Investigating the rate of electrolysis of zinc sulfate with a copper anode and a zinc cathode
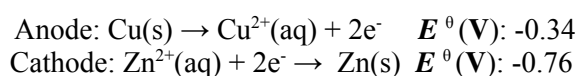
Research Question

Comparing the rate of electrolysis of zinc sulfate with a copper anode and a zinc cathode using colorimetry against a model based on current flow.

Background Information

As per the Pearson Higher Level Chemistry book[1], electrolysis is the process during which an "external source of electrical energy" is used "to bring about a redox reaction that would otherwise be non-spontaneous". Therefore, we can electrolyze zinc sulfate using a copper cathode and a zinc anode, plugged into a standard 6V power source.

We know that cations - in this case, Zn^{2+}- will be attracted to the negative electrode: the zinc cathode. Additionally, from section 24 of the data booklet and the Person book, we can assume that the following reactions are happening at the electrodes:

| |
|---|
| Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ E^{θ}(**V**): -0.34 |
| Cathode: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ E^{θ}(**V**): -0.76 |

As the copper electrode is an active anode, we can deduce that copper ions will be dissolved into the electrolyte solution (as copper is lower than hydrogen in the electrochemical series)[2] while pure zinc will be deposited onto the zinc plate, while. As zinc sulfate is mostly colorless, we can expect the electrolyte to turn bluer as copper ions dissolve into the solution and form $CuSO_4$ (a blue compound). Accordingly, we can investigate the rate of electrolysis by measuring the amount of copper sulfate that replaced the original zinc sulfate in the electrolyte.

Colorimetry is used to measure this rate. Then, the results are compared to the expected concentration based on Faraday's law of electrolysis, stating that "the amount of product formed by an electric current is chemically equivalent to the amount of electrons supplied"[3].

The Model

Combining the equations for the reactions happening at the electrodes mentioned above, we can derive the redox equation of the electrolysis, without the spectator sulfate ions:

$$Zn^{2+}(aq) + Cu(s) \rightarrow Zn(s) + Cu^{2+}(aq) \qquad \textit{(Equation 1)}$$
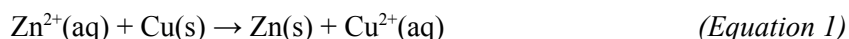
Therefore, using Faraday's law, we can derive the number of copper ions produced:

$$n(Cu^{2+}) = \frac{1}{2} \times n(e^-)$$

In addition, as the electrolyte solution is of a 100mL, we can calculate the concentration of the solution based on the moles of electrons passing through:

$$C(Cu^{2+}) = \frac{n(Cu^{2+})}{V} = \frac{\frac{1}{2} \times n(e^-)mol}{0.1\ L} \qquad \textit{(Equation 2)}$$

Finally, we can calculate the moles of electrons passing through using the following equations.
First, we calculate the total charge by measuring the current every 10 minutes and summing them up.

$$\text{charge (Q)} = \text{current (I)} \times \text{time (t)} \qquad \textit{(Equation 3)}$$

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[1] Brown, C., & Ford, M. (2014). 07 Equilibrium. In *Chemistry: Higher level: Supporting every learner across the IB continuum* (pp. 442-443). Pearson Education Limited.
[2] "Electrolysis Using Inert and Reactive Electrodes - MLC." Miracle Learning Centre, 17 Oct. 2022, https://miraclelearningcentre.com/chemistry-lessons-topic-electrolysis-using-inert-and-reactive-electrodes/.
[3] Ibid. (pp. 595-596)

Then, dividing the total charge in a period of time by the charge of 1 mol of electrons (otherwise known as Faraday's constant, or the charge of one electron multiplied by Avogadro's constant), we get the number of moles of electrons going through the system in the same period:

$$N(e^-) = \frac{charge\ (Q)}{1.602189 \times 10^{-19}\ C}$$

$$n(e^-) = \frac{charge\ (Q)}{1.602189 \times 10^{-19}\ C \times 6.02 \times 10^{23}\ mol^{-1}} = \frac{charge\ (Q)}{96\,500\ C\ mol^{-1}} = \frac{charge\ (Q)}{Faraday's\ constant} \qquad (Equation\ 4)$$

Combining all the above equations into 1 (to avoid inaccuracy due to rounding values repeatedly), the concentration of copper based on charge is:

$$C(Cu^{2+}) = \frac{\frac{1}{2} \times n(e^-)}{0.1} = \frac{\frac{charge}{2 \times 96\,500}}{0.1} = \frac{charge}{0.2 \times 96\,500} = \frac{charge}{19\,300}\ mol.\,L^{-1} \qquad (Equation\ 5)$$

We can thus make a model predicting the concentration of the electrolyte at any point. This will then be compared to the actual concentration, which will be measured using the color of the electrolyte. The color of the solution will be quantified by colorimetry through computer software and will be compared to the color of "control" samples of chosen concentrations (made by diluting solid copper sulfate into distilled water, a colorless solution) to find the concentration of copper in the electrolyte at a specific moment in time.

In this essay, the word "model" will thus refer to the excel model that will take current values over time and return the expected concentration of the electrolyte solution.

Preliminary Investigations

Preliminary investigations were centered around figuring out how to quantify the color of the electrolyte and then compare it. The original idea was to use a colorimeter; however, the reaction takes too long, and it is difficult to take regular samples over an extended period of time, as class time is limited. In consequence, I decided to film the reaction using a timelapse on a mobile phone and leaving aqueous copper sulfate solutions of different concentrations around the beaker where the electrolysis was taking place, in order to compare the color of those "control" solutions and the electrolyte over time. To measure the colors, RGB values originally seemed like the best option as the difference between two colors could be computed through the euclidean distance formula, considering them as points in a 3D environment with R, G, and B as the axis of said environment.

The main problem with this method is that the distance given by the formula above does not "match the human-perceived distance between colors" [4]. To correct this, the L*a*b color environment will be used to compare colors. As it is a perceptually uniform colorspace[5] (except when saturation is involved, which isn't for our experiment as the phone camera quality is high enough), it will not only account for the problem explained above, but is also easier to recalibrate the color of the electrolyte when lighting changes. Finally, using the L*a*b color space allows me to use the Delta-E measure, which is defined as "the measure of change in visual perception of two given colors"[6]. All RGB values retrieved from the camera will thus be converted into L*a*b values before being compared by computer software.

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[4] Bukys, Liudvikas. "How to Compare Two Colors for Similarity/Difference." *Stack Overflow*, 26 Jan. 2012, https://stackoverflow.com/a/9019461.
[5] Schuessler, Zachary. "Delta E 101". *Delta E*, 25 Sept. 2019, https://zschuessler.github.io/DeltaE/learn/.
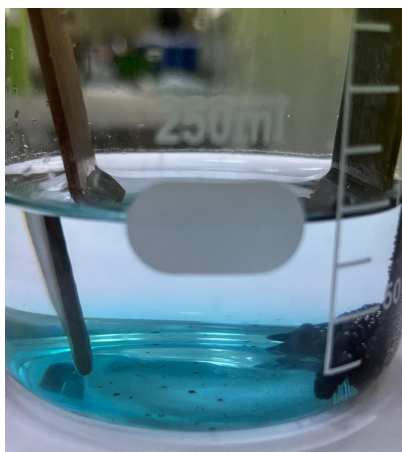[6] Ibid.

Figure 1: *Picture of the electrolytic cell after an hour of electrolysis*, author

After running a trial run, it was noticed that zinc accumulated at the zinc electrode, and solid copper was be deposited at the bottom of the beaker and slowly grow from the cathode to the anode. To prevent a short circuit, the copper electrode was wrapped with paper. However, the experiment could thus not be run for more than 2-3 hours, as the copper solidifying at the bottom of the beaker would either get too close to creating a short circuit or cover too much surface, preventing me from measuring the color of the electrolyte.

The experiment had to be run in 1 day and could not be carried over from one day to another as control solutions would evaporate, get contaminated, and some of the solutions would solidify and stick to the side of the beaker. Once the current was stopped, the electrolyte would turn colorless again as copper would be deposited at the bottom of the beaker. This is due to the fact that copper is displaced by the solid zinc near the cathode, as zinc is higher than copper in the activities series (as per the data booklet).

Variables

❖ The **independent** variable is the concentration of the solution
❖ The **dependent** variable is time
❖ The **controlled** variables are as follows:

Figure 2: Table of the controlled variables of the experiment, their impact on it, and how they were controlled.

Controlled Variable	Its impact on the experiment	How was it controlled?
Lighting	Different lightings change the color of the solutions: sunlight/ambient light can vary over time	The setup was placed into a deep box, with white paper taped to the sides and the bottom so the color of the box wouldn't impact the color of the solutions. Color values of the electrolytes were recalibrated using the colors of the control solutions at the same moment.
Depth of the beakers	When looking at the beaker from the side, the color of the electrolyte isn't uniform; the closer to the bottom of the beaker, the darker the tone of blue (figure 1). However, when looking from above, the color of the solution is uniform	Each control solution was adjusted to have the same depth as the electrolyte (by removing some of the solutions until the meniscus of all beakers was at the same height, through observation using the eye)
Shading	Shading can change the color of the solutions: sunlight/ambient light can vary over time	A lamp was added so it could light up all solutions equally, in a constant manner. While there was still shading in some areas, the shading didn't cover entire beakers meaning there were still areas that could be analyzed using colorimetry.
Surface area of the electrodes in contact with the electrolyte & Distance between the electrodes	The surface area in contact and the distance between the electrodes can affect internal resistance and the current of the cell	The electrodes were pushed perpendicularly into the cell until they hit the bottom. A piece of foam was added between the electrodes, keeping them perpendicular to the electrolyte and the distance between them constant

Equipment list

DC 6V Power source with cables and clips, ±0.01A multimeter/ammeter, stopwatch, camera, a pure copper, and a pure zinc electrode, a 250mL beaker, 6 100mL beakers, a 100 ± 1mL measuring cylinder, a lamp and a scale (± 0.01g)

Safety and Environmental concerns

Zinc sulfate, and copper sulfate can be corrosive and cause serious eye damage or irritation[7]. Therefore, gloves, goggles, and a long coat are required when manipulating the substances during this experiment. Additionally, it cannot be poured down the drain and has to be taken care of in accordance with local laws. Dispose of the above substances "to an approved waste disposal plant" [8].

Method

Electrolysis

- ❖ Measure 100.00mL of 0.50M zinc sulfate using a measuring cylinder.
- ❖ Place the zinc sulfate into the 250mL beaker.
- ❖ Wrap a copper electrode with a layer of paper.
- ❖ Place a zinc electrode and the copper electrode into the solution, separating the two using a piece of polystyrene foam board, in order to keep the distance between them constant over trials.
- ❖ Create a circuit with the two electrodes, the power source, and the ammeter as shown in the figure above (also refer to figures 3 and 4), using cables and crocodile clips.
- ❖ Place the beaker and ammeter in a box as discussed in the "controlled variable" section above.

Control solutions

- ❖ Beaker 1: Place 100.00mL (measured with a measuring cylinder) of distilled water into a 100mL beaker.
- ❖ To make the other 5 control solutions of concentration 0.10M, 0.20M, 0.30M, 0.40M, and 0.50M, weigh out 2.50, 4.99, 7.49, 9.99, and 12.49g of solid copper sulfate pentahydrate using a scale, and place them in separate 100mL beakers. Those values were found by doing the following calculations: (using 249.7g/mol as the molecular mass of copper sulfate pentahydrate[9])

$$C(Cu^{2+}) = C(CuSO_4.5H_2O)$$

$$C = \frac{n}{v} \quad n = Cv \quad n(CuSO_4.5H_2O) = \frac{m(CuSO_4.5H_2O)}{M(CuSO_4.5H_2O)} = \frac{m(CuSO_4.5H_2O)}{249.7}$$

$$m(CuSO_4.5H_2O) = 249.7 \times n(CuSO_4.5H_2O) = 249.7 \times Cv = 249.7 \times 0.1L \times C = 24.97\,C$$

- ❖ Measure 100.00mL of distilled water and add it to each of the 5 beakers containing copper sulfate.
- ❖ Stir the solutions so the solid copper is diluted into the solution.
- ❖ Remove some solution from the 6 100mL beakers using a pipette, until the height of the solution is around the same height of the electrolyte in the 250mL beaker (in order to keep the depth of all 7 beakers the same; refer to figure 3).
- ❖ Place the control solutions right next to the 250mL beaker in the box.

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[7] "Safety Data Sheet - Uprm.edu." *Zinc Sulfate*, Global Safety Management, 19 Mar. 2015, https://www.uprm.edu/citai/wp-content/uploads/sites/222/2021/06/Zinc-Sulfate.pdf.

[8] "Safety Data Sheet - Fisher SCI." *Copper(II) Sulfate*, ThermoFisher Scientific, 26 Dec. 2021, https://www.fishersci.com/msds?productName=AC422870050.

[9] "Copper Sulfate Pentahydrate." National Center for Biotechnology Information. PubChem Compound Database, U.S. National Library of Medicine, https://pubchem.ncbi.nlm.nih.gov/compound/Copper-sulfate-pentahydrate#section=Chemical-and-Physical-Properties.

Setup

- ❖ Attach a phone or camera to a retort stand and place it into the box, to film the beakers from above.
- ❖ Place a lamp directly above the box in order to light all the beakers in the box.
- ❖ Place a stopwatch and the ammeter in a position where it is visible when filming.
- ❖ Start recording in timelapse
- ❖ Start recording the time using the stopwatch as soon as the power cell is turned on.
- ❖ Let the reaction take place until the copper accumulating at the bottom is close to the copper electrode (as contact would create a short circuit). (Usually around 2-3 hours).
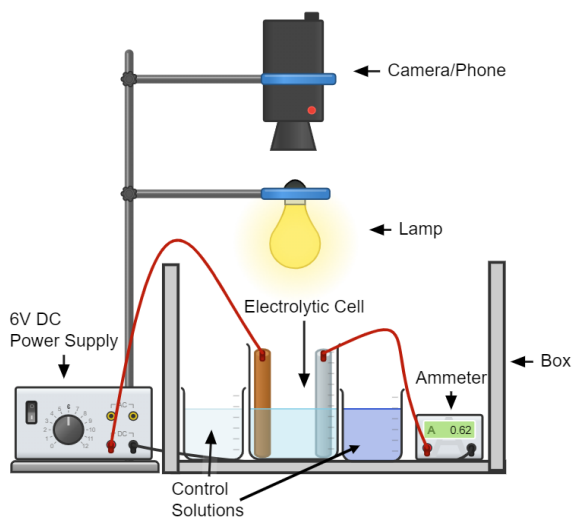
Repeat the experiment twice.

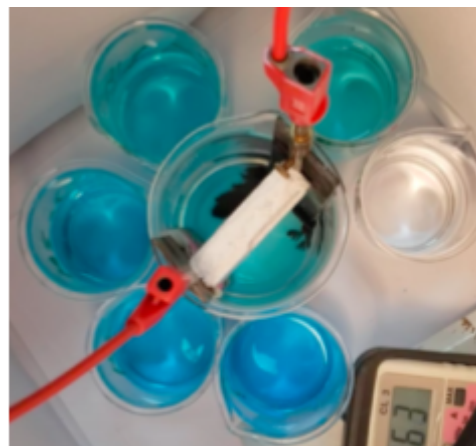
Figure 3, *Diagram of the setup*, Author using Chemix

Figure 4, *Picture of the setup,* Author

Expected concentration over time of the solution based on current flow, using an Excel Model

As the current was recorded every 10 minutes, we can use *Equation 5* to predict the expected concentration of the solution at a given time. This is calculated using an excel model, where the concentration of the electrolyte is given by the total charge of the system until that point is divided by 19300. The total charge will be calculated using a graph plotting current over time; when integrating the equation of the line of best fit, we can find the area under the graph, which is, in this case, equal to the total charge until that point, as per *Equation 3*. For example, for run 1, the graph of current over time is as follows:
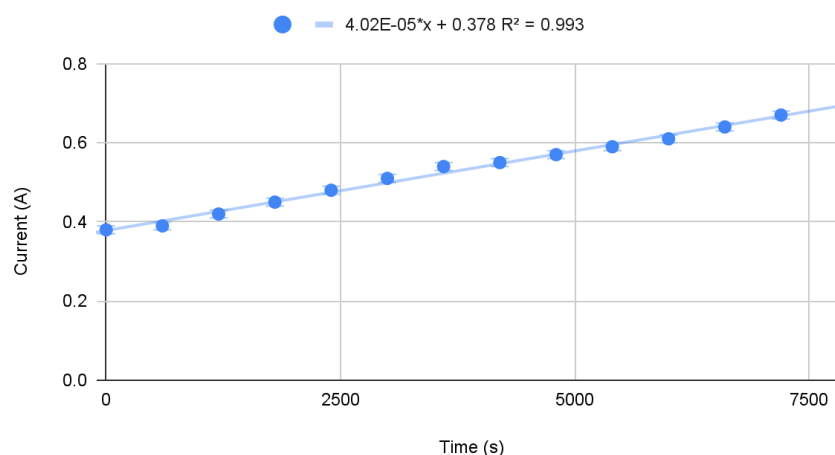

Current over Time for run 1

4.02E-05*x + 0.378 R² = 0.993

Figure 5: *Graph of the current over time during run 1*, author

5

$$charge(t) = \int_0^t (current)\, dx = \int_0^t (4.02 \times 10^{-5}\, x + 0.378)\, dx = \left[2.01 \times 10^{-5}\, x^2 + 0.378x \right]_0^t$$

$$charge(t) = 2.01 \times 10^{-5}\, t^2 + 0.378\, t$$

$$C(Cu^{2+})(t) = \frac{2.01 \times 10^{-5}\, t^2 + 0.378\, t}{19\,300}\, mol.\, L^{-1} \qquad \textit{(Equation 6)}$$

However, this model assumes there is no backward reaction during the electrolysis and that the current increases constantly as we are using a linear line of best fit (for simplicity, but also because the correlation coefficient suggests the linear correlation is very strong).

Actual results based on colorimetry

To find the concentration of the electrolyte using colorimetry, the following steps will be followed:

1. Save images from the video recorded in intervals of 10 minutes (starting from minute 10).
2. Split those images into 7 pictures so every control solution and the electrolyte has its own file
3. Crop those 7 pictures in a way that only the color of the solution is seen (remove any spaces where shadow, reflections, saturation, or else is visible. Crop the pictures in the same location for every image throughout the experiment). Here is the example for the first run (two runs were conducted):

Setup at t = 80 minutes

0M 0.1M 0.2M 0.3M 0.4M 0.5M

Electrolyte

The image at a specific time (on the left) is split into 7 pictures (on the right).
Those pictures are then cropped (red rectangles).
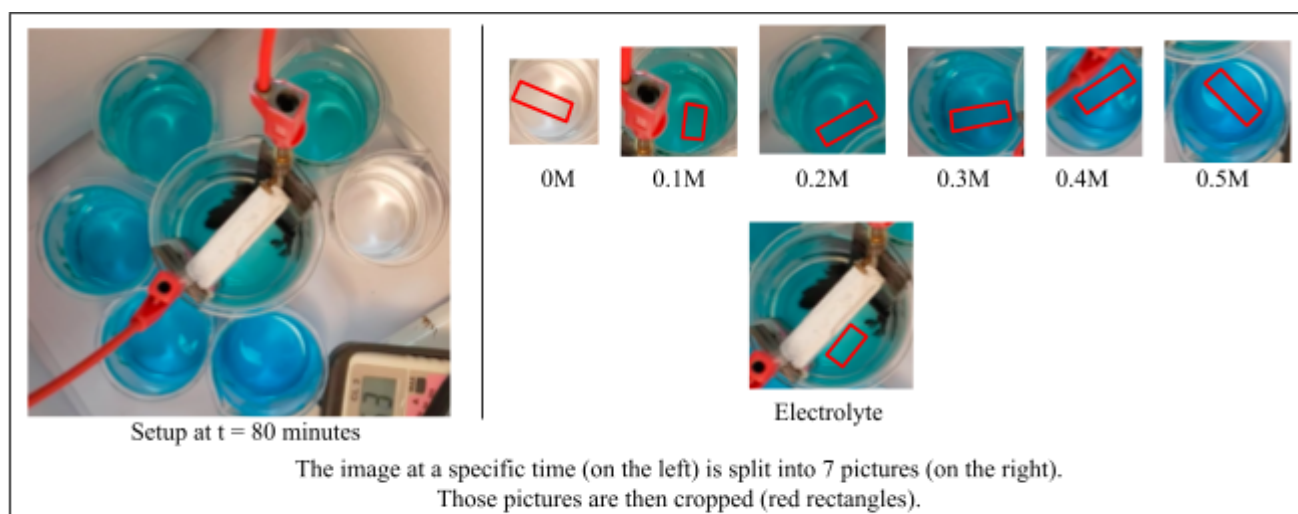
Figure 6, *How the pictures are cropped,* Author

4. First, using the Python program 3 [10], find the average color of the cropped areas of the control solutions at t = 0 min. The average values are as follows:

Concentration of solution(M)	R	G	B	Color	Concentration of solution(M)	R	G	B	Color
0.00	190	199	196		0.30	69	152	193	
0.10	134	195	216		0.40	67	154	205	
0.20	97	178	206		0.50	51	106	168	

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[10] All programs area available at https://anonymous.4open.science/r/Chemistry-IA-Programs/README.md

5. Make a gradient that is 200 pixels long with those colors. The color of the first pixel will thus be the same color as the 0.00M control solution, and the color of the last pixel will be the same color as the 0.50M solutions. This assumes that the x-coordinate of the pixel is proportional to the concentration of the solution, and follows the following equation: (the size of 200 pixels was selected as it takes a very short time to run through in the programs and any more precision is unnecessary).

$$concentration(pixel) = (x\ coordinate\ of\ the\ pixel\ -\ 1) \times \frac{0.50M}{199} \quad (Equation\ 7)$$

The gradient is as follows:

6. Using program 3, find the color of the electrolyte for each image, and recalibrate those values if the lighting changed.
7. Plug those values into program 2, which will convert them to L*a*b values, and find the pixel in the gradient above that is the closest to the inputted value. It will run through all the pixels of the gradient above and return the one with the smallest difference in the shade (through Delta-E values) when compared with the imputed color of the electrolyte. Using equation 6, we can thus find the concentration of the solution at that time.

For example, for run 1, the values are as follows:

Figure 7: Table of the RGB values of the electrolyte, the position of the pixel closest to the color of the electrolyte, and the concentration of the electrolyte over time, Run 1

Time (min)	R	G	B*	Position of the pixel closest to the color of the electrolyte	Concentration of the electrolyte (M)
10	148	157	159	4	0.008
20	139	158	167	13	0.030
30	134	165	159	12	0.028
40	120	152	165	22	0.053
50	118	155	161	25	0.060
60	109	151	158	34	0.083
70	99	143	155	41	0.101
80	101	151	166	41	0.101
90	92	140	158	59	0.146
100	89	139	157	66	0.163
110	87	133	152	72	0.178
120	83	130	142	79	0.196
130	75	128	147	89	0.221

*(Note: although the average color is given in RGB format, when the data is analyzed, they are converted to L*a*b values.)

Qualitative data

As the electrolysis occurred, the electrolyte turned bluer and bluer. However, the bottom of the beaker seemed to turn bluer than the surface, as seen in figure 2. Once the current was stopped, a backward reaction could be seen as copper amassed at the bottom of the beaker, growing from the zinc electrode to the copper electrode. Leaving the cell untouched for a couple of hours would result in the electrolyte turning completely colorless again with all the copper in the solution depositing at the bottom. The temperature of the beaker also seemed to rise as the experiment took place, which could justify the reason why the current increased.

Comparing expected concentrations based on current flow to concentrations found using colorimetry

To make the control solutions, we weighed out an amount of solid copper sulfate before adding 100mL of distilled water. Those values had an uncertainty of 0.01g and 1mL respectively. However, these uncertainties don't directly affect the values given by colorimetry or by the model. Therefore, we will take the largest uncertainty for the control solution and add it to all our data points in the form of vertical error bars:

relative uncertainty of concentration = relative uncertainty of mass + relative uncertainty of volume

$$= \frac{0.01}{1.60} + \frac{1}{100} = 1.63\%$$

For the uncertainty of the values given by the model, we can include 0.01A as the uncertainty of the ammeter. As the charge isn't constant, we can again take the moment at which the uncertainty is the largest and add it to all of our data points:

relative uncertainty of charge = relative uncertainty of current + relative uncertainty of time

$$= \frac{0.01}{0.36} + \frac{1}{600} = 2.94\%$$

Therefore, that makes a total of 4.57% uncertainty for the values given by the model, and uncertainty of 1.63% for the values given by colorimetry. We can also add horizontal error bars of length 1 to both series as the stopwatch has an uncertainty of 1s, although it is insignificant and is thus negligible.

Figure 8: Table of the expected concentrations and concentrations found using colorimetry over time

Time (s)	Concentration from colorimetry (M)	Concentration from model (M)	Time (s)	Concentration from colorimetry (M)	Concentration from model (M)
600	0.008	0.012	4800	0.101	0.118
1200	0.030	0.025	5400	0.146	0.136
1800	0.028	0.039	6000	0.163	0.155
2400	0.053	0.053	6600	0.178	0.175
3000	0.060	0.068	7200	0.196	0.195
3600	0.083	0.084	7800	0.221	0.216
4200	0.101	0.101			

Concentration from colorimetry and Concentration from model over time

- Concentration from Colorimetry $\quad y = 3 \ 10^{-5} \ t \quad | \quad R^2 = 0.990$
- Concentration from Model $\quad y = (2.01 \ 10^{-5} \ t^2 \ + \ 0.378 \ t) / 19 \ 300 \quad | \quad R^2 = 1$
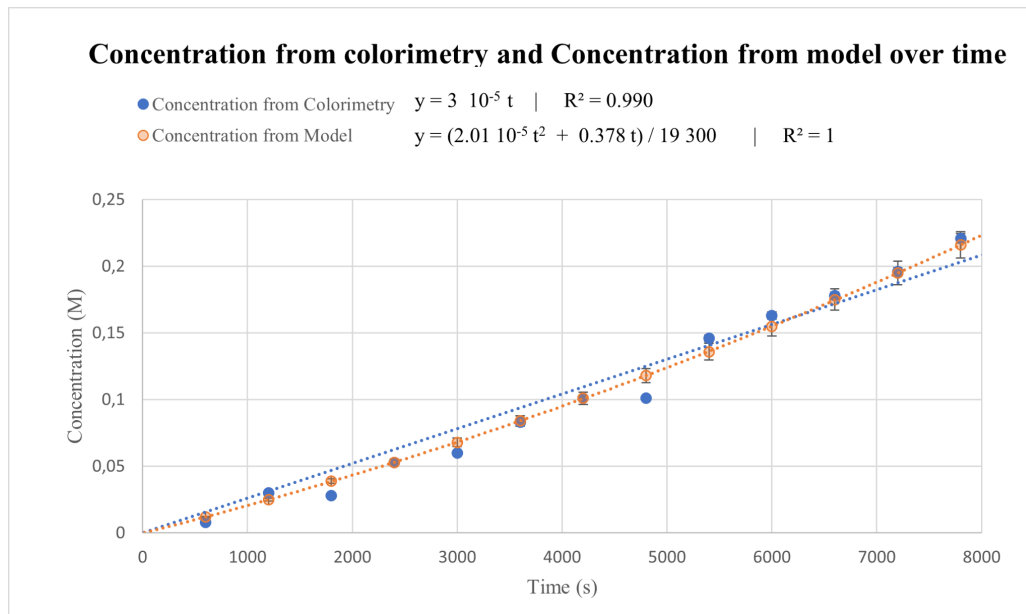
Figure 9: *Graph comparing the concentration of the electrolyte found by colorimetry and found by the model over time: **Trial 1**, author*
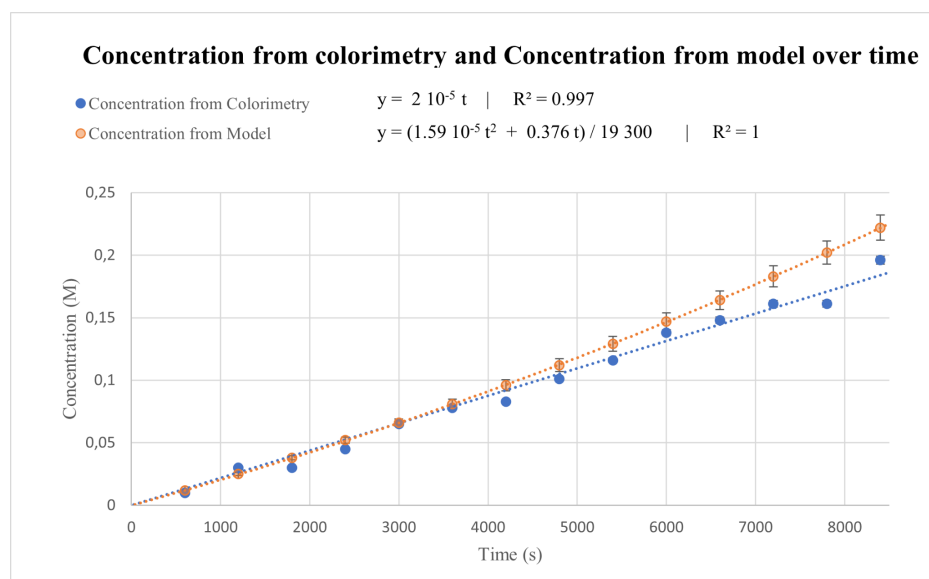
The final graph for the other run is as follows:[11]

Concentration from colorimetry and Concentration from model over time

- Concentration from Colorimetry $\quad y = \ 2 \ 10^{-5} \ t \quad | \quad R^2 = 0.997$
- Concentration from Model $\quad y = (1.59 \ 10^{-5} \ t^2 \ + \ 0.376 \ t) / 19 \ 300 \quad | \quad R^2 = 1$

Figure 10: *Graph comparing the concentration of the electrolyte found by colorimetry and found by the model over time: **Trial 2**, author*

In both graphs, both the curve based on colorimetry and the curve based on the model seems to show that concentration does increase in a near-constant manner over time, as the electrolysis takes place. All four curves have correlation coefficients near 1, indicating a strong correlation, thus showing the results are consistent. The same can be said about the current against time graphs, producing linear lines going through nearly all data points. The two lines of best fit in each respective concentration over time graph seem similar, with some data points overlapping. However, the values given by the model form a polynomial curve while the curve based on the values retrieved from colorimetry is linear for both trials.

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[11] The calculations and data can be found at the following links: https://postimg.cc/FYXkvBCb
https://postimg.cc/3WWKcF1K

Evaluation

To investigate whether there actually was a backward reaction, letting the reaction run for a longer period of time would have been interesting, as it would have been more apparent whether the colorimetry values were actually smaller than the model values, or if it only was for our current data sets. The reason I couldn't run the experiment for longer is that the zinc accumulating next to the zinc electrode would eventually reach the copper electrode and create a short circuit. This could have been solved by using a bigger beaker, or by separating the two electrodes with a plate that would prevent the zinc from going through but would still let the electrolyte solution circulate. Some other sources of errors in the experiment, their effect, and how they could have been prevented are summarized in the following table:

Figure 11: Table of the sources of errors in the experiment, their type and effect, and how they could have been prevented

Source of error	Type	Effect	How to prevent
Lighting isn't completely constant	Systematic	Some electrolytes might be lighted more strongly than others, affecting the "lighting" value of the colors.	Flashing a stronger light directly above the solutions (the lamp was slightly to the side in my experiment as the camera was taking up a lot of space).
Corrosion on the ends of the cable plugged into the electrodes	Random	The corrosion on the metal ends of the cables could have caused a change in resistance of the wires, affecting the potential difference across the wire	Ensuring the metal parts are not rusted, and only cleaning the end of the electrode in contact with the solution, so the voltage is the same across runs
Positioning of the electrodes	Random	The electrodes were positioned slightly differently across drops, changing the voltage entering the cell	While a foam board was placed to keep the distance between the electrodes constant, that distance was only constant at the top of the electrodes, and those might have been inclined towards the bottom. Making markings where they should be to control them would have been useful.
Delta-E comparison between 2 colors not being exact	Random	Two colors compared with Delta-E might not give the expected value that the human eye perceives	Using better cameras and technology to have more precise and accurate readings
Color not being exactly the same on every pixel of the solutions	Random	Some pixels might be a different color compared to other pixels in a solution, thus not measuring the color values exactly	Using bigger areas instead of rectangles when measuring the color of a solution, or select a single pixel that will be used for every recording to keep it consistent across recordings

Additionally, random error could have been reduced by recording values more often instead of every 10 minutes; for example, current might not increase as linearly as what was measured. However, as shown by the small uncertainties and correlation coefficient values, this would not have a huge impact.

The temperature might have had an impact on the current of the cell; however, as the current was tracked, this was accounted for.

Furthermore, impurities could have had an impact on the color of the solutions. For the control solutions, slight impurities such as sweat, manufacturing impurities (of the copper sulfate pentahydrate that was used), or improperly cleaned equipment by the previous user could impact their color. This is because impurities either reduce the concentration of the control solutions and/or the colors of the impurities may have conflicted with that of the solutions. However, it is safe to assume that the copper in the electrolyte solution was relatively pure, as it was made by electrolytic refining. Therefore, only the impurity of the control solutions could have had a meaningful impact on the experiment.

Conclusion

As expected, the electrolyte did turn bluer over time. We saw that in both cases, the model and the colorimetry values were relatively similar, although the model predicted a polynomial relationship while the colorimetry values seemed to show a linear relationship. The reason the model predicted a polynomial relationship is because the current of our electrolyte cell increased over time, meaning the concentration increased faster and faster. The colorimetry values were often smaller than the model values on the graphs, meaning the actual concentration of the electrolyte was lower than what was predicted. This could be due to the fact that as zinc accumulated on the zinc electrode, a reverse reaction occurred where zinc displaced copper ions in the copper sulfate electrolyte (as zinc is higher than copper in the activities series); this was even more apparent when the current was stopped and the solution turned colorless again but might have happened during the electrolysis too.

Using the "linest" function in excel, we can find the uncertainty of the gradient of both the model and the colorimetry curves. If we consider the model as the control values (as they are more precise and accurate, for reasons we will explore in the Evaluation section), we can also find the average experimental error of the colorimetry values, by applying the following formula:

$$Experimental\ Error\ (t)\ =\ \frac{Model\ Value\ (t)\ -\ Colorimetry\ Value\ (t)}{Model\ Value\ (t)}$$

Figure 12: Table of the average experimental error and the uncertainty of the gradient of the concentration over time graphs for the Excel Model and the colorimetry values, trial 1 and 2

Trial	Avg. Experimental Error (%)	Uncertainty of the gradient (%)	
1	3.99	Model	1.80
		Colorimetry	2.96
2	9.27	Model	1.54
		Colorimetry	1.60

The uncertainties of the gradient of all 4 lines range between 1 to 3%, meaning the uncertainty of this experiment was relatively small. An interesting observation is that the model values always had less uncertainty than the colorimetry values; this was also shown by the correlation coefficient values in the graphs: the model curves had R^2 values of 1, while the colorimetry curve had R^2 values between 0.990 and 0.997. This makes sense as the Delta-E comparison isn't perfect, and the areas of the solutions we used to measure color were imprecise.

We can also see the average experimental error values of the colorimetry recordings are below 10%, which shows the model accurately predicted the concentration of the electrolyte. This thus shows our results were coherent and reliable. As mentioned above, if the experiment was repeated for longer periods of time, it would have been possible to investigate the backward reaction.

Bibliography

1. Brown, C., & Ford, M. (2014). 07 Equilibrium. In Chemistry: Higher level: Supporting every learner across the IB continuum (pp. 442-443). Pearson Education Limited.

2. "Electrolysis Using Inert and Reactive Electrodes - MLC." Miracle Learning Centre, 17 Oct. 2022, https://miraclelearningcentre.com/chemistry-lessons-topic-electrolysis-using-inert-and-reactive-electrodes/.

3. Bukys, Liudvikas. "How to Compare Two Colors for Similarity/Difference." Stack Overflow, 26 Jan. 2012, https://stackoverflow.com/a/9019461.

4. Schuessler, Zachary. "Delta E 101". Delta E, 25 Sept. 2019, https://zschuessler.github.io/DeltaE/learn/.

5. Phillips, Ken. "What Is CIELAB Color Space?" *Color Measurement Spectrophotometer Supplier & Manufacturer*, HunterLab, 16 Oct. 2022, https://www.hunterlab.com/blog/what-is-cielab-color-space/.

6. "Copper Sulfate Pentahydrate." National Center for Biotechnology Information. PubChem Compound Database, U.S. National Library of Medicine, https://pubchem.ncbi.nlm.nih.gov/compound/Copper-sulfate-pentahydrate#section=Chemical-and-Physical-Properties.

7. "Safety Data Sheet - Uprm.edu." *Zinc Sulfate*, Global Safety Management, 19 Mar. 2015, https://www.uprm.edu/citai/wp-content/uploads/sites/222/2021/06/Zinc-Sulfate.pdf.

8. "Safety Data Sheet - Fisher SCI." *Copper(II) Sulfate*, ThermoFisher Scientific, 26 Dec. 2021, https://www.fishersci.com/msds?productName=AC422870050.

9. Moraes, Edgar P., and Mario R. Confessor. "Integrating Mobile Phones into Science Teaching To Help Students Develop a Procedure To Evaluate the Corrosion Rate of Iron in Simulated Seawater." *ACS Publications*, 2015, https://pubs.acs.org/doi/10.1021/acs.jchemed.2c00633.

10. Smith, James. "Ways to Work out the Average Colour of an Image." *Golang Project Structure*, 8 May 2022, https://golangprojectstructure.com/calculating-average-colour-from-images/.

11. "Color-Convert NPM." *Npm.io*, 2019, https://npm.io/package/color-convert.