

Name \_\_\_\_\_

## Molar Mass of Butane Lab

*Purpose:* To determine the molar mass of butane

*Pre-lab Discussion:* In this experiment you will collect a volume of butane gas and use the ideal gas law equation ( $PV = nRT$ ) to determine the molecular mass of the gas. In this equation P represents pressure, V is volume, n is the number of moles of the gas, R is the ideal gas constant, and T is the Kelvin temperature. The number of moles (n) can be rewritten as being equal to the mass of the gas (m) divided by the molar mass of the gas (M). A little algebra enables us to rewrite the equation as  $M = mRT/PV$ . There are many values of the gas constant R since there are many allowable units for pressure and volume, but one value is  $62.4 \times 10^3$  torr-ml / mole-K.

The pressure of the gas will be adjusted to atmospheric pressure and the mass of the butane will be measured. Our source of butane will be a cigarette lighter and we will collect the gas by water displacement. This is possible because butane is not very soluble in water. Dalton's Law of Partial Pressures will enable us to determine the pressure of the butane gas alone in the cylinder.

### *Procedure*

1. Weigh the lighter to the nearest 0.01 grams or better. Attach a plastic tube to the gas opening on the top of the lighter.
2. Submerge a 100. ml graduated cylinder in water so that the cylinder fills completely with water. Invert the cylinder. Make sure there are no air bubbles remaining in the graduated cylinder.
3. Take the plastic tube and insert it in the graduated cylinder under the water. Carefully release the butane from the lighter and collect it in the cylinder. Release enough butane to fill the tube to within 5 ml of its graduated capacity. (95 ml for a 100 ml cylinder). Remove the plastic tube from the graduated cylinder and the lighter.
4. Allow the butane in the cylinder to reach room temperature (about five minutes). Then adjust the level of the water inside and outside the tube to be the same. With the pressure inside the same as the pressure outside, record the volume to the nearest ml. This makes the pressure of the combined water vapor and butane gas equal to the pressure of the atmosphere.
5. Measure the mass of the lighter again.
6. Record the air temperature in Kelvin.
7. Record the water temperature and barometric pressure.
8. Use the derivation of the ideal gas law equation ( $M = mRT/PV$ ) to find the molar mass.
9. Repeat the entire procedure 2 more times for a total of 3 trials. Find the average and the % error.

*Lab report:*

This lab is to be written up as a real lab report. The write-up will be worth 20 pts. See the handout for details of how to write the report. The handout can also be found at [http://www.apsu.edu/ROBERTSONR/chem1110-20/chem\\_1110-20.htm](http://www.apsu.edu/ROBERTSONR/chem1110-20/chem_1110-20.htm).

## Data Sheet

	Trial 1	Trial 2	Trial 3
Mass of lighter assembly and contents before collection			
Mass of the lighter assembly and contents after collection			
Mass of the butane gas evolved			
Volume of gas produced at room temperature and atmospheric pressure			
Barometric pressure			
Water temperature			
Vapor pressure of water at water temperature (from table)			
Partial pressure of dry butane (show calculation) $P_{\text{tot}} = P_{\text{butane}} + P_{\text{water vapor}}$			
Room temperature			
Molar mass of the butane (show calculation) $M = mRT/PV$			

True value for the molar mass \_\_\_\_\_

Average molar mass for the three trials \_\_\_\_\_

Experimental error for the average (show calculation) \_\_\_\_\_