3.11 First we need to find the final temperature using conservation of energy. Note that 50 l of water is the same as $50 \times 10^3$ cm$^3$ of water. Thus 50 l of water has a mass of $50 \times 10^3$ g, since the density of water is 1 g/cm$^3$.

\[-Q_{\text{hot}} = Q_{\text{cold}}\]
\[(55^\circ C - T_f)(50 \times 10^3 \text{ g})(4.2 \text{ J/g})\]
\[= (T_f - 10^\circ C)(25 \times 10^3 \text{ g})(4.2 \text{ J/g})\]

$110^\circ C - 2T_f = T_f - 10^\circ C$

$3T_f = 120^\circ C$

$T_f = 40^\circ C$.

So the heat gained (lost) by the cold (hot) water is:

\[Q = (40^\circ C - 10^\circ C)(25 \times 10^3 \text{ g})(4.2 \text{ J/g})\]

\[Q = 3.15 \times 10^6 \text{ J} \quad \text{(a useless fact for this problem, though)}\]

Since the heat capacity of water is relatively constant over the temperature range when water is liquid, we can use Eq. 3.19 to calculate the entropy change and pull $C_v$ out of the integral:

\[\Delta S = C_v \ln \frac{T_f}{T_i} = (4.2 \text{ J/g}^\circ C)(25 \times 10^3 \text{ g}) \ln \left(\frac{40 + 273}{10 + 273}\right)\]

\[= 10,600 \text{ J/K}\]
\[ \Delta S_{\text{hot}} = C_v \ln \frac{T_f}{T_c} = \left( \frac{4.2 \text{ J}}{\text{g} \cdot \text{K}} \right) \left( 50 \times 10^3 \text{ g} \right) \left( \ln \frac{40 + 273}{55 + 273} \right) \]

\[ = -9830 \ \frac{\text{J}}{\text{K}} \]

\[ \Rightarrow \Delta S_{\text{total}} = 770 \ \frac{\text{J}}{\text{K}} \]