Chemistry IA 2

How does increasing the concentration of copper (II) nitrate solution in a voltaic cell affect the voltage produced?

Design

Introduction

Voltaic cells are electrochemical cells that use a reduction reaction to produce electrical energy.[1] They are used in many rechargeable electronic devices such as batteries, as well as in solar panels and electric cars. Batteries are used to create electrical energy, due to the potential differences between the two electrodes and electrolytes inside the battery. As batteries mimic voltaic cells this investigation shall aim to investigate whether the change in one electrolytes, copper (II) nitrate, has an effect on the potential difference. The cell consists of a cathode and an anode which are submerged into solutions which contain cations and anions of the respective electrode metals. In this investigation the voltaic cell consist of a potassium salt bridge (KNO_3), an zinc anode $Zn(NO_3)_2$, and a copper cathode $Cu(NO_3)_2$. In this experiment, copper will undergo reduction as it is the less reactive of the two metals, meaning that the copper ions in the copper nitrate solution will start forming on the copper metal electrode. The reaction that occurs during a voltaic cell is as follows:
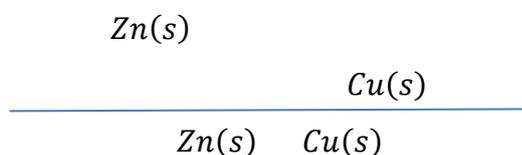
The half reaction that occurs in the anode is the oxidation of Zn ions to Zn^{2+}:

$$Zn(s)$$

The half reaction that occurs in the cathode is the reduction of Cu^{2+} ions to Cu:

$$Cu(s)$$

Therefore the overall reaction is as follows:

$$Zn(s)$$
$$Cu(s)$$
$$\overline{}$$
$$Zn(s) \quad Cu(s)$$

This experiment aims to determine how a change in concentration of copper (II) nitrate solutions will affect the voltage produced in the cell. This will be done by adding in the mass

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[1] "Voltaic Cells." *Voltaic Cells*. N.p., n.d. Web. 19 Mar. 2014.
http://www.chem.wisc.edu/deptfiles/genchem/netorial/rottosen/tutorial/modules/electrochemistry/03voltaic_cells/18_31.htm

of 0.1M copper (II) nitrate. This experiment is significant as it allows us to deduce what concentration copper solution will yield the highest voltage, which is an experiment which could have industrial uses.

Background information

Voltage or potential difference is the electrical force that drives an electric current between two points.[2] Voltage is measured through the cell's potential difference; which can be calculated using the equation below:

$$\mathbf{E^o_{cell} = E^o_{reduction} + E^o_{oxidation}}$$

Usually when we measure the voltage it is under a standard state, whereby the concentration 1 mole per liter (1 M) and has a pressure of 1 atmosphere at 25°C.[3] If we were to calculate the voltage of a zinc and copper nitrate reaction under standard cell potential we would receive, 1.101V.[4] This is shown below.

Oxidation: $Zn(s)$

$(- \quad V)$

Reduction: $Cu(s)$

Overall: $Zn(s) \quad Cu(s)$

[2]"Current vs Voltage." *Diffen*. N.p., n.d. Web. 19 Mar. 2014.
http://www.diffen.com/difference/Current_vs_Voltage
[3] "Electrochemical_Cell_Potentials." *Electrochemical_Cell_Potentials*. N.p., n.d. Web. 19 Mar. 2014.
http://www.chem.purdue.edu/gchelp/howtosolveit/electrochem/electrochemical_cell_potentials.htm
[4] "Electrochemical_Cell_Potentials." *Electrochemical_Cell_Potentials*. N.p., n.d. Web. 19 Mar. 2014.
http://www.chem.purdue.edu/gchelp/howtosolveit/electrochem/electrochemical_cell_potentials.htm

However as this investigation analyses the the impact of an increase in concentration we need to use the non-standard state cell potential, whereby the concentration of copper is increasing. The non- standard state cell potential is when the conditions are other than standard state (concentrations not 1 molar and/or pressures not 1 atmosphere).[5] Usually in this situation to calculate the voltage produced under non-standard conditions, the following equation would be used:

$$(RT \quad nF)$$

Whereby:

Ecell = cell potential at non-standard state conditions

E°cell = standard state cell potential

R = constant (8.31 J/mole K)

T = absolute temperature (Kelvin scale)

F = Faraday's constant (96,485 C/mole e-)

n = number of moles of electrons transferred in the balanced equation for the reaction occurring in the cell

Q = reaction quotient for the reaction. aA + bB --> cC + dD,

However by using a water bath I was able to keep the temperature at a constant of 25°C, in which case the equation simplifies to:

$$(0 \qquad n)$$

Furthermore the reaction quotient Q is now represented as:

$$\overline{\qquad}$$

Therefore in relation to this investigation's voltaic cell, Zn^{2+} the anode is divided by Cu^{2+} the cathode:

$$\overline{\qquad}$$

[5] "Electrochemical_Cell_Potentials." *Electrochemical_Cell_Potentials*. N.p., n.d. Web. 19 Mar. 2014. http://www.chem.purdue.edu/gchelp/howtosolveit/electrochem/electrochemical_cell_potentials.htm

So when we plug these values in we see that the value of Q decreases, due to an increase in cathode concentration. Therefore there is less potential cell difference, meaning that there is less subtracted from the Ecell, thereby increasing the voltage produced.

Risk Assessment

- As a variety of chemicals will be used, a lab coat, goggles and gloves must be worn in all times in order to protect clothing, eyes and skin.
- Furthermore chemicals should not be inhaled, in order to prevent any harmful chemicals from entering the body. To prevent this the chemicals should be kept arm's length away from the face at all times.
- When using the water bath it is very important to keep the voltmeter and wires away from the distilled water inside the tub.

Research Question

How does a change in the concentration of zinc and iron solution affect the voltage produced in a voltaic cell?

Variables

Variable	Variable Description	Method of Controlling Variable
Independent	Concentration of copper (II) nitrate solution used in each trial of the experiment.	The initial 0.1mol/dm^3 (M) of copper (II) nitrate solution will be made by adding (3.751±0.001g) to $200cm^3$ of distilled water into a 500 cm beaker. Then after each trial an extra 0.1M of copper (II) nitrate crystals will be added to the solution, in order to increase its concentration. The copper (II) nitrate crystals will be measured out by using a 3 decimal scale, and then the $200cm^3$ of distilled water will be measured out using a $250cm$ ±$5cm^3$ graduated cylinder. This solution will then be poured into a $500cm^3$ beaker.

Dependent	The voltage being produced by the two electrodes which is measured by the voltmeter	The cell potential (voltage) produced by the two voltaic cell will be measured using a voltmeter.
Controlled	Concentration of the zinc nitrate solution.	The concentration of zinc nitrate was kept at the same concentration of 0.1M. The solution was made by measuring out 3. g on a 3 decimal scale, and then placing them in a $500cm$ beaker. Then the $200cm$ of distilled water was measured out using a $250cm$ $\pm 5cm^3$ graduated cylinder. This was then poured into the $500cm$ beaker and dissolved.
Controlled	Concentration of potassium nitrate in the salt bridge	A concentration of 0.1M potassium nitrate was made to be used for the paper towel salt bridge. This was done by measuring out 2. g of potassium crystals onto a 3 decimal scale, and adding 200 cm^3 of distilled water which were measured out using a $250cm$ $\pm 5cm^3$ graduated cylinder. This solution was then dissolved in a $500cm^3$ beaker. Three paper towels were then twisted together and soaked in the potassium nitrate solution. It was then placed as the salt bridge across the two beakers which constituted the voltaic cell.
Controlled	Material of cathode and anode	Copper was used for the cathode and zinc was used for the anode. The same pieces were reused for each trial.
Controlled	A constant temperature of 25℃ was kept during the reaction	By placing all three beakers inside the water bath during the reaction, this ensured that the temperate was held constant at 25℃.
Controlled	Voltmeter used	The same voltmeter was used for all 5 trials, ensuring that the equipment used to collect

		the data was held constant .
Controlled	Volume of distilled water	The same $250cm$ $\pm 5cm^3$ graduated cylinder, was used to measure out all of the distilled water.

Materials

- Copper metal

- Zinc metal

- Distilled water (   )

- Copper $Cu(NO_3)_2$ powder (g)

- Zinc $Zn(NO_3)_2$ powder (g)

- KNO_3 (g)

- 3 x 500   beakers

- 250 $\pm 5cm$ graduated cylinder

- Voltmeter

- Wires

- Alligator clips

- 3 plastic weighing boats

- 3 paper towels

- 3 stirring rod

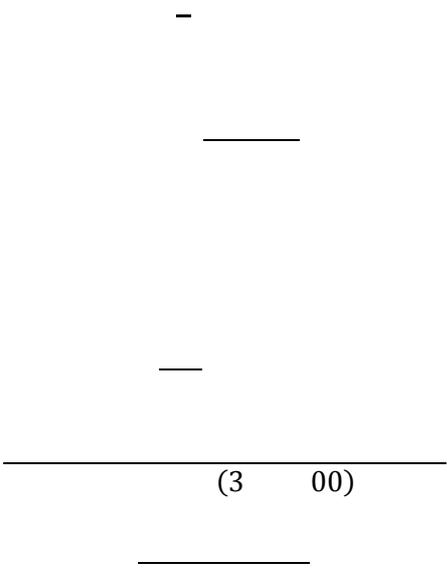- 3 decimal scale (accuracy of $\pm 0.001g$)

- 3 scoopula

Calculations

KNO$_3$ Solution:

The chosen concentration for the KNO_3 solution in the salt bridge was 0.1 mol/dm^3(M) , as this is the concentration of the $Zn(NO_3)_2$ solution and starting concentration of the $Cu(NO_3)_2$ solution.

In order to find the mass of KNO_3 crystals needed to make $200cm$ of 0.1M solution:

Where:
C = The concentration of the
solution (mol/dm^3 or M)
V = The volume of the solution
measured (dm^3)
n = The molarity of the solution
(mol)

Where:
n= The molarity of the solution
(mol)
m = The mass of the compound (g)
Mr = The molar mass of the
compound (g/mol)

(3   00)

___

Procedure:

1. Place a plastic weigh boat on the 3 decimal scales, and then tare the scale.

2. Then using a scoopula, measure out   of KNO$_3$. Add this into the
   beaker.

3. Using the 250 $\pm 5cm$ graduated cylinder measure out $200cm$ of distilled water.

4. Dissolve the contents of the beaker using the stirring rod.

Uncertainties:

The mass of the KNO$_3$ crystals were measured using a three point decimal scale, therefore the
uncertainty in the measurements is $\pm 0.001g$. The $200cm^3$ of water was measured out in a
$250cm$ graduated cylinder, which has holds an uncertainty of $\pm 5cm^3$.

 Furthermore the percentage uncertainty values are as follows:

___

—— x $100 = \pm 0.049\%$

—— x $100 = \pm 0.250\%$

<u>Zn(NO$_3$)$_2$ Solution:</u>

0.1M was chosen as the starting concentration for the Zn(NO$_3$)$_2$ solution, the same starting concentration as the copper (II) nitrate solution. In order to calculate the mass required to make a $200cm^3$ of 0.1M solution of Zn(NO$_3$)$_2$, the number of moles of the compound required must first be calculated. To calculate the number of moles required, the following equation was used:

Where:
C = The concentration of the solution (mol/dm^3 or M)
V = The volume of the solution measured (dm^3)
n = The molarity of the solution (mol)

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$\overline{}\quad\overline{}$

$\overline{}$

$(2(14\qquad(3\qquad00))$

$\overline{}$

Where:
n= The molarity of the solution (mol)
m = The mass of the compound (g)
Mr = The molar mass of the compound (g/mol)

<u>*Procedure:*</u>

1. Place a plastic weigh boat on the 3 decimal scales, and then tare the scale.
2. Then using a scoopula, measure out   of Zn(NO$_3$)$_2$. Add this into the beaker.
3. Using the 250 $\pm5cm$ graduated cylinder measure out $200cm$ of distilled water.
4. Briefly dissolve the contents of the beaker using the stirring rod.

The mass of the $Zn(NO_3)_2$ crystals were measured using a three point decimal scale, therefore the uncertainty in the measurements is $\pm0.001g$. The $200cm^3$ of water was measured out in a $250cm^3$ graduated cylinder, which has holds an uncertainty of $\pm5cm^3$.

Furthermore the percentage uncertainty values are as follows:

$$\overline{}$$

$$\overline{} \text{ x } 100 = \pm\,0.026\%$$

$$\overline{} \text{ x } 100 = \pm\,0.250\%$$

$Cu(NO_3)_2$ Solution:

To produce the mass of copper (II) nitrate required in order to make a $200cm$ of 0.1M solution, the following calculations need to be performed.

Where:
C = The concentration of the solution (mol/dm^3 or M)
V = The volume of the solution measured (dm^3)
n = The molarity of the solution (*mol/M*)

$$-$$

$$\overline{} \quad \overline{}$$

Where:
N= The molarity of the solution (*mol*)
m = The mass of the compound (*g*)
Mr = The molar mass of the compound (g/mol)

$$\overline{}$$

$$\overline{}$$
$$(63 \qquad (2(14 \qquad (3 \quad 16))$$

$$\overline{}$$

1. Place a plastic weigh boat on the 3 decimal scales, and then tare the scale.

2. Then using a scoopula, measure out   of $Cu(NO_3)_2$. Add this into the
   beaker.

3. Using the 250 $\pm 5cm$ graduated cylinder measure out $200cm$ of distilled water.

4. Briefly dissolve the contents of the beaker using the stirring rod.

Uncertainties:

Once again, the same uncertainty ratio's need to be applied as the same equipment was used.
So the mass of the $Cu(NO_3)_2$ crystals were measured using a three point decimal scale,
therefore the uncertainty in the measurements is $\pm 0.001g$. The $200cm^3$ of water was
measured out in a $250cm^3$ graduated cylinder, which has holds an uncertainty of $\pm 5cm^3$.

$$\overline{}$$

$$\overline{} \text{ x } 100$$

$$\overline{} \text{ x } 100 =$$

This uncertainty will, however, not have a significant impact on the outcome of this
experiment, as it will affect every trial equally.

Method

1. Using the above calculations prepare the solution of KNO_3:
   a. Begin by measuring out g of KNO_3 crystals by using a 3
      decimal scale and placing it in a 500   beaker.
   b. Then measure out 200   of distilled water in a $250cm$ graduated
      cylinder and add to the beaker. Stir until the KNO_3 has dissolved.

2. Using the above calculations prepare the $Zn(NO_3)_2$ solution:
   a. The $Zn(NO_3)_2$ is a powder, which must be measured out using a weighing boat and a 3 decimal scale. Measure out of $3.788g \pm 0.001g$ of $Zn(NO_3)_2$.
   b. Then measure out 200   of distilled water into a $250cm$ graduated cylinder, and then place into a 500   beaker and dissolve the   of $Zn(NO_3)_2$.

3. Using the above calculations prepare the solution of $Cu(NO_3)_2$:
   a. Begin by measuring out g of $Cu(NO_3)_2$ crystals by using a 3 decimal scale and placing it in a 500   beaker.
   b. Then measure out 200   of distilled water into a $250cm$ graduated cylinder, and then place into a $500cm$ beaker and dissolve the   of $Zn(NO_3)_2$.

4. Then place the zinc electrode into the 0.1M zinc nitrate solution, and the copper electrode into the 0.1M copper nitrate solution.

5. Set up the water bath:
   a. Fill the water bath with distilled water and then set the temperature to 25   (allow 5 minutes for warming up).
   b. Then place both beakers (half-cells) inside the water bath.

6. Create the salt bridge by twisting the three paper towels together and submerge them inside the KNO_3 solution for 15 seconds.
   a. Then remove the paper towels from the solution, and place the salt bridge across both cells, making sure that each end was submersed in each of the two half-cells.

7. Set up the voltmeter by connecting the necessary wires with alligator clips to the zinc anode and copper cathode.
   a. Make sure to not let the wires and voltmeter touch the water.

8. Then record the voltage produced by 0.1M of copper (II) nitrate.

9. Then add in another _____ of Cu(NO$_3$)$_2$ crystals in order to increase the concentration from 0.1M to 0.2M. Once again record the voltage produced.

10. Repeat this procedure until a concentration of 0.5M copper (II) nitrate solution has been achieved.

11. After this remove the contents of the Cu(NO$_3$)$_2$ beaker, and recreate a 0.1M solution.

12. Repeat steps 8-11 for another 4 trails, continually increasing the concentration by 0.1M and recording the results.

13. After all necessary data was collected; all equipment was thoroughly cleaned and stored.

Illustration
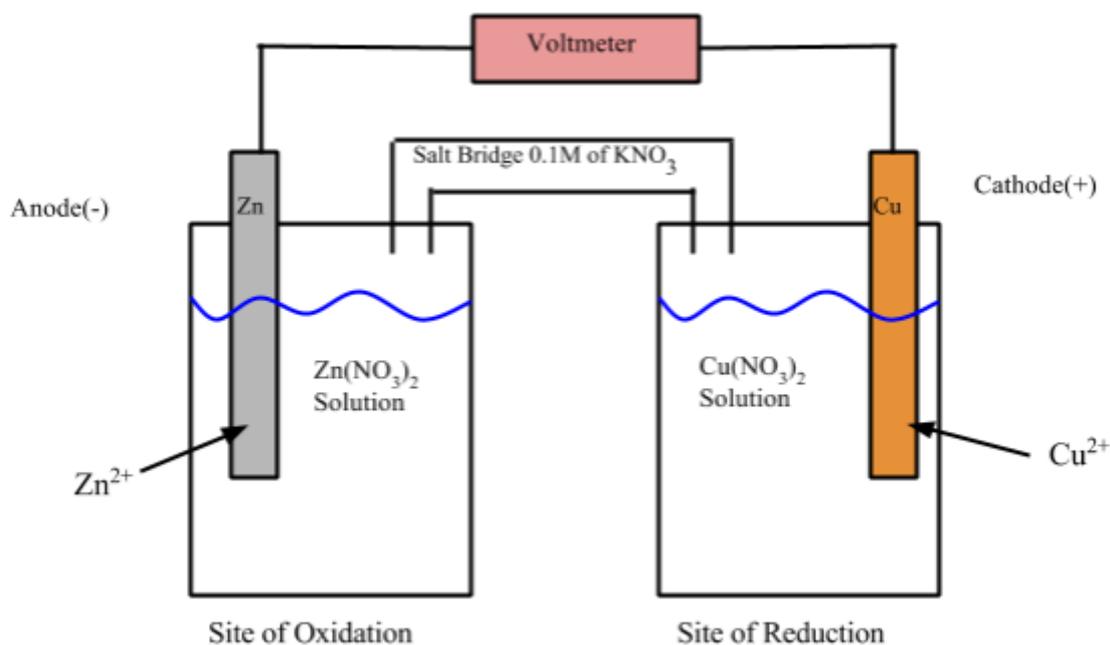
Data Collection and Presentation

Raw Data

Table 1: Cell potential/ voltage potential (V) produced from 0.1 mol/dm3 concentration of copper (II) nitrate.

Trial	Concentration of KNO_3 (mol/dm^3) ±0.299%	Concentration of $Zu(NO_3)_2$ (mol/dm^3) ±0.276%	Concentration of $Cu(NO_3)_2$ (mol/dm^3) ±0.276%	Voltage produced ±0.001 V
1	0.1	0.1	0.1	0.665
2	0.1	0.1	0.1	0.663
3	0.1	0.1	0.1	0.664
4	0.1	0.1	0.1	0.667
5	0.1	0.1	0.1	0.666

In order to calculate the average voltage produced by 0.1 M concentration of copper (II) nitrate, the following calculation must be used:

$$\bar{x}(me\ n) \quad \underline{\hspace{4cm}}$$

$$\bar{x} \quad \underline{\hspace{5cm}}$$

$$\bar{x}$$

Uncertainty Value:

$$\underline{\hspace{3cm}}$$

$$Cu(NO\ ) \quad \underline{\hspace{2cm}}$$

$$\underline{\hspace{2cm}}$$

$$Cu(NO\ ) \quad \underline{\hspace{1.5cm}}$$

Table 2: Cell potential/ voltage potential (V) produced from 0.2 mol/dm3 concentration of copper (II) nitrate.

Trial	Concentration of KNO_3 (mol/dm^3) $\pm0.299\%$	Concentration of $Zu(NO_3)_2$ (mol/dm^3) $\pm0.276\%$	Concentration of $Cu(NO_3)_2$ (mol/dm^3) $\pm0.276\%$	Voltage produced ±0.001 V
1	0.1	0.1	0.2	0.699
2	0.1	0.1	0.2	0.695
3	0.1	0.1	0.2	0.689
4	0.1	0.1	0.2	0.690
5	0.1	0.1	0.2	0.666

Average voltage produced by 0.2M of copper (II) nitrate:

$$\bar{x} \quad \overline{}$$

$$\bar{x}$$

<u>Uncertainty Value:</u>

$$Cu(NO\ ) \quad \overline{}$$

$$Cu(NO\ ) \quad \overline{}$$

Table 3: Cell potential/ voltage potential (V) produced from 0.3 mol/dm3 concentration of copper (II) nitrate.

Trial	Concentration of KNO_3 (mol/dm^3) $\pm0.299\%$	Concentration of $Zu(NO_3)_2$ (mol/dm^3) $\pm0.276\%$	Concentration of $Cu(NO_3)_2$ (mol/dm^3) $\pm0.276\%$	Voltage produced ±0.001 V
1	0.1	0.1	0.3	0.701
2	0.1	0.1	0.3	0.705
3	0.1	0.1	0.3	0.704
4	0.1	0.1	0.3	0.708
5	0.1	0.1	0.3	0.709

Average voltage produced by 0.3 M of copper (II) nitrate:

$$\bar{x} \quad \underline{\hspace{8cm}}$$

$$\bar{x}$$

Uncertainty Value:

$$Cu(NO\ ) \quad \underline{\hspace{4cm}}$$

$$Cu(NO\ ) \quad \underline{\hspace{2cm}}$$

Table 4: Cell potential/ voltage potential (V) produced from 0.4 mol/dm3 concentration of copper (II) nitrate.

Trial	Concentration of KNO₃ (mol/dm³) ±0.299%	Concentration of Zu(NO₃)₂ (mol/dm³) ±0.276%	Concentration of Cu(NO₃)₂ (mol/dm³) ±0.276%	Voltage produced ±0.001 V
1	0.1	0.1	0.4	0.714
2	0.1	0.1	0.4	0.718
3	0.1	0.1	0.4	0.719
4	0.1	0.1	0.4	0.720
5	0.1	0.1	0.4	0.718

Average voltage produced by 0.4 M of copper (II) nitrate:

$$\bar{x} \quad \underline{\hspace{8cm}}$$

$$\bar{x} \quad 718V$$

Uncertainty Value:

$$Cu(NO\ ) \quad \underline{\hspace{4cm}}$$

$$Cu(NO\ ) \quad \underline{\hspace{2cm}}$$

Table 5: Cell potential/ voltage potential (V) produced from 0.5 mol/dm3 concentration of copper (II) nitrate.

Trial	Concentration of KNO_3 (mol/dm^3) ±0.299%	Concentration of $Zu(NO_3)_2$ (mol/dm^3) ±0.276%	Concentration of $Cu(NO_3)_2$ (mol/dm^3) ±0.276%	Voltage produced ±0.001 V
1	0.1	0.1	0.5	0.737
2	0.1	0.1	0.5	0.729
3	0.1	0.1	0.5	0.723
4	0.1	0.1	0.5	0.721
5	0.1	0.1	0.5	0.739

Average voltage produced by 0.5 M of copper (II) nitrate:

$$\bar{x} \quad \frac{}{}$$

$$\bar{x}$$

Uncertainty Value:

$$Cu(NO\ ) \quad \frac{}{}$$

$$Cu(NO\ ) \quad \frac{}{}$$

Processed Data

Table 6: Average voltage produced by increasing concentration of copper (II) nitrate in a voltaic cell.

Concentration of KNO_3 (mol/dm^3) ±0.299%	Concentration of $Zu(NO_3)_2$ (mol/dm^3) ±0.276%	Concentration of $Cu(NO_3)_2$ (mol/dm^3) ±0.276%	Voltage produced ±0.001 V
0.1	0.1	0.1	0.665
0.1	0.1	0.2	0.6878
0.1	0.1	0.3	0.7054
0.1	0.1	0.4	0.7178
0.1	0.1	0.5	0.7298

Graph 1: Effect on increasing concentration of copper (II) nitrate on cell potential produced by voltaic cell
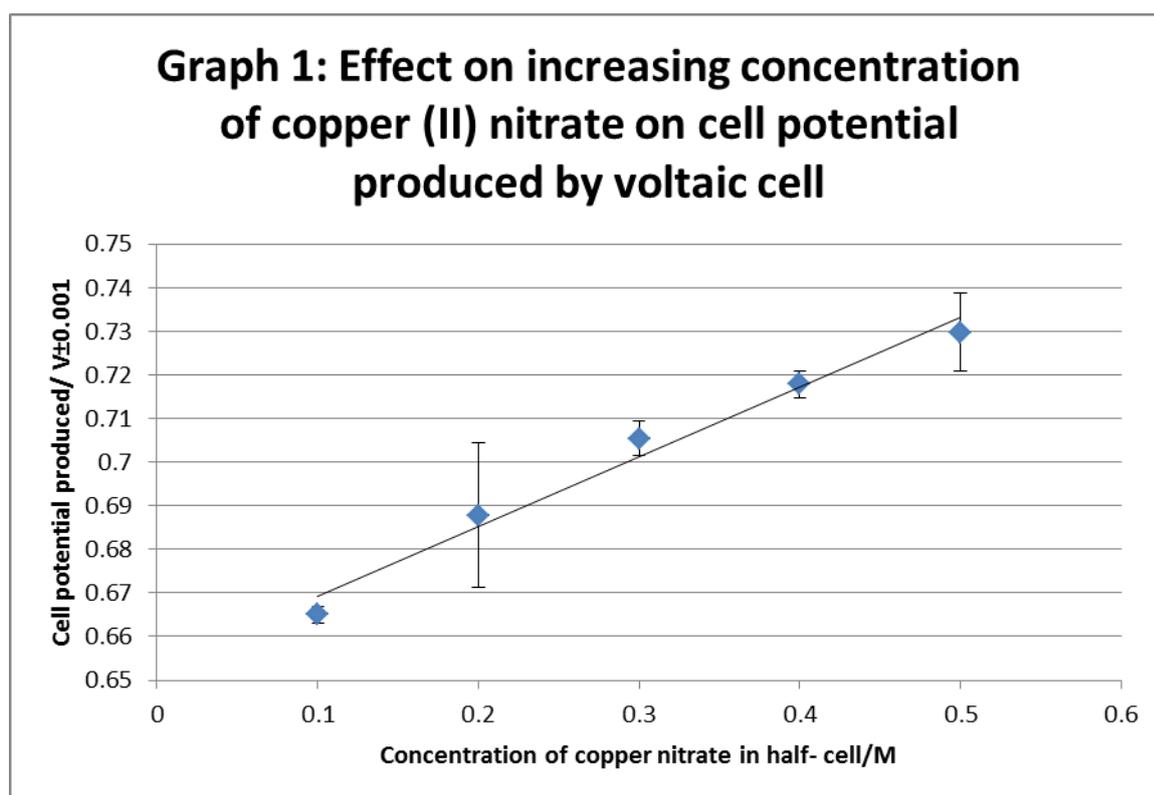
Conclusion and Evaluation

The graph above shows the clear relationship between the concentrations of the copper (II) nitrate solution vs. the voltage produced. There is a clear positive and proportional relationship between the two variables. The first concentration of copper (II) nitrate (0.1M) resulted in cell potential of *0.665± 0.001 V,* while the second concentration of copper (II) nitrate (0.2M) resulted in a cell potential of *0.688 ± 0.001 V,* and so on. Thus it is apparent that as the concentration increases, the cell potential does as well. This is justified due to the Nerst equation, which was explained before in the background. Whereby decreasing the value of Q, due to an increase in cathode concentration, results in is less potential cell difference, thereby increasing the voltage produced.

Another interesting relationship to comment on is how the results related to the theoretical standard state cell potential. Whereby the larger the concentration (0.5 > 0.1) and the closer the concentration gets to 1.0M, the closer the voltage gets towards 1.101 V. Even though these results did follow the theory of an increasing voltage, due to an increase in the concentration of a copper cathode, there are still many errors which may have altered the results. These possibilities are discussed below:

Evaluation of Weaknesses, Limitations and Errors:

Type of Error	**Error**	**Solution**
Systematic	*Potential inaccuracy of voltmeter* Even though the same voltmeter was used in order to measure the amount of volts being produced by the voltaic cell, often there was never one significant value. The readings would often alternate and vary, in other words there was no stable reading. This error is somewhat significant as there was not a set time for measuring the reading.	Firstly by setting a specific time for when the results should be measured would be a big improvement. Secondly by either cross-referencing the weight with another voltmeter (testing the same concentration but with another voltmeter) would be a simple way to check the result. Or finally use a more accurate voltmeter in order to reduce the variance.

Systematic	*Misreading of equipment* As readings relied on the human eye this may have easily happened when using the graduated cylinder, to measure out the distilled water. This is an issue as it may have slightly diluted the solution, which would have slightly altered the concentration.	Even though all the equipment was measured on a flat, even surface, another way to ensure accuracy would be reading the graduated cylinders from the same angle.
Systematic	*Re-use of weighing boats* Whilst I assigned separate weighing boats towards individual chemicals, I did reuse the same boat for the measurement of copper (II) nitrate. This is an issue as there may have excess residue from the last measurement, and seeing as each time I tared the weighing boat before measurement this would have gone unnoticed- and would have not been calculated into that trials measurement.	In future I would simply measure out each sample on an unused weighing boat.
Systematic	*Lack of timing* Instead of using a timer to keep track of how long the voltaic cell had been conducting, I simply waited until a relatively steady volt reading was reached. Although most trials were measured around 3 minutes	This proved as a significant error, as the results are to an extent unmonitored. By allowing certain trial a longer amount of time to produce energy, this could have allowed a larger amount of zinc deposit to build up.
Random	*Re-use of the same salt-bridge* It is possible that the salt bridge dried up over time as many trials were carried out. This could have affected the voltage as it would have affected its functioning as a salt bridge.	I would create a new salt bridge every trial with the same amount of paper towels and reused the same solution from the original. However, this would create the problem of the salt bridge sometimes having more or less solution after being wringed after each trial.
Random	*Excess produce from the salt bridge* The most important part of the salt bridge is that each end of	If available I would use a glass salt bridge. If this was not available I would try to ensure that there was no

	the paper towel is in each separate solution. However if the salt bridge had excess solution on it from not being properly rinsed – this could have contaminated the two solutions as the KNO3 solution dripped into the two beakers. This would have affected the voltage produced as it would have slightly affected the molarity in each solution.	excess solution on the salt bridge by hanging it in the air for 10 seconds. This would ensure that there was no excess solution available to contaminate the two other solutions.
Random	*Surface area of the metals* In this experiment I did not hand cut the metals to the exact same size, I simply used two pieces of metal available. With this is mind the pieces of metal seemed to be the same size. This is an issue as it could have drastically altered the amount of volume produced.	In future I would measure and hand-cut out two identical pieces of zinc and copper metal. In doing so this would there would be improved accuracy and reliability from the data as both electrodes would have almost identical surface area, reducing any possible change in voltages produced by that.
Random	*Re-use of zinc solution* Since the concentration of copper was increasing within the same half- cell, resulting oxides on the zinc electrode may still be present. This is error is somewhat significant.	In future I would for each new trail I would create a new zinc half-cell. In doing so this would ensure that the measurements were an accurate representation of the increasing concentration of copper (II) nitrate. Furthermore after each individual reading I would wipe down the electrodes with a paper towel.
Random	*Re-use of copper beaker* When beginning a new trial I simply drained the previous solution and gave it a quick rinse. In doing so this allowed for either copper (II) nitrate residue to be present in the beaker or even water. Both of which had the potential to alter the new concentration of copper.	In future I would either thoroughly cleanse the beaker by rinsing, cleaning and drying the beaker before reuse. Or simply use a new beaker for each trial.

Evaluation

In future in order to ensure more accurate and reliable results, all of the above improvements would have been made. However after reflecting upon the weaknesses, limitations and errors of this experiment, it becomes evident which errors held the most significance. These errors were in relation to the salt bridge. As described beforehand the salt bridge was a key weakness as it was continually reused throughout the entire experiment. Furthermore it also had the potential to contaminate both electrodes. The reused salt-bridge may have dried up, therefore constricting the quantity of electron transfer between the two solutions. Meaning that the earlier trials would have held more accurate results, as the later would have experienced the impact of the reduced quantity of ions being transferred. Secondly the contamination of the cells would have led to slightly different concentrations than originally recorded. However if this were the case, all trials would have been equally affected, therefore the error is less significant.

Finally this investigation was successful as it not only was successfully conducted, but it also followed the chemical theory. This was not only seen by using the Nerst equation, but also by Le Chateliers Principle[6]. As the concentration of the copper (II) nitrate cathode electrolyte was increased, this resulted in more electrons being produced, thus resulting in a higher cell potential.

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[6] "IB Chemistry Review." *: Topic 7: Equilibrium*. N.p., n.d. Web. 19 Mar. 2014. http://ibchemistry-review.blogspot.ch/p/topic-7-equilibrium.html

<u>Works Cited:</u>

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