

The Effect of Sodium Sulfate Concentration on the Rate of Production of Hydrogen Gas in the Electrolysis of Water

IB HL Chemistry IA

Introduction

Hydrogen has a myriad of applications, specifically in rocketry and fuel cell research (Heiney, 2015). Hydrogen undergoes a combustion reaction with oxygen, allowing it to act as a fuel source. Due to a lack of carbon, hydrogen combustion does not produce unwanted carbon based byproducts that form in hydrocarbon combustion (such as smoke and carbon dioxide gas). This makes hydrogen a much cleaner, and less harmful fuel source than other hydrocarbon based fuels (such as ethanol or gasoline).

The production of hydrogen gas can be done through multiple methods (Office of Energy Efficiency and Renewable Energy, n.d., *Hydrogen Production Processes*). One of these methods, the electrolysis of water, produces no greenhouse gasses on its own (Office of Energy Efficiency and Renewable Energy, n.d., *Hydrogen production: Electrolysis*). If powered through clean energy, electrolysis is a prime contender in providing a greenhouse gas free source of Hydrogen.

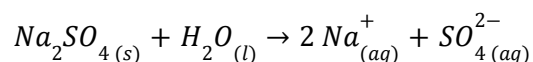
Electrolysis is a reaction in which current is run through a compound in order to drive a dissociation reaction (Kognity, 2016, *Redox processes*). Electrolytes are compounds that dissociate in water to produce ions that are capable of carrying current, decreasing resistance, and increasing conductance. In the electrolysis of water, electrolytes are necessary, as water is not a good conductor of electricity (Kognity, 2016, *Redox processes (AHL)*). The aim of this paper will be to investigate how electrolyte concentrations affect the rate of an electrolysis reaction. Specifically, the electrolyte used in this paper will be an aqueous solution of Sodium Sulfate, and the reaction investigated will be the electrolysis of water.

Research Question

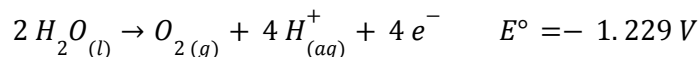
How does the concentration of sodium sulfate as an electrolyte affect the rate of production of Hydrogen gas in an electrolysis reaction?

Background

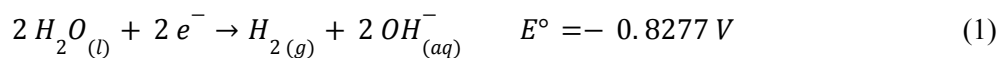
The electrolysis of water involves a constant current applied across a dilute aqueous solution of an electrolyte (Kognity, 2016, *Redox processes (AHL)*). In this paper, the electrolyte is Na_2SO_4 , which dissociates as follows,



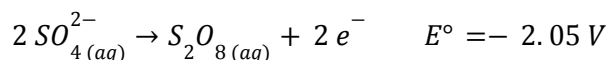
At the anode, water undergoes an oxidation reaction, producing oxygen gas (Libretexts, n.d., *The Electrolysis of Water*),



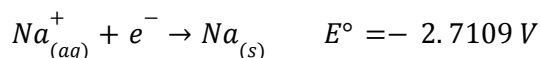
At the cathode, water undergoes an reduction reaction, producing hydrogen gas,



At the anode, sulfate anions also get oxidized,

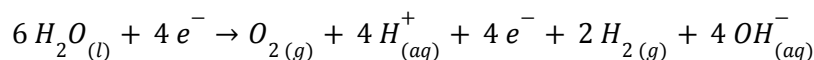


At the cathode, sodium cations also gets reduced,

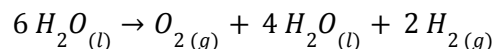


Since the standard oxidation potential of water is greater than the standard oxidation potential of sulfate, and the standard reduction potential of water is greater than the standard reduction potential of sodium cations, the water redox reactions are more likely to occur, and are thus more prevalent (Gho & Singh, 2022, *Connection between cell potential, ΔG , and K*).

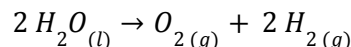
We can now combine the two half reactions for water into one (Libretexts, n.d., *The Electrolysis of Water*; Kognity, 2016, *Redox processes (AHL)*),



Canceling out the electrons on both sides, and bonding the hydrogen cations with hydroxide anions,



Canceling out the excess water on both sides, we achieve a final full formula,



Theoretical Relationship

To find the amount of hydrogen and oxygen gas produced in our electrolysis reaction, we can start by calculating the charge supplied by the battery over a certain period of time (Kognity, 2016, *Redox processes (AHL)*). We can use

$$Q = It$$

where Q is charge, I is current, and t is time. Current can be computed as,

$$I = \frac{E}{R}$$

where I is current, E is voltage, and R is resistance. Substituting back into $Q = It$,

$$Q = \frac{Et}{R}$$

Charge can be expressed in terms the number of moles of electrons, multiplied by the charge of each electron,

$$Q = Fn$$

where n is the number of moles electrons, and F is Faraday's constant. Substituting into $Q = \frac{Et}{R}$,

$$Fn = \frac{Et}{R}$$

$$n = \frac{Et}{FR}$$

According to Equation 1, for every 2 moles of electrons supplied, one mole of hydrogen gas is produced. We can express this as,

$$2n = \text{moles } H_2$$

Substituting in $n = \frac{Et}{FR}$,

$$\text{moles } H_2 = \frac{2Et}{FR}$$

We can now find a formula for the rate of production of hydrogen gas, by dividing both sides by time.

$$\text{Rate of } H_2 \text{ Production} = \frac{2E}{FR}$$

In an ideal system, conductivity, and thus conductance is directly proportional to electrolyte concentration (Lower, 2022, *Weak and strong electrolytes*). This can be expressed as,

$$\sigma \propto G \propto [Na_2SO_4]$$

where σ is conductivity, and G is conductance. It is important to note that at high concentrations, conductivity decreases as electrolyte concentration increases. With that being said, the concentrations tested in this experiment will be relatively low, thus resembling a more ideal situation. Since conductance is the inverse of resistance,

$$R \propto \frac{1}{[Na_2SO_4]}$$

Substituting back into $\text{Rate of } H_2 \text{ Production} = \frac{2E}{FR}$,

$$\text{Rate of } H_2 \text{ Production} \propto \frac{2[Na_2SO_4]E}{F} \propto [Na_2SO_4]$$

Thus the expected relationship between the concentration of sodium sulfate as an electrolyte, and the rate of production of Hydrogen gas in an electrolysis reaction, is expected to be linear.

Hypothesis

If an electrolysis reaction is carried out across an dilute aqueous solution of sodium sulfate, then the rate of production of hydrogen gas will be found to be linearly proportional to the concentration of sodium sulfate in the solution, because an increase in sodium sulfate results in an increase of sodium and sulfate ions, which increase the conductance of the electrolyte solution.

Methodology

In order to complete this experiment, an electrolysis machine is required. For this paper, I built my own custom electrolysis machine, using a container, a 9 volt battery, alligator clips, and two graphite drawing pencils to act as graphite leads. This setup is illustrated in figure 1.

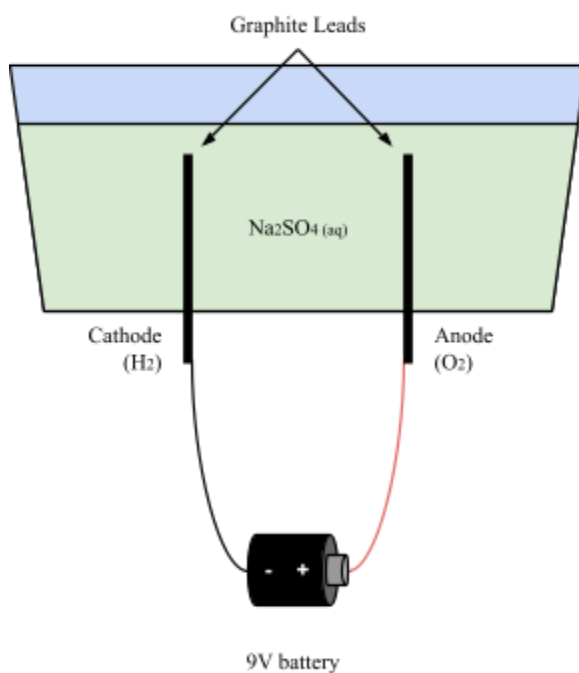


Figure 1. Illustrated electrolysis apparatus.

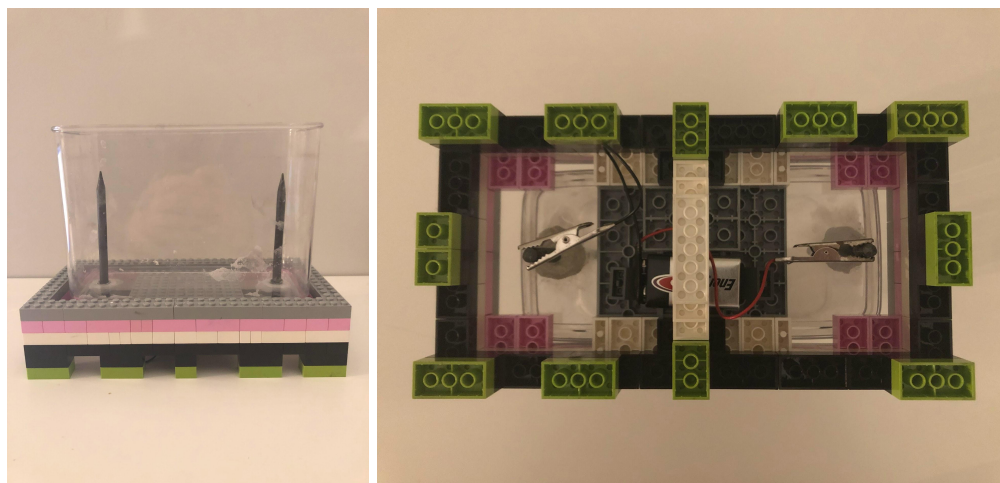


Figure 2. Real electrolysis machine. On the left, a forward facing view. On the right, a bottom-up view.

The graphite leads are sourced from pure graphite sketching pencils. They are sanded down to remove the plastic coating. The plastic container has a 2 L volume. Holes for the graphite leads were done by heating a paperclip, and poking it through the plastic. Afterwards, the graphite leads were run through the goals, and secured with J-B Weld epoxy. The lego stand

Variables

Controlled Variables:

1) Volume of solution

1.8 L of water was initially used. This same volume of water was used throughout the whole experiment. Different sodium sulfate concentrations were achieved by repeatedly adding solid sodium sulfate to the same volume of water. Evaporation was minimized using saran wrap. This solution did eventually become contaminated because of the timespan in which it was left standing. This is a potential source of systematic error.

2) Voltage

9V batteries were used throughout the experiment. The battery was replaced three times throughout the experiment, when they ran out. With that being said, data points collected towards the end of the batteries life may have received lower voltage. This is another source of systematic error.

3) Distance between leads, and height of leads

Single holes were drilled at the start of the experiment, through which all leads were inserted. Lead height was kept constant throughout the experiment.

4) Resistance of wires and graphite leads

The same alligator clips and wires are used for every data point to connect the 9V batteries to the graphite leads. The graphite leads, due to constantly being in contact with water, are extremely fragile, and broke two times. Since the graphite leads had to be replaced, a source of random error was introduced. With that being said, a constant length is used between experiments.

5) Time interval of reaction

The rate of production of hydrogen gas is measured through finding the volume of hydrogen produced in five minutes, and dividing it by the elapsed time. This five minute interval is kept constant between data points and trials.

Independent Variable: Concentration of sodium sulfate

In this experiment, solid sodium sulfate was repeatedly added in intervals of 5 g to 1.8 L, to make solutions with 10.42 g, 15.66 g, 20.93 g, 25.93 g, and 30.95 g sodium sulfate. These solutions have sodium sulfate concentrations of 0.041 M, 0.061 M, 0.082 M, 0.10 M, and 0.12 M.

Dependent Variable: Rate of production of hydrogen gas

This was determined by running the electrolysis reaction for five minutes per data point, and collecting the hydrogen gas in an inverted graduated cylinder filled with water. The volume of hydrogen gas was then divided by the five minutes elapsed to determine the rate of hydrogen production.

Materials

- Electrolysis apparatus (see Figures 1 and 2)
- Graduated cylinder (10 mL)
 - Note: for this paper, larger graduated cylinders with less precision were initially used.
Because of this, the uncertainties for earlier data points are significantly larger than those for later data points.
- Ring stand and burette clamp for ring stand
- Test tube clamp
- 2 beaker (one 25 mL beaker and one 2 L beaker)
- Balance scale
- Spoon

Chemicals

- Water (1.8 L)
- Sodium sulfate powder (30.95 g)

Procedure

- 1) Connect the 9 V battery to the two graphite leads using the alligator clips.
- 2) Set up the ring stand and burette clamp as seen in Figure 3.

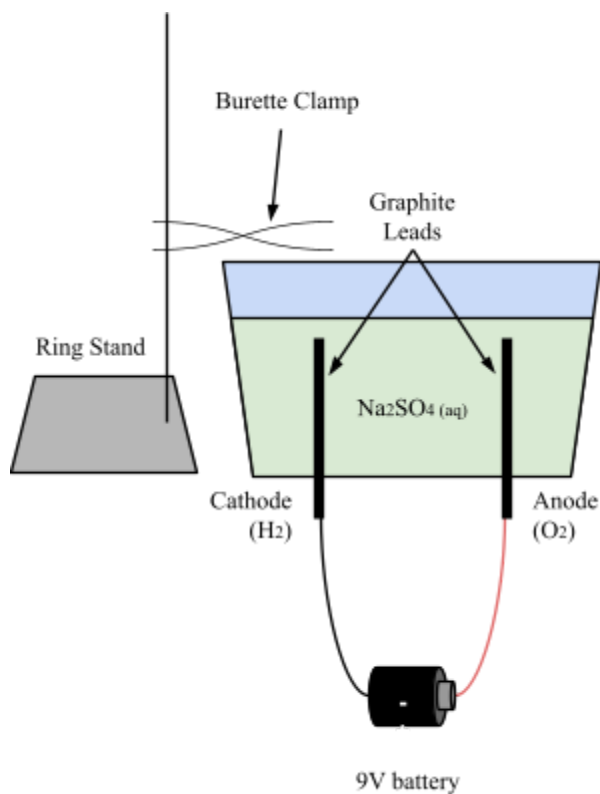


Figure 3. Illustrated experimental setup.

- 3) Fill a 2 L beaker with 1.8 L of water.
- 4) Place the 25 mL beaker on the centigram balance. Scoop the sodium sulfate powder into the beaker, until you have 10 g sodium sulfate.
- 5) Pour the sodium sulfate into the 2 L beaker and stir to form the first, 0.041 M sodium sulfate solution.
- 6) Transfer the contents of the 2 L beaker into the electrolysis machine.
- 7) Use the test tube clamp to fill the graduated cylinder with sodium sulfate solution. Ensure that there are no bubbles in the solution. Then, clamp the graduated cylinder upside down to the burette clamp. Make sure that the opening of the graduated cylinder remains submerged in the solution, so that suction is maintained.
- 8) Start a stopwatch, and measure the volume by which the water is displaced in the graduated cylinder, after five minutes. This is the volume of hydrogen gas produced.
- 9) Repeat step eight, five times, once for each trial.
- 10) Empty out the electrolysis machine to the 2 L beaker. Add another 5 g of sodium sulfate, and stir. Return the solution to the electrolysis machine. Repeat steps 8 to 10, three more times, to collect all 25 data points (5 trials \times 5 concentrations).
- 11) Dispose sodium sulfate solution down the sink.

Safety and Ethical Considerations

Anhydrous sodium sulfate is a relatively safe chemical to handle (Fisher Scientific, 2008, *Sodium sulfate anhydrous*). Contact with the eyes and skin should be avoided. For this reason, goggles were worn at all times during the lab.

Likewise, a 0.1 M solution of sodium sulfate is also relatively safe to handle (Fisher Scientific, 2015, *Sodium Sulfate Solution, 0.1M*). Contact with the eyes and skin should also be avoided, however the risk of skin contact can be mitigated by washing your hands after experiment completion.

In order to not waste sodium sulfate and water, the same 1.8 L solution was used throughout all data points and trials. This saved around 7.2 L of water, and 70 g of sodium sulfate total.

Data and Observations

Qualitative Data

There is no visually discernible change to the reaction when the concentration of sodium sulfate is increased. With that being said, the reaction does take some time to reach its maximum rate, and is a potential source of systematic error. Additionally, there are noticeable hydrogen bubbles that escape the graduated cylinder. This means that not all the hydrogen gas produced is being measured, and is another source of systematic error.

The most concerning observation of all however, is the quality of the water. As seen in Figure 4, over the duration of the experiment, the water got significantly contaminated. This is a major source of systematic error.

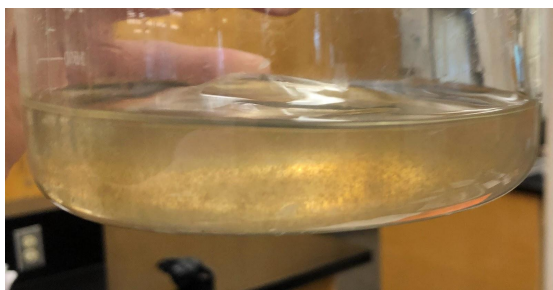


Figure 4. The water had significant contamination precipitating at the bottom of the beaker by the end of the experiment, manifesting in brown cloudy particles. This could be potentially mitigated by keeping the beaker in a sterile environment, or not recycling the solution.

When measuring the volume of hydrogen gas produced, some bubbles formed in the hydrogen pocket of the graduated cylinder. This leads to error and uncertainty in the measurements of the volumes of hydrogen gas produced. To account for this, all uncertainties for the volume of hydrogen gas produced will be increased by 0.1 on top of their usual value.



Figure 5. Bubbling at the top of the graduated cylinder.

Table 1 - Volumes of Hydrogen Gas Produced for Variable Masses of Sodium Sulfate

Sodium Sulfate Masses Dissolved (g)		10.42 ± 0.02	15.66 ± 0.03	20.93 ± 0.04	25.93 ± 0.05	30.95 ± 0.06
Volume of Hydrogen Gas Produced (mL)	Trial 1	3.5 \pm 0.6	4.5 \pm 0.6	4.4 \pm 0.3*	4.1 \pm 0.2	4.3 \pm 0.2*
	Trial 2	3.5 \pm 0.6	4.4 \pm 0.3	4.9 \pm 0.2	4.0 \pm 0.2	5.9 \pm 0.2
	Trial 3	4.0 \pm 0.6	4.0 \pm 0.3	5.0 \pm 0.2	4.0 \pm 0.2	5.9 \pm 0.2
	Trial 4	4.0 \pm 0.6	3.8 \pm 0.3	4.8 \pm 0.2	6.0 \pm 0.2*	6.1 \pm 0.2
	Trial 5	3.0 \pm 0.6	3.6 \pm 0.3	5.1 \pm 0.2	3.5 \pm 0.2	5.9 \pm 0.2

*Not included in later graphs and analysis, due to being significant outliers

Table 2 - Controlled Variables

Time Elapsed for Each Datapoint	$300 \pm 20 \text{ s}$
Volume of Water Used	$1.8 \pm 0.1 \text{ L}$

Data Analysis

First, we can convert the mass of sodium sulfate dissolved into the sodium sulfate molarity, and the volume of hydrogen produced in 5 minutes to a rate of production. The following is a sample calculation for converting $10.42 \pm 0.02 \text{ g Na}_2\text{SO}_4$ into a molarity,

$$\begin{aligned}
 [\text{Na}_2\text{SO}_4] &= (10.42 \pm 0.02 \text{ g}) \left(\frac{\text{mol}}{2(22.99 \text{ g}) + (32.07 \text{ g}) + 4(16.00 \text{ g})} \right) (1.8 \pm 0.1 \text{ L})^{-1} \\
 &= (10.42 \pm 0.02 \text{ g}) \left(\frac{\text{mol}}{142.05 \text{ g}} \right) (1.8 \pm 0.1 \text{ L})^{-1} = \frac{10.42 \pm 0.02 \text{ mol}}{(142.05)(1.8 \pm 0.1 \text{ L})} = \frac{0.07335 \pm 0.00014 \text{ mol}}{1.8 \pm 0.1 \text{ L}} \\
 &= \frac{0.07335}{1.8} \text{ M} \pm \left(\left(\frac{0.1}{1.8} + \frac{0.00014}{0.07335} \right) \times 100\% \right) = 0.041 \text{ M} \pm 6\% = 0.041 \pm 0.00246 \text{ M} \\
 &= 0.041 \pm 0.003 \text{ M}
 \end{aligned}$$

And the sample calculation for converting $3.5 \pm 0.6 \text{ mL H}_2$ into a production rate,

$$\begin{aligned}
 \text{Rate of H}_2 \text{ Production} &= \frac{3.5 \pm 0.6 \text{ mL}}{300 \pm 20 \text{ s}} = \frac{3.5}{300} \text{ mL s}^{-1} \pm \left(\left(\frac{0.6}{3.5} + \frac{20}{300} \right) \times 100\% \right) \\
 &= 0.012 \text{ mL s}^{-1} \pm 24\% = 0.012 \pm 0.003 \text{ mL s}^{-1}
 \end{aligned}$$

Table 3 - Rates of Production of Hydrogen Gas for Variable Concentrations of Sodium Sulfate

Sodium Sulfate Concentration (M)		0.041 ± 0.003	0.061 ± 0.004	0.082 ± 0.005	0.10 ± 0.01	0.12 ± 0.01
Rate of Production of Hydrogen Gas (mL s^{-1})	Trial 1	0.012 ± 0.003	0.015 ± 0.003	$0.015 \pm 0.002^*$	0.014 ± 0.002	$0.014 \pm 0.002^*$
	Trial 2	0.012 ± 0.003	0.015 ± 0.002	0.016 ± 0.02	0.013 ± 0.002	0.020 ± 0.002
	Trial 3	0.013 ± 0.003	0.014 ± 0.002	0.017 ± 0.02	0.013 ± 0.002	0.020 ± 0.002
	Trial 4	0.013 ± 0.003	0.013 ± 0.002	0.016 ± 0.02	$0.020 \pm 0.002^*$	0.020 ± 0.003
	Trial 5	0.010 ± 0.003	0.012 ± 0.002	0.017 ± 0.02	0.012 ± 0.002	0.020 ± 0.002

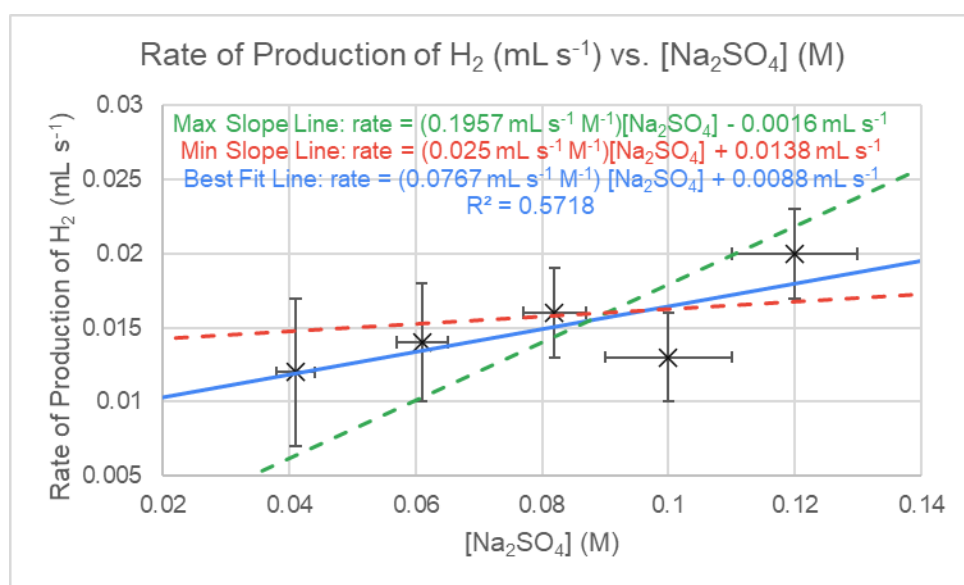
*Not included in later graphs and analysis, due to being significant outliers

Next, we can find the average production rate of hydrogen gas, for every given sodium sulfate concentration. Following, is the sample calculation for the average production rate of hydrogen gas for a $[Na_2SO_4]$ of $0.041 \pm 0.003 M$.

$$\begin{aligned} \text{Average Rate of } H_2 \text{ Production} &= \frac{0.012+0.012+0.013+0.013+0.010}{5} \text{ mL s}^{-1} \\ &\pm \frac{\text{Max Average Rate of } H_2 \text{ Production} - \text{Min Average Rate of } H_2 \text{ Production}}{2} \\ &= 0.012 \pm \frac{0.016-0.007}{2} \text{ mL s}^{-1} = 0.012 \pm \frac{0.009}{2} \text{ mL s}^{-1} = 0.012 \pm 0.005 \text{ mL s}^{-1} \end{aligned}$$

Table 4 - Average Rate of Production of Hydrogen Gas for Variable Concentrations of Sodium Sulfate

Sodium Sulfate Concentration (M)	0.041 ± 0.003	0.061 ± 0.004	0.082 ± 0.005	0.10 ± 0.01	0.12 ± 0.01
Rate of Production of Hydrogen Gas (mL s^{-1})	0.012 ± 0.005	0.014 ± 0.004	0.016 ± 0.003	0.013 ± 0.003	0.020 ± 0.003



It was found, following a linear relationship, using the Least Mean Squared (LMS) algorithm on excel, and fitting to the data in Table 4, that the line of best fit without significant figures or uncertainties is

$$\text{rate} = (0.0767 \text{ mL s}^{-1} \text{ M}^{-1})[Na_2SO_4] + 0.0088 \text{ mL s}^{-1}$$

The R^2 was found to be 0.5718, showing a weak fit to the data. The line with maximum slope, that intercepts all error bars, is

$$\text{rate} = (0.1957 \text{ mL s}^{-1} \text{ M}^{-1})[Na_2SO_4] + 0.0016 \text{ mL s}^{-1}$$

The minimum slope line that intercepts all error bars is

$$rate = (0.025 \text{ mL s}^{-1} \text{ M}^{-1})[Na_2SO_4] + 0.0138 \text{ mL s}^{-1}$$

Since the deviation between the best fit line and min slope line, and the deviation between the best fit line and max slope line, are different, we will have asymmetric uncertainties in our best fit line. This will be notated using superscripts for positive uncertainties, and subscripts for negative uncertainties. This method is the standard method for representing asymmetric uncertainties. Below, are the uncertainty calculations for the best fit line,

$$\text{let } E_{m+} = \text{positive uncertainty on slope of best fit line}$$

$$\text{let } E_{m-} = \text{negative uncertainty on slope of best fit line}$$

$$\text{let } E_{b+} = \text{positive uncertainty on y intercept of best fit line}$$

$$\text{let } E_{b-} = \text{negative uncertainty on y intercept of best fit line}$$

$$E_{m+} = (0.1957 - 0.0767) \text{ mL s}^{-1} \text{ M}^{-1} = 0.119 \text{ mL s}^{-1} \text{ M}^{-1} = 0.2 \text{ mL s}^{-1} \text{ M}^{-1}$$

$$E_{m-} = (0.0767 - 0.025) \text{ mL s}^{-1} \text{ M}^{-1} = 0.0517 \text{ mL s}^{-1} \text{ M}^{-1} = 0.1 \text{ mL s}^{-1} \text{ M}^{-1}$$

$$E_{b+} = (0.0138 - 0.0088) \text{ mL s}^{-1} = 0.005 \text{ mL s}^{-1}$$

$$E_{b-} = (0.0088 - 0.0016) \text{ mL s}^{-1} = 0.0072 \text{ mL s}^{-1} = 0.008 \text{ mL s}^{-1}$$

Now that we have calculated our asymmetrical uncertainties, we can use the previously mentioned superscript/subscript notation to express our final best fit line,

$$rate = (0.1^{+0.2}_{-0.1} \text{ mL s}^{-1} \text{ M}^{-1})[Na_2SO_4] + 0.009^{+0.005}_{-0.008} \text{ mL s}^{-1}$$

Conclusion

Experimentally, it was found that the relationship between the concentration of a sodium sulfate solution, to the rate of hydrogen gas production at the cathode of an electrolysis reaction, is

$$rate = (0.1^{+0.2}_{-0.1} \text{ mL s}^{-1} \text{ M}^{-1})[Na_2SO_4] + 0.009^{+0.005}_{-0.008} \text{ mL s}^{-1}$$

Theoretically, our hypothesis stated that there would be that the rate of hydrogen gas production would be directly proportional to the concentration of sodium sulfate, or

$$rate \propto [Na_2SO_4] \text{ or } rate = m[Na_2SO_4], \text{ where } m \text{ is a positive constant}$$

The full range of possible values for the slope of the experimental relationship is positive. Additionally, the relatively small size of the intercept makes it negligible. Because of this, the experimental relationship supports the hypothesis.

With that being said, this experiment is inconclusive. The experimental relationship has an R^2 of 0.5718, suggesting an extremely weak fit. Additionally, the uncertainties on the experimental relationship's slope are extremely high ($0.1^{+200\%}_{-100\%} \text{ mL s}^{-1} \text{ M}^{-1}$). The inconclusiveness of this paper is due to the large number of significant sources of error in this experiment.

Sources of Error

- 1) Water contamination - In an attempt to save on chemicals, the same solution was used throughout the whole experiment (see Figure 4). Due to the amount of time the solution was left to stand still, it likely began growing some kind of contaminant organic matter. This contaminant likely affected the conductance of the solution, thus affecting the final volume of hydrogen gas produced. This is a systematic source of error. This can be mitigated by preparing a new solution for each trial, or working in a sterile environment, with sterile water.
- 2) Hydrogen bubbles in the graduated cylinder - After the hydrogen is collected in the inverted graduated cylinder, some bubbles are left over (see Figure 5). These bubbles make it difficult to take an accurate reading from the graduated cylinder. This systematic source of error can be mitigated by waiting a period of time after each data point to wait for the bubbles to dissipate.
- 3) Hydrogen escaping the graduated cylinder - Noticeable bubbles can be seen escaping the graduated cylinder at the sides. This means that there is a constant loss of hydrogen to the surroundings. This can be mitigated by attaching an inverted funnel to the graduated cylinder to capture as much hydrogen as possible.

Future Work and Improvements

One major improvement that should be made to this experiment is swapping out the battery for a constant power supply. As the battery loses charge, it supplies less voltage, which is a huge source of systematic error.

One major strength of this paper is the lack of material usage, allowing for easy and cheap replication. Using one solution throughout the whole experiment saves around 7.2 L of water, and 70 g of sodium sulfate total. Because of this, as an avenue for future work, significantly higher concentrations of sodium can be tested economically. Increasing the concentration would allow you to observe the theoretical drop in conductance, and thus drop in hydrogen production rate (Lower, 2022, *Weak and strong electrolytes*), as well as investigate the concentration required for peak production rate.

Another future avenue of exploration is the investigation of other electrolytes. Potassium hydroxide is commonly used in real world applications for hydrogen production (ScienceDirect, n.d., *Water Electrolysis*). Testing for the production of other gasses using electrolysis (e.g. chlorine gas in the electrolysis of aqueous sodium chloride) is also a potential avenue for future work.

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