

3. (a) Since the gas is ideal, its pressure p is given in terms of the number of moles n , the volume V , and the temperature T by $p = nRT/V$. The work done by the gas during the isothermal expansion is

$$W = \int_{V_1}^{V_2} p dV = nRT \int_{V_1}^{V_2} \frac{dV}{V} = nRT \ln \frac{V_2}{V_1} .$$

We substitute $V_2 = 2V_1$ to obtain

$$W = nRT \ln 2 = (4.00 \text{ mol}) \left(8.31 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (400 \text{ K}) \ln 2 = 9.22 \times 10^3 \text{ J} .$$

- (b) Since the expansion is isothermal, the change in entropy is given by $\Delta S = \int (1/T) dQ = Q/T$, where Q is the heat absorbed. According to the first law of thermodynamics, $\Delta E_{\text{int}} = Q - W$. Now the internal energy of an ideal gas depends only on the temperature and not on the pressure and volume. Since the expansion is isothermal, $\Delta E_{\text{int}} = 0$ and $Q = W$. Thus,

$$\Delta S = \frac{W}{T} = \frac{9.22 \times 10^3 \text{ J}}{400 \text{ K}} = 23.1 \text{ J/K} .$$

- (c) $\Delta S = 0$ for all reversible adiabatic processes.