UNIT 10 – CHAPTER 5 STUDENT NOTES: GASES

With gases, volume must be considered with temperature, pressure, and amount of material (moles).

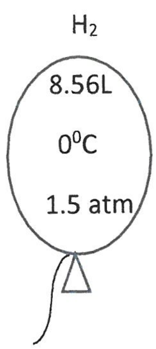
Volume – Liters

Temperature – Kelvin ($K = {}^{0}C + 273$)

Pressure – 3 common units: atm/torr, mmHg, also kPa

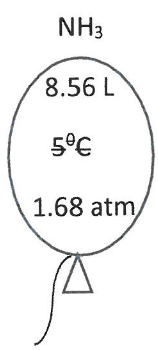
Ideal gas law – used in gas law calculations

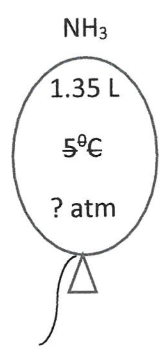
EX 1: **Ideal Gas Law: $P \cdot V = n \cdot R \cdot T$**



n = ?

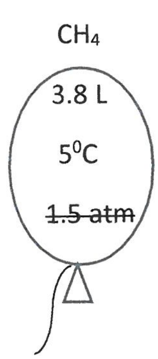
EX 2: **Boyle's Law – $P_1 \cdot V_1 = P_2 \cdot V_2$**



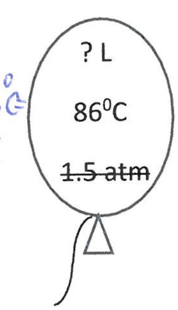


EX 3: **Charles' Law – $\frac{V_1}{T_1} = \frac{V_2}{T_2}$**

273+5°C =
278K



273+86°C =
359K

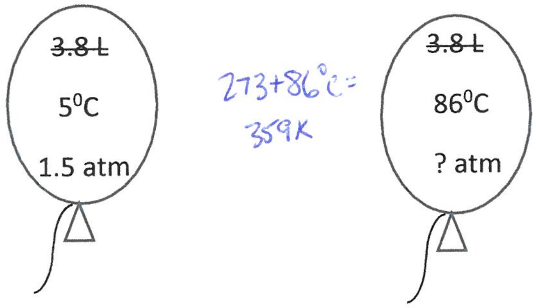


$\frac{3.8L}{278K} = \frac{V_2}{359K}$

$\boxed{V_2 = 4.9L}$

EX 4: **Gay-Lussac's Law - $\frac{P_1}{T_1} = \frac{P_2}{T_2}$**

CH_4



273+5°C = 278K

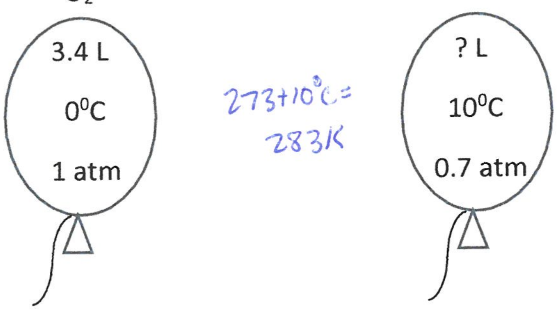
273+86°C = 359K

$\frac{1.5}{278} = \frac{P_2}{359}$

$\boxed{P_2 = 1.9\text{ ATM}}$

EX 5: **Combined Gas Law - $\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$**

O_2



273+0°C = 273K

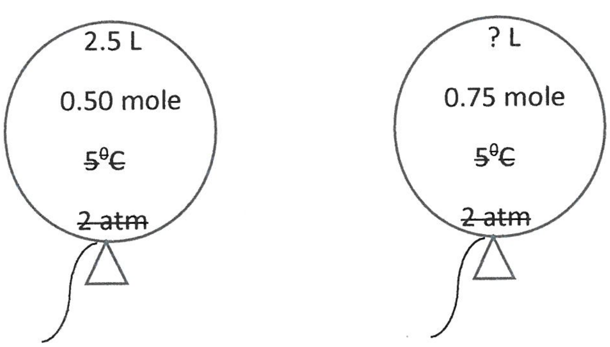
273+10°C = 283K

$\frac{(1)(3.4L)}{273} = \frac{(.7)(V_2)}{283}$

$\boxed{V_2 = 5.0L}$

EX 6: **Avagadro's Law – $\frac{V_1}{n_1} = \frac{V_2}{n_2}$**

N_2



$\frac{2.5L}{.50mol} = \frac{V_2}{.75mol}$

$\boxed{V_2 = 3.75L}$

Problems:

A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L. If the balloon is filled with 2.0 L of helium at sea level (760 mm Hg), is released, and it rises to an altitude at which the atmospheric pressure is only 500.0 mm Hg, will the balloon burst? (Assume the temperature is constant)

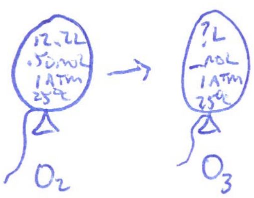
$P_1V_1 = P_2V_2$

$(760\text{ TORR})(2.0L) = (500\text{ TORR})(V_2)$

$\boxed{V_2 = 3.0L}$

PRESSURE CAN BE IN ANY UNITS IF NOT USING PV=nRT

Suppose you have a 12.2-liter sample containing 0.50 moles of oxygen gas at a pressure of 1 atm and a temperature of 25°C. If all this O_2 was converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone?



$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

$\frac{12.2L}{.50MOL} = \frac{V_2}{.33MOL}$

$\boxed{V_2 = 8.1L}$

STOICH

$3O_2 \rightarrow 2O_3$

.50MOL ?MOL

$.50MOL\ O_2 \times \frac{2MOL\ O_3}{3MOL\ O_2} = .33MOL\ O_3$

$PV=nRT \quad (1ATM)(V) = (.33MOL)(.08206)(298K)$

$\boxed{V = 8.1L}$

Molar volume – volume of 1 mole of a gas at standard conditions

EX 7: A 1.75-liter volume of gas exists at STP conditions. How many moles of gas are there?

1. Calculate the grams of calcium carbonate needed to produce 3.2 L of carbon dioxide at

   a) STP

   b) 10°C and 0.7 atmospheres of pressure

2. 2.80 liters of methane gas at 25°C and 1.65 atm of pressure reacts with 35.0 liters of oxygen at 31.0°C and 1.25 atm of pressure. What volume of carbon dioxide is produced at 125°C and 2.50 atm of pressure?

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

CH_4	O_2	CO_2
2.80L	35.0L	X
25°C	31°C	125°C
1.65ATM	1.25ATM	2.50ATM

LR

$CH_4 \rightarrow (1.64)(2.80L) = (n)(.08206)(298K) = .0189MOL\ CH_4 \times \frac{1MOL\ CO_2}{1MOL\ CH_4} = .0189\ MOL\ CO_2$

$O_2 \rightarrow (2.50)(35.0L) = (n)(.08206)(304K) = 1.75MOL\ O_2 \times \frac{1MOL\ CO_2}{2MOL\ O_2} = .875MOL\ CO_2$

$(2.50ATM)(V) = (.0189MOL)(.08206)(398K)$

$\boxed{V = 2.47L\ CO_2}$

3. A student prepares a sample of hydrogen gas by electrolyzing water at 25⁰C. She collects 152 mL of H_2 at a total pressure of 758 mm Hg (the vapor pressure of water at 25⁰C is 23.76 mm Hg). Calculate the moles of H_2 collected.

$\underset{(TOTAL)}{758\ mm\ Hg} - \underset{(H_2O)}{23.76\ mm\ Hg} = \underset{(DRY\ GAS)}{734.24\ mm\ Hg}$

$\frac{734.24\ mm\ Hg \mid 1\ ATM}{\mid 760\ mm\ Hg} = (.965\ ATM)(.152\ L) = (n)(.08206)(298K)$

$\boxed{n = .006\ mol\ H_2}$

4. Air bags are activated when a severe impact causes a steel ball to compress a spring and electrically ignite a detonator cap. This causes sodium azide (NaN_3) to decompose explosively according to the following reaction:

$$\overset{Xg}{2\ NaN_{3(s)}} \rightarrow 2\ Na_{(s)} + \overset{70.1L}{3\ N_{2(g)}}$$

What mass of $NaN_{3(s)}$ must be reacted to inflate an air bag to 70.1 L at STP?

$PV = nRT$

$(1.00\ ATM)(70.1\ L) = (n)(.08206)(273K)$

$n = 3.13\ mol\ N_2 \left| \frac{2\ mol\ NaN_3}{3\ mol\ N_2} \right| \frac{65.02g\ NaN_3}{1\ mol\ NaN_3} = \boxed{135.7g\ NaN_3}$

5. Consider the reaction between 50.0 mL of liquid methyl alcohol, CH_3OH (density = 0.850 g/mL), and 22.8 L of O_2 at 27⁰C and a pressure of 2.00 atm. The products of the reaction are $CO_{2(g)}$ and $H_2O_{(g)}$. Calculate the number of moles of H_2O formed if the reaction goes to completion.

6. In a chemical reaction, calcium reacts with hydrogen bromide. How many grams of calcium are needed to produce 24.2 liters of hydrogen by water displacement at 17⁰C and 738 mm Hg?

$\underset{Xg}{Ca} + 2\ HBr \rightarrow CaBr_2 + \underset{24.2L}{H_2}$

$\underset{(TOTAL)}{738\ mm\ Hg} - \underset{(H_2O)}{14.5\ mm\ Hg} = \underset{(DRY\ GAS)}{723.5\ mm\ Hg}$

$\frac{723.5\ mm\ Hg \mid 1\ ATM}{\mid 760\ mm\ Hg} = (.952\ ATM)(24.2\ L) = (n)(.08206)(290K)$

$n = .972\ mol\ H_2 \left| \frac{1\ mol\ Ca}{1\ mol\ H_2} \right| \frac{40.1g\ Ca}{1\ mol\ Ca} = \boxed{38.9g\ Ca}$

7. In a chemical reaction, 2.3 grams of lithium chloride reacts with bromine to produce lithium bromide and chlorine. Calculate the volume of chlorine gas collected at 10⁰C and 768.6 mm Hg by water displacement.

Dalton's Law of Partial Pressure – for a mixture of gases in a container, the total pressure exerted is the sum of the pressures that each gas would exert if it were alone.

N_2
O_2
CH_4

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

1. A mixture of 1.00 g H_2 and 1.00 g He is placed in a 1.00-L container at 27^0C. Calculate the partial pressure of each gas and the total pressure of the system.

1.00g H_2
1.00g He

$n_{H_2} = 1.00g\,H_2 \left| \frac{1\,mol\,H_2}{2.02g\,H_2} \right. = .496\,mol\,H_2$

$n_{He} = 1.00g\,He \left| \frac{1\,mol\,He}{4.003g\,He} \right. = .250\,mol\,He$

$P_{H_2} \rightarrow (P_{H_2})(1.00L) = (.496\,mol)(.08206)(300K) = 12.2\,ATM\,H_2$

$P_{He} \rightarrow (P_{He})(1.00L) = (.250\,mol)(.08206)(300K) = 6.2\,ATM\,He$

+

$\boxed{18.4\,ATM\ TOTAL}$

$(P_{H_2} + P_{He})(1.00L) = (.496 + .250)(.08206)(300)$

$\boxed{P_{H_2} + P_{He} = 18.4L}$

2. A piece of solid carbon dioxide with a mass of 7.8 grams is placed in a 4.0-liter, otherwise empty, container at 27^0C. What is the pressure in the container after all the carbon dioxide vaporizes? If this carbon dioxide were placed in the same container but it already contained air at 740 torr, what would be the partial pressure of the carbon dioxide and the total pressure in the container after the carbon dioxide vaporizes?

Molar mass of a gas – can be calculated using the density of a gas

$$MM = \frac{dRT}{P} = \frac{mRT}{vP}$$

1. A certain gas with a pressure of 1.50 atm and 27^0C has a density of 1.96 g/liter. What is the molar mass and most probable identity of the gas?

$MM = \frac{(1.96g)(.08206)(300K)}{L\ (1.50\,ATM)} = \boxed{\frac{32.2g}{L} \rightarrow O_2}$

Mole fraction – ratio of the number of a given component in a mixture to the total number of moles.

$$X = \frac{n_1}{n_{total}} = \frac{P_1}{P_{total}}$$

1. 2.0 liters of He at 46^0C and 1.2 atm pressure was added to a vessel that contains 4.5 liters of N_2 at STP. What is the total pressure and partial pressure of each gas at STP after the He is added?

2.0L He
46°C
1.2 ATM

4.5L
N_2
STP

$n_{He} = (1.2\,ATM)(2.0L) = (n)(.08206)(319K) = .091\,MOL\,He$

$n_{N_2} = (1\,ATM)(4.5L) = (n)(.08206)(273K) = .2\,mol\,N_2$

$(P)(4.5L) = (.291\,mol_{TOTAL})(.08206)(273) = \boxed{1.45\,ATM}$

2. A balloon is filled with 2.1 grams of oxygen gas and 5.2 grams of chlorine gas, producing a total pressure of 748 torr. What is the partial pressure of each gas in the balloon?

3. When one mole of methane, CH_4, is heated with four moles of oxygen, the following reaction occurs:

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

a) Assuming all the methane is converted to CO_2 and H_2O, what are the mole fractions of O_2, CO_2, and H_2O in the resulting mixture?

b) If the total pressure of the mixture is 1.26 atm, what are the partial pressures of the gases?

Root mean square velocity – an expression dealing with the average velocity of gas particles – a derived formula

r = 8.3145 j/mole

$$V_{rms} = \sqrt{3rt/m}$$

t = K temperature

m = molar mass in kg

EX 8: Calculate the root mean velocity of He gas and O_2 gas at 25^0C.

Effusion and Diffusion

Effusion – relates to the passage of a gas through a hole in an evacuated chamber

Diffusion – relates to mixing of gases; directly related to effusion

Graham's Law of Effusion and Diffusion

$$\frac{\text{Rate}_{\text{effusion 1}}}{\text{Rate}_{\text{effusion 2}}} = \sqrt{\frac{m_2}{m_1}}$$

EX 9: How many times faster will He effuse than NO_2 gas?

He – NO_2?

NH_3 – HCl?

Real gases – different than ideal gases in two important ways

1) The space a molecule takes up – Real gas molecules take up actual volume that is used for movement. The actual free volume is less in real gases than ideal gases.

2) The pressure because of attractive forces – Because the molecules are attracted to each other, the pressure on the container is less than ideal. The pressure in the container will be less for a real gas than an ideal gas.

Volume correction $\qquad$ **$V' = V - n \cdot b$**

***moles (amount of material) affects volume; more moles, more effect**

***b is a constant that differs for each gas**

Pressure correction $\qquad$ **$P_{observed} = P' - (an^2/v^2)$**

***moles (amount of material) important**

***a is a constant that differs for each gas**

***depends on the volume of the container; more volume, more effect**

***since two molecules interact, the effect must be squared**

Ideal Gas Law $\qquad$ **$P \cdot V = n \cdot R \cdot T$**

Real Gas Law $\qquad$ **$P_{observed} + (an^2/v^2) \cdot (V - nb) = nRT$** $\qquad$ **(van der Waals equation)**

***a and b are determined by experimentation (Table 5.3, pg. 210)**

***different for each gas**

***a depends on both size and polarity**

***once given, plug and chug**

EX 10: Calculate the pressure exerted by 0.5000 moles N_2 in a 1.0000-liter container at 25.0°C

a) using the ideal gas law $PV = nRT \rightarrow (P)(1.0000) = \underbrace{(.5)(.08206)(298)}_{12.23} = \boxed{12.23 \text{ ATM (IDEAL)}}$

b) using the van der Waals equation

c) compare the results by finding the % error between the two

$$b)\ P + \left(\frac{(1.39 \cdot .5000)^2}{(1)^2}\right) \cdot [1.0000L - (.5)(.039)] = 12.23$$

$$(P + .348) \cdot (.9805) = 12.23 \rightarrow \boxed{P = 12.13 \text{ (VDW)}}$$

$$c)\ \%ERROR = \left|\frac{P_{VDW} - P_{IDEAL}}{P_{IDEAL}}\right| \times 100\% \qquad \frac{12.13 - 12.23}{12.23} \times 100\% = \boxed{.82\% \text{ ERROR}}$$

EX 11: Calculate the above experiment if it were to occur in a 10.000-liter container and compare.