Effect of electrode electronegativity on standard cell potential

λ 1. Problem/Purpose

Determining the effect of increasing the difference in electrode electronegativity on the voltage produced by a galvanic cell.

λ 2. Background

In Chemistry Class, we have always studied electronegativity as just another periodic table trend, but I have always considered the concept to be the gold nugget of the periodic table. As a numerical value assigned to each element using the Pauling scale, electronegativity represents the tendency of an atom to attract electrons towards it, thus the higher the value, the greater the attractive tendency. How attractive is that! This tendency is facilitated by not only the element's inherent affinity for nuclear attraction, but also the distance between their valence electrons and nucleus. For example, beryllium has a higher electronegativity than lithium because beryllium has a larger nucleus, more attractive force on its valence electrons, and an overall smaller atomic/ionic radius. Generally, what I have always loved about this concept is its simplicity, but even more than that, how it can shape our understanding of generating electrical energy in our increasingly industrial world.

Here's the logic: when used to create a galvanic cell, cathodes tend to have higher electronegativity. Since they attract electrons more strongly it makes sense that they will get reduced (gain electrons) because they are essentially attempting to pull electrons off of the anode and into the other side of the cell to plate the cathode. Hence, if the difference in electronegativity of the two electrodes is increased, with the anode always having the lower electronegativity, then it is possible to infer that the standard cell potential will increase with the increase in attractive force.

Galvanic cells http://butane.chem.uiuc.edu/pshapley/GenChem2/C6/1.html

Galvanic cells are electrochemical cells made up of two half-reactions involving an anode (positively charged due to the absence of electrons) and a cathode (negatively charged due to the receiving of electrons), which are both considered electrodes (metals that participate in the generation of electricity in a galvanic cell). When the anode gets oxidized (loses electrons) and the cathode gets reduced (gains electrons), electrons can flow from the anode to the cathode via a conducting wire. In a galvanic cell specifically, the redox reaction is spontaneous because the anode is a stronger reducing agent than the cathode. In fact, it is often the strongest reducing agent, whereas the electrolytic solution of the cathode contains the strongest oxidizing agent in ionic form. It is for this reason that the cathode becomes increasingly plated and increases in mass with the progression of the redox reaction that eats up the anode. Generally, the function of the electrolytic solutions in a galvanic cell is to initiate the oxidation of the anode as a reducing agent (not the strongest though), and similarly, to initiate the reduction of the cathode as an oxidizing agent. In the case of a $Zn_{(s)}|Zn^{2+}{}_{(aq)}\|Cu^{2+}{}_{(aq)}|Cu_{(s)}$ cell, zinc metal is the strongest reducing agent that gets oxidized by zinc (II) ion. This releases two electrons per mole of zinc metal such that when copper (II), the strongest oxidizing agent, receives the electrons given by zinc, copper is plated on itself at the cathode. A final aspect of galvanic cells is that the material of the cathode does not affect the rate of redox reaction, or the voltage produced by the cell. While the cathode electrolyte

and anodic half-cell must be controlled, the cathode itself can be replaced by an inert element such as carbon or platinum, which would yield the same voltage.
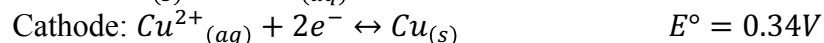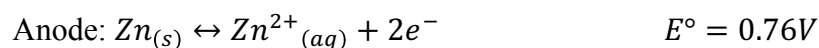
Figure 1. zinc-copper galvanic cell
http://www.periodni.com/gallery/galvanic_cell.png

$$Zn(s) \mid ZnSO_4(aq) \parallel CuSO_4(aq) \mid Cu(s)$$
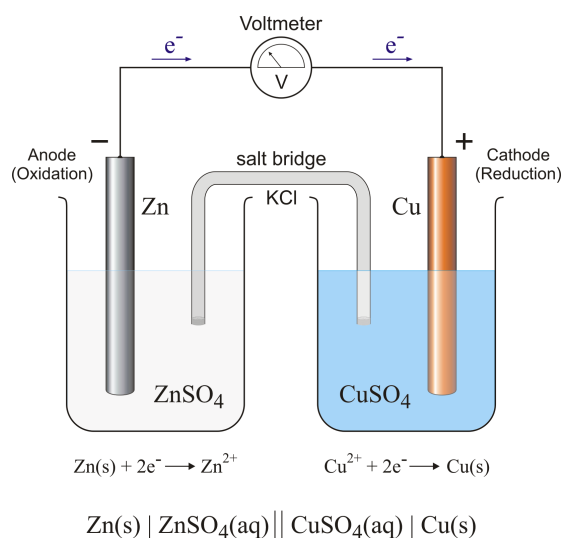
The voltage produced by a galvanic cell is calculated by adding the standard cell potentials (standard refers to the concentration of electrolytic solution at 1.0mol*L⁻ as well the temperature of the reaction at 298K) of each half reaction.

Calculating the theoretical Standard Cell Potential of a $Zn_{(s)}|Zn^{2+}{}_{(aq)}\|Cu^{2+}{}_{(aq)}|Cu_{(s)}$ cell,

Anode: $Zn_{(s)} \leftrightarrow Zn^{2+}{}_{(aq)} + 2e^-$ $E° = 0.76V$
Cathode: $Cu^{2+}{}_{(aq)} + 2e^- \leftrightarrow Cu_{(s)}$ $E° = 0.34V$

$E°_{Total} = E°_1 + E°_2 = 0.76V + 0.34V = 1.10V$

Note: Standard Cell Potentials are taken from *Chemistry IB Data Booklet 2016*

Figure 2: Galvanic Cell, and standard cell potential

| Galvanic Cell	Standard Cell Potential, $E°$, (V)
$Zn_{(s)}|Zn^{2+}{}_{(aq)}\|Cu^{2+}{}_{(aq)}|Cu_{(s)}$	1.10
$Mg_{(s)}|Mg^{2+}{}_{(aq)}\|Al^{3+}{}_{(aq)}|Al_{(s)}$	0.71
$Mg_{(s)}|Mg^{2+}{}_{(aq)}\|Zn^{2+}{}_{(aq)}|Zn_{(s)}$	1.61
$Mg_{(s)}|Mg^{2+}{}_{(aq)}\|Fe^{2+}{}_{(aq)}|Fe_{(s)}$	1.92
$Mg_{(s)}|Mg^{2+}{}_{(aq)}\|Cu^{2+}{}_{(aq)}|Cu_{(s)}$	2.71

Nernst Equation

Although Nernst Potential is not explicitly used in this experiment, the theory surrounding the Nernst Equation, which describes the proportional relationship between standard cell potential and the concentration of cathodic electrolyte as $E° \propto log \frac{\{Anode\ Electrolyte\}}{\{Cathode\ Electrolyte\}}$ (. Therefore, as the concentration of the anode electrolyte is increased, so will the voltage produced because there will be more frequent oxidations of the anode metal. As for the cathode electrolyte, if its concentration decreases, then the voltage of the reaction will also increase.

https://chem.libretexts.org/Core/Analytical_Chemistry/Electrochemistry/Nernst_Equation

Prediction

Considering that the electronegativity of an element is defined as the strength of its attractive force exerted on electrons either in a bond or by itself, then it is possible to infer that galvanic cells with a greater difference in electronegativity will also yield a greater voltage because an increase in ΔEN means an increase in the attractive force on electrons exerted by the cathode. As this force increases, the greater the rate at which electrons will flow towards the cathode, and the greater the electromotive force pushing the electrons will be. In essence, $E° \propto \Delta EN$ in a linear proportionality with a positive slope.

Figure 3:

Difference in electronegativity, ΔEN
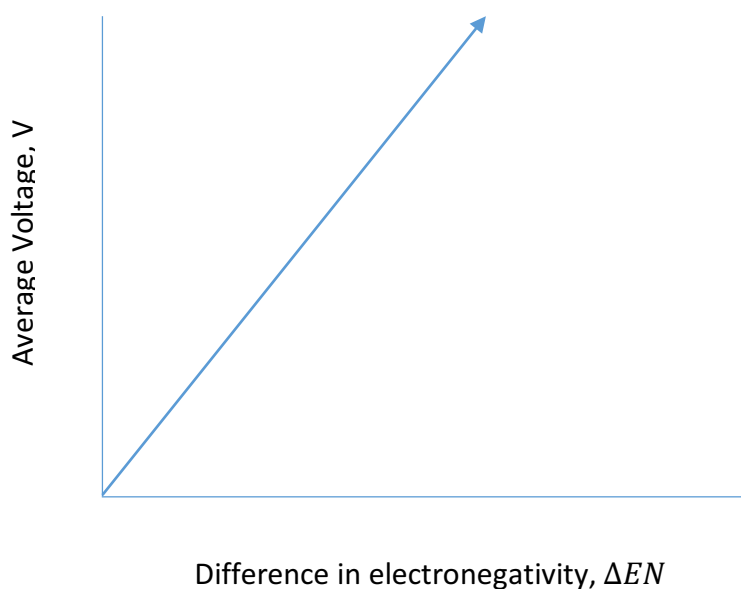
Note: The y-intercept will be (0,0) because if there is no difference in electronegativity between the two electrodes, meaning the electrodes will be made of the same material then the cell will produce no voltage because neither electrode will be a stronger reducing agent.

λ 3. Manipulated, Responding, Controlled Variables

Manipulated Variable: Difference in electronegativity of the electrodes. For electrodes (cathode, anode respectively) I will be using:

1. copper (II), zinc (II)
2. aluminum(III), magnesium (II)
3. zinc (II), magnesium (II)
4. iron (II), magnesium (II)
5. copper (II), magnesium (II)

RV: Electrical Potential $E°$ of galvanic cell, measured by voltmeter in Volts.
CV: volume and concentration of each electrolytic solution, cathodes selected always have greater electronegativity than their anodes

Note: Sodium and nitrate ions do not precipitate.

λ 4. Assumptions

1. Salt solution does not spill into either electrode solution and does not affect the redox reactions
2. Circuit does not lose voltage through heating and resistance.
3. All electrolytic solutions are standard molar concentration (1.0mol/L)
4. Mass of manipulated metals (which varies) does not affect its redox reaction or the Standard Cell Potential measured.
5. Variance in the mass and surface area of each metal will not affect Voltage readings
6. The way the metals are submerged in their respective electrolytic solutions (For example, if one metal is more submerged) will not affect voltage values.
7. Switching anodes from zinc to magnesium will not affect the observed trend as long as electronegativity difference is maintained.

λ 6. Materials List

Table 1: Equipment, Glassware and Chemicals

Equipment	Glassware	Chemicals
magnesium metal strip	6*100mL beakers	17.532 g $NaCl_{(s)}$
zinc metal strip	6*100mL volumetric flask (3)	24.652 g $MgSO_4*7H_2O_{(s)}$
iron metal strip	U-tube	34.222 g $AlSO_4*8H_2O_{(s)}$
copper metal strip	Waste Beaker (labeled)	28.759 g $ZnSO_4*7H_2O_{(s)}$
aluminum metal strip		24.972 g $CuSO_4*5H_2O_{(s)}$
Voltmeter		27.806 g $FeSO_4*7H_2O_{(s)}$
2*Cotton balls		600mL distilled water
2*Alligator clip wires		
Masking tape		

Sharpie		
Electronic Balance		
6*Weighing Boat		
6*Funnels and Corks		
Eyedropper		
Paper Towels		

λ 7.Procedure

Safety: The solutions used in this lab are all meant to have standard molarity ($1.0mol*L^{-1}$) and can be seriously harmful to the eyes, skin, and organs if swallowed or come into contact with. Solutions are also toxic and can be corrosive, so avoid direct contact by wearing gloves, aprons and goggles when handling chemicals. Inspect and discard any chipped or cracked beakers, no matter how small the damage. Sweep up broken glass with brush and dustpan; do not use fingers. Ensure that all metals or heavy materials are handled carefully such that they do not cut the skin, or cause substantial bodily harm (dropping electronic balance on the toes). Dispose of all solutions in a labeled waste beaker and do not wash them down the sink.

A: Preparing standard concentration solutions

1. Prepare six 100mL volumetric flasks and label them with the names of each of the five metal hydrate solutes, and a sixth one for sodium chloride.
2. Measure the masses of each solute needed to prepare a 1.0 $mol*L^{-1}$ solution using a weigh boat and an electronic balance (use the masses calculated in the background).
3. Measure the mass of sodium chloride required to prepare a 3.0 $mol*L^{-1}$ solution.
4. Transfer each metal hydrate solute into their respective flasks, using funnels to assure no material is lost during transfer.
5. Rinse the sides of the funnel down with distilled water to ensure that the flask captures all residual material.
6. Fill each flask with roughly 40mL of distilled water, then cap the flask and swirl it to dissolve the solutes.
7. Once each material has dissolved, fill each flask to slightly below the etched line with distilled water.
8. Pause and let liquid drain down from the sides of the flask.
9. Use an eyedropper to carefully raise the base of the meniscus to the etched line.
10. Finally, cap, mix, swirl the flasks to make sure that all solute is dissolved.
11. Record qualitative observations of the dissolving process.

B: Constructing galvanic cell circuits

1. Label each of the six 100mL beakers with the names of the 6 solutions prepared previously.
2. Fill each beaker with their respective chemicals by pouring directly from the volumetric flasks.
3. Create a salt bridge by pouring roughly 40mL of sodium chloride solution into a U-tube, then carefully placing two cotton balls in each opening to block the solution from leaving the tube. Make sure there are no air bubbles in the U-tube.

4. Set up first galvanic cell circuit $(Zn_{(s)}|Zn^{2+}_{(aq)}||Cu^{2+}_{(aq)}|Cu_{(s)})$ by connecting each alligator clip wire to a respective electrode, and placing each electrode in their corresponding solutions (zinc metal into ZnSO4*7H2O$_{(aq)}$).
5. Place the two beakers containing the electrolytic solutions close to each other, and bridge the U-tube between them, making sure that each end of the U-tube is submerged.
6. Connect a voltmeter to the ends of the alligator clip wires and read the voltage with regard to uncertainty. If the voltage is negative, switch the wires around.
7. Record voltage readings for this cell 5 times as well as qualitative observations of each metal's properties.
8. Repeat steps 4-7 for the remaining 4 galvanic cells, making sure to clip the wires to the metals the same way each time, and submerge the metals in the solution containing their oxidized ions.
9. Make sure to refill the salt-bridge with salt solution if air bubbles become substantial, and try to prevent solutions from spilling. Also wipe off each metal between different circuits.
10. Record the volumes of each solution used throughout the experiment.
11. Dispose of used solutions in waste beaker and throw away wet cotton balls. Rinse and keep solid metals and glassware for future use.

λ 8.OBSERVATIONS

Table 2: Mass of each solid salt dissolved in 100mL distilled water to create 1.0 mol*L^{-1} (standard concentration) electrolyte solutions, volume of each salt solution used, and qualitative observations.

Salt name	Salt mass, m_{salt}, $(\pm 0.01g)$	Volume of salt solution (electrolyte), $V_{electrolyte}$, $(\pm 0.08mL)$	Qualitative observations
aluminum (III) sulfate octahydrate	34.22	100.00	Salt is white, powdery, solid. Solution is translucent, colorless, liquid. Note: a mass of undissolved solute (roughly 0.5 g) is left in volumetric flask upon transfer of solution to a separate 100mL beaker.
magnesium(II) sulfate heptahydrate	24.65	>100.00	Salt is white, chunky, solid. Solution is clear, colorless, liquid. Note: when filling the volumetric flask with distilled water, the meniscus exceeded the 100mL mark by roughly 0.4 cm.
zinc (II) sulfate heptahydrate	28.76	100.00	Salt is white, powdery, solid. Solution is clear, colorless, liquid

Table 2 (Continued): Mass of each solid salt dissolved in 100mL distilled water to create 1.0 mol*L^{-1} (standard concentration) electrolyte solutions, volume of each salt solution used, and qualitative observations.

Salt name	Salt mass, m_{salt}, $(\pm 0.01g)$	Volume of salt solution (electrolyte), $V_{electrolyte}$, $(\pm 0.08mL)$	Qualitative observations
copper (II) sulfate pentahydrate	24.97	100.00	Salt is blue, opaque, grainy, solid. Solution is blue, translucent, liquid.
Iron (II) sulfate heptahydrate	27.80	100.00	Solid is white, powdery, solid. Solution is translucent, light green, liquid.
Sodium chloride	17.53	100.00	Solid is white, powdery, opaque, solid. Solution is transparent, colorless, liquid.

Table 3: Galvanic cell name (Standard Cell notation), measured voltage of each galvanic cell, and qualitative observations.

Galvanic Cell Name	Voltage, V, $(\pm 0.01V)$	Qualitative Observations
$Zn_{(s)}|Zn^{2+}_{(aq)}||Cu^{2+}_{(aq)}|Cu_{(s)}$	0.95	
	0.95	$Zn_{(s)} - Grey, shiny, solid$
	0.95	
	0.95	$Cu_{(s)} - bronze, shiny, solid$
	0.95	
$Mg_{(s)}|Mg^{2+}_{(aq)}||Al^{3+}_{(aq)}|Al_{(s)}$	0.69	
	0.69	$Al_{(s)} - grey, shiny, solid$
	0.69	
	0.69	$Mg_{(s)} - thin, grey, shiny, solid$
	0.69	
$Mg_{(s)}|Mg^{2+}_{(aq)}||Zn^{2+}_{(aq)}|Zn_{(s)}$	1.35	
	1.35	$Zn_{(s)} - Grey, shiny, solid$
	1.35	
	1.35	$Mg_{(s)} - thin, grey, shiny, solid$
	1.35	

Table 3: Galvanic cell name (Standard Cell notation), measured voltage of each galvanic cell, and qualitative observations.

Galvanic Cell Name	Voltage, V, $(\pm 0.01V)$	Qualitative Observations
$Mg_{(s)}|Mg^{2+}{}_{(aq)}||Fe^{2+}{}_{(aq)}|Fe_{(s)}$	1.88	
	1.88	$Fe_{(s)} - grey, shiny, solid$
	1.88	
	1.88	$Mg_{(s)} - thin, grey, shiny, solid$
	1.88	
$Mg_{(s)}|Mg^{2+}{}_{(aq)}||Cu^{2+}{}_{(aq)}|Cu_{(s)}$	2.45	
	2.45	$Cu_{(s)} - bronze, shiny, solid$
	2.45	
	2.45	$Mg_{(s)} - thin, grey, shiny, solid$
	2.45	

λ 9. Analysis

In order to graphically determine the relationship between the ΔEN of each pair of electrodes and their produced voltages, ΔEN (Pauling Scale) for each cell must first be calculated by subtracting the electronegativity of the cathode by the electronegativity of the anode to yield a positive difference (the difference is positive because the cathodes consistently have a greater electronegativity than the anodes used in this experiment).

Calculating ΔEN for $Zn_{(s)}|Zn^{2+}{}_{(aq)}||Cu^{2+}{}_{(aq)}|Cu_{(s)}$,

$$\Delta EN = EN_{anode} - EN_{cathode} = EN_{zinc} - EN_{copper} = 1.90 - 1.65 = 0.25$$

Note: Electronegativity values are taken from:
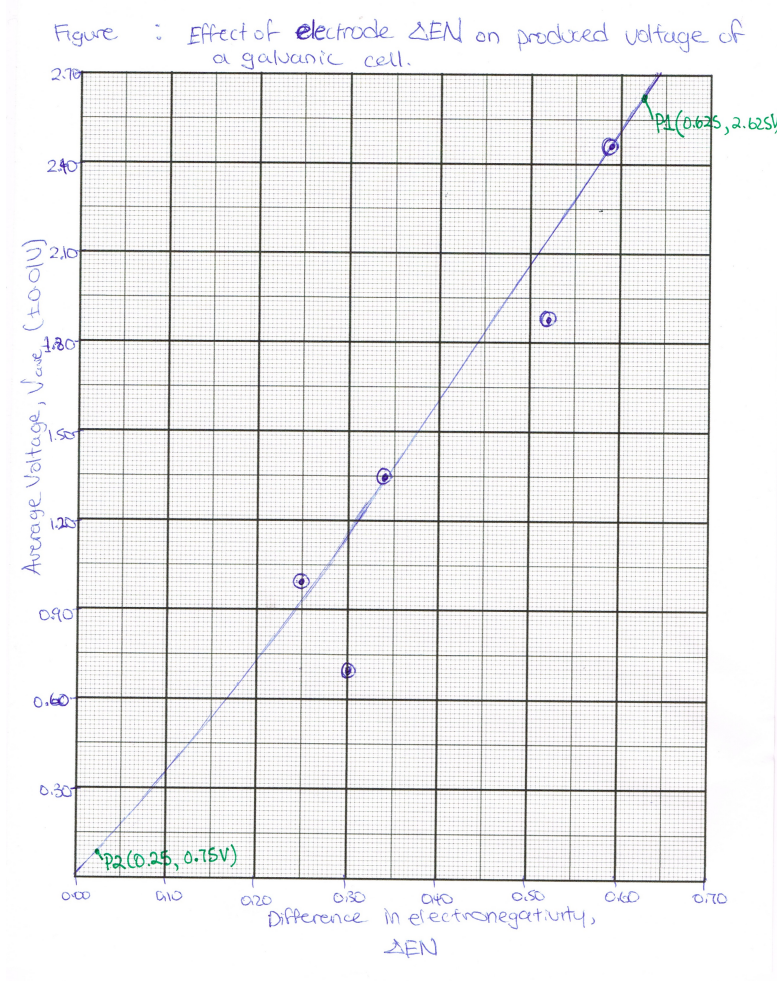https://www.webelements.com/electronegativity.html

After calculating ΔEN, the 5 replicated voltage values for each cell must be averaged in preparation for the graphing of 5 responding voltage values only.

Averaging voltage replicates for $Zn_{(s)}|Zn^{2+}{}_{(aq)}||Cu^{2+}{}_{(aq)}|Cu_{(s)}$,

$$V_{zn||cu} = \frac{(V_1 + V_2 + V_3 + V_4 + V_5)}{5}$$

$$= \frac{(0.95 \pm 0.01V + 0.95 \pm 0.01V + 0.95 \pm 0.01V + 0.95 \pm 0.01V + 0.95 \pm 0.01V)}{5}$$

$$= 0.95 \pm 0.01V$$

Table 4: Galvanic cell name (Standard Cell notation, difference in electronegativity between electrodes, and average voltage produced by each galvanic cell)

Galvanic Cell Name	Difference in electronegativity between electrodes, ΔEN	Average Voltage, V_{ave}, $(\pm 0.01V)$
$Zn_{(s)}|Zn^{2+}_{(aq)}||Cu^{2+}_{(aq)}|Cu_{(s)}$	0.25	0.95
$Mg_{(s)}|Mg^{2+}_{(aq)}||Al^{3+}_{(aq)}|Al_{(s)}$	0.30	0.69
$Mg_{(s)}|Mg^{2+}_{(aq)}||Zn^{2+}_{(aq)}|Zn_{(s)}$	0.34	1.35
$Mg_{(s)}|Mg^{2+}_{(aq)}||Fe^{2+}_{(aq)}|Fe_{(s)}$	0.52	1.88
$Mg_{(s)}|Mg^{2+}_{(aq)}||Cu^{2+}_{(aq)}|Cu_{(s)}$	0.59	2.45

Figure : Effect of electrode ΔEN on produced voltage of a galvanic cell.

With the relationship between electrode ΔEN and V_{ave} graphically represented, it is possible to define the relationship as a linear proportionality $V_{ave} \propto \Delta EN$ for a specific range of ΔEN as shown in the figure. Clearly, the line of best fit passes through the maximum number of data points possible, for ΔEN of 0.34 and 0.59. This choice of a linear line of best fit with a positive slope is justified by:

1. The requirement that the line of best fit must pass through (0,0) because if $\Delta EN = 0$, then the electrodes must be made of the same metal element, which would yield no voltage if connected in a galvanic cell.
2. The lack of experimental data for $\Delta EN < 0.25$, which classifies the data point at $\Delta EN = 0.25$ as an anomaly, and justifies straightening the line of best fit to pass through (0,0) instead of curving the line upward in a V-shape starting from $(0.30, 0.69V)$ and passing through $(0.25, 0.95V)$.
3. The switch from $Zn_{(s)}|Zn^{2+}{}_{(aq)}\|Cu^{2+}{}_{(aq)}|Cu_{(s)}$ cell to $Mg_{(s)}|Mg^{2+}{}_{(aq)}\|Al^{3+}{}_{(aq)}|Al_{(s)}$ cell causes the anode to change from zinc to magnesium, and since magnesium is used for the rest of the cells, the data point with zinc anode must be inconsistent with the trend using magnesium as a controlled anode. Hence point 1 rather than point 2 is considered anomalous.

To define this relationship algebraically, two points, P1 $(0.625, 2.625\pm0.01V)$ and P2 $(0.25, 0.75\pm0.01V)$ have been denoted on the graph in order to calculate the equation of the line.

Determining the slope of the equation,

Slope$=\frac{y_2-y_1}{x_2-x_1}=\frac{2.625\pm0.01V-0.75\pm0.01V}{0.625-0.25}=\frac{1.875\pm0.02V}{0.375}=5.00V\pm\left(\frac{0.02V}{1.875V}\right)=5.00V\pm1\%$

Determining the equation of the line in slope-intersect (y=mx+b) form, and considering $V_{ave} \propto \Delta EN$,

$$V_{ave} = (5.00V \pm 1\%) * \Delta EN$$

Graphically, the significance of the slope, as shown by this equation, is that for every unit increase in ΔEN, V_{ave} increases by 5.00 V. Put in another way, the greater the potential an atom of cathode has to attract electrons with a greater force in bonds between it and other atoms with lower electronegativity, the greater the kinetic energy of the electrons flowing towards the cathode will be. This makes sense because a cell produces electromotive force to move electrons in a circuit and it explains why cathodes tend to get reduced (gain electrons) rather than anodes which become oxidized because compared to anodes, cathodes have higher electronegativity values (Pauling Scale) and therefore attract electrons with greater force. In a circuit, the electrons will be more strongly attracted by the cathode than the anode. Hence, the anode always gets oxidized, sending electrons through a connecting wire to reduce the cathode. Of course, these theories are all based on the fundamental assumption that the method for choosing to straighten the line of best fit and making it pass through (0,0) is scientifically justified and accurate.

Finally, addressing the anomaly in the data points, particularly (0.25, 0.95V), it is evident that this point presents a contradiction to the main trend of the rest of the 4 data points, which is linear proportionality. If the anomaly is considered, then the resultant v-shaped graph, as mentioned before, would imply:

1. That in the sudden and abrupt switch from $Zn_{(s)}|Zn^{2+}{}_{(aq)}||Cu^{2+}{}_{(aq)}|Cu_{(s)}$ cell to $Mg_{(s)}|Mg^{2+}{}_{(aq)}||Al^{3+}{}_{(aq)}|Al_{(s)}$ cell, the change in electronegativity of the anode (from $\Delta EN = 1.65$ to $\Delta EN = 1.31$) had already disrupted the trend with $Mg_{(s)}$ as a controlled anode, which resulted in the decrease in voltage from $\Delta EN = 0.25$ to $\Delta EN = 0.30$.
2. The point (0.25, 0.95V) represents the beginning of a separate increasing trend using $Zn_{(s)}$ as the controlled anode. For example, if $Zn_{(s)}$ were to be used consistently without switching to $Mg_{(s)}$, then the resultant relationship would also be a linear proportionality, but with greater slope. This is shown in figure 4.
3. The reason for such an irregular trend may possibly be periodicity. Because magnesium and zinc are in periods 2 and 3 respectively, like atomic radii trends (figure 5), their respective trends for cell potential may exhibit a similar pattern. As the data moves up increasing periods, there is a substantial increase in atomic size of each atom, therefore, the nucleus of a period 3 atom such as zinc does not attract valence electrons as strongly as the nucleus of a period 2 atom such as magnesium. This discrepancy makes it easier for an opposing cathode to attract and pull electrons off of zinc rather than magnesium. Even with approximately equal ΔEN in cells $Zn_{(s)}|Zn^{2+}{}_{(aq)}||Cu^{2+}{}_{(aq)}|Cu_{(s)}$ and $Mg_{(s)}|Mg^{2+}{}_{(aq)}||Al^{3+}{}_{(aq)}|Al_{(s)}$, since it is easier for copper to pull electrons from zinc, there is correspondingly a greater voltage produced, creating the anomaly.

λ 10. Evaluation

Throughout this experiment, measurement uncertainty, and unforeseen procedural errors such as the failure to dissolve portions of solid hydrates due to limited utility created most of the major errors that affect the validity of the resulting data. Specific errors included:

1. One major error in this experiment was during the volumetric procedure conducted on magnesium(II) sulfate heptahydrate, when distilled water was used to fill the volumetric flask approximately 0.4 cm above the meniscus, resulting in excess water being used to create the concentration solution. This is a major error because despite the fact that an eyedropper was quickly used to remove the excess water, loss of solute was not guaranteed, and therefore the experimental concentration of magnesium(II) sulfate heptahydrate electrolytic solution for all trials involving magnesium as a cathode were lower than expected. Overall, it is a major, procedural and systematic error that, according to the Nernst Equation (refer to background), lowered the expected voltage of all cells containing magnesium solid. To fix this error, more attentive and careful procedural steps should be taken to ensure that no excess distilled water is added to the volumetric flask.
2. The second major error of this lab was caused by resistance in each galvanic circuit. The resistance of the connecting wires and voltmeter, the resistance of the electrolytic solutions, the resistance provided by the salt bridge, and finally the resistance from the oxidized surfaces of all the electrodes contributed wholly to the producing of lower voltages than theoretically predicted in the background. This was a major, instrumental, and systematic error that affected all voltage values by lowering them a few tenths of a Volt. To reduce this error, a porous cup design can be used so that wires or a salt bridge are not needed equipment to perform the experiment.

3. While attempting to create standard concentration solutions of electrolytic solutions, the solute aluminum (III) sulfate octahydrate failed to dissolve completely in the 100mL volumetric flask (it was stuck to the side of the glass), leaving behind approximately (estimated by eye) 0.5 grams of solute left in the flask after the solution was transferred to the 100mL beaker for creation of the half-cell. This resulted in a minor, procedural and random error which caused the experimental concentration of aluminum (III) sulfate octahydrate solution to be lower than theoretically predicted, and hence, according to the Nernst Equation, lowering the expected voltage of the $Mg_{(s)}|Mg^{2+}_{(aq)}||Al^{3+}_{(aq)}|Al_{(s)}$ cell. Justifiably, the lower the concentration of a cathode's electrolytic solution, the less frequently reduction will occur and the less voltage will be produced. To resolve this error, the solute being dissolved should be crushed into finer pieces so that no undissolved chunks gather and adhere to the side of the volumetric flask.

λ 11. Conclusion

In response to the research statement: Determining the effect of increasing the difference in electrode electronegativity on the voltage produced by a galvanic cell, it is evident, according to the relationship $V_{ave} = (5.00V \pm 1\%) * \Delta EN$, that $V_{cell} \propto \Delta EN$. This means that the produced voltage of a cell varies proportionately with the electronegativity difference between the anode and the cathode, which supports the hypothesis made at the beginning of this experiment. Overall, considering that both of the major errors were systematic, the trend of proportionality between voltage and ΔEN is a fairly reliable conclusion although individual data points may not have as accurate voltage values.

λ 12. Applications

A potential application for this experiment is that it adds another criterion in the pool of criteria chemical engineers use to test the power of galvanic cells. Now, it is possible to choose the electrode with the highest electronegativity as the cathode, and the electrode with the lowest electronegativity as the anode in order to create the greatest difference in electronegativity and therefore the greatest voltage. Although one cautionary advice is to make sure that the anode is always controlled as a variable so the upward increasing trend remains continuous with no anomalies.

λ 13. Applications