

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

Pressure = Force / Area

Standard Atmospheric Pressure = 14.7 psi

(at sea level)

= 101325 Pa

= 1 atm

= 760 mm Hg*

= 760 torr*

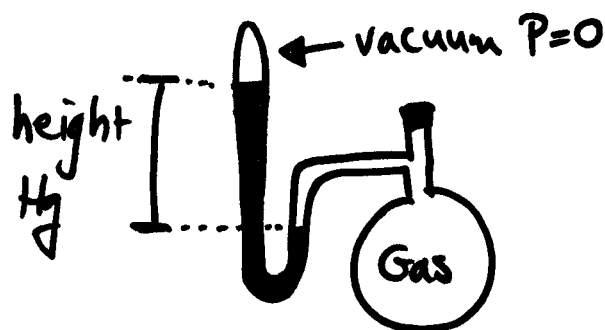
(1b./ in² = force / area)

(pascals) { 1013.25 millibars }

(atmospheres)

*equivalent

A. Closed Manometers:



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

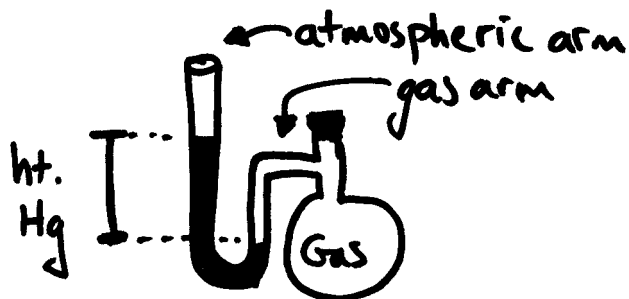
Height of Mercury = the Pressure

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

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

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.



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

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

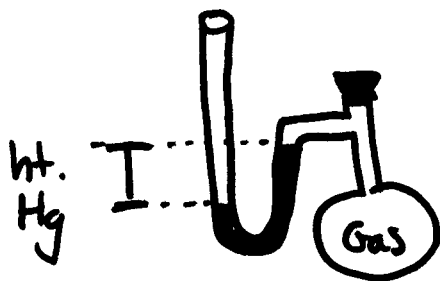
$$100000. \text{ Pa} + 6.0 \times 10^3 \text{ Pa} = P_{\text{gas}}$$

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

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

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

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



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

$P_{\text{atm}} = P_{\text{Hg}} + P_{\text{gas}} \quad \leftarrow \text{The air pushes the hardest (strongest), so it is alone!}$
 $100000. \text{ Pa} = 6.0 \times 10^3 \text{ Pa} + P_{\text{gas}}$
 $94000. \text{ Pa} = P_{\text{gas}}$
 $9.4 \times 10^4 \text{ Pa}$

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

II. Ideal Gas Law:

In this **ideal gas law**, the gas molecules are spread out. There will not be any attractions, so P_{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.

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

Ex. 1) What pressure is exerted by 0.622 mol of CO_2 contained in a 922 ml vessel at 16°C ?

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

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

$PV = nRT$

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

$P = 16.0 \text{ atm}$

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

Ex. 1) A gas in a container at 25 °C, having a volume of 300 cm³ 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 H₂ that at 9.00 °C has a volume of 613 cm³.

**Standard Temperature = 273 K

$$V_1 = ?$$

$$V_2 = 613 \text{ cm}^3$$

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

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

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

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

***Temperatures must be in Kelvin!**
(Adding conversions will not cancel.)

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

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

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

Charles' Law

$$\frac{V_1}{(273\text{K})} = \frac{(613 \text{ cm}^3)}{(282 \text{ K})}$$

**Cross Multiply!!

$$\mathbf{V_1 = 593 \text{ cm}^3}$$

Ex. 3) 45.3 mol of CH₄ gas occupies 916 L of space. What volume would 0.214 mol occupy?

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

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

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

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

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

$$\frac{(916 \text{ L})}{(45.3 \text{ mol})} = \frac{V_2}{(0.214 \text{ mol})}$$

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

Ex. 4) CO₂ 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 CH_2 . A 200 ml flask at 298 K and 755 torr contains 0.57 g of the gaseous compound. Give the molecular formula.

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

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

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

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

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

$$\text{Empirical Formula} = \text{CH}_2 = 14.027 \text{ g/mol}$$

$$\frac{70. \text{ g/mol}}{14.027 \text{ g/mol}} = 5$$

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

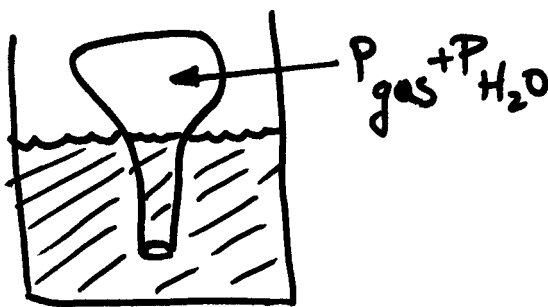
#31 Notes V. Dalton's Law of Partial Pressures

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



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

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

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

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

$$n = ?$$

$$\begin{aligned} PV &= nRT \\ (0.580\text{ atm})(0.0200\text{ L}) &= n(0.08206\text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K}))(283\text{ K}) \end{aligned}$$

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

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

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

Ex. 2) 6.00 dm³ of H₂ and 2.00 dm³ of O₂ 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

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

$$P_1 V_1 = P_2 V_2$$

$$\underline{\text{H}_2}$$

$$P_1 V_1 = P_2 V_2$$

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

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

$$\underline{\text{O}_2}$$

$$P_1 V_1 = P_2 V_2$$

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

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

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)

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

$$PV = nRT$$

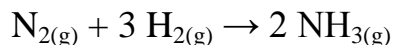
$$(1 \text{ atm}) V = (1 \text{ mol}) (0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K})) (273 \text{ K})$$

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

**** 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 H₂ are produced from 39 L NH₃?

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

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.

#33 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

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

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

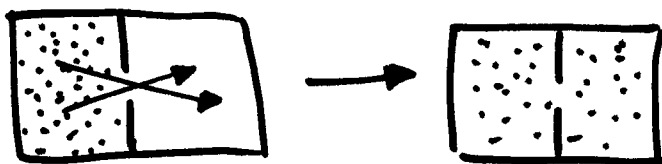
VIII. Root Mean Square Velocity

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

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

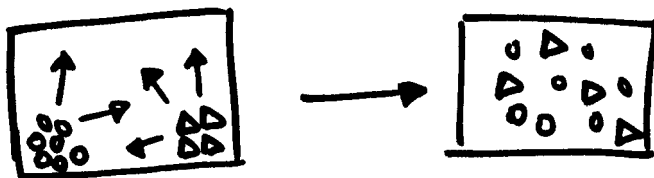
** molar mass must be in **kg/mol**, velocity in m/s

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.)

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

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

Ex. 1) Find KE and v_{rms} for F_2 at 25°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 CH_4 has a rate of 45 ml/min, what is the molar mass of the other gas?

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

$$\frac{\text{Rate ? gas}}{\text{Rate CH}_4} = \frac{\sqrt{m_{\text{CH}_4}}}{\sqrt{m_{\text{?gas}}}}$$

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

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

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

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

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

$PV = nRT$ is for Ideal Gases

The particle volume was assumed to be zero, but molecules really do take up space.

$$V_{\text{volume around the gas}} = V_{\text{volume of the container}} - nb_{\text{volume of the gas particles}}$$

Where $n = \text{mols}$ and

$b = \text{an experimentally found empirical constant}$

The attractive forces were also assumed to be zero, but attractive forces between molecules will keep them from hitting the container as hard as predicted.

$$P_{\text{observed pressure}} = P'_{\text{the real pressure}} - a \left(\frac{n}{V}\right)^2 \text{ decrease in pressure caused by the attractive forces}$$

Where $a = \text{an experimentally found proportionality constant}$ and

$n = \text{mols}$ and V is volume

$$\text{so } P'_{\text{real}} = P_{\text{observed}} + a \left(\frac{n}{V}\right)^2$$

$$P_{\text{real}} V_{\text{around the gas}} = nRT$$

$$[P_{\text{observed}} + a \left(\frac{n}{V}\right)^2] (V - nb) = nRT \quad a, b: \text{ see Table in textbook}$$

Van der Waals Equation for Real Gases

Ex.1) Calculate the pressure exerted by 0.2500 mol of Krypton in a 1.000 L container at 25.0 °C.

$$P = ?$$

$$n = 0.2500 \text{ mol}$$

$$V = 1.000 \text{ L}$$

$$T = 25 \text{ }^\circ\text{C} = 298 \text{ K}$$

$$\text{Look up for Kr: } a = 2.32 \text{ atm}\cdot\text{L}^2/\text{mol}^2$$

$$b = 0.0398 \text{ L/mol}$$

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

$$[P + (2.32 \text{ atm}\cdot\text{L}^2/\text{mol}^2) (0.2500 \text{ mol}/ 1.000 \text{ L})^2] [1.000 \text{ L} - (0.2500 \text{ mol})(0.0398 \text{ L/mol})] = (0.2500 \text{ mol})(0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K}))(298 \text{ K})$$

$$[P + 0.145 \text{ atm}] [0.99005 \text{ L}] = (6.11347 \text{ L}\cdot\text{atm})$$

$$[P + 0.145 \text{ atm}] = 6.1749 \text{ atm}$$

$$\mathbf{P = 6.03 \text{ atm}}$$

End of Notes (Assignments #35-36 are Review Assignments. There are no notes for these assignments.)