Dalton's Law of Partial Pressure and Graham's Law of Effusion

Read from **Lesson 2 of Gas Laws** in the **Chemistry Tutorial Section, Chapter 10** of **The Physics Classroom**:

Part g: Dalton's Law of Partial Pressure Part h: Graham's Law of Effusion

Part 1: Dalton's Law of Partial Pressure

Unlike the other gas laws, Dalton's Law of Partial Pressures ONLY refers to pressure. (The other variables are all held constant!)

According to John Dalton's Law, the total pressure of a gas mixture is the sum of the individual gas pressures:

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

The individual pressure of each gas is called its partial pressure. All pressure units in the equation must be the same.

Example Problem: Mark Eury is generating oxygen gas in the chem lab. The oxygen gas is collected over water at 22°C. The total pressure in the container is 127 kPa. What is the partial pressure of the dry oxygen gas? Assume the water vapor pressure at 22°C is 2.65 kPa.

$$P_{total} = P_{O2} + P_{H2O}$$

$$127 \text{ kPa} = P_{O2} + 2.65 \text{ kPa} \qquad P_{O2} = 124 \text{ kPa}$$

Mole Fraction

The pressure of a gas in a given volume and temperature is also dependent on the quantity of gas particles present. If you have a mixture of gases A, B, and C (n_A of gas A, n_B of gas B, and n_C of gas C) at a given volume and temp:

$$n_{total} = n_A + n_B + n_C + \ldots$$

In a mixture of gases at a given volume and temperature, the mole fraction of gas A would be $\chi_A = \dfrac{n_A}{n_{Total}}$

This can also be expressed in terms of pressure. $\dfrac{P_A}{P_{Total}} = \dfrac{n_A}{n_{Total}}$

Each individual gas contributes to the total pressure and the extent of the contribution is proportional to the mole fraction (**X**) of that individual gas in the mixture.

$$P_A = X_A \cdot P_{Total} \qquad P_B = X_B \cdot P_{Total} \qquad P_C = X_C \cdot P_{Total}$$

or

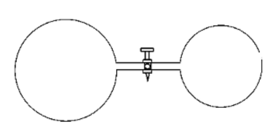
In general, $P_i = X_i \cdot P_{Total}$

Problems: Show all work as you solve these problems.

1. A sample of butane gas (C_4H_{10}) is collected over water at 50.0°C. The pressure in the lab is 98.8 kPa. The water vapor pressure at 50.0°C is 12.3 kPa.
   a. What is the partial pressure of the dry butane gas (in kPa)?

   b. What is the mole fraction of the butane gas?

2. Two flasks of different volumes are separately filled with CO_2 and N_2 in a chem lab at 22.0°C. The left flask is 2.5 L and contains 4.40 g CO_2. The right flask is filled with N_2, has a volume of 1.5 L and an internal pressure of 550 mm Hg. A closed valve prevents the gases from mixing. Once the valve is opened, the two gases diffuse to fill both containers.
   a. What is the partial pressure of the CO_2 gas in the final state?



b. What is the partial pressure of the N_2 gas in the final state?

c. What is the pressure in the two flasks after the valve was opened?

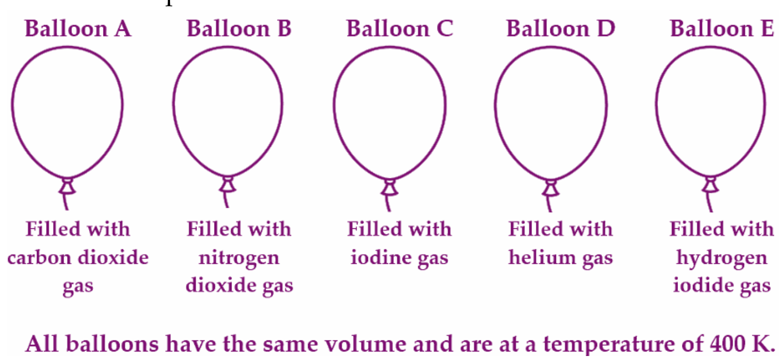
d. What is the mole fraction of the CO_2 gas and the N_2 gas?

Part 2: Graham's Law of Effusion

Scottish chemist Thomas Graham noted that less massive gases effuse faster than heavier ones. **Graham's Law of Effusion** says that the rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses. In comparing Gas A to Gas B, where **R** represents the **effusion rate,** and **MM** represents the **molar mass,** the equation is:

$$\frac{R_A}{R_B} = \sqrt{\frac{MM_B}{MM_A}}$$

Problems: Show all work as you solve these problems.



1. How much faster is the effusion rate of Balloon D compared to Balloon A?

2. Rank the balloons from lowest to highest effusion rates.

3. An additional "Balloon F" is observed to effuse at a rate that is 1.348 times faster than Balloon C. What is the molar mass of the gas in Balloon F? (Bonus: Which diatomic elemental gas is contained in Balloon F?)