CHEMICAL REACTIONS

Introduction

This is an introduction to chemical reactions. The goal is to demonstrate chemical reactions, reinforce formula writing, introduce students to writing and balancing chemical equations, and to present the reasons why chemical reactions go to completion. This can be reinforced by microscale or small scale laboratory experiment on chemical reactions.

An introduction activity to chemical reactions is Chemical Reactions which can be found at http://www.chymist.com/chemical%20reactions.pdf This is a microscale experiment on observing changes when chemical reactions take place. This can be performed in a classroom or lecture hall where students are seated at desks or tables. It is not intended as an exercise in writing chemical equations.

A follow-up experiment is Synthesis of Zinc Iodide: Tracking a Chemical Reaction which can be found at http://www.chymist.com/zinc%20iodide.pdf This small scale experiment examines the properties of the reactants and products and the visual change that occurs in a chemical reaction. The aim of this experiment is to visualize a chemical change.

A complete overview of chemical reactions can be covered in a microscale experiment where multiple substances, elements and compounds, are reacted and students are required to observe changes and to write balanced equations for those reactions that occur. Such experiments are generally found in standard laboratory manuals.

In this demonstration, chemical reactions are classified as direct union, decomposition, single replacement, metathesis (or double replacement) and combustion. This follows a worksheet I utilize in my classes. http://www.chymist.com/Equations.pdf

The reactions presented in this package are only a small possibility of the possible reactions that can be demonstrated in class. There is a wide variety of elements and compounds to choose from depending on what you want your students to observe. When choosing reactants for demonstrations, select those that will produce visible and, when possible, colorful reactions. Avoid fire, smoke and explosion reactions – this is meant as a learning exercise, not entertainment.

Safety and Disposal

This is not a detailed set of safety instructions for performing these reactions. There is a significant amount of safety information that is left to the individual who attempts these reactions. It is important that the demonstrator knows and understands the safety involved with every substance and procedure that they choose to run. Consult the SDS for each chemical along with safety manuals that document potentially hazardous reaction.

The demonstrator must wear proper eye protection at all times. Gloves are essential for handling any reactive elements or compounds. Use tongs and spatulas when obtaining materials from storage containers.

All the reactions must be carried out in an area where there is sufficient ventilation.

Counter tops or table tops must be chemically resistant or protected by resistant coverings or trays.

If any flammable materials are used, or if a reaction produces sparks or flames, a fire extinguisher must be readily available.

If anything sparks, or has the possibility of explosion, adequate safety shielding must be provided.

Use the smallest amount of material possible in performing these reactions. Drops of reactant may be sufficient to show a reaction taking place. Use a document camera or similar device to project the reactions onto a large screen or a wall to make the reactions visible to a class.

All flammable materials must be in small containers with secure caps. They must be removed from the demonstration area before doing any chemical reactions.

When possible, solutions should be at a concentration of 0.1 M.

Acids and bases may have to be at concentrations ranging from 0.1 M to 3 M, depending on the reaction being demonstrated. Use the lowest concentration that works.

Do not pass around open containers of liquids where the contents can be spilled. Line them up on the demonstration table so the class can observe them before leaving the classroom. Dismiss the class a few minutes early to accomplish this.

Disposal must be in accordance with all local regulations.

Materials Needed and Procedures

Direct Union or Combination Reactions

Any reaction in which two or more substances combine to form a single product is a *direct union* or *combination* reaction. The general form of a direct union reaction is

$$A + B \rightarrow AB$$

Some examples of this type of reaction:

a. A metal + non-metal

$Zn + I_2 \rightarrow ZnI_2$ (This can be performed in a zip-lock bag. Add water to initiate.)

$2\,Fe + O_2 \rightarrow 2\,FeO$ (use powder Fe or burn steel wool)

$2\,Mg + O_2 \rightarrow 2\,MgO$ (burn Mg ribbon)

$Fe + S \rightarrow FeS$ (heat in a crucible)

b. Metal oxide + non-metal oxide

$CaO + CO_2 \rightarrow CaCO_3$ (filter solution of CaO, bubble in CO_2)

c. Non-metal + non-metal

$C + O_2 \rightarrow CO_2$ (burn graphite or a stick of charcoal)

$2 H_2 + O_2 \rightarrow 2 H_2O$ (premix gases in a pop bottle or ignite a balloon filled with H_2 – loud noise)

Decomposition Reactions

Decomposition is the reverse of combination. That is, a single reactant is broken down into two or more products either elements or compounds. A decomposition reaction will take place because the compound is unstable or as a result of heating or electrical decomposition (electrolysis). The general form for a decomposition reaction is:

$$AB \rightarrow A + B$$

Some examples of decomposition reactions are: (Note: any production of O_2 can be confirmed by a glowing wood splint test.)

$2 HgO \rightarrow 2 Hg + O_2$ (Perform in large test tube. Ventilation needed. Must be able to dispose of Hg)

$2 KClO_3 \rightarrow 2 KCl + 3 O_2$ (Perform in large test tube. Take extreme care if testing for oxygen, this can be an explosion hazard)

$CaCO_3 \rightarrow CaO + CO_2$ (Do this in a flask or beaker with a loose cover. Pour CO_2 over a small tea candle.)

$2 H_2O_2 \rightarrow 2 H_2O + O_2$ (Use yeast or MnO_2 as catalyst)

To understand how to predict products of decomposition reactions, see Chemical Equations worksheet, Section V. The Effect of Heat on Metallic Compounds.

Single Replacement Reactions

(Recent textbooks refer to these as oxidation-reduction reactions.)

A displacement reaction involves an element reacting with a compound whereby the element displaces a second element from the compound. The general form of this type reaction is:

$$A + BC \rightarrow AC + B$$

Displacement reactions usually occur between the following combinations: (Refer to the Chemical Equations worksheet, Refer to Section IV, The Electromotive Series.)

a. An active metal + water

Active metals in the alkali and alkaline earth groups of the periodic table react with water to liberate hydrogen and form a hydroxide.

$$2 Li + 2 H_2O \rightarrow 2 LiOH + H_2$$

Safety Note: Wear gloves when working with alkali and alkaline earth metals.

Perform this reaction using a very small piece of the active metal in 200 mL of room temperature water in a 600-mL or 800-mL beaker with a fine mesh wire gauze covering the beaker. Use a document camera, or similar, to project the reaction on a screen.

Repeat this using small pieces of Na and K. When cutting Li, Na and K from a larger piece, do that under the document camera to show the silvery surface of the metal. Add one or two drops of phenolphthalein to the solution after the reaction to show that it is alkaline.

Repeat this reaction using Sr granules, Ba granules, Ca and Mg. Group II elements can be reacted in a large test tube. (Note: Mg does react with room temperature water. Add two drops of phenolphthalein indicator to the water before adding the Mg. The phenolphthalein will turn pink around the Mg as it reacts. Bubbles of H_2 will be visible on the surface of the Mg.)

Collect any hydrogen from the Ca reaction only by placing a second test tube upside down over the reaction tube. Move the test tube of hydrogen over a flame for a "popping" sound. Depending on the amount of hydrogen collected, a flame should be observed.

Have the class observed the difference between the solutions obtained from the reactions of the Group I and the Group II metals. The "Earth" (or precipitates) from the Alkaline Earth elements should be obvious.

b. An active metal + an acid

When a metal which is above hydrogen in the activity series is reacted with an acid, hydrogen is liberated and a salt is formed.

$$Zn + 2\,HCl \rightarrow ZnCl_2 + H_2$$ (Perform in large test tube)

$$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$$

Use other metals such as Ag, Cu, Sn, Pb, Ni, Co, Fe, Al, and Ca.

c. A metal + a salt

Each metal in the activity series displaces any metals below it to form a salt in solution.

$$Cu + 2\,AgNO_3 \rightarrow Cu(NO_3)_2 + 2\,Ag$$ (Perform in a large test tube. Crystals of Ag will form over time.)

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

Use other metals such as Ag, Sn, Pb, Ni, Co, Fe, Al, Ca, and Mg with solutions of chloride, sulfate, or nitrate salts.

d. A Halogen + halide salt

A halogen (F, Cl, Br, I, At) will displace any less active halogen from a halide salt. The order of activity decreases going from top to bottom down the halogen family in the periodic table.

$$Cl_2 + 2\ NaI \rightarrow 2\ NaCl + I_2$$

$$Cl_2 + 2\ NaBr \rightarrow 2\ NaCl + Br_2$$

$$Br_2 + 2\ NaI \rightarrow 2\ NaBr + I_2$$

Use chlorine water (pale yellow-green) or bromine water (brown) with the clear, colorless Na or K halogen salt solution.

A quick summary of the activity series of the metals: (This is an approximation.)

Most reactive	Group I
	Group II
	Group III (Al is the most common element we use.)
	Transition elements (Mainly the first row except for Cu)
	Group IV (Sn and Pb are the most common elements)
	Hydrogen
	Cu
	Hg (used for dental amalgams)
Least reactive	Precious metals used for jewelry (Ag, Au, Pt)

Metathesis Reactions

A metathesis reaction is a double displacement reaction that usually occurs in water solution. The general form of a metathesis reaction is:

$$AB + CD \rightarrow AD + CB$$

In order to have any appreciable degree of completion of metathesis reactions, one or both of the products must become unavailable for the reverse reaction. the principal conditions that favor the completion of these reactions are:

(1) Formation of an insoluble compound - a precipitate
(2) Formation of a gas
(3) Formation of water

Metathesis reactions are generally classified as precipitation reactions or as neutralization reactions.

a. Precipitation Reactions

In this type of reaction, two compounds which are water soluble react to form two new compounds, one of which is a precipitate (i.e. insoluble in water). The precipitate is often indicated by an arrow pointing downward, ↓, written next to its formula.

$$AgNO_3 + NaCl \rightarrow AgCl\downarrow + NaNO_3$$

$$BaCl_2 + K_2SO_4 \rightarrow BaSO_4\downarrow + 2\ KCl$$

Repeat these reactions using different salt solutions. Choose combinations that produce colored precipitates such as $Pb(NO_3)_2$ and KI. In cases where a lead salt, or other heavy metal salt considered "hazardous", only 1 or 2 drops of that salt solution are required reducing the need for expensive hazardous waste disposal.

In order to determine which one of the products will be the precipitate requires a knowledge of the solubilities of salts in water. The rules governing the solubility of common salts are given below:

THE SOLUBILITY RULES

1. All sodium, potassium, and ammonium salts are soluble in water.
2. The nitrates, chlorates, and acetates of all metals are soluble in water. Silver acetate is sparingly soluble.
3. The chlorides, bromides, and iodides of all metals except lead, silver, and mercury(I) are soluble in water. $PbCl_2$, $PbBr_2$, and PbI_2 are soluble in hot water.
4. The sulfates of all metals except lead, mercury(I), barium, and calcium are soluble in water. Ag_2SO_4 is slightly soluble.
5. The carbonates, phosphates, borates, sulfites, chromates, and arsenates of all metal except sodium, potassium, and ammonium are insoluble in water.
6. The sulfides of all metals except barium, calcium, magnesium, sodium, potassium, and ammonium are insoluble in water.
7. The hydroxides of sodium, potassium, and ammonium are very soluble in water. The hydroxides of calcium and barium are moderately soluble. The oxides and hydroxides of all other metals are insoluble.

A gaseous product can be observed by reacting sodium carbonate or sodium bicarbonate with an acid.

$$Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2$$

The product, H_2CO_3 , decomposes to H_2O and CO_2

b. Neutralization Reactions (sometimes called acid-base reactions)

A neutralization reaction occurs between an acidic compound and a basic compound to form a chemical salt and water.

1. **Reaction between an acid and a base**

$$HCl + NaOH \rightarrow NaCl + H_2O$$

$$H_2SO_4 + Mg(OH)_2 \rightarrow MgSO_4 + H_2O$$

Note: Use phenolphthalein indicator to show end point.

2. **Reaction between a metal oxide and an acid**.

When oxides of many metals are added to water, bases are formed.

$$CaO + H_2O \rightarrow Ca(OH)_2$$

Generally, these metal oxides are called basic anhydrides and they act like bases when mixed with acids.

$$CaO + 2\ HCl \rightarrow CaCl_2 + H_2O$$

$$Na_2O + 2\ HNO_3 \rightarrow 2\ NaNO_3 + H_2O$$

3. **Reaction between a non-metal oxide and a base**.

Many non-metal oxides are classified as acid anhydrides.
These form acids when mixed with water.

$$SO_2 + H_2O \rightarrow H_2SO_3$$

Non-metal oxides act as acids when mixed with a base.

$$SO_2 + 2\ NaOH \rightarrow Na_2SO_3 + H_2O$$

$$CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$$

Combustion Reactions

Combustion reactions generally apply to organic compounds, such as hydrocarbons, which are used as fuels. In these cases, the compound is being burned in air (or oxygen) and producing carbon dioxide and water as products. A general form for a combustion reaction is:

$$C_nH_{2n+2} + \left(\frac{3n+1}{2}\right) O_2 \rightarrow n\ CO_2 + (n+1)\ H_2O$$

Note: The actual coefficients will vary based on the composition of the starting compound.

Some examples of combustion reactions are:

$$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

$$2\ C_4H_{10} + 9\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

$$C_2H_5OH + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$$

Use several hydrocarbon liquids. B small quantities on watch glasses. The longer the chain, the more yellow the flame color and the more soot will be produced. Alcohols are considered to be "oxygenated" fuels and burn with a bluer flame.

Safety note: Keep containers of the hydrocarbons sealed and away from the combustion reactions. Do not try to add additional hydrocarbon to any watch glass.