

Introduction

Sodium chlorate is a compound composed of sodium, chlorine, and oxygen. When sodium chlorate is heated in the presence of a catalyst, all the oxygen is released from the compound as a gas, the other salt product will remain in the reaction container. The reaction must be catalyzed with manganese dioxide so that it will proceed at a reasonable rate. The gas will be collected to determine a percent yield.

A second method of processing the data will be done by calculating the % oxygen in the original compound by determining the mass lost from the original compound. This can be compared to the theoretical % of oxygen in the sodium chlorate compound.

A third method of processing the data will be done in an attempt to confirm the Law of Conservation of mass.

PreLAD: *Questions can be answered in the space below. Work must be presented clearly.*

1. Make a data/results table. It must include a line by line listing of all measurements to be made, and the subsequent calculation results.
 - Don't forget the processing the data and indicate how the calculations will be made using just letters (variables or words.)
2. Write out a balanced equation for the decomposition of sodium chlorate.
 - Is this a redox reaction? Write out the oxidation numbers to determine if any element (which one) is oxidized, and which if any is reduced.
3. Let's determine the maximum amount of sodium chlorate that is feasible to use when collecting gas in a 1L soda bottle.
 - a. Use the ideal gas law to calculate the number of moles of oxygen gas that can be collected in a 1.0 L container at 740 mmHg pressure and at 298°C.
 - b. Use the stoichiometry of the equation above to determine the mass of sodium chlorate that will produce the moles of gas just calculated in part a.
4. Our Bunsen burners will not heat objects much hotter than between 400–600°C. Do a search and report the melting temperatures of sodium chlorate and sodium chloride.

Materials – at each lab bench

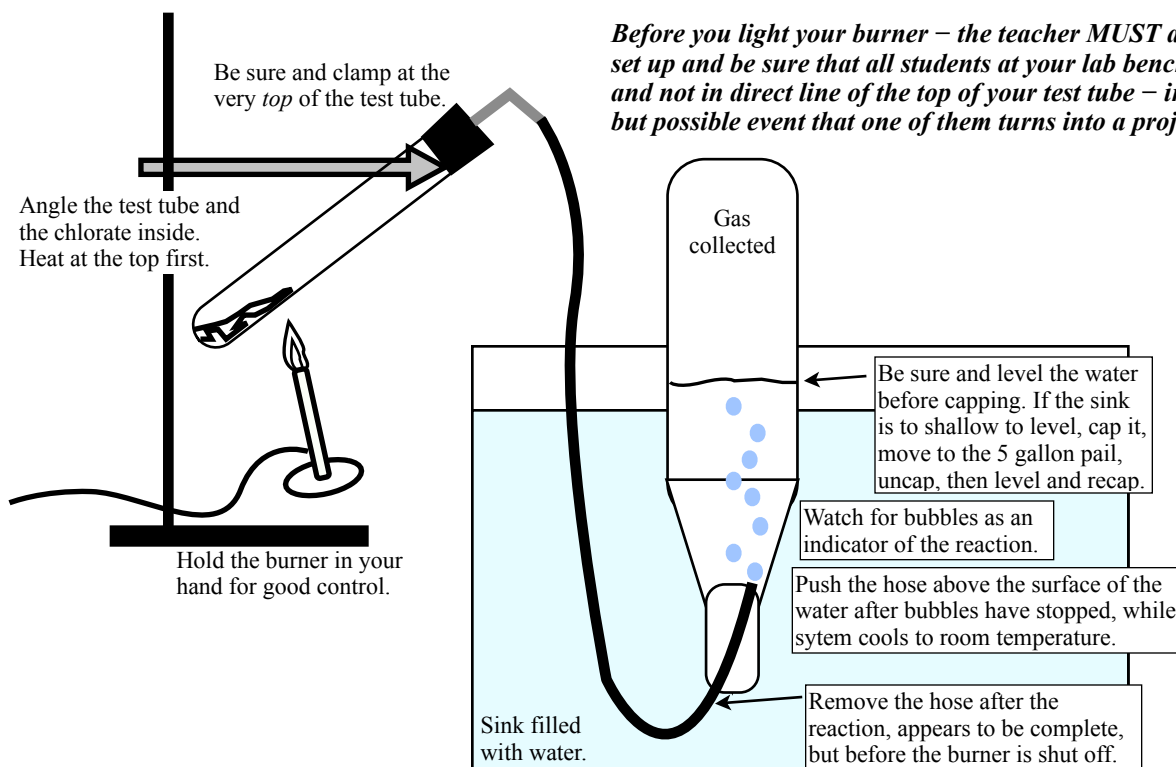
- 2x large test tubes
- 2x stopper with glass elbow
- 2x flexible tubing
- vial of sodium chlorate with scoop
- vial of manganese(IV) oxide with scoop
- 2x bunsen burners
- 2x ring stands and clamps
- 2x gas collecting bottle (1 L or 2L bottle)
- 2x 1000 ml graduated cylinder
- thermometer
- sink drain stop

Procedure *Goggles must be worn at all times when anywhere in the lab area.
You may take them off only if in the classroom section of the lab.*

Extra Warning: *This lab has extra hazards. The possibility of projectile rubber stoppers and the danger of any combustible material which would react in the same manner as demonstrated by the gummy bear sacrifice.*

- Note that there are two dimples in the balance pans that will keep the test tube from rolling off the pan. Record the mass of the empty PYREX test tube and the rubber stopper/glass elbow, tare, then use the metal scoop to add no more than 2 g of the sodium chlorate (recording the mass of the sodium chlorate exactly exactly). Be sure and wipe any scoops used so that you do not contaminate the chemicals. Then add approximately 1 g of manganese dioxide (recording the mass of it exactly). (*Report to the class data table.*) Shake/twirl the test tube gently to mix the two chemicals completely. Complete contact between the reactant and the catalyst is critical to successful results.
- Set up the apparatus as demonstrated in the diagram on the next page. Do not shove the tube all the way up into the 1 L collection bottle. It is helpful to be able to observe that gas is still being generated by the observation of bubbles. An added benefit, is that the gas will cool as it bubbles through the water. Do not light the Bunsen burner until your set-up is OK'd by the instructor and you have checked that everyone at your lab bench is wearing protective eyewear.
- Heat the test tube from the top of the solid mixture downward. Hold the Bunsen burner in your hand so you can move the flame around and gradually ease the sodium chlorate into melting. Do NOT heat only at the bottom of the tube - if you heat the bottom, you run the risk of vaporizing the oxygen underneath while unmelted solid chlorate sits on top of it. The oxygen will seek to expand about a thousand times its original volume (in the salt) and will surely drive anything above it out of the test tube. So heat gently from the top of the solid, moving the flame around and easing the edges of the chlorate into melting and decomposing.
- Once the chlorate has appeared to completely decompose, and no more bubbles of gas are being generated. (Check to see that the end of your hose is still below the surface of the water in the 1 L collection bottle.) The sodium chloride and the manganese (IV) dioxide will remain in the tube. These chemicals are no longer explosive and the test tube can be heated strongly for just a couple more minutes to drive out any remaining oxygen.
- When the reaction is complete, before turning off the burner, push the hose all the way into the 1 L collection bottle so the hose is above the surface of the water, then shut off the burner, and allow the entire apparatus to cool. While waiting for the apparatus to cool check your calculation plans with your lab-mates.

Schematic Diagram for Apparatus Set-up



- F. Lift or lower the bottle until the level of water inside the bottle is even with the level of water outside the bottle. If the sink is too shallow, and leveling is not possible, cap the bottle and carry it to the 5 gallon pail at the back of the lab where you can submerge, uncap, and re-level. When level, cap the bottle and lift out of the sink and turn upright. At this point you can do the glowing splint test to confirm the presence of oxygen. DO NOT LEAVE the reignited splint in the plastic bottle – it will melt. Measure the volume of water remaining in the bottle. Then refill the bottle and measure the total volume of water that fits in the bottle.
- G. When the test tube portion of apparatus is cool to the touch, disassemble and determine the mass of the test tube and the rubber stopper-glass elbow with the product and catalyst in it.
- H. Measure the temperature of the air in the classroom and the temperature of the water in the sink. We hope they are the same - if not, take an average to use in you gas law calculations. (*Report to the class data table.*)
- I. Look up the water vapor pressure for your lab temperature. (*Report to the class data table.*)

Disposal

Dump the remaining chemicals into the trash, then put the test tube back on the tray. Put all the other materials back on the tray.

Processing the Data – *Put these results and method of calculation on your data/results table.*

Method A - Percent Yield

1. Calculate the experimental volume of oxygen collected. (*Report to the class data table.*)
2. Calculate the moles of sodium chlorate that were decomposed.
3. Use the moles of sodium chlorate to calculate to the *theoretical* number of moles of oxygen that should be produced.
4. Remember that there were two gases in the collection bottle. Use the water vapor pressure table in the Pre LAD D.1 to calculate the pressure of the oxygen gas that you collected.
5. Use the *theoretical* moles of oxygen (previously calculated) and the ideal gas law to determine the *theoretical* volume of oxygen that should have been produced.
6. Using the *experimental* volume of oxygen measured and the *theoretical* volume of oxygen (previously calculated) calculate the percent yield of the *experimental* volume of oxygen gas produced.

Method B - Percent Oxygen in the Compound

7. Calculate mass of the NaCl remaining in the test tube. (*Report this to the class data table.*)
8. Use the mass of the NaCl remaining to calculate the *experimental* mass of oxygen removed from the sodium chlorate.
9. Calculate the *experimental* mass percent of oxygen in the sodium chlorate.
10. Knowing the chemical formula for sodium chlorate, calculate the *theoretical* mass percent of oxygen. Compare this to your experimental result.

Method 3 - Completing the Stoichiometric Picture, Confirming the Law of Conservation of Mass

11. Calculate the mass of sodium chloride remaining in the tube after the reaction.
12. Sum the mass of the products formed.

Post LAD Questions – these questions can be answered in the space provided.

1. How did you know when the reaction was done? Why was it helpful to keep the end of the hose under water and not stuck far up into the collection bottle?
2. As you started to heat the sodium chlorate, you may have seen some melt near the edges, but as you continue to heat, any observed melted material quickly turns to solid. Why is it that solid forms while heating with a hot flame - this seems counterintuitive?
3. It is important that after the reaction is completed, and before the Bunsen burner is turned off, the hose should be pulled out from the bottle and out of the sink of water. Why? (Hint: consider the effect of the gases in the test tube cooling down while the hose is still underwater.)
4. Comment on the following sources of error and how they would affect the percent yield of oxygen collected. Justify your answer by stating the error's effect on any measurements and following those effects through the resulting calculations. Be specific by stating if the measurements and calculations would be higher or lower as a result of the source of error. What if.....
 - a. some of the oxygen is still locked up in the crystals of the remaining solid in the test tube.
 - b. oxygen gas was quite soluble in water?
 - c. the temperature of the oxygen gas did not cool down to the water/room temperature that was measured?
 - d. You forgot to "correct for water vapor" and calculated the theoretical volume of "dry" oxygen gas.
5. Cooling the oxygen gas formed was important as you mentioned in your error analysis above, what procedural technique was used in this lab to help cool the oxygen gas quickly?
6. Calculate the theoretical mass percentage of oxygen in sodium chlorate. Compare this to your experimental value. Would the loss of some of the dusty manganese dioxide up out of the test tube have been a valid source of error for your data?