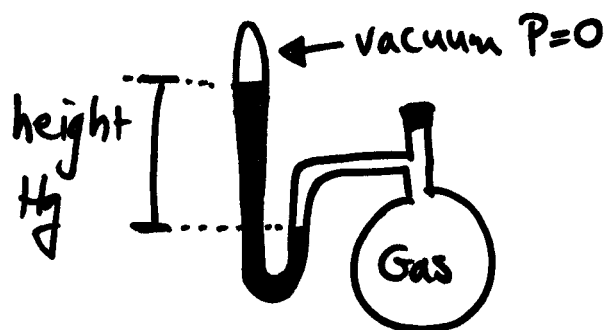


Gas particles move in straight line paths. As they collide, they create a force, pressure.

Pressure = Force / Area

Standard Atmospheric Pressure = 14.7 psi	(1b./ in <sup>2</sup> = force / area)
(at sea level) = 101325 Pa	(pascals) { 1013.25 millibars }
= 1 atm	(atmospheres)
= 760 mm Hg*	
= 760 torr*	*equivalent

### A. Closed Manometers:



Ex. 1) Find the pressure in Pa and atm, if the height of the mercury is 45 mm.

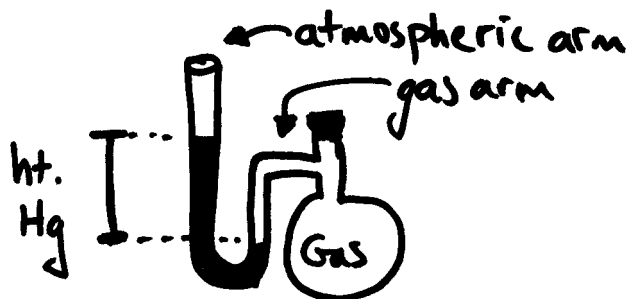
**Height of Mercury = the Pressure**

$$\frac{45 \text{ mm Hg}}{760 \text{ mm Hg}} \left| \frac{101325 \text{ Pa}}{760 \text{ mm Hg}} = 6.0 \times 10^3 \text{ Pa} \right.$$

$$\frac{45 \text{ mm Hg}}{760 \text{ mm Hg}} \left| \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 5.9 \times 10^{-2} \text{ atm} \right.$$

### B. Open Manometers:

Ex. 2) The mercury is 45 mm higher in the atmospheric arm. Find the pressure in Pa and atm, if the atmospheric pressure is 100000. Pa.



$$\frac{45 \text{ mm Hg}}{760 \text{ mm Hg}} \left| \frac{101325 \text{ Pa}}{760 \text{ mm Hg}} = 6.0 \times 10^3 \text{ Pa} \right.$$

$P_{\text{atm}} + P_{\text{Hg}} = P_{\text{gas}}$  ← **The gas pushes the hardest (strongest), so it is alone!**

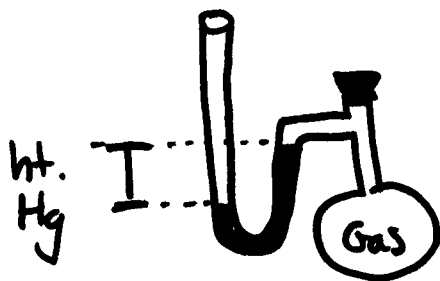
$$100000. \text{ Pa} + 6.0 \times 10^3 \text{ Pa} = P_{\text{gas}}$$

$$106000 \text{ Pa} = P_{\text{gas}}$$

$$1.1 \times 10^5 \text{ Pa} = P_{\text{gas}}$$

$$\frac{1.1 \times 10^5 \text{ Pa}}{101325 \text{ Pa}} \left| \frac{1 \text{ atm}}{101325 \text{ Pa}} = 1.1 \text{ atm} \right.$$

Ex. 3) Same as Ex. 2, except the mercury is higher in the gas arm.  
 (height of Hg = 45mm,  $P_{\text{atm}} = 100000. \text{ Pa}$ )



$$45 \text{ mm Hg} \left| \frac{101325 \text{ Pa}}{760 \text{ mm Hg}} = 6.0 \times 10^3 \text{ Pa} \right.$$

$P_{\text{atm}} = P_{\text{Hg}} + P_{\text{gas}}$  ← **The air pushes the hardest (strongest), so it is alone!**

$$100000. \text{ Pa} = 6.0 \times 10^3 \text{ Pa} + P_{\text{gas}}$$

$$94000. \text{ Pa} = P_{\text{gas}}$$

$$\mathbf{9.4 \times 10^4 \text{ Pa}}$$

$$\frac{9.4 \times 10^4 \text{ Pa}}{101325 \text{ Pa}} \left| \frac{1 \text{ atm}}{101325 \text{ Pa}} \right. = \mathbf{0.93 \text{ atm}}$$

Van der Waals Equation for Real Gases:

This accounts for the attractions of the gas particles.

This subtracts out the volume of the gas particles.

$$[\text{P}_{\text{observed}} + a \left(\frac{n}{V}\right)^2] (V - nb) = nRT$$

a, b are constants for the specific gas

n is mols, R is the gas constant

By assuming the gases are far away from each other, the equation can be simplified to

$$PV = nRT \quad \text{where } R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K}) \quad ^\circ\text{C} + 273 = \text{K}$$

\*\*Temperatures always in Kelvin!

In this **ideal gas law**, the gas molecules are spread out. There will not be any attractions, so  $P_{\text{observed}}$  will not need any adjustments. The molecules will be much smaller than the distances separating them. They are treated as point masses. They have mass, but negligible volume. (This works at low pressures and high temperatures.) Because of this the volume of the container can be used without subtracting the volume of the gas particles.

$$PV = nRT \quad \text{where } R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K}) \quad ^\circ\text{C} + 273 = \text{K}$$

Ex. 1) What pressure is exerted by 0.622 mol of CO<sub>2</sub> contained in a 922 ml vessel at 16 °C?

$$P = ?, \quad n = 0.622 \text{ mol}, \quad V = 922 \text{ ml}, \quad T = 16 \text{ }^\circ\text{C}$$

$$\frac{922 \text{ ml}}{1 \text{ ml}} \left| \frac{1 \times 10^{-3} \text{ L}}{1 \text{ ml}} \right. = 0.922 \text{ L} \quad 16 \text{ }^\circ\text{C} + 273 = 289 \text{ K}$$

$$PV = nRT$$

$$P (0.922 \text{ L}) = (0.622 \text{ mol}) (0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})) (289 \text{ K})$$

$$P = 16.0 \text{ atm}$$

Ex. 2) What number of mols exerts 234 torr in a 63.1 ml vessel at 25 °C?

$n = ?$  mols,  $P = 234$  torr,  $V = 63.1$  ml,  $T = 25$  °C

$$\frac{234 \text{ torr}}{760 \text{ torr}} \left| \frac{1 \text{ atm}}{760 \text{ torr}} \right. = 0.308 \text{ atm} \quad 25^\circ\text{C} + 273 = 298 \text{ K}$$

$$\frac{63.1 \text{ ml}}{1 \text{ ml}} \left| \frac{1 \times 10^{-3} \text{ L}}{1 \text{ ml}} \right. = 0.0631 \text{ L}$$

$PV = nRT$

$$(0.308 \text{ atm})(0.0631 \text{ L}) = n (0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K}))(298 \text{ K})$$

$$0.019435 = n (24.454)$$

$$\mathbf{7.95 \times 10^{-4} \text{ mol} = n}$$

#40 Notes III. More Gas Laws (Derived from the Ideal)

Ex. 1) A gas in a container at 25 °C, having a volume of 300 cm<sup>3</sup> exerts a pressure of 520 torr. What is its volume at standard atmospheric pressure?

$$T_1 = 25 \text{ }^\circ\text{C}$$

$$V_1 = 300 \text{ cm}^3$$

$$P_1 = 520 \text{ torr}$$

$$V_2 = ?$$

$$P_2 = 760 \text{ torr} = 1 \text{ atm etc.}$$

$$PV = \underline{nRT} \quad n, R, T \text{ are constant}$$

$$PV = \text{constant} \quad \text{so, } P_1V_1 = \text{constant} \quad \text{and} \quad P_2V_2 = \text{constant}$$

$$P_1V_1 = P_2V_2 \quad \text{Boyle's Law}$$

$$(520 \text{ torr})(300 \text{ cm}^3) = (760 \text{ torr}) V_2$$

$$\mathbf{205 \text{ cm}^3 = V_2}$$

Ex. 2) Find the volume at standard temperature for a sample of H<sub>2</sub> that at 9.00 °C has a volume of 613 ml.

$$\text{**Standard Temperature} = 273 \text{ K}$$

$$V_1 = ?$$

$$V_2 = 613 \text{ ml}$$

$$T_1 = 273 \text{ K}$$

$$T_2 = 9.00 \text{ }^\circ\text{C}$$

$$9.00 \text{ }^\circ\text{C} + 273 = 282 \text{ K}$$

$$\underline{PV} = \underline{nRT} \quad P, n, R \text{ are constant}$$

**\*Temperatures must be in Kelvin!**

(Adding conversions will not cancel.)

$$V = \frac{\underline{nRT}}{\underline{P}}$$

$$\frac{\underline{V}}{\underline{T}} = \frac{\underline{nR}}{\underline{P}}$$

$$\frac{\underline{V}}{\underline{T}} = \text{constant} \quad \text{so, } \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

**Charles' Law**

$$\frac{V_1}{(273\text{K})} = \frac{(613 \text{ ml})}{(282 \text{ K})}$$

$$(V_1)(282 \text{ K}) = (273 \text{ K})(613 \text{ ml}) \quad \text{**Cross multiply!!}$$

$$\mathbf{V_1 = 593 \text{ ml}}$$

Ex. 3) 45.3 mol of CH<sub>4</sub> gas occupies 916 L of space. What volume would 0.214 mol occupy?

$$\begin{array}{ll} n_1 = 45.3 \text{ mol} & n_2 = 0.214 \text{ mol} \\ V_1 = 916 \text{ L} & V_2 = ? \end{array}$$

$$\underline{PV} = n\underline{RT} \quad P, R, T \text{ are constant}$$

$$V = \frac{n\underline{RT}}{\underline{P}}$$

$$\frac{V}{n} = \frac{\underline{RT}}{\underline{P}} \quad \frac{V}{n} = \text{constant}$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{Avogadro's Law}$$

$$\frac{(916 \text{ L})}{(45.3 \text{ mol})} = \frac{V_2}{(0.214 \text{ mol})} \quad \text{**Cross Multiply!}$$

$$\mathbf{4.33 \text{ L} = V_2}$$

Ex. 4) CO<sub>2</sub> in a 45 ml vessel at 0.98 atm and 25 °C is transferred to a 0.067 L container at 35 °C. What is its new pressure?

$$\begin{array}{lll} V_1 = 45 \text{ ml} & V_2 = 0.067 \text{ L} & \frac{45 \text{ ml} \mid 1 \times 10^{-3} \text{ L}}{\mid 1 \text{ ml}} = 0.045 \text{ L} \\ P_1 = 0.98 \text{ atm} & P_2 = ? & \\ T_1 = 25 \text{ }^\circ\text{C} & T_2 = 35 \text{ }^\circ\text{C} & \begin{array}{l} 25 \text{ }^\circ\text{C} + 273 = 298 \text{ K} \\ 35 \text{ }^\circ\text{C} + 273 = 308 \text{ K} \end{array} \end{array}$$

$$PV = n\underline{RT} \quad n, R \text{ are constant}$$

$$\frac{PV}{T} = n\underline{R} \quad \frac{PV}{T} = \text{constant}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{(0.98 \text{ atm})(0.045 \text{ L})}{(298 \text{ K})} = \frac{(P_2)(0.067 \text{ L})}{(308 \text{ K})}$$

$$\mathbf{0.68 \text{ atm} = P_2}$$

$$PV = nRT \quad \text{mols} = \frac{\text{mass}}{\text{Molar mass}}$$

$$PV = \frac{(\text{mass})}{(\text{molar mass})} RT$$

$$(\text{molar mass}) = \frac{(\text{mass}) RT}{PV}$$

$$(\text{molar mass}) = \frac{(\text{mass})}{V} \cdot \frac{RT}{P}$$

$$\text{Molar mass} = \frac{dRT}{P} \quad \text{but density is in g/L !!!}$$

Ex. 1) A compound has the empirical formula  $\text{CH}_2$ . A 200 ml flask at 298 K and 755 torr contains 0.57 g of the gaseous compound. Give the molecular formula.

$$d = m/V \quad \frac{200 \text{ ml}}{1 \text{ ml}} \cdot 1 \times 10^{-3} \text{ L} = 0.200 \text{ L}$$

$$d = \frac{0.57 \text{ g}}{0.200 \text{ L}} \quad \frac{755 \text{ torr}}{760 \text{ torr}} \cdot 1 \text{ atm} = 0.993 \text{ atm}$$

$$d = 2.85 \text{ g/L}$$

$$mm = dRT/P \quad mm = \frac{(2.85 \text{ g/L}) (0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K})) (298 \text{ K})}{(0.993 \text{ atm})}$$

$$\text{mm} = 70. \text{ g/mol}$$

$$\text{Empirical Formula} = \text{CH}_2 = 14.027 \text{ g/mol} \quad \frac{70. \text{ g/mol}}{14.027 \text{ g/mol}} = 5$$

$$\text{Molecular Formula} = \text{C}_5\text{H}_{10}$$

\*\*argon = Ar, noble gases are monoatomic, others are diatomic ( $\text{N}_2$ ,  $\text{O}_2$ ,  $\text{H}_2$ )

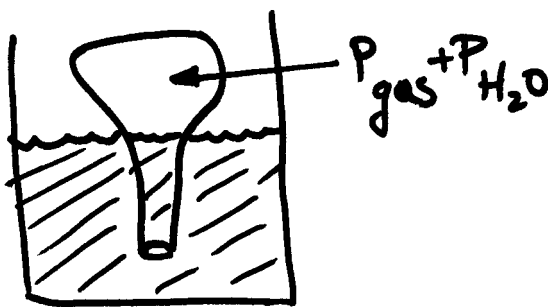
## #42 Notes V. Dalton's Law of Partial Pressures

In a mixture of gases, each particular gas will exert the same pressure it would, if it were alone. In this room  $P_{O_2} = 0.20$  atm,  $P_{N_2} = 0.79$  atm, What is  $P_{tot}$  ?

$$P_{tot} = P_1 + P_2 + P_3 + \dots$$

The total pressure is the sum of the partial pressures.

Ex. 1)  $CO_2$  is collected over  $H_2O$  at  $10^\circ C$  in a  $20.0\text{ cm}^3$  vessel. The manometer indicates a pressure of  $60.0\text{ kPa}$ . What is the mass of gas collected? ( $P_{H_2O}$  at  $10^\circ C = 1.2\text{ kPa}$ )



$$\begin{aligned} P_{tot} &= P_{gas} + P_{H_2O} \\ 60\text{ kPa} &= P_{gas} + 1.2\text{ kPa} \\ 58.8\text{ kPa} &= P_{gas} \end{aligned}$$

$$P = \frac{58.8\text{ kPa}}{1\text{ kPa}} \left| \frac{1000\text{ Pa}}{1\text{ kPa}} \right| \left| \frac{1\text{ atm}}{101325\text{ Pa}} \right| = 0.580\text{ atm}$$

$$T = 10^\circ C = 283\text{ K}$$

$$V = \frac{20.0\text{ cm}^3}{1\text{ cm}^3} \left| \frac{1\text{ ml}}{1\text{ cm}^3} \right| \left| \frac{1 \times 10^{-3}\text{ L}}{1\text{ ml}} \right| = 0.0200\text{ L}$$

$$n = ?$$

$$PV = nRT$$

$$(0.580\text{ atm})(0.0200\text{ L}) = n(0.08206\text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K}))(283\text{ K})$$

$$4.995 \times 10^{-4}\text{ mol} = n$$

$$\frac{4.995 \times 10^{-4}\text{ mol } CO_2}{1\text{ mol}} \left| \frac{44.00915\text{ g}}{1\text{ mol}} \right| = 2.2 \times 10^{-2}\text{ g } CO_2$$

\*\* Remember: All gases diatomic ( $N_2$ ,  $H_2$ ,  $O_2$ ) except noble gases ( $He$ ,  $Ne$ ,  $Ar$ )



Ex. 2) 6.00 dm<sup>3</sup> of H<sub>2</sub> and 2.00 dm<sup>3</sup> of O<sub>2</sub> at 25 °C and 760. torr are pumped into a 0.250 L container. What is the pressure in the container?

$$\begin{aligned} & \underline{\text{H}_2} \\ V &= 6.00 \text{ dm}^3 \\ T &= 25 \text{ }^\circ\text{C} \\ P &= 760. \text{ torr} \end{aligned}$$

$$\begin{aligned} & \underline{\text{O}_2} \\ V &= 2.00 \text{ dm}^3 \\ T &= 25 \text{ }^\circ\text{C} \\ P &= 760. \text{ torr} \end{aligned}$$

$$\begin{aligned} & \underline{\text{new container}} \\ V &= 0.250 \text{ L} \\ \\ P &= ? \end{aligned}$$

Constant T, so ignore

$$PV = nRT \quad n, R, T \text{ are constant}$$

$$P_1 V_1 = P_2 V_2$$

$$P_1 V_1 = P_2 V_2$$

$$\begin{aligned} (760 \text{ torr}) (6.00 \text{ dm}^3) &= P_2 (0.250 \text{ L}) \\ 18240 \text{ torr} &= P_2 \text{ of H}_2 \end{aligned}$$

$$**1\text{L} = 1\text{dm}^3$$

$$\begin{aligned} & \underline{\text{O}_2} \\ P_1 V_1 &= P_2 V_2 \\ (760 \text{ torr}) (2.00 \text{ dm}^3) &= P_2 (0.250 \text{ L}) \\ 6080 \text{ torr} &= P_2 \text{ of O}_2 \end{aligned}$$

$$P_{\text{tot}} = P_{\text{H}_2} + P_{\text{O}_2} = 18240 \text{ torr} + 6080 \text{ torr} = 24320 = \mathbf{2.43 \times 10^4 \text{ torr}}$$

Derive a conversion factor between mols and volume at STP for 1 mol of any gas.

STP = standard Temperature (273 K) and standard Pressure (1atm)

$$P = 1 \text{ atm} \quad V = ? \quad n = 1 \text{ mol} \quad T = 273 \text{ K}$$

$$PV = nRT$$

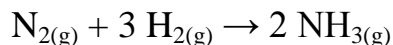
$$(1 \text{ atm}) V = (1 \text{ mol}) (0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K})) (273 \text{ K})$$

$$V = 22.4 \text{ L}$$

**\*\* 1 mol of any gas = 22.4 L at STP**

Ex. 1) An excess of nitrogen gas reacts with 5.0 g of hydrogen gas. What volume of ammonia is produced at STP?

Write the reaction:



$$\frac{5.0 \text{ g H}_2}{2.016 \text{ g H}_2} \left| \frac{1 \text{ mol H}_2}{3 \text{ mol H}_2} \right| \frac{2 \text{ mol NH}_3}{1 \text{ mol NH}_3} \left| \frac{22.4 \text{ L NH}_3}{1 \text{ mol NH}_3} \right| = 37 \text{ L NH}_3$$

How many grams of H<sub>2</sub> are produced from 39 L NH<sub>3</sub>?

$$\frac{39 \text{ L NH}_3}{22.4 \text{ L NH}_3} \left| \frac{1 \text{ mol NH}_3}{2 \text{ mol NH}_3} \right| \frac{3 \text{ mol H}_2}{1 \text{ mol H}_2} \left| \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right| = \mathbf{5.3 \text{ g H}_2}$$

If it is not at STP, do stoichiometry, but stop at mols (after the mol ratio) and then use PV = nRT.

{OR Change given amounts to STP by using gas equations (T = 273 K, P = 1atm), then do stoichiometry.}

If told 25 L of a gas, but not told the compound, look in the reaction to find which is a gas.

## #44 Notes VII. Kinetic Molecular Theory of Gases

- 1) The gas particles are so small compared to the distances separating them that we can assume their volume is negligible (zero). **Ideal Gases**
- 2) The gas particles are in constant motion and their collisions cause the pressure exerted by the gas.
- 3) The particles are assumed to exert no attractive/ repulsive forces on each other.  
**Ideal Gases**
- 4) The average Kinetic Energy (KE) is directly proportional to temperature.

$$KE = \frac{1}{2} m v^2$$

↑ ↑  
mass velocity

\*high temperature, high velocity of particles, high KE

$$KE = \frac{3}{2} RT \quad \text{where } R = 8.31 \text{ J}/(\text{mol} \cdot \text{K})$$

$$1 \text{ joule} = 1 \text{ Newton} \cdot \text{meter}$$
$$1 \text{ Newton} = 1 (\text{kg} \cdot \text{m}) / \text{s}^2$$

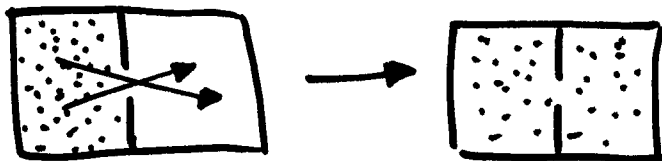
## VIII. Root Mean Square Velocity

$$KE = \frac{1}{2} m v^2 = \frac{3}{2} RT$$
$$m v^2 = 3RT$$
$$v^2 = \frac{3RT}{m}$$

$$v_{\text{rms}} = \sqrt{\frac{3RT}{m}}$$

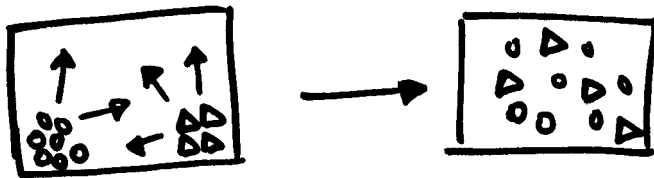
\*\* molar mass must be in **kg/mol**, velocity in m/s

## IX. Effusion/ Diffusion



Effusion:

-filling of a gas into a chamber by a hole.



Diffusion:

-mixing of 2 gases.

Graham's Law of Effusion (works for diffusion too.)

$$\begin{aligned} \text{Rate of effusion for gas \#1} &= v_{\text{rms}} \text{ \#1} = \sqrt{3RT/m_1} = \sqrt{(1/m_1)} = \sqrt{m_2} \\ \text{Rate of effusion for gas \#2} &= v_{\text{rms}} \text{ \#2} = \sqrt{3RT/m_2} = \sqrt{(1/m_2)} = \sqrt{m_1} \end{aligned}$$

$$\begin{aligned} \text{Rate gas \#1} &= \sqrt{m_2} \\ \text{Rate gas \#2} &= \sqrt{m_1} \end{aligned} \quad \text{**Rate must be in volume/time, but units don't matter if they agree.}$$

Ex. 1) Find KE and  $v_{\text{rms}}$  for  $F_2$  at  $25^\circ\text{C}$ .

$$\text{KE} = 3/2 RT = 3/2 (8.31 \text{ J}/(\text{mol}\cdot\text{K})) (298 \text{ K}) = \mathbf{3.71 \times 10^3 \text{ J/mol}}$$

$$v_{\text{rms}} = \sqrt{3RT/m} = \sqrt{\frac{(3) (8.31 \text{ J}/(\text{mol}\cdot\text{K})) (298 \text{ K})}{(0.03800 \text{ kg/mol})}} = \mathbf{442 \text{ m/s}}$$

$$F_2 = \frac{38.00 \text{ g}}{\text{mol}} \left| \frac{1 \text{ kg}}{1 \times 10^3 \text{ g}} \right. = 0.03800 \text{ kg/mol } F_2$$

↑↑↑

Ex. 2) A gas has an effusion rate of 15 ml/min. If  $\text{CH}_4$  has a rate of 45 ml/min, what is the molar mass of the other gas?

$$\frac{R_1}{R_2} = \frac{\sqrt{m_2}}{\sqrt{m_1}} \quad \text{CH}_4 = \frac{16.04 \text{ g}}{\text{mol}} \left| \frac{1 \text{ kg}}{1 \times 10^3 \text{ g}} \right. = 0.01604 \text{ kg/mol}$$

$$\begin{aligned} \text{Rate ? gas} &= \sqrt{m_{\text{CH}_4}} \\ \text{Rate CH}_4 &= \sqrt{m_{\text{?gas}}} \end{aligned}$$

$$\frac{15 \text{ ml/min}}{45 \text{ ml/min CH}_4} = \frac{\sqrt{0.01604 \text{ kg/mol CH}_4}}{\sqrt{m_{\text{?gas}}}}$$

$$0.333 = \frac{0.1266}{\sqrt{m_{\text{?gas}}}}$$

$$0.333 \sqrt{m_{\text{?gas}}} = 0.1266$$

$$\sqrt{m_{\text{?gas}}} = 0.380$$

$$m_{\text{?gas}} = \mathbf{0.144 \text{ kg/mol}}$$

**\*End of Notes\*** (Assignments #45-46 are Review Assignments. There are no notes for these assignments.)