7.8: Dalton’s law of partial pressure

Calculations in examples 2 and 3 of section 7.7 prove that 1.50 mol He in 5.1 L chamber exerts a pressure of 7.2 atm; 0.6 mol He in the same chamber exerts a pressure of 2.9 atm, and a mixture of the two in the same chamber at the same temperature exerts a pressure equal to the sum of the pressures that each fraction exerts if it is alone in the chamber, i.e., $P_{total} = 7.2 atm + 2.9 atm = 10.1 atm$. What if one of the gas was hydrogen, and the other helium? The answer is: that the calculations using the ideal gas law remain the same because it is the number of molecules, not the type of molecules that are involved in the calculations. Properties of gases depend on the number of moles of gas n, and not on the nature of the gas, as illustrated in Fig. 7.8.1.

Dalton's law of partial pressure

The total pressure of a mixture of gases equals the sum of the pressure that each component gas in the mixture would exert if it was present alone.

The mathematical form of Dalton’s law is:

$$P_{\text{tatal}} = P_{\text{He}} + P_{\text{exygen}} = 9.3 \text{ atm} + 2.4 \text{ atm} = 11.5 \text{ atm}$$

, where P_1, P_2, P_3 are partial pressures of individual gas#1, #2, #3 in the mixture.

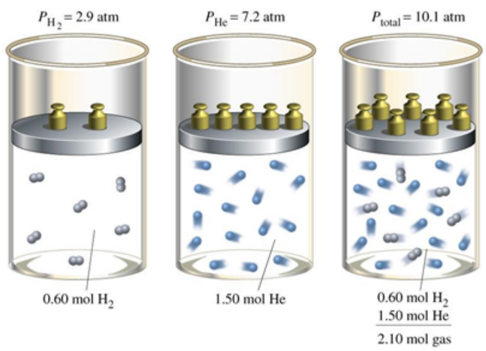

Figure 7.8.1: Illustration of Dalton’s law of partial pressures: the total pressure of a H_2+He mixture equals the sum of the pressure that of H_2 and He exert alone the same size chamber at the same temperature. Source: Dr. Blair Jesse Ellyn Reich / CC BY-SA (https://creativecommons.org/licenses/by-sa/3.0)

Atmospheric air is a mixture of nitrogen, oxygen, argon, carbon dioxide, water vapors, and trance amount of some other gases. The atmospheric pressure is the sum of the partial pressures of components of air, as illustrated in Fig. 7.8.2.

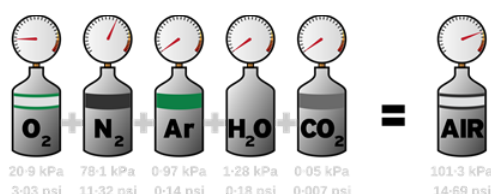

Figure 7.8.2: Partial pressures of main atmospheric constituents at sea level. Source: Andrew Jarvis / CC BY-SA (https://creativecommons.org/licenses/by-sa/4.0).

Example 7.8.1

A 46 L He and 12 L O_2 sample, both at 1.0 atm and 25 ^oC were pumped into a 5.0 L scuba diving tank at 25 ^oC. Calculate the partial pressure of each gas and the total pressure in the tank?

Solution

For He before mixing: V = 46 L, T = 25 ^oC + 273 = 298 K, P = 1.0 atm, n_{He}=?, $R = 0.08206 \frac{L\text{−atm}}{\text{mol}-K}$

Formula and calculations: $n_{\text{He}} = \text{PV/RT} = \frac{1.0 \text{ atm} \times 46 \text{ L}}{0.08206 \frac{\text{Latm}}{\text{mol}-\text{K}} \times 298 \text{ K}} = 1.9 \text{ mol}$

For O_2 before mixing: V=12 L, T =25 ^oC + 273 =298 K, P=1.0 atm, n_{O_2} = ?, $R = 0.08206 \frac{\text{L−atm}}{\text{mol}-\text{K}}$

Formula and calculations: $n_{H_2} = \mathrm{PV/RT} = \frac{1.0\text{ atm}\times 12\text{ L}}{0.08206\frac{\text{L-atm}}{\text{mol}-K}\times 298\text{ K}} = 0.49\text{ mol}$

After mixing: V=5 L, T =25 ^{o}C+273 = 298 K, n_{He}=1.9 mol, n_{O_2} =0.49 mol, $R = 0.08206\frac{L-\text{atm}}{\text{mol}-K}$.
Requried: P_{He} = ?, P_{O_2}=?, and P_{total}=?

Formula: $P_{He} = \frac{n_{He}RT}{V} = \frac{1.9\text{ mol}\times 0.08206\frac{L-atm}{\text{mol}-K}\times 298\text{ K}}{5.0\text{ L}} = 9.3\text{ atm}$, $P_{O_2} = \frac{n_{\text{oxygen}}RT}{V} = \frac{0.49\text{ mol}\times 0.08206\frac{\text{L-atm}}{\text{mol}-K}\times 298\text{ K}}{5.0\text{ L}} = 2.4\text{ atm}$,

$$P_{\text{total}} = P_{He} + P_{O_2} = 9.3\text{ atm} + 2.4\text{ atm} = 11.5\text{ atm}.$$

Hyperbaric chamber -a medical tool

The hyperbaric chamber is an air chamber that is at two to three atmospheric pressure, as shown in Fig. 7.8.3. The solubility of gases increases with an increase in pressure. A patient placed in a hyperbaric chamber has a higher concentration of oxygen dissolved in blood because the partial pressure of oxygen is two to three times higher than in the atmospheric air. The higher concentration of oxygen is toxic to many strains of bacteria. Therefore the hyperbaric chambers are used to treat burn patients, in surgeries, and to treat some cancers.

The hyperbaric chambers are also used to treat carbon monoxide (CO) poisoning because the higher concentration of oxygen in the chamber can displace the CO bound with hemoglobin faster than atmospheric oxygen does. Another use of hyperbaric chambers is to treat scuba divers suffering from the bends. If a diver ascends too quickly, the nitrogen dissolved in blood makes bubbles in the vessels that block the blood flow –a condition called bends. The divers suffering from the bends are placed in a hyperbaric chamber at high pressure, and then the pressure is slowly decreased to atmospheric pressure. The nitrogen dissolves in blood under higher pressure and slowly diffuses out through the lungs as the pressure is gradually decreased.

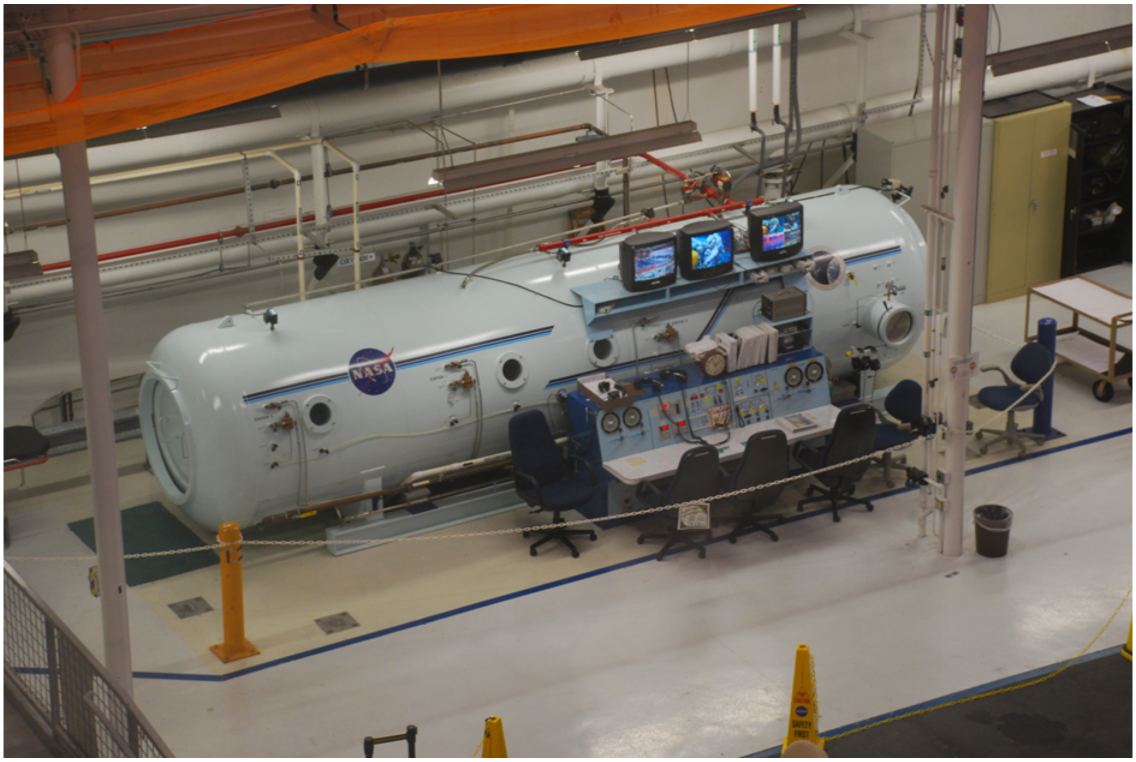

Figure 7.8.3: The hyperbaric chamber at the Neutral Buoyancy Lab. Source: Mike / CC BY (https://creativecommons.org/licenses/by/2.0)