20-4. (a) This may be considered a reversible process (as well as isothermal), so we use $\Delta S = Q/T$ where $Q = Lm$ with $L = 333$ J/g from Table 18-4. Consequently,

$$\Delta S = \frac{(333 \text{ J/g})(12.0 \text{ g})}{273 \text{ K}} = 14.6 \text{ J/K} .$$

(b) The situation is similar to that described in part (a), except with $L = 2256$ J/g, $m = 5.00$ g, and $T = 373$ K. We therefore find $\Delta S = 30.2$ J/K.

20-11. The connection between molar heat capacity and the degrees of freedom of a diatomic gas is given by setting $f = 5$ in Eq. 19-51. Thus, $C_V = \frac{5}{2}R$, $C_p = \frac{7}{2}R$, and $\gamma = \frac{7}{5}$. In addition to various equations from Chapter 19, we also make use of Eq. 20-4 of this chapter. We note that we are asked to use the ideal gas constant as R and not plug in its numerical value. We also recall that isothermal means constant-temperature, so $T_2 = T_1$ for the $1 \rightarrow 2$ process.

(a) The gas law in ratio form (see Sample Problem 19-1) as well as the adiabatic relations Eq. 19-54 and Eq. 19-56 are used to obtain

$$p_2 = p_1\left(\frac{V_1}{V_2}\right) = \frac{p_1}{3} ,$$

$$p_3 = p_1\left(\frac{V_1}{V_3}\right)^{\gamma} = \frac{p_1}{3^{1.4}} \approx 0.215p_1 ,$$

$$T_3 = T_1\left(\frac{V_1}{V_3}\right)^{\gamma-1} = \frac{T_1}{3^{0.4}} \approx 0.644T_1 .$$

(d) process $1 \rightarrow 2$:

The work is given by Eq. 19-14: $W = nRT_1 \ln(V_2/V_1) = RT_1 \ln 3$ which is approximately $1.10nRT_1$.

The energy absorbed as heat is given by the first law of thermodynamics: $Q = \Delta E_{\text{int}} + W = 0 + W \approx 1.10nRT_1$.

The internal energy change is $\Delta E_{\text{int}} = 0$ since this is an ideal gas process without a temperature change (see Eq. 19-45).

The entropy change is $\Delta S = Q/T_1 = 1.10nR$.

(h) process $2 \rightarrow 3$:

The work is zero since there is no volume change.

The internal energy change is

$$Q = nC_V(T_3 - T_2) = n\left(\frac{5}{2}R\right)\left(\frac{T_1}{3^{0.4}} - T_1\right) \approx -0.889nRT_1 .$$

This $(-0.889nRT_1)$ is also the value for ΔE_{int} (by either the first law of thermodynamics or by the definition of C_V).

For the entropy change, we obtain

$$\begin{aligned}\Delta S &= nR \ln\left(\frac{V_3}{V_1}\right) + nC_V \ln\left(\frac{T_3}{T_1}\right) \\ &= nR \ln(1) + n\left(\frac{5}{2}R\right)\ln\left(\frac{T_1/3^{0.4}}{T_1}\right) \\ &= 0 + \frac{5}{2}nR \ln\left(3^{-0.4}\right) \approx -1.10nR .\end{aligned}$$

(l) process $3 \rightarrow 1$:

$W = Q - \Delta E_{\text{int}} = 0 - \frac{5}{2}nR(T_1 - T_3) = -\frac{5}{2}nRT_1(1 - 3^{-0.4}) = -0.889nRT_1$

By definition, $Q = 0$ in an adiabatic process, which also implies an absence of entropy change (taking this to be a reversible process).

The internal energy change must be $-W$ (see above), so $\Delta E_{\text{int}} = +0.889nRT_1$.

20-13. (a) We refer to the copper block as block 1 and the lead block as block 2. The equilibrium temperature T_f satisfies $m_1 c_1(T_f - T_{i,1}) + m_2 c_2(T_f - T_{i,2}) = 0$, which we solve for T_f:

$$
\begin{aligned}
T_f &= \frac{m_1 c_1 T_{i,1} + m_2 c_2 T_{i,2}}{m_1 c_1 + m_2 c_2} \\
&= \frac{(50\,\text{g})(.386\,\text{J/g·K})(400\,\text{K}) + (100\,\text{g})(.128\,\text{J/g·K})(200\,\text{K})}{(50\,\text{g})(.386\,\text{J/g·K}) + (100\,\text{g})(.128\,\text{J/g·K})} \\
&= 320\,\text{K} .
\end{aligned}
$$

(b) Since the two-block system in thermally insulated from the environment, the change in internal energy of the system is zero.

(c) The change in entropy is

$$
\begin{aligned}
\Delta S &= \Delta S_1 + \Delta S_2 = m_1 c_1 \ln\left(\frac{T_f}{T_{i,1}}\right) + m_2 c_2 \ln\left(\frac{T_f}{T_{i,2}}\right) \\
&= (50\,\text{g})(.386\,\text{J/g·K}) \ln\left(\frac{320\,\text{K}}{400\,\text{K}}\right) + (100\,\text{g})(.128\,\text{J/g·K}) \ln\left(\frac{320\,\text{K}}{200\,\text{K}}\right) \\
&= +1.72\,\text{J/K} .
\end{aligned}
$$

20-16. In coming to equilibrium, the heat lost by the $100\,\text{cm}^3$ of liquid water (of mass $m_w = 100$ g and specific heat capacity $c_w = 4.190\,\text{J/g·K}$) is absorbed by the ice (of mass m_i which melts and reaches $T_f > 0°\text{C}$). We begin by finding the equilibrium temperature:

$$
\sum Q = 0
$$

$$
Q_{\text{warm water cools}} + Q_{\text{ice warms to } 0°} + Q_{\text{ice melts}} + Q_{\text{melted ice warms}} = 0
$$

$$
c_w m_w (T_f - 20°) + c_i m_i (0° - (-10°)) + L_F m_i + c_w m_i (T_f - 0°) = 0
$$

which yields, after using $L_F = 333$ J/g and values cited in the problem, $T_f = 12.24°$ which is equivalent to $T_f = 285.39$ K. Sample Problem 20-2 shows that

$$
\Delta S_{\text{temp change}} = mc \ln\left(\frac{T_2}{T_1}\right)
$$

for processes where $\Delta T = T_2 - T_1$, and Eq. 20-2 gives

$$
\Delta S_{\text{melt}} = \frac{L_F m}{T_o}
$$

for the phase change experienced by the ice (with $T_o = 273.15$ K). The total entropy change is (with T in Kelvins)

$$
\begin{aligned}
\Delta S_{\text{system}} &= m_w c_w \ln\left(\frac{285.39}{293.15}\right) + m_i c_i \ln\left(\frac{273.15}{263.15}\right) + m_i c_w \ln\left(\frac{285.39}{273.15}\right) + \frac{L_F m_i}{273.15} \\
&= -11.241 + 0.663 + 1.469 + 9.753 = 0.644\,\text{J/K} .
\end{aligned}
$$

20-26. (a) Eq. 20-11 leads to

$$
\varepsilon = 1 - \frac{T_L}{T_H} = 1 - \frac{333\,\text{K}}{373\,\text{K}} = 0.107 .
$$

We recall that a Watt is Joule-per-second. Thus, the (net) work done by the cycle per unit time is the given value 500 J/s. Therefore, by Eq. 20-9, we obtain the heat input per unit time:

$$
\varepsilon = \frac{W}{|Q_H|} \implies \frac{0.500\,\text{kJ/s}}{0.107} = 4.66\,\text{kJ/s} .
$$

(b) Considering Eq. 20-6 on a per unit time basis, we find $4.66 - 0.500 = 4.16$ kJ/s for the rate of heat exhaust.

20-27. (a) Energy is added as heat during the portion of the process from a to b. This portion occurs at constant volume (V_b), so $Q_{in} = nC_V \Delta T$. The gas is a monatomic ideal gas, so $C_V = \frac{3}{2}R$ and the ideal gas law gives $\Delta T = (1/nR)(p_b V_b - p_a V_a) = (1/nR)(p_b - p_a)V_b$. Thus, $Q_{in} = \frac{3}{2}(p_b - p_a)V_b$. V_b and p_b are given. We need to find p_a. Now p_a is the same as p_c and points c and b are connected by an adiabatic process. Thus, $p_c V_c^{\gamma} = p_b V_b^{\gamma}$ and

$$p_a = p_c = \left(\frac{V_b}{V_c}\right)^{\gamma} p_b = \left(\frac{1}{8.00}\right)^{5/3} (1.013 \times 10^6 \, \text{Pa}) = 3.167 \times 10^4 \, \text{Pa} \, .$$

The energy added as heat is

$$Q_{in} = \frac{3}{2}(1.013 \times 10^6 \, \text{Pa} - 3.167 \times 10^4 \, \text{Pa})(1.00 \times 10^{-3} \, \text{m}^3) = 1.47 \times 10^3 \, \text{J} \, .$$

(b) Energy leaves the gas as heat during the portion of the process from c to a. This is a constant pressure process, so

$$
\begin{aligned}
Q_{out} &= nC_p \Delta T = \frac{5}{2}(p_a V_a - p_c V_c) = \frac{5}{2}p_a(V_a - V_c) \\
&= \frac{5}{2}(3.167 \times 10^4 \, \text{Pa})(-7.00)(1.00 \times 10^{-3} \, \text{m}^3) = -5.54 \times 10^2 \, \text{J} \, .
\end{aligned}
$$

The substitutions $V_a - V_c = V_a - 8.00V_a = -7.00V_a$ and $C_p = \frac{5}{2}R$ were made.

(c) For a complete cycle, the change in the internal energy is zero and $W = Q = 1.47 \times 10^3 \, \text{J} - 5.54 \times 10^2 \, \text{J} = 9.18 \times 10^2 \, \text{J}$.

(d) The efficiency is $\varepsilon = W/Q_{in} = (9.18 \times 10^2 \, \text{J})/(1.47 \times 10^3 \, \text{J}) = 0.624$.

20-32. (a) Using Eq. 19-54 for process $D \rightarrow A$ gives

$$
\begin{aligned}
p_D V_D^{\gamma} &= p_A V_A^{\gamma} \\
\frac{p_0}{32}(8V_0)^{\gamma} &= p_0 V_0^{\gamma}
\end{aligned}
$$

which leads to

$$8^{\gamma} = 32 \implies \gamma = \frac{5}{3}$$

which (see §19-9 and §19-11) implies the gas is monatomic.

(b) The input heat is that absorbed during process $A \rightarrow B$:

$$Q_H = nC_p \Delta T = n\left(\frac{5}{2}R\right)T_A \left(\frac{T_B}{T_A} - 1\right) = nRT_A \left(\frac{5}{2}\right)(2-1) = p_0 V_0 \left(\frac{5}{2}\right)$$

and the exhaust heat is that liberated during process $C \rightarrow D$:

$$Q_L = nC_p \Delta T = n\left(\frac{5}{2}R\right)T_D \left(1 - \frac{T_L}{T_D}\right) = nRT_D \left(\frac{5}{2}\right)(1-2) = -\frac{1}{4}p_0 V_0 \left(\frac{5}{2}\right)$$

where in the last step we have used the fact that $T_D = \frac{1}{4}T_A$ (from the gas law in ratio form — see Sample Problem 19-1). Therefore, Eq. 20-10 leads to

$$\varepsilon = 1 - \left|\frac{Q_L}{Q_H}\right| = 1 - \frac{1}{4} = 0.75 = 75\% \, .$$

20-35. A Carnot refrigerator working between a hot reservoir at temperature T_H and a cold reservoir at temperature T_L has a coefficient of performance K that is given by $K = T_L/(T_H - T_L)$. For the refrigerator of this problem, $T_H = 96° \, \text{F} = 308.7 \, \text{K}$ and $T_L = 70° \, \text{F} = 294.3 \, \text{K}$, so $K = (294.3 \, \text{K})/(308.7 \, \text{K} - 294.3 \, \text{K}) = 20.4$. The coefficient of performance is the energy Q_L drawn from the cold reservoir as heat divided by the work done: $K = |Q_L|/|W|$. Thus, $|Q_L| = K|W| = (20.4)(1.0 \, \text{J}) = 20.4 \, \text{J}$.

1. This is an isothermal process so: $Q = T\Delta S = (273 + 342)(5.3098 - 3.6848) = 1000$ kJ

2. These are isobaric processes so $W = p\Delta V$, and $Q = \Delta E + p\Delta V$

   (a) 1b→2: $3209 - 2456 + 15 \times 10^3 \cdot (.02491 - .01034) = 753 + 219 = 972$ kJ

   (b) 3→4: $3297 - 2782 + 10^3 \cdot (.4011 - .2042) = 515 + 197 = 712$ kJ

3. The heat required 6b→1a approximated as a straight-line pV process: first work: $W = (15 \times 10^3 + 15) \cdot (.00166 - .00101-)/2 = 5$ kJ then heat $Q = \Delta E + W = 1586 - 226 + 5 = 1365$ kJ,

4. These are adiabatic processes so $W = -\Delta E$

   (a) 2→3: $3209 - 2782 = 427$ kJ

   (b) 4→5: $3297 - 2578 = 719$ kJ

5. The only useful work occurs in the turbine: 427+719=1146. The required heat: 1365+1000+972+712=4049, for an efficiency of .28. (Note: this is well below actual: much of the 6b→1a heating is in fact done with 'used' steam.) Carnot gives: $1 - 327/873 = .63$.

6. $Q = \Delta E + p\Delta V = 226 - 2449 + 15(.00101 - 10.0228) = -2220 - 150 = -2370$ kJ

7. $m = Q/c\Delta T = 2370/4.186 * 20 = 28$ kg