**Speeding Things Up At the Campground**

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While you are camping you and your friends go for a hike. The hike takes longer than you expected and it is past dinner time when you get back to your campsite. You are all very hungry after the hike but you still have to cook the food you brought before it can safely be eaten. To make the food cook faster one of your friends suggests turning up the flame on the camp stove. Another of your friends questions whether this will really make the food cook faster and thinks you would just be wasting fuel. While it seems reasonable that this would increase the temperature on the outside of the pan, it is clear from the discussion that follows that many of you have heard conflicting information about whether or not turning up the heat can be an effective method for getting your dinner ready faster. Fortunately, while you all are trying to decide the best course of action your food finishes cooking. You have been having a hard time getting chemistry off your mind. As you eat your dinner you wonder if there is anything related to the chemistry of cooking that could help you answer the question.

Later on you need to use the "facilities". The campground you are staying in does not have flush toilets. When the friend who made the arrangements for this trip first told you about the toilets, you were dreading using them because you have strong memories of smelly pit toilets from trips with your family. To your surprise the toilets here have very little odor. On this trip to the facilities, you read a sign on the wall that says that the low odor is due to the fact that they are enzymatic toilets. It goes on to explain that enzymatic toilets are designed so enzymes are always present which quickly decompose the waste before the odor causing bacteria can grow. As you read you wonder, could chemistry be playing a role here as well?

**Background**

Often chemists will want to speed up the rate of a reaction for a variety of reasons including getting it to fit a normal work schedule, avoiding decomposition of product over time, and freeing up equipment for the next reaction. Even our own bodies like to improve the favorability of many of the reactions we need to survive by increasing the reaction rate. While we can increase rate by increasing the concentration of the reactants as we saw last week, often this is not practical. This week we will explore whether or not there are other factors which can influence the reaction rate including changing the temperature and adding a compound that is not in the overall reaction.

To help us figure this out we are going to look at how the rate of decomposition hydrogen peroxide changes as we vary concentration and temperature. As hydrogen peroxide decomposes, it produces oxygen and water as seen in equation 1.

2 H2O2(aq) 2 H2O(l) + O2(g) (1)

As we have already seen, we can define our rate as either the decrease in concentration of a reactant (H2O2) or the increase in concentration of a product (H2O or O2). Measuring this poses a bit of a challenge. For H2O2, our reactant, it is difficult to measure changes in concentration directly without affecting the concentration. On the product side, H2O is the solvent for which concentration effectively does not change and O2 is a gas which escapes and so is not part of the aqueous solution. What we can measure is the increase in the volume of gas over time by trapping the oxygen produced. So how can we relate volume of O2 produced back to the decrease in H2O2 concentration?

$n = \frac{PV}{RT}$ (2)

Remembering back to last semester, once we have volume, if we also measure the pressure and temperature, the ideal gas law can be rearranged to solve for moles of oxygen (equation 2). Once we have the moles of O2, we can calculate the decrease in moles of H2O2 using the stoichiometric ratio between O2 and H2O2 given in equation 1 which is then subtracted from our initial moles. Since our solution volume does not change, we can get concentration of hydrogen peroxide by dividing the moles left by the volume of the solution. While we can get the rate for a single time point by dividing the change in concentration by the time it took (equation 3), it would be better to have an average.

$\frac{-∆\left[H\_{2}O\_{2}\right]}{∆s} = rate$ (3)

Once we begin collecting O2 from the decomposition reaction, we can record volume and time at regular intervals. Rather than calculating the rate at each interval, we are going to graph the calculated concentration of H2O2 on the y axis and time on the x axis (figure 1). After adding a trendline, the slope now has the same units as our rate and so represents the average rate. Note: because we have defined our rate as the decrease in concentration of H2O2, the slope is negative so we report the absolute value as the rate.

Hydrogen peroxide as a 3% aqueous solution is used for first aid and can be kept for a long time on store shelves in spite of the fact that it decomposes because the rate of decomposition is so slow. If we are going to be able to measure the rate of decomposition within our lab period, we need to speed the reaction up substantially. The first thing that we will try is adding iodide ions from KI to the reaction to see if the decomposition goes faster. In addition to varying the amount of H2O2 added, we will vary the amount of KI added and thus the amount of I−. If iodide has any effect on the rate, when we determine the rate law (equation 4), we should get an order with respect to I− that is greater than zero even though I− is not part of the overall reaction.

$\frac{-∆\left[H\_{2}O\_{2}\right]}{∆s}=k\left[H\_{2}O\_{2}\right]^{x}[I^{-}]^{y}$ (4)

The second thing we want to try is changing the temperature. Food is stored in a refrigerator to reduce the rate of spoiling. When we are cooking, we sometimes will cook things at a higher temperature to get it done faster. Since cooking food and spoiling food are chemical reactions, it stands to reason that temperature may play a role in the rate of a reaction. As we investigate the effect of changing temperatures on the rate, how will we know whether the change is due to temperature or the amount of KI and H2O2 present?

When good scientists want to isolate the effects from multiple sources on a single event, they establish controls so that only one variable changes at a time. To separate the effect of reactant concentration from the effect of changing temperatures, we will control the temperature for all of our reactions by placing them in a water bath. Keeping the water temperature constant for the first two solutions will allow us to isolate the effect of concentration as the amounts of KI and H2O2 are varied. To isolate the effect of temperature, we will repeat the reaction using the same concentrations of reactants as the first reaction but at a different temperature. Since all the concentrations are the same, any changes in rate will be due to the change in temperature.

Once we have determined the effects of concentration and temperature there is still one more interesting question we can try to answer. If iodide is part of the rate law, what happens to I− since it is not part of equation 1? Does it become something else in a secondary reaction or is it still there when the decomposition is complete? In order to find out, we are going to allow one of the reactions to go all the way to completion and then titrate the solution with silver nitrate (equation 5) to form silver iodide (a fine yellow precipitate).

Ag+(aq) + I−(aq) AgI(s) (5)

Since it is difficult to measure the amount of AgI formed and to know for sure when all the iodide has reacted, potassium chromate (K2CrO4) will be added as an indicator. When all the I− has been consumed, the chromate ion reacts with silver ions to form a red solution. Just as with other titrations you have done in the past, by measuring the volume of silver nitrate added at the point when the indicator permanently changes color and taking into account the concentration of AgNO3 and the stiochiometric ratio between Ag+ and I−, we can determine the number of moles of I− left in solution after the decomposition. This can be compared to the number of moles we started with to see if anything happened to the iodide.

**Procedure**

Part A: Equipment Set-Up (Figure 2)

1. Attach a 60 mL syringe to the tip of a 50 mL burette. With the stopcock open, add room temperature water to the burette until there is a small amount of water in the bottom of the syringe when it is held above the 15 mL mark on the burette.
2. Place the small stopper of the stopper-tube assembly in the top of the burette. Fill a 125 mL flask with d.i. water and stopper the flask with the other stopper. Do not let water get into the tubing.
3. Check for leaks by either raising or lowering the syringe as far as you can and holding it steady. If the water levels stop changing before they become even, there are no leaks. If leaks are found they need to be eliminated before proceeding.
4. Empty the flask and use it for the next part.

Part B: Concentration Effect

1. Fill a small container approximately halfway with room temperature water to create a water bath and place it on top of a magnetic stirplate.
2. Place the 0.1 M KI and water for solution 1 in the flask. Fill a graduated cylinder with the required amount of hydrogen peroxide.
3. Add a magnetic stirbar to the flask. Put the flask in water bath. Stir the solution with the magnetic stirrer at a medium speed for a few minutes to equilibrate the temperature in the flask with that of the water bath. Measure the temperature of the water bath.
4. Hold the syringe next to the burette and equalize the water levels (figure 3). There should only be a small amount of water in the syringe and the water level in the burette should be near the top.
5. While maintaining the speed on the magnetic stirrer, quickly pour in the H2O2 and stopper the flask.
6. Immediately equalize the water levels, read the initial volume in the burette and start the timer.
7. Move the syringe to keep the water levels equal as O2 is produced to keep the pressure constant. Record the new volume and the time at 2 mL intervals until 14 mL of oxygen gas has been produced.
8. Measure the temperature of the bath again. Discard the solution and clean out the flask.
9. Repeat steps 2-8 with your second solution.

Part C: Temperature Effect

1. Fill the water bath approximately halfway with water at your assigned temperature and place it on top of a magnetic stirplate.
2. Repeat steps 2-7 from part B with a freshly prepared solution 1.
3. Measure the temperature of the bath again. Replace the water in the bath with hot water and let the reaction go to completion (no more O2 produced). Save the solution for part D.

Part D: Does I- get consumed during the reaction?

1. Remove the tubing from the burette & empty out the water. Rinse the burette with a small amount of 0.1 M AgNO3. Fill the burette with 0.1 M AgNO3 so that the level in the burette is around the 37 mL mark and read the exact volume.
2. Add 1-2 drops of K2CrO4 indicator to the flask from part C.
3. Titrate the solution with AgNO3 until the endpoint is reached. At the endpoint, the fine yellow precipitate that forms as AgNO3 is added coagulates and the solution becomes red from Ag2CrO4 formation. Note: This often appears more like an orange precipitate.
4. Read the final volume on the burette.

**Data Collection**

*Create a table in your notebook for the preparation of your solutions.* Be sure that it is numbered and titled appropriately. Include the initial concentrations for I− and H2O2.

Copy the following tables into your notebook filling in the appropriate information for your solutions.

**Table 2.** Decomposition rate for solution 1 at room temperaturea.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| burette reading (mL) | volume of O2 (mL) | time (min:s) | time (s) | moles of O2 | moles of H2O2 | [H2O2] |
|  | 0.00 | 0:00 | 0 | 0 |  |  |
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|  |  |  |  |  |  |  |

a) initial temperature \_\_\_\_\_\_\_\_°C, final temperature \_\_\_\_\_\_\_\_°C, barometric pressure \_\_\_\_\_\_\_

*For tables 3 & 4 create identical tables to table 2 including footnotes using the titles below.*

**Table 3.** Decomposition rate for solution (*insert your second solution*) at room temperaturea.

**Table 4.** Decomposition rate for solution 1 at (*insert your second target temperature*)a.

**Table 5.** Titration of iodide with AgNO3 to determine amount remaining.

|  |  |
| --- | --- |
| initial burette reading (mL) |  |
| final burette reading (mL) |  |

**Calculations and In Class Analysis**

1. **Prior to coming to lab, COMPLETE the following calculations and include one example of each in the calculation section of your notebook**
	1. intitial concentration (moles/liter) of I− **for each solution**
	2. initial moles of H2O2 assuming 3.00% wt/vol stock solution **for each solution**
	3. initial concentration of H2O2 **for each solution**
	4. *In class you will be creating a spreadsheet to do the calculations for Tables 2-4. We will assume that the vapor pressure of water has no effect and that the temperature of the O2 collected is the same as room temperature.* Prior to lab for solution 1, assuming that the temperature in the lab is 23 °C, the barometric pressure is 754 mm Hg and that exactly 2.00 mL of O2 are collected for the second row calculate:
		1. moles of O2 collected
		2. moles of H2O2 remaining
		3. new concentration of H2O2

*You may want to set-up a spreadsheet in advance for in lab calculations.*

1. In Class
	1. Create a graph for each solution with time on the x axis and the concentration of H2O2 on the y axis. Include the equation of the trendline and the R2 value. The units of the slope will mole/(L s) so the absolute value of the slope will be equal to the rate of the reaction.
	2. Report the average temperature, rate and [H2O2]0 and [I-]0 for each of your solutions and the order of reaction that you obtained for either H2O2 or I-.