The Ideal Gas Law: How Can a Value of R for the Ideal Gas Law Be Accurately Determined Inside the Laboratory? Adapted from Argument Driven Inquiry

**Guiding Question: How can a value of *R* for the ideal gas law be** **accurately determined inside the laboratory?**

**Introduction**

A *gas* is the state of matter that is characterized by having neither a fixed shape nor a fixed volume. Gases exert pressure, are compressible, have low densities, and diffuse rapidly when mixed with other gases. On a submicroscopic level, the molecules in a gas are separated by large distances and are in constant, random motion. A gas can be described using four measurable properties: pressure (*P*), defined as the force exerted by a gas per unit area; volume (*V*), defined as the quantity of space a gas occupies; temperature (*T*), defined as the average kinetic energy of the molecules that make up a gas; and the number of moles of gas (*n*). The relationships among these properties are summarized by the gas laws, as shown in Table L15.1.

TABLE L15.1

**The gas laws**

|  |  |  |
| --- | --- | --- |
| **Gas law** | **Relationship** | **Equation** |
| Boyle’s law | *V* α 1/*P* (*T* and *n* are held constant). As gas pressure increases, gas volume decreases. | *P1V1 = P2V2* |
| Charles’ law | *V* α *T* (*P* and *n* are held constant). As gas temperature increases, gas volume increases. | *V1/T1 = V2/T2* |
| Gay-Lussac’s law | *P* α *T* (*V* and *n* are held constant). As gas pressure increases, gas temperature increases. | *P1*/*T1* = *P2*/*T2* |
| Avogadro’s law | *V* α *n* (*P* and *T* are held constant). As the number of moles of gas increase, gas volume increases. | *V1*/*n1* = *V2*/*n2* |
| Combined law | *V* α *T*/*P* (*n* is held constant); obtained by combining  Boyle’s law, Charles’ law, and Gay-Lussac’s law. | (*P1V1*)/*T1* = (*P2V2*)/*T2* |

The ideal gas law combines Boyle’s law, Charles’ law, Gay-Lussac’s law, and Avogadro’s law to describe the relationship among the pressure, volume, temperature, and number of moles of gas. Émile Clapeyron is often given the credit for developing this law. The ideal gas law provides chemists with a powerful predictive tool that helps them understand how gases will react in different systems. The ideal gas law is expressed mathematically as *PV* = *nRT*. *P* is pressure in atmospheres (atm); *V* is volume in liters (L); n is the number of moles of gas (mol); and *T* is absolute temperature in Kelvin (K). The remaining component of the ideal gas law is *R*, which is called the ideal gas constant. The theoretical value of *R* that is often reported in textbooks and handbooks is 0.0821 L•atm/mol•K or 8.314 L•kPa/mol•K.

As chemists worked to determine an exact value for *R* in the mid-1800s, they were faced with numerous challenges. First, they had to develop a method that they could use to generate the experimental data they needed to calculate a value for *R* from the ideal gas law. Second, they needed to improve the precision of their measurements. The French chemist Henri Victor Regnault was able to overcome many of these challenges and generate some of the most precise experimental data about the properties of gases at that time. Rudolf Clausius, a German physicist, then used Regnault’s data to calculate the earliest published value for *R*. This value for *R*, however, was not very precise by current standards. Fortunately, there have been numerous advancements in the methods and tools that chemists use to measure the properties of gases, and the value of *R* has become increasingly precise over time. There is, however, always room for improvement. In this investigation, you will have an opportunity to follow in the footsteps of Clapeyron, Regnault, and Clausius by designing, conducting trials of, refining, and then evaluating a method that can be used to calculate a precise value for the ideal gas constant.

**Materials**

You may use any of the following materials during your investigation:

**Consumables**

• 6 M hydrochloric acid (HCl) solution

• Magnesium (Mg) ribbon

**Equipment**

• Side-arm Erlenmeyer flask with stopper (50 ml)

• Pneumatic trough

• Plastic or rubber tubing (50 cm long)

• Electronic or triple beam balance

• Graduated s (one each 50 ml, 100 ml, 250 ml, and 500 ml)

• Utility clamp

• Ring stand

**Safety Precautions**

Follow all normal lab safety rules. Hydrochloric acid is corrosive to eyes, skin, and other body tissues. Your teacher will explain relevant and important information about working with the chemicals associated with this investigation. In addition, take the following safety precautions:

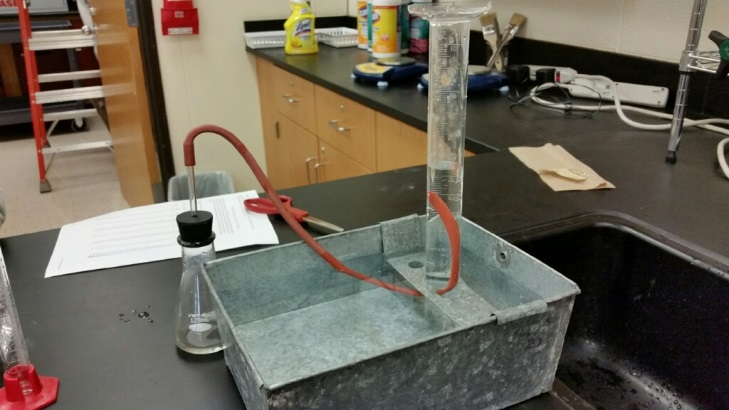
• Wear indirectly vented chemical-splash goggles and chemical-resistant gloves and apron while in the laboratory.

• Handle all glassware with care.

• Wash your hands with soap and water before leaving the laboratory.

**Procedure**

# Partner 1

1. Fill your pneumatic trough with tap water, making sure the water level rises above the tray.
2. Fill your gas collecting cylinder to the rim with tap water. Cover the cylinder with the petri dish, making sure no air (little air) has been trapped; water may squirt out the sides as you do so. Holding onto the petri dish, invert the cylinder and submerge it in the trough. Once submerged, remove the petri dish and slide the cylinder onto the tray. You may need to lift the cylinder a bit to do so, **BE CAREFUL NOT TO PULL THE CYLINDER COMPLETELY OUT OF THE WATER (TROUGH) OR YOU WILL HAVE TO RESTART**. Position the cylinder so that its opening is over the center hole in your tray.
3. Insert the tubing into the cylinder through the hole in the tray, be sure it is inserted enough that the tubing will not fall out.
4. Make sure the other end of the rubber tubing fits tightly onto the glass tubing inserted into a rubber stopper.

# Partner 2

1. Obtain a 2 inch piece of magnesium ribbon and sand until shiny to remove any oxidation.
2. Find the mass of the magnesium ribbon and record.
3. Measure 10. mL of 6 M HCl. Record the **actual** volume to **ONE** uncertain digit.

# Both Partners

1. Pour the 6.0 M HCl into the flask and drop in the magnesium ribbon while **quickly capping the flask with the rubber stopper.**
2. Once the flask stops fizzing read the final volume to ONE uncertain digit.
3. Record the temperature of the water and atmospheric pressure.
4. Repeat for a second trial.

**Name:**

**Reactants Data**: Mg (s) + 2 HCl (aq) 🡪 MgCl­2 (aq) + H2 (g)

|  |  |  |
| --- | --- | --- |
| **Measurement** | **Trial 1** | **Trial 2** |
| Mass of magnesium ribbon (g) |  |  |
| Volume 6.0 M HCl (g) |  |  |

1. Calculate the mols of magnesium used.
2. Convert volume of HCl used to liters.

1. Calculate the mols of HCl used using the following equation:
2. Determine the limiting reactant and calculate the theoretical yield of H2 in liters.

# Data

**Volume Data:**

|  |  |  |
| --- | --- | --- |
| **Measurement** | **Trial 1** | **Trial 2** |
| Volume of gas (mL) |  |  |

# Reactants Data

# Pressure Data

|  |  |
| --- | --- |
| Atmospheric Pressure (mmHg) |  |
| Partial Pressure of H2O (mmHg) |  |
| Partial Pressure of H2 (mmHg) |  |

Temperature H2O = \_\_\_\_\_\_\_\_\_\_\_ºC vvvvvvvvvvvv \_\_\_\_\_\_\_\_\_\_\_K

# Calculations

1. Using your data calculate R for the best trial.
2. Calculate your percent error of R.
3. Explain how the following would affect the measured value of R.
   1. Student did not cap flask in a timely manner.
   2. Student did not subtract the partial pressure of water.
   3. Student spilled a three drops of HCl.