DETERMINING THE MOLECULAR MASS OF BUTANE

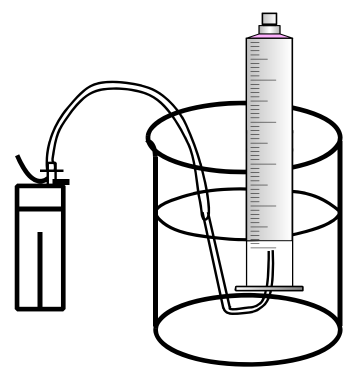
An Italian chemist, Amadeo Avogadro, declared in 1811 that "*equal volumes of gases at the same temperature and pressure contain an equal number of molecules."* Given that information, it did not take long for early scientists to realize that a standard volume of each different gas had a different mass. Since the number of molecules in that standard volume is a constant, the "mole concept" was soon born. The mass of that fixed number of molecules (Avogadro's number) is called the molecular mass. In this experiment you are going to determine the molecular mass of the gaseous hydrocarbon, butane, from the mass of a measured volume of gas.

Unfortunately, the experiment is not that simple, when an insoluble gas is collected by the downward displacement of water, the water evaporates into the butane until the butane becomes saturated with water molecules. The amount of water which evaporates is related to the force exerted by the water molecules as they push themselves between the butane molecules. This force is called the vapor pressure of water. To correct your gas volume for the added water vapor, simple subtract this vapor pressure from the atmospheric pressure. The difference, the partial pressure of the butane alone, is then substituted in the ideal gas equation when calculating its molecular mass.

Materials:

Source of butane such as a lighter or cartridge, plastic tubing to fit the source, "J" tube, 60 -ml or 120 ml Syringe with tip cap, 400-ml or 600-ml beaker, thermometer.

Procedure:



1. Mass a dry, butane filled, lighter or cartridge.

2. Fill the inverted Syringe with water, seal the open end with the tip cap.

3. Force the open end of the plastic tubing onto a "J" tube and attach the tubing to the butane source. Be certain that all the joints are tight to avoid lose of gas. Place the "J" tube into the beaker of water and slip the Syringe over the open end of the tube without admitting any air.

4. Release butane gas from its source into the syringe until the meniscus of the gas inside the syringe is at the maximum calibration scribe when the level inside the flask just matches the level of the water in the beaker. Remember to add an additional 1.2 ml to your measured value to account for the tip of the syringe.

5. Remove the tubing and mass the butane source. Be certain that the butane source in thoroughly dry before massing it. Record the water temperature to the nearest 1°C.

6. Blow the water out of the tubing and repeat the experiment as time permits.

Calculations and Questions:

For each trial, calculate questions Q1 through Q3.

Q1. (a) Determine the vapor pressure of water at the water temperature from a table of the vapor pressures of water at various temperatures.

(b) Calculate the partial pressure of butane in the volumetric flask from Dalton's Law of Partial Pressures. Where, $P_{atmosphere} = P_{water\ vapor} + P_{butane}$

or $P_{butane} = P_{atmosphere} - P_{water\ vapor}$

(c) Convert the partial pressure of the butane from units of torr to atmospheres.

$$P_{butane} = \frac{\text{Partial pressure in torr(mm of Hg)}}{\text{760 torr/atm}}$$

Q2. Calculate the number of moles, n, of butane gas contained in the volumetric flask from a form of the ideal gas equation

$$PV = nRT \qquad \text{or} \qquad n = \frac{PV}{RT}$$

Where
P = partial pressure of the butane gas in atmospheres (Q1c)
V = volume of gas in the Syringe in liters (Adjusted)
n = number of moles of butane in the flask
R = ideal gas constant, 0.0821 ℓ·atm/deg·mol
T = the temperature of the gas in degrees Kelvin, K

Q3. (a) Calculate the mass of butane released from the butane source by subtracting the final mass of the butane source from the initial mass of the butane source.

(b) Calculate the molecular mass for butane in grams/mol, where

$$\text{Formula mass} = \frac{\text{mass of butane gas in grams (Q3a)}}{\text{moles of butane gas in the flask (Q2)}}$$

(c) Calculate the average molecular mass for all trials.

Q4. By another experiment, butane was found to contain 82.8% carbon and 17.2% hydrogen. Using this information, calculate the empirical formula for butane.

Q5. Find the molecular formula for butane from its average molecular mass (Q3c) and its empirical formula (Q4).

Q6. Butane fuel is rarely pure butane but a mixture of butane and either propane, C_3H_8, or pentane, C_5H_{12}. According to the molecular formula you calculated in question #Q5, is your sample almost pure butane, a mixture of butane and propane, a mixture of butane and pentane, or some other mixture? Explain your rationale.

DATA:

Atmospheric pressure:____________ torr

	Trial #1	Trial #2	Trial #3
Initial mass of the butane source	g.	g.	g.
Final mass of the butane source	g.	g.	g.
Mass of butane	g.	g.	g.
Water temperature	°C	°C	°C
Volume of the syringe	ml	ml	ml
Vapor pressure of water at the water temperature	torr	torr	torr
Partial pressure of butane gas	torr	torr	torr
Moles of butane in the flask	mol	mol	mol
Molecular mass of butane	g/mol	g/mol	g/mol
Average molecular mass of butane			g/mol

Vapor Pressure of Water			
°C	torr (mm Hg)	°C	torr (mm Hg)
0	4.6	26	25.2
5	6.5	27	26.7
10	9.2	28	28.3
15	12.8	29	30.0
16	13.6	30	31.8
17	14.5	40	55.3
18	15.5	50	92.5
19	16.5	60	149.4
20	17.5	70	233.7
21	18.7	80	355.1
22	19.8	90	525.8
23	21.1	100	760
24	22.4	105	906.1
25	23.8	110	1074.6

TYPICAL DATA:

Atmospheric pressure: 756 torr

	Trial #1	Trial #2	Trial #3
Initial mass of the butane source	14.27 g.	14.06 g.	13.84 g.
Final mass of the butane source	14.06 g.	13.84 g.	13.63 g.
Mass of butane	0.21 g.	0.22 g.	0.21 g.
Water temperature	30°C	30 °C	29 °C
Volume of the volumetric flask	100 ml	100 ml	100 ml
Vapor pressure of water at the water temperature	32 torr	32 torr	30 torr
Partial pressure of butane gas	724 torr	724 torr	726 torr
Moles of butane in the flask	0.0038 mol	0.0038 mol	0.0039 mol
Molecular mass of butane	55 g/mol	57 g/mol	56 g/mol
Average molecular mass of butane			56 g/mol
Percent error based on a formula mass of 58.1 g/mol			4.4 %

PREPARATION OF THE "J" TUBE:

Heat 150 ml of water in a 250-ml beaker to boiling. Cut the stem from a thin stem pipet and stretch roughly one inch from the end to reduce its outside diameter. Bend a thin stem pipet into a "J" shape with its unstretched end about 2-3 cm from the bend. Remove the beaker from its heat source and dip the bent portion of the pipet into the hot water. Remove the pipet from the hot water after 15-20 seconds of heating. Holding the open end and stem parallel, dip the heated portion into a beaker of cold water to fix the bend. Cut off the excess stretched potion.