Activity 15

Name ________________________________

Gases

Part I. Pressure, Temperature, Volume and the Gas Laws

Matter can exist in four phases (or states), solid, liquid, gas, and plasma plus a few other extreme phases, like critical fluids and degenerate gases. Generally, as a solid is heated (or as pressure decreases), it will change to a liquid form, and will eventually become a gas. For example, ice (frozen water) melts into liquid water when it is heated. As the water boils, the water evaporates and becomes water vapor. An ideal gas state is determined by the pressure, temperature, amount (mass or moles) and volume. In a scientific manner, we can freeze any two of the four primary properties and study the effect of one remaining property on the other. Pressure may be the only measurement that is unfamiliar, but we work with pressure in balloons, in car tires and combustion engines. Pressure is a measure of the force of molecules hitting a surface within a container and is affected by the energy, speed and number of particles. Pressure has common units where

1 atm (atmosphere of pressure) = 760 torr = 760 mmHg (millimeters of liquid mercury) = 29.9 inHg = 14.7 psi (pounds per square inch (lb/in^2)) = 101 kPa (kilo Pascal (Newton per square meter)), plus others.
Chemists tend to use atm and mmHg or torr most commonly.

1. What quantities pertaining to gases can be measured? How would you measure these quantities, what pieces of equipment would you use?

   Pressure with a barometer and manometer in any of the units listed below
   Volume – using by displacement or using a known volume container – generally in Liters or mL
   Amount either in mass weighed on a balance or in moles calculated using molar mass
   Temperature using a thermometer in Kelvin

2. Which one of the following represents the highest pressure and Temperature?
   1 atm (atmosphere of pressure) = 760 torr = 760 mmHg (millimeters of liquid mercury) = 29.9 inHg = 14.7 psi (pounds per square inch (lb/in^2)) = 101 kPa (kilo Pascals (Newtons per square meter)), plus others. Chemists tend to use atm and mmHg most commonly.

   a) 800 torr $\times \frac{1\ atm}{760\ torr} = 1.05\ atm$

   b) 15 lb/in^2 $\times \frac{1\ atm}{14.7\ \frac{lb}{in^3}} = 1.02\ atm$

   c) 30.00 in Hg $\times \frac{1\ atm}{29.92\ inHg} = 1.00\ atm$

   d) 81.0 cm Hg $\times \frac{1\ atm}{760\ cmHg} \times \frac{100\ cm}{1\ m} \times \frac{1\ m}{1000\ mm} = 1.07\ atm$

   a) 320°C – highest temperature

   b) 205°F $(205\ ^oF - 32)\frac{5}{9}$ = 96°c

   c) 580 K – 273 = 307°C

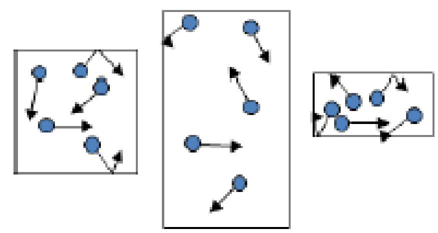
Since gases are highly compressible and will expand when heated, these properties have been studied extensively. The relationships between volume, pressure, temperature and moles are referred to as the gas laws.

We study the relationships or changes that occur to one quantity when two are held constant and only one is changes. For example, when pressure and amount is held constant, we can study the effect on volume when the temperature is changed. The following relationships are studies and called the gas laws.

Model 1. Pressure to Volume relationship at constant Temperature and amount

Trial 2 Trial 3 Trial 1

Trial	Volume	Pressure
1	0.5 L	2.0 atm
2	1.0 L	1.0 atm
3	2.0 L	0.5 atm
4	5.0 L	0.2 atm
5	10 L	0.1 atm



Trial 2 Trial 3 Trial 1

3. In the above diagram, what has occurred to the pressure when the volume increased?
   The particles are hitting the surface of the container less often and therefore the pressure has decreased.

4. What has happened to the pressure when the volume decreased?
   The particles are hitting the surface of the container more often and therefore the pressure has increased.

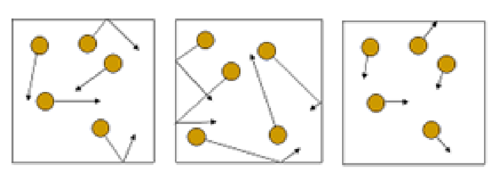
5. If you multiply the pressure and volume, what value do you get for each situation?
   P x V will equal a constant for example in the above 0.5 L x 2 atm = 1 L atm

   or 10 L x 0.1 atm = 1 L atm

6. Can you write an equation between pressure and volume that would reflect the relationship between P and V?
   P x V will equal a constant or $P_1V_1 = P_2V_2$

Model 2: Pressure to Temperature Relationship at constant Volume and amount.

Trial	Temperature	Pressure
1	1200 K	3.0 atm
2	800 K	2.0 atm
3	600 K	1.5 atm
4	400 K	1.0 atm
5	200 K	0.5 atom



Trial 4 Trial 2 Trial 5

7. In the above diagram, what has occurred to the pressure when the temperature increased?

   The particles are hitting the surface of the container more often and will greater energy due to the increase in Kinetic energy of the particles and therefore the pressure has increased.

8. What occurred when the temperature decreased?

   The particles are hitting the surface of the container less often and will less energy due to the decrease in Kinetic energy of the particles and therefore the pressure has decreased.

9. If you multiply the pressure and temperature, what value do you get for each situation? Does this work? What if you divide pressure by temperature?

   P x T is not a constant but P / T is a constant

10. Can you write an equation between pressure and temperature that would reflect the relationship?

$$\frac{P}{T} = constant \text{ or } \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

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Gases

Part II. Ideal Gas Law

The theory by which we define gases is called the Kinetic Molecular theory (KMT). The basic statements of this theory are: All gases at the same temperature have the same kinetic energy but as temperature changes, kinetic energy changes proportionately; particles of gases are significantly far apart, such that, the volume of the individual particle is significant small compared to the volume between the particles; when particles of a gas collide there is no change in the energy of each particle due to its collision and finally, there is no interaction or attraction between particles of gas. If gas particles obey these statements of the KMT, then we can say gases at low pressures, in large volumes and relatively high temperatures obey the ideal gas law,

$p\,V = n\,R\,T$ - *The ideal gas law*

where R is a constant (known as the *gas constant*) that has the value R = 0.08206 L atm K^{-1} mol^{-1}. Appropriate units to use for p, V, n, and T in the ideal gas equation are those used for R above. Thus, the pressure (p) should be in atmosphere, the volume (V) in Liters, the temperature (T) in Kelvin, and the amount of gas (n) should be in moles. Useful conversion factors are

<u>Pressure:</u> 1 atm = 760 Torr = 760 mmHg = 101.3 kPa = 29.92 inHg

<u>Temperature:</u> K = 273 + °C

<u>Volume:</u> 1 L = 1000 mL = 1000 cm^3

1. What is the pressure of 2.50 moles of an ideal gas that occupies 5.0 L at 40°C?
   a. What information do you know? Which variable is unknown?
      n = 2.50 moles P = ?
      V = 5.0 L
      T = 40°C
   b. Are the variables in the proper units? If not, put the variables in the proper units.
      All except Temperature – all temperatures must be in Kelvin
   c. Which relationship contains all the above variables which will allow you to solve for pressure?
      PV = nRT
   d. Determine the pressure. = 12.8 atm

$$P = \frac{nRT}{V} = \frac{(2.50\,mol)\left(0.08206\,\frac{L\,atm}{mol\,K}\right)((40°C + 273)K)}{5.0\,L} =$$

2. What volume will 0.450 moles of Hydrogen gas occupy at 30.4 in Hg and 27.5°C? V = 10.9L

$$V = \frac{nRT}{P} = \frac{(0.450\,mol)\left(0.08206\,\frac{L\,atm}{mol\,K}\right)((27.5°C + 273)K)}{30.4\,in\,Hg \times \frac{1\,atm}{29.92\,in\,Hg}} =$$

STP

Often you will see gas volumes reported at STP *(standard temperature and pressure)*. STP is defined as T = 273 K (0°C) and p = 1 atm. Substitution of these values into the ideal gas law shows that the *volume of 1 mol* of any gas is *approximately 22.4 L at STP.*

3. How many moles of an ideal gas are present in a 5.00 L container at STP?

$$n = \frac{PV}{RT} = \frac{(1\ atm)(5.00\ L)}{\left(0.08206\,\frac{L\ atm}{mol\ K}\right)((0°C + 273)K)} =$$

= 0.223 moles

Gas Density (d) and Molar Mass (M)

Rearranging the ideal gas equation, $pV{=}nRT$, and using the definitions of density, d, and molar mass, MM, gives

$$\frac{n(mol)}{V(L)} = \frac{p\ (atm)}{R\,T(K)}\ \ and\ \ d\left(\frac{g}{L}\right) = \frac{n(mol)}{V(L)} \times MM\left(\frac{g}{mol}\right) = \frac{p \times MM}{R\,T}\ \ or\ MM = d\,\frac{RT}{p}$$

4. 10 grams of an ideal gas occupies 2 L at STP, what is the molar mass of the ideal gas?

$$n = \frac{PV}{RT} = \frac{(1\ atm)(2\ L)}{\left(0.08206\,\frac{L\ atm}{mol\ K}\right)((0°C + 273)K)} = 0.08928\ moles$$

$$\mathcal{MM} = \frac{mass\ (g)}{moles} = \frac{10\ g}{0.08928\ mol} = 112\ g/mol$$

5. What is the volume occupied by 35.4 g of nitrogen gas at 35°C and 735 torr?

$$moles = \frac{mass\ (g)}{\mathcal{MM}} = \frac{35.4\ g}{28.0\ g/mol} = 1.26\ mol$$

$$V = \frac{nRT}{P} = \frac{(1.26\ mol)\left(0.08206\,\frac{L\ atm}{mol\ K}\right)((35°C + 273)K)}{735\ torr\ \times \frac{1\ atm}{760\ torr}} =$$

32.9 L

Since $\frac{pV}{nT} = R$ where R is the ideal gas constant, it follows that

$$\frac{p_1V_1}{n_1T_1} = \frac{p_2V_2}{n_2T_2}\quad \text{The combined gas equation}$$

where the subscript "1" represents one set of conditions, and the subscript "2" represents different set of conditions. More specialized equations may be derived from the above combined gas equation when one or more of the variables is held constant. For example, you can easily derive the familiar equations given below in this manner (convince yourself that this works!):

<u>Boyle's Law</u>: $p_1V_1 = p_2V_2$ (obtained when $n_1{=}n_2$ and $T_1 = T_2$)

$$\underline{\textbf{Charles' Law}}: \quad \frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ (obtained when } n_1=n_2 \text{ and } p_1 = p_2\text{)}$$

$$\underline{\textbf{Avogadro's Law}}: \quad \frac{V_1}{n_1} = \frac{V_2}{n_2} \text{ (obtained when } T_1=T_2 \text{ and } p_1 = p_2\text{)}$$

6. A sample of nitrogen gas, N_2, occupies 3.0 L at a pressure of 3.0 atm. What volume will the gas occupy when the pressure is changed to 0.50 atm and the temperature remains constant?

a. What information do you know? Which variable is unknown?

V_1 = 3.0 L
P_1 = 3.0 atm
P_2 = 0.5 atm

b. Which relationship should be used to solve for the unknown variable?

$$\frac{p_1V_1}{n_1T_1} = \frac{p_2V_2}{n_2T_2} \text{ where moles and temperature are considered to be constant}$$

c. Determine the new volume. a. If the pressure decreases, volume will increase, P and V are inversely proportional

$$V_2 = \frac{V_1P_1}{P_2} = \frac{(6.75\ L)\ (3.0\ atm)}{(0.5\ atm)} =$$

40.5 L

7. A 6.75 L flask contains a fixed amount of gas at 31 °C and a constant pressure. If the temperature is increased to 125 °C, what will the volume of the gas be?

If temperature increases, volume should increase, direct proportionality.

$$\frac{p_1V_1}{n_1T_1} = \frac{p_2V_2}{n_2T_2} \text{ where moles and pressure are considered to be constant}$$

$$V_2 = \frac{T_2V_1}{T_1} = \frac{\big((31°\text{C} + 273)K\big)(6.75\ L)}{(125°\text{C} + 273)K} =$$

5.16 L

8. Many gases are produced during industrial chemical production. If 173 L of the poisonous gas nitric oxide at 300°C and 8809 torr into the atmosphere where the conditions are at STP, what volume would the gas occupy?

$$V_1 = 173L; P_1 = 889torr \times \frac{1atm}{760torr} = 1.17atm; T_1 = 30^o C + 273 = 303K;$$

$$V_2 = ???; P_2 = 1.0atm; T_2 = 273K =$$

$$\frac{p_1V_1}{n_1T_1} = \frac{p_2V_2}{n_2T_2} \text{ where moles are considered to be constant}$$

$$V_2 = \frac{V_1P_1T_2}{P_2T_1} = \frac{173L \times 1.17atm \times 273K}{1atm \times 303K}$$

= 182 L

$$V = \frac{173\ L\ x\ 8809\ torr\ x\ 273\ K}{760\ torr\ x\ 573\ K} = 955\ L$$

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Gases

Part III. Dalton's Law

Model 3. Partial Pressure

Moles of Gas A	Pressure of Gas A (atm)	Moles of Gas B	Pressure of Gas B (atm)	Total Pressure (atm)
2	0.5	0	0.0	0.5
2	0.5	2	0.5	1.0
0	0.0	2	0.5	0.5
2	0.5	4	1.0	1.5
4	1.0	2	0.5	1.5
4	1.0	4	1.0	2.0
6	1.5	6	1.5	3.0

1. In the above model, what occurs to the pressure when the number of moles of A increases?
   As the moles of an individual gas increases, the pressure of the gas increases.

2. Is the pressure similarly affected when the moles of B increases?
   yes

Since $\frac{p_1V_1}{n_1T_1} = \frac{p_2V_2}{n_2T_2}$, when Volume and Temperature are held constant, then $\frac{p_1}{n_1} = \frac{p_2}{n_2}$

Where condition 1 can be the pressure and number of moles of A only and the condition 2 can be the pressure and number of moles of gas B.

3. If the pressure of gas A is independent of the pressure of gas B, what can you say about the total pressure in the container?
   The total pressure in the container is a sum of the individual pressures of each gas present.

Dalton's Law shows that since pressure is directly proportional to the number of gas particles in a container but not dependent on the mass or identity of the gas, then the total pressure of the container should be directly proportional to the total number of moles.

$$p_{total} = The\ sum\ of\ the\ pressure\ of\ each\ gas = \sum p_i = n_{total}\frac{RT}{V}$$

where R, T and V are constant. Simplifying gives

$$\frac{p_{total}}{n_{total}} = \frac{RT}{V} = \frac{p_i}{n_i} \ or \ p_i = \frac{n_i}{n_{total}} p_{total}$$

P_i is the partial pressure of an individual gas, which is proportional to the number of moles of that gas in the mixture. According to Dalton's Law, the total pressure is a sum of the individual partial pressures of each gas in the mixture.

4. What is the partial pressure of oxygen in a mixture both nitrogen and oxygen if the partial pressure of nitrogen is 535 mm Hg and the total pressure is 1.2 atm?

P_{total} = 1.2 atm x 760 mmHg/1 atm = 912 mm Hg

$P_{nitrogen}$ = 535 mmHg; P_{oxygen} = 912 – 535 = 377 mm Hg

5. A 35 L container at a temperature of 25°C contains 3 moles of N_2 and 4 moles of O_2, what will the pressure of the resulting mixture of gases be?

$$n_{total} = n_{N_2} + n_{O_2} = 3 + 4 = 7 \ moles$$

$$p_{total} = \frac{n_{total} RT}{V} = \frac{(7\ mol)\left(0.08206 \frac{L\ atm}{mol\ K}\right)(298\ K)}{35\ L} \qquad = 4.89 \text{ atm}$$

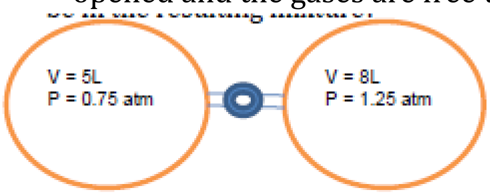
6. What's the partial pressure of carbon dioxide in a container that holds 5 moles of carbon dioxide, 3 moles of nitrogen, and 1 mole of hydrogen and has a total pressure of 1.05 atm?

$$p_{individual} = \frac{n_i}{n_{total}} p_{total}$$

$$p_{CO_2} = \frac{n_{CO_2}}{n_{total}} p_{total} = \frac{5\ mol}{(5 + 3 + 1\ moles)} 1.05\ atm = 0.583\ atm$$

$P_{nitrogen}$ = 0.350 atm; $P_{hydrogen}$ = 0.117 atm

7. Two flasks are connected with a stopcock. The first flask has a volume of 5 liters and contains nitrogen gas at a pressure of 0.75 atm. The second flask has a volume of 8 L and contains oxygen gas at a pressure of 1.25 atm. When the stopcock between the flasks is opened and the gases are free to mix, what will the pressure be in the resulting mixture?



Assume the volume of the stopcock is minimal, the total V = 13 L, solve for moles of N_2 and O_2, assuming a contant temperature (273K)
n_{N2} = 0.75 atm*5L/R*273K = 0.167 mol
n_{O2} = 1.25 atm*8L/R*273K = 0.446 mol

n_{total} = 0.167 + 0.446 = 0.614 mol
V_{total} = 5 + 8 = 13 L
P_{total} = 0.614 mol * R * 273K / 13L = 1.05 atm

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Gases

Part IV. Stoichiometry

Stoichiometry is the relationship of the amounts of compounds that are involved in a chemical change. Mass Stoichiometry is the process of determining the molar and mass relationships between substances in a chemical reaction. Volume stoichiometry is the process of relating the volume and the number of moles. Molarity is defined as the number of moles of a substance dissolved (solute) in a Liter of solution. Finally in the ideal gas equation, the number of moles is related to the pressure, volume and temperature of the gas. The coefficients in a balanced chemical reaction determine the molar ratios between two substances. If we combine these concepts, stoichiometry allows for a simple calculation in which given the quantities can determine the number of moles of substance A which can be used to determine the moles of substance B within a chemical reaction.

$mass \div molar\ mass\ =$
$molarity \times volume =$
$\frac{pressure \times volume}{R \times Temperature} =$

$$mol_A \times \frac{coefficient\ of\ B\ in\ reaction}{coefficient\ of\ A\ in\ reaction} = mol_B$$

1. The reaction of Nitrogen gas added to hydrogen gas produces ammonia gas. What is the balanced chemical reaction?

$$3\ H_2 + 1N_2 \rightleftarrows\ 2NH_3$$

2. According to Avogadro's Law, the moles of a gas are proportional to the volume occupied by the gas. If 10 moles of nitrogen are present in 5 L, what volume of Hydrogen would be needed to react with this volume of nitrogen? Assume the Temperature and pressure remain constant.

$$mol_{N_2} \times \frac{coefficient\ of\ H_2\ in\ reaction}{coefficient\ of\ N_2\ in\ reaction} = mol_{H_2}$$

$$10\ mol_{N_2} \times \frac{3\ moles\ H_2}{1\ mol\ N_2} = 30\ mol_{H_2}$$

$$\frac{p_1V_1}{n_1T_1} = \frac{p_2V_2}{n_2T_2}\quad where\ temperature\ and\ pressure\ are\ considered\ to\ be\ constant$$

$$\frac{V_{N_2}}{n_{N_2}} = \frac{V_{H_2}}{n_{H_2}}$$

$$V_{H_2} = \frac{n_{H_2}V_{N_2}}{n_{N_2}} = \frac{3\ mol_{H_2} \times 5\ L}{1\ mol_{N_2}}$$

3. Nitrogen gas occupying a 10 L vessel at 25°C and 875 torr is mixed with 10.5 grams of hydrogen, how many grams of ammonia will be produced in the reaction?

$$10\ g_{H_2} \times \frac{1\ mole\ H_2}{2\ g\ H_2} = 5\ mol_{H_2} \times \frac{2\ moles\ NH_3}{3\ mol\ H_2} = 3.33\ mol_{NH_3}$$

$$mol_{N_2} = \frac{PV}{RT} = \frac{\left(875\ torr \times \frac{1\ atm}{760\ torr}\right)(10\ L)}{\left(0.08206\frac{L\ atm}{mol\ K}\right)((25°C + 273)K)} = 0.471\ mol \times \frac{2\ moles\ NH_3}{1\ mol\ N_2} = 0.941\ mol_{NH_3}$$

The moles of nitrogen are the limiting reactant; therefore, only 0.941 mol of ammonia will be produced. The mass is the moles x molar mass of ammonia = 16.0 g

4. What volume would the ammonia occupy at STP?

$$V = \frac{nRT}{P} = \frac{(0.941\ mol)\left(0.08206\frac{L\ atm}{mol\ K}\right)(273\ K)}{1\ atm} =$$

21.08 L