

$$1.25 \text{ L He} \rightarrow \Delta T = 100^\circ\text{C}; \Delta U = ?$$

**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:*  $V = 1.25 \text{ L}; \Delta T = 100^\circ\text{C}$ .

**SOLVE** From the ideal gas law  $PV = N k_B T$  we find that internal energy increases by:

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

$$\rightarrow N k_B = \frac{PV}{T}$$

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

Insert known values:

$$\Delta U = \frac{3}{2} (1 \text{ atm})(1.25 \text{ L}) \frac{(100^\circ\text{C})}{(20^\circ\text{C})} = \frac{3}{2} (101 \text{ kPa})(1.25 \times 10^{-3} \text{ m}^3) \frac{(373 \text{ K})}{(293 \text{ K})} = 64.8 \text{ J}$$

**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.

$$\uparrow 45 \text{ W} = \frac{\Delta U}{\Delta t} \leftarrow \Delta U \uparrow \quad \text{Work rate } 165 \text{ W} \quad \left| \quad \frac{\Delta Q}{\Delta t} = ?$$

**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:

$$\frac{\Delta U}{\Delta t} = \frac{\Delta Q}{\Delta t} + P$$

$$\leftarrow \Delta U = \Delta Q + W / \Delta t$$

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

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

and

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

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

$$P [W]$$

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

# Sample Problem 18-5

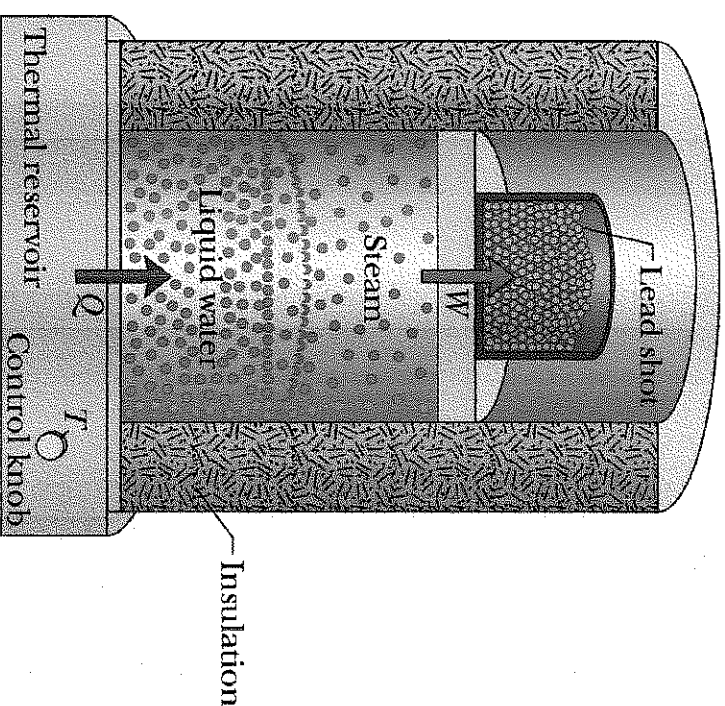
Let 1.0 kg of liquid water at 100°C be converted to steam at 100°C by boiling at standard atmospheric pressure (which is 1.0 atm or  $1.01 \times 10^5$  Pa) in the arrangement of the figure. The volume of the water changes from an initial value of  $1.0 \times 10^{-3}$  m<sup>3</sup> as a liquid to 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

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

$$W = -p(V_f - V_i)$$

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

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



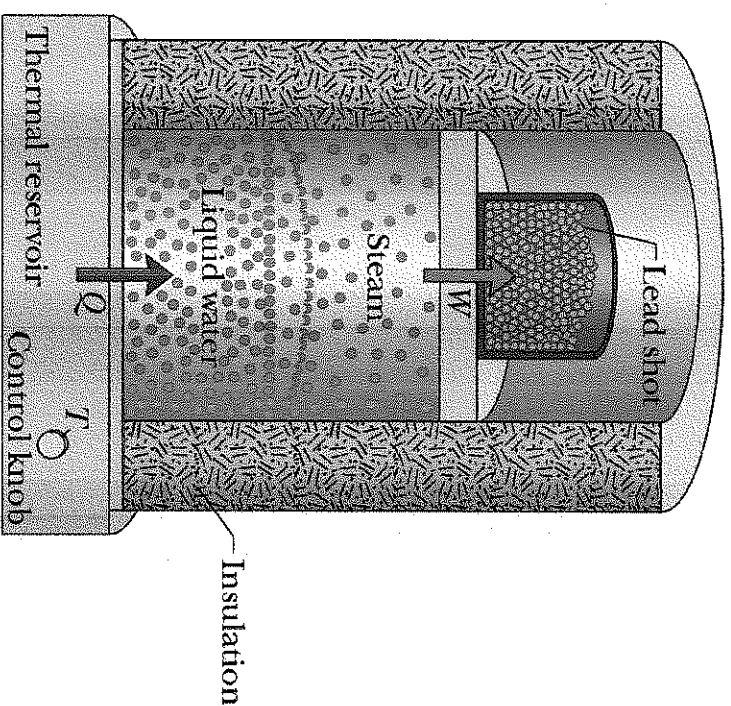
# Sample Problem 18-5 (cont)

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

$$Q = L_V m = (2256 \text{ kJ/kg})(1.0 \text{ kg}) = 2,256 \text{ kJ}$$

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

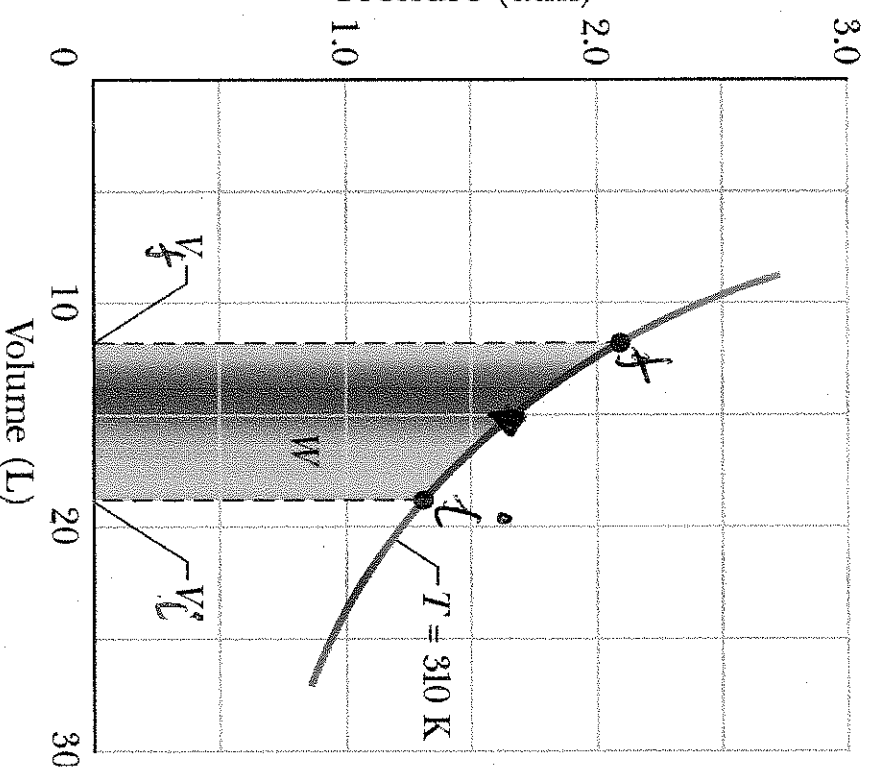
$$\begin{aligned} \Delta U &= Q + W, \text{ heat } W = -169 \text{ kJ} \\ \Delta U &= Q + W = 2256 \text{ kJ} - 169 \text{ kJ} \\ &= 2,087 \text{ kJ} \end{aligned}$$




# Sample Problem 19-2

One mole of oxygen (assume to be an ideal gas) expands 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?

$$\begin{aligned}
 W &= -nRT \ln \frac{V_f}{V_i} \\
 &= nRT \ln \frac{V_i}{V_f} \\
 &= (1 \text{ mol})(8.31 \text{ J/mol}\cdot\text{K})(310 \text{ K}) \ln \left( \frac{12 \text{ L}}{19 \text{ L}} \right) \\
 &= 1183 \text{ J}
 \end{aligned}$$



0.3 mol He 

$$V_f = 5V_i \text{ at } 300 \text{ K} \quad | \quad W_{\text{gas}} = ?$$

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

*Known:*  $n = 0.30 \text{ mol}$ ;  $T = 300 \text{ K}$ ;  $V_f = 5V_i$ .

**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,  $\nabla P$  at  $T = \text{const}$   
 $Q = ?$  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:*  $n = 0.10 \text{ mol}$ ;  $T = 300 \text{ K}$ ;  $V_f = 3V_i$ .

**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.

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

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

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

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

$Q = -W$

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

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

Expansion  $Q=0$ , 2 moles  $H_2$  do 750 J of work

$\Delta T = ?$

**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:*  $n = 2.0 \text{ mol}$ ;  $W = -750 \text{ J}$ .

**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:

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

$$\begin{aligned} P_f V_f &= nRT_f \\ P_i V_i &= nRT_i \end{aligned}$$

Solve for the temperature difference and insert know values:

$$\Delta T = \frac{(\gamma - 1)W}{nR} = \frac{\left(\frac{7}{2} - 1\right)(-750 \text{ J})}{(2.0 \text{ mol})(8.31 \text{ J/(mol}\cdot\text{K)})} = -18 \text{ K}$$

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

**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.