Estimation of the relative molecular mass, M_r, of a volatile liquid

Theory

According to the equation of state of an ideal gas, the relationship between the pressure, P, volume occupied, V, temperature, T, and number of moles, n, of an ideal gas is:

$$PV = nRT$$

where R is a constant called the gas constant. Measuring pressure in Pa, volume in m^3, and temperature in Kelvins, the value of the gas constant is 8.314 J K^{-1} mol^{-1}.
Thus if the pressure, volume and temperature of the vapour of a sample of a volatile liquid are measured, the number of moles of it present can be calculated.
If the mass of the sample is known then the relative molecular mass, M_r, can then be calculated from the relationship between mass, m, relative molecular mass, M_r, and the number of moles, n:

$$n = m/M_r$$

In this experiment a small amount of a volatile liquid is allowed to vaporise by placing it in a container surrounded by hot water. The temperature of the water is measured, and therefore the temperature of the vapour is thus known. The vapour is under atmospheric pressure and the volume it occupies at this pressure is recorded. The mass of the vapour is also measured. The data is used to calculate the number of moles of vapour and hence the value of the relative molecular mass, M_r, using the two formulas above.

Method

Chemicals and Apparatus

- ✓ Sample of propanone 
- ✓ (or unknown sample provided)
- ✓ Water

- ✓ 250 cm^3 conical flask
- ✓ 600 cm^3 beaker (into which the 250 cm^3 conical flask can be easily fitted)
- ✓ Aluminium foil
- ✓ Clamp
- ✓ Dropping pipette
- ✓ Pin
- ✓ Rubber band
- ✓ Bunsen burner (or hotplate)
- ✓ Tripod and Gauze
- ✓ Thermometer
- ✓ Barometer
- ✓ Electronic balance
- ✓ 100 cm^3 graduated cylinders

Procedure

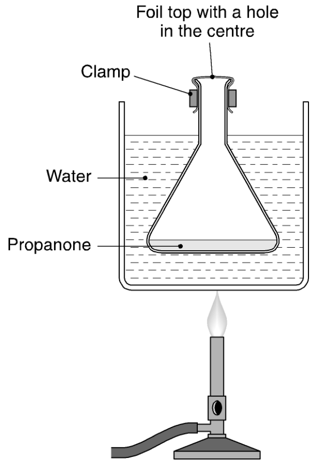
<u>NB</u>: Wear your safety glasses.

Two-thirds fill the beaker with water, place on tripod stand and heat to almost boiling with the Bunsen burner. Control the flame so that the temperature remains at 95 ^{0}C.

Cut a circle of aluminium foil large enough to cover the mouth of the conical flask and fold down a little around the sides of the flask.

Find the total mass of the clean dry conical flask, the aluminium foil and the rubber band.

Using a dropping pipette, add 3 – 4 cm^3 of the volatile liquid to the flask. You need not worry about the exact quantity at this stage because some of this liquid will be lost as a vapour during the experiment.



Cover the mouth of the flask with the aluminium foil. Hold in place tightly with the rubber band so that no vapour can escape between the foil and the glass. With the pin, prick one small hole in the centre of the aluminium foil cap. Attach the clamp to the neck of the flask.

<u>Carefully</u> immerse the conical flask into the boiling water (Fig. 1). Holding the clamp, move the flask up and down periodically to check the liquid level in the flask.

All of the volatile liquid vaporises and some of it will escape out through the hole in the cap until the pressure inside the flask is equal to atmospheric pressure. When the flask appears to be empty (i.e. all the liquid appears to have evaporated), this stage has been reached. Immediately and with care, remove the flask from the beaker. Record the exact temperature of the hot water.

Record the value of atmospheric pressure using the barometer.

Allow the flask to cool. Then, thoroughly dry the outside of the flask, including the foil. You may now notice that there is a small quantity of liquid in the flask. This is the volatile liquid, which has cooled and condensed. Find the mass of the flask, cap, rubber band and contents. By subtraction of the first mass recorded, the mass of the vapour, which occupied the flask at the temperature of the boiling water, is now known.

Remove the cap and rubber band. Find the volume of the flask by completely filling it with water, and then transferring all of the liquid from it to graduated cylinders. Record the volume of the liquid transferred.

Calculate the value of the relative molecular mass, M_r, of the volatile liquid.

Table of Results

(Note that 760 mmHg = 101325 Pa)

Mass of flask, cap and rubber band	g
Mass of condensed vapour, flask, cap and rubber band	g
Mass of condensed vapour	g
Atmospheric pressure	mmHg
Atmospheric pressure	Pa
Temperature of boiling water	^{0}C
Temperature of boiling water	K
Volume of flask	cm^3
Volume of flask	m^3

Gas Constant, R = 8.314 J K^{-1} mol^{-1}

Specimen Results (Method 1)

Mass of flask, cap and rubber band	115.15 g
Mass of condensed vapour, flask, cap and rubber band	115.67 g
Mass of condensed vapour	0.52 g
Atmospheric pressure	756 mmHg
760 mmHg = 101325 Pa	
Atmospheric pressure	100792 Pa
Temperature of boiling water	100 ^{0}C
Temperature of boiling water	373 K
Volume of flask	284 cm^3
Volume of flask	2.84 x 10^{-4} m^3

Gas Constant, R = 8.314 J K^{-1} mol^{-1}

Sample calculations

PV = nRT

⇨ n = PV / RT

⇨ n = 100792 x 2.84 x 10^{-4} / 8.314 x 373

⇨ n = 9.23 x 10^{-3} moles

n = m / M_r

⇨ M_r = m / n

⇨ M_r = 0.52 / 9.23 x 10^{-3}

⇨ Mr = 56.33

Student Questions

Why is it necessary that the liquid used in this experiment is volatile?

The liquid must be capable of forming a vapour at the temperature of the experiment so that the ideal gas equation can be applied. The equation applies with greatest accuracy to those vapours that are most like ideal gases. Gases are least ideal when on the point of condensing, so the greater the difference between the boiling point of the liquid and the temperature of the reaction the more accurate the results will be.

Name another technique for measuring relative molecular mass.

Mass spectrometry

Since the vapour is not an ideal gas, which quantity measured in the experiment is most likely to introduce inaccuracy in the result and why?

The temperature and pressure of the surroundings were measured. These measurements are independent of whether the gas is ideal or not. The gas constant is given. The measurement of volume is the measurement that would be expected to be least accurate when used in the ideal gas equation, when compared to the volume that would be obtained if an ideal gas could be used. A real gas occupies a smaller volume than an ideal gas because of intermolecular attractions in real gases, e.g. hydrogen bonding, van der Waals' forces etc.

If a small drop of water were present in the flask used in Method 1 or the gas syringe used in Method 2, how would this affect the results?

The small drop of water would vaporise during the experiment and occupy quite a large volume. The reading for the volume of the volatile liquid's vapour would be far too large and the result calculated for M_r very inaccurate as a result. It would be too small.

From your results calculate the density of the vapour of the volatile liquid at the temperature of boiling water.

Density = mass / volume. For example, using the data in the sample results:

Method 2: density = 0.12/72 = 0.0017 g/cm^3