CHAPTER 19: Heat and the First Law of Thermodynamics

Responses to Questions

3. (*a*) No. Because the internal energies of solids and liquids are complicated and include potential energies associated with the bonds between atoms and molecules, two objects may have different internal energies but the same temperature. Internal energy will also vary with the mass of the object. If two objects that are at different temperatures are placed in contact, there will be a net energy transfer from the hotter object to the colder one, regardless of their internal energies.

 (*b*) Yes. Just as in (*a*), the transfer of energy depends on the temperature difference between the two objects, which may not be directly related to the difference in internal energies.

5.   Because the specific heat of water is quite large, water can absorb a large amount of energy with a small increase in temperature. Water can be heated, then easily transported throughout a building, and will give off a large amount of energy as it cools. This makes water particularly useful in radiator systems.

7.   When water at 100ºC comes in contact with the skin, energy is transferred to the skin and the water begins to cool. When steam at 100ºC comes in contact with the skin, energy is transferred to the skin and the steam begins to condense to water at 100ºC. Steam burns are often more severe than water burns due to the energy given off by the steam as it condenses, before it begins to cool.

8.   Energy is needed to convert water in the liquid state to the gaseous state (latent heat). Some of the energy needed to evaporate molecules on the surface comes from the internal energy of the water, thus decreasing the water temperature.

9.   No. The water temperature cannot go above 100ºC, no matter how vigorously it is boiling. The rate at which potatoes cook depends on the temperature at which they are cooking.

18.   Metabolism is a biochemical process by which living organisms get energy from food. If a body is doing work and losing heat, then its internal energy would drop drastically if there were no other

source of energy. Metabolism (food) supplies this other source of energy. The first law of thermodynamics applies because the energy contributions from metabolism are included in Q.

20. An adiabatic compression is one that takes place with no exchange of heat with the surroundings. During the compression, work is done on the gas. Since no heat leaves the gas, then the work results in an increase in the gas's internal energy by the same amount, and therefore an increase in its temperature. $\Delta E_{int} = Q - W$, so if $Q = 0$ then $\Delta E_{int} = -W$.

Solutions to Problems

2. Find the mass of warmed water from the volume of water and its density of $1000\,\text{kg}/\text{m}^3$. Then use the fact that 1 kcal of energy raises 1 kg of water by $1\text{C}°$, and that the water warms by $25\text{C}°$.

$$V = At = \frac{m}{\rho} \quad \rightarrow \quad m = \rho At = \left(1025\,\text{kg}/\text{m}^3\right)\left(1.0\,\text{m}^2\right)\left(0.50\times10^{-3}\,\text{m}\right) = 0.5125\,\text{kg}$$

$$\left(0.5125\,\text{kg}\right)\left(25\text{C}°\right)\frac{\left(1\,\text{kcal}\right)}{\left(1\,\text{kg}\right)\left(1\text{C}°\right)} = 12.8\,\text{kcal} \;\; ; \;\; 12.8\,\text{kcal}\left(\frac{1\,\text{bar}}{300\,\text{kcal}}\right) = \boxed{0.043\,\text{bars}}$$

3. (a) $2500\,\text{Cal}\left(\dfrac{4.186\times10^3\,\text{J}}{1\,\text{Cal}}\right) = \boxed{1.0\times10^7\,\text{J}}$

   (b) $2500\,\text{Cal}\left(\dfrac{1\,\text{kWh}}{860\,\text{Cal}}\right) = \boxed{2.9\,\text{kWh}}$

   (c) At 10 cents per day, the food energy costs $\boxed{\$0.29\ \text{per day}}$. It would be practically impossible to feed yourself in the United States on this amount of money.

6. The wattage rating is 350 Joules per second. Note that 1 L of water has a mass of 1 kg.

$$\left[\left(2.5\times10^{-1}\,\text{L}\right)\left(\frac{1\,\text{kg}}{1\,\text{L}}\right)\left(60\text{C}°\right)\right]\frac{1\,\text{kcal}}{\left(1\,\text{kg}\right)\left(1\text{C}°\right)}\left(\frac{4186\,\text{J}}{\text{kcal}}\right)\left(\frac{1\,\text{s}}{350\,\text{J}}\right) = \boxed{180\,\text{s} = 3.0\,\text{min}}$$

9. (a) The heat absorbed can be calculated from Eq. 19-2. Note that 1 L of water has a mass of 1 kg.

$$Q = mc\Delta T = \left[(1.0\,\text{L}) \left(\frac{1 \times 10^{-3}\,\text{m}^3}{1\,\text{L}} \right) \left(\frac{1.0 \times 10^3\,\text{kg}}{1\,\text{m}^3} \right) \right] (4186\,\text{J/kg·C°})(100°\text{C} - 20°\text{C})$$

$$= 3.349 \times 10^5\,\text{J} \approx \boxed{3.3 \times 10^5\,\text{J}}$$

(*b*)   Power is the rate of energy usage.

$$P = \frac{\Delta E}{\Delta t} = \frac{Q}{\Delta t} \quad \rightarrow \quad \Delta t = \frac{Q}{P} = \frac{3.349 \times 10^5\,\text{J}}{100\,\text{W}} \approx \boxed{3300\,\text{s}} \approx 56\,\text{min}$$

10. The heat absorbed by all three substances is given by Eq. 19-2, $Q = mc\Delta T$. Thus the amount of mass can be found as $m = \dfrac{Q}{c\Delta T}$. The heat and temperature change are the same for all three substances.

$$m_{\text{Cu}} : m_{\text{Al}} : m_{\text{H}_2\text{O}} = \frac{Q}{c_{\text{Cu}}\Delta T} : \frac{Q}{c_{\text{Al}}\Delta T} : \frac{Q}{c_{\text{H}_2\text{O}}\Delta T} = \frac{1}{c_{\text{Cu}}} : \frac{1}{c_{\text{Al}}} : \frac{1}{c_{\text{H}_2\text{O}}} = \frac{1}{390} : \frac{1}{900} : \frac{1}{4186}$$

$$= \frac{4186}{390} : \frac{4186}{900} : \frac{4186}{4186} = 10.7 : 4.65 : 1 \approx \boxed{11 : 4.7 : 1}$$

12. The heat lost by the horseshoe must be equal to the heat gained by the iron pot and the water. Note that 1 L of water has a mass of 1 kg.

$$m_{\text{shoe}} c_{\text{Fe}} \left(T_{\text{shoe}} - T_{\text{eq}} \right) = m_{\text{pot}} c_{\text{Fe}} \left(T_{\text{eq}} - T_{\text{pot}} \right) + m_{\text{H}_2\text{O}} c_{\text{H}_2\text{O}} \left(T_{\text{eq}} - T_{\text{H}_2\text{O}} \right)$$

$$(0.40\,\text{kg})(450\,\text{J/kg·C°})(T_{\text{shoe}} - 25.0°\text{C}) = (0.30\,\text{kg})(450\,\text{J/kg·C°})(25.0°\text{C} - 20.0°\text{C})$$
$$+ (1.05\,\text{kg})(4186\,\text{J/kg·C°})(25.0°\text{C} - 20.0°\text{C})$$

$$T_{\text{shoe}} = 150.8°\text{C} \approx \boxed{150°\text{C}}$$

15.   The heat lost by the iron must be the heat gained by the aluminum and the glycerin.

$$m_{\text{Fe}} c_{\text{Fe}} \left(T_{i\,\text{Fe}} - T_{\text{eq}} \right) = m_{\text{Al}} c_{\text{Al}} \left(T_{\text{eq}} - T_{i\,\text{Al}} \right) + m_{\text{gly}} c_{\text{gly}} \left(T_{\text{eq}} - T_{i\,\text{gly}} \right)$$
$$(0.290\,\text{kg})(450\,\text{J/kg·C°})(142\text{C°}) = (0.095\,\text{kg})(900\,\text{J/kg·C°})(28\text{C°}) + (0.250\,\text{kg}) c_{\text{gly}} (28\text{C°})$$
$$c_{\text{gly}} = 2305\,\text{J/kg·C°} \approx \boxed{2300\,\text{J/kg·C°}}$$

17. We assume that all of the kinetic energy of the hammer goes into heating the nail.

$$KE = Q \;\rightarrow\; 10\left(\tfrac{1}{2}m_{hammer}v_{hammer}^2\right) = m_{nail}c_{Fe}\Delta T \;\rightarrow$$

$$\Delta T = \frac{10\left(\tfrac{1}{2}m_{hammer}v_{hammer}^2\right)}{m_{nail}c_{Fe}} = \frac{5(1.20\text{ kg})(7.5\text{ m/s})^2}{(0.014\text{ kg})(450\text{ J/kg}\cdot C^\circ)} = 53.57\,C^\circ \approx \boxed{54\,C^\circ}$$

19. Assume that the heat from the person is only used to evaporate the water. Also, we use the heat of vaporization at room temperature (585 kcal/kg), since the person's temperature is closer to room temperature than 100°C.

$$Q = mL_{vap} \;\rightarrow\; m = \frac{Q}{L_{vap}} = \frac{180\,\text{kcal}}{585\,\text{kcal/kg}} = 0.308\,\text{kg} \approx \boxed{0.31\,\text{kg}} = 310\,\text{mL}$$

20. Assume that all of the heat lost by the ice cube in cooling to the temperature of the liquid nitrogen is used to boil the nitrogen, and so none is used to raise the temperature of the nitrogen. The boiling point of the nitrogen is $77\,\text{K} = -196^{\circ}\text{C}$.

$$m_{ice}c_{ice}\left(T_{ice\;initial} - T_{ice\;final}\right) = m_{nitrogen}L_{vap} \;\rightarrow$$

$$m_{nitrogen} = \frac{m_{ice}c_{ice}\left(T_{ice\;initial} - T_{ice\;final}\right)}{L_{vap}} = \frac{\left(3.5\times10^{-2}\text{kg}\right)\left(2100\text{ J/kg}\cdot C^\circ\right)\left(0^\circ C - -196^\circ C\right)}{200\times10^3\text{ J/kg}} = \boxed{7.2\times10^{-2}\text{kg}}$$

22.   (a)   The heater must heat both the boiler and the water at the same time.

$$Q_1 = Pt_1 = \left(m_{Fe}c_{Fe} + m_{H_2O}c_{H_2O}\right)\Delta T \;\rightarrow$$

$$t_1 = \frac{\left(m_{Fe}c_{Fe} + m_{H_2O}c_{H_2O}\right)\Delta T}{P} = \frac{\left[(180\text{ kg})(450\text{ J/kg}\cdot C^\circ) + (730\text{ kg})(4186\text{ J/kg}\cdot C^\circ)\right](82\,C^\circ)}{5.2\times10^7\text{ J/h}}$$

$$= 4.946\,\text{h} \approx \boxed{4.9\,\text{h}}$$

(b) Assume that after the water starts to boil, all the heat energy goes into boiling the water, and none goes to raising the temperature of the iron or the steam.

$$Q_2 = Pt_2 = m_{H_2O}L_{vap} \rightarrow t_2 = \frac{m_{H_2O}L_{vap}}{P} = \frac{(730\,\text{kg})(22.6\times10^5\,\text{J/kg})}{5.2\times10^7\,\text{J/h}} = 31.727\,\text{h}$$

Thus the total time is $t_1 + t_2 = 4.946\,\text{h} + 31.727\,\text{h} = 36.673\,\text{h} \approx \boxed{37\,\text{h}}$

24. The heat lost by the aluminum and the water must equal the heat needed to melt the mercury and to warm the mercury to the equilibrium temperature.

$$m_{Al}c_{Al}\left(T_{Al} - T_{eq}\right) + m_{H_2O}c_{H_2O}\left(T_{H_2O} - T_{eq}\right) = m_{Hg}\left[L_{fusion} + c_{Hg}\left(T_{eq} - T_{melt}\right)\right]$$

$$L_{fusion} = \frac{m_{Al}c_{Al}\left(T_{Al} - T_{eq}\right) + m_{H_2O}c_{H_2O}\left(T_{H_2O} - T_{eq}\right)}{m_{Hg}} - c_{Hg}\left(T_{eq} - T_{melt}\right)$$

$$= \frac{\left[(0.620\text{ kg})(900\,\text{J/kg}C°) + (0.400\text{ kg})(4186\,\text{J/kg}C°)\right](12.80°C - 5.06°C)}{1.00\text{ kg}}$$

$$- (138\,\text{J/kg}C°)\left[5.06°C - (-39.0°C)\right]$$

$$= \boxed{1.12\times10^4\,\text{J/kg}}$$

25. The kinetic energy of the bullet is assumed to warm the bullet and melt it.

$$\tfrac{1}{2}mv^2 = Q = mc_{Pb}\left(T_{melt} - T_{initial}\right) + mL_{fusion} \rightarrow$$

$$v = \sqrt{2\left[c_{Pb}\left(T_{melt} - T_{initial}\right) + L_{fusion}\right]} = \sqrt{2\left[(130\,\text{J/kg}C°)(327°C - 20°C) + (0.25\times10^5\,\text{J/kg})\right]}$$

$$= \boxed{360\,\text{m/s}}$$

27. Segment *A* is the compression at constant pressure. Since the process is at a constant pressure, the path on the diagram is horizontal from 2.0 L to 1.0 L.

Segment *B* is the isothermal expansion. Since the temperature is constant, the ideal gas law says that the product *PV* is constant.
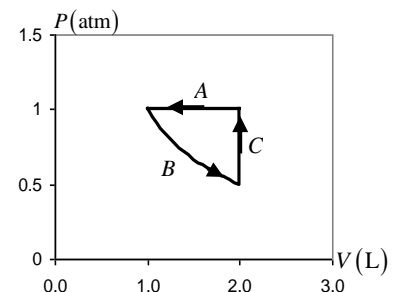
Since the volume is doubled, the pressure must be halved, and so the final point on this segment is at a pressure of 0.5 atm. The path is a piece of a hyperbola.

Segment C is the pressure increase at constant volume. Since the process is at a constant volume, the path on the diagram is vertical from 0.5 atm to 1.0 atm.

31. (*a*)   No work is done during the first step, since the volume is constant. The work in the second step

is given by $W = P\Delta V$.

$$W = P\Delta V = (1.4 \text{ atm})\left(\frac{1.01 \times 10^5 \text{ Pa}}{1 \text{ atm}}\right)(9.3\text{L} - 5.9\text{L})\left(\frac{1 \times 10^{-3}\text{m}^3}{1\text{ L}}\right) = \boxed{480\text{J}}$$

(*b*) Since there is no overall change in temperature, $\Delta E_{\text{int}} = \boxed{0\text{J}}$

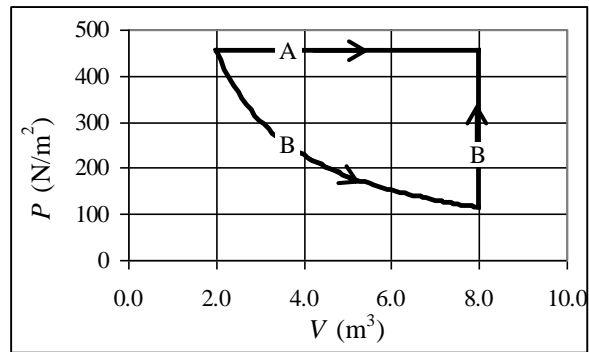(*c*) The heat flow can be found from the first law of thermodynamics.

$$\Delta E_{\text{int}} = Q - W \;\;\rightarrow\;\; Q = \Delta E_{\text{int}} + W = 0 + 480\text{J} = \boxed{480\text{J (into the gas)}}$$

32.   (*a*)   See the diagram. The isobaric expansion

is just a horizontal line on the graph.

(*b*)   The work done is found from Eq. 19-9a.

$$W = P\Delta V$$
$$= (455 \text{ N/m}^2)(8.00\text{m}^3 - 2.00\text{m}^3)$$
$$= \boxed{2730\text{J}}$$

The change in internal energy depends on the temperature change, which can be related to the ideal gas law, $PV = nRT$.

$$\Delta E_{\text{int}} = \tfrac{3}{2}nR\Delta T = \tfrac{3}{2}(nRT_2 - nRT_1)$$

$$= \tfrac{3}{2}\left[(PV)_2 - (PV)_1\right] = \tfrac{3}{2}P\Delta V = \tfrac{3}{2}W = \tfrac{3}{2}(2730\text{J}) = \boxed{4.10 \times 10^3\text{J}}$$

(c) For the isothermal expansion, since the volume expands by a factor of 4, the pressure drops by a factor of 4 to $114\,\mathrm{N/m^2}$. The spreadsheet used for this problem can be found on the Media Manager, with filename "PSE4_ISM_CH19.XLS," on tab "Problem 19.32."

(d) The change in internal energy only depends on the initial and final temperatures. Since those temperatures are the same for process (B) as they are for process (A), the internal energy change is the same for process (B) as for process (A), $\boxed{4.10\times10^3\,\mathrm{J}}$.

33. (a) The work done by an ideal gas during an isothermal volume change is given by Eq. 19-8.

$$W = nRT\ln\frac{V_2}{V_1} = (2.60\,\mathrm{mol})(8.314\,\mathrm{J/mol\cdot K})(290\,\mathrm{K})\ln\frac{7.00\,\mathrm{m^3}}{3.50\,\mathrm{m^3}} = 4345.2\,\mathrm{J} \approx \boxed{4350\,\mathrm{J}}$$

(b) Since the process is isothermal, there is no internal energy change. Apply the first law of thermodynamics.

$$\Delta E_{\mathrm{int}} = Q - W = 0 \;\rightarrow\; Q = W = \boxed{4350\,\mathrm{J}}$$

(c) Since the process is isothermal, there is no internal energy change, and so $\boxed{\Delta E_{\mathrm{int}} = 0}$.

35. Since the expansion is adiabatic, there is no heat flow into or out of the gas. Use the first law of thermodynamics to calculate the temperature change.

$$\Delta E_{\mathrm{int}} = Q - W \;\rightarrow\; \tfrac{3}{2}nR\Delta T = 0 - W \;\rightarrow$$

$$\Delta T = -\tfrac{2}{3}\frac{W}{nR} = -\frac{2(7500\,\mathrm{J})}{3(1.5\,\mathrm{mol})(8.315\,\mathrm{J/mol\cdot K})} = -401\,\mathrm{K} = \boxed{-4.0\times10^2\,\mathrm{K}}$$

38. For the path ac, use the first law of thermodynamics to find the change in internal energy.

$$\Delta E_{\mathrm{int}_{ac}} = Q_{ac} - W_{ac} = -63\,\mathrm{J} - (-35\,\mathrm{J}) = -28\,\mathrm{J}$$

Since internal energy only depends on the initial and final temperatures, this ΔE_{int} applies to any path that starts at a and ends at c. And for any path that starts at c and ends at a,
$\Delta E_{\mathrm{int}_{ca}} = -\Delta E_{\mathrm{int}_{ac}} = 28\,\mathrm{J}$.

(*a*) Use the first law of thermodynamics to find Q_{abc} .

$$\Delta E_{\underset{abc}{int}} = Q_{abc} - W_{abc} \quad \rightarrow \quad Q_{abc} = \Delta E_{\underset{abc}{int}} + W_{abc} = -28\,\text{J} + (-54\,\text{J}) = \boxed{-82\,\text{J}}$$

(*b*) Since the work along path bc is 0, $W_{abc} = W_{ab} = P_b \Delta V_{ab} = P_b\left(V_b - V_a\right)$. Also note that the work

along path da is 0.

$$W_{cda} = W_{cd} = P_c \Delta V_{cd} = P_c\left(V_d - V_c\right) = \tfrac{1}{2}P_b\left(V_a - V_b\right) = -\tfrac{1}{2}W_{abc} = -\tfrac{1}{2}(-54\,\text{J}) = \boxed{27\,\text{J}}$$

(*c*) Use the first law of thermodynamics to find Q_{abc} .

$$\Delta E_{\underset{cda}{int}} = Q_{cda} - W_{cda} \quad \rightarrow \quad Q_{cda} = \Delta E_{\underset{cda}{int}} + W_{cda} = 28\,\text{J} + 27\,\text{J} = \boxed{55\,\text{J}}$$

(*d*) As found above, $E_{int,a} - E_{int,c} = \Delta E_{\underset{ca}{int}} = -\Delta E_{\underset{ac}{int}} = \boxed{28\,\text{J}}$

(*e*) Since $E_{int,d} - E_{int,c} = 12\,\text{J},\ E_{int,d} = E_{int,c} + 12\,\text{J}$ and so $\Delta E_{\underset{da}{int}} = E_{int,a} - E_{int,d} = E_{int,a} - \left(E_{int,c} + 12\ \text{J}\right)$

which then gives $\Delta E_{\underset{da}{int}} = \Delta E_{\underset{ca}{int}} - 12\,\text{J} = 28\,\text{J} - 12\,\text{J} = 16\,\text{J}$. Use the first law of thermodynamics to

find Q_{da} .

$$\Delta E_{\underset{da}{int}} = Q_{da} - W_{da} \quad \rightarrow \quad Q_{da} = \Delta E_{\underset{da}{int}} + W_{da} = 16\,\text{J} + 0 = \boxed{16\,\text{J}}$$

51.   For a diatomic gas with no vibrational modes excited, we find the γ parameter.

$$\gamma = \frac{C_P}{C_V} = \frac{C_V + R}{C_V} = \frac{\tfrac{5}{2}R + R}{\tfrac{5}{2}R} = \tfrac{7}{5}$$

For an adiabatic process, we have $PV^\gamma = \text{constant}$. Use this to find the final pressure.

$$P_1 V_1^\gamma = P_2 V_2^\gamma \quad \rightarrow \quad P_2 = P_1\left(\frac{V_1}{V_2}\right)^\gamma = (1.00\,\text{atm})\left(\frac{1}{1.75}\right)^{1.4} = 0.4568\,\text{atm} \approx \boxed{0.457\,\text{atm}}$$

Use the ideal gas law to find the final temperature.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \rightarrow \quad T_2 = T_1\left(\frac{P_2}{P_1}\right)\left(\frac{V_2}{V_1}\right) = (293\,\text{K})(0.4568)(1.75) = 234\,\text{K} = \boxed{-39°\text{C}}$$

53.   (a)   We first find the final pressure from the adiabatic relationship, and then use the ideal gas law to

find the temperatures. For a diatomic gas, $\gamma = 1.4$.

$$P_1V_1^{\gamma} = P_2V_2^{\gamma} \quad \rightarrow \quad P_2 = P_1\left(\frac{V_1}{V_2}\right)^{\gamma} = (1.00\,\text{atm})\left(\frac{0.1210\,\text{m}^3}{0.750\,\text{m}^3}\right)^{1.4} = 7.772\times10^{-2}\,\text{atm}$$

$$PV = nRT \quad \rightarrow \quad T = \frac{PV}{nR} \qquad T_1 = \frac{P_1V_1}{nR} = \frac{\left(1.013\times10^5\,\text{Pa}\right)\left(0.1210\,\text{m}^3\right)}{(3.65\,\text{mol})\left(8.314\,\dfrac{\text{J}}{\text{molgK}}\right)} = 403.9\,\text{K} \approx \boxed{404\,\text{K}}$$

$$T_2 = \frac{P_2V_2}{nR} = \frac{\left(7.772\times10^{-2}\right)\left(1.013\times10^5\,\text{Pa}\right)\left(0.750\,\text{m}^3\right)}{(3.65\,\text{mol})\left(8.314\,\dfrac{\text{J}}{\text{molgK}}\right)} = 194.6\,\text{K} \approx \boxed{195\,\text{K}}$$

(b)

$$\Delta E_{\text{int}} = nC_C\Delta T = n\left(\tfrac{5}{2}R\right)\Delta T = (3.65\,\text{mol})\tfrac{5}{2}\left(8.314\,\frac{\text{J}}{\text{molgK}}\right)(194.6\,\text{K} - 403.9\,\text{K}) = \boxed{-1.59\times10^4\,\text{J}}$$

(c)   Since the process is adiabatic, no heat is transferred. $\boxed{Q = 0}$

(d)   Use the first law of thermodynamics to find the work done by the gas. The work done ON the gas is the opposite of the work done BY the gas.

$$\Delta E_{\text{int}} = Q - W \quad \rightarrow \quad W = Q - \Delta E_{\text{int}} = 0 - \left(-1.59\times10^4\,\text{J}\right) = 1.59\times10^4\,\text{J}$$

$$W_{\substack{\text{on}\\\text{gas}}} = \boxed{-1.59\times10^4\,\text{J}}$$

89. (a) Since the pressure is constant, the work is found using Eq. 19-9a.

$$W = P\Delta V = \left(1.013\times10^5\,\text{Pa}\right)\left(4.1\,\text{m}^3 - 2.2\,\text{m}^3\right) = 1.9247\times10^5\,\text{J} \approx \boxed{1.9\times10^5\,\text{J}}$$

(b) Use the first law of thermodynamics.

$$\Delta E_{int} = Q - W = 6.30 \times 10^5 \text{J} - 1.92 \times 10^5 \text{J} = 4.38 \times 10^5 \text{J} \approx \boxed{4.4 \times 10^5 \text{J}}$$

(*c*) See the accompanying graph.

P (atm)

1.0

0   2.2   4.1

V (m ³)