Developing the Activity Series

Overview: The students will develop their own activity series of metals based on lab results. Qualitative observations will be used. Tie-ins include single-replacement reactions (SRR), oxidation-reduction, practical applications such as galvanization, and predicting reactions. Students work in teams to determine the relative reactivity of six different metals in a solution of copper sulfate. My students have already covered and have an understanding of atomic structure and bonding before performing this lab.

Goal/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.

Here is the website where all of the NACE corrosion labs can be found.

http://www.nace-foundation.org/programs/images_programs/cKit_experiments.pdf

Materials/Equipment (per team):

test tube rack with six test tubes

small samples of 6 different metals:

(1) zinc - zinc metal strips from Flinn

(2) lead - lead metal strips from Flinn

(3) aluminum - aluminum metal strips from Flinn

(4) tin - tin metal strips from Flinn

(5) magnesium - an inch of ribbon

(6) iron - flattened wire

copper sulfate solution (0.2 M works well)

steel wool for cleaning metal samples of oxidation

Basic Procedure:

1. Pour approximately 5 mL (or 1 inch) of "blue stuff" (copper sulfate solution) into each test tube.

2. Clean the 6 metal samples with steel wool to remove any oxidation.

3. Put a different metal sample in each of six test tubes one at a time so initial observations can be made.

4. Record observations in your journal for approximately 5 to 10 minutes. Look for signs that a chemical reaction is occurring:
   - gas bubbles being produced
   - temperature changes
   - changes in color
   - a solid precipitate forming
   - solid disintegrating

5. Rank the metals in order of reactivity.

6. Clean-up.

Class Discussion - see notes below.

Teacher demo: Place a strip of copper foil in a test tube with 7 mL of 0.1M silver nitrate solution. Have students write observations. Have them place silver in their activity series.

Notes:

Aluminum is a very active metal but is slow to start in this lab. Aluminum oxide bonds so tightly to the surface of the metal that it is difficult to clean off. Adding sodium chloride jump starts the reaction like a catalyst. 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 .2M copper 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, you may wish to show them two different aluminum trials, one with salt added to the copper sulfate solution and one without.

As your students perform the lab and write observations in their journals the teacher will need to do a little "troubleshooting". The students usually think the metals are "rusting". The teacher will need to give guidance to get them to realize that it is copper metal being formed - not rust.

Following is sort of a guideline as to how we lead the class discussion about the lab results.

Ask the students which metal they think is most reactive. They usually get this one right. It is magnesium. So I write it at the top of the chalkboard. Then we talk about which one is next. I often have to mediate because all the groups don't always agree on the order of ranking. I use "majority rules" as a guideline and it almost always works. Lead and tin can be difficult to distinguish because neither one reacts much. Here is the final ranking to be agreed upon:

Magnesium
Aluminum
Zinc
Iron
Tin
Lead

After you agree on this list, tell the students that they have left out a metal. This usually confuses them because they had six metal samples and there are six elements on their list. At this point you can tell them what the "blue stuff" is - copper sulfate. With encouragement they come up with the missing metal - copper. Ask them where it goes on the list. Most are usually not sure. Thus, you begin your discussion of what really happened in the test tubes: a single-replacement reaction involving oxidation-reduction.

Put a sample equation on the board such as:

$$Fe + CuSO_4 \longrightarrow$$

Ask them to fill in the products. They realize that copper was formed, so somehow iron and copper switched places.

Finish the equation using ions this time:

$$Fe + Cu^{+2}SO_4^{-2} \longrightarrow Cu + Fe^{+2}SO_4^{-2}$$

The sulfate is unchanged. Iron went from an atom to a positive ion (define oxidation at this time) and copper went from a positive ion to a neutral atom (define reduction). It may be helpful to define a single-replacement reaction at this point.

Now try to make a connection that they will understand. You use a 1 to 10 scale to rate the metals. Here is how you explain what is occurring in the oxidation-reduction reaction.

The lovely Miss Sulfate attends a school dance with Mr. Copper. Now Mr. Copper isn't exactly the finest guy. He is only about a 3 on the 1 to 10 rating scale. When Mr. Iron (who happens to be about a 6 on the scale - definitely a more handsome, popular fellow) walks into the dance without a date, Miss Sulfate wastes no time in getting rid of Mr. Copper and getting together with Mr. Iron. Who wouldn't trade a 3 for a 6? Mr. Copper is "reduced" to being alone - no charge - no attraction - no date. Mr. Iron became "charged" when he walked in the door (was oxidized) by giving his unwanted electrons to Mr. Copper which is what reduced him. Now Miss Sulfate is attracted to Mr. Iron because of his opposite charge and they are now a couple while Mr. Copper has to wait in the singles line all alone (elemental atoms of copper). Iron can "force" his unwanted electrons onto the copper ions because iron is more reactive than copper.

This gets the students' attention. You can then use another equation to see if they are getting a handle on the concept. It is good to use zinc or magnesium. The students help you finish the equation once you write the reactants on the board. Now you can finish "rating" the metals. Here are the numbers to use:

10 magnesium
9 aluminum
7 zinc
6 iron
5 tin
4 lead

Now the students are ready to place copper on the chart in the proper location. Since copper got "replaced" (dumped as the students will tell you) in every test tube - he must be at the bottom of the list. So we add copper to the bottom as a "3".

The next question you can pose to the students is this:

"Is Mr. Copper doomed 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 we can find a "2" or a "1". At this point you can do the teacher demonstration using a copper strip and silver nitrate. The students love this reaction. Good news for Mr. Copper. Mr. Silver is less reactive and will be reduced (replaced) by Mr. Copper. So, copper can steal silver's date (Miss Nitrate) and he is left alone to hang around by himself. We add silver to the list as a "2". By now of course the students want to know who "1" is (the real loser!!). And the answer is gold or platinum. Usually someone will think to ask if there is anyone 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. You can find clips on YouTube that show various alkali metals being placed into water. Flinn Scientific has a 37 minute video available entitled "Sodium: A Spectacular Event" that is very good.

At this point you 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 you can throw in an equation such as this one:

$$Fe + MgSO_4 \longrightarrow$$

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".

Now you can try to interject practical applications of this concept. One is the galvanization of steel or iron metal. Ask the students to explain how a more reactive metal (zinc) can protect a less reactive metal (iron). The fact that zinc is more reactive IS the reason why it works. Zinc acts as a "sacrificial". Zinc will react with the 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. Oxidized iron does not make a protective coating because it flakes off exposing new iron to be oxidized. The use of sacrificial zinc, aluminum, or magnesium in a hot water heater is another practical example. This concept can also lead to a discussion of methods used to extract (reduce) metals from their ores. We follow up this lab by doing a copper oxidation/reduction lab that simulates reclaiming copper from an ore.

Metals are usually reclaimed from their ores in one of two ways: chemical reduction (using carbon in the form of coke) or electrolytic reduction (using electricity). Carbon is a relatively cheap way to reduce metals from their ores. Carbon fits on the activity series scale between aluminum and zinc – you can assign it a rating of "8". Thus chemical reduction using coke is a viable option for metals below carbon on the activity series. Metals that are more active than carbon use electrolytic reduction. Aluminum is the most abundant metallic element in Earth's crust but has become widely used only in the past century because of the difficulty in extracting it from its ore – it requires an abundant source of cheap hydroelectric power. It is much cheaper to recycle aluminum than to extract it from its ore.

Another practical example that is easy to use involves a coffee can and a tomato sauce can. The inside of the tomato sauce can looks different than the inside of the coffee can. The tomato sauce can is galvanized to protect the steel from the acid in the tomato sauce (hydrogen is less reactive than iron). The coffee can is not galvanized because it does not contain an acid. Pineapple juice cans are also galvanized. Other contextual examples include galvanized grain bins, nails, and zinc tabs on ships or boats. Corrosion Experiment #4 – Poly Coat – investigates the need for a polymer liner inside soda cans to protect the aluminum metal from the carbonic acid in the soda.

Another quick demonstration that also illustrates the different activities of metals is to "burn" a piece of magnesium ribbon, a sample of steel wool, and a copper strip or sheet using either a propane torch or a Bunsen burner.

Heat a copper sheet with a propane torch and it will oxidize. Touch the tip of the inner blue cone of the flame to the hot copper sheet and it will actually reduce the copper oxide on the surface – it is "stealing" the oxygen away to help the propane burn. Pull the tip of the flame away from the surface and the copper will immediately oxidize as the air hits it. The different colors in the oxidation are due to the thickness of the oxidation layers.
As a final write-up for this lab, have the students discuss the following in their journals:

1. Explain oxidation and reduction using an example from the lab. Include these terms: oxidation, reduction, atom, and ion. Also, include an equation.

2. Explain why putting zinc into magnesium sulfate would NOT produce a reaction.

$$Zn + MgSO_4 \longrightarrow \text{no reaction}$$

Additional assessment questions may include the following:

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

   a. $Al + AgNO_3 \longrightarrow$

   b. $Cu + FeSO_4 \longrightarrow$

   c. $Zn + PbSO_4 \longrightarrow$

2. Explain (justify) your answers to question #2.

3. List 3 indications that a chemical reaction is occurring.

4. 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.

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

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

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

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

1. 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

2. Explain (justify) your answers to question #2.

   $Al + AgNO_3$ = 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

   $Cu + FeSO_4$ = 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

   $Zn + PbSO_4$ = 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 lead is lower on the activity series

3. 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

4. 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 that 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 either 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.

5. 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

6. 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.

7. 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. Most copper is found in ore however.

8. 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.

Here is the website where all of the NACE corrosion labs can be found:

http://www.nace-foundation.org/programs/images_programs/cKit_experiments.pdf

Debbie Goodwin
Chillicothe High School
Chillicothe, MO
nywin@hotmail.com