

The Ideal Gas Law

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Introduction:

The ideal gas law is an approximation that enables scientists to treat all gases the same and eliminates the need for gas specific variables. The law describes the proportional relationship between pressure, temperature, volume, and number of particles. It has been crucial for numerous advancements in science and technology, because it provided critical understanding of gases by proving that they act alike, especially at low density and high temperature. Since all gases behave in similar manners regardless of their identity, they can be generalized as “ideal,” which allows scientists to make quicker and easier calculations when predicting their behavior. In order to classify a gas as ideal, scientists developed a molecular model to explain the observed macroscopic behavior of gas called the kinetic theory of gases. The kinetic theory defines a gas as a collection of particles in random perpetual motion that are constantly colliding. By the model, when a gas is characterized as “ideal” the gas is inert, has fully elastic collisions, moves randomly, and occupies no volume. The ideal gas equation uses the universal gas constant R , the product of Avagadro’s number and Boltzman’s constant, to allow scientists to equate the pressure and volume of a gas to its temperature and number of moles for any ideal gas. The law is the result of collaboration between numerous scientists beginning in the 1600s, including Robert Boyle, Joseph Louis Gay-Lussac, and Jaques Charles. We intend to see how accurately the term “ideal” applies to gases, regardless of the molecular interactions between the atoms, by exploring how the change in pressure and volume on a gas are related.

The combined knowledge of various scientists was very important to establish the current understanding of gases at a molecular level. Scientists use the principals of the gas law to more accurately predict how gases will behave in different situations and fully understand how

changes in pressure influence how the gas acts on its surroundings. Leading the quest to fully understand what gas is and how it can exert a force was Robert Boyle. Boyle experimented with pressure and volume while maintaining constant temperature, and found the relationship $P \propto 1/V$, known as Boyle's Law. Understanding the properties of gases made him one of the first chemists to try to understand matter at the molecular level and his work prompted many future scientists to further question the properties of matter.

Boyle's discovery played a key role in history, by paving the way for other scientists to employ pressure to do useful work. Boyle's experiments were conducted during a time of scientific enlightenment and were especially important in the invention of the steam engine in 1698, which played a critical role in the industrial revolution. The inventor, Thomas Savery, is thought to have worked with Boyle through the Royal Society, where Boyle's ideas could be shared and be a foundation for other research. (Musson; 47). Discovering the relationship between volume and pressure was only part of the puzzle.

After the principal ideas used by the Royal Society began to take hold, Joseph Louis Gay-Lussac discovered the relationship between volume and temperature. He was able to experimentally prove that a direct relationship exists between volume and temperature when the mass and the pressure remain constant. He explored how atmospheric air changed in density and contributed to the overall understanding of gas behavior. Gay-Lussac was known for his research on how gases combine and interact with one another. He also realized the relevance of Jacques Charles' work, who discovered the direct relationship between pressure and temperature when volume remains constant (Crosland; 27,56) and named the equation for the relationship Charles' law. Gay-Lussac helped develop the public's understanding of gas behavior by

publishing both Charles' law, relating volume and temperature, and his own law, relating pressure and temperature in 1802.

It wasn't until Benoit-Paul Emile Clapeyron, a French engineer exploring the Carnot engine, tied together Boyle's and Gay-Lussac's law to form the combined gas law, $PV = RT$; where P and V represent pressure and volume, respectively, T is the absolute temperature, and R is the constant determined from the experimental evidence by Henri Victor Regnault, a French chemist. Regnault's data supported the notion of an "ideal" gas because in his experiments, all gases behaved in a similar manner when the pressure or temperature was far from the point of condensation but as gases came closer to condensing, the ideal gas law held less true. Additionally, the law did not hold true for all gases, only water vapor, because it was based on mass rather than number of particles. In order to correct the equation so it holds true for all gases in their ideal state, Regnault integrated Avogadro's law to use moles, rather than mass, for the calculation. In 1873, the combined gas law was revised into the version scientists currently accept as the ideal gas law. The final ideal gas law was a result of the combined efforts, contributions and experimental results from many of the first physicists, chemists, and engineers. It was an important stepping stone that helped establish the kinetic theory of gases, sparked the scientific revolution, helped to progress industry, and supported the first models proposed for molecular interactions.

From the observed behavior of gases, scientists were able to come up with the kinetic theory to explain gases on a microscopic level. The kinetic theory was designed to support the scientific data and relationships explained by the ideal gas law. It explains the relationship by defining a gas as particles that are constantly moving and colliding. The collisions of particles with the side of the container make up pressure, which we can measure with a pressure sensor

that detects the number of collisions with a sensor surface. The kinetic theory also explains the relationship between pressure and volume. If volume increases and the number of particles remain constant, the density of particles would decrease and the likelihood of collisions with the walls of the container will also decrease, resulting in a decrease in the measured pressure. The kinetic theory provides a simple model that allows scientists to explain molecular movement at a very basic level however, a number of assumptions must be made if the model is used. The kinetic theory treats every particle as a point charge and does not take into account mass and or volume even though real gases must have both mass. By the kinetic theory, gases have no attractive forces between one another and therefore have no interactions aside from the fully elastic random collisions. Kinetic theory also defines temperature as the average kinetic energy of the collection of particles; if energy is added to the system, the average kinetic energy will increase and particles will move around faster. Faster particles will result in more collisions and a greater pressure. In our project we wanted to keep the kinetic energy constant and used only water in equilibrium with the air so that the speed of particles would not change. A constant temperature allows us to ensure that any changes in volume are a result of the changes in pressure and cannot be attributed to a change in the kinetic energy. This helps us to eliminate possible sources of error and have overall better results since we know that the change in volume observed is a result of the difference in the area which causes a density difference leading to a change in the number of collisions with any surface (Brucat). With so many assumptions in order for the gas law to hold true, it is hard to imagine the behavior of gases acting in accordance to the gas law while considering the real structure of gases to predict a better model for the behavior of the gases.

In order to confirm that the value of R is constant with any gas, we wanted to prove for ourselves that the ideal gas law holds true regardless of interactions such as Vander Waals forces, dipole-dipole interactions, bonding arrangements and size. To do this, we observed 3 gases of different molecular identities; Helium, Carbon Dioxide, and Nitrogen. Helium is a member of the inert group of noble gases on the periodic table and does not have many intermolecular interactions since it is very small, consists of a single molecule without any dipole, has a full octet in its valence shell so it does not form bonds, and experiences very little or no Vander Waals forces. We propose helium will act the most similarly to an ideal gas out of the three we are testing. Nitrogen is a diatomic element, meaning its most stable state is when bound to another molecule identical to itself. Nitrogen naturally exists in the N_2 form and does not interact with the surrounding molecules. We suspect it will act closely to an ideal gas since it has no dipole moment or free electrons, but it will not hold to the gas constant as closely as helium because Van der Waals forces will be slightly stronger on the bigger molecule. Carbon dioxide is an even larger molecule with two partially negative oxygen atoms bound to a central partially positive carbon. We suspect that due to its bigger mass and partially charged regions, carbon dioxide will be the furthest from the accepted R value of all of the three gases tested, and would need an addition to the ideal gas law to account for its intermolecular interactions. When we explore the relationship between volume and pressure we will get a better understanding of the gas law and be able to determine if the law should be modified to account for a real gas that may not act ideal.

By comparing 3 gases with different identities and characteristics, we will determine how large of a role, if any, intermolecular interactions play in gases at a constant temperature. If we observe results that are not consistently significantly different from the ideal gas law, our data

will support the kinetic theory of gases and the assumptions that must be made to employ it. If we see trends where any gas differs from the ideal, we will know that the ideal gas law is an optimistic tool used for very basic calculations and the assumptions of the kinetic theory do not hold true; in order to have more accurate calculations another constant would need to be included to account for molecular interactions based on the specific gas.

Objective:

The ideal gas constant, R , is a number that is used regularly, especially by introductory chemistry and physics students when learning about kinetic theory and gas behavior. To prove the ideal gas law and help establish our own opinion of the relevance of the interactions of gas molecules and their adherence to the ideal model, our goal is to experimentally determine the R value of various gases with different known characteristics by comparing a change in volume as a result of a pressure change while maintaining constant temperature and number of moles.

Through this investigation, we hope to achieve a better understanding of the properties of gases and their behavior and confirm for ourselves the relationship between pressure and volume of a real gas vs. an ideal gas and determine if a noticeable difference exists.

Methods:

In order to determine the ideal gas constant, R , for each individual gas we will find two volumes (one at the pressure in the balloon, V_b , and the other at the pressure in the room, V_{STP}), the pressure of the gas in the room, the temperature of the room, and the number of moles of gas.

To do so, we used:

- Pressure Sensor
- 2 Large Plastic Containers
- Logger Pro Software & Computer
- 1 Liter Beaker
- Bag of small Balloons
- Plastic Tubing

- Access to Gases:
 - Helium
 - Nitrogen
 - Carbon Dioxide
- Large Graduated Cylinder
- Digital Thermometer



(1)



(2)



(3)



(4)



(5)



(6)

Figures 1-6 detail the procedure used in the experiment.

To begin the experiment a single balloon was stretched and filled with a specific gas to a point that the balloon was expanded enough to exert a pressure on the gas inside of it (Figure 1). The pressure of the gas inside of the balloon was then measured using the pressure sensor and data acquisition system (Logger Pro Software, Figure 2). Little to no gas was lost in this process. A 1 Liter beaker was then completely filled and placed in the large plastic container. The volume of the gas while inside of the balloon, V_1 , was measured using displacement by submerging the balloon in the beaker using a hook to push it down (Figure 3). Another large plastic container was filled with water and allowed to settle to room temperature. The large graduated cylinder was also filled with room temperature water and submerged upside down in

the large plastic container (Figure 4). Careful attention was paid to making sure that no air was left inside of the graduated cylinder. Plastic tubing was then filled with water and inserted up into the graduated cylinder (Figure 5). A second volume, V_{room} , was recorded under STP as the gas in the balloon was attached to the tubing and allowed to displace the water in the cylinder (Figure 6). Attention was paid to making sure that all of the gas left the tubing and entered the cylinder. This volume was recorded as the volume under standard pressure in the room and was the volume used to calculate the number of moles of gas. With all variables pressure (P), volume (V), moles (n), and temperature (T) present, the Ideal Gas Law, $PV=nRT$, was used to calculate the R value for the gas in question. This process was repeated for each of the three gases: Nitrogen, Helium, and Carbon Dioxide.

Data:

Calculations of individual gas constants were based around three values for our inflated balloon, one measured directly (pressure, P_b) and two calculated from directly measureable phenomena (volume, V_b , and number of moles, n). In order to calculate these two, temperature was measured (to ensure it stayed approximately 297 K), the volume of water the filled balloon displaced when submerged completely was measured, and finally the volume of the gas when it was allowed to come to room conditions was measured. In order to calculate volume and number of moles of gas, we needed the following equations:

$$V_b = V_l \quad (\text{eq. 1})$$

$$n = V_{STP} / V_m \quad (\text{eq. 2})$$

$$V_{STP} = V_{\text{room}} * P_{STP} / P_{\text{room}} \quad (\text{eq. 3})$$

where V_b , is the volume inside the balloon, V_l is the volume of the displaced water, n is the number of moles, V_{STP} is the volume our gas would take up at standard temperature and pressure

(STP), V_m is the molar volume at STP (about $0.0244 \text{ m}^3/\text{mol}$)(Hooker), V_{room} is the volume of the gas at room conditions P_{STP} , is standard pressure ($1.01 \times 10^5 \text{ N/m}^2$) and P_{room} is room pressure (at our elevation, 1320 m (SLC Classic), about $8.7 \times 10^5 \text{ N/m}^2$)(Engineering Toolbox). Equation 1 is used to calculate the volume of the balloon and equations 2 and 3 are used to calculate the number of moles of the gas inside the balloon. Equation 2 would be enough to calculate number of moles at sea level; however this experiment was not conducted at sea level, therefore we also used equation 3, which relates the volume of gas at STP to the volume of a gas at our pressure, They can be combined and simplified to:

$$n = (V_{room} * P_{STP}) / (V_m * P_{room}) \quad (\text{eq. 4})$$

Or, with values substituted where appropriate:

$$n = V_{room} / (.0283 \text{ m}^3/\text{mol}) \quad (\text{eq. 5})$$

Once V_b , P_b , T and n were obtained they were used to calculate each individual gas constant, R , by rearranging the ideal gas law as:

$$R = (P_b * V_b) / (n * T) \quad (\text{eq. 6})$$

Measured Values:	P_b (N/m ²)	V_1 (m ³)	V_{room} (m ³)	T (K)
	91500	7.21E-04	7.25E-04	295.25
	mol density (m ³ /mol):	0.0283		
	n (mol)= $V_{room} / (.0283) =$	$7.25E-04 \text{ m}^3 / (.0283 \text{ m}^3/\text{mol}) =$.0256 mol	
	R (J/(mol*K))= $(P_b * V_b) / (n * T) =$	$(91500 \text{ (N/m}^2) * 7.21E-04 \text{ m}^3) / (.0256 \text{ mol} * 295.25 \text{ K}) =$	8.72 J/mol*K	

Table 1. Sample calculations for molarity, n , and gas constant, R

	P_b (N/m ²)	V_1 (m ³)	V_{room} (m ³)	n (mol)	T (K)	R (J/(mol*K))
Helium						
Trial 1	91500	7.21E-04	7.25E-04	0.0256	295.25	8.72
Trial 2	91300	7.03E-04	7.50E-04	0.0265	295.25	8.20
Trial 3	90400	3.32E-04	3.65E-04	0.0129	297.35	7.83
Trial 4	90500	3.96E-04	4.25E-04	0.015	297.35	8.03
Average						8.2
Carbon Dioxide						
Trial 1	91800	5.02E-04	5.25E-04	0.0186	295.25	8.41
Trial 2	91600	5.25E-04	5.92E-04	0.0209	295.25	7.79
Average						8.1
Nitrogen						
Trial 1	91500	7.00E-04	7.72E-04	0.0273	295.25	7.95
Trial 2	92600	5.80E-04	6.70E-04	0.0237	295.25	7.68
Trial 3	90600	4.47E-04	4.75E-04	0.0168	297.35	8.11
Trial 4	90700	5.20E-04	5.61E-04	0.0198	297.35	8.00
Average						7.9
mol density (m³/mol):	0.0283					
	Accepted	He	CO₂	N₂		
R (J/(mol*K))	8.31	8.2	8.1	7.9		
SDM (+/-)	N/A	0.2	0.3	0.1		

Table 2. Data and Calculations

Results/Discussion:

The goal set to be achieved by this project was to test three different gases under the same and “as ideal” circumstances as possible to see if they behave like an ideal gas. To do this, the ideal gas constant was calculated from measured values of pressure and volume. The moles found of each gas were calculated by placing the gas under STP conditions. Using the calculated

value for the ideal gas constant, R , it was then compared to the accepted value for an ideal gas to see if the tested gases indeed did behave like ideal gases. Multiple trials were conducted for each gas to give a good set of data where the average R value and standard deviation of the mean could be calculated and analyzed.

What can be observed from the data is that all three gases that were tested came out to behave similarly to an ideal gas. Each calculated value had a seemingly minimal difference from the actual ideal gas constant. The difference seen between the calculated values and the accepted value arises from systematic error throughout the procedures for the values being measured. The measurements for the volumes were taken by the displacement of water into a container that was then poured into a large graduated cylinder. This caused the loss of small amounts of water that were left behind in the larger container as water drops. Another minor error with this procedure was in adding enough mass to displace the water by breaking the surface tension of the water. This meant that more mass of the gas had to be added before the water began to spill over the edge of the container. It further caused the measurements to be a smaller volume than there actually was of the gas. The balloon, however, added a small amount of mass to displace the water as well which would only in part counteract a small portion of the volume of water remaining in the container after displacing the water. Error could have also come from the pressure readings when attaching the balloon to the pressure sensor. The mouth of the balloon was held firmly in place around the pressure tube, but small amounts of gas were able to leak through and the pressure reading fluctuated within a range of 500 N/m^2 . Our best judgment was used in acquiring what we thought to be the accurate reading of the pressure for each gas. Minor discrepancies should be noted throughout the procedures when considering the accuracy of the tools used. Under the circumstances, the pressure gauge sensor was self-calibrated when

connected to the logger pro software. The graduated cylinder used to measure the volumes of the gases was deemed accurate to give correct volumes, and STP conditions were used to calculate the moles of the gases. The STP conditions are a way to find out how many moles of a gas there are when the pressure a gas is exerting on water inside a container is the same as the pressure outside of the container.

For every trial conducted using each gas, the gas was kept at room temperature, as was the water. The volume for each gas happened to be larger for the second reading than it was for the first reading. This was partially due to the fact that they were at different pressures. The first volume measured was under pressure when submerged to get the pressurized volume of the gas and the hook apparatus, V_1 (see image 3). This did include the volume of the hook used to submerge the balloon as well as the volume of the balloon containing the gas. In response, necessary measurements were taken to eliminate the volume of the hook from the total volume. The first volume measured then was used in our calculation for the ideal gas constant value, R , while the second volume was used in acquiring a completely separate value, number of moles, n , to use in our equation for calculating the R value. The second volume is the volume at standard temperature and room pressure such that the pressure inside the container holding the gas was the same as the pressure outside of the container, in the room. It was further used to identify the number of moles from the volume measured because at standard temperature and room pressure at our elevation, there are 0.0283 meters cubed for every mole (see equation 5).

The accepted value for the ideal gas constant is $8.31 \text{ J}/(\text{mol}\cdot\text{K})$ and using the standard deviation of the mean for each gas we determined whether they behaved like an ideal gas, within experimental uncertainty. Helium and carbon dioxide behaved like ideal gases because their average R values, $8.19 \text{ J}/(\text{mol}\cdot\text{K})$ and $8.1 \text{ J}/(\text{mol}\cdot\text{K})$ respectively, were within one standard

deviation of the mean ($\pm 0.2 \text{ J}/(\text{mol}\cdot\text{K})$ and $0.3 \text{ J}/(\text{mol}\cdot\text{K})$) of the accepted ideal gas constant ($8.31 \text{ J}/(\text{mol}\cdot\text{K})$). Nitrogen's calculated R value ($7.94 \text{ J}/(\text{mol}\cdot\text{K})$), however, was not within one standard deviation of the mean ($\pm 0.1 \text{ J}/(\text{mol}\cdot\text{K})$) of the accepted value. It therefore does not behave like an ideal gas whereas helium and carbon dioxide do (see Table 2).

What was noticed in particular about the three gases tested (helium, carbon dioxide and nitrogen) was that helium behaved more like an ideal gas than nitrogen or carbon dioxide. This was expected because helium is a very light gas with weak intermolecular forces; collisions between helium atoms would behave more elastically than gases that are larger and have greater intermolecular forces. What was then expected for the behavior of the other two gases was that nitrogen would behave more like an ideal gas than carbon dioxide even though it is a diatomic molecule because carbon dioxide is more massive and has a greater bond polarity; however, our results showed the opposite, as carbon dioxide behaved, within experimental uncertainty, like an ideal gas, while nitrogen did not. An ideal gas has no interactive forces between the molecules of a gas or its container and that its volume is finite. Since neither are true for any real gas, no gas would be expected to behave exactly like an ideal gas. Gases with smaller atoms would be expected to behave much more like an ideal gas than any larger sized gas. The data acquired by the measurements and calculations, although containing minor systematic errors, is good data that gave excellent ideal gas constant values. What can be concluded about helium and carbon dioxide is that they do behave like ideal gases, while nitrogen does not. The behavior of helium and carbon dioxide is how other gases are expected to behave which is what allows for the equation $PV=nRT$ to be used. By knowing any three of the four variables used in the ideal gas law, any other variable can be calculated for some gas.

Conclusion:

Based on our data, we are able to conclude that the ideal gas constant is an appropriate figure used to complete the Ideal Gas Law. While none of the calculated values matched up exactly with the expected value, our % discrepancies were all under 10%. Such small discrepancies lead us to believe that the differences were due to the errors discussed in the previous section. Within error, our goal was achieved by verifying that the equation used in our chemistry and physics classes works for a variety of gases.

References:

[1] A.E. Musson and Eric Robinson. Science and Technology in the Industrial Revolution.

Gordon and Breach. New York. Reprinted in 1994

[2] Brucat, PJ. "Kinetic-Molecular Theory." University of Florida."

<http://itl.chem.ufl.edu/2041_u00/lectures/lec_d.html>

[3] "The Engineering Toolbox" <http://www.engineeringtoolbox.com/air-altitude-pressure-d_462.html>

[4] Hooker, P (2012, March) Personal Interview.

[5] Maurice P. Crosland. Gay-Lussac: Scientist and Bourgeois. Cambridge University Press.
2004

[6] "SLC Classic" <http://www.slcclassic.com/info/area_info/faq_new.htm>

[7] Smith, B (2012, March) Personal Interview.

[8] William B. Jensen. The Universal Gas Constant. J. Chem. Educ., 2003, 80, 731-732