Last time we went through a series of exercises that explored relationships between variables that affect pressure. Recalling that Pressure is related to the number & force of collisions, three variables are important:

PV = constant	P = constant (T)	P = constant (n)
Boyle's Law	Charles Law	Avogadro's Law
$P_1V_1 = P_2V_2$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Combined: $PV = nRT$

Ideal Gas Law

R can take on lots of forms depending on which units are used

$R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$

$R = 8.314 \frac{J}{mol \cdot K}$ ($P = Pa$ $V = m^3$)

($Pa = \frac{J}{m^3}$)

We can also use the Ideal Gas Law to learn about properties of Ideal Gases:

Determine the molarity of Carbon dioxide at 2 atm and 0°C.

- First, get rid of °C → work w/ less units.

$T = 0°C + 273.15K$

$T = 273.15K$

- now rearrange the ideal gas law so that we have the units of concentration:

$PV = nRT \rightarrow \frac{n}{V} = \frac{P}{RT}$

$\frac{n}{V} = \frac{2\ atm}{0.08206 \frac{L \cdot atm}{mol \cdot K} (273.15\ K)}$

$\frac{n}{V} = 0.0892 \frac{mol}{L} = 0.0892\ M$

What if we wanted to determine the density (g/L) of CO_2 under the same conditions?

$\frac{mol}{L} \left| \frac{MW\ g}{mol} \right. = \frac{g}{L}$

Taking $\frac{mol}{L} \times MW$ gives us g/L

$\frac{n}{V} \times MW = \frac{g}{L} = \frac{P \cdot MW}{RT}$

$density = \frac{2\ atm\ (44.01\ g/mol)}{0.08206 \frac{L \cdot atm}{mol \cdot K} \cdot 273.1\ K} = 3.93 \frac{g}{L}$

* note here that Molarity of an ideal gas is completely independent of the identity of a gas. This is because we treat all ideal gases the same... all we care is that it is a gas. Density, however, is a property that IS dependent on what the gas is

If I mix NO_2 (yellow gas) together with Xe (colorless gas) in equal proportions, the Molarity will be identical, but the two gases would separate in the flask based on their density

← NO_2 MW = 46.01 g/mol

← Xe MW = 131.3 g/mol

This idea of gas Ideality is important when we consider mixtures of gases.

If two gases are mixed together, they don't react or interact (unless they are reactive toward each other). So, gases within a mixture of non-reactive gases can be treated independently.

If 1 mol of Ar and 3 mol of Kr are combined in a 1.8 L flask at 298 K, determine:

① the total pressure in the flask
② the partial pressure of Ar
③ the partial pressure of Kr

① $P = \frac{nRT}{V} = \frac{4(0.08206)(298)}{1.8} = 54.32$ atm (4 = 3+1 mol)

② $P = \frac{1(0.08206)(298)}{1.8} = 13.56$ atm

③ $P = \frac{3(0.08206)(298)}{1.8} = 40.76$ atm

using the $\frac{P_x}{P_{total}}$: ② $\frac{P_{Ar}}{54.32} = \frac{1}{4}$

$P_{Ar} = 13.56$ ✓ yup, it works

Things to note

① $P_{Kr} + P_{Ar} = P_{total} \Rightarrow$ Partial pressures add up to the total pressure

② The only thing changing in the equation is moles

P = constant (n)

but we knew this already

$$\frac{P_x}{P_{total}} = \frac{n_x}{n_{total}}$$

What about when gases do react?

3 L of H_2 + 1 L $N_2 \rightarrow$ 2 L NH_3

$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

So If I start with 17 L of H_2 and 5 L of N_2 @ STP (1 atm, 273.15 K), what volume of NH_3 will be made? What volume of H_2 & N_2 are left? What is the partial pressure of each gas?

Two possible approaches:

① Work through moles

OR

② Apply gas laws

$PV = nRT$

$n = \frac{PV}{RT}$

① H_2: $V = 17\ L$ $P = 1\ atm$ $T = 273.15\ K$

$$n_{H_2} = \frac{(17\ L)(1\ atm)}{0.08206\ \frac{L \cdot atm}{mol \cdot K}(273.15\ K)} = 0.758\ mol\ H_2 \left| \frac{2\ mol\ NH_3}{3\ mol\ H_2} \right. = 0.505\ mol\ NH_3$$

N_2: $5\ L$ $P = 1\ atm$ $T = 273.15\ K$

$$n_{N_2} = \frac{(5\ L)(1\ atm)}{0.08206\ \frac{L \cdot atm}{mol \cdot K}(273.15\ K)} = 0.223\ mol\ N_2 \left| \frac{2\ mol\ NH_3}{1\ mol\ N_2} \right. = 0.446\ mol\ NH_3$$

N_2 is L.R. Need Volume of NH_3 $V = \frac{nRT}{P} = \frac{0.446\ mol\ (0.08206)(273.15)}{1\ atm}$

$\boxed{V = 10\ L}$

② OR apply gas laws

P & T are constant (constant) $V = \left(\frac{RT}{P}\right) n$ $V = (\text{constant})\, n$

ALL gas @ the same T + P, so you can treat volume like moles:

$$17\ L\ H_2 \left| \frac{2\ L\ NH_3}{3\ L\ H_2} \right. = 11.33\ L \qquad 5\ L\ N_2 \left| \frac{2\ L\ NH_3}{1\ L\ N_2} \right. = 10\ L\ NH_3$$

Volume N_2 left over = Ø. It is the L.R.

H_2: $10\ L\ NH_3\ made \left| \frac{3\ L\ H_2}{2\ L\ NH_3} \right. = 15\ L\ H_2\ used$ $17\ L - 15\ L = \boxed{2\ L\ left}$

Partial Pressures: N_2 = Ø b/c nothing left

NH_3 + H_2 → need moles still @ 273.15 K

NH_3: 0.446 moles made (see above)

$P = \frac{0.446\ mol\ (0.08206)(273.15\ K)}{10\ L} = 1\ atm$

This should not be surprising, It was stated that the pressure doesn't change

H_2: $\frac{2\ L\ left}{17\ L\ started} = \frac{x}{0.758\ started}$ $x = 0.089\ mol\ H_2$

$P = \frac{0.089\ (0.08206)(273.15)}{2\ L}$

$P = 1.00\ atm$