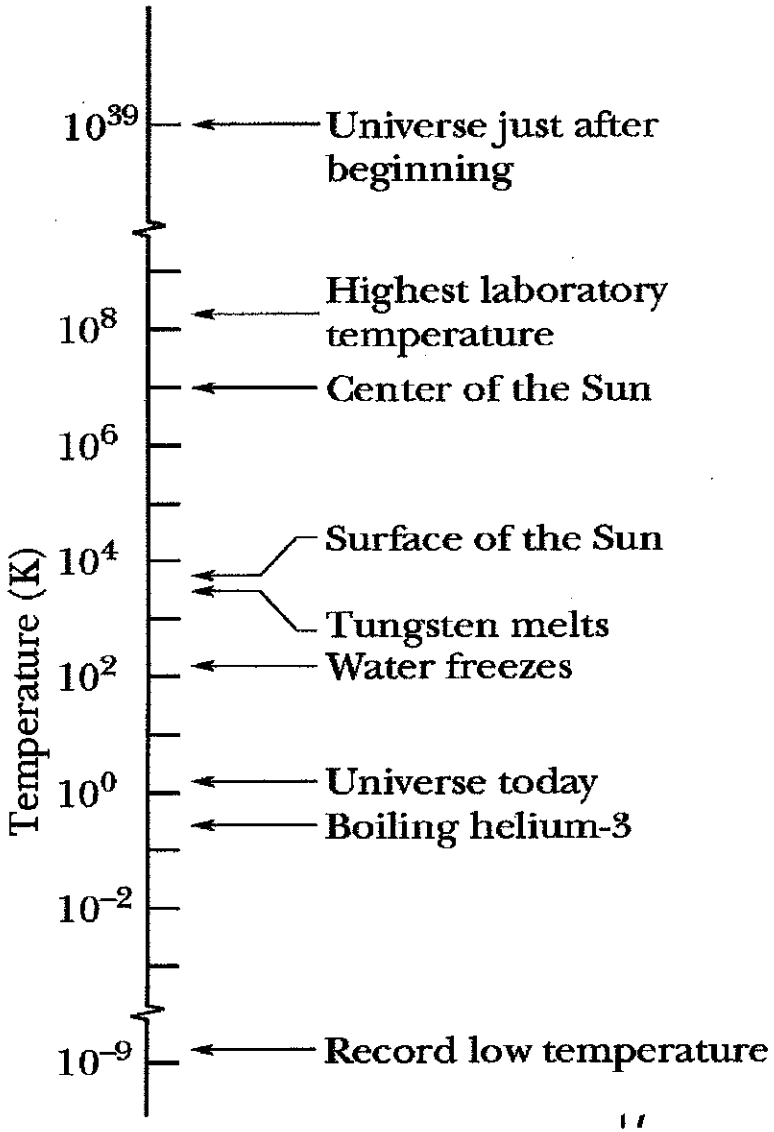
Thermodynamics & Temperature

- **Mechanics**: mechanical energy of systems governed by Newton's laws
- **Thermodynamics:** internal energy of bodies – *thermal energy*
- **Temperature**: central concept - sometimes a measure of the internal energy
  - Kelvin temperature - absolute temperature (see Fig 18-1)

T [K]



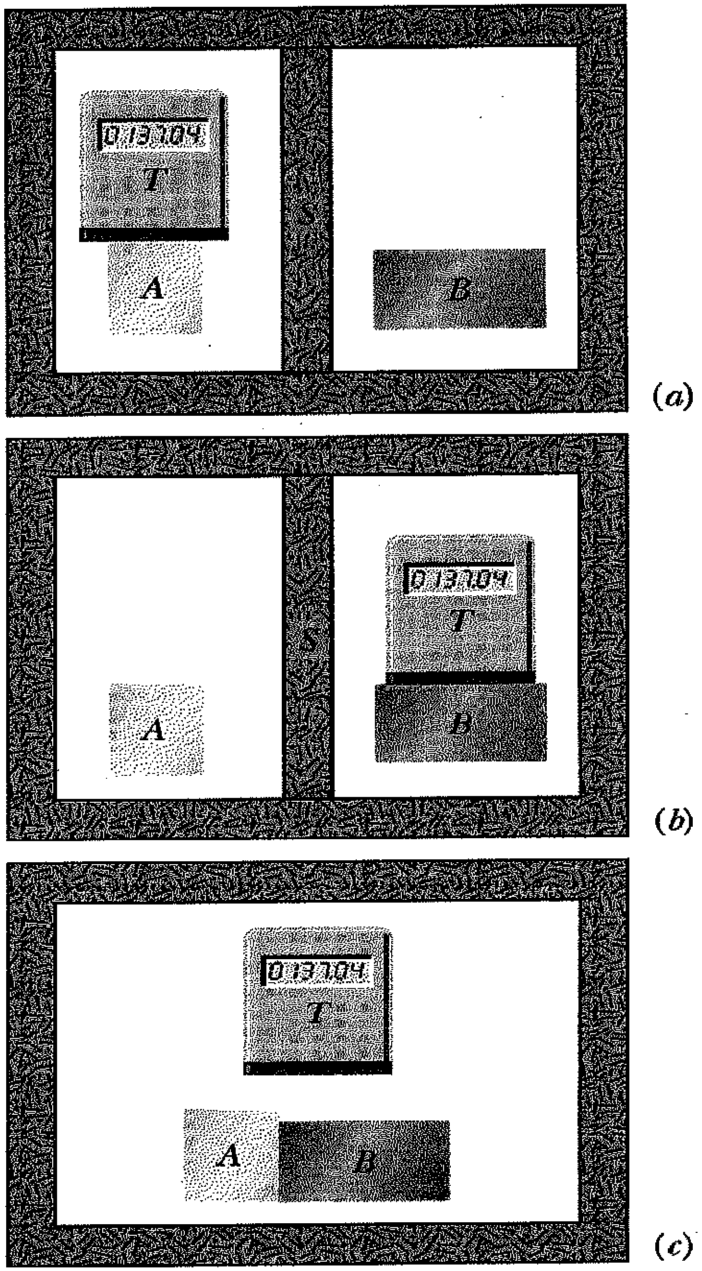
Celsius and Fahrenheit Scales

- Celsius scale: used worldwide

  $$T_C = T - 273.15\,K$$

  - 1° C = 1 K
  - 0° C => ice-water phase transition, normal pressure
  - 100° C => water-vapor phase transition
  - 37° C => normal human body temperature
  - 20° C => normal room temperature (68° F)

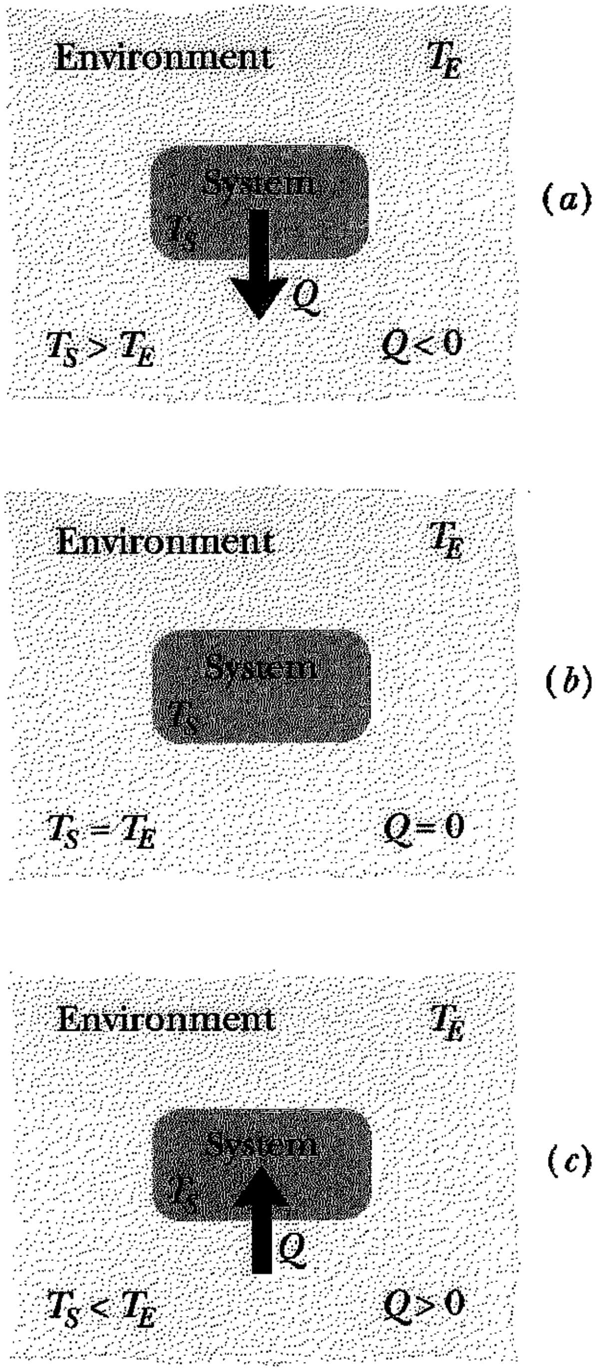
The Zeroth Law of Thermodynamics



- **Thermal equilibrium**: no net exchange of thermal energy
- **Thermoscop** (see Fig. 18-2)
- **Zeroth Law** (see Fig 18-3 =>):
  - *If bodies A and B are each in thermal equilibrium with a third body T, then A and B are in equilibrium with each other*
  - *Each body characterized by its* <u>*temperature*</u>

Temperature and Heat (Q)

- System vs Environment
- Change in temperature: transfer of internal energy, called ***heat***
- (see Fig 18-12 =>)
- ***Heat*** is the energy that is transferred between a system and its environment because of the temperature difference that exists between them



- **Units of heat (Q):**
  - *calorie* (cal): heat 1 gram of water from 14.5° C to 15.5° C
  - *British thermal unit* (Btu): heat 1 lb of water from 63° F to 64° F   1Btu = 1055,05 J
  - *Joule* (J): SI unit ; 1 cal = 4.186 J

$$1 \text{ cal} = 3.969 \times 10^{-3} \text{ Btu} = 4.186 \text{ J}$$

1 Food Calorie = 1,000 cal = 4186 J

Review: Absorption of Heat

- Heat Capacity $Q = C(T_f - T_i)$
- Specific Heat $Q = c\,m\,(T_f - T_i)$
- Molar Specific Heat $Q = \underset{C_V \;\; C_P}{C}\,n\,(T_f - T_i)$
- Heats of Transformation: $Q = L\,m$

L_V, L_F

Absorption of Heat

- **Heat Capacity**
  - capacity of a body to absorb heat
  - specific to one body

$$Q = C(T_f - T_i)$$

$$C = \frac{Q}{T_f - T_i} \quad \begin{matrix} \rightarrow [J] \\ \rightarrow [K] \end{matrix}$$

  - Units: cal/K, Btu/K, J/K

- **Specific Heat**
  - specific to units of mass

$$Q = c\,m(T_f - T_i)$$

Q [J], m [kg], c - given
ΔT -> in [K] or [°C]
does not matter

$$c = \frac{Q}{m\Delta T} = \frac{Q}{m(T_f - T_i)}$$

  - specific heat of water (for other Table 18-3)

$$c_w = 1 cal / g \cdot K = 1 Btu / lb \cdot F = 4190 J / Kg \cdot K$$

[redacted]

- **Heats of transformation**

phase changes

  - *phases* or *states*: solid, liquid, gas
  - When a *phase* transform into another (*phase transformation*) the temperature is constant
  - Needs ***heat of transformation***

$$Q = L m \qquad L = \frac{Q}{m}$$

; Q [J], m [kg]

  - Units for L: J/Kg
  - Examples (Table 18-4); Heat of vaporization $L_{V;}$ Heat of fusion L_F

TABLE 19-4 **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

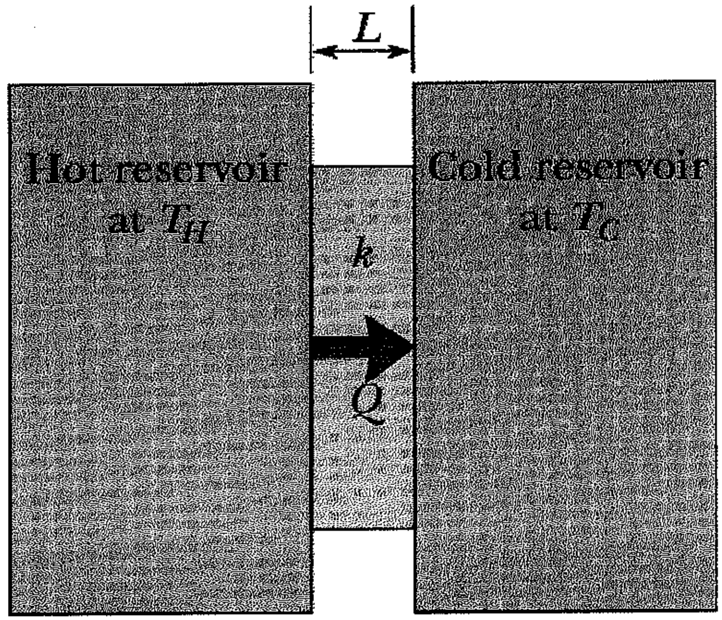
Heat Transfer Mechanisms

- ***Conduction***: through solids (metals) due to atomic vibration
  - – Conduction through a slab (Fig. 19-18)

$$P_{cond} = \frac{Q}{t} = k\,A\,\frac{T_H - T_C}{L}$$

J/s -> [W]

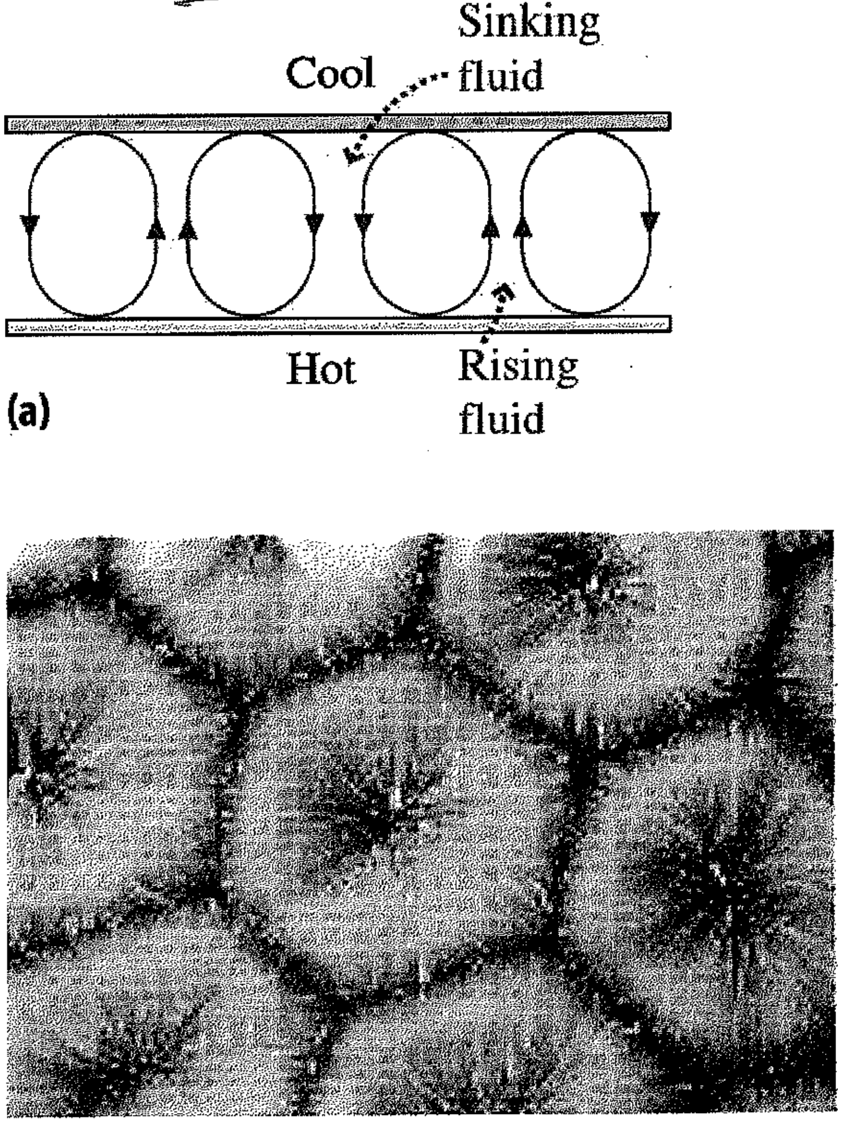
- P_{cond} – conduction rate (per unit time)
- k - thermal conductivity
- T_H (T_C) – hot(cold) reservoir temperature



$T_H > T_C$

Figure 13.19

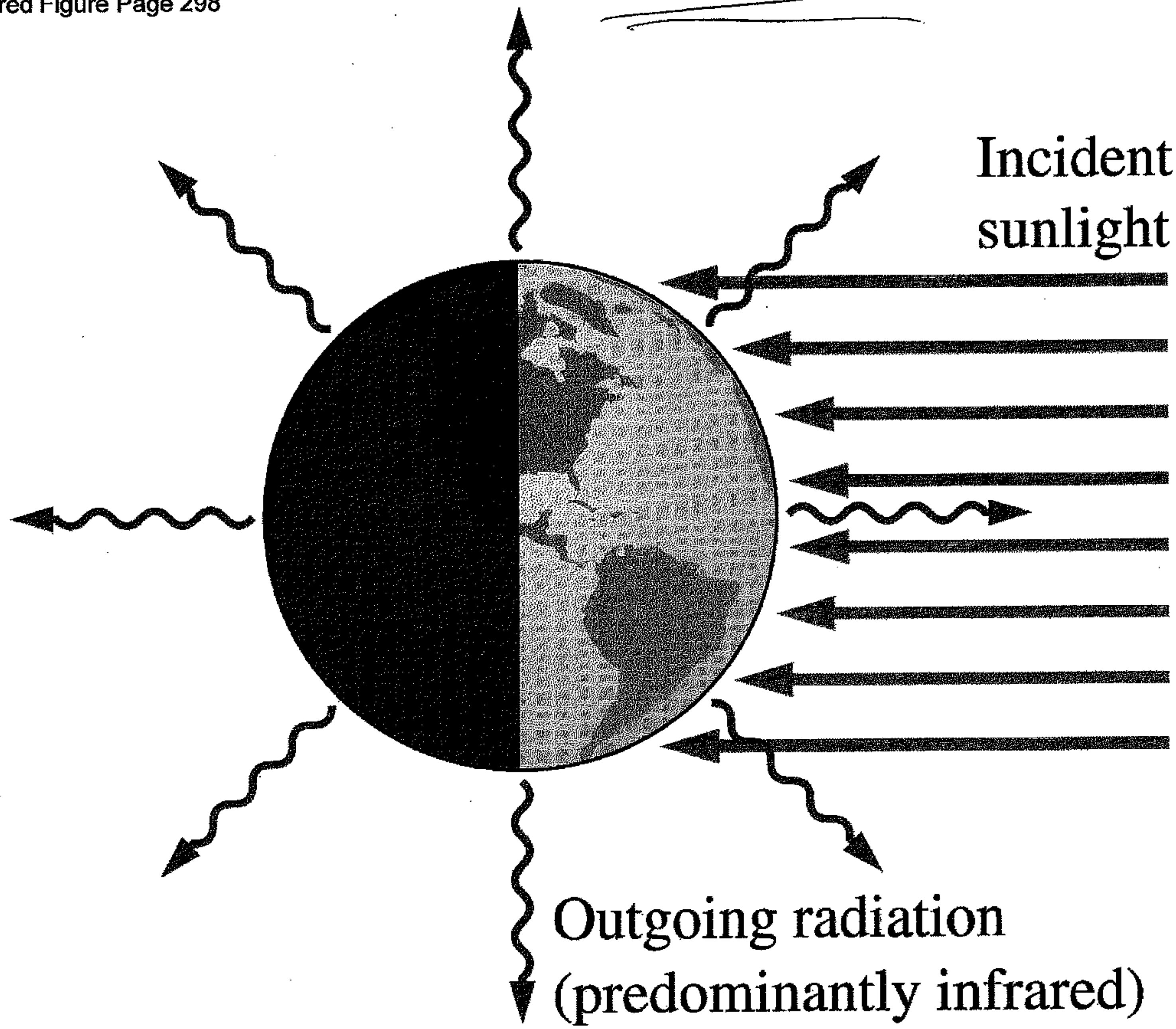
Convection



(b)

Radiation

Unnumbered Figure Page 298



- Convection – atmospheric convection
- Radiation –

T [K]

A -> area [m^2]

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

where $\sigma = 5.673 \times 10^{-8}$ W/m^2 K

(Stefan-Boltzmann)

$$P_{abs} = \sigma \varepsilon A T^4_{(env)}$$

$$P_{net} = P_{abs} - P_{rad} = \sigma \varepsilon A \left(T^4_{(env)} - T^4\right)$$

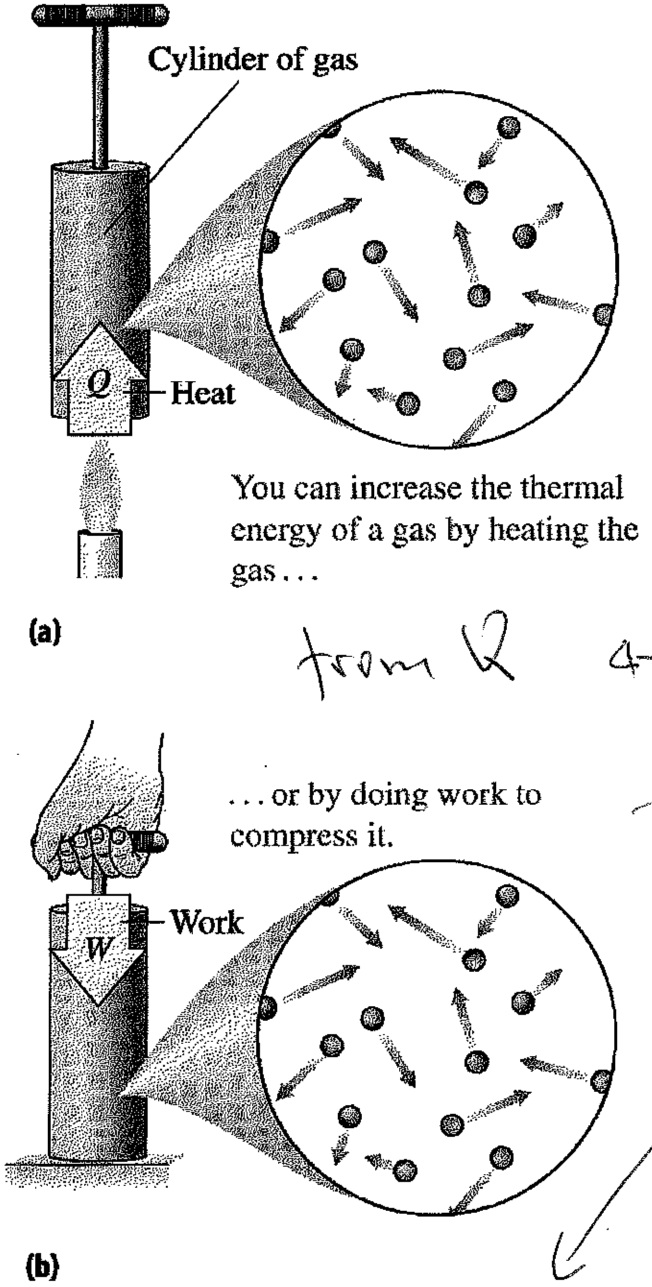
ε – emissivity [0 - 1]

Laws of Thermodynamics

$U_{internal} \to E_{th}$

$\to E_{bonds}$

Figure 14.1



1st Law

$$\boxed{\Delta U = Q + W}$$

no bonds

Monoatomic gas

$$\Delta U = \Delta E_{th} = \frac{3}{2} N k_B \Delta T$$

from Q ←

Diatomic gas

single bond

+2 degrees of freedom

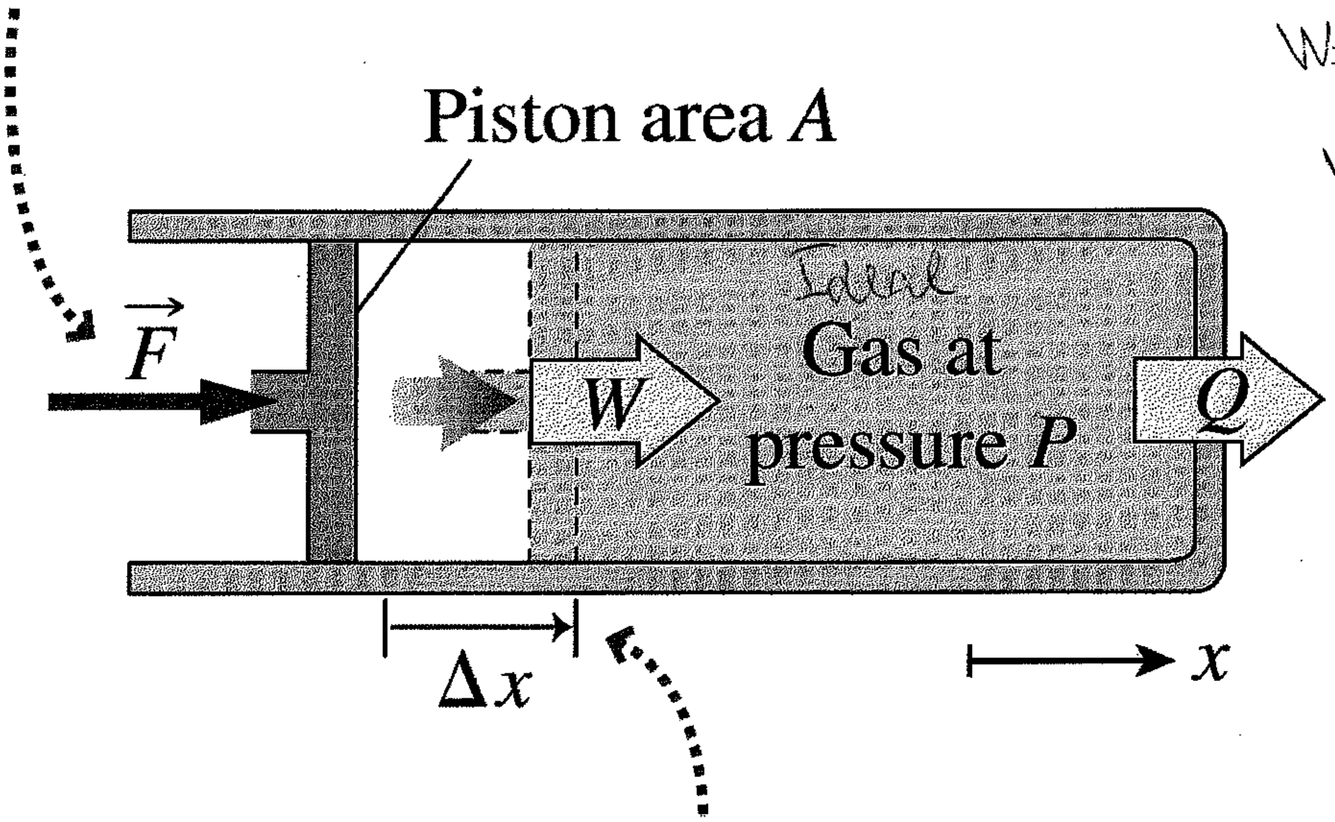
$$\Delta U = \Delta E_{th} = \frac{5}{2} N k_B \Delta T$$

from Q

$P =$ const. (Isobaric process) $P =$ const

Figure 14.3

A constant force is applied and heat is allowed to escape, so the pressure remains constant as the gas is compressed.



$W = F_x . \Delta X = P . A . \Delta X$

$V = A . X$

$\Delta V = - A . X$

$W = - P . \Delta V$

where

$\Delta V = V_f - V_i$

$\Delta V < 0$

The piston moves through displacement Δx, resulting in work $W = F_x \Delta x = PA\Delta x$. so $W > 0$ for fluid

Here $\boxed{\Delta U = Q + W}$

T= const. Isothermal process

Figure 14.4

In an isothermal process, the temperature of the gas is held constant while the gas is compressed (or expands).

From $p = nRT/V$

$p \cdot V = nR \cdot T$

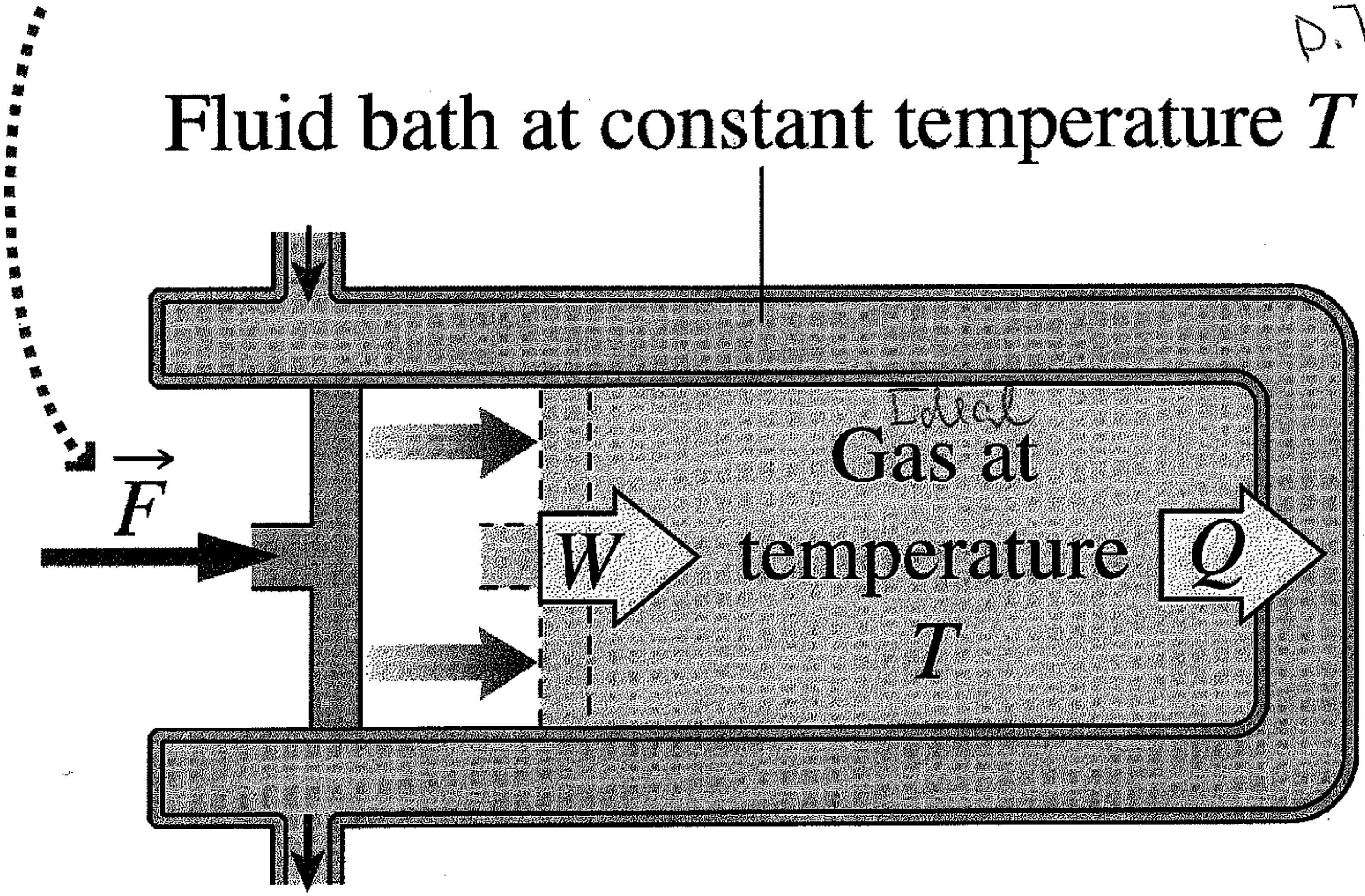
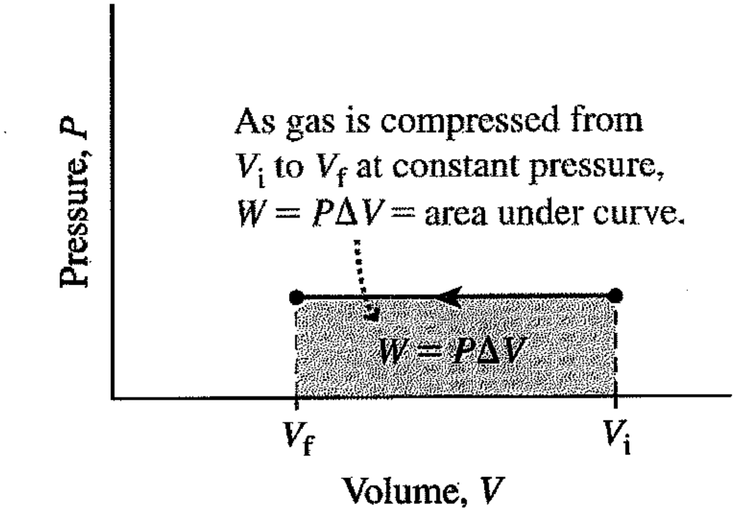


Figure 14.5



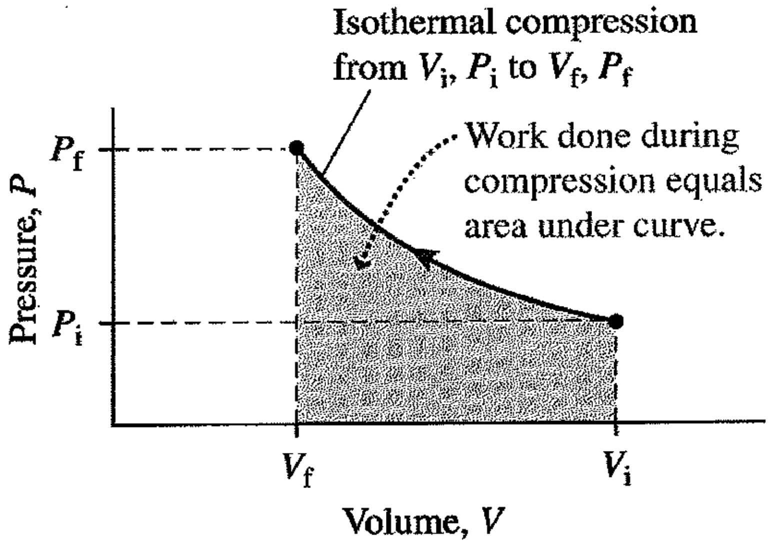
(a) Compression at constant pressure

For $P = \text{const}$

$W = -P \cdot \Delta V$

$\Delta V = V_f - V_i < 0$

so $W > 0$ for fluid



(b) Isothermal compression

$P \cdot V = nRT$

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

$W = nRT \ln\left(\frac{V_i}{V_f}\right)$

for compression $V_i > V_f$ so $W > 0$

for expansion $V_i < V_f$ so $W < 0$

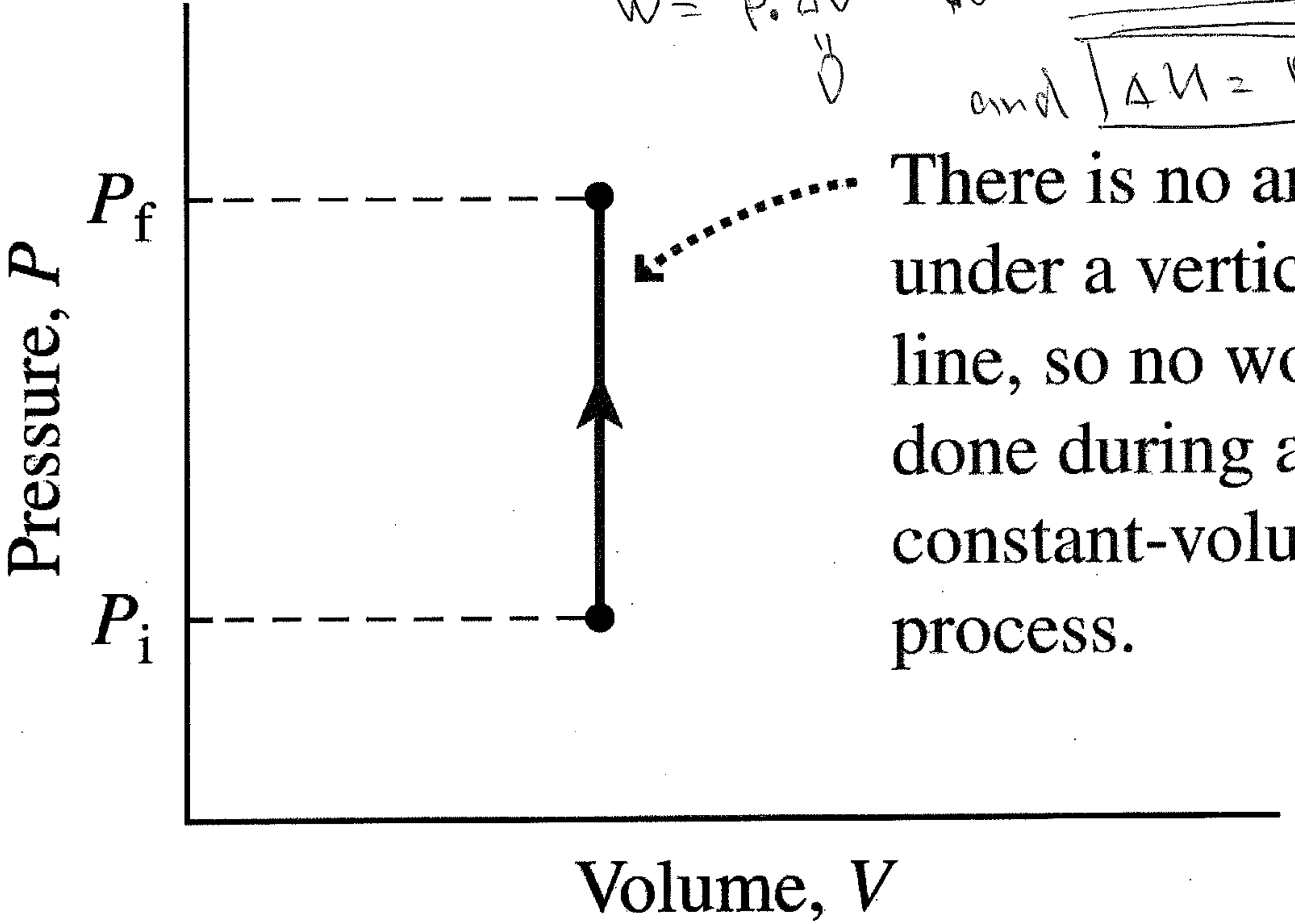
For ideal gas T const $\rightarrow \Delta U = 0$ so $Q + W = 0$

Figure 14.7

Constant Volume

$W = P \cdot \underset{0}{\Delta V}$ so $W = 0$

and $\boxed{\Delta U = Q}$



There is no area under a vertical line, so no work is done during a constant-volume process.

Figure 14.8

Adiabatic process $Q=0$

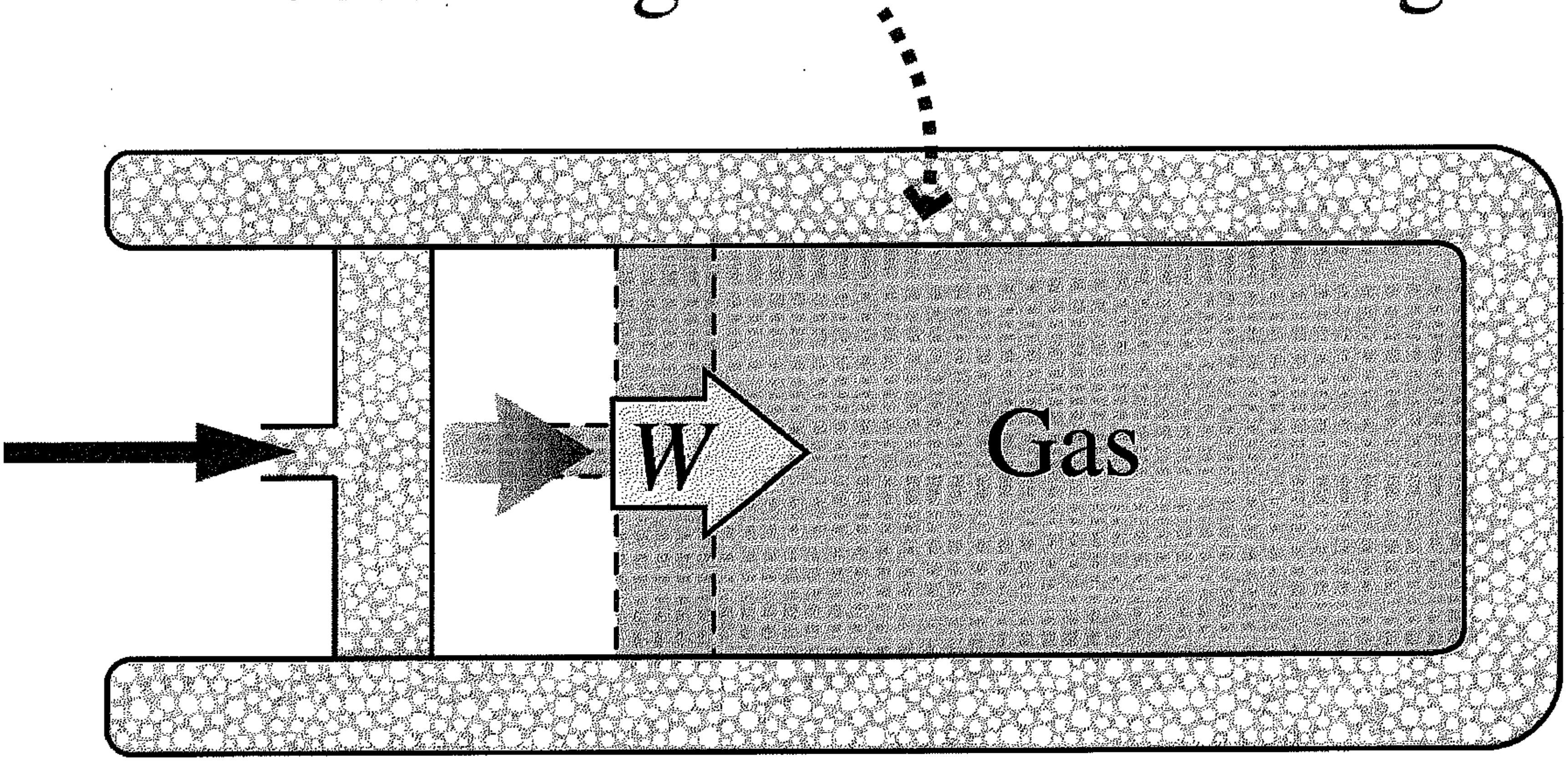
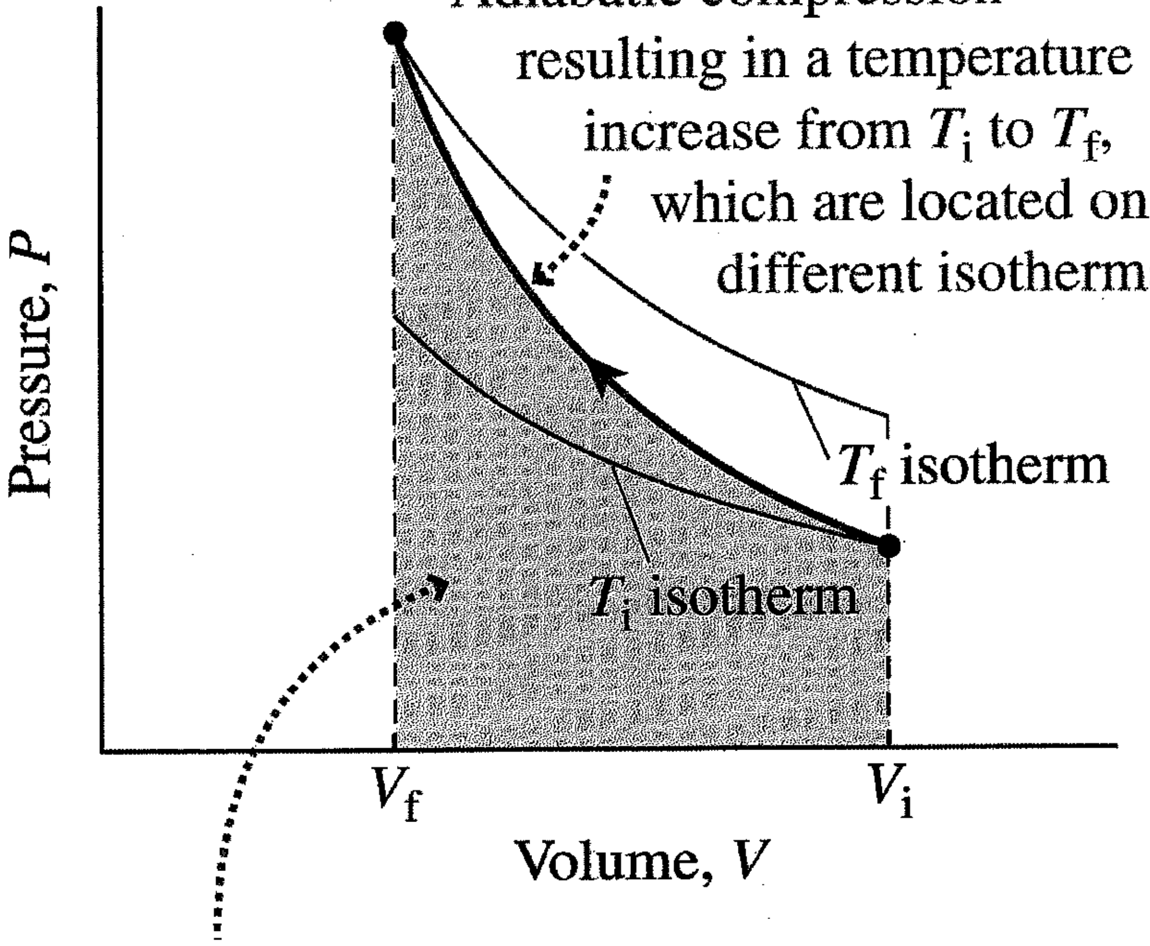

Insulation prevents heat flow between gas and surroundings.

Figure 14.9



Adiabatic compression resulting in a temperature increase from T_i to T_f, which are located on different isotherms

Area under adiabatic curve is greater than area under T_i isotherm, so adiabatic compression requires more work than equivalent isothermal compression.

$Q = 0$

so $\boxed{\Delta U = W}$

$P \cdot V^{\gamma} = \text{const}$

$\gamma = C_p / C_v$

For monoatomic

$C_p = 5R/2$

$C_v = 3R/2$

so $\gamma = 5/3$

For diatomic

$\gamma = 7/5$

$$W = \frac{P_f V_f - P_i V_i}{\gamma - 1}$$

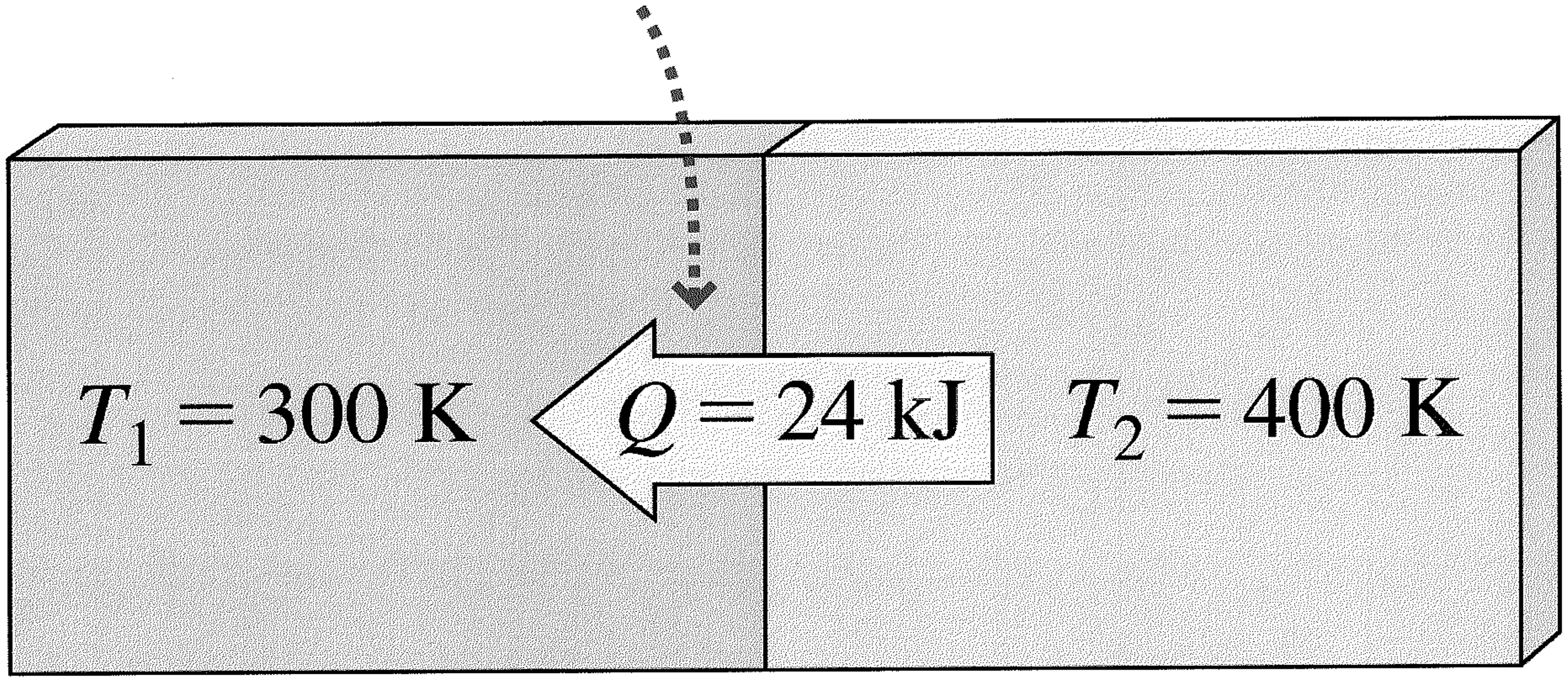
Reviewing New Concepts: The First Law of Thermodynamics in Thermal Processes

Process	Work W	First law accounting, with $\Delta U = Q + W$
Constant pressure	$W = -P\Delta V$	$\Delta U = Q - P\Delta V$
Constant temperature (isothermal)	$W = nRT \ln\left(\frac{V_i}{V_f}\right)$	$\Delta U = 0$ $Q = -W = -nRT \ln\left(\frac{V_i}{V_f}\right)$
Constant volume	$W = 0$	$\Delta U = Q$
Adiabatic ($Q = 0$)	$W = \frac{P_fV_f - P_iV_i}{\gamma - 1}$	$\Delta U = W = \frac{P_fV_f - P_iV_i}{\gamma - 1}$

Figure 14.10

Second Law (Direction of physical processes, also Q.)

Heat Q flows from warmer to cooler object.



2) **Second law of thermodynamics:** Heat flows spontaneously only from hotter to cooler objects.

$\left[\frac{J}{K}\right] \leftarrow$ $\Delta S = \frac{Q}{T}$ (Entropy change; SI unit: J/K)

ΔS always increases

ΔS can be ± depending on ±Q

$$\Delta S_{\text{cool}} = \frac{Q}{T_{\text{cool}}} = \frac{24{,}000\ \text{J}}{300\ \text{K}} = 80\ \text{J/K}$$

Heat flows *out* of the hot object, so $Q = -24{,}000$ J, and

$$\Delta S_{\text{hot}} = \frac{Q}{T_{\text{hot}}} = \frac{-24{,}000\ \text{J}}{400\ \text{K}} = -60\ \text{J/K}$$

The net entropy change of the whole system is

$$\Delta S_{\text{total}} = \Delta S_{\text{cool}} + \Delta S_{\text{hot}} = 80\ \text{J/K} - 60\ \text{J/K} = +20\ \text{J/K}$$

Second law of thermodynamics: Heat flow is accompanied by an increase in the entropy of the universe.

Second law of thermodynamics: Natural processes evolve toward a state of maximum entropy.

Heat Engines

- Elements of an engine (Fig 20-7)
  - Heat Q_H is transferred from the hot reservoir of temperature T_H to the working substance
  - Heat Q_L is transferred from the working substance to the cold reservoir T_L

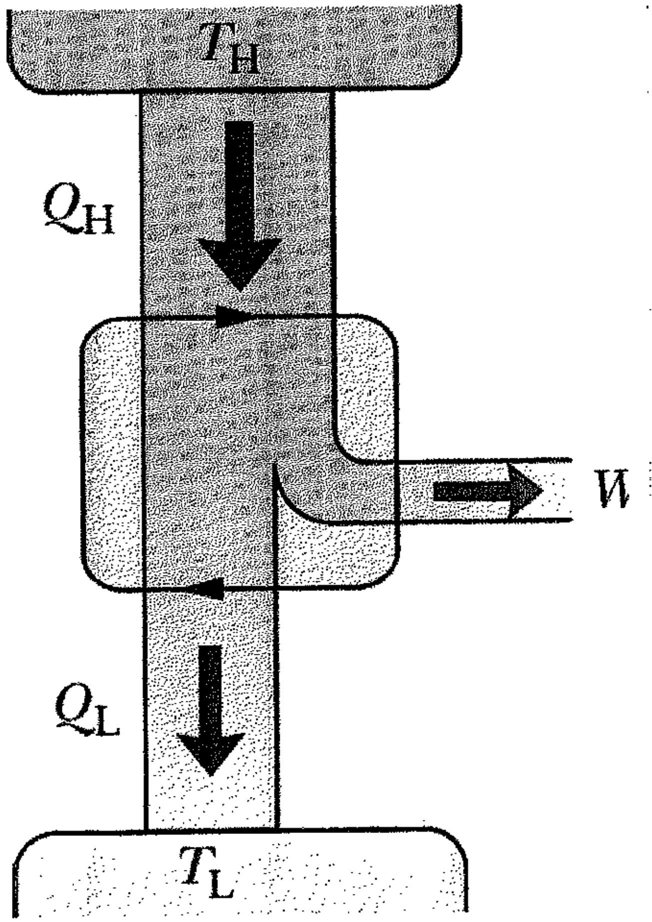
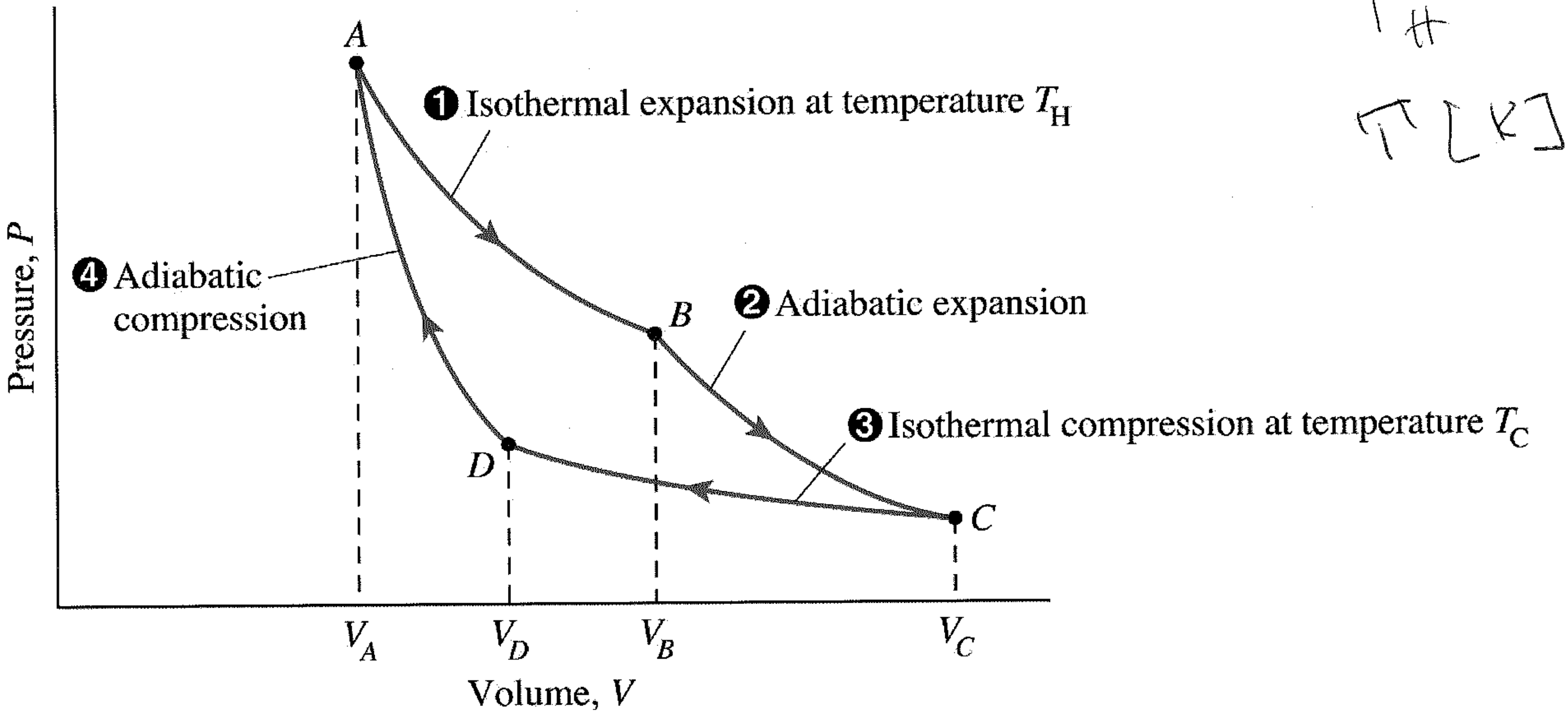


Figure 14.12

Carnot Engine (How good engines could be?)

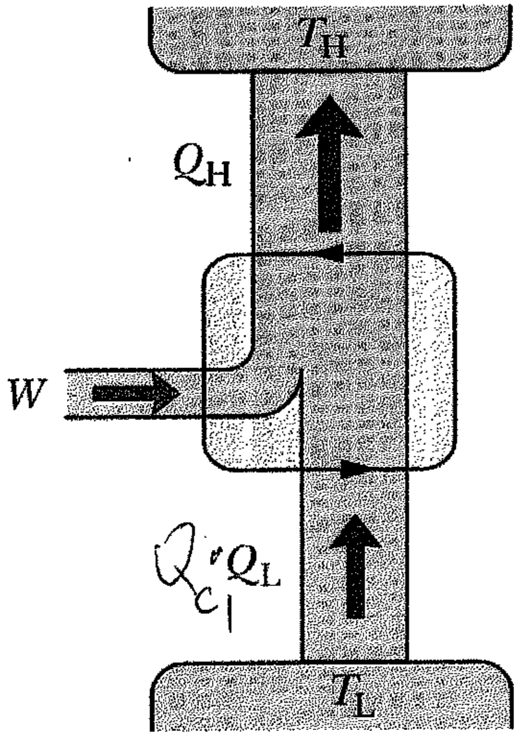
$$e_{Carnot} = 1 - \frac{T_C}{T_H}$$

T [K]



Refrigerators

- **Refrigerator:** device that uses work to transfer thermal energy from the low-temperature reservoir to the high-temperature reservoir (Fig 20-13)
- **Ideal refrigerator:** processes involved in the refrigerator's operations are reversible



- Coefficient of performance:

$$K = \frac{what\,we\,want}{what\,we\,pay\,for} = \frac{|Q_L|}{|W|}$$

$$|W| = |Q_H| - |Q_L|$$

$$K = \frac{|Q_L|}{|Q_H| - |Q_L|}$$

- Carnot (ideal) refrigerator:

$$COP = K_C = \frac{T_L}{T_H - T_L}$$; $COP > 1$; T [K]

200 Cal → J 1 Food C = 4186 J

22. SOLVE The conversion between food calories to Joules gives:

$$200\text{ Cal}\left[\frac{4186\text{ J}}{1\text{ Cal}}\right]=837\text{ kJ}$$

A piece of metal absorbs 3.6 kJ of heat, increasing its temperature by 33 °C. What is its heat capacity ?

By definition $Q = C\Delta T$ so heat capacity $C = Q / \Delta T$

In our case $Q = 3.6$ kJ and $\Delta T = 33$ °C

Therefore $C = (3600\ \text{J})/(33\ °\text{C}) = 109.1\ \text{J}/°\text{C}$

An iron rod of mass 0.5 kg is at temperature of 20 °C. How much heat, Q, in Joules must it absorb so that its temperature raises to 80 °C ?

By definition $Q = m.c.\Delta T$ where specific heat of iron $c_{Fe} = 449$ J/(kg. °C)

Therefore, $Q = (0.5 \text{ kg}).\ (449).(80 - 20) = 13470$ J

2.3mL Hg ~~&~~

at 0°C

| Q=? so T↑ 100°C

55. **ORGANIZE AND PLAN** This is a straightforward use of Equation 13.2: $Q = mc\Delta T$, where the specific heat of mercury is from Table 13.1: $c = 140\ \text{J/kg}\cdot{}^\circ\text{C}$. → c_{Hg} The one thing we will need is the density of liquid mercury from Table 10.1: $\rho = 13{,}600\ \text{kg/m}^3$.

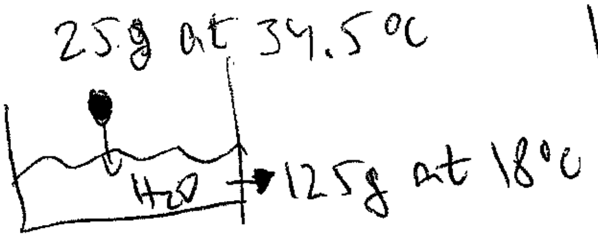
$1L = 10^{-3}\ m^3$, $1mL = 10^{-3} L \rightarrow 1mL = 10^{-6}\ m^3$

Known: $V = 2.30$ mL, $\Delta T = 100^\circ$C.

SOLVE Plugging the mass of mercury $(m = \rho V)$ into Equation 13.2:

$$Q = \rho V c \Delta T = (0.0136\ \text{kg/mL})(2.30\ \text{mL})(140\ \text{J/kg}\cdot{}^\circ\text{C})(100^\circ\text{C}) = 438\ \text{J}$$

REFLECT This heat causes the mercury to expand slightly, which results in the liquid rising inside the thermometer. Because this rise is uniform, we can use it to measure the temperature. In this way, the thermometer works simply by absorbing heat (or losing heat) to the environment.



$T_{eq} = 22.9\ °C$

$C_M = ?$

Boon

53. ORGANIZE AND PLAN We know that the sum of the heat lost by the material and gained by the water is zero: $Q_M + Q_W = 0$, so we'll use that to solve for the unknown specific heat: c_M. The temperature changes for the material and the water are:

$$\Delta T_M = 22.9°\text{C} - 34.5°\text{C} = -11.6°\text{C}$$
$$\Delta T_W = 22.9°\text{C} - 18°\text{C} = 4.9°\text{C}$$

$$-Q_M = +Q_{H_2O}$$

Known: $m_M = 25.0$ g, $m_W = 125$ g.

SOLVE Using the heat exchange and Equation 13.2:

$$c_M = -\frac{m_W c_W \Delta T_W}{m_M \Delta T_M} = -\frac{(0.125\text{ kg})(4186\text{ J/kg°C})(4.9°\text{C})}{(0.0250\text{ kg})(-11.6°\text{C})} = 8841\text{ J/kg°C}$$

$$\boxed{Q = mC\Delta T}$$

REFLECT Looking through Table 13.1, there's no material that matches this specific heat. But of course this list is not exhaustive, so we shouldn't be concerned.

$$-m_M C_M \Delta T_M = +m_W C_W \Delta T_W$$

$$C_M = -\frac{m_W C_W \Delta T_W}{m_M \Delta T_M}$$

$C_W = 4186\ \text{J/kg}\cdot°\text{C}$

↓

Table 13.1

45. You mix 18 kg of water at 25 °C with 6 kg of water at 2 °C, what is the final temperature ?

All we can say is that the hotter water changes temperature by: $\Delta T_{\text{hot}} = T_f - 25^\circ\text{C}$, while the colder water changes temperature by: $\Delta T_{\text{cold}} = T_f - 2.0^\circ\text{C}$. We will be able to solve for T_f using Equation 13.2,

$$\text{i.e. } Q = c.m.\,\Delta T$$

and the fact that the heat lost by the hot water is gained by the cold water: $Q_{\text{hot}} = -Q_{\text{cold}}$, assuming of course that no heat is lost to the surroundings.

Known: $m_{\text{hot}} = 18\text{ kg}$, $m_{\text{cold}} = 6\text{ kg}$. and $c_{\text{water}} = 4186\text{ J/(kg.}^\circ\text{C)}$ (Table 13.1)

SOLVE The equal but opposite heat exchange implies:

$$Q_{\text{hot}} = -Q_{\text{cold}} \Rightarrow m_{\text{hot}} c \Delta T_{\text{hot}} = -m_{\text{cold}} c \Delta T_{\text{cold}}$$

Solving for the final temperature:

$$T_f - 25^\circ\text{C} = -\frac{6\text{ kg}}{18\text{ kg}}(T_f - 2.0^\circ\text{C}) \Rightarrow T_f = 19^\circ\text{C}$$

56g N_2 at 25°C + 12g He at 45°C

$T_{eq} = ?$

57. **ORGANIZE AND PLAN** The nitrogen starts off colder, so it will gain heat from the helium: ($Q_N = -Q_{He}$). We'll assume that the gases are combined under fixed pressure, so that the heat gained or lost will come from Equation 8.4: $Q = nc_P\Delta T$. The molar specific heats can be taken from Table 13.2 for nitrogen ($c_{N_2} = 29.1$ J/mol·C) and for helium ($c_{He} = 20.8$ J/mol·C). We'll need to convert the given masses into moles, and write the temperature change for the nitrogen as: $\Delta T_N = T_f - 25°\text{C}$, and the helium as: $\Delta T_{He} = T_f - 45°\text{C}$.
Known: $m_N = 56$ g, $m_{He} = 12$ g.

SOLVE The equal but opposite heat exchange implies:

$$-Q_N = +Q_{He} \Rightarrow n_N c_N \Delta T_N = -n_{He} c_{He} \Delta T_{He}$$

56g N_2 = n · 28 $\frac{g}{mol}$

12g He = n · 4 $\frac{g}{mol}$

The molar masses are 28 g/mol for nitrogen gas and 4 g/mol for helium gas, so the number of moles are 2 mol of nitrogen and 3 mol of helium. Solving for the final equilibrium temperature:

$$\overbrace{T_f - 25°\text{C}}^{\Delta T\ N_2} = -\frac{(3\text{ mol})(20.8\text{ J/mol}°\text{C})}{(2\text{ mol})(29.1\text{ J/mol}°\text{C})}\overbrace{(T_f - 45°\text{C})}^{\Delta T\ He.} \Rightarrow \underline{\underline{T_f = 35°\text{C}}}$$

$n_{N_2} = 2$

$n_{He} = 3$

REFLECT The answer makes sense, since the final temperature is halfway between the initial temperatures of the nitrogen and the helium. If you assumed that the gases were mixed with constant volume, the result would be practically the same: $T_f = 34°\text{C}$. This is because the ratio of the molar specific heats (c_{He}/c_N) is practically the same for constant volume and constant pressure.

Table 13.2 | He → $c_p = 20.8$ mol · °C

N_2 → $c_p = 29.1$ mol · °C

How much heat, Q, is required to melt 500 g of ice at 0 °C ?

By definition $Q = mL_f$ where the heat of fusion of ice is $L_{ice} = 3.33 \times 10^5$ J/kg

Therefore, $Q = (0.5\text{ kg}).(3.33 \times 10^5) = 166.5\text{ kJ}$

Sample Problem 18-3 (cont)

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

$Q_{12} = 15.98$ kJ, $Q_F = 239.8$ kJ, $Q_{23} = 45.25$ kJ

Final state: ICE and WATER, $T_f = 0^o$ C

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

$Q_{rem} = m_w L_F$

$m_w = Q_{rem} / L_F = 194 \text{ kJ}/333 \text{ kJ/kg} = 0.583 \text{ kg}$

$m_{ice} = 0.720 \text{ kg} - 0.583 \text{ kg} = 0.137 \text{ kg}$

Sample Problem 18-3

(a) How much heat must be absorbed by ice of mass $m = 720$ g at $T_1 = -10^oC$ to take it to liquid state at $T_3 = 15^oC$?

Let $T_2 = 0^oC$. Then

$Q_{12} = c_{ice}m(T_2 - T_1) = (2{,}220 \text{ J/kg K})(0.72 \text{ kg})[0^oC - (-10^oC)]$

$= 15{,}984 \text{ J} = 15.98 \text{ kJ}$

Table 13.3

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

Table 13.1

$Q_{23} = c_w\, m\, (T_3 - T_2) = (4{,}190 \text{ J/kg K})(0.720 \text{ kg})(15^oC - 0^oC)$

$= 45{,}252 \text{ J} = 45.25 \text{ kJ}$

$Q = Q_{12} + Q_F + Q_{23} = 15.98 \text{ kJ} + 239.8 \text{ kJ} + 45.25 \text{ kJ} = 300 \text{ kJ}$

P 25. Calculate the amount of energy, in joules, required to completely melt 130 g silver initially at 15 C°.

$T\,K = T\,°C + 273.15$

- 25. The melting point of silver is 1235 K, so the temperature of the silver must first be raised from 15.0° C (= 288 K) to 1235 K. This requires heat

$$Q = cm(T_f - T_i) = (236\,\text{J/kg}\cdot\text{K})(0.130\,\text{kg})(1235°\text{K} - 288°\text{K}) = 2.91\times10^4\ \text{J}.$$

- Now the silver at its melting point must be melted. If *LF* is the heat of fusion for silver this requires

$$Q = mL_F = (0.130\,\text{kg})(105\times10^3\ \text{J/kg}) = 1.36\times10^4\ \text{J}.$$

- The total heat required is (2.91 x 10^4 J + 1.36 x 10^4 J) = 4.27 x 10^4 J.

A window with area of 2 m^2 is made of 4 mm thick glass. If it is 20 °C colder outside than inside, what is the heat flow rate, *H*, through the window ?

By definition $H = k.\ A.\ (\Delta T/\Delta x)$

In our case thermal conductivity of glass $k = 0.8\ W/(C^{o}.m)$, $A = 2\ m^2$, $\Delta T = 20\ ^{o}C$ and $\Delta x = 0.004$ m

Therefore, $H = (0.8).(2).(20)/(0.004) = 8000$ W

Boou $\rightarrow | ? | \leftarrow$ concrete   So $R_c = R_{wood}$ (1.8 cm)

94. **ORGANIZE AND PLAN** We are asked to compare the insulation provided by two materials. In Problems 13.78 through 13.81, we were introduced to the *R*-value of a material, which is defined as the thickness divided by the thermal conductivity: $R = \Delta x / k$. To find the R-values for concrete and wood walls, we'll need their respective thermal conductivities from Table 13.4: $k_C = 1.28\ \text{W/°C}\cdot\text{m}$ and $k_W = 0.12\ \text{W/°C}\cdot\text{m}$.

$$\frac{\Delta x_c}{K_c} = \frac{\Delta x_{wood}}{K_{wood}} = R$$

Known: $\Delta x_W = 1.8$ cm.

SOLVE For a concrete wall to have the same R-value of a piece of wood, its thickness would need to be:

$$\Delta x_C = \Delta x_W \frac{k_C}{k_W} = (1.8\ \text{cm}) \frac{(1.28\ \text{W/°C}\cdot\text{m})}{(0.12\ \text{W/°C}\cdot\text{m})} = 19.2\ \text{cm}$$

REFLECT To get the same insulation as a piece of wood, the thickness of a concrete wall has to be over ten times thicker. That's because concrete lets heat escape faster than wood does. Think of a cold concrete floor vs. a warm wooden one.

$$K_c \cancel{A} \frac{\cancel{\Delta T}}{\Delta x_c} = H = K_w \cancel{A} \frac{\cancel{\Delta T}}{\Delta x_w}$$

$$\frac{K_c}{\Delta x_c} = \frac{K_w}{\Delta x_w}$$

$$\Delta x_c = \Delta x_w \frac{K_c}{K_w}$$

A sphere of surface area 2 m^2 and emissivity of 0.5 is at temperature of 300 °C. What is the rate at which the sphere radiates heat into empty space ?

The rate, P, at which an object at temperature T radiates energy is:

$P = e.\ \sigma.\ A.\ T^4$, where $\sigma = 5.67 \times 10^{-8}\ W/(m^2.K^2)$

In our case $e = 0.5$, $A = 2\ m^2$ and $T = 300 + 273.15 = 573.15$ K

Therefore,

$P = (0.5).\ (5.67 \times 10^{-8}).(2).\ (573.15)^4 = 6118$ W

$R_{sun} = 6.96 \times 10^8$ m

$T_{sun} = 5800$ K

Book

82. **ORGANIZE AND PLAN** The Stefan-Boltzmann law (Equation 13.8) tells us the rate at which a body radiates energy: $P = e\sigma AT^4$. Saying the Sun is a blackbody means that its emissivity is one, i.e.: $e = 1$. The Stefan-Boltzmann constant is: $\sigma = 5.67\times10^{-8}\ \text{W/m}^2\text{K}^4$, and the surface area of a sphere is: $A = 4\pi r^2$.
Known: $r = 6.96\times10^8$ m, $T = 5800$ K.
SOLVE Substituting the values into the Stefan-Boltzmann law:

$$P = e\sigma AT^4 = (1)(5.67\times10^{-8}\ \text{W/m}^2\text{K}^4)\left(4\pi(6.96\times10^8\ \text{m})^2\right)(5800\ \text{K})^4 = 3.91\times10^{26}\ \text{W}$$

REFLECT Currently, the world uses somewhere around 15 TW (1.5×10^{13} W) of power. In comparison, the Sun emits over 10 quadrillion times the energy we use.

1.25 l He → $\Delta T = 100°C$; $\Delta U = ?$

27. ORGANIZE AND PLAN Helium is a monatomic gas so the internal energy difference when warmed is $\Delta U = \frac{3}{2} N k_B \Delta T$. The number of atoms in 1.25 L is given by the ideal gas law.
Known: $V = 1.25\ \text{L}$; $\Delta T = 100°\text{C}$.
SOLVE From the ideal gas law $PV = Nk_BT$ we find that internal energy ~~increases by~~:

$$\to \quad Nk_B = \frac{PV}{T}$$

$$\Delta U = \frac{3}{2} N k_B \Delta T = \frac{3}{2} PV \frac{\Delta T}{T}$$

$$1\ \text{l} = 10^{-3}\ \text{m}^3$$

Insert known values:

$$\Delta U = \tfrac{3}{2}(1\ \text{atm})(1.25\ \text{L})\frac{(100°\text{C})}{(20°\text{C})} = \tfrac{3}{2}(101\ \text{kPa})(1.25\times10^{-3}\ \text{m}^3)\frac{(373\ \text{K})}{(293\ \text{K})} = 64.8\ \text{J}$$

REFLECT We will assume room temperature and atmospheric pressure unless stated otherwise (or unless such an assumption appears obviously erroneous). Note that it's important to convert the room temperature to Kelvin because it's not a temperature difference.

$\uparrow 45\,W = \frac{\Delta U}{\Delta t} \leftarrow \Delta U \uparrow$, W rate $165\,W$ | $\frac{\Delta Q}{\Delta t} = ?$

29. ORGANIZE AND PLAN Heat, internal energy, and work are related through the first law of thermodynamics. We will divide the quantities in this law by time to get powers and heat flow rates.

Known: $\Delta U/\Delta t = 45\text{ W}$; $P = -165\text{ W}$.

SOLVE The first law of thermodynamics divided by time is:

$$\frac{\Delta U}{\Delta t} = \frac{\Delta Q}{\Delta t} + P \quad \leftarrow \quad \Delta U = \Delta Q + W / \Delta t$$

which we can rewrite to get an expression for the heat flow into the system:

and $\frac{W}{\Delta t} = P$

$$\frac{\Delta Q}{\Delta t} = \frac{\Delta U}{\Delta t} - P = (45\text{ W}) - (-165\text{ W}) = 210\text{ W}$$

REFLECT The power is negative because the system is doing work.

$P[W]$

$[W] = \frac{J}{s}$

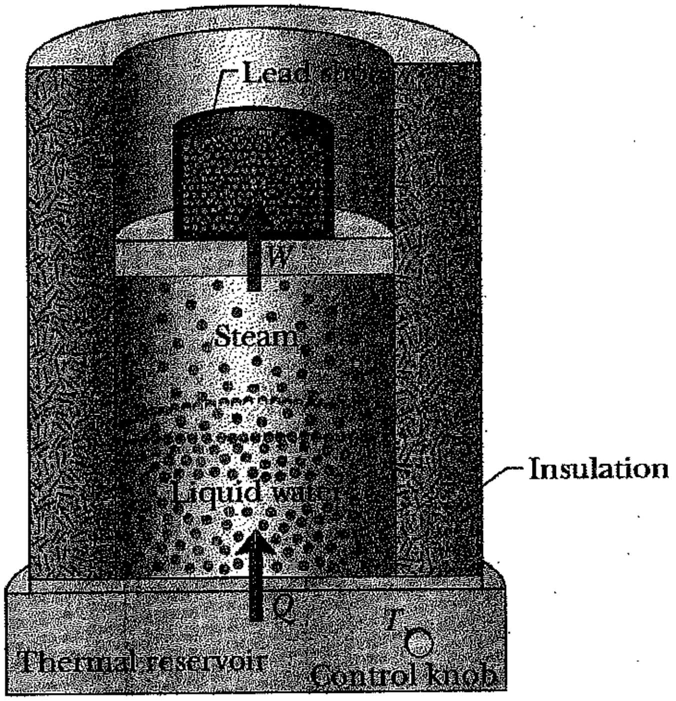
Example: A flexible container contains 2.42 x 10^{-5} m^3 of fluid at room T. Somebody pushes on the container, maintaining a constant 1-atm pressure, and reduces its volume by 25 %. How much work is done on the fluid ?

Work done on the fluid: $W = -P(V_f - V_i) = 1\text{atm}(0.75V_i - V_i) =$

$= -(1.013\text{x}10^5\ \text{Pa})V_i(-0.25) = (1.013\text{x}10^5\ \text{Pa})(2.42 \text{ x } 10^{-5}\ \text{m}^3)(0.25) = 0.61\ \text{J}$

Sample Problem 18-5

Let 1.0 kg of liquid water at 100°C be converted to steam at 100°C by boiling at standard atmospheric pressure (which is 1.0 atm or 1.01×10^5 Pa) in the arrangement of the figure, The volume of the water changes from an initial value of 1.0×10^{-3} m³ as a liquid to 1.672 m³ as steam.



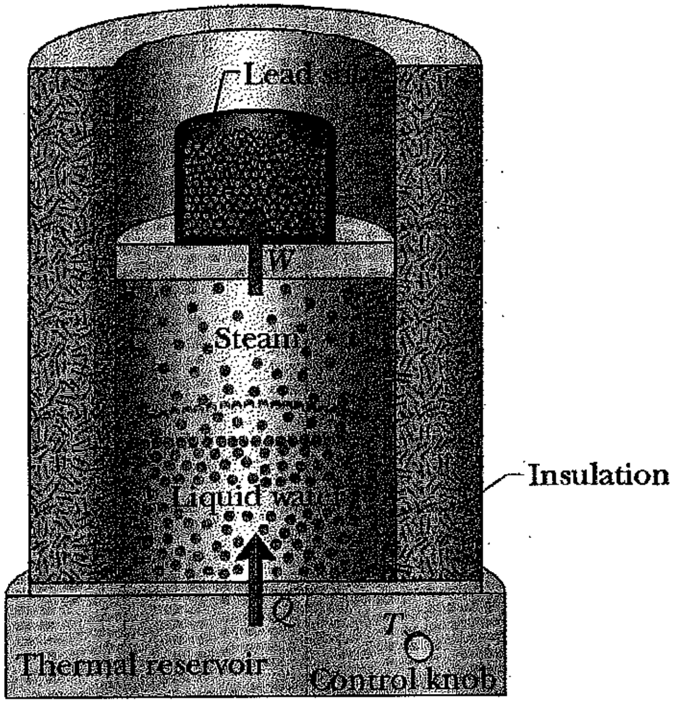
(a) How much work is done by the system?

$$W = -p_o(\underbrace{V_f - V_i}_{\Delta V})$$

$$= -(1.01 \times 10^5\ Pa)(1.67\,m^3 - 1.0 \times 10^{-3}\ m^3)$$

$$= -1.69 \times 10^5\ J = -169\ kJ$$

Sample Problem 18-5 (cont)



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

$Q = L_V m = (2256 \text{ kJ/kg})(1.0 \text{ kg})$

$= 2{,}256 \text{ kJ}$

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

$\Delta U = Q + W$, but $W = -169 \text{ kJ}$

$\Delta u = Q + W = 2256 \text{ kJ} - 169 \text{ kJ}$

$= 2{,}087 \text{ kJ}$

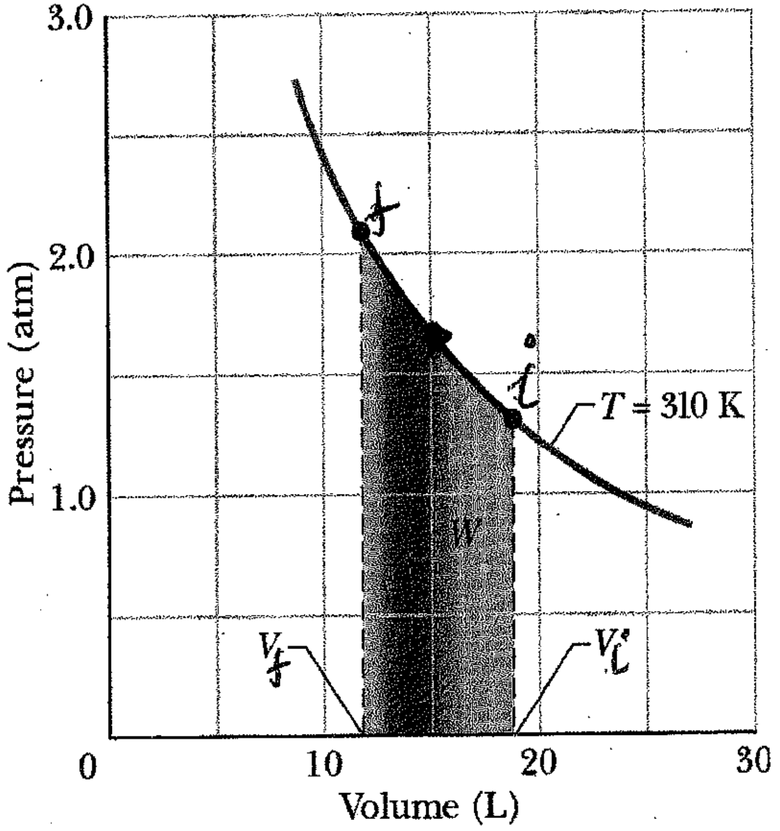
Sample Problem 19-2

One mole of oxygen (assume to be an ideal gas) shrinks at a constant temperature T of 310 K from an initial volume V_i of 12 L to a final volume V_f of 19 L. How much work is done on the gas during the expansion?

$$W = \qquad = nRT\ln\frac{V_i}{V_f}$$

$$= (1\,mol)(8.31\,J/mol\cdot K)(310\,K)\ln\left(\frac{19\,L}{12\,L}\right)$$

$$= 1183\ J$$



0.3mol He  $V_f = 5V_i$ at 300 K | $W_{gas} = ?$

38. **ORGANIZE AND PLAN** We will use the formula for work done on the gas in an isothermal process.

Known: $n = 0.30$ mol; $T = 300$ K; $V_f = 5V_i$.

SOLVE The work done on the gas is:

$$W = nRT\ln\left(\frac{V_i}{V_f}\right) = (0.30\text{ mol})(8.31\text{ J/(mol}\cdot\text{K)})(300\text{ K})\ln\left(\frac{1}{5}\right) = -1.2\text{ kJ}$$

The work done by the gas is positive 1.2 kJ.

REFLECT It does not matter in an isothermal process whether the gas is monatomic or diatomic.

0.1 mol gas at 300K , V↑ at T= const
Q=? so $V_f = 3V_i$

42. ORGANIZE AND PLAN The work done on an expanding gas is negative. The sum of the work done to and the heat added to an isothermal system is zero. To calculate a precise value we use the formula for work done on an isothermal system.

Known: $n = 0.10$ mol; $T = 300$ K; $V_f = 3V_i$.

SOLVE (a) When the gas is expanding the work done on the gas is negative, so the heat added is positive, i.e., heat flows into the gas.

$0 = \Delta U \sim E_{th} = Q + W$
W is (−w) so Q>0
for T= const

(b) The work done on an isothermal system is:

$$W = nRT \ln\left(\frac{V_i}{V_f}\right)$$

↖ on the gas

and $W + Q = 0$, so the required heat is:

$Q = -W$

$$Q = -W = -nRT \ln\left(\frac{V_i}{V_f}\right) = -(0.10\ \text{mol})(8.31\ \text{J}/(\text{mol}\cdot\text{K}))(300\ \text{K})\ln\left(\frac{1}{3}\right) = 0.27\ \text{kJ}$$

REFLECT It is straightforward to see that the answer is reasonable, because $\ln(3) \approx 1$.

$Q=0$; 2 mols H_2 do 750 J of work | $\Delta T = ?$

expansion

46. **ORGANIZE AND PLAN** The work done on the gas in an adiabatic process is proportional to the difference between two products: the final pressure times volume minus the initial pressure times volume. From the ideal gas law this means that the work done on the gas is proportional to the difference between final and initial temperature.

Known: $n = 2.0$ mol; $W = -750$ J.

SOLVE (a) The gas does positive work. This means the work done on the gas is negative, and from our argument above that means the final temperature is less than the initial temperature, so the gas temperature decreases.

The reason for this is that to do work the gas must expend its internal energy, which is proportional to temperature for a diatomic gas.

(b) The work done in an adiabatic process is:

$$W = \frac{P_f V_f - P_i V_i}{\gamma - 1}$$

$\gamma = \frac{7}{2}$ for diatomic

Rewrite this expression using the ideal gas law:

$$W = \frac{P_f V_f - P_i V_i}{\gamma - 1} = \frac{nR}{\gamma - 1}(T_f - T_i) = \frac{nR\Delta T}{\gamma - 1}$$

$P_i V_i = nRT_i$

$P_f V_f = nRT_f$

$P_f V_f - P_i V_i = nR(T_f - T_i)$

Solve for the temperature difference and insert know values:

$$\Delta T = \frac{(\gamma - 1)W}{nR} = \frac{\left(\frac{7}{5} - 1\right)(-750\text{ J})}{(2.0\text{ mol})(8.31\text{ J/(mol}\cdot\text{K)})} = -18\text{ K}$$

REFLECT The expression we derived here between the work done and the temperature difference holds true for all adiabatic processes if the gas is an ideal gas.

Problem # 2 How much energy must be transferred as heat for a reversible isothermal expansion of an ideal gas at 132° C if the entropy of the gas increases by 46 J/K ?

- From Eq. 20-2, $\Delta S = S_f - S_i = \frac{Q}{T}$

we obtain

$$Q = T\Delta S = (405\ \text{K})(46.0\ \text{J/K}) = 1.86\times10^4\ \text{J}.$$

100g ice → H_2O at 0°C ; 100g H_2O $\xrightarrow{100^\circ C}$ steam

57. **ORGANIZE AND PLAN** The entropy change is the heat added to the ice or the water divided by the temperature. The heat added can be calculated using the heats of transformation from Table 13.3.
Known: $m = 100\text{ g}$; $T_f = 0°\text{C}$; $L_f = 3.33\times10^5\text{ J/kg}$; $T_v = 100°\text{C}$; $L_v = 2.26\times10^6\text{ J/kg}$.
SOLVE When the ice melts the entropy change is:

$$\Delta S_f = \frac{Q_f}{T_f} = \frac{mL_f}{T_f} = \frac{(100\text{ g})(3.33\times10^5\text{ J/kg})}{(0°\text{C})} = 122\text{ J/K}$$

When the water boils the entropy change is:

$$\Delta S_v = \frac{Q_v}{T_v} = \frac{mL_v}{T_v} = \frac{(100\text{ g})(2.26\times10^6\text{ J/kg})}{(100°\text{C})} = 606\text{ J/K}$$

$\Delta S = ?$ when 75g steam $\xrightarrow{100°C}$ H_2O

Revw

56. ORGANIZE AND PLAN The entropy change is the heat removed from the steam divided by the temperature. The heat removed can be calculated using the heat of vaporization from Table 13.3.

Known: $m = 75\text{ g}$; $T = 100°\text{C}$; $L_v = 2.26\times10^6\text{ J/kg}$.

SOLVE The heat removed from the steam to condense it is:

$$Q = -mL_v = -\underset{\overset{''}{0.075\text{kg.}}}{(75\text{ g})}(2.26\times10^6\text{ J/kg}) = -1.7\times10^5\text{ J}$$

* Table 13.3

The entropy change is:

$$\Delta S = \frac{Q}{T} = \frac{(-1.7\times10^5\text{ J})}{(100°\text{C})} = \frac{(-1.7\times10^5\text{ J})}{(373\text{ K})} = -4.5\times10^2\text{ J/K}$$

However $S\uparrow$ in the Universe since, for example, refrigerator does work to remove the heat and dump it in the Universe

Problem A Carnot engine has an efficiency of 22 %. It operates between constant-temperature reservoirs differing in temperature by 75 K. What is the temperature of the (a) lower-temperature and (b) higher-temperature reservoir ?

25. We solve (b) first $\varepsilon = 1 - \frac{T_L}{T_H}$ *Carnot cycle efficiency* $= \frac{\overset{\text{" }\Delta T}{\overbrace{T_H - T_L}}}{T_H} = \varepsilon$

(b) For a Carnot engine, the efficiency is related to the reservoir temperatures by Eq. Therefore,

$$T_H = \frac{\overset{\text{" }\Delta T}{\overbrace{T_H - T_L}}}{\varepsilon} = \frac{75\text{ K}}{0.22} = 341\text{ K}$$

which is equivalent to 68°C. $\leftarrow$ $T\,°C = T\,K - 273$

(a) The temperature of the cold reservoir is $T_L = 341 - 75 = 266\text{ K}$

$\Delta T = 75\text{K} = 341\text{K} - T_L$

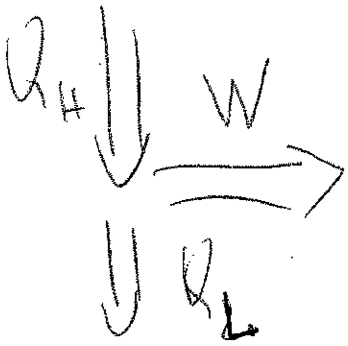
$W = 650\text{ J} \rightarrow Q_L = 1270\text{ J};\quad \varepsilon = ?$

63. . **ORGANIZE AND PLAN** The efficiency is the ratio between work done and heat used.
Known: $W = 650\text{ J}$; $Q_H = 1270\text{ J}$.
SOLVE The heat engine's efficiency is:

$$Q_H = W + Q_L = 1920\text{ J}$$

$$e = \frac{W}{Q_H} = \frac{(650\text{ J})}{1920\text{ J}} = 0.338 \sim 34\%$$

P



Jet engine ; $\varepsilon = ?$

96. ORGANIZE AND PLAN The maximum efficiency is that of a Carnot engine, one minus the temperature ratio between the cold and hot reservoirs.

Known: $T_H = 1050°\text{C}$; $T_C = 590°\text{C}$.

SOLVE The maximum efficiency is:

$$e_{\text{Carnot}} = 1 - \frac{T_C}{T_H} = 1 - \frac{(590°\text{C})}{(1050°\text{C})} = 0.348 \neq 1 - \frac{(590 + 273)}{(1050 + 273)}$$

REFLECT If you could invent a way of operating a jet engine such that the exhaust temperature is that of the surrounding air (typically −40°C at the cruising height of commercial airliners), you would raise the maximum efficiency to better than 0.8!

Sample Problem 20-4

Imagine a Carnot engine that operates between the temperatures T_H = 850 K and T_L = 300 K. The engine performs 1200 J of work each cycle, which takes 0.25 s.

(a) What is the efficiency of this engine?

$$\varepsilon = 1 - \frac{T_L}{T_H} = 1 - \frac{300\,K}{850\,K} = 0.647$$

(b) What is the average power of this engine?

$$P = \frac{W}{t} = \frac{1200\,J}{0.25\,s} = 4800\,W = 4.8\,kW$$

$$K = \frac{what\ we\ want}{what\ we\ pay\ for} = \frac{|Q_L|}{|W|}$$

$$|W| = |Q_H| - |Q_L|$$

$$K = \frac{|Q_L|}{|Q_H| - |Q_L|}$$

Problem

A Carnot refrigerator does 200 J of work to remove 600 J from its cold compartment. *(a)* What is the refrigerator's coefficient of performance? *(b)* How much energy per cycle is exhausted to the kitchen as heat?

34. (a) We use Eq. 21-12,

$$\frac{Q_L}{Q_H - Q_L} = K = \frac{|Q_L|}{|W|} = \frac{600}{200} = 3 .$$

(b) Energy conservation for a refrigeration cycle requires $|Q_L| + |W| = |Q_H|$, so that the result is 800 J.

$Q_H - Q_L = W$

$Q_H = W + Q_L = 200J + 600J = 800J$

$T_H = ?$ if $T_L = 20°C$

68. **ORGANIZE AND PLAN** In a Carnot cycle, the efficiency is one minus the temperature ratio between the cold and hot reservoirs.

Known: $T_C = 20°\text{C}$; $e_{\text{Carnot}} = 0.5$.

SOLVE The Carnot efficiency is:

$$e_{\text{Carnot}} = 1 - \frac{T_C}{T_H}$$

$$\rightarrow \frac{T_c}{T_H} = 1 - \varepsilon \rightarrow T_H = \frac{T_c}{1-\varepsilon}$$

which we can rewrite for to calculate the maximum temperature:

$$T_H = \frac{1}{1 - e_{\text{Carnot}}} T_C = \frac{1}{1 - 0.5} 20°\text{C} =$$

$\rightarrow 293\text{K} = (20 + 273)$

$\approx 600\ \text{K}$

REFLECT When the Carnot efficiency is 0.5, the hot reservoir is twice the temperature of the cold reservoir. (in K)

$\frac{Q_L}{W} = COP = 3.8$; $P = 600\,W = \frac{W}{t}$ | $\frac{Q_L}{t} = ?$

71. ORGANIZE AND PLAN The coefficient of performance of a refrigerator is the heat removed divided by the required work.

Known: COP = 3.8; $P = 600$ W.

SOLVE The refrigerator can remove heat at a rate:

$\frac{Q_L}{t} = COP \times W / t \longrightarrow$ $$\frac{Q_L}{t} = \text{COP}\frac{W}{t} = \text{COP} \times P = (3.8)(600\text{ W}) = 2.3\text{ kJ/s}$$

REFLECT The higher the COP the larger the amount of heat removed.

per 1sec $Q_H = 1700\text{ MJ}$; $Q_C = 1100\text{ MJ}$ | $P = ?$ $e = ?$

65. ORGANIZE AND PLAN The efficiency is the ratio between work done and the thermal energy removed from the uranium fuel. The work done is the difference between the thermal energy and the waste heat.

Known: $Q_H = 1700\text{ MJ}$; $Q_C = 1100\text{ MJ}$; $t = 1\text{ s}$.

$$\boxed{Q_H = W + Q_C}$$

SOLVE (a) The work done is:

$$W = Q_H - Q_C = (1700\text{ MJ}) - (1100\text{ MJ}) = 600\text{ MJ}$$

The efficiency is the power plant is:

$$e = \frac{W}{Q_H} = \frac{(600\text{ MJ})}{(1700\text{ MJ})} = 0.353 \ \%$$

(b) The rate of electrical energy produced, i.e., the electrical power is:

$$P = \frac{W}{t} = \frac{(600\text{ MJ})}{(1\text{ s})} = 600\text{ MW}$$

REFLECT If the nuclear power plant is close to a city, the waste heat could instead be used to heat thousands of homes.

Sample Problem 20-5

An inventor claims to have constructed an engine that has an efficiency of 75% when operated between the boiling and the freezing points of water. Is this possible?

$$\varepsilon \leq \varepsilon_{Carnot} = 1 - \frac{T_L}{T_H} = 1 - \frac{(0+273)\,K}{(100+273)\,K} = 0.268 \approx 27\%$$

NO !

Air conditioner consumes $\underset{\text{power}}{1200\text{ W}}$; $COP = 3.2$; $Q_c = ?$ for $t = 24$ hours

Book

72. **ORGANIZE AND PLAN** The coefficient of performance of a refrigerator (to which we count air-conditioning systems) is the heat removed divided by the required work.

Known: $COP = 3.2$; $P = 1200\text{ W}$; $t = 24\text{ h}$.

SOLVE The electrical energy (required work) consumed in 24 h is:

$$W = Pt = (1200\text{ W})(24\text{ h}) = 1.0\times10^8\text{ J} \quad \leftarrow ;\ P = \frac{W}{t}$$

The amount of heat removed from the house is:

$$Q_c = \text{COP}\times W = (3.2)(1.0\times10^8\text{ J}) = 3.3\times10^8\text{ J} \quad ;\ COP = \frac{Q_c}{W}$$

REFLECT The higher the COP, the larger the amount of heat removed.

$W_{old} \cdot \xrightarrow{e^{\ominus}} \$180$; If pump with COP = 3.1 | Bill = ?

94. **ORGANIZE AND PLAN** The coefficient of performance (COP) of a heat pump equals the heat delivered divided by the electrical energy used. This means that with a heat pump, the required electrical energy is reduced by a factor that equals the COP.

Known: $c = \$180$; COP = 3.1.

SOLVE The monthly heating bill after installing a heat pump is:

$$\frac{c}{\text{COP}} = \frac{(\$180)}{(3.1)} = \$58$$

REFLECT We have assumed that the cost per kWh is constant. This is not always true: some electrical utility companies charge higher prices per kWh once a certain level has been surpassed. In such a situation, the savings would be even greater.

$$W_{old} = Q_H = \$180$$

$$\text{COP} = \frac{Q_H}{W_{new}} \rightarrow W_{new} = \frac{Q_H}{\text{COP}} = \frac{W_{old}}{\text{COP}} = \frac{\$180}{\text{COP}}$$