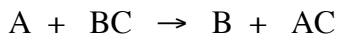


Experiment 9 - Single Replacement Reactions

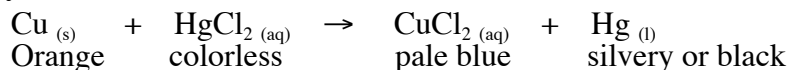
A single replacement reaction is a type of oxidation-reduction reaction. In a single replacement reaction (also called a single displacement reaction), an element reacts with an ionic compound to give a different free element and a different ionic compound. The general form of a single replacement reaction looks like this:



Where A is an element, BC is an ionic compound consisting of positively charged B ions and negatively charged C ions, B is an element, and AC consists of positively charged A ions and negatively charged C ions. It is called a “single replacement” reaction because it appears as though the element A is replacing B in its compound. (In actuality, BC is usually a soluble compound, so it consists of separated B^{x+} ions and C^{x-} ions. Since they are already separated, it isn't really accurate to say that any replacement is occurring.) If this reaction ($A + BC \rightarrow B + AC$) occurs as written, it is said that A is more active than B. If it does not occur, B must be more active than A. (We are assuming, of course, that A and B do not have the same exact activity.)

Recall that oxidation-reduction reactions involve a transfer of electrons. The substance that gets oxidized loses electrons and its oxidation number increases. On the other hand, the substance that is getting reduced is gaining electrons and its oxidation number is therefore decreasing. In a reaction of this type, the element A is going from a zero oxidation number to a positive oxidation number. Its oxidation number is therefore increasing, so it is being oxidized. The substance B, however, is going from a positive to a zero oxidation number, and is therefore being reduced. Therefore, **the more active element is the one that is oxidized more easily.**

To use a specific example, if a strip of copper metal is immersed in a solution of mercury (II) chloride, a reaction occurs:

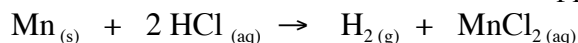


Two electrons are taken from each copper atom, and transferred to a mercury (II) ion, converting it to a neutral mercury atom. The process simultaneously converts Cu to Cu^{2+} , the copper (II) ion. Copper is thus more active than mercury.

The reverse is not the case. That is, Cu^{2+} cannot oxidize Hg (it cannot take electrons from Hg). This can be proven by adding $\text{Hg}_{(l)}$ to a solution of $\text{CuCl}_{2(aq)}$: nothing happens. If a certain reaction occurs spontaneously – in our example, the transfer of 2 electrons from Hg^{2+} to Cu – the reverse process will NOT be spontaneous.

The purpose of this experiment is to determine relative activities of different elements by combining elements with aqueous solutions of ionic compounds. By looking at the surface of the metal during the next few minutes, we can decide whether or not the oxidation-reduction reaction has occurred. If the reaction is happening, a coating will form on the surface of the metal. (Just as a coating of mercury formed on the surface of the copper in the previous example.) In cases where reaction does occur, it will mean that the metal of the strip is undergoing oxidation and that the other metal's ions are gaining electrons (being reduced), forming a coating of that element on the surface of the original metal.

Acidic solutions can also undergo single replacement reactions with some metals. For example, when a piece of manganese metal is placed in a solution of HCl, there is visual evidence of a reaction. The metal begins to dissolve and a gas is formed, which bubbles out of the solution. The reaction that is happening is the following:



The gas being produced is hydrogen gas, H_2 . The manganese dissolves to become manganese ions. In this case, the manganese is displacing the hydrogen from its compound, so the manganese is more active than hydrogen. Metals that react with acid are referred to as “active” metals.

Whether or not the reactions occur, each combination will tell you which of the two elements is more active. You will test five different metals and an acid in this experiment. When all of the data is obtained, you will rank the substances in order of activity.

Safety Precautions:

- Wear safety goggles.
- Lead metal and solutions containing lead ions are poisonous. Make sure to wash your hands after handling any lead or lead compounds.
- Silver nitrate ($AgNO_3$) solutions will stain skin and clothes. If you suspect you may have spilled $AgNO_3$ on yourself, rinse it off immediately. The stains are dark brown and they don't show up right away. You'll know the next day whether or not you spilled $AgNO_3$ on yourself.

Waste Disposal:

- While you are doing the experiment, pour your liquid waste into a beaker. Separate the solid metal pieces from the waste solution.
- Silver pieces must be saved, rinsed off, and given back to the instructor.
- Lead pieces should be collected in a beaker in one of the fume hoods.
- Other metal pieces should be rinsed with water, and then they may be thrown in one of the regular garbage cans.
- When you are finished with the experiment, pour the contents of the waste beaker (liquid waste only) into the **inorganic waste** container (with a blue label) in the fume hood.

Procedure

With some of the combinations used in these experiments the reactions may be slow or difficult to detect. If you see no immediate evidence of reaction, set the tube aside and allow it to stand for about 10 minutes, then reexamine it.

Evidence of reaction will be either evolution of a gas or appearance of a metallic deposit on the surface of the metal strip. Metals deposited from a solution are often black or gray (in the case of copper, very dark reddish brown) and do not resemble commercially prepared metals.

Note that it is not necessary to measure the volumes of the solutions. Find out in the beginning of the experiment what 1 mL looks like in a test tube, and fill all of the tubes to about the same depth.

See the following page for a sample data table. Record detailed observations for each reaction in the data table.

1. Obtain five pieces each of zinc, copper, lead, and magnesium. Obtain one piece of silver from the instructor. Clean the metal pieces with fine sandpaper or steel wool until their surfaces are clean and shiny. This is important, because coatings on their surfaces can inhibit chemical reactions.

- Rinse four test tubes with deionized water and shake out the excess water. Set them up in your test tube rack.
- In each of the four test tubes, place about 1 mL of copper (II) nitrate solution. Add a different metal piece to each test tube (but do not use copper metal). Observe the metal pieces for evidence of reaction. Record your results. (You will need to save and rinse off the piece of silver for re-use.) Dispose of the solution and metals as previously directed (see "Waste Disposal", above).
- Rinse four test tubes with deionized water and shake out the excess water. This time, place about 1 mL of lead (II) nitrate solution into each tube. Add a different metal piece to each tube (but do not use lead metal). Record your results.
- Rinse and shake dry four test tubes, as before. Place about 1 mL of magnesium sulfate solution into each tube. Add a different metal piece to each tube (but do not use magnesium metal). Record your results.
- Rinse and shake four test tubes. Place about 1 mL of zinc sulfate solution into each tube. Add a different metal piece to each tube (but do not use zinc metal). Record your results.
- Rinse and shake four test tubes. Place about 1 mL of silver nitrate solution into each tube. Add a different metal piece to each tube (but do not use silver metal). Record your results.
- Rinse and shake **five** test tubes. Place about 1 mL of hydrochloric acid solution into each tube. Add a different metal piece to each tube. Record your results.
- Dispose of all waste as previously directed. Rinse off the silver metal and return it to the instructor.
- On the basis of your observations, rank all of the metals and hydrogen in order from the most to the least active. Explain your reasoning thoroughly.

Suggested Data Table:

	Cu _(s)	Pb _(s)	Mg _(s)	Zn _(s)	Ag _(s)
Cu(NO ₃) ₂ (aq)					
Pb(NO ₃) ₂ (aq)					
MgSO ₄ (aq)					
ZnSO ₄ (aq)					
AgNO ₃ (aq)					
HCl _(aq)					

Questions:

1. Write the net ionic equation for each reaction that occurred in this experiment.
2. It is observed that tin (Sn) will dissolve in acid, but it will not react with ZnSO_4 solution. Where does tin fit in the activity series that you determined? What additional test(s) would be needed to determine the precise position of tin in the activity series relative to the other metals?
3. Aluminum is oxidized more easily than zinc. What would happen if you mixed together $\text{Zn}_{(s)}$ and $\text{AlCl}_{3(aq)}$? What would happen if you mixed together $\text{Al}_{(s)}$ and $\text{ZnSO}_{4(aq)}$? Write the net ionic equation for the reaction that would occur.
4. Would aluminum be considered an “active” metal? Explain why or why not.