# Lab 9

Using the Ideal Gas Law

# LAB 9

## Using the Ideal Gas Law

## Learning Objectives

- Determine the relationship between pressure and temperature

- Understand how to use Charles's Law

- Understand how to use the Ideal Gas Law

## INTRODUCTION

Earth, fire, water, and air (Figure 1) made up the known elements in the ancient world that philosophers such as Aristotle and Plato observed. Aristotle believed that different combinations of these elements filled all of space, and that air, being a light substance, moved outward away from the universe's center. Today, scientists know that Earth's gravity attracts particles and gases in the air toward the planet, making the air denser as it approaches the surface.

## PROPERTIES OF GASES

Isaac Beeckman (17th century) was one of the first people to realize that air had real substance. He speculated that air had weight, pressed down on objects on Earth, and was expandable. In this way, gases have several unique properties. For example, a gas expands spontaneously to fill its container, and the volume of a gas equals the volume of the container in which it is held. Gas is also highly compressible. When pressure is applied to a gas, its volume readily decreases. To further understand the behavior of gases, it is important to discuss the properties of temperature, pressure, and volume.
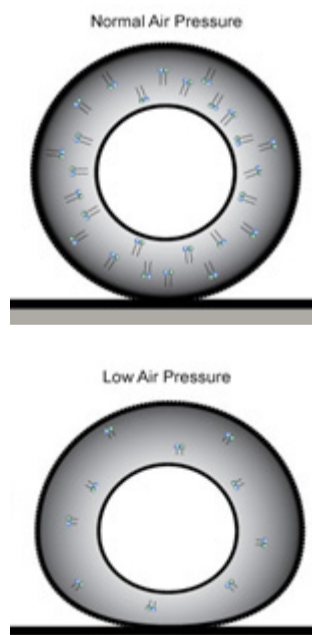


**Figure 1:** Air consists primarily of nitrogen (78%) and oxygen (21%).



**Figure 2:** When temperature decreases, molecules move slower. This decrease in velocity reduces the force applied to the tire, causing a less inflated tire.

### TEMPERATURE

**Temperature (T)** is the property of matter that reflects the kinetic energy of the particles. There are several standardized scales used to measure this value (e.g., Kelvin, Celsius, Fahrenheit). An increase in temperature means an increase in the motion of atoms in a material (Figure 2). Gas particles move past one another with little restriction; this is why gases occupy an entire container or space, while liquids and solids have a confined volume and surface area.

### PRESSURE

**Pressure (P)** is the amount of force that gas molecules exert on the area of space they fill. If you have ever had low tire pressure due to cold weather, it was because the gas molecules that fill the tire were moving slower. Therefore, they did not exert as much force on the inside of the tire as they normally would (Figure 2). Changing a container's volume can also change pressure.

### VOLUME

**Volume (V)** is the amount of space a gas occupies. Pressure decreases when volume increases because the molecules have more space in which they can travel. Pressure increases when volume decreases because the molecules have less space in which they can travel (Figure 3).

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## ATMOSPHERE

Atmosphere refers to the gases surrounding a star or planet held in place by gravity. Atmosphere is also a **unit of pressure**. One atmosphere is defined as the pressure caused by the weight of air above a given point. Normal atmospheric pressure at sea level is 14.7 pounds per square inch (lb/in²).  In chemistry and in various industries, the standard pressure that is used is 1 atmosphere (**1 atm**).



**Figure 3:** When volume decreases, the pressure increases. Conversely, when volume increases, the pressure decreases. This is what is known as an inverse relationship and is demonstrated in Boyle's Law (Figure 5).

## GAS LAWS

There are three fundamental gas laws that describe the relationship between pressure, temperature, volume, and the amount of gas. **Charles's Law** (Figure 4) tells us that the volume of gas increases as the temperature increases. **Boyle's Law** (Figure 3) tells us that the volume of gas increases as the pressure decreases. Finally, **Avogadro's Law** tell us that the volume of gas increases as the amount of gas increases. The **Ideal Gas Law** is the combination of these three gas laws.
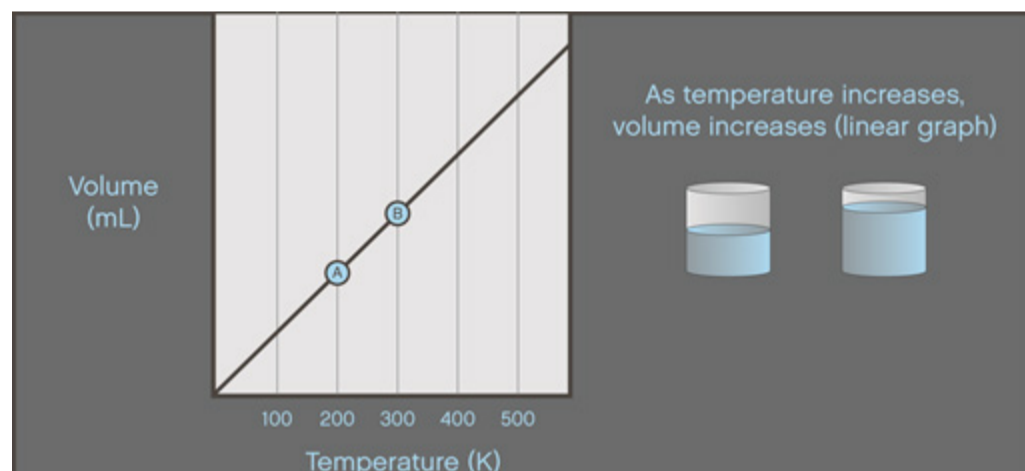


**Figure 4:** Point A on the graph represents the initial volume ($V_1$), whereas Point B represents the final volume ($V_2$).  Notice that as the temperature increases, the gas expands, thus increasing its volume.

## CHARLES'S LAW

**Charles's Law** was developed in the 19th century by Jacques Charles, a French chemist. Charles investigated the variation of volume with temperature for a fixed mass of gas at constant pressure. He found that as temperature decreases, the volume of a gas also decreases (Figure 4). Further, as temperature increases, the volume of a gas also increases. The two variables are directly proportional. The mathematical expression of this is:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad or \quad V_1 T_2 = V_2 T_1$$

An example of how to use Charles's Law is shown on the next page. Jacques Charles is also credited with discovering that all gas law equations must be calculated in Kelvin. **Kelvin** (K) is a temperature scale designed so that zero K is defined as **absolute zero.** Absolute zero is the temperature in which all molecular movement stops. To convert from degrees Celsius to Kelvin, simply add 273:

$$K = °C + 273$$

Note that you do not use the degree symbol when temperature is expressed in Kelvin.

# Using the Ideal Gas Law

## EXAMPLE USING CHARLES'S LAW:

The volume of a gas sample is 746 mL at 20°C. What is its volume at body temperature (37°C)? Assume the pressure remains constant.

### Step 1:

Using Charles's Law formula, write down all of the known values of the variables.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad or \quad V_1 T_2 = V_2 T_1$$

**Initial Volume** ($V_1$) = 746 mL
**Initial Temperature** ($T_1$) = 20 °C
**Final Temperature** ($T_2$) = 37 °C
**Final Volume** ($V_2$) = unknown variable to solve for

### Step 2:

Unit conversion.

$$T_1 = 20°C + 273 = 293\ K$$

$$T_2 = 37°C + 273 = 310\ K$$

### Step 3:

Convert the formula to solve for the desired variable, plug in the known quantities, and calculate the unknown variable.

$$V_2 = V_1 \frac{T_2}{T_1}$$

$$V_2 = 746\ mL \times \frac{310\ K}{293\ K}$$

$$V_2 = 786\ mL$$

### Step 4:

Check the logic of the answer. Is the answer reasonable? The final volume is larger than the initial volume. Since the temperature increased from 293 K to 310 K, it is reasonable that the volume of the gas also increased.

# Using the Ideal Gas Law

## BOYLE'S LAW

**Boyle's Law** is the gas law that relates pressure and volume. Boyle's Law states that when temperature and amount of chemical are held constant, pressure and volume are inversely proportional. In other words, as the pressure increases, the volume decreases (Figure 5). This is mathematically expressed as:

$$P_1V_1 = P_2V_2$$

Boyle's Law also states that the product of pressure and volume is a constant. This constant is referred to as **k**:

$$PV = k$$

For example, suppose an experiment begins with a pressure of 1 atmosphere (atm) and a volume of 10 L. In this case, k is equal to 1 atm x 10 L = 10 atm x 1 L. As the experiment proceeds, the pressure increases to 2 atm and the volume decreases to 5 L. As this happens, k will remain the same because 2 atm x 5 L is still equal to 10 atm x 1 L. Therefore, $P_1V_1 = P_2V_2$ is just another way of expressing k = k.

## AVOGADRO'S LAW

**Avogadro's Law** describes the relationship between volume (V) and the amount of gas (n) (measured in moles) when pressure and temperature are held constant (Figure 6). For example, 1 L of carbon dioxide will have just as many molecules in it as a liter of nitrogen. The mathematical expression of this is:

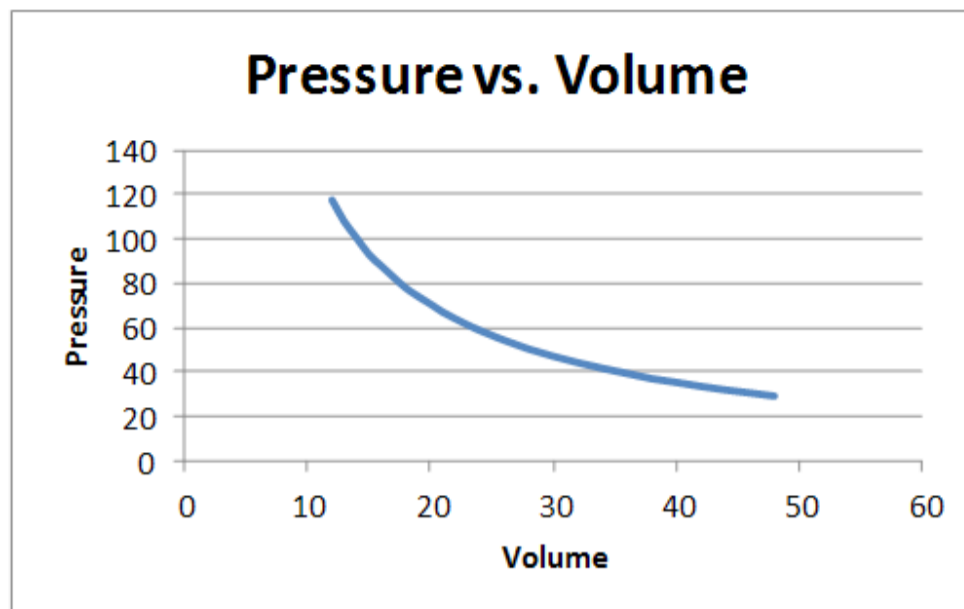$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$



**Figure 5:** The relationship between pressure and volume of an ideal gas. With a constant temperature and number of moles, pressure will decrease as volume increases.
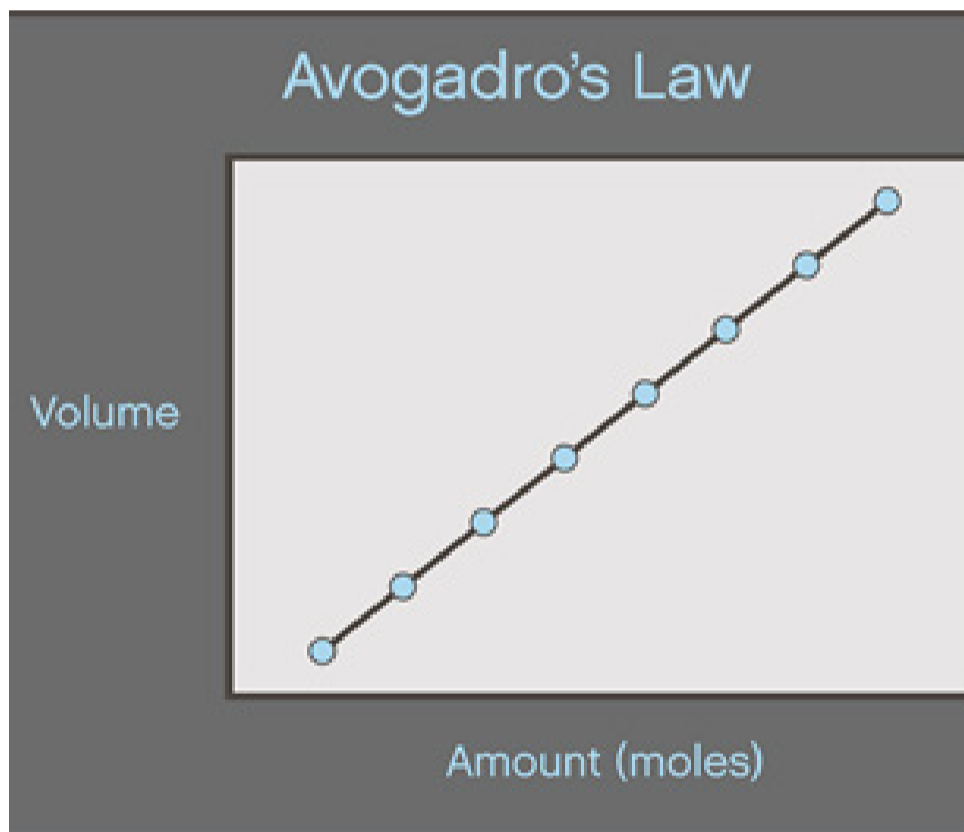


**Figure 6:** There is a direct relationship between volume and amount of a gas. As the volume increases, so will the number of moles of the gas.
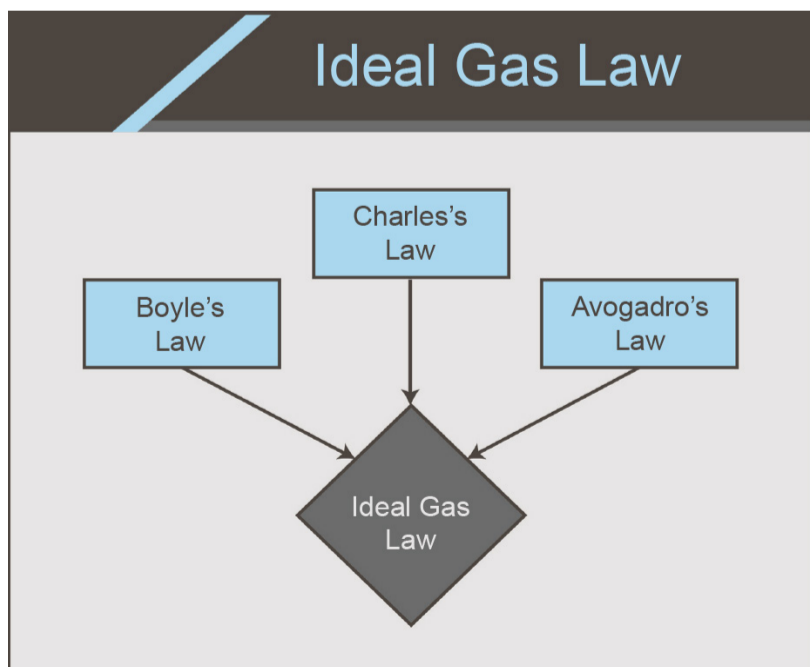


**Figure 7:** An ideal gas is an idealized model of a real gas.

## IDEAL GAS LAW

The Ideal Gas Law is a combination of all the simple gas laws (Figure 7). The Ideal Gas Law describes the behavior of real gases as long as the pressure and temperature are not too extreme. The Ideal Gas Law is very valuable when dealing with gases since it establishes a relationship between temperature, pressure, volume, and amount of a gas. The behavior of an ideal gas—that is, the relationship of pressure (P), volume (V), and temperature (T)—can be mathematically expressed as:

$$PV = nRT$$

Where:

- **P** = the gas pressure in atmospheres (atm)
- **V** = the volume of the gas in liters (L)
- **n** = the number of moles (mol) of the gas
- **R** = the constant value of 0.0821 L x (atm/mol x K)
- **T** = the temperature of the gas in Kelvin (K)

# Using the Ideal Gas Law

## EXAMPLE USING THE IDEAL GAS LAW:

5.0 g of neon is at 256 mm Hg and at a temperature of 35 °C. What is the volume?

### Step 1:

Using the Ideal Gas Law formula, and write down all of the known values of the variables.

$$P = 256 \ mm \ Hg$$

$$V = unknown$$

$$n = 5.0g$$

$$R = 0.0820574 \ L \cdot atm \cdot mol^{-1} \cdot K^{-1}$$

$$T = 35°C$$

### Step 2:

Unit conversion (round up).

$$P = 256 \ mm \ Hg \times \frac{1 \ atm}{760 \ mm \ Hg} = 0.3368 \ atm$$

$$n = 5.0 \ g \ Ne \times \frac{1 \ mol}{20.1797 \ g} = 0.25 \ mol$$

$$T = 35°C + 273 = 308 \ K$$

### Step 3:

Convert the formula to solve for the desired variable, plug in the known quantities, and calculate the unknown variable (round up).

$$V = \frac{nRT}{P}$$

$$V = \frac{0.25 \ mol \times 0.08206 \frac{L \ atm}{K \ mol} \times 308 \ K}{0.3368 \ atm} = 19 \ L$$

### Step 4:

Check the logic of the answer. Imagine 9.5 two-liter bottles of a soft drink stacked on top of each other. That is the volume (19 L) that was calculated.

# Using the Ideal Gas Law

## HYDROGEN PEROXIDE

Now that we know how to use the Ideal Gas Law, let us take a look at a couple of real-world reagents that can be used in the application of the Ideal Gas Law. **Hydrogen peroxide** ($H_2O_2$) is a chemical used in industry and research for its oxidative properties and in medicine as an antiseptic. You may have a bottle of 3% hydrogen peroxide in your medicine cabinet. At room temperature, a water solution of hydrogen peroxide ($H_2O_2$) will undergo an extremely slow decomposition reaction to form molecular oxygen gas ($O_2$) and liquid water ($H_2O$). Since hydrogen peroxide forms oxygen gas when it decomposes, we can use the Ideal Gas Law to observe this decomposition reaction. However, since the reaction will decompose very slowly (weeks) we will use a catalyst.



**Figure 8.** Catalase catalyzes the decomposition of hydrogen peroxide to water and oxygen.

## CATALASE

A **catalyst** is a substance that speeds up a reaction without being part of the reaction. Yeast cells contain a catalyst called **catalase.** Catalase speeds up the decomposition of hydrogen peroxide to water and oxygen (Figure 8). Catalase is very effective at decomposing hydrogen peroxide. In fact, one molecule of the enzyme can catalyze the conversion of more than 6,000,000 hydrogen peroxide molecules into water and oxygen every second.
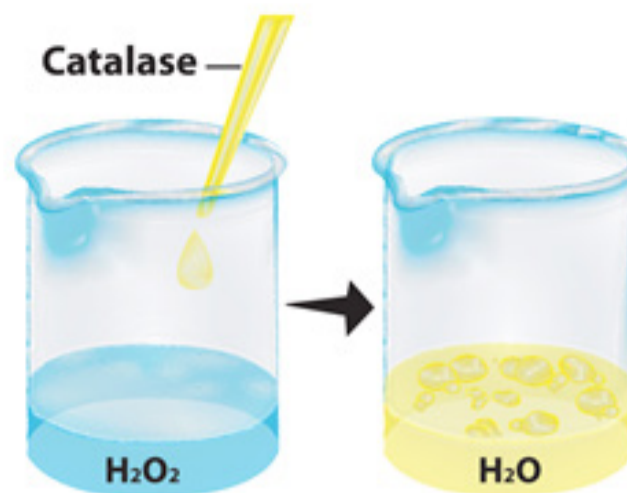
# Experiment 1

Charles's Law

## Experiment Inventory

### Materials

(1) 8 oz. Styrofoam® Cup

Graph Paper

**\*Ice**

**\*Pencil**

**\*Tap Water**

**Note: You must provide the materials listed in \*red.**

### Labware

(1) 10 mL Sealable Syringe with Cap

(1) Thermometer

(1) Ruler

**\*Timer (stopwatch, clock, internet, etc.)**

## EXPERIMENT 1: CHARLES'S LAW

In this experiment, you will explore the relationship between temperature and volume, and connect this to Charles's Law.

## PROCEDURE

1. Fill the syringe halfway with air, and seal the syringe by screwing the cap on. Record the volume (mL) of gas in the cylinder in Table 1 on the Experiment 1 Data Sheet.

2. Using a timer, hold the top of the thermometer up in the air for 30 seconds. Record the temperature of the room in Table 1 on the Data Sheet.

3. Fill one Styrofoam® cup with hot tap water. Place the thermometer in the cup.

   **LAB SAFETY:** *Avoid using extremely hot water. Handle hot water carefully.*

4. Fully submerge and hold the air-filled syringe in the hot water for 30 seconds. Use the timer to monitor time.

5. Gently push the piston down until the plunger cannot move any farther. Observe the new volume inside of the syringe. Record this volume and the water temperature in Table 1 on the Data Sheet.

6. Prepare a cup of water mixed with ice. Repeat Steps 4 and 5 for the cold water bath. Record the volume and temperature in Table 1 on the Data Sheet.

7. Convert all measured temperatures from degrees Celsius to Kelvin, and record your results in Table 1 on the Data Sheet.

### Table 1: Temperature and Volume Data

| Temperature Conditions | Temperature (ºC) | Temperature (K) | Volume (mL) |
|---|---|---|---|
| Room Temperature | | | |
| Hot Water | | | |
| Ice Water | | | |

## Experiment Inventory

### Materials

10 mL Hydrogen Peroxide ($H_2O_2$)

Yeast Packet

Paper Clip

Rubber Band

*Clear Tape

*Internet Access

*100 mL Warm Tap Water

*500 mL Tap Water

### Labware

(1) 10 mL Graduated Cylinder

(1) 100 mL Graduated Cylinder

(1) 250 mL Beaker

(1) 250 mL Erlenmeyer Flask

(1) 500 mL Beaker

(1) 24 in. Flexible Tubing

(1) 3 in. Rigid Tubing

(2) Pipettes

(1) Glass Stir Rod

(1) Thermometer

(1) Stopper with 1-hole

*Timer (stopwatch, clock, internet, etc.)

Note: You must provide the materials listed in *red.

## EXPERIMENT 2: USING THE IDEAL GAS LAW

In this experiment, you will use a catalyst to observe a reaction that obeys the Ideal Gas Law.

## PROCEDURE

1. Prepare the materials for the apparatus as shown in Figure 9. Insert the smaller rigid tubing into one end of the larger, flexible tubing. Insert the free end of the rigid tubing securely into the rubber stopper hole.

   Caution: *Be careful when working with rigid tubing. This tubing is made of glass and can break if excess force is applied.*
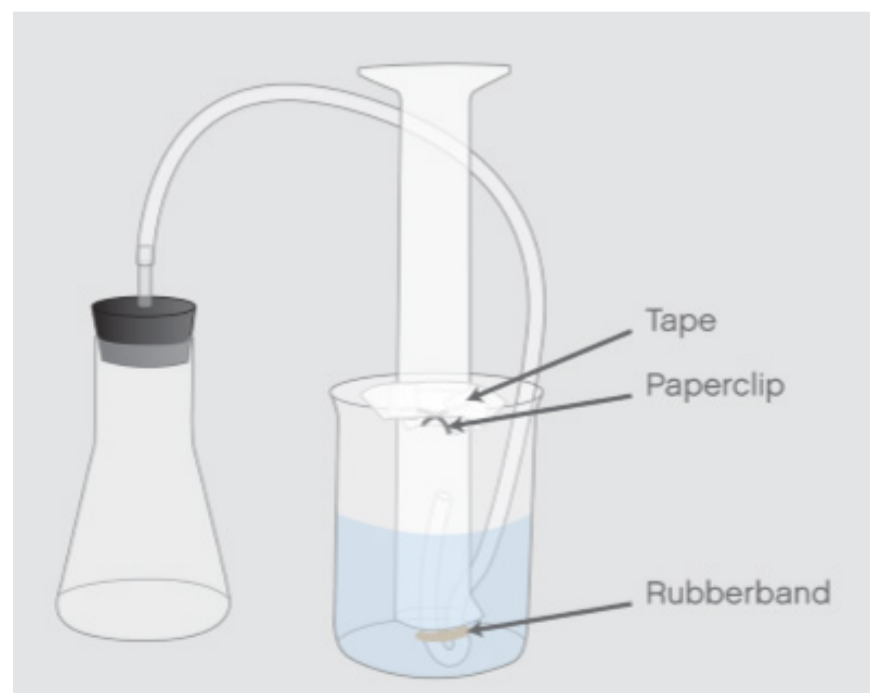


Figure 9: Setup of the gas-collection apparatus.

- [ ] 2. Bend the free end of the flexible tubing into a U-shape, and use a rubber band to hold this shape in place. This will allow you to more easily insert the end of the flexible tubing into the inverted graduated cylinder. Make sure the tubing is not pinched and that gas can flow freely through it.
- [ ] 3. Fill the 100 mL graduated cylinder with tap water to the 100 mL mark.
- [ ] 4. Fill the 500 mL beaker with 400 mL of tap water.
- [ ] 5. Using the thermometer, measure the temperature of the water in the 500 mL beaker and record it in Table 2 on the Experiment 2 Data Sheet.
- [ ] 6. Use the internet to determine the barometric pressure in the room, and record it in Table 2 on the Data Sheet.

    **Hint:** If necessary, use the regional pressure as a close substitute to the room pressure; this can easily be found online. You may need to convert this value from inches of mercury (inHg) or bars (bar) to atmospheres (atm). For reference, 1 inHg = 0.033 atm and 1 bar = 0.987 atm.

- [ ] 7. Mix 100 mL of warm water (45 °C) and one packet of baker's yeast in a 250 mL beaker. The warm water will activate the yeast from the dormant (dry) state. Be sure to mix the solution well with a glass stir rod until the yeast is completely dissolved.

    **Hint:** The water must be between 42–47°C, otherwise, you will kill the yeast (i.e., no activation).

- [ ] 8. Use a 10 mL graduated cylinder and pipette to measure out 5 mL of hydrogen peroxide ($H_2O_2$). Pour the hydrogen peroxide into the Erlenmeyer flask, and securely place the stopper with the attached stopper tube into the top of the Erlenmeyer flask.
- [ ] 9. Clean the 10 mL graduated cylinder by rinsing it at least three times with tap water. Dispose of the rinse down the drain.
- [ ] 10. Cover the opening of the 100 mL graduated cylinder with two or three fingers and quickly turn it upside down into the 500 mL beaker already containing 400 mL of water. DO NOT remove your fingers from the opening until the graduated cylinder is fully submerged under the water. If the amount of trapped air exceeds 50 mL, refill the cylinder and try again. You want as little air as possible to be in the graduated cylinder.
- [ ] 11. Insert the U-shaped flexible tubing into the beaker, and carefully snake it into the submerged opening of the graduated cylinder.
- [ ] 12. Secure the graduated cylinder to the beaker by bending a paper clip around the graduated cylinder and using a piece of clear tape to hold it in place.
- [ ] 13. With the cylinder vertical, record the volume of air inside (the water line) in Table 2 on the Data Sheet.
- [ ] 14. Using a pipette, measure out 5 mL of yeast solution into the rinsed 10 mL graduated cylinder.
- [ ] 15. Remove the stopper (still connected to the hose) from the Erlenmeyer flask. Get your timer ready.
- [ ] 16. Quickly pour the 5 mL of yeast solution into the Erlenmeyer flask. Immediately place the stopper securely in the opening of the Erlenmeyer flask by gently twisting it down into the flask.
- [ ] 17. Start timing the reaction with the your timer as soon as you seal the flask with the stopper. If necessary (i.e., if you are using a clock or a watch), record the start time of the reaction in Table 3 on the Experiment 2 Data Sheet.
- [ ] 18. Carefully swirl the Erlenmeyer flask to mix the two solutions together. Bubbles should soon begin to form in the 100 mL graduated cylinder.

    **Hint:** If gas bubbles are not immediately visible, make sure the stopper is on tight and the tubing is not leaking.

19. Continue to swirl the Erlenmeyer flask, and let the reaction run until no more bubbles form (to assure the reaction has gone to completion).

> **Hint:** Catalase in the yeast works best around the temperature of the human body. You can speed up the reaction by warming the Erlenmeyer flask with your hands.

20. Record the time when the reaction finished, and then the total reaction time in Table 3 on the Data Sheet. Also record the final volume of air in Table 2. Remember to read the flask content at eye-level and measure from the bottom of the meniscus.

21. Pour all other liquids down the drain, and clean the labware.

## Table 2: Temperature, Pressure, and Volume Data

| Temperature of Tap Water (°C) | Room (or regional) Pressure (atm) | Initial Volume of Air (mL) | Final Volume of Air (after reaction) (mL) | Volume of Oxygen Collected (Final Volume - Initial Volume) |
|---|---|---|---|---|
|  |  |  |  |  |

## Table 3: Reaction Time Data

| Time Reaction Started | Time Reaction Ended | Total Reaction Time |
|---|---|---|
|  |  |  |