Ideal Gas Law and Gas Constant Lab Report

Introduction: Of the three standard states of matter, solids, liquids and gases, gases are the most simple and difficult to generalize. In order to better understand the patterns of behavior of gases, a theory has been developed known as the Ideal Gas Law, PV=nRT, which allows one to fairly accurately predict what a gas will do under certain circumstances, although the law itself is derived by derivation from an "ideal" setting of high temperature and low pressure. The Ideal Gas Law is a derivative of three simpler gas laws that came before it, Charles' Law, Boyle's Law, and Avogadro's Law, which relate volume and temperature, volume and pressure, and volume and moles, respectively, with Charles' and Avogadro's Laws being direct relationships and Boyle's Law being indirect.

Objective: To verify the applicability of the ideal gas law and to determine the value of the gas constant.

Procedure: Utilizing the paragraph provided, which displays a manometer and an accompanying 24 - flask of adjustable temperature to which nitrogen gas may be added, varying amounts of gas were input at a constant temperature to observe the subsequent changes in pressure. The relationship of pressure to mole concentration in Mole/L was automatically graphed to the side, the slope of which could be manipulated to give, in an ideal situation, the ideal gas constant.

Calculations/Results: Attached

Discussion: In the experimental setup conducted at 25 degrees Celsius, or 298 degrees Kelvin, the slope of the graph, the gas constant, was found to be approximately .08212 atm * L / mol * K, which is nearly identical to the standard ideal gas constant of .08206 atm * L / mol * K, if only a little different. The second setup, conducted at 100 degrees Celsius, 373 degrees Kelvin, however, showed significant deviation with a value of around .06000 atm * L / mol * K. The most likely explanation would be found through the studying the concepts of the Ideal Gas Law. In an ideal gas, all collisions between gas particles are purely elastic, gas particles have no volume, and there are no intermolecular forces of attraction whatsoever. Real gases behave most closely to ideal gases at high temperatures and low pressures because at high temperatures molecules are so mobile that the effect of their intermolecular forces is dampened and at low pressures molecules are not forced into contact with each other or the walls of the container mutually either. A situation of high temperature and low pressure is paradoxical, though, according to Gay Lussac's Law, as temperature increases so does the pressure. Considering the case of the second setup, even though the high temperature should have caused that the gas would behave more ideally than the first setup, the fact that the volume was so constrained caused that the drastically high pressure had more influence to bring intermolecular forces into play.