Chem 111 – Experiment 6 – Simulation – Apparent Molecular Weight of Air

Background

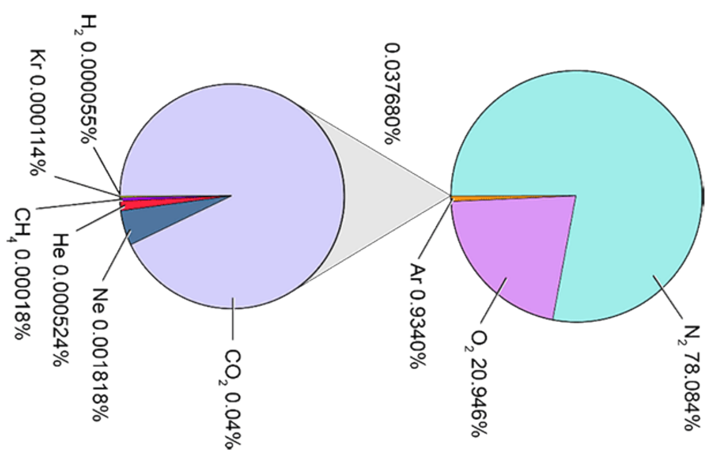
This final experiment once again returns us to gases and the Ideal Gas Law: -

$$PV = nRT$$

We are going to use this law to attempt to determine the Molecular Weight of Air. However, in the process we are also going to have a closer look at this Ideal Gas Law as a prelude to most of you going into Chem 112, as this Ideal Gas Law has some flaws associated with it that only allows its use under certain conditions. As a teaser the Ideal Gas Law could not be used to predict the pressure in a scuba divers gas tanks, there would be a huge discrepancy between what the gas gauge on the tanks read and what you would have calculated it to be using the Ideal Gas Law – more on that later.

First let us take a look at air.

The Composition of Air



Ideal Gas Law

As is typical of all Laws, the Ideal Gas Law, came about by experimental observations that date quite some time back. The three names that you are probably most familiar with are: -

Robert Boyle, 1627-1691, Boyle's Law.
Jacques Alexandre César Charles, 1746-1823, Charles Law – he was an avid balloonist.
Joseph Louis Gay-Lussac, 1778-1850, Gay-Lussac's Law.
Two of these Laws, Boyle's and Charles, you will be using in this Simulation

Others were involved but the combination led to what we now know as the Ideal Gas Law. However at that time the mathematical equations were a result of experiments but why they worked did not come about until much later with the development off **The Kinetic and Molecular Theory of Gases.**

In this Simulation we are first going to check that Air behaves as an Ideal Gas under the conditions of this Simulation. Once that has been established we will use the Ideal Gas Law to determine the Molecular Weight of Air (Molar Mass of Air). How? By a little manipulation of **PV = nRT.**

$$\mathbf{n\ (\#\ moles)} = \frac{\mathbf{PV}}{\mathbf{RT}}$$

Where

P:	Pressure – measured in atmospheres
V:	Volume – measured in liters
n:	# moles
R:	Ideal Gas Constant – for this experiment = 0.08205 $L.atm.mol^{-1}.K^{-1}$
T:	Temperature – measured in K

This allows us to determine the number of moles of air, and if we also know the mass in grams of air then using: -

$$\textbf{\# moles} = \frac{\textbf{\# g}}{\textbf{Molar Mass}}$$

Rearranging this a little

$$\textbf{Molar Mass} = \frac{\textbf{\#g}}{\textbf{\# moles}}$$

That's all that is to it! However, a question begs to be answered, why the need to check that we can use the Ideal Gas Law?

The Kinetic and Molecular Theory of Gases

With the Kinetic and Molecular Theory of Gases came an understanding of the various gas laws, why they worked but more importantly when the Ideal Gas Law does not work! The following is but a brief description. Let's look at just two of the postulates: -

a) The volume occupied by gas molecules is negligible in comparison to the volume of the container they are in.
This allows us to consider the volume that the gas molecules occupy is equal to the volume of the container that the gas is in.

b) Collisions between gas molecules are totally elastic.
Simply meaning no loss of energy when the collisions occur which infers no force of attraction between the gas molecules. In Chem 112 you will deal with Intermolecular Forces (IMF's), which in simplistic terms deal with the glue that binds molecules together. Take bottle of water, if there were no glue to bind the individual water molecules then it would be a gas. This glue that binds molecules together can be extremely strong (consider super glue) or it can be very weak (consider the simple post-it).

No rocket science needed to see potential flaws in just these two: -

Postulate a). Well if there is a really large quantity of gas then surely the volume that the gas can occupy must be the volume of the container minus the volume of all the gas molecules. Mentioned earlier, scuba divers, for obvious reasons their gas tanks contain a seriously large quantity of gas. Thus the Ideal Gas Law would be seriously underestimate the actual pressure that the gas gauge would display.

Postulate b). We have all seen demos involving liquid nitrogen. Nitrogen is a gas but at -196^{o}C it is a liquid. In order for it to be a liquid there has to be come glue that binds them together. This glue is relatively weak but at -196^{o}C, the nitrogen molecules are moving so slow that the collisions allow them to bind together and thus the collisions are not totally elastic.

Let's leave it at that. PV=nRT works extremely well at normal temperatures and relatively low pressures (think not a huge quantity of gas). Deviate from these conditions and PV=nRT, starts to break down.

Just for the fun of it do a web search, but instead of using the words "Ideal Gas Equation" in your search, try using "Real Gas Equations" and see what you come up with.

About This Lab

In this lab, you will assay a sample of air to determine whether it behaves as an ideal gas. If it does, you will then use the ideal gas equation to find its apparent molar mass. (Note that the molar mass is called apparent – this is because even though air is a mixture of gases it behaves as if it were a single gas with a single molar mass).

Open the simulation by clicking on the virtual lab icon shown on the left on the Hayden-McNeil Web Site. The simulation will launch in a new window.

You may need to move or resize the window in order to view both the Procedure and the simulation at the same time.

Procedure

*Go to the **General Chemistry web site** and **download the Report file**, this when completed is what you should send to your TA.*

Exp 6a: Study the Relationship Between Volume and Temperature for a Sample of Air

1. Take a **150mL Erlenmeyer flask** from the **Containers shelf** and place it on the workbench. **Double-click** to select the option to **close the flask**.

2. Take **air** from the **Materials shelf** and **add 1.25atm** to the flask.

3. Take a **gas syringe** from the **Instruments shelf** and **attach it to the Erlenmeyer flask**. Watch as air from the Erlenmeyer flask fills the gas syringe to equalize the pressure inside to the room air pressure of 1.000atm.

4. **The volume of a 150mL Erlenmeyer flask is 182.00mL.**
   Note: The volume of the flask is larger than 150mL due to the space above the top volume marking. The total gas volume is the sum of the volume of the Erlenmeyer flask plus the volume inside the gas syringe.
   Double-click on the **gas syringe** and **read the volume**. **Record** the **volume of air in the syringe** and the **total air volume**.

5. Take a **thermometer** from the **Instruments shelf** and **attach it to the Erlenmeyer flask**. **Record** the **temperature of the gas**.

6. Take a **constant temperature bath** and place it on the workbench. **Run** the constant **temperature bath** at **0°C**.

7. **Move** the Erlenmeyer **flask** into the **constant temperature bath**. **Wait** until the **temperature stabilizes at 0.0°C**. **Record** the **volume of air in the syringe**, **the total air volume**, and the **temperature**.

8. **Run** the bath at **30°C**.

9. Wait until the **temperature of the air stabilizes at 30.0°C**. **Record** the **volume of air in the gas syringe**, the **total air volume**, and the **temperature**.

10. **Repeat** the measurement **at 40°C, 50°C, 60°C, and 80°C**.

11. **Clear your station by** dragging all instruments back to the shelf, emptying the containers into the waste, then placing the containers in the sink.

Exp 6b: Study the Relationship Between Volume and Pressure for a Sample of Air

1. Take a **250mL Erlenmeyer flask** from the **Containers shelf** and place it on the workbench.

2. **Double-click on the Erlenmeyer** flask and **select the option to close the flask**.

3. Take a **pressure gauge** from the **Instruments shelf** and **attach it to the Erlenmeyer flask**. **Make sure** the **pressure gauge reads 1.000atm**.

4. There are no solids or liquids in the Erlenmeyer flask, so any volume in this container is available for the air to fill. Therefore, the air takes up the entire Erlenmeyer flask volume of **314.00mL**.
   Note: This volume is greater than 250mL due to the space above the top volume marking on the flask.

5. **Record** the first data pairs, **total volume of air** (**314.00mL**) and **pressure** (**1.000atm**).

6. Take a **graduated cylinder** from the **Containers shelf**. **Add 20mL water** from the **Materials shelf** to the **graduated cylinder**. **Double-click the graduated cylinder** to read and **record the actual volume at the meniscus**. **Transfer** the **water to the Erlenmeyer flask**, then **record** the **gas volume** and **pressure**.
   *Although it is not shown to you, the water is actually injected into the flask via the stopper to maintain the airtight seal on the flask. The **gas volume is the total volume of the Erlenmeyer** flask **minus** the **volume of the water**.*

7. **Continue** to **add water in 20mL increments** until the liquid **volume of water reaches 100mL**. **Record** the **gas volume** and the **pressure** for **each increment** of water added.

8. **Clear your station** by dragging all instruments back to the shelf, emptying the containers into the waste, then placing the containers in the sink.

Exp 6c: Measure the Mass of Air

1. Take a **250mL Erlenmeyer flask** from the **Containers shelf** and a **balance** from the **Instruments shelf** and place them on the workbench.

2. **Close the flask** and **evacuate the flask by adding 0.0001atm of air** from the **Materials shelf**.

3. Move the **flask** onto the **balance**. **Record** the **mass of the evacuated flask**.

4. Move the **flask back to the workbench** and **add 1.0atm air**.

5. Take a **pressure gauge** from the **Instruments shelf** and **attach it to the Erlenmeyer flask**. **Record** the **pressure**, then **remove the pressure gauge**.

6. **Weigh** the **closed Erlenmeyer flask** using the balance. **Record the mass** of the **closed Erlenmeyer flask and the air inside**.

7. **Double-click the flask** and **select the option to open the container** to release the gas. **Then reseal** the flask.

8. **Repeat** steps **4–7** **four additional times** **adding air** at a pressure of **1.1atm, 1.2atm, 1.3atm, and 1.4atm**.

9. Clear the bench of all materials, containers, and instruments, then use the **Report File** that you downloaded from the **General Chemistry web site**. When completed **send it to your TA.**