

\*\* If you need any help or more examples email me or find me or go to student hours. ☺

## I. Gases

- Kinetic Molecular Theory (KMT): How we model (predict) the behavior of gases in our calculations.
- A gas that follows the assumptions of KMT is called an ideal gas.
- Many gases don't follow KMT, those gases are called real gasses.
- Assumptions of KMT:
  1. Gases consist of a large number of molecules in constant, random motion.
  2. The volume (size) of molecules is negligible (irrelevant) compared to the total volume of the container.
  3. The attractiveness and repulsive forces between molecules is negligible.
  4. Collisions of gases are perfectly elastic (no sticking together, no energy lost)
  5. The average Kinetic Energy (KE) of the molecules is proportional to the temperature of the sample (definition of temperature)
- At low temperature and high pressure the assumptions of KMT stop working (#2 and #3 are the most important)
- Relating Pressure (P) to KMT
  - As long as the volume is held constant, pressure is dependent on only two things:
    1. Temperature (avg. KE)
    2. # of molecules
  - Pressure is defined as the force exerted over and area ( $P = F/a$ )
  - What are the only two ways to increase the force?
    1. Hit the sides more often (increasing the number of collisions/molecule)
    2. Hit the sides harder (increasing KE/T)
  - Pressure doesn't depend on the mass of the molecule, if it did all the gas laws wouldn't work

## II. Gas Laws

- 3 Gas Laws
  - Boyle's Law: relates Pressure and volume.  $V \propto \frac{1}{P}$  (volume is inversely related to pressure)
  - Charles Law: relates volume and temperature.  $V \propto T$  (volume is directly related to temperature)
  - Avogadro's Law: relates volume and amount (moles).  $V \propto n$  (volume is directly related to amount)
- Pressure Conversions
  - 1atm= 760 mmHg = 760 torr = 101.352 kPa
  - Pascals (Pa) is the SI unit for pressure
  - Atmospheres (atm) is the most common SI Unit.
- Temperature conversions: ALWAYS CONVERT TO KELVIN (K)!!!!
  - $K = ^\circ C + 273.15$
  - Kelvin is always positive because absolute zero (0K) is the lowest possible temperature.
- Standard temperature Pressure (STP)
  - An arbitrarily chosen temperature and pressure used to compare gases.

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- SI units: **273 K** and 101.325 kPa
- Non-SI units: °C and **1 atm**.
- **BOLDED**: what we use.
- If you have 1 mole of gas (any gas) if the gas is at STP it occupies a volume of 22.4 L  
1 mol. = 22.4 L @ STP
- Combined gas law
  - Combines the three individual gas laws, and is used to determine what the new conditions will be after a change.
  - $\frac{P_1V_1}{T_1n_1} = \frac{P_2V_2}{T_2n_2}$
  - **TIP**: Cancel out the constants.
  - **TIP**: Units don't matter as long as they are the same on both sides of the equation. (exception: temperature must be K (Kelvin))

### III. Ideal Gas Law

- $PV=nRT$ 
  - P- pressure (units- atm)
  - V- volume (units- L)
  - n- number of moles (units- moles)
  - R- gas constant (Units- L\*atm/mol.\*K)
  - T- temperature (units- K)
- There are many R's depending on the units (for this class:  $R= 0.0821 \text{ L*atm/mol.*K}$ )
- We use the ideal gas law to calculate all the variables of a gas.  
Ex 1) What is the volume occupied by 0.118 moles of He at a pressure of 0.97 atm. And a temperature of 305 K?

$$PV=nRT$$

$$V = \frac{nRT}{P} = \frac{(0.118 \text{ mol.})(0.0821)(305\text{K})}{0.97\text{atm.}} = \mathbf{3.0L}$$

**TIP: solve for the desired variable before solving for the answer.**

Ex 2) Would the volume be different (referring to the previous problem) if the gas were Ar instead?

No, the identity and mass of a molecule doesn't change the ideal gas law.

Ex 3) How many moles of air are in a 0.5L breath on top of Mt. Everest if the pressure is 0.33 atm and the temperature is 254 K?

$$PV=nRT$$

$$n = \frac{PV}{RT} = \frac{(0.33\text{atm.})(0.5\text{L})}{(0.0821)(254 \text{ K})} = \mathbf{0.008 \text{ mol of air}}$$

### IV. Molar Volume

- We know 1 mole of any gas occupies 22.4 L at STP, but why?
  - Ideal gas law:  $V = \frac{nRT}{P} = \frac{(1 \text{ mol})(0.0821)(273\text{K})}{1 \text{ atm}} = 22.4 \text{ L}$
- Example:  
Assuming ideal behavior which gas occupies the largest volume at STP?  
**(a) 1g of H<sub>2</sub> - lowest molar mass**

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(b) 1g of O<sub>2</sub>

(c) 1g of Ar

A, because H<sub>2</sub> has the lowest molar mass so it has the most moles. The more moles you have, the larger the volume.

#### V. Density of Gasses

- Density= mass/volume. The density of a gas at STP can be defined as D= mass/volume.
- For gases the units are g/L
- Examples

Ex 1) what is the density of helium at STP?

$$d_{\text{He}} = \frac{4 \text{ g/mol}}{22.5 \text{ L/mol}} = \mathbf{0.179 \text{ g/L}}$$

Ex 2) what is the density of nitrogen gas at STP? (remember diatomic elements: H,N,F,O,I,Cl,Br)

$$d_{\text{N}_2} = \frac{28.02 \text{ g/mol}}{22.4 \text{ L/mol}} = \mathbf{1.25 \text{ g/L}}$$

- The density of a gas is directly proportional to its molar mass. (i.e. the larger the mass the more dense it is!)
- Less dense gases float on top of more dense gases (air= 28.8 g/mol)
- What if the gas isn't at STP?

$$\text{density} = \frac{\text{pressure} * \text{mass}}{\text{gas constant} * \text{temperature}} \rightarrow d = \frac{PM}{RT}$$

- We can use the above equation to do the following
  - (1) Calculate the density of a gas under any conditions
  - (2) Calculate the molar mass of an unknown gas just by measuring its density.

#### VI. Gas Mixtures and Partial Pressures

- Air is made up of many things (N<sub>2</sub>(78%), O<sub>2</sub>(21%), Ar(0.9%), CO<sub>2</sub>(0.04%))
- Does this make calculating the pressure of a gas sample difficult?  
No, since we assume ideal behavior and pressure is only based upon the collisions (not the type of molecule), calculating total and individual pressures is easy.
- Dalton's Law of Partial Pressures:  $P_T = P_1 + P_2 + P_3 + \dots$

$P_T$ : total pressure  $P_1$ : pressure of gas 1

- $P_1 = n_1 * \frac{RT}{V}$
- $P_T = n_T * \frac{RT}{V}$
- For a given sample the temperature and volume of all gases are the same, since they are under the same conditions
- The partial pressure of a gas is only dependent on the number of molecules of that gas.

- Mole fraction: useful if calculating a partial pressure from the total pressure.

$$X_a = \frac{n_a}{n_T} \quad X_a = \text{mole fraction of A} \quad n_a = \text{moles of a} \quad n_T = \text{moles total}$$

$$P_a = X_a P_T$$

- Since the partial pressure of a gas is solely dependent on the amount of gas, we can use a ratio.

#### VII. Gas Stoichiometry

- Stoichiometry with gases involves converting volume of a gas to moles using:

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1.  $PV=nRT$  (for any conditions)

2. 1 mol. = 22.4 L (if at STP)

- If at STP you can convert from the volume of a to the volume of b just using the mole ratio
- For examples see HW (the really long one ☺)

#### VIII. Percent Yield

- Why do we do stoichiometry?
  - Stoichiometry allows us to predict how much product we should make, which is called the theoretical yield.
- Theoretical yield assumes there is no human error or side reactions
- In the lab we rarely have no error so we often calculate the percent yield:
$$\%Yield = \frac{\text{actual yield}}{\text{theoretical yield}} * 100$$
- Actual yield is the amount actually made (given in the problem)
- When you are asked to calculate the percent yield, do normal stoichiometry and that gives you the theoretical yield. Take the given yield (actual) in the problem and plug both into the formula.

#### IX. Limiting Reactants

- What happens in a chemical reaction, if there is an insufficient amount of one reactant.
- Limiting reactant: the reactant that is completely used up in a chemical reaction (aka what you run out of)
- Excess reactant: the reactant not completely used up in a chemical reaction (aka what you have left over)
- How to solve:
  - Use the moles of each of the reactant to calculate one of the products (convert both reactants to the desired product)
  - The reactant that gave the smaller calculated value of the product is the limiting reactant.
  - The reactant that gave the larger calculated value of the product is the excess reactant.
  - The actual yield of the amount of product is the smaller of the calculated values.
- To calculate excess reactant, convert the maximum yield back to the excess reactant and subtract that from the original amount.