Molar mass of a volatile gas - lab report example

Science, Chemistry

Molar Mass of a Volatile Gas

Chemistry Lab Report 08 April Molar Mass of a Volatile Gas Objective The main objective of this experiment was to determine the molar mass of a volatile gas using the ideal gas equation.

Materials and Apparatus

Volatile liquid

A pair of goggles

A beaker

Aluminum foil

Bunsen burner

Thermometer

Weighing balance

Procedure

The weight of an empty flask together with its aluminum foil cover measured carefully and recorded. A tiny pinhole was made on the outer surface of the aluminum foil. Five milliliters of a volatile liquid were measured and added to the flask. The weight of the flask, aluminum foil and the unknown volatile liquid was then measured and recorded. Thereafter, the volatile liquid was heated in a water bath. The flask was tilted slightly to make it easy to see when all the volatile liquid had vaporized. Subsequently, the temperature of the water bath was measured when the liquid in the flask vaporized. The atmospheric pressure, which was assumed to be equal to the pressure of the volatile gas, was also measured and recorded.

When all the liquid had evaporated, cold water was run over the flask to facilitate the cooling of the vapor. The mass of the flask, aluminum foil as well as the condensed vapor was then determined. It was assumed that the mass of the condensed fluid was equivalent to the vapor that filled the flask. The molar mass of the gas was then computed using the ideal gas law.

Results

Table 1: Raw data

Parameter	Quantity
Mass of flask and foil	40 g
Mass of flask, foil and unknown liquid	45. 93 g
Mass of unknown liquid	5. 93 g
Mass of unknown gas	0. 07g
Temperature of the gas	97 oC (370K)
Volume of the gas	0. 1255 L
Pressure of the gas	756 mmHg (0. 994 atm)

Calculations

Mass of the unknown sample= 45. 93-40. 00

= 5. 93 g

The number of moles in the unknown sample was calculated from the formula n= PV/RT where n was the number of moles, P was the pressure of the gas, V was the volume of the gas, R was the gas constant (8. 21x10-2 L atm mol-1 K-1), and T was the temperature in Kelvin (Slowinski, Wosley and Rossi 55).

Number of moles= (0. 994 atm×0. 1255L)/ 8. 21x10-2 L atm mol-1 K-1×370 K

= 4. 1066×10-3 moles

The molar mass of the unknown sample= Mass of the unknown/ number of moles

= 0. 07 g/4. 1066×10-3 moles

= 17. 045 moles

Discussion and Conclusion

The experimental molar mass was lower than 32, which was the actual molar mass of the unknown gas. The low experimental value could be due to experimental errors. For example, it was possible that there was condensed vapor in the foil cover, which interfered with the accuracy of the measurements. One other possibility that led to the disparities in the experimental molar mass of the unknown gas and the actual value was deviations from the ideal gas law. It was possible that the gas did not behave as described by the ideal gas law hence leading to the disparities in the two values. It was also possible that excess vapor escaped from the flask leading to an underestimation of the mass of the condensed liquid and the subsequent molar mass.

Overall, the experiment gave an estimation of the molar mass of the unknown gas using the ideal gas law. Therefore, it was concluded that the ideal gas law was a useful equation in describing the behavior of gases.

Work Cited

Slowinski, Emil, Wayne Wosley and Robert Rossi. Chemical Principles in the Laboratory, 10th ed. 2011. Belmont, CA: Brookes/Cole. Print.