CHEM 101 Answer to Homework Problems (4th edition)

Chapter 10

10.17

Strategy: This problem is very similar to Sample Problem 10.1. We can use the equation derived in the sample problem to solve for the pressure exerted by the column of toluene. The equation is

\[ \text{pressure} = \text{height} \times \text{density} \times \text{gravitational constant} \]

The gravitational constant is 9.80665 m/s\(^2\).

Solution: Recall that for units to cancel properly, density must be expressed in kg/m\(^3\).

\[ \text{density} = \frac{0.867 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left( \frac{100 \text{ cm}}{1 \text{ m}} \right)^3 = 867 \text{ kg/m}^3 \]

Remember that when a unit is raised to a power, any conversion factor used must be raised to the same power. Substituting into the equation to solve for pressure gives

\[ \text{pressure} = 87 \text{ m} \times \frac{867 \text{ kg}}{\text{m}^3} \times \frac{9.80665 \text{ m}}{\text{s}^2} = 7.40 \times 10^4 \text{ kg/m} \cdot \text{s}^2 = 7.40 \times 10^4 \text{ Pa} \times \frac{1 \text{ atm}}{101,325 \text{ Pa}} = 7.3 \text{ atm} \]
10.25

**Strategy:** Pressure is held constant in this problem. Only volume and temperature change. This is a Charles’s law problem. We use Equation 10.5 to solve for the unknown volume.

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

**Solution:** Initial Conditions | Final Conditions
--- | ---
\(T_1 = 35^\circ C + 273 = 308 \text{ K}\) & \(T_2 = 72^\circ + 273 = 345 \text{ K}\)
\(V_1 = 28.4 \text{ L}\) & \(V_2 = ?\)

\[
V_2 = \frac{V_1 T_2}{T_1} = \frac{(28.4 \text{ L})(345 \text{ K})}{308 \text{ K}} = \mathbf{31.8 \text{ L}}
\]

10.37

In this problem, the moles of gas and the volume the gas occupies are constant. Temperature and pressure change. We use the equation derived in Problem 10.53:

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

Initial Conditions | Final Conditions
--- | ---
\(P_1 = 1.00 \text{ atm}\) & \(P_2 = ?\)
\(T_1 = 273 \text{ K}\) & \(T_2 = 210^\circ C + 273 = 483 \text{ K}\)

Solving for the final pressure gives

\[
P_2 = \frac{P_1 T_2}{T_1} = \frac{(1.00 \text{ atm})(483 \text{ K})}{273 \text{ K}} = \mathbf{1.8 \text{ atm}}
\]
From the moles of CO\textsubscript{2} produced, we can calculate the amount of calcium carbonate that must have reacted. We can then determine the percent by mass of CaCO\textsubscript{3} in the 3.00 g sample.

The balanced equation is:

\[
\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CO}_2(g) + \text{CaCl}_2(aq) + \text{H}_2\text{O}(l)
\]

The moles of CO\textsubscript{2} produced can be calculated using the ideal gas equation.

\[
n_{\text{CO}_2} = \frac{PV_{\text{CO}_2}}{RT}
\]

\[
n_{\text{CO}_2} = \left(\frac{792 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) (20 + 273 \text{ K})}\right) = 2.84 \times 10^{-2} \text{ mol CO}_2
\]

The balanced equation shows a 1:1 mole ratio between CO\textsubscript{2} and CaCO\textsubscript{3}. Therefore, 2.84 \times 10^{-2} mole of CaCO\textsubscript{3} must have reacted.

\[
? \text{ g CaCO}_3 \text{ reacted} = (2.84 \times 10^{-2} \text{ mol CaCO}_3) \times \frac{100.1 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 2.84 \text{ g CaCO}_3
\]

The percent by mass of the CaCO\textsubscript{3} sample is:

\[
\% \text{ CaCO}_3 = \frac{2.84 \text{ g}}{3.00 \text{ g}} \times 100\% = 94.7\%
\]

10.73

\[P_i = X_i P_T\]

We need to determine the mole fractions of each component in order to determine their partial pressures. To calculate mole fraction, write the balanced chemical equation to determine the correct mole ratio.

\[
2\text{NH}_3(g) \rightarrow \text{N}_2(g) + 3\text{H}_2(g)
\]

The mole fractions of N\textsubscript{2} and H\textsubscript{2} are:
\[ \frac{1 \text{ mol}}{3 \text{ mol} + 1 \text{ mol}} = 0.250 \]

\[ \frac{3 \text{ mol}}{3 \text{ mol} + 1 \text{ mol}} = 0.750 \]

The partial pressures of N\(_2\) and H\(_2\) are:

\[ P_{N_2} = X_{N_2} P_T = (0.250)(866 \text{ mmHg}) = 217 \text{ mmHg} \]

\[ P_{H_2} = X_{H_2} P_T = (0.750)(866 \text{ mmHg}) = 650 \text{ mmHg} \]

10.85

\[
\text{RMS speed} = \sqrt{\frac{(2.0^2 + 2.2^2 + 2.6^2 + 2.7^2 + 3.3^2 + 3.5^2)(m/s)^2}{6}} \approx 2.8 \text{ m/s}
\]

\[
\text{Average speed} = \frac{(2.0 + 2.2 + 2.6 + 2.7 + 3.3 + 3.5)m/s}{6} = 2.7 \text{ m/s}
\]

The root-mean-square value is always greater than the average value, because squaring favors the larger values compared to just taking the average value.

10.97

When \(a\) and \(b\) are zero, the van der Waals equation simply becomes the ideal gas equation. In other words, an ideal gas has zero for the \(a\) and \(b\) values of the van der Waals equation. It therefore stands to reason that the gas with the smallest values of \(a\) and \(b\) will behave most like an ideal gas under a specific set of pressure and temperature conditions. Of the choices given in the problem, the gas with the smallest \(a\) and \(b\) values is Ne (see Table 10.6).