

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 NH₃**?
(R = 0.0821 L atm/ mol K)



Step 1: Covert grams of NH₃ to Moles of H₂

$$n = \frac{0.895 \text{ g NH}_3}{17 \text{ grams NH}_3} \left| \frac{1 \text{ mole NH}_3}{2 \text{ mole NH}_3} \right| \left| \frac{3 \text{ Mole H}_2}{2 \text{ mole NH}_3} \right| = 0.079 \text{ mole H}_2$$

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

$$R = 0.0821 \text{ L atm/ mol K}$$

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

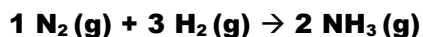
$$T_{\text{H}_2} = 127^\circ\text{C} + 273 = 400 \text{ Kelvin}$$

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

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

$$V = \frac{nRT}{P} = \frac{(0.079 \text{ moles}) \left(0.0821 \frac{\text{atm L}}{\text{mol K}}\right) (400 \text{ K})}{(0.763 \text{ atm})} = 3.40 \text{ L H}_2$$

- 2.) What **volume of N₂** would be required to produce the NH₃ under the same conditions of temperature and pressure? (R = 0.0821 L atm/ mol)



Step 1: Covert grams of NH₃ to Moles of N₂. (Note everything stays the same except for moles N₂.)

$$n = \frac{0.895 \text{ g NH}_3}{17 \text{ grams NH}_3} \left| \frac{1 \text{ mole NH}_3}{2 \text{ mole NH}_3} \right| \left| \frac{1 \text{ Mole N}_2}{2 \text{ mole NH}_3} \right| = 0.026 \text{ mole N}_2$$

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

$$R = 0.0821 \text{ atm L / mol K}$$

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

$$T_{\text{H}_2} = 127^\circ\text{C} + 273 = 400 \text{ Kelvin}$$

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

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

$$V = \frac{nRT}{P} = \frac{(0.026 \text{ moles}) \left(0.0821 \frac{\text{atm L}}{\text{mol K}}\right) (400 \text{ K})}{(0.763 \text{ atm})} = 1.12 \text{ L N}_2$$

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

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

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

$$P = 0.967 \text{ atm} \quad V = 0.830 \text{ L} \quad T = 295 \text{ K}$$

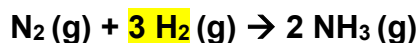
Step 2: Since molar mass = grams/ mole, rearrange the Ideal Gas Law to solve for moles.

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

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.

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

- 4.) What **volume** of hydrogen gas at **STP** is required to produce **85.0 g** of NH_3 ? ($R = 8.314 \text{ kPa L/ mol K}$)



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

$$n = \frac{85 \text{ g NH}_3}{17 \text{ grams NH}_3} \left| \frac{1 \text{ mole NH}_3}{17 \text{ grams NH}_3} \right| \left| \frac{3 \text{ Mole H}_2}{2 \text{ mole NH}_3} \right| = 7.5 \text{ mole H}_2$$

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

$$R = 8.314 \text{ kPa L / mol K}$$

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

$$\text{Temperature @ STP} = 273 \text{ Kelvin}$$

$$\text{Pressure @ STP} = 101.325 \text{ kPa}$$

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

$$V = \frac{nRT}{P} = \frac{(7.5 \text{ moles}) \left(8.314 \frac{\text{kPa L}}{\text{mol K}} \right) (273 \text{ K})}{(101.325 \text{ kPa})} = 168 \text{ L H}_2$$

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- 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? ($R = 8.314 \text{ kPa L / mol K}$)

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

$$R = 8.314 \text{ kPa L / mol K}$$

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

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

$$T = 565 \text{ °C} + 273 = 838 \text{ Kelvin}$$

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

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

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

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

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

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

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

$$R = 8.314 \text{ kPa L / mol K}$$

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

$V = \text{This is a little sneaky.}$ Since the density is given in g/L, you can assume $V = 1 \text{ Liter}$.

$$T = 66 \text{ °C} + 273 = 339 \text{ Kelvin}$$

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

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

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

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

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

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 $35.453 \times 2 = 70.906 \text{ g/mol Cl}_2$.

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. ($R = 0.0821 \text{ atm L/ mol K}$ or $R = 8.314 \text{ kPa L/ mol K}$) Just make sure your units match.

If $R = 8.314 \text{ kPa L/ mol K}$ @ STP, then

Temperature = 273 K and Pressure = 101.325 kPa

$$V = \frac{nRT}{P} = \frac{(1 \text{ mole Cl}_2) \left(8.314 \frac{\text{kPa L}}{\text{mol K}} \right) (273 \text{ K})}{(101.325 \text{ kPa})} = 22.4 \text{ L Cl}_2$$

Step 2: Density = mass/ volume

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

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

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

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