

(H) Chemistry
Molar Volume of a gas
(Butane Lab)

Name: _____
Blk: _____ Date: _____
Lab# _____

Introduction:

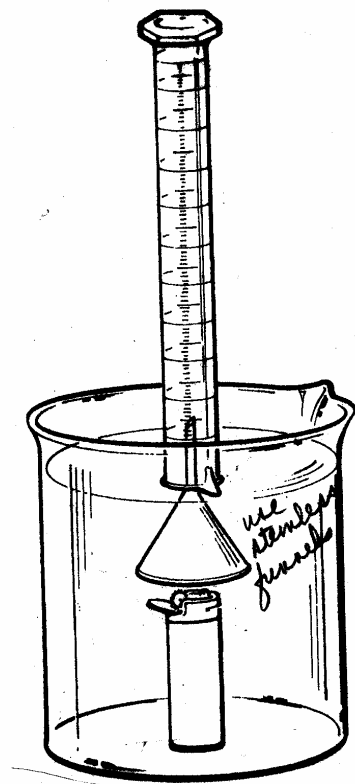
Avogadro hypothesized that equal number of molecules of different gases would occupy the same volume under the identical conditions of temperature and pressure. These findings were substantiated for any number of molecules (as long as equal units were used) and at any set of conditions (as long as the conditions for different gases were the same conditions of temperature and pressure). However, scientists like to use a standard point of reference and comparison. The most logical one is to compare exactly one mole of every gas at STP. Recall that STP is 0°C and 1 atm (convert to Kelvin). This is the concept of molar volume whereby a mole of any gas will occupy 22.4 L of space at STP. We will collect butane gas (C₄H₁₀). Our calculations will use the Ideal Gas Equation ($PV=nRT$).

Procedure:

1. Fill the bucket (sink) almost full. Do this immediately to allow the water to come to room temperature (equilibrate the temperature)
2. Obtain a butane lighter (UNDER NO CIRCUMSTANCES IS THIS LIGHTER TO BE LIT DURING THIS LAB!!) Immerse the lighter briefly in the water (allows water to replace any accumulated gas). Remove the lighter, shake off the water and dry it thoroughly.
3. Mass the lighter and record the mass.
4. One lab partner should fill a 100 mL graduated cylinder with water. **Do this under water!**
5. The other lab partner should hold the lighter at the bottom of the bucket (sink) and place the funnel over the lighter. The cylinder is carefully tilted over the tip of the funnel. The entire assembly is carefully turned upright. (see the diagram) Make sure there are no air bubbles trapped in the graduated cylinder!

AT ALL TIMES, KEEP THE GAS COLLECTION ASSEMBLY SUBMERGED!!!

6. Partners will work together – one holding the collection assembly (a gas collection tube) while the other depresses the lever near the flint wheel. This releases the butane gas. It will bubble into the funnel and up into the cylinder. Release the gas under water being careful that all of it is collected in the 'eudiometer' (gas collection tube) by water displacement.



7. Release enough butane gas to fill the graduated cylinder to within 3 mL of its calibrated capacity. (this may take ~ 5 min. so have a good grip on all the pieces of assembly)
8. Allow the butane to reach room temperature (~ 5 more min.) Then **adjust the level of the water inside and outside the cylinder until they are the same by raising and lowering the cylinder in the sink.** Be careful that you don't raise the cylinder too high and let air leak in.
9. This equilibrates the pressure in the cylinder with the pressure in the room (pressure on surface of water in container) You may now read the volume of gas collected using the calibration markings on the side of the cylinder (try to read to the nearest 0.5 mL)
10. Record this volume.
11. Record the temperature of the room air and the water. (they should be the same)
12. Remove the lighter from the beaker. Shake off any excess water, and dry thoroughly.
13. Mass the lighter and record this new mass.
14. Record the barometric pressure (instructor will tell you)

Data:

- | | |
|--|---|
| A) mass of lighter + butane BEFORE experiment (g) | |
| B) mass of lighter + contents AFTER experiment (g) | |
| C) mass of butane collected (A - B) (g) | * |
| D) volume of gas collected (#10) (mL)
(@ R.T. & R.P.) | |
| Conversion to liters | * |
| E) atmospheric pressure (R.P.)(mm Hg) | |
| Conversion to atm (" wet gas ") | ! |
| F) room temperature (R.T.) (°C) | |
| Conversion to Kelvin | * |
| G) v.p. of water at R.T. (use chart) | ! |
| H) partial pressure of "dry" butane
(E - G) (atm)
($P_{butane} = P_{atm} - P_{water}$) | * |

Analysis:

1. Using the ideal gas equation **and your experimental data**, determine the **experimental molecular mass** of butane.

(hint: $PV = nRT$, or since $n(\# \text{mols}) = \text{mass} \cdot \frac{\text{moles}}{\text{mass}}$ can transform into :

$$PV = \frac{mRT}{M} \text{ where } M \text{ is "molecular mass", therefore: } M = \frac{mRT}{PV})$$

2. Using the formula for butane (C_4H_{10}) , calculate its **theoretical molar mass**.

3. Calculate your % error

4. Using **your lab data** (g/L), calculate the **experimental density** of your gas

5. Using the **theoretical** $\frac{\text{molar mass}}{\text{molar volume}}$ calculate the **theoretical density** of butane

6. Calculate a % error on this value

7. Write a balanced equation for the complete combustion of the gas (butane).

8. Use **your** data and the ideal gas equation to calculate an “n” value (hint: transform ideal law for “n”); use this to calculate a molar volume for butane based on **your** experiment. $\left(\frac{L}{mol} \right)$

Compare **your** value with the established molar volume of gases from Avogadro’s work and calculate a % error for this value.

Conclusions:

1. Explain your understanding (complete sentences) of how a more massive gas molecule could still occupy the same molar volume as a less massive one..

2. Offer 2 possible reasons why **collection of a gas** over water might not be an acceptable method for collecting and calculating these types of data. (what conditions might make the data spurious and suspect).