**CHEM 101: LECTURE NOTES: CHAPTER 12 "THE GASEOUS PHASE OF MATTER"**

## *PROPERTIES OF GASES*

**a. Low density (the lowest of all phases of matter)**

   **Units of density = g/L (usually) as opposed to g/mL for liquids or solids**

**b. Mostly empty space: large distances exist between atoms/molecules; hence, gases are easily compressed.**

**c. Atoms and molecules of gases travel at high velocities**

   $H_2$ = 1600 m/sec = 1 mi/sec

**d. Diffuse (move from area of greater concentration to lesser concentration) readily and rapidly**

**e. Effuse: move through small pores**

**Example: A helium balloon eventually sinks to the ground due to this.**

**f. 1 mole of a gas at STP ($0^o$C and 1 atmosphere pressure) = 22.4 L**

## *The Kinetic-Molecular Theory*

**This is used to explain the behavior and properties of gases.**

1. Molecules/atoms of a gas are in constant, random motion.

2. Gases consist of tiny (submicroscopic) particles.

3. Volume of a gas = mostly empty space; large distance between particles

4. Gas particles move in straight lines and have no attraction for one another.

5. Gas particles collide frequently; these collisions result in no loss in energy (termed *perfectly elastic).*

6.  The higher the temperature, the greater the average kinetic energy.

> Kinetic Energy (K.E.) = ½ mv$^2$, where m = mass of particle (in kg), v = velocity (meters/sec).  These units when multiplied give units of joules.

In the kinetic-molecular theory, gases are assumed to be ideal.  This means that the volume changes as expected with temperature (gives a linear or straight line relationship).

*NO GAS IS IDEAL:  ALL GASES ARE REAL.*

> **Real (actual) gases liquefy at low temperatures and high enough pressures.**

## Graham's Law of Effusion

The greater the molar mass (atomic mass) of a gas, the slower it effuses.

This is an indirect relationship. (Increase one, decrease the other)

Equation:  $\dfrac{\text{rate of effusion of A}}{\text{rate of effusion of B}} = \dfrac{\sqrt{\text{molar mass of B}}}{\sqrt{\text{molar mass of A}}}$

## Pressure

**Pressure is the force per unit area an object exerts.**

$$P = \frac{F}{A} \text{ where F is the force and A is the area.}$$

**For gases, pressure is due to collisions of particles with a surface; the higher the temperature, the more collisions per unit time and greater the pressure.**

**Pressure (P) also depends on the number of particles in a gas sample and their temperature.**

*As particles increases, P increases; as T increases, P increases*

*(assuming the volume of the sample to be constant)*

Pressure can have the following units: torr, mm Hg, kPa, bar, atm, in. Hg, $lb/in^2$

1 atm (atmosphere) is the same as all the below:

**1 atm = 760 torr = 760 mm Hg = 101.325 kPa = 1.013 bar =**

**29.92 in Hg = 14.7 lb/in²**

The instrument used to measure P is a barometer.

The P we are exposed to due to the atmosphere pushing down is termed atmospheric P. Normal atmospheric P at sea level = 1 atm.

As one goes up in elevation, P decreases as less gas is present to exert a force.

NOTE: P is exerted on all sides of an object uniformly INCLUDING THE BOTTOM.

## Conversion Problems Examples

**1. Atmospheric P = 768 mm Hg, how many atm is this?**

**760 mm Hg = 1 atm**

**768 mm Hg x 1 atm          = 1.01 atm**
                **760 mm Hg**

2.  30 lb/in$^2$ = ? kPa

30 lb/in$^2$ x $\underline{101.325 \text{ kPa}}$   = 206.79 kPa
            14.7 lb/in$^2$

## *Gas Laws Relating P, V, and T*

1.  BOYLE'S LAW

First stated by Robert Boyle:  the P of a gas is indirectly related to the V of a gas and vice versa *IF* T is constant.

   ↑P   ↓V if T = constant and vice versa


  Boyle's Law can be written mathematically
                      as
         $P_1V_1 = P_2V_2$ if T = constant

Example:  What V will 200 mL of $CO_2$ at 800.00 mm Hg occupy at 750.00 mm Hg?

Boyle's Law states as ↓ P  ↑ V.  P decreases in this example from 800 to 750 mm Hg. So V ↑.

New V = initial V x *P ratio* greater than 1 in this example so V increases.

*New V = 200 mL x (800 mm Hg/750 mm Hg) = 213.33 mL of carbon dioxide*

**Example 2:  What V will 200 ml of $H_2$ at 750 mm Hg occupy at 800 mm Hg?**

# Boyle's Law :  ↑ P     ↓ V
# P increases from 750 to 800 mm Hg.   So V ↓.

**New volume = initial V x P ratio less than 1.**
**New volume =** *200 ml x (750 mm Hg/800 mm Hg) = 187.5 mL hydrogen*


**<u>Example 3</u>:  What will the P be if a gas initially is at 100.25 kPa has a V of 35.1 L and its V is changed to 25.6 L?**

**Boyle's Law states:  ↓ V ↑ P.**
**So, V ratio used is > 1.**

**New P = 100.25 kPa x (35.1 L/25.6 L) = 137.45 kPa**

# <u>Charles' Law</u>

**Proposed by French physicist Jacque Charles.**

**He determined experimentally that a gas' volume is dependent upon T if P is constant.  This dependence is a direct relationship.  (↑ T ↑ V)**

*The lowest T possible is termed ABSOLUTE 0.  Absolute 0 would be 0 Kelvin (recall Kelvin has no degree sign).*

**THERE ARE NO GASES AT 0 K AS THEY WOULD LIQUEFY BEFORE REACHING IT!**

**Charles' Law:  As T ↑ V ↑ and vice versa IF the P is a constant.**

**Can be written mathematically as:  $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$**
- **T must HAVE UNITS OF KELVIN NOT CELCIUS!**

**Example 1:** The T of a gas is 100 K; its V is 0.28 mL. If T increases to 200 K, what would its new V be?

T increases from 100 to 200 K. As T ↑ V ↑, so fraction to multiply by would be >1.

New V = initial V x T ratio

New V = 0.28 mL x (200 K/100 K) = 0.56 mL. So, V simply doubles!

**Example 2:** $T_{Ne}$ = 289 K, $V_{Ne}$ = 5.12 x 10$^5$ dm$^3$, $V_{final}$ = 3.89 x 10$^4$ dm$^3$
$T_{final}$ = ?

V decreases from 5.12 x 10$^5$ dm$^3$ to 3.89 x 10$^4$ dm$^3$. As V ↓ T ↓, so use ratio less than 1.

New T = 289 K x 3.89 x 10$^4$ dm$^3$/5.12 x 10$^5$ dm$^3$ = 21.96 K

# Gay-Lussac's Law

Proposed by French chemist J. L. Gay-Lussac. Relates P and T of a gas if V = a constant.

As T ↑ P ↑ (a direct relationship as in Charles' Law) and vice versa.

Expressed mathematically as: $\dfrac{P_1}{T_1} = \dfrac{P_2}{T_2}$

**T must also be in Kelvin.**

**Example 1:** The P of neon is 2.8 atm at 30$^o$C. What would its P be at –10$^o$C?

T ↓ P ↓, Convert T to Kelvin.

30$^o$C = 303.15 K (273.15 + 30); -10$^o$C = 263.15 K (-10 + 273.15)

Use ratio less than 1 for T ratio as P would decrease.

*2.8 atm x (263.15 K/303.15 K) = 2.43 atm*

**Example 2:  P = 90 torr, T$_1$ = 1,580.15 K, P$_2$ = 175 torr, T$_2$ = ?**
**T  = 1,580.15 K x (175 torr/90 torr) = 3,072.51 K**

## *Standard Temperature and Pressure*

*These are selected T's and P's which provide a reference point when comparing volumes of gases.*

*Standard T = 0.00°C or 273.15 K*

*Standard P = 1 atm or 760 torr or 760 mm Hg or 101.325 kPa or 14.7 lbs./in$^2$ or 29.92 in. Hg*

## *The Combined Gas Law*

**What would happen if you combined Boyle's, Charles', and Gay-Lussac's Laws?  Answer:  The Combined Gas Law!**

**The use of the combined gas law simply involves an additional fraction (ratio) to multiply by (a ratio of P, V, or T) where ' represents a different value.**

**FOR V:  new V = initial V x (P'/P) x (T'/T)**

**FOR T:  new T = initial T x (V'/V) x (P'/P)**

**FOR P:  new P = initial P x (T'/T) x (V'/V)**

**Example 1:  A 41.56 cm$^3$ sample of water vapor at 300.15 K and 600 mm Hg has its T changed to 373.15 K and its P changed to 692 mm Hg.  What would its new V be?**

**T↑ V↑ and P ↑ V↓**

**Hence, multiply by T ratio >1 and P ratio <1.**

**41.56 cm$^3$ x (373.15 K/300.15K) x (600 mm Hg/692 mm Hg) = 44.80 cm$^3$**

**Example 2:** A sample of water vapor at 50$^o$C and 27.80 mm Hg has a volume of 4.83 L. The final P is 31.92 in. Hg and final V = 3.72 L. What would the final T be?

**First, convert T to Kelvin: K = 50 + 273.15 = 323.15 K**

**P↑ T↑ and V↓ T↓**

**323.15 K x ( 31.92 in. Hg/27.80 in. Hg) x (3.72 L/4.83 L) = 285.77 K**

*Example 3:* A 40.15 dm$^3$ volume of chlorine has a P of 301.5 torr at 486.15 K. What would its P be if the final V = 50.28 dm$^3$ at 393.15K?

*V↑ P↓ and T↓ P↓*

**301.5 kPa x (40.15 dm$^3$/50.28 dm$^3$) x (393.15 K/486.15 K) = 194.70 kPa**

# Dalton's Law of Partial Pressures

Each gas in a mixture of gases exerts its own P. The total P is a sum of each of these gases individual P's.

**THE PRESSURE OF ONE OF THESE GASES (A, B, C, etc.) IS TERMED A PARTIAL PRESSURE.**

Mathematically this can be written as: $P_{total} = P_A + P_B + P_C + ..$

If a gas is collected over water, the P of that gas <u>when dry</u> would be: $P_{dry\ gas} = P_{total} - P_{H2O}$. *To find $P_{H2O}$ a table would be used.*

<u>Example:</u> A gas is collected over water. If the P of the gas is 801 torr and $P_{H2O} = 34.1$ torr, What is $P_{total}$?

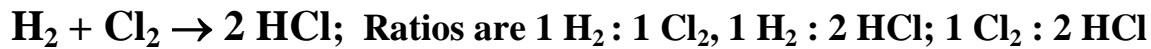$$P_{total} = P_{dry\ gas} + P_{H2O} = 801\ torr + 34.1\ torr = 835.1\ torr$$

**Example 2:** $P_{total} = 931\ torr$, $P_{He} = 201\ torr$, $P_{Ne} = 53\ torr$, $P_{Ar} = 484\ torr$, $P_{Kr} = ?$

$$P_{Kr} = P_{total} - P_{He} - P_{Ne} - P_{Ar} = (931 - 201 - 53 - 484)torr = 193\ torr$$

# Gay-Lussac's Law of Combining Volumes

*The ratios of volumes of reacting gases are small whole numbers if measured at the same T's and P's.*

$H_2 + Cl_2 \rightarrow 2\ HCl$;  Ratios are $1\ H_2 : 1\ Cl_2$, $1\ H_2 : 2\ HCl$; $1\ Cl_2 : 2\ HCl$

Two years later Avogadro stated:  If the T and P are the same, equal volumes of different gases have the same number of molecules or atoms.

**This important law allowed for the following:**

   a. **Explained Gay-Lussac's Law of Combining Volumes and the diatomic nature of some elemental gases**
   b. **Allowed for the determination of molar masses of gases.**
   c. **Helped provide the basis of the Kinetic-Molecular Theory**

## Mole-Mass-Volume Relationships

RECALL: 1 mole of a gas at STP = 22.4 L = molar

mass of gas in grams = $6.022 \times 10^{23}$ atoms or

molecules

**Example 1:**   10 moles of Ne = ? L Ne at STP

10 moles Ne x 22.4L/1 mole Ne = 224 L

**Example 2:**   10 g of Ne = ? L Ne at STP

10 g Ne x 1 mole Ne/20.18 g x 22.4 L/mole =
11.10 L

Example 3:  10 molecules Ne = ? L Ne at STP

10 molecules Ne x 1 mole Ne/$6.022 \times 10^{23}$ molecules Ne x 22.4 L/1 mole Ne = $3.72 \times 10^{-22}$ L

# Density of Gases

Since gases are mostly empty space, they have low densities (units of g/L rather than g/mL for liquid and solids).

See Table 12.4 p. 273 of text for Density of common gases.

To determine the density of a gas at STP, simply divide the molar mass by 22.4 L.

Example 1:  The density of $Cl_2$ at STP = ?

D= m/V = 70.90 g/22.4 L = 3.17 g/L

Example 2:  $D_{Iodine}$ = ?  D = 253.8 g/22.4 L = 11.33 g/L

Example 3:  The molar mass of a gas is 64.07 g/mole.  What is its identity and density?  Possible gases = Ne, HCl, $SO_2$

Answer:  $SO_2$ , D = 64.07 g/22.4 L = 2.86 g/L

# The Ideal Gas Law

This allows for the determination of moles or V of a gas at ANY T and P (NOT just STP).

## *Equation*:  PV = nRT, where P = pressure, V = volume, n = moles, R = a constant, and T = temperature of sample

**NOTE:  The units of the constant R depend upon units of P, V, and T used.**

One possibility: $R = \dfrac{PV}{nT} = \dfrac{1\,atm \times 22.4L}{1\,mole \times 273.15K}$ = 0.0821 atm-L/mol-K

Another possibility: $R = \dfrac{760\,torr \times 22.4L}{1\,mole \times 273.15K} =$

*62.32 torr-L/mol-K or 62.32 mm Hg-L/mol-K*

Example 1:   A gas at 20°C and 2.8 atm has a V or 10.8 L.   How many moles of the gas are present?

Solving for n, $n = \dfrac{PV}{RT} =$

(2.8 atm*10.8 L)/(0.0821 atm-L/mol-K*293.15 K) = 1.26 moles

Example 2:  P = 91 torr, n = 0.25 moles, T = 101.15 K, V = ?

$V = \dfrac{nRT}{P} =$ (0.25 moles*62.32 torr-L/mol-K*101.15 K)/91 torr

## V = 17.32 L

# Volume-Volume Calculations

At constant P and T when gases react in order to convert from one quantity to another, the mole ratios can be used in the balanced equation for that reaction.

**Example:** $2\ NO_2 \rightarrow 2\ NO + O_2$

Suppose 4 L of nitrogen dioxide decompose. How many L of oxygen are produced?

$$4\ L\ NO_2 \times \frac{1\ L\ O_2}{2\ L\ NO_2} = 2\ L\ O_2$$

# REAL GASES

*Remember that there is no such thing as an ideal gas. An ideal gas serves as a model for gas behavior. Gases do not behave as expected all the time because of intermolecular forces of attraction between their molecules and atoms as well as liquefying at high enough pressures and/or low enough temperatures.*