University of Toronto, Engineering Science

CHE260: Thermodynamics and Heat Transfer **Lab Report #1**

Ideal Gas Law

Investigating the Path Independency and Equilibrium Behavior of Temperature and Pressure in Ideal Gases

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Abstract

The purpose of the lab is to examine the behaviour of state properties by manipulating the temperature and pressure of two gas chambers. Our experiment confirmed that both these properties will stabilize at roughly the same value regardless of their path to equilibrium, given identical starting conditions. These results verify that air behaves according to the ideal gas law.

Introduction

The ideal gas law, $PV = nR_uT$, is a fundamental concept used in thermodynamics and other forms of science which closely defines the relationship between variables in real gases [1]. Properties such as P, V, and T are state functions, meaning they depend only on the current state of the system [2]. These traits allow thermodynamicists to perform calculations using only the initial and final state of a gas, and their relationship as defined in the ideal gas law equation. Their independence from the path of the function means experiments are only required to collect specific data points, as they will be consistent regardless of how the system changes.

The goal of the experiment is to show the path independency of the state properties by observing the values at which temperature and pressure stabilize when allowing the variables to reach equilibrium through two different methods. In thermodynamics, equilibrium is defined as when properties such as temperature, pressure, or density remain constant, and the internal energy is stable [3].

Experimental Method

The experiment utilized pressure gauges to observe how two pressurized containers of air come to equilibrium in two different ways. The temperature and pressure of each tank were recorded using LabVIEW technology, and the ambient pressure of the laboratory was 1007.0±0.5 mbar, or

14.605±0.007 psi, which was added to gauge values to present absolute pressure. The uncertainty is based on the increment of the manometer. Thus, with unit conversion, pressure in psi is presented to three decimal places [4]. For both experimental methods, the containers were set up by opening the appropriate valves & solenoids, and allowing the air to flow at a rate of 50g/min. The desired pressure in the left tank was 40 psi, and -6 psi for the right tank.

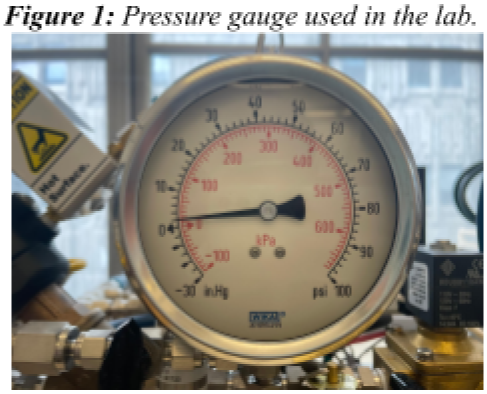
To perform the first method, the center solenoid was opened and data was recorded until the temperature of both tanks were roughly the same. The tank was emptied and reset to the previously indicated pressures. Secondly, the experiment was performed by opening the small ball valve and micrometer valve. Data was recorded until the pressure of both tanks were roughly the same, elapsing around 20 minutes.

Prior to the second portion of the lab, the tanks were emptied with their pressures returning to 0 psi. Lastly, the left tank was pressurized to 40 psi and monitored until temperature and pressure stabilized. LabVIEW was then paused and the tanks were emptied to conclude the collection of experimental data for this lab.

Results

After conducting the first portion of the lab, we compared the final values of temperature and pressure in the two tanks to analyze the differences between methods. In the first method, where the solenoid between tanks was opened at around 270s, Figure 2 depicts how the temperatures [°C] of two tanks evolve over time [s], demonstrating a gradual convergence of temperatures to 26.00±0.05°C. The uncertainty comes from the one decimal place data in LabVIEW's measurement scale. Furthermore, Figure 3 illustrates how at around 270s, the pressure of both tanks abruptly

Figure 1: *Pressure gauge used in the lab.*



converge and stabilize at 23±1psi. The uncertainty is 1 psi, as shown on the gauge in Figure 1, seen as 0 psi in LabVIEW.

Figure 2: *Temperature [°C] of both tanks during the first method*

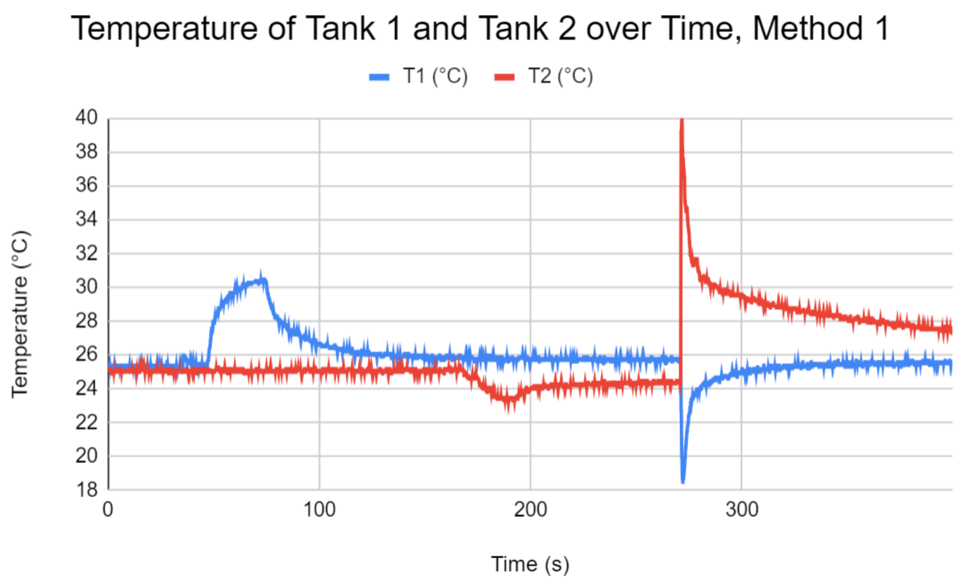
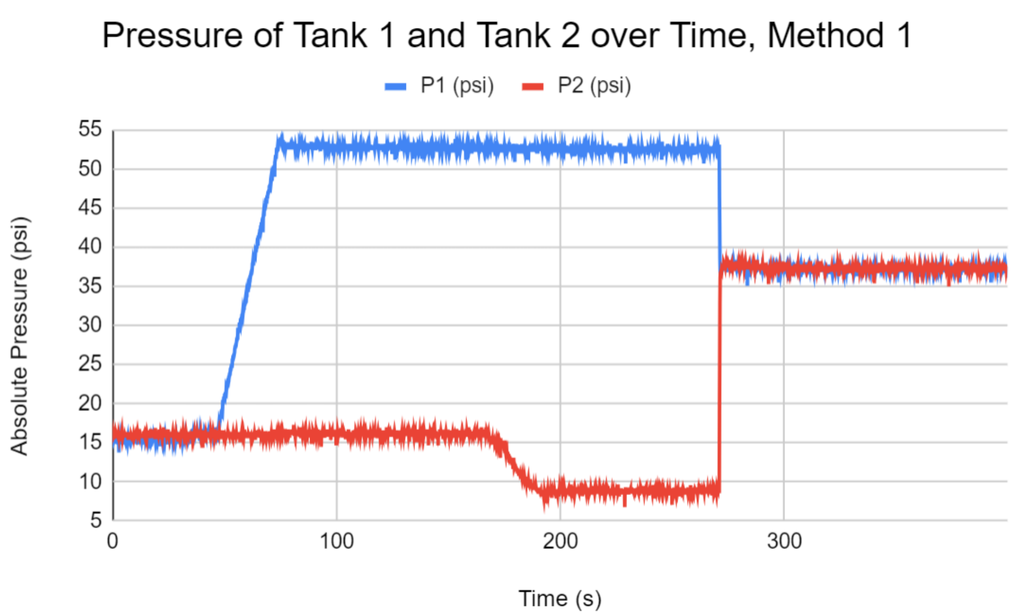


Figure 3: *Pressure [psi] of both tanks during the first method*



In the second method, when opening the micrometer permits gas to flow from the left to the right tank, the pressure of both tanks converges to 23±1psi (see Fig. 4). Both lines exhibit a comparable pattern to method one, indicating that they gradually approach the same pressure as time progresses. At the same moment when the micrometer opens, the temperature of both tanks are observed converging toward 26.00±0.05°C (see Fig. 5).

Figure 4: *Pressure [psi] of both tanks during the second method*

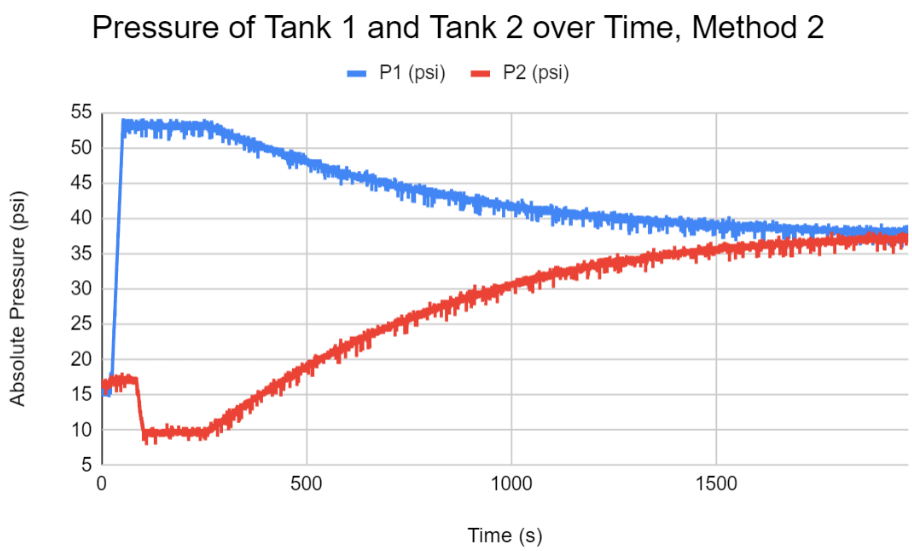
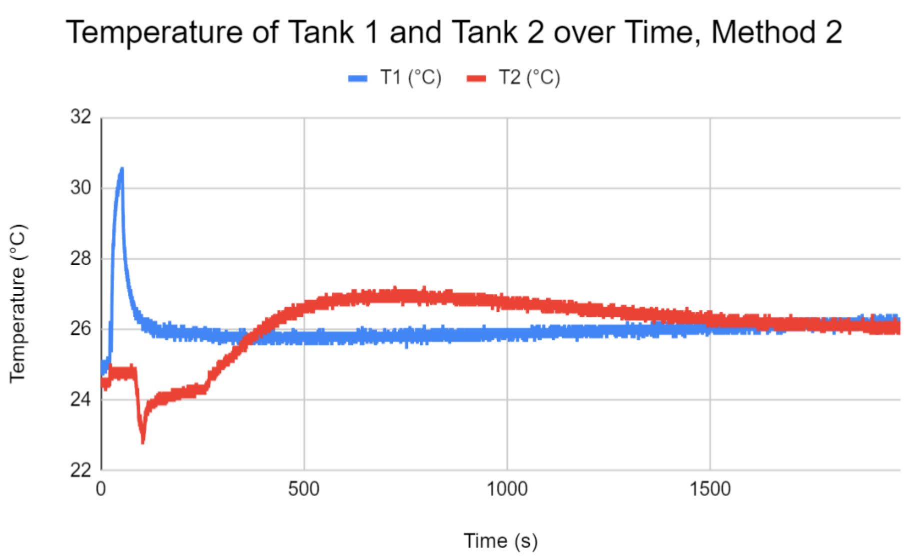
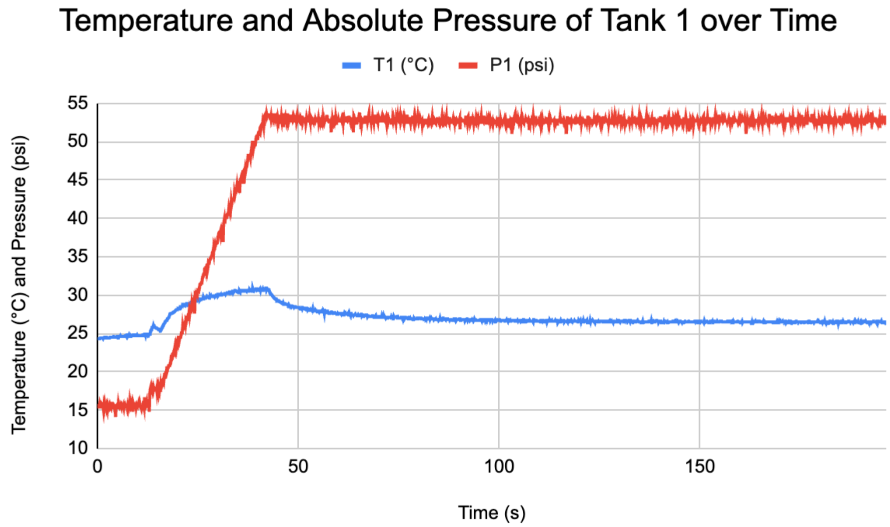


Figure 5: *Temperature [°C] of both tanks during the second method*



Furthermore, Figure 6 displays the stabilization of temperature and pressure over time in the second experiment. As depicted, both parameters reached a stable state after approximately 100s, and match the other experimental initial values of temperature (26.00±0.05°C) and pressure (39.5±1psi) in Tank 1, shown as the blue line in Figures 2-5.

Figure 6: *Temperature [°C] and Pressure [psi] of Tank 1 stabilizing over time*



Discussion

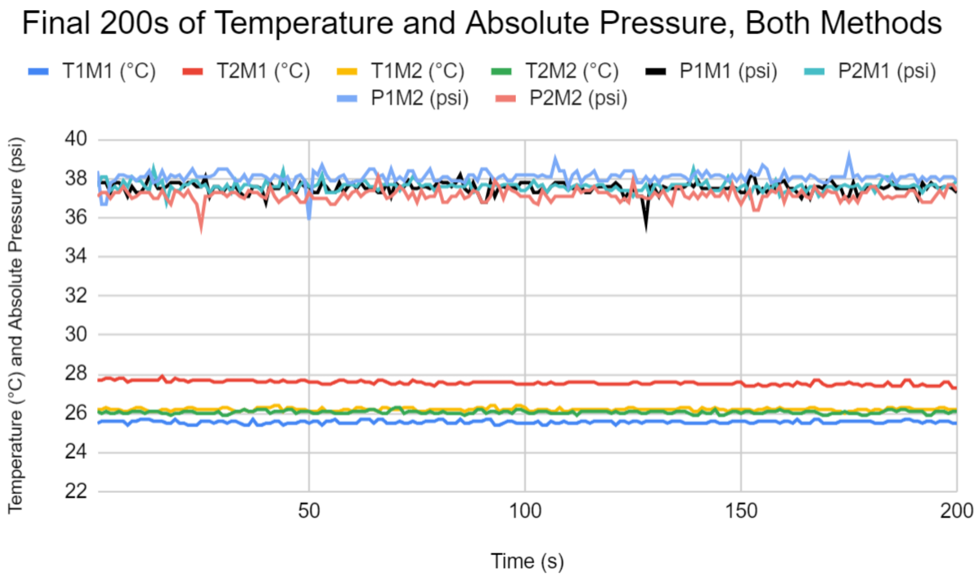
Although the second path took significantly longer to stabilize, our data points show that the temperature and pressure between both methods reach equilibrium at the same value. Knowing our initial conditions were identical, we conclude that these properties naturally returned to the same values regardless of the procedure. Therefore, as illustrated in Figures 2-5, the results of our experiment confirm the hypothesis that temperature and pressure are properties of gases that are independent of the system's path, and solely dependent on the current state of the system. The experiment also confirms the theory presented in known gas laws, like Boyle's Law, which is one of several equations relating state properties such as temperature and pressure [5]. The results proved that initial and final conditions of state properties are sufficient to calculate changes in the system since they obtain the same values no matter the path.

Using the Ideal Gas Law and that mass is the product of the number of moles and atomic weight, the principle of mass conservation implies that the quantity of moles remains constant. Firstly, the Ideal Gas Law equation was rearranged: $PV = NR_uT \rightarrow N = PV/R_uT$. From conservation of mass, it is clear that: $N_{Li} + N_{Ri} = N_{Lf} + N_{Rf}$, with subscripts indicating initial and final states of the left and right tanks. By substituting the equations together, the following equation is obtained: $\frac{P_{Li}V_L}{R_u T_{Li}} - \frac{P_{Lf}V_L}{R_u T_{Lf}} = \frac{P_{Rf}V_R}{R_u T_{Rf}} - \frac{P_{Ri}V_R}{R_u T_{Ri}}$. P_f represents the final pressure of both tanks as they reach the same equilibrium values, similarly for T_f. Therefore, the ratio is $\frac{V_L}{V_R} = \frac{\frac{P_f}{T_f} - \frac{P_{Ri}}{T_{Ri}}}{\frac{P_{Li}}{T_{Li}} - \frac{P_f}{T_f}}$, calculating to a volume ratio of 1.69 for method 1 and 1.67 for method 2.

Before the solenoid opened in method 1, it is acknowledged that V, N, R_u in the ideal gas law equation don't change. Thus, when the left tank was pressurized, it did work on the adjacent tank and the temperature increased. When the two tanks were connected by the centre solenoid, both temperatures and pressures experienced a steep change. This explains the expansion of the heat from the left tank to the right tank. Then a gradual convergence between the pressure curves is observed (see Fig. 3), meaning the pressure does not change anymore. Figure 4 shows the temperature in two tanks going to their equilibrium state, and accounts for heat transfer related to the surroundings. During the second method, instead of a sudden change, both of the temperature and pressure curves are smoothly converging towards their equilibrium state after the needle was opened (see Fig. 4&5). Therefore, all the heat transfer comes from the surroundings.

Pressure and temperature are state properties, meaning they depend on the current state of the system, regardless how this state was achieved [2]. The first part of the experiment successfully demonstrates that regardless of which method is used, with identical initial conditions, the temperature and pressure will stabilize at the same values. Figure 7 overlays all recorded paths of temperature and pressure from the experiment, demonstrating how the values were within the range of uncertainty despite following different paths.

Figure 7: Final values of temperature [°C] and pressure [psi] from both methods



Using the data points we collected from the second experiment, we can identify that it took 28.9s for the tank to reach 40±1psi (gauge) using a flow rate of 50g/min. Therefore, there initially is 24.08g of air in the left tank. At the time when this mass has been fully added to the tank, the temperature of the tank is 29.5°C. Assuming the air by the ideal gas law, $V = \frac{mR_u T}{MP}$, the volume of the left tank is therefore $5.56 \times 10^{-3} m^3$.

Compressibility can be defined as the comparison between the real volume of a gas under specific temperature and pressure conditions and the volume that it would occupy if it behaved ideally under those conditions [6]. The compressibility factor for natural gas results in approximately a 0.5% volume correction per 100 psi in a flow rate measuring instrument at standard conditions [7]. In the lab, the maximum pressure is in the range of 50 psi. Therefore, around 0.25% percent error in the compressibility factor has only a minimal impact on volume measurements and it is negligible.

Although this experimental set-up and procedure produced reasonable conclusions as predicted by the law regarding path independency of ideal gases, there are still sources of error that limit the accuracy of the results obtained. Firstly, according to the ideal gas law, air is assumed to behave like an ideal gas under the assumption that the gas particles move randomly, have negligible volume, are equally sized and do not experience intermolecular forces (attractive or repulsive) between other gas particles and have perfectly elastic collisions with no energy loss [1]. However these assumptions don't hold at low temperatures and high pressures. Examining the results from Figure 4, it is observed that the behavior of this air slightly deviates from an ideal gas because the air doesn't reach thermal equilibrium at the same pressure. Secondly, the molar mass of air was assumed to hold constant which may have varied due to small changes in the composition and proportions of gases of air.

Furthermore, the instruments used in the experiment were not accurately calibrated to one another. For instance, when the tank was completely emptied, the pressure gauge had a reading of 1±psi, while the reading on LabVIEW was 0 psi. As such, a source of systematic error was the discrepancy between the real and measured readings. This was likely the largest source of error, but was unavoidable based on the manufacturing of the device, and being unable to calibrate the system ourselves. The magnitude of systematic error is estimated to be 10^{-1} from the difference between initial pressure gauge and LabVIEW readings. Another critical source of error may have stemmed from the experimental design, where data was produced from a single trial of each method. If multiple trials were conducted, then anomalies could have been identified and eliminated by taking the average reading to minimize random errors. Additionally this set-up consists of a manual lever to release the gas from the tanks. However this lever mechanism was loose and not tightly sealed, which could have potentially allowed a negligible amount of gas to leak affecting data accuracy.

Despite the limitations addressed above, it is essential to consider the strengths of this practical, which corroborate the fact that air approximates to an ideal gas and path independence holds. A major strength is that the instrumental uncertainties are significantly low because tracking software like labVIEW have high precision for monitoring temperature and pressure at small intervals. Correspondingly, utilizing two different instruments to record pressure allowed a validation of data, by verifying the readings on the pressure gauge and tracking pressure on labVEIW to account for calibration errors and zero-offsets and consistency of readings. Ultimately this experimental set-up was sufficient to collect relevant data, perform calculations and examine the state properties of air as an ideal gas.

Conclusion

The experimental results confirm the predicted behavior of temperature and pressure when allowed to reach equilibrium in different manners by the Ideal Gas Law, which was to stabilize at the same values despite following different paths. The qualities of these state properties permit for ideal gas calculations solely by obtaining their initial and final values. Acknowledging the limitations aforementioned, the experimental set-up and precision of calculations can be potentially improved in future by repeating more trials to reduce the impact of uncertainties and calibrating the pressure gauge correctly to reduce systematics errors and conducting the lab at even higher temperatures and lower pressure than ambient pressure and temperature so that air better approximates to an ideal gas. Nevertheless, the results corroborate the predictions and verify the ideal gas behavior by observing path independency of air.

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