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:

$$A+BC \rightarrow B+AC \qquad \text{Equation 1}$$

or

$$A+BC \rightarrow C+BA \qquad \text{Equation 2}$$

In the equation, 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 if element A is replacing element B in its compound in **Equation 1**. (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 accurate to say that any replacement is occurring.) This reaction can also occur with the anion 'replaced' instead of the cation as in **Equation 2.**

When the reactions occur as written, the more active element is on the reactant side, forming the less reactive element on the product side. This is the spontaneous direction. The reaction is not spontaneous in the opposite direction.

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 Single Replacement Reaction, elements will form products based on their metallic character. Metals are oxidized to form cations; non-metals are reduced to form anions. Therefore, **the more active METAL element is the one that is oxidized more easily**. **The more active NON-METAL element is the one that is reduced more easily.**

Metals: a strip of copper metal is immersed in a solution of mercury (II) chloride, a reaction occurs:

$$Cu_{(s)} \quad + \quad HgCl_{2\,(aq)} \quad \rightarrow \quad CuCl_{2\,(aq)} \quad + \quad Hg_{(l)}$$

Orange colorless pale blue silvery or black

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. Copper will give up electrons in the presence of mercury(II) ions.

The reverse reaction of copper(II) ions and mercury metal does not go spontaneously. Cu^{2+} cannot oxidize Hg (it cannot take electrons from Hg). This can be proven by adding $Hg_{(l)}$ to a solution of $CuCl_{2\ (aq)}$: 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.

Non-metals: When a halogen acts as an oxidizing agent in a solution, it takes an electron from the reducing agent, becomes a halide, and is, subsequently, hydrated by water. There is two parts to this process that stabilizes the solution formed: ease of electron removal and stability of the hydrated ion. Small ions tend to form good hydrated ions because water can surround the ion with good spatiality. We would expect that the oxidizing ability of the halogens should decrease down the family.

A solution of sodium chloride is mixed with a solution of aqueous fluorine (sometimes called 'fluorine water')

$F_{2(aq)}$	+	$NaBr_{(aq)}$	$\rightarrow$	$NaF_{(aq)}$	+	$Br_{2(aq)}$
colorless		colorless		colorless		brown to yellow

An electron is taken from each bromide ion, and transferred the fluorine atoms in the fluorine molecule, converting them to fluoride ions. The process simultaneously converts F_2 to two F^{-} ions and two Br^{-} ions to Br_2. Bromine is less reactive than fluorine. Why? Bromide is easily oxidized to form bromine, and fluorine is easily reduced to form fluoride. Fluorine is more reactive and would rather exist as an anion instead of an element in the presences of an oxidizing agent. We will generate the halogen in chlorine water, then extract the halogen into cyclohexane. Halogens in water and in cyclohexane have distinctive colors.

The reverse is not the case. That is, $Br_{2(aq)}$ cannot oxidize fluoride. (it cannot take electrons from F^{-}). This can be proven by adding $Br_{2(aq))}$ to a solution of $NaF_{(aq)}$: nothing happens. If a certain reaction occurs spontaneously – in our example, the transfer of 2 electrons from F^{-} to Br_2 – the reverse process will NOT be spontaneous.

Acidic solutions can also undergo single replacement reactions with some metals. Although hydrogen is not a metal, it will readily take electrons from metals that are more active. 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 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. Not all metals react with acids. Also, as with all these reactions, we are not considering other possible controlling factors: surface area of the pieces being compared, complete oxide removal, obfuscation of the observation, etc.

The purpose of this experiment is to determine relative activities of different elements by combining elements with aqueous solutions of ionic compounds and verify the known activity series. By looking at the surface of the metal during the next few minutes, we can decide if 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. For the reactions involving solutions, we will look for evidence of a reaction by color change, or gas formation. Most ionic solutions are colorless. The halogen waters (halogens dissolved in water) have some color: Fluorine (no color), Chlorine (pale yellow-green), Bromine (red-orange), iodine (red-brown).

If the reactions occur, each combination will tell you which of the two elements is more active. This experiment asks you to think of reactions occurring in two directions. If the reaction does not work in one direction, think about the reverse reaction. You will test four different metals, four halogens, and an acid in this experiment. When all 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 if you spilled $AgNO_3$ on yourself.
- Halogen waters are basic and can burn. Handle them carefully.

WASTE DISPOSAL:

- While you are doing the experiment, pour your liquid waste into a beaker. Separate the solid metal pieces from the waste solution. Keep the halogen waste in a separate beaker.
- 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. You might also have to place the tube in a warm water (not boiling!!) bath.

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, and/or a color change in the solution.

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 the tubes to about the same depth.

Part 1 Activity of metals

Obtain four pieces each of zinc, copper, lead, magnesium. 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.

1. Rinse three test tubes with deionized water and shake out the excess water. Set them up in your test tube rack.
2. **Testing copper(II) ions with metals:** In each of the three 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). Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Dispose of the solution and metals as previously directed (see "Waste Disposal", above).
3. **Testing magnesium ions with metals:** Rinse and shake dry three 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). Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Record your results.
4. **Testing zinc ions with metals:** Rinse and shake dry three test tubes, as before. 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). Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Record your results.
5. **Testing lead(II) ions with metals:** Rinse and shake dry three test tubes, as before. 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). Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Record your results.
6. **Testing silver ions with metals:** 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). Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Record your results.
7. **Testing hydrogen ions with metals:** Rinse and shake **four** test tubes. Place about 1 mL of hydrochloric acid solution into each tube. Add a different metal piece to each tube. Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Record your results.
8. **Testing potassium ions with metals:** Rinse and shake **four** test tubes. Place about 1 mL of potassium chloride solution into each tube. Add a different metal piece to each tube. Warm test tubes in hot water bath. Record your results. Observe the metal pieces for evidence of reaction. Record your results,
9. Dispose of all waste as previously directed.

10. Based on your observations, rank all the metals and hydrogen (molecular) in order from the most to the least active.

RESULTS:

1. Fill out the report sheet for this lab. It is found on line.
2. In the report sheet you will:
    a. Give correct NIE for each reaction. If a reaction did not occur, write the reactant, arrow, NR.
    b. Rank each of your metals as greater or less active for the given test

QUESTIONS

See the report sheet for questions