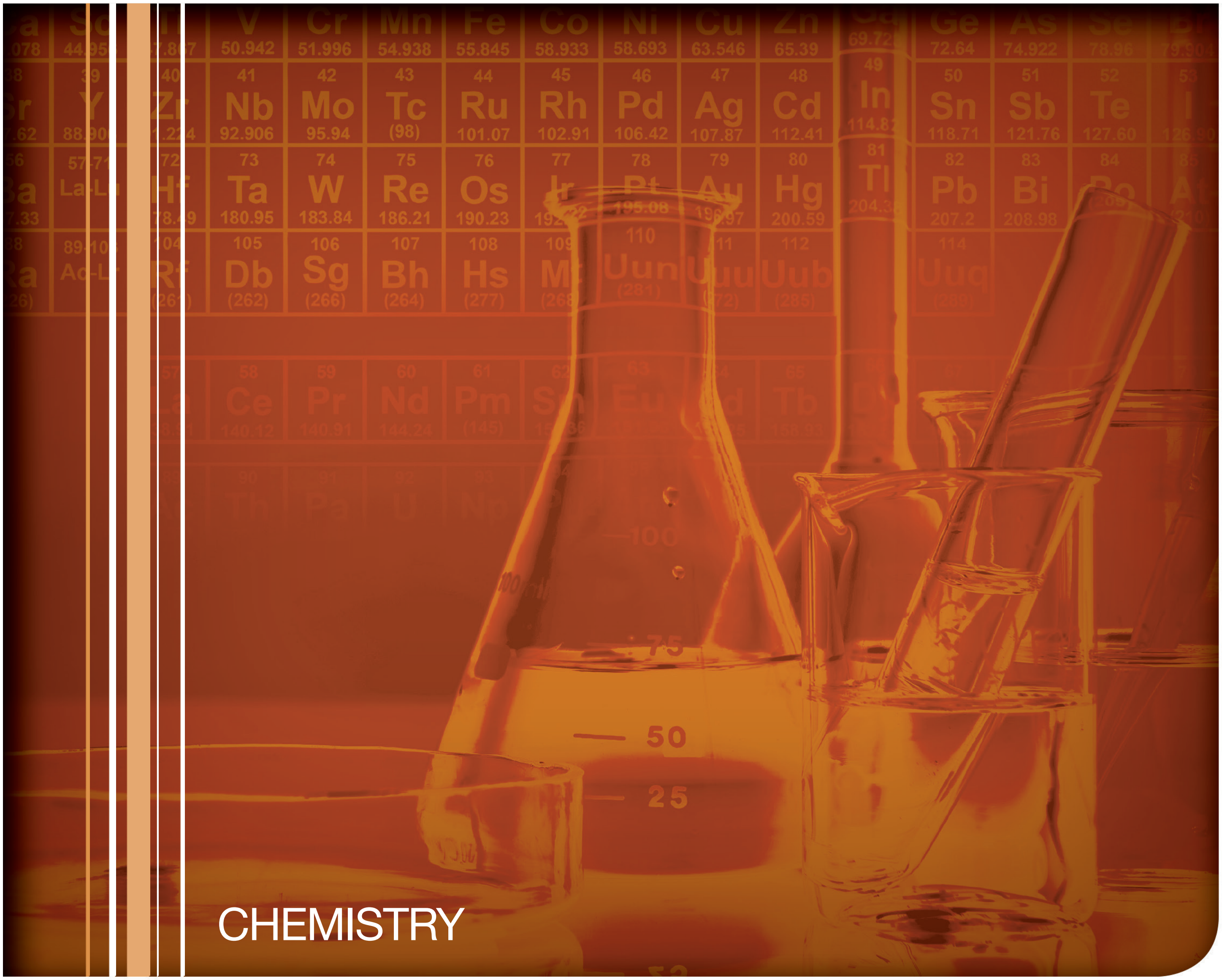


# Determination of Ideal Gas Law Constant


Investigation Manual

# DETERMINATION OF IDEAL GAS LAW CONSTANT

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## Overview

Students will test the relationship between pressure and the number of moles of gas as described by the ideal gas law. Specifically, they will generate and collect oxygen gas from the decomposition of hydrogen peroxide in an enzyme-catalyzed reaction. They will then determine the volume and the number of moles of oxygen generated. From these quantities, they will predict the constant that governs the behavior of ideal gases.

## Outcomes

- Determine the ideal gas law constant.
- Use a catalyst to initiate a gas-forming reaction.
- Use the ideal gas law to determine the number of moles of gas generated in a reaction.
- List the kinetic theory assumptions about ideal gases.

## Time Requirements

Preparation ........................................ 5 minutes
Activity 1 ........................................ 60 minutes
Activity 2 ........................................ 30 minutes

## Key

Personal protective equipment (PPE)

goggles gloves apron

follow link to video

photograph results and submit

stopwatch required

warning corrosion flammable toxic environment health hazard

## Background

Gases are one of the three states of matter on earth and an integral part of the physical world. Understanding the behavior of gases is important to understanding the environment and the biological systems that it supports. An obvious example is the air that supplies oxygen, essential to all living things, and carbon dioxide, essential to plants and, therefore, our source of food. In this investigation, you will gain some insight into how gases, which are generally invisible, can be measured and how their behavior can be assessed.

In this investigation, oxygen gas will be used. The atmosphere is about 20% oxygen, but this experiment requires pure oxygen gas. Pure oxygen in this investigation will be acquired from the decomposition of hydrogen peroxide to form oxygen gas, as shown below:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

However, this reaction is relatively slow, and the rate of decomposition must be increased if the investigation is to be completed in a reasonable time. A **catalyst** can be introduced into a reaction to reduce the activation energy, effectively making it easier for a reaction to occur and to allow it to proceed at a reasonable rate. Importantly, a catalyst is not consumed and is not altered in the reaction, so it has no effect on the products or reactants and does not need to be included as either a product or reactant in the molecular equation. An enzyme from yeast, **catalase**, will serve as the catalyst here. Catalase is present in most organisms and enables their cells to decompose peroxides that are produced as part of normal cellular function. The presence of a catalyst is indicated in the molecular equation as follows:

$$2H_2O_2 \xrightarrow{catalase} 2H_2O + O_2$$

As the gas is produced, it must be measured. The gas will be formed and collected in the apparatus shown in Figure 1.

**Figure 1.**

Hydrogen peroxide and yeast will be added to the dropper bottle, shown on the left in Figure 1. This combination will generate oxygen gas, which will flow through the tubing and into the inverted graduated cylinder. The graduated cylinder will initially be full of water, and as gas is generated, the water will be displaced. The volume of gas in the graduated cylinder at any time will be equal to the volume of oxygen gas generated in the decomposition reaction. The apparatus is designed so that the pressure in the reaction vessel will remain constant. This means that the pressure of the gas in the graduated cylinder will equal the pressure in the bottle, and the pressure in the apparatus will not change over the course of the experiment. The pressure in the entire reaction apparatus is designed to be equal to the ambient air pressure.

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## Background continued

**Kinetic Theory of Gases**

The **kinetic theory of gases** describes how the movement of gas particles is modeled under several assumptions below. These assumptions mean that kinetic theory very closely approximates the properties of a real gas, but the assumptions mean that the kinetic theory of gases models a so called **ideal gas**, and is not impacted by the assumptions below that would apply to a real gas.

a. Gases consist of particles in constant, random motion and will move in a straight line until a collision occurs.

b. Particles are point masses with no volume. The particles are so small compared with the space between them, that we do not consider their size in ideal gases.

c. No molecular forces are at work. This means that there is no attraction or repulsion between the particles.

d. Gas pressure is due to the molecules colliding with the walls of the container. No energy is lost or gained from collisions.

e. The time during a collision is negligible compared with the time between collisions.

f. The temperature is the average kinetic energy of all of the gas molecules. Individual gas molecules may move at different speeds, but the temperature and kinetic energy of the gas refer to the average of these speeds.

g. The kinetic energy of a gas particle is directly proportional to the temperature. An increase in temperature increases the speed at which the molecules move.

h. All gases at the same temperature have the same average kinetic energy.

i. Lighter gas molecules move faster than heavier molecules.

The behavior of gases is described by the **Ideal Gas Law**, which incorporates several parameters that describe the gases under consideration and the interrelationships of these parameters. These parameters are pressure (p), volume (V), moles of gas (n), and its temperature (T). The Ideal Gas Law is:

$$PV = nRT$$

Each of these parameters is discussed below. When studying gases, it can be helpful to visualize the individual atoms or molecules that make up the gas, and this discussion employs this useful tactic.

**Pressure (P)** is the force a gas exerts on its surroundings. This force is the result of the gas molecules colliding with a surface. Pressure can be expressed in pascals (Pa), millimeters mercury (mm Hg), or atmospheres. The SI unit for pressure is a pascal, equal to 1 newton per square meter; however, most people are more familiar with atmospheres (atm) as 1 atm is equivalent to atmospheric pressure at sea level.

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**Volume** is the three-dimensional space that a gas fills and is most commonly measured in liters (L) or mL.

The variable **n** is the number of moles of gas in the system and has a unit of mole (mol).

**Temperature** is a measurement of the average speed of the molecules of gas and is expressed in units of Kelvin (K). A sample of gas with molecules that have a high average speed will have a higher temperature than a sample of gas with molecules with a lower average speed. Temperature measured in Celsius can be converted to Kelvin through the following equation:

$$T_{Kelvin} = T_{Celsius} + 273.15$$

The constant **R** is the actual gas law constant and has a value of 0.08206 (atm * L)/(mol * K).

In this experiment, the relationship known as **Avogadro's Law** will be investigated. Avogadro's Law states that the volume of a gas is directly proportional to the number of moles of the gas. The pressure and temperature will be held constant in this investigation, allowing the experimental gas constant **R** to be calculated if the moles of gas are graphed versus their volume:

$$V = \frac{RT}{P} n$$

This can be compared to the equation of a line.

$$y = mx+b$$

In this situation, volume (V) is the y, and x is the number of moles (n).

$$RT/P = m$$

Where m is the slope of the line and b is equal to zero, since, if there are no moles of a gas, they would not occupy any volume. To solve for R, the equation can be rearranged as follows.

$$R = mP/T$$

This experimental R value can then be compared to the actual R value through a percent error calculation.

$$\% \text{ error} = (|\text{actual} - \text{experimental}|/\text{actual}) * 100$$

This calculation helps to show how close the experimental value is to the known value of R.

# DETERMINATION OF IDEAL GAS LAW CONSTANT

## Materials

**Included in the materials kit:**

Bottle, 15 mL

Plastic tubing


Yeast packet

**Needed from the chemical kit:**


Hydrogen peroxide, 3%

**Needed from the equipment kit:**

Graduated cylinder, 10 mL

Graduated cylinder, 50 mL

Beaker, 250 mL

Weigh boat

**Needed but not supplied:**
- Graphing program
- Tap water

**Reorder Information:** Replacement supplies for the Determination of Ideal Gas Law Constant investigation can be ordered from Carolina Biological Supply Company, item number 580344.

**Call:** 800-334-5551 to order.

## Safety

Wear your safety goggles, chemical apron, and gloves at all times while conducting this investigation.

Read all the instructions for this laboratory activity before beginning. Follow the instructions closely and observe established laboratory safety practices, including the use of appropriate personal protective equipment (PPE) described in the Safety and Procedure sections.

Hydrogen peroxide causes skin irritation.

Hydrogen peroxide is corrosive and can cause serious eye damage.

Do not eat, drink, or chew gum while performing this investigation. Wash your hands with soap and water before and after the investigation, and clean up the work area with soap and water after finishing. Keep pets and children away from lab materials and equipment.

## Preparation

1. Read procedure thoroughly.
2. Locate and clean work area.
3. Gather all needed materials.

# ACTIVITY

## ACTIVITY 1

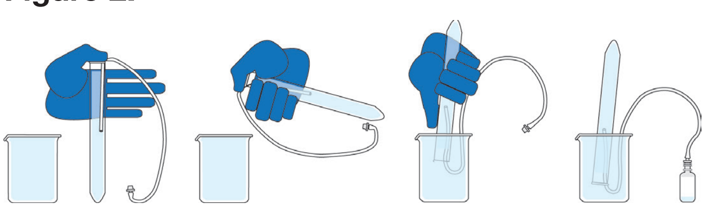
1. Determine the current air temperature using the thermometer and record it in Data Table 1.
2. Fill the 50-mL graduated cylinder completely with water.
3. Fill the 250-mL beaker completely with water.
4. Attach the dropper top to one end of the clear tubing.
5. Insert the opposite end of the tubing into the 50-mL graduated cylinder.

**Note:** The next step will cause the water in the beaker to overflow; have paper towels or a baking pan ready to collect any spills.

6. Place thumb over opening of the 50-mL graduated cylinder and invert into the beaker. See Figure 2 below for an illustration.

   See how to assemble the collection apparatus.
   http://players.brightcove.net/17907428001/HJ2y9UNi_default/index.html?videoId=4573398321001

**Figure 2.**



7. Record the initial volume of gas in the graduated cylinder in Data Table 1. If no gas is present, the volume will be 0.
8. Turn on the balance.
9. Place a weigh boat on the balance and tare it.
10. Weigh 0.1 gram of yeast into the weigh boat.
11. Pour the yeast into the empty 15-mL bottle.
12. Measure 1 mL of hydrogen peroxide in the 10-mL graduated cylinder. Record the exact volume in Data Table 1.
13. Prepare the end of the tubing with the dropper top.

**Note:** The top will need to be snapped in place immediately after the hydrogen peroxide is added.

14. Pour the hydrogen peroxide into the 15-mL bottle.
15. Snap the dropper top with tubing onto the bottle.
16. Allow the reaction to proceed until no additional gas has been generated for at least one minute. At this point, the reaction can be considered complete.
17. Record the final volume of gas in Data Table 1.
18. Calculate the change in volume in mL and record in Data Table 1.
19. Remove the top from the bottle, pour the contents down the sink, and rinse the bottle three times with water.
20. Repeat the activity using 1 mL hydrogen peroxide.
21. Repeat the activity varying the volume of hydrogen peroxide in **step 12** as listed in the table below.

| Trials 3 and 4 | Trials 5 and 6 | Trials 7 and 8 |
|---|---|---|
| 2 mL $H_2O_2$ | 3 mL $H_2O_2$ | 4 mL $H_2O_2$ |

## ACTIVITY 2

1. Calculate the molarity of (mol/L) of 3% hydrogen peroxide and record it in Data Table 2. The term “3% concentration” is a weight-to-weight measurement, which means that there are 3 g solute per 100 g solvent. The molar mass of hydrogen peroxide is 34.01 g/mol.
2. Calculate the moles of hydrogen peroxide present in each trial.
3. Calculate the moles of oxygen generated in each trial.
4. Convert the change in volume from Activity 1 to liters.
5. Convert the air temperature from Activity 1 from Celsius to Kelvin.
6. Using a computer graphing program, create an XY scatter plot of the moles of oxygen vs. volume of gas generated.
7. Create a best-fit line on your graph. Set the y-intercept equal to zero and display the equation on the graph.
8. Calculate the gas constant **R** from the equation of your graph; assume that the pressure = 1.00 atm.
9. Calculate the percentage error of the calculated **R** value with the actual value.

## Disposal and Cleanup

1. Dispose of solutions down the drain with the water running. Allow the faucet to run a few minutes to dilute the solutions.
2. Rinse and dry the lab equipment and return the materials to your equipment kit.
3. Sanitize the workspace.

**Data Table 1.**

| | Trial 1 | Trial 2 | Trial 3 | Trial 4 | Trial 5 | Trial 6 | Trial 7 | Trial 8 |
|---|---|---|---|---|---|---|---|---|
| Air Temperature | | | | | | | | |
| Volume $H_2O_2$ Liquid (mL) | | | | | | | | |
| Initial Volume Gas (mL) | | | | | | | | |
| Final Volume Gas (mL) | | | | | | | | |
| $\Delta V$ (mL) | | | | | | | | |

**Data Table 2.**

| | Trial 1 | Trial 2 | Trial 3 | Trial 4 | Trial 5 | Trial 6 | Trial 7 | Trial 8 |
|---|---|---|---|---|---|---|---|---|
| Concentration $H_2O_2$ | | | | | | | | |
| Moles $H_2O_2$ | | | | | | | | |
| Moles $O_2$ | | | | | | | | |
| $\Delta V$ (L) | | | | | | | | |
| Air Temperature (K) | | | | | | | | |
| Air Pressure (atm) | | | | | | | | |
| Equation of the Line | | | | | | | | |
| Gas Constant R | | | | | | | | |
| Percent Error | | | | | | | | |

# NOTES

CHEMISTRY
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