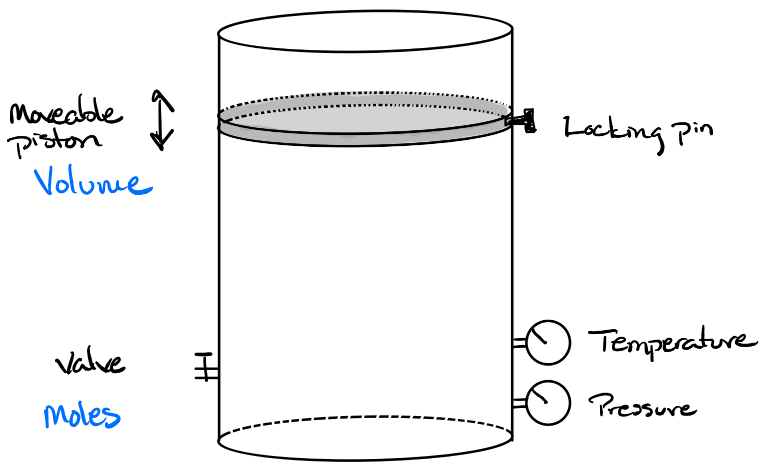
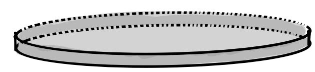
Gas Laws

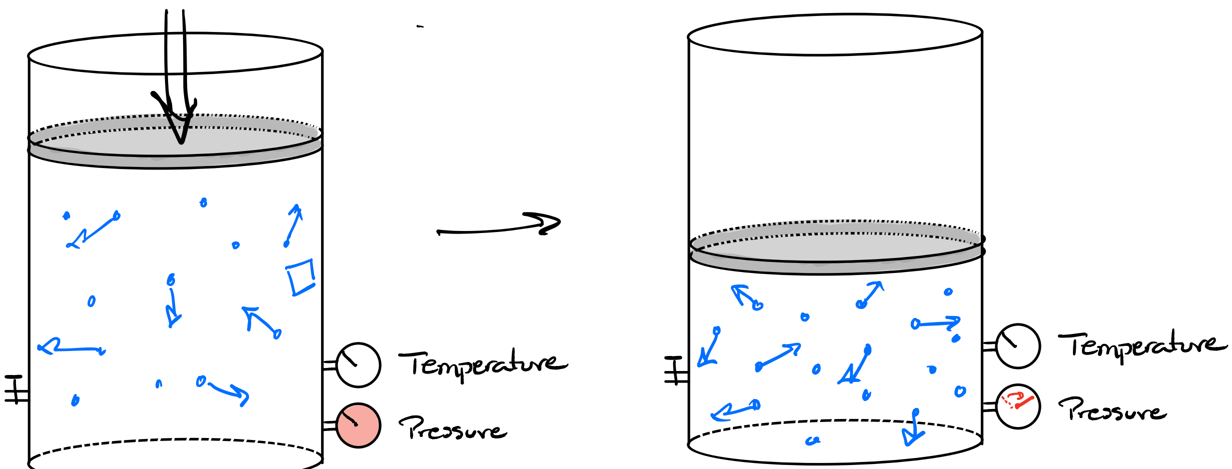




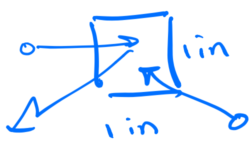
Assumptions "Ideal Gas"

- Travel in straight lines unless they hit something.
  - unaffected by gravity
  - No intermolecular attractive forces
- Energy transferred in collisions
- point mass (particles have no volume)
- On average N_2 molecule at room temp moving 500 mph

① Look at Pressure & Volume
Hold Temp & moles constant



- less Collisions w/ unit area when gas more diffuse



pressure/in^2

lbs/in^2

Voume goes down
pressure goes up

- As gas gets Concentrated
# of Collisions per unit area goes up

More collisions = More force

More force per unit area = higher pressure

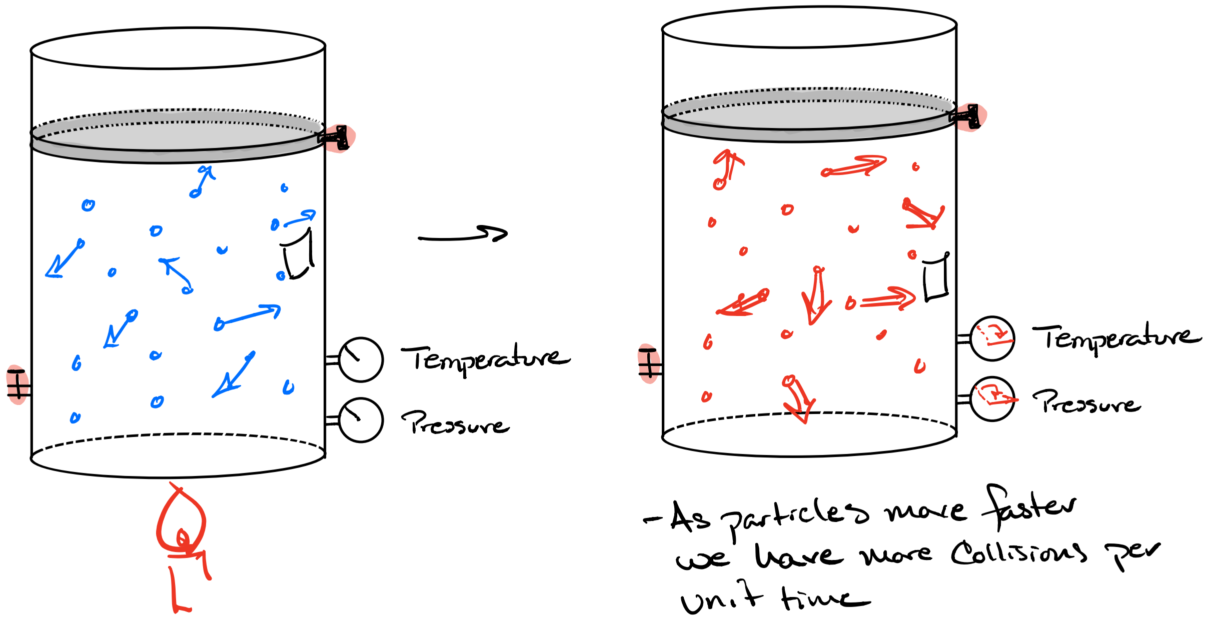
α = Proportional to

$V \downarrow \quad P \uparrow$

Inversely Proportional

$\uparrow P \propto \frac{1}{V \downarrow} \downarrow$

Pressure Vs. Temp
(moles & Volume Constant)



- As particles move faster we have more Collisions per unit time
- More Collisions = more force
- Faster particles = Stronger Collisions
  - Increase in force per unit area

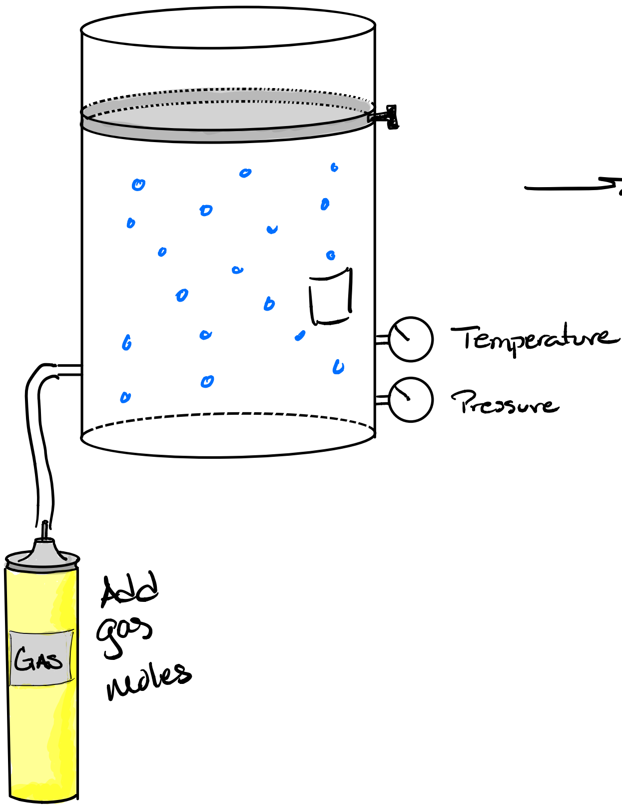
$\Rightarrow$ Increase in pressure

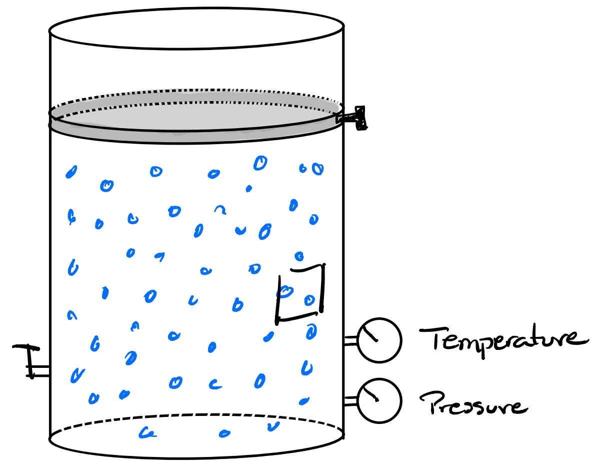
$$\uparrow T \propto P \uparrow$$

Directly Proportional

Pressure vs. Moles

Temperature & Volume Constant





As more gas particles are added ⇒ The number of collisions goes up.

\# of collisions $\propto$ pressure

pressure goes up

$$\uparrow P \propto n \text{ (moles)} \uparrow$$

Directly proportional

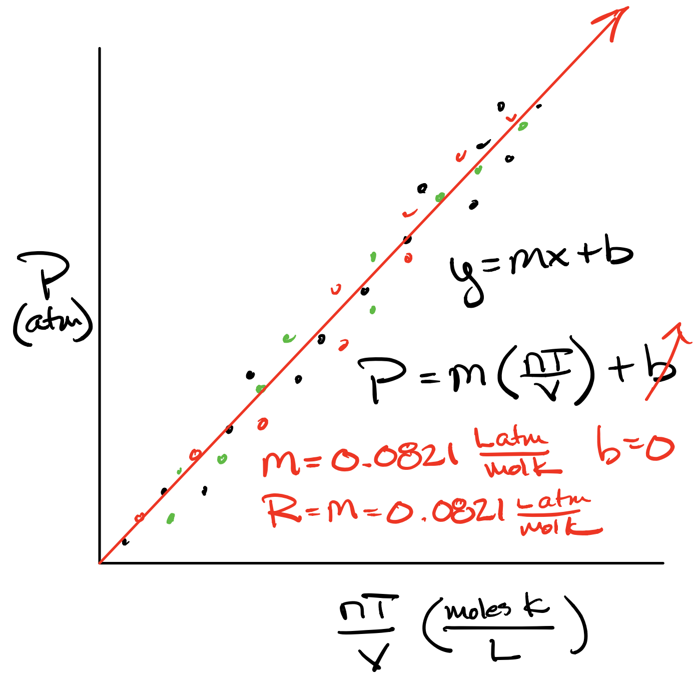
$$P \propto \frac{1}{V}$$

$$P \propto T$$

$$+\ P \propto n$$

$$P \propto \frac{nT}{V}$$

?



$$V \times P = R\,\frac{nT}{\cancel{V}} \times \cancel{V}$$

$$PV = nRT$$

R is a $\propto$ constant. It is experimentally measured

$$R = 0.0821\ \frac{L\,atm}{mol\,K}$$

$PV = nRT$ Ideal Gas law

Ideal Gas Law problems

A container at 25°C has a volume of 0.75 L and contains 0.023 g of nitrogen gas (N_2). What is the pressure in the container?

$$R = \underbrace{0.0821}_{3\text{ sig figs}} \frac{L\ atm}{mole\ K}$$

$P = ?$

$V = 0.75\ L$

$n = 0.023\ g\ N_2 \times \frac{1\ mole\ N_2}{28.02\ g\ N_2} = 8.208 \times 10^{-4}\ moles\ N_2$

$R = 0.0821\ \frac{L\ atm}{mol\ K}$

$T = 25°C + 273.15 = 298.15\ K$

$N = 14.01\ g/mole$

$N_2 = 2(14.01) = 28.02\ g/mole$

$PV = nRT$

$$P = \frac{nRT}{V} = \frac{(8.208 \times 10^{-4}\ \cancel{moles})(0.0821\ \frac{\cancel{L}\ atm}{\cancel{mol}\ \cancel{K}})(298.15\ \cancel{K})}{0.75\ \cancel{L}}$$

2, 3, 3, 2

$= 0.026790\ atm$

$= \boxed{0.027\ atm}$

Units

$\frac{L\ atm}{mol\ K}$

Volume in Liters

Moles in Moles (mass g → moles)

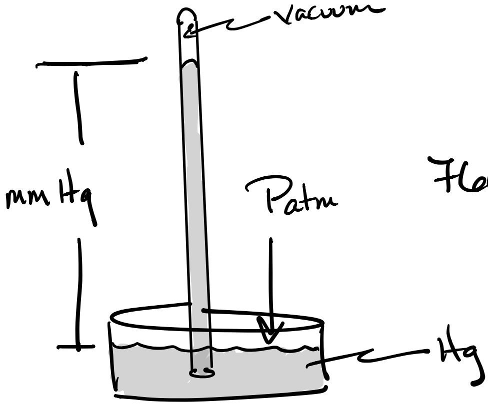
Temp K (°C + 273.15) $(°F - 32) \times \frac{100°C}{180°F} = °C$

Pressure atm

atm
Torr
mmHg

lbs/in²
pascals

1 atm = 760 mmHg = 760 torr
1 mmHg = 1 torr



$760 \text{ mmHg} = 1 \text{ atm}$

Molar Mass of a gas

Measure

mass
Vol
Temp
Pressure

Calc Molar Mass of a gas

$$PV = nRT \qquad Eq\ 1$$

$$\text{Molar Mass} = \frac{\text{Mass}}{\text{Moles}} \qquad Eq\ 2$$

$$\text{moles} = n$$

$$\text{Molar Mass} = \frac{\text{mass}}{n}$$

Solve for n
& Sub into Eq 1

$$n = \frac{\text{mass}}{\text{Molar Mass}}$$

$$PV = nRT$$

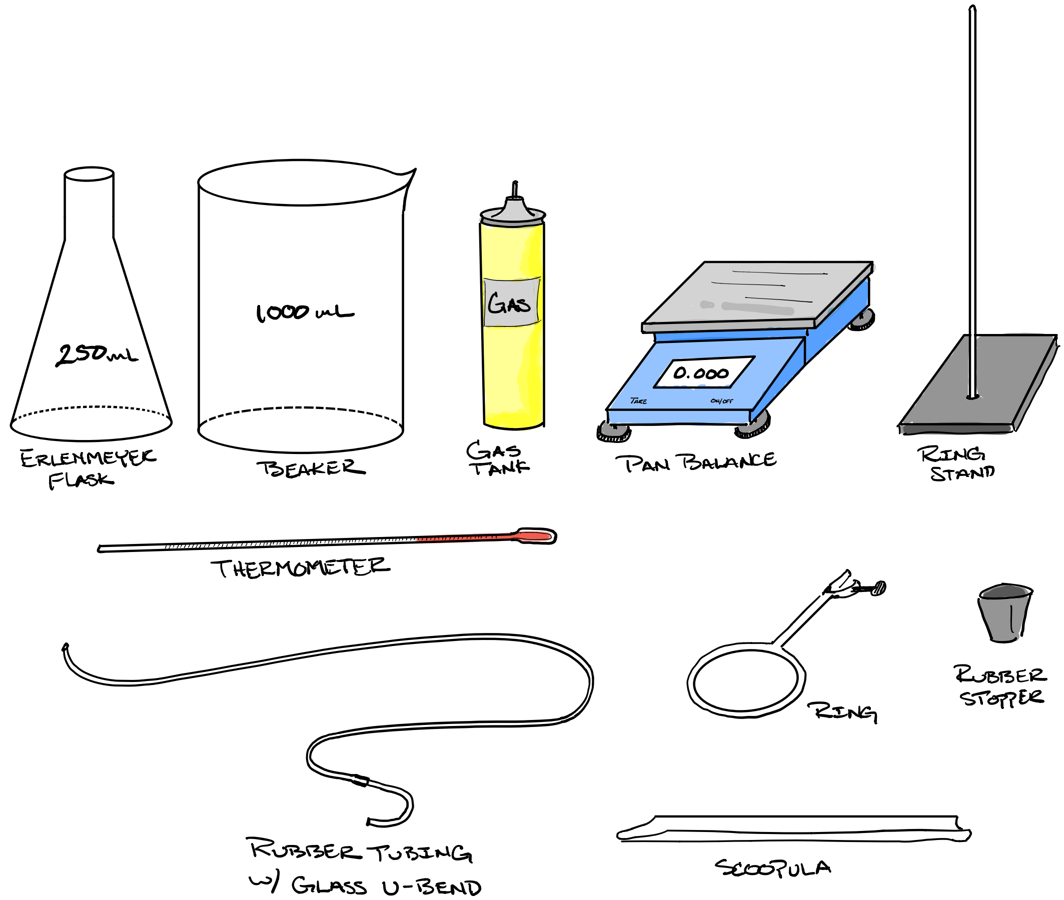
$$PV = \frac{\text{mass}\,RT}{\text{Molar Mass}}$$

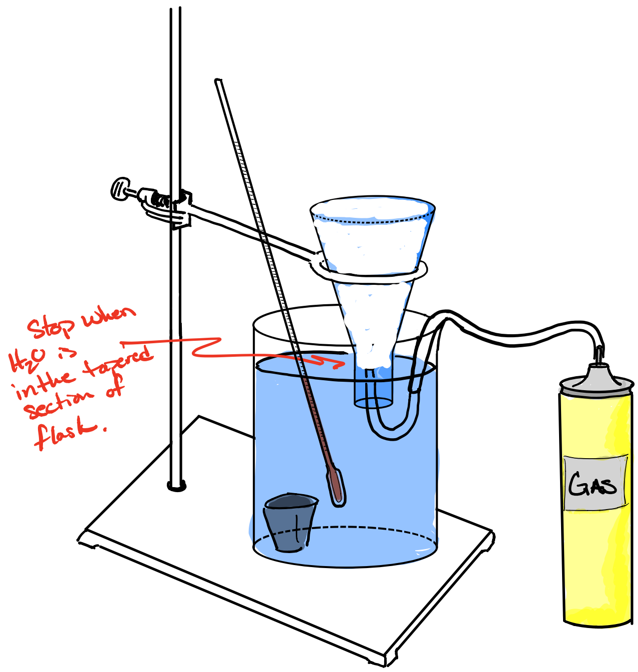
$$\text{Molar Mass}\,PV = \text{mass}\,RT$$

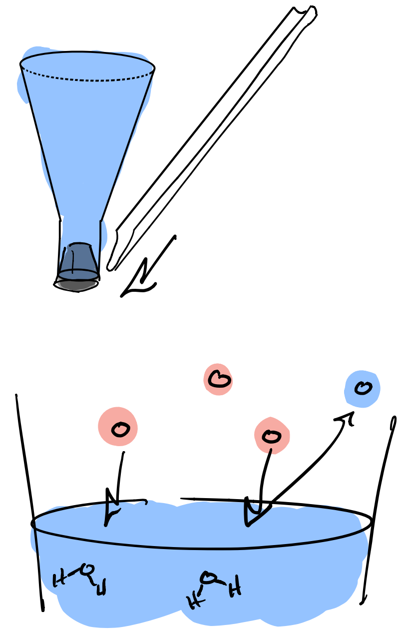
$$\text{Molar Mass} = \frac{\text{mass}\,RT}{PV}$$

$$= \frac{g\left(\frac{\cancel{L}\,\cancel{atm}}{mol\,\cancel{K}}\right)\cancel{K}}{\cancel{atm}\,\cancel{L}} = \frac{g}{mole}$$

= molar Mass


250 mL
ERLENMEYER FLASK
1000 mL
BEAKER
GAS
GAS TANK
0.000
PAN BALANCE
RING STAND
THERMOMETER
RING
RUBBER STOPPER
RUBBER TUBING W/ GLASS U-BEND
SCOOPULA





1) fill flask w/ water.
2) invert into beaker full of H_2O ⇒ 700 mL
3) measure mass gas tank

   mass $Tank_0$ 826.2 g

4) use the U-tube to put gas into flask displacing the H_2O

5) Reweigh gas tank

   mass $Tank_f$ = 825.6 g

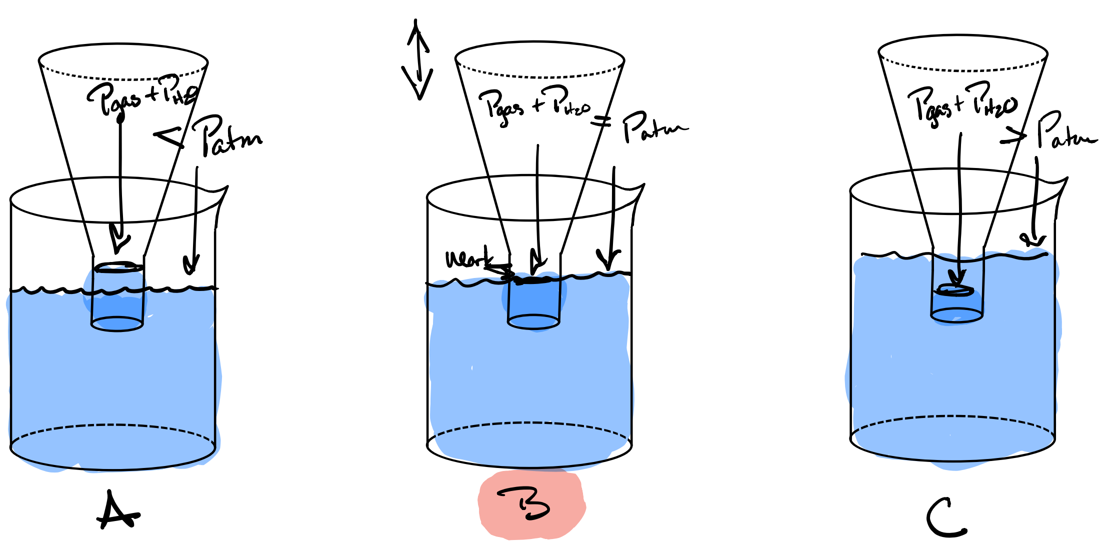
6) Difference between mass = mass gas

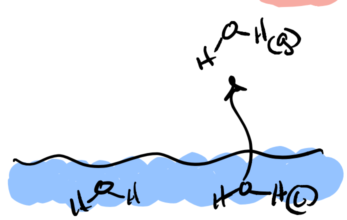
$$\begin{array}{r} 826.20g \\ 825.60g \\ \hline 0.60g \end{array} = \text{mass gas}$$

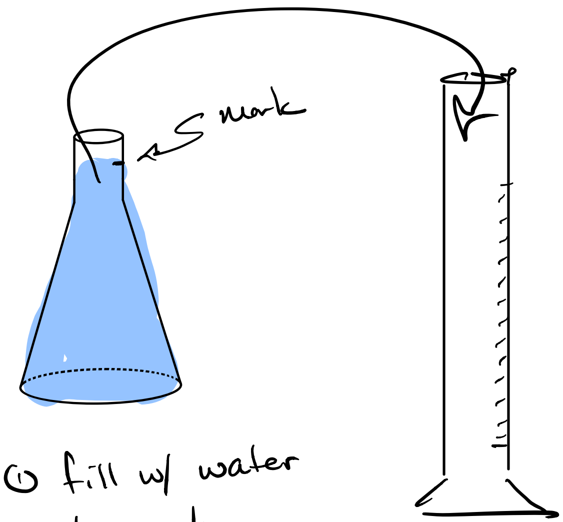
$$\text{Molar Mass} = \frac{\text{mass}\,RT}{PV}$$

7) measure temp of the H_2O → temp gas

21.2°C







① fill w/ water to mark

$V = 237\,mL$

Pressure

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

Dalton's Law of Partial Pressure

$P_T = P_1 + P_2 + P_3 + \dots + P_n$

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

related to temp found on table 18.8 torr

Measured using a barometer 752.56 mmHg

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

Mass (g) 0.60 g

Volume 237 mL

Pressure atm 752.56 mmHg

Pressure H_2O 18.8 torr

Temperature 21.2 °C

Mass (g) 0.60 g

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

Pressure atm 752.56 mmHg

Pressure H_2O 18.8 torr

Temperature 21.2 °C + 273.15 = 294.35 K

$P_{gas} = P_{atm} - P_{H_2O} = 752.56\ mmHg - 18.8\ torr$

$P_{gas} = 733.76\ mmHg$

$P_{gas} = 733.76\ mmHg \times \frac{1\ atm}{760\ mmHg} = 0.96547\ atm$

$$\text{Molar Mass} = \frac{\text{mass}\ RT}{PV} =$$

$$\frac{(\overset{2}{0.60\,g})(\overset{3}{0.0821}\,\frac{L\,atm}{mol\,K})(\overset{5}{294.35\,K})}{(\underset{5}{0.96547\,atm})(\underset{3}{0.237\,L})}$$

$= 63.368$ g/mole → $\boxed{63\ \text{g/mole}}$

$\boxed{55\text{ g/mole} - 64\text{ g/mole}}$

Butane $C_4H_{10} = 58.125$ g/mole

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:

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

   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.

Ex ? molar mass = $\frac{\downarrow gRT}{PV}$

don't forget that $P_{gas} = P_{atm} - P_{H_2O}$

$\downarrow \text{value} = \frac{\downarrow \text{num}}{\text{denom}}$

$\text{value} \propto \text{num}$

$\text{value} \propto \frac{1}{\text{denom}}$