Chapter 5: The Gaseous State

**These Notes are to <u>SUPPLIMENT </u>the Text, They do NOT Replace reading the Text Material.
Additional material that is in the Text will be on your tests!**

**To get the most information, <u>READ THE CHAPTER</u> prior to the Lecture, bring in these lecture notes
and make comments on these notes. These notes alone are NOT enough to pass any test!**

Gas Pressure and Measurement
The Difference between Solids and Liquids and Gasses – U can compress a gas

U can relate Pressure – Volume – Temperature and Molar Voume in 1 Equation!!

PRESSURE – the force exerted per unit area of surface

ACCELERATION – The change of speed per unit of time
          Acceleration of Gravitiy (AOG) = 9.81 m / sec^2

FORCE   = Mass * AOG
PRESSURE   = Force / Area = Force / [Pi r^2]

A coin 9.3 mm in radius weighs 2.5 g is on a table.
     Force = Mass * AOG   = 2.5 x 10^{-2} kg * 9.81 m/s^2 = 2.45 x 10^{-1} **= 2.5 * 10^{-2} kg m/s^2**
     Pressure = Force / [Pi r^2]   = 2.5 x 10^{-3} kg m/s^2 / 3.14 * (9.2 x 10^{-3} m)^2 **= 93 kg / ms^2**

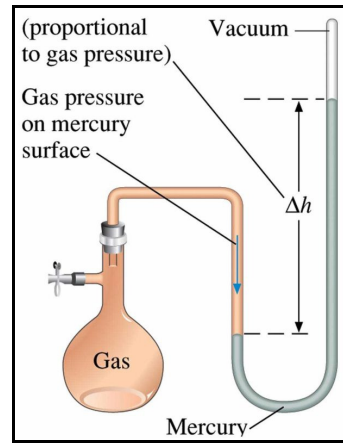**Units of Pressure:**   **Pascal (Pa)** kg / ms^2
                    Atm   1 atm = 1.01325 x 10^5 Pa = approx 100 k Pa
                    Mm Hg   = TORR = 760 mm Hg = 1 Atm

<u>**Barameter **</u>is a device for measuring atmospheric pressure. At sea level it's 760 mm Hg

(proportional to gas pressure) Vacuum

Gas pressure on mercury surface

Δh

Gas

Mercury
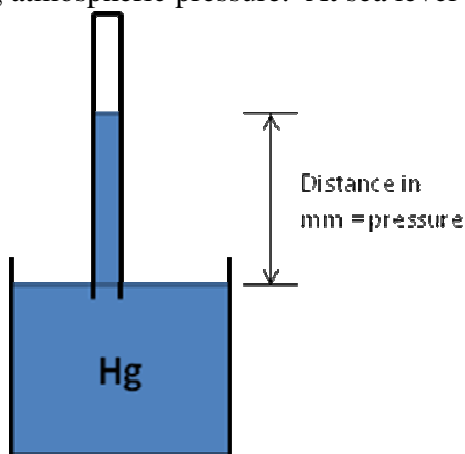
Manometer is a device to measure the pressure of a gas.

Example 5.1 The pressure in a flask is 797.7 mm Hg. Find the pressure in Pa and Atm.

1. 797.7 mm Hg * 1.01325 x 105 Pa / 760 mm Hg = **1.064 x 10⁵ Pa**

2. 797.7 mm Hg * 1 atm / 760 mm Hg = **1.050 atm**

CONCEPT CHECK: 5.1 Two Barameters, one with Hg, one with water. Which has the higher column of water?

H = P / (g * d)   H = height   P = Pressure   g = AOG   d = density of liquid

H $_{water}$ = P / g * D $_w$ = P / g * 1 g/cc   **H** $_{Hg}$ = P / g D $_{Hg}$ = P / g * 13 g/cc

H $_{water}$ **>>** **H** $_{Hg}$

BOYLES LAW   Gas is compressible. The volume of a gas at a given temperature varies inversely with the applied pressure.

V = 1 / P   P V = constant   V = Volume   P = Pressure

For O_2 at 0° C the constant is approximately 0.7

1 L   **1/2**   **2 L**
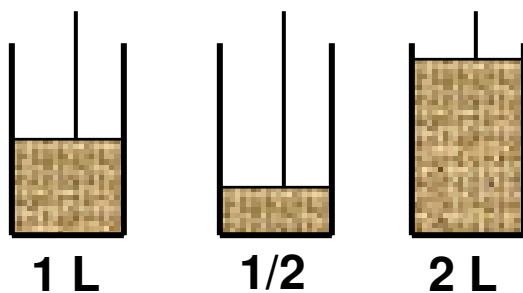
A more useful Boyles Law Equation:

P₁ V₁ = P₂ V₂   Calculate volume when the pressure changes →   **V₂ = V₁ P₁ / P₂**

PROBLEM   50.0 L of O_2 at 15.7 atm at 21°C. What is the volume at 1.00 atm and 21° C?

V₂ = V₁ P₁ / P₂ = 50.0 L * [15.7 atm / 1.00 atm] = **785. L**
Note: The Volume vs Pressure rations [numbers in brackets]

Example 5.2 12.0 dm³ at 98.9 kPa → 119.0 kpa. What is the new volume?

V₂ = V₁ P₁ / P₂ = 12.0 dm³ * [98.9 kPa / 119.0 k Pa] = **9.97 dm³**

Note: The constant and these equations are **approximately** true at low to moderate pressure

CHARLES LAW   A gas contracts when cooled and expands when heated.
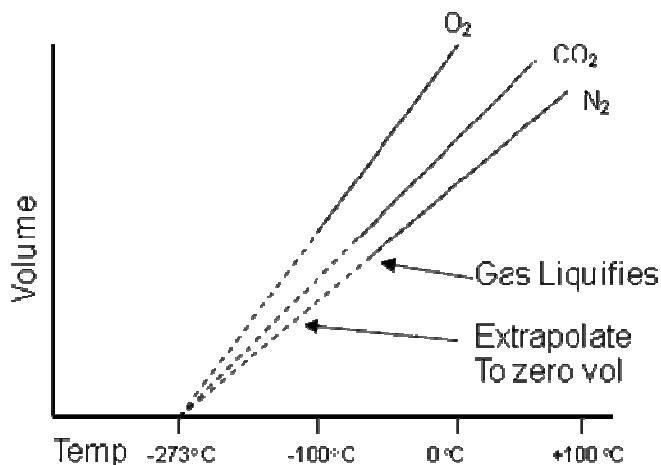
Plot the volume of a gas at various temperatures. At low temp, the gas liquifies.

Extrapolate the values to zero volume and the temp is **– 273.15 degrees C** [or 0 deg K] !

Plot shows the volume of a gas at constant pressure is directly proportional to the ABSOLUTE temp [$^\circ$ K].

$$V / T = CONST \qquad \text{-or-} \qquad V_1 / T_1 = V_2 / T_2$$

Example 5.3 A 785 L tank of Oxygen is at 1 atm and 21° C. What is the volume at 28° C?

$T_1 = 21^\circ$ C $= 294^\circ$ K $T_2 = 28^\circ$ C $= 301^\circ$ K

$V_2 = V_1 * T_2 / T_1 = 785$ L $* [301^\circ$ K $/ 294^\circ$ K $] = 803.68$ L $= $ **804 L**

Note: Watch the ratio of TEMPS. Temp goes up, Vol goes up. Temp goes down, Vol goes down

COMBINED GAS LAW

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

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

Example 5.4 C, H, N analysis uses hot CuO to oxidize organic compounds to CO_2, H_2O and convert Nitrogen in nitrogen containing compounds to N_2. A 39.8 mg sample of caffeine gave 10.1 cm^3 of Nitrogen [N_2] at 23° C and 746 mm Hg. What is the volume at **STP $\rightarrow$ 0° C and 760 mm Hg**?

$T_1 = 23^\circ$ C $+ 273 = 296^\circ$ K $T_2 = 0^o$ C $+ 273 = 273^\circ$ K

The book sets up the data as below. I'd recommend that you also do it this way:

$V_1 = 10.1$ cm^3 $P_1 = 746$ mm Hg $T_1 = 296^\circ$ K

$V_2 = ?$ $P_2 = 760$ mm Hg $T_2 = 273^\circ$ K

Then plug the data into the equation:

$V_2 = V_1 * [P_1 / P_2] * [T_2 / T_1] = 10.1$ $cm^3 * [$ 746 mm Hg $/$ 760 mm Hg $] * [273^\circ$ K $/ 296^\circ$ K$] = $

9.1436 $cm^3 = $ **9.14 cm^3**

CONCEPT CHECK: 5.2   We have a 10.0 L flask at 20 atm.

        A. Decrease the temp by 10° C   What happens?

        B. Increase the volume   What happens?

        C. Heat it by 10° C   What happens?

        D. Decreast the volume   What happens?

AVOGADRO's LAW: Equal volume of any 2 gases at the same temp and pressure contains the same number of molecules!

Molar Gas Volume: The volume of 1 mole of a gas at STP [0° C & 1 atm] = **22.4 L / Mole** and contains 6.02×10^{23} molecules of the gas.

***** IDEAL GAS LAW *****   **P V = n R T**

P = Pressure   V = Volume   n = moles = g / Mw   R = constant   T = Temp ^{o}K

    R =   **0.082058**   **L atm / K Mole**   *********
                              kg m^2 / s^2 K Mole
                              k Pa dm^3 / K Mole
             1.9872   **cal / K Mole**   *********

Example 5.6   50.0 L of Oxygen at 21° C and 15.7 atm. Find the number of grams of Oxygen?

    PV = nRT   →   n = g / Mw   →   PV = g R T / Mw

    g = P V Mw / R T = 15.7 atm * 50.0 L * 32.00 g/Mole / 0.082058 L atm / K Mole * 294° K =

    g =   **1.04 x 10^3 g Oxygen**

GAS DENSITY   Density = Mass / Volume = grams / Liter

    BUT Volume changes with temperature, so the gas density changes with temperature!

Example 5.7   What is the density of O_2 in g/L at 25° C and 0.850 atm?

    # Note the density above is in grams per Liter, so the volume is One Liter

    PV = g R T / Mw

    g = P V Mw / R T = 0.850 atm * 1 L$^\#$ * 32.00 g/Mole / 0.082058 L atm / K Mole * 298° K =

    g =   **1.11 g**

    Density = mass / vol = g / L = 1.11 g / 1 L = **1.11 g/L**

Example 5.7   Use Chlorine instead of Oxygen?

Example 5.7   Calculate the density of Helium in g/L at 21° C and 752 mm Hg?

Example 5.7   Air is 1.188 g/L. What is the difference between air and Helium?

DENSITY AND MW   The density of a gas is proportional to the Mw [see equations above].

Bromine vs Air → Bromine is much denser.

Helium vs Air → Helium balloons float.

This is how the Mw of gases was determined in the olden days!

Pg 192 Halothane at 71° C and 768 mm Hg has a density of 7.05 g/L. What is the Mw?

$$PV = n R T \rightarrow P V = \frac{g R T}{Mw}$$

$$Mw = g R T / PV = \frac{7.05 \text{ g} * 0.082058 \text{ L atm} / \text{K Mole} * 344° \text{ K}}{[760 \text{ mm Hg} / 733 \text{ mm Hg}] / \text{atm} * 1^{\#} \text{L}} = 191.937 = \textbf{192 g / Mole}$$

Example 5.8 200 ml flask contains a gas with a mass of 0.970 g, the temp Is 99° C and pressure 733 mm Hg.

What is the Mw?

$$Mw = g R T / PV = \frac{0.970 \text{g} * 0.082058 \text{ L atm} / \text{K Mole} * 372° \text{ K}}{[733 \text{ mm Hg} / 760 \text{ mm Hg} / \text{Atm}] * 0.2 \text{ L}} = 153.5 = \textbf{154 g/Mole}$$

CONCEPT CHECK: 5.3:   Three 1 Liter Flask contain Helium [Mw 4.00], Argon [Mw 39.95], and Xenon [Mw 131.20] at 878 mm Hg.

A. Which has the most atoms of gas?

B. Which flask has the greatest density?

C. The Helium flask is heated, the Argon cooled. Which of the three has the highest pressure?

D. The Helium flask is cooled, the Xenon heated. Which has the lowest pressure?

E. Which of the three flasks has the most moles of gas?

Pg 195 $2 KClO_3 \rightarrow 2 KCl + 3 O_2$ With an MnO_2 Cat and heated.

0.0100 moles of Potassium Chlorate is heated at 298° K and 1.02 atm.

What's the volume of Oxygen generated?

1st calculate the moles of oxygen $\dfrac{0.0100 \text{ Mole } KClO_3}{2 \text{ Moles } KClO_3} = \dfrac{x}{3 \text{ Moles } O_2}$ x = 0.0150 Mole

Then: **PV = nRT**

V = nRT / P = 0.0150 Mole * 0.082058 L atm / K Mole * 298° K / 1.02 atm = 0.3595 = **0.360 L**

Example 5.9 Auto air bags use Sodium Azide to fill the bag. How many grams of Sodium Azide are required to generate 75.0 L of Nitrogen at 25° C and 748 mm Hg?

$6 NaN_3 + Fe_2O_3 \rightarrow 2 Na_2O + 2 Fe + 9 N_2$ (gas)

   ?   **75.0 L**

First – determine the number of moles of Nitrogen gas at STP.   **PV = nRT**

n = PV / RT = [748 mm Hg / 760 mm Hg / atm] * 75.0 L / 0.082058 L atm / K Mole * 298° K =

        3.02 Moles

Based on the reaction where it takes 6 moles of NaN_3 to 9 moles of N_2, calculate the moles of NaN_3:

$$\frac{3.02\ \text{Moles}\ N_2}{9\ \text{Moles}\ N_2} = \frac{x}{6\ \text{Moles}\ NaN_3} \qquad x = 2.01\ \text{Moles}\ NaN_3$$

2.01 Moles * 65.01 g / Moles = **131 g NaN_3**

Example 5.9 How many Liters of chlorine gas are obtained from reacting 9.41 g HCl at 40° C and 787 mm Hg

$$2\ KMnO_4\ +\ 16\ HCl\ \rightarrow\ 8\ H_2O\ +\ 2\ KCl\ +\ 2\ MnCl_2\ +\ 5\ Cl_2$$

DALTONS LAW OF PARTIAL PRESSURE   The sum of the partial pressure of all of the different gasses in a mixture equals the total pressure of he mix:

$$P_t\ =\ P_a\ +\ P_b\ +\ P_c\ +\ ...$$

The mole fraction of a component gas equals the fraction of moles of that component in the total moles of the gas mixture   Mole Fract$_a$ = n_a / n_t = P_a / P_t

Ex 5.10   1.00 L of dry air at 786 mm Hg contains 0.925 g of Nitrogen at 25° C. What is the partial

pressure of the Nitrogen? What is the mole fraction of the Nitrogen?

   PV = nRT   We need to calculate the pressure of the Nitrogen in the 1.00 L mixture.

   P = nRT / V = [0.925 g / 28.0 g/Mole N_2] * 0.082058 L atm / K Mole * 298° K / 1.00 L = 0.808 atm

   0.808 atm * 760 mm Hg / Atm = **614 mm Hg pressure of the Nitrogen**.

   Mole Fraction of the Nitrogen = 614 mm Hg / 786 mm Hg = **0.81**

GASES OVER WATER
As a gas is bubbled through water, it picks up some water vapour per the chart.

Temp °C	0	10	15	27	29	32	23	25	27	30	40
Pressure mm Hg	4.6	9.2	12.8	14.5	165	18.7	21.1	23.8	26.7	31.8	55.3

Ex 5.11   Hydrogen is collected over water from the following reaction:

$$2\ HCl\ +\ Zn\ \rightarrow\ ZnCl_2\ +\ H_2$$
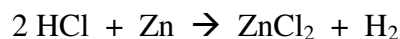
   How much gas is collected over water, if 156 ml of gas is collected at 19° C and 769 mm Hg. What is the mass of H_2 collected?
   1. Find the partial pressure of the Hydrogen. $P_t = P_h + P_w$.

   $P_h = P_t – P_w$ = 769 mm Hg – 16.5 mm Hg = **752 mm Hg for only the Hydrogen**

   2. Determine the number of Moles of Hydrogen: **PV = nRT**

   n = PV / RT = $\dfrac{[\ 752\ \text{mm Hg} / 760\ \text{mm Hg} / \text{Atm}\ ] * 0.156\ L}{0.082058\ L\ \text{atm} / K\ \text{Mole} / 293°\ K}$ = **6.42 x 10^{-3} Mole**

   Mass of Hydrogen = 6.42 x 10^{-3} Mol * 2.02 g/Mole = **1.30 x 10^{-2} g H_2**

Example 5.11 Oxygen can be prepared from heating Potassium Chlorate with a catalyst. How many moles of

O_2 is obtained from 1.300 g Potassium Chlorate, collected over water at 23° C at 745 mm Hg and what volume

would it occupy?

KINETIC THEORY

Kinetic Energy = Energy associated with an object in motion = $E_k = \frac{1}{2} m V^2$

POSTULATES: [Test Questions can be how these relate to PV=nRT]

1. **Gases are composed of molecules whose size is negligible compared with the average distance between them.**

2. **Molecules move randomly in a straight lines in all directions and various speeds.**

3. **The forecs of attraction or repulsion between two molecules of a gas are negligible, except when they collide.**

4. **When molecules collide, the collisuions are elastic [no energy is lost].**

5. **The average Kinetic Energy of a molecule is proportional to the absolute temperature.**

Discuss these various variable and their relationships to PV = nRT

MOLECULAR SPEED

The RMS Molecular speed, the average molecular speed, is equal to the speed of a molecule having the average kinetic energy.

Root-mean-square (rms) molecular speed = $u = \sqrt{\dfrac{3\,R\,T}{M_m}}$ R = Gas Constant, T = Temp ^{o}K, M_m = Molar Mass

Note: The gas with the higher M_m has the lower speed!

Hydrogen at 20° C has a speed of 1.90×10^3 m/s => 4000 mph! – **Prove it!**

Example 5.12 Oxygen in a cylinder is at 21° C and 15.7 atm. What is the rms speed of the molecules:

$u = [\ 3 * \textbf{8.31 kg m}^2\textbf{/ s}^2\textbf{ K Mole} * 294° \text{ K} / 32.0 \times 10^{-3} \text{ kg / Mole}\ ]^{1/2} = \textbf{15.1 m / s [Note book says 479]}$

DIFFUSION / EFFUSION

Diffusion is the process whereby a gas spreads out through another gas to occupy the space uniformly.

Effusion is the process where a gas flows through a hole in a container.

Grahm's Law of Effusion: The rate of effusion of gas molecules from a particular hole is inversely proportional to the square root of the molecular weight of the gas at constant T & P.

Rate   *is proportional to* $1 / [\text{ Mw }]^{1/2}$

Example 5.13 Calculate the ratio of effusion rates of molecules of carbon dioxide vs sulfur dioxide in the same container through a small hole at the same T and P.

$$\frac{\text{Rate of effusion of } CO_2}{\text{Rate of effusion of } SO_2} = \sqrt{\frac{64.1 \text{ g/mole } SO_2}{44.0 \text{ g/mole } CO_2}} = 1.21$$

Example 5.13   If it takes 3.52 seconds for 10.0 ml of helium to effuse throught a hole in a container at T and P, how long will it take 10.0 ml of oxygen to effuse through the same hole?

VAN DER WAALS EQUATION for pressure

Previous we assumed gases obeyed certain Postulates. Two of these may not be true at higher pressures:

1. The volume of space occupied by a molecule of gas may not be negligable.   ($P + n^2 A / V^2$)

2. The attraction between molecules are weak.   ($V - n B$)

$\quad$ ($P + n^2 A / V^2$) * ($V - n B$) = $n R T$ [similar to $P V = n R T$]

A & B are listed in the text book for several gases.

Example 5.14 As an ideal gas, 1.000 moles of Sulfur Dioxide at 0.0^o C occupies 22.41 L at 1 atm. Calculate the pressure using Van Der Waals Equation

$\quad$ **P = [n R T / (V – n B)] – [n² A / V2]**   n = 1 Mole   R = 0.08206 L atm / K Mole

$\qquad\qquad$ T = 273.2° K   V = 22.41 L   b = 0.05679 L/Mole   A = 6.865 L^2 atm/Mole^2

$\quad$ **P = 0.989 atm**

See page 213 for "A Chemist looks at the Carbon Dioxide Gas and the Greenhouse Effect.

Practice Questions:

Review Questions	**P 216**	**5.9, 5.10**
Concept Questions	**P 217**	**5.23, 5.25, 5.27**
Units of Pressure	**P 218**	**5.31**
Empirical Gas Laws	**P 218**	**5.33, 5.35, 5.39, 5.41, 5.45**
Ideal Gas Law	**P 218**	**5.49, 5.51, 5.53, 5.57, 5.59, 5.61, 5.63**
Stoichiometry with GV	**P 219**	**5.67, 5.69**
Gas Mixtures	**P 220**	**5.75, 5.77, 5.81,**
Molecular Speeds, Effusion	**P 220**	**5.83, 5.85, 5.89, 5.91**
Van Der Waals Eq	**P 220**	**5.95**
General Problems	**P 221**	**5.101, 5.103, 5.109, 5.111**