

61. The aim of this problem is to emphasize what it means for the internal energy to be a state function. Since path 1 and path 2 start and stop at the same places, then the internal energy change along path 1 is equal to that along path 2. Now, during isothermal processes (involving an ideal gas) the internal energy change is zero, so the only step in path 1 that we need to examine is step 2. Equation 19-28 then immediately yields -20 J as the answer for the internal energy change.

45. Argon is a monatomic gas, so $f = 3$ in Eq. 19-51, which provides

$$C_V = \frac{3}{2}R = \frac{3}{2}(8.31 \text{ J/mol} \cdot \text{K}) \left(\frac{1 \text{ cal}}{4.186 \text{ J}} \right) = 2.98 \frac{\text{cal}}{\text{mol} \cdot \text{C}^\circ}$$

where we have converted joules to calories, and taken advantage of the fact that a Celsius degree is equivalent to a unit change on the Kelvin scale. Since (for a given substance) M is effectively a conversion factor between grams and moles, we see that c_V (see units specified in the problem statement) is related to C_V by $C_V = c_V M$ where $M = mN_A$, and m is the mass of a single atom (see Eq. 19-4).

(a) From the above discussion, we obtain

$$m = \frac{M}{N_A} = \frac{C_V / c_V}{N_A} = \frac{2.98 / 0.075}{6.02 \times 10^{23}} = 6.6 \times 10^{-23} \text{ g.}$$

(b) The molar mass is found to be

$$M = C_V / c_V = 2.98 / 0.075 = 39.7 \text{ g/mol}$$

which should be rounded to 40 g/mol since the given value of c_V is specified to only two significant figures.

38. (a) From the graph we see that $v_p = 400$ m/s. Using the fact that $M = 28$ g/mol = 0.028 kg/mol for nitrogen (N_2) gas, Eq. 19-35 can then be used to determine the absolute temperature. We obtain $T = \frac{1}{2} M v_p^2 / R = 2.7 \times 10^2$ K.

(b) Comparing with Eq. 19-34, we conclude $v_{\text{rms}} = \sqrt{3/2} v_p = 4.9 \times 10^2$ m/s.

53. (a) Since the process is at constant pressure, energy transferred as heat to the gas is given by $Q = nC_p \Delta T$, where n is the number of moles in the gas, C_p is the molar specific heat at constant pressure, and ΔT is the increase in temperature. For a diatomic ideal gas $C_p = \frac{7}{2}R$. Thus,

$$Q = \frac{7}{2}nR\Delta T = \frac{7}{2}(4.00 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(60.0 \text{ K}) = 6.98 \times 10^3 \text{ J}.$$

(b) The change in the internal energy is given by $\Delta E_{\text{int}} = nC_V \Delta T$, where C_V is the specific heat at constant volume. For a diatomic ideal gas $C_V = \frac{5}{2}R$, so

$$\Delta E_{\text{int}} = \frac{5}{2}nR\Delta T = \frac{5}{2}(4.00 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(60.0 \text{ K}) = 4.99 \times 10^3 \text{ J}.$$

(c) According to the first law of thermodynamics, $\Delta E_{\text{int}} = Q - W$, so

$$W = Q - \Delta E_{\text{int}} = 6.98 \times 10^3 \text{ J} - 4.99 \times 10^3 \text{ J} = 1.99 \times 10^3 \text{ J}.$$

(d) The change in the total translational kinetic energy is

$$\Delta K = \frac{3}{2}nR\Delta T = \frac{3}{2}(4.00 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(60.0 \text{ K}) = 2.99 \times 10^3 \text{ J}.$$

27. (a) We use $\varepsilon = L_V/N$, where L_V is the heat of vaporization and N is the number of molecules per gram. The molar mass of atomic hydrogen is 1 g/mol and the molar mass of atomic oxygen is 16 g/mol, so the molar mass of H_2O is $(1.0 + 1.0 + 16) = 18$ g/mol. There are $N_A = 6.02 \times 10^{23}$ molecules in a mole, so the number of molecules in a gram of water is $(6.02 \times 10^{23} \text{ mol}^{-1})/(18 \text{ g/mol}) = 3.34 \times 10^{22}$ molecules/g. Thus

$$\varepsilon = (539 \text{ cal/g})/(3.34 \times 10^{22}/\text{g}) = 1.61 \times 10^{-20} \text{ cal} = 6.76 \times 10^{-20} \text{ J}.$$

(b) The average translational kinetic energy is

$$K_{\text{avg}} = \frac{3}{2}kT = \frac{3}{2}(1.38 \times 10^{-23} \text{ J/K})[(32.0 + 273.15) \text{ K}] = 6.32 \times 10^{-21} \text{ J}.$$

The ratio $\varepsilon/K_{\text{avg}}$ is $(6.76 \times 10^{-20} \text{ J})/(6.32 \times 10^{-21} \text{ J}) = 10.7$.

17. When the valve is closed the number of moles of the gas in container A is $n_A = p_A V_A / RT_A$ and that in container B is $n_B = 4p_B V_A / RT_B$. The total number of moles in both containers is then

$$n = n_A + n_B = \frac{V_A}{R} \left(\frac{p_A}{T_A} + \frac{4p_B}{T_B} \right) = \text{const.}$$

After the valve is opened, the pressure in container A is $p'_A = Rn'_A T_A / V_A$ and that in container B is $p'_B = Rn'_B T_B / 4V_A$. Equating p'_A and p'_B , we obtain $Rn'_A T_A / V_A = Rn'_B T_B / 4V_A$, or $n'_B = (4T_A / T_B)n'_A$. Thus,

$$n = n'_A + n'_B = n'_A \left(1 + \frac{4T_A}{T_B} \right) = n_A + n_B = \frac{V_A}{R} \left(\frac{p_A}{T_A} + \frac{4p_B}{T_B} \right).$$

We solve the above equation for n'_A :

$$n'_A = \frac{V}{R} \frac{(p_A / T_A + 4p_B / T_B)}{(1 + 4T_A / T_B)}.$$

Substituting this expression for n'_A into $p'_A V_A = n'_A R T_A$, we obtain the final pressure:

$$p' = \frac{n'_A R T_A}{V_A} = \frac{p_A + 4p_B T_A / T_B}{1 + 4T_A / T_B} = 2.0 \times 10^5 \text{ Pa.}$$