

The Effect of The Concentration of Zinc Nitrate in a Redox Reaction With Copper (II) Sulfate on the Voltage Produced by a Voltaic Cell



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Introduction

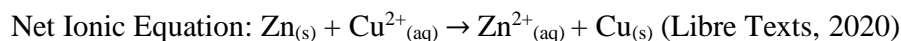
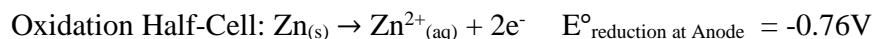
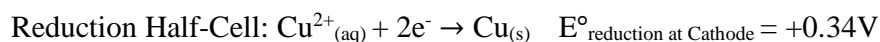
Aims/Objective

This Internal Assessment aims to investigate the relationship between the concentration of zinc nitrate in a redox reaction with 0.50M copper (II) sulfate, and the voltage produced by the voltaic cell in which the reaction occurs.

Introduction/Background

While learning about enthalpies and heat transfer in class, I always wondered how the energy released by reactions could be harnessed as a source of power. This led me to ponder how different types of energy, especially electrical energy, could be produced through chemical reactions. Doing more research, I came across the concept of voltaic (Galvanic) cells, which used spontaneous redox reactions to produce electrical energy (Libretexts, 2020). After wrapping my head around voltaic cells and redox reactions, I wanted to know how varying the concentration of the anode solution in a voltaic cell affected its voltage. This is because I noticed that minimal experimental research was conducted, as most used the cathode solution as their independent variable, not the anode solution. Furthermore, having not yet covered this topic in standard-level chemistry makes it a completely new exploration. This led to the question, how does changing the concentration of zinc nitrate in a redox reaction with copper sulfate affect the voltage produced by a voltaic cell?

A voltaic cell consists of two electrodes, the anode and cathode, which are solid metals ($\text{Zn}_{(s)}$ and $\text{Cu}_{(s)}$) placed into solutions of $\text{Zn}^{2+}_{(aq)}$ and $\text{Cu}^{2+}_{(aq)}$ respectively (for this experiment specifically). Before the voltaic cell is complete, a salt bridge soaked in an electrolyte solution connects the two solutions to complete the circuit, preventing a build-up of charge at each electrode. The oxidation part of the reaction occurs at the anode, and the reduction part occurs at the cathode, which can be represented as half-reactions.



The difference between the reduction potentials of the two cells is the cell potential (E°), which can be measured as a voltage (cell potential and voltage are used as interchangeable terms). Note that the reduction potential of the oxidation half cell is negative, because at its core, reduction potential is the ability of an element to undergo reduction (gain electrons). Since $\text{Zn}_{(s)}$ oxidizes (loses electrons), its reduction potential is negative. This gives us,

$$E^{\circ} = E^{\circ}_{\text{reduction at cathode}} - E^{\circ}_{\text{reduction at anode}}. \quad (\text{Libre Texts, 2020}).$$

Therefore, $E^{\circ} = 0.34\text{V} - (-0.76\text{V}) = 1.10\text{V}$. This is the cell potential of our reaction at standard conditions and will be used in further calculations, specifically in the Nernst's Equation.

Moving on, this investigation is directly related to the **Nernst Equation**, which is derived through the relationship between Gibbs Free Energy (G) and Cell Potential (E). Gibbs Free Energy is the amount of energy that can be used for work. The formula governing this relationship is $\Delta G = -nFE$ (1). n represents the number of electrons transferred in the reaction based on the balanced chemical equation (2 for this reaction); ΔG denotes the change in free energy; E is the cell potential of the reaction (V); and F is the Faraday Constant (96,485 C/mol), the charge associated with one mole of electrons (Key, 2014). Under standard conditions, equation (1) becomes $\Delta G^{\circ} = -nFE^{\circ}$ (2). Finally, non-standard Gibbs Free Energy, and Standard Gibbs Free Energy, are related through the following formula: $\Delta G = \Delta G^{\circ} + RT \ln(Q)$ (3), where:

$$R = \text{Gas Constant } (8.314 \frac{J}{mol \cdot K})$$

$$T = \text{Temperature } (298 K)$$

$$Q = \text{Reaction Quotient } \left(\frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} \text{ in our case} \right)$$

Substituting equation (1) and (2) into (3), we get:

$$\Rightarrow -nFE = -nFE^\circ + RT \ln(Q) \quad (\text{divide both sides by } -nF)$$

$$\Rightarrow E = E^\circ - \frac{RT}{nF} \ln(Q) \quad (\text{Libretexts, 2020})$$

There we have it, we have derived the Nernst equation, which allows us to find the cell potential of the reaction at non-standard conditions using its standard cell potential.

Next, let's fit the concept of equilibrium (where $\Delta G = 0$ and $E = 0$) into the big picture of cell potentials. Our reaction eventually produces an equilibrium, where the products side is **heavily** favoured (Libre Texts, 2019). Consequently, as the reaction proceeds, the concentration of products ($Zn^{2+}_{(aq)}$) increases, and the concentration of reactants ($Cu^{2+}_{(aq)}$) decreases, meaning the reaction quotient increases and the system gets closer to equilibrium. Since cell potential decreases as the reaction quotient increases (according to the Nernst equation); we can deduce that the further away the system is from equilibrium, the larger the cell potential measured is expected to be. In fact, this has numerous applications to the real world; Certain batteries (e.g lead-acid cells in automobiles) work utilizing the same principle. They are made of at least one voltaic cell, and the reason these batteries die is due to the products and reactants reaching an equilibrium, at which $E = 0$, so no energy is produced (Honsberg & Bowden, n.d).

Lastly, there is an important piece non-chemistry information to consider; rearranging the Nernst equation to linearize it in terms of E and $[Zn(NO_3)_2]$. Firstly, $\frac{RT}{nF}$ is a constant, so substituting in the values results in:

$\frac{(8.314 \frac{J}{mol \cdot K})(298K)}{(2)(96,485 \frac{C}{mol})} = 0.0128V$. Moreover, $E^\circ = 1.10 V$, which was calculated earlier. Therefore, the equation simplifies to $E = 1.10 V - 0.0128V \cdot \ln(Q)$. Additionally, it is known that:

$$\begin{aligned} \ln(Q) &= \ln \left(\frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} \right) \\ &= \ln \left(\frac{[Zn(NO_3)_2]}{0.50 M} \right) \\ &= \ln[Zn(NO_3)_2] - \ln(0.50M) \end{aligned}$$

This gives us: $E = 1.10 V - 0.0128 (\ln[Zn(NO_3)_2] - \ln(0.50 M))$

$$\therefore E = -0.0128 \cdot \ln[Zn(NO_3)_2] + 1.109V \rightarrow y = m \cdot (x) + b$$

Therefore, plotting E against $\ln[Zn(NO_3)_2]$ is expected to yield a linear relationship, with the slope being -0.0128 , the negative of the value that results from evaluating $\frac{RT}{nF}$.

Research Question

How does changing the concentration of the zinc nitrate solution (0.25M, 0.50M, 0.75M, 1.00M, 1.25M) in a redox reaction with a 0.50M copper (II) sulfate solution at 298 K affect the initial voltage (V) that is produced by the voltaic cell in which the reaction proceeds?

Theoretical Values

To begin, a comparison between the experimental values found and the theoretical values that are expected should be made to construct a good analysis. Therefore, using the Nernst equation determined earlier, the theoretical voltages that should be produced can be calculated.

Concentration of zinc nitrate (M)	Theoretical Cell Potential (V)	Theoretical Cell Potential Sample Calculation (0.25M)
0.25	1.109	zinc nitrate Concentration = 0.25M copper (II) sulfate Concentration = 0.50M $E^\circ = 1.10\text{V}$ $R = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}$ $n = 2$ $F = 96,485 \frac{\text{C}}{\text{mol}}$ $T = 298 \text{ K}$ $E = 1.10 - \frac{8.314(298)}{2(96485)} \ln(0.25\text{M}/0.50\text{M})$ $E = 1.109 \text{ V}$
0.50	1.100	
0.75	1.095	
1.00	1.091	
1.25	1.088	

Figure 1: Table of theoretical cell potentials for each concentration change with a sample calculation

Methods and Materials

Materials and Safety

Item	Quantity Used	Safety: Risks and Precautions
Copper Sulfate Pentahydrate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)	$312.100 \pm 0.001\text{g}$	Long exposure can cause eye irritation and skin irritation. Wash out eyes at the eyewash station in case of exposure to the eyes, and make sure to wash skin thoroughly in case of skin exposure for 15 minutes. If ingested, rinse out the mouth and call poison control immediately; do not induce vomiting as it will increase irritation. Wear gloves and eye protection at all times, and clothing that covers as much skin as possible. Wear a mask to ensure no dust particulates are inhaled. Please dispose of the solution in an inorganic waste bin as copper sulfate is a major hazard to aquatic life. Do not flush into surface water or a sanitary sewer system. (Fisher Scientific, Material Safety Data Sheet, 2008).
Zinc Nitrate Hexahydrate ($\text{Zn}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$)	$441.563 \pm 0.001\text{g}$	Long exposure can cause eye and skin irritation. Wash out eyes at the eyewash station in case of exposure to the eyes, and make sure to wash skin thoroughly in case of skin exposure for 15 minutes. If ingested, rinse out the mouth and call poison control immediately. Do not induce vomiting. Wear gloves and eye protection at all times, and clothing that covers as much skin as possible. Wear a mask to ensure no dust particulates are inhaled. Please dispose of the solution in an inorganic waste bin as zinc nitrate is a major hazard to aquatic life. Do not flush into surface water or sewer system. (Fisher Scientific, Material Safety Data Sheet, 2021).
Potassium Nitrate (KNO_3)	$10.110\text{g} \pm 0.001\text{g}$	Long exposure can cause eye and skin irritation. Wash out eyes at the eye wash station. However, one measure is different from copper sulfate and zinc nitrate. In the case of ingestion, do induce vomiting to remove all the potassium nitrate from the body. Conduct all experiments in a well-ventilated area to prevent a build up of potassium nitrate dust particulates. Do not flush into surface water or sanitary sewer system as potassium nitrate is a major biohazard to aquatic life. Dispose waste solution in an inorganic waste bin (EPA.gov, Materials Safety Data Sheet, 2008).

Item	Quantity Used	Safety: Risks and Precautions
250 mL beakers	2	Use care in handling and storing glassware. If an item is chipped or broken, discard it, do not attempt to reuse it. Ensure that the flasks and beakers are cleaned out thoroughly so that no solution is left in the glassware. Finally, when holding the Erlenmeyer flask, always have one hand on the neck, and another hand on the base of the flask. (Marquette University, 2022)
100 mL Graduated Cylinder ($\pm 0.5\text{mL}$)	1	
1000 mL Erlenmeyer Flask	2	
Copper Metal Strips (2) Zinc Metal Strips (2) Distilled Water ($5000.0 \pm 0.5 \text{ mL}$) Alligator Clips (2) Multimeter ($\pm 0.01 \text{ V}$) (1) Weighing boat (2) Electronic balance ($\pm 0.001\text{g}$) (1) Scoopula (2) Safety Glasses (1) Stirring Rod or Magnetic Stirrer (2) Filter Papers or Tissue Papers (25 strips)		There were <i>no ethical concerns</i> that needed to be taken into consideration, as no animal or human subjects were used in the experiment.

Variables

Variable	Description	How it will be controlled, measured, or manipulated
<u>Independent</u>	Concentration of $\text{Zn}(\text{NO}_3)_2$ solution	A volume of 100.0 mL will be used with the following concentrations of solution: 0.25M, 0.50M, 0.75M, 1.00M, 1.25M. These values correspond to adding: 5.888g, 11.775g, 17.663g, 23.550g, 29.438g to 100.0 mL of distilled water to achieve the respective concentrations listed. The solution concentrations are placed in equal intervals of 0.25M to ensure an observable and consistent change in initial voltage.
<u>Dependent</u>	Voltage produced by the voltaic cell	Just like measuring the voltage of a circuit, a multimeter will be attached to the anode and cathode using alligator clips. The initial voltage displayed on the multimeter will be recorded.
<u>Controlled</u>	Concentration of CuSO_4 solution	The concentration of this solution (0.50M) needs to be kept constant as changing it would change the reaction quotient (Q). This would lead to changes in the voltage produced, thereby increasing the percent error of this experiment. The concentration will be controlled by added precisely 12.484g of copper sulfate to 100.0mL of distilled water measured out using a graduated cylinder. The experiment will be carried out as soon as possible after producing the solution to ensure that minimal water is lost through evaporation.
<u>Controlled</u>	Concentration of potassium nitrate Salt Bridge	Speaking in terms of electricity, ions are moving through the salt bridge to balance out the charges that start building up at each electrode. Increasing the concentration of ions in the salt bridge increases conductivity, thereby decreasing the resistance. Since voltage and resistance are connected through $V = IR$, a change in resistance could possibly hinder the voltage reading (Elert, n.d). Therefore, it is crucial to keep the concentration of the solution used for the salt bridge constant.

Variable	Description	How it will be controlled, measured, or manipulated
<u>Controlled</u>	Temperature	Temperature needs to be kept constant. As seen in the Nernst equation, changing temperature has an overall effect on the cell potential. This would lead to changes in the data, reducing the accuracy of the experiment. Temperature will be controlled by conducting all the trials at laboratory temperature to limit temperature changes as much as possible. The solutions will also be allowed to rest for a short period of time to reach room temperature since dissolving can be either <i>slightly</i> exothermic or endothermic. Therefore, it is crucial to allow the solutions to get back to room temperature. Luckily, fluctuations of a few degrees Celsius has minimal impact on the voltage produced, to the extent that its possible that the multimeter is unable to pick up a change in voltage.
<u>Controlled</u>	Time after which voltage will be measured	The voltage will be measured right after the two solid metals ($\text{Cu}_{(s)}$ and $\text{Zn}_{(s)}$) are added to their respective solutions. The alligator clips will be pre-attached to the solids to ensure that the voltage measured is as close to the initial voltage as possible. The instant the metals make contact with the solutions, the redox reaction will start since it is spontaneous, and the first voltage reading seen will be recorded.
<u>Controlled</u>	Water Source Used to Produce Solutions	This variable will be controlled by using the same type of distilled water each trial. The company from which the distilled water will be sourced is named Equate. Even though this factor would not have a significant impact on the data, it is still good to take precautionary measure since it is very easy to keep this variable constant anyways.

Method

Part A: Preparing 0.50M copper sulfate Solution

1. 250.0 mL beaker was filled with 100.0 mL of distilled water (measured out using a graduated cylinder)
2. 12.484g of copper sulfate were measured out using a scoopula and weigh boat, and then completely dissolved into the water to produce the desired solution
3. Solution was set aside to be used in the experiment once Parts B and C were completed

Part B: Preparing 0.25M zinc nitrate Solution

1. Followed the same procedure as part A, except 5.888g of zinc nitrate was used as the solute, with the same volume of water
2. Solution was set aside to be used in the experiment once Part C was completed

Part C: Preparing Salt Bridge

1. The tissue paper was folded into an approximately 20cm long strip, and then folded into a “U-shape”
2. The strip was soaked in 1.00M potassium nitrate solution, which was produced by dissolving 10.110g of potassium nitrate into 100.0 mL of distilled water
3. Soaked strip was set aside for a short period of time as the experiment was being set up

Experiment:

1. Beakers from Part A and Part B were set 5cm apart from each other

- The red probe on the multimeter was attached to a copper metal strip, and the black probe on the multimeter was attached to a zinc metal strip, using alligator clips
- Multimeter was set to the 20V DV setting
- The copper metal strip was placed into the copper sulfate solution made in Part A, and the zinc metal strip was placed into the zinc nitrate solution made in Part B
- The filter paper salt bridge was placed into the solution, such that one end of the strip was in the copper sulfate solution, and the other end was in the zinc nitrate solution
- Temperature was controlled by conducting all experiments at standard laboratory temperature between 23°C and 25°C (this small fluctuation however was uncontrollable)
- The voltage displayed on the multimeter instantly after adding the salt bridge was noted, along with the concentration of the solutions
- Both solutions were disposed into an Erlenmeyer flask to hold the waste and the beakers were washed out with distilled water.
- Repeated Parts A, B and C, as well as steps 1-8 of the experiment four more times to achieve five trials
- Prepared zinc nitrate solutions with solute masses of 11.775g, 17.663g, 23.550g, 29.438g using the steps outlined in Part B to produce 0.50M, 0.75M, 1.00M, 1.25M solutions respectively
- For each zinc nitrate solution, a 0.50M copper sulfate solution was made using the steps outlined in Part A
- Steps 1-8 were repeated for every pair of copper sulfate and zinc nitrate solution. Five trials were done for each change in concentration of zinc nitrate solution, summing to a total of 25 trials
- After all 25 trials were completed, the Erlenmeyer flasks were emptied out into an inorganic waste bin (Therefore, 0.50M copper sulfate was experimented against 0.25M, 0.50M, 0.75M, 1.00M, and 1.25M solutions of zinc nitrate, with five trials being done for each change in concentration of zinc nitrate)

How mass needed for each solution was determined (Using 0.50M solution of zinc nitrate as an example)

$$C \cdot V = n$$

$$\rightarrow n = 0.50\text{M} \cdot 0.1000\text{L}$$

$$= 0.050 \text{ mol}$$

The molar mass of zinc nitrate Hexahydrate is 235.50 g/mol, so using $m = n \cdot M$

$$m = 0.050 \text{ mol} \cdot 235.50 \text{ g/mol}$$

$$m = 11.775 \text{ g}$$

Figure 2: Sample calculation for how mass of zinc nitrate needed for 0.50M solution was determined

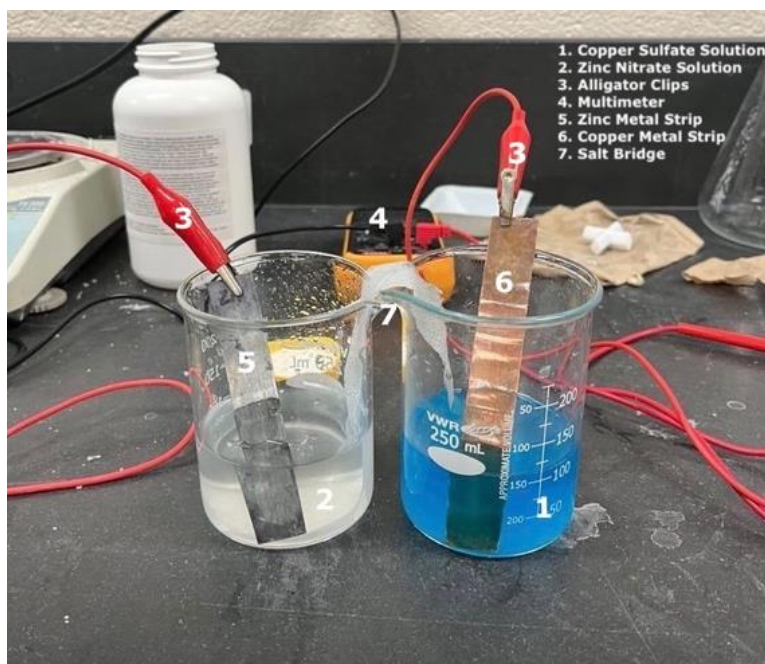


Figure 3: Apparatus of Experiment with Labels

Results and Discussion

Raw Data

Concentration of zinc nitrate (M)	Initial Voltage ($\pm 0.01V$)					
	Trial #1	Trial #2	Trial #3	Trial #4	Trial #5	Average
0.250 ± 0.001 ($\pm 0.4\%$)	1.14	1.11	1.13	1.14	1.16	1.14
0.500 ± 0.003 ($\pm 0.6\%$)	1.03	1.04	1.11	1.07	1.05	1.06
0.750 ± 0.004 ($\pm 0.5\%$)	1.02	1.04	1.04	1.01	0.98	1.02
1.000 ± 0.005 ($\pm 0.5\%$)	1.00	0.97	0.98	0.97	0.99	0.99
1.250 ± 0.006 ($\pm 0.5\%$)	1.01	0.99	0.96	1.02	0.86	0.98
Qualitative Observations	In a normal trial, when replacing the copper sulfate solution with a new one, brown solid copper particles were visible in the solution showing how the reaction was working successfully and the copper was being reduced into $Cu_{(s)}$. Another visible change was blackening of the Zinc metal strip, showing that the $Zn_{(s)}$ was successfully being oxidized into $Zn^{2+}_{(aq)}$. As a further extension out of curiosity, the cell was kept intact for one extra trial and was not disassembled after the initial voltage was measured to see what occurs when the reaction was kept continuous. The observation that was made was that the colour intensity of the copper sulfate solution slowly turned from a bright blue into a paler blue, and the zinc nitrate solution became cloudier. This was because the copper ions which gave the blue colour were precipitating into copper metal, reducing the intensity of the colour, and the zinc metal was releasing zinc ions, which caused cloudiness.					

Figure 4: Table showing quantitative results and qualitative observations of the experiment

<p>Uncertainty for 1.25M zinc nitrate</p> <p>The grams (m) of zinc nitrate used were $29.438 \pm 0.001g$</p> <p>The molar mass (M) of zinc nitrate hexahydrate is $235.50 \frac{g}{mol}$</p> <p>Using $n = \frac{m}{M} \rightarrow n = 0.125002 \pm \left(\frac{0.001g}{29.438g}\right) \cdot 0.125002 \text{ mol}$ $n = 0.125002 \pm 0.000004 \text{ mol}$ (*extra decimal place kept to maintain the uncertainty*)</p> <p>Next, $C = \frac{n}{V}$ can be used to determine the concentration uncertainty.</p> <p>The volume used to produce the solutions was $0.1000 \pm 0.0005L$.</p> $C = \frac{0.125002 \text{ mol}}{0.1000 L} \pm \left(\frac{0.000004 \text{ mol}}{0.125002 \text{ mol}} + \frac{0.0005 L}{0.1000 L}\right) \cdot 1.25002 M$ <p>$C = 1.250 \pm 0.006 M$ (Only 4 significant figures and 3 decimal places for the concentration can be kept)</p>	<p>Percent Uncertainty Calculation</p> <p>Using the absolute uncertainty calculated for 1.25M zinc nitrate as an example:</p> $\% \text{Uncertainty} = \frac{\text{absolute uncertainty}}{\text{value}} \cdot (100\%)$ $\% \text{Uncertainty} = \frac{0.006M}{1.25M} \cdot (100\%)$ $\% \text{Uncertainty} = 0.48\% \rightarrow \underline{0.5\%}$ <p>(Rounded to one significant figure)</p> <p>Average Voltage Calculation (1.25M)</p> $\text{Average Voltage} = \frac{1.01 + 0.99 + 0.96 + 1.02 + 0.84}{5}$ $= 0.97 V$ <p>The uncertainty of the average is still $\pm 0.01V$ since all of the voltages have an uncertainty of $\pm 0.01V$.</p>
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Figure 5: Sample calculations for how concentration uncertainty was determined, how average voltage was determined, and how percent uncertainty was determined

Processed Data

There are three different manipulations of data that will be conducted to see trends between the variables. The first is a graph comparing the experimental values and theoretical values of the experiment. The second is a graph that demonstrates the linear relationship between the natural logarithm of the concentration of zinc nitrate and the initial voltage produced. Finally, the third is a table showing the percent error of the experiment in comparison to the theoretical values.

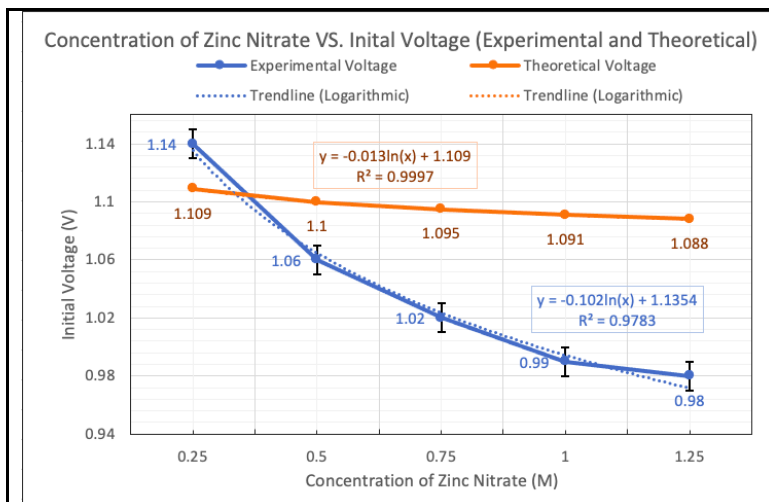


Figure 6: Graph of the concentration of zinc nitrate against the initial voltage, showing experimental and theoretical data

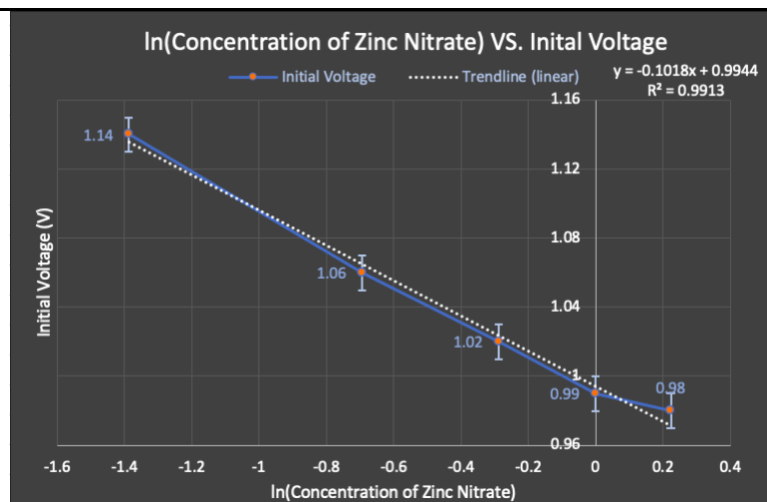


Figure 7: Graph of the $\ln(\text{concentration zinc nitrate})$ against the initial voltage produced in the experiment

Concentration (M)	Percent Error	Sample Calculation for 1.25M zinc nitrate $\% \text{ Error} = \frac{ \text{Experimental} - \text{Theoretical} }{\text{Theoretical}} \cdot 100\%$ $\% \text{ Error} = \frac{ 0.98V - 1.088V }{1.088V} \cdot 100\%$ $\% \text{ Error} = 2.8\%$ <i>*Numerator evaluates to 0.108, however only 2 decimal places can be kept, so it is 0.11. Therefore only 2 significant digits can be kept in the final answer*</i>
0.25M	2.8%	
0.50M	3.6%	
0.75M	6.8%	
1.00M	9.3%	
1.25M	9.9%	
Average	6.5%	

Figure 8: Percent error for each concentration and a sample calculation using the 1.25M zinc nitrate Concentration

Discussion

In Figure 6, the relationship between the concentration of zinc nitrate and the initial voltage produced can be seen. The overall trend that is demonstrated through both the theoretical and experimental data is that as the concentration of zinc nitrate increases, the voltage produced by the voltaic cell decreases according to a logarithmic trend. The R^2 value for the experimental line is 0.9783, providing evidence that there is a strong negative logarithmic correlation between the initial voltage and the concentration of zinc nitrate, due to how close the value is to 1. This relationship is expressed through the equation $-0.102\ln(x) + 1.1354$. The theoretical data of course has an R^2 value that essentially equals 1, and the equation of its trendline is $-0.013\ln(x) + 1.109$. From the graph, there is a clear discrepancy however between the experimental and theoretical data. Even though the experimental data follows an almost perfect logarithmic relationship, the decrease in voltage for each concentration increment is much greater for the experimental values than it is for the theoretical values. This is supported by the fact that the absolute value of the slope for the experimental trend line is significantly larger (by almost a factor of 10) than the absolute value of the slope for the theoretical trend line. Another thing to consider is the presence of an outlier, in trial 5 of the 1.25M concentration (Figure 4). The value of 0.86V is notably lower than the other values in the 1.25M trials, which was

likely caused due to an error in producing the solutions for that specific trial. Moving on to *Figure 7*, a very strong negative linear relationship can be seen between the natural logarithm of the concentration of zinc nitrate and the initial voltage produced. Mathematically speaking, this was an expected result since in the introduction when the Nernst equation was linearized, it was noted that plotting $\ln[\text{Zn}(\text{NO}_3)_2]$ against initial voltage would produce a linear correlation due to the nature of the relationship between $[\text{Zn}(\text{NO}_3)_2]$ and the cell potential. The R^2 value is exceptionally close to 1, with a value of 0.9913, providing experimental evidence that the relationship between $\ln[\text{Zn}(\text{NO}_3)_2]$ and cell potential is indeed a strong negative linear correlation, supporting the mathematical expectations. However, the equation of the trendline is $-0.1018x + 0.9944$, whereas it was expected to be $-0.0128x + 1.109$, showing a significant disparity. This difference is because the theoretical data decreases as a slower rate than the experimental data as stated earlier, but the reason as to why this occurs will be discussed later. Finally, *Figure 8* shows the average percent error of the experiment; the extent by which the experimental data differed from the theoretical data that was calculated. On average, the percent error between the experimental and theoretical data was 6.5%, which is relatively low, but in terms of this experiment specifically is quite high. This experiment used the Daniell cell, whose variables can be controlled more easily compared to other voltaic cells such as the porous cup cell, which runs into problems such as a mixing of the two electrolyte solutions (Foundation, 2014). Therefore, there must have been some substantial errors that were made either in producing solutions or in keeping one of the controlled variables properly controlled. It is also seen how the percent error gets larger and larger as the concentration increases which further supports the argument that the experimental data decreases at a more significant rate than the theoretical data.

Using the processed data, the research question can be answered. As the concentration of zinc nitrate increases, the initial voltage produced decreases. First let's take a look at the mathematical reason for this relationship. It is known that $E = E^\circ - \frac{RT}{nF} \ln(Q)$, where:

$$E^\circ = 1.10\text{V}$$

$$R = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

$$n = 2$$

$$F = 96,485 \frac{\text{C}}{\text{mol}}$$

$$T = 298 \text{ K}$$

Additionally, the $\frac{RT}{nF}$ term is known to be 0.0128V. This reduces the equation to $E = 1.10\text{V} - 0.0128 \text{ V} \cdot \ln(Q)$.

Recalling that $Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$ ($[\text{Cu}^{2+}] = 0.50\text{M}$), as $[\text{Zn}(\text{NO}_3)_2]$ increases, $[\text{Zn}^{2+}]$ increases, meaning Q increases. This means the overall E (voltage) decreases accordingly, since the $0.0128\text{V} \cdot \ln(Q)$ term gets larger. Therefore, speaking strictly in mathematical terms, the results that were shown make complete sense. However, there is a rigorous chemistry-related reason as to why this trend is observed. The reasoning lies in the concept of equilibrium. An increasing Q value means that the system is getting closer to equilibrium since the products concentration is increasing and the reactants concentration is decreasing. Essentially, as the concentration of Zn^{2+} is increased, the reaction starts at a point closer to equilibrium. Referring to the background information, at equilibrium, $\Delta G = 0$, so at a point before equilibrium, $\Delta G < 0$ since the reaction is spontaneous. Specifically, ΔG is the change in free energy of a system as it goes from the initial state, the start of the reaction, to the final state, the system reaching equilibrium ($\Delta G = G_{\text{Final}} - G_{\text{Initial}}$). Therefore ΔG tells us the maximum amount of usable free energy in the system as it proceeds from the initial to final state. G_{final} is constant for this reaction since the final state is always equilibrium. However, the magnitude of ΔG for a reaction decreases as it gets closer to equilibrium, meaning the amount of free energy available in the initial state, G_{initial} , must decrease (Bodner Research Web, n.d). What this means is that since increasing the concentration of Zn^{2+} makes the reaction start closer to equilibrium, increasing the concentration of the $\text{Zn}(\text{NO}_3)_2$ solution decreases the initial voltage due to there being less free energy available to be used in the initial state of the system. This logic adds up in terms of Le Chateliers principle as well, which states that if a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium shifts to counteract the change to re-establish an equilibrium (Anne Marie Helmenstine, 2019). Assuming the redox reaction is held at equilibrium, the impact of the stresses of increasing or decreasing the concentration of products (Zn^{2+}) can be determined. If the concentration of Zn^{2+} is decreased in the equilibrium constant expression, the reaction would be driven in forward direction, meaning spontaneity and therefore cell potential increases. On the other hand, if the concentration of Zn^{2+} is increased in the equilibrium constant expression, the reaction would be driven in the reverse

direction, meaning spontaneity and therefore cell potential *decreases*, suggesting that the experimental data and trend observed was perfectly reasonable. Moving on, it was stated earlier that the experimental data decreased faster than the theoretical data, but a reason was not given as to why this occurs. This discrepancy was due to the flaws in the method that were originally not considered while doing this experiment. It was missed in the calculations that zinc nitrate and copper sulfate are both hydrated materials. What should have been done was either a calculation of how much less water was needed so that when the powders dissolved, the solution volume ended up back at 100mL, or the powders should have been desiccated so that they were anhydrous. Additionally, the data could have been affected by a change in the salt bridge concentration due to the potassium nitrate precipitating out, or by the dimensions of the salt bridge which were not kept constant. These factors are most likely the reason why the experimental data and theoretical data had different rates of change and later in the conclusion, the extent of these factors impact on the data will be evaluated.

The final thing to consider is the uncertainties of this experiment, and how significantly they affect the seen results. The uncertainties that were found in raw data section came out to a maximum of $\pm 0.006\text{M}$ for the concentration uncertainty, and $\pm 0.01\text{V}$ for the voltage uncertainty. These are miniscule uncertainties, and they have a negligible impact on the experimental values. For example, the concentration uncertainty being just $\pm 0.006\text{M}$ means that any error related to solution concentration has to be caused by systematic or random errors, as the uncertainty is too low to impact results. Additionally, the voltmeter uncertainty can only be used to explain an error of $\pm 0.01\text{V}$, and since none of the experimental values are within the $\pm 0.01\text{V}$ of the theoretical values, it can be deduced that there were other errors outside of equipment-related errors involved. The uncertainty themselves were shocking and show how precise the measuring equipment used was, however, it also shows that there was a large percentage of error that was due to flaws in the experimental method. In *Figures 6* and *7*, the absolute uncertainties of the concentrations were used to form the horizontal error bars, and the absolute uncertainties of the voltages were used to produce the vertical error bars. The small size of the uncertainties is why only vertical error bars can be seen; the horizontal error bars were too small to be visible so they provided no meaning even though they were present. The vertical error bars however were crucial to the graph as this experiment produced a range of data limited to 0.16V , so an uncertainty of $\pm 0.01\text{V}$ was clearly visible and important to consider. Overall, the affect on the processed and raw data due to the uncertainties was negligible, meaning that the uncertainties were *not large enough to account for the percent error of 6.5%*.

Conclusion and Evaluation

Conclusion

This investigation aimed to explore how the concentration of zinc nitrate in a redox reaction with 0.50M copper sulfate, affected the voltage produced by a voltaic cell. In *Figure 6*, although it was clear that the experimental values were lower and changed at a different rate than the theoretical values, both sets of data indicated a strong negative logarithmic correlation between concentration and initial voltage, with an R^2 value of 0.9783 and 0.9997 respectively. Additionally, in *Figure 7*, there was a strong negative linear correlation between the natural logarithm of the concentration of zinc nitrate and the initial voltage of the voltaic cell, supported by an R^2 value of 0.9913 . Therefore, all the data collected points to the same conclusion; that as the concentration of zinc nitrate increased, the voltage produced by the cell decreased according to a logarithmic trend. This trend was successfully explained both mathematically using the Nernst equation, and chemically using the concept of equilibrium and Le Chateliers principle, where it was concluded that increasing the concentration of Zn^{2+} ions would lead to the system starting closer to equilibrium, decreasing the amount of free energy available to produce a voltage.

A paper written by UK Essays in 2021 titled Factors Affecting the Voltage of Electrochemical Cells does a similar experiment in which one of the sections is how changing the concentration of the cathode solution affects the voltage produced by the voltaic cell. For whatever reason, there was quite literally no research that could be found online that used the anode solution as an independent variable, and the single one that did use the anode solution as the independent variable was blocked by a paywall and did not make use of zinc nitrate. Regardless, one point that was useful from UK Essays' paper was an important factor that was overlooked when doing the experiment, the surface area of the salt bridge. While doing the experiment, the paper strips were made without taking any

measurements of length or width, they just needed to be long enough to fit. However according to UK Essays, an increase in surface area can change the voltage by up to 0.05V, which is significant on the scale of this experiment.

Finally, with all of the effort put into this experiment, are the end results trustworthy, exactly how reliable are these results, and what could have been done to reduce the errors that occurred? On a positive note, there were many strengths of this experiment, such as the tiny concentration uncertainties and how easy it was to repeat the experiment in an almost identical manner to a previous trial. These strengths shine when looking at the ridiculously accurate R^2 values this experiment produced (0.9783 and 0.9913 in *Figures 6* and *7*). Additionally, there were essentially no errors that could be accounted to the equipment. The equipment uncertainty was so low that a difference in voltage due to these errors may not have even been picked up by the multimeter since it only measured to two decimal places. The multimeter itself also had a small uncertainty ($\pm 0.01V$) which further increased the precision of this experiment. In terms of the weaknesses that were present in the method, the table below summarizes all of the possible *random* and *systematic* errors that may have occurred during the experiment.

Sources of Error	Severity	Possible Improvements
Systematic Errors		
Not taking into consideration the fact that zinc nitrate and copper sulfate are hydrated	High Severity: Using 1.25M as an example, the sample calculation below shows how much the concentration was impacted by this error. The molar mass of $Zn(NO_3)_2$ anhydrous is $189.36 \frac{g}{mol}$, and the molar mass of the hydrated form is $235.50 \frac{g}{mol}$, meaning that the mass percent of the hydrated version that is actually zinc nitrate is 80.41%. Since 29.438g of hydrated powder was used, this means that 23.671g of it was zinc nitrate. This is 0.125mol of zinc nitrate, the amount that was required for a 0.100L solution. However, the remaining 5.767g amounts to 5.767mL of water, which changes the volume to $\sim 0.106L$. Using $C = \frac{0.125mol}{0.106L}$, the concentration of the solution was in fact 1.18M, which is a 5.6% error from the expected solution which was supposed to be 1.25M. This error slightly increased the cell potential measured.	There are two ways to solve this problem. First, by calculating how much less water needs to be added so that the water in the zinc nitrate molecule brings the solution volume back to 100mL. Second, by desiccating the powders through heating to evaporate off all the water. The same improvement should be applied to the hydrated copper sulfate powder as well.
The Surface area of the tissue salt bridge used	Medium Severity: According to the study done by UK Essays, the doubling of surface area increases the voltage by 0.05V. Since the tissue paper strips were just ripped by hand, without any width measurements taken, an error as large as a tripling of the surface area is possible. The size of the salt bridge used is a potential explanation to why the experimental data was lower than the theoretical data, as the theoretical formula does not account for salt bridge flaws in the form of resistance.	This error can be easily mitigated by simply being more careful while preparing the tissue strips and cutting them out into equal dimensions for length and width. For example, a 20cm length and a 3cm width.
The salt bridge solution precipitating out some potassium nitrate	Low Severity: In the potassium nitrate solution, precipitation of the potassium nitrate would decrease the concentration of the solution. This has a direct impact on the conductivity of the salt bridge. Once again referencing UK Essays, a decrease in the salt bridge concentration decreases its conductivity, which increases the resistance. This thereby decreases the voltage produced, which is likely to be one of the reasons why the voltage for the experimental values were lower than those of the theoretical data.	A way to solve this error is by carefully watching the solution before every trial and stirring if there is precipitate. Another less efficient way to reduce the impact of this error is to remake the solutions after every 5 trials to ensure that a fresh solution is used when starting each new concentration interval.

Sources of Error	Severity	Possible Improvements
Random Errors		
Temperature Fluctuations	Low Severity: As seen in the Nernst equation, temperature does impact the cell potential. Regardless, the temperature was always in the range of 23 - 25°C, which would only slightly affect the hundredths decimal place of the data, so these fluctuations do not hold much significance.	If an extremely precise temperature was needed for each trial, a temperature controller could be used, but that is out of the scope of this experiment
Drift of Electronic Measuring Devices	Very Low Severity: Drift is essentially the change in measurements over time where readings become higher or lower due to external factor such as temperature and even electrical noise (Anne Marie Helmenstine, 2020). This could impact the mass that was measured out; however, this error is already considered in the uncertainty.	There is not a good way to solve this problem other than getting even more precise equipment. However, for an experiment of this scale, drift error does not matter too much.

Figure 9: Table showing the possible sources of error, their severity and impact, and how to improve the method to mitigate these errors if the experiment was to be repeated

Future Work and Extensions

Firstly, a basic extension of this experiment would be to see the effect of other factors on the voltage produced by the voltaic cell. Examples include changing the concentration of the cathode solution, changing the surface area of the salt bridge, changing the temperature of solution, or changing the concentration of the salt bridge. However, something else that was found to be intriguing while conducting the experiment was the idea of connecting this system to a lightbulb or a motor in a circuit, actually harnessing the electrical energy. For a light bulb, it would be interesting to experiment with different factors, such as concentration changes. Of course, given a voltage and current, the light bulb will glow, and using a photometer its possible to see how different concentrations of electrolyte affect the intensity of the light emitted by the light bulb. The rate at which the brightness of the light bulb decreases could also be used to find the rate law expression of the redox reaction under investigation. There are also an innumerable amount of different applications of electrochemistry in real life as well, other than the application regarding batteries that was already discussed. For example, electroplating. Anyone with jewellery has most likely had it electroplated, which is a coating of the metal with a thin layer of another metal. This process makes use of an electrolytic cell, which is essentially the opposite of the voltaic cell in the sense that it converts electrical energy to chemical energy, and this starts a reaction where the anode metal is plated (deposited) onto the cathode metal (Siyavula, n.d). Finally, arguably one of the most ground-breaking uses of electrochemistry in general is with regards to medicine. Electrochemical treatment of tumours (EChT), also known as electrochemical therapy, is a type of treatment where tumorous tissue is treated with a current using two electrodes that are placed within the tumour. The electrical energy that is produced when the tissues is electrolysed is converted into the chemical energy, and this chemical energy in simple terms destroys the cancer cells. Although it lacks clinical data, it is a simple, safe, and non-invasive anti-tumour therapy according to a paper written by Nilsson et. Al published in the national centre for biotechnology information (NCBI). Overall, this experiment helped me expand my understanding of electrochemical cells, redox reactions, and the scientific method considerably. I truly believe that the interest in electrochemistry this experiment has developed in me will be of great benefit in my future pursuits of being a biotechnology researcher, as many applications of electrochemistry in medicine such as the EChT can already be seen.

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