

How does increasing the concentration of Magnesium $[Mg^{2+}]$ affect the cell potential, E_{cell} (voltage) in a Magnesium-Silver Voltaic Cell?

Justification

Batteries are a fundamental aspect of how the substantial amounts and usage of modern devices and technologies critical to that undergo daily use function. Investigating and researching the relationships arising from changing variables and reactions in the application of Voltaic cells can lead to improvements of its performance and elevate its efficiency, thus transforming how the surrounding environment functions. Hence it fulfills the statement of inquiry, "Chemical reactions bring out the changes that can be used to transform our immediate environment". Additionally, it is a direct real life implication that pertains to this unit, chemical reactions and redox reactions in particular, allowing the implementation of knowledge and theory into applied usage, modifying the physical environment.

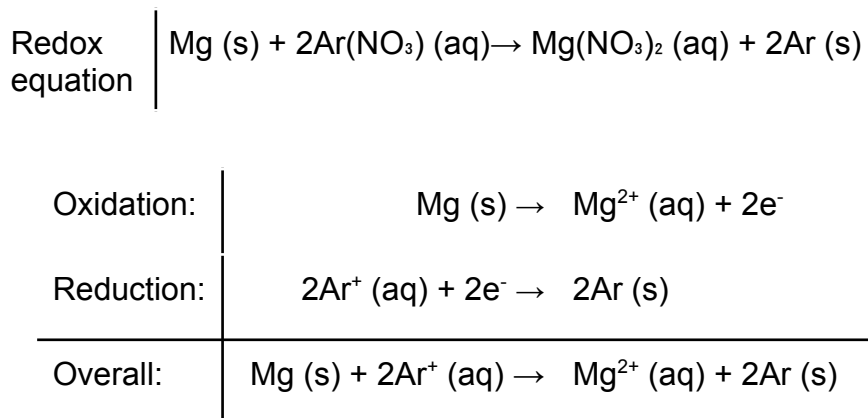
Introduction

Electrochemical cells are composed of two half-cells in which oxidation and reduction half-reactions occur in each (Ciobaru, "17.2"). In this experiment, a specific type of electrochemical cell, the voltaic cell, containing magnesium nitrate solution and silver nitrate solution with its respective electrodes and a salt bridge will be used to investigate the relationship that arises from changing concentrations of the former. Each solution is filled in two separate beakers - the half-cells, and contains a rod composed of its respective metal, in this case a magnesium and a silver rod. A salt bridge will be placed between the beakers and submerged in each solution to neutralize the half-cells as ions build up over time and electric charge differences increase, allowing the electron flow to continue ("Electrochemistry Basics").

According to Zephyrus, magnesium is more reactive than silver, therefore has a higher tendency to lose its electrons and form an ion ("The Reactivity"). Thus the direction of current flow will originate from magnesium and flow to silver. In order to measure the voltage of said current flow, otherwise known as the electric cell potential, a voltmeter will be attached between the magnesium anode and the sodium cathode with a black and red wire respectively ("Electrochemistry").

The movement of electrons within voltaic cells is ensured as the elements within this redox equation are not in equilibrium due to the usage of metals with different reactivity. In generalised terms, chemical species move from high to low electrochemical potential, hence an electron current will always be present so long as there is an imbalance in the electrochemical potential ("Equilibrium"). Therefore, as long as both the more reactive and less reactive metals remain present in the cell and are not enough to satisfy equilibrium, there will exist a current.

The following are the chemical equations associated with the reaction:



Hypothesis

The results of this experiment will be hypothesized according to the Nernst equation (Ciobanu, “Equilibrium Constant”, “Nernst”):

$$E = E^0 - \frac{RT}{zF} \ln Q$$

where:

E = reduction potential

E^0 = standard potential

R = universal gas constant

T = temperature (kelvin)

z = ion charge (moles of electrons)

F = Faraday’s constant

Q = reaction quotient

From observation and analysis, it is derived that the values of the standard potential, E^0 , the universal gas constant, R, the temperature, T, the ion charge, z, and the Faraday constant, F will all remain constant in the context of this experiment. The universal gas constant, R, and the faraday constant, F, are completely independent of the experiment and reaction as a whole as they are universal constants. The temperature, T, also occurs independently of the contents of the reaction, although it is alterable. The standard potential, E^0 , and the ion charge, z, is not independent from the experiment, however it will remain constant as it is only affected by the types of metal used and the number of electrons interchanged during the reaction. That is, regardless of the molarity of the two solutions, those values remain unchanged. Thus, the reaction quotient, Q, will be the only independent factor affected by the independent variable in this experiment.

The reaction quotient, Q, is equivalent to the quotient of the molarity of the reactants over the molarity of the products ("The Reaction Quotient"). It can be derived accordingly, with considerations of molar concentration of the reactant, $[Ag^+]$, being 1.00M factored in, resulting in the following:

$$Q = \frac{[\text{ions oxidized}]^p}{[\text{ions reduced}]^r} \quad \left| \quad Q = \frac{[Mg^{2+}]}{[Ag^+]^2} \quad \left| \quad Q = \frac{[Mg^{2+}]}{1} \quad \left| \quad Q = [Mg^{2+}] \right.$$

In context of the Nernst equation, it is evident that the resulting reduction potential, E, will be affected according to Q in a logarithmic relationship,

$$\ln Q = \ln[Mg^{2+}]$$

$$E = E^0 - \frac{RT}{zF} \ln[Mg^{2+}]$$

Upon further investigation, it is observed that increasing the molarity of the reactant, $[Mg^{2+}]$, will result in an increasing natural log, which decreases the value of the entire second term while turning it positive. As a result, the reduction potential, E, would decrease as the standard potential, E^0 , is added by a gradually decreasing value for $M < 1$ or subtracted by a gradually increasing value for $M > 1$. Overall, this results in a decreasing trend for the reduction potential, E.

Another observation pertaining to equivalence is also evident based on the computations in figure_, where when $M = 1$, or $M_r = M_p$ in a general case, $\ln Q = 0$. This leaves the entire second term as 0, thus the reduction potential, E, to be equivalent to the standard potential, E^0 . This signifies the equivalence between the reduction and the standard potential when an equivalence between the concentrations of both the products and reactants takes place.

Overall, plotting the reduction potential, E, against the molar concentration of Magnesium should result in a decreasing pattern, likely an exponential decay due to the subtraction of a logarithmic function. If plotting against the natural logarithm of the concentration instead, it would result in a decreasing linear relationship, as the logarithm would cancel out the exponential decay. In both cases, it is expected that the reduction potential at the concentration of 1M or $\ln 1M$ should be equivalent or close to the standard potential of magnesium and silver reaction.

Variables

Variables	Procedure/Method
Independent variable	
The independent variable in this experiment is the molar	The independent values used are the following: 0.0001M, 0.001M, 0.01M, 0.1M, 1M. Each of these concentrations will be formulated

concentration of Magnesium Nitrate $[Mg^{2+}]$ that will be used in a Voltaic cell system of Silver and Magnesium.

from a 1M solution through the following method:

90:10 mixture of distilled water and the previous solution of 1 order of magnitude higher in molar concentration. Using this calculation, a minimum of 277.775ml of 1M solution is needed to formulate the rest.

Figure 1. Dilution calculation for magnesium nitrate.

Molarity (M)	Distilled water (90%, ml)	Previous solution (10%, ml)	Total volume (ml)
0.0001	225	25	250
0.001	247.5	27.5	275
0.01	249.75	27.75	277.5
0.1	249.975	27.775	277.75
1	N/A	N/A	277.775

However these are not a requirement, as long as the total volume does not fall below the values on the table, and the ratio of distilled water against the previous solution remains 90:10, there will be a sufficient amount of accurate solutions for the experiment.

Dependent variable	
The dependent variable is the cell potential, E_{cell} measured in voltage	The voltmeter will be attached to both electrodes with a red and a black wire, measuring the current from the anode rod (Mg) to the cathode rod (Ag) in Volts (V).
Controlled variables	
Temperature	All trials will be conducted within the same environment with a 25°C room temperature, ensuring that both solutions are at 25°C each trial using a thermometer. Additional temperature manipulation may be required if the room temperature does not satisfy the conditions - ie, the use of water baths.
Concentration of Silver Nitrate	The magnesium nitrate will be kept at a concentration of 1.00M throughout the experiments, to ensure all trials have the same starting concentration, the solutions can not be used for repeated trials. Each must be newly obtained from the large batch every trial.
Ionic compound and concentration in salt bridge	The salt bridge will be created by dipping a rolled up cellulose paper in a solution of 2M sodium nitrate $[Na(NO_3)]$ solution for 1 minute. A timer is used to ensure consistent dip duration throughout.
Length of salt bridge	The cellulose paper 5cm in width will be rolled and cut at a length of 10cm, using a ruler to ensure it remains consistent throughout all 25 pieces.
Surface area of metal	The magnesium and silver rods are identical in size and the same

electrode	rods are used throughout. Each time sanding them with 120 grade sandpaper as equally distributed across the submerged area as possible.
Volume and depth of solution in each half-cell	50ml of both the magnesium and silver nitrate solutions are used for each trial, filling in the same dimensions of beaker. 50 ml is measured with a graduated cylinder. Therefore both volume and depth is conserved.
Duration of reaction	Take note of the voltage after 5 seconds of placing both rods in their respective solutions. Use a timer for accuracy.

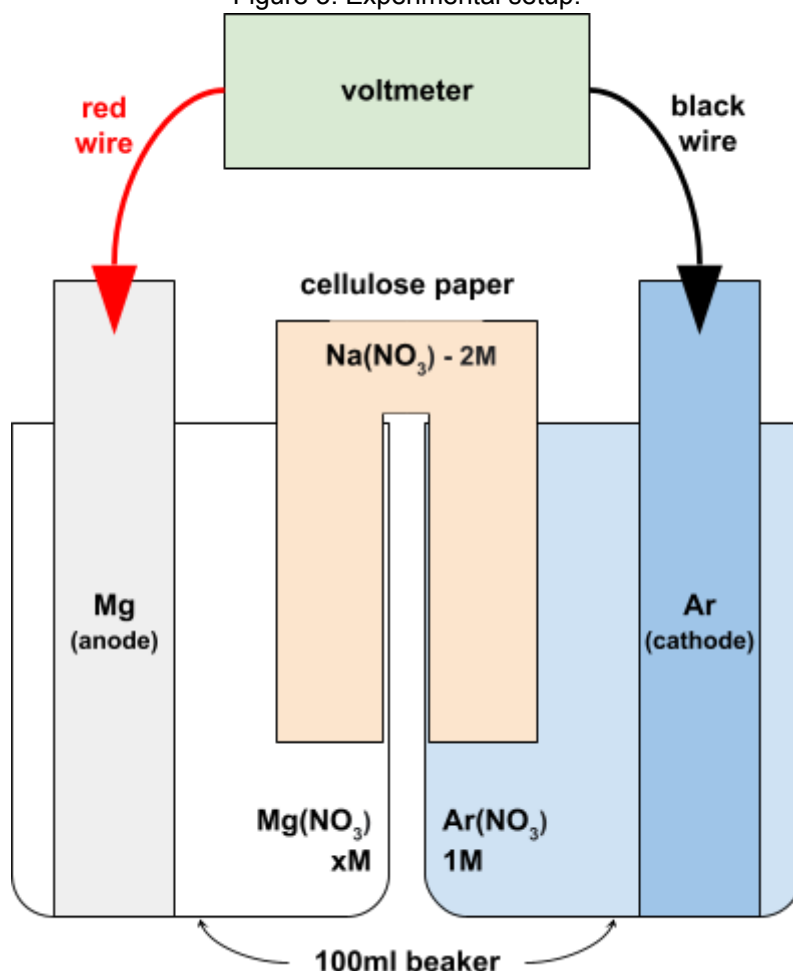
Apparatus

Figure 2. Apparatus list

<i>Apparatus</i>	<i>Type/Size</i>	<i>Quantity</i>
Voltmeter	Volts, DC	1
Magnesium rod	8 by 1 cm	1
Copper rod	8 by 1 cm	1
Silver Nitrate solution	1.00M	1,250ml
Magnesium Nitrate solution	1.00M	300ml
Sodium Nitrate solution	2.00M	200ml
Water	Distilled	1,000ml
Wire	with crocodile clips and connector leads	2 (one red, one black)
Beaker	100ml	2
Sandpaper	120 grade	2
Cellulose paper	5cm width	25
Graduated cylinder	100ml	1
Thermometer	Degrees celsius (°C)	1
Timer	Seconds, milli seconds	1
Bowl	At least 10cm wide	1
Scissors	N/A	1
Ruler	Cm	1

Set up

Figure 3. Experimental setup.



Method

1. Prepare 25 rolls of cellulose paper with 5cm width, cut all to 10cm length.
2. Connect the voltmeter, metal rods and wires according to the diagram above.
3. Sand the magnesium and silver rods with 120 grade sandpaper, ensure that all areas are sanded equally to the best of abilities.
4. Place in one roll of cellulose paper in a bowl with 200 ml of 2M Na(NO₃) solution for 1 minute, use a timer for accuracy.
5. Set up two 100ml beakers close together and add in 50ml of 0.0001M Mg(NO₃) solution into the first beaker and 1M Ar(NO₃) solution in the second beaker. Use a graduated cylinder to measure.
6. Place the cellulose paper in between the two beakers like the diagram shows. Try to ensure equal lengths of the cellulose paper are submerged in both beakers.
7. Place in the metal rods in their respective beakers at the same time, start the timer and record the voltmeter reading at 5 seconds.
8. Repeat steps 3-7 4 more times to reach 5 trials total

9. Repeat steps 3-8 4 more times, however replace the 0.0001M $\text{Mg}(\text{NO}_3)$ solution with 0.001M, 0.01M, 0.1M, and 1M solutions. Formulation of each solution has been detailed in the variables sections and figure 1.

Risk Assessment

Standard safety measures apply.

Chemical risks:

Magnesium nitrate, magnesium ("Magnesium")

- Magnesium nitrate in solid form can be a fire hazard, serious skin, eye and respiratory irritation. In solutions with 0.5M or higher, magnesium nitrate is irritating to eyes and skin, at 0.5M and lower, it has low hazards.

Silver nitrate, silver ("Silver")

- Silver nitrate in solid form can explode when coming in contact with magnesium powder and water, leading to many injuries. In concentrations of 0.3M and more, it's very corrosive, causing severe eye damage and skin burn, and internal damage due to absorption in the blood, deposition of silver in various tissues if swallowed. Between 0.18 to 0.3M, it can cause eye damage and skin irritation, producing black stains on the skin which eventually wears off. Between 0.06M to 0.18M, it is irritating to eyes and skin. Any lower than 0.06M is low hazard.

Sodium nitrate, sodium ("Sodium")

- Sodium nitrate in both solid and solution form at 3M and more are harmful when swallowed, at 1M or more is irritating to the eye and skin

To reduce the described risks, avoid raising dust if solid substances are used, avoid directly handling silver nitrate solids and do not create solutions with concentrations higher than the experiment conditions. Since many of these elements are irritating to the skin and eye, using gloves on top of goggles will be beneficial.

Ethical/Environmental considerations

- Silver nitrate between 0.18M to 0.3M is extremely toxic to aquatic life, therefore it is a necessity to conduct proper disposal of it as well as the magnesium and sodium nitrate ("Silver").
- The experiment conditions already utilize large amounts of the 3 different solutions as well as apparatus such as cellulose paper, contributing to resource wastage and the eventual loss of metals on earth, which are a finite resource ("Metal"). Thus, it's crucial to avoid excessively using more of the apparatus than necessary.
- Likewise, large amounts of distilled water is used in this experiment, especially to formulate the different magnesium nitrate solutions. As a general precaution, limit any excessive use and waste of water.

Data Collection

Figure 4.The Reduction Potential (V) of a Magnesium-Silver Voltaic Cell System vs. Molar Concentration of Magnesium Nitrate Solution.

Mg(NO ₃) Concentration (M)	ln[Mg ²⁺]	Electric Cell Potential (V)					
		1	2	3	4	5	Average
0.0001	-9.2103	3.284	3.281	3.274	3.276	3.285	3.280
0.001	-6.9078	3.259	3.258	3.247	3.246	3.263	3.255
0.01	-4.6052	3.212	3.219	3.194	3.227	3.228	3.224
0.1	-2.3026	3.189	3.188	3.194	3.190	3.199	3.192
1	0	3.166	3.166	3.176	3.165	3.175	3.170

Figure 5.

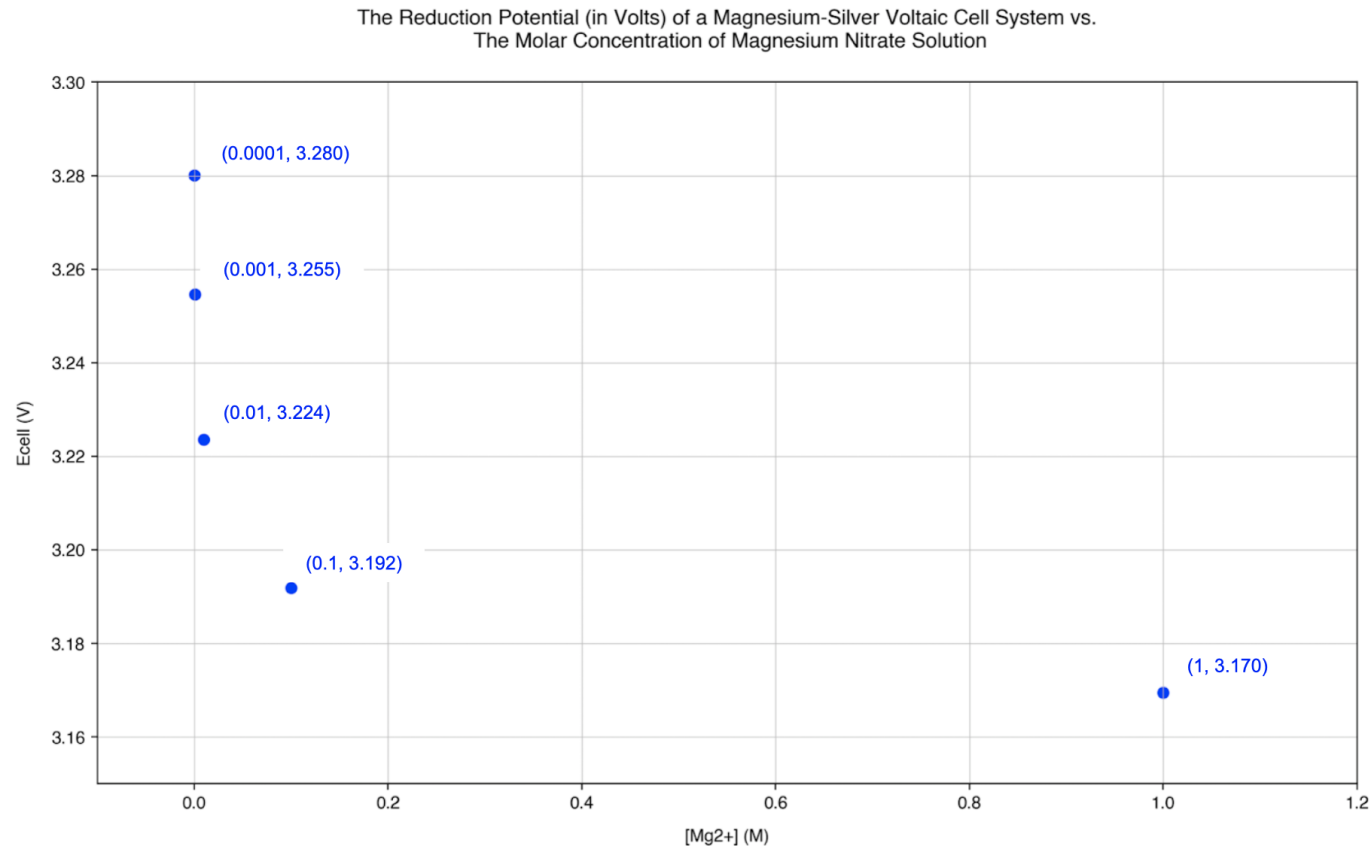
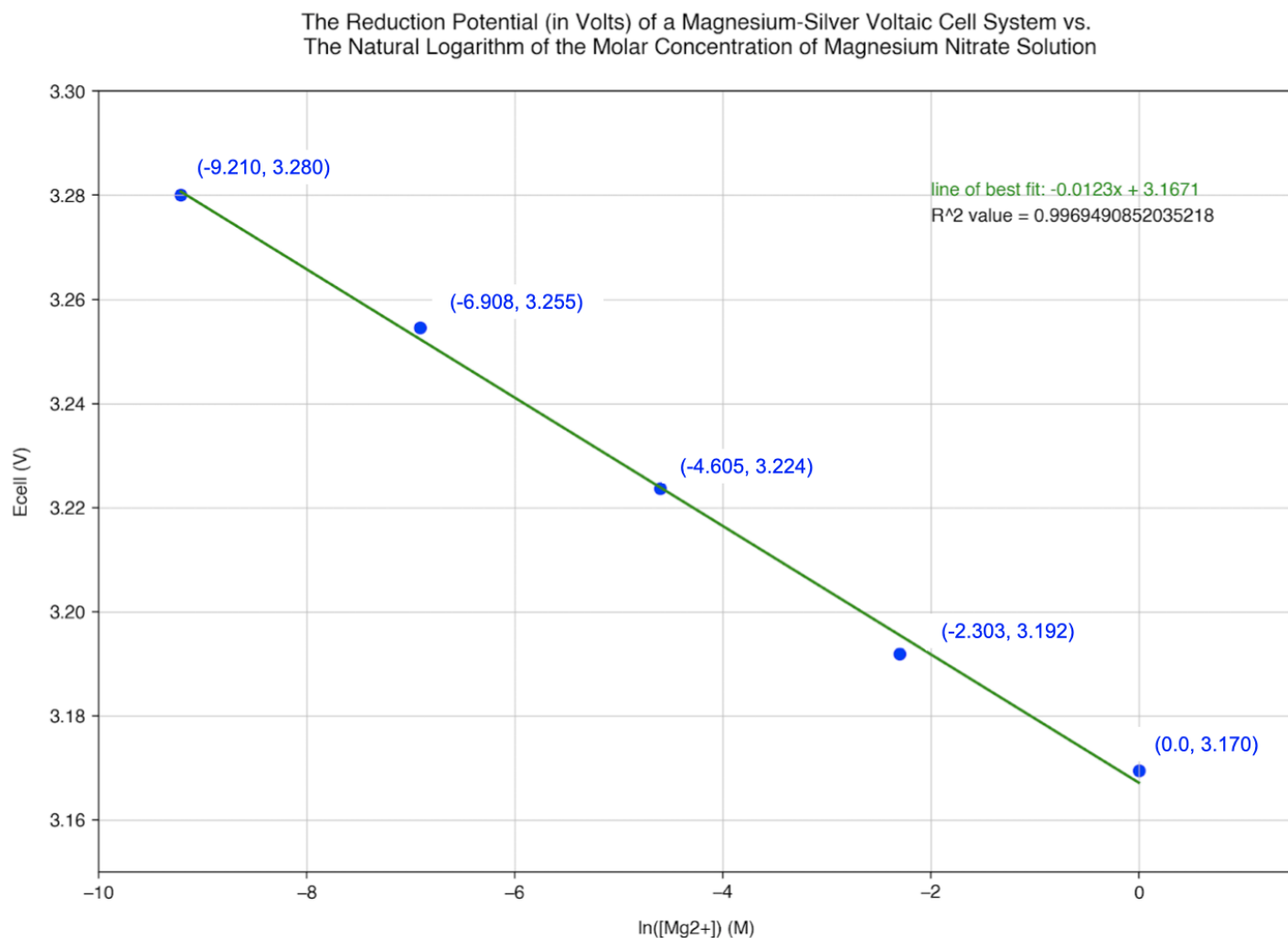


Figure 6.



Conclusion

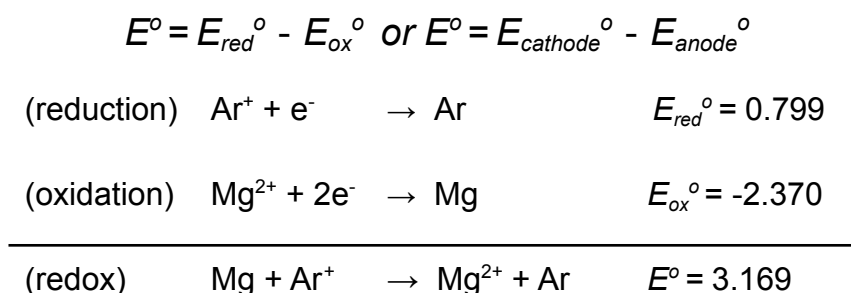
Increasing the concentration of magnesium nitrate in a Magnesium-Silver Voltaic cell system resulted in a decreasing electric cell potential that follows an exponential decay pattern as apparent in figure 5. However due to difficulty with generating accurate exponential decay models to fit data, plotting the y values to the natural log of the concentration of magnesium nitrate provided for a more efficient analysis.

Overall, the data demonstrates that concentrations of 0.0001, 0.001, 0.01, 0.1, 1M generated an average electric cell potential of 3.280, 3.255, 3.224, 3.192, 3.170 Volts respectively. In the natural log plot, the r^2 value of the data against the linear regression function of $y = -0.0123x + 3.1617$ is 0.9969, a relatively high precision. Small deviations from the trend line according to figure 6 is evident, namely for molar concentrations of 0.001M, 0.1M and 1M. However there are no major anomalies present that greatly affected the accuracy or the precision of the results as evident in the r^2 value.

If there had existed larger anomalies and outliers that cause more deviation from the trend, it could be a result of the many inconsistencies surrounding the controlled variables as well as systematic errors within the apparatus itself, many of which are discussed in the evaluation. A major

method that could potentially result in lack of precision is the formulation of the 5 different magnesium nitrate solutions. Producing each of the solutions, required multiple accumulative usage of the measuring cylinder which has an uncertainty of $\pm 0.5\text{ml}$. One or two cumulative uses alone would not have a great effect, but in this case, the 0.0001M would have had around 20 uses to arrive at the final solution. The rest of the factors that can affect deviations are discussed in the evaluation.

The hypothesis concludes that the results should follow an exponential decay trend and a decreasing linear trend when plotted against the molar concentration of magnesium nitrate and its natural logarithm respectively. The data collected clearly supports the hypothesis as discussed above and as apparent in figure 5 and 6 respectively. Additionally, a prediction of the out of 1M concentration was also proven true in the experimental data, where the reduction potential of magnesium and silver, 3.169 , is equivalent to the one achieved in the result, 3.16947 (Bertrand). The reduction potential, E^0 is calculated by subtracting the standard potential of the oxidized element from the reduced element ("Equilibrium Constant", "17.2", "Electrochemistry Basics"):



However, one thing that was not considered in the hypothesis was the gradient of the linear function plotted with the natural logarithm as the x axis. According to ("Equilibrium Constant", "Lab"), when generalising the Nernst equation and its constant variables (universal gas constant, R , temperature, T , and Faraday's constant, F), it would result in the following:

$$\frac{RT}{zF} = \frac{(8.314)(298.15)}{z(96485)} = \frac{0.0257}{z}$$

$$\therefore E = E^0 - \frac{0.0257}{z} \ln Q$$

Applying the conditions of this experiment, z would be equivalent to 2, hence the equation becomes the following, the latter of it rearranged:

$$E = E^0 - 0.01285 \ln Q$$

$$E = -0.01285 \ln Q + E^0$$

Plugging in E^0 as derived above results in:

$$E = -0.01285 \ln Q + 3.169$$

Unsurprisingly, the rearranged equation closely mirrors the function given by the linear regression:

$$y = -0.0123x + 3.1617$$

where $x = \ln Q = \ln[\text{Mg}^{2+}]$. The gradient and vertical shift of the linear regression resembles those of the rearranged nernst equation, only off by 0.0062V and 0.0073V respectively. This suggests that the experiment was conducted fairly accurately and precisely.

The following contains the results of a zinc-copper voltaic cell system included in a lecture from the University of Florida ("an"). In this case, the molar concentration of zinc is the independent variable and is plotted against the logarithm of the reaction quotient, as described above (molar concentration of zinc over that of copper). Like magnesium with silver, zinc is more reactive than copper, hence the resulting trend is similar to that of this experiment, they are both decreasing linear equations. Furthermore, it also identifies the equivalence between the standard potential of zinc and copper and the resulting reduction potential at 1M, exactly as observed in this experiment. Unfortunately, the source did not present the equation of the corresponding linear trend.

The source also comments on what the demonstrated trend fundamentally represents, which was not thoroughly acknowledged and discussed in the hypothesis and this section thus far. The lower the concentration of the oxidizing element relative to the reducing element, the more work the cell does. Otherwise, when the reaction quotient, $Q < 1$, the more work the cell does. The opposite also applies - when $Q > 1$, the less work the cell does, that is until the cell decreases until equilibrium and no longer does work ("Electrochemistry", "an"). This is where the electrochemical potential is equivalent in both half-cells ("Equilibrium"). For a magnesium-silver voltaic cell, this can be calculated by equating the linear equation to 0 where $x = \ln Q$, then solving for Q:

$$y = -0.0123 \ln Q + 3.1617$$

$$0 = -0.0123 \ln Q + 3.1617$$

$$0.0123 \ln Q = 3.1617$$

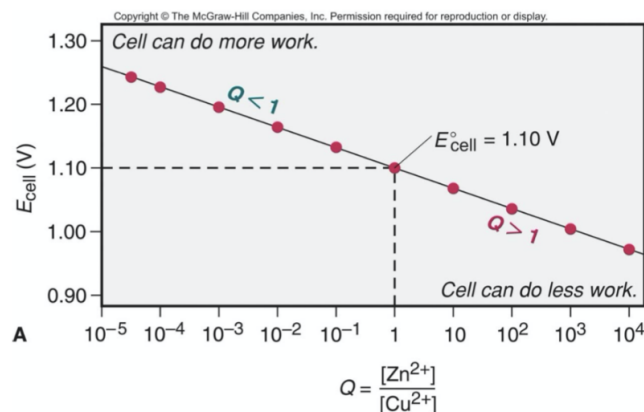
$$\ln Q = \frac{3.1617}{0.0123}$$

$$Q = e^{257.0488}$$

$$\therefore Q = 4.3139 \cdot 10^{111}$$

This implies the magnesium-silver voltaic cell will eventually reach equilibrium and achieve no work whatsoever at a molar concentration in the ranges of the 111th place, provided that the concentration of silver remains at 1M. Evidently, it is highly unrealistic that this can be tested due to safety considerations. It does, however, lead to another aspect of the experiment that was not discussed previously.

Figure 7. Results from the University of Florida ("an")



The fundamental function behind this experiment is Le Chatelier's principle, which states the position of a dynamic equilibrium will shift to reestablish equilibrium, counteracting the disturbance caused by changing conditions (Clark). Additionally, it implies that increasing the concentration of the reactant drives the reaction towards the products, and increasing that of the product drives the reaction towards the reactant, essentially creating more of the reactants ("The Reaction Quotient", "factors"). Therefore, when the magnesium nitrate increases in concentration, there will be a greater 'push' for more silver to dissociate from its nitrate compound and become solid according to Le Chatelier's principle. As a result, the current will flow at a higher voltage to satisfy the increased demand of the right leaning equilibrium.

Evaluation

<i>Limitations of experiment</i>	<i>Effect on experiment</i>	<i>Improvements/Extensions</i>
The distribution of the independent variable values are not well spaced out, and it could have included magnesium concentrations that were higher than that of silver.	Firstly, it was more difficult to observe the raw exponential decay trend as seen in figure 5 since the first 4 points are condensed with each other. Secondly, not having values that exceed the silver concentration prevents gaining a holistic view of the relationship, and only represents the cell producing more work, equal work, but not less work than the reduction potential, as represented in figure 7.	Choosing a more distributed range of iv values that include one or two where where $[Mg^{2+}] > [Ag^+]$ to extrapolate the results and make more observations and conclusions on the overall relationship.
<u>Random errors</u>		
The salt bridge is subject to inconsistency since it's heavily predisposed to human error in the preparation and usage.	Relying on human methods to roll and cut the cellulose paper to size, equally distribute the submersion, and place it between two beakers so it is submerged equally is prone to errors and disrupt variables that should be controlled. This creates discrepancies in the precision of the data collected, as more data points are more scattered apart.	Instead of creating the alt bridge by hand, there are already existing salt bridges that are used for voltaic cells. The relative advantage is its non changing shape, therefore it's much easier to control and keep consistent throughout.
Inability to control the room temperature and therefore the overall temperature of the reaction.	This would affect the temperature variable in the nernst equations, thus altering the expected results and trend line by making them shift higher or lower vertically. Not to mention, the temperature wouldn't also be constant across all trials, leading to fluctuating data. Thus,	To ensure better control over the temperature, use a water bath, set it to 25°C and use it for the experiment. This would result in more accurate and consistent temperatures.

	the results will be inaccurate throughout.	
Sanding the surfaces of both metal rods proved to be prone to inconsistencies as it is near impossible to ensure they are sanded the exact same amount and have the same resulting surface area.	This in turn generates errors and irregularities in the final data. Overall making the data contain more imprecision and inaccuracy.	Although it goes against the notion of resource saving for these experiments, replacing the metal rods with new ones every trial is the best option to minimize surface area differences to a minimum
Build up and corrosion on the metal rods can lead to drastic changes in surface area. Sanding helps limit the change but even sanding itself is inconsistent.	The result of these changes are the inconsistent amount of electrons being released from Mg and bonded with Ar^+ . Subsequently, it will cause more deviation from the trend and affect the precision of the experiment	
<u>Systematic errors</u> The following are some systematic errors that occurred due to the uncertainties in the tools and apparatus used to conduct the experiment. Most of the time, there are not many available ways to improve on them, their uncertainties are simply factored into the results.		
<u>Measuring cylinder ($\pm 5\text{ml}$)</u> Generating the solutions for each concentration involves multiple steps that are prone to systematic error - mixing 10% of the previous solution with 90% distilled water using a graduated cylinder.	Even if the molar concentration is off by a relatively small amount, it could result in significant changes in the cell potential. Therefore as these uncertainties accumulate, the unwanted changes do to.	
<u>Thermometer ($\pm 0.5^\circ\text{C}$)</u> Ensuring the temperature of each solution is at 25°C will have uncertainty of its own.	Small changes and inconsistencies in the temperature would not have a great effect on the overall data since it would only change by less than 1-2% in kelvin, assuming the differences do not exceed 4°C . However they still have the ability to generate minor discrepancies and outliers in the results which can affect the overall trend.	
<u>Ruler ($\pm 5\text{mm}$)</u> Measuring the cellulose paper and creating cuts based on the measurements.	Combined with the random error from manually conducting the process of cutting the salt bridge, uncertainty in the measurements itself can lead to fluctuating errors in the results.	

Overall, there are many factors that can lead to discrepancies and inconsistencies in the data collection and therefore affect the quality of the results. In the case of this experiment, they did not cause too much deviation and overall the trends observed were accurate to the research and other published data. That being said, there is an extension to the experiment that could improve the validity of the result, namely adding in more independent variable values that will provide for a more well rounded representation of the relationship, allowing better observation of the exponential decay trend and encompassing the entirety of the connection with work.

Bibliography

- "An Application: the Reaction of Metals with Acids." *Chem University of Florida*,
<https://www.chem.ufl.edu/wp-content/uploads/sites/43/2017/03/Chapter-21-Electrochemistry-Week-2.pdf>. Accessed May 2021.
- Bertrand, Gary L. *ELECTROCHEMICAL CELLS*. online voltaic cell simulation.
<https://web.mst.edu/~gbert/Electro/Electrochem.html>. Accessed May 2021.
- Ciobanu, Madalina, et al. "1 - Fundamentals." *Handbook of Electrochemistry 2007*, 2007, pp. 3-29. *Elsevier*,
<https://www.sciencedirect.com/science/article/pii/B9780444519580500021>. Accessed May 2021.
- Clark, Jim. "Le Chatelier's Principle Fundamentals." *Chem LibreTexts*, 15 Mar 2021,
[https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_\(Physical_and_Theoretical_Chemistry\)/Equilibria/Le_Chateliers_Principle/Le_Chatelier's_Principle_Fundamentals](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Equilibria/Le_Chateliers_Principle/Le_Chatelier's_Principle_Fundamentals). Accessed May 2021.
- "Electrochemistry Basics." *Chem LibreTexts*, Chem LibreTexts,
[https://chem.libretexts.org/Bookshelves/Analytical_Chemistry/Supplemental_Modules_\(Analytical_Chemistry\)/Electrochemistry/Basics_of_Electrochemistry](https://chem.libretexts.org/Bookshelves/Analytical_Chemistry/Supplemental_Modules_(Analytical_Chemistry)/Electrochemistry/Basics_of_Electrochemistry). Accessed May 2021.
- "Electrochemistry: Galvanic Cells and the Nernst Equation." *TheChemCollective*, 11 July 2016,
http://chemcollective.org/chem/electrochem/step2_cell.php. Accessed May 2021.
- "Equilibrium Constant and Cell Potential." *Lumen Learning*,
<https://courses.lumenlearning.com/introchem/chapter/equilibrium-constant-and-cell-potential/>. Accessed May 2021.
- "Equilibrium Constant and Free Energy Change for an Electrochemical Cell." *Chem Purdue*, Chem Purdue,
https://www.chem.purdue.edu/gchelp/howtosolveit/Electrochem/Equilibrium_constant_free_energy.htm. Accessed May 2021.
- "Factors that Affect Chemical Equilibrium." *Lumen Learning*, Lumen Learning,
<https://courses.lumenlearning.com/boundless-chemistry/chapter/factors-that-affect-chemical-equilibrium/>. Accessed May 2021.
- "Lab 13 - Electrochemistry and the Nernst Equation." *Web Assign*, Web Assign, 2011,
https://www.webassign.net/labsgraceperiod/ucscgenchem11/lab_13/manual.html. Accessed May 2021.
- "Magnesium and Calcium salts." *CLEAPSS*, CLEAPSS,
<http://science.cleapss.org.uk/resource/SSS036-Magnesium-and-calcium-salts.pdf>. Accessed May 2021.
- "Metals." *BBC*, BBC, 2021, <https://www.bbc.co.uk/bitesize/guides/zv4g4qt/revision/3>. Accessed May 2021.
- "Nernst Equation Calculator." *Physiology Web*, Physiology Web, 12 June 2011,
https://www.physiologyweb.com/calculators/nernst_equation_calculator.html. Accessed May 2021.
- "The Reaction Quotient." *Chem LibreTexts*, MindTouch, 16 August 2020,
[https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_\(Physical_and_Theoretical_Chemistry\)/Equilibria/Chemical_Equilibria/The_Reaction_Quotient](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Equilibria/Chemical_Equilibria/The_Reaction_Quotient). Accessed May 2021.
- "The Reactivity Series." *Zephyrus*, Zephyrus, <https://www.zephyrus.co.uk/thereactivityseries.html>. Accessed May 2021.
- "17.2 GALVANIC CELLS." *Open Text*, <https://opentextbc.ca/chemistry/chapter/17-2-galvanic-cells/>. Accessed May 2021.
- "Silver and its compounds." *CLEAPSS*, CLEAPSS, 2019,
<http://science.cleapss.org.uk/Resource/SSS046-Silver-and-its-compounds.pdf>. Accessed May 2021.
- "Sodium and potassium salts." *CLEAPSS*, CLEAPSS,
<http://science.cleapss.org.uk/resource/SSS034-Sodium-and-potassium-salts.pdf>. Accessed 2019 May 2021.

Appendix

Raw code data

0.0001

[, 3.2807496010966934, 3.2742176422484577, 3.2762274892605032, 3.2846110182246333]

16.40021095231817

0.001

[3.2592718633528017, 3.2577511105735613, 3.246928661095459, 3.2460380985014865,

3.2627992001431307]

16.27278893366644

0.01

[3.2118822641584237, 3.2193703093834394, 3.231380128047131, 3.2271502615669285,

3.228343089644301]

16.11812605280022

0.1

[3.18913396698001, 3.1877840154299255, 3.193589488105983, 3.189526377361286,

3.1991198798487477]

15.959153727725953

1

[3.1656012671640466, 3.1659646380372597, 3.1760981139190876, 3.164944378727197,

3.174717770180586]

15.847326168028179

[3.2800421904636337, 3.254557786733288, 3.2236252105600443, 3.1918307455451904,

3.169465233605636]

[0.0001, 0.001, 0.01, 0.1, 1] [3.2800421904636337, 3.254557786733288, 3.2236252105600443,

3.1918307455451904, 3.169465233605636]

Code used to generate and graph data

```
import math
```

```
import numpy as np
```

```
from numpy import log as ln
```

```
import matplotlib.pyplot as plt
```

```
import scipy.optimize
```

```
def nernst(h):
```

```
    y2_list = []
```

```
    s1 = 0.799 # standard potential of first element
```

```
    s2 = -2.37 # standard potential of second element
```

```
    e0 = s1 - s2 # standard potential
```

```
    z = 2 # ion
```

```
    r = 1
```

```
    for element in h:
```

```
        y_list = []
```

```

    print(element)
    for i in range(5):
        o = element
        E = e0 - ((8.314 * 298.15) / (z * 96485)) * ln(o / r)
        E = E + np.random.normal(-0.005, 0.005)
        y_list.append(E)
    print(y_list)
    print(sum(y_list))
    y2_list.append(sum(y_list) / 5)
print(y2_list)
return(h, y2_list)

```

```

def plotgraph(M, y):
    plt.figure()
    plt.rcParams['font.sans-serif'] = 'helvetica'
    plt.title("The Reduction Potential (in Volts) of a Magnesium-Silver Voltaic Cell System vs.\n The Natural Logarithm of the Molar Concentration of Magnesium Nitrate Solution", pad=20, ha='center')
    plt.xlabel("ln([Mg2+]) (M)", labelpad=10)
    plt.ylabel("Ecell (V)", labelpad=10)
    plt.xlim([-10, 1.5])
    plt.ylim([3.15, 3.3])
    M = ln(M)

    m, b = np.polyfit(M, y, 1)
    plt.plot(M, m * M + b, 'g')
    best = "line of best fit: " + str(round(m, 4)) + "x " + "+" + str(round(b, 4))
    plt.text(-2, 3.28, best, c='g')

    correlation_matrix = np.corrcoef(M, y)
    correlation_xy = correlation_matrix[0, 1]
    r_squared = correlation_xy**2
    rrr = "R^2 value = " + str(r_squared)
    plt.text(-2, 3.275, rrr)

    Mr = [round(M[i], 5) for i in range(len(M))]
    yr = [round(y[i], 5) for i in range(len(y))]

    for ij in zip(Mr, yr):
        plt.annotate('%s, %s' %ij, xy=ij, xytext=(10, 10), textcoords='offset points', c='b')

    plt.scatter(M, y, c='b')
    plt.grid(linestyle='-', linewidth=0.5)
    plt.show()

```

```

def func(x, a, b, c):
    return a * np.exp(-b * x) + c

```

```

def plotgraphex(M, y):
    plt.figure()
    plt.rcParams['font.sans-serif'] = 'helvetica'
    plt.title("The Reduction Potential (in Volts) of a Magnesium-Silver Voltaic Cell System vs.\n The Molar Concentration of Magnesium Nitrate Solution", pad=20, ha='center')
    plt.xlabel("[Mg2+] (M)", labelpad=10)

```



```
plt.ylabel("Ecell (V)", labelpad=10)
plt.xlim([-0.1, 1.2])
plt.ylim([3.15, 3.3])
```

```
Mr = [round(M[i], 5) for i in range(len(M))]
yr = [round(y[i], 5) for i in range(len(y))]
```

```
for ij in zip(Mr, yr):
    plt.annotate('%s, %s' %ij, xy=ij, xytext=(10, 10), textcoords='offset points', c='b')
```

```
plt.scatter(M, y, c='b')
plt.grid(linestyle='-', linewidth=0.5)
plt.show()
```

```
a = [0.0001, 0.001, 0.01, 0.1, 1]
M, y = nernst(a)
```

```
print(M, y)
plotgraph(M, y)
plotgraphex(M, y)
```