Single Replacement Reactions & Batteries

Objectives

The objectives of this laboratory are to:

a) Perform and observe the results of a variety of single replacement reactions,
b) Become familiar with some of the observable signs of these reactions,
c) Predict and identify the products formed in each of these reactions,
d) Write balanced chemical equations for each single replacement reaction.
e) Make three voltaic cells by using a salt bridge, a citrus fruit and potato.
f) Understand the function of a salt bridge.

Background

During a chemical reaction both the form and composition of matter are changed. Old substances are converted to new substances, which have unique physical and chemical properties of their own. Some of the observable signs that a chemical reaction has occurred include the following:

- A metallic deposit appears
- Bubbles appear
- A temperature change occurs
- A color change occurs
- A precipitate (cloudy, tiny particles) appears

Note that there are other observable signs for chemical reactions, but these are most likely to be seen in this lab.

Single Replacement Reactions

In Part A of this lab we will examine **single replacement reactions**. This is one type of **oxidation-reduction reaction**, or **redox reaction**, because it occurs via a transfer of electrons.

Single replacement reactions have the general form:

$$\mathbf{A + BC \rightarrow B + AC}$$

Here, A is an element and BC is usually an aqueous ionic compound or an acid (consisting of B^+ and C^- aqueous ions). Element A replaces element B in the compound BC; this results in the formation of a new element B and a new ionic compound or acid, AC. If the new element B is a metal, it will appear as a **metallic deposit**. If it is a gas, it will appear as bubbles.

An **activity series** of elements is often used to determine if A will displace B in a single replacement reaction. The activity series is provided on the following page. As a rule, if A has a higher activity that B, a single replacement reaction will occur. However, if A has lower activity than B, a single replacement reaction will not occur.

Example 1: magnesium metal + aqueous aluminum chloride

Since Mg is more active than Al, a single replacement reaction will occur. The predicted products are aluminum metal and aqueous magnesium chloride

Reaction Equation: $\underset{\mathbf{A}}{3\ Mg} + \underset{\mathbf{B\ C}}{2\ AlCl_3} \rightarrow \underset{\mathbf{B}}{2\ Al} + \underset{\mathbf{A\ C}}{3\ MgCl_2}$

Example 2:	aluminum metal + aqueous magnesium chloride Since Al is **not** more active than Mg, a replacement will **NOT** occur. Reaction Equation: Al + $MgCl_2$ $\rightarrow$ **NO REACTION** **A B C**

ACTIVITY SERIES FOR METALS (and HYDROGEN)		
highest activity	**Li**	
	K	
	Ca	
	Na	
	Mg	
	Al	
	Zn	**→ Zn^{+2}**
	Cr	**→ Cr^{+3}**
	Fe	**→ Fe^{+2}**
	Cd	**→ Cd^{+2}**
	Ni	**→ Ni^{+2}**
	Sn	**→ Sn^{+2}**
	Pb	**→ Pb^{+2}**
	H_2	
	Cu	**→ Cu^{+2}**
	Ag	**→ Ag^{+1}**
	Hg	**→ Hg^{+2}**
lowest activity	**Au**	**→ Au^{+3}**

Oxidation numbers indicate how many electrons each atom in a compound possesses relative to the free atom. It serves as a "book-keeping" system so that the flow of electrons can be observed. If the oxidation number of any atom changes during a chemical reaction, a transfer of electrons has occurred.

When an atom loses electrons, it is being oxidized (or has undergone oxidation). **Oxidation is a loss of electrons.** Conversely, when an atom gains electrons, it is been reduced (or has undergone reduction). **Reduction is a gain of electrons.**

When describing redox reactions relative to the entire molecule, we use the terms **oxidizing agent** and **reducing agent**. A reducing agent is a substance that is being oxidized and thus causes another substance to be reduced. An oxidizing agent is a substance that is being reduced and thus causes another substance to be oxidized. We can remember these concepts with the helpful mnemonic "OIL RIG":

O	*xidation*	
I	*s*	***The REDUCING AGENT***
L	*oss of Electrons*	
R	*eduction*	
I	*s*	***The OXIDIZING AGENT***
G	*ain of Electrons*	

Example 1: Revisited	magnesium metal + aqueous aluminum chloride
	Oxidation Number: **0** **+3 -1** **0** **+2 -1**
	Reaction Equation: $3\ Mg\ (s) + 2\ AlCl_3\ (aq) \rightarrow 2\ Al\ (s) + 3\ MgCl_2\ (aq)$
	A **B C** **B** **A C**

In the above example, magnesium is oxidized (the reducing agent) because it has lost electrons. In addition, the aluminum atom in aluminum chloride is reduced because it has gained an electron. Thus, aluminum chloride is the oxidizing agent.

Electricity can be described as a flow of electrons through a wire. This form of energy is caused by the motion of electrons. A device that creates electrical current from redox reactions is called an **electrochemical cell**, **voltaic cell**, **galvanic cell** or ***battery***.

Batteries serve as a source of energy for flashlights, radios, as well as car motors. In Part B we will build an electrochemical cell. It will consist of two metals (called ***electrodes***) connected by a **salt bridge** between individual half-cells. A salt bridge, which contains a strong electrolyte, will join the two reactions and complete the circuit. It allows for the mixing of the two solutions.

The metal strip where oxidation occurs is the ***anode*** and is labeled with a negative (–) sign. The metal strip where reduction occurs is called the ***cathode*** and is labeled with a (+) sign. These symbols can be observed on every day batteries. Electrons flow from ***anode*** to the ***cathode***:

A (anode) = O (oxidation)

↓ electron flow

C (cathode) = R (reduction)

Procedure

Safety

Be especially cautious when using the 6M HCl, and 3M H_2SO_4. These substances can burn your skin. Also be aware that skin discoloration will result from contact with $AgNO_3$. If you feel any tingling sensations or see any color changes on your skin, flush with water immediately for a minimum of 15 minutes. Inform your instructor of any chemical contact as soon as possible.
Personal Protective Equipment (PPE) required: safetygoggles, lab coat, closed-toe shoes

Materials and Equipment

At front bench: copper, zinc, magnesium, sodium bicarbonate, coins, paper towels, matzo or bread, oranges, potatoes, string soaked in potassium sulfate, battery kit (containing voltmeter, alligator clips, copper and zinc electrodes), 5 medium test tubes

Solutions (in hood): 3M sulfuric acid, 6M hydrochloric acid, 1M sodium chloride, all other solutions are 0.1M and include silver nitrate, lead (II) nitrate, copper (II) sulfate, zinc nitrate, nickel (II) nitrate, aluminum sulfate

Obtain from your locker: large beaker, 2 small beakers

Part A: Single Replacement Reactions

1. Use the medium sized test tubes; place them in a large beaker. Use clean test tubes that have been rinsed with ***distilled water***. *The test tubes do not have to be dry.*

2. For 3-mL quantities of all solutions, use two full dropper squirts of each solution.

3. Place one piece of metal in the test tube first, and then add the solution. The metal should be completely immersed in the solution used. If results are not obtained immediately, give the reaction some time. Some reactions take longer than others.

4. Perform the following reactions, and record your observations for each on the data sheet. ***All waste is to be disposed of in the plastic container in the hood*!**

    a. Zinc metal + hydrochloric acid
    b. Copper metal + aqueous silver nitrate
    c. Copper metal + aqueous zinc nitrate
    d. Zinc metal + aqueous lead (II) nitrate
    e. Magnesium metal + sulfuric acid

Part B: Batteries

Citrus and Potato Batteries

1. Obtain a cut piece of orange from front bench.
2. Place the polished piece of zinc metal plate into the fruit. Connect to black lead on voltmeter using alligator clips.
3. Place a polished piece of copper metal plate into the fruit. Connect to red lead on voltmeter using alligator clips.
4. Turn voltmeter to 2V and record the voltage.
5. Repeat 1- 4 with a potato. Record your observations.

Conductivity of Matzo or Bread

1. Obtain a piece of matzo or bread from front bench.
2. Connect alligator clips to either side of sample. Other end of the alligator clips should still be connected to the voltmeter.
3. Turn voltmeter to 2V and record the voltage.

Salt Bridge Battery

1. Into a 250-mL beaker add approximately 25 mL of 1M copper (II) sulfate. (This is the cathode.) Label as Beaker #1.
2. Into another 250-mL beaker, add approximately 25 mL of 1M zinc sulfate. (This is the anode.) Label as Beaker #2.
3. Connect the solutions in the beakers by placing one end of a 10- inch piece of cotton twine into Beaker #1 and the other end into the Beaker #2. Obtain twine which has been soaked in a concentrated potassium sulfate solution from your instructor. (The twine is the salt bridge.)
4. Place a polished piece of copper metal plate into Beaker #1. Connect to voltmeter using wire clips.

5. Place a polished piece of zinc metal plate into Beaker #2. Connect to voltmeter using wire clips.
6. Turn voltmeter to 2V and record the voltage.
7. Remove salt bridge. Record the voltage.

Coin Battery

1. Obtain four coins from the front bench: penny, 2 nickels, and dime. Cut 3 circles of paper towel about the same diameter or slightly larger than the coins.
2. Soak the paper towel pieces in a beaker containing 20 mL of a 1M sodium chloride solution.
3. Make a stack alternating coins and NaCl-soaked paper towel pieces: penny, paper, dime, paper, nickel, paper, nickel. Make sure that the penny is on one end, and the nickel on the other end. **Do not allow the coins or the edges of the paper towels to touch.**
4. Touch red probe to penny and the black probe to the nickel.
5. Turn voltmeter to 2V and record the voltage.