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Investigation: Calculating the R-Value of an Ideal Gas Using Boyle's Law

Research question Under constant temperature and amount, how does the pressure of a physical gas depend on its volume, and how does it differ from an ideal gas?

Boyle's law states that the pressure of an ideal gas is inversely proportional to its volume at a constant temperature. In other words, $PV = \text{constant}$. From the ideal gas equation, we see that this constant is equal to nRT, the product of amount of gas, temperature and the ideal gas constant R. Therefore, by knowing the amount of gas in moles n, the pressure P, the volume V and the temperature of the gas T, it is possible to calculate the constant R experimentally.

The purpose of this investigation is to determine this constant R through a simulation (*Highered MHEducation Gas Laws Simulation*). This will be done by recording the change in pressure as a result of a change in volume, while keeping all other variables constant. Therefore, the independent variable of this investigation is volume, which will be measured in dm^3 and will be incremented in intervals of $0.1\,\mathrm{dm}^3$. The dependent variable is the pressure, which will be measured in atm. The controlled variables of this experiment are temperature, which will be kept at room temperature ($298\,\mathrm{K}$), the type of gas, which will be helium gas (due to it being available in the simulation), and the amount of gas, which will be kept constant at $0.5\,\mathrm{mol}$.

I chose to conduct this investigation on Boyle's law because I find it very elegant. It makes perfect conceptual sense to me that the pressure and the volume of the gas are inversely proportional because there are so many every day examples of it. For instance, trying to compress a balloon (reducing its volume) will result in an increase in pressure, which is why one finds resistance when attempting it. There are other instances of this, such as bicycle and car tyres, pumps and even pressurised gas containers, which become less pressurized as the amount of air inside them decreases. With this investigation, I wanted to discover just how much do real gasses behave like ideal gasses and how close the ideal gas constant is to its experimental counterpart.

The reason I have chosen to conduct a simulation instead of attempting a physical experiment for this investigation is that the very nature of simulations allows for virtually no limitations when it comes to things like shortage of equipment or measuring instruments. Hence, simulations tend to have less potential for errors. In addition, simulations make the experiment much more flexible and allow for experiments which would otherwise be very difficult or impossible in the real world. In the case of this investigation, accurately measuring the volume and pressure of a gas container, as well as ensuring that there are no gas leaks would be very tricky to say the least.

On the other hand, simulations do have the inherent limitation that they are just that—simulations. This means that conclusions drawn from results do not necessarily reflect results that one would get in the real world and are merely theoretical. They are a valuable tool nonetheless.

Preparation

The apparatus used within the simulation is listed below:

1. Gas container with the ability to expand in volume.
2. 0.5 mol of helium gas.
3. Barometer
4. Thermometer

The barometer and the thermometer were attached to the gas container so that the temperature and pressure of the gas could be measured.

The first step of the experiment was to fill the gas container with 0.5 mol of helium gas and set its volume to $1.0\,\mathrm{dm}^3$. Then, the pressure displayed on the barometer, along with the volume was recorded in a table. The volume of the container was then increased by by $0.1\,\mathrm{dm}^3$. The new pressure and volume were recorded on the table. This process continued until a volume of $2.0\,\mathrm{dm}^3$ was reached. Throughout the process, the temperature of the gas remained constant close to room temperature ($297.30\,\mathrm{K} \pm 0.05$).

Due to the fact that the experiment was conducted through a simulation, there was no need to consider ethics or safety precautions.

Results

Overall, qualitatively speaking, the behaviour of the gas in the simulation was in accordance to Boyle's law, i.e. increasing the volume of the gas resulted in a decrease of the pressure. The table below displays the raw data for pressure and volume that was collected during the experiment, as well as uncertainty values for their corresponding instruments of measurement.

Table 1: Raw data from the experiment.

Volume (dm^3) ± 0.05	Pressure (atm) ± 0.005
1.00	11.890
1.10	10.810
1.20	9.910
1.30	9.150
1.40	8.490
1.50	7.930
1.60	7.430
1.70	7.000
1.80	6.610
1.90	6.260
2.00	5.950

The table shows clearly the gradual decrease in pressure as a result of the increase in volume. Using this data, we can now calculate the R value of the ideal gas law. The ideal gas law states the following:

$$PV = nRT$$

We can substitute the experimental values into this equation. First of all, we have to convert into base SI units, as the constant R itself is in base units. For the first data point as an example,

$$P_1 = 11.890\,\text{atm} \pm 0.005 = 1\,204\,457.0\,\text{Pa} \pm 506.5$$

$$V_1 = 1.00\,\text{dm}^3 \pm 0.05 = 1.00 \times 10^{-3}\,\text{m}^3 \pm 0.05 \times 10^{-3}$$

Substituting into the equation with values for $T = 297.3\,\text{K} \pm 0.05$ and $n = 0.50\,\text{mol} \pm 0.05$, we get a value for R:

$$R = 8.103\,\text{JK}^{-1}\text{mol}^{-1} \pm 0.151$$

In comparison, the expected value for an ideal gas is $8.31\,\text{JK}^{-1}\text{mol}^{-1}$. The percentage error $(\frac{|\text{expected}-\text{experimental}|}{\text{expected}})$ works out to be 2.49%. This indicates a moderate accuracy to the expected value.

Indeed, all of the data points from the experiment yield a value of R that is well within the uncertainty of ± 0.151. Below is a plot of the R values versus the volume readings.

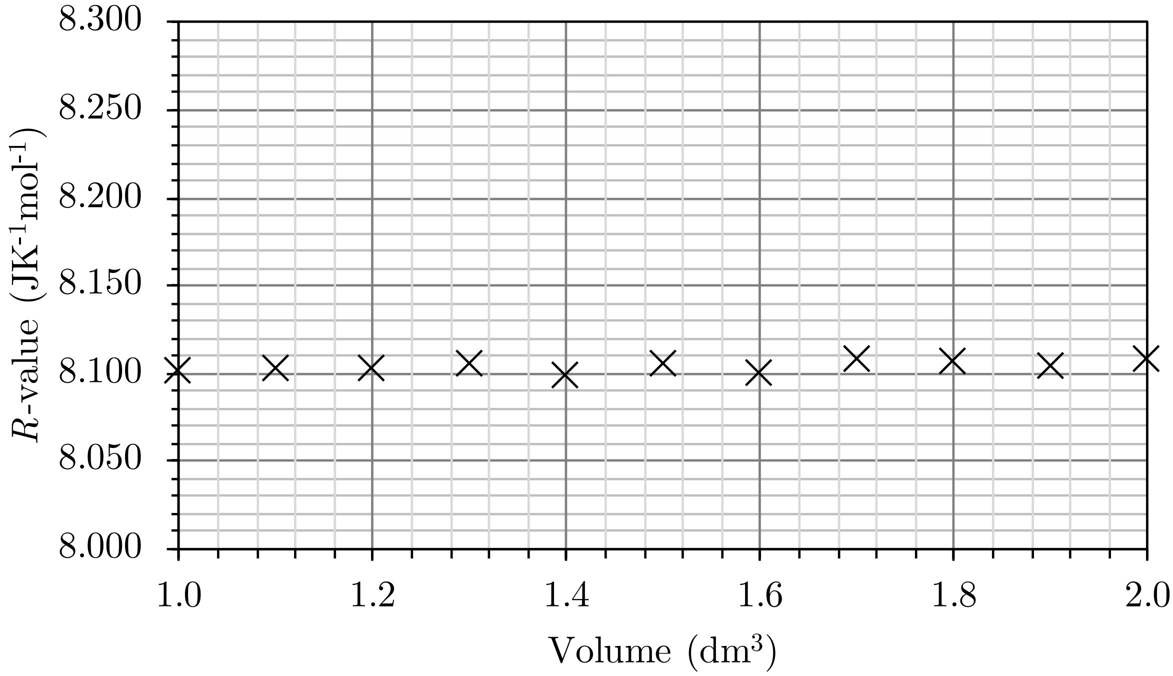

Figure 1: Plot of experimental R value versus volume.

From the plot above we can see that all points yield a value around 8.1, with the highest being 8.109 and the lowest being 8.100. From this, we can work out a percentage error of the experimental R values ($\frac{\text{maximum}-\text{minimum}}{2}$) that turns out to be 0.45%. This indicates very high experimental precision.

Plotting a P-V diagram can give a better idea of the relationship between pressure and volume.

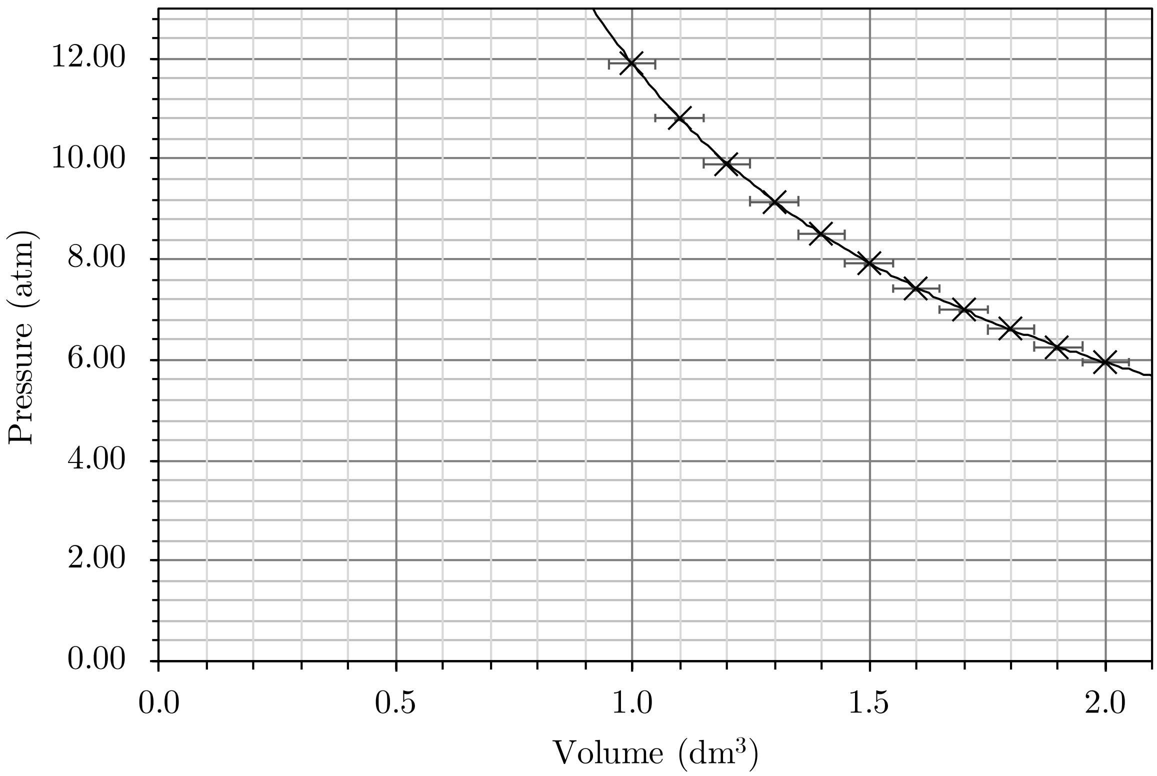

Figure 2: P-V diagram of the raw data table.

From the P-V diagram, we can observe the isotherm that has formed from the data points from the exponential nature of the trend. This graph very well depicts the inverse proportionality of pressure and volume, as predicted by Boyle's law. The curve of best fit that was drawn aims to predict the position of the isotherm line. Due to the fact that the points do not fall on a linear trajectory, minimum and maximum gradient lines have not been drawn, but the percentage error in the R value was still calculated through numerical means (see above).

Evaluation & Conclusion

The value R value that was obtained from the simulation (average of $8.105\,\mathrm{JK^{-1}mol^{-1}}$) is reasonably close to the expected value of $8.31\,\mathrm{JK^{-1}mol^{-1}}$. The fact that the error in readings was less than 0.5% suggests that the experiment itself was very precise in determining the value of R for the gas (helium) that was under simulation. Therefore, the larger 2.49% error when comparing to the expected value can be attributed to the fact that helium, although very light in terms of mass per particle, is not an ideal gas and therefore it would not be correct to expect it to behave like one. A different R value than that of the ideal gas law means that the constant of proportionality of pressure and volume is different for helium than for an ideal gas, but the fact remains that the two variables are inversely proportional (visible on the P-V plot, as well as the results table).

Overall, I believe that this experiment went well. The fact that it was a simulation and not a real experiment meant that the data collection stage was much more painless and did not require too much time to complete. In addition, the simulation allowed for measurements that would be quite difficult to achieve in reality without advanced equipment, such as a gas container that could maintain constant temperature, provide a reading for the pressure of the gas, as well as eliminate any gas leaks and other sources of error. Due to the fact that the data came from the simulation, there weren't any major anomalies and all readings were normal and expected. This also helped a great deal during data processing because the data fitted the theory very easily.

Despite all this, this investigation had some limitations. Primarily, the only gas that was tested was helium and this significantly constricted the scope of the investigation. Had the I investigated the behaviour of other gasses, more general conclusions could have been drawn that would be addressed at physical gasses as a whole instead of just helium.

Unfortunately, the simulation that I used only had support for helium (in fact, it had support for argon too but the feature did not seem to be functional). Also, another limitation was the fact that the simulation provided pressure values that were perfectly consistent every time no matter how many trials were done. This does not really reflect the reality of a physical experiment, in which one would expect to get slightly different values due to subtle measurement inconsistencies et.c. Hence, only one measurement of pressure was taken per volume reading and this somewhat decreased the reliability of the results.

Last but most certainly not least, there is a fundamental problem to using a simulation for such an investigation. As was briefly mentioned in the introduction, the conclusion that helium has an R value reasonably close to the ideal gas R value and that pressure and volume are inversely proportional cannot be used reliably as a real-world truly experimental conclusion due to the fact that the software running the simulation is itself based on the very same ideal gas laws that were the topic of this investigation. Arguably this does create a sort of paradox of using a theory to prove itself and could therefore somewhat invalidate the conclusion that was made.

References

http://highered.mheducation.com/olcweb/cgi/pluginpop.cgi?it=swf::100%25::100%25::/sites/dl/free/0023654666/117354/Ideal_Nav.swf::Ideal%20Gas%20Law%20Simulation.