MUST SHOW ALL WORK FOR CREDIT

1.) What volume of hydrogen gas at 580. mm Hg and 127°C is required to produce 0.895 g of NH_3? (R = 0.0821 L atm/ mol K)

1 N_2 (g) + 3 H_2 (g) → 2 NH_3 (g)

Step 1: Covert grams of NH_3 to Moles of H_2

$$n = \frac{0.895\ g\ NH_3}{} \left| \frac{1\ mole\ NH_3}{17\ grams\ NH_3} \right| \frac{3\ Mole\ H_2}{2\ mole\ NH_3} \right| = 0.079\ mole\ H_2$$

Step 2: Look at the Gas Law constant to determine units.

R = 0.0821 L atm/ mol K

Step 3: Convert Temperature and Pressure to match the units of the Gas Law constant.

T_{H2} = 127°C + 273 = **400 Kelvin**

$$\frac{580\ mm\ Hg}{760\ mm\ Hg} = \frac{P}{1\ atm} \qquad P = 0.763\ atm$$

Step 4: Now rearrange the Ideal Gas Law equation to solve for Volume and substitute in the values.

$$V = \frac{nRT}{P} = \frac{(0.079\ moles)\left(0.0821\ \frac{atm\ L}{mol\ K}\right)(400\ K)}{(0.763\ atm)} = \boxed{3.40\ L\ H_2}$$

2.) What volume of N_2 would be required to produce the NH_3 under the ***same conditions*** of temperature and pressure? (R = 0.0821 L atm/ mol)

1 N_2 (g) + 3 H_2 (g) → 2 NH_3 (g)

Step 1: Covert grams of NH_3 to Moles of N_2. (Note everything stays the *same* except for moles N_2.)

$$n = \frac{0.895\ g\ NH_3}{} \left| \frac{1\ mole\ NH_3}{17\ grams\ NH_3} \right| \frac{1\ Mole\ N_2}{2\ mole\ NH_3} \right| = 0.026\ mole\ N_2$$

Step 2: Look at the Gas Law constant to determine units.

R = 0.0821 atm L / mol K

Step 3: Convert Temperature and Pressure to match the units of the Gas Law constant.

T_{H2} = 127°C + 273 = **400 Kelvin**

$$\frac{580\ mm\ Hg}{760\ mm\ Hg} = \frac{P}{1\ atm} \qquad P = 0.763\ atm$$

Step 4: Now rearrange the Ideal Gas Law equation to solve for Volume and substitute in the values.

$$V = \frac{nRT}{P} = \frac{(0.026\ moles)\left(0.0821\ \frac{atm\ L}{mol\ K}\right)(400\ K)}{(0.763\ atm)} = \boxed{1.12\ L\ N_2}$$

3.) What is the molar mass of a gas if 1.45 g occupies 830. mL at 735 mm Hg and 22.0 °C? (R = 0.0821 L atm/ mol K)

Step 1: Since R has units of atm, L, and K. You must convert Pressure, Volume, and Temperature.

$\frac{735\ mm\ Hg}{760\ mm\ Hg} = \frac{Pressure}{1\ atm}$ $\quad$ $\frac{830\ ml}{} \left| \frac{1\ Liter}{1000\ ml} \right| = Volume$ $\quad$ $22.0\ °C + 273 = \text{Temperature}$

P = 0.967 atm $\quad$ **V = 0.830 L** $\quad$ **T = 295 K**

Step 2: Since molar mass = grams/ mole, rearrange the Ideal Gas Law to solve for moles.

$$PV = nRT \qquad n = \frac{PV}{RT} = \frac{(0.967\ atm)(0.830\ L)}{\left(0.0821\ \frac{atm\ L}{mol\ K}\right)(295\ K)} = 0.0331\ moles$$

Step 3: Now take the number of grams from the word problem and divide by the number of moles to calculate the grams/mole for Molar Mass.

$$\frac{1.45\ g}{0.0331\ moles} = 44.0\ g/mol$$

4.) What volume of hydrogen gas at STP is required to produce 85.0 g of NH_3? (R = 8.314 kPa L/ mol K)

N_2 (g) + 3 H_2 (g) → 2 NH_3 (g)

Step 1: Covert grams of NH_3 to Moles of H_2. (Note everything stays the *__same__* except for moles N_2.)

$$n = \frac{85\ g\ NH_3}{} \left| \frac{1\ mole\ NH_3}{17\ grams\ NH_3} \right| \frac{3\ Mole\ H_2}{2\ mole\ NH_3} \right| = 7.5\ mole\ H_2$$

Step 2: Look at the Gas Law constant to determine units.

R = 8.314 kPa L / mol K

Step 3: Convert Temperature and Pressure to match the units of the Gas Law constant.

Temperature @ STP = 273 Kelvin

Pressure @ STP = 101.325 kPa

Step 4: Now rearrange the Ideal Gas Law equation to solve for Volume and substitute in the values.

$$V = \frac{nRT}{P} = \frac{(7.5\ moles)\left(8.314\ \frac{kPa\ L}{mol\ K}\right)(273\ K)}{(101.325\ kPa)} = \boxed{168\ L\ H_2}$$

Scroll down to next page to continue

5.) A sample of nitrogen gas, N_2, is collected in a 100 mL container at a pressure of 688 mm Hg and a temperature of 565 °C. How many grams of nitrogen gas are present in this sample? (R = 8.314 kPa L / mol K)

Step 1: Look at the Gas Law constant to determine units.

R = 8.314 kPa L / mol K

Step 2: Convert the Volume, Temperature, and Pressure to match the units of the Gas Law constant.

$$V = \frac{100\ ml}{} \left| \frac{1\ Liter}{1000\ ml} \right| = 0.100\ L$$

T = 565 °C + 273 = **838 Kelvin**

$$\frac{688\ mm\ Hg}{760\ mm\ Hg} = \frac{P}{101.325\ kPa} \qquad P = 91.7\ kPa$$

Step 3: Solve Ideal Gas Law for number of moles.

$$PV = nRT \qquad n = \frac{PV}{RT} = \frac{(91.7\ kPa)(0.100\ L)}{\left(8.314\ \frac{kPa\ L}{mol\ K}\right)(838\ K)} = 0.00132\ moles$$

Step 3: Use the molar mass of nitrogen gas to covert to grams.

$$\frac{0.00132\ moles\ N_2}{} \left| \frac{28\ grams\ N_2}{1\ mole\ N_2} \right| = \boxed{0.0370\ grams\ N_2}$$

6.) Determine the molar mass of a gas that has a density of 2.18 g/L at 66°C and 720 mm Hg. (R = 8.314 kPa L / mol K)

Step 1: Look at the Gas Law constant to determine units.

R = 8.314 kPa L / mol K

Step 2: Convert the Volume, Temperature, and Pressure to match the units of the Gas Law constant.

V = *<u>This is a little sneaky</u>*. Since the density is given in g/L, you can assume V = 1 Liter.

T = 66 °C + 273 = **339 Kelvin**

$$\frac{720\ mm\ Hg}{760\ mm\ Hg} = \frac{P}{101.325\ kPa} \qquad P = 95.99\ kPa$$

Step 3: Solve Ideal Gas Law for number of moles.

$$PV = nRT \qquad n = \frac{PV}{RT} = \frac{(95.99\ kPa)(1\ L)}{\left(8.314\ \frac{kPa\ L}{mol\ K}\right)(339\ K)} = 0.0341\ moles$$

Step 3: Use the grams from the density and the moles from above to solve for molar mass.

$$\text{Molar Mass} = \frac{2.18\ g}{0.0341\ moles\ N_2} = \boxed{63.9\ g/mol}$$

Challenge Problem: What is the density of chlorine gas at STP?

Prior Knowledge: **Density = g/ml** **Molar mass = g/mol** **PV = nRT**

Step 1: Use grams of chlorine gas from molar mass.

Chlorine gas is diatomic, so 35.453 x 2 = 70.906 g/mol Cl_2.

Remember this means 70.906 grams = 1 mole Cl_2.

Step 2: If you are lucky, you remembered that 1 mole of gas = 22.4 L at STP. Assuming you were not so lucky, rearrange PV=nRT to solve for Volume. (Since they didn't give you an R value, chose whichever one you like. (R = 0.0821 atm L/ mol K or R = 8.314 kPa L/ mol K) Just make sure your units match.

If R = 8.314 kPa L/ mol K @ STP , then

Temperature = 273 K and Pressure = 101.325 kPa

$$V = \frac{nRT}{P} = \frac{(1\ mole\ Cl_2)\left(8.314\ \frac{kPa\ L}{mol\ K}\right)(273\ K)}{(101.325\ kPa)} = 22.4\ L\ Cl_2$$

Step 2: Density = mass/ volume

$$D = \frac{mass}{volume} = \frac{70.906\ grams}{22.4\ L} = 3.165\ g/L$$

Or if you want to write it as grams/milliliters, then

$$V = \frac{22.4\ L}{} \left|\frac{1000\ ml}{1\ L}\right| = 22,400\ ml$$

$$D = \frac{mass}{volume} = \frac{70.906\ grams}{22,400\ L} = 3165\ g/ml$$