

Lab report on ideal gas law



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Ideal Gas Law Lab 1. Procedure: First, we used a balance to weigh the canister of gas, and recorded that mass as the original weight. Then, we filled a large bucket with water and recorded the temperature. We then filled a small test tube with water at the same temperature and poured that water into a graduated cylinder to measure the original volume of water in the tube. We then poured the water back into the test tube and placed the tube into the bucket with the opening upwards, turning the open end downwards after the tube was fully submerged beneath the surface. We then placed the canister directly below the opening of the test tube, and released the gas so that the bubbles rose into the test tube. Next, we used a cork whose head was larger than the opening of the test tube to block off the opening (without changing the pressure inside of the tube), so that we could transport the remaining water to a graduated cylinder. When doing this, it was very important that the water level inside of the tube was equal to that of the surrounding water in the bucket, because that ensured that since the water pressure in the tube was the same as that of the surrounding water, the pressure of the gas would be the same as that of the surrounding air. Thus, we recorded the gas pressure to be the same as the pressure in the room, which was calculated to be 763.0 mmHg. We poured the remaining water from the test tube into a graduated cylinder to calculate the difference between the original water volume and the volume remaining, because this difference was equal to the volume of the gas released. Next, we blow-dried and shook the canister to get rid of the extra mass that would have been added by any water that had clung onto it while it was in the bucket, and weighed the canister again. We subtracted this mass from the original mass of the canister to find out the mass of gas released. Now that we had collected our

data, we used Dalton's law of partial pressures to calculate the pressure of the gas from the canister alone, so we subtracted the pressure of the water vapor at the recorded water temperature (15.5°C) from the total pressure of the gas (763.0 mmHg). With the pressure of the unknown gas alone, we were ready to plug our values into the $PV = nRT$ equation. Solving for n , we were able to determine how many moles of the gas we had used, so we only had to divide the mass of gas used by the moles to calculate the molar mass of the gas (in grams per mole). With the molar mass of the gas, we were able to identify the gas.

2. Results: Below are the values we recorded for each step in the experiment, as well as the calculations we made to come up with the molar mass of the unknown gas. Thus, the unknown gas was butane, C_4H_{10} , whose molar mass is about 58.1 g/mol.

3. Discussion: The theory behind our experiment was to find each necessary factor in the simplest, most accurate way. The most difficult part was figuring out how to catch all of the gas that we released, and to measure the volume of that gas precisely. The method we used was very effective because we were able to see the movement of the gas, and we were therefore able to control its entry into the test tube. Our results were pretty accurate, but there was some room for error. Below are the calculations for the percent error of our molar mass measurement. Some of the sources of error came from possible procedural problems, while others came due to the theoretical limitations of the experiment. One of the procedural difficulties we may have encountered was that there might have still been some water left on the canister when we weighed it the second time. Because the second mass of the canister would have the added mass of the water, the difference between the original mass and the mass after the gas had been released would have been smaller.

Thus, we would have recorded a lower mass of gas released, so the molar mass calculation would have been lower because the numerator of the equation would have been smaller. Perhaps we could have blow-dried the canister for a longer period of time until we were absolutely sure that no water remained. Another procedural problem was the balance we used. Because it only measures to the hundredths place, we recorded the mass of the gas released to only one significant figure. Because of this, we were limited to only one significant figure in our calculation of the molar mass, so although we would have had an answer of 52.6 g/mol, we had to round down to 50 g/mol, which led to a greater percent error. The theoretical difficulties arose because we were applying the ideal gas law to a real gas. So, for example, while ideal gas particles have elasticity and don't lose speed or energy when they collide, the particles of a real gas do lose energy and speed, and this decrease leads to lower pressure. Also, ideal gases do not have any intermolecular forces, and the volume of the particles of an ideal gas can be ignored, but with a real gas, the particle size makes a difference. Because butane molecules are so large and exist in a state very close to liquid form, we know that there are very strong dispersion forces holding the molecules together. Larger molecules have more momentary dipoles, and thus a stronger attraction, so because the molecules are held so close together, the volume we recorded was smaller than it would have been if butane were an ideal gas. However, despite these errors, our calculated molar mass was not very far from the actual molar mass of butane gas, so we did a relatively good job of controlling these variables that could have greatly affected our results.