Gas Laws Equations

The Ideal Gas Law $PV = nRT$ | $R = 0.082058 \frac{L\,atm}{mol\,K}$ (5 SF)

or

$0.0821 \frac{L\,atm}{mol\,K}$ (3 SF)

— used to calc one variable
Given other 3

Combined Gas Law $\dfrac{P_1 V_1}{n_1 T_1} = \dfrac{P_2 V_2}{n_2 T_2}$ | T must be in K

— used for changing conditions
— Conditions that do not change (held constant)
Can be dropped.

— Boyle's Law
— Charle's Law
— Avogadros' Law

Dalton's Law of Partial Pressure

$\Rightarrow$ $P_{Total} = P_1 + P_2 + P_3 + \ldots + P_n$

$P_{Total} = \dfrac{n_1 RT}{V} + \dfrac{n_2 RT}{V} + \dfrac{n_3 RT}{V} + \ldots + \dfrac{n_n RT}{V}$

$\Rightarrow$ $P_{Total} = (n_1 + n_2 + n_3 + \ldots + n_n) \dfrac{RT}{V}$ ← Sum of all moles gas

Some questions combine

$$PV = nRT \quad \text{with} \quad \text{molar mass} = \frac{g}{mole}$$

or

$$\text{with} \quad \text{density} = \frac{g}{L}$$

Molar Mass

$$\text{mole} \times \text{molar mass} = \frac{g}{mole} \times mole$$

$$\frac{\text{mole} \times \text{molar mass}}{\text{molar mass}} = \frac{g}{\text{molar mass}}$$

$$\text{mole} = \frac{g}{\text{molar mass}}$$

$$n = \frac{g}{\text{molar mass}} \qquad \text{Substitute into } PV = nRT$$

$$\boxed{PV = \left(\frac{g}{\text{molar mass}}\right) RT}$$

$$PV = \frac{gRT}{\text{molar mass}}$$

$$\boxed{\text{molar mass} = \frac{gRT}{PV}}$$

Density

$$L \times \text{density} = \frac{g}{L} \times L$$

$$\frac{L \times \text{density}}{\text{density}} = \frac{g}{\text{density}}$$

$$L = \frac{g}{\text{density}}$$

$$V = \frac{g}{\text{density}} \qquad \text{Substitute in for } PV = nRT$$
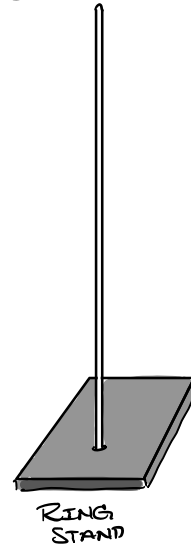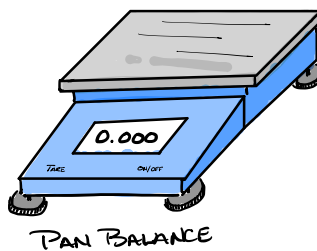
$$\boxed{P\left(\frac{g}{\text{density}}\right) = nRT}$$

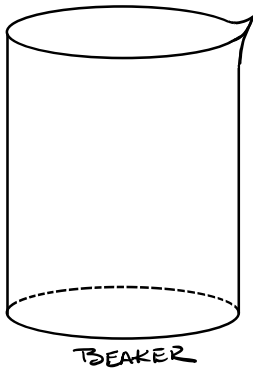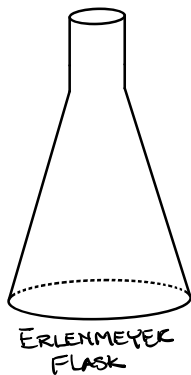$$\text{density} \times \frac{Pg}{\text{density}} = nRT \times \text{density}$$

$$\frac{Pg}{nRT} = \frac{nRT \times \text{density}}{nRT}$$

$$\boxed{\text{density} = \frac{Pg}{nRT}}$$

Molar Mass of a Gas

Goal: Experimentally determine the molar mass of an unknown gas.

ERLENMEYER FLASK

BEAKER

GAS TANK

PAN BALANCE

RING STAND

THERMOMETER

RUBBER TUBING w/ GLASS U-BEND

RING

RUBBER STOPPER

SCOOPULA
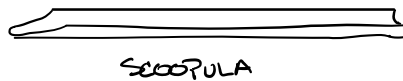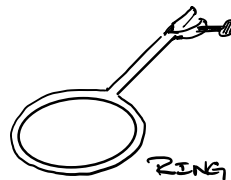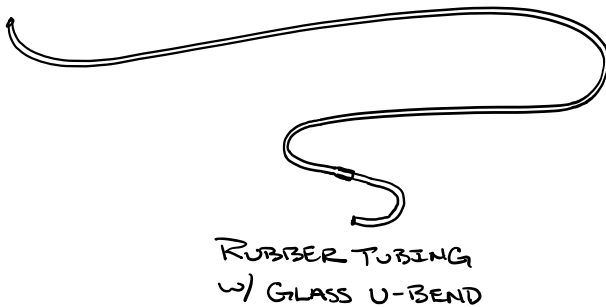
The Molar Mass of a Gas

Goals

❑ The purpose of this experiment is to determine the number of grams per mole of a gas by measuring the pressure, volume, temperature, and mass of a sample.

Terms to Know

Molar Mass - The number of grams per mole of a substance.

Ideal Gas Law - The relation between pressure, volume, number of moles, and temperature of a gas: $PV = nRT$

Vapor Pressure of Water - The pressure exerted by water vapor as it saturates the environment above a sample of liquid water. Water vapor pressures depend upon the temperature, and tabulated values are found in reference tables.

Introduction

The molar mass of a gaseous compound can be determined by experiment even though the formula or composition of the compound is not known. In other words, it is possible to find the molar mass of a gas even if we do not know its identity. The molar mass is determined by using the fact that the number of moles of a gas sample relates to the pressure, volume, and temperature by the gas laws. The mass, pressure, volume and temperature of a gas sample are measured experimentally and, used to calculate the molar mass.

To determine the molar mass of a gaseous substance from the mass, volume, temperature, and pressure of a sample, the number of moles in the sample is first calculated using the ideal gas law, $PV = nRT$. Recall that n represents the number of moles. The number of moles of a gas is found from the pressure, volume, and temperature of a sample: $n = PV/RT$ where R is the universal gas constant.

As an example, consider a 0.508 g sample of a gas that occupies 522 mL volume at 100 oC and 0.960 atmospheres. First summarize the data as:

mass	P	V	T
0.508 g	0.960 atm	0.522 L	373 K (100 + 273)

The number of moles of gas is found using the ideal gas law. (Fill in the following blanks and calculate the number of moles.)

$$n = \frac{\text{atm} \times \text{L}}{0.0821 \frac{\text{L atm}}{\text{mole K}} \times \text{K}} = \text{mole}$$

The molar mass of the gas is found from the ratio of the mass of the sample to the moles in the sample. Divide the mass of the sample by the number of moles using the value calculated above.

$$\frac{0.508 \text{ g}}{\text{mole}} = \frac{\text{g}}{1 \text{ mole}}$$

In this experiment, a sample of gas is collected by water displacement. That is, the gas bubbles into a container of water and as the gas accumulates, it displaces the water. The volume of the sample is the volume the gas occupies in the container. The temperature is found by measuring the temperature of the water in contact with the gas. The mass of the gas sample is measured by weighing a small tank or cylinder of gas, delivering the sample, and reweighing the cylinder.

The pressure of the sample is found by making sure that the pressure of the gas sample is the same as the atmospheric pressure. The atmospheric pressure is measured with a barometer. However, when a gas is collected by water displacement, it becomes saturated with water vapor. This means that once the gas sample is collected, it will be a mixture of the gas and water vapor. This does not affect the volume or temperature of the gas but the measured pressure is the pressure of the mixture. That is, the measured pressure is the total pressure of the gas and the water vapor. To determine the pressure of the gas sample, the pressure of the water vapor can be subtracted from the total pressure. Appendix 2 lists the vapor pressure of water at various temperatures. To find the pressure of the gas sample in the experiment, look up the vapor pressure of water at the measured temperature and subtract this pressure from the measured atmospheric pressure:

$$P_{GAS} = P_{ATM} - P_{WATER}$$

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

Experimental Procedure

For this experiment you will need a 1-L beaker, a 250-mL Erlenmeyer flask, and a thermometer.

Place a piece of magic tape or a length of a gummed label along the neck of the flask near the top. Fill the flask to the brim with tap water and put about 500 mL of water into the beaker. Allow the water to stand for a while to be sure that it is at room temperature. Set up the apparatus as shown in Figure 19-1 below.
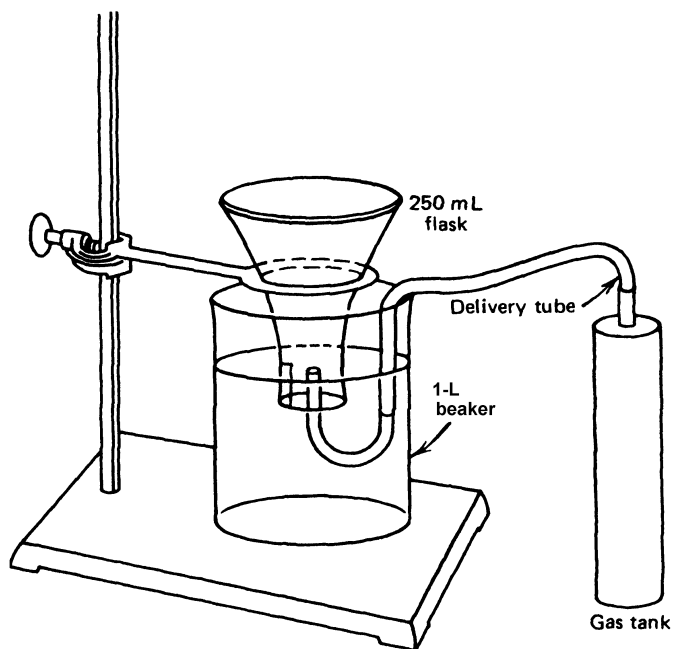

Figure 19.1   Apparatus for collecting gas samples

Gently stopper the flask with a rubber stopper and do not allow any air to become trapped in the flask under the stopper. Invert the flask through the iron ring which will hold it in position. Lower the ring with the flask so that the mouth of the flask is below the water level in the beaker. Secure

the ring to support the flask. Use a spatula to remove the rubber stopper from the mouth of the flask. The stopper can remain in the beaker since it will not interfere with the experiment.

Obtain a small tank of gas. Use a paper towel to gently wipe the tank to make sure that it is clean. Weigh the tank to 0.01 g.

Attach a plastic delivery tube to the valve on the tank and insert the glass tube on the other end into the mouth of the flask in the beaker of water. Push the valve on the tank straight down to deliver gas to the flask. Continue to deliver gas until the flask is nearly full of gas. That is, deliver gas until the gas level is near the neck of the flask. Do not allow any gas to escape from the flask. If any gas escapes, you will have to start the experiment over.

Carefully remove the delivery tube from the gas tank, gently clean and dry the tank with a paper towel and weigh the tank to 0.01 g. Place a thermometer in the beaker.

Move the flask of gas so that the level of water in the flask is the same as the water level in the beaker. This will adjust the pressure in the flask to atmospheric pressure. Use a pen or pencil to place a mark on the tape to indicate the water level.

Remove the flask from the water and fill it with water to the level corresponding to the mark on the tape. This volume of water will correspond to the volume of the gas sample. Carefully pour the water from the flask into a graduated cylinder and measure the volume to the nearest 1 mL. Record this as the volume of the gas sample.

Record the temperature of the water in the beaker. Assume that the temperature of the gas sample is the same as the water.

Record the atmospheric pressure and look up the vapor pressure of water at the temperature of the gas in the table at the end of this experiment.

Repeat the experiment to obtain a second set of data. Calculate the molar mass of the gas using each set of data. If the results are within 10 percent of one another, express your answer as the average of the two results. If your two results seem to be too far apart, consult your instructor.

Vapor Pressure of Water

Temperature °C	Vapor Pressure torr	Temperature °C	Vapor Pressure torr
0	4.6	25	23.8
5	6.5	26	25.2
10	9.2	27	26.7
15	12.8	28	28.3
16	13.6	29	30.0
17	14.5	30	31.8
18	15.5	40	55.3
19	16.5	50	92.5
20	17.5	60	149.4
21	18.6	70	233.7
22	19.8	80	355.1
23	21.2	90	525.8
24	22.4	100	760.0

Questions for Analysis

Answer the following questions in your laboratory notebook:

1. Referring to the experimental determination of the molar mass, explain how and why each of the following factors would affect your calculated molar mass. That is, would the calculated value be greater than it should be, less than it should be, or not changed? Note that the relationships between the various factors involved in the calculation are:

   molar mass = g/mol = g/(PV/RT) = $\dfrac{gRT}{PV}$

   a. The measured temperature is a lower value than the actual temperature.
   b. The measured volume is a higher value than the actual volume.
   c. Some of the gas sample escapes from the tank before it reaches the flask.

2. The molar mass of a gas is determined by collecting a gas sample by water displacement.

   a. Using the following data, calculate the molar mass of the gas: Sample volume, 163 mL; temperature, 21.0 °C; mass, 0.281 g; total (i.e., barometric) pressure, 752 torr.
   b. The gas in part (a) contains 85.5% C and 14.5% H. First, determine the empirical formula, and then use the result from part (a) to determine the actual formula.

$$\text{Molar Mass} = \frac{g\,RT}{PV}$$
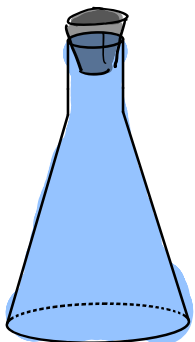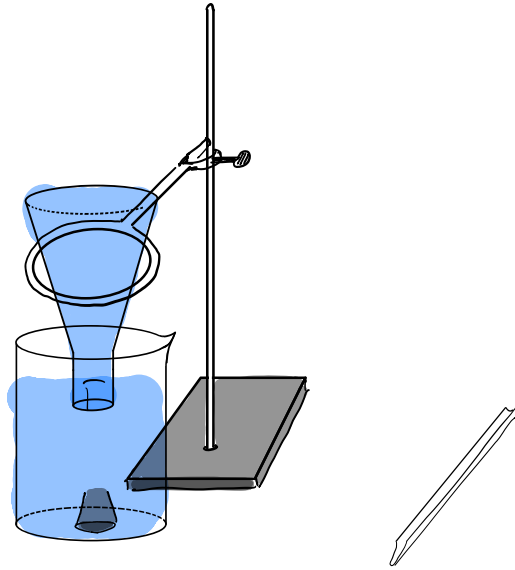
mass gas = grams

Temperature = K

Pressure = atm

Volume = L

$$R = 0.082058 \frac{L\,atm}{moles\,K}$$

$$= \frac{(grams)\left(\frac{L\,atm}{mol\,K}\right)(K)}{(atm)(K)}$$
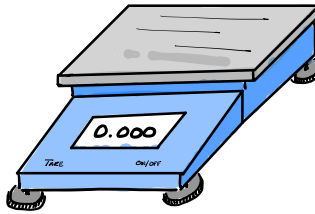
$$= \frac{grams}{moles} = g/mole \quad \checkmark$$
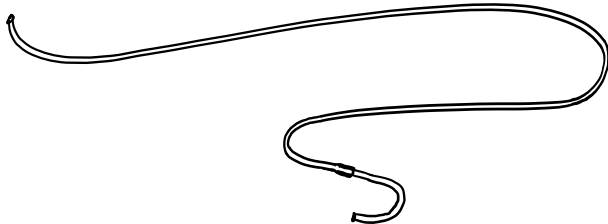
ERLENMEYER
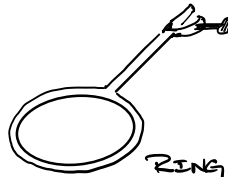FLASK

BEAKER

GAS
TANK

Gas

0.000

PAN BALANCE

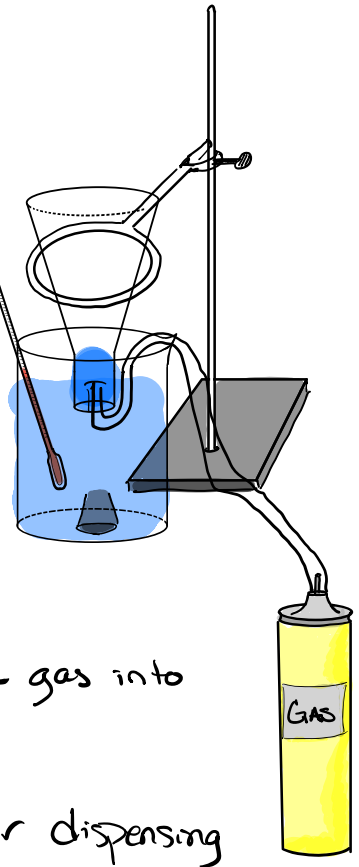RING
STAND

THERMOMETER

RUBBER TUBING
W/ GLASS U-BEND

RING

RUBBER
STOPPER

SCOOPULA

① Fill erlynmeyer flask w/ H_2O

② Stopper flask

③ place ~ 700 mL H_2O into beaker

④ Invert erlenmeyer flask into beaker & Remove Stopper

⑤ Weigh the gas tank
   — 826.20 g

⑥ Use the U-tube to dispense gas into erlenmeyer flask.
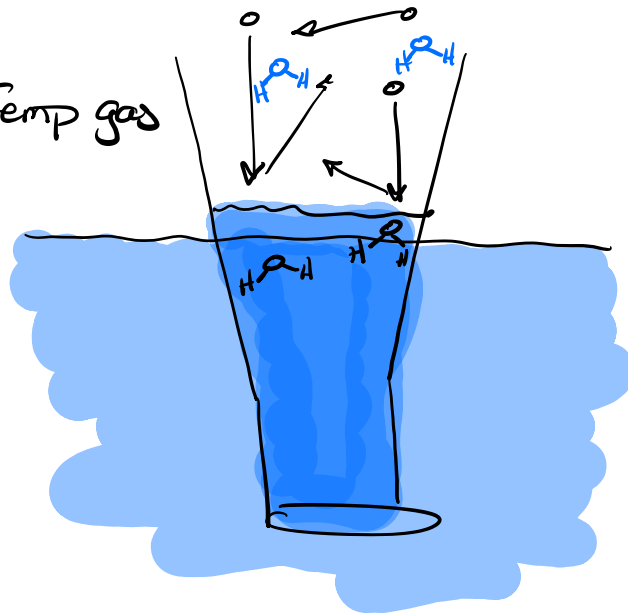
⑦ Re-weigh gas tank after dispensing gas.

   — 825.60 g

⑧ Mass of gas used in experiment

   826.20 g
 − 825.60 g
 _____
   0.60 g used

$$\text{molar mass} = \frac{\overset{\vee}{g} RT}{PV}$$

⑨ use thermometer to take temperature
of H_2O

Temp H_2O = Temp gas

Temp = 21.2 °C

molar mass = $\dfrac{gRT}{PV}$

?

⑩ Volume & Pressure

$P_{atm} > P_{gas} + P_{H_2O}$

flask
up
&
down

P_{gas}

P_{atm}

← equality

$P_{atm} = P_{gas} + P_{H_2O}$

$P_{gas} + P_{H_2O}$

P_{atm}

Goal

$P_{atm} < P_{gas} + P_{H_2O}$

P_{atm}

P_{gas}

$$P_{atm} = 752.56 \text{ mmHg}$$

← mark

fill to
mark
w/ H_2O

← mark

$V = 237 mL$
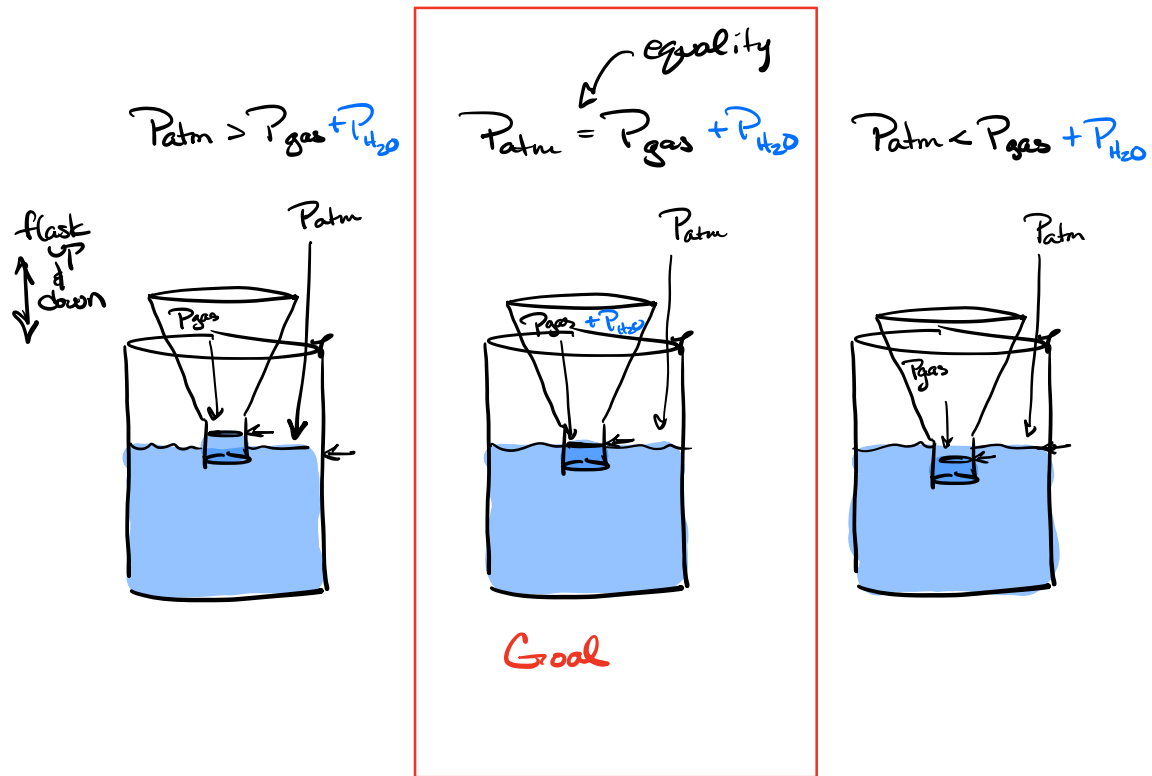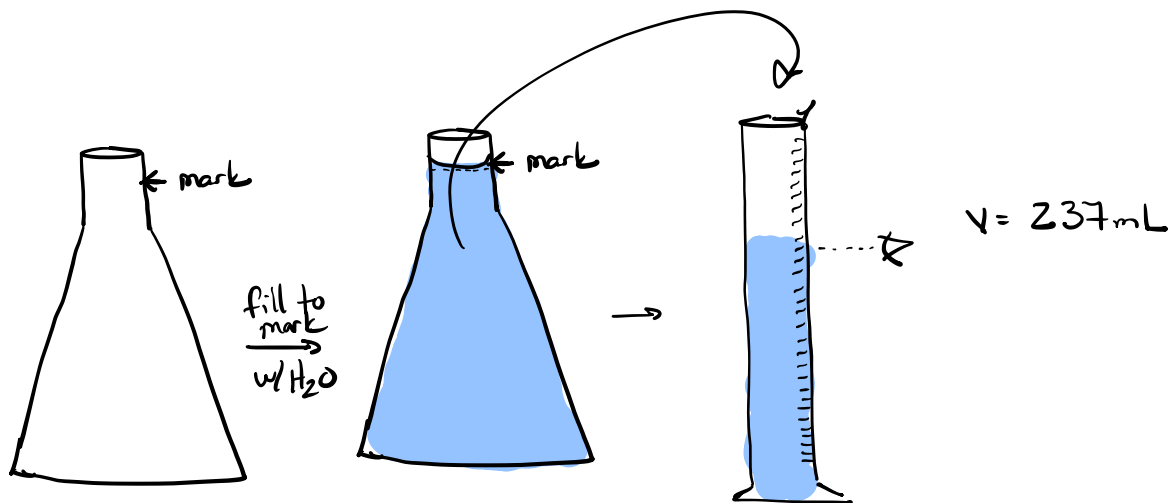
P_{H_2O} is based on temperature

Vapor Pressure of Water

Temperature °C	Vapor Pressure torr	Temperature °C	Vapor Pressure torr
0	4.6	25	23.8
5	6.5	26	25.2
10	9.2	27	26.7
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16	13.6	29	30.0
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23	21.2	90	525.8
24	22.4	100	760.0

Temp = 21.2 °C

$$19.8 - 18.6 = \frac{1.2}{10} = 0.12$$

$$2 \times 0.12 = 0.24$$

$$\begin{array}{r} 18.6 \\ 0.24 \\ \hline 18.84 \end{array} = 18.8 \text{ torr} = P_{H_2O}$$

$$P_{atm} = P_{gas} + P_{H_2O}$$

$$P_{gas} = P_{atm} - P_{H_2O}$$

$$P_{atm} = 752.56 \, mmHg \qquad 1 \, mmHg = 1 \, torr$$

$$P_{H_2O} = -18.8 \, torr$$

$$\underline{\phantom{P_{gas} = 733.76 \, mmHg}}$$

$$P_{gas} = 733.76 \, mmHg$$

$$P_{gas} = 733.76 \, mmHg \times \frac{1 \, atm}{760 \, mmHg} = 0.965473684 \, atm$$

$$T = 21.2°C + 273.15 = 294.35 \, K$$

$$V = 237 \, mL \times \frac{1 \, L}{1000 \, mL} = 0.237 \, L$$

$$mass = 0.60 \, g$$

$$Molar \, Mass = \frac{g \, RT}{PV}$$

$$= \frac{(0.60 \, g)(0.082058 \frac{L \cdot atm}{mol \cdot K})(294.35 \, K)}{(0.965473684 \, atm)(0.237 \, L)}$$

$$= 63.335533 \, g/mole$$

$$\boxed{= 63 \, g/mole}$$

Repeat experiment 3x
and average Result

We have trial # 1 here

I'll put up data for other 2 trials

- Calc other 2 results
  & Average the 3 answers

- Questions at end to answer