cKit™

The Teacher's Guide

to the Corrosion Toolkit 2nd Edition

AMPP™

AMPP™

Welcome

AMPP (The Association for Materials Protection and Performance) is a global community of professionals dedicated to materials protection through the advancement of corrosion control and protective coatings, ensuring the maximum performance, integrity, and durability of the assets our society depends on every day.

If you have never given corrosion much thought, consider this: the effects of corrosion is estimated to cost US$2.5 trillion globally, or 3.4% of the global GDP.

With the input from experts in the corrosion industry, we have developed a teaching kit, Corrosion Toolkit (cKit™), for math and science teachers to provide a fun and innovative hands-on learning experience to students in materials science. Teachers will learn what experiments to conduct and how to incorporate them into their classroom curriculum. Listed below are items contained in the cKit™:

- A digital voltmeter
- An electrode metal set that includes metal strips of aluminum, copper, lead, magnesium, tin, and zinc
- A AA battery and holder
- Alligator clips
- A teacher's guide to the Corrosion Toolkit
- A comic book that tells a story about combating corrosion

It is our hope that when your students learn more about the science behind corrosion, they will be intrigued by it and develop interests in that field. After you perform these experiments, I encourage you and your students to visit our Web site: **ampp.org/emerg** to find valuable information about future careers in corrosion engineering and scholarship opportunities available to qualified students.

On behalf of AMPP, we would like to thank you for helping to fuel the inquisitive minds of your

students. Sincerely,

The AMPP EMERG Student Outreach Team

ATTENTION TEACHERS:

The experiments in the cKit™ have been aligned to the Performance Expectations and the 3 dimensions of the Next Generation Science Standards (NGSS). The focus is on High School, but the experiments are also appropriate for Middle School. The 3 dimensions are: disciplinary core ideas, crosscutting concepts and science and engineering practices. The emphasis is on the Physical Science standards, but Engineering Design is also supported by the experiments. Connections can also be made to the Nature of Science and to Engineering, Technology, and Applications of Science.

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CORROSION EXPERIMENT #1
Corrosion Exposure

INTRODUCTION

The corrosion of metals is the result of a chemical reaction called oxidation. When exposed to oxygen, iron will rust or oxidize. Large amounts of energy are required to produce pure metal from its ore since metal is more stable in an oxidized form (ore). Therefore, metals corrode and deteriorate through chemical reactions between it and the surrounding environment in order to return to a preferred lower energy state. The surrounding environment affects the rate at which iron or steel corrodes. Factors such as the humidity, presence of salts, amount of oxygen, etc. play a role in how quickly iron will rust.

In the experiment, students will investigate the effects of different environments on the corrosion of steel wool (iron). Pieces of steel wool will be placed in various combinations of water, salt and 3% hydrogen peroxide. Four "ingredients" are necessary for corrosion to occur. They are represented by the acronym "ACME": anode, cathode, metallic pathway and electrolyte. All 4 must be present for oxidation to occur. In this lab, the steel wool is the anode, cathode and metallic pathway. The salt is the electrolyte. The Teacher Notes Section provides more information about "ACME." This lab alters the amount of oxygen available for corrosion and the presence of salt in water forming an electrolyte solution.

cKit™ Lab #2 investigates the reactivity of different metals in the same solution – copper (II) sulfate. cKit™ Lab #4 investigates the effect of the pH of a solution on the corrosion rate of aluminum.

PRACTICAL APPLICATIONS

Steel is the second most used building material in the world. The rusting of steel is one of the most obvious corrosion processes known to mankind. Corrosion rates vary based on a multitude of factors including the type of metal and the environment it is in.

Metal corrodes much quicker at the ocean's surface than deep in the ocean. There is not as much oxygen at lower depths and it is often much colder. Water at the surface contains more oxygen due to wave action. Many man-made objects are located in areas that contain higher levels of oxygen. Some examples are: steel pilings supporting a bridge over water; a pier at the beach; offshore oil platforms and offshore windmills.

The corrosion of steel can be greatly reduced by controlling its surrounding environment (i.e. humidity, oxygen and the presence of an electrolyte such as salt). Anything around the house that is made of steel (from steel tools, to gardening tools, to cars and even water pipelines) can corrode.

The Teacher Notes section provides additional information on the practical applications of "ACME" related to the everyday world – this includes ships and barges in the ocean, steel posts and towers on land and rebar in concrete.

OBJECTIVE

The students will investigate the components needed for corrosion to occur and use "ACME" to explain the effect of different environments on the corrosion rate of a metal.

NEXT GENERATION SCIENCE STANDARDS (NGSS) CORRELATIONS

- Performance Expectations: HS-PS1-5 and HS-PS1-6
- Disciplinary Core Ideas: PS1.B
- Crosscutting Concepts: Stability and Change, Cause and Effect
- Science and Engineering Practices: Carrying Out Investigations

MATERIALS AND EQUIPMENT LIST

MATERIALS LIST:

- Small plastic cups or beakers
- Steel wool pieces – each about the size of a sugar cube
- Water
- Table salt
- 3% Hydrogen Peroxide
- Graduated cylinder (50 mL or 100 mL)
- Plastic spoons

**Note: Additional trials using different solutes may be included.*

PER STUDENT GROUP:

- **4** x Small plastic cups or beakers
- **4** x Steel wool pieces
- **100 mL** x water
- **1 tbsp.** x table salt
- **100 mL** x 3% hydrogen peroxide
- **1** x Graduated cylinder
- **4** x Plastic spoons

SAFETY NOTE

A lab apron and goggles should be worn when performing this lab. Students should wash their hands after the lab. It is the sole responsibility of the teacher conducting this Experiment to ensure that disposal methods are properly implemented. Dispose of the remnants of the Experiment in accordance with school district, state and federal environmental regulations. If in doubt, refer to the chemical disposal procedures as found in the most recent edition of the Flinn Scientific Catalog/Reference Manual or contact **flinnsci.com** for specific instructions.

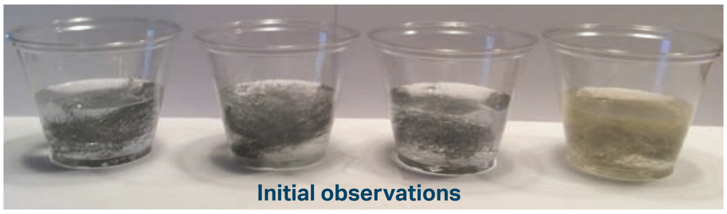
PROCEDURE

1. Label the plastic cups 1 through 4.
2. Add the following ingredients to each cup and stir to mix:
   - Cup #1 – 50 mL of water
   - Cup #2 – 50 mL of water and ½ tablespoon of salt
   - Cup #3 – 50 mL of 3% hydrogen peroxide
   - Cup #4 – 50 mL of 3% hydrogen peroxide and ½ tablespoon of salt
3. Place one piece of steel wool in each cup.
4. Use a plastic spoon to submerge the steel wool in each cup.
5. Make initial observations.
6. Make additional observations after 5 minutes and again after 15 or 20 minutes.
7. ***Optional:*** Let stand overnight and make final observations.
   Note that observations can be made continuously over the period of one week.
8. ***Optional:*** Additional trials using different solutes may be included.

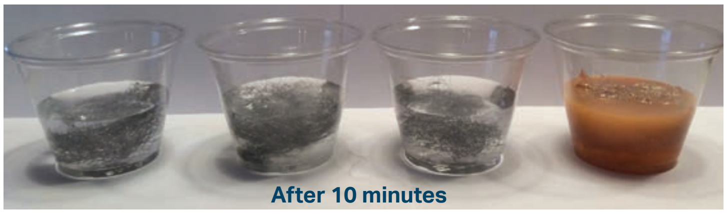
STUDENT OBSERVATIONS

Examples of observations made by students include:

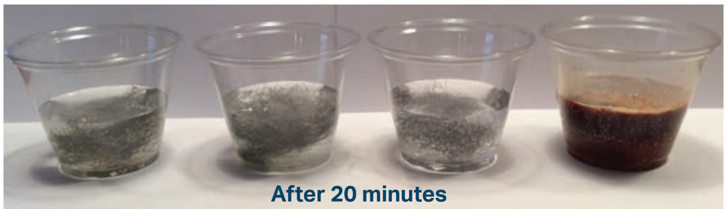
- **After 15-20 minutes:**
  - Cup #1 – No visible change; may be a slight color change depending on the minerals in the local tap water
  - Cup #2 – Slight color change; solution looks murky, minute rust spots on steel wool
  - Cup #3 – No visible change
  - Cup #4 – Steel wool is disintegrating and both it and the solution are rust colored
- **After 24 hours:**
  - Cups #1 & #2 – Red rust is visible on steel wool
  - Cup #3 – No visible change
  - Cup #4 – Steel wool will be completely disintegrated and/or become very dark, indicating iron oxide and/or iron hydroxide.



Initial observations



After 10 minutes



After 20 minutes

TEACHER NOTES, BACKGROUND AND EXTENSIONS

All supplies for the experiment can be purchased locally. Use scissors to cut the steel wool into equal sized pieces about 1" in diameter. Pulling on the steel wool can cause metal splinters. It is recommended that the teacher cut the steel wool prior to class and have it ready for the students to use. If the students are to use a 5th trial (cup) of their own choice have extra solutes available for them. Suggestions include: sugar, borax, vinegar, ammonia, etc.

The acronym "**ACME**" refers to the four requirements for a corrosion cell to occur: anode, cathode, metallic pathway and electrolyte. In this experiment the steel wool is the cathode, anode and metallic pathway. The same piece of metal can have thousands of tiny corrosion cells. Impurities in the metal and physical stress in the metal can cause one area to behave as an anode while another area behaves as a cathode.

Very little if any corrosion occurs in the cup with just tap water (cup #1) because the electrolyte is missing in "ACME." If there are enough ions dissolved in the water, some corrosion will occur over time.

The steel wool in the salt water (cup #2) gradually shows significant corrosion over time. The salt in the solution acts as the electrolyte to complete "ACME" for the corrosion cell.

Cup #3 typically does not show corrosion of the steel wool even after a longer period of time. The electrolyte (salt) is missing, and the 3% hydrogen peroxide does not have the dissolved minerals that tap water will have.

The steel wool quickly and massively corrodes in cup #4. Salt is present to provide the electrolyte, and the hydrogen peroxide breaks down into water and oxygen. The added oxygen in the solution increases the corrosion rate dramatically. The gaseous bubbles observed in cup #4 are oxygen gas. Some of the oxygen produced by the breakdown of hydrogen peroxide is used in the rapid corrosion of the steel wool, and some of the oxygen is released at the surface of the liquid.

Corrosion is an exothermic reaction. Usually the reaction occurs slowly enough that the heat dissipates to the surroundings without a noticeable change in temperature. The rapid rate of reaction in cup #4 produces enough heat that the cup will feel noticeably warmer than the other three.

A fifth cup can be added with each team choosing another solution to test, such as sugar, vinegar, ammonia, etc. Solutions without ions should not show much corrosion if any, since the electrolyte will be missing. Solutions with ions that will act as an electrolyte should show appreciable corrosion over time. A comparison of tap water and distilled water should be interesting to the students if the local water supply has a high concentration of minerals.

After performing the basic experiment, extensions using other variables such as temperature and coarseness of the steel wool could be investigated. The temperature can be varied by putting the cups under a heat lamp, on a window sill, in a refrigerator, etc. Other mechanisms for increasing the amount of oxygen in the solution could be explored; for example using an aquarium aerator or agitating the cup.

"ACME"

An **anode** is the site where oxidation takes place. The metal atoms lose electrons and become positive ions. An example is $Fe \rightarrow Fe^{+2} + 2e^-$ **or** $Fe \rightarrow Fe^{+3} + 3e^-$. Metals oxidize to become more stable; it puts them in a lower energy state. In nature, metals are found in ore in their oxidized form.

A **cathode** is the site where reduction takes place. An element (often a metal) gains electrons. In terms of a metal, a positive ion becomes a neutral atom. Examples include $Cu^{+2} + 2e^- \rightarrow Cu$ and $2H^+ + 2e^- \rightarrow H_2$. Metals become less stable when they are reduced, meaning they are in a higher energy state. Therefore, pure metals are always trying to oxidize (corrode).

The **metal connector** (pathway) is where the electrons pass from the anode to the cathode. A battery is a corrosion cell that operates on the principles of "ACME." The metal connector runs between the positive and negative terminals of the battery external to the battery. The metal connector in this instance can include an appliance such as a light bulb or radio.

The **electrolyte** allows the circuit to be completed. It allows ions to flow. Examples include saltwater, moist paste and acid.

REAL WORLD APPLICATIONS

Corrosion is the gradual destruction of materials (usually a metal) by chemical or electrochemical reaction with its environment. The term corrosion is sometimes also applied to the degradation of plastics, concrete and wood, but generally refers to metals.

Salt water corrodes steel faster than fresh water due to the increased concentration of ions in the water (electrolyte). Steel objects corrode faster at the surface of water due to larger amounts of oxygen dissolved in the water. For example, steel objects at the bottom of the ocean such as the Titanic take many years to corrode away. Ships floating on the surface experience significant corrosion in a short amount of time unless measures are taken to protect the metal. These measures can include paint, zinc anodes or the use of impressed current.

A field trip can be arranged in order to view metal objects partially submerged in flowing water. Teachers in coastal areas will have numerous opportunities to view the differential corrosion taking place at the waterline versus in the atmosphere or where an object is permanently submerged.

Steel fence posts, lamp posts, flag poles and tower supports that are partially buried in soil for extended periods of time will corrode at a faster rate near the surface of the soil. Moisture in the ground and oxygen from the air combine to complete a corrosion cell at the interface.

Pipeline/utility service risers are underground gas lines that extend out of the ground, connecting to buildings. The transition from soil to air of a pipeline can be extremely sensitive to corrosion. The soil to air interface is an area that is greatly affected by moisture and oxygen levels and other factors. The risers need protection from contact with the soil to prevent a corrosion cell from forming. Material choice, coatings and wrappings can help mitigate the problem.

Steel rebar increases the tensile strength of concrete. Rust occupies a greater volume than steel which leads to increased stress and eventual cracking of the concrete when the rebar corrodes. Over time, chloride ions from de-icing road salt or ocean water spray can penetrate through concrete to the rebar damaging roads and bridges. Other mechanisms that lead to the corrosion of rebar in concrete include carbonation and the use of dissimilar metals.

OTHER RESOURCES

- **Website: The Corrosion Doctors – Corrosion of the Sunken Titanic**
  http://www.corrosion-doctors.org/Landmarks/titan-corrosion.htm
  A description of the corrosion caused by bacteria in the ocean depths.
  The Corrosion Doctors website is extensive with many pages of information related to corrosion.
- **Website: Corrosion Technology Laboratory – NASA: Kennedy Space Center**
  http://corrosion.ksc.nasa.gov/corr_fundamentals.htm
  Fundamentals of Corrosion and Corrosion Control. Lots of information and pictures.
  Includes the main topics: Why Metals Corrode; Electrochemistry; Forms of Corrosion; and Corrosion Control
- **Article: "Corrosion of Embedded Metals" – Portland Cement Association**
  http://www.cement.org/learn/concrete-technology/durability/corrosion-of-embedded-materials
  Excellent explanation of the factors affecting the corrosion of metal rebar in concrete.
  Includes a description of the components of "ACME."
- **Video: Modern Marvels – Corrosion and Decomposition**
  https://www.youtube.com/watch?v=pFJLWCE2QAU
  44 minute video available through the History Channel or the GeoChannel

Scan above QR code to quickly access external links!

SOURCE

This experiment was inspired by: Gus Schmucker; 4th Grade Student at Ligonier Middle School; Ligonier, Pennsylvania.

CORROSION EXPERIMENT #1: STUDENT ASSESSMENT QUESTIONS

1) Write a paragraph summarizing your observations. Note signs of chemical reaction or change after 5 minutes and after 20 minutes.

2) Describe the chemical reaction that forms rust.

Iron atoms combine with oxygen to form iron oxide, also called rust. This is a type of reaction known as oxidation.

3) What did the hydrogen peroxide contribute to the corrosion of the steel wool?

Hydrogen peroxide increases the amount of available oxygen for the oxidation of iron.

4) How does the concentration of oxygen affect oxidation rate?

The higher the oxygen concentration, the greater the rate and amount of oxidation.

5) How does the presence of salt affect oxidation rate?

The presence of salt increases the rate and amount of oxidation.

6) Using the acronym AMCE, which role does the salt perform in the corrosion of the steel wool in this lab?

Electrolyte

7) In what other ways could O_2 be added?

Possible answers include agitation of the water (such as wave action) and aeration as in an aquarium.

8) Why did the temperature change in Cup #4?

The chemical reaction producing the iron oxide is an exothermic reaction which means that it gave off energy in the form of heat.

9) What were the bubbles that formed in Cup #4 and why/how did they form?

The bubbles formed in Cup #4 are oxygen gas. Salt is present to provide the electrolyte, and the bubbles are formed when hydrogen peroxide breaks down into water and oxygen.

10) Write the basic chemical equation for the formation of rust.

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$

OR

$$2Fe + O_2 \rightarrow 2FeO$$

**Note: A portion of the questions in this experiment were taken from the Science NetLinks website at sciencenetlinks.org.*

11) What is the chemical name of rust?

Iron Oxide.

12) What is another name for the process of rusting?

Oxidation or corrosion

13) Why do metals submerged at great depths not corrode quickly?

Less oxygen dissolved in the water at greater depths and usually colder temperatures.

14) How could you protect the steel from corrosion?

Use a protective coating such as oil, paint or other polymer or galvanize it with zinc so that the metal does not come in contact with the electrolyte.

15) What is a chemical reaction?

The process by which one or more substances may be transformed into one or more new substances.

16) What is the definition of an anode, and electrolyte and a cathode?

- An **anode** is a piece of metal that readily gives up electrons
- An **electrolyte** is a liquid that helps ions move
- A **cathode** is a piece of metal that readily accepts electrons

17) Can you think of anything used in everyday life that rusts and corrodes if you leave it outside to be exposed to rain?

Examples may include gardening tools, bicycles, lawn chairs…anything with exposed iron.

18) Do other metals corrode or oxidize?

Yes. Silver tarnishes and loses its shine, copper oxidizes to a greenish color.

Additional Notes

CORROSION EXPERIMENT #2

Mr. Copper & Miss Sulfate

INTRODUCTION

Students will develop their own activity series of metals based on lab results. Qualitative observations will be used to rank metals from most to least reactive. Tie-ins include single-replacement reactions (SRR), oxidation-reduction, predicting reactions and practical applications such as galvanization, sacrificials and reducing metal from ore.

Students work in teams to determine the relative reactivity of six different metals in a solution of copper (II) sulfate. It is helpful if the students already have covered and have an understanding of atomic structure and bonding before performing this lab.

cKit™ Lab #4 uses one metal (aluminum) in different solutions, and shows that the type of environment makes a difference as well as the type of metal.

PRACTICAL APPLICATIONS

This experiment guides the students in the construction of a galvanic series of metals. Different environments will produce a different activity series of metals. A very common one is the galvanic series in seawater. The activity series produced in this lab correlates to the electro-chemical series (electro-motive series) which is a list of metals arranged in order of their standard potentials to the hydrogen electrode.

The selection of metallic materials should be made in light of their reactivity towards the specific environment where they are going to be used. The applications of these principles are shown by the world of electronic components where copper, aluminum, gold, carbon and silicon are constructed to last in humid and polluted environments. The combination of many of the metals used in this experiment is quite common in practical systems. Aluminum sleeves are crimped over steel, copper is clad onto stainless steel, gold is plated on aluminum and all these metals have to behave in close proximity in environments which are often very humid.

The Real World Applications section (page 15) included in this experiment contains practical applications of the activity series related to the everyday world. This includes hot water heaters, pipelines, boats, food cans and even the Washington Monument!

OBJECTIVE

The students will develop an activity series based on their own lab observations and will use it to predict and explain single replacement reactions and oxidation-reduction.

NEXT GENERATION SCIENCE STANDARDS (NGSS) CORRELATIONS

- Performance Expectations: HS-PS1-1 and HS-PS1-2
- Disciplinary Core Ideas: PS1.A, PS1.B and PS2.B
- Crosscutting Concepts: Patterns, Stability and Change, Structure and Function
- Science and Engineering Practices: Carrying Out Investigations, Engaging in Argument from Evidence, Communicating Information

MATERIALS AND EQUIPMENT LIST

MATERIALS LIST:

- 0.2 M copper (II) sulfate solution
- Metal strips, cut into 1" pieces
  - Aluminum (Al)
  - Lead (Pb)
  - Tin (Sn)
  - Zinc (Zn)
  - Magnesium (Mg)
- Iron nails – 2½" bright nails, smooth
- Preforms or Test tubes
- Preform rack or Test tube rack
- Non-iodized salt
- Steel wool
- Silver nitrate solution – *For teacher demonstration only*

PER STUDENT GROUP:

- **25-50 mL** x Copper (II) sulfate solution
- **1** x 1" Aluminum (Al) metal strip
- **1** x 1" Lead (Pb) metal strip
- **1** x 1" Tin (Sn) metal strip
- **1** x 1" Zinc (Zn) metal strip
- **1** x 1" Magnesium (Mg) strips
- **2** x 2½" Iron (Fe) bright nails
- **6** x Preforms or Test tubes
- **1** x Preform or Test tube rack
- **1** x Steel wool piece

TEACHER PREPARATION

1) Metal Sources:

The metal strip packed provided in the cKit™ includes Copper (Cu) Aluminum (Al), Lead (Pb), Magnesium (Mg), Tin (Sn), and Zinc (Zn). Bright nails (1½" to 2½", smooth finish) will represent the iron pieces and can be purchased at any hardware store.

The metal strips will need to be cut into approximately 1" pieces using scissors or sheet metal shears. Additional metal strips can be purchased at Flinn Scientific (**flinnsci.com**):

- Aluminum strips, pkg. of 10, item #: A0178
- Lead strips, pkg. of 6, item #: L0065
- Magnesium ribbon, 25 grams, item #: M0001
- Tin strips, pkg. of 6, item #: T0087
- Zinc strips, pkg. of 10, item #: Z0024

2) Distinguishing Metal Types:

In order to distinguish the metal strips, the dimension of the Zinc strip is 5" x ½"; and Magnesium is 6" x ⅛". The Copper, Aluminum, Lead and Tin strips are all 6" x ½". Copper is most easily differentiated, due to its color. Of the remaining three (Aluminum, Lead and Tin), Lead is the darkest color and heaviest strip; and the Tin strip is approximately three times as thick and heavier than the Aluminum strip.

3) Preparation of 0.2M copper (II) sulfate solution:

Dissolve 32 grams of copper (II) sulfate (CuSO4) in 1 liter of distilled water.

An alternative method is to dilute 1 M $CuSO_4$ solution to 0.2 M $CuSO_4$ by adding distilled water. For every 100 mL of 0.2 M $CuSO_4$ solution add 80 mL of distilled water to 20 mL of 1 M $CuSO_4$.

Add 1.5 to 2.0 grams of non-iodized salt (NaCl) to every 100 mL of copper (II) sulfate solution. Additional information can be found in the Teacher Notes section.

SAFETY NOTE

A lab apron and goggles should be worn when performing this lab. Students should wash their hands after the lab. It is the sole responsibility of the teacher conducting this Experiment to ensure that disposal methods are properly implemented. Dispose of the remnants of the Experiment in accordance with school district, state and federal environmental regulations. If in doubt, refer to the chemical disposal procedures as found in the most recent edition of the Flinn Scientific Catalog/Reference Manual or contact **flinnsci.com** for specific instructions.

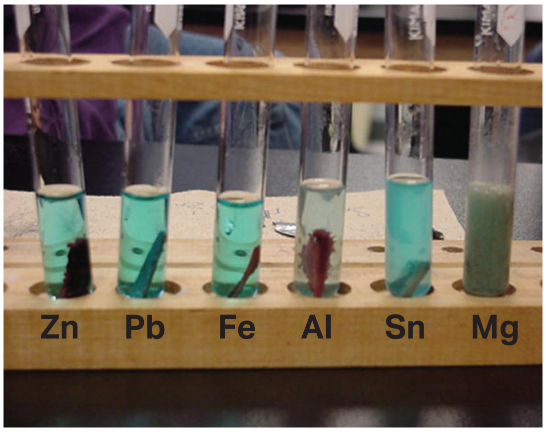
PROCEDURE

1. Pour approximately 5 mL (or 1 inch) of "blue solution" into each test tube.
   NOTE: The "blue solution" is the prepared copper (II) sulfate solution. Depending on how the teacher desires to process the lab with the students, the students can be told before or after they perform the experiment that it is copper (II) sulfate.
2. Clean the 6 metal samples with steel wool to remove any oxide.
3. Label each test tube according to the metal sample that will be placed in the tube.
4. Place a different metal sample in each of the test tubes one at a time so initial observations can be made.
   Do not cap or cover the openings of the test tubes.
5. Record observations for 5-10 minutes (look for signs that a chemical reaction is occurring: i.e. gas bubbles being produced; temperature changes; changes in color; a solid precipitate forming; solid disintegrating, etc.).
6. Rank the metals in order of reactivity.

STUDENT OBSERVATIONS

Examples of observations made by students include:

- **Magnesium:** begins reacting almost immediately, many gas bubbles are produced, reddish specks form (precipitate), solution turns cloudy, surface of metal turns green
- **Aluminum:** bubbles slowly form on the metal surface, rust-colored powdery substance forms (precipitate), test tube feels warm to the touch, blue color disappears
- **Zinc:** surface of metal turns black almost immediately, reddish precipitate (solid) builds up on surface of the metal
- **Iron:** surface of nail turns rusty almost immediately, rust slowly builds up
- **Tin:** no immediate reaction, slight color change on edges of metal, small amount of fine reddish particles on bottom of test tube, solution turns cloudy or whitish
- **Lead:** no obvious reaction, might see a slight color change on edge of the metal



TEACHER NOTES, BACKGROUND AND EXTENSIONS

Aluminum is a very active metal but is slow to start reacting in this lab. Aluminum oxide forms so quickly and bonds so tightly to the surface of the metal that it is difficult to have a clean surface when it is put into the solution in the test tube. Adding sodium chloride to the copper (II) sulfate solution jump starts the reaction like a catalyst – it allows the solution to reach the surface of the metal. Add a little NaCl to the copper sulfate solution prior to the start of the lab (approximately 1.5 to 2 grams per 100 mL of 0.2M copper (II) sulfate solution). This allows the aluminum to "behave" according to the actual activity series and does not affect the placement of the other metals.

Teachers can decide whether or not to discuss the addition of the salt with their students based on their goals and objectives and the level of the students. After the students perform the lab, the teacher may wish to show them two different aluminum trials, one with salt added to the copper sulfate solution and one without.

Based on the level of the students, a possible lab extension is to have the students clean only one half of each metal sample with the steel wool before placing them in the solution. This can also be done as a teacher demonstration after the students perform the lab. The oxide that forms on the surface of most metals serves to protect the metal underneath. If the oxidation is not removed with steel wool the reaction with the copper (II) sulfate solution will be reduced or will not occur. Students can compare the reaction rates of the two different surfaces – cleaned and oxide coated.

It is helpful to tap the sides of the test tubes (near the bottom) after the reactions have been occurring for a few minutes. This will break some of the precipitate loose from the surface of the metal. This allows the students to see the product better and also exposes the metal surface to keep the reaction going. It is beneficial for the students to use a magnifier or jeweler's loupe to make observations of the reactions occurring in the test tubes if they are available.

As the students perform the lab and write observations in their journals, the teacher can circulate around the room and listen to the group discussions as they make their rankings. The students usually think the metals are "rusting". Copper metal is actually being formed…not rust. The misconception will be addressed during the class discussion of results after the lab is completed. Students often ask what criteria they are supposed to use to make their group ranking of metal reactivity: How quickly do they start to react? How many signs of reaction are observed? How much precipitate forms? The students are purposely NOT given a direct answer to this question. It is up to each group to decide the criteria used to rank the metals.

After each lab group has completed their ranking of the activity of the metals, a class discussion is held to develop a single final ranking of the metals for the whole class. The teacher may have to mediate because all the groups do not always agree on the order of ranking. The teacher can use "majority rules" as a guideline and it almost always works.

HOW TO LEAD THE CLASS DISCUSSION ABOUT LAB RESULTS

Ask the students which metal they think is most reactive. They usually get this one right. It is magnesium. Write it at the top of the board. Then talk about which one is next until all the metals have been listed. The answers may vary from group to group based on the criteria they used to make their ranking. It will take some discussion to come to a final agreement. Lead and tin can be difficult to distinguish because neither one reacts much.

Here is the final ranking to be agreed upon:

1. Magnesium
2. Aluminum
3. Zinc
4. Iron
5. Tin
6. Lead

After the class has agreed on this list, tell the students that a metal was left out. This usually confuses them because they had six metal samples and there are six elements on their list. At this point tell the students what the "blue stuff" is – copper sulfate solution. With encouragement they come up with the missing metal – Copper. Ask the class where copper should go on the activity of metals list. Most are usually not sure and many will want to place it at the top of the list and others will want to place it somewhere in the middle. Thus, you begin your discussion of what really happened in the test tubes: a single-replacement reaction involving oxidation-reduction. An analogy will be used to help the students figure out where copper belongs in the list instead of the teacher giving the students a direct answer.

Put a sample equation on the board such as:

$$Zn + CuSO_4$$

Ask the students to fill in the products:

$$Zn + CuSO_4 \rightarrow Cu + ZnSO_4$$

The students realize that copper was formed, so somehow iron and copper switched places. Rewrite the equation using ions this time:

$$Zn + Cu^{+2}\,SO_4^{-2} \rightarrow Cu + Zn^{+2}SO_4^{-2}$$

The sulfate is unchanged. Zinc went from an atom to a positive ion (describe oxidation at this time) and copper went from a positive ion to a neutral atom (describe reduction at this time). It is helpful to define a single-replacement reaction at this point. The following bullet points describing oxidation and reduction are derived through analyzing the equation and a teacher-led class discussion.

$$Zn + Cu^{+2}\,SO_4^{-2} \rightarrow Cu + Zn^{+2}SO_4^{-2}$$

reduced

oxidized

Zn OXIDATION	Cu REDUCTION
element → compound	compound → element
neutral atom → (+) ion	(+) ion → neutral atom
loses electrons	gains electrons
less stable → more stable	more stable → less stable
metal in ore form	makes pure metal

MR. COPPER AND MISS SULFATE GO TO THE DANCE

An analogy will be used to help the students place copper in the correct position in the activity series. The following scenario has been useful in aiding all levels of students develop a basic understanding of how the activity series works. A 1 to 10 scale is used to rate the metals.

Here is the original rating or ranking:

10 Magnesium
9 Aluminum
7 Zinc
6 Iron
5 Tin
4 Lead

The lovely Miss Sulfate decides to attend the homecoming dance with Mr. Copper. Even though Mr. Copper is at about a 3 on the scale, he was willing to pay for his date's ticket to the dance so Miss Sulfate agreed to go with him.

Mr. Copper and Miss Sulfate are having a great time at the dance until Mr. Zinc arrives. Although he's at about a 7 on the scale, Mr. Zinc was unwilling to pay for an extra ticket and arrives to the dance alone. Seeing him without a date, Miss Sulfate wastes no time in dumping Mr. Copper and getting together with Mr. Zinc.

Where would this place copper on the scale? Is Miss Sulfate going to trade a higher number for a lower number or vice versa? Mr. Copper is "reduced" to being alone - no charge - no attraction - no date. Mr. Zinc became "charged" when he walked in the door (was oxidized) by giving his unwanted electrons to Mr. Copper which is what reduced him.

Miss Sulfate is attracted to Mr. Zinc because of his opposite charge and they are now dancing while Mr. Copper has to wait in the singles line all alone (elemental atoms of copper). Zinc can "force" his unwanted electrons onto the copper ions because zinc is more reactive than copper.

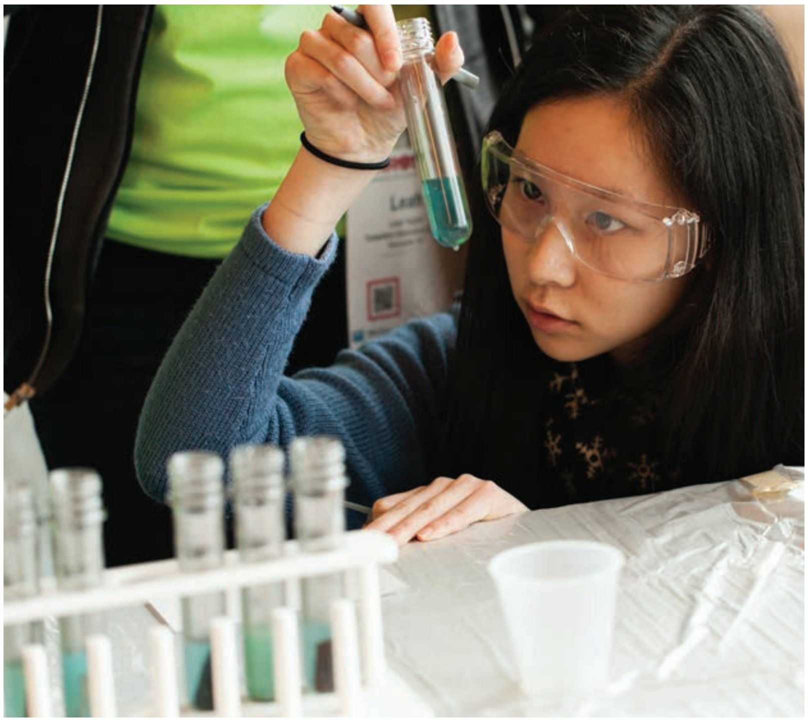
Now the students are ready to place copper on the activity scale in the proper location. Since copper got "replaced" (dumped as the students will often say) in every test tube, he must be at the bottom of the list. So copper is added to the bottom as a "3".

10 Magnesium
9 Aluminum
7 Zinc
6 Iron
5 Tin
4 Lead
3 Copper

The teacher can then use another equation to see if the students are getting a handle on the concept. It is good to use iron or magnesium. The students help complete the equation once the teacher writes the reactants on the board. Then have the students identify which metal was oxidized and which was reduced.

The next question the teacher can pose to the students is this: "Is Mr. Copper destined to be in the singles line forever? Can he ever hope to get a date while the other 'guys' (metals) are around?"

The answer is **YES** if a "2" or a "1" can be found. At this point the teacher can do an optional demonstration using a copper strip and silver nitrate.



TEACHER DEMONSTRATION: COPPER IN SILVER NITRATE SOLUTION

Place a strip of copper metal in a test tube with approximately 7 mL of 0.1M silver nitrate solution*. Silver nitrate solution can be made one of two ways:

- Dissolve 1.7 grams of silver nitrate ($AgNO_3$) in 100 mL of distilled water
- Purchase 0.1 M silver nitrate solution

Both solid silver nitrate and 0.1 M silver nitrate solution are available from Flinn Scientific.

Have students make observations and place silver in their activity series. Silver is replaced by the copper – the silver ions are reduced and silver metal forms on the surface of the copper. The copper is oxidized and copper ions go into solution eventually turning the solution a pale blue color.

The classroom discussion of the activity series and the dating analogy continues at this point. Good news for Mr. Copper: Mr. Silver is less reactive and will be reduced (replaced) by Mr. Copper. Therefore, copper can steal silver's date (Miss Nitrate) and he is left alone to hang around by himself.

$$Cu + AgNO_3 \rightarrow Ag + Cu(NO_3)_2$$

Silver is added to the activity series list as a "2". The students now want to know who "1" is (the real loser)! The answer is gold or platinum. Now ask the students if there is a metal higher than a 10. The answer is YES. Examples are calcium, sodium and potassium. This can lead to a discussion of the periodicity of elements on the periodic table. The one who can steal a date from anyone else? Mr. Francium.

At this point, the teacher can put a few more equations on the board that represent the reactions that occurred in class and ask the students to label which "participants" are being reduced and which are being oxidized. Then the students can be given an equation such as this one:

$$Fe + MgSO_4 \rightarrow$$

Ask the students to predict what will happen. The correct answer is "no reaction." Iron is less reactive than magnesium (lower on the scale) and therefore cannot steal away the date from magnesium. A "6" will not take away the "girl" from a "10." A less reactive metal cannot replace a more reactive metal in a compound.

A quick demonstration that also illustrates the different reactivity of metals is to "burn" a small piece of magnesium ribbon, a sample of steel wool and a copper strip or sheet using either a propane torch or a Bunsen burner. The metal is held with needle nose pliers or a pair of tongs. This should be done by the teacher in a lab setting where it can be controlled. The magnesium will ignite and burn producing a white smoke which is magnesium oxide. The white light given off by the burning magnesium is so bright that it should be observed out of the corner of the eye and not stared at directly. The steel wool will spark and glow red as it oxidizes. The copper will oxidize and change color on the surface but will not spark or burn.

An expanded version of the activity series is given to the students to include carbon and hydrogen even though they are not metals. Hydrogen is a part of acids and therefore important when considering the environment in which the metal will be used. Carbon in the form of cola is important for the reduction of metal from ore. There is more information about this in the section on Real World Applications.

10 Magnesium
9 Aluminum
8 Carbon
7 Zinc
6 Iron
5 Tin
4 Hydrogen
3 Copper
2 Silver
1 Gold

NONMETALS

FIGURE 1 shows the electro-chemical series (electro-motive series), which is a list of metals arranged in order of their standard potentials to the hydrogen electrode.

Metals higher up in the electrochemical series displaces metals lower in the series. This means that when connecting two metals with different potentials the metal with the lowest potential corrodes.

Element	Electrode Potential (volts)
Gold	+1.50
Chlorine	+1.36
Platinum	+1.20
Bromine	+1.07
Mercury	+0.85
Silver	+0.80
Iodine	+0.54
Copper	+0.34
Antimony	+0.10
Hydrogen	+0.00
Lead	-0.13
Tin	-0.14
Nickel	-0.24
Cobalt	-0.28
Cadmium	-0.40
Iron	-0.44
Tungsten	-0.58
Chromium	-0.74
Zinc	-0.76
Manganese	-1.19
Aluminum	-1.67
Beryllium	-1.85
Magnesium	-2.37
Sodium	-2.71
Barium	-2.80
Calcium	-2.87
Potassium	-2.92
Rubidium	-2.92
Lithium	-3.04

FIGURE 1

REAL WORLD APPLICATIONS

Practical applications of the metal activity series are numerous. Several examples are discussed below that work well at the middle school and high school level.

Galvanized metal is steel or iron coated with zinc. Zinc is more reactive than iron and will react with an oxidizing agent before iron. Oxidized zinc will form a tough, protective coating on the outside of the steel or iron, thus keeping the oxygen (or oxidizer) away from the iron underneath. Oxidized iron does not make a protective coating because it flakes off exposing new iron to be oxidized. Examples of uses of galvanized metal that students will recognize include: guardrails, grain bins, furnace ductwork, galvanized buckets and watering cans, and some nails. Cans containing acidic foods such as tomato sauce and pineapple juice are often made of galvanized metal. However, many food cans are now polymer coated on the inside instead of galvanized. Paints and other polymer coatings are often used to protect metals from corrosion. cKit Lab™ #4 in this toolkit – Poly Coat – investigates the need for a polymer liner inside soda cans to protect the aluminum metal from the phosphoric acid in the soda.

A **sacrificial** is another common example of the activity series in action. This involves purposely using a more reactive metal to protect a less reactive metal. The use of a sacrificial zinc, aluminum or magnesium rod (anode) in a hot water heater is an everyday example of this. Other examples of sacrificial use include zinc tabs and plates on boats and ships. Zinc is more reactive than iron and will be attacked first by an oxidizing agent such as oxygen in seawater. Buried gas and water pipelines are connected to large blocks of zinc. As the zinc corrodes, electrons are passed along the metal connector to the steel/iron pipeline thus preventing the pipeline from corroding. This happens because zinc is more reactive than iron. Once all of the zinc is oxidized, the pipeline will begin to corrode if the anode is not replaced.

The **extraction of metals** from their ores is also related to the activity series. Metals are in oxidized form in their ores. They need to be reduced to get pure metal. Metals are usually reclaimed from their ores in one of two ways: chemical reduction (using carbon in the form of cola) or electrolytic reduction (using electricity).

In **chemical reduction**, the oxidized metal in ore is melted or dissolved and exposed to a more reactive element. Carbon is a relatively cheap way to reduce metals from their ores. Carbon fits on the activity series scale between aluminum and zinc – in the lab discussion it was assigned a rating of "8". Thus chemical reduction using cola is a viable option for metals below carbon on the activity series.

Electrolytic reduction is used on metals that are more reactive than carbon. Electricity is used to reduce the oxidized metal in the molten or dissolved ore. Aluminum is the most abundant metallic element in the earth's crust, but has become widely used only in the past century because of the difficulty in extracting it from its ore. Electrolytic reduction requires an abundant source of cheap hydroelectric power. It is much cheaper to recycle aluminum than to extract it from its ore.
An interesting history connection is the Washington Monument. The capstone at the top of the monument is made of six pounds of aluminum. At the time of its completion in the 1880's, aluminum was more precious than gold, silver or platinum. Now we make soda cans out of it.

OTHER RESOURCES

- **Video: Alkali Metals in Water, Accurate! (Periodicity of Alkali Metals)**
  http://www.youtube.com/watch?v=uixxJtJPVXk
  Small pieces of the first five alkali metals dropped into water. Approximately two and a half minutes.
- **Video: How to Change a Water Heater Anode – from This Old House**
  https://www.youtube.com/watch?v=2lUNlUZz4Os
  Excellent everyday practical example of a sacrificial anode. Approximately four and a half minutes.
- **Article: Pressure-Treated Wood: The Next Generation from Fine Homebuilding**
  http://www.finehomebuilding.com/pdf/021160082.pdf
  Practical example of corrosion caused by dissimilar metals in new decking and the remedies. CCA treated wood was banned from residential use in 2004. The new preserved wood has a much higher copper content which required the use of different hangers and fasteners. Excellent graphics.
- **Article: Montana Resources mines the water**
  http://www.pitwatch.org/montana-resources-mines-the-water/
  Recovering copper from mine wastewater. Showcases a practical use of the activity series.
- **Website: The Corrosion Doctors – Why Metals Corrode**
  http://www.corrosion-doctors.org/Definitions/why-corrosion.htm
  A description of the true driving force (energy) of various common metals in their quest to become more stable. The Corrosion Doctors website is extensive with many pages of information related to corrosion.

Scan above QR code to quickly access external links!

SOURCE

This experiment was inspired by: Debbie Goodwin; Science teacher at Chillicothe High School; Home of the Hornets; Chillicothe, Missouri.

CORROSION EXPERIMENT #2: STUDENT ASSESSMENT QUESTIONS

1) Write a final summary of oxidation/reduction using the following bullet points as a guide:

- Summarize oxidation and reduction by using an example from the lab.
- Be sure to include an equation and state which element was oxidized, and which element was reduced in the reaction.
- Write at least one paragraph explaining oxidation, and at least one paragraph explaining reduction.
- Use the bullet points from the class discussion to write your paragraphs.
- Use complete sentences and put them in a logical order to write your summaries.
- Explain why putting zinc into magnesium sulfate would NOT produce a reaction.

$Zn + MgSO_4$ $\rightarrow$ no reaction

2) Use the activity series of metals list to predict whether or not the following reactions will occur (Answer "Yes" or "No" for each reaction).

$Al + AgNO_3$	Yes
$Cu + FeSO_4$	No
$Zn + PbSO_4$	Yes

3) Explain (justify) your answers to question #1:

$Al + AgNO_3$
Yes, reaction will occur. Aluminum is more reactive than silver therefore it will oxidize as it causes the reduction of silver, the aluminum will replace the silver since silver is lower on the activity series. A more reactive metal can replace a less reactive metal in a compound.

$Cu + FeSO_4$
No, a reaction will not occur. Copper is less reactive than iron therefore it will not replace it, copper cannot cause the reduction of iron. A less reactive metal cannot replace a more reactive metal in a compound.

$Zn + PbSO_4$
Yes, a reaction will occur. Zinc is more reactive than silver therefore it will oxidize as it causes the reduction of lead, the zinc will replace the lead since it is lower on the activity series. A more reactive metal can replace a less reactive metal in a compound.

4) List 3 indications that a chemical reaction is occurring:

- Bubbles (gas) produced
- Temperature change
- Precipitate forms (new solid appears)
- Solid disintegrates (is "eaten" away as opposed to dissolved)
- Color change
- Odor produced
- Light emitted

5) Explain how a more reactive metal like zinc can be used to "protect" a less reactive metal like iron. Give a practical example of this.

Since zinc is more reactive than iron, zinc will be "attacked" and oxidized first. Once all of the zinc is oxidized, then the iron will be oxidized. Zinc can be used to coat iron or used as a sacrificial. Galvanized metal is an example of using a protective coat. Iron or steel is coated with zinc. The zinc oxidizes and forms a tough ceramic coating that protects the metal underneath from being oxidized as long as it isn't scratched through to expose the metal beneath. Iron does not make a protective coating when it oxidizes, the oxidation flakes off and exposes new metal to be oxidized. A water heater has a rod of magnesium, aluminum, or zinc inserted into the tank. This rod acts as a sacrificial to protect the steel case forming the outside of the tank. The rod made out of a more reactive metal is preferentially oxidized thus protecting the steel tank.

6) Zinc metal could be used in the process to claim (reduce) many metals from their ores. List three metals that zinc can replace (reduce):

- Iron
- Tin
- Lead
- Copper
- Silver

7) Do metals prefer to be oxidized or reduced? Defend your answer.

Metals prefer to be oxidized. Metals obtain a stable number of electrons in their outermost shell (achieve the octet) when they are oxidized. Metals have an unstable number of electrons (1, 2, or 3 valence electrons) when they are reduced and in their elemental form.

8) Are most metals found in the ground in a pure state (as a metallic element) or as compounds? Explain why.

Most metals are found in the ground as a compound – in ore. Metals in compound form are oxidized which is their preferred state. Therefore most metals have to be "won" from their ore in a reduction process. Reduced metals are pure elements and only a few are found this way in nature.

Metals found in the elemental state in nature are referred to as "native metals" and can include gold, silver, and copper. However, most copper is found in ore.

9) Predict whether magnesium is more likely to be reclaimed (reduced) from its ore by chemical reduction or electrolytic reduction. Justify your answer.

Magnesium is more likely to be reclaimed from its ore by an electrolytic reduction process. Metals which are less reactive than carbon can be cheaply reduced from their ore by chemical reduction using coke (a form of carbon). Metals that are more reactive than carbon are reduced from their ore by using electricity on molten or dissolved ore. Since magnesium is more reactive than carbon (as shown by the activity series), it is reclaimed from ore by electrolytic reduction.

CORROSION EXPERIMENT #3

Fruit "Juice" Objective

INTRODUCTION

Batteries operate on the principles of electrochemistry and illustrate a useful outcome of corrosion. They are based on the concepts of "ACME" (the components of a corrosion cell) and the activity series of metals (dissimilar metals). Batteries provide their voltage and current by the oxidation of the more easily corrodible (most reactive) of two metals in a couple.

In this experiment, students will use a voltmeter to determine the voltage produced using different pairs of metal strips embedded in an orange. This will simulate a simple battery. The voltage reading will depend on which metals are used. The more dissimilar the metals in terms of electric potential, the higher the voltage reading.

Students will continue to develop an understanding of "ACME" and the activity series of metals, which were introduced in cKit™ Labs #1 and #2. The more reactive metal will act as the anode while the less reactive metal serves as the cathode. The acidic juice in the fruit is the electrolyte. The results of this experiment should mirror the results of cKit™ Lab #2 – Mr. Copper and Miss Sulfate. The students will be able to create an activity series based on the voltage readings from the experiment.

PRACTICAL APPLICATIONS

In today's society, people are voracious users of batteries of all kinds. These batteries all deliver their power through the process called corrosion, also known as oxidation-reduction. In this case, the corrosion process is used to man's advantage.

Corrosion that occurs when two dissimilar metals are brought into electrical contact under water (or a corrosive electrolyte) is known as **galvanic corrosion**. When a galvanic couple forms, one of the metals in the couple becomes the anode and corrodes faster than it would all by itself, while the other becomes the cathode and corrodes slower than it would alone.

OBJECTIVE

The students will investigate the reactivity of different metals by measuring voltages produced using citrus fruit and metal strips, and use "ACME" to explain that a battery is a corrosion cell.

NEXT GENERATION SCIENCE STANDARDS (NGSS) CORRELATIONS

- Performance Expectations: HS-PS1-1, HS-PS1-2 and HS-PS3-3
- Disciplinary Core Ideas: PS1.A, PS3.A and PS3.B
- Crosscutting Concepts: Patterns, Energy and Matter, Structure and Function,
- Science and Engineering Practices: Carrying Out Investigations, Analyzing and Interpreting Data, Developing and Using Models, Using Mathematics and Computational Thinking

MATERIALS AND EQUIPMENT LIST

MATERIALS LIST:

- Oranges (or other Citrus Fruits)
- Metal strips, cut into 1½" or 2" pieces
  - Aluminum (Al)
  - Copper (Cu)
  - Lead (Pb)
  - Tin (Sn)
  - Zinc (Zn)
  - Magnesium (Mg)
- Iron nails – 2½" bright nails, smooth
- Extra magnesium strips
- Extra copper strips
- Alligator cords – *insulated electrical wires with alligator clips on each end*
- Voltmeter, with leads
- LED light (Optional)

PER STUDENT GROUP:

- **4** x Oranges (or other Citrus Fruits)
- **1** x 1½"-2" Aluminum (Al) metal strip
- **1** x 1½"-2" Lead (Pb) metal strip
- **1** x 1½"-2" Tin (Sn) metal strip
- **1** x 1½"-2" Zinc (Zn) metal strip
- **1** x 1½"-2" Magnesium (Mg) strips
- **1** x Iron (Fe) 2½" bright nails
- **3** x Extra magnesium strips
- **5** x Alligator cords
- **1** x Voltmeter, with leads
- **1** x LED light (Optional)

TEACHER PREPARATION

1) Teacher Demonstration Materials:

If the circuit extension is done only as a teacher demonstration, please note that only one piece of fruit and 2 alligator cords are needed per team.

2) Metal Sources:

The metal strip packed provided in the cKit™ includes Copper (Cu) Aluminum (Al), Lead (Pb), Magnesium (Mg), Tin (Sn), and Zinc (Zn). Bright nails (1½" to 2½", smooth finish) will represent the iron pieces and can be purchased at any hardware store.

3) Distinguishing Metal Types:

*In order to distinguish the metal strips, the dimension of the Zinc strip is 5" x ½"; and Magnesium is 6" x ⅛". The Copper, Aluminum, Lead and Tin strips are all 6" x ½". Copper is most easily differentiated, due to its color. Of the remaining three (Aluminum, Lead and Tin), Lead is the darkest color and heaviest strip; and the Tin strip is approximately three times as thick and heavier than the Aluminum strip.

4) Voltmeter Instructions:

Instructions dealing with voltmeter (VOM) operation and obtaining proper readings are addressed on the following website: **bcae1.com/vomillia.htm**. Note that Electrons flow (-) to (+). Attach black or negative to the MOST reactive metal to get a (+) reading on the voltmeter. Use 200mV or 2 Volt scale. Always use the scale that gives the highest meter deflection without going beyond the limits of the meter.

SAFETY NOTE

Safety goggles should be worn. Hands should be washed at the end of the experiment. Caution students not to place metal strips in their mouths. Caution students to handle equipment (voltmeters) with care to prevent breakage. It is the sole responsibility of the teacher conducting this Experiment to ensure that disposal methods are properly implemented. Dispose of the remnants of the Experiment in accordance with school district, state and federal environmental regulations. If in doubt, refer to the chemical disposal procedures as found in the most recent edition of the Flinn Scientific Catalog/ Reference Manual or contact **flinnsci.com** for specific instructions.

IMPORTANT: *If no copper electrode is used in this experiment, hydrogen gas is given off as a by-product of the reactions taking place. Therefore, ensure adequate ventilation and no open flames nearby.*

PROCEDURE

1. Set-up the voltmeter, and insert the leads if they are not already attached. Choose the proper setting – 200mV or 2 Volt scale. (Check with the teacher if there are any questions or you are unsure about the settings.)
2. Attach an alligator cord to each lead of the voltmeter.
3. Roll the fruit on the tabletop to break up some of the tissue inside and create more juice.
4. Insert a copper strip in the center of the piece of fruit to a depth of about 1.25 cm to 2.5 cm. The copper strip acts as the cathode.
5. Insert a different type of metal strip (the anode) about 2.5 cm away from the copper strip. Be careful that the two strips of metal do not touch. The metal strips act as the electrodes for the battery.
6. Connect a lead to each metal strip (via the alligator cords). If a negative reading appears, switch the leads/cords.
7. Record the type of metals used and the voltage output.
8. Repeat steps #5 through #7 using different combinations of metals. Record all data.
9. Rank the metals from most to least reactive based on the voltage readings. The larger the voltage reading, the more dissimilar the metals in terms of reactivity.

EXTENSION 1: BUILDING A SERIES CIRCUIT

1. Using copper and magnesium strips, hook up two oranges in a series circuit. Record the voltage output.
2. Predict the voltage output if three oranges are hooked up in a series circuit using copper and magnesium. Make the series circuit and record the actual voltage output.
3. Repeat using four oranges.

EXTENSION 2: LIGHTING AN LED (OPTIONAL)

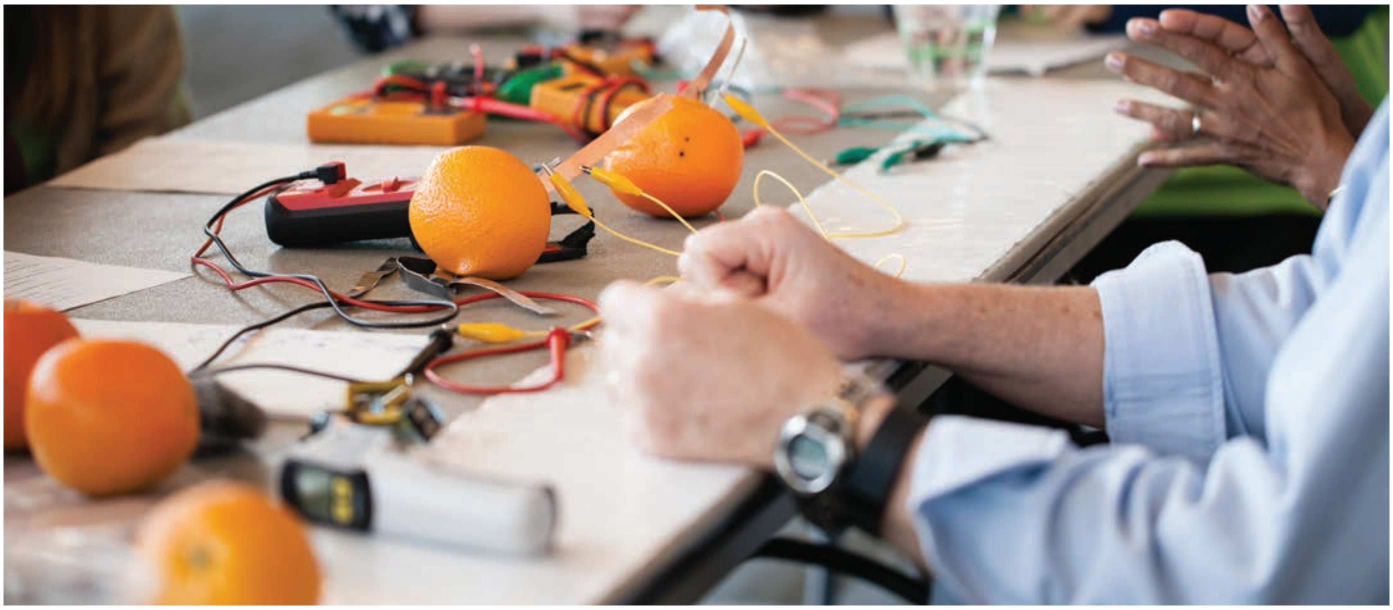
Using what they have learned, challenge students to design a fruit battery that will light an LED!

STUDENT OBSERVATIONS

The sample data table shows voltage readings using a copper strip as the cathode, and a variety of other metals as the anode (these are not absolute values). The readings will vary in a middle school or high school lab situation. Different combinations of metals may be used.

SAMPLE DATA TABLE

Metal #1	Metal #2	Fruit	Voltage
Copper	Magnesium	Orange	1.800
Copper	Aluminum	Orange	.920
Copper	Zinc	Orange	.820
Copper	Iron	Orange	.550
Copper	Tin	Orange	.400
Copper	Lead	Orange	.280
Copper	Glass stirring rod	Orange	.000



TEACHER NOTES, BACKGROUND AND EXTENSIONS

Students can investigate voltages produced by various pairs of metals and a single orange. Each group can then use multiple oranges and metals strips to create a series circuit. This can also be done as an extension demonstration after the lab. Ultimately, a series circuit could be built to light an LED.

Juicy, soft, ripe fruit works the best. It helps to roll the fruit firmly on the lab tabletop to break up some of the tissue inside and release more juice. Any type of fruit or vegetable can work. Acidic fruit works especially well. The more acidic the fruit and the more dissimilar the metals, the stronger the voltage that will be produced. Cleaning the metal samples with steel wool can also improve results. Using alligator cords with clips allows more consistent contact than using just the voltmeter leads. Make sure that the metal strips do not touch each other inside the fruit. Try to keep the distance between the strips consistent with each trial.

This experiment works well on its own but is especially effective in combination with cKit™ Labs #1 and #2. Building a corrosion cell that acts as a battery will reinforce the concept of "ACME":

- More reactive metal = anode
- Less reactive metal = cathode
- Alligator cords = metallic pathway
- Fruit juice = electrolyte

In cKit™ Lab #2, students develop an activity series based on observations of signs of a chemical reaction. Energy released by chemical reactions can be harnessed as electrical energy. In cKit™ Lab #3, the students develop an activity series based on voltage readings. The energy produced from a corrosion cell or battery (series of corrosion cells) can be used to generate light (flashlight), sound (radio) or to do mechanical work.

TEACHER EXTENSION: BUILDING A CIRCUIT

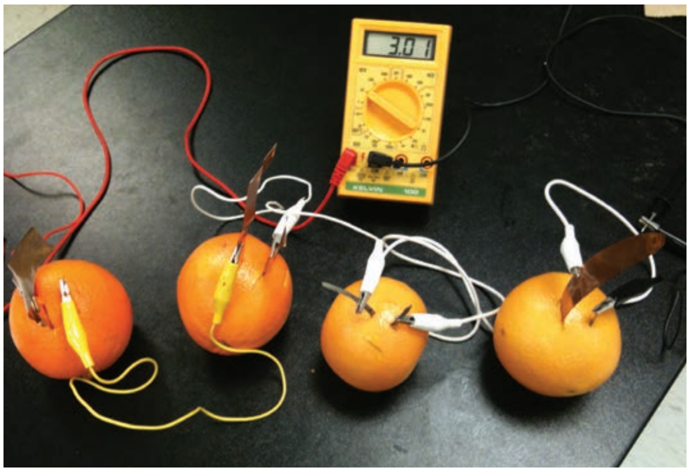
It is easy to demonstrate a series circuit using multiple pieces of fruit and magnesium and copper strips. The metal strips must be connected with the alligator cords in the proper sequence. The black lead of the voltmeter attaches to the most reactive metal – in this case magnesium. Use alligator cords to connect the copper strip in the orange to the magnesium strip in the next orange. The final copper strip will be attached to the red lead of the voltmeter. Hooking up two oranges in a series in this manner will give approximately double the voltage of the single orange. Using the sample data table this means that a voltage of approximately 3.6V could be expected. It would be approximately triple using 3 oranges, etc.

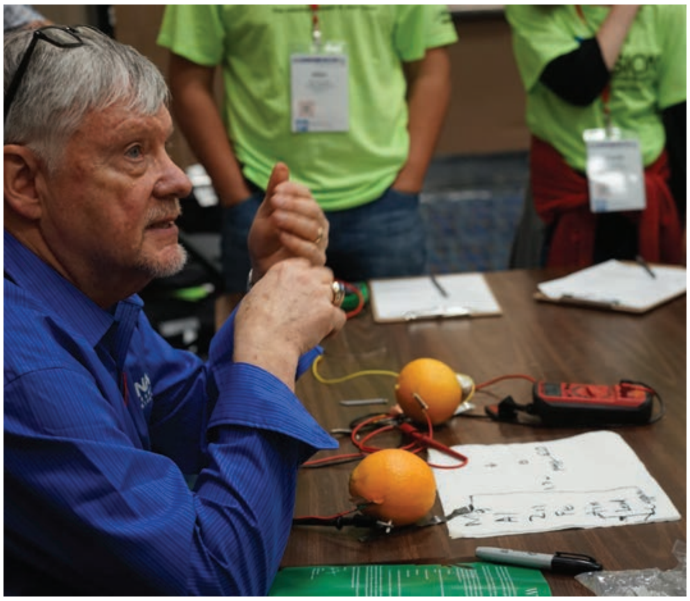
As an extension, once a series circuit has been made using 4 oranges, switch the magnesium strip from one orange with a copper strip of a neighboring orange. There is no electric potential difference between 2 strips of the same metal, so there will be no voltage increase in the orange with 2 magnesium strips nor the orange with 2 copper strips. The two oranges with a magnesium strip and a copper strip will produce a voltage. Thus, even though four oranges are connected, only two will produce voltage and thereading on the voltmeter will be approximately double that of a single orange (cell).

TEACHER EXTENSION: LIGHTING AN LED

An optional extension to the experiment is using a fruit battery to light an LED. In looking for LEDs to use, be aware that both voltage and amperage requirements will need to be met in order for the LED to light up. Select an LED with the lowest activation amperage and voltage possible. Also note that polarity issues are involved, LEDs conduct current in one direction. The LED will light up only if it is connected to the circuit correctly.

Students should choose the most reactive metal combination (copper and magnesium) from the initial procedure to make the individual fruit cells. Add individual fruit cells wired in series until enough voltage is produced to light the LED.





REAL WORLD APPLICATIONS

The use of dissimilar metals plays a part in many corrosion stories that had negative outcomes. One example is a new U.S. Navy program called the Littoral Combat Ship (LCS). Due to design flaws, major corrosion occurred after one year of operation. Concentrated corrosion happened where steel housings came in contact with the aluminum structure. More information can be found at NACE International's "Corrosion Failure: LCS 2 USS Independence Naval Ship Engine Corrosion" in the Other Resources section.

The Statue of Liberty was also a victim of corrosion caused by the use of dissimilar metals and a failure to keep a barrier between them. The internal supports were made of iron and the outer skin was made of copper. The original insulating barrier between the metals actually wicked moisture and deteriorated instead of protecting the metal. The interior of the statue was coated with many layers of paint over the years which also trapped water between the copper and the iron. Over time, the internal iron skeleton corroded away necessitating a massive restoration of the statue. This was done in the early 1980s and was completed in time for the centennial celebration in 1986.

History connections can be made with a discussion of the 2,200 year old clay Baghdad battery and Volta's invention of 1800. See the link to Battery University in the Other Resources section for an excellent article on the history of batteries. Primary batteries (non-rechargeable) and secondary batteries (rechargeable) provide their voltage and current by the corrosion of the more easily corrodible of the two metals in a couple. Students will be able to relate the content of the article to the experiment they performed. As an extension, students can construct other types of batteries such as a Volta Pile or a wet cell using various electrolyte solutions rather than a piece of fruit.

OTHER RESOURCES

- **Website: Battery University (www.batteryuniversity.com)**
  Information about different types of batteries with a great article about the history of batteries.
  Article: http://batteryuniversity.com/learn/article/when_was_the_battery_invented
- **Article: Corrosion Failure: LCS 2 USS Independence Naval Ship Engine Corrosion**
  https://www.nace.org/CORROSION-FAILURE-LCS-2-USS-Independence-Naval-Ship-Engine-Corrosion.aspx
  Article about a major corrosion problem caused by the use of dissimilar metals.
- **Book: *Rust: The Longest War* by Jonathan Waldman, Chapter 1: "A High-Maintenance Lady"**
  © 2015 ISBN: 978-1-4516-9159-7
  Excellent narrative of the corrosion problems and restoration of The Statue of Liberty in the 1980s.
  The entire book is an excellent resource about corrosion – entertaining and informative for everyone.
- **Website: Explain That Stuff!**
  http://www.explainthatstuff.com/batteries.html
  Basics of a how a battery works.
- **Video: Mocomi Kids – How do Batteries Work?**
  https://www.youtube.com/watch?v=KkRwuM4S8BQ
  Simple but good explanation of how batteries work.
- **Video: TED-Ed: Adam Jacobson – How Batteries Work**
  https://www.youtube.com/watch?v=9OVtk6G2TnQ
  Covers the basics of how a battery works and the different types of batteries (approximately 4 minutes). Companion lesson to go along with the video: http://ed.ted.com/lessons/why-batteries-die-adam-jacobson#discussion

Scan above QR code to quickly access external links!

SOURCE

This experiment was inspired by: Andy Nydam's 2004-2005 Materials Science Classes; Olympia, Washington.

CORROSION EXPERIMENT #3: STUDENT ASSESSMENT QUESTIONS

1) Which pair of metals produced the highest voltage reading?

Magnesium and copper

2) Which pair of metals produced the lowest voltage reading?

The answer can vary but should be two metals next to each other on the activity series. The answer according to the sample data given would be **aluminum** and **zinc**.

3) Develop an activity series of metals based on your voltage readings.

Most to least reactive:

- Magnesium
- Aluminum
- Zinc
- Iron
- Tin
- Lead
- Copper

4) How does this activity series compare to the activity series developed in Experiment #2?

The two series should be the same.

5) What would happen if both metal strips were the same (for example, two copper strips)?

The voltage reading would be 0.0 since no potential difference is present.

6) Do you think other fruits or vegetables could produce voltage? Why or why not?

Yes, as long as they are acidic or moist with ions present, they will be able to serve as the electrolyte.

7) What factors other than the types of metal strips used may affect the voltage readings?

- Temperature of the fruit
- Distance between the metal strips
- Oxidation on the surface of the metals
- Type of fruit

8) Identify the components of the experiment that relate to each part of the acronym "ACME".

- A (anode) = more reactive metal
- C (cathode) = less reactive metal
- M (metallic pathway) = alligator cords
- E (electrolyte) = fruit (juice of the fruit)

CORROSION EXPERIMENT #4

Poly Coat

INTRODUCTION

Aluminum is a very reactive metal and easily oxidizes. However, aluminum oxide forms a tough coat on the surface of the aluminum which protects the metal underneath from further corrosion. Certain environments, such as low pH or high pH, disrupts the protective oxide coating and allows extensive corrosion of the underlying metal. In certain applications, aluminum must be protected from its environment via coatings.

A common use of aluminum that students will be familiar with is beverage cans. Soda is acidic and can attack the aluminum can; therefore, it has a polymer liner on the inside of the can to protect it from corrosion.

In this experiment students will investigate the effects of pH on the corrosion of aluminum. Paint is removed from the surface of the can and it is cut into strips. The strips of aluminum are placed into solutions with different pH levels. Some of the solutions will oxidize the aluminum, leaving only the polymer liner. This experiment is an extension of cKit™ Lab #2 in which students develop an activity series of metals and learn that aluminum is a reactive metal.

OBJECTIVE

The students will investigate the effects of common solutions and pH on the oxidation of aluminum, and the importance of coatings in protecting consumer products from corrosion.

NEXT GENERATION SCIENCE STANDARDS (NGSS) CORRELATIONS

- Performance Expectations: HS-PS1-2
- Disciplinary Core Ideas: PS1.B
- Crosscutting Concepts: Stability and Change, Cause and Effect
- Science and Engineering Practices: Carrying Out Investigations, Constructing Explanations and Designing Solutions

MATERIALS AND EQUIPMENT LIST

MATERIALS LIST:

- Aluminum soda can
- Scotch Brite pad or Steel Wool
- Scissors
- Small containers, such as test tubes, preforms, small beakers or plastic cups
- Solutions
  - Water
  - Soda (Cola)
  - 2M or 3M Hydrochloric Acid (HCl)
  - Vinegar
  - 0.2M Copper Sulfate Solution (with a pinch of salt added)
  - 2M or 3M Sodium Hydroxide Solution (NaOH)
- ***Optional:*** Other solutions such as soap, hydrogen peroxide, ammonia, lemon juice, etc.

PER STUDENT GROUP:

- **1** x Aluminum soda can
- **1** x Scotch Brite pad or Steel Wool
- **1** x Scissors
- **6** x Small containers
- **2-10 mL** x Water
- **2-10 mL** x Soda (Cola)
- **2-10 mL** x 2M or 3M Hydrochloric Acid (HCl)
- **2-10 mL** x Vinegar
- **2-10 mL** x 0.2M Copper Sulfate Solution (with a pinch of salt added)
- **2-10 mL** x 2M or 3M Sodium Hydroxide Solution (NaOH)

SAFETY NOTE

A lab apron and goggles should always be worn when working with solutions. Caution should be emphasized when handling cut metal strips so as not to injure hands. It is the sole responsibility of the teacher conducting this Experiment to ensure that disposal methods are properly implemented. Dispose of the remnants of the Experiment in accordance with school district, state and federal environmental regulations. If in doubt, refer to the chemical disposal procedures as found in the most recent edition of the Flinn Scientific Catalog/Reference Manual or contact **flinnsci.com** for specific instructions.

PROCEDURE

1. Remove paint from an aluminum soda can with a Scotch Brite pad or steel wool. It is easier to sand an unopened can.
2. Empty the contents of the can into a cup and rinse the can with water.
3. Use scissors to cut the can into strips narrow enough to fit into test tubes or cups.
4. Pour solutions into labeled test tubes or cups to a depth of approximately 2.5 cm.
5. Add an aluminum strip to each test tube or cup.
6. Make and record initial observations.
7. After 10 to 15 minutes, make and record final observations.
8. ***Optional:*** Let sit overnight and make observations after 24 hours.

STUDENT OBSERVATIONS

The HCl and NaOH will react vigorously with the aluminum – gas bubbles will be evident and the test tube will feel very warm to the touch. After approximately 15 minutes, no solid aluminum will be left and only the polymer liner that was on the inside of the can will be found. Note, the polymer strip can be hard to see at times. Use a glass stirring rod to remove it from the test tube to examine.

The aluminum in the vinegar and soda might show slight signs of degradation after 15 minutes. If allowed to sit for 24 to 48 hours, signs corrosion will be more obvious. The aluminum strip in the copper sulfate solution (with salt) will show obvious signs of reaction. Copper ions will reduce and appear as powdery copper atoms while the aluminum atoms oxidize and become ions in solution, leaving the polymer liner behind. No detectable reaction is observed in the test tube with water.

TEACHER NOTES, BACKGROUND AND EXTENSIONS

A modification to the experiment is to have students use strips of aluminum that are sanded on one end, and still have paint on the other end. Submerge the strips so that both the painted and non-painted sections of the aluminum are in the solution. Students can carefully remove the strips from the solutions using a glass stirring rod and examine them with jeweler's loupes or magnifiers.

Acids and bases below and above a certain pH level will show a reaction. Other solutions will not affect the aluminum until enough time has lapsed, though adding salt will speed up the reaction rate of some of the solutions such as copper sulfate. The chloride ions in salt will interfere with the protective oxide layer and allow the solution to react with the aluminum.

Aluminum beverage cans have a polymer coating/liner on the inside to protect the aluminum from the contents of the can. Soda contains both phosphoric acid and carbonic acid, which will attack the aluminum. Lemon-lime sodas also contain citric acid. By removing the paint from the outside, the aluminum is exposed and can be tested for a reaction with various solutions. The inside of the can does not have paint, therefore it is coated with a polymer to prevent a reaction on the inside of the can. This experiment allows students to discover the need for the polymer lining and understand that paint isn't just "for looks".

Aluminum is amphoteric – it reacts with both acids and bases. Figure 2 shows how you might expect corrosion rates of aluminum to vary with pH. It also compares this behavior with the behavior of iron (carbon steel), which has higher corrosion rates than aluminum in neutral environments, but lower corrosion rates in most alkali (basic or high-pH) environments.

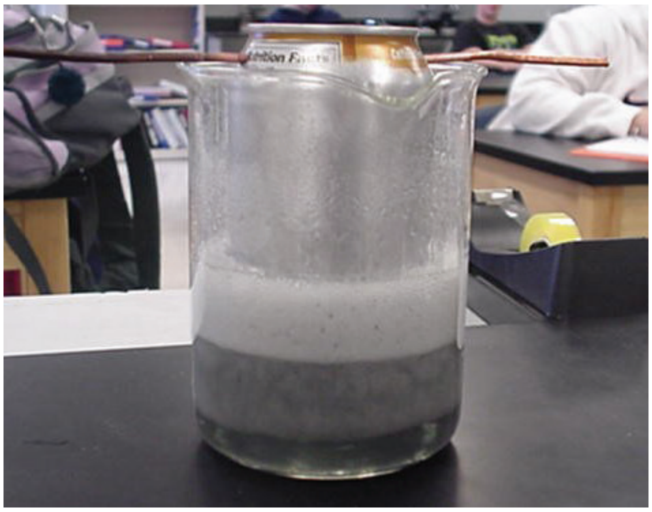


FIGURE 1

TEACHER EXTENSION: EXPOSE THE LINER

It is possible to expose the liner of the whole can instead of just strips.

1) Use a Scotch Brite pad or steel wool to remove the paint from the sides of an unopened aluminum can.

2) Pour out the soda, rinse the can and fill it with cold tap water. IMPORTANT: the tab of the can must be opened before proceeding. Do NOT leave the can sealed.

3) Insert a pencil or glass rod through the tab and place the can in a glass beaker.

4) Gradually add one of the following solutions to the beaker:
   - Copper chloride solution
   - Copper sulfate solution with sodium chloride
   - 3M HCl
   - Add between ½" – 1" at a time. Once the reaction stops, add another ½" - 1" of acid (or solution).

5) Once the desired amount of polymer liner is exposed, remove the can from the beaker and gently rinse the outside of the can with running water.

Safety Note: If 3M HCl is used in this extension, hydrogen gas will be produced and should be kept away from any open flames.

After the reaction is complete, the aluminum will have "disappeared" and you will be left with the polymer "sac" holding the water. If you use a copper sulfate or copper chloride solution, a large amount of reduced copper metal will collect in the bottom of the beaker.

Students often think this is "rusted" metal and will benefit from a class discussion about oxidation/reduction and the activity series. If cKit™ Lab #2 "Mr. Copper and Ms. Sulfate" has been performed, this extension can be used to reinforce the concepts taught previously. A similar lab is "The Can Ripper" from Flinn Scientific (see in Other Resources section).

FIGURE 2

REAL WORLD APPLICATIONS

This experiment demonstrates the effectiveness of **organic coatings** at protecting the most vulnerable metals against corrosion. Organic coatings of various forms (paints, lacquers, wrappings) and chemical formula are used in all sorts of applications to protect metals and alloys against a great variety of environments.

Cars, bicycles, trains, airplanes, boats and ships all have protective coatings. if the coating is breached by scratching, however, the corrosion process may start and rust, or another type of corrosion product shows its ugly face.

Another method of protecting iron or steel is galvanization. Galvanization is a zinc coating over iron or steel. Acidic foods such as tomato sauce and pineapple juice are often found in galvanized cans.

OTHER RESOURCES

Scan above QR code to quickly access external links!

- **Website: Flinn Scientific – The Can Ripper**
  https://www.flinnsci.com/can-ripper/dc91465/
  A lab write-up available from Flinn Scientific. The lab involves scoring the inside of a soda can to cut through the polymer liner and expose aluminum to a reactive solution. A free PDF of the lab can be downloaded from this site.

- **Website: Corrosion Technology Laboratory – NASA: Kennedy Space Center**
  https://corrosion.ksc.nasa.gov/corr_control_procoat.htm
  Corrosion Fundamental – Protective Coatings.

- **Book: *Rust: The Longest War* by Jonathan Waldman, Chapter 4: "Coating the Can"**
  © 2015 ISBN: 978-1-4516-9159-7
  Excellent narrative of the history of beverage and food cans, and the coatings they require.
  The entire book is an excellent resource about corrosion – entertaining and informative for everyone.

- **Article: The Secret Life of the Aluminum Can, A Feat of Engineering**
  https://www.wired.com/2015/03/secret-life-aluminum-can-true-modern-marvel/
  Excerpt from Rust: The Longest War – Chapter 4 – on Wired.com.

- **Video: YouTube – How It's Made: Aluminum Cans**
  https://www.youtube.com/watch?v=V7Y0zAzoggY
  The video shows the steps in making an aluminum beverage can including the polymer coating and its purposes.
  Video length is 4:45.

SOURCE

This experiment was inspired by Andy Nydam, Materials Science teacher; Olympia, Washington.

CORROSION EXPERIMENT #4: STUDENT ASSESSMENT QUESTIONS

1) Why aren't all aluminum structures or products (such as window frames) coated with paint or a polymer liner?

   Aluminum oxidizes rapidly when exposed to air, and the oxidation layer bonds very tightly to the surface of the aluminum. This provides a protective coating of aluminum oxide that protects the aluminum underneath from corrosion.

2) Why do you think some aluminum structures or products (such as soda cans) need to be coated with paint or a polymer liner?

   The structures or products are exposed to an environment (such as a fairly strong acid or base) that can dissolve or eat through the aluminum oxide on the surface, and thus attack or corrode the aluminum underneath.

3) Write a chemical equation for a reaction that happened in one of the test tubes or cups.

   Sample answers include:

$$2Al + 3CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$$

$$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$$

4) Which solutions seemed to corrode the aluminum the quickest?

   Acids and bases (solutions with a high pH or a low pH) will react the fastest.

5) Hypothesize what might happen if the polymer liner on the inside of a pop can was accidentally scratched before it was filled with cola.

   If the liner is scratched all the way through to the aluminum, the phosphoric and carbonic acids in the cola will react with (oxidize) the aluminum at the scratch. The corrosion could go all the way through the aluminum to the paint on the outside. With just the paint holding the can together at that point, the pressure of the fluid inside the can could cause the cola to burst through the paint.

Additional Notes

CORROSION EXPERIMENT #5
Quick Silver

INTRODUCTION

Silver reacts with sulfur compounds in the air such as hydrogen sulfide, which results in silver sulfide, better known as tarnish. Tarnish starts as a slightly dark film on the surface of the silver and can eventually turn black if allowed to continue long enough. Silver also reacts with oxygen, but the main culprit in the corrosion of silver is sulfur.

Tarnish can be mechanically removed from the surface of the silver piece, using abrasive polishes or creams that will remove some of the silver atoms along with the tarnish. Some tarnish removers work by dissolving the silver sulfide in a solution, but this also removes some of the silver atoms. The third option is to reverse the oxidation process by chemically reducing the tarnish to pure silver using a more reactive metal. In this experiment, aluminum, hot water, baking soda and salt will be used to reverse the reaction that created the tarnish.

This experiment may be performed by teams of students or as a teacher demonstration. It shows a practical application of cKit™ Labs #2 and #4 (the activity series of metals and how pH affects the corrosion of aluminum).

OBJECTIVE
Students will utilize the activity series of metals and everyday materials to reverse the corrosion process that tarnishes silver.

NEXT GENERATION SCIENCE STANDARDS (NGSS) CORRELATIONS
- Performance Expectations: HS-PS1-2
- Disciplinary Core Ideas: PS1.B
- Crosscutting Concepts: Patterns, Stability and Change, Cause and Effect
- Science and Engineering Practices: Carrying Out Investigations, Constructing Explanations and Designing Solutions

MATERIALS AND EQUIPMENT LIST

MATERIALS LIST:

- Aluminum foil pan, disposable
- Aluminum foil pieces – approximately 6" square, crumpled
- Electric hot water kettle
- Table salt
- Baking soda
- Tarnished silver pieces
- Plastic spoon, for stirring
- Pot holders or oven mitts
- Paper towels

PER STUDENT GROUP:

- **1** x Aluminum foil pan, disposable
- **6 pieces** x Aluminum foil – approximately 6" square, crumpled
- **1** x Electric hot water kettle
- **1-3 tsp** x Table Salt
- **1-3 tsp** x Baking soda
- **1 piece** x Tarnished Silver Item
- **1** x Plastic spoon, for stirring
- **1 pair** x Pot holders or oven mitts
- Paper towels

TEACHER PREPARATION

1) Tarnished Silver:

This experiment must be performed on silver, not stainless steel or plated objects. Students can bring in pieces of tarnished silver – jewelry, silverware, candlesticks, trays, etc. Pieces can often be found at thrift stores, such as Goodwill. Schools often have old trophies that are tarnished silver that can be cleaned using this method.

2) Boiling Water:

A microwave or hot plate can be used to boil water instead of an electric hot water kettle. All supplies for the experiment can be purchased locally.

3) Other Supplies:

All supplies for the experiment can be purchased locally. The size of the aluminum foil pan and amount of baking soda and salt depends on the size of the sterling silver pieces that are being cleaned.

4) Self-Tarnish Silver:

In the event you need to tarnish the silver, you can do so by mixing 1g of sulfur in 2 mL of water (Soluble sulfides would only take seconds. An alternate to sulfide would be using egg salad). Immerse the silver item into the solution. Remove the silver item from the mixture after two days and let dry.

SAFETY NOTE

Use pot holders or oven mitts to pour boiling water and handle hot silver. A lab apron and goggles should be worn if there is a potential for boiling water to splash. The water with baking soda and salt may be poured down the drain. The aluminum foil pieces and pan may be placed in metal recycling bins or the trash.

PROCEDURE

1. Boil water using the electric hot water kettle.
2. Place several pieces of loosely crumpled aluminum foil into the foil pan.
3. Add boiling water to the foil pan. (Increased temperature speeds up the reaction.)
4. Mix 1-3 teaspoons of salt and 1-3 teaspoon of baking soda (sodium bicarbonate) into the boiling water. Stir until dissolved.
5. Place tarnished silver on top of the aluminum foil. Add more boiling water to cover the piece of tarnished silver if needed.
6. Make observations. Take note of any odor produced.
7. After several minutes, remove the piece of silver and rub with a clean cloth or paper towel. Remove a piece of aluminum foil and note any changes.
8. Rinse the silver with tap water to remove any remaining baking soda or salt.

STUDENT OBSERVATIONS

Examples of observations made by students include:

- The silver will appear shiny and clean – the oxidized silver has been reduced to pure silver.
- The aluminum foil and pan will turn dark due to oxidation by sulfur and oxygen.
- A sulfur odor will be apparent during the lab.

TEACHER NOTES, BACKGROUND AND EXTENSIONS

This experiment reinforces one of the main concepts learned in cKit™ Lab #2 – more reactive metals can replace less reactive metals in a compound. A single replacement reaction occurs where the aluminum is oxidized while the silver is reduced:

$$3Ag_2S + 2Al \rightarrow 6Ag + Al_2S_3$$

Variations of the "recipe" to try include using just baking soda without the salt. Salt dissolved in the water acts as an electrolyte for the electrochemical reaction. The aluminum is the anode and the silver is the cathode. The silver and the aluminum must be in contact in order for the reaction to occur.

Baking soda also forms ions when dissolved in water, therefore the salt may not be necessary for the reaction to occur. The presence of salt may speed up the reaction, however. As aluminum oxidizes, the surface forms a protective coating which resists further corrosion. In other words, it passivates. The chloride ions from the dissolved salt will help keep the surface of the aluminum clean so that the reaction may continue (see more information about this in cKit™ Lab #2).

Baking soda is needed to provide a basic pH. Aluminum reacts (oxidizes) at pH's approximately below 4 and above 8.5 (see more information about this in cKit™ Lab #4).

Another variation of the "recipe" to try is using washing soda (sodium carbonate) instead of baking soda (sodium bicarbonate). The pH of a baking soda solution is around 9 and washing soda is about 10 to 11. Therefore the washing soda will be more aggressive.

Sometimes the tarnish is loosened, but has to be rubbed with a clean towel to completely remove it from the surface of the silver. If the silver is heavily tarnished, it may take several treatments to thoroughly remove it. Rubbing with a clean cloth can help. Also wrapping the silver in aluminum foil can help, as it improves contact between the metal surfaces.

The water must be very hot for the reaction to occur noticeably. If the reaction appears to slow down, add some boiling water to the pan. Also, using steel wool to rough up the surface of the aluminum foil pan can be beneficial. This would remove some of the thin oxide layer on the surface of the aluminum and would also break through any coating that may be on the pan.

Since copper is also much less reactive than aluminum, this method may be used to reverse the oxidation of copper and brass.

REAL WORLD APPLICATIONS

Sterling silver cleaning kits may be purchased via home shopping networks or online stores. These kits basically contain an aluminum metal plate and baking soda or washing soda (sodium carbonate), though that is not what they are called in advertisements.

The ingredients, however, can be purchased much more economically at a local store. Understanding basic chemistry, the activity series of metals, and the principles of corrosion can help consumers make wise choices.

The cleaning of silverware with the recipe described in this experiment can be useful at home. Though this experiment reverses corrosion that has already occurred, the sacrificial corrosion of aluminum (or another easily corroded metal) to protect a less reactive metal is even more useful. This has been practiced industrially in a process called **cathodic protection** for almost two centuries.

Many industrial appliances are protected by using sacrificial metals such as zinc, magnesium and aluminum. Ships and boats are usually equipped with anodes made of these metals for protection in the areas most prone to corrosion. Most domestic water heaters contain an internal magnesium or aluminum anode to provide added durability. More details about cathodic protection can be found in cKit™ Lab #2.

OTHER RESOURCES

A web search for silver cleaning products will reveal several websites and YouTube videos that promote kits for cleaning silver. Many of these contain an aluminum plate and a packet of baking soda or washing soda though they are often not identified as such. The sites often use the terms "metallic plate" or "cleaning plate" for the aluminum and "activator" for the baking soda or washing soda. For a lesson in consumer science, have students read a promotional website or view a video and have them critique it in terms of scientific accuracy and evaluate the costs of the materials provided in the kit.

Scan above QR code to quickly access external links!

Some suggested sites include:

- **Website: Hammacher Schlemmer – The Museum Precious Metals Cleaning Plate**
  http://bit.ly/2ra2WzP

- **Website: International Jewelry Marketing – The Speedy Plate**
  http://www.internationaljewelrymarketing.com/speedyplate.htm

- **Website: Metal Brite Products – Metal Brite Super Power Plate**
  http://www.metalbrite.net/products

- **Video: Metal Brite Demo – "Chemically Free Cleaning Process"**
  https://www.youtube.com/watch?v=5B92SpwiUrQ

SOURCE

This experiment was inspired by: Kehlen Sachet, Olympia High School Materials Science Class; Olympia, Washington.

CORROSION EXPERIMENT #5: STUDENT ASSESSMENT QUESTIONS

1) What happened to the tarnish on the silver?
   Tarnish is mostly silver sulfide with some silver oxide. Because aluminum is more reactive than silver, the silver is reduced and the aluminum is oxidized.

2) What caused the aluminum foil to change color?
   The formation of aluminum sulfide (possibly mixed with aluminum oxide) made the surface of the aluminum darker and less lustrous.

3) Some silver polishes are abrasive; they rub or scratch the tarnish off the silver. What is a disadvantage to using this method of tarnish removal?
   Some of the silver will be removed along with the tarnish. Over time, the silver will become thinner.

4) Why is aluminum used?
   Aluminum is more reactive than silver, readily available and inexpensive.

5) What is the purpose of the baking soda/salt mixture?
   The solution provides ions to transport the electrons and the baking soda raises the pH so the aluminum will oxidize.

6) Write the equation for the chemical reaction that is occurring.

$$3Ag_2S + 2Al \rightarrow 6Ag + Al_2S_3$$

7) What is tarnish?
   Silver sulfide (Ag_2S) which is a form of oxidized (corroded) silver.

8) Is silver the only metal that tarnishes?
   Tarnish is another name for corrosion (usually associated with silver like "rust" is the corrosion of iron). All metals will undergo corrosion to some extent – some more than others. The noble metals (such as gold and platinum) are difficult to corrode or tarnish, whereas the more reactive metals such as magnesium and aluminum corrode or tarnish very easily and readily.

9) What is the odor given off during the lab?
   Hydrogen sulfide gas (H_2S) produced as a by-product of the reaction between aluminum sulfide, water and baking soda.

10) Would this experiment work with a different liquid?
   It should work with other electrolytic solutions that also provide the proper pH. A suggestion would be to try vinegar instead of baking soda.

11) Explain the results of the lab in terms of the activity series of metals.
   Silver is less reactive than aluminum so it can be replaced in a compound by aluminum. The less reactive metal, silver, is reduced and the more reactive metal, aluminum, is oxidized. Therefore the tarnish which is silver sulfide became pure silver while the aluminum formed aluminum sulfide on its surface.

CORROSION EXPERIMENT #6
Silver Pennies

INTRODUCTION

Many metal objects and structures are coated to protect them from corrosion and to improve other surface properties. The coatings industry is vast and of extreme importance. It includes organic coatings, such as paint and metallic coatings or galvanization. Electroplating is a huge part of this and many common objects encountered daily have been through this process.

Electroplating is the deposition of a metal coating onto an object by using a negative charge and a solution that contains an electrolyte. Metal cations in the solution are attracted to a negatively charged object (cathode) and are reduced to form a thin metal coating on the object.

In this experiment, students will electroplate the surface of a penny with zinc using a weak sodium hydroxide solution, a zinc strip and an AA battery. Students will be able to correlate electroplating with the components of “ACME”: an anode (A) which oxidizes and releases metal cations into the electrolyte (E), a power source (battery) which also acts as the metallic pathway (M) and the cathode (C) where metal ions are reduced and form a metal coating on an object.

OBJECTIVE

The students will develop a basic understanding of electroplating by coating a penny with zinc using an AA battery as the driving force. The concept of stray current corrosion will also be introduced.

NEXT GENERATION SCIENCE STANDARDS (NGSS) CORRELATIONS

- Performance Expectations: HS-PS1-2
- Disciplinary Core Ideas: PS1.B
- Crosscutting Concepts: Stability and Change, Cause and Effect
- Science and Engineering Practices: Carrying Out Investigations, Constructing Explanations and Designing Solutions

MATERIALS AND EQUIPMENT LIST

MATERIALS LIST:

- Pennies, clean and shiny
- Small glass bowl
- 1M Sodium hydroxide (NaOH)
- Granular Zinc
- Zinc metal strip
- AA battery holder with alligator clips
- AA battery
- Plastic spoon with holes melted into the bowl of the spoon
- Plastic spoon
- Thick copper wire
- Distilled water in beakers or cups
- Plastic gloves
- Sandpaper or steel wool

PER STUDENT GROUP:

- **1** x Penny (clean and shiny) per student
- **1** x Small glass bowl
- **100 mL** x 1M Sodium hydroxide (NaOH)
- **1 tsp** x Granular Zinc
- **1** x Zinc metal strip
- **1** x AA battery holder with alligator clips
- **1** x AA battery
- **1** x Plastic spoon with holes melted into the bowl of the spoon
- **1** x Plastic spoon
- **1** x Thick copper wire, 4" long
- **200 mL** x Distilled water beaker or cup
- **1 pair** x Plastic gloves per student
- **1 piece** x Sandpaper or steel wool

TEACHER PREPARATION

1) Zinc-Plating Solution Set-Up:

The set-up for zinc-plating includes:

- **Small glass bowl**
- **Zinc strip** - Bend to lay flat on bottom of the bowl, run up the side, and hang over down the outside of the bowl.
- **1M NaOH (sodium hydroxide)** - Fill the dish just enough to submerge the bowl of the spoon and penny, and not have the spoon touch the granular zinc on the bottom of the glass bowl.
- **Granular zinc** - Add about 1 tsp or enough to cover the bottom of the dish.

It takes time for zinc ions to form in the solution. Prepare the electroplating station the day before the lab allowing the granular zinc and zinc strip to sit in the NaOH solution overnight.

Alternatively, plate 2 or 3 pennies before the students do the lab to get the ions into solution. The first penny may take up to a minute or two to plate. The second penny will take approximately 30 seconds. After the 3rd or 4th penny, it typically takes 10 seconds or less to plate each one.

2) Materials Sources:

The sodium hydroxide solution, zinc strip and granular zinc can be purchased from Flinn Scientific (**flinnsci.com**). Small glass bowls can be purchased at discount stores such as Walmart or found at thrift stores such as Goodwill. Beakers may also be used.

3) Making Plastic Spoon Ladles:

To make the spoon (ladle) for holding the penny during the plating process, melt holes in the bottom of a plastic spoon using a heated copper wire. Heat the spoon with a hair dryer or heat gun (paint stripper) and bend it upwards to make a ladle, so it is easier to manipulate in and out of the solution.

SAFETY NOTE

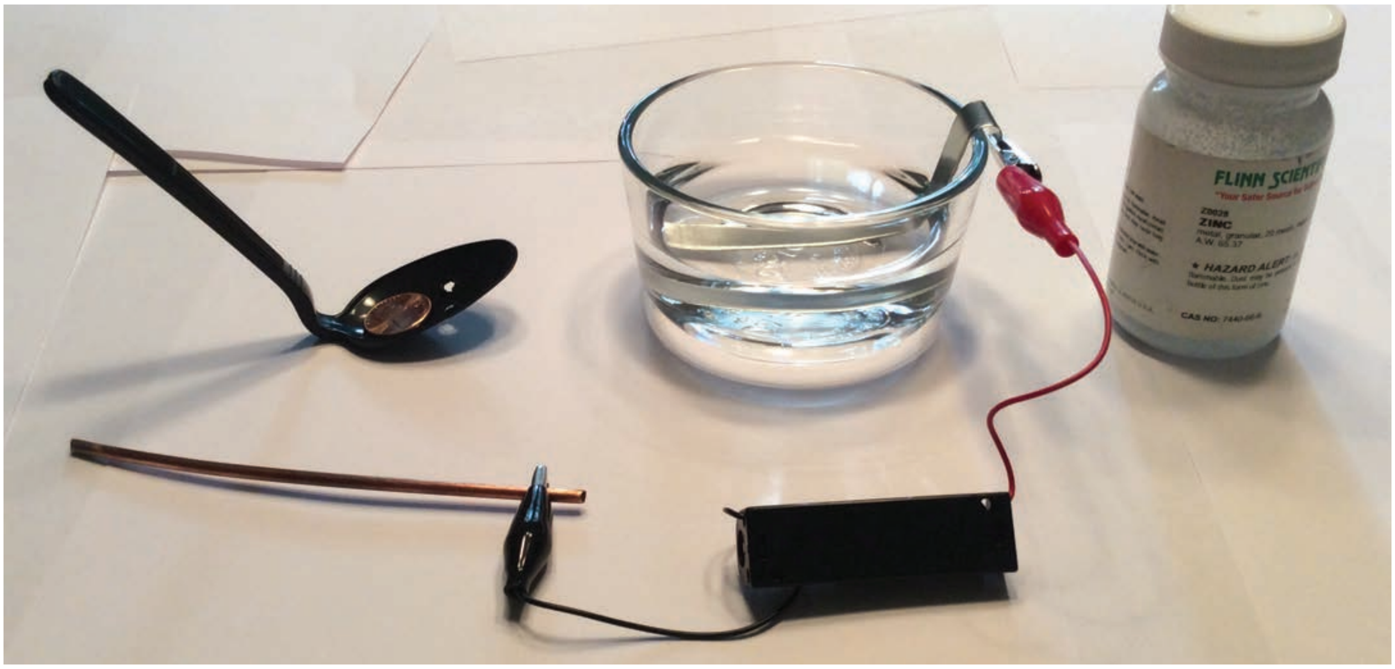
A lab apron, goggles and gloves should always be worn when working with solutions. It is the sole responsibility of the teacher conducting this Experiment to ensure that disposal methods are properly implemented. Dispose of the remnants of the Experiment in accordance with school district, state and federal environmental regulations. If in doubt, refer to the chemical disposal procedures as found in the most recent edition of the Flinn Scientific Catalog/Reference Manual or contact **flinnsci.com** for specific instructions.

PROCEDURE

1. Clean the penny.
   - Shiny new penny – Simply rub with a towel to remove oils from your skin.
   - Penny with oxidation – Use a cleaner such as Tarnex to remove all oxidation. Wash with soap and water and dry with a towel.
   - Do not touch the penny with your skin after cleaning. Hold the penny with a paper towel to place into the bowl of the spoon.
2. Go to a zinc-plating station.
   - Mix 100mL of 1M NaOH and 1 tsp of granular zinc in the glass bowl.
   - Submerge zinc metal strip beneath the mixture so it extends over the edge of the dish.
   - Connect the DC voltage source (AA battery) by placing the positive alligator clip (red wire) on the zinc strip.
   - Connect the negative alligator clip (black wire) to the copper wire.
3. Place the penny in the “bowl” of the plastic spoon (the spoon with holes in the bottom).
4. Submerge the penny in the 1M sodium hydroxide solution containing granular zinc. Do not touch the spoon to the zinc strip or the granular zinc on the bottom of the dish.
5. Touch the copper wire (attached to the negative/black alligator clip) to the penny for a few seconds. Start with about 8 seconds, and make firm contact between the copper wire and the penny.
6. Remove the copper wire from the penny AND the solution while keeping the penny submerged in the solution.
7. After all bubbling has stopped, remove the penny from the solution and check the plating.
8. Repeat steps #3 through #7 as necessary to get full coverage.
9. Flip the penny over in the bowl of the spoon to check the plating on the other side. Repeat steps #3 through #7 as necessary.
10. After the penny is completely covered with zinc, quickly rinse the penny in a cup or beaker of distilled water. Do not allow the penny to have prolonged contact with air until it has been thoroughly rinsed with water.
11. Rinse the zinc-plated penny under running water, gently rubbing it to make sure the sodium hydroxide solution is removed. Gently dry the penny. The plating is very thin and will come off if rubbed too vigorously.

STUDENT OBSERVATIONS

Gas bubbles appear on the surface of the penny when the circuit is connected. The penny turns silver colored as the zinc is deposited onto the surface of the penny.



TEACHER NOTES, BACKGROUND AND EXTENSIONS

Each student should be allowed to electroplate a penny. One electroplating station can be set up for the students to rotate through, or multiple stations can be set up with a small group of students at each station. Older and more advanced students can set up the electroplating station themselves. The instructor should set up the station for younger and inexperienced students. See "Teacher Preparation" section for detailed instructions on how to set up an electroplating station.

Holes should be melted in the bottom of a plastic spoon using a heated copper wire. This allows the sodium hydroxide solution to drain from the bowl of the spoon after plating, and also allows better contact between the bottom of the penny and the solution. The spoon is easier to manipulate in and out of the solution if you heat the handle with a hair dryer or heat gun (paint stripper) and bend it upwards to make a ladle.

Hydrogen gas is released during the plating process. Keep open flames away from the electroplating station. The hydrogen gas can get trapped between the spoon and the bottom of the penny, and this will prevent plating from occurring on the bottom surface. Use another plastic spoon to flip the penny over after the top is completely covered with zinc and resume plating until both sides are zinc-plated. The copper wire will need to be sanded occasionally to remove the buildup of zinc on the end.

This experiment works very well but can take some finessing. It works best if you submerge the penny in the solution first and then firmly touch the copper wire to the penny. Also, the circuit needs to be broken before bringing the penny out of the solution – in other words, the penny should not be connected to the battery unless it is submerged in the NaOH/zinc solution.

A quick decision needs to be made as to whether enough plating has been done. The "wet" penny needs to either go back into the NaOH solution for more plating or into a cup/beaker of distilled water quickly and not be exposed to air for any length of time with the NaOH still on it. After breaking the connection, do not remove the penny from the solution until all bubbles have dissipated.

Do not let the steel alligator clips touch the NaOH solution. The solution becomes contaminated and will not plate properly. It his happens, start anew with fresh solution. Use fairly thick gauge copper wire (approximately 16 or 18 gauge) for the electrode. Best results are obtained when using brand new uncirculated pennies. It is crucial that all oxidation be removed from the pennies prior to the zinc-plating. Corrosion is an ionically bonded ceramic material and the zinc metal will not bond to it. A product called Tarnex works great for removing oxidation. It can be purchased at local stores.

Caution the students about rubbing the pennies too hard after they are zinc-plated. Excess handling will cause the zinc-plating to rub off. The zinc layer is only a few atoms thick – it is nanoscale. Making a thicker layer of zinc by plating for a longer time will result in a gray, non-lustrous coat because the surface becomes uneven. Also have students avoid handling the pennies after cleaning them. Oil from their fingers can leave a coating on the pennies that interferes with the plating.

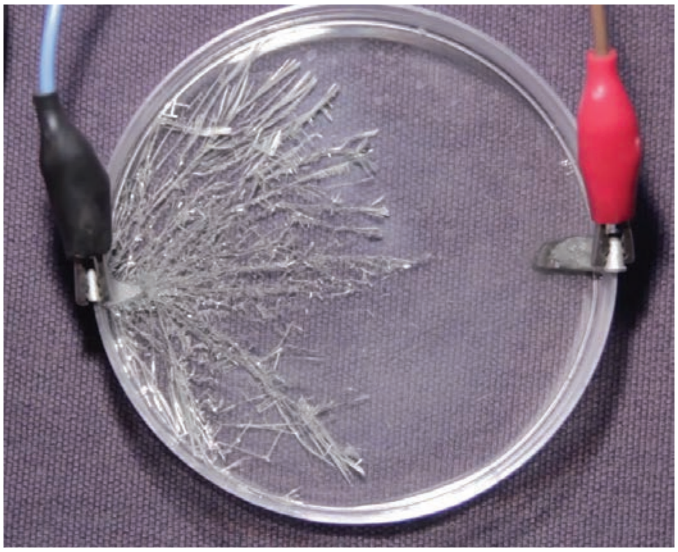
A very thin layer of zinc (several atoms thick) is plated onto the copper surface of the penny. In electroplating, positively charged particles from one material are removed and transferred through an electrolyte solution until they are deposited on a negatively charged surface of another material. At the positive electrode (zinc strip), electrons leave the material creating positive Zn ions in the solution. Electrons then travel through the power source to the negative electrode (copper penny). The positive zinc ions travel through the solution to the negative electrode, and are reduced to zinc atoms on the penny's surface.

The leads must also be connected in the proper order. If the red and black alligator clips are reversed plating will not occur. Over time, the zinc strip will appear pitted. The zinc atoms give their electrons to the battery and form zinc ions (oxidation), making the zinc strip the anode which will eventually corrode away. These ions travel through the electrolyte solution to the cathode (the copper penny) where they are reduced, creating a shiny silver surface.

TEACHER EXTENSION: STRAY CURRENT CORROSION

After all the students have had an opportunity to electroplate their penny, an interesting extension to the experiment is to demonstrate stray current corrosion. Use the existing set-up for zinc-plating and a clean penny. Once the top surface of the penny is plated with zinc, pull the copper wire slightly away from the surface of the penny but keep the penny and the tip of the wire submerged in the solution. The zinc will "disappear" below the tip of the wire. Move the copper wire around above the surface of the penny. The zinc will be removed from the area directly below the wire, but will redeposit in areas to the side. It looks like the zinc is being "vacuumed" from the surface. When the gap between the penny (cathode) and copper wire is narrow enough, the current will pass through the electrolyte making the top surface of the penny the anode and the copper wire the cathode.

Stray current corrosion can be a major problem for pipelines and other infrastructure. The stray currents can originate from direct-current distribution lines, substations, or electrical powered rail systems. Pipelines are often buried underneath powerlines. If stray current is not controlled or minimized, normal corrosion that could take a lifetime to accumulate is accelerated to just days or weeks.



TEACHER EXTENSION: THE REVERSIBLE TIN MAN

Teachers may perform an additional experiment called the Reversible Tin Man to reinforce the concepts of electrolysis, oxidation-reduction, anode, and cathode. Tin (II) chloride solution is placed in a Petri dish with steel paper clips acting as anode and cathode. A 9-V battery is the power source. At the cathode, tin crystals appear and grow beautifully toward the anode. When the leads are switched (black and red), the original cathode now becomes the anode and the tin "disappears" as it is oxidized back to ions in solution. The former anode now becomes the cathode and tin crystals start growing outward. A PDF document of the experiment, titled "Tin Man Electrolysis," is free to download from the Flinn Scientific website. The materials for the experiment are also available.

REAL WORLD APPLICATIONS

The plating of metals is an important industry that allows easily fabricated and low cost steel objects to become protected by a thin layer of another metal, thus changing its easily corroded behavior. Many metallic surfaces used in our modern world have been coated by another metal, i.e. galvanized fences, posts, nails, screws, sheet metal, etc.

Electroplating is used widely in many industries. Electroplating increases the life of metal and prevents corrosion as well as altering other surface properties.

Easily observed examples of electroplated objects include chrome wheels on a car, bathroom fixtures and gold or silver-plated jewelry.

Besides protection from corrosion, other properties affected by electroplating include lower friction, better electrical conductivity, increased hardness, enhanced appearance, increased wear resistance and heat resistance. In a process called **electroforming**, the thickness of undersized parts may be built up.

OTHER RESOURCES

- **Website: Flinn Scientific – Tin Man Electrolysis**
  https://www.flinnsci.com/tin-man-electrolysis/dc91646/
  Additional experiment demonstrating electrolysis of tin (II) chloride that reinforces additional concepts of electrolysis, oxidation-reduction, anode, and cathode.
  Video Demo: https://www.youtube.com/watch?v=coPGazAaVBE

- **Website: Sciencing – The Uses for Electroplating**
  http://sciencing.com/uses-electroplating-8100658.html
  The website contains many articles on corrosion, electroplating, metals, etc.

- **Website: The Corrosion Doctors – Stray Current in Transit Systems**
  http://www.corrosion-doctors.org/StrayCurrent/Transit-Systems.htm
  An example of how stray current can cause major corrosion problems. The Corrosion Doctors website is extensive with many pages of information related to corrosion.

- **Video: Electroplating – How It's Done**
  https://www.youtube.com/watch?v=z7f7dQF2KLA
  A great explanation of the electroplating process. The video describes the plating of chrome, copper, nickel and gold, as well as plating over both metal and plastic bases. materials.

Scan above QR code to quickly access external links!

SOURCE

This experiment was inspired by Andy Nydam and Debbie Goodwin.

CORROSION EXPERIMENT #6: STUDENT ASSESSMENT QUESTIONS

1) Using the zinc and copper electrodes, does it matter how you connect the battery to zinc plate the penny? Explain.

   Yes. The positive alligator clip (red wire) must be connected to the zinc strip, and the negative alligator clip (black wire) must be connected to the copper wire.

2) Where did the zinc come from that coated the penny?

   From the zinc strip.

3) Is the reaction occurring at the zinc strip (the anode) oxidation or reduction?

   Oxidation.

4) Is the reaction occurring at the penny (the cathode) oxidation or reduction?

   Reduction.

5) Would you be able to electroplate the penny if you used water instead of sodium hydroxide solution?

   No. An electrolyte must be present to complete corrosion cell.

6) What are the signs of a chemical reaction?

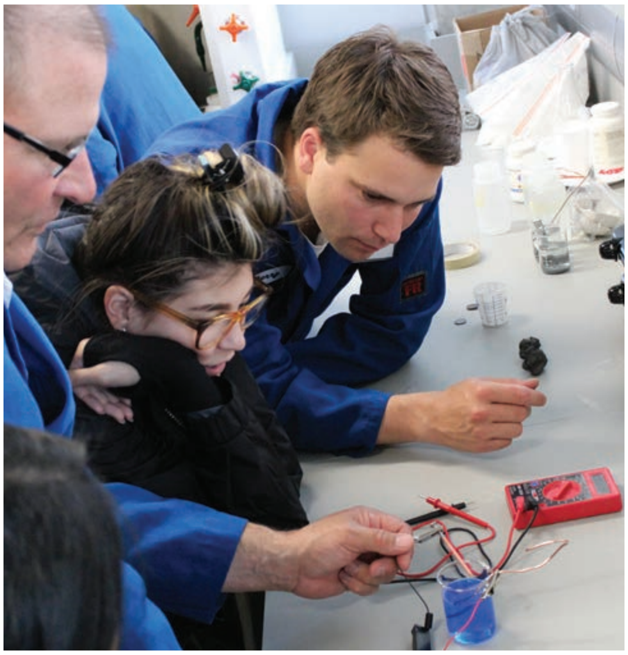
   - Gas bubbles produced
   - Temperature change
   - Color change
   - Solid precipitate formed
   - Solid disintegration
   - Odor produced

7) What would happen if this experiment ran for 3 days?

   The zinc strip will appear pitted. Zinc atoms are oxidized, and enter the solution as zinc ions to replenish the ions being reduced at the cathode (the penny).

8) List practical applications of electroplating.

   Coating base metals with gold or silver to make inexpensive jewelry that looks good. Corrosion protection. Circuitry in electronic devises. Increased surface hardness.



Additional Notes

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Special Thanks

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