Unit 8: Gas Behavior

In this unit, you will learn: (1) the nature of gases on the particle level, (2) the physical behavior of gases in terms of pressure, volume and temperature and (3) gas stoichiometry.

Model of a Gas - Kinetic Molecular Theory Of Gases

Some common observations you may have made about gases can lead us to a model. First, they are readily compressible. For example, you can pump up a tire by forcing or compressing gas into it. By contrast, it is difficult to squeeze a solid or a liquid. They can only be slightly compressed. This must mean that the **distance between the gas particles is very large** in relation to the actual size of the particles. Second, gases **completely occupy the volume** available to them, in contrast to liquids and solids. This is evidence that the forces of attraction between gas molecules are weak, and the molecules are easily separated and free to move. Third, our concept of a gas is that of **particles in motion** (The average N_2 molecule in the air is moving about 1000 mi/hr at room temperature). If your friend is wearing a perfume or your pizza smells good, how do you know it? Molecules of the food odor or the perfume enter the gas phase and drift through space until they reach the cells in your nose that react to odors. **If the pizza is hot, then these molecules move faster** and it takes less time for them to reach your nose than if the pizza is cold.

The scientific theory that formally presents the previous ideas as its principle tenets is called the Kinetic Molecular Theory of gases (KMT). Right now we are using the KMT to provide a *description* of gases, but ultimately, the kinetic molecular theory provides an *explanation* for the behavior of gases at the molecular level.

1. Gases consist of molecules whose separation is much larger than the molecules themselves.
2. The molecules of a gas are in continual, random and rapid motion.
3. The average kinetic energy of the gas molecules is determined by the temperature of the gas. [KE = ½ (mass)$(\text{speed})^2$]
4. The molecules collide with each other and with the walls of their container without the net loss of energy.
5. The molecules do not exert any attractive force on each other.

A gas that strictly obeys all the points above is said to be an **"ideal gas."** In practice, there is no such thing as an ideal gas, but some gases come very close to this behavior.

Molecular Speeds

The kinetic energy of a molecule is dependent on its mass and speed. Specifically,

$$KE = \frac{1}{2} mv^2$$

Where "KE" equals the kinetic energy of the molecule, "m" equals its mass and "v" its speed. Gas molecules move faster at higher temperatures. It can be shown that the average kinetic energy of a sample of gas molecules is directly related to its temperature,

KEY CONCEPT: As <u>temperature</u> of a substance increases, the <u>average KE</u> of its particles increases and the <u>average speed</u> of its particles increases.

The example on the next page shows one application of this concept.

Molecular Speeds (continued)

Let's consider two samples of different gases, hydrogen and oxygen, both at the same temperature. Do the molecules of each of these gas samples have the same average speed?

$$\text{Temperature of } H_2 = \text{Temperature of } O_2$$

$$\text{KE of } H_2 = \text{KE of } O_2$$

$$\tfrac{1}{2}mv^2\ H_2 = \tfrac{1}{2}mv^2\ O_2$$

$$mv^2\ H_2 = mv^2\ O_2$$

Since the mass of an oxygen molecule is sixteen times as great as the mass of a hydrogen molecule, then it must be true that **the average speed of the hydrogen molecules must be greater than the average speed of the oxygen molecules**. With just a bit of algebra, it can be shown that the average speed of a hydrogen molecule is four times as great as the average speed of an oxygen molecule if both gases are at the same temperature.

Quantities that Describe a Sample of Gas

Moles – Moles is an indication of the number of particles present. *In equations: moles =* ***n***

Volume – The volume of a sample of gas is always equal to the volume of its container. Therefore, the volume of a given mass or mole amount of gas is variable. This in turn means that the density of a gas is variable (unlike solids and liquids).

In equations: Volume = ***V*** *and will be measured in Liters (L) = cubic decimeter (dm^3)*
(remember 1 mL = 1 cm^3 = .001 L)

Temperature – According to the Kinetic Molecular Theory (KMT), the temperature of a sample of gas is interpreted as a measure of the average kinetic energy (KE) of the particles. We have to speak in terms of the average KE because within a given sample of gas there is a distribution of kinetic energies, and, therefore, a distribution of speeds. In other words, there are some particles that are moving faster than others, so technically we should speak in terms of the average speed.

In equations: Temperature = ***T*** *and will be measured in Kelvin (K)* *($K = {}^oC + 273$)*

The Kelvin scale is an absolute temperature scale. According to the KMT, at a temperature of absolute zero all particle KE = 0, i.e. all particle motion ceases at 0 K.

By definition: **Standard Temperature** = 273 K (= 0 ^oC)

Pressure – On the particle level, pressure is interpreted as the result of the collisions between the particles and the walls of the container. Macroscopically it indicates a force per unit area.

In equations: Pressure = ***P*** *and will be measured in millimeters of mercury (mmHg) (a.k.a. torr)*
or in atmospheres (atm)
or in kilopascals (kPa)
or (rarely) in pounds per square inch (lb/in^2) (a.k.a. psi)

To convert from one pressure unit to another we will use the corresponding values of an (arbitrarily chosen)

Standard Pressure = 760 mm Hg = 1.00 atm = 101.3 kPa = 14.9 lb/in^2

You do not need to memorize these values, they will be given. These values are approximately equal to atmospheric (or barometric) pressure at sea level - more about that below.

<u>More about Pressure</u>

Where does that odd unit of **mm Hg** come from? It comes from measuring pressures in terms of the **height** of a column of mercury they can support against atmospheric pressure and/or the pull of gravity. Mercury is used because it is very dense.

The **barometer** is used for measuring (barometric) atmospheric or "room" pressure. You are probably familiar with the sensation of pressure you feel on your ears when you dive to the bottom of the deep end of a swimming pool or if you scuba dive or snorkel. This is due to the weight of the water above you. If you make the analogy with our atmosphere, you could say that we live at the bottom of an ocean of air, and the pressure effects are of similar origin. **The barometer measures the pressure the atmosphere exerts in terms of the height (in mm) of a column of Hg it can support against the pull of gravity.** *(In class) Draw a labeled diagram below.* The atmospheric pressure decreases as altitude increases. At any given altitude, it changes slightly due to atmospheric weather conditions.

Manometers allow for the measurement of the pressure of a trapped sample of gas. A **closed end** manometer measures the gas pressure relative to a vacuum in the other arm and is relatively straight forward (Text p.200). In an **open end** manometer, the other arm is open to the atmosphere and requires a little more thought. *It is more efficient to draw the diagrams below than to try to write descriptions.*

<u>BAROMETER</u>

<u>OPEN END MANOMETER</u>
<u>(Hg level in both arms equal)</u>

<u>OPEN END MANOMETER</u>
<u>(Hg level higher in arm open to atmosphere)</u>

<u>OPEN END MANOMETER</u>
<u>(Hg level higher in arm open to gas sample)</u>

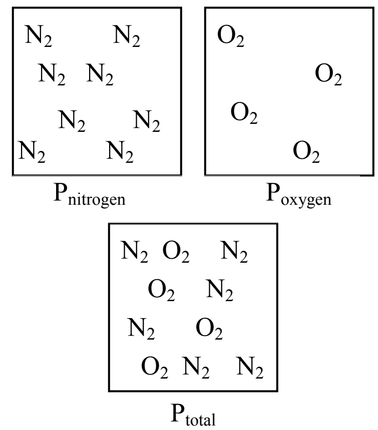
Dalton's Law of Partial Pressures

This fundamental law of gas behavior addresses the pressure conditions in a **mixture of gases**. There are two concepts involved.

1. The **partial pressure** of a given gas is equal to the pressure that that gas would exert if it were alone in the container.
2. The **total pressure** is the sum of all the partial pressures.

Algebraically, for a mixture of gases A, B, and C:

$$P_{TOT} = P_A + P_B + P_C$$



The diagrams to the right show three situations involving 2 moles of N_2, 1 mole of O_2 and a rigid container.

- If only the 2 moles of N_2 were in the container, it would exert a pressure called the "partial pressure of N_2". ($P_{nitrogen}$)
- If only the 1 mole of O_2 were in the container, it would exert a pressure called the "partial pressure of O_2". (P_{oxygen})
- As a mixture, the gases combine to exert a total pressure equal to the sum of the two partial pressures. (P_{total})

$$P_{total} = P_{oxygen} + P_{nitrogen}$$

By definition, in a mixture, the two gases are at the same temperature and occupy the same volume. As you will see shortly, this indicates that the partial pressure of a given gas is directly related to the number of moles of that gas in the container.

Problems on Kinetic Molecular Theory Pressure and Dalton's Law

1. A sample of Xenon gas is heated from 25 ^oC to 100 ^oC.

a. How does the kinetic energy of the atoms at 100 ^oC compare to the kinetic energy at 25 ^oC?

A. more KE at 100 ^oC B. less KE at 100 ^oC C. same KE at both temperatures

b. In this sample of Xenon, how does the speed of the atoms at 100 ^oC compare to the speed at 25 ^oC?

A. greater speed at 100 ^oC B. slower speed at 100 ^oC C. same speed at both temperatures

2. Two containers are filled with air which is a mixture of nitrogen and oxygen. The small one has a volume of 500 mL and the large one has a volume of 1500 ml. Both containers are at the same temperature and pressure.

a. In which container are the nitrogen molecules moving faster?

A. 500 mL container B. 1500 mL container C. same speed in both containers

b. In the 500 mL container, are the oxygen molecules or the nitrogen molecules moving faster?

A. oxygen molecules faster B. nitrogen molecules faster C. both have the same speed

(A, A, C, B)

3. Convert the following temperatures to Kelvin.

a. 25 ^oC b. 110 ^oC

(a. 298 K b. 383 K)

4. A force is a push or pull. What is pressure? (a single phrase answer)

5. Convert a pressure of 97.5 kPa to:

a. atmospheres

(0.962 atm)

b. mmHg

(731 mmHg)

6. A gas is put into an open end manometer. In the U-tube, the mercury level is 150 mm lower on the side exposed to the atmosphere. Atmospheric pressure is measured at 740 mmHg. What is the pressure exerted by the gas?

(590 mmHg)

7. A balloon contains a mixture of helium and neon. The partial pressure of the helium is 93.0 kPa and the total pressure inside the balloon is 101.0 kPa. What is the partial pressure of the neon?

(8.0 kPa)

Vapor Pressure, Evaporation and Boiling Point

Vapor Pressure

- A vapor is a gas produced by evaporation. When water is in an open container, it will evaporate to dryness. If it is in a closed container, the evaporation appears to stop after a certain amount of water has turned to vapor.
- The vapor pressure of a liquid is defined as the pressure exerted by its vapor when it is in a closed container.

Liquids or solids that have significant vapor pressures at room temperature are often said to be "volatile". This basically just means that they evaporate easily. Nail polish remover, for example, is very volatile – you can smell the fumes very quickly after you open the bottle and a bottle left open will quickly evaporate. Water, on the other hand, evaporates very slowly at room temperature, so it is much less volatile. A few solids are considered volatile: moth balls, for example.

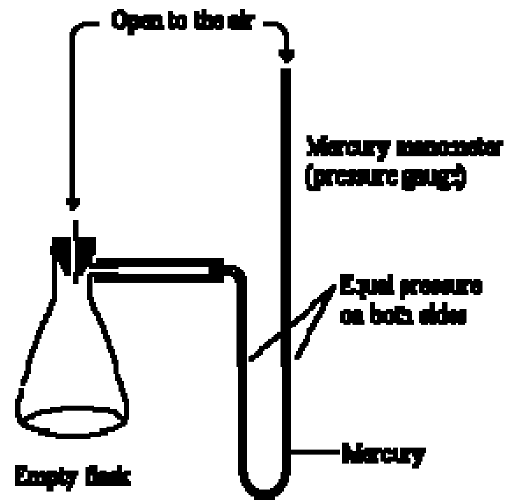
The sequence of three diagrams on the following pages shows how the vapor pressure of a volatile organic liquid called acetone is generated and measured at 25 ^oC.

All of the diagrams in this section are courtesy of Purdue University

http://www.chem.purdue.edu/gchelp/liquids/vp2.gif&imgrefurl=http://www.chem.purdue.edu/gchelp/liquids/vpacetn.html&h=321&w=335&sz=3&hl=en&start=19&tbnid=eL7cPx1onOwmBM:&tbnh=114&tbnw=119&prev=/images%3Fq%3Dvapor%2Bpressure%26svnum%3D10%26hl%3Den%26sa%3DN

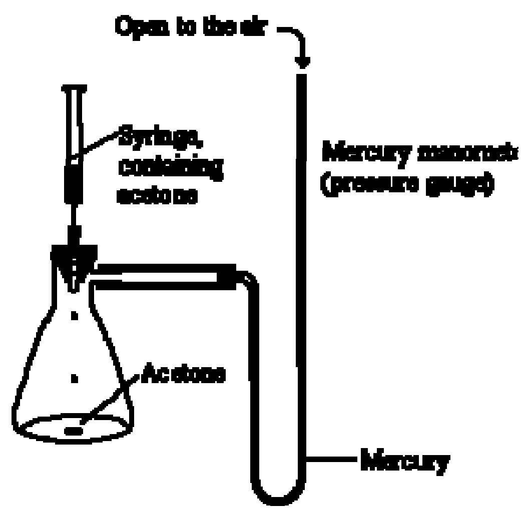
Vapor Pressure (continued)

1) Initially, both sides of the manometer are open to the air. The level of mercury on both sides will equal because the pressure on both sides is equal.



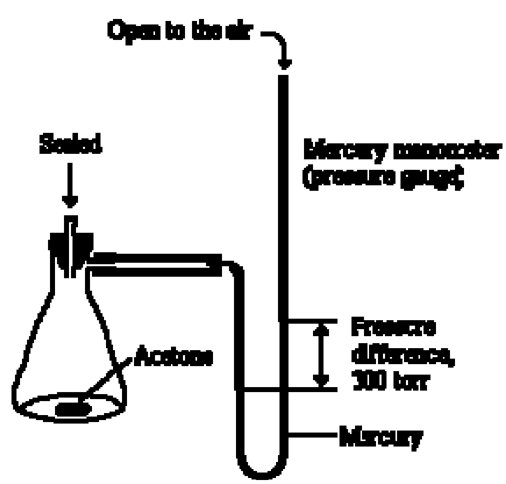
2) A little bit of liquid acetone (CH_3COCH_3) is then injected into the sealed flask at 25 ^oC.

As the acetone evaporates, the pressure in the flask increases and the increased pressure pushes on the column of mercury, so that the side exposed to the acetone vapor is now lower than the side exposed to the atmosphere.



3) When the levels in the column of mercury stop changing, the difference in the heights of the two sides of the manometer is measured. This difference in heights is used to calculate the increase in pressure due to evaporation of the liquid in the flask. This value is called the "equilibrium vapor pressure", or usually just the "vapor pressure".

The example indicates that the vapor pressure of acetone at 25 ^oC is 300 torr (= 300 mmHg).



Vapor Pressure and Evaporation

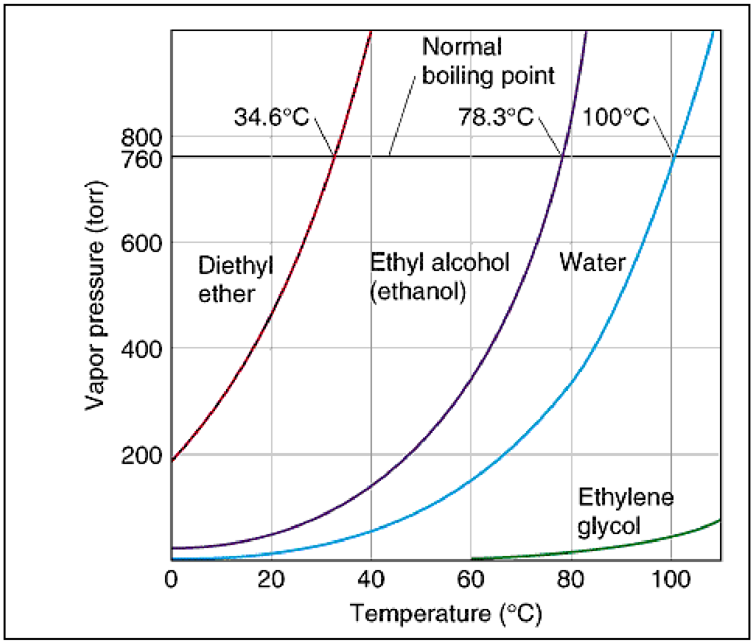
At any given temperature, a certain number of molecules in a liquid possess sufficient kinetic energy to escape from the surface. This process is called **vaporization** or **evaporation**. Evaporation only takes place at the *surface* of the liquid.

If the liquid is in an open container, evaporation occurs until the liquid is completely gone. If the container is closed, evaporation occurs to a state of **dynamic equilibrium**. On the particle level the process of evaporation continues, but condensation occurs at the same rate. Because these two opposing processes occur at the same rate, evaporation appears to stop.

Vapor Pressure, Temperature and Force of Attraction Between Particles

The vapor pressure of any liquid is a function of temperature and the force of attraction that exists between the particles in the liquid phase.

- Temperature



All substances tend to vaporize more at a higher temperature. For any given liquid, as temperature increases, vapor pressure increases. Pictured to the right is a graph that illustrates the dependence of vapor pressure (torr = mmHg) on temperature for four different liquids. Notice that while the vapor pressures of the 4 different liquids are quite different, all 4 liquids show an increase in vapor pressure as a function of temperature. This is because as the temperature is increased, more molecules have sufficient kinetic energy escape the surface of a liquid (or solid). There are more moles of the substance in the vapor phase, and the particles in the vapor phase have greater kinetic energy. These two factors combine to produce more collisions per unit area and collisions of greater impact between gas particles and the walls of the container. This results in an increase in vapor pressure. The values of the vapor pressure at specific temperatures are characteristic of a given liquid and apply to the liquid whether it is in an open container or a closed container.

- Strength of intermolecular forces

As you learned earlier in the unit, groups of molecules which are held together with strong attractive forces are less likely to evaporate. Therefore, when comparing different liquids at the same temperature, we can conclude that the liquid with the higher vapor pressure has the weaker attractive force acting between the molecules. In the vapor pressure graph above notice that the vapor pressure of ethyl alcohol is higher than that of water at any given temperature. This indicates that the intermolecular forces in ethyl alcohol are weaker than those in water. So, among molecular substances at a given temperature, nonpolar liquids (or solids) will have the highest vapor pressures; especially if they have a relatively small mass. Polar substances will have lower vapor pressures than nonpolar substances (of similar mass), and substances with hydrogen bonding will tend to have the lowest vapor pressures of all molecular substances (of similar mass). Metals, network covalent substances and ionic compounds have low vapor pressures unless heated to very high temperatures.

Vapor Pressure and Boiling Point

As the temperature continues to increase eventually even molecules *within* the liquid are sufficiently energetic to vaporize. A vapor bubble will form within a liquid only if the pressure of the vapor (i.e. the vapor pressure) is equal to or greater than the pressure of the air pushing down on the surface of the liquid or the **external pressure** (a.k.a. surrounding or ambient pressure**)** Therefore:

1) **a liquid boils only when its vapor pressure is equal to or greater than the external pressure.**
2) **a "boiling point" of a liquid is the temperature at which its vapor pressure is equal to the external pressure.**
3) **the normal boiling point of a liquid is the temperature at which its vapor pressure is equal to 760 mmHg (1 atm).**

The fact that boiling point is a function of vapor pressure has two important consequences:

► Different liquids reach a vapor pressure of 760 mm Hg at different temperatures and thus have different normal boiling points.

► The boiling point of a particular liquid depends on the pressure of the air above the liquid, and therefore any liquid will boil at a variety of temperatures. For example water boils at 100°C at sea level, at 85°C on Mount Whitney and at temperatures above 100°C in a pressure cooker where the external pressures are greater than one atmosphere.

Vapor Pressure Questions (Answers on p.39)

8. Explain the vapor pressure data in the table to the right. Include the classification of each substance and the attractive force as part of your explanation.

substance	vapor pressure (@ 30 °C)
C_3H_8	.921 atm
C_6H_6	.145 atm
NH_3	.043 atm

Use the graph on the previous page to answer the following questions

9. Why is 78.3 °C called the "normal boiling point" of ethyl alcohol?

10.a. What is the approximate boiling point of diethyl ether if the external pressure is raised to 1000 mmHg? _________

b. Water will boil at 60 °C if the external pressure is (approximately) _____________.

Collecting a Gas Over Water - a Useful Application of Dalton's Law and Vapor Pressure

Often the most efficient way to collect a gas is by water displacement (you will do this with hydrogen gas in the Ideal Gas Law Lab). At first glance, it would seem that pure H_2 had been collected as a result of the reaction between Mg and HCl. However, when a gas is collected over water, the evaporation of the water is inevitable, and **a mixture of water vapor and the collected gas results**. The partial pressure of the collected gas is usually the desired quantity, and so a correction must be made.

$$P_{TOT} = P_{H_2} + P_{H_2O}$$

$$P_{TOT} - P_{H_2O} = P_{H_2}$$

(In class) Draw a diagram of H_2 gas being collected over water.

11. A sample of oxygen is collected by bubbling it through water at 15° C. The total pressure of the mixture of gases was 745.0 mm Hg. The vapor pressure of water at 15°C is 12.8 mm Hg. What is the pressure of the oxygen gas?

Gas Laws

SUMMARY TABLE OF INDIVIDUAL GAS LAWS

Avogadro's Hypothesis Equal volumes of gas at the same temperature and pressure contain equal numbers of molecules $V \propto n$ $\frac{V}{n} = k$	**Boyle's Law** The volume of a fixed amount of gas at constant temperature is inversely proportional to the gas pressure $P \propto 1/V$ $PV = k$
Charles' Law The volume of a fixed amount of gas at constant pressure is proportional to the absolute temperature of the gas (absolute = Kelvin temperature) $V \propto T$ (K) $\frac{V}{T} = k$	**Gay-Lussac's Law** The pressure of a fixed amount of gas of constant volume is proportional to the absolute temperature of the gas. (fill this in after Lab 7-2)

Relationships Between Quantities that Describe a Sample of Gas

Individual Gas Laws

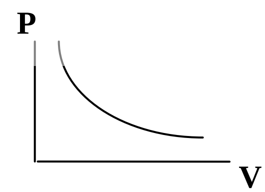
The gas laws are mathematical relationships which were discovered empirically by European scientists during the 17th and 18th centuries. Historically they are important, and to be able to refer to them by name shows a certain degree of scientific literacy. They are useful for understanding gas behavior because they each isolate two quantities at a time. However, we will soon have two equations that will incorporate all of these individual relationships and be more useful for calculation purposes. (We can fill in the results for Gay-Lussac's Law after we do the lab. You may know the answer from the reading.) In order of discovery, they are

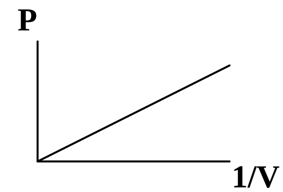
1. **Boyle's Law (1660)**:
   (assumes that n and T are fixed)

   $$P \propto \frac{1}{V} \quad \left(\text{or } V \propto \frac{1}{P}\right)$$

   or $PV = \text{constant}$

   or $P_1V_1 = P_2V_2$

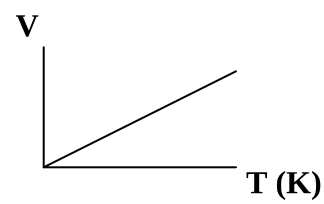




2. **Charles' Law (late 1700s)**:
   (assumes n and P are fixed)

   $V \propto T$ (in K)

   or $\frac{V}{T} = \text{constant}$ or $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

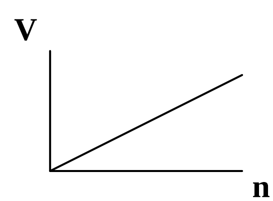


*3. **Gay-Lussac's Law (early 1800s)**:
   (assumes n and V fixed)

4. **Avogadro's' Principle (1811)**:
   (assumes P and T fixed)

   $V \propto n$

   or $\frac{V}{n} = \text{constant}$ or $\frac{V_1}{n_1} = \frac{V_2}{n_2}$



Implications of Avogadro's Principle – The fact that the volume of a sample of gas is directly proportional to the number of moles in the sample (assuming identical pressure and temperature conditions) is a valuable quantitative tool. This is primarily because the analysis of the situation does not depend on the type of gas involved – it applies to all types of gases.

Consider the three containers of identical volume below.
All are at the same pressure and temperature.

CO_2	H_2	N_2
P and T	P and T	P and T

According to Avogadro's Principle, all three contain the **same number of moles** of gas. To find the exact number of moles, we will have to develop the relationships a little further.
The **mass** of gas in each container would be different

Avogadro's Principle is also useful when we want to relate quantities of gases involved in a reaction using a balanced chemical equation (yes, stoichiometry is lurking out there). Remember that the coefficient ratio equals the mole ratio. Now we know that for gases at the same P and T the volume ratio equals the mole ratio. Therefore, in this situation, **the coefficient ratio equals the volume ratio**. This can shortcut some stoichiometry calculations.

Derivations of Major Gas Laws

Ideal Gas Law

Boyle's Law, Avogadro's Principle and Charles' Law give us the following relationships.

$$V \propto \frac{1}{P} \qquad V \propto n \qquad \text{and} \qquad V \propto T \text{ (in K)}$$

Mathematics tells us that if a quantity is directly proportional to two or more other quantities, the first quantity is proportional to the product of the other quantities.

Therefore: $$V \propto \frac{nT}{P}$$

In order to change an expression from a proportionality to an equation we add a **constant of proportionality**. In the Ideal Gas Law an "**R**" is used.

Therefore: $$V = \frac{RnT}{P} \quad \text{or} \quad \frac{PV}{nT} = R$$

or, in its more conventional form: $$\boxed{PV = nRT}$$ (**V** in **Liters** and **T** in **Kelvin)**

MEMORIZE THIS!

THE IDEAL GAS LAW gives us ANOTHER WAY TO CACULATE MOLES of a substance!

The value of R depends on the pressure units.

If the given or desired pressure is in:	The value of R is:
mm Hg	$62.4\ \frac{L \cdot mmHg}{mol \cdot K}$
atm	$.0821\ \frac{L \cdot atm}{mol \cdot K}$
kPa	$8.31\ \frac{L \cdot kPa}{mol \cdot K}$

You do not need to memorize these. They will be given (**with the units)**.

Derivations of Major Gas Laws (continued)

Combined Gas Law

This is useful for performing calculations that pertain to situations of changing conditions on a trapped sample of gas, and the moles (or mass) of gas are not desired. In other words, in situations where the number of moles of gas are held constant, but P and/or T and/or V are changed.

Under the first set of P, V, and T conditions: $P_1V_1 = nRT_1$

or: $$\frac{P_1V_1}{T_1} = nR$$

Similarly, at the second set of P, V, and T conditions:

$$\frac{P_2V_2}{T_2} = nR$$

Therefore: $$\boxed{\frac{\mathbf{P_1V_1}}{\mathbf{T_1}} = \frac{\mathbf{P_2V_2}}{\mathbf{T_2}}}$$ (**T** in **Kelvin**)

MEMORIZE THIS!

The P_1 & P_2 pressure units and the V_1 & V_2 volume units must be consistent.
If a quantity is not indicated as changing, simply remove it from the equation.
This equation gets its name from the fact that it incorporates all of the relationships derived by Boyle, Charles and Gay-Lussac.

Applications of the Major Gas Laws

Molar volume of any gas at Standard Temperature and Pressure (STP)

(In class) Perform the following calculation.

What is the volume of 1.00 mole of **any** gas at STP?

(STP = standard temperature and pressure = 1.00 atm and 0^oC)

(answer: 22.4 L)

This is an important quantity because STP conditions are (as the name implies) often used as standard reference conditions for gases. You will be given the number as a scientific constant without any units. You will need to remember its significance.

This produces another relationship you may want to memorize (optional*), and another way to calculate moles.

$$\boxed{n = \frac{V_{STP}}{22.4\ L/mol}}$$ (**V_{STP}** in **Liters**)

*It's actually just an algebraically rearranged variation of the Ideal Gas Law with standard T & P substituted and arithmetically combined with R.

Molar volume of any gas at STP (continued)

We are now ready to add <u>one more conversion to our "mole map"</u>. The rule is:

> **1 mole** (6.02 x 10^{23} molecules) of ANY GAS at standard temperature and pressure (STP) occupies a volume of **22.4 liters**.

REMEMBER:

- it **only applies at STP**.
- this new conversion **only applies to gases**.

So,

1 mole of CO_2 gas occupies 22.4 liters at STP
1 mole of O_2 gas occupies 22.4 liters at STP
1 mole of CH_4 gas occupies 22.4 liters at STP
2 mole of CH_4 gas occupies 44.8 liters at STP, etc.

<u>Problems on Combined Gas Law, Ideal Gas Law, and Molar Volume of a Gas at STP</u>

12. A 3.0 liter balloon is heated from 25^{o}C to 100^{o}C.

a. Does the volume of the balloon increase or decrease?_________________
b. What is the new volume?

(a. inc, b. 3.8L)

13. A 30.0 mL sample of gas is trapped in a syringe at a pressure of 750 mm Hg. A force is applied to the piston on the syringe to increase the pressure to 1.5 atm.

a. Does the volume of the gas increase or decrease? ______________________________
b. What is the new volume of the gas?

(a. dec., b. 19.7 mL)

14. Samples of sulfur dioxide gas and a sample of krypton gas are in separate 10.0L containers at 20.0 ^{o}C and 742 mmHg. Which of the following is true about the two samples of gas?

A. They contain a different number of moles, and have a different mass.
B. They contain a different number of moles, but have the same mass.
C. They contain the same number of moles, but have a different mass.
D. They contain the same number of moles, and have the same mass.

15. Which of the following samples of gas occupies the largest volume?

A. 50.0 g Ne B. 50.0 g Ar C. 50.0 g Xe D. They all occupy the same volume.

(Answers to 14 & 15 on p.?)

16. The gas in a particular tank had a pressure of 85.3 kPa at 23^oC. When placed in the sunlight the temperature rose to 48 ^oC.
a. does the pressure of the gas sample increase or decrease? ______________________
b. what is the new pressure of the gas sample?

(a.inc, b. 92.5 kPa)

17. 250.0 mL of oxygen was collected at 21^o C and 785 mm Hg. The next day the temperature was 37 ^oC and the pressure was 770 mm Hg. What was the new volume of the gas?

(269 mL)

18 How many moles of gas are contained in a 5.0 L balloon at room temperature (25^oC) and a pressure of 740 mm Hg?

(0.20 mol)

19. What is the pressure (measured in atm) in a 25.0 L vessel containing 1.00 kg of oxygen gas at 300 K?

(30.8 atm)

20. What mass of H_2 is required to fill an 80.0 L tank to a pressure of 150 atm at 27.0^oC?

(974 g)

21. What is the volume occupied by 3.57 moles of carbon dioxide gas at STP?

(80.0 L)

22. How many moles of oxygen gas occupy 5.0 liters at STP?

(0.22 mol)

23. How many grams of nitrogen gas occupy 6.0 liters at STP?

(7.5 g)

24. How many methane gas molecules are there in 4.0 liters of methane at STP?

(1.08 x 10^{23} molecules)

Calculations Involving Density of Gases

The basic concept of density for gases is the same as it is for solids and liquids:

$$\text{Density} = \frac{\text{mass}}{\text{Volume}} \quad \text{or} \quad \boxed{D = \frac{m}{V}}$$

The units are usually **g/L**.

As mentioned previously, the density of a gas is variable because the volume is variable. For a given mass of gas, volume depends on pressure and temperature conditions.

Problems involving density can be a little tricky, but remember a few things:

- When you are given the density of a gas, it is the same as being given both a **mass and a volume**. For example, if it is given that a gas has a density of .857 g/L under a certain set of P and T conditions, then a mass of .857 g has a volume of 1.00 L at those conditions.
- Density is an **intensive property**, which means it does not does not depend on the size of the sample. In order to get started on a problem that seems to be under prescribed in terms of given information, you can **assume a sample size (e.g. 1.00 mole, 1.00 L, 1.00 g)**.
- You should then be able to use the Ideal Gas Law and/or the Combined Gas Law along with $n = \frac{m}{MM}$ to find the quantities that you need.

Calculations Involving Density of Gases (continued)

STP conditions are often used as reference conditions for densities.

By definition: $D_{STP} = \frac{m}{V_{STP}}$

If we assume a 1.00 mole amount, then m = MM and V_{STP} = 22.4 L/mol.

Substituting gives: $\boxed{D_{STP} = \frac{MM}{22.4 \text{ L/mol}}}$ another equation you may want to memorize

optional

Problems on Density of Gases

25. Calculate the density of helium gas at STP.

(0.18 g/L - assume 1 mole He)

26. Calculate the density of helium gas at a temperature of 25° C and pressure of 560 mm Hg.

(0.120 g/L - assume 1 mole He)

Problems on Density of Gases (continued)

27. The density of an unknown gas is measured to be 2.03 g/L at 99.8 kPa and 120 ^oC.
(i.e. A mass of 2.03 g occupies a volume of 1.00 L at 99.8 kPa and 120 ^oC.)

HINT: It is possible to start this problem with either part a or part b.

a. How many moles of gas are in a 1.00 L sample?

(0.0306 mol)

b. What volume would the sample occupy at STP?

(0.686 L)

c. What would the density of the gas be at STP?

(2.96 g/mol)

d. What is the molar mass of this gas?

(66.3 g/mol)

Gas Stoichiometry

The central concept is the same as when you did stoichiometry in Unit 3 – **the coefficients** in a balanced equation are used to relate the **moles** of one substance to the moles of another substance.

When solving problems, first look for the shortcut based on Avogadro's Principle. **If both substances are gases at the same temperature and pressure, the coefficient ratio equals the volume ratio**.

You may have to use the Ideal Gas Law or 22.4 L/mol @ STP to calculate the moles of a gas from the given information or to calculate the final answer.

Problems on Gas Stoichiometry

Problems 24-28 refer to the equation for the combustion of ethane (C_2H_6):

$$2\ C_2H_{6(g)}\ +\ 7\ O_{2(g)}\ \rightarrow\ 4\ CO_{2(g)}\ +\ 6\ H_2O_{(l)}$$

28. How many moles of ethane (C_2H_6) are required to produce 10.0 L of carbon dioxide at STP?

(0.223 mol)

29. What volume of oxygen measured at STP is needed to produce 3.20 g of water?

(4.65 L)

Problems 24-28 refer to the equation for the combustion of ethane (C_2H_6):

$$2\ C_2H_{6(g)} + 7\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 6\ H_2O_{(l)}$$

30. What mass of CO_2 is produced when 15.8 L of O_2 measured at 742 mm Hg and 20°C are consumed?

(16.1 g)

31. What volume of oxygen measured at 25 °C and 1.02 atm is required to react completely with 0.317 L of ethane at the same pressure and temperature?

(Hint: You can save time if you remember how moles and volumes are related when P and T are the same.)

(1.11 L)

32. What volume of CO_2 is produced at 35°C and 111 kPa when 18.5 grams of O_2 are consumed?

(7.62 L)

33. When calcium carbonate is heated, it decomposes to produce calcium oxide solid and carbon dioxide gas:

$$CaCO_{3(s)} + \text{energy} \rightarrow CaO_{(s)} + CO_{2(g)}$$

a. What volume of carbon dioxide gas is produced at STP when 10.0 grams of calcium carbonate is decomposed?

(2.24 L)

b. What mass of $CaCO_3$ will be needed to generate 150.0 liters of $CO_{2(g)}$ on day when the temperature is 37° C and the pressure is 95 kPa?

(554 g)

34. Zinc metal reacts with hydrochloric acid.
a. Write a balanced equation for this reaction.

____

b. A strip of zinc is placed in 50.0 mL of 6.00 M hydrochloric acid. When the reaction stops, most of the zinc strip remains. What volume of gas is produced if it is measured at a temperature of 18°C and pressure of 0.955 atm?

($Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$)
(3.75 L)

UNIT 8 GRADED HOMEWORK ASSIGNMENT

DO THIS ASSIGNMENT ON A SEPARATE SHEET OF PAPER

G1. A 302 mL flask contains 0.00893 mol of neon. In order for the neon to exert a pressure of 715 mmHg it must be heated to a specific temperature. Calculate the temperature.

G2. A 5.50 L tank sample of carbon dioxide is at 5.00 ^{o}C. If the gas is exerting a pressure of 75.0 kPa, calculate the mass of carbon dioxide in the tank.

G3. A 250.0 L sample of propane at a temperature of 38.0 ^{o}C exerts a pressure of 125 kPa. What volume will this sample occupy at 95.0 ^{o}C and 750.0 mmHg?

G4. In a reaction involving gases, what conditions must be met in order to say that the coefficient (or mole) ratio of the gases is equal to the volume ratio of the gases.

G5. One step in the commercial production of sulfuric acid is:

$$SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)}$$

If all the gases are at the same temperature and pressure conditions of 30.0 ^{o}C and 0.953 atm:

a) What volume of sulfur dioxide is required to make 1000.0 L of sulfur trioxide?

b) What volume of oxygen is required to make 1000.0 L of sulfur trioxide?

G6. When making bread, the fermentation of glucose produces carbon dioxide which helps the bread rise:

$$C_6H_{12}O_{6(s)} + 2\ O_{2(g)} \rightarrow 2\ CH_3COOH_{(l)} + 2\ CO_{2(g)} + 2\ H_2O_{(l)}$$

What mass of glucose is required to produce 150.0 mL of carbon dioxide measured at STP.

G7. Your body obtains energy from the metabolic breakdown of glucose ($C_6H_{12}O_6$).

$$C_6H_{12}O_{6(aq)} + 6\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 6\ H_2O_{(l)}$$

What volume of oxygen, measured at 37 ^{o}C and 1.00 atm, is needed to react with 1.00 g of glucose?

Lab 8-1: DALTON'S LAW, VAPOR PRESSURE and IMF's

BACKGROUND

If a liquid evaporates in an open container, eventually all the liquid will disappear because the molecules that have escaped from the liquid into the vapor phase diffuse readily into the atmosphere. ***If the same quantity of liquid at the same temperature is placed in a closed container, what will happen?*** In this case, the volume of the liquid will initially decrease and then eventually become constant. If we monitored the pressure of the gas above the liquid, we would find that it initially increases and then it too levels off at a constant value. These observations can be explained in the following way.

The molecules with higher kinetic energies begin to leave the liquid, evaporating into the vapor phase. In time the space above the liquid becomes occupied with more and more gaseous molecules, and the pressure of the vapor increases. With the increasing number of chaotically moving gas molecules the number of collisions with the walls in this restricted volume also increases. One of these walls is the surface of the liquid itself, which will trap any bombarding molecules having low kinetic energies. Thus, condensation as well as evaporation of molecules in the vapor becomes large enough so that the rate at which the gas condenses exactly equals the rate at which the liquid evaporates. No further change in either the volume of the liquid or the pressure exerted by its vapor is observed. To emphasize this point again, vaporization and condensation are still taking place, but with no change in the liquid volume of the vapor pressure.

At this point, the liquid is said to be in dynamic equilibrium with its vapor. The pressure exerted by the quantity of vapor above the liquid when equilibrium is established is called the (equilibrium) vapor pressure of the liquid.

- The vapor pressure of a liquid depends on the ease with which its molecules can leave the liquid and enter the vapor phase. **In liquids where the intermolecular attractive forces are strong, the vapor pressure will be low. In liquids where the attractive forces are weak, the vapor pressure will be high.**
- Since increasing the temperature of a liquid increases the number of molecules possessing sufficient energy to overcome the attractive forces, **the vapor pressure must increase with increasing temperature.** Thus, whenever vapor pressure is given, the temperature at which it is measured must always be specified.

METHOD

In this experiment, you will measure the vapor pressure of two organic liquids over a range of temperatures. The liquids are **hexane** (C_6H_{14}) a hydrocarbon component of gasoline and **1-pentanol** ($C_5H_{11}OH$) an alcohol. The idea is to place a test tube containing either the hexane or 1-pentanol into a water bath and raise the temperature of the bath slowly, measuring the vapor pressure generated by the liquid in the tube and simultaneously recording the temperature.

In your prelab prepare a data table to record the vapor pressure (mm Hg) of hexane and 1-pentanol at one degree increments between 22°C and 45°C. Don't fill in the temperatures yet, but make a table for about 25 readings.

PROCEDURE

You and your partner will have a LabQuest unit. A temperature sensor will be attached to Channel 1, and a pressure sensor will be attached to Channel 2.

1. Turn on the LabQuest. After a few seconds the temperature and pressure should appear on the display.
2. The pressure will probably be in kPa, and you want to change it to mmHg.
   a. Touch **Sensor**
   b. Touch **Change Units**
   c. Touch **Gas Pressure Sensor**
   d. Touch **mmHg**

When the thermometer is reading in ^{o}C and the pressure sensor is reading in mmHg you are ready to begin the experiment.

3. Fill the beaker about ¾ full of tap water, put the stir bar in the beaker and set the beaker on the hot plate/stirrer, but **DO NOT TURN ON THE HEAT YET!**

4. Remove the cork from the test tube containing hexane and attach the pressure sensor stopper to the tube, lower the tube into the water bath and clamp it into the ring stand assembly.

5. Start the magnetic stirrer.

6. Turn on the heating element of the hot plate to moderate heat (~3). **BE SURE THE CABLES AND PLASTIC TUBING DO NOT COME IN CONTACT WITH THE HOT PLATE.**

7. When the water temperature reaches 22 ^{o}C, record the temperature and corresponding pressure. Continue to record temperature and pressure at one degree increments up to 45°C. Work as a team to record the data. Be accurate, but don't worry if you miss an entry; you are collecting data that will be plotted on a graph

8. After you have recorded the temperatures and pressures for the hexane, **turn off the hot plate** and the stirrer.

9. Remove the hexane tube, stopper it and refill the beaker with tap water. **Do not put the beaker on the hot plate until you are ready to assemble the set up for the second run**. **CAUTION: The hot plate will still be warm.**

10. When you are ready for the second run, set the beaker on the hot plate, repeat steps 3-7 with the tube containing 1-pentanol.

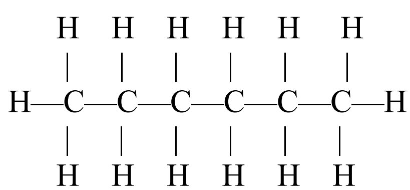
FYI - When you take the cork off of the test tube, the pressure inside the tube is equal to atmospheric pressure. When the rubber stopper is inserted, there may be a small increase in pressure due to the slightly decreased volume. For our purposes, we can treat the pressure in the tube as unchanged and, therefore, equal to the atmospheric pressure. The atmospheric pressure then should equal the pressure reading at 22 ^{o}C which is also the total pressure in the tube (P_T). Don't be concerned if your readings at 22 ^{o}C are slightly different for the two gases.

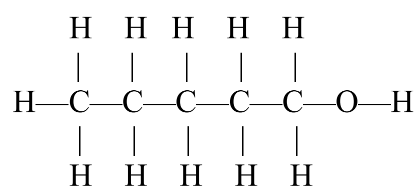
11. After you have recorded the temperatures and pressures for the 1-pentanol, **turn off the hot plate.** Remove the 1-pentanol tube and stopper it.

12. Touch **File** on the LabQuest screen, touch **Quit** and hold down the power button until the LabQuest is **off**

ANALYSIS AND CONCLUSIONS:

1. Hexane is C_6H_{14}, 1-pentanol is $C_5H_{11}OH$. Their structures are shown below.





a. Identify the significant intermolecular force acting between hexane molecules.

b. Identify the significant intermolecular force acting between 1-pentanol molecules.

c. Based on the type of IMF, which liquid has stronger intermolecular forces?

NOTE: You must use a spreadsheet/graphing program like Excel to do the calculations and to generate results tables and graphs for this analysis.

2. For each substance: Use Dalton's Law to convert the total pressure at each temperature to the vapor pressure of the sample. (Keep reading for help with these calculations.)

Calculation Background: Using Dalton's Law to Correct for Air Pressure in the Test Tube

The gas in the tube consists of the vapor generated by evaporation of the sample and the air that was in the tube when it was stoppered.

Dalton's Law of Partial Pressures states that the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures. For this lab, the total pressure in the tube (P_T) is equal the vapor pressure generated by the sample (P_{vap}) *plus* the pressure generated by the air (P_{air}).

$$P_T = P_{vap} + P_{air}$$

You measured the total pressure in the tube, but you need to determine the vapor pressure of the samples. **At 22°C, the vapor pressure of hexane is 120 mm Hg and the vapor pressure of 1-pentanol is approximately 6 mmHg**. With this information, you can calculate the partial pressure of the air. Assuming the partial pressure of the air remains constant throughout the experiment; you can make an adjustment to the total pressure readings and calculate the approximate vapor pressures of the samples.

In the case of hexane:
First, using your $\mathbf{P_T}$ **reading at 22 °C**, the fact that $\mathbf{P_{vap} = 120}$ **mm Hg** and the equation:

$$\mathbf{P_T = P_{vap} + P_{air}}$$

Calculate P_{air} . In other words:

$$\mathbf{P_T - 120\ mm\ Hg = P_{air}}$$

Second, for each of your **other** $\mathbf{P_T}$ **readings** calculate $\mathbf{P_{vap}}$ at each temperature by using:

$$\mathbf{P_T - P_{air} = P_{vap}}$$

For 1-pentanol you would do the same calculation for, but $\mathbf{P_{vap}}$ **would be 6 mmHg** at 22 °C.

3. Prepare a single graph that illustrates the relationship between the vapor pressure (vertical) and temperature (horizontal) for these two liquids. GRAPH MUST INCLUDE: Graph title, label both axes, values on both axes, scale so that plot lines "fill" graph, etc.

4.a. How does the vapor pressure of hexane at any given temperature differ from the vapor pressure of 1-pentanol?

b. Which of these liquids is more **volatile**?

5.a. Based on vapor pressure data, which liquid has stronger intermolecular forces?

b. Explain your answer in terms of the process of evaporation at the molecular level.

6. Are your answers to 1c and 5a consistent?

Using your graphs:

7. What would the surrounding pressure have to be in order to make 1-pentanol boil at 35 ^{o}C?

8. At what temperature would hexane boil if the surrounding pressure is lowered to 150 mm Hg?

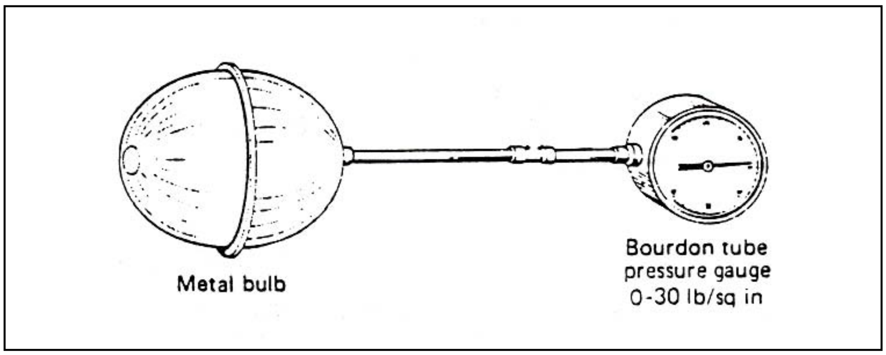
NAME____

Lab 8-2: THE RELATIONSHIP BETWEEN PRESSURE AND TEMPERATURE OF A GAS

In the late 1700's, a French chemist named Joseph Gay-Lussac studied how the pressure of a gas depends on its temperature when the volume is held constant.

We will carry out an experiment similar to Gay-Lussac's. We will heat and cool a fixed amount of gas in a closed container and measure the pressure of the gas at each temperature. We will graph our data, and then make some conclusions about the relationship of the pressure and temperature of a gas.

The apparatus we will use for these measurements is pictured below. The closed metal sphere at one end is filled with a gas. The pressure gauge at the top (calibrated in pounds per square inch) measures the pressure of the gas enclosed in the sphere.



DEMO PROCEDURE

Your teacher will immerse the sphere in four water baths; each having a different temperature. You are to record the temperature of each water bath and the corresponding pressure of the gas.

DATA TABLE

Temperature (^oC)	Pressure (lb/sq. in.)

ANALYSIS AND CONCLUSIONS:

NOTE: YOU MUST USE A SPREADSHEET/GRAPHING PROGRAM LIKE EXCEL FOR THE GRAPH PREPARATION AND GRAPHICAL ANALYSIS REQUESTED BELOW.

1. Use the class data to prepare a graph of Pressure (y axis) vs. Temperature (x axis) for this gas sample. Include the true origin (0, 0) in your graph. Assume the plot is linear and include the trend line (i.e. line of best fit).

2. Prepare a second graph of Pressure vs. Temperature using the same data you have collected for this gas; however, make the smallest number on the x-axis “-350°C” and the smallest number on the y-axis “0 lbs/in^2”. Extend the line back until it crosses the T axis. Read the graph and determine the temperature at which the pressure of the gas equal to 0 lbs/in^2 ?

3. Using the first graph, determine the equation for the line of best fit in “y = mx + b” form. Since y corresponds to P and x corresponds to T, you have actually generated an equation for the line as:

$$P = mT + b$$

Use this equation to calculate the temperature at which P = 0 lbs/in^2.

4.a. In reference to your answers to question 2, what does it mean in terms of KMT (i.e. particle collisions and particle motion) for the pressure to equal 0?

b. Your answers to 2 and 3 are experimental values for "absolute zero" expressed in ^{o}C. What is the accepted value for (absolute) 0 K in ^{o}C?

5. Calculate the percent difference between the average of your experimentally determined values for "absolute zero" and the accepted value in centigrade.

$$\text{\% difference} = \frac{\text{experimental value - actual value}}{\text{actual value}}(100)$$

6. Write the correct answer in your analysis.

a. The mathematical relationship between the pressure of a gas and its centigrade temperature is:

A. inverse B. direct C. directly proportional d. exponential

b. The mathematical relationship between the pressure of a gas and its Kelvin temperature is:

A. inverse B. direct C. directly proportional d. exponential

7. In the design of every experiment it is important to compare only two variables while holding the others constant. In this case, we varied the temperature and then measured the pressure. Think about the apparatus that was used and describe how the volume and number of moles of gas are held constant in this experiment?

8. There is a warning on the side of every aerosol can, such as a spray paint can, against “incinerating” the can. Why is it dangerous to heat an aerosol can?

Lab 8-3: THE IDEAL GAS LAW

BACKGROUND

We have considered four laws that describe the behavior of gases as revealed by experimental observations:

Boyle's Law $PV = k$ (at constant T and n)

Charles's Law $\frac{V}{T} = k$ (at constant P and n)

Gay-Lussac's Law $\frac{P}{T} = k$ (at constant V and n)

Avogadro's Law $\frac{V}{n} = k$ (at constant P and T)

These relationships can be combined to show that the volume, temperature, pressure and number of moles of a gas are related in the following way:

$$\frac{PV}{nT} = R$$ where R is the Ideal Gas Law constant.

It is important to recognize that the ideal gas law is an empirical equation, i.e., it is based on experimental measurements of the properties of gases. A gas that obeys this equation is said to behave *ideally*. Most gases closely obey this equation **at or below pressures of 1 atm**.

To verify that this is true, you will carry out an experiment in which you will calculate the pressure, volume, temperature and number of moles of a gas. Then by calculation, you will determine whether the gas you have collected is behaving as an "ideal" gas.

To generate the gas, you will react a known mass of magnesium with an excess of hydrochloric acid to produce hydrogen gas:

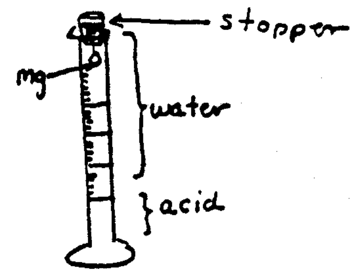
$$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

The hydrogen gas produced will be collected by the displacement of water. Since you are working with very small quantities of materials, it is essential that you work with great care.

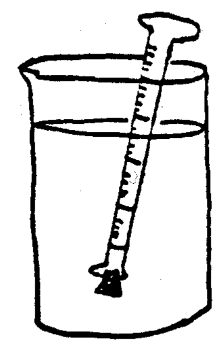
PROCEDURE

You should complete the entire experiment twice.

1. Fill a 1 liter beaker with about 900 mL of water.
2. Obtain a short piece of magnesium ribbon. Make sure it is within the recommended range of lengths and record the length to the nearest ½ mm. Also record the mass of 1 mm of magnesium ribbon (linear density). This data will be provided by your instructor.
3. Pour 3 mL of 3M HCl into your 10-mL graduated cylinder.
4. Tilt the graduated cylinder and slowly fill it **completely** with distilled water. While adding the water, rinse any acid that may be on the sides so that the final liquid in the top of the cylinder will contain very little acid. Avoid stirring up the acid layer in the bottom of the tube so the reaction does not begin prematurely.

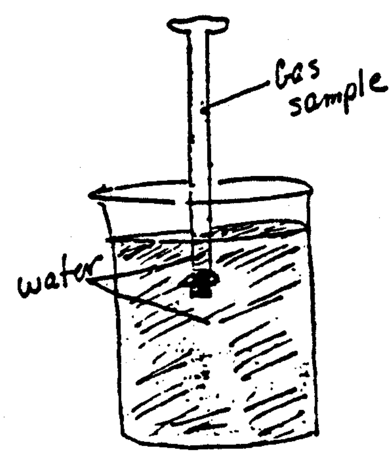


Perform steps 5-7 working quickly, but not rushing:
5. Put the piece of magnesium ribbon in the graduated cylinder.
6. Firmly insert a one-hole rubber stopper. If there is an air space, add water through the hole in the stopper.
7. Place your index finger over the hole in the stopper, invert the graduated cylinder and lower the end of the graduated cylinder into the beaker of water.



8. Note any evidence of a chemical reaction, and watch the magnesium to make sure it doesn't stick to the cylinder.
9. Allow the apparatus to stand for a minute or two after the magnesium has completely reacted.
10. Tap the sides gently to dislodge any gas bubbles that may have become attached to the sides of the graduated cylinder.

11. To accurately measure the volume of gas in the graduated cylinder, move it vertically (while keeping the open end submerged) until the water level inside the cylinder is the same as the water level in the beaker. **This is done to equalize the pressure of the gas trapped inside the cylinder with the atmospheric pressure**. Read the volume of gas in the cylinder as precisely as possible. Remember, you are interested in the volume of the hydrogen gas collected and **the cylinder scale is upside down.**
12. Your instructor will post room temperature and room pressure on the board. Record this information in your data
13. Repeat the entire experiment.



CLEAN UP
14. Empty the cylinder and place the stopper in the top of the cylinder.
15. Pour the water into the sink, refill the beaker.
16. Wipe your countertop with a wet sponge.

ANALYSIS AND CONCLUSIONS:

Show the calculations for questions 1-4 for your first trial.
Your answers for both trials (questions 1-4) should be presented in a results table.

1.a. What was the mass of the magnesium you reacted?
b. How many moles of magnesium reacted?

2. Using the balanced equation and the moles of magnesium that reacted, determine the moles of hydrogen produced in the reaction.

3. Since you collected the H_2 over water, the gas contains a certain amount of water vapor which adds to the total pressure. Apply Dalton's Law of Partial Pressures (total pressure of a mixture of gases equals the sum of the partial pressures of its component gases) to find the partial pressure of the H_2. Remember that the *total* pressure of the gas in the graduated cylinder is equal to the room pressure (aka atmospheric/barometric pressure). The vapor pressures of water at various temperatures are given below. Assume that the temperature of the water is equal to room temperature.

Temperature (°C)	Vapor Pressure of Water (mm Hg)
15	12.8
16	13.6
17	14.5
18	15.5
19	16.5
20	17.5
21	18.6
22	19.8
23	21.0
24	22.4
25	23.7
26	25.2
27	26.7
28	28.3
29	30.0

4. Calculate the values of R from your data. (one for each trial)

5. Average your two values of R and compare your experimental value to the accepted value of the Ideal Gas Law constant by calculating the percent difference between the accepted value and your calculated value?

6. What are the most likely sources of discrepancy? Be specific and remember that the error you describe must be consistent with the discrepancy between your value and the known value.

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UNIT 8 PRACTICE TEST

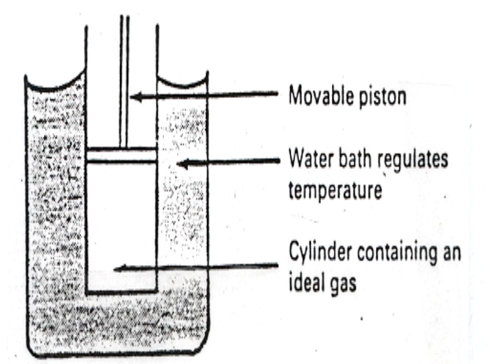
Part 1: Multiple Choice

1. The average kinetic energies of different samples of gases can best be compared by measuring their:
   A. molar masses B. pressures C. volumes D. temperatures

2. Two gases of different molar mass are at the same temperature. The heavier particles have a:
   A. greater average kinetic energy B. faster average speed
   C. lower a average kinetic energy D. slower average speed

3. Substances with a high vapor pressure at room temperature have:
   A. weak intermolecular forces and high normal boiling points
   B. weak intermolecular forces and low normal boiling points
   C. strong intermolecular forces and high normal boiling points
   D. strong intermolecular forces and low normal boiling points

Base your answers to the next three questions (4-6) on the diagram below.



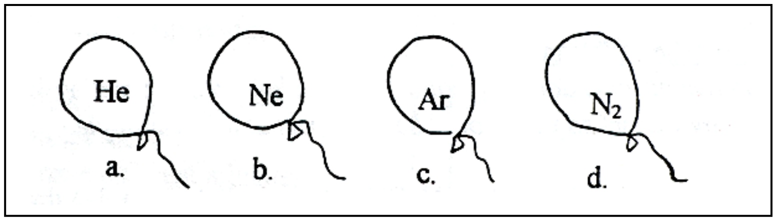
4. One way to <u>decrease the volume</u> of the gas in the container is to:
   A. increase the pressure
   B. replace the gas with the same number of moles of a different gas that has a smaller molar mass
   C. increase the temperature of the water bath
   D. none of the above will decrease the volume

5. If the temperature of the water bath is increased, and the piston is allowed to move (i.e. the pressure on the gas remains constant). Which of the following will occur?
   A. the mass of gas in the cylinder will remain the same
   B. volume will increase
   C. density of the gas will decrease
   D. all of the above will occur

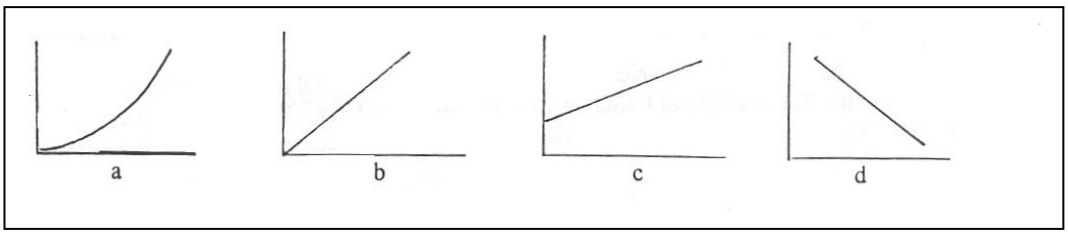
6. If the piston is held in position so that it cannot move, and the temperature of the bath is lowered:
   A. the average speed of the gas molecules decreases and the gas pressure increases
   B. the average speed of the gas molecules increases and the gas pressure increases
   C. the average speed of the gas molecules decreases and the gas pressure decreases
   D. the average speed of the gas molecules increases and the gas pressure decreases

7. If the pressure on a sample of gas is tripled at a constant temperature, the volume of the gas will be:
   - A. reduced to ⅓ its original volume
   - B. reduced to ⅔ its original volume
   - C. increased to 3 times its original volume
   - D. increased to 1½ it original volume

8. All of the 3L balloons pictured to the right are at room temperature and room pressure) Which has the greatest mass?



9. Which of the following represents the graph of pressure vs. temperature **<u>in ^oC</u>** for a given sample of gas held at a constant volume?



<u>Part 2: Problems</u>

10. What volume does 35.0 moles of N_2 occupy at STP?

11. A sample of hydrogen gas is collected by water displacement from a graduated cylinder at 35^{o}C. The cylinder is adjusted so that the liquid levels in the cylinder and the water bath are equal. The pressure of the atmosphere is 748.0 mm Hg. The vapor pressure of water at 35^{o}C is 42.2 mm Hg. What is the partial pressure of hydrogen in the flask?

12. A weather balloon is filled with 150 L of helium at 23^oC and 1.00 atm. What volume does the balloon have when it has risen to a point in the atmosphere where the pressure is 220 mm Hg and the temperature is -31^oC?

13. To what temperature (^oC) must 2.00 liters of gas at 22^oC be cooled, at constant pressure, so that the volume of the gas is reduced to 350 mL?

14. What mass of hydrogen gas, H_2, is needed to fill an 80.0L tank to a pressure of 150.0 atm at 27^oC?

15. A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L. The balloon is filled with 2.0 L of helium at sea level (pressure = 1.00 atm), is released and rises to an altitude at which the atmospheric pressure is only 500.0 mm Hg. Assuming that the temperature remains constant, will the balloon burst? Show your work.

16. What is the density of water vapor $H_2O_{(g)}$ at 110°C and 99.0 kPa?

17. The mass of a 50.0 ml sample of a particular gas is 0.0658 g. The temperature and pressure of the gas are 0.958 atm and 23.0°C respectively.

a. What is the molar mass of this gas?

b. What is the density of this gas at STP?

18. Consider the following balanced equation:

$$4\ Cu_2S_{(s)} + 6\ O_{2(g)} \rightarrow 4\ Cu_2O_{(s)} + 4\ SO_{2(g)}$$

a. What volume of oxygen gas, measured at 27.5°C and 0.998 atm, is required to react with 25.0 g of copper (I) sulfide?

b. What volume of sulfur dioxide gas is produced when an excess of copper(I) sulfide reacts with 42.0 liters of oxygen at STP?

19. A student carried out an experiment in which he produced some hydrogen gas by the following reaction:

$$2\ K_{(s)} + 2\ HCl_{(aq)} \rightarrow 2\ KCl + H_{2(g)}$$

He reacted a piece of potassium with an excess of hydrochloric acid and collected 29.2 mL of hydrogen gas by water displacement (temp = 20 °C). The barometric pressure was 745 mm Hg. The vapor pressure of water at 20 °C is 17.5 mm Hg.

a. How many moles of hydrogen gas were produced?

b. What was the mass of the piece of potassium?

Part 3: Essay

There may be an essay question on the test. The question will be taken from the following topic unless you are notified otherwise.

E1. Kinetic Molecular Theory

ANSWERS to UNIT 6 PRACTICE TEST

Part 1: Multiple Choice

1. D
2. D
3. B
4. A
5. D
6. C
7. A
8. C
9. C

Part 3: Problems

10. 784 L
11. 706 mm Hg
12. 424 L
13. -222 ^oC
14. 974 g H_2
15. V_2 = 3.0 L so yes it will burst
16. 0.56 g/L

17.a. 33.4 g/mol
   b. 1.49 g/L

18.a. 5.8 L
   b. 28.0 L

19.a. 0.00116 moles
   b. 0.0907 g

Answers to Unit 8 Homework Assignments

Assignment 2

p.8 problems 8-10

8. Vapor pressure results from molecules having sufficient KE to overcome the attractive forces in a liquid and evaporate into the vapor phase.

NH_3 (ammonia) molecules are attracted to one another by hydrogen bonds. C_3H_8 (propane) and C_6H_6 (benzene) are both nonpolar hydrocarbons in which weaker dispersion forces are the significant intermolecular force acting. NH_3 has the strongest intermolecular force among the three liquids. As a result, fewer NH_3 molecules escape from the surface of the liquid into the vapor phase, and, therefore, NH_3 will have a lower vapor pressure than C_3H_8 or C_6H_6 at any given temperature.

Similarly, C_3H_8 has a MM = 44 g/mole and benzene has a MM = 78 g/mole. As a result, C_3H_8 has a weaker dispersion force and therefore the weakest intermolecular force among the three liquids. More C_3H_8 molecules escape from the surface of the liquid into the vapor phase, and, therefore, C_3H_8 will have a higher vapor pressure than NH_3 or C_6H_6 at a given temperature.

9. Because vapor pressure of ethyl alcohol equals 760 mmHg (1 atm) at 78.3 °C.

10.a.. 40 ^{o}C

b. 150 mm Hg

p.9 problem 11

11. 732.2 mm Hg

Assignment 3

B: Avogadro's Hypothesis problems

p.13 problems 14, 15

14. C. Equal volumes of different gases at the *same temperature and pressure* contain the same number of moles. Since the gases have different molar masses, the samples will have different masses.

15. A. Since neon has the smallest molar mass, a 50.0g sample will contain the greatest number of moles of gas. Therefore it will occupy the greatest volume if all samples are at the same temperature and pressure.

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