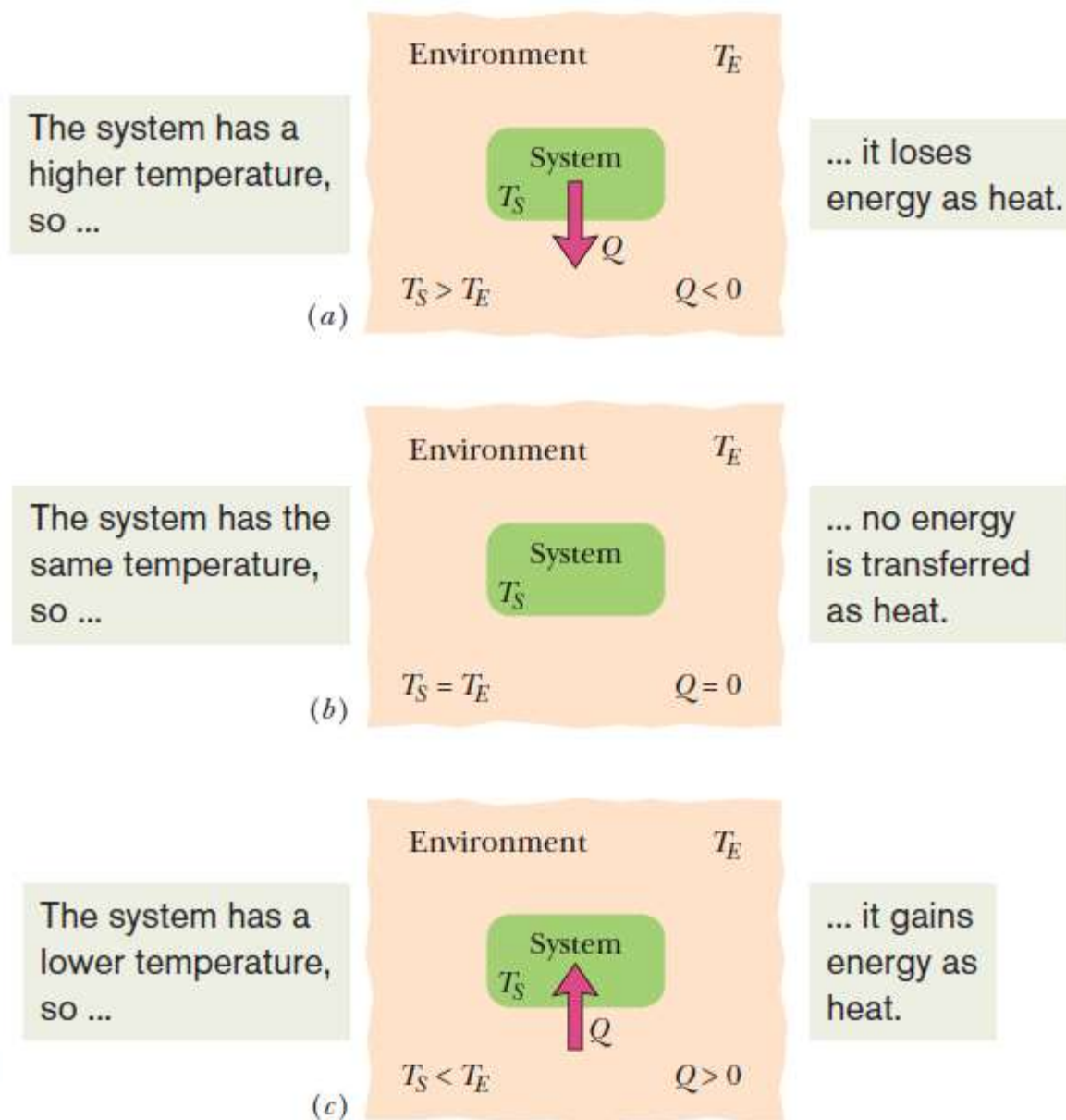


Fig. 18-12 If the temperature of a system exceeds that of its environment as in (a), heat Q is lost by the system to the environment until thermal equilibrium (b) is established. (c) If the temperature of the system is below that of the environment, heat is absorbed by the system until thermal equilibrium is established.

19.7: Temperature and Heat: Units

- The **calorie (cal)** was defined as the amount of heat that would raise the temperature of 1 g of water from 14.5°C to 15.5°C.
- In the British system, the corresponding unit of heat was the **British thermal unit (Btu)**, defined as the amount of heat that would raise the temperature of 1 lb of water from 63°F to 64°F.
- The SI unit for heat is the **joule**.
- The calorie is now defined to be 4.1868 J.
- $1 \text{ cal} = 3.968 \times 10^{-3} \text{ Btu} = 4.1868 \text{ J}$.

19.8: The Absorption of Heat

❖ The **heat capacity** C of an object is the proportionality constant between the heat Q that the object absorbs or loses and the resulting temperature change ΔT of the object

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

in which T_i and T_f are the initial and final temperatures of the object.

❖ Heat capacity C has the unit of energy per degree or energy per kelvin.

19.8: The Absorption of Heat: Specific Heat

The **specific heat**, c , is the heat capacity per unit mass.

It refers not to an object but to a unit mass of the material of which the object is made.

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

When specific heats are expressed in units of moles (rather than a mass unit), they are called **molar specific heats**.

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

It is important to distinguish the heat transfer is done with constant volume or constant pressure. The specific heat is different for different processes, particular for gases.

Table 18-3

Some Specific Heats and Molar Specific Heats at Room Temperature

Substance	Specific Heat		Molar Specific Heat
	$\frac{\text{cal}}{\text{g} \cdot \text{K}}$	$\frac{\text{J}}{\text{kg} \cdot \text{K}}$	$\frac{\text{J}}{\text{mol} \cdot \text{K}}$
<i>Elemental Solids</i>			
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
<i>Other Solids</i>			
Brass	0.092	380	
Granite	0.19	790	
Glass	0.20	840	
Ice (-10°C)	0.530	2220	
<i>Liquids</i>			
Mercury	0.033	140	
Ethyl alcohol	0.58	2430	
Seawater	0.93	3900	
Water	1.00	4187	

19.8: The Absorption of Heat by Solids and Liquids: Heat of Transformation

The amount of energy per unit mass that must be transferred as heat when a sample completely undergoes a phase change is called **the heat of transformation L (or latent heat)**. When a sample of mass m completely undergoes a phase change, the total energy transferred is:

$$Q = Lm.$$

(# sublimation: transition from solid directly to gas phases)

When the phase change is between solid to liquid, the heat of transformation is called the **heat of fusion L_F** .

When the phase change is between liquid to gas, the heat of transformation is called the **heat of vaporization L_V** .

Table 18-4

Some Heats of Transformation

Substance	Melting		Boiling	
	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

So, some key words:

Heat capacity

Specific heat (at constant volume, or, at constant pressure)

Molar specific heat

Latent heat

Example, Hot Slug in Water:

A copper slug whose mass m_c is 75 g is heated in a laboratory oven to a temperature T of 312°C. The slug is then dropped into a glass beaker containing a mass $m_w = 220$ g of water. The heat capacity C_b of the beaker is 45 cal/K. The initial temperature T_i of the water and the beaker is 12°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.

Calculations: To relate the transfers to the temperature changes, we can use Eqs. 18-13 and 18-14 to write

$$\text{for the water: } Q_w = c_w m_w (T_f - T_i); \quad (18-19)$$

$$\text{for the beaker: } Q_b = C_b (T_f - T_i); \quad (18-20)$$

$$\text{for the copper: } Q_c = c_c m_c (T_f - T). \quad (18-21)$$

Because the total energy of the system cannot change, the sum of these three energy transfers is zero:

$$Q_w + Q_b + Q_c = 0. \quad (18-22)$$

Substituting Eqs. 18-19 through 18-21 into Eq. 18-22 yields

$$c_w m_w (T_f - T_i) + C_b (T_f - T_i) + c_c m_c (T_f - T) = 0. \quad (18-23)$$

Temperatures are contained in Eq. 18-23 only as differences. Thus, because the differences on the Celsius and Kelvin scales are identical, we can use either of those scales in this equation. Solving it for T_f , we obtain

$$T_f = \frac{c_c m_c T + C_b T_i + c_w m_w T_i}{c_w m_w + C_b + c_c m_c}.$$

Using Celsius temperatures and taking values for c_c and c_w from Table 18-3, we find the numerator to be

$$(0.0923 \text{ cal/g} \cdot \text{K})(75 \text{ g})(312^\circ\text{C}) + (45 \text{ cal/K})(12^\circ\text{C}) \\ + (1.00 \text{ cal/g} \cdot \text{K})(220 \text{ g})(12^\circ\text{C}) = 5339.8 \text{ cal},$$

and the denominator to be

$$(1.00 \text{ cal/g} \cdot \text{K})(220 \text{ g}) + 45 \text{ cal/K} \\ + (0.0923 \text{ cal/g} \cdot \text{K})(75 \text{ g}) = 271.9 \text{ cal/C}^\circ.$$

We then have

$$T_f = \frac{5339.8 \text{ cal}}{271.9 \text{ cal/C}^\circ} = 19.6^\circ\text{C} \approx 20^\circ\text{C}. \quad (\text{Answer})$$

From the given data you can show that

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

Apart from rounding errors, the algebraic sum of these three heat transfers is indeed zero, as Eq. 18-22 requires.

Example, Heat to Change Temperature:

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

Warming the ice: The heat Q_1 needed to increase the temperature of the ice from the initial value $T_i = -10^\circ\text{C}$ to a final value $T_f = 0^\circ\text{C}$ (so that the ice can then melt) is given by Eq. 18-14 ($Q = cm \Delta T$). Using the specific heat of ice c_{ice} in Table 18-3 gives us

$$\begin{aligned} Q_1 &= c_{\text{ice}} m (T_f - T_i) \\ &= (2220 \text{ J/kg} \cdot \text{K})(0.720 \text{ kg})[0^\circ\text{C} - (-10^\circ\text{C})] \\ &= 15\,984 \text{ J} \approx 15.98 \text{ kJ.} \end{aligned}$$

Melting the ice: The heat Q_2 needed to melt all the ice is given by Eq. 18-16 ($Q = Lm$). Here L is the heat of fusion L_F , with the value given in Eq. 18-18 and Table 18-4. We find

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

Warming the liquid: The heat Q_3 needed to increase the temperature of the water from the initial value $T_i = 0^\circ\text{C}$ to the final value $T_f = 15^\circ\text{C}$ is given by Eq. 18-14 (with the specific heat of liquid water c_{liq}):

$$\begin{aligned} Q_3 &= c_{\text{liq}} m (T_f - T_i) \\ &= (4186.8 \text{ J/kg} \cdot \text{K})(0.720 \text{ kg})(15^\circ\text{C} - 0^\circ\text{C}) \\ &= 45\,217 \text{ J} \approx 45.22 \text{ kJ.} \end{aligned}$$

Total: The total required heat Q_{tot} is the sum of the amounts required in the three steps:

$$\begin{aligned} Q_{\text{tot}} &= Q_1 + Q_2 + Q_3 \\ &= 15.98 \text{ kJ} + 239.8 \text{ kJ} + 45.22 \text{ kJ} \\ &\approx 300 \text{ kJ.} \end{aligned}$$

(Answer)

Example, Heat to Change Temperature, cont.:

(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?

KEY IDEA

From step 1, we know that 15.98 kJ is needed to raise the temperature of the ice to the melting point. The remaining heat Q_{rem} is then 210 kJ – 15.98 kJ, or about 194 kJ. From step 2, we can see that this amount of heat is insufficient to melt all the ice. Because the melting of the ice is incomplete, we must end up with a mixture of ice and liquid; the temperature of the mixture must be the freezing point, 0°C.

Calculations: We can find the mass m of ice that is melted by the available energy Q_{rem} by using Eq. 18-16 with L_F :

$$m = \frac{Q_{\text{rem}}}{L_F} = \frac{194 \text{ kJ}}{333 \text{ kJ/kg}} = 0.583 \text{ kg} \approx 580 \text{ g}.$$

Thus, the mass of the ice that remains is 720 g – 580 g, or 140 g, and we have

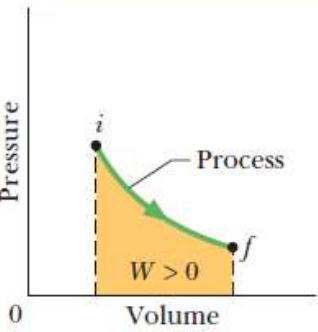
580 g water and 140 g ice, at 0°C. (Answer)

19.9: A Closer Look at Heat and Work

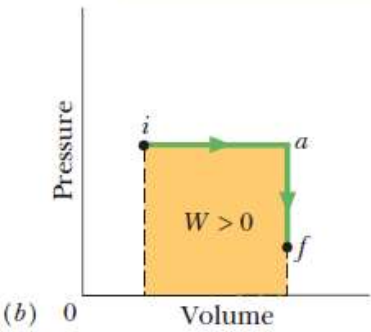
$$dW = \vec{F} \cdot d\vec{s} = (pA)(ds) = p(A ds)$$
$$= p dV,$$

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

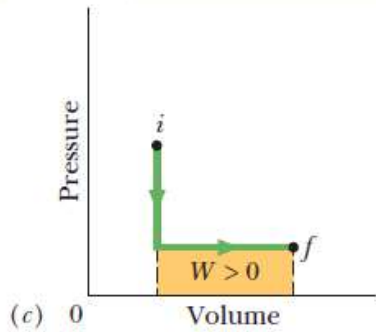
Gas moves from i to f ,
doing positive work.



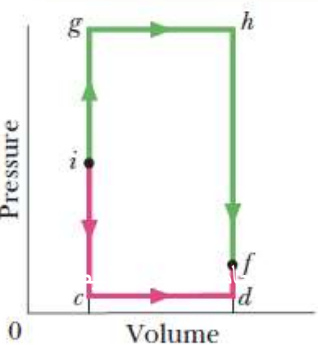
It still goes from i to f ,
but now it does *more* work.



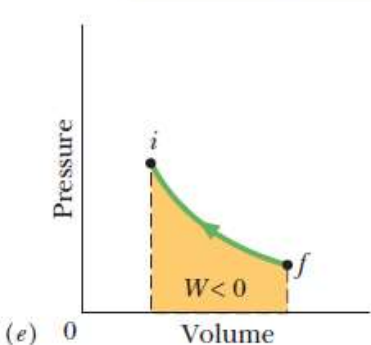
It still goes from i to f
but now it does *less* work.



We can control how
much work it does.



Moving from f to i ,
it does negative work.



Cycling clockwise
yields a positive net
work.

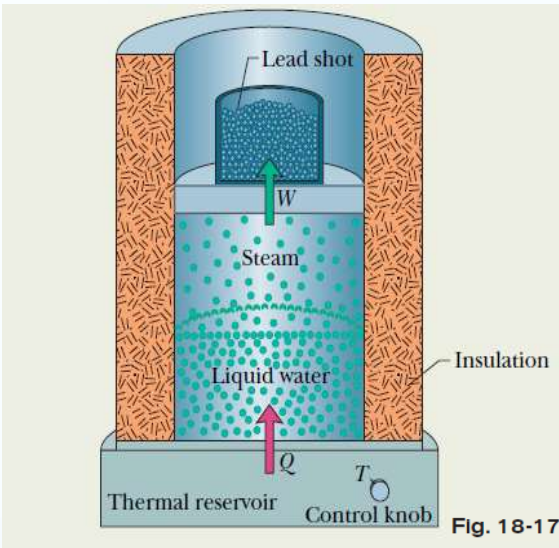
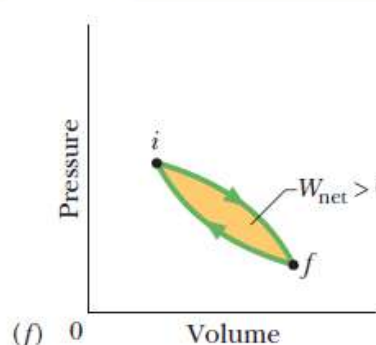


Fig. 18-17

19.10: 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.

$$dE_{\text{int}} = dQ - dW \quad (\text{first law}).$$

(Q is the heat absorbed and W is the work done by the system).

✓ The quantity $(Q - W)$ is the same for all processes. It depends only on the initial and final states of the system and does not depend at all on how the system gets from one to the other.

✓ This is simply conservation of energy.

19.11: Some Specific Cases of the First Law of Thermodynamics

- 1. Adiabatic processes.** An adiabatic process is one that occurs so rapidly or occurs in a system that is so well insulated that *no transfer of thermal energy occurs* between the system and its environment. Putting $Q=0$ in the first law,
$$\Delta E_{\text{int}} = -W \quad (\text{adiabatic process})$$
- 2. Constant-volume processes.** If the volume of a system (such as a gas) is held constant, so that *system can do no work*. Putting $W=0$ in the first law,
$$\Delta E_{\text{int}} = Q \quad (\text{constant-volume process})$$
- 1. Cyclical processes.** There are processes in which, after certain interchanges of heat and work, the system is restored to its initial state. No intrinsic property of the system—including its internal energy—can possibly change. Putting $E_{\text{int}} = 0$ in the first law
$$Q = W \quad (\text{cyclical process})$$
- 2. Free expansion:** which no heat transfer occurs between the system and its environment and no work is done on or by the system. Thus, $Q = W = 0$, and the first law requires that
$$\Delta E_{\text{int}} = 0 \quad (\text{free expansion})$$

19.11: Some Specific Cases of the First Law of Thermodynamics

Table 18-5

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	$Q = W = 0$	$\Delta E_{\text{int}} = 0$

Example, First Law of Thermodynamics:

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 Fig. 18-17. The volume of that water changes from an initial value of 1.00×10^{-3} m³ as a liquid to 1.671 m³ as steam.

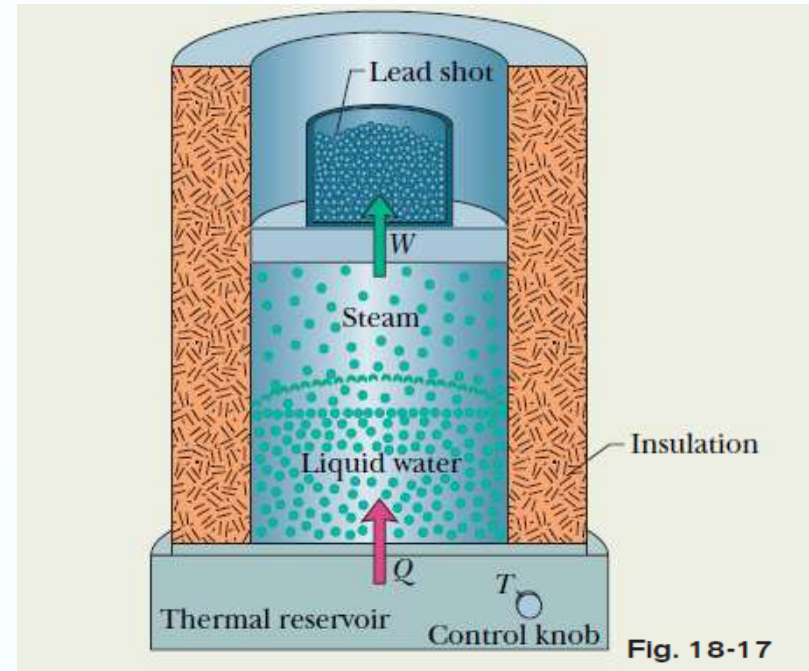
(a) How much work is done by the system during this process?

KEY IDEAS

(1) The system must do positive work because the volume increases. (2) We calculate the work W done by integrating the pressure with respect to the volume (Eq. 18-25).

Calculation: Because here the pressure is constant at 1.01×10^5 Pa, we can take p outside the integral. Thus,

$$\begin{aligned} W &= \int_{V_i}^{V_f} p \, dV = p \int_{V_i}^{V_f} dV = p(V_f - V_i) \\ &= (1.01 \times 10^5 \text{ Pa})(1.671 \text{ m}^3 - 1.00 \times 10^{-3} \text{ m}^3) \\ &= 1.69 \times 10^5 \text{ J} = 169 \text{ kJ.} \end{aligned} \quad (\text{Answer})$$



Example, First Law of Thermodynamics, cont.:

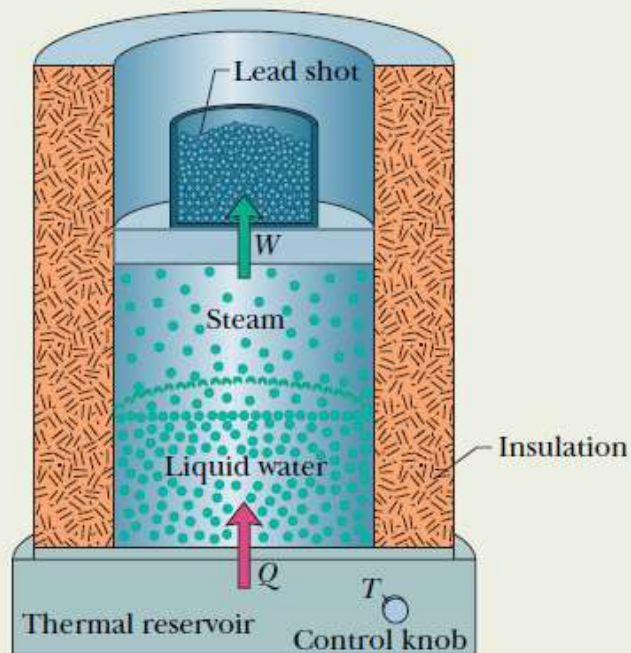


Fig. 18-17 Water boiling at constant pressure. Energy is transferred from the thermal reservoir as heat until the liquid water has changed completely into steam. Work is done by the expanding gas as it lifts the loaded piston.

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

KEY IDEA

Because the heat causes only a phase change and not a change in temperature, it is given fully by Eq. 18-16 ($Q = Lm$).

Calculation: Because the change is from liquid to gaseous phase, L is the heat of vaporization L_V , with the value given in Eq. 18-17 and Table 18-4. We find

$$\begin{aligned} Q &= L_V m = (2256 \text{ kJ/kg})(1.00 \text{ kg}) \\ &= 2256 \text{ kJ} \approx 2260 \text{ kJ}. \end{aligned} \quad (\text{Answer})$$

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

KEY IDEA

The change in the system's internal energy is related to the heat (here, this is energy transferred into the system) and the work (here, this is energy transferred out of the system) by the first law of thermodynamics (Eq. 18-26).

Calculation: We write the first law as

$$\begin{aligned} \Delta E_{\text{int}} &= Q - W = 2256 \text{ kJ} - 169 \text{ kJ} \\ &\approx 2090 \text{ kJ} = 2.09 \text{ MJ}. \end{aligned} \quad (\text{Answer})$$

This quantity is positive, indicating that the internal energy of the system has increased during the boiling process. This energy goes into separating the H_2O molecules, which strongly attract one another in the liquid state. We see that, when water is boiled, about 7.5% ($= 169 \text{ kJ}/2260 \text{ kJ}$) of the heat goes into the work of pushing back the atmosphere. The rest of the heat goes into the system's internal energy.

19.12: Heat Transfer Mechanisms: Conduction

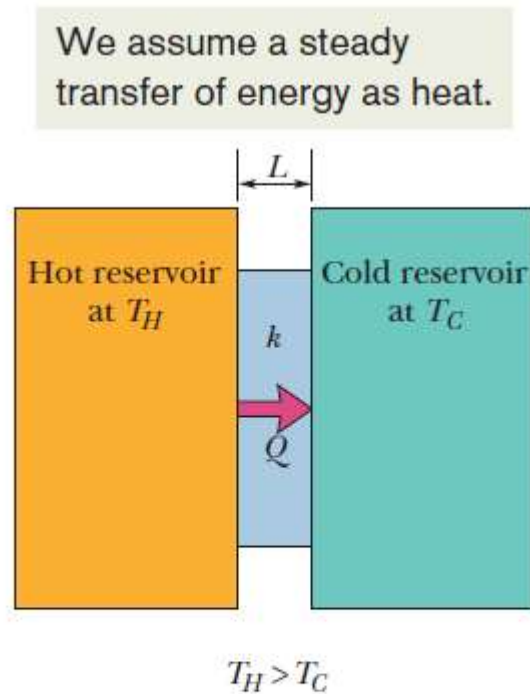


Fig. 18-18 Thermal conduction. Energy is transferred as heat from a reservoir at temperature T_H to a cooler reservoir at temperature T_C through a conducting slab of thickness L and thermal conductivity k .

A slab of face area A and thickness L , have faces maintained at temperatures T_H and T_C by a hot reservoir and a cold reservoir. If Q be the energy that is transferred as heat through the slab, from its hot face to its cold face, in time t , then the conduction rate P_{cond} (the amount of energy transferred per unit time) is

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

Here k , called the **thermal conductivity**, is a constant that depends on the material of which the slab is made.

The **thermal resistance** R of a slab of thickness L is defined as:

$$R = \frac{L}{k}.$$

1912: Heat Transfer Mechanisms: Conduction

Table 18-6

Some Thermal Conductivities

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

1912: Heat Transfer Mechanisms: Conduction

Letting T_X be the temperature of the interface between the two materials, we have:

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

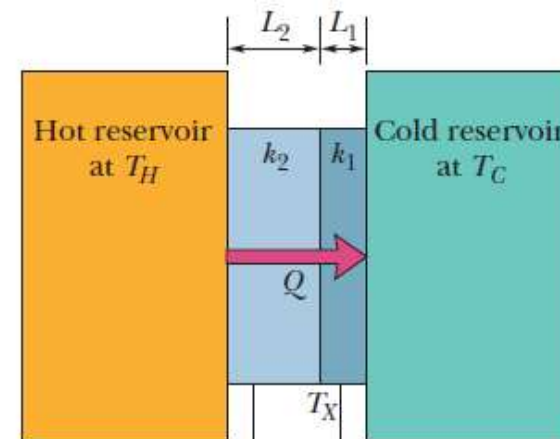
$$T_X = \frac{k_1 L_2 T_C + k_2 L_1 T_H}{k_1 L_2 + k_2 L_1}.$$

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

For n materials making up the slab,

$$P_{\text{cond}} = \frac{A (T_H - T_C)}{\Sigma (L/k)}.$$

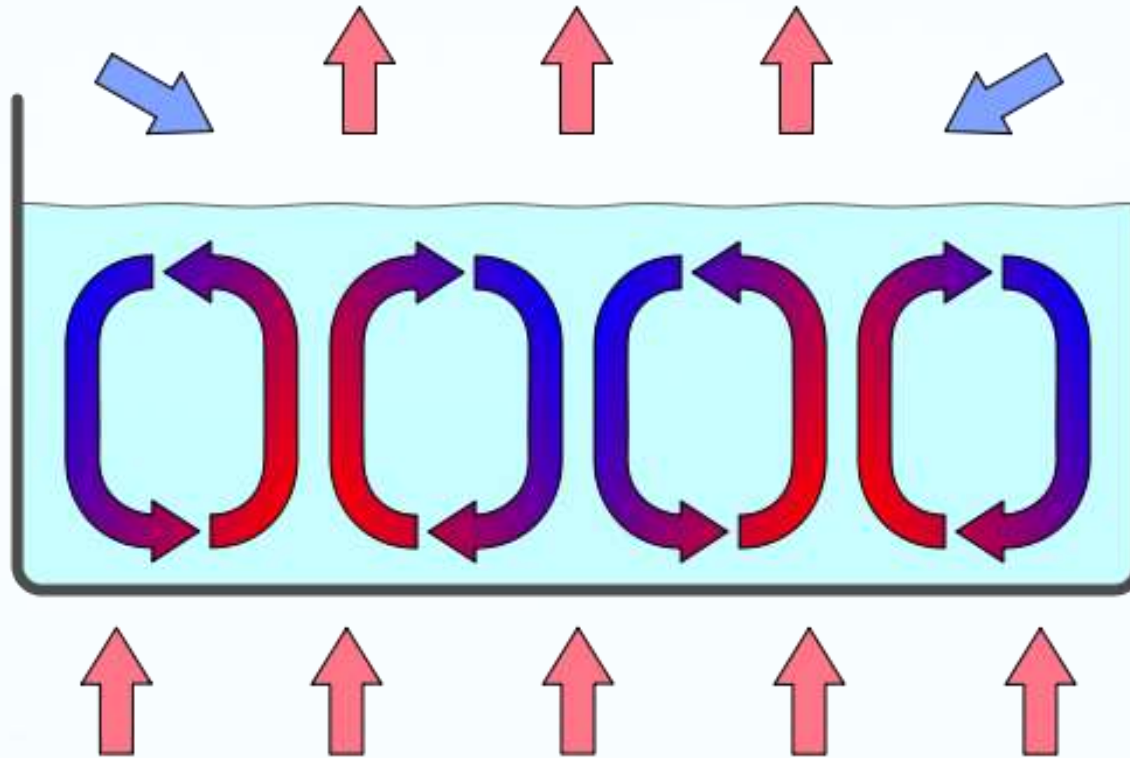
Fig. 18-19 Heat is transferred at a steady rate through a composite slab made up of two different materials with different thicknesses and different thermal conductivities. The steady-state temperature at the interface of the two materials is T_X .



The energy transfer per second here ...

... equals the energy transfer per second here.

19.12: Heat Transfer Mechanisms: Convection



In convection, thermal energy is transferred by bulk motion of materials from regions of high to low temperatures. This occurs when in a fluid a large temperature difference is formed within a short vertical distance (the temperature gradient is large).

19.12: Heat Transfer Mechanisms: Radiation



Fig. 18-20 A false-color thermogram reveals the rate at which energy is radiated by a cat. The rate is color-coded, with white and red indicating the greatest radiation rate. The nose is cool. (*Edward Kinsman/Photo Researchers*)

In radiation, an object and its environment can exchange energy as heat via electromagnetic waves. Energy transferred in this way is called **thermal radiation**.

The rate P_{rad} at which an object emits energy via electromagnetic radiation depends on the object's surface area A and the temperature T of that area in K, and is given by:

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

Here $\sigma = 5.6704 \times 10^{-8} \text{ W/m}^2 \text{ K}^4$ is called the **Stefan–Boltzmann constant**, and ε is the **emissivity**.

If the rate at which an object absorbs energy via thermal radiation from its environment is P_{abs} , then the object's net rate P_{net} of energy exchange due to thermal radiation is:

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

Example, Thermal Conduction Through a Layered Wall:

Figure 18-22 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 $T_1 = 25^\circ\text{C}$, $T_2 = 20^\circ\text{C}$, and $T_5 = -10^\circ\text{C}$. What is interface temperature T_4 ?

KEY IDEAS

(1) Temperature T_4 helps determine the rate P_d at which energy is conducted through the brick, as given by Eq. 18-32. However, we lack enough data to solve Eq. 18-32 for T_4 . (2) Because the conduction is steady, the conduction rate P_d through the brick must equal the conduction rate P_a through the pine. That gets us going.

Calculations: From Eq. 18-32 and Fig. 18-22, we can write

$$P_a = k_a A \frac{T_1 - T_2}{L_a} \quad \text{and} \quad P_d = k_d A \frac{T_4 - T_5}{L_d}.$$

Setting $P_a = P_d$ and solving for T_4 yield

$$T_4 = \frac{k_d L_d}{k_a L_a} (T_1 - T_2) + T_5.$$

Letting $L_d = 2.0L_a$ and $k_d = 5.0k_a$, and inserting the known temperatures, we find

$$\begin{aligned} T_4 &= \frac{k_a(2.0L_a)}{(5.0k_a)L_a} (25^\circ\text{C} - 20^\circ\text{C}) + (-10^\circ\text{C}) \\ &= -8.0^\circ\text{C}. \end{aligned} \quad (\text{Answer})$$

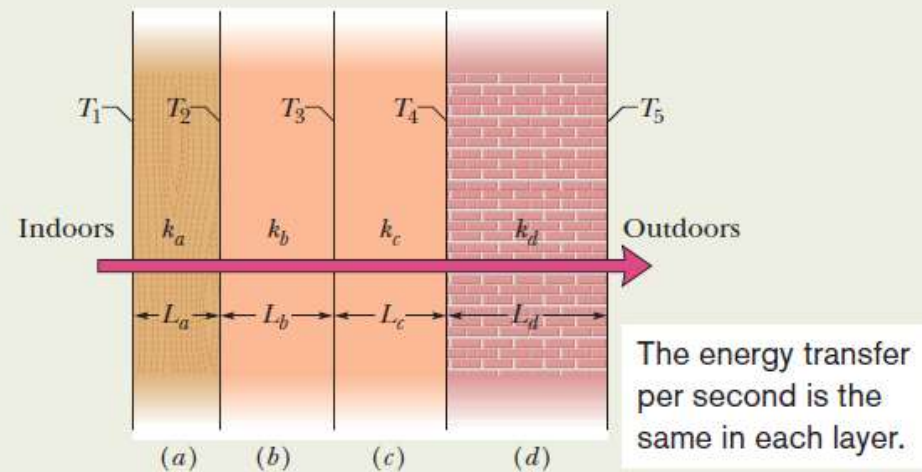


Fig. 18-22 Steady-state heat transfer through a wall.

Homework:

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