Physics 1401
Homework Solutions - Walker, Chapter 17

Problems

1. Ideal Gas Law says:

$$PV = nRT\ . \qquad (1)$$

Note that I'm using the "chemist's form" of the Ideal Gas Law since we're given the number of *moles* of the gas, not the number of *particles*.

Solving for the volume, I get:

$$V = \frac{nRT}{P}$$

Putting in numbers:

$$V = \frac{\left(1\text{ mol}\right)\left(8.31\ \frac{\text{J}}{\text{mol}\cdot\text{K}}\right)\left(273.15\text{ K}\right)}{\left(1.013\times10^{5}\ \text{Pa}\right)} = 0.0224\text{ m}^3$$

Note that I've converted the temperature 0°C to *kelvins* before substituting it into the ideal gas law. I've also expressed the pressure in *pascals*, not *kilo*pascals.

Finally, I'll point out here that, when expressed in *liters*, this volume becomes:

$$\left(0.0224\text{ m}^3\right)\left(\frac{1\text{ L}}{0.001\text{ m}^3}\right) = 22.4\text{ L}\ ,$$

which you may recognize from your chemistry classes as the well-known result for the volume of one mole of an ideal gas at standard temperature and pressure in the units usually used in chemistry problems.

4. Ideal gas law says:

$$PV = nRT$$

We're told that $V = 0.0185\text{ m}^3$ when $T = 294\text{ K}$ and $P = 212\text{ kPa} = 2.12\times10^5\ \text{Pa}$. So the number of moles initially in the tire is:

$$n_i = \frac{P_i V}{RT} = \frac{\left(2.12\times10^5\ \text{Pa}\right)\left(0.0185\text{ m}^3\right)}{\left(8.31\ \frac{\text{J}}{\text{mol}\cdot\text{K}}\right)\left(294\text{ K}\right)} = 1.61\text{ mol}$$

If we want to increase the pressure to $2.52\times10^5\ \text{Pa}$, then the final number of moles in the tire must be:

$$n_f = \frac{P_f V}{RT} = \frac{\left(2.52\times10^5\ \text{Pa}\right)\left(0.0185\text{ m}^3\right)}{\left(8.31\ \frac{\text{J}}{\text{mol}\cdot\text{K}}\right)\left(294\text{ K}\right)} = 1.91\text{ mol}$$

So the number of additional moles that must be pumped into the tire is:

$$\Delta n = n_f - n_i = 0.30\text{ mol}$$

59. Since this process is assumed to take place in a well insulated container, all the heat that comes out of the iron block must go into the water. But will this amount of heat be enough to change some or all of the liquid water to steam? Well, let's first calculate the amount of heat required to bring the liquid water from 20°C to 100°C:

$$Q_1 = m_w c_w \left(\Delta T\right)_w = \left(0.0400\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(100^\circ\text{C} - 20.0^\circ\text{C}\right) = 1.33952\times10^4\text{ J}.$$

(The specific heat of water, $c_w = 4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}$, can be looked up in Table 16-2. And notice that I've kept many sig figs in this intermediate step to minimize round-off error. I'll round off at the end of the last calculation.)

The amount of heat just calculated just gets the water up to its boiling point. To change all of the water to steam, an additional amount of heat must go into it:

$$Q_2 = m_w \left(L_v\right)_{water},$$

in which $\left(L_v\right)_{water}$ is the latent heat of vaporization of water. Looking this up in Table 17-4, I find:

$$\left(L_v\right)_{water} = 22.6\times10^5\text{ J/kg}.$$

So:

$$Q_2 = \left(0.0400\text{ kg}\right)\left(22.6\times10^5\ \frac{\text{J}}{\text{kg}}\right) = 9.04\times10^4\text{ J}.$$

So the total amount of heat required to convert all the liquid water at 20.0°C to steam at 100°C is:

$$Q_3 = Q_1 + Q_2 = 1.33952\times10^4\text{ J} + 9.04\times10^4\text{ J} = 1.037952\times10^5\text{ J}.$$

Now we must decide whether there is enough heat available to do this. If the iron block cools down to 100°C, the heat liberated will be:

$$Q_{\substack{\text{out of}\\\text{block}}} = m_{block} c_{iron} \left(\Delta T\right)_{block} = \left(0.825\text{ kg}\right)\left(560\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(352\ ^\circ\text{C} - 100\ ^\circ\text{C}\right) = 1.16424\times10^5\text{ J}$$

Since $Q_{\substack{\text{out of}\\\text{block}}} > Q_3$, there will be enough heat available to convert all the liquid water to steam at 100°C. The heat left over after this is done will be:

$$Q_{\substack{\text{left}\\\text{over}}} = Q_{\substack{\text{out of}\\\text{block}}} - Q_3 = 1.26288\times10^4\text{ J}$$

This heat goes into raising the temperature of the steam and block from 100°C to the final equilibrium temperature T_f. So:

$$1.26288\times10^4\text{ J} = m_w c_w \left(\Delta T\right)_w + m_{block} c_{iron} \left(\Delta T\right)_{block}$$

$$1.26288\times10^4\text{ J} = \left(0.0400\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(T_f - 100^\circ\text{C}\right) + \left(0.825\text{ kg}\right)\left(560\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(T_f - 100^\circ\text{C}\right)$$

Solving for T_f, I get:

$$T_f = \frac{\left(1.26288\times10^4\text{ J}\right) + \left(0.0400\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(100^\circ\text{C}\right) + \left(0.825\text{ kg}\right)\left(560\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(100^\circ\text{C}\right)}{\left(0.0400\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right) + \left(0.825\text{ kg}\right)\left(560\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)}$$

$$T_f = 123^\circ\text{C}$$

This is the final equilibrium temperature of the steam and iron block. And *all* of the water has been converted to steam.

62. (a.) The liquid water will cool, liberating some heat, and all of this heat must go into the ice. Since the ice is *already* at 0.0°C, any heat going into it will go into *melting* some (or all) of it. The heat required to melt *all* of the ice is:

$$Q_1 = m_{ice}\left(L_f\right)_{water} = \left(0.075\text{ kg}\right)\left(33.5\times10^4\text{ J/kg}\right).$$

$$Q_1 = 2.5125\times10^4\text{ J}$$

Now the question we must answer is, "Is there enough heat available to melt all the ice?" Well, if the liquid water at 14°C cools all the way down to 0.0°C, the heat liberated will be:

$$Q_2 = m_w c_w\left(\Delta T\right)_w = \left(0.33\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(14^\circ\text{C} - 0.0^\circ\text{C}\right)$$

$$Q_2 = 1.933932\times10^4\text{ J}$$

Comparing Q_2 with Q_1, we see that $Q_2 < Q_1$, so there will not be enough heat to melt all of the ice. But if all of Q_2 goes into melting the ice, as it must, then the mass of ice that *will* be melted can be found from:

$$1.933932\times10^4\text{ J} = m\left(L_f\right)_w$$

$$1.933932\times10^4\text{ J} = m\left(33.5\times10^4\text{ J/kg}\right)$$

$$m = \frac{1.933932\times10^4\text{ J}}{33.5\times10^4\text{ J/kg}} = 0.05773\text{ kg}$$

So the mass of ice remaining will be:

$$m_{remaining} = 0.075\text{ kg} - 0.05773\text{ kg} = 0.017\text{ kg}$$

And the final temperature of the system will be 0.0°C, since the liquid water has now cooled down to 0.0°C and there is still ice remaining.

(b.) If we want the heat coming out of the liquid water as it cools to be just enough to melt all the ice, then we want exactly $Q_1 = 2.5125\times10^4\text{ J}$ of heat to come out of the liquid water as the liquid water cools from its initial temperature T_{wi} down to 0.0°C. Thus, we want:

$$m_w c_w\left(\Delta T\right)_w = 2.5125\times10^4\text{ J}$$

$$\left(0.33\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)\left(T_{wi} - 0.0^\circ\text{C}\right) = 2.5125\times10^4\text{ J}$$

Solving for T_{wi}, I find:

$$T_{wi} = \frac{2.5125\times10^4\text{ J}}{\left(0.33\text{ kg}\right)\left(4186\ \frac{\text{J}}{\text{kg}\cdot^\circ\text{C}}\right)} = 18^\circ\text{C}$$