Name:______________________ Date:________

H. Chemistry

Ch. 12.6-12.7 Lecture Guide

- **Dalton's Law of Partial Pressures**
  - Most gases are ___mixtures___ of several gases.
    - Ex) air (oxygen, carbon dioxide, nitrogen, water vapor)
    - Ex) scuba tanks (oxygen + helium)
  - **Partial Pressure**
    - the pressure that the gas would exert if it were alone in the container
  - **Dalton's Law**
    - $P_{total} = P_1 + P_2 + P_3 + \ldots$

Assume that each gas behaves ideally → $P_1 = \frac{n_1 RT}{V}$, $P_2 = \frac{n_2 RT}{V}$, $P_3 = \frac{n_3 RT}{V}$

$$P_{total} = P_1 + P_2 + P_3 = \frac{n_1 RT}{V} + \frac{n_2 RT}{V} + \frac{n_3 RT}{V}$$

  - n_{total} $P_{total} = P_1 + P_2 + P_3 = \underbrace{(n_1 + n_2 + n_3)}_{n_{total}}\left(\frac{RT}{V}\right)$
    - the sum of the numbers of moles of the gases in the mixture.

*Thus, it is the total # of moles of the gases, not the identity of the gases that matters.

  - Practice Problems
    - Mixtures of helium and oxygen are used in "air" tanks of underwater divers for deep dives. For a particular dive, 12 L of O_2 at 25 °C and 1.0 atm and 46 L of He at 25 °C and 1.0 atm were both pumped into a 5.0 L tank. Calculate the partial pressure of each gas and the total pressure in the tank at 25 °C.

P_{O_2} = 1.0 atm
V_{O_2} = 12 L
T_{O_2} = 25 °C = 298 K

P_{He} = 1.0 atm
V_{He} = 46 L
T_{He} = 298 K

V_{tank} = 5.0 L
T_{temp} = 298 K

- Step 1: Use Ideal Gas Law to calculate the number of moles of each gas.
- Step 2: Use Ideal Gas Law to calculate the partial pressure of each gas in the 5.0 L tank.
- Step 3: Add the two pressures together to get the total pressure in the tank.

$$n_{O_2} = \frac{PV}{RT} = \frac{(1.0\ atm)(12\ \cancel{L})}{(0.08206\ \frac{\cancel{L}\cdot \cancel{atm}}{\cancel{K}\cdot mol})(298\ \cancel{K})} = 0.49\ mol\ O_2$$

$$n_{He} = \frac{PV}{RT} = \frac{(1.0\ \cancel{atm})(46\ \cancel{L})}{(0.08206\ \frac{\cancel{L}\cdot \cancel{atm}}{\cancel{K}\cdot mol})(298\ \cancel{K})} = 1.9\ mol\ He$$

$$P_{O_2} = \frac{n_{O_2} RT}{V} = \frac{(0.49\ \cancel{mol})(0.08206\ \frac{\cancel{L}\cdot atm}{\cancel{K}\cdot \cancel{mol}})(298\ \cancel{K})}{(5.0\ \cancel{L})} = \boxed{2.4\ atm\ O_2}$$

$$P_{He} = \frac{n_{He} RT}{V} = \frac{(1.9\ \cancel{mol})(0.08206\ \frac{\cancel{L}\cdot atm}{\cancel{K}\cdot \cancel{mol}})(298\ \cancel{K})}{(5.0\ \cancel{L})} = \boxed{9.3\ atm\ He}$$

$$P_{total} = P_{O_2} + P_{He} = 2.4\ atm + 9.3\ atm = \boxed{11.7\ atm\ total}$$

- A sample of solid potassium chlorate was heated in a test tube and decomposed according to the reaction below:

$$2KClO_3\,(s) \rightarrow 2KCl\,(s) + 3O_2\,(g)$$

The oxygen produced was collected by displacement of water at 22 °C. The resulting mixture of O_2 and H_2O vapor had a total pressure of 754 torr and a volume of 0.650 L. Calculate the partial pressure of O_2 in the gas collected and the number of moles of O_2 present. The vapor pressure of water at 22 °C is 21 torr.

- Step 1: Use the total pressure and the vapor pressure of water to calculate the pressure of O_2.
- Step 2: Use the Ideal Gas Law to solve for the number of moles of O_2.

$$P_{total} = P_{O_2} + P_{H_2O}$$

$$754 \text{ torr} = P_{O_2} + 21 \text{ torr}$$

$$\boxed{P_{O_2} = 733 \text{ torr}}$$

$P_{O_2} = 733 = 0.964 \text{ atm}$

$V = 0.650 \text{ L}$

$T = 22°C = 295 \text{ K}$

$$n_{O_2} = \frac{P_{O_2} V}{RT} = \frac{(0.964 \cancel{\text{atm}})(0.650 \cancel{\text{L}})}{(0.08206 \frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\cancel{\text{K}}\cdot\text{mol}})(295 \cancel{\text{K}})} = \boxed{2.95 \times 10^{-2} \text{ mol } O_2}$$