

Gases and Gas Laws

- Pressure is the force that a gas applies over the area of the container

- Units

- SI: (Pa)scal = $\text{kg m}^{-1} \text{s}^{-2}$
- (atm)osphere
- mmHg or torr (for Torricelli)
 - 1 atm = 760 mmHg (torr) = 101325 Pa

Pressure and its units

- Boyle's Law $P_1V_1 = P_2V_2$

- Charles' Law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

- Gay-Lussac's Law $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

- Avogadro's Law $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

- Combined Gas Law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

This leads to the molar volume which is 22.414 L at STP (1 atm and 0°C).

Individual Gas Laws

- All of the individual gas laws can be combined into one overall gas law.
- This is the *Ideal Gas Law*.

$$PV = nRT$$

- $R = 0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1}$
 $= 62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1}$
 $= 8.314 \text{ J mol}^{-1} \text{ K}^{-1} = 8.314 \text{ L kPa mol}^{-1} \text{ K}^{-1}$

- Remember that when using any gas law the temperature must be in Kelvin.

Gas Laws

- 15.03 g of Sulfur Hexafluoride gas are added to a 1.085 L container at 46.8°C. What is the pressure in the container in mmHg?

$$PV = nRT \quad P = \frac{nRT}{V}$$

$$= \frac{(15.03 \text{ g SF}_6) \left(\frac{1 \text{ mol SF}_6}{146.052 \text{ g SF}_6} \right) (62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1}) (320.0 \text{ K})}{1.085 \text{ L}}$$

$$= 1893 \text{ mmHg}$$

Example

- If we multiply both sides of the ideal gas law by mass and divide by volume and moles we get $MP = dRT$

- This is the relationship between the molar mass of a gas and its density.

Molar mass of gases

- The density of air at STP is 1.2 g L^{-1} . What is the average molar mass of air?

$$M = \frac{dRT}{P}$$

$$= \frac{(1.2 \text{ g L}^{-1})(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(273.15 \text{ K})}{1.000 \text{ atm}}$$

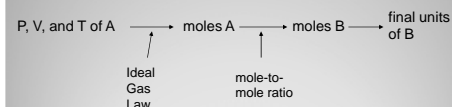
$$= 27 \text{ g mol}^{-1}$$

Example

- We still need to have a balanced chemical equation in order to complete the problem correctly.
- The difference here is in how are presented the data and how we calculate moles at the beginning.
- We will be given P, V and T for the gas and will use the ideal gas law to get moles.

Stoichiometry and gases

- P, V, and T gives us moles and we then use the mole-to-mole ratio to find the moles of product.



Stoichiometry

- Acetone vapor ($\text{C}_3\text{H}_6\text{O}$) burns in oxygen to give carbon dioxide gas and water. How many grams of water can be produced when 46.7 L of acetone vapor at 156.3°C and 0.9934 atm burns?

Example 1

- Balanced equation
 - $\text{C}_3\text{H}_6\text{O} + 4 \text{ O}_2 \rightarrow 3 \text{ CO}_2 + 3 \text{ H}_2\text{O}$
- Set up and calculate:

$$? \text{ g H}_2\text{O} = \frac{PV}{RT}$$

$$= \frac{(0.9934 \text{ atm})(46.7 \text{ L})}{(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(429.5 \text{ K})}$$

$$\times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol C}_3\text{H}_6\text{O}} \times \frac{18.0153 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 71.1 \text{ g H}_2\text{O}$$

Answer

- Phosphorus gas (P_4) reacts with oxygen gas to make diphosphorus pentoxide. What mass of diphosphorus pentoxide can be made from 35.4 L of Phosphorus at 775 mmHg and 301 K and 168 L of oxygen at 748 mmHg and 462 K?

Practice

- Balanced Equation:
 - $P_4 + 5 O_2 \rightarrow 2 P_2O_5$
- Setup and calculate:

$$? \text{ g } P_2O_5 = \frac{(35.4 \text{ L})(775 \text{ mmHg})}{(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(301 \text{ K})} \times \frac{2 \text{ mol } P_2O_5}{1 \text{ mol } P_4} \times \frac{141.9446 \text{ g } P_2O_5}{1 \text{ mol } P_2O_5} = 415 \text{ g } P_2O_5$$

$$? \text{ g } P_2O_5 = \frac{(168 \text{ L})(748 \text{ mmHg})}{(62.364 \text{ L mmHg mol}^{-1} \text{ K}^{-1})(462 \text{ K})} \times \frac{2 \text{ mol } P_2O_5}{5 \text{ mol } O_2} \times \frac{141.9446 \text{ g } P_2O_5}{1 \text{ mol } P_2O_5} = 248 \text{ g } P_2O_5$$

- 248 g of P_2O_5 can be made.

Solution

- When gases are mixed in a container they behave as if they are the only gas in the container.
- They will each have their own pressure called the *partial pressure*.
- All of the partial pressures will add up to the total pressure in the container.

$$P_T = P_1 + P_2 + \dots + P_n = \sum_{i=1}^n P_i$$

Gas mixtures

- Gas mixtures are often defined in terms of their mole fractions.

$$\chi_i = \frac{n_i}{n_T} = \frac{P_i}{P_T}$$

- Because the pressure is proportional to the number of moles we also get the above
- which can be rearranged to

$$P_i = P_T \chi_i$$

Gas mixtures

- 1.54 moles of Neon, 2.08 moles of Argon and 1.11 moles of Nitrogen gas are placed into a 10.00 L container at 43.3°C. Calculate the partial pressure of each gas and the total pressure.

Example

- Calculate the partial pressure of each gas.

$$P_{Ne} = \frac{nRT}{V} = \frac{(1.54 \text{ mol})(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(316.5 \text{ K})}{10.00 \text{ L}} = 4.00 \text{ atm}$$

$$P_{Ar} = \frac{nRT}{V} = \frac{(2.08 \text{ mol})(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(316.5 \text{ K})}{10.00 \text{ L}} = 5.40 \text{ atm}$$

$$P_{N_2} = \frac{nRT}{V} = \frac{(1.11 \text{ mol})(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(316.5 \text{ K})}{10.00 \text{ L}} = 2.88 \text{ atm}$$

Answer

- Calculate the total pressure

$$P_T = P_{Ne} + P_{Ar} + P_{N_2} = 4.00 \text{ atm} + 5.40 \text{ atm} + 2.88 \text{ atm} = 12.28 \text{ atm}$$

Answer continued

- Calculate the total pressure based on the number of moles

$$n_T = 1.54 \text{ mol} + 2.08 \text{ mol} + 1.11 \text{ mol} = 4.73 \text{ mol}$$

$$P_T = \frac{n_T RT}{V}$$

$$= \frac{(4.73 \text{ mol})(0.082058 \text{ L atm mol}^{-1} \text{ K}^{-1})(316.5 \text{ K})}{10.00 \text{ L}}$$

$$= 12.3 \text{ atm}$$

Alternate answer

- Find moles fractions and partial pressures

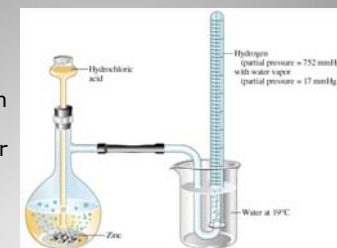
$$X_{\text{H}_2} = \frac{n_{\text{H}_2}}{n_T} = \frac{1.54 \text{ mol}}{4.73 \text{ mol}} = 0.3256 \quad P_{\text{H}_2} = X_{\text{H}_2} P_T = (0.3256)(12.3 \text{ atm}) = 4.00 \text{ atm}$$

$$X_{\text{Ar}} = \frac{n_{\text{Ar}}}{n_T} = \frac{2.08 \text{ mol}}{4.73 \text{ mol}} = 0.4397 \quad P_{\text{Ar}} = X_{\text{Ar}} P_T = (0.4397)(12.3 \text{ atm}) = 5.41 \text{ atm}$$

$$X_{\text{N}_2} = \frac{n_{\text{N}_2}}{n_T} = \frac{1.11 \text{ mol}}{4.73 \text{ mol}} = 0.2347 \quad P_{\text{N}_2} = X_{\text{N}_2} P_T = (0.2347)(12.3 \text{ atm}) = 2.89 \text{ atm}$$

Alternate answer continued

- Partial pressures can be used when collecting gases over water.



Gas mixtures

- The gas in collected in the tube is a mixture of the product gas and water vapor. $P_T = P_{\text{product}} + P_{\text{H}_2\text{O}(g)}$

- The pressure of water vapor is a tabulated number called the vapor pressure of water. (Table 5.6 in your text)

Collecting gases over water

- Volume of particles is negligible
- Particles are in constant motion
- No inherent attractive or repulsive forces
- The average kinetic energy of a collection of particles is proportional to the temperature (K)

Kinetic Molecular Theory

- The root-mean-square (rms) molecular speed, u ,** is an average molecular speed, equal to the speed of a molecule having the average molecular kinetic energy. It is given by the following formula:

$$u = \sqrt{\frac{3RT}{M_m}}$$

In this formula R is in $\text{J mol}^{-1} \text{K}^{-1}$
the temperature is in K and the molar mass is in kg mol^{-1} .

Molecular speeds

- What is the rms speed of hydrogen molecules at standard temperature?

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{\frac{3(8.314 \text{ J mol}^{-1} \text{ K}^{-1})(273.15 \text{ K})}{(2.016 \times 10^{-3} \text{ kg mol}^{-1})}} = 1838 \text{ m s}^{-1}$$

Example

- **Diffusion** is the movement of a gas through space or another gas over time.
- **Effusion** is the movement of a gas through a small opening into a vacuum.
- The equation for the rms velocity of gases shows the following relationship between rate of effusion and molecular mass

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{M_m}}$$

Diffusion and Effusion

- **Graham's law** relates the rate of effusion or diffusion to the molecular mass.

$$\frac{\text{Rate of effusion of gas "A"}}{\text{Rate of effusion of gas "B"}} = \sqrt{\frac{M_m \text{ of Gas B}}{M_m \text{ of gas A}}}$$

Graham's Law

- A particular gas effuses 3.67 times slower than helium. What is the molar mass of the gas?

$$\frac{R_{He}}{R_U} = \sqrt{\frac{M_U}{M_{He}}}$$

$$R_{He} = 3.67 R_U$$

$$M_U = (3.67)^2 M_{He} = (3.67)^2 (4.0026 \text{ g mol}^{-1}) = 53.9 \text{ g mol}^{-1}$$

Example