Experiment 8 - Single Replacement Reactions

Introduction

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:

$$A(s) + BC(aq) \rightarrow B(s) + AC(aq)$$

where A is a pure element, BC is an ionic compound consisting of positively charged B ions and negatively charged C ions, B is a pure 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. We can represent this with a net ionic equation, such as (assuming both A and B have a +1 oxidation state):

$$A(s) + B^{+}(aq) \rightarrow B(s) + A^{+}(aq)$$

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, and the reverse reaction should occur ($B + AC \rightarrow A + BC$). It is unlikely that A and B have the same exact activity if they are not the same element.

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 gets reduced gains electrons and its oxidation number decreases. In a reaction of this type, the element A goes from a zero oxidation number (pure elements are always neutral and uncharged) to a positive oxidation number. Its oxidation number increases, so it is being oxidized. The substance B, however, goes 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 gold (III) chloride, a reaction occurs:

$$\underset{\text{Orange}}{3Cu(s)} + \underset{\text{colorless}}{2AuCl_3(aq)} \rightarrow \underset{\text{pale blue}}{3CuCl_2(aq)} + \underset{\text{black}}{2Au(s)}$$

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

The reverse is not the case. That is, Cu^{2+} cannot oxidize Au (it cannot take electrons from Au). This can be proven by adding solid gold to a solution of $CuCl_2$: nothing happens. If a certain reaction occurs spontaneously (in our example, the transfer of 2 electrons to Au^{3+} from Cu metal) the reverse process (transfer of 2 electrons from solid Au to Cu^{2+}) 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, 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. 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. Redox reactions are not always immediate, so check after 10 minutes to see if anything has happened. Note that the elemental metal formed in these reactions may not look shiny like "normal" 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:

$$Mn(s) + 2\,HCl(aq) \rightarrow H_2(g) + MnCl_2(aq)$$

The gas is hydrogen gas, H_2. The manganese dissolves to become manganese(II) ions. In this case, the manganese is displacing the hydrogen from its compound, so the manganese is more active than hydrogen. The hydrogen goes from a +1 oxidation state to a zero oxidation state in the elemental form. Metals that react with any acid are referred to as **active metals**. (Some metals react only with specific acids, such as nitric acid or aqua regia ("king's water"), a mix of HCl and HNO_3 which is the only acid that dissolves gold.)

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 acid in this experiment. When all of the data is obtained, you will rank the substances in order of activity.

<u>Safety Precautions</u>:

- Wear safety goggles.
- Lead (Pb) 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.

<u>Waste Disposal</u>:

- While you are doing the experiment, pour your liquid waste into a beaker. Separate the solid metal pieces from the waste solution and sort them by type.
- Silver pieces must be saved, rinsed off, and given back to the instructor.
- Other metal pieces should be placed in the appropriate container in the waste hood.
- When you are finished with the experiment, pour the contents of the waste beaker (liquid waste only) into the **inorganic waste** container 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 (such as many bubbles on the metal surface) 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.

See the following page for a sample data table. Use a whole page for the table so you can record detailed observations for each reaction.

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 (slow down) chemical reactions.
2. Rinse a well plate with deionized water and shake out the excess water.
3. In separate wells, place one piece of each metal except copper. Fill the well almost to the top with copper (II) nitrate solution. Observe the metal pieces for evidence of reaction. Record your results.
4. In separate wells, place one piece of each metal except lead. Fill the well almost to the top with lead (II) nitrate solution. Observe the metal pieces for evidence of reaction. Record your results.
5. In separate wells, place one piece of each metal except magnesium. Fill the well almost to the top with magnesium sulfate solution. Observe the metal pieces for evidence of reaction. Record your results.
6. In separate wells, place one piece of each metal except zinc. Fill the well almost to the top with zinc (II) sulfate solution. Observe the metal pieces for evidence of reaction. Record your results.
7. In separate wells, place one piece of each metal except silver. Fill the well almost to the top with silver nitrate solution. Observe the metal pieces for evidence of reaction. Record your results.
8. Rinse and shake **five** test tubes and set them up in a test tube rack. Place about 1 mL of hydrochloric acid solution into each tube. Add a different metal piece to each tube. Record your results.
9. Dispose of all waste as previously directed. Rinse off the silver metal and return it to the instructor.
10. Write net ionic equations for each reaction that happened in your notebook. You can assume that ions formed by the reactions have the same charge as ions of that element in the starting solutions. (For example, if we start with a lead(II) solution, then lead ions formed will be Pb^{2+}). Also write one molecular and complete ionic equation for each type of solution that reacted.
11. On the basis of your observations, rank all of the metals and hydrogen in order from the most to the least active.

Suggested Data Table (but use a whole page of your notebook so you have space for detailed observations!)

	Cu(s)	Pb(s)	Mg(s)	Zn(s)	Ag(s)
copper(II) nitrate (aq)					
lead(II) nitrate (aq)					
magnesium sulfate (aq)					
zinc(II) sulfate (aq)					
silver nitrate (aq)					
HCl (aq)					

Name: Section:

Experiment 8 Pre-Lab Sheet

1. (2 pts) Does the precise volume of solution used matter? How much should you use for each test?

2. (1 pts) How long should you wait to see if a reaction happens?

3. (1 pts) Why is polishing the metals pieces useful?

4. (2 pts) If a reaction is observed between Al and $Sn(NO_3)_2$ and between Fe and $Sn(NO_3)_2$, but no reaction is observed between Fe and $Al(NO_3)_3$, what are the relative activities of the three metals?

   (*least*) < < (*most*)

5. (4 pts) Write the balanced complete and net ionic equations for each reaction (assume aqueous ionic compounds are strong electrolytes, although this may not be completely true since the cations are transition metals):

$$Cu(s) + AuCl_3(aq) \rightarrow CuCl_2(aq) + Au(s)$$

$$Mn(s) + 2HCl(aq) \rightarrow H_2(g) + MnCl_2(aq)$$