An Activity Series
AP Chemistry Laboratory #20

Introduction

In this experiment, a series of metals and a series of nonmetal halogens are studied to find their relative reactivities. The reactivity of the metals is determined by combining the metals with a complementary series of metal ions in solution. The reactivity of three halogens is found by mixing each with a halide ion solution. Using the observed reactions, an activity series, from most reactive to least reactive, is developed for the metals and for the halogens.

Concepts

- Activity series
- Oxidation–reduction
- Half-cell reaction

Background

A ranking of elements according to their reactivity is called an *activity series*. For example, an activity series containing the elements calcium, gold, and iron would put the reactive calcium at the top, iron in the middle, and the unreactive gold at the bottom. If a piece of iron metal is placed in a solution of gold nitrate, the iron dissolves forming positive ions in solution while solid gold metal appears. The more reactive metal (iron) displaces ions of the less reactive metal [gold(III)] from solution. The less reactive element appears as the solid element.

Reactions such as these are examples of *oxidation–reduction reactions*. Oxidation is defined as the process of losing electrons and substances that lose electrons during chemical reactions are said to be oxidized. Substances that gain electrons during chemical reactions undergo reduction and are said to be reduced. If one reactant gains electrons, another must lose electrons. Oxidation and reduction reactions occur simultaneously, and there must be an equal number of electrons lost and gained during the two reactions. In the reaction of iron metal with gold ions, the iron metal is oxidized and the gold ions are reduced. The more reactive metal is the one that is more easily oxidized.

$3Fe(s) \rightarrow 3Fe^{2+}(aq) + 6e^-$	Iron loses electrons.	*Oxidation*
$2Au^{3+}(aq) + 6e^- \rightarrow 2Au(s)$	Gold gains electrons.	*Reduction*
$3Fe(s) + 2Au^{3+}(aq) \rightarrow 3Fe^{2+}(aq) + Au(s)$		*Oxidation–reduction*

Figure 1. Reduction of gold ions by iron metal.

When writing oxidation–reduction reactions, it is customary to break the reaction into the two parts or *half-cell reactions.* These half-cell reactions represent the separate oxidation and reduction processes that occur simultaneously. The electrons within the two half-cell reactions must be equal so there is no net gain or loss of electrons for the overall reaction.

When a substance readily loses electrons (and is oxidized), it acts as a good reducing agent. When a substance has a strong tendency to gain electrons (and be reduced), it acts as a good oxidizing agent. Gold ions, $Au^{3+}(aq)$, have a strong tendency to acquire electrons to form neutral gold atoms, $Au(s)$. Gold ions are thus easily reduced and act as good oxidizing agents.

Experiment Overview

The purpose of this experiment is to determine the activity series for five metals and for three halogens. The first part of this experiment derives an activity series for metals and uses a microscale technique. The second part derives an activity series for halogens. It makes use of a solvent extraction technique.

The series of metals to be studied are copper, zinc, magnesium, lead, and silver. Solutions of metal nitrates for each of these metals are placed in reaction wells. A piece of each metal is then placed in the other metals' nitrate solutions and observed to see if any reaction occurs. If a metal reacts with another metal nitrate, then the solid metal has reduced the other metal ion and is, therefore, the more reactive metal of the two. By comparing the results of 16 different reactions, the five metals are ranked from most reactive to least reactive.

In Part 2, tests are performed to determine the activity series of the halogens. Chlorine (Cl_2), bromine (Br_2), and iodine (I_2) are placed in solutions containing chloride (Cl^-), bromide (Br^-), or iodide (I^-). An activity series of the nonmetallic halogens places the most reactive halogen at the top. In the reaction of a free halogen (X_2) with a different halide ions (Y^-), the free halogen gains electrons and is then reduced to its corresponding halide ions (X^-). The original halide ions lose electrons and therefore are oxidized to their corresponding free halogen (Y_2). The more reactive halogens displaces ions of the less reactive halides from solution. In an activity series of halogens, the most reactive halogen is the one most easily reduced.

$$X_2(aq) + 2e^-(aq) \rightarrow 2X^-(aq) \qquad \textit{Reduction}$$

$$2Y^-(aq) \rightarrow Y_2(aq) + 2e^- \qquad \textit{Oxidation}$$

$$X_2(aq) + 2Y^-(aq) \rightarrow 2X^-(aq) + Y_2(aq) \qquad \textit{Oxidation--reduction}$$

Figure 2. Reduction of a free halogen X_2 by halide ions Y^-.

To determine if a reaction occurs, a method is needed to identify which halogen is present. Halogens dissolve in the nonpolar solvent mineral oil forming different colored solutions. Mineral oil does not dissolve in water, but when shaken with an aqueous halogen solution, the halogen is extracted from the water into the mineral oil. The color of the mineral oil layer indicates which halogen is present.

Materials

Copper foil, 6 × 6 mm pieces, 4
Zinc foil, Zn, 6 × 6 mm pieces, 4
Magnesium ribbon, Mg, 6-mm pieces, 4
Lead foil, Pb, 6 × 6 mm pieces, 4
Silver nitrate solution, $AgNO_3$, 0.1 M, 4 mL
Copper(II) nitrate solution, $Cu(NO_3)_2$, 0.1 M, 4 mL
Zinc nitrate solution, $Zn(NO_3)_2$, 0.1 M, 4 mL
Magnesium nitrate solution, $Mg(NO_3)_2$, 0.1 M, 4 mL
Lead nitrate solution, $Pb(NO_3)_2$, 0.1 M, 4 mL
24-well reaction plate
Beral-type pipets, 5
Forceps
Stirring rod

Procedure

Part 1. Determine an Activity Series for Metals.

1. Place the 24-well plate on top of a piece of white paper so that there are 6 wells across (columns) and 4 wells down (rows). Refer to Figure 3 to see how the wells are arranged. Note that each well is identified by an unique combination of a letter and a number, where the letter refers to a horizontal row and the number to a vertical column.

2. Put one dropper-full (about 15 drops or 1 mL) of copper(II) nitrate solution in wells B1, C1, and D1 in the first column.

3. Put one dropper-full of magnesium nitrate solution in wells A2, C2, and D2 of the second column.

4. Put one dropper-full of lead nitrate solution in wells A3, B3, and D3 of the third column.

5. Put one dropper-full of zinc nitrate solution in wells A4, B4, and C4 of the fourth column.

6. Put one dropper-full of silver nitrate solution in each of the wells A5 through D5 in the fifth column.

7. Put a small piece of copper metal in each of the wells containing a solution in the first row.

8. Add magnesium metal to the solutions in the second row, add lead metal to the solutions in the third row, and add zinc metal to the solutions in the fourth row. Use a stirring rod to submerge each metal in the solutions. Allow to stand at least 5 minutes.

9. Determine if a reaction has occurred in each well by observing if a new metal has deposited or if the surface of the metal has become coated.

Record each observation as either *coating forms* or *no reaction* in the Part 1 Data Table.

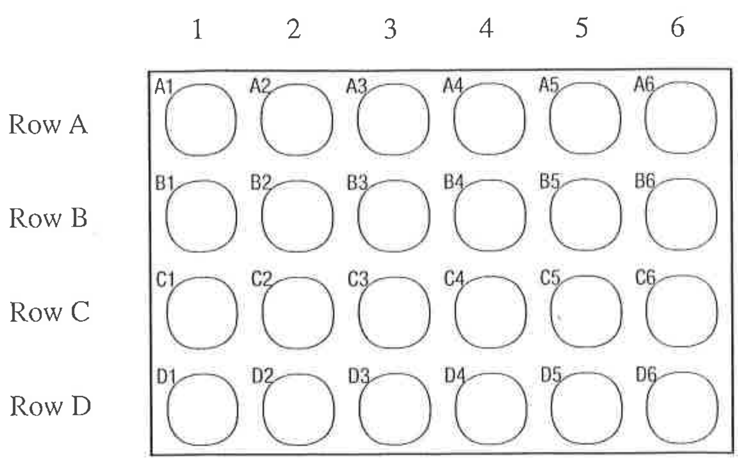

Figure 3.

Data Tables

Name ______________________________

Part 1. An Activity Series for Some Metals.

Record your observations in the data table below:

	$Cu^{2+}(aq)$	$Mg^{2+}(aq)$	$Pb^{2+}(aq)$	$Zn^{2+}(aq)$	$Ag^{+}(aq)$
Cu(s)	×				
Mg(s)		×			
Pb(s)			×		
Zn(s)				×	

Post-Lab Questions

1. Write balanced net ionic equations for all the reactions that occurred with the metals.
2. List the metals in order of decreasing ease of oxidation. Compare this list with an activity series found in a textbook. How do the two lists correlate?
3. Write reduction half-reactions for each of the metal ions. Arrange the reaction list in order of decreasing ease of reduction. Compare the order with a listing found in a table of standard reduction potentials. How do the two lists correlate?