A Determination of the Molecular Mass of an Unknown Gas

In 1811, Amadeo Avogadro, an Italian scientist, published his now famous theorem, called **Avogadro's Hypothesis.** Avogadro postulated that "equal volumes of gases contain an equal number of molecules at the same conditions of temperature and pressure." In this experiment, a large hypodermic syringe will be filled with two different gases at the "same conditions of temperature and pressure," allowing the experimenter to readily calculate the molecular weight of an unknown gas using the ideal gas equation:

$$PV = nRT$$

Where: P = pressure in atm., $\left[\text{pressure in atm.} = \dfrac{\text{barometric pressure}}{760 \text{ torr/atm or } 29.92 \text{ in/atm}} \right]$

V = volume of gas in liters , $\left[\text{volume in liters} = \dfrac{\text{volume in milliliters}}{1000 \text{ ml/liter}} \right]$

n = number of moles of gas in the syringe
R = ideal gas constant, 0.0821 liter @atm/deg. @mol
T = the gas temperature in degrees K, $\left[°K = °C + 273.15 \right]$

The calculations will be completed in several steps; in the first step, the experimenter will determine number of moles of gas contained in the syringe. Next, the mass of air in the syringe will be calculated from the moles of air in the syringe and its molar mass. The mass of the empty syringe can be determined from the mass of the syringe containing only air. Next, the molecular mass of the unknown gas will be calculated from both the mass and moles of unknown gas in the syringe. Given the percentage composition, the empirical formula for the gas is determined, which leads to its molecular formula.

In a typical experiment, a student was assigned a gas labeled "butylene." The student determined the mass of 130 ml of air at a barometer reading of 30.35 inches of Hg and a room temperature of 19°C to be 80.42 grams. The same syringe filled with "butylene" had a mass of 80.58 grams. What is the molecular mass of "butylene?"

The first step in the calculation would be to determine the number of moles, n, of the gas in the syringe from the relationship, PV = nRT. Where

$$P = \frac{30.35 \text{ in}}{29.92 \text{ in/atm}} = 1.014 \text{ atm} \qquad\qquad R = 0.0821 \frac{\text{liter} \cdot \text{atm}}{\text{deg K} \cdot \text{mol}}$$

$$V = \frac{130 \text{ ml}}{1000 \text{ ml / liter}} = 0.13 \text{ liter} \qquad\qquad T = 19°C + 273° = 292 \text{ K}$$

Therefore,   number on moles of gas, $n = \dfrac{1.014 \text{ atm x } 0.13 \text{ liter}}{\left(0.0821 \; \frac{\text{liter} \cdot \text{atm}}{\text{deg} \cdot \text{mol}} \text{ x } 292 \text{ K} \right)} = 0.0055 \text{ mol}$

The next step in the calculations would be to find the mass of air in the syringe allowing one to calculate the mass of the empty syringe. The average molecular mass of the molecules in air is 29 grams per mole.

$$\text{number of moles of air} = \frac{\text{mass of air in the syringe}}{29 \text{ g/mol}}$$

or mass of air in the syringe = 29 g/mol x number of moles

mass of air in the syringe = 29 g/mol x 0.0055 mol = 0.16 gram of air

mass of the empty syringe = mass of the syringe plus air - mass of air

mass of the empty syringe = 80.42 grams - 0.16 grams = 80.26 grams

The mass of unknown gas is determined from the mass of the empty syringe, 80.26 grams, and the mass of the syringe plus unknown gas, 80.58 grams.

The mass of unknown gas in the syringe = mass of the syringe + unknown - mass of the empty syringe

The mass of unknown gas in the syringe = 80.58 grams - 80.26 grams = 0.32 gram

Since the molecular mass of a substance is the number of grams per one mole, then the molecular mass of the unknown gas may be calculated from the moles of gas in the syringe, 0.0055 mole, and the mass of unknown, 0.32 gram.

$$\text{Molecular mass} = \frac{\text{mass of unknown gas in the syringe}}{\text{moles of gas in the syringe}} = \frac{0.32 \text{ gram}}{0.0055 \text{ mol}} = 58 \text{ g/mol}$$

The instructor informed the experimenter that the unknown gas, "butylene," was composed of 85.71% carbon and 14.29% hydrogen. Since the percentage composition is given, the empirical formula or simple whole number ratio of the atoms of the substance in the compound may now be calculated assuming a 100 gram sample making the percentage the same as the mass.

$$\text{number of moles of carbon} = \frac{85.71 \text{ grams}}{12 \text{ grams/mol}} = 7.14 \text{ mol carbon}$$

$$\text{number of moles of hydrogen} = \frac{14.29 \text{ grams}}{1 \text{ gram/mol}} = 14.19 \text{ mol hydrogen}$$

$$\text{empirical formula} = \text{ratio of } \frac{\text{moles hydrogen}}{\text{moles carbon}} = \frac{14.29 \text{ mol H}}{7.14 \text{ mol C}} = \frac{2 \text{ mol H}}{1 \text{ mol C}}$$

empirical formula = CH_2

Since the molecular formula is a simple whole number ratio of the empirical formula, it may be calculated from both of these quantities. The first step is to find the empirical formula mass. Since the empirical formula contains one carbon atom and two hydrogen, then it would have a mass of 14 grams (1 x 12 g/mol for carbon plus 2 x 1 g/mol for hydrogen = 14 grams). Now the molecular formula may be calculated.

$$\text{Ratio} = \frac{\text{molecular mass, 58 grams}}{\text{empirical formula mass, 14 grams}} = 4.1 \approx 4$$

$$\text{Molecular formula} = \text{empirical formula x ratio} = 4 \text{ x } CH_2 = C_4H_8$$

The last step in the calculations in this experiment is too determine the percentage deviation from the expected molecular mass. Since the molecular formula is C_4H_8, the substance should have a molecular mass of 56 grams/mole (4 atoms C x 12 g/mol + 8 atoms H x 1 g/mol). However, the experimenter determined a molecular mass of 58 grams per mole, a 2 gram difference. Therefore, the experimental results differ by 2 parts in 56 parts or a 3.6% difference.

$$\text{\% difference} = \frac{|\text{Difference between the experimental and expected values}|}{\text{expected value}} \text{ x } 100$$

$$\text{\% difference} = \frac{|58 \text{ g/mol - 56 g/mol}|}{56 \text{ g/mol}} \text{ x } 100 = 3.6\% \text{ difference}$$

Materials:

Unknown gas, large hypodermic syringe assembly, nail, tip cap, thermometer.
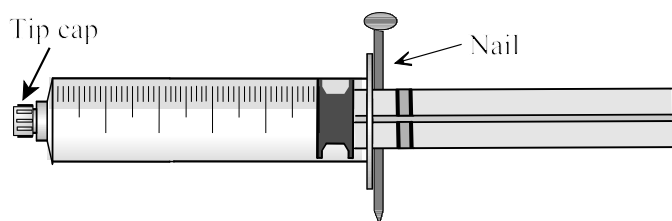
Hypodermic Syringe Assembly

Procedure:

1. Remove the tip cap from the syringe, pull back on the piston filling the syringe with air. Place a nail through the hole in the syringe piston and push the piston forward until its motion is blocked by the syringe body. The nail fixes the volume of the "equal volumes" portion of Avogadro's Hypothesis. Place the tip cap on the syringe opening to seal the apparatus, and mass the apparatus to the precision of your balance. Record the atmospheric pressure, the temperature of the air in the room, and the volume of air in the syringe.

2. Remove the nail and tip cap from the apparatus. Fill the syringe with your assigned unknown gas in the same manner as it was filled with air. Be certain to use the nail to control the volume and tip cap to seal the apparatus. Determine the mass of the syringe assembly containing the unknown gas to the precision of your balance. Repeat the trial as time permits.

Calculations and Questions:

Q1. Convert the barometer reading of the atmospheric pressure to atmospheres, the gas temperature in degrees K, and the volume of the gas in the syringe in liters.

Q2. Calculate the number of moles of gas in your syringe assembly from the ideal gas equation, $PV = nRT$.

Q3. (a) Calculate the mass of air in the syringe from the number of moles of gas in the syringe and the average molecular mass of air, 29.0 grams/mole.

    (b) Calculate the mass of the empty syringe by subtracting the mass of air in the syringe from the mass of air plus the syringe assembly.

Q4. (a) In each trial, calculate the mass of unknown gas in the syringe by subtracting the mass of the empty syringe assembly from the mass of the syringe assembly plus the gas.

    (b) In each trial, calculate the molar mass of your unknown gas from the mass of gas in the syringe (Q4a) and the number of moles of gas in the syringe (Q2).

    (c) Calculate the average molar mass of your unknown.

Q5. Given the percentage composition for you unknown gas, determine its empirical formula. A common gas which may be studied in this experiment is Freon 134a, an environmentally friendly refrigerant. It is composed of 23.53% carbon, 1.96% hydrogen, and 74.51% fluorine. Another fluorocarbon which may be studied is an electronic parts duster, Freon 152a, has a composition of 36.37% carbon, 6.10% hydrogen, and 57.53% fluorine. Your instructor will give you the percentage composition of other gases studied.

Q6. (a) Determine the molecular formula of your unknown gas from its empirical formula (Q5) and its average molecular mass (Q4). The molecular formula is an integer multiple of the empirical formula. To find the ratio, calculate the mass of the substance according to its empirical formula determined above and divide that mass into the molecular mass. Round the resulting quotient to the nearest whole number.

    (b) Determine the percentage difference between your experimental molecular mass and the value expected according to the molecular formula found in question Q6a.

DATA:

Identity of the unknown gas:_____

Atmospheric Pressure:_____Units:_____

Room Temperature:_____°C

Volume of gas in the syringe:_____ ml

Mass of the syringe assembly plus air:_____ g.

Moles of air in the syringe:_____ mol

Mass of air in the syringe:_____ g.

Mass of the empty syringe assembly:_____ g.

Trial	1	2	3
Mass of syringe plus gas	g.	g.	g.
Mass of gas in the syringe	g.	g.	g.
Molecular mass of the gas	g/mol	g/mol	g/mol

Average molecular mass of the unknown gas: _____ g/mol

Molecular formula for you unknown gas: _____

Notes for Teachers

1. There are a number of reasonably pure gases which are readily available locally for this experiment. The higher the molecular mass, the better the results. Two inexpensive gases are 1,1-difluoroethane, called Freon R152a with a molar mass of 66 g/mol, and 1,1,1,2-tetrafluoroethane, known as Freon R134a with a molar mass of 102 g/mol. Both gases may be purchased as an electronic duster at office supply stores. R134a is also used as a refrigerant and may be purchased in the auto supply section of discount stores. A Luer Loc to $^1/_{16}$" tubing adapter is needed to attach a syringe to the container of Freon. This adapter fits the duster can perfectly. Wet the serrated end of the adapter with hot glue, airplane cement, or even epoxy and push the adapter into the nozzle. Both the trigger assembly and the Luer Loc adapter are made of polyethylene which does glue easily but the glue acts as a sealant.

    Some other gases which made be purchased at a local hardware store or home repair center are propane, propylene, butane, and oxygen. However, the two Freon's have a higher molecular mass making them better for this experiment.

2. The larger the syringe, the lower the error in this experiment. The largest syringe currently available is a 140-ml syringe. To prepare the syringe, position the piston so that it reads 140-ml and either drill a hole for the nail or heat the nail in a Bunsen burner and push it through the syringe.
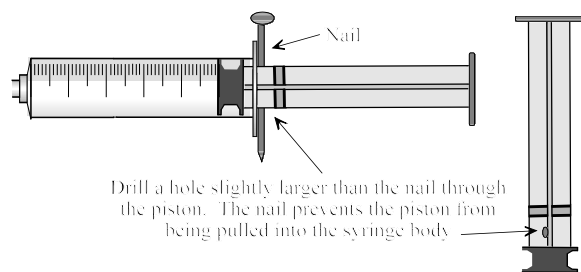
3. Source of materials:

    Nails - 8 penny common nails

    All of the specialized materials are available from Micro Mole Scientific Company, P.O. Box 847, Pasco, WA 99301. Phone/fax: (509) 545-4904. Micro Mole sells the 140-ml syringes, the tip caps, and Luer Loc to tubing adapters.

    Each pair of students would need one 140-ml syringe, one tip cap, and a nail. A Luer Loc to $^1/_{16}$" tubing adapter is needed for each can of gas. Only one can of each gas is required in this experiment.