Name ________________________________ Date ________________

GAS DENSITY, REACTION STOICHIOMETRY and MOLAR MASS

1. What is the density of helium gas at STP?

$d = \frac{m}{V} = \frac{4.00 \frac{g}{mol}}{22.4 \frac{L}{mol}} = 0.178 \frac{g}{L}$

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

$d = \frac{MM \cdot P}{RT} = \frac{(4.00 \frac{g}{mol})(1.00 atm)}{(0.0821 \frac{atm \cdot L}{mol \cdot K})(273 K)} = \boxed{0.178 \frac{g}{L}}$

2. What mass of nitrogen gas is required to fill a 1.5 L balloon at STP?

$PV = nRT$

$(1.00 atm)(1.5 L) = x (0.0821 \frac{atm \cdot L}{mol \cdot K})(273 K)$

$x = 0.0669$ mole N_2 $\frac{28.02 g}{1 mole N_2} = \boxed{1.87 g N_2}$

3. A student adds 4.00 g of dry ice (solid CO_2) to an empty balloon. What will be the volume of the balloon at STP after all the dry ice sublimes?

$4.00 g CO_2 \frac{1 mole CO_2}{44.01 g CO_2} \frac{22.4 L CO_2}{1 mole CO_2} = \boxed{2.04 L CO_2}$

4. Oxygen gas can be obtained by the thermal decomposition of mercury II oxide into its elements. Write the balanced equation for the reaction. What volume of oxygen gas, measured at 30.°C and 725 torr, can be produced from the complete decomposition of 4.10 g of mercury II oxide?

$2HgO \rightarrow 2Hg + O_2$

4.10g

x L

30.0C = 300 K

725 torr

$4.10 g HgO \frac{1 mole HgO}{216.59 g HgO} \frac{1 mole O_2}{2 mole HgO} = 0.00947 mole O_2$

$PV = nRT$

$(725 torr)(x L) = (0.00947 mole)(62.4 \frac{torr \cdot L}{mol \cdot K})(303 K)$

$\boxed{x = 0.247 L}$

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.

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

CH_3OH: 50.0 ml, d = 0.850 g/ml

O_2: 22.8 L, 27°C = 300K, 2.00 atm

H_2O: x mole

CH_3OH: $(0.850 \frac{g}{ml})(50.0ml) = 42.5g \cdot \frac{1 mole}{32.05g} = 1.33$ mole $\frac{CH_3OH}{2} = 0.665$

O_2: $PV = nRT$ $\quad (2.00atm)(22.8L) = x(0.0821 \frac{atm \cdot L}{mol \cdot K})(300K)$

$\frac{1.85 \text{ mole } O_2}{3} = 0.617$ limits

1.85 mole $O_2 \cdot \frac{4 \text{ mole } H_2O}{3 \text{ mole } O_2} = \boxed{2.47 \text{ mole } H_2O}$

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.

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

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

$CH_2 = 14.03 \frac{g}{mol}$

$\frac{42.1}{14.03} = 3$ $\quad \boxed{C_3H_6}$

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.

$PV = nRT$

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

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

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

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

$\frac{99.3}{48.47} = 2$ $\quad \boxed{C_2H_2Cl_2}$

1. Consider the reaction between 100.0 mL of liquid methyl alcohol (CH_3OH) (density = 0.850 g/mL) and 44.8 L of O_2 at 25°C and a pressure of 1.00 atm. The products of the reaction are $CO_{2(g)}$ and $H_2O_{(g)}$. Calculate the number of moles of H_2O formed if the reaction goes to completion.

$2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$

100.0 ml, 0.850 g/ml — 44.8 L, 25°C = 298 K, 1.00 atm

CH_3OH: $(100.0\ ml)(0.850\ g/ml) = 85.0\ g \cdot \frac{1\ mole}{32.05\ g} = \frac{2.65\ mole\ CH_3OH}{2} = 1.33$

O_2: $PV = nRT$

$(1.00\ atm)(44.8\ L) = x\,(0.0821\ \frac{atm \cdot L}{mol \cdot K})(298\ K)$

$x = \frac{1.83\ mole\ O_2}{3} = 0.61$*

$1.83\ mol\ O_2 \cdot \frac{4\ mole\ H_2O}{3\ mole\ O_2} = \boxed{2.44\ mole\ H_2O}$

2. The partial pressure of CH_4 gas is 0.175 atm and that of O_2 gas is 0.250 atm in a mixture of two gases.

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

$P_T = 0.175\ atm + 0.250\ atm = 0.425\ atm$

$X_{CH_4} = \frac{0.175}{0.425} = 0.412$

$X_{O_2} = \frac{0.250}{0.425} = 0.588$

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.

$PV = nRT$

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

$x = 0.161$ mole total

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

$0.161\ mole\,(0.412) = 0.0663\ mol\ CH_4 \cdot \frac{16.05\ g}{1\ mole} = \boxed{1.06\ g\ CH_4}$

$0.161\ mole\,(0.588) = 0.0947\ mole\ O_2 \cdot \frac{32.00\ g}{1\ mole} = \boxed{3.02\ g\ O_2}$

3. A mixture of 1.00 g H_2 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.

atm

$1.00 g\ H_2 \frac{1 mole\ H_2}{2.02 g\ H_2} = 0.495$ mole H_2

$PV = nRT$

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

$\boxed{X = 12.2 atm = P_{H_2}}$

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

$PV = nRT$

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

$\boxed{X = 6.16 atm = P_{He}}$

$\boxed{P_T = 18.36 atm}$

4. At 0°C a 1.0 L flask contains 5.0×10^{-2} mol N_2, 1.5×10^2 mg O_2, and 5.0×10^{21} molecules NH_3. What is the partial pressure of each gas, and what is the total pressure in the flask?

5.0×10^{-2} mol N_2

$1.5 \times 10^{2}\ mg \frac{1 g}{1000 mg} \frac{1 mole\ O_2}{32.00 g\ O_2} = 0.00469$ mole O_2

5.0×10^{21} molecules NH_3 $\frac{1 mole\ NH_3}{6.02 \times 10^{23} molecules\ NH_3} = 0.00831$ mole NH_3

N_2

$PV = nRT$

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

$\boxed{X = 1.12 atm}$

O_2

$PV = nRT$

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

$X = \boxed{0.105 atm}$

NH_3 $PV = nRT$

$x(1.0 L) = (0.00831 mol)(0.0821)(273 K)$

$= 0.186$ atm

$\boxed{P_T = 1.41 atm}$

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 $NaClO_3$ in the original sample. (At 22°C the vapor pressure of water is 19.8 torr)

$2 NaClO_3 \rightarrow 2 NaCl + 3 O_2$

57.3 mL = 0.0573 L

22°C = 295 K

734 torr − 19.8 torr = 714.2 torr

$PV = nRT$

$(714.2 torr)(0.0573 L) = x (62.4 \frac{torr \cdot L}{mol K})(295 K)$

$X = 0.00222$ mole $NaClO_3$ $\frac{106.44 g}{1 mole} = 0.2363 g$

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