5.1 A heat engine based on a Carnot cycle does 2.50 kJ of work per cycle and has an efficiency of 45.0%.
What are \(q_H\) and \(q_C\) for one cycle?

**Solution**

Since the engine does work on the surroundings it loses that amount of energy so \(w = -2.50 \text{ kJ}\). The efficiency is given as \(\varepsilon = -w/q_H\), so we can solve for \(q_H\):

\[
q_H = \frac{-w}{\varepsilon} = \frac{-(2.50 \text{ kJ})}{0.45} = 5.6 \text{ kJ}
\]

Now we can find \(q_C\) from the expression for the efficiency:

\[
\varepsilon = 1 + \frac{q_C}{q_H}
\]

Solve for \(q_C\):

\[
q_C = q_H (\varepsilon - 1) = (5.6 \text{ kJ})(0.55) = -3.1 \text{ kJ}
\]

5.2 A heat engine running in reverse is called a heat pump (e.g., a refrigerator is a type of heat pump). A heat pump uses an input of work, \(w_{cyc}\), to cause the system to absorb heat, \(q_C\), from a cold reservoir at \(T_C\) and emit heat, \(q_H\), into a hot reservoir at \(T_H\). The efficiency of a heat pump, \(\varepsilon_{HP}\), is called the coefficient of performance and is defined by \(-q_H/w_{cyc}\). The efficiency of a refrigerator, \(\varepsilon_R\), is called the coefficient of performance and is defined by \(q_C/w_{cyc}\). For a reversible Carnot cycle heat pump and refrigerator, express (a) \(\varepsilon_{HP}\) and (b) \(\varepsilon_R\) in terms of \(T_C\) and \(T_H\). (c) Show that \(\varepsilon_{HP}\) is always greater than 1. (d) Suppose a reversible heat pump transfers heat from the outdoors at 273 K to a room at 293 K. For each joule of work input to the heat pump, how much heat will be deposited in the room? (e) What happens to \(\varepsilon_R\) as \(T_C\) goes to 0 K?

(Note that \(w_{cyc}\) is the algebraic sum of all the work steps around the cycle. Whereas \(w_{cyc}\) < 0 for heat engines, \(w_{cyc}\) > 0 for heat pumps and refrigerators.)

**Solution**

a. For the Carnot cycle,

\[
\Delta U_{cyc} = q_C + q_H + w_{cyc}
\]

\[
\downarrow \text{ internal energy is a state function}
\]

\[
= 0
\]

so we can write the work in terms of the heat:

\[
w_{cyc} = -(q_C + q_H)
\]

Substituting this into the expression for \(\varepsilon_{HP}\):

\[
\varepsilon_{HP} = -\frac{q_H}{w_{cyc}} = \frac{q_H}{q_C + q_H}
\]

\[
= \frac{1}{1 + \frac{q_C}{q_H}}
\]

\[
\downarrow \text{ in class we showed that } \frac{q_C}{q_H} = \frac{T_C}{T_H}
\]

\[
= \frac{1}{1 - \frac{T_C}{T_H}}
\]

\[
= \frac{T_H}{T_H - T_C}
\]

b. Now do the same for \(\varepsilon_R\):

(continued on next page)
\[ \varepsilon_R = \frac{q_c}{w_{cyc}} \]
\[ = -\frac{q_c}{q_c + q_H} \]
\[ = -\frac{1}{1 + \frac{q_H}{q_c}} \]
\[ \downarrow \text{ substitute } \frac{q_H}{q_c} = -\frac{T_H}{T_C} \text{ from } \frac{q_c}{q_H} = \frac{T_C}{T_H} \]
\[ = -\frac{1}{1 - \frac{T_H}{T_C}} \]
\[ = \frac{T_C}{T_C - T_H} \]
\[ = \frac{T_C}{T_C - T_H} \]
\[ \frac{T_H}{T_C} \]
\[ \frac{1}{T_C - T_H} \]
\[ \frac{T_C}{T_C - T_H} \]
\[ c. \] Since, by definition, \( T_H > T_C \), the denominator in \( \varepsilon_{HP} \) from (a) is always positive and less than \( T_H \) so that \( \varepsilon_{HP} \) will always be greater than 1.
\[ d. \] For this part, we want to relate \( q_H \) (the heat deposited into the room) to \( w_{cyc} \) (the net work input). This is exactly what the coefficient of performance describes. Equating the expression derived in (a) with the definition of \( \varepsilon_{HP} \) gives
\[ \varepsilon_{HP} = -\frac{q_H}{w_{cyc}} = \frac{T_H}{T_H - T_C} \]
\[ \downarrow \text{ using } T_H = 293 \text{ K and } T_C = 273 \text{ K} \]
\[ q_H = \frac{T_H}{T_H - T_C} w_{cyc} \]
\[ \downarrow \]
\[ = -14.7 w_{cyc} \]
According to this result, if \( w_{cyc} = +1 \text{ J} \), then \( q_H = -14.7 \text{ J} \). In other words, for each joule of work input, 14.7 J of heat are deposited into the room.
\[ e. \] With \( \varepsilon_R = T_C/(T_H - T_C) \), it’s clear that \( \varepsilon_R \to 0 \text{ as } T_C \to 0 \).

5.3 The enthalpy of vaporization of argon at 87.3 K and 1 atm (its normal boiling point) is \( 6.53 \times 10^3 \text{ J mol}^{-1} \). At this temperature and pressure, what is (a) the entropy of vaporization, \( \Delta_{vap}S^\circ \) and (b) the entropy of condensation, \( \Delta_{cond}S^\circ \)?

**Solution**
\[ a. \] \( \Delta_{vap}S^\circ \) is related to \( \Delta_{vap}H^\circ \) according to:
\[ \Delta_{vap}S^\circ = \frac{\Delta_{vap}H^\circ}{T_{vap}} \]
\[ = \frac{6.53 \times 10^3 \text{ J mol}^{-1}}{87.3 \text{ K}} \]
\[ = +74.8 \text{ JK}^{-1} \text{ mol}^{-1} \]
\[ b. \] Because entropy is a state function and condensation is the opposite of vaporization,
\[ \Delta_{cond}S^\circ = -\Delta_{vap}S^\circ = -74.8 \text{ JK}^{-1} \text{ mol}^{-1} \]
5.4 At 1 atm and between 300 K and 400 K, the constant pressure molar heat capacity of O₂ gas is given by the empirical relationship \( C_{p,m} = a + bT \) where \( a = 25.75 \text{ J K}^{-1} \text{ mol}^{-1} \) and \( b = 0.013 \text{ J K}^{-2} \text{ mol}^{-1} \). What is \( \Delta S \) when 2 mol of O₂ is heated from 300 K to 400 K at a constant pressure of 1 atm?

**Solution**
The entropy change associated with a change in temperature from \( T_1 \) to \( T_2 \) is

\[
\Delta S = N \frac{C_{p,m}}{T} \ln \frac{T_2}{T_1}
\]

\( \downarrow \) substitute the given expression for \( C_{p,m} \)

\[
\Delta S = N \left[ a \frac{1}{T} dT + b \left( \frac{T}{T_1} \right) \right]
\]

\[
= N \left( a \ln \frac{T_2}{T_1} + b \Delta T \right)
\]

\[
= (2 \text{ mol}) \left[ (25.75 \text{ J K}^{-1} \text{ mol}^{-1}) \ln \frac{400K}{300K} + (0.013 \text{ J K}^{-2} \text{ mol}^{-1})(400K - 300K) \right]
\]

\[= 17.4 \text{ JK}^{-1}\]

5.5 Determine whether the following are true or false. If a statement is false, explain why or cite a counter example. (a) For a closed system, \( \Delta S_{sys} \) can never be negative. (b) For a reversible process in a closed system, \( \Delta S_{sys} \) must be zero. (c) For a reversible process in a closed system, \( \Delta S_{tot} \) must be zero. (d) For an adiabatic process in a closed system, \( \Delta S_{sys} \) cannot be negative. (e) For a process in an isolated system, \( \Delta S_{tot} \) cannot be negative. (f) For an adiabatic process in a closed system, \( \Delta S_{sys} \) must be zero. (g) An adiabatic process cannot decrease the entropy of a closed system. (h) For a closed system, equilibrium has been reached when \( S_{sys} \) has been maximized.

**Solution**

a. F. Cooling a gas that is held in an impermeable container would make \( \Delta S_{sys} < 0 \).

b. F. Reversible isothermal expansion of an ideal gas corresponds to \( \Delta S_{sys} > 0 \).

c. T

d. T

e. T

f. This is true if the process is reversible and false if it is irreversible.

g. T

h. F. Entropy is maximized at equilibrium for an isolated system (not a closed system). In a closed system, equilibrium is a balance between maximizing the entropy and minimizing the energy.

5.6 For each of the following processes deduce whether each of the quantities \( \Delta S_{sys} \) and \( \Delta S_{tot} \) is positive, zero, or negative. (a) Reversible melting of solid benzene at 1 atm and the normal melting temperature. (b) Reversible melting of ice at 1 atm and 273 K. (c) Reversible adiabatic expansion of an ideal gas. (d) Reversible isothermal expansion of an ideal gas. (e) Joule-Thomson expansion of an ideal gas. (f) Reversible adiabatic expansion of an ideal gas into a vacuum. (g) Reversible cooling of an ideal gas at constant pressure. (h) Reversible heating of an ideal gas at constant temperature. (i) Combustion of benzene in a sealed container with rigid, adiabatic walls. (j) Adiabatic expansion of a nonideal gas into vacuum.

**Solution**

The answers are on the next page. Keep the following in mind:

\[
dS_{rev} = \frac{\delta q_{rev}}{T}, \quad dS_{err} > \frac{\delta q_{rev}}{T}, \quad \Delta S_{tot} = \Delta S_{sys} + \Delta S_{sur} \begin{cases} = 0 \text{ if reversible} \\ > 0 \text{ if irreversible} \end{cases}
\]

and for adiabatic processes \( \Delta S_{sys} \begin{cases} = 0 \text{ if reversible} \\ > 0 \text{ if irreversible} \end{cases} \)

(continued on next page)
a. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} = 0$

b. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} = 0$

c. $\Delta S_{\text{sys}} = 0$ and $\Delta S_{\text{tot}} = 0$

d. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} = 0$

e. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} > 0$

f. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} > 0$ (Comment: Joule-Thomson expansion is adiabatic and irreversible)

g. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} = 0$

h. $\Delta S_{\text{sys}} < 0$ and $\Delta S_{\text{tot}} = 0$

i. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} > 0$ (Comment: combustion is irreversible)

j. $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{tot}} > 0$

5.7 (a) What is $\Delta S$ for each step of a Carnot cycle? (b) What is $\Delta S_{\text{tot}}$ for each step of a Carnot cycle?

**Solution**

a. For both adiabatic steps, $\Delta S = 0$. For the isothermal compression at $T_H$:

$$\Delta S = q_i = NRT\ln\frac{V_i}{V_f} = NRT\ln\frac{P_f}{P_i}$$

For the isothermal expansion at $T_C$:

$$\Delta S = \frac{q_i}{T_C} = NRT\ln\frac{V_i}{V_f} = NRT\ln\frac{P_i}{P_f}$$

b. All the steps of a Carnot cycle are reversible, so $\Delta S_{\text{tot}} = 0$ for each.

5.8 One mol of an ideal gas is taken through the cycle shown (this is a diesel cycle). Step $1 \rightarrow 2$ is an adiabatic expansion to a volume of 20 L from a pressure of 5 atm and an initial volume of 10 L. During step $2 \rightarrow 3$, the gas cools by releasing 1 kJ of heat. The gas then undergoes an adiabatic compression back to the original pressure of 5 atm during step $3 \rightarrow 4$. The final step, $4 \rightarrow 1$, is an expansion to the original volume of 10 L against a constant external pressure of 5 atm. For each step and the overall cycle, calculate $\Delta S$ for the system, surroundings, and total. Sketch what this cycle would look like if $T$ vs. $S$ were plotted instead of $P$ vs. $V$. For this problem, assume that $C_{V,m} = \frac{5R}{2}$ and that all the steps are reversible.

**Solution**

All the steps are reversible so $\Delta S_{\text{tot}} = 0$ and $\Delta S_{\text{sur}} = -\Delta S_{\text{sys}}$. All we have to do is calculate $\Delta S_{\text{sys}}$ for each step.

1 $\rightarrow$ 2:

For a reversible adiabatic expansion $\Delta S_{\text{sys}} = 0$ and $\Delta S_{\text{sur}} = 0$.

2 $\rightarrow$ 3:

For a constant volume decrease in temperature from $T_2$ to $T_3$ we use

$$\Delta S_{\text{sys}} = NC_{V,m} T_3\ln\frac{T_3}{T_2}$$

We need $C_{V,m}$, $T_2$, and $T_3$. We are given $C_{P,m}$, so $C_{V,m}$ is just $C_{P,m} - R = \frac{3R}{2}$. To find $T_2$ and $T_3$, we will need $T_1$. The gas is ideal, so this is just:

$$T_1 = \frac{PV}{NRT}$$

$$= \frac{(5\text{ atm})(10\text{ L})}{(1\text{ mol})(8.206 \times 10^{-2} \text{ L atm K}^{-1} \text{ mol}^{-1})}$$

$$= 609.33\text{ K}$$

(continued on next page)
Since the process from 1 to 2 is adiabatic, we can now find \( T_2 \) using

\[
T_2 = T_1 \left( \frac{V_2}{V_1} \right)^{\gamma - 1}
\]

\[
= (609.33K) \left( \frac{20L}{10L} \right)^{\frac{5}{3}}
\]

\[
= 383.85K
\]

To find \( T_3 \), we need to be a little more creative. We’re told that going from 2 to 3, the system releases 1000 J of heat. Since this is a constant volume process, the heat is equal to the change in internal energy which can be calculated from \( C_{V,m} \), \( T_2 \), and \( T_3 \):

\[
q_{23} = \Delta U_{23} = \frac{\partial U}{\partial V} \left|_{V_2} \right| (T_3 - T_2) = -1000J
\]

Note that \( q \) is negative because this is the heat released by the system.

Solve for \( T_3 \):

\[
T_3 = \frac{T_2}{C_{V,m}} + q_{23}
\]

\[
= \frac{383.85K}{(1mol)\left( \frac{3}{2} \times 8.314 \text{JK}^{-1}\text{mol}^{-1} \right)} + (-1000J)
\]

\[
= 303.67K
\]

Now we have what we need to calculate the entropy change for the system:

\[
\Delta S_{sys} = (1mol)\left( \frac{3}{2} \times 8.314 \text{JK}^{-1}\text{mol}^{-1} \right) \ln \frac{303.67K}{383.85K}
\]

\[
= -2.92\text{JK}^{-1}
\]

3 \( \rightarrow \) 4:

This is a reversible adiabatic process (compression in this case) so \( \Delta S_{sys} = 0 \).

4 \( \rightarrow \) 1:

The gas expands reversibly against a constant external pressure of 5 atm and its temperature drops from \( T_4 \) to \( T_1 \). The entropy change for the system is

\[
\Delta S_{sys} = \frac{\partial S}{\partial V} \left|_{P_4} \right| \ln \frac{T_1}{T_4}
\]

However, we first need to determine \( T_4 \) using

\[
T_4 = T_1 \left( \frac{V_4}{V_3} \right)^{\gamma - 1}
\]

which requires \( V_4 \). We can relate the pressures and volumes of states 3 and 4 using:

\[
V_4 = V_3 \left( \frac{P_3}{P_4} \right)^{\frac{1}{\gamma}}
\]

Substituting this into the formula for \( T_4 \) gives:

\[
T_4 = T_1 \left( \frac{P_3}{P_4} \right)^{\frac{1}{\gamma} - 1}
\]

where \( P_3 \) is easily calculated using the ideal gas equation:

\[
P_3 = \frac{NRT_3}{V_3}
\]

\[
= \frac{(1mol)(8.206 \times 10^{-2} \text{Latm}^{-1}\text{mol}^{-1})(303.67K)}{20L}
\]

\[
= 1.25\text{atm}
\]

Now we can calculate \( T_4 \):

\[
T_4 = (303.66K) \left( \frac{1.25\text{atm}}{5\text{atm}} \right)^{\frac{3}{5} - 1}
\]

\[
= 529.41K
\]

(continued on next page)
Finally, we can calculate the entropy change for the system:

$$\Delta S_{\text{sys}} = (1 \text{ mol}) \left( \frac{5}{2} \times 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \right) \ln \left( \frac{609.33 \text{ K}}{529.41 \text{ K}} \right)$$

$$= +2.92 \text{ J K}^{-1}$$

As a final check, we can note that the sum of the entropy changes for the system around the cycle is zero which it should be since the entropy is a state function.

The second part of the question asks us to sketch the Diesel cycle on a $T$ vs. $S$ plot:

5.9 The constant pressure heat capacity for some solid substance at 10 K is 0.62 J K$^{-1}$ mol$^{-1}$. (a) Calculate its third law entropy, $S\,_{010}$, at this temperature, (b) then find $C_{p,m}$ and $S\,_{06}$ at $T = 6$ K.

**Solution**

a. At 10 K, the third law entropy is given by

$$S_{10} = S_{0} + \int_{0}^{10K} \frac{C_{p,m}}{T} \, dT$$

$$= \int_{0}^{10K} \frac{C_{p,m}}{T} \, dT \quad \text{since} \quad S_{0} = 0$$

To integrate, we need to know how the heat capacity varies with temperature between 0 K and 10 K. Because the temperature is so low, we can use the Debye approximation, $C_{p,m} = aT^3$, and use the given heat capacity at 10 K to determine the constant, $a$:

$$C_{p,m} = aT^3$$

↓ solve for $a$

$$a = \frac{C_{p,m}}{T^3} = \frac{0.62 \text{ J K}^{-1} \text{ mol}^{-1}}{(10 \text{ K})^3} = 6.2 \times 10^{-4} \text{ J K}^{-4} \text{ mol}^{-1}$$

The entropy at 10 K can now be calculated:

$$S_{10} = a \int_{0}^{10K} T^2 \, dT$$

$$= \frac{1}{3} a (10 \text{ K})^3$$

$$= \frac{1}{3} (6.2 \times 10^{-4} \text{ J K}^{-4} \text{ mol}^{-1})(10 \text{ K})^3$$

$$= 0.207 \text{ J K}^{-1} \text{ mol}^{-1}$$

b. At $T = 6$ K, we use the same expression in (a), but with 6 K as the upper integration limit:

$$S_{6} = \frac{1}{3} a (6 \text{ K})^3$$

$$= \frac{1}{3} (6.2 \times 10^{-4} \text{ J K}^{-4} \text{ mol}^{-1})(6 \text{ K})^3$$

$$= 0.045 \text{ J K}^{-1} \text{ mol}^{-1}$$
5.10 Consider the reaction $2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g)$. For temperatures between 273 K and 1000 K, the heat capacity for each species is described by the empirical formula $C_{P,m}(T) = a + bT + cT^2 + dT^3$ where $a$, $b$, $c$, and $d$ are empirically determined constants summarized in the following table:

<table>
<thead>
<tr>
<th>Species</th>
<th>$a$ (J·K$^{-1}$·mol$^{-1}$)</th>
<th>$b$ (J·K$^{-2}$·mol$^{-1}$)</th>
<th>$c$ (J·K$^{-3}$·mol$^{-1}$)</th>
<th>$d$ (J·K$^{-4}$·mol$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{O}_2$</td>
<td>25.67</td>
<td>1.33×10$^{-2}$</td>
<td>-3.76×10$^{-6}$</td>
<td>-7.31×10$^{-11}$</td>
</tr>
<tr>
<td>$\text{CO}$</td>
<td>28.74</td>
<td>-1.79×10$^{-3}$</td>
<td>1.05×10$^{-5}$</td>
<td>-4.29×10$^{-9}$</td>
</tr>
<tr>
<td>$\text{CO}_2$</td>
<td>21.64</td>
<td>6.36×10$^{-2}$</td>
<td>-4.06×10$^{-5}$</td>
<td>9.70×10$^{-9}$</td>
</tr>
</tbody>
</table>

(a) Evaluate $\Delta S_{298}$ for this reaction. The third law entropy, $S_{298}$, for $\text{O}_2(g)$, $\text{CO}(g)$, and $\text{CO}_2(g)$ are $205.14$ J·K$^{-1}$·mol$^{-1}$, $197.67$ J·K$^{-1}$·mol$^{-1}$, and $213.74$ J·K$^{-1}$·mol$^{-1}$, respectively. (b) Using the given expression for $C_{P,m}(T)$, evaluate $\Delta S_{1000}$. (c) Calculate $\Delta C_{P,m}(298 \text{K})$ using the given data, then use that to reevaluate $\Delta S_{1000}$ assuming that the heat capacities are temperature independent.

**Solution**

a. The reaction entropy at 298 K is

$$\Delta S_{298}^\circ = 2S_{298}^\circ(\text{CO}_2, g) - 2S_{298}^\circ(\text{CO}, g) - S_{298}^\circ(\text{O}_2, g)$$

$$= 2(213.74 \text{ J·K}^{-1} \text{·mol}^{-1}) - 2(197.67 \text{ J·K}^{-1} \text{·mol}^{-1}) - (205.14 \text{ J·K}^{-1} \text{·mol}^{-1})$$

$$= -173.00 \text{ J·K}^{-1} \text{·mol}^{-1}$$

b. The reaction entropy at $T_2 = 1000$ K is related to the reaction entropy at $T_1 = 298$ K by

$$\Delta S_{T_2}^\circ = \Delta S_{T_1}^\circ + \int_{T_1}^{T_2} \frac{\Delta C_{P,m}(T)}{T} dT$$

where

$$\Delta C_{P,m}(T) = 2C_{P,m}(\text{CO}_2) - 2C_{P,m}(\text{CO}) - C_{P,m}(\text{O}_2)$$

$$= 2(a + bT + cT^2 + dT^3)_{\text{CO}_2} - 2(a + bT + cT^2 + dT^3)_{\text{CO}} - (a + bT + cT^2 + dT^3)_{\text{O}_2}$$

$$= 2a_{\text{CO}_2} + 2b_{\text{CO}_2}T + 2c_{\text{CO}_2}T^2 + 2d_{\text{CO}_2}T^3 - 2a_{\text{CO}} - 2b_{\text{CO}}T - 2c_{\text{CO}}T^2 - 2d_{\text{CO}}T^3 - a_{\text{O}_2} - b_{\text{O}_2}T - c_{\text{O}_2}T^2 - d_{\text{O}_2}T^3$$

$$= \left(2a_{\text{CO}_2} - 2a_{\text{CO}} - a_{\text{O}_2}\right) + \left(2b_{\text{CO}_2} - 2b_{\text{CO}} - b_{\text{O}_2}\right)T + \left(2c_{\text{CO}_2} - 2c_{\text{CO}} - c_{\text{O}_2}\right)T^2 + \left(2d_{\text{CO}_2} - 2d_{\text{CO}} - d_{\text{O}_2}\right)T^3$$

For convenience, use $\Delta a$, $\Delta b$, $\Delta c$, and $\Delta d$ for the quantities in parentheses

$$= \Delta a + \Delta bT + \Delta cT^2 + \Delta dT^3$$

Substitute this into the expression for the reaction entropy at 1000 K:

$$\Delta S_{T_2}^\circ = \Delta S_{T_1}^\circ + \Delta a \ln \frac{T_2}{T_1} + \Delta b \left(\frac{T_2}{T_1} - 1\right) + \Delta c \left(\frac{T_2^2}{T_1^2} - 1\right) + \frac{1}{3} \Delta d \left(\frac{T_2^3}{T_1^3} - 1\right)$$

With $T_1 = 298$ K, $T_2 = 1000$ K, $\Delta S_{298}^\circ$ from (a), and $\Delta a$, $\Delta b$, $\Delta c$, and $\Delta d$ calculated from the data in the table:

$$\Delta S_{1000}^\circ = (-172.00 \text{ J·K}^{-1} \text{·mol}^{-1}) + (-39.87 \text{ J·K}^{-1} \text{·mol}^{-1}) \ln \frac{1000}{298} + (0.1175 \text{ J·K}^{-2} \text{·mol}^{-1})(1000 - 298)$$

$$= -174.54 \text{ J·K}^{-1} \text{·mol}^{-1}$$

c. Use the expression from (b) for $\Delta C_{P,m}(T)$ with $T = 298$ K:

(continued on next page)
If the heat capacities are independent of temperature, the reaction entropy at $T_2$ is related to the entropy at $T_1$ through:

$$
\Delta S_{T_2} = \Delta S_{T_1} + \Delta C_{P,m} \int_{T_1}^{T_2} \frac{1}{T} dT
$$

$$
= \Delta S_{T_1} + \Delta C_{P,m} \ln \frac{T_2}{T_1}
$$

Using $\Delta C_{P,m}(T)$ we just calculated, with $T_2 = 1000 \text{ K}$ and $T_1 = 298 \text{ K}$:

$$
\Delta S_{1000} = (-173.00 \text{ JK}^{-1} \text{ mol}^{-1}) + (-12.86 \text{ JK}^{-1} \text{ mol}^{-1}) \ln \frac{1000}{298}
$$

$$
= -188.57 \text{ JK}^{-1} \text{ mol}^{-1}
$$