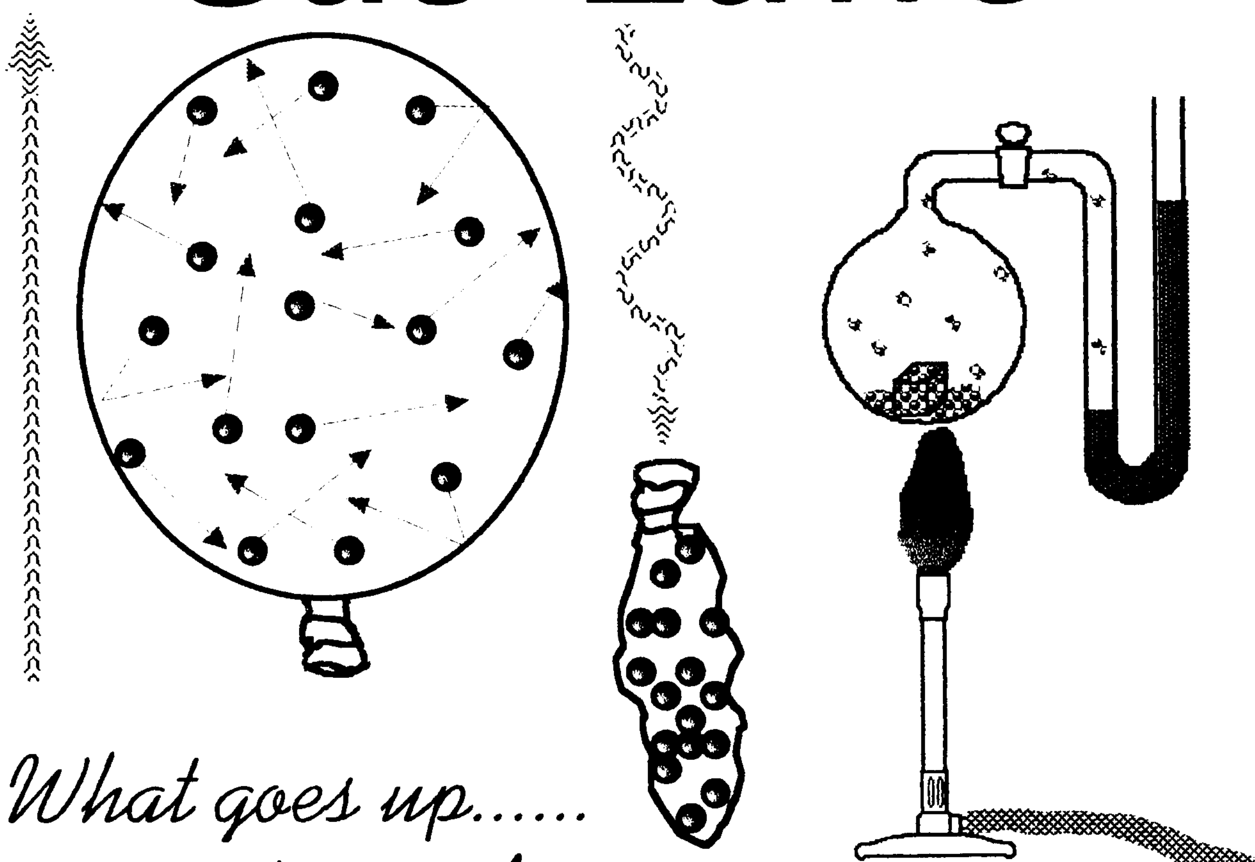
The Gas Laws



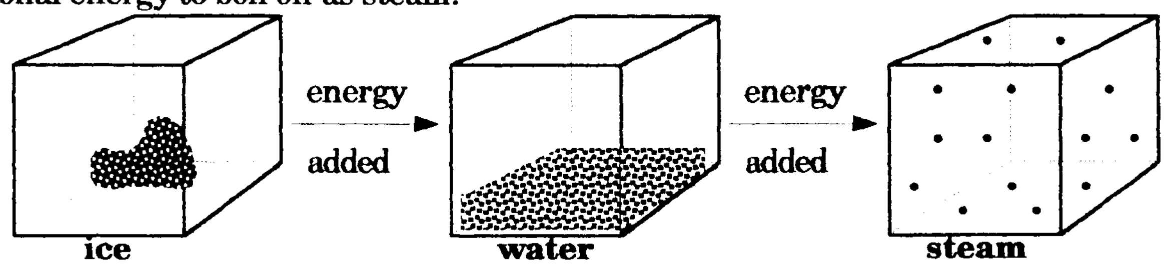
What goes up......
must come down.

Unit 9

Unit 9: Gas Relationships and the Kinetic Molecular Theory

I. A Comparison of Solids, Liquids and Gases

The strength of the attractive forces between the particles of a substance determines the temperature at which melting and boiling occurs. The stronger the attractions, the higher the melting and boiling points. For example, nonpolar butane (C_4H_{10}) has only weak dispersion interaction between its molecules and will boil at 0°C, whereas water with its strong hydrogen bonding will freeze solid at this same temperature! As the temperature of a substance increases, its particles gain kinetic energy and move more rapidly. Phase changes occur when the amount of vibrations and movement become greater than the forces of attraction. In the solid state, the attractive forces are the strongest and the particles are held in fixed positions, having only **vibrational motion**. As the solid melts, the particles gain the freedom to tumble and roll around within the liquid(**rotational motion**), but they still remain in contact with each other. When the boiling temperature is reached, the forces of attraction are no longer strong enough to hold the particles together, and the liquid changes into a gas, causing the particles to expand outwards. The **translational motion** of the gas particles is limited only by collisions with the walls of its container and with other gas particles. The diagrams below illustrate how the particle arrangements vary as ice absorbs enough heat energy to melt and then the water gains additional energy to boil off as steam.



Problem: Compare the structures of ice, water and steam shown above and describe the general characteristics of each in the following chart:

	Solids	Liquids	Gases
shape?			
volume occupied by phase?			
amount & type of particle movement?			
phase density?			
effectiveness of attractive forces?			

II. The Kinetic Molecular Theory and an Ideal Gas

To be able to explain the properties of solids, liquids and gases, chemists use a kinetic molecular model that describes the behavior of the sub microscopic particles in terms of their degree of motion. In the solid and liquid phase, the attractive forces are very strong, which limits the movement of the particles. In the gaseous phase, the kinetic energy of the particles allows them to break away from their attractive forces and act independently of each other. The characteristics of pressure, volume, expansion and diffusion of gases relative to their temperature can be explained using the **kinetic molecular theory** (KMT), proposed by Boltzmann and Maxwell around 1860.

Five assumptions of the KMT to describe an ideal gas are:

- *Molecules of an ideal gas are dimensionless points with no volume.* Since the volume of the gas molecules themselves is very small in comparison to the total volume of gas in the container, this assumption is valid.
- *Molecules of a gas are in constant, random straight-line motion.* The molecules continue in straight line trajectories until they hit another molecule and are deflected. The average distance between two successive collisions is called the **mean free path**.
- *The collisions that occur between molecules or with the walls of the container are perfectly elastic.* That is to say that there is no total kinetic energy lost or gained by the collisions. Energy lost by one molecule in a collision is transferred to another molecule.
- *The average kinetic energy of the gas molecules is proportional to the absolute temperature.* As the temperature increases, the average speed of the molecules increases, as does their kinetic energy. At colder temperatures, the molecules move slower and have less kinetic energy.
- *The molecules of an ideal gas do not exert any attractive forces on each other.* This assumption is valid at temperatures well above the boiling point of the liquid and at low atmospheric pressures. As the temperature drops and/or the pressure exerted on a gas increases, the gas would eventually condense into a liquid.

Mathematically, the kinetic energy of a gas can be determined by the equation,

$\mathbf{KE = 1/2mv^2}$, where m = molar mass and v = velocity of the gas.

Graham's Law of Diffusion

The movement of the particles of one material through the particles of some other material(a medium) is called **diffusion**. The rates of diffusion for different gases in the same medium are different. If the temperature is the same for two gases, a and b, they must have the same average kinetic energy.

Therefore..... $\mathbf{1/2(m_{gas\ a})(v_{gas\ a})^2 = 1/2(m_{gas\ b})(v_{gas\ b})^2}$

Solving algebraically for the velocity ratio.............

$$\frac{v_{gas\ a}}{v_{gas\ b}} = \sqrt{\frac{m_{gas\ b}}{m_{gas\ a}}}$$

This relationship is called Graham's Law of Diffusion.

Problems:

1. Calculate how much faster NO_2 will diffuse compared to Cl_2 under similar conditions.

ANS: NO_2 diffuses 1.24 times faster

2. A chemist wants to separate a mixture of gases containing H_2, O_2 and H_2O vapor by allowing the sample to diffuse through a long series of pipes. After 5 minutes, a high concentration of $H_2(g)$ is detected by a gas chromatograph at the end of the piping. How much time should it take until the chemist detects the O_2 and H_2O vapor?

ANS: 20 min for O_2, 15 min for H_2O

III. Measuring Pressure

Pressure is the force exerted upon some given area. For solids and liquids, the force is the weight of the object caused by its gravitational attraction for the Earth. Mathematically, the pressure is determined by dividing the object's weight by the area over which the weight is distributed. A liquid spreads the pressure of its weight evenly to the walls of the container supporting it. Whereas a solid exerts pressure only in a downward direction over the area upon which it rests.

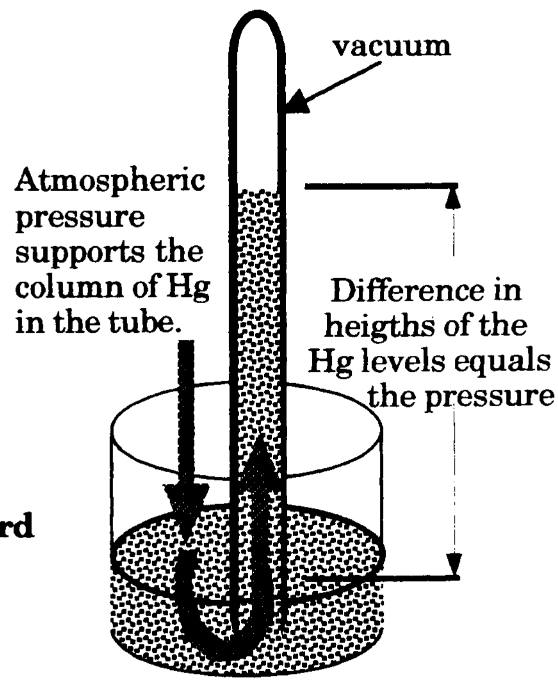
Problem:
Consider the pressure caused by a 2 kg brick with the dimensions of 25 cm x 10 cm x 5 cm. Explain how the pressure will vary depending on which side the brick is resting upon.

ANS: compared to pressure on 25 x 10 side, 2x greater on 25 x 5 side, 5x greater on 10 x 5 side

For gases, the pressure is caused by the molecules striking the walls of its container. Gas pressure can be increased by 3 methods:

- Adding more gas(molecules) to the container, which increases the number of collisions.
- Compressing the given amount of trapped gas molecules into a smaller space. This reduces the mean free path distance, resulting in more collisions.
- Increasing the temperature, causing the molecules to move faster, collide more frequently and with greater force.

The pressure of the gases in the atmosphere can be determined using a barometer, like the one at the right. In 1640, an Italian mathematician/ physicist, named Torricelli, sealed a long tube at one end and filled it with mercury. Then he covered the open end of the tube and inverted it in a beaker filled with Hg. When he uncovered the open end, the pressure of the atmosphere supported the weight of the column of Hg in the tube to a height of about 3/4 of a meter. The level of the mercury will rise and fall as the atmospheric pressure increases or decreases , so the difference in Hg levels can be used as a measure of air pressure. The average atmospheric pressure at sea level, called **standard pressure (SP)**, supports a column of Hg that is 760.mm in height. This value can be expressed in other units of pressure, as well. Equivalent values include 760 torr, 101.3 kPa (kilopascals), 1 atm (atmosphere), 29.9 inches Hg, and 14.7 lbs / in^2 (psi).

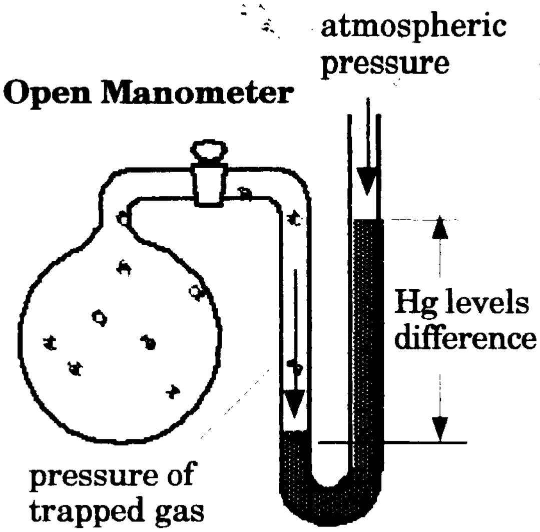


beaker and tube filled with Hg

Standard temperature (ST) is defined as the freezing point of water, which is 0°C.

The barometric pressure will change with weather conditions, partly based on the humidity levels. The atmosphere is 78% $N_2(g)$ and 21% $O_2(g)$ by mass. Since water vapor is less dense than either of these gases, as the humidity level increases, air pressure decreases and wet weather is on the way! High pressure weather systems bring fair skies.

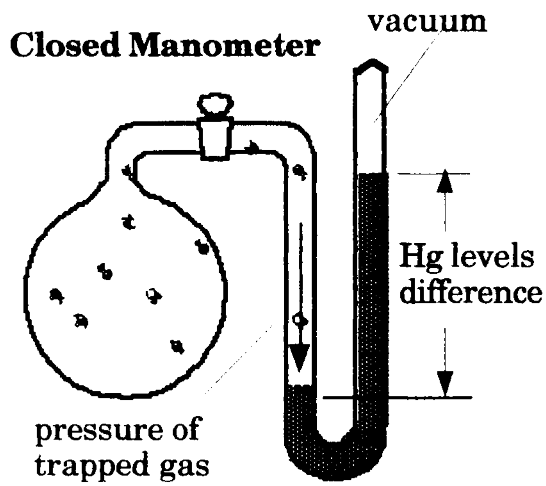
Another device used to measure the pressure of a gas sample in the lab is the manometer.



In the open manometer at the left, the trapped gas pushes down on one side of the mercury in the bent tubing while the atmosphere pushes down on the other side of the tube. The gas pressure equals the atmospheric pressure ± difference in Hg levels.

- If the **Hg levels are equal**, then the pressure of the trapped gas equals the barometer reading.
- If the **Hg level in the open end of the tubing is higher** than in the end with the trapped gas, then the pressure of the trapped gas is greater than the atmospheric pressure. Add the mm difference in Hg levels to the barometer reading.
- If the **Hg level in the open end of the tubing is lower,** then the pressure of the trapped gas is less than the atmospheric pressure. Subtract the mm difference in Hg levels from the barometer reading.

If the barometric pressure in the lab is 755 mm and the Hg level difference in the open manometer above is 65 mm Hg, then the pressure of the trapped gas sample would be equal to 820 mm Hg (Patm + difference in Hg levels)



In the closed manometer at the left, the trapped gas pushes down on one side of the mercury in the bent tubing while there is **no gas** pushing down on the other side of the tube. The sealed end of the tubing contains a vacuum, which exerts no pressure by itself. Only the weight of the Hg column creates any pressure on the side of the sealed tube. Therefore, the pressure of the trapped gas equals the height of the Hg column that it can support(the difference in the Hg levels).

A barometer is also a type of closed manometer, where the "trapped gas" is the Earth's atmosphere.

If the difference in Hg levels in the closed manometer above is 65 mm, then that is also the pressure of the trapped gas sample. Atmospheric pressure has no effect on the sealed gas.

Problems:

1. If the barometer reading in the lab is 745 mm Hg and the difference in Hg levels in an open manometer is 125 mm higher on the side of the tubing open to the air, then what is the pressure of the gas sample? Draw a diagram of the open manometer and solve the problem.

ANS: 870 mm Hg

2. A gas sample in a closed manometer causes a Hg level difference of 190 mm Hg. If this sample was transferred to an open manometer when the atmospheric pressure was 760 mm Hg, what would be the new difference in Hg levels? Draw a diagram of both manometers and solve the problem.

ANS: 570 mm lower on open side

Since there is so much empty space between gas molecules, one gas can readily share the same space as another gas. Each gas behaves independently of the other and adds to the total pressure.

Dalton's Law of Partial Pressure states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of each individual gas.

$$P_{total} = P_a + P_b + P_c \cdots + P_z$$

Since the temperature and volume conditions will be the same for all gases in the same mixture, the partial pressure differences will be due to the different number of moles of each gas that is present. Therefore, *the mole ratio between the gases in the mixture will be equal to the ratio of their partial pressures.*

For example, when $CO_2(g)$ is collected by water displacement during an experiment, there will be some water vapor mixed in with the gas sample. The total pressure of the mixture equals the sum of the partial pressures of the $CO_2(g)$ and $H_2O(g)$ present. The partial pressure of the water vapor increases with temperature, as shown in the chart below:

Temp °C	0	10	20	30	40	50	60	70	80	90	100
Press kPa	0.61	1.23	2.34	4.24	7.38	12.4	20.0	31.2	47.3	70.1	101.3

Sample Problems:

1. What is the partial pressure of a sample of $CO_2(g)$ collected by water displacement at 60°C, if the total pressure of the mixture is 100.0 kPa?

Answer: From the chart above, at 60°C the water vapor pressure is 20.0 kPa. The $P_{total} = P_{CO_2} + P_{H_2O}$ therefore.....100.0 kPa $= P_{CO_2} +$ 20.0 kPa solving for $P_{CO_2} =$ 80.0 kPa

2. If the reaction in this experiment produced 11.0g $CO_2(g)$ collected by water displacement, how many moles of water molecules are present in the mixture?

Answer: Remember that the mole ratio of gases = partial pressure ratio. First, calculate the moles of $CO_2(g)$ that are equal to 11.0g $CO_2(g)$ using the molar mass as a conversion factor:

$$x \text{ mole } CO_2(g) = 11.0\text{g } CO_2(g) \left(\frac{1 \text{ mole } CO_2(g)}{44.0\text{g } CO_2(g)}\right) = 0.25 \text{ mole } CO_2(g)$$

Then determine how many moles of water vapor are needed to have the mole ratio with $CO_2(g)$ equal the pressure ratio:

$$\frac{x \text{ mole } H_2O(g)}{0.25 \text{ mole } CO_2(g)} = \frac{20.0 \text{ kPa } H_2O(g)}{80.0 \text{ kPa } CO_2(g)} \quad \text{solving for } x = 0.063 \text{ mole } H_2O(g)$$

Problems:

1. What is the partial pressure of methane gas collected by water displacement at a temperature of 20.0°C and an atmospheric pressure of 94.4 kPa?

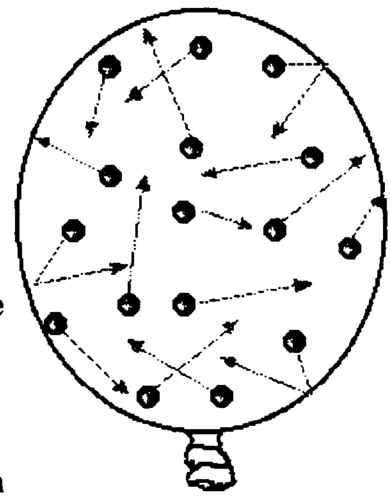
ANS: 92.1 kPa;

2. A mixture of 4.0g H_2 and 16.0g O_2 exerts a pressure of 506 kPa. What is the partial pressure exerted by each individual gas?

ANS: H_2 = 405 kPa; O_2 = 101 kPa

IV. Pressure, Volume and Temperature Relationships

To understand how pressure, volume and temperature changes are inter-related for a fixed number of moles of gas molecules, consider the diagram of the balloon at the right. The balloon is an equilibrium system that has a balance between 2 opposing forces......the pressure of the gas molecules inside of the balloon and the atmospheric pressure on the outside of the balloon. If the temperature of the gas inside the balloon is doubled, the molecules move twice as fast and the number of collisions with the inner balloon walls are doubled. To alleviate this increase in internal pressure, the balloon expands. The volume will increase until it is twice the original size. At this point, the gas pressure will once again equal the atmospheric pressure, because even though the molecules are moving faster, they have a greater distance between collisions. When the equilibrium is re-established, the number of molecular collisions inside the balloon will be the same as before heat was added.



A. Boyle's Law: volume versus pressure

Definition: If the temperature of a definite amount of dry gas remains constant, then the volume of the gas varies inversely with the pressure exerted on it.

Simply stated, this means that as the pressure exerted on a gas is increased, the volume will decrease; and as the pressure is decreased the volume increases. When expressed as a mathematical relationship: $P_1V_1 = P_2V_2$ **if T is constant**

Explanation: As the pressure exerted on a gas is increased, the gas molecules are pushed closer together and the empty space between them is reduced. This causes more collisions with the container walls and an increased internal pressure. The gas volume will continue to contract until the internal gas pressure equals the external pressure again.

Problems:

1. A 500 mL sample of oxygen gas is collected in a manometer under 750 mm of pressure at constant temperature. What volume will the gas occupy at 500 mm of pressure?

ANS: 750 mL

2. If a scuba diver goes 100 feet underwater where her lung capacity is 25% of the volume at sea level, how much pressure are her lungs being subjected to at this depth?

ANS: 4x standard pressure

B. Charles' Law: volume versus temperature

Definition: If the pressure exerted on a definite amount of dry gas remains constant, then the volume of the gas varies directly with its Kelvin(absolute) temperature.

Simply stated, this means that as the temperature is increased the volume will increase(expand), and as the temperature is decreased the volume decreases(contracts).

The Kelvin temperature scale has only positive values, beginning at the coldest known temperature, -273°C or 0°K, which is called **absolute zero.** $K = {}^0C + 273$

Mathematically,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \textbf{if P is constant}$$

Explanation: As the temperature is increased, the gas molecules will move more rapidly causing more collisions with the container walls. This increased internal pressure will cause the container to expand until the greater distance between molecules results in the same number of collisions(and gas pressure) as before heating. The gas pressure will again equal the external pressure.

Remember: All temperatures **must** be converted to the Kelvin scale when problem solving.

Problems:

1. What volume would 500 mL of a gas at 173^{o}K occupy at a temperature of 0^{o}C?

ANS: 789 mL

2. If 5 mL of CO_2(g) at -140°C expands to a volume of 15 mL, what is the new temperature of the gas?

ANS: 399°K

C. Gay-Lussac's Law: pressure versus temperature

Definition: If the volume of a definite amount of dry gas remains constant, then the pressure exerted by the gas varies directly with the Kelvin temperature. (Note the similarity to Charles' law)

Simply stated, this means that as the temperature of a confined gas is increased, it will exert more pressure on the walls of the fixed-volume container. The mathematical relationship is:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \textbf{if V is constant}$$

Explanation: As the temperature is increased, the gas molecules will move more rapidly causing more collisions with the container walls. This causes an increased internal pressure. Since the volume of the container is constant, the internal gas pressure will continue to increase, possibly causing the container to explode.

Problems:

1. The pressure of a sample of gas in a gas cylinder tank is 800. kPa at standard temperature. What pressure would the gas exert at $100.^{o}$C?

ANS: 1093 kPa

2. An ideal gas exerts a pressure of 760 mm Hg at 0°C. At what temperature would this gas exert a pressure of only 1 mm Hg on the walls of a fixed-volume container?

ANS: 0.4°K

D. The Combined Gas Law

The three gas laws can be combined into one single equation that shows the relationship between pressure, temperature and volume for a definite amount of gas:

Combined Gas Law: $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$ **if moles of gas are constant**

Problems:

1. The volume of a sample of gas is 200. mL at 275°K and 92.1 kPa. What will the new volume be at 350°K and 98.5 kPa?

ANS: 238 mL

2. If a gas occupies 10 liters at STP, what would the Kelvin temperature have to be in order to have a volume of 30 liters at one-half of the original pressure?

ANS: 410°K

3. If a balloon filled with a gas occupies 6.0 liters at 250°K and 1 atmosphere of air pressure, what would the gas pressure be if the balloon is compressed to a volume of 2.0 liters and the temperature is raised to 1000°K ?

ANS: 12 atm

V. The Ideal Gas Law

The ideal gas law is an application of the combined gas law that includes the number of moles of gas as one last variable. Both the pressure the gas exerts and the volume it occupies are directly proportional to the number of moles of gas molecules in the container. The variable, **n**, is used to represent the moles, and is placed in the denominator of the Combined Gas Law equation. If you would solve the combined gas law using 1 mole of a gas at standard temperature and pressure(STP)..........

$$\frac{PV}{Tn} = \frac{(101.325\text{kPa})\,(22.4\text{L})}{(\ 273\text{ K}\ )(\ 1\text{ mole}\ )} = \mathbf{8.31\ kPa{\bullet}L/K{\bullet}mol} = \mathbf{R}$$ **ideal gas law constant**

If the pressure or volume is expressed in units other than those listed above, then the value for R can be recalculated. For example, if 1 atmosphere is used as the value for standard pressure instead of 101.325 kPa, then R will be **0.0821 atm**•L/K•mol (if pressure is in mm Hg, then R = **62.4 mm Hg**•L/K•mol)

Algebraically re-arranging this expression, the **Ideal Gas Law** becomes **PV= nRT**

The Ideal Gas Law can be used to determine the value of any of the four variables, given the other three. This equation is used when the experimental conditions remain constant. When problem solving, make sure that the units for pressure, volume and temperature are the same as those used for the R value.

Problems:

1. Calculate the pressure of 1.65g of helium gas at 16^oC that occupies a volume of 3.25L.

ANS: 304 kPa

2. What will be the volume of 0.70g of nitrogen gas(N_2) at standard temperature and 150kPa of pressure?

ANS: 0.378 L

3. How many moles of argon are there in a 27.3 L sample of gas at 85.9kPa and 50.5^oC?

ANS: 0.872 mole

VI. Molar Volume, Molar Mass and Density Calculations

Standard molar volume is defined as the volume occupied by one mole of gas at STP. This value is 22.4 L and can be derived from the ideal gas law equation. The molar volume is a unitary rate that can be used to find the number of moles of a different volume of gas at STP. The molar mass can then be used to determine the mass of the gas sample.

Sample Problem:

What is the mass of 5.6 L of oxygen gas at STP? What is the density of this gas sample?

$$x\,g\,O_2 = 5.6\,L\,O_2 \left(\frac{1\text{ mole }O_2}{22.4\,L\,O_2}\right)\left(\frac{32\,g\,O_2}{1\text{ mole }O_2}\right) = 8.0\,g\,O_2$$

$$\text{Remember that density(D)} = \text{mass / volume} = \left(\frac{8.0\,g\,O_2}{5.6\,L\,O_2}\right) = 1.4\,g/L\,O_2$$

$$\textbf{At STP}\text{, density of a gas} = \left(\frac{\text{molar mass}}{\text{molar volume}}\right) = \left(\frac{32\,g/\text{mole }O_2}{22.4\,L/\text{mole }O_2}\right) = 1.4\,g/L\,O_2$$

Problems:

1. What is the molar mass of a gas that has a density of 0.0899 g/L at STP?

ANS: 2.01 g/mole

2. What is the density of $CO_2(g)$ at STP?

ANS: 1.96 g/L

For Conditions Other than STP:

Molar volumes, molar mass and density can be calculated at conditions other than at STP, using PV = nRT. The number of moles(n) is equal to the mass of the gas sample(m) divided by its molar mass(M). Substituting this ratio into the ideal gas equation gives:

$$\mathbf{PV = \frac{mRT}{M}} \quad \textbf{or...} \quad \mathbf{M = \frac{mRT}{PV}} \quad \textbf{or...} \quad \mathbf{density(m/v) = \frac{MP}{RT}}$$

Sample Problems:

1. What is the density of a sample of xenon gas that is collected at 1.0 atm of pressure and 20°C?

$$\text{density of a gas} = \frac{MP}{RT} = \left(\frac{131\text{ g Xe / mole Xe}}{0.0821\text{ atm}\bullet\text{L / K}\bullet\text{mole}}\right)\left(\frac{1\text{ atm}}{293\text{ K}}\right) = 5.45\text{ g / L}$$

2. What is the molar mass of a gas if a 0.50 g sample occupies a volume of 250 mL at a temperature of 27°C and atmospheric pressure of 100. kPa?

$$\text{molar mass} = \frac{mRT}{PV} = \frac{(0.50\text{ g sample})(8.31\text{ kPa}\bullet\text{L / K}\bullet\text{ mole})(300\text{ K})}{(100.\text{ kPa})(0.250\text{ L})} = 50.\text{ g / mole}$$

Problems:

1. What is the mass of He gas in a balloon with a volume of 3.0 L at 22°C and barometric pressure of 95.0 kPa?

ANS: 0.47 g

2. What is the molar mass of an unknown gas if a 0.52 g sample occupies a volume of 210 mL at 25°C and a barometric pressure of 764 mm Hg?

ANS: 60. g/mol

VII. Chemical Reactions and Gases......A Review of Stoichiometry

Chemists use the gas laws to determine the number of moles, masses, and volumes of gaseous reactants and products in a chemical reaction.

Recall **Avogadro's hypothesis** which states that equal volumes of gases at the same temperature and pressure contain the same number of molecules(moles) of gas. This means the coefficients from a balanced equation give *both* the mole ratios and the volume ratios between gases.

Sample Problem:

How many liters of oxygen gas are needed to completely burn 4.0 L of methane gas (CH_4) in a Bunsen burner? How many liters of carbon dioxide gas will be formed? (Assume that all gases are measured at the same temperature and pressure.)

First write a balanced chemical equation to represent the reaction:

$$CH_4(g) + 2O_2(g) \text{ --> } CO_2(g) + 2H_2O(g)$$

Use the ratio of coefficients from the equation as a conversion factor for the volumes.

$$x\text{ L }O_2(g) = 4.0\text{ L }CH_4(g)\left(\frac{2\text{ L }O_2(g)}{1\text{ L }CH_4(g)}\right) = 8.0\text{ L }O_2(g)$$

Since the mole ratio between CH_4 and CO_2 is 1:1, the volumes must also be the same. Therefore, 4.0 L $CO_2(g)$ will be formed by burning 4.0 L of CH_4.

Problem:

How many liters of carbon dioxide gas are formed when 2.0 L of propane gas (C_3H_8) burns completely? How many liters of oxygen gas are needed for complete combustion?

ANS: 6.0 L CO_2, 10. L O_2

For experiments conducted at STP conditions:

If some of the reactants and/or products are solids or liquids, then the molar volume at STP can be used as a conversion factor.

Sample Problem:
What volume of hydrogen gas at STP can be produced from the reaction of 1.25 g of Zn with 25.0 mL of 2.00M HCl(aq)?

The balanced equation is: Zn(s) + 2HCl(aq) ---> H_2(g) + $ZnCl_2$(aq)

Remember that the mole, mass and volume ratios based on this equation will always remain constant. Therefore, you must first check to see if either of the reactants are in excess. Find out how many mL of HCl(aq) are actually needed to react completely with 1.25 g Zn, and compare this value to the amount used in the experiment. If no excess occurs, the amounts needed and available will be the same.

$$x \text{ mL HCl} = 1.25\text{g Zn}\left(\frac{1 \text{ mole Zn(s)}}{65.4 \text{ g Zn(s)}}\right)\left(\frac{2 \text{ moles HCl}}{1 \text{ mole Zn(s)}}\right)\left(\frac{1000 \text{ mL HCl}}{2.00 \text{ mole HCl}}\right) = 19.1 \text{ mL HCl needed}$$

conversions used: molar mass mole ratio molarity

Since there is more HCl available(25.0 mL) as compared to the 19.1 mL needed, the HCl is in excess and some will remain unreacted. Therefore, the amount of H_2(g) that can be formed is limited by the 1.25 g Zn used in the experiment.

$$x \text{ L } H_2(g) = 1.25\text{g Zn(s)}\left(\frac{1 \text{ mole Zn(s)}}{65.4 \text{ g Zn(s)}}\right)\left(\frac{1 \text{ mole } H_2(g)}{1 \text{ mole Zn(s)}}\right)\left(\frac{22.4 \text{ L } H_2(g)}{1 \text{ mole } H_2(g)}\right) = 0.428 \text{ L } H_2(g)$$

Problems:

1. How many grams of $CaCO_3$ must be decomposed to produce 9.0 L of CO_2 gas measured at STP?

ANS: 40.2 g $CaCO_3$

2. What mass of sulfur is produced by 4.11 g of iodine reacting with 317 mL of hydrogen sulfide at STP? The equation for the reaction is: I_2(aq) + H_2S(g) ---> 2HI(aq) + S(cr)

ANS: 0.45 g S

For experiments conducted at conditions other than STP:

If the experimental conditions are other than STP, then the ideal gas law can be used to determine the moles of gaseous reactants and products. This information can then be used to solve the stoichiometry problem.

Sample Problem:
What mass of mercury (II) chloride will react with 0.567 L of ammonia at 27°C and 102.7 kPa according to the equation:

$HgCl_2$(aq) + $2NH_3$(g) -----> $Hg(NH_2)Cl$(cr) + NH_4Cl(aq)

Using the ideal gas law, PV = nRT, find the number of moles of ammonia(NH_3) that reacts: (102.7 kPa) (0.567 L NH3) = n (8.31 kPa•L / K•mole) (300 K)
Solving for **n** = 0.0234 mole NH_3, which can then be used in a stoichiometry problem to

find the mass of mercury (II) chloride that will react.

$$x\ g\ HgCl_2 = 0.0234\ \text{mole}\ NH_3 \left(\frac{1\ \text{mole}\ HgCl_2}{2\ \text{moles}\ NH_3}\right)\left(\frac{271\ g\ HgCl_2}{1\ \text{mole}\ HgCl_2}\right) = 3.16\ g\ HgCl_2$$

Problems:

1. What volume of chlorine gas at 24°C and 99.2 kPa would be required to react completely with 2.51 g of Ag to form silver chloride?

ANS: 0.291 L Cl_2

2. What mass of potassium chlorate would need to be decomposed by heating to produce 500. mL of $O_2(g)$ at 900.°C and 97.5 kPa? (The solid residue remaining in the reaction tube after heating is potassium chloride.)

ANS: 0.407 g $KClO_3$

Objectives:

After completing the study guide and solving the problems you should be able to:

1. Summarize the differences between a solid, liquid and a gas.
2. Use the postulates of the kinetic molecular theory to explain the differences between the 3 phases of matter.
3. Describe how the velocity of gases is effected by changes in temperature, pressure and molar mass.
4. Determine the relative velocity or molar masses of gases using Graham's Law of Diffusion.
5. Calculate pressure in kilopascals, atmospheres, millimeters of Hg, and psi using data from a barometer or manometer.
6. Calculate the partial pressure of a gas collected by water displacement.
7. Calculate the partial pressure of a gas in a mixture given the mole ratio of the mixture, or find the number of moles present given the partial pressure ratio.
8. Describe the effect of a change in temperature on a gas and determine the new volume or gas pressure at the new temperature.
9. Calculate the new volume of a gas when the atmospheric pressure exerted on the gas changes.
10. Use the combined gas law to determine either the pressure, temperature or volume of a gas when experimental conditions change.
11. Apply the ideal gas law to determine the pressure, volume, temperature or number of moles in a gas sample.
12. Determine the molar mass, molar volume or density of a gas from experimental data.
13. Use the gas laws to solve stoichiometry problems involving gaseous reactants or products.