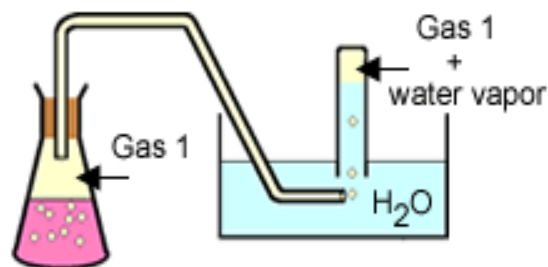


Background:

One of the common ways to collect and measure a gas is by water displacement. The volume of the gas can be measured by collecting the gas in a calibrated cylinder. The temperature of the gas will be the temperature of the water which can be measured with a thermometer. The sample will contain both the gas collected and some water vapor. The pressure of the mixture will be equal to atmospheric pressure, but to determine the pressure of the gas alone, the vapor pressure of the water will need to be subtracted from the total pressure. (An application of Dalton's Law of Partial Pressures)



$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water}} \quad (\text{Formula 1})$$

Once temperature, volume and pressure of the gas are determined, the Ideal Gas Law ($PV=nRT$) can be used to determine the number of moles of gas present. If the mass of the gas can be measured, there is enough information to figure out the grams per mole (g/mol) or molar mass of the gas.

In this lab, the gas from a butane lighter will be collected and measured. The molar mass of butane can then be determined.

Materials:

graduated cylinder
gas collecting trough
thermometer
lighter
balance

Procedure:

- 1- Measure the mass of a butane lighter.
- 2- Using the procedure demonstrated by your instructor, collect a sample of the gas from the lighter and record the volume of the gas.
- 3- Measure and record the temperature of the water/gas.
- 4- Using available resources, determine atmospheric pressure.
- 5- Dry the butane lighter and measure its mass after the gas has been released.
- 6- Dry the butane lighter again and measure its mass again. If this value agrees with the first value, accept that value. If not, dry and mass again and repeat until the value is constant.

Table 18-2

Vapor Pressure of Water					
Temperature (°C)	Pressure (kPa)	Temperature (°C)	Pressure (kPa)	Temperature (°C)	Pressure (kPa)
0	0.6	21	2.5	30	4.2
5	0.9	22	2.6	35	5.6
8	1.1	23	2.8	40	7.4
10	1.2	24	3.0	50	12.3
12	1.4	25	3.2	60	19.9
14	1.6	26	3.4	70	31.2
16	1.8	27	3.6	80	47.3
18	2.1	28	3.8	90	70.1
20	2.3	29	4.0	100	101.3

Pre-lab Questions:

- 1- A student collected a sample of gas at 23°C on a day when the atmospheric pressure was 101.6 kPa. What was the partial pressure of the gas? (Show work)

- 2- WRAL reports the atmospheric pressure in units called millibars (mb). One mb is equal to 100 Pascals. Perform the following conversions (show all work including the conversion factors):
 - a) 1015 mb = _____ kPa

 - b) 1000 mb = _____ kPa

 - c) 1050 mb = _____ kPa

- 3- A student collected a sample of gas on a day when the atmospheric pressure was listed as 1020 mb and the temperature of the water over which the gas was collected was 20.0°C. What was the pressure of the dry gas? (Show work)

- 4- An overexcited student was performing this lab activity and before collecting the gas, he couldn't resist "flicking his Bic" and lighting the lighter. How will this affect his lab results?

5- Another impatient student did not want to re-weigh her lighter until the mass was constant, so she took the second value she measured. If a water droplet was caught in the lighter mechanism, how would that affect her lab results?

6- A very conscientious student obtained the following data:

Water temperature	22.0°C
Atmospheric pressure	102.0 kPa
Volume of gas	50.0 mL

How many moles of gas did the student collect? (Show work)

7- Why do you measure the mass of the lighter before and after collecting the gas?

8- One lab group collected 2.5×10^{-3} moles of a gas. They measured the mass of the sample to be 0.11 g. What was the molar mass of the gas? (Show work)

9- The chemical formula of butane is C_4H_{10} . If a student determined the mass to be 62.0 g/mol, what was his/her percent error? (Show work)

Data:

Initial mass of the lighter		g
Final mass of the lighter	g (1 st weighing)	g (2 nd weighing)
Mass of sample gas		g
Volume of sample gas	mL	L
Atmospheric pressure	in Hg	kPa
Dry gas pressure		kPa
Water temperature	°C	K

Calculations: (Show all work in a logical fashion. Use units and labels.)

- 1) What two gases are in the graduated cylinder after step #4 of the experiment?
- 2) What is the vapor pressure of water (P_{water}) at the water bath temperature?
- 3) Use Dalton's Law of partial pressures (Formula 1) to determine the partial pressure of butane in the graduated cylinder (P_{but}).
- 4) Use the Ideal gas law to find the number of moles of butane released from the lighter.

Conclusions:

What did you determine the molar mass of butane to be by laboratory investigation?

_____ g/mol

What was your percent error?

_____ %