

The Effects of the Temperatures of Anode and Cathode Solutions on Cell Potential and the Equilibrium Constant

1. Introduction

1.1: Scientific Question

How does the temperature of each half-cell solution of 0.5M CuSO₄ and 0.5M ZnSO₄ at 303K (30°C), 313K (40°C), 323K (50°C), 333K (60°C), and 343K (70°C), affect the cell potential, E_{cell}, thereby affecting the equilibrium constant?

1.2: Motivation & Aims

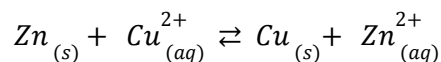
From power grids to large scale infrastructure, electricity has become evolutionary. These innovations always stem back to the root invention that started it all, which is the battery, a process which uses redox reactions to create electrical energy. Sparked by the chemistry of electricity as well as my passion for electrical engineering, I decided to investigate more deeply in this field.

Understanding that batteries have limited power supply made me curious about how we can create a more efficient cell that produces a greater voltage. To further my understanding, I consolidated my passions for geography to question whether hotter climates have better or worse conductance of electricity. This information can be used to create more efficient and powerful cells through a simple variation in temperature. As a result, I have decided to study the effects of the temperature of each of the half-cells (copper and zinc) on the cell potential of a Daniell cell and the equilibrium constant.

2. Background

2.1: Daniell Cell

A Daniell Cell is composed of two half-cells: one being a cathode, the positive terminal, and one being an anode, the negative terminal. For this experiment, the cathode is the copper half-cell and the anode is the zinc half-cell. The two half-cells are connected by a circuit and for this experiment, a voltmeter. The circuit is completed by placing a salt bridge between the two electrolyte solutions. The salt bridge is a filter paper soaked in a salt solution, in this instance KCl, where ions flow to certain sides to prevent the build-up of electrical charge. The anode half-cell is where oxidation occurs, which is the loss of electrons. Electrons will then travel through the circuit to the cathode where reduction occurs (Bylikin et al., 2014/2014). The equation for the cell being used is shown below.

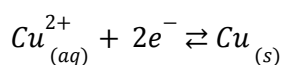


2.2: Nernst's Equation

Nernst's equation is used to determine the cell potential under non-standard conditions, which is when the temperatures are not constant at 298 K and the electrolyte solutions are not 1 M. This is given by the equation:

$$E = E_{cell}^{\theta} - \frac{RT}{nF} \ln Q$$

Where R is the universal gas constant (8.314 J K⁻¹ mol⁻¹), T is the absolute temperature, n is the moles of electrons that are lost/gained in the cell, F is the Faraday's constant (96500 C mol⁻¹), E_{cell}^θ is standard reduction potential at standard conditions and Q is the reaction quotient, which is the ratio of the product concentrations to the reactants when the equation is not at equilibrium (Bylikin et al., 2014/2014). Since the temperature of both solutions are different, we must first calculate the standard reduction potential except with a change in temperature for each half cell, and then sum them together in order to find the cell potential. Note that Q is equal to a different expression depending on the half-cell reactions. Looking at the half-cell reaction of the cathode, copper, the reaction quotient is given based on the equation below.



$$Q = \frac{[\text{Cu}_{(s)}]}{[\text{Cu}_{(aq)}^{2+}]}$$

However, since the concentration of the solid copper will not change as it reaches equilibrium, we omit it from the equation. This leaves us with the reaction quotient of $Q = \frac{1}{[Cu_{(aq)}^{2+}]}$.

2.3: Relationship Between K_c and the Cell Potential

Using the equations for Gibbs free energy, we can identify the expression that relates the cell potential to the equilibrium constant. The equilibrium constant is the ratio of the product concentrations to the reactant concentrations when the forward and reverse rates of reactions are occurring at the same rate. Using these two equations, we can determine an expression to isolate and find the K_c value (LibreTexts Chemistry, 2013).

$$\Delta G = -nFE_{cell}^{\theta} \quad (1)$$

$$\Delta G = -RT \ln(K_c) \quad (2)$$

Rearranging Nernst's equation with respect to E_{cell}^{θ} :

$$E_{cell}^{\theta} = E + \frac{RT}{nF} \ln Q \quad (3)$$

Substituting equation (3) into (1)

$$\Delta G = -nF(E + \frac{RT}{nF} \ln Q) \quad (4)$$

Substituting equation (4) into (2) and isolating for K_c

$$-RT \ln(K_c) = -nF(E + \frac{RT}{nF} \ln Q)$$

$$RT \ln(K_c) = nFE + RT \ln(Q)$$

$$\ln(K_c) = \frac{nFE}{RT} + \ln(Q)$$

$$K_c = e^{\frac{nFE}{RT}} Q$$

Looking at the equation, we can ignore the reaction quotient Q , as the product concentrations are equal to the reaction concentration, that being 0.5M $CuSO_4$ and 0.5M $ZnSO_4$, which results in 1. When using this equation with my experimental data, I will need to find an average of the temperatures at each-half cell in order to determine the equilibrium constant. Therefore, the equilibrium constant may not be accurate due to such averaging.

2.4: Hypothesis

Looking at the Nernst's equation, in the copper half-cell, as the standard reduction potential is positive and $Q > 1$, as the temperature of the copper half-cell increases, the cell potential decreases. For the zinc half-cell, as the standard reduction potential is negative and $Q < 1$, as the temperature of the zinc half-cell increases, the cell potential also decreases. Therefore, when either solution is increased, the cell potential will decrease. As far as the equilibrium constant is concerned, the value is dependent upon the ratio of $\frac{E}{T}$. Since the magnitude of the cell potential is less than the magnitude of the temperature, as the cell potential decreases, and the temperature increases, the equilibrium constant will decrease, indicating an equilibrium shift to the left.

3. Variables

3.1: Independent Variables

Independent	Possible Impact of Investigation	Method of Changing and Measuring Variable
Temperature (K)	Increasing the temperature of the solutions, whether that is the anode or cathode, decreases the cell potential and the equilibrium constant.	Place the half-cell whose temperature will be varied on a hot plate until the temperature reaches 70°C. Take the solution off the hot plate, make the circuit, and slowly lower the half-cell in an ice bath to lower the temperature. Use a thermometer to keep track of measurements.

3.2: Dependent Variables

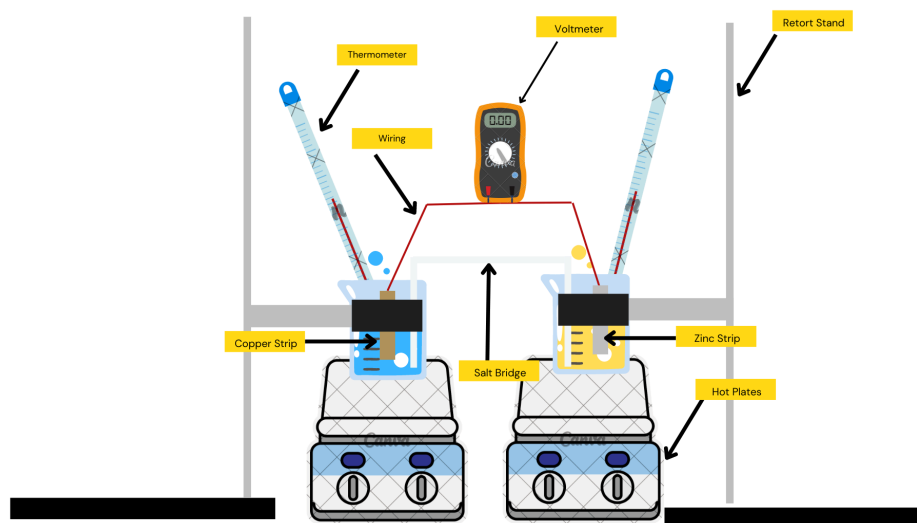
Dependent	Link to Research Question	Method of Measurement
Voltage (V)	By measuring the voltage, we can compare the relationship between the temperature and the cell potential, and through the equation proven in <i>section 2.4</i> , we can calculate the equilibrium constant as well.	Use alligator clips to attach wires to the metals being submerged into the half-cells and connect both wires to a voltmeter to measure the voltage. Record at temperature intervals (-5°C from 70°C).

3.3: Control Variables

Controlled	Possible Impact on Investigation	Method of Control
Wiring & Internal Resistance	Electrical wires have different lengths and thicknesses that have varying internal resistances, which affects the voltage.	Use the same electrical wires for each trial, label each wire connection for easier assembly and to keep all the results accurate with each other.
Concentrations of Solutions	When the concentration of a solution increases, the voltage increases as there will be more forward reactions occurring due to a greater amount of particles.	Concentrations of both the CuSO ₄ solution and the ZnSO ₄ solutions will be constant at 0.5 M.
Position of the Salt Bridge	When the salt bridge is submerged far into the solution, there is greater exposure from the ions in the salt bridge to the free floating ions in the electrolyte solutions. This may mean that there are higher voltages due to no build up of charge, and if the bridge is not submerged, a lower voltage will be shown, therefore, the position of the salt bridge affects the experiment effectively.	Using labelling tape on the half cell beakers, identify the point at which the edge of the salt bridge must stop. Precisely put a salt bridge edge over the label.
The Contamination of Electrolyte Solutions with Salt Bridge	Leaving the electrolyte solutions that are already used for a trial will result in less voltage since there are less Cu ²⁺ ions due to their bonds with the ions in the salt bridge.	Replace solutions after every trial.
The Contamination of Salt Bridge with Electrolyte Solutions	The concentration of the ionic solution in the salt bridge will also be contaminated by the electrolyte solutions, affecting the amount of voltage that is produced.	Replace salt bridge after every trial.
The temperature of the solution that is not being changed	To measure the effects of the anode or cathode when the temperature is changed, the other half-cell has to have a constant temperature of 303K (30°C).	Instead of heating up the solution directly, heat up 200ml of water to 303K (30°C) on a hot plate, and lower the solution into the water until the desired temperature is reached, then start the trial.

4. Methodology

4.1: Diagram



4.2: Apparatus/Materials

- Voltmeter (± 0.0005 V)
- Alligator Clips & Wiring
- Copper and Zinc Strips
- Filter Paper
- Scale (± 0.0005 g)
- 2 Graduated Cylinders (± 0.5 ml)
- 2 100ml beakers (± 25 ml)
- 3 250ml beakers (± 25 ml)
- 2 400ml beakers (± 25 ml)
- 2 Thermometer (± 0.5 °C)

4.3: Chemicals

- 100mL of 1 M KCl
- 480mL of 0.5M CuSO_4
- 480mL of 0.5M ZnSO_4

4.4 Methodology

1. Prepare solutions that are being tested with appropriate concentrations for required volumes.
 - a. Solutions will be split over two days, create 280 ml of CuSO_4 & ZnSO_4 on Day 1, then create the same volume of solutions on Day 2. Pour into 400ml beakers.
2. Set up two retort stands with their beaker clamps.
3. For the first trial, put both the copper strip and the zinc strip in the 250ml beakers and use labelling tape to indicate the position at which the strips need to be for each trial.
4. Put a filter paper and mark where the salt bridge must be positioned for all trials, once again using labelling tape.
5. Pour 80 ml of CuSO_4 into one beaker and pour the other 80 ml of ZnSO_4 into the other beaker. Put both beakers on their respective retort stands under the hot plates.
6. Put a 250ml beaker of water on the hot plate above the solution where the temperature is going to be constant, with a thermometer inside the water. Turn the hot plate on until the temperature reaches 30°C .
 - a. Once temperature is reached, push the hot plate back, put the beaker of water on the lab bench, and slowly lower the solution into the beaker of water, taking the thermometer from the water and putting it in the solution. Wait until the solution reaches the desired temperature of 30°C .

7. In the meantime, put the solution that will be varied in temperature onto the surface of the hotplate with the thermometer inside the solution until the temperature reaches 70°C.
 - a. Once desired temperature is reached, raise the beaker clamp up so that the solution is off the hot plates, push the hot plate back.
8. Take another 250ml beaker of water with ice, which will act as an ice bath, and set it under the solution that will vary in temperature.
9. Carefully soak a piece of filter paper for 10 seconds in a 250ml beaker of KCl, then place the salt bridge according to the labels marked at the start of the trial.
10. Place both metal strips into the solution with alligator clips attached to the strips which are connected to the voltmeter.
11. Carefully lower the solution that will vary in temperature into the ice bath. Record the voltage starting at 70°C, incrementing downwards by 5°C until the temperature reaches 30°C
12. Once done, clean out beakers and complete 3 trials for varying the anode solution temperature and 3 trials for varying the cathode solution temperature.

4.5: Safety, Environmental & Ethical Considerations

Consideration	Significance	How To Prevent?
Heat Generated from Hot Plates and Solutions	<p>The heat that will be generated by the hot plates and the solutions will be significantly hot. This can burn your skin, cause irritation and blister.</p> <p>Moreover, if used incorrectly, the hot plate can cause environmental damage, through an electrical shock and potentially a fire, destroying the work environment.</p>	<p>Do not touch the hot plate, and use gloves to protect against the heat. When moving solutions, make sure to touch the clamps, rather than the solution itself.</p> <p>Prior to experiment, be well aware of how to use a hot plate safely and effectively without causing damage to myself, others or the environment (Wayne State University, 2023).</p>
Health Risks of Chemicals	The chemicals, CuSO ₄ , ZnSO ₄ and KCl that when touched or in the eyes can cause skin irritation and when consumed can cause significant damage to organs and overall well being (New Jersey Department of Health, 2008).	Use gloves and goggles when measuring and dealing with the chemicals.
Proper Disposal of Chemicals	The chemicals, CuSO ₄ , ZnSO ₄ and KCl if not disposed properly can damage the ecosystem and environment in a multifaceted way. This has an impact upon animals, human health, etc (ASTS, 2018; (Carl Roth, 2016)).	Ensure that chemicals are put inside the inorganic chemicals bin, instead of pouring it out in the sink (ASTS, 2018; Teck, 2018; (Carl Roth, 2016)).
Use of Chemicals	To ensure that chemicals are not wasted, specific measurements for the specific volume will be done. Moreover, the process does involve any harm to animals.	Prior to experiment, calculate masses needed to dissolve and measure the exact or close to the amount of water needed.

5. Results

5.1: Raw Data

Below, you can find the data for the voltage that is recorded when the temperature of either the anode or cathode solution was adjusted or decreased. Note that when the anode solution or cathode solution was changing in temperature, the other solution was at constant temperature at 30°C. The uncertainty of the temperature was determined based on the thermometer, which was taken by subtracting the smallest possible measurement from the largest possible measurement, which was $\pm 0.5^\circ\text{C}$. The same process was repeated to find the uncertainty of the cell potential.

Table I. Raw Data For All Trials

Temperature of Solution (°C) (± 0.5 °C)	Cell Potential (V) (± 0.0005 V) When cathode temperature was varied			Cell Potential (V) (± 0.0005 V) When anode temperature was varied		
	Trial 1	Trial 2	Trial 3	Trial 1	Trial 2	Trial 3
70	1.129	1.109	1.155	1.053	1.049	1.063
65	1.130	1.106	1.155	1.053	1.048	1.062
60	1.131	1.108	1.155	1.052	1.047	1.061
55	1.131	1.109	1.154	1.052	1.046	1.060
50	1.130	1.110	1.152	1.052	1.046	1.058
45	1.129	1.109	1.148	1.052	1.046	1.057
40	1.128	1.107	1.143	1.051	1.046	1.055
35	1.126	1.105	1.134	1.048	1.046	1.055
30	1.119	1.103	1.130	1.046	1.046	1.046

5.2: Qualitative Analysis

After each trial was completed, when I removed the copper and zinc strips, I noticed that the surface of the copper strip was very pale, and a bit rough which can be as a result of the slight increase or decrease in the mass from the voltaic cell. Moreover, on the zinc strip, I noticed that some of the metal started wearing off, as when I wiped off the excess solution on the zinc strip, the paper towel had a black stain, which may be resulting from the loss of mass of the zinc strips.

5.3 Sample Calculations

First, I calculated the mass needed for each solution in each trial, which is 80 ml of 0.5 M CuSO_4 and ZnSO_4 . Using the mass, I then calculated the absolute uncertainty on the concentration shown on the right of the table. The uncertainty comes from the maximum measurement possible minus the minimum measurement possible, and as the scale is 3 digits, precise, the uncertainty becomes $\pm 0.0005\text{g}$. The graduated cylinder which was used for measuring the volume of water needed to dissolve the solid has an uncertainty of $\pm 0.5\text{ml}$ using the same process.

Mass of Solution Needed	Uncertainty of Concentration
$C_{\text{CuSO}_4} = \frac{m}{VM}$ $m_{\text{CuSO}_4} = C_{\text{CuSO}_4} VM$ $m_{\text{CuSO}_4} = (0.5) (0.08)(63.55 + 32.07 + 4(16))$ $m_{\text{CuSO}_4} = 6.3848\text{g} \pm 0.0005\text{g}$ $C_{\text{ZnSO}_4} = \frac{m}{VM}$ $m_{\text{ZnSO}_4} = C_{\text{ZnSO}_4} VM$ $m_{\text{ZnSO}_4} = (0.5) (0.08)(65.38 + 32.07 + 4(16))$ $m_{\text{ZnSO}_4} = 6.4580\text{g} \pm 0.0005\text{g}$	$\frac{\Delta C_{\text{CuSO}_4}}{C_{\text{CuSO}_4}} = \frac{\Delta m_{\text{CuSO}_4}}{m_{\text{CuSO}_4}} + \frac{\Delta V}{V}$ $\frac{\Delta C_{\text{CuSO}_4}}{0.5} = \frac{0.0005}{6.3848} + \frac{0.0005}{0.080}$ $\Delta C_{\text{CuSO}_4} = \pm 0.0032\text{ M}$ $\frac{\Delta C_{\text{ZnSO}_4}}{C_{\text{ZnSO}_4}} = \frac{\Delta m_{\text{ZnSO}_4}}{m_{\text{ZnSO}_4}} + \frac{\Delta V}{V}$ $\frac{\Delta C_{\text{ZnSO}_4}}{0.5} = \frac{0.0005}{6.458} + \frac{0.0005}{0.080}$ $\Delta C_{\text{ZnSO}_4} = \pm 0.0032\text{ M}$

Then, I averaged out the cell potential for when the temperatures of the anode and cathode varied using all three trials in *Table 1*. For the uncertainty, I decided to use the half range method, which takes the difference between the maximum value and the minimum value and divides it by 2 to give the uncertainty. A sample calculation of the average voltage and uncertainty is provided below for the changing of temperature of the cathode solution at 70°C.

$$E_{70^{\circ}\text{C at cathode}} = \frac{1.129 + 1.109 + 1.155}{3} = 1.131\text{ V}$$

$$\Delta E_{70^{\circ}\text{C at cathode}} = \frac{1.155 - 1.109}{2} = \pm 0.023\text{ V}$$

To use the equation proven in *section 2.4*, we must find the average temperature between the anode and the cathode. The average temperature and the uncertainty is depicted below.

$$T_{cell} = \frac{70 + 30}{2} = 50.0^{\circ}\text{C}$$

$$\Delta T_{cell} = \frac{0.5 + 0.5}{2} = \pm 0.5^{\circ}\text{C}$$

After determining the cell potential, I proceeded to insert the cell potential into the equation that was proven in *section 2.4*. The uncertainty and the value of the equilibrium constant is shown below. Note that the reaction quotient, Q, is ignored as the concentration of the anode and cathode solutions are 1. However, in the uncertainty, the Q constant must be considered.

Equilibrium Constant Value	Equilibrium Constant Uncertainty
$K_c = e^{\frac{nFE}{RT}}$ $K_c = e^{\frac{(2)(96500)(1.131)}{(8.314)(50+273)}}$ $K_c = e^{\frac{(2)(96500)(1.131)}{(8.314)(50+273)}}$ $K_c = 2.00 \times 10^{35}$	$\frac{\Delta K_c}{K_c} = e\left(\frac{\Delta T}{T} + \frac{\Delta E}{E}\right) + \frac{\Delta C_{CuSO_4}}{C_{CuSO_4}} + \frac{\Delta C_{ZnSO_4}}{C_{ZnSO_4}}$ $\frac{\Delta K_c}{2.002 \times 10^{35}} = e\left(\frac{0.5}{50} + \frac{0.023}{1.131}\right) + \frac{0.0032}{0.5} + \frac{0.0032}{0.5}$ $\Delta K_c = \pm 0.19 \times 10^{35}$

To further my analysis, I will also be calculating the theoretical values to calculate the percentage error. Using the Nernst’s equation, we can find the half cell potential for the anode solution at 30°C and the cathode solution at 70°C.

Zinc Anode Half Cell Potential	Copper Cathode Half Cell Potential
$E_{cell}^{\theta} = -0.76 - \frac{(8.314)(30 + 273)}{(2)(96500)} \ln 0.5$ $E_{cell}^{\theta} = -0.751\text{ V}$	$E_{cell}^{\theta} = 0.34 - \frac{(8.314)(70 + 273)}{(2)(96500)} \ln 2$ $E_{cell}^{\theta} = 0.330\text{ V}$

Using these two values, we can then use the cell potential formula to determine the cell potential when both half-cells are connected together. Note that since the zinc anode half cell has a negative reduction potential, it therefore gets oxidized, hence the sign will flip.

$$E_{cell}^{\theta} = E_{ox}^{\theta} + E_{red}^{\theta}$$

$$E_{cell}^{\theta} = 0.751 + 0.330$$

$$E_{cell}^{\theta} = 1.081V$$

Then, we plug the theoretical cell potential into the equilibrium constant equation alongside the average temperature.

$$K_{c \text{ theoretical}} = e^{\frac{(2)(96500)(1.081)}{(8.314)(50 + 273)}}$$

$$K_{c \text{ theoretical}} = 5.39 \times 10^{33}$$

Using the theoretical values and the experimental values, we can then find out the experimental error for the cell potential and the equilibrium constant.

$$\text{Experimental Error for Cell Potential} = \frac{1.131 - 1.081}{1.081} \times 100 = 4.65\%$$

$$\text{Experimental Error for } K_c = \frac{2.00 \times 10^{35} - 5.39 \times 10^{33}}{5.39 \times 10^{33}} \times 100 = 3612.35\%$$

5.4: Processed Data

Using the calculations above, here are the tables with the processed data.

Table 2. Experimental Cell Potential and Equilibrium Constant

Temperature (°C) (± 0.5 °C)	Average Cell Potential (V)		Equilibrium Constant	
	Variation in Cathode Temperature	Variation in Anode Temperature	Variation in Cathode Temperature	Variation in Anode Temperature
70	1.131 ± 0.023	1.055 ± 0.007	2.00 × 10 ³⁵ ± 0.19 × 10 ³⁵	8.50 × 10 ³² ± 0.49 × 10 ³²
65	1.130 ± 0.025	1.054 ± 0.007	3.61 × 10 ³⁵ ± 0.36 × 10 ³⁵	1.46 × 10 ³³ ± 0.09 × 10 ³³
60	1.131 ± 0.024	1.053 ± 0.007	7.36 × 10 ³⁵ ± 0.73 × 10 ³⁵	2.48 × 10 ³³ ± 0.15 × 10 ³³
55	1.131 ± 0.023	1.053 ± 0.007	1.42 × 10 ³⁶ ± 0.15 × 10 ³⁶	4.34 × 10 ³³ ± 0.27 × 10 ³³
50	1.131 ± 0.021	1.052 ± 0.006	2.62 × 10 ³⁶ ± 0.25 × 10 ³⁶	7.67 × 10 ³³ ± 0.48 × 10 ³³
45	1.128 ± 0.020	1.052 ± 0.006	4.44 × 10 ³⁶ ± 0.43 × 10 ³⁶	1.40 × 10 ³⁴ ± 0.09 × 10 ³⁴
40	1.126 ± 0.018	1.051 ± 0.005	7.19 × 10 ³⁶ ± 0.68 × 10 ³⁶	2.46 × 10 ³⁴ ± 0.16 × 10 ³⁴
35	1.122 ± 0.015	1.050 ± 0.005	1.04 × 10 ³⁷ ± 0.09 × 10 ³⁷	4.36 × 10 ³⁴ ± 0.29 × 10 ³⁴
30	1.117 ± 0.014	1.046 ± 0.001	1.50 × 10 ³⁷ ± 0.14 × 10 ³⁷	6.36 × 10 ³⁴ ± 0.38 × 10 ³⁴

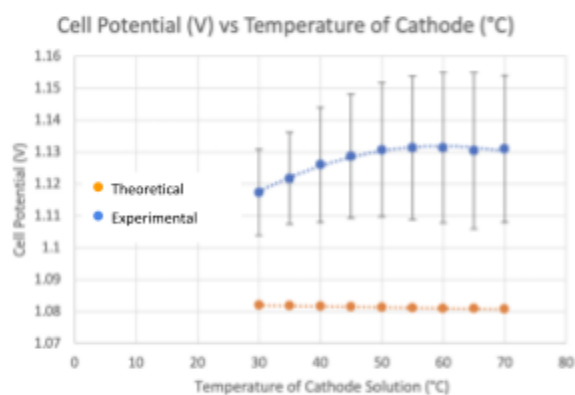
Table 3. Theoretical Cell Potentials, Equilibrium Constant and Percentage Error

Temperature (°C) (± 0.5 °C)	Theoretical Cell Potential (V)		Theoretical Equilibrium Constant		Cell Potential Percentage Errors (%)		Equilibrium Constant Percentage Errors (%)	
	Variation in Cathode Temperature	Variation in Anode Temperature	Variation in Cathode Temperature	Variation in Anode Temperature	Variation in Cathode Temperature	Variation in Anode Temperature	Variation in Cathode Temperature	Variation in Anode Temperature
70	1.08071	1.08071	5.39×10^{33}	5.39×10^{33}	4.65	2.38	3612.35	84.24
65	1.08086	1.08086	9.99×10^{33}	9.99×10^3	4.58	2.45	3499.36	85.36
60	1.08101	1.08101	1.87×10^{34}	1.87×10^{34}	4.66	2.56	3839.40	86.74
55	1.08116	1.08116	3.53×10^{34}	3.53×10^{34}	4.64	2.64	3911.45	87.71
50	1.08131	1.08131	6.74×10^{34}	6.74×10^{34}	4.56	2.71	3788.89	88.62
45	1.08146	1.08146	1.30×10^{35}	1.30×10^{35}	4.37	2.75	3310.68	89.22
40	1.08161	1.08161	2.53×10^{35}	2.53×10^{35}	4.10	2.86	2738.63	90.29
35	1.08176	1.08176	4.99×10^{35}	4.99×10^{35}	3.69	2.97	1975.25	91.27
30	1.08191	1.08191	9.95×10^{35}	9.95×10^{35}	3.27	3.32	1409.34	93.61

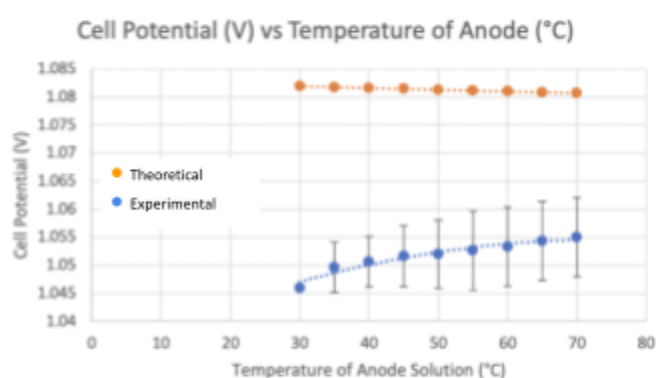
5.5: Graphical Representation

For the first graph, I decided to compare the theoretical and experimental values for what happens to the cell potential when the temperature of the anode (zinc) or cathode (copper) solution is changed with the other solution being constant..

Graph 1



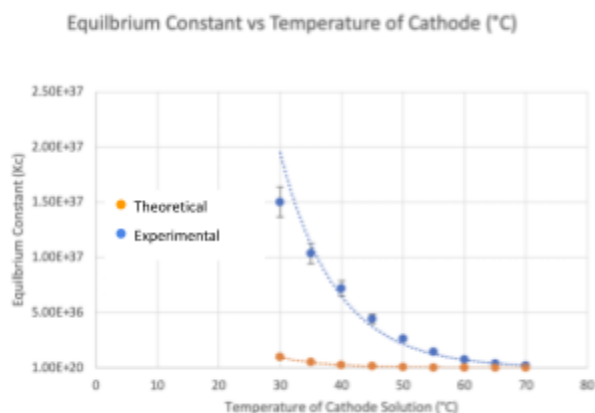
Graph 2



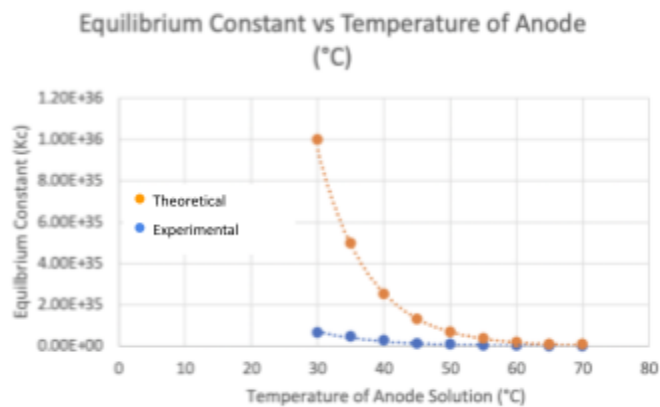
The error bars for the temperature were ignored, as they are quite small relative to the scale that is given. However, the cell potential error bars were included as they do have some significance to the line of best fit. It is important to note that while the theoretical line of best fit appears to be linear, the experimental line of best fit is quadratic, which may be as a result of the lack of ions that are present as the solution reaches, which is explained in the analysis section of the paper.

Afterwards, I decided to create two graphs based on the equilibrium constant with reference to the temperature of either the anode or cathode solutions.

Graph 3



Graph 4



The vertical error bars for *Graph 3* were much more significant, hence why I decided to include some of the bars. However, the points, notably those towards the higher temperatures, were ignored due to there being less uncertainty that does not fit the y-axis scale. Through the same reasoning, all of the vertical error bars in *Graph 4* were ignored as well. As predicted, the experimental data follows the exponential curve that is expected due to the exponential function used to calculate the equilibrium constant.

6: Analysis, Evaluation & Conclusions

6.1: Analysis

As shown in *Graph 1 and 2*, the experiment that was conducted showcases that as the temperature of either the cathode or anode half-cell increases, the cell potential subsequently decreases. While the trend does not fully align with the trend expected, there is some slight indication that there is some theoretical alignment. In *Graph 1*, I noticed that after the increase in cell potential from 30°C - 50°C, after 50°C, the cell potential had a decrease based on the regression line, which follows the general trend that exists from the theoretical values.

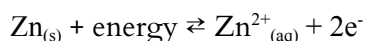
Since the galvanic cell is both exothermic and spontaneous, using Le Chatelier's principle, the equilibrium will shift to the left, favouring the endothermic reaction as temperature is increased, therefore the reactants. In doing so, not only will the equilibrium constant decrease, however, the amount of electrons that are lost and gained will also decrease, thereby decreasing the cell potential. This explains the reasoning behind *Graph 3 and 4*, which is when the equilibrium constant is lower when the temperature is increased, as the endothermic direction is favoured to maintain an equal distribution of energy in the reaction.

The experiment may have indicated the voltage going up as the temperature increases due to the fact that as the ions complete the redox reactions, the concentrations will change. In the zinc anode, there is an increase in the concentration of the Zn^{2+} and there is a decrease in the concentration of Cu^{2+} ions available in the copper cathode. As there is a lack of ions available, the voltage will gradually decrease. This is supported by the fact that it took a longer time for the solution to reach lower temperatures, therefore giving more time for the voltaic cell to deplete its ions to react.

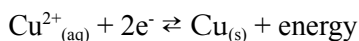
Another reason why the voltage may decrease when the temperature increases lies in the buildup of charge in the solution. In the anode, oxidation results in the solution having a positive charge due to electrons being released to travel through the circuit, leaving the positive ions of Zn^{2+} . On the other hand, at the copper cathode, the solution becomes negative, as the SO_4^{2-} will break their bond with the Cu^{2+} , so that the copper ions can take place with the reduction reaction. While the salt bridge aims to create neutral solutions by forming bonds with the negative or positive ions in the solutions with the ions in the salt bridge, as the ions in the salt bridge start to deplete, the solutions become much more negative or positive. This was apparent when I took out the salt bridge from each trial, where the copper cathode side was considerably blue. As the cathode becomes more negative, electrons will repel the negative solution, preventing the movement of electrons within the circuit. Subsequently, there is less voltage that occurs in the circuit.

Something else that was interesting to note when looking at the graphs is that when I adjusted the cathode's temperature, the expected cell potential and equilibrium constant, in *Graphs 1 and 3* were lower than the experimental values. This may be due to the fact that I used pre-made 0.5 M solutions of the CuSO_4 solutions, which may be contaminated or have more concentration than

the 0.5M. As a result, this may have increased the cell potential, which therefore increased the equilibrium constant. As I ran out of ZnSO_4 pre-made solutions, I was required to make my own solutions prior to commencing the experiment. Due to internal resistance of the wires, the build of charge and other external factors, the cell potential is relatively lower when changing the temperature of the anode compared to the theoretical value in *Graphs 2 and 4*. It is also prominent to understand the effect of temperature on each half-cell, which provides a good indication to the trends that are expected on the graph. Notice that an oxidation at the anode is the second ionization energy, hence it is endothermic. Below is the half-cell equation for the anode which is zinc:



Using Le Chatelier's principle, we understand that as the temperature increases, the number of electrons that are produced increases, as the equilibrium is shifted to the right, which explains why when the anode temperature increases, the voltage does as well. The same principles can be applied to the copper cathode, below is the half-cell equation for the cathode which is copper:



Note that the reduction is similar to electron affinity which is exothermic. Hence as temperature increases, the reaction favors the endothermic way, which shifts equilibrium to the left. As a result, the voltage decreases, as the solid is not formed, and there is a buildup of electric charge. This trend is depicted after the 50°C mark on the first graph. It is important to note that experimentally, it appears that the cell potential increases as the temperature increases, but this is most likely stemming from the fact that the build up of charge as the cell continues to run, prevents the movement of electrons, thereby decreasing voltage automatically.

6.2: Evaluation

Reflecting upon the lab process, I realized that there are numerous strengths and weaknesses to how I have constructed a methodology to answer the question. One prominent strength in the consideration of certain controlled variables that could potentially deviate the results. Specifically, the replacement of the salt bridge and the solutions after every trial is heavily important as it prevents the buildup of charge, which can significantly reduce the cell potential after every trial. Furthermore, the methodology and the experiment follows similar trends to secondary sources and formulas. However, there are still some deviations from the expected values which were previously explained within section 6.1.

However, there are still numerous weaknesses associated with regards to the experiment. First of all, there are numerous explanations to why there is a certain correlation between the data points. There isn't a specific reason that explains the trends, specifically whether it increases or decreases. In order to focus on one reasoning, there lies importance in understanding and controlling the variables to see the impact of simply one explanation. For example, by using only solutions made by myself, we could immediately eliminate one of the reasons that I have provided. Furthermore, when using Nernst's equation at the point where both solutions are at 30°C , the value of the cell potential is around 1.081 V, which slightly deviates from the voltage at standard potential which should be at 1.1 V. Even though there is a slight difference between the temperatures at standard (25°C) and 30°C generally speaking, the value should be more specific at around 1.1 V. Hence, this methodology and my theoretical calculations are slightly deviated from the actual standard cell potential expected.

It was very challenging to ensure that the position of the salt bridge was constant and in place. As the filter paper was fairly light, the ends of the paper kept moving around, which changes the surface area that is exposed, changing the amount of ions that are allowed to pass through to the solution. This can greatly affect the voltage, as if there are more ions readily available, there is less build-up of electric charge therefore increasing the voltage. To further improve this methodology, it may be necessary to purchase a glass salt bridge within cotton plugs at the end to ensure that the salt bridge does not move, varying the exposure. Moreover, another major weakness was the fact that it was hard to maintain the temperature of the solution that was not being varied. While the water bath did help, as the trial was running, the temperature of the solution started decreasing, which can be attributed to the heat losses within the solution. To fix this issue, it may be beneficial to get an electrical water bath, to ensure that there is a more accurate way to get a constant temperature of the solution. Lastly, a major weakness in general was the immense amount of percentage error in the equilibrium constant. This may have arisen from the large magnitude involved when calculating the value of the equilibrium constant or from the numerous random errors that greatly affected the experiment such as the build-up of electric charge, the maintaining of temperature and other reasons listed in the table below.

Error	Type of Error	Significance	Correction
Position of Salt Bridge	Random	As the filter paper was fairly light, the ends of the paper kept moving around in the solution changing the amount of ions that are allowed to enter the solution.	There are two possibilities to correct this error. As it is a random error, there can be further repetitions done to reduce the error. Moreover, rather than using filter paper, to avoid the error, a glass tube with cotton plugs with the solutions inside the tube can prevent the movement of the paper.
Maintaining Constant Temperature of Solution	Random	It was hard to maintain the constant temperature of the solution that was not being varied, as even though the water bath did help, the temperature of the solution started decreasing, which can be attributed to the heat losses within the solution.	Either further the repetitions, or an electric water bath can be used in order to ensure the water is at constant heat or stove, preventing further heat losses.
Internal Resistance of Wiring And Effect of Temperature	Random	As the temperature of the solution increases, the internal resistance in the wires increases, which results in the decrease in cell potential, thereby deviating from the results.	Use a material that has a relatively conductive surface without resistance, and a thinner wire.
Lack of Time Left For Solution At Particular Temperature	Random	When I put the solution in the ice bath, the temperature, at first, rapidly decreased, which resulted in the cell not being able to rest and set at that particular temperature. This may have resulted in a lack of accuracy in the reading of the voltage, as the salt bridge or other external factors may have been moving or changing.	Try to use a water bath and leave the cell to rest at a constant temperature for each trial, rather than conducting all temperatures in one go.

6.3: Conclusion

All in all, as the temperature increases between the anode or cathode, the cell potential decreases due to the fact that the cell is an exothermic process, and an increase in temperature simply reduces the concentrations of each solution in a cell, thereby decreasing the cell potential. Experimentally, there were major errors that showed different trends from theoretical values and trends and there were numerous reasons for the experimental trends, posing a challenge to the validity of the experiment. As a further extension, I would like to investigate more into renewable energy and the effect of temperature on such production mechanisms, such as PV cells. This extension could potentially aid us in understanding how temperature changes favour certain countries' energy production mechanisms.

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