**Reaction Types**

*Introduction*

Four types of reactions in aqueous solution presented in most general chemistry textbooks: gas-forming reaction, precipitation reaction, acid-base (neutralization reaction), and oxidation–reduction (redox) reaction. Each reaction type can readily be identified. A gas-forming reaction requires that at least one product exist in the gas phase. A precipitation reaction requires that at least one product, called a precipitate, be insoluble in water and thus exist as a solid. In an acid-base reaction, an acid reacts with a base to generate a weak electrolyte (usually water) and a salt. (An acid is a source of H+ in water; a base is a source of OH- in water.) A redox reaction involves the transfer of electrons from one reactant to another. One sign that a reaction is a redox reaction is that an element occurs as either a product or reactant in its elemental form.

Three of the reaction types are covered in this demonstration: gas-forming, precipitation, and oxidation-reduction. Acid-base reactions are covered in different demonstrations. The reactions involve some reactive and/or toxic species; however, these are the reactions popularly chosen to be used in illustrations in physical sciences textbooks. Thus, seeing in person the reaction in photos in the textbook should be a valuable experience.

*Gas-forming Reaction*

The gas forming reaction involves the reaction of copper metal with concentrated nitric acid to produce nitrogen dioxide and copper(2+) ions (if the acid is diluted, another reaction occurs) (Equation 1). The resulting solution is initially green from the formation of cupric nitrate while a red-brown gas, NO2, is released. The gas should not be breathed so that the reaction must be performed outdoors or in a well-functioning fume hood.

Cu(s)  +  4 HNO3(aq) 🡪 Cu(NO3)2(aq)  +  2 NO2(g)  +  2H2O(l) (Eqn. 1)

The addition of water to the solution, diluting the nitrate ions present, will turn the solution the blue color of hydrated Cu2+ ions as the nitrate bound to the cupric ions is displaced by water. This reaction is also an oxidation reduction reaction. Note that elemental copper occurs in only one side of the equation. Reactions need not be only one of the four types.

Most notably, the observation of this reaction was an inspiration for the career of the well-known chemist Ira Remsen (1846-1927), who was the second president of John Hopkins University and the co-discoverer of the artificial sweetener saccharin.

“While reading a textbook of chemistry I came upon the statement, "nitric acid acts upon copper." I was getting tired of reading such absurd stuff and I was determined to see what this meant. Copper was more or less familiar to me, for copper cents were then in use. I had seen a bottle marked nitric acid on a table in the doctor's office where I was then "doing time." I did not know its peculiarities, but the spirit of adventure was upon me. Having nitric acid and copper, I had only to learn what the words "act upon" meant. The statement "nitric acid acts upon copper" would be something more than mere words. All was still. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table, opened the bottle marked nitric acid, poured some of the liquid on the copper and prepared to make an observation. But what was this wonderful thing which I beheld? The cent was already changed and it was no small change either. A green-blue liquid foamed and fumed over the cent and over the table. The air in the neighborhood of the performance became colored dark red. A great colored cloud arose. This was disagreeable and suffocating. How should I stop this? I tried to get rid of the objectionable mess by picking it up and throwing it out of the window. I learned another fact. Nitric acid not only acts upon copper, but it acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and another fact was discovered. Nitric acid acts upon trousers. Taking everything into consideration, that was the most impressive experiment and relatively probably the most costly experiment I have ever performed. . . . It was a revelation to me. It resulted in a desire on my part to learn more about that remarkable kind of action. Plainly, the only way to learn about it was to see its results, to experiment, to work in a laboratory.”[1]

*Precipitation Reaction*

A precipitate is a solid, water-insoluble product formed in a reaction. A reaction that forms a precipitate from soluble reactants is a precipitation reaction. The formation of the precipitate lead chromate from the mixing of a solution of potassium chromate and a solution lead nitrate (Equation 2) is an excellent example as the precipitate is a bright yellow color. The color should be familiar to students as lead chromate was used previously as the pigment for yellow highway lines (before toxic concerns arose) as the solid has extremely limited solubility in addition to its bright color.

K2CrO4(aq) + Pb(NO3)2(aq) 🡪 PbCrO4(s) + 2 KNO3(aq) (Eqn. 2)

*Oxidation-reduction (redox) Reaction*

In an oxidation-reduction reaction or redox reaction, electrons are transferred from one reactant to another; thus, one reactant loses electrons while another gains electrons. When a substance loses electrons, the process is called an oxidation. The substance losing the electrons is said to be oxidized. In turn, when a substance gains electrons, the process is called a reduction, and the substance accepting or gaining the electrons is said to be reduced. Equation 3 is an example of an oxidation reduction reaction.

2 Al(s) + 3 Cu2+(aq) 🡪 3 Cu(s) + 2Al3+(aq) (Eqn. 3)

A redox reaction can be separated in two components: the reduction half-reaction and the oxidation half reaction. For Equation 3 describing the reaction between aluminum metal and cupric ions, the half reactions are

2 Al(s) 🡪 2 Al3+(aq) + 6 e-  oxidation (electrons lost)

3 Cu2+(aq) + 6 e- 🡪 3 Cu(s) reduction (electrons gained)

The sum of the two half reactions gives the overall reaction as the electrons gained and lost in the half-reactions cancel. In this reaction, aluminum is the reducing agent, the species that loses electrons, while cupric ions are the oxidizing agent, the species that is reduced.

The reaction occurs only in the direction as written; the reverse reaction (the reaction of copper metal with Al3+) is not spontaneous. Which direction the reaction will proceed in can be predicted. The ease with which several metals can be oxidized has been established, and the metals have been arranged in order in order of their decreasing ease of oxidation in the activity series (Table 1). Metals that appear at the top of the table are easily oxidized and generally exist as compounds of their ions. Metals that appear at the bottom of the table are not readily oxidized and generally occur as the metal. Note that aluminum appears in the table above copper. Hence, aluminum is the more easily oxidized. Thus, cupric ions will react with aluminum to result in the oxidation of the metal atoms to aluminum ions. In turn, aluminum ions will not react with copper metal.

Table 1. Activity series of metals in aqueous solution (ease of oxidation increases from bottom to top)

Metal Oxidation Reaction

Lithium Li(s) 🡪 Li+(aq) + e-

Potassium K(s) 🡪 K+(aq) + e-

Barium Ba(s) 🡪 Ba2+(aq) + 2 e-

Calcium Ca(s) 🡪 Ca2+(aq) + 2 e-

Sodium Na(s) 🡪 Na+(aq) + e-

Magnesium Mg(s) 🡪 Mg2+(aq) + 2 e-

Aluminum Al(s) 🡪 Al3+(aq) + 3 e-

Manganese Mn(s) 🡪 Mn2+(aq) + 2 e-

Zinc Zn(s) 🡪 Zn2+(aq) + 2 e-

Chromium Cr(s) 🡪 Cr3+(aq) + 3 e-

Iron Fe(s) 🡪 Fe2+(aq) + 2 e-

Cobalt Co(s) 🡪 Co2+(aq) + 2 e-

Nickel Ni(s) 🡪 Ni2+(aq) + 2 e-

Tin Sn(s) 🡪 Sn2+(aq) + 2 e-

Lead Pb(s) 🡪 Pb2+(aq) + 2 e-

Hydrogen H2(g) 🡪 2 H+(aq) + 2 e-

Copper Cu(s) 🡪 Cu2+(aq) + 2 e-

Silver Ag(s) 🡪 Ag+(aq) + e-

Mercury Hg(s) 🡪 Hg2+(aq) + 2 e-

Platinum Pt(s) 🡪 Pt2+(aq) + 2 e-

Gold Au(s) 🡪 Au3+(aq) + 3 e-

*Demonstration*

Materials

1. Gas forming reaction

5 g Cu metal (fillings or wire)

40 mL of nitric acid (concentrated)

100 mL H2O

500-mL Erlenmeyer flask

100-mL graduated cylinder

Waste bottle

1. Precipitation reaction

200 mL of 0.1 M K2CrO4

200 mL of 0.1M Pb(NO3)2

1 L beaker

Stirring rod

Waste bottle

1. Redox reaction

Aluminum foil (sheet ~12” x 12”, waded loosely into a ball)

300 mL of 0.1 M CuCl2 (or CuSO4 and NaCl)

Copper metal (filings or wire) (enough to nearly cover bottom of beaker or be seen by audience)

300 mL of 0.1 M AlCl3

Water

2 600-mL or 1-L beakers

Waste bottle

Procedure

**Gas-forming Reaction**

Place the copper into the bottom of the Erlenmeyer flask. Outside or in a well-functioning hood, add the nitric acid. As the reaction slows with time, the container can be swirled. When the reaction is complete or close to completion, add the water slowly.

**Precipitation Reaction**

The solutions are made by dissolving 6.62 g of Pb(NO3)2 in water to give a total volume of 200 and by dissolving 3.88 g K2CrO4 in water to give a total volume of 200 mL.Add one of the solutions to the 1-L beaker. Slowly pour the other solution into the beaker. Stir with a stirring rod to guarantee completion.

**Oxidation-reduction Reaction**

Place the copper chloride (or copper sulfate and sodium chloride) solution and the aluminum chloride solution into separate beakers. (The solutions are made by dissolving 5.11 g CuCl2.2H2O in water to give a total volume of 300 mL and by dissolving 7.24 g AlCl3.6H2O in water to give a total volume of 300 mL. Alternatively for the cupric ion solution, 7.49 g of CuSO4.5H2O and 1.75 g NaCl can be dissolved in water to give a total volume of 300 mL). Add the waded ball of aluminum foil to the cupric ion solution. Add the copper fillings or wire to the aluminum solution. In the beaker containing the cupric ion solution, the blue color of hydrated cupric ion should diminish as copper metal forms at the aluminum foil and the aluminum foil is consumed as Al3+ ions are formed. No reaction occurs in the beaker with the copper metal in the AlCl3 solution. Note that chloride ions must be present for the reactions to occur. The surface of the aluminum is covered by a layer of aluminum oxide, which protects it against reaction; the chloride can penetrate this layer.

**Safety Notes:** Safety goggles are required for all experiments.

Nitric acid is a corrosive chemical and should be handled with care. Wearing gloves is recommended when handling these chemicals. Spills should be cleaned with large volumes of water. If contact is made with your skin, rinse thoroughly with water. The reaction of copper metal and nitric acid produces nitrogen oxides that should not be inhaled; thus, this procedure should be performed in a fume hood or outside.

Lead(II) nitrate, potassium chromate, and the copper salts are toxic chemicals. Gloves should be worn when handling these chemicals.

Unwanted chemicals are not to go down the drain but should be placed in appropriate waste containers.

**References**

1. from F. H. Getman, "The Life of Ira Remsen"; Journal of Chemical Education: Easton, Pennsylvania, 1940; pp 9-10; quoted in Richard W. Ramette, "Exocharmic Reactions" in Bassam Z. Shakhashiri, *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 1.*Madison: The University of Wisconsin Press, 1983, p. xiv.