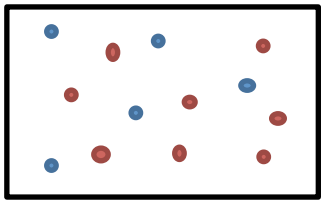
Department of Chemistry
University of Texas at Austin

Mixtures of Gases – Supplemental Worksheet

1. Container 1 has a pressure of 1.5 atm and 5 molecules of O_2 floating around. Container 2 has a pressure of 3.5 atm and 8 molecules of N_2 floating around. Both are containers have the same volume. Given the information what would Container 2 look like if Container 1 was added to it and what would be the total pressure?



Total pressure 5 atm, mixture of O_2 and N_2.

2. A. A piece of solid carbon dioxide, with a mass of 19.4g, is placed in an otherwise empty 4.00 L container at 29°C. What is the pressure in the container after all the carbon dioxide vaporizes?
   B. If 19.4g of solid carbon dioxide was place in a similar container already containing air at 745 torr what would be the partial pressure of carbon dioxide and the total pressure in the container after the carbon dioxide had vaporized?

$$A.\quad P_{CO_2} = \frac{n_{CO_2}RT}{V} = \frac{\left(19.4g \times \frac{1\ mol}{44.01\ g}\right) \times \frac{0.08206\ L\ atm}{K\ mol} \times 302\ K}{4.00\ L} = 2.73\ atm = P_{Total}$$

$$B.\quad P_{Total} = P_{CO_2} + P_{air} = 2.73\ atm + (745\ torr \times \frac{1\ atm}{760\ torr}) = 2.73 + 0.980 = 3.71\ atm$$

3. Helium is collected over water at 25°C and 1.00atm total pressure. What total volume of gas must be collected to obtain 0.492g of helium? (At 25°C the vapor pressure of water is 23.8 torr)

$$P_{He} + P_{H_2O} = 1.00\ atm = 760\ torr = P_{He} + 23.8\ torr \qquad P_{He} = 736.2\ torr$$

$$n_{He} = 0.492\ g \times \frac{1\ mol}{4.003\ g} = 0.123 mol\ He$$

$$V = \frac{n_{He}RT}{P_{He}} = \frac{0.123 mol \times \frac{0.08206\ L\ atm}{K\ mol} \times 298\ K}{736.2\ torr \times \frac{1\ atm}{760\ torr}} = 3.12\ L$$

4. A. In a mixture of the two gases, the partial pressures of CH_4 (g) and O_2 (g) are 0.175 atm and 0.250 atm, respectively. What is the mole fraction of each gas in the mixture?
   B. Calculate the number of moles of each gas in the mixture if the total number of moles of gas is 0.161mol.

$$A.\ mole\ fraction\ CH_4\ \chi_{CH_4} = \frac{P_{CH_4}}{P_{total}} = \frac{0.175\ atm}{0.175+0.250\ atm} = 0.412 \qquad \chi_{O_2} = 1.000 - 0.412 = 0.588$$

$$B.\ \chi_{CH_4} = \frac{n_{CH_4}}{n_{total}};\ n_{CH_4} = \chi_{CH_4} \times n_{total} = 0.412 \times 0.161\ mol = 6.63\times10^{-2} mol\ CH_4$$

$$6.63\times10^{-2} mol\ CH_4 \times \frac{16.04\ g\ CH_4}{mol\ CH_4} = 1.06\ g\ CH_4$$

$$n_{O_2} = 0.588 \times 0.161\ mol = 9.47\times10^{-2} mol\ O_2 \times \frac{32.00\ g\ O_2}{mol\ O_2} = 3.03 g\ O_2$$

5. The mole fraction of nitrogen in air is 0.7808. Calculate the partial pressure of N_2 in air when the atmospheric pressure is 760 torr.

$$P_{N_2} = \chi_{N_2} \times P_{total} = 0.7808 \times 760\ torr = 593\ torr$$

6. Natural gas is a mixture of hydrocarbons, primarily methane (CH_4) and ethane (C_2H_6). A typical mixture might have $\chi_{methane} = 0.923$ and $\chi_{ethane} = 0.077$. What are the partial pressures of the 2 gases in the 14.00 L container of natural gas at 21°C and 1.48 atm? Assuming complete combustion of both gases in the natural gas sample, what is the total mass of water formed?

$$P_{methane} = P_{total} \times \chi_{methane} = 1.48 \times 0.923 = 1.37\ atm$$

$$P_{ethane} = 1.48 - 1.37 = 0.11\ atm \qquad n_{natural\ gas} = \frac{PV}{RT} = \frac{1.48 atm \times 14.00 L}{\frac{0.08206\ L\ atm}{K\ mol} \times 294 K} = 0.859\ mol$$

$$n_{methane} = n_{natural\ gas} \times \chi_{methane} = 0.859 \times 0.923 = 0.793\ mol$$

$$n_{ethane} = 0.859\ mol - 0.793\ mol = 0.066 mol\ ethane$$

CH_4 (g) + $2O_2$ (g) → CO_2 (g) + $2H_2O$ (l); $2C_2H_6$ + $7O_2$ (g) → $4CO_2$ + $6H_2O$ (l)

$$0.793\ mol\ CH_4 \times \frac{2 mol\ H_2O}{1 mol\ CH_4} \times \frac{18.02\ g\ H_2O}{mol\ H_2O} = 28.6\ g\ H_2O$$

$$0.066\ mol\ C_2H_6 \times \frac{6 mol\ H_2O}{2 mol\ CH_4} \times \frac{18.02\ g\ H_2O}{mol\ H_2O} = 3.6\ g\ H_2O$$

Total H_2O = 28.6 g + 3.6g = 32.2g H_2O

7. A sample of solid $KClO_3$ was heated in a test tube and decomposed according to the following reaction:

$KClO_3$ (s) → KCl (s) + O_2 (g)

The oxygen produced was collected by displacement of water at 22°C at a total pressure of 749 torr. The volume of the gas collected was 0.650 L, and the vapor pressure of water at 22°C is 21 torr. Calculate the partial pressure of O_2 in the gas collected and the mass of $KClO_3$ in the sample that was decomposed.

$2KClO_3$ (s) → $2KCl$ (s) + $3O_2$ (g)

$$P_{total} = P_{O_2} + P_{H_2O} = P_{O_2} + 21\ torr = 749\ torr \qquad P_{O_2} = 728\ torr \times \frac{1 atm}{760\ torr} = 0.958\ atm$$

$$n_{O_2} = \frac{0.958\ atm \times 0.650 L}{\frac{0.08206\ L\ atm}{mol\ K} \times 295 K} = 2.57 \times 10^{-2} mol\ O_2$$

$$2.57 \times 10^{-2}\ mol\ O_2 \times \frac{2 mol\ KClO_3}{3 mol\ O_2} = 1.71 \times 10^{-2}\ mol\ KClO_3$$

$$1.71 \times 10^{-2}\ mol\ KClO_3 \times \frac{122.6 g\ KClO_3}{1 mol\ KClO_3} = 2.10\ g\ KClO_3$$

8. Sulfur dioxide reacts with oxygen in the presences of platinum to give sulfur trioxide:

$$2\ SO_2\ (g) + O_2\ (g) \rightarrow 2\ SO_3\ (g)$$

Suppose that at one stage of the reaction, 24.1 mol SO_2, 79.0 mol O_2, and 23 mol SO_3 are present in the reaction vessel at a total pressure of 0.923 atm. Calculate the mole fraction of SO_3 and its partial pressure.

$$n_{total} = 24.1 + 79.0 + 23 = 126.1\ mol$$

$$\chi_{SO_3} = \frac{n_{SO_3}}{n_{total}} = \frac{23 mol}{126.1\ mol} = 0.182$$

$$P_{SO_3} = \chi_{SO_3} \times P_{total} = 0.182 \times 0.923 atm = 0.168\ atm$$