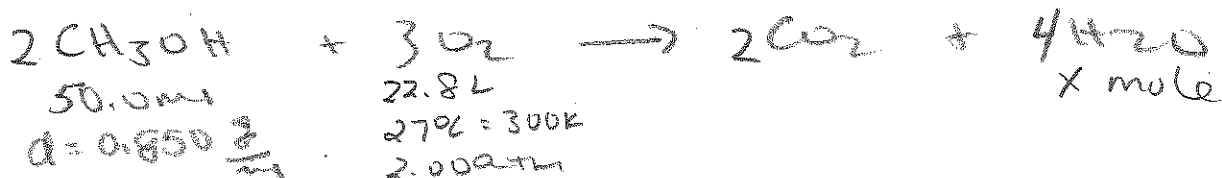


5. Consider the combustion reaction between 50.0 ml of liquid methanol, CH_3OH (density = 0.850 g/ml) and 22.8 L of O_2 at 27°C and a pressure of 2.00 atm. Write the balanced equation for the combustion reaction. Calculate the number of moles of H_2O formed if this reaction goes to completion.



$$\frac{\text{CH}_3\text{OH}}{(0.850 \frac{\text{g}}{\text{ml}})(50.0 \text{ ml})} = 42.5 \text{ g} \frac{1 \text{ mole}}{32.03 \text{ g}} = 1.33 \text{ mole CH}_3\text{OH} = 0.665$$

$$\text{O}_2 \quad PV = nRT \quad (2.00 \text{ atm})(22.8 \text{ L}) = x (0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}})(300 \text{ K})$$

$$1.85 \text{ mole O}_2 = 0.617 \times \text{limit}$$

$$1.85 \text{ mole O}_2 \frac{4 \text{ mole H}_2\text{O}}{3 \text{ mole O}_2} = \boxed{2.47 \text{ mole H}_2\text{O}}$$

6. A gas consisting of only carbon and hydrogen has an empirical formula of CH_2 . The gas has a density of 1.65 g/L at 27°C and 734 torr. Determine the molar mass and molecular formula of the gas.

$$MM = \frac{dRT}{P} = \frac{(1.65 \frac{\text{g}}{\text{L}})(62.4 \frac{\text{torr} \cdot \text{L}}{\text{mol} \cdot \text{K}})(300 \text{ K})}{734 \text{ torr}}$$

$$MM = 42.1 \frac{\text{g}}{\text{mol}}$$

$$\text{CH}_2 = 14.03 \frac{\text{g}}{\text{mol}}$$

$$\frac{42.1}{14.03} = 3$$



7. A compound has the empirical formula CHCl . A 250 mL flask, at 373 K and 750 torr, contains 0.800 g of the gaseous compound. Give the molecular formula.

$$PV = nRT$$

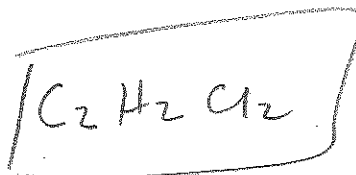
$$(750 \text{ torr})(0.250 \text{ L}) = x (62.4 \frac{\text{torr} \cdot \text{L}}{\text{mol} \cdot \text{K}})(373 \text{ K})$$

$$x = 8.06 \times 10^{-3} \text{ mole}$$

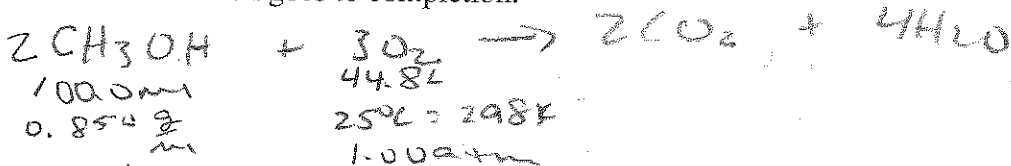
$$MM = \frac{0.800 \text{ g}}{8.06 \times 10^{-3} \text{ mol}} = 99.3 \frac{\text{g}}{\text{mol}}$$

$$\text{CHCl} = 48.47 \frac{\text{g}}{\text{mol}}$$

$$\frac{99.3}{48.47} = 2$$



1. Consider the reaction between 100.0 mL of liquid methyl alcohol (CH₃OH) (density = 0.850 g/mL) and 44.8 L of O₂ at 25°C and a pressure of 1.00 atm. The products of the reaction are CO_{2(g)} and H₂O(g). Calculate the number of moles of H₂O formed if the reaction goes to completion.



CH₃OH $(100.0 \text{ mL}) (0.850 \frac{\text{g}}{\text{mL}}) = 85.0 \text{ g} \frac{1 \text{ mole}}{32.05 \text{ g}} = 2.65 \text{ mole CH}_3\text{OH} = 1.33$

O₂ $PV = nRT$
 $(1.00 \text{ atm})(44.8 \text{ L}) = x (0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}})(298 \text{ K})$
 $x = 1.83 \text{ mole O}_2 = 0.61^*$

2. The partial pressure of CH₄ gas is 0.175 atm and that of O₂ gas is 0.250 atm in a mixture of two gases.

- a) what is the mole fraction of each gas in the mixture

$$P_T = 0.175 \text{ atm} + 0.250 \text{ atm} = 0.425 \text{ atm}$$

$$X_{\text{CH}_4} = \frac{0.175}{0.425} = 0.412$$

$$X_{\text{O}_2} = \frac{0.250}{0.425} = 0.588$$

- b) if the mixture occupies a volume of 10.5 L at 65°C calculate the total number of moles of gas in the mixture.

$$PV = nRT$$

$$(0.425 \text{ atm})(10.5 \text{ L}) = x (0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}})(338 \text{ K})$$

$$x = 0.161 \text{ mole total}$$

- c) calculate the number of grams of each gas in the mixture

$$0.161 \text{ mole} (0.412) = 0.0663 \text{ mole CH}_4 \frac{16.05 \text{ g}}{1 \text{ mole}}$$

$$0.161 \text{ mole} (0.588) = 0.0947 \text{ mole O}_2 \frac{32.00 \text{ g}}{1 \text{ mole}} = 3.02 \text{ g O}_2$$

$$\boxed{3.02 \text{ g O}_2}$$

3. A mixture of 1.00 g H₂ and 1.00 g He is placed in a 1.00 L container at 27°C. Calculate the partial pressure of each gas and the total pressure.

$$1.00 \text{ g H}_2 \times \frac{1 \text{ mole H}_2}{2.02 \text{ g H}_2} = 0.495 \text{ mole H}_2$$

$$PV = nRT$$

$$(x)(1.00 \text{ L}) = (0.495 \text{ mole})(0.0821 \frac{\text{atm L}}{\text{mole K}})(300 \text{ K})$$

$$x = 12.2 \text{ atm} = P_{\text{H}_2}$$

$$1.00 \text{ g He} \times \frac{1 \text{ mole He}}{4.00 \text{ g}} = 0.250 \text{ mole He}$$

$$PV = nRT$$

$$(x)(1.00 \text{ L}) = (0.250 \text{ mole})(0.0821 \frac{\text{atm L}}{\text{mole K}})(300 \text{ K})$$

$$x = 6.16 \text{ atm} = P_{\text{He}}$$

$$P_T = 18.36 \text{ atm}$$

4. At 0°C a 1.0 L flask contains 5.0 x 10⁻² mol N₂, 1.5 x 10⁻² mol O₂, and 5.0 x 10²¹ molecules NH₃. What is the partial pressure of each gas, and what is the total pressure in the flask?

$$5.0 \times 10^{-2} \text{ mol N}_2$$

$$1.5 \times 10^{22} \text{ molecules O}_2 \times \frac{1 \text{ mole O}_2}{6.02 \times 10^{23} \text{ molecules O}_2} = 0.00469 \text{ mole O}_2$$

$$5.0 \times 10^{21} \text{ molecules NH}_3 \times \frac{1 \text{ mole NH}_3}{6.02 \times 10^{23} \text{ molecules NH}_3} = 0.00831 \text{ mole NH}_3$$

$$\frac{\text{N}_2}{PV = nRT}$$

$$x(1.0 \text{ L}) = (5.0 \times 10^{-2} \text{ mole})(0.0821 \frac{\text{atm L}}{\text{mole K}})(273 \text{ K})$$

$$x = 1.12 \text{ atm}$$

$$\frac{\text{O}_2}{PV = nRT}$$

$$x(1.0 \text{ L}) = (0.00469 \text{ mole})(0.0821 \frac{\text{atm L}}{\text{mole K}})(273 \text{ K})$$

$$x = 0.105 \text{ atm}$$

$$\frac{\text{NH}_3}{PV = nRT}$$

$$x(1.0 \text{ L}) = (0.00831 \text{ mole})(0.0821 \frac{\text{atm L}}{\text{mole K}})(273 \text{ K})$$

$$x = 0.186 \text{ atm}$$

5. At elevated temperatures, sodium chlorate decomposes to produce sodium chloride and oxygen gas. A 0.8765 g sample of impure sodium chlorate was heated until the production of oxygen ceased. The oxygen gas collected over water occupied 57.2 mL at a temperature of 22°C and a pressure of 734 torr. Calculate the mass percent of NaClO₃ in the original sample. (At 22°C the vapor pressure of water is 19.8 torr)



$$57.3 \text{ mL} = 0.0573 \text{ L}$$

$$22^\circ \text{C} = 295 \text{ K}$$

$$734 \text{ torr} - 19.8 \text{ torr} = 714.2 \text{ torr}$$

$$PV = nRT$$

$$(714.2 \text{ torr}) (0.0573 \text{ L}) = x \left(62.4 \frac{\text{torr L}}{\text{mole K}} \right) (295 \text{ K})$$

$$x = 0.00222 \text{ mole NaClO}_3 \times \frac{106.44 \text{ g}}{1 \text{ mole}} = 0.2363 \text{ g}$$

$$\frac{0.2363}{0.8765} \times 100 = 26.96 \%$$