Investigating the effect of concentration of the electrolyte on the electrode potential of a Copper-Zinc voltaic cell

Introduction

My fascination with Elon Musk's ideas of energy triggered me to investigate the change in electrode potential with different concentrations of the electrolyte. However, it wasn't Tesla or SpaceX that truly intrigued me. Though Musk is seen to be leading the path of alternative energy, William Robert Grove, who came up with the Grove voltaic cell *(Mengel, 2020)*, is taking the world by storm. The fascinating concept about this cell was that it did not use any scarce resources to do this. It was a simple zinc and platinum electrode. People all around the world are amazed by this gas voltaic cell, and how competitive it is against Musk's Tesla. With Bill Gates and Jeff Bezos investing millions of dollars into this technology, it is clear that it is extremely powerful. I wondered how Grove could maximize the amount of energy that was made, using such simple resources.

From this, I formulated the research question: **What effect does the decrease in the concentration of zinc sulphate and copper sulphate electrolytes have on the electrode potential of the Copper-Zinc voltaic cell, and how does this affect Gibbs free energy?** This relates to the IB Chemistry topic 1 Quantitative Chemistry and topic 9 Redox Processes, specifically Voltaic cells. I am able to use theoretical data using the Nernst equation.

Background Research

Daniell Cell
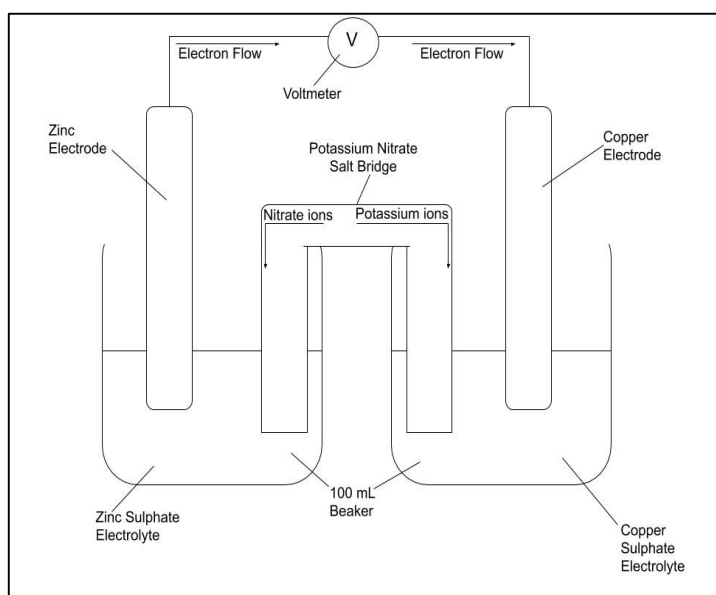
The zinc electrode (Anode) loses electrons from zinc atoms and forms Zn^{2+} ions, while the copper electrode (Cathode) provides electrons to Cu^{2+} ions to form copper metal. As electrons flow from an area of high concentration to an area of low concentration, the electron flow is from the anode to the cathode. This is shown in the diagram I made to the left *(figure 1)*

The two half equations for this voltaic cell are:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \ (0.76V)$$
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \ (0.34V)$$
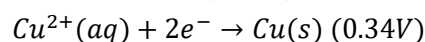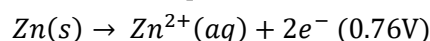
Thus, the redox reaction is:

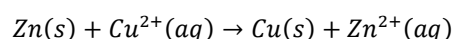$$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$$

Figure 1 showing the Copper-Zinc voltaic cell

Factors that Affect Voltage of a Battery

Resistance in the internal and external Circuit *(Bylikin et al., 2014)*

Both of the electrodes used in the experiment are made out of metal. This means they consist of a regular lattice of metal ions with a sea of delocalized, randomly moving, electrons around them. When the copper and zinc electrode are connected to each other, there is another more regular motion added to the random motion of the electrons. Through this, a general 'drift' of electrons through the external circuit from the anode to the cathode occurs. There are collisions between these delocalized electrons and positive ions which obstruct the electrons' path and hinder their movement through the metal. This creates an external resistance in the circuit.

Within the battery, the current is carried by ions. These ions experience a similar internal resistance as it takes a certain amount of time for them to diffuse in the electrolytes

The potential difference between two points in a circuit is a measure of energy given out as each unit of charge passes between two points:

$$Potential\ Difference = \frac{Energy}{Charge}$$

Current is known as the rate of flow of electric charge:

$$Current = \frac{Charge}{Time}$$

The resistance of a component in a circuit is the ratio between the potential difference and the current:

$$External\ Resistance = \frac{Potential\ Difference}{Current}$$

The electromotive force (Voltage) is a measure of the total energy made available by the chemical reactions in the cell per unit charge transferred:

$$Voltage = \frac{Work\ done}{Charge}$$

Not all of this energy is made available to the charge as it passes through the external circuit as some energy is wasted through internal resistance as the ions pass through the cell. The amount that this drops by (voltage drop) is the difference between electromotive force and terminal potential.

In the external circuit:

$$Terminal\ Potential = Current \cdot External\ Resistance$$

In the complete circuit:

$$Voltage = Current \cdot (External\ Resistance + Internal\ Resistance)$$

$$Voltage = Terminal\ Potential + (Current \cdot Internal\ Resistance)$$

Thus:

$$Voltage\ Drop = Current \cdot Internal\ Resistance$$

Through this equation, it can be found that the maximum current will be delivered when the external resistance is 0.

$$Current = \frac{Voltage}{External\ Resistance + Internal\ Resistance}$$

Thus:

$$I_{max} = \frac{Voltage}{Internal\ Resistance}$$

Concentration of the Electrolytes

The equation of which relates Gibbs free energy (ΔG) with the equilibrium constant (K) is:

$$\Delta G^{\theta} = -R \cdot T \cdot \ln(K)$$

$R(molar\ gas\ constant) = 8.314 Jmol^{-1}K^{-1}$

The temperature (T) of the cell was taken as 24°C, thus T=297K

The equation which describes the behaviour of ΔG as the reaction proceeds is:

$$\Delta G = -(R \cdot T \cdot \ln(K)) + (R \cdot T \cdot \ln(Q))$$

$$\Delta G = \Delta G^{\theta} + (R \cdot T \cdot \ln(Q))$$

Furthermore, in topic 9, ΔG^{θ} can be found through the equation:

$$\Delta G^{\theta} = -n \cdot F \cdot E^{\theta} cell$$

This can be represented in non-standard conditions as:

$$\Delta G = -n \cdot F \cdot Ecell$$

Where n is the moles of the electrons in the reaction and F is the Faraday's constant.

$F = 9.65 \cdot 10^{4} Cmol^{-1}$

Using this and the equation above:

$$-n \cdot F \cdot Ecell = \left(n \cdot F \cdot E^{\theta} cell\right) + (R \cdot T \cdot \ln(Q))$$

This can be used to derive the Nernst equation:

$$Ecell = E^{\theta} cell - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$$

Using the IB chemistry data booklet *(Haynes and Lide, 2012)*, the standard electrode potential for this voltaic cell can be calculated to 1.10V.

As the moles of electrons is 2 in the equation for the voltaic cell, n=2

$$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

Thus, as I change the concentration of the electrolytes, Q changes. As the concentration of the zinc sulphate ($ZnSO_4$) increases, Q increases, however, the relationship is inversely correlated between Q and copper sulphate ($CuSO_4$).

Hypothesis

My hypothesis is that as the concentration of the $CuSO_4$ decreases, the electrode potential will decrease, however, as the concentration of the $ZnSO_4$ decreases, the electrode potential will increase. This is because, as I increase the concentration of the zinc ions, the equilibrium of the Zinc half equation to the left and cause lesser ions to form. This will decrease charge separation, and thus the electrode potential. However, as I increase the concentration of the copper ions, the equilibrium of the Copper half-cell will shift to the right, thus creating a higher concentration of ions at the cathode, compared to the anode. Thus, lesser electrons will flow through the wire, allowing for a lower electrode potential.

Variables

Independent

I will change the concentration of Cu^{2+} and Zn^{2+} electrolyte by creating 200cm^3 of different concentration standard solutions (1.0, 0.8. 0.6. 0.4, 0.2moldm^{-3}). The method used for making these standard solutions was adapted from Quantitative Chemistry (topic 1). There will be a sufficient volume for 4 repeats per concentration. In order to prepare solutions of a higher concentration, a higher mass of the solutes has to be dissolved in the same volume of the solvent. This solvent may not be able to dissolve all of the solute to prepare higher concentrations. Thus, it is more difficult to prepare solutions of higher concentrations, as the solute will not completely dissolve in the solvent.

Dependent

I will measure the electrode potential of the voltaic cell using a voltmeter connected to the electrodes in my cell. This voltage will represent the electrode potential under non-standard conditions. Hence, I will be able to compare this voltage to the theoretical voltage obtained using the Nernst equation. Furthermore, I can calculate the ultimate variable: the Gibbs free energy for each differing concentration of electrolyte.

Controlled

- Once the $ZnSO_4$ and $CuSO_4$ solutions are prepared, the solutions will be kept for 15 minutes in a temperature-regulated room (at 24°C) before the start of the experiment, to make sure the temperature of both the solutions were the same (297K), which will be measured using a thermometer.
- The dimensions and the type of electrode were kept the same. This was done by cutting the strip of the metal the same dimensions and using the same strips throughout the entire experiment by storing them in a basket labelled with my name and asked to not be touched by other students. This would help keep the moles electrons the same (2 mol) in the Nernst equation.
- I will control the size of the salt bridge used by cutting the filter paper in the same dimensions. This will allow me to effectively control the number of ions on the salt bridge to the best of my abilities. This is because I will have soaked the filter paper in excess saturate concentration potassium nitrate
- I will immerse more than half of the electrode and salt bridge to ensure that the same surface area of the electrode and salt bridge was in the solution each time.

- I kept the voltmeter used throughout the experiment the same by keeping my voltmeter in a basket that was labelled with my name and asked not to be touched by other students. This was done to consider systematic error.
- The volume of electrolyte used in every repeat was kept at $50cm^3(\pm0.05cm^3)$. This was done by using a burette to measure and pour the electrolytes into the beaker. This would have helped me ensure the same number of ions were provided to each electrode. However, both the solutions were in excess of ions.
- I used the same mass balance to measure the solid $CuSO_4$ and solid $ZnSO_4$. This was done by measuring all the masses needed for my solutions on the same day with the same mass balance. Furthermore, the surface that the mass balance was on was kept the same. This was to make sure there was no systematic error when measuring the mass.

Method

Chemicals	Apparatus
• 226.05g solid Zinc Sulphate $CuSO_4$) • 223.45g solid Copper Sulphate ($ZnSO_4$) • $3600cm^3$ Distilled Water • $500cm^3$ Saturated solution of Potassium Nitrate (KNO_3)	• (8cmx2cm) (0.25mm thickness) Zinc Metal Strip (99.99% pure) • (8cmx2cm) (0.1mm thickness) Copper Metal Strip (99.9% pure) • Sandpaper • Electronic Mass Balance ($\pm0.001g$) • $1000cm^3$ Volumetric Flasks ($\pm0.4cm^3$) • $100cm^3$ Volumetric Flasks ($\pm0.08cm^3$) • 50mL Burette ($\pm0.05mL$) • 100mL Beakers • 250mL Beakers • Conducting wires • Crocodile Clips (X2) • Voltmeter (X1) ($\pm0.002V$) • (8cmx2cm) Absorptive Paper (x36) (used as a salt bridge) • Thermometer (X1) (±0.05ºC) • Stirring Rod • Spatula

Table 1 showing the equipment and chemical list

There were no ethical factors in this experiment. However, the volume of distilled water required to make the standard solutions was quite high. Solid $CuSO_4$ is known to be very toxic to aquatic life (CLEAPSS, 2019), thus the solution was fully dissolved and disposed of, with plenty of water. The same method was used for $ZnSO_4$, as Zn^{2+} is a heavy metal ion, which can cause environmental and health issues if in excess.

Risk Assessment adapted from CLEAPSS (CLEAPSS, 2019)

Hazard	Risk	Comment
$1moldm^{-3}$ $CuSO_4$	CORR. IRRITANT ENVIR*.	Skin irritant; cause serious eye damage; harmful if swallowed. *Solid copper sulphate is dangerous to aquatic life.
$1moldm^{-3}$ $ZnSO_4$	CORR. IRRIT.	Causes severe skin burns and eye damage.

Saturate concentration KNO_3	OXID. IRRITANT	Harmful if swallowed (more than 3moldm⁻³) and irritating to eyes and skin. (more than 1moldm⁻³)

Table 2 showing the hazards and risks of this experiment

Thus, safety goggles, gloves and an apron will be worn when these chemicals are being used. If chemicals do come in contact, the area is immediately rinsed.

Preparing The Salt Bridge
1. Cut out 36 pieces of absorptive paper with the dimensions 8cmx2cm
2. The saturated solution of Potassium Nitrate was taken in a 250 mL beaker
3. Place the absorptive paper pieces in the beaker

Preparing The Standard Electrolyte Solutions
1. Tare the mass balance with a plastic weighing boat on it
2. Weigh out different masses of the solid $CuSO_4$/$ZnSO_4$ to prepare solutions of known concentration, as shown below:

Electrolyte	Volume wanted (cm³)	Concentration wanted (moldm⁻³)	Mass used (g)
	1000	1.0	161.47
	200	0.8	12.916
$ZnSO_4$	200	0.6	9.678
	200	0.4	6.687
	200	0.2	3.229
	1000	1.0	159.61
	200	0.8	12.768
$CuSO_4$	200	0.6	9.576
	200	0.4	6.687
	200	0.2	3.229

Table 3 showing the mass used with volume and concentration of each electrolyte in the experiment

3. Place the weighed amount of the solid into a 250 mL beaker and pour some distilled water into it
4. Wash out the plastic weighing boat with distilled water into the beaker to ensure all the solid is transferred into the beaker
5. Use a stirring rod, stir the solution to dissolve the solute
6. Add minimum quantities of distilled water, while stirring, to make sure all the solid is dissolved
7. Transfer (using a funnel and glass rod) the solution into the volumetric flask (200cm³ or 1dm³ dependent on volume wanted)
8. Wash out the beaker, stirring rod and funnel with minimum quantities of distilled water into the volumetric flask to ensure that the solution is fully transferred into the volumetric flask
9. Add sufficient distilled water and then, using a teat pipette, the solution is made up to the mark

Preparing The Voltaic Cell
1. Cut out and clean zinc and copper strips with the dimensions 8cmx2cm
2. Place each strip on a 100mL beaker as shown in *figure 2*
3. Transfer 50cm³ of 1moldm⁻³ $ZnSO_4$ solution using a burette into the beaker with the Zinc strip and make sure that more than half of the strip is immersed in the solution
4. Transfer 50cm³ of a prepared 1moldm⁻³ $CuSO_4$ solution using a burette into the beaker with the Copper strip and make sure that more than half of the strip is immersed in the solution

5. Place one of the absorptive paper pieces between the two beakers and make sure that a part of the absorptive paper is immersed in the electrolyte, as shown in *figure 2*
6. The zinc strip and copper strip were connected using conducting wires through a voltmeter
7. Measure the voltage of the cell using the voltmeter
8. Repeat steps 3-7 with the 0.8, 0.6, 0.4, 0.2moldm^{-3} of ZnSO$_4$, keeping the concentration CuSO$_4$ as 1moldm^{-3}
9. Repeat steps 3-7 with 0.8, 0.6, 0.4, 0.2moldm^{-3} of CuSO$_4$, keeping the concentration ZnSO$_4$ as 1moldm^{-3}
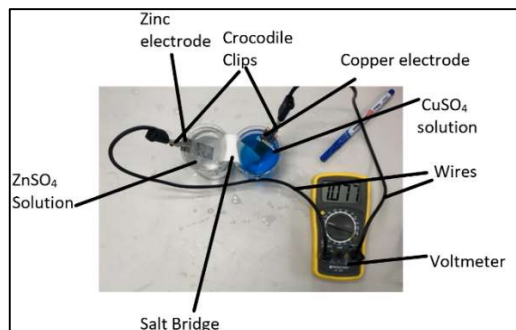

Figure 2 showing the voltaic cell with electrolytes 1moldm^{-3} ZnSO$_4$ and 0.8moldm^{-3} CuSO$_4$

Method Rationale

My method for setting up my voltaic cell was adapted from what I learned in Redox (topic 9) of IB Chemistry. The equation to find moles is:

$$Moles\ (mol) = \frac{Mass\ (g)}{Molar\ Mass\ (gmol^{-1})}$$

Moles can also be found using the concentration and volume required:

$$Moles\ (mol) = Concentraion\ (moldm^{-3}) \cdot Volume(dm^3)$$

Thus, the mass required to make the different concentration solution can be found using:

$$Mass(g) = Concentraion\ (moldm^{-3}) \cdot Volume(dm^3) \cdot Molar\ Mass\ (gmol^{-1})$$

Sample calculation to prepare the standard solution of 1dm^3 1moldm^{-3} of ZnSO$_4$:

$$Mass = 1 \cdot 1 \cdot 161.47$$
$$= 161.47g$$

Thus, I needed 161.47g of solid zinc to get 1mol of zinc.

I decided to use a 100mL beaker and fill it up halfway (using 50cm^3 of solution) to ensure that half of the electrode would be fully submerged in the solution and that the salt bridge could be connected to both solutions. I decided to make standard solutions using solid CuSO$_4$ and ZnSO$_4$, rather than using stock solutions provided to us, to ensure that I am using the right concentration.

I found that the uncertainty of the voltmeter was ±0.001V, however, upon further inspection of my results, I found that some of my repeats had the half range of 0.002V (For example, the voltmeter readings for 0.6moldm^{-3} CuSO$_4$ and 1.0moldm^{-3} ZnSO$_4$). Thus, I decided to use the uncertainty of ±0.002V to ensure that any random error could be considered.

Results

Raw Data

Conc of CuSO$_4$ (±0.2 moldm^{-3})	Conc of ZnSO$_4$ (±0.2 moldm^{-3})	Voltmeter reading (± 0.001V)				
		Repeat 1	Repeat 2	Repeat 3	Repeat 4	Average
1.0	1.0	1.080	1.081	1.081	1.079	1.080
0.8	1.0	1.077	1.076	1.076	1.077	1.077
0.6	1.0	1.073	1.071	1.074	1.072	1.073
0.4	1.0	1.068	1.067	1.067	1.069	1.068
0.2	1.0	1.059	1.060	1.058	1.058	1.059

1.0	0.8	1.082	1.082	1.083	1.083	1.083
1.0	0.6	1.085	1.087	1.088	1.086	1.087
1.0	0.4	1.093	1.092	1.092	1.091	1.092
1.0	0.2	1.102	1.099	1.101	1.099	1.100

Table 4 showing Electrode potential when electrolyte concentration is changed

Qualitative Data

1. Before the experiment, the colour of the solutions were noted:
   a. $CuSO_4$- blue in colour
   b. $ZnSO_4$- colourless
2. During the experiment it was observed that:
   a. $CuSO_4$- the colour of the solution faded, indicating a decrease in concentration of Cu^{2+} ions
   b. $ZnSO_4$- the solution remained colourless
3. At the end of the experiment, the observations were the same as during the experiment
4. There were fluctuations in the voltmeter reading while measuring the electrode potential

Processed Data

Sample Calculation of Uncertainties

Percentage Uncertainty in Mass of 1moldm^{-3} $CuSO_4$ and $ZnSO_4$:

$CuSO_4$- $\frac{0.001}{159.61} \cdot 100 = 6.265 \cdot 10^{-4}\%$ $ZnSO_4$- $\frac{0.001}{161.47} \cdot 100 = 6.193 \cdot 10^{-4}\%$

Percentage Uncertainty in Volumetric Flask of 1dm^3 of $CuSO_4$ and $ZnSO_4$:

$CuSO_4$- $\frac{0.8}{1000} \cdot 100 = 0.08\%$ $ZnSO_4$- $\frac{0.8}{1000} \cdot 100 = 0.08\%$

Percentage Uncertainty in Temperature of solutions

$$\frac{0.05}{24} \cdot 100 = 0.208\%$$

Percentage Uncertainty of Voltmeter

$$\frac{0.002}{1.080} \cdot 100 = 0.185\%$$

Total Percentage Uncertainty of 1moldm^{-3} $ZnSO_4$ and 1moldm^{-3} $CuSO_4$

$$6.265 \cdot 10^{-4} + 6.193 \cdot 10^{-4} + 0.08 + 0.08 + 0.208 + 0.185 = 0.555$$

Absolute Uncertainty of Electrode Potential

$$\frac{0.462}{100} \cdot 1.080 = 0.006$$
$$1.080 \pm 0.006V$$

Concentration of $ZnSO_4$ (moldm^{-3})	Concentration of $CuSO_4$ (moldm^{-3})	Electrode Potential of cell (V) (3d.p)	Percentage Uncertainty (%) (1s.f)	Absolute Uncertainty (V) (1s.f)
1.0	1.0	1.080	0.6	0.006
1.0	0.8	1.077	0.5	0.005
1.0	0.6	1.073	0.5	0.005
1.0	0.4	1.068	0.5	0.005
1.0	0.2	1.059	0.5	0.005
0.8	1.0	1.083	0.5	0.005
0.6	1.0	1.087	0.5	0.005
0.4	1.0	1.092	0.5	0.005
0.2	1.0	1.100	0.5	0.005

Table 5 showing the percentage uncertainty and absolute uncertainty of the electrode potential
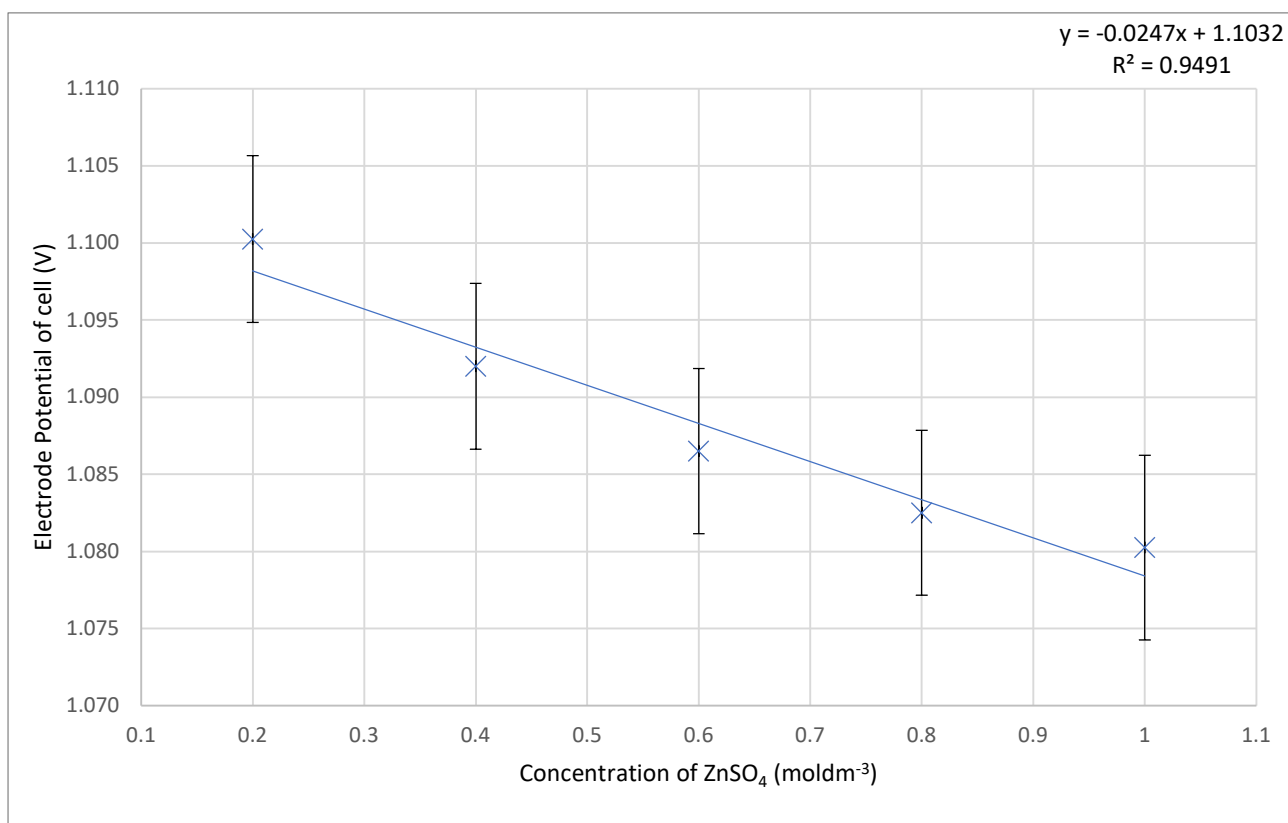
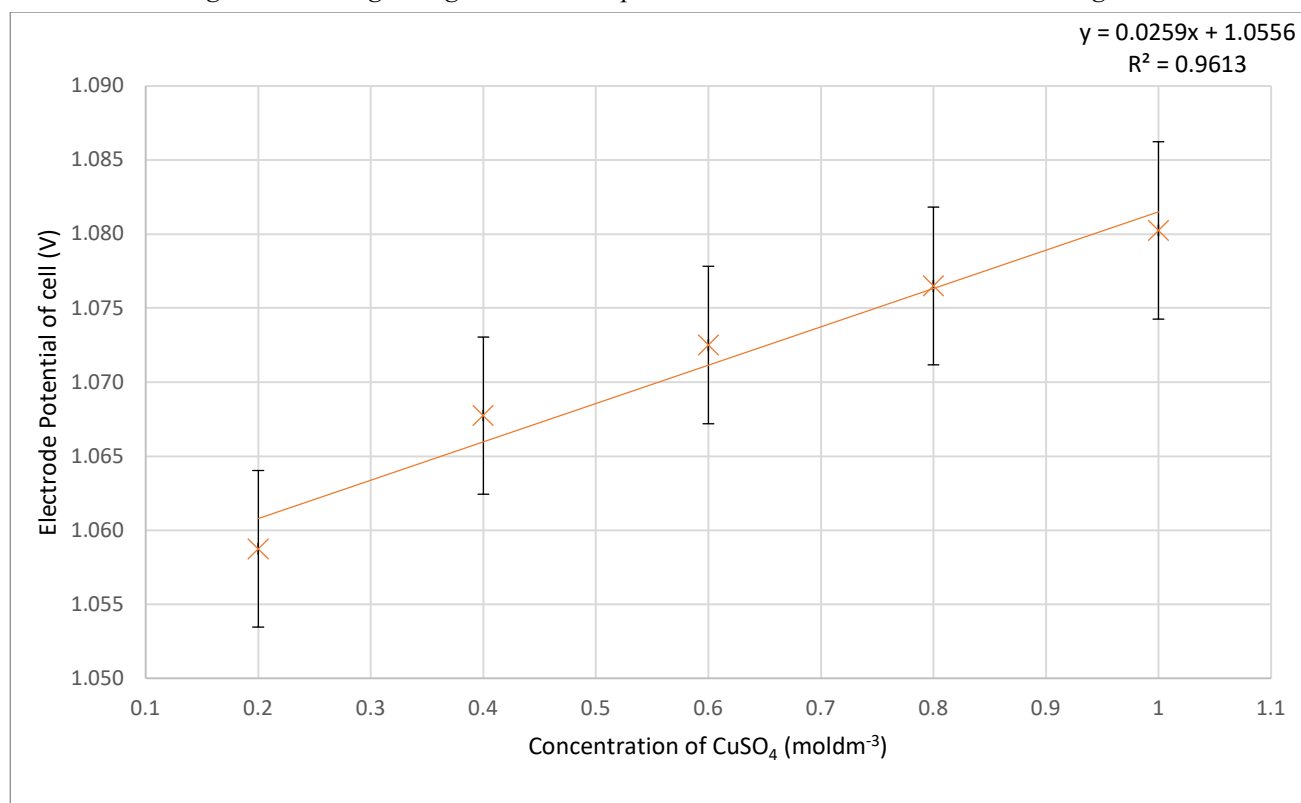*Figure 3 showing change in electrode potential as ZnSO₄ concentration is changed*

Figure 4 showing change in electrode potential as CuSO₄ concentration is changed

Nernst Equation Theoretical Data

Using the previously mentioned Nernst Equation, I am able to obtain theoretical data for my voltaic cells.

$$E^\theta cell = E^\theta ox + E^\theta red$$
$$E^\theta cell = 0.76 + 0.34$$
$$= 1.100V$$

Calculating the electrode potential of the voltaic cell with 1moldm^{-3} ZnSO$_4$ and 1moldm^{-3} CuSO$_4$:

$$Ecell = 1.100 - \left(\frac{8.314 \cdot 297}{2 \cdot 9.65 \cdot 10^4}\right)\left(\ln\left(\frac{1}{1}\right)\right)$$

Thus, Ecell= 1.10V

Concentration of ZnSO$_4$ (moldm^{-3})	Concentration of CuSO$_4$ (moldm^{-3})	Theoretical Electrode Potential of cell (V) (3d.p)
1.0	1.0	1.100
1.0	0.8	1.097
1.0	0.6	1.093
1.0	0.4	1.088
1.0	0.2	1.079
0.8	1.0	1.103
0.6	1.0	1.107
0.4	1.0	1.112
0.2	1.0	1.121

Table 6 showing the theoretical electrode potentials for the different concentrations of electrolytes

$$Percentage\ Error = \frac{Theoretical\ Data - Practical\ Data}{Theoretical\ Data} \cdot 100$$

Using the values for the voltaic cell with 1moldm^{-3} ZnSO$_4$ and 1moldm^{-3} CuSO$_4$ I am able to

$$\frac{1.100 - 1.080}{1.100} \cdot 100 = 1.795\%$$

Thus, I am able to calculate the percentage error for all my data:

Concentration of ZnSO$_4$ (moldm^{-3})	Concentration of CuSO$_4$ (moldm^{-3})	Practical Electrode Potential of cell (V) (3d.p)	Theoretical Electrode Potential of cell (V) (3d.p)	Percentage Error (%) (3d.p)
1.0	1.0	1.080	1.100	1.795
1.0	0.8	1.077	1.097	1.869
1.0	0.6	1.073	1.093	1.876
1.0	0.4	1.068	1.088	1.861
1.0	0.2	1.059	1.079	1.877
0.8	1.0	1.083	1.103	1.859
0.6	1.0	1.087	1.107	1.852
0.4	1.0	1.092	1.112	1.799
0.2	1.0	1.100	1.121	1.851

Table 7 showing the percentage error for each electrode potential

From the electrode potential, I am able to find the Gibbs free energy of the cell.

As seen in the background information, the equation for Gibbs free energy is:

$$\Delta G(kJ) = -n(mol) \cdot F(Cmol^{-1} \cdot Ecell(V)$$
$$\text{Where } n = 2 \text{ } and \text{ } F = 9.65 \cdot 10^4$$

Thus, the calculation for the Gibbs free energy of my voltaic cell with 1moldm^{-3} ZnSO$_4$ and 1moldm^{-3} CuSO$_4$ is:

$$-2 \cdot 9.65 \cdot 10^4 \cdot 1.080 = -208488J$$
$$\Delta G = -208.488kJ$$

From this I am able to form a table for the Gibbs free energy for each voltaic cell:

Concentration of ZnSO$_4$ (moldm^{-3})	Concentration of CuSO$_4$ (moldm^{-3})	Electrode Potential of cell (V) (3d.p)	Gibbs free energy of cell (kJ) (4 s.f)
1.0	1.0	1.080	-208.5
1.0	0.8	1.077	-207.8
1.0	0.6	1.073	-207.0
1.0	0.4	1.068	-206.1
1.0	0.2	1.059	-204.4
0.8	1.0	1.083	-208.9
0.6	1.0	1.087	-209.7
0.4	1.0	1.092	-210.8
0.2	1.0	1.100	-212.3

Table 8 showing the Gibbs free energy as the electrolyte concentration changes

Conclusion

My research question is: **What effect does the decrease in the concentration of zinc sulphate and copper sulphate electrolytes have on the electrode potential of the Copper-Zinc voltaic cell, and how does this affect Gibbs free energy?** My data supports the hypothesis, which stated that, as the concentration of the CuSO$_4$ decreases, the electrode potential will decrease, however, as the concentration of the ZnSO$_4$ decreases, the electrode potential will increase.

The data I found can be seen to be precise, as the R^2 values were both very close to 1, with figure 3 being 0.949 and figure 4 being 0.961. The graphs show a clear difference in correlation between electrodes; thus, I am confident in my conclusion. This is also supported by the theoretical data I found, as my percentage error was all below 1.900%. This suggests that the correlation I found was true.

Figure 3 (page 8) shows an inverse relationship between ZnSO$_4$ concentration and electrode potential. This is due to increased charge separation between electrodes. The half equation suggests more Zn^{2+} ions. As electrons from the anode to the cathode, they require a higher concentration gradient from the copper electrode to the zinc electrode. This decreases the gradient, thus decreasing the electron movement, thus decreasing the electrode potential. Figure 4 (page 8) shows a direct relationship between CuSO$_4$ concentration and electrode potential. This is due to the decrease in the ion concentration gradient between the anode and cathode. This would increase electron movement, and thus electrode potential. My data is supported by the Nernst equation. The values shown in table 3 relate to the same trends found in figure 3 and 4. Even though the electrode potential decreased, the Gibbs free energy remains negative, thus the cells remain spontaneous. However, as the concentration of ZnSO$_4$ decreases, the reaction becomes more spontaneous, and as the concentration of CuSO$_4$ decreases, the reaction becomes less spontaneous.

From this data, I can conclude that it is possible to maximize electrode potential by keeping the concentration of the ZnSO$_4$ electrolyte low and keeping the concentration of the CuSO$_4$ electrolyte high. This could be one of the methods used by companies to ensure maximum efficiency. I am not able to confirm whether Grove used this in industrial work, as his cell is presented in standard conditions.

Evaluation

<u>Strengths</u>

My percentage error was relatively low compared to the theoretical value provided by the Nernst equation. This would suggest that my data was fairly accurate. I had no anomalies in my data, thus, I made sure to change the filter paper (salt bridge) after every repeat to ensure there were an equal number of ions provided in each repeat. This avoided any systematic error, which would have potentially decreased the electrode potential.

<u>Weaknesses</u>

Type of Error	Limitation of Method	Effect on Results	Possible Modification
Systematic	The wire was old and had poor conduction	The 1st repeat of my experiment seemed way too low, and once I changed the wires, I saw a significant increase in electrode potential. This was much closer to the theoretical value, and it was clear that the wire was a better conductor, compared to the other wires. However, these wires could have still had some hindrance in conduction, which would increase my percentage error, as there would be an increase in resistance on the electrons through the wires	Use the same new, conductive wires, rather than the older wires that may be worn out
Systematic	The temperature of the solution was measured before the experiment started, however, the temperature may have varied as the reaction occurred	The Nernst equation states that temperature is a factor that affects the electrode potential. Thus, as the reaction takes place, and heat is released (causing an increase in temperature), the electrode potential would increase	Keep the circuit in a water bath, controlled at 24°C, to ensure that the temperature remained constant throughout the experiment
Random	There was a large fluctuation in data, as visible in my graph. This may have been because of imprecise equipment	This could cause anomalous results or more random errors. It is visible in figure 3 and 4 that many of the voltage uncertainties overlap each other. This overlap could decrease the correlation strength of my data, as it suggests that some of my data points could have been higher, or lower, than the other data points	Use more precise equipment, such as a voltage probe, temperature probe or a mass balance with a higher readability, and use 5 repeats to spot any anomalies
Systematic	The voltmeter may have changed values if left for longer. The voltage of the voltmeter was fluctuating, as I left it to read the circuit for longer	The electrode potential value may have been different from the one I recorded. It would have been higher and allowed for a closer value to the theoretical value	Wait longer than 1 minute, until the fluctuations stop occurring, to ensure that the voltmeter value was the true electrode potential of the circuit
Systematic	The electrodes were not switched between repeats, as there weren't enough copper or zinc strips for 36 repeats	The copper electrode would have had more copper and the zinc electrode would have had more zinc ions. This	Cut out 36 different strips of zinc and copper each, to ensure that the initial conditions of

		could have changed the electrode potential. This would have affected the percentage purity of the electrode. This would have decreased the electrode potential of the circuit	each voltaic cell remained the same
Systematic	I used two different voltmeters in my experimentation. The calibration of this voltmeter may have been different.	The electrode potential values would not have been able to be compared, as it may not be calibrated the same way. This would mean that the voltmeters would measure a different value for the same circuit	Keep the same voltmeter in the basket, and label it to let people know that it is being used by me

Table 9 showing the evaluation comments of this experiment

<u>Extension</u>

It could be possible to see how other factors, such as the number of salt bridges or temperature, affect electrode potential to see how else Grove could maximize electrode potential on the Grove voltaic cell. I could also find the thermodynamic efficiency of the cell, by finding the enthalpy change of the cell, using the formula:

$$\eta = \frac{\Delta G}{\Delta H}$$

Through this, I will be able to further compare whether it is truly most efficient to increase the concentration of the anode and decrease the concentration of the cathode.

<u>Bibliography</u>

Bylikin, S., Horner, G., Murphy, B. and Tarcy, D., 2014. Chemistry: course companion. 2014 edition ed. Oxford IB Diploma programme. Oxford: Oxford University Press.

Chempages Netorials, 2020. *Electrochemical Cells and Voltage*. [Online] Chempages Netorials. Available at: <https://www.chem.wisc.edu/deptfiles/genchem/netorial/rottosen/tutorial/modules/electrochemistry/04voltage/18_41.htm > [Accessed 11 Feb. 2020].

CLEAPSS, 2019a. *Student Safety Sheets*. [Online] Available at: <http://science.cleapss.org.uk/Resource/SSS040-Copper-and-its-compounds.pdf> [Accessed 10 Feb. 2020].

CLEAPSS, 2019b. *Student Safety Sheets*. [Online] Available at: <http://science.cleapss.org.uk/Resource/SSS049-Zinc-and-its-compounds.pdf> [Accessed 10 Feb. 2020].

CLEAPSS, 2019c. *Student Safety Sheets*. [Online] Available at: <http://science.cleapss.org.uk/Resource/SSS034-Sodium-and-potassium-salts.pdf> [Accessed 10 Feb. 2020].

Haynes, W.M. and Lide, D.R. eds., 2012. *CRC handbook of chemistry and physics: a ready-reference book of chemical and physical data*. 93. ed., 2012th–2013th ed. Boca Raton, Fla.: CRC Press.

Mengel, J., 2020. *Has Elon Musk Lost His Mind?* [Online] Available at: <https://www.outsiderclub.com/has-elon-musk-lost-his-mind-/94654> [Accessed 14 Feb. 2020].