

## The Molar Mass of a Condensable Vapor

Reading assignment: Chang, Chemistry, 10<sup>th</sup> edition, pp. 85-87; 185-187.

### Goals

To determine the molar mass of a condensable vapor through use of the ideal gas law.

### Equipment and Materials

Hot plate, ring stand, analytical balance, 1-L beaker, 250-mL round bottom flask, tap water, 5-cm by 5-cm piece of aluminum foil, 5-mL of a volatile organic liquid (methyl ethyl ketone or ethyl acetate).

### Safety

Safety glasses or goggles are required for this experiment. Heating should take place in a fume hood. Waste must be disposed of in an organic chemicals waste container.

### Discussion

The molar mass of a condensable vapor can be determined by adding an excess of the corresponding liquid to a weighed bulb, heating the assembly to volatilize all the liquid and to expel the excess vapor, cooling, and then weighing the bulb with the condensed vapor. We will use one of two organic liquids, either methyl ethyl ketone (MEK) or ethyl acetate (EA) to study.

In this experiment we place a liquid that possesses a low boiling point in a flask that possesses a small opening. The flask is then placed into a beaker of boiling water (this is sometimes called a water bath). The water bath ensures that the temperature of the gas is the same as the boiling point of the water (100°C if the barometric pressure is 1 atm). The small opening in the flask allows the vapor from the liquid to escape at a controlled rate and ensures that the pressure of the vapor in the container is equal to the barometric pressure. The ideal gas law ( $PV = nRT$ ) can then be used to determine the number of moles of a gas that occupies the container:

$n = \frac{PV}{RT}$  where P is the barometric pressure, V is the volume of the gas occupying the container, R is the universal gas constant, and T is the temperature of the gas and is the same temperature as the water.

If we rapidly cool the vapor, then the gas will be liquefied and can be weighed. The mass of this condensed vapor is equal to the mass of vapor that fills the flask at the elevated temperature and barometric pressure.

The molar mass of the liquid/vapor is then  $\mathcal{M} = \frac{\text{mass}}{\text{moles}} = \frac{m}{n}$ . We can compare the results of this procedure to the known molar mass of the compound. A result within 20% of the actual value is not unusual.

### Procedure

1. Place a hot plate squarely at the bottom of a ring stand. Place a one-liter beaker, containing about 700 mL of tap water, on the plate and use setting "HI" to achieve quick heating. When the water boils, reduce the setting to 5 or 6 to maintain gentle boiling until you complete all mass measurements in this experiment. Place a small vessel nearly full of water on one corner of the plate, for later use in maintaining the water level in the big beaker.
2. Obtain a square of aluminum foil (about 5 cm on a side), and fashion a cap for your 250 mL flask by laying the foil on the mouth and folding down the sides. With a pin, poke a tiny hole (as small as possible) in the center of the cap. Determine the mass of the unwashed flask with cap to the nearest 0.1 mg using the analytical balance. From your instructor obtain about 3 mL of an unknown liquid (which contains a few

milligrams of dye to help you see the liquid's disappearance) in the flask. Replace the cap securely. Do not weigh the flask at this point.

3. Clamp the flask on its lip and place the flask as far down in the beaker as the clamp permits. Fix the clamp to the pole so that the flask is tilted enough to concentrate its liquid in one "corner." Also record the barometric pressure.

4. As the water boils, watch the liquid in the flask. One minute after all of the liquid appears to have evaporated, remove the flask and clamp from the pole. Allow cool tap water to run over the flask, being careful that no water gets on or under the foil cap. Use paper towels to dry the flask and to dab the foil in case any water droplets are on it. Weigh the flask, cap, and condensed unknown liquid. Be sure to readjust the water level in the beaker to about 700 mL.

5. Without pouring the liquid out of the flask, add more of the liquid to the flask (perhaps 1 mL) so that you have a similar amount of liquid as you did earlier. Then proceed as you did with the first sample (re-do steps 3 and 4). If your two weights of condensed vapor differ by more than 0.05 g, you must obtain more liquid and repeat the procedure. Your two closest weights, those within 0.05 g of each other, should be used to find the average mass of condensed vapor.

6. Dispose of the liquid (MEK or EA) in the organic waste container provided. Rinse the flask and fill it completely with tap water. The round bottom flask does not have a volume of 250 mL. Measure the volume of the flask all the way to the top with a graduated cylinder to the nearest 1 mL.

**Calculations**

1.  $1 \text{ atm} = 760 \text{ torr} = 760 \text{ mm Hg}$
2.  $1 \text{ mL} = 10^{-3} \text{ L}$ . The ideal gas law uses volume in liters.
3.  $T_K = T_C + 273.15$
4. Use the ideal gas law for this calculation:  $PV = nRT$
5.  $\text{molar mass} = \frac{\text{mass of vapor}}{\text{moles of vapor}}$
6. Using the actual molar mass of your unknown (calculated from its formula), calculate the actual percent error of your molar mass determination.

$$\% \text{ error} = \left| \frac{\text{actual value} - \text{measured value}}{\text{actual value}} \right| \times 100$$

After you have reported your calculated molar mass to the instructor, ask for the name or the formula of your unknown.

1. Calculate the molar mass of the condensable vapor. From the known molar mass calculate a percentage error. Show all calculations clearly on a separate piece of paper.
2. On a separate piece of paper describe (in one paragraph) the evidence that the liquid exists in gaseous form at room temperature.
3. Describe two or three ways in which the measurements you made could lead to errors in the results. Explain how each error would affect the answer of molar mass. In other words, would the molar mass go up or down and why?

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**Experiment 9 Molar Mass of a Condensable Vapor**  
**Data and Calculations Sheet**

Name \_\_\_\_\_

Partner's Name (s) \_\_\_\_\_

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**Data Table**

Run Number	1	2	3
Mass of flask and cap (same each run)			g
Mass of flask, cap, and volatile liquid	g	g	g
Mass of condensed vapor	g	g	g
Average mass of condensed vapor			g
Temperature of boiling water			°C
Barometric pressure			torr
Volume of flask			mL

**Calculations** (show all calculations in your report)

Barometric pressure in atm (1)	atm
Volume in liters (2)	
Temperature in kelvin (3)	
Gas constant R	0.08206 L·atm/mol·K
Average mass of condensed vapor (in grams)	
Moles of vapor (4)	mol
Molar mass of vapor (5)	g/mol
Chemical Formula (from instructor)	
Molar Mass from Chemical Formula	
Percent Error in Molar Mass (6)	