JOURNAL OF
EMERGING INVESTIGATORS

Article

Analysis of reduction potentials to determine the most efficient metals for electrochemical cell alternatives

Peyton Carroll[1] and Jason Carroll[1]

[1] Fremont High School, Sunnyvale, California

SUMMARY

The aim of this work is to determine what metals make the most efficient electrochemical cells. Specifically, we measured the reduction potentials of several metals and then calculated the theoretical voltage that could be conducted by cells constructed with these metals. We tested our theoretical results by constructing an electrochemical cell and then measuring the actual voltage conducted by the cell. We hypothesized that an electrochemical cell constructed of metals with the greatest oxidation states would generate the greatest voltage. Based on measured reduction potentials and our subsequent calculations, we determined that an electrochemical cell made of magnesium and iron would be the most efficient cell with a theoretical voltage of 3.17 volts. Upon construction of an electrochemical cell made of magnesium and iron, we measured actual voltages of 3.19 V and 3.24 V, respectively. Determining the types of metals that will make the most efficient electrochemical cells is important; if we can increase the amount of energy that batteries can hold, then we can increase their lifetime and decrease their size.

INTRODUCTION

Constructing more efficient and durable batteries will decrease the number of times people need to change their batteries and the amount of waste used in making batteries. Smaller batteries would decrease the size of electronics and increased battery life is important for life-saving, portable medical devices such as insulin pumps.

One of the largest classes of chemical reactions are oxidation-reduction reactions. Many of these reactions include sugars, fats, and proteins to generate the energy needed for life to flourish. Electrochemical cells are batteries which use the difference in electrical potential to generate electricity.

Oxidation is the process of an atom losing an electron, while reduction is the process of an atom gaining an electron. In electrochemical reactions, one electron is lost and another is gained. The only way for this to be possible is if the loss and gain occurs at the same time. Thus, in an electrochemical cell, oxidation and reduction must occur at the same time (2) These processes occur by electrons transferring over a wire between the oxidized and reduced metals. These reactions (called half reactions) occur separately. These half reactions, which occur in half cells, can easily be conducted through submerging a metal rod into a solution with dissolved ions. Because metals have unique oxidation states and atomic numbers, they have different affinities for their electrons and release different amounts of electrons. This causes metals to have unique electrical potentials, or the ability to cause an electric current. Additionally, different arrangements of these metals cause different amounts of electricity to be conducted by the electrochemical cell.

An electrochemical cell is built with two half cells, attaching the dissolved ions with a salt bridge, and the metal rods with a metal electrically conductive strip (6). If we do not measure our values in standard conditions, we will use the Nernst equation to calculate our voltage. However, if the result is not in standard conditions, you can convert that result to standard condition through the Nernst equation:

$$E = E0 - (R * T)/(n * F) * \ln(Q)$$

where, E is the reduction potential, T is temperature, E^0 is the standard reduction potential, R is the universal gas constant (8.314 J/mol * K), n is the number of moles of electrons transferred in an electrochemical cell, and F is Faraday's constant (9.65*104 C/mol). Q is the reaction quotient, which gives the ratio of reactants and products and helps determine which direction the reaction will go.

The amount of electricity conducted when a half cell combines with a hydrogen electrode is called the standard reduction potential. The standard refers to the fact that the surroundings are at 1 atm and 25 degrees Celsius. The hydrogen half cell is used as a standard for all reduction potentials. The standard reduction potential for a zinc electrode is 0 volts (V). In this experiment, we will measure the standard reduction potentials for a specific group of metals (4). Once we have measured the reduction potentials of each metal, we can use electrochemistry to add the potential to determine the most efficient battery. This potential refers to the pull on the electrons by the atom. This pull is related to the amount of electricity generated in the reduction oxidation reaction. For example, let us consider the electrochemical reaction between copper and aluminum. We will substitute random values in for the reduction potentials, because we have not determined them yet. E is defined as electrochemical potential.

Cu2+ + 2 e- ----> Cu E = 0.5 volts
Al3+ + 3 e- ----> Al E = 0.8 volts

Here, we see the reaction presented by copper releases 0.5 volts, while the reaction for aluminum releases

0.8 volts. First, we must note reduction here is occurring in both reactions. But, as seen earlier, oxidation must be accompanied by reduction, so one of these reactions must be flipped (5). Because aluminum has the highest reduction potential, its reaction is favored and therefore the aluminum reaction stays the same and the copper reaction is flipped. By flipping the reaction, the reduction potential sign is also flipped.

$$E = 0.8 - 0.5 = 0.3 \text{ volts}$$

Therefore, the reduction potential is 0.3 volts for this reaction. We will use this logic to determine the most efficient metals to use in the battery. We have compiled a list of the oxidation states and reaction equations for all metals used in this experiment (**Table 1**).

The goal is to create a list of reduction potentials for various metals. These experiments were conducted by placing each of these metals in a 1 molar (M) solution of its associated ion. Then, we connected this half cell to the zinc half cell, and used a potentiometer to measure the current generated. We used zinc for all of these metals for consistency. The metals we tested with reference to zinc were copper, iron, magnesium, and silver. We hypothesize an electrochemical cell constructed of copper and iron will produce the strongest battery because the oxidation states of copper and iron are both 3+ (the largest of the supplied metals), and therefore the most amount of electrons will be transferred. Because by definition electricity is the movement of electrons, the more electrons transferred, the stronger the electrical current. We then tested the predicted best cell to determine if our prediction was correct.

This will allow us the ability to make electrochemical cells smaller and more efficient, developing portable electronics which are smaller and have longer battery life. In addition, we will be able to increase the lifespan of medical devices which rely on battery power such as insulin pumps. In this experiment, we will measure the various voltages of these reactions by constructing many combinations of electrochemical cells. With these measurements, we will be able to construct a reduction series for the metals. With these series, we can determine which cells and which metals that compose them will generate the most amount of electricity, and therefore determine the most efficient battery.

RESULTS

In this experiment, we tested different electrochemical cells in order to experimentally determine the reduction potentials of copper, zinc, magnesium, and silver. We created four electrochemical cells, with each of the four aforementioned metals. We measured the voltage produced by each of these cells and used that information to calculate the reduction potentials of each of these four metals. With the knowledge that zinc has a reduction potential of -0.76 volts and the experimentally determined reduction potentials of the other four metals, we predicted the cell which produces the most volts. We then tested to see if this prediction of the cell voltage was correct by testing the predicted cell twice.

Table 1. Oxidation states and reduction equations for all metals used in this experiment

Metal name	Metal oxidation state	Metal reduction equation
Zinc	Zn^{2+}	$Zn^{2+} + 2e^- \leftrightarrow Zn(s)$
Silver	Ag^{1+}	$Ag^+ + e^- \leftrightarrow Ag(s)$
Magnesium	Mg^{2+}	$Mg^{2+} + 2e^- \leftrightarrow Mg(s)$
Iron	Fe^{3+}	$Fe^{3+} + e^- \leftrightarrow Fe2+(s)$
Copper	Cu^{2+}	$Cu^{2+} + 2e^- \leftrightarrow Cu(s)$

Table 2. Copper and zinc electrochemical cell

Electrochemical cell measurements for the copper and zinc cell		
Trial	1	2
Metals used	copper and zinc	copper and zinc
Reduction potential of zinc	-0.76 volts	-0.76 volts
Measured voltage with potentiometer	1.07 volts	1.13 volts
Calculated reduction potential of copper	0.31 volts	0.37 volts

We created electrochemical cells with 2 cm long pieces of copper, iron, magnesium, and silver in one half cell, and zinc in the other half cell (**Tables 2-5**). It is important to note that the reduction potential of zinc is -0.76 volts (1).

We conducted two trials with the copper and zinc electrochemical cell (**Table 2**). From these results, we observed that the total voltage generated in trial 1 was 1.07 volts. Because the reduction potential of zinc ic -0.76 volts, we can calculate the reduction potential of copper:

$$1.07 - 0.76 = 0.31 \text{ volts}$$

In addition, for the second trial, the voltage generated was 1.13 volts. This means the calculated reduction potential of copper was 0.37 volts. On average, we determined the reduction potential of copper as 0.34 volts.

Table 3 reports the measurements made in two trials for the iron and zinc electrochemical cell. From these results, the total voltage generated was 1.57 volts in the first trial. Because the reduction potential of zinc is -0.76 volts, we can calculate the reduction potential of copper:

$$1.57 - 0.76 = 0.81 \text{ volts}$$

The total voltage generated by the cell was 1.52 volts in the second trial, resulting in a calculated reduction potential

Table 3. Iron and zinc electrochemical cell

Electrochemical cell measurements for the iron and zinc cell		
Trial	1	2
Metals used	iron and zinc	iron and zinc
Reduction potential of zinc	-0.76 volts	-0.76 volts
Measured voltage with potentiometer	1.57 volts	1.52 volts
Calculated reduction potential of iron	0.81 volts	0.76 volts

Table 4. Magnesium and zinc electrochemical cell.

Electrochemical cell measurements for the Magnesium and zinc cell		
Trial	1	2
Metals used	Magnesium and zinc	Magnesium and zinc
Reduction potential of zinc	-0.76 volts	-0.76 volts
Measured voltage with potentiometer	1.64 volts	1.61 volts
Calculated reduction potential of Magnesium	-2.4 volts	-2.37 volts

of 0.76 volts. Over the course of both trials, the average reduction potential of iron was 0.785 volts.

We then measured the voltage generated in the magnesium and zinc electrochemical cell in trials 1 and 2 (**Table 4**). From these results, the total voltage generated was 1.64 volts in trial 1. Because the reduction potential of zinc is -0.76 volts, we can calculate the reduction potential of Magnesium:

$$-1.64 - 0.76 = -2.4 \text{ volts}$$

Here, the reduction potential of Magnesium if 1.64 volts and not -1.64 volts. This is due to the nature of an electrochemical cell. Because electrochemical cells must have a reduction and oxidation component, one of the metals must go through the reduction phase, and another must go through the oxidation phase. The metal with the higher reduction potential generates more of an electrical charge when being reduced, and thus goes through the reduction phase, and the other metal goes through the oxidation phase. Thus, we can create an equation:

$$E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode}$$

Table 5 reports the measurements in the silver and zinc electrochemical cell in trials 1 and 2. From these results, the total voltage generated was 1.07 volts in trial 1. Because the reduction potential of zinc is -0.76 volts, we can calculate the reduction potential of copper:

$$1.52 - 0.76 = 0.76 \text{ volts}$$

From the second trial, the total voltage of the cell was 1.147 volts. Therefore, the reduction potential of silver was calculated as 0.71 volts. The average reduction potential for copper was calculated as 0.735 volts.

Table 5. Silver and zinc electrochemical cell.

Electrochemical cell measurements for the silver and zinc cell(Trial 1 and 2)		
Trial	1	2
Metals used	silver and zinc	silver and zinc
Reduction potential of zinc	-0.76 volts	-0.76 volts
Measured voltage with potentiometer	1.52 volts	1.47 volts
Calculated reduction potential of silver	0.76 volts	0.71 volts

Table 6. Electrochemical cell measurements for ideal cell of iron and magnesium

Electrochemical cell measurements for ideal cell of iron and magnesium		
Metals used	iron and magnesium	iron and magnesium
Measured voltage with potentiometer	3.19 volts	3.24 volts

Finally, we created the iron and magnesium cell, and tested it to determine if our predictions that the iron and magnesium cell was indeed the cell which produced the greatest voltage.

Our data showed that overall, the predictions of the best electrochemical cell were correct. The average voltage from these two trials was 3.215 volts.

DISCUSSION

The data from the experiments showed that the magnesium and iron electrochemical cell was the most efficient electrochemical cell. It had an average voltage of 3.125 volts. This was the largest of all the voltages. This was due to the fact that magnesium had the lowest electrochemical potential and iron had the highest electrochemical potential. From this experiment, we concluded that the iron-magnesium cell was the most efficient. We also discovered that the electrochemical cells that produced the greatest voltage contained two metals; one had an extremely low electrochemical potential and another had an extremely high electrochemical potential. The more extreme these values, the more electricity generated. We found electrochemical cells out of magnesium and iron unless the conditions where the cell is being used does not permit such an action. For example, someone should not use this type of cell near damp areas, because iron rusts. Below is a table of the costs of all the metals we used (**Table 7**).

Because we concluded that an electrochemical cell made of iron and magnesium was the most efficient, we must examine the prices of these metals. These two metals are the least expensive (iron) and third least expensive (Magnesium) of the metals tested. We conclude cost is no obstacle in the cell we propose.

Some qualitative observations were that we did not observe any change visually in the system. There was no color change or distinct odor. All the metals were in the same rod-shaped shape. The surface area of the metals were all the same because the metal rods were all the same size and dunked in the solution at equal levels.

In this experiment, we determined the reduction

Table 7. Costs of the metals we used.

Zinc	1.05 USD per pound
Iron	0.24 USD per pound
Copper	2.80 USD per pound
Silver	209.35 USD per pound
Magnesium	2.10 USD per pound

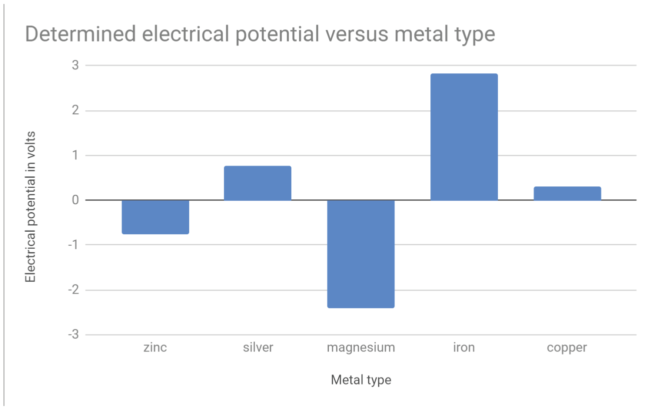

Figure 1. Bar chart exhibiting the experimentally measured electrical potentials.

potentials of five different metals. Because oxidation and reduction is what occurs in electrochemical cells, and the larger the reduction potential the more volts generated, we wanted to select the metal with the highest reduction potential for one of the half cells. The metal chosen was iron (reduction potential of 0.785 volts). Because the reduction potential is the opposite of the oxidation potentials (i.e. if reduction potential = 1, oxidation potential = -1), the other half cell should have the lowest reduction potential. The metal that satisfied this was magnesium (-2.385 volts reduction potential). Simple calculations produce the electrical capacity of this cell:

0.785 volts - (-2.385 volts) = 3.17 volts

Theoretically, 3.17 volts will be produced from the electrochemical cell comprising of magnesium and iron. This was experimentally supported when we determined an average reduction potential of 3.215 volts after two trials. We conducted every electrochemical cell trial twice, in identical conditions to increase the strength of our results.

Some errors in this experiment include using copper for the conduction wire in the electrochemical cells. Because the copper wire was not insulated perfectly and copper is not a perfect conductor, the voltage readings were not 100% accurate. However, because all the electrochemical cells used copper, the drop in electrical charge would be uniform and not cause a statistically significant jump in the voltage. Some future experiments would be to experimentally create the other electrochemical cells and measure the voltage generated to possibly find an electrochemical cell which produces more voltage.

METHODS

All the metals used in this experiment were obtained from the company, Fisher Scientific. All the metals used were advertised as 99.9% pure, and for the purposes of this experiment, we considered them pure. None of the metals were alloys. We conducted each trial twice, and averaged the results.

Firstly, we polished the metals used in the experiment (zinc, copper, magnesium, iron, and silver) with steel wool. We polished the metals to ensure the metals were as pure as possible and to scrape off any oxidized metal. We then acquired five zinc pieces of metal, one copper piece of metal, one silver piece of metal, one magnesium piece of metal, and one iron piece of metal. We then dropped four droplets of the correct solution into each well of a 24-well plate (**Figure 2**).

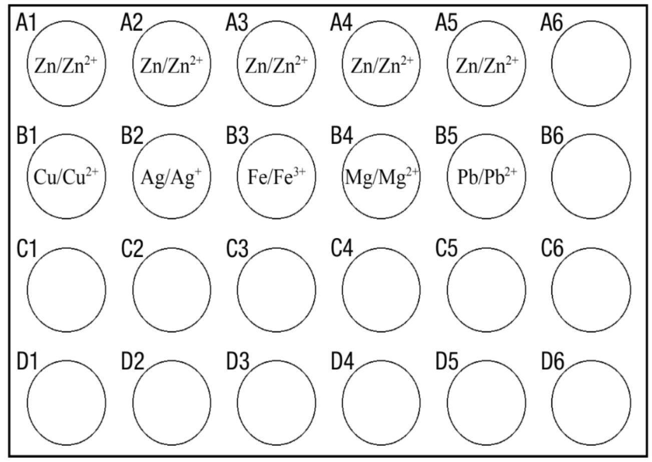

Figure 2. Experimental set-up in a 24-well plate that demonstrates where each ionic solution was placed. Generated with the drawing tool from Google Docs.

With the well plate set up, we then placed each piece of polished metal in the respective cell. Next, we prepared the salt bridge. We dipped a piece of filter paper in KNO3 solution. We then took out our voltmeter, and placed the negative terminal on the zinc electrode, and the positive terminal on the other electrode (copper, magnesium, iron, or silver). We connected the two metal ion solution with the filter paper soaked in KNO_3, and watched the voltmeter. We recorded the measurements made by the voltmeter.

Received: June 30, 2019
Accepted: May 5, 2020
Published: July 10, 2020

REFERENCES

1. Libretexts. "6.2: Standard Electrode Potentials." Chemistry LibreTexts, Libretexts, 23 Feb. 2019, https://chem.libretexts.org/Courses/Mount_Royal_University/Chem_1202/Unit_6:_Electrochemistry/6.2:_Standard_Electrode_Potentials.
2. Clark, Jim. "DEFINITIONS OF OXIDATION AND REDUCTION (REDOX)." Definitions of Oxidation and Reduction (Redox), Jan. 2013, www.chemguide.co.uk/inorganic/redox/definitions.html.
3. Elements Oxidized, www.occc.edu/kmbailey/Chem1115Tutorials/Element_Oxidized.htm.
4. Redox Reactions, www.chemistry.wustl.edu/~coursedev/Online tutorials/Redox.htm.
5. Libretexts. "Oxidation-Reduction Reactions."Chemistry LibreTexts, Libretexts, 5 June 2019, chem.libretexts.org/Bookshelves/Analytical_Chemistry/Supplemental_Modules_(Analytical_Chemistry)/Electrochemistry/Redox_Chemistry/

Oxidation-Reduction_Reactions.
6. "Redox Reaction."Redox Reaction - an Overview | ScienceDirect Topics, www.sciencedirect.com/topics/engineering/redox-reaction.
7. Libretexts. "6.2: Standard Electrode Potentials." Chemistry LibreTexts, Libretexts, 23 Feb. 2019, https://chem.libretexts.org/Courses/Mount_Royal_University/Chem_1202/Unit_6:_Electrochemistry/6.2:_Standard_Electrode_Potentials.