A vessel contains 1.15 g liq H₂O in equilibrium with water vapor at 30.°C. At this temperature, the vapor pressure of H₂O is 31.82 torr. What volume increase is necessary for all the water to evaporate?

We calculate the volume at which the gas pressure is equal to the partial pressure. For smaller volumes, some of the water will remain as a liquid.

\[
V = \frac{nRT}{P} = \frac{1.15 \text{ g}}{18.02 \text{ g}} \times 8.206 \times 10^{-2} \text{ L atm mol}^{-1} \text{ K}^{-1} \times (273.15 + 30)\text{K} \\
= \frac{31.82 \text{ Torr} \times \frac{1 \text{ atm}}{760 \text{ Torr}}}{37.9 \text{ L}}
\]

Consider a 31.0 L sample of moist air at 60.°C and one atm in which the partial pressure of water vapor is 0.131 atm. Assume that dry air has the composition 78.0 mole percent N₂, 21.0 mole percent O₂, and 1.00 mole percent Ar.

a. What are the mole percentages of each of the gases in the sample?

b. The percent relative humidity is defined as %RH = \( P_{H₂O}^* \)/\( P_{H₂O} \) where \( P_{H₂O} \) is the partial pressure of water in the sample and \( P_{H₂O}^* \) = 0.197 atm is the equilibrium vapor pressure of water at 60.°C. The gas is compressed at 60.°C until the relative humidity is 100%. What volume does the mixture contain now?

c. What fraction of the water will be condensed if the total pressure of the mixture is isothermally increased to 810 atm?
(a) \[ \text{mol} \% \text{N}_2 = 100 \times \frac{P_{N_2}}{P_{total}} = 100 \times \frac{0.78 \times 0.869 \text{ atm}}{1 \text{ atm}} = 67.8\% \]

\[ \text{mol} \% \text{O}_2 = 100 \times \frac{P_{O_2}}{P_{total}} = 100 \times \frac{0.21 \times 0.869 \text{ atm}}{1 \text{ atm}} = 18.2\% \]

\[ \text{mol} \% \text{Ar} = 100 \times \frac{P_{Ar}}{P_{total}} = 100 \times \frac{0.01 \times 0.869 \text{ atm}}{1 \text{ atm}} = 0.869\% \]

\[ \text{mol} \% \text{H}_2\text{O} = 100 \times \frac{P_{H_2O}}{P_{total}} = 100 \times \frac{0.131 \text{ atm}}{1 \text{ atm}} = 13.1\% \]

(b) \[ P'_{H_2O} = \frac{n_{H_2O}RT}{V} \]

\[ P'_{H_2O}V' = P_{H_2O}V \] where the primed quantities refer to 100% RH

\[ V' = \frac{P'_H_{2O}V}{P'_{H_2O}} = \frac{0.131 \text{ atm} \times 31.0 \text{ L}}{0.197 \text{ atm}} = 20.6 \text{ L} \]

(c) If all the water remained in the gas phase, the partial pressure of water at a total pressure of 81.0 atm would be

\[ P_{H_2O} = P_{total} \times \text{mol fraction} \text{ H}_2\text{O} = 81.0 \text{ atm} \times 0.131 = 10.6 \text{ atm} \]

However, the partial pressure of water cannot be greater than 0.197 atm, and the excess will condense. The fraction that condenses is given by

\[ \text{fraction condensed} = \frac{10.6 \text{ atm} - 0.197 \text{ atm}}{10.6 \text{ atm}} = 0.981 \]
A mixture of oxygen and hydrogen is analyzed by passing it over hot copper oxide and through a drying tube. Hydrogen reduces the CuO according to the reaction \( \text{CuO}(s) + \text{H}_2(g) \rightarrow \text{Cu}(s) + \text{H}_2\text{O}(l) \), and oxygen reoxidizes the copper formed according to \( \text{Cu}(s) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{CuO}(s) \). At 25°C and 750. Torr, 172.0 cm³ of the mixture yields 77.5 cm³ of dry oxygen measured at 25°C and 750. Torr after passage over CuO and the drying agent. What is the original composition of the mixture?

Two consecutive reactions must be considered. In the following reactions alpha (\( \alpha \)) represents the number of moles of \( \text{H}_2(g) \) that react with \( \text{CuO}(s) \) and beta (\( \beta \)) represents the number of moles of \( \text{Cu} \) that react with \( \text{O}_2(g) \).

\[
\begin{align*}
\text{CuO}(s) + \text{H}_2(g) & \rightarrow \text{H}_2\text{O}(l) + \text{Cu} \\
\text{moles present after reaction} & \quad n_{\text{H}_2}^* - \alpha \quad \alpha - \beta \\
\text{Cu}(s) + \frac{1}{2}\text{O}_2(g) & \rightarrow \text{CuO} \\
\text{moles present after reaction} & \quad \alpha - \beta \quad n_{\text{O}_2}^* - \frac{1}{2}\beta
\end{align*}
\]

In the final state, only \( \text{O}_2 \) is present. Therefore \( \alpha = n_{\text{H}_2}^* \). In an excess of \( \text{O}_2 \), all the copper is oxidized. Therefore \( \alpha - \beta = 0 \) or \( \beta = n_{\text{H}_2}^* \). We conclude that \( n_{\text{O}_2} = n_{\text{O}_2}^* - \frac{1}{2}n_{\text{H}_2}^* \).

Let \( V_1 \) and \( V_2 \) be the initial and final volumes.

\[
V_1 = \left( n_{\text{H}_2}^* + n_{\text{O}_2}^* \right) \frac{RT}{P} \quad V_2 = \left( n_{\text{O}_2}^* - \frac{1}{2}n_{\text{H}_2}^* \right) \frac{RT}{P}
\]

Dividing the second equation by the first yields

\[
\frac{V_2}{V_1} = \frac{n_{\text{O}_2}^* - \frac{1}{2}n_{\text{H}_2}^*}{n_{\text{H}_2}^* + n_{\text{O}_2}^*} = x_{\text{O}_2}^* - \frac{1}{2}x_{\text{H}_2}^* = 1 - x_{\text{H}_2}^* - \frac{1}{2}x_{\text{H}_2}^* = 1 - \frac{3}{2}x_{\text{H}_2}^*
\]

\[
x_{\text{H}_2}^* = \frac{2}{3} \left( 1 - \frac{V_2}{V_1} \right) = \frac{2}{3} \left( 1 - \frac{77.5 \text{ cm}^3}{172 \text{ cm}^3} \right) = 0.366; \quad x_{\text{O}_2}^* = 1 - x_{\text{H}_2}^* = 0.634
\]
Consider the oxidation of the amino acid glycine \( \text{NH}_2\text{CH}_2\text{COOH} \) to produce water, carbon dioxide, and urea \( \text{NH}_2\text{CONH}_2 \):

\[
\text{NH}_2\text{CH}_2\text{COOH}(s) + 3\text{O}_2(g) \rightarrow \text{NH}_2\text{CONH}_2(s) + 3\text{CO}_2(g) + 3\text{H}_2\text{O}(l)
\]

Calculate the volume of carbon dioxide evolved at \( P = 1.00 \text{ atm} \) and \( T = 305. \text{ K} \) from the oxidation of 0.022 g of glycine.

From the stoichiometric equation,

\[
n_{\text{CO}_2} = 3n_{\text{glycine}} = 3 \times \frac{0.022 \text{ g}}{75.07 \text{ mol g}^{-1}} = 8.79 \times 10^{-4} \text{ mol}
\]

\[
V_{\text{CO}_2} = \frac{n_{\text{CO}_2}RT}{P} = \frac{8.79 \times 10^{-4} \times 8.206 \times 10^{-2} \text{ L atm mol}^{-1} \text{ K}^{-1} \times 305 \text{ K}}{1.00 \text{ atm}} = 2.20 \times 10^{-2} \text{ L}
\]
A compressed cylinder of gas contains \(2.74 \times 10^3\) g of \(N_2\) gas at a pressure of \(3.75 \times 10^7\) Pa and a temperature of 18.7°C. What volume of gas has been released into the atmosphere if the final pressure in the cylinder is \(1.80 \times 10^5\) Pa? Assume ideal behavior and that the gas temperature is unchanged.

Let \(n_i\) and \(n_f\) be the initial and final number of mols of \(N_2\) in the cylinder.

\[
\frac{n_iRT}{P_i} = \frac{n_fRT}{P_f}
\]

\[
n_f = n_i \frac{P_f}{P_i} = \frac{2.74 \times 10^3 \text{ g}}{28.01 \text{ g mol}^{-1}} \times \frac{1.80 \times 10^5 \text{ Pa}}{3.75 \times 10^7 \text{ Pa}} = 0.470 \text{ mol}
\]

\[
n_i = \frac{2.74 \times 10^3 \text{ g}}{28.01 \text{ g mol}^{-1}} = 97.8 \text{ mol}
\]
The volume of gas released into the atmosphere is given by

\[ V = \frac{(n_f - n_i)RT}{p} = \frac{(97.8 - 0.470) \text{ mol} \times 8.2057 \times 10^{-2} \text{ L atm mol}^{-1} \text{K}^{-1} \times (273.15 + 18.7) \text{ K}}{1 \text{ atm}} = 2.33 \times 10^3 \text{ L} \]
A sample of propane \((\text{C}_3\text{H}_8)\) is placed in a closed vessel together with an amount of \(\text{O}_2\) that is 2.15 times the amount needed to completely oxidize the propane to \(\text{CO}_2\) and \(\text{H}_2\text{O}\) at constant temperature. Calculate the mole fraction of each component in the resulting mixture after oxidation, assuming that the \(\text{H}_2\text{O}\) is present as a gas.

The reaction is \(\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})\). If \(m\) mol of propane are present initially, there must be \(2.15 \times 5 m\) mol of \(\text{O}_2\). After the reaction is complete, there are \(3m\) mol of \(\text{CO}_2\), \(4m\) mol of \(\text{H}_2\text{O}\), and \((5 \times 2.15 - 5)m = 5.75 m\) mol of \(\text{O}_2\).

Given these relationships, we can calculate the mol fractions.

\[
x_{\text{CO}_2} = \frac{3m}{3m + 4m + (5 \times 2.15 - 5)m} = 0.235; \quad x_{\text{H}_2\text{O}} = \frac{4m}{3m + 4m + (5 \times 2.15 - 5)m} = 0.314;
\]

\[
x_{\text{O}_2} = \frac{(5 \times 2.15 - 5)m}{3m + 4m + (5 \times 2.15 - 5)m} = 0.451
\]
Yeast and other organisms can convert glucose \((C_6H_{12}O_6)\) to ethanol \((CH_3CH_2OH)\) by a process called alcoholic fermentation. The net reaction is

\[
C_6H_{12}O_6(s) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)
\]

Calculate the mass of glucose required to produce 2.25 L of \(CO_2\) measured at \(P = 1.00\ \text{atm}\) and \(T = 295\ \text{K}\).

\[
m_{\text{glucose}} = M_{\text{glucose}} \times \frac{PV}{RT} \times \frac{n_{\text{glucose}}}{n_{\text{CO}_2}}
\]

\[
= 180.16 \ \text{g mol}^{-1} \times \frac{1.00 \ \text{atm} \times 2.25 \ \text{L}}{8.206 \times 10^{-2} \ \text{L atm mol}^{-1} \ \text{K}^{-1} \times 295 \ \text{K}} \times 0.5
\]

\[
= 8.37 \ \text{g}
\]
An athlete at high performance inhales \( \sim 3.75 \) L of air at 1.00 atm and 298 K. The inhaled and exhaled air contain 0.50 and 6.2% by volume of water, respectively. For a respiration rate of 32 breaths per minute, how many moles of water per minute are expelled from the body through the lungs?

\[
V_{H_2O} = \text{number breaths} \times V_{\text{air}} \times (\text{fraction inhaled} - \text{fraction exhaled})
\]

\[
= 32 \times 3.75 \text{ L} \times (0.062 - 0.0050) = 6.84 \text{ L}
\]

\[
\text{mol } H_2O = \frac{PV_{H_2O}}{RT} = \frac{1 \text{ atm} \times 6.84 \text{ L}}{8.206 \times 10^{-2} \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}} = 0.280 \text{ mol}
\]