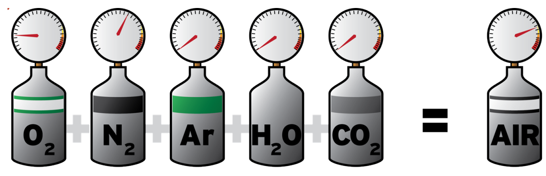
5.5 Dalton's Law of Partial Pressure



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Deadline for registration AND payment is Wednesday, 11/6!!

Dalton's Law of Partial Pressure

John Dalton's experiments with mixtures of gases contributed to his postulates on atomic theory.

He stated: "For a mixture of gases in a container, the total pressure exerted is the sum of the pressures that each gas would exert if it were alone."

From this, it is the number of moles of gas that matters, rather than the type of gas.

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

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

Diving Example

A mixture of helium and oxygen can be used in scuba diving to prevent "the bends." If 46 L of He at 25°C and 1.0 atm and 12 L of O_2 (same conditons) are pumped into a tank with a volume of 5.0 L:

1. What are the partial pressures of each gas?
2. What's the total pressure in the tank?

Diving Answer: Partial Pressures

1. What are the partial pressures of each gas?
Use Ideal Gas Law to find moles of each:

$$n_{He} = \frac{PV}{RT} = \frac{1.0\,atm \cdot 46\,L}{0.0821\frac{L \cdot atm}{K \cdot mol} \cdot 298\,K} = 1.9\,mol\,He$$

$$n_{O_2} = \frac{PV}{RT} = \frac{1.0\,atm \cdot 12\,L}{0.0821\frac{L \cdot atm}{K \cdot mol} \cdot 298\,K} = 0.49\,mol\,O_2$$

Diving Answer: Partial Pressures

Use the Ideal Gas Law again to find pressures of each (in the tank):

$$P_{He} = \frac{nRT}{V} = \frac{1.9\,mol \cdot 0.0821\frac{L \cdot atm}{K \cdot mol} \cdot 298\,K}{5.0\,L} = 9.3\,atm$$

$$P_{O_2} = \frac{nRT}{V} = \frac{0.49\,mol \cdot 0.0821\frac{L \cdot atm}{K \cdot mol} \cdot 298\,K}{5.0\,L} = 2.4\,atm$$

Finally, add up partial pressures:

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

$$= 9.3\,atm + 2.4\,atm = \boxed{11.7\,atm}$$

Mole Fractions

It is useful to make a ratio of the number of moles of a gas in a mixture vs. number of moles total.
This ratio is called a mole fraction, and is mathematically shown:

$$\chi_1 = \frac{n_1}{n_{Total}} = \frac{n_1}{n_1 + n_2 + n_3 + ...}$$

Mole Fraction Expression

The mole fraction of each component in a gas mixture is directly related to its partial pressure:

$$\chi_1 = \frac{n_1}{n_{Total}} = \frac{P_1}{P_{Total}}$$

Partial Pressure Expression

Mole Fraction Examples

2. The partial pressure of oxygen was observed to be 156 torr in air with a total atmospheric pressure of 743 torr. Calculate the mole fraction of oxygen.

$$\chi_1 = \frac{P_{O_2}}{P_{Total}} = \frac{156\,torr}{743\,torr} = \boxed{0.210}$$

Mole Fraction Examples

3. The mole fraction of nitrogen in the air is 0.7808. How many moles of nitrogen are in a 12.3 L sample of air at 952 torr, and 28.0ºC?
4. What mass is this?

First, find partial pressure of nitrogen:

$$\chi_{N_2} = \frac{P_{N_2}}{P_{Total}}$$

$$P_{N_2} = \chi_{N_2} \bullet P_{Total} = 0.7808 \bullet 952\,torr = 743.3\,torr$$

Mole Fraction Examples

Next, use the Ideal Gas Law to compute moles (after converting to atm and kelvins):

$$743.3\,torr \bullet \frac{1.00\,atm}{760\,torr} = 0.978\,atm \quad 28.0^\circ C + 273\,K = 301\,K$$

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{0.978\,atm \bullet 12.3\,L}{0.0821\frac{L \bullet atm}{K \bullet mol} \bullet 301\,K} = \boxed{0.487\,mol\,N_2}$$

Then mass: $0.487\,mol\,N_2 \bullet \frac{28.02\,g\,N_2}{1\,mol\,N_2} = \boxed{13.6\,g\,N_2}$

Partial Pressure of Water

Any time a gas is collected via water displacement, that gas will have vaporous water molecules in it. Test tube in beaker demo.

This happens because water molecules evaporate at the surface of the liquid, and enter the analyte gas.

One must account for water in this resulting mixture, realizing that water vapor pressure is directly related to the temperature of the system.

5. Vapor Pressure Problem

Potassium chlorate ($KClO_3$) is heated in a test tube, and oxygen is produced, and collected by water displacement.
The oxygen is stored at 22ºC at a total pressure of 754 torr, the volume of the gas collected is 0.650 L, and the vapor pressure of water at this temperature is 21 torr.
Calculate the partial pressure of the gas collected and the mass of $KClO_3$ in the sample that was decomposed.

5. Vapor Pressure Solution

First, determine the partial pressure of oxygen:

$$P_{Total} - P_{H_2O} = 754\,torr - 21\,torr = 733\,torr$$

Convert: $733\,torr \bullet \frac{1\,atm}{760\,torr} = 0.964\,atm$

Next determine how many moles of oxygen there are, using the Ideal Gas Law:

$$PV = nRT$$

$$n_{O_2} = \frac{PV}{RT} = \frac{0.964\,atm \bullet 0.650\,L}{0.0821\frac{L \bullet atm}{K \bullet mol} \bullet 295\,K} = \boxed{0.0259\,mol\,O_2}$$

5. Vapor Pressure Solution

Finally, a balanced equation followed by a mole ratio and stoichiometry delivers the answer:

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2$$

$$0.0259\,mol\,O_2 \bullet \frac{2\,mol\,KClO_3}{3\,mol\,O_2} \bullet \frac{122.55\,g\,KClO_3}{1\,mol\,KClO_3}$$

$$= \boxed{2.12\,g\,KClO_3}$$

Homework:

Read 5.6 to end of Chapter 5 in your book.

5.5 Booklet Problems
Due: Next Class