**Molar Mass of Butane**

**Part 1: Pre-Lab**

 The early gas laws were developed at the end of the 18th century, when scientists began to realize that relationships between the pressure, volume and temperature of a sample of gas could be obtained which would hold for all gases. An ideal gas can be characterized by those three variables because gasses behave in a similar way over a wide variety of conditions. The relationship between the three variables may be deduced the **Ideal Gas Law**.

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$$Gas Constant$$

$$R= .080206 \frac{atm\*L}{moles\*K}$$

Temperature needs to be in the units of Kelvins because that is what is used in the gas constant. The lab does not have a thermometer that measures Kelvins. It is important to convert to Kelvins. To do this you use the following equation: Kelvin= OC + 273

Butane is a common fuel used in disposable lighters. In this investigation we will use a disposable lighter and collect butane gas “over” water. This will provide us with a gaseous sample of butane to experiment with. Using this sample we will measure the volume, temperature and pressure of butane gas and the relationship between them. From your data you will be able to determine the molar mass of butane. The molar mass will be used to find the chemical formula for butane.

**Pre-Lab Questions**

1. Looking at the gas constant R, what units do we need to be in for

P:

V:

n:

T:

1. If I have 4 moles of a gas at pressure of 5.6atm and a volume of 12 liters, what is the temperature? (PUT UNITS ON YOUR ANSWER!!)
2. Convert 25 OC to Kelvins.
3. If I have an unknown quantity of gas at a pressure of 1.2atm, a volume of 31 liters and a temperature of 87 OC, how many moles do I have? HINT: Pay attention to units!!
4. How many moles of oxygen must be placed in a 3.00 liter container in order to exert a pressure of 2.00 atmospheres at 25 °C?

**Part 2: Procedure**

1. Obtain a new disposable lighter and remove the flint, wheel and spring as demonstrated by your teacher. Once they are removed weigh the lighter to the nearest 0.1g and record your measurement.
2. Fill a 250mL graduated cylinder all the way to the top, so that there is no room for air when you invert it in the plastic trough. If there are air bubbles just remove the graduated cylinder and refill it with water and try again.
3. Submerge the lighter under the water. Hold it below and inside the mouth of the 250mL graduated cylinder. Press the button on the lighter so that the gas escapes. Watch carefully to see that all of the bubbles of the gas are going up inside the graduated cylinder. Let the gas escape until you have about 200mL of gas.
4. Record the **exact** volume of gas in the graduated cylinder.

|  |  |
| --- | --- |
| **Temperature (°C)** | **Water Vapor Pressure (atm)** |
| 10 | .012 |
| 12 | .014 |
| 14 | .016 |
| 16 | .018 |
| 18 | .021 |
| 19 | .022 |
| 20 | .023 |
| 21 | .025 |
| 22 | .026 |
| 23 | .028 |

1. When done, shake excess water from the lighter and dry with a paper towel. Let it sit to dry while you finish with your measurements. **It needs to be dry before you reweigh it.**
2. Use a temperature probe to measure the temperature of water in the trough. We will assume that the gas will be at the same temperature as the water. Look up the vapor pressure for water at this temperature in the table to the right and record it in the table below.
3. To calculate the partial pressure of butane, you will need to know the current barometric pressure (air pressure). Your teacher will either provide it for you from a weather website. Record the pressure in the table below.
4. Empty your trough and clean up your work area.
5. Now that it is dry, reweigh the disposable lighter. Record the mass in the data table below.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  | Volume ofbutane (mL) | Temperature of Water (°C) | Barometric Pressure (atm) | Water VaporPressure (atm) | Mass of lighterbefore (g) | Mass of lighter after (g) |
| Trial #1 |  |  |  |  |  |  |
| Trial #2 |  |  |  |  |  |  |

You will use the ideal gas equation to calculate the moles (n) of butane you must first obtain each of the variables in the formula.

**Part 3: Calculating the Molar Mass** Follow the following steps:

Pressure, P, for butane must be calculated using Dalton’s Law of Partial Pressures. By making the water level inside the tube equal to the water level in the tub, the pressure inside the tube is equal to the pressure outside. In the end your PButane needs to be in the unit of atmospheres (atm). This is so you can have your units cancel, using the ideal gas constant given on the last page.



**Dalton’s Law of Partial Pressures**

* 1. Calculate the pressure for butane. Show your work:

1. Now that you have the pressure of butane, calculate the number of moles of butane (n) using the ideal gas equation:
2. Calculate the *experimental* molar mass of the butane using the mass of the gas (difference in mass of lighter before and after) and dividing it by the number of moles (n) you obtained in step b above. Show your work.
3. What is the *actual* molar mass of butane using the periodic table? Butane’s formula is: C4H10.
4. Calculate your percent error using the following equation: $\% error= \frac{\left|experimental-actual\right|}{actual} x 100$.

**Part 4: Thinking it through**

1. Was your percent error too high or too low? What do you think the largest source of error that made your error too high or low was in this experiment? Support your answer with evidence.
2. Is butane a liquid or a gas when it is inside the lighter? ***Explain*** your reasoning.
3. Why was it important to account for the water vapor pressure?