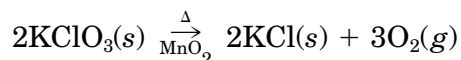
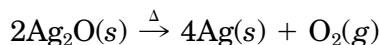


# 1. The Decomposition of Potassium Chlorate

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## Introduction

Small quantities of molecular oxygen ( $O_2$ ) can be obtained from the thermal decomposition of various oxygen-containing compounds. Some examples of these reactions are



The last reaction includes manganese dioxide ( $MnO_2$ ) as a catalyst. This reaction is particularly important in this experiment.

## Purpose

You will study the thermal decomposition of  $KClO_3$  (potassium chlorate) in the absence of a catalyst. You will be able to identify the solid that remains after the decomposition from the quantity of oxygen that is evolved. You will verify this identification by comparing the measured mass of the solid product with a calculated value. Will the solid be  $KCl$ , or will the absence of a catalyst affect the identity of the solid?

## A word about catalysts

A catalyst is a substance that causes an increase in the rate of a chemical reaction without being used up in the reaction. Moreover, a catalyst does not *usually* alter the end result of a reaction. The same products are usually obtained in the presence or the absence of a catalyst.

## Concept of the experiment

The solid product that results from the thermal decomposition of  $KClO_3$  is  $KCl$  when the catalyst  $MnO_2$  is present. What will happen if the catalyst is not added? Will the loss of oxygen be less extensive? If so, the solid product will be either  $KClO_2$  (potassium chlorite) or  $KClO$  (potassium hypochlorite). How does this experiment answer these questions?

You will heat a sample of  $KClO_3$  of known mass in the absence of a catalyst until the evolution of oxygen is complete. Oxygen will be collected in a flask by the displacement of water (Ebbing/Gammon, Section 5.5). In order to determine the correct stoichiometry of this reaction, you will need to obtain the number of moles of  $O_2$  that have been evolved. You can calculate this quantity from the rearranged form of the ideal gas law,  $n = PV/RT$ . Clearly, you will need to measure the pressure, volume, and temperature of the oxygen.

Because the oxygen is collected by displacing water, water vapor will also be present in the gas. The experiment is designed so that the total pressure of the oxygen and water vapor will be equal to the atmospheric pressure; you can easily measure this quantity with a barometer. The partial pressure of oxygen in the flask can be calculated from the total pressure and the vapor pressure of water. Table 1.1 gives the vapor pressure of water at various temperatures.

Figure 1.1 shows a suitable apparatus for this experiment. The sample of  $\text{KClO}_3$  is placed in the test tube and the flask is filled with water. Some of the water is displaced by oxygen and is pushed into the beaker. The volume of water in the beaker will be identical to the volume of oxygen in the flask.

## Procedure

### *Getting started*

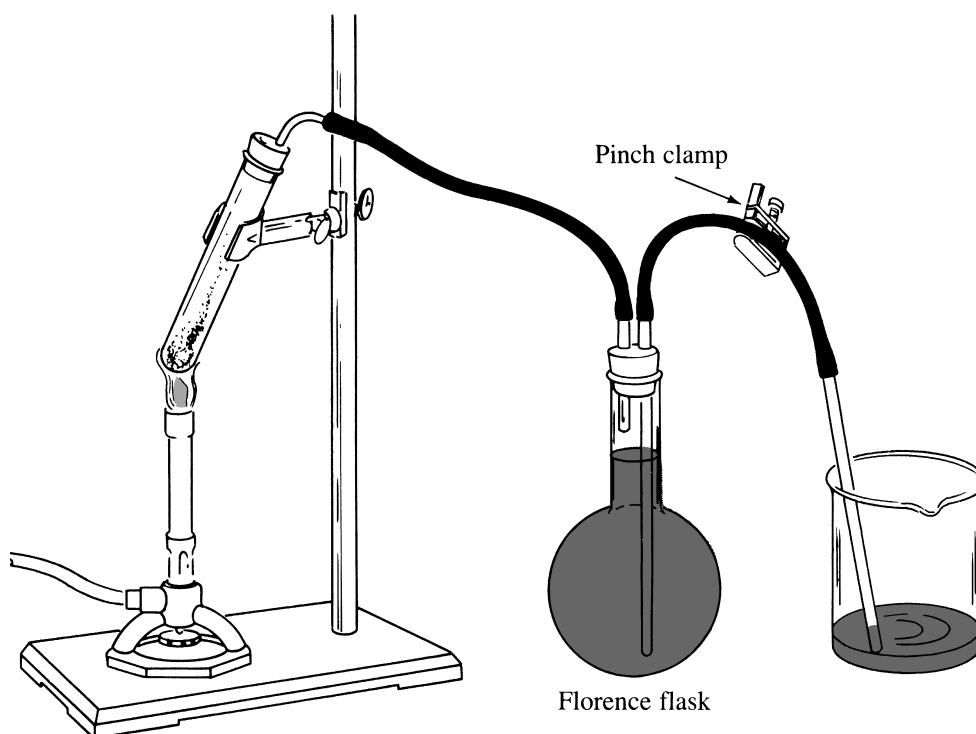
1. Your laboratory instructor may ask you to work with a partner.
2. Obtain and clean the glassware for the apparatus in this experiment.
3. Obtain a wooden splint (a small piece of wood).
4. Obtain and record the atmospheric pressure.
5. Obtain directions for discarding the solid product that will be produced in this experiment.
6. *Without fail*, observe the following safety precaution throughout this experiment:

**CAUTION:  $\text{KClO}_3$  is a very strong oxidizing agent. Do not under any circumstances let this substance contact paper or the rubber stopper in the test tube belonging to the apparatus. A fire or explosion can occur.**

### *Setting up the apparatus*

1. Assemble the apparatus as shown in Figure 1.1.
2. Test the apparatus for leaks, using the method in Steps 3 through 6. Remember that a leak will ruin your results.
3. Fill the flask almost completely with water and open the pinch clamp.
4. Remove the test tube from its stopper. Blow through the glass tube until the rubber tube is filled with water. Water will begin to flow from the flask to the beaker.
5. Replace the test tube while the water is flowing.
6. If no leak is present, no more water will flow and the water level in the flask will no longer change. If a leak is present, water will continue to flow.
7. If a leak is indicated by this test, it is almost certainly located around one of the rubber stoppers. Press each one firmly into place. If water continues to flow, ask your laboratory instructor for assistance.
8. Close the pinch clamp. Water should stay in the rubber tube. Do not drain it or try to shake it out.

FIGURE 1.1  
A suitable  
apparatus for  
the thermal  
decomposition  
of  $\text{KClO}_3$ .



### *Doing the experiment*

1. Make sure that the test tube is clean and dry. Measure and record its mass to the nearest 0.001 g (or any other precision stipulated by your laboratory instructor).
2. Your laboratory instructor will put about 2.0 g of  $\text{KClO}_3$  into the test tube.
3. Measure and record the mass of the test tube and its contents with the same precision as in Step 1. Calculate and record the mass of  $\text{KClO}_3$ .
4. Do not connect the test tube to the apparatus yet.
5. Fill the beaker with water. Open the clamp and raise the beaker above the flask. Water will siphon back into the flask. Allow the water in the flask to rise until it is about 1 inch below the glass tube leading to the test tube. Close the pinch clamp.
6. Connect the test tube containing  $\text{KClO}_3$  to its rubber stopper. None of the solid material should touch the rubber stopper or even be near it.
7. Equalize the pressure in the apparatus and atmospheric pressure. To do so, open the pinch clamp and adjust the height of the beaker until the water levels in the flask and in the beaker are identical. Close the clamp.
8. Empty the beaker but do not dry it. The volume of the water drops that remain in the beaker will be roughly equal to the volume that will remain after the displaced water is poured into a graduated cylinder for measurement. Do not empty or shake out the rubber tube when you empty the beaker.

9. Place the tube back in the beaker and *open the clamp*.

**CAUTION: If the clamp is not opened at this point, the build-up of gas could cause an explosion, although it is more likely that a stopper would be forced to loosen.**

10. Use your laboratory burner to heat the test tube. Be cautious at first and brush the flame over the test tube. After a few minutes, you can heat the test tube more strongly. The solid will melt, oxygen will be evolved, and water from the flask will be displaced into the beaker.
11. Heat until no more gas is evolved. The contents of the test tube may solidify, because the melting point of the product is greater than that of  $\text{KClO}_3$ .
12. Remove the flame and allow the system to come back to room temperature. About 15–20 min will be required.
13. Equalize the pressure, using the same procedure as in Step 7. The total pressure of the gas (oxygen plus water vapor) will then be equal to the atmospheric pressure.
14. With the clamp closed, remove the tube from the beaker. Measure the volume of water in the beaker by pouring it into a 1-L graduated cylinder. Record the volume.
15. Measure the temperature of the water to the nearest degree. This temperature will be taken as the temperature of the gas. You should also use it to determine the appropriate vapor pressure of water from Table 1.1.

**Table 1.1 Vapor Pressure of Water**

| Temperature<br>(°C) | Vapor Pressure<br>(mmHg) | Temperature<br>(°C) | Vapor Pressure<br>(mmHg) |
|---------------------|--------------------------|---------------------|--------------------------|
| 20                  | 17.5                     | 25                  | 23.7                     |
| 21                  | 18.6                     | 26                  | 25.2                     |
| 22                  | 19.8                     | 27                  | 26.7                     |
| 23                  | 21.1                     | 28                  | 28.3                     |
| 24                  | 22.4                     | 29                  | 30.0                     |

16. Obtain the mass of the test tube and its contents. Calculate and record the mass of the product.
17. Ignite the wooden splint, let it burn for a short time, and blow out the flame. A glowing ember should remain. Thrust the glowing end of the splint into the flask and observe the results. Record your observation.
18. If time permits, repeat Steps 1 through 16 with a clean test tube and a fresh sample of  $\text{KClO}_3$ .

**CAUTION: Before you leave the laboratory, make sure that your gas outlet and those of your neighbors are closed.**





# *The Decomposition of Potassium Chlorate*

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Date: ..... Student name: .....  
Course: ..... Team members: .....  
Section: .....  
Instructor: .....

## **Results**

Atmospheric pressure: ..... mmHg

| <b>Trial</b>                                | <b>1</b> | <b>2</b> |
|---|----------|----------|
| Mass of test tube + $\text{KClO}_3$ (g)     | .....    | .....    |
| Mass of empty test tube (g)                 | .....    | .....    |
| Mass of $\text{KClO}_3$ (g)                 | .....    | .....    |
| Volume of water displaced (mL)              | .....    | .....    |
| Temperature of water ( $^{\circ}\text{C}$ ) | .....    | .....    |
| Mass of test tube + product (g)             | .....    | .....    |
| Mass of empty test tube (g)                 | .....    | .....    |
| Mass of product (g)                         | .....    | .....    |
| Observation with the glowing splint:        |          |          |

## Questions

1. Using the data that you have accumulated, complete the following table. Show your calculations.

| <b>Trial</b>  | <b>1</b> | <b>2</b> |
|---|----------|----------|
| Pressure of gas (mmHg)                              | .....    | .....    |
| Vapor pressure of water (mmHg)                      | .....    | .....    |
| Pressure of O <sub>2</sub> (mmHg)                   | .....    | .....    |
| Volume of O <sub>2</sub> (mL)                       | .....    | .....    |
| Temperature of O <sub>2</sub> (°C)                  | .....    | .....    |
| Moles of O <sub>2</sub>                             | .....    | .....    |
| Moles of KClO <sub>3</sub> taken originally         | .....    | .....    |
| Moles of O <sub>2</sub> /mol KClO <sub>3</sub>      | .....    | .....    |
| Mean moles of O <sub>2</sub> /mol KClO <sub>3</sub> | .....    |          |

Calculations:

Student name: ..... Course/Section: ..... Date: .....

2. a. Write a chemical equation for each possible product in this experiment.

b. Using the masses of  $\text{KClO}_3$  that you took in this experiment, calculate the expected masses of each solid product from these equations.

3. a. What is the chemical equation for the reaction that occurred, based on the ratio mean moles of  $O_2$ /mol  $KClO_3$ ? Why? Is this conclusion in accord with the mass of the solid product that you obtained? Explain.
- b. What do you conclude about the effect of the absence of the catalyst on the outcome of the thermal decomposition of  $KClO_3$ ?
4. What does your observation with the glowing splint indicate about the nature and reactivity of oxygen?