Lecture 44

(Ch13:2)

- Fahrenheit scale: used in United States

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

0° C => 32° F

10° C => 36.5 ° F

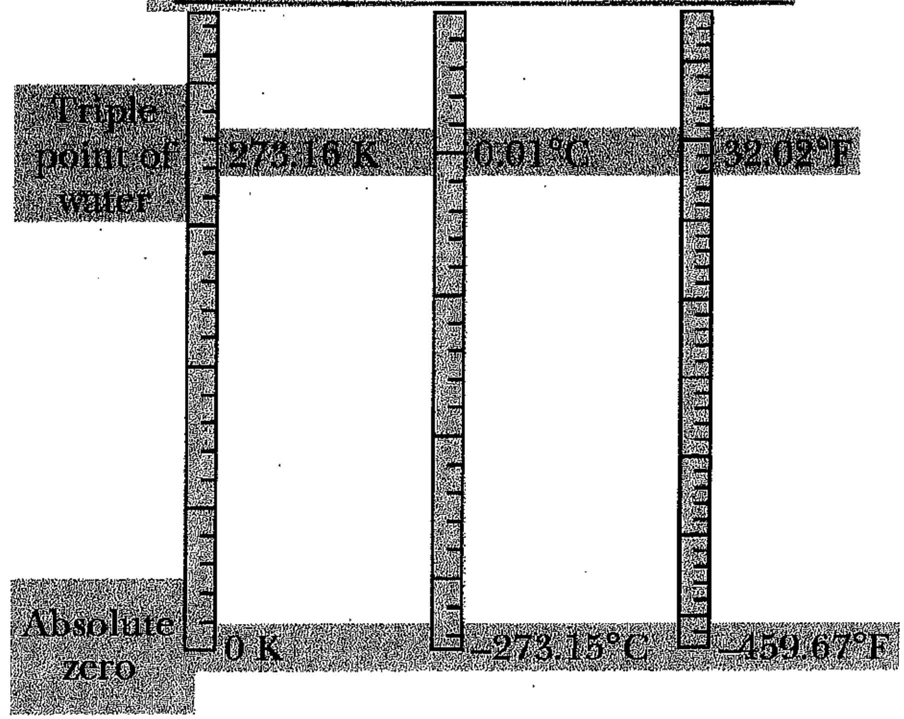
– Some corresponding temperatures (Table 18-1 and Fig 18-7)

$T_C = \frac{5}{9}(T_F - 32°)$

$T_F = T_C - 273.15$

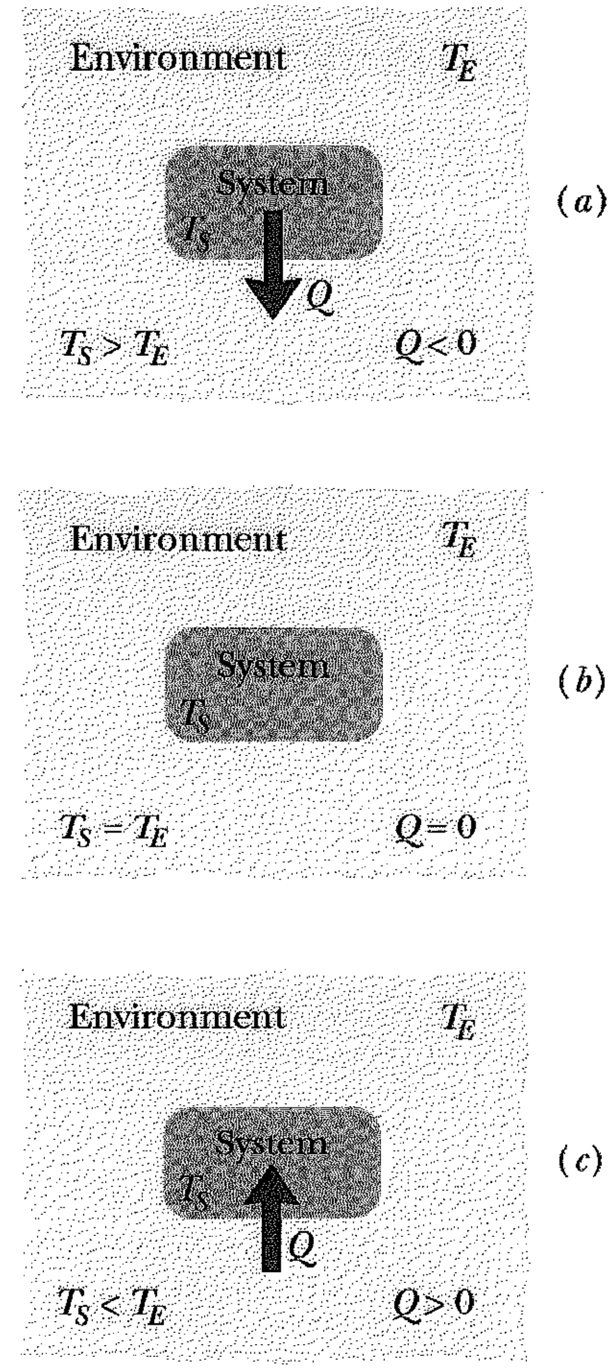
TABLE 19-1 Some Corresponding Temperatures

Temperature	°C	°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



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
  - *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

Absorption of Heat

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

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

$$C = \frac{Q}{T_f - T_i}$$

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

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

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

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

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

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

Checkpoint 18-3

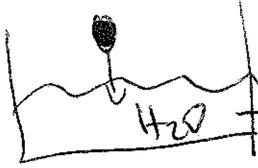
A certain amount of heat Q will warm 1 g of material A by 3° C and 1 g of material B by 4° C. Which material has the greatest specific heat?

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

25g at 34.5 °C



125g at 18°C

$T_{eq} = 22.9\ °C$

$C_M = ?$

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:

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

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

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.

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

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

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

↓

Table 13.1

You have 300 g of coffee at 55 °C. How much 10 °C water do you need to add in order to reduce the coffee's temperature to a more bearable 49 °C ? Note the specific heat, c, of coffee and water are the same ($c = 4186$ J. kg.$^{-1}$.°C^{-1})

Heat is exchanged between the hot coffee and the cold water, but the whole system does not lose or receive heat. Therefore, $Q_{\text{hot}} + Q_{\text{cold}} = 0$. The temperature of the coffee drops, while that of the added water rises:

$$\Delta T_{\text{hot}} = 49°\text{C} - 55°\text{C} = -6°\text{C}$$
$$\Delta T_{\text{cold}} = 49°\text{C} - 10°\text{C} = 39°\text{C}$$

Known: $m_{\text{hot}} = 300$ g

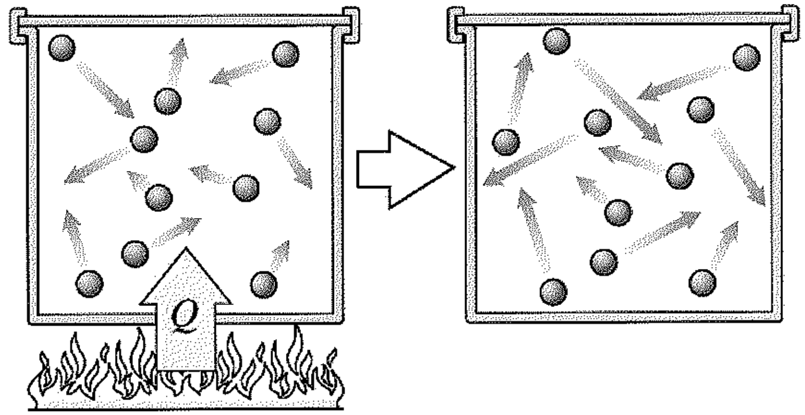
SOLVE The heat exchange is written:

$$Q_{\text{hot}} + Q_{\text{cold}} = m_{\text{hot}} c \Delta T_{\text{hot}} + m_{\text{cold}} c \Delta T_{\text{cold}} = 0$$

The specific heat of coffee is the same as water, so the c's will cancel out of the equation. Solving for the cold water mass:

$$m_{\text{cold}} = -m_{\text{hot}} \frac{\Delta T_{\text{hot}}}{\Delta T_{\text{cold}}} = -(300\text{ g}) \frac{(-6°\text{C})}{(39°\text{C})} = 46\text{ g}$$

Figure 13.7

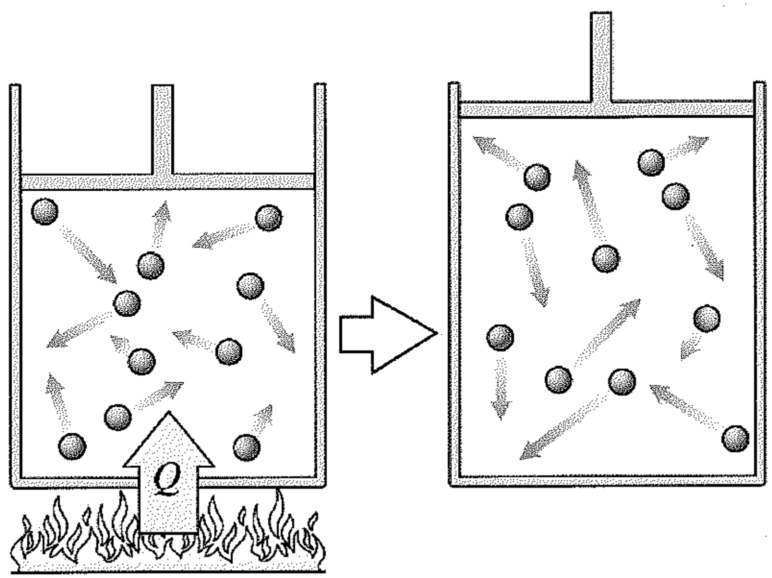


(a) When heat is added at constant volume, all the energy goes into thermal motion.

$Q = m c_v \Delta T$

2 Why

$c_p > c_v$



(b) When heat is added at constant pressure, some of the energy goes into thermal motion and some into expanding the container.

$Q = m c_p \Delta T$

Specific Heat of Gases

Gases have heat capacity and specific heat, but they're expressed differently than for solids and liquids. That's because a gas can change pressure and volume when heated, while solids and liquids undergo much less of those changes.

When a gas is heated, its temperature change depends on how much the pressure and volume change. For this reason, there are two measures of a gas's specific heat: at constant volume and at constant pressure. Another difference is that gas specific heats are usually given on a molar basis, rather than per unit mass. Thus, the equations relating heat flow Q to temperature change contain the number of moles n of the gas. For a constant-volume process,

$$Q = nc_V \Delta T \qquad \text{(Specific heat of a gas at constant volume; SI unit: J/(mol}\cdot{}^\circ\text{C))} \tag{13.3}$$

where c_V is the **molar specific heat at constant volume**. Similarly, for a constant-pressure process,

$$Q = nc_P \Delta T \qquad \text{(Specific heat of a gas at constant pressure; SI unit: J/(mol}\cdot{}^\circ\text{C))} \tag{13.4}$$

where c_P is the **molar specific heat at constant pressure**. Table 13.2 shows some values of c_V and c_P for selected gases. The units for both are J/(mol · °C). Therefore, with n in mol and ΔT in °C, the heat Q is in J.

Ideal Gas Laws

State variables

$P \cdot V = \text{const}$; $n, T = \text{const}$

$V \sim T$; $n, P = \text{const}$

$P \sim n$; $T, V = \text{const}$

$P \sim T$; $n, V = \text{const}$

$P\,[Pa]$

$1\,atm = 1.03 \times 10^5\,Pa$

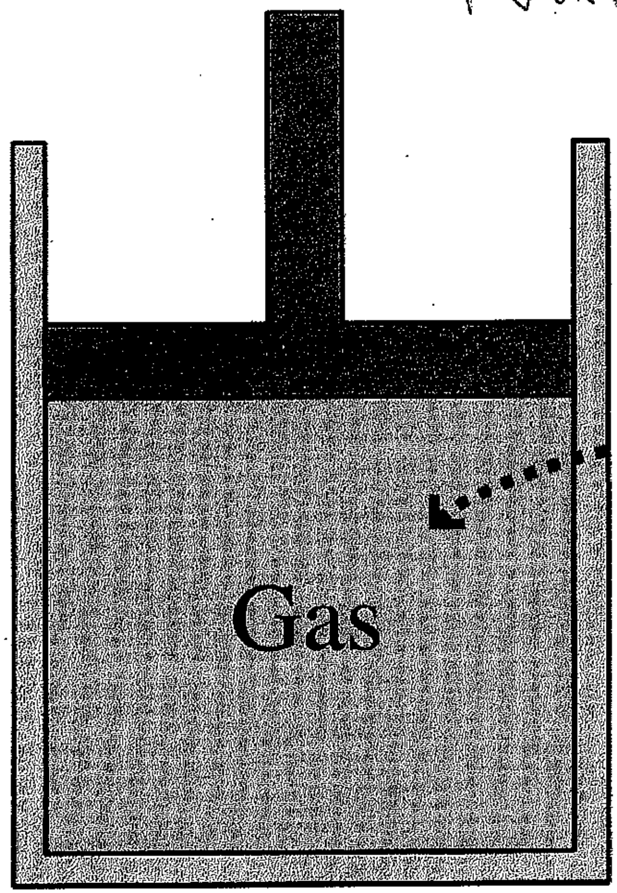
$V\,[m^3]$

$T \to [K]$

$$\boxed{P \cdot V = n \cdot R T}$$

$R = 8.31\,[mol^{-1} \cdot J \cdot K^{-1}]$ molar gas constant

Figure 12.8



The amount of gas remains constant, but its volume, temperature, and pressure may change.

$N = n \cdot N_A \to n = \frac{N}{N_A}$

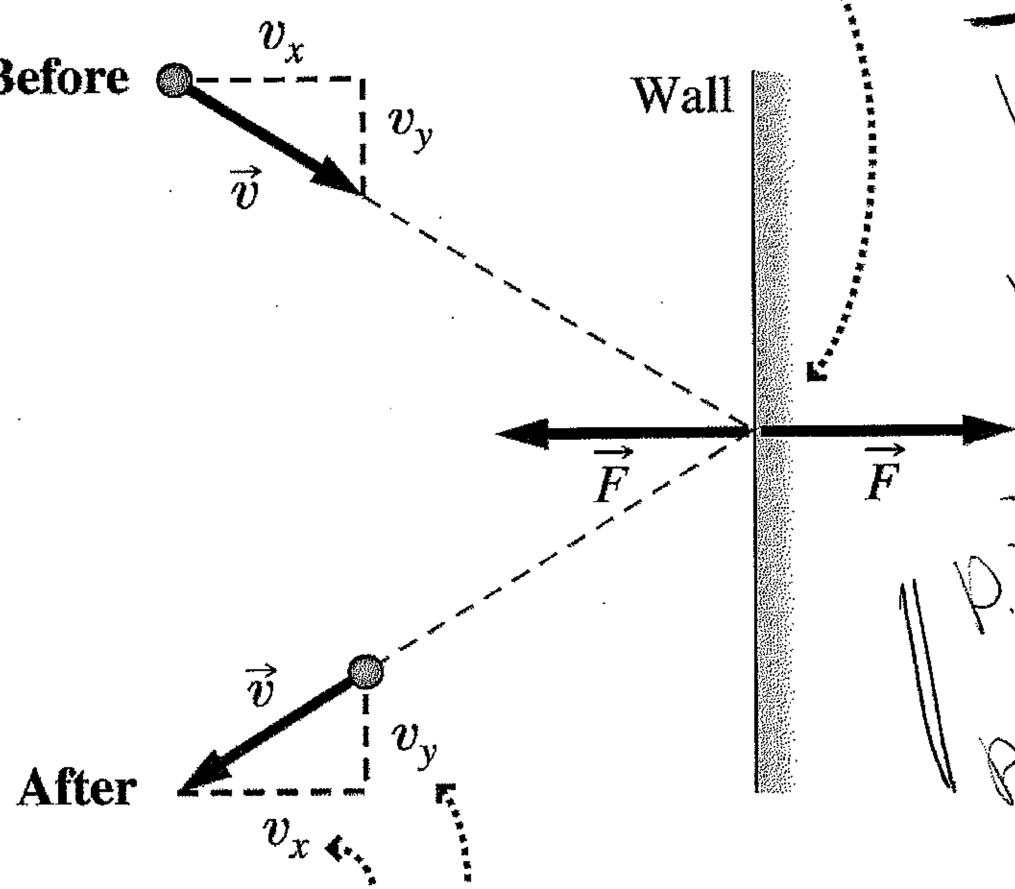
$P \cdot V = N \frac{R}{N_A} \cdot T = N k_B T$, where

$k_B = 1.38 \times 10^{-23}\ \frac{J}{K}$

molecular gas constant

Figure 12.10

When a molecule collides elastically with a container wall, the molecule and wall exert forces on each other.



The force exerted on the molecule reverses the sign of the x-component of the velocity but does not change the y-component.

$E_k = const$

$\frac{\Delta p}{\Delta t} = F \neq 0$

$p = m.v$

P – pressure

$P = N \frac{m \overline{v^2}}{3V}$

$\overline{v^2} = \frac{1}{N} \sum v_i^2$

$v_{rms} = \sqrt{\overline{v^2}}$

$P.V = N \frac{m \overline{v^2}}{3}$

$P.V = N.k_B T$

and

$\overline{K} = \frac{1}{2} m \overline{v^2}$

$\boxed{\overline{K} = \frac{3}{2} k_B T}$

one atom/single atomic molecule ← Average kinetic

Total $E_{th} = N.\overline{K} = \frac{3}{2} N k_B T$

$\overline{K}$ [J] ; T [K]

$Q = mc_m \Delta T$ | solids

Table 13-2

$Q = nC_V \Delta T$ | gas

$Q = nC_p \Delta T$

TABLE 13.2 Molar Specific Heats of Selected Gases

Gas	C_V in J/(mol · °C)	C_P in J/(mol · °C)
Monatomic gases		
He	12.5	20.8
Ne	12.5	20.8
Ar	12.5	20.8
Diatomic gases		
H_2	20.4	28.7
N_2	20.8	29.1
O_2	20.9	29.2
Air (a predominantly diatomic mixture)	20.8	29.1

$E_{tr} = \frac{3}{2} N k_B T$

$Q = \frac{3}{2} N k_B \Delta T = \frac{3}{2} nR\Delta T$

where $N = nN_A$

$N_A k_B = R$

$R = 8.31\ \mathrm{J/(mol \cdot °C)}$

But

$Q = n C_V \Delta T = \frac{3}{2} nR\Delta T$

so

$C_V = \frac{3}{2} R = \frac{3}{2}\left(8.31 \frac{\mathrm{J}}{\mathrm{mol \cdot °C}}\right)$

$= 12.5 \frac{\mathrm{J}}{\mathrm{mol \cdot °C}}$

usually

$C_p > C_V$

$C_V = 37.5\ \text{J/°C}$; $n = ?$ when $V = \text{const.}$

$C_p = ?$ $Q = C_p \Delta T = n c_p \Delta T$

58. ORGANIZE AND PLAN The molar specific heat for a $\boxed{\text{monatomic gas}}$ is $c_V = 12.5\ \text{J/mol·°C}$ for constant volume, and $c_P = 20.8\ \text{J/mol·°C}$ for constant pressure. To find the number of moles, we divide the given heat capacity by c_V and use that answer to find the heat capacity at constant pressure.

From Table 13.2

Known: $C_V = 37.5\ \text{J/°C}$.

SOLVE (a) The number of moles is just:

$$n = \frac{C_V}{c_V} = \frac{37.5\ \text{J/°C}}{12.5\ \text{J/mol·°C}} = 3.00\ \text{mol}$$

(b) The heat capacity at constant pressure is:

$$C_P = n c_P = (3.00\ \text{mol})(20.8\ \text{J/mol·°C}) = 62.4\ \text{J/°C}$$

REFLECT Recall that the heat capacity at constant pressure is greater than the heat capacity at constant volume. This is because — in the constant pressure case — a given amount of delivered energy goes into both increasing the thermal energy and increasing the volume of the gas (i.e., the energy does work). In the constant volume case, the delivered energy is only used to increase the thermal energy.

Figure 13.8

Equipartition | Each degree of freedom $\rightarrow \frac{1}{2}R$ molar c

Rotation occurs around the two axes perpendicular to the molecular bond.

$K = \frac{1}{2}mv_x^2 + \frac{1}{2}mv_y^2 + \frac{1}{2}mv_z^2$

$C_v = 3 \times \frac{1}{2}R = \frac{3}{2}R \sim 12 \frac{J}{mol \cdot {}^\circ C}$

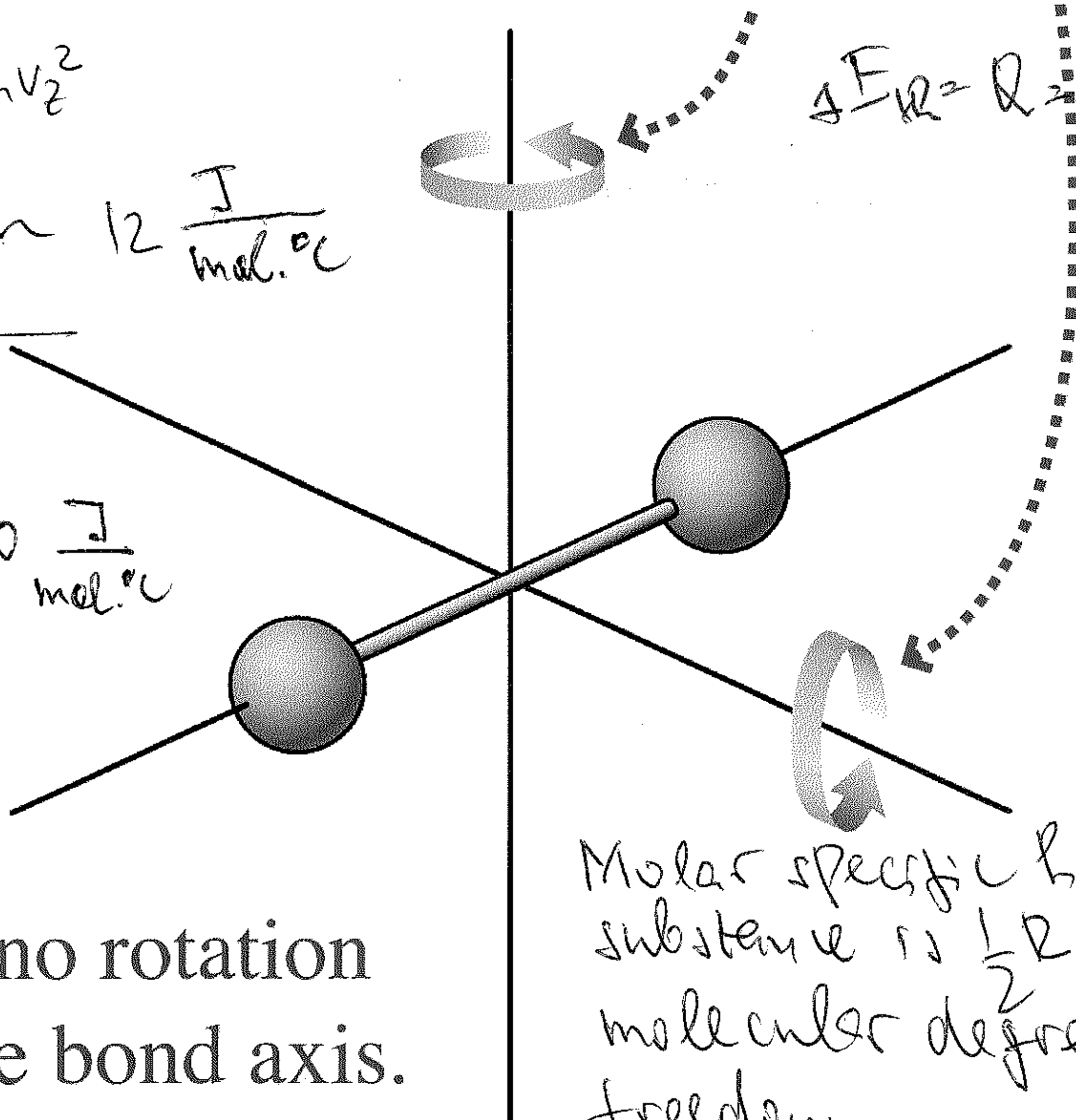
for Binary

$C_v = \frac{5}{2}R \sim 20 \frac{J}{mol \cdot {}^\circ C}$

$\Delta E_{KE} = Q = \frac{3}{2}nR\Delta T = C_v \Delta T \cdot n$

for 1 mole

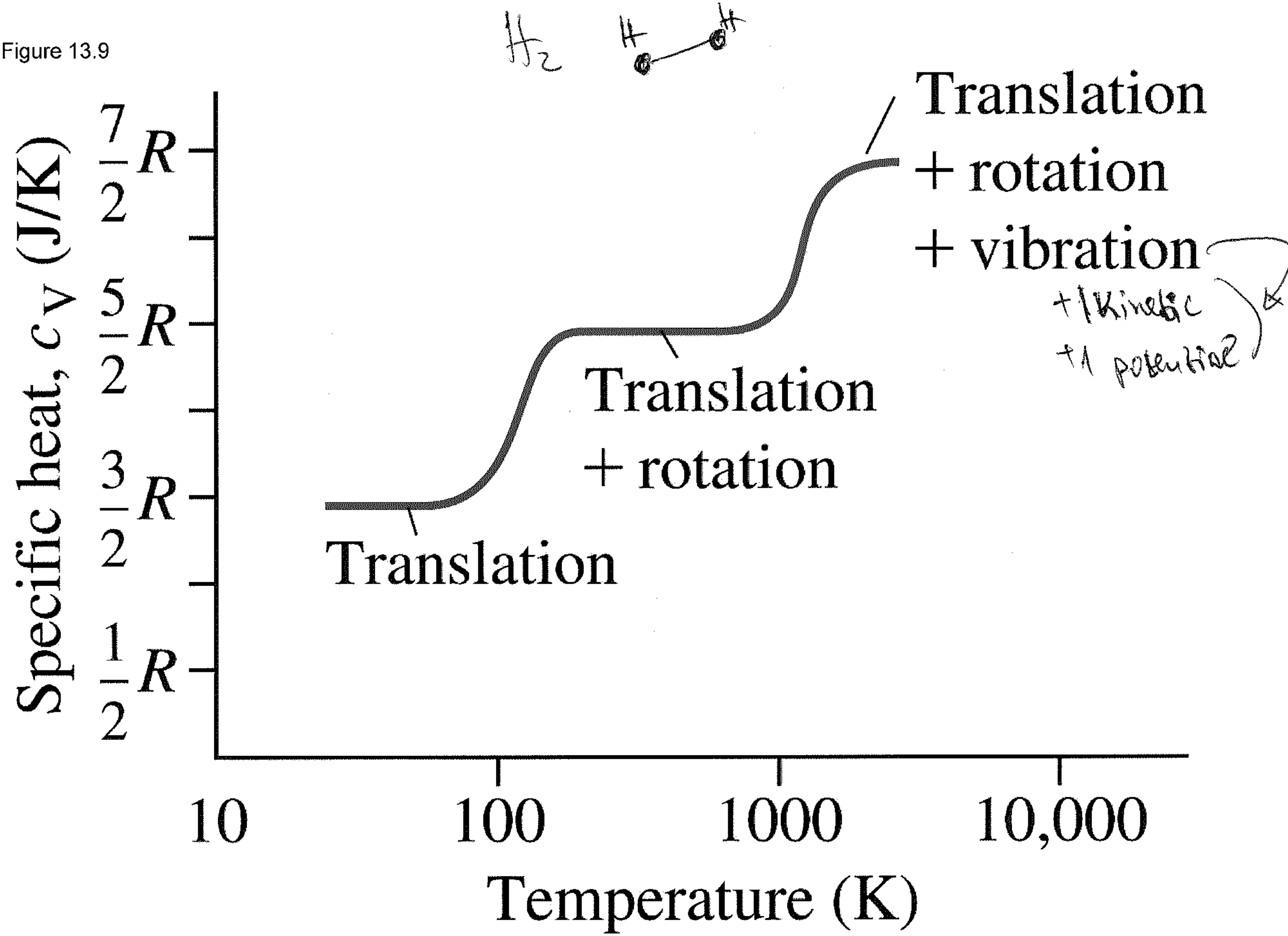
$C_v = \frac{3}{2}R$

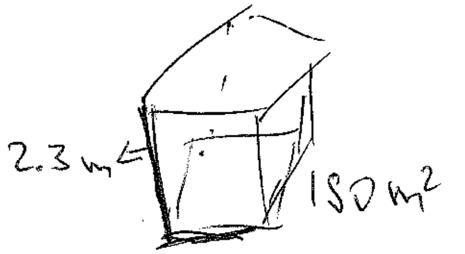


There's no rotation about the bond axis.

Molar specific heat of a substance is $\frac{1}{2}R$ for each molecular degree of freedom

Figure 13.9


H_2
H
H
Translation
+ rotation
+ vibration
+1 kinetic
+1 potential
Translation
+ rotation
Translation
Specific heat, c_V (J/K)
$\frac{7}{2}R$
$\frac{5}{2}R$
$\frac{3}{2}R$
$\frac{1}{2}R$
10
100
1000
10,000
Temperature (K)



air = 79% N_2 + 21% O_2
Q = ? T↑ 1°C
$J = ? if 1 kWh = 16¢

59. **ORGANIZE AND PLAN** We'll assume a constant volume of air in the house. We can find the total number of moles using the ideal gas law ($PV = nRT$). We could then find the number of moles of nitrogen and the number of moles of oxygen using the percentages given. But they are both diatomic gases, so they have essentially the same molar specific heat. In Table 13.3, the molar specific heat of air is given ($c_V = 20.8$ J/mol·°C), so we'll use that to find the energy needed to raise the air temperature by one degree Celsius (Equation 13.3).

Known: $A = 190\text{ m}^2$, $h = 2.3\text{ m}$, $\Delta T = 1°\text{C}$, $r = 16¢/\text{kWh}$.

SOLVE (a) The volume is the area times the height ($V = Ah$), and we assume that the pressure is 1atm (or 101,325 Pa in SI units) and the room temperature is 20°C (or 293 K). So the total number of moles is:

$$n = \frac{PV}{RT} = \frac{(101{,}325\text{ Pa})(190\text{ m}^2)(2.3\text{ m})}{(8.315\text{ J/mol}\cdot\text{K})(293\text{ K})} = 18{,}200\text{ mol}$$

PV = nRT
V = const

You could also come to a similar result by using the fact that the molar volume of an ideal gas is 22.4 L/mol at standard temperature and pressure, but that would be assuming the house is a very chilly 0°C to begin with. Plugging the total moles into Equation 13.3:

$$Q = nc_V\Delta T = (18{,}200\text{ mol})(20.8\text{ J/mol}°\text{C})(1°\text{C}) = 3.79\times10^5\text{ J}$$

(b) The cost of this much energy is:

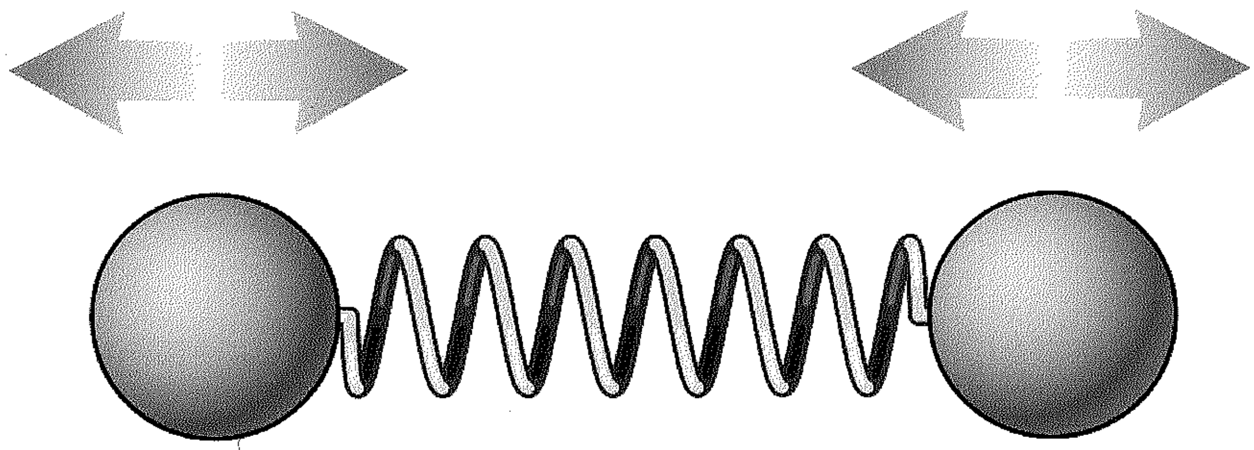
$$C = rQ = (16¢/\text{kWh})(3.79\times10^5\text{J})\left[\frac{1\,\cancel{\text{W}}\cdot\cancel{\text{s}}}{1\text{ J}}\right]\left[\frac{1\text{ kWh}}{(1000\,\cancel{\text{W}})(60\cdot60\,\cancel{\text{s}})}\right] = 1.7¢$$

$[W]\ P = \frac{J}{s} \rightarrow J = W.s$

Figure 13.10

Vibration in molecule

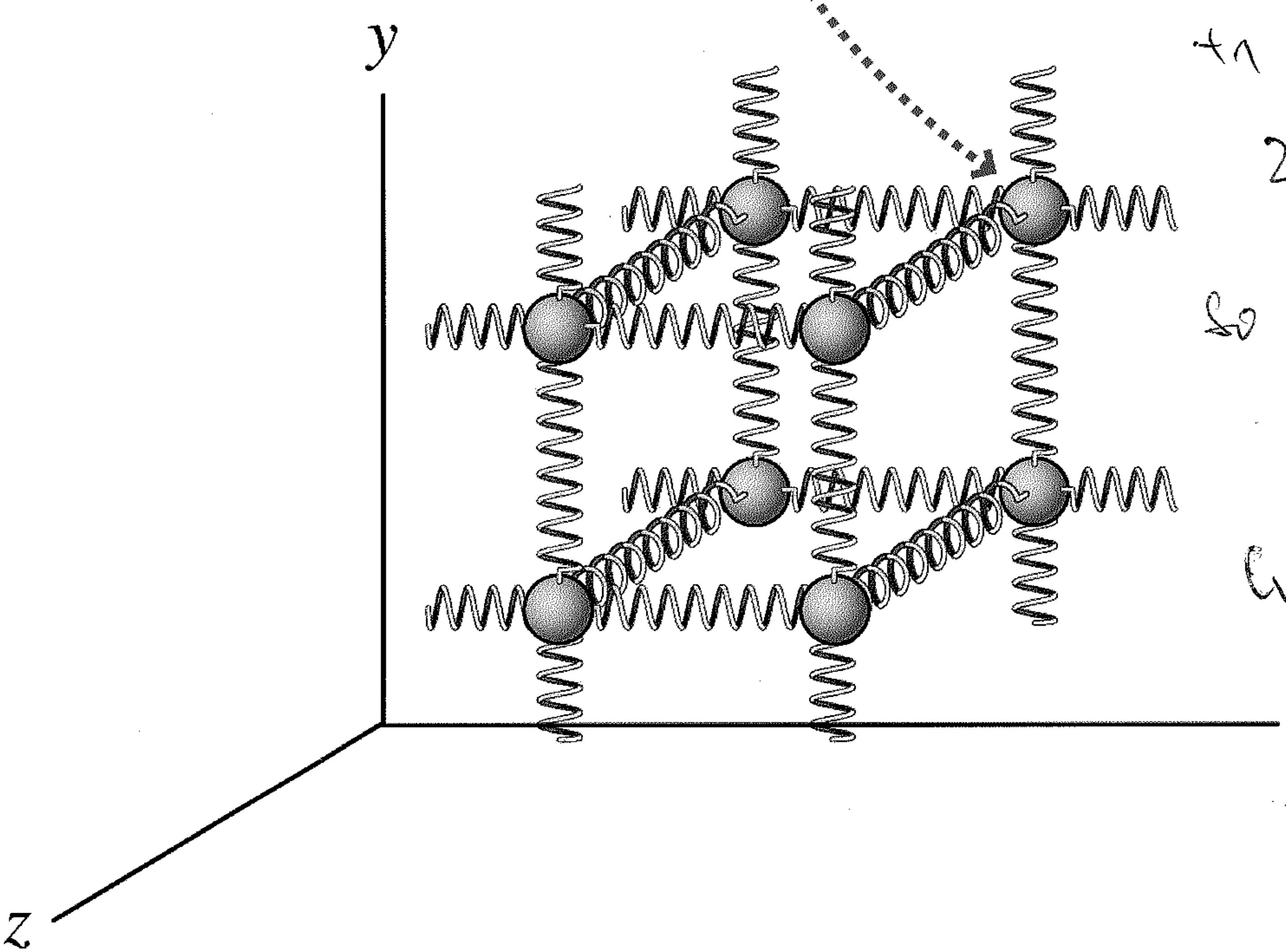
The atoms in a diatomic molecule can vibrate back and forth like masses joined by a spring.



Vibration in solids

Figure 13.11

Each atom can move in three directions (x, y, z) and has both kinetic and potential energy due to each direction's motion.



+1 kinetic
+1 potential
$\}$ 2

$2 \times 3_{\text{directions}} = 6$

So for solids

$C_V = 6 \times \frac{1}{2} R = 3R$
(cubic)

$C_v(Cu) = 3R$ (per mole) $63.6 \frac{g}{mol}$ mass of one mole of Cu

$C_m = ?$

56. **ORGANIZE AND PLAN** The equipartition theorem states that each degree of freedom in a substance adds $R/2$ to the molar specific heat, where $R = 8.315$ J/mol·C is the molar gas constant. Ideally, the vibrations of copper atoms inside a solid piece of copper account for 6 degrees of freedom, so the specific of heat should be $3R$. We are simply asked to convert this into a specific heat using the molar mass of copper: $M = 63.6$ g/mol.

SOLVE Using the values given, the specific heat of copper is:

$$c = \frac{3R}{m} = \frac{3(8.315\ \text{J/mol}\cdot\text{C})}{(0.0636\ \text{kg/mol})} = 392\ \text{J/kg}\cdot\text{C}$$

$C_m m \cancel{\Delta T} = Q = n C_v \cancel{\Delta T}$, $n = 1$

This is slightly higher than the value in Table 13.1: $c = 385$ J/kg·C.

REFLECT The fact that the derived specific heat is slightly higher than the actual value implies that copper atoms have slightly less than 6 degrees of freedom. They may be partly constrained in their vibrations.

$$C_m = \frac{C_v}{m}$$

1250W ← → 1ℓ H_2O at 20°C
t=? to boil

51. **ORGANIZE AND PLAN** Water boils at 100°C, so we'll use Equation 13.2 ($Q = mc\Delta T$) to calculate how much heat is needed to bring the water up to that temperature. The time it takes to reach boiling is just the heat delivered divided by the kettle's power: $t = Q/P$, where recall that W = J/s.

Known: $P = 1250$ W, $V_{H2O} = 1.0$ L , $T_i = 20°$C, $T_f = 100°$C.

SOLVE The temperature change in the water is: $\Delta T = T_f - T_i = 80°$C, and given that the density of water is 1.00 g/mL, the mass of 1.0 L is 1.0 kg. The heat needed to bring this much water to boil is:

$$Q = mc\Delta T = (1.0 \text{ kg})(4186 \text{ J/kg}°\text{C})(80°\text{C}) = 3.3 \times 10^5 \text{ J}$$

The time it takes the kettle to deliver this much heat is:

$$t = \frac{Q}{P} = \frac{(3.3 \times 10^5 \text{ J})}{(1250 \text{ W})} = 268 \text{ s}$$

$P = \frac{W}{t} = \frac{Q}{t} \rightarrow t = \frac{Q}{P}$

REFLECT This is about 4 and half minutes, which seems reasonable for bringing that much water to boil. Notice that the time could be longer because some of the heat may escape.