CHM 130LL: Chemical Reactions

Introduction

We often study chemistry to understand how and why chemicals (reactants) can be transformed into different chemicals (products) via a chemical reaction:

Reactants $\rightarrow$ **Products**

the starting materials,
also called "reagents"

TYPES OF REACTIONS

Many aspects of our lives actually involve chemical reactions—from batteries that power our cars and radios to the thousands of processes occurring within our bodies. We cannot even begin to identify the millions of chemical reactions occurring around us all the time. For now, it will be sufficient for us to focus on the following six types of chemical reactions: **combination reactions**, decomposition reactions, **single-replacement reactions, double-replacement/ precipitate reactions, acid-base neutralization reactions**, and **combustion reactions**.

Combination Reactions

Combination Reactions: A + Z $\rightarrow$ AZ

In a combination reaction, two or more reactants combine to form a single product. The most common type of combination reaction involves a metal reacting with a nonmetal to produce a solid ionic compound. For example, sodium metal reacts with chlorine gas to produce solid sodium chloride:

$$Na(s) + Cl_2(g) \rightarrow NaCl(s)$$

To **balance the equation**, we get the same number of Cl atoms by putting a coefficient of 2 in front of NaCl*(s)*, then we put a 2 in front of the Na*(s)* to balance Na:

$$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$$

If you were carrying out this reaction in a laboratory, you will observe shiny sodium metal reacting with yellow chlorine gas to produce a white solid.

Decomposition Reactions

Decomposition: AB $\rightarrow$ A + B

In a decomposition reaction, a single compound breaks down into two or more substances. In general, decomposition reactions occur when a solid compound is heated. This type of reaction almost always produces a gas. For example, heating mercury (II) oxide produces oxygen gas:

$$2\ HgO\ (s) \rightarrow 2\ Hg\ (l) + O_2\ (g)$$

Similarly, heating sodium hydrogen carbonate produces sodium carbonate, carbon dioxide gas, and steam, H_2O (g):

$$2\ NaHCO_3\ (s) \rightarrow Na_2CO_3\ (s) + CO_2\ (g) + H_2O\ (g)$$

Aqueous Solutions(aq)

Note that many reactions occur in an aqueous environment (i.e., in a solution of ions and compounds dissolved in water). When a reactant or product has the physical state "*(aq)*", that substance is dissolved in water. An ionic compound in aqueous solution exists as individual ions in water; for example, NaCl*(aq)* exists as Na^+ and Cl^- ions moving around in water.

Single-Replacement Reactions

Single-Replacement Reactions: $\mathbf{A + BZ \rightarrow AZ + B}$

In a single-replacement reaction, a more "active" metal displaces metal cation or hydrogen ion in an aqueous solution of water. In general, the greater its metallic character, the more active the metal. To determine the relative reactivity of a specific metal, we refer to the **Activity Series**:

Activity Series

Activity Series for Metals

Li > K > Ba > Sr > Ca > Na > Mg > Al > Mn > Zn > Fe > Cd > Co > Ni > Sn > Pb > (H) > Cu > Ag > Hg > Au

In the Activity Series, the metals are arranged in order of their ability to undergo reaction. The most reactive metals appear first, and the least active metals are at the end of the list. We can predict if a reaction will occur by comparing the solid metal with the metal ion (or hydrogen) in solution. If the solid metal is more active, it will displace the metal ion or hydrogen in solution, so a reaction occurs. If the solid metal is less active than the metal ion or hydrogen in solution, then no reaction occurs.

Let us consider the reaction between aluminum metal and copper (II) chloride. We can represent the reactants as follows:

$$Al(s) + CuCl_2(aq) \rightarrow$$

If we refer to the Activity Series, we see that aluminum is more active—i.e., is higher on the Activity Series—than copper, so aluminum can displace copper from the solution. The solid aluminum goes into solution as aluminum ion, Al^{3+}, and the copper (II) ions, Cu^{2+}, come out of solution as solid copper, Cu*(s)*. We can write the complete equation as

$$Al(s) + CuCl_2(aq) \rightarrow Cu(s) + AlCl_3(aq)$$

Note that the ionic compound formed by aluminum and chloride ions is $\mathbf{AlCl_3}$ because aluminum ion is Al^{3+} and chloride ion is Cl^-, so based on their charges, aluminum chloride is $AlCl_3$. **Remember that the formula for the compounds formed in a chemical reaction is always based on the charges of the ions that make up the compound!**

Finally, we balance the reaction by getting the same number of each atom on both sides of the equation:

$$2\ Al(s) + 3\ CuCl_2(aq) \rightarrow 3\ Cu(s) + 2\ AlCl_3(aq)$$

If we carry out this reaction by placing a piece of aluminum metal in blue copper (II) chloride solution, we will see dark granules of copper forming at the bottom of the beaker.

Let us consider a different reaction, this time between silver metal and copper (II) chloride. We can represent the reactants as follows:

$$Ag(s) \quad + \quad CuCl_2(aq) \quad \rightarrow$$

If we refer to the Activity Series, we find that silver is less active—i.e., is lower on the Activity Series—than copper, so it cannot displace copper.

Consequently, no reaction occurs. We represent this by indicating **NR** after the reaction arrow:

$$Ag(s) \quad + \quad CuCl_2(aq) \quad \rightarrow \quad NR$$

Let us consider yet another reaction, this time between cadmium metal and hydrochloric acid. We can represent the reactants as follows:

$$Cd(s) \quad + \quad HCl(aq) \quad \rightarrow$$

If we refer to the Activity Series, we see that cadmium is more active—i.e., is higher on the Activity Series—than hydrogen, so cadmium can displace hydrogen from the solution. The solid cadmium goes into solution as cadmium ion, Cd^{2+}, and the hydrogen ions, H^+, come out of solution as hydrogen gas, $H_2(g)$, since hydrogen is one of the elements that exists as a diatomic molecule.

Diatomic Molecules

Note: Seven elements exist naturally as ***diatomic molecules***. These are **H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2**. Thus, in a chemical equation hydrogen is written as H_2, not simply as H.

The complete equation for the reaction is written as:

$$Cd(s) \quad + \quad HCl(aq) \quad \rightarrow \quad CdCl_2(aq) \quad + \quad H_2(g)$$

Note that the ionic compound formed by cadmium and chloride ions is **$CdCl_2$** because cadmium ion is Cd^{2+} and chloride ion is Cl^-, so based on their charges, cadmium chloride is $CdCl_2$.

Finally, we balance the reaction by getting the same number of each atom on both sides of the equation:

$$Cd(s) \quad + \quad 2\,HCl(aq) \quad \rightarrow \quad CdCl_2(aq) \quad + \quad H_2(g)$$

If we carry out this reaction by placing a piece of cadmium metal in colorless hydrochloric acid, we will see bubbles of hydrogen gas forming from the cadmium sitting in the solution.

Consider yet another reaction, one between copper metal and hydrochloric acid. We can represent the reactants as follows:

$Cu(s) + HCl(aq) \rightarrow$

If we refer to the Activity Series, we see that copper is less active—i.e., is lower on the Activity Series—than hydrogen, so copper cannot displace hydrogen from the solution. Consequently, no reaction occurs. We represent this by indicating **NR** after the reaction arrow:

$Cu(s) + HCl(aq) \rightarrow$ NR

If we were to carry out this experiment in the laboratory and place a piece of copper in a test tube containing hydrochloric acid, the copper would simply sit there, and nothing would happen.

Many ionic compounds are ***soluble*** in water—i.e., dissolve in water; for example, table salt (NaCl) is soluble in water. To determine if an ionic compound is soluble—i.e., will dissolve—in water, we use the Solubility Rules:

Solubility Rules

Solubility Rules for Ionic Compounds in Water

The **compound is SOLUBLE** if it has:	The **compound is INSOLUBLE** if it has:
1. Li^+, Na^+, K^+, NH_4^+ ions (**ALWAYS!**)	6. Carbonate ion, CO_3^{2-}, **but** Rule 1 ions
2. Acetate ion, $C_2H_3O_2^-$	7. Chromate ion, CrO_4^{2-}, **but** Rule 1 ions
3. Nitrate ion, NO_3^-	8. Phosphate ion, PO_4^{3-}, **but** Rule 1 ions
4. Halide ions (X^-): Chloride ion (Cl^-), bromide ion (Br^-), or iodide ion (I^-), ***but*** AgX, PbX_2, Hg_2X_2 are **insoluble**	9. Sulfide ion, S^{2-}, **but** Rule 1 ions, CaS, SrS, BaS are **soluble**
5. Sulfate ion (SO_4^{2-}), ***but*** $CaSO_4$, $SrSO_4$, $BaSO_4$ are **insoluble**	10. Hydroxide ion, OH^-, **but** Rule 1 ions, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$ are **soluble**

Compounds containing ions included in the left column are generally **soluble**, with a few exceptions. Compounds containing ions included in the right column are generally insoluble, again with a few exceptions. The Solubility Rules allow us to determine which compounds are soluble, and we represent these compounds as aqueous: e.g., $KI(aq)$, $BaCl_2(aq)$, $NaOH(aq)$, etc. The Solubility Rules also tell us which compounds are **insoluble**—i.e., do not dissolve in water and remain as solids: e.g. $BaSO_4(s)$, $AgCl(s)$, $CaCO_3(s)$, etc.

Double-Replacement/ Precipitation Reactions

Double-Replacement/Precipitation Reactions: AX + BZ → AZ + BX

In a double-replacement/precipitation reaction, two ionic compounds in aqueous solution swap metals to form a ***precipitate***, a solid formed when two ionic solutions are mixed.

Let us consider the reaction between aqueous lead (II) nitrate with aqeous potassium bromide. We represent the reactants as follows

$Pb(NO_3)_2(aq)$ + $KBr(aq)$ →

When the metals are swapped, lead (II) ion, Pb^{2+}, combines with bromide ion, Br^-, to form lead (II) bromide, $PbBr_2$, and potassium ion, K^+ combines with nitrate ion, NO_3^-, to form potassium nitrate, KNO_3. We can represent the reactants and products as follows:

$$Pb(NO_3)_2(aq) + KBr(aq) \rightarrow PbBr_2 + KNO_3$$

Note that the chemical formulas for the products formed are based on their charges, not how they appear on the reactant side of the chemical equation.

Next, we refer to the Solubility Rules to determine if $PbBr_2$ and KNO_3 are aqueous or solid. Based on Rule #4, $PbBr_2$ is insoluble, so we indicate it as $PbBr_2(s)$. Based on Rule #1, any compound with K^+ is soluble, so we represent potassium nitrate as $KNO_3(aq)$. Thus, the complete unbalanced equation is:

$$Pb(NO_3)_2(aq) + KBr(aq) \rightarrow PbBr_2(s) + KNO_3(aq)$$

Because a precipitate formed—the $PbBr_2(s)$ in this case, then a reaction occurred. We balance the equation by reconciling the number of atoms and/or polyatomic ions on both sides of the equation. The balanced equation is:

$$Pb(NO_3)_2(aq) + 2\ KBr(aq) \rightarrow PbBr_2(s) + 2\ KNO_3(aq)$$

Note that only the lead (II) ions and bromide ions react to form the lead (II) bromide precipitate. The potassium and nitrate ions exist as individual ions before and after the reaction; because they themselves did not react, they are called **"*spectator ions*,"** as if they simply watched the other ions reacting.

The lead (II) nitrate and potassium bromide solutions are both clear and colorless. When they are combined in the lab, the resulting mixture is cloudy, indicating a solid is present and small particles of the solid are suspended in the solution. When enough solid has formed, it will begin to settle at the bottom of the beaker. Thus, a clear solution becoming cloudy when another solution is added is often taken as experimental evidence of a precipitate forming.

Acids and Bases

Acids are compounds that ***release hydrogen ions (H^+),*** also called ***protons***, when dissolved in water. The chemical formulas for acids are most often given with the H's at the beginning, so acids are usually easy to recognize. A few common acids are hydrochloric acid, $HCl(aq)$, nitric acid, $HNO_3(aq)$, and sulfuric acid, $H_2SO_4(aq)$.

For now, ***bases*** are compounds that ***release hydroxide ions, OH^-,*** when dissolved in water. A few common bases are sodium hydroxide, $NaOH(aq)$, potassium hydroxide, $KOH(aq)$, and calcium hydroxide, $Ca(OH)_2$.

Acid-Base Neutralization Reactions

Acid-Base Neutralization Reactions

In an acid-base neutralization reaction, an acid reacts with a base to produce water and a salt:

$$\underset{\text{acid}}{HX(aq)} + \underset{\text{base}}{MOH(aq)} \rightarrow H_2O(l) + \underset{\text{salt}}{MX(aq)}$$

The ***hydrogen ion (H^+)*** from the acid reacts with the ***hydroxide ions (OH^-)*** from the base to form the water. The **salt** forms from the ***cation from the base*** and the ***anion from the acid***. Because water is always formed, acids will always react with bases, regardless of whether the salt is soluble or insoluble.
Let us consider the reaction between phosphoric acid and potassium hydroxide. We represent the reactants as follows

$$H_3PO_{4(aq)} + KOH(aq) \rightarrow$$

We know that the hydrogen ions, H^+, from the phosphoric acid, $H_3PO_{4(aq)}$, combine with the hydroxide ions, OH^-, from the KOH*(aq)* to produce water. The remaining ions, the potassium ions, K^+, from the KOH*(aq)* and the phosphate ions, PO_4^{3-}, from the $H_3PO_{4(aq)}$ combine to form the salt, K_3PO_4. We refer to the Solubility Rules to determine if K_3PO_4 is soluble or insoluble, and based on Rule #1, it is soluble, so we represent it as $K_3PO_4(aq)$. Thus, the complete unbalanced chemical equation is shown as:

$$H_3PO_{4(aq)} + KOH(aq) \rightarrow H_2O(l) + K_3PO_4(aq)$$

We balance the equation by reconciling the number of atoms and/or polyatomic ions on both sides of the equation. Thus, the balanced chemical equation is:

$$H_3PO_{4(aq)} + 3\ KOH(aq) \rightarrow 3\ H_2O(l) + K_3PO_4(aq)$$

When carrying out an acid-base neutralization reaction in the laboratory, one observes that most acid solutions and base solutions are colorless, and the resulting water and soluble salt solutions are also colorless. Thus, it is impossible to determine if a reaction has occurred. To help us monitor acid-base reactions, we use an ***acid-base indicator***, a solution that changes color depending on the pH (or acid-content) of the solution. One commonly used indicator is *phenolphthalein*, which is colorless in acidic and neutral solutions and pink in basic (or alkaline) solution. Thus, one can determine whether a solution is acidic or basic using phenolphthalein.

One can also use litmus paper to determine if a solution is acidic or basic. Blue litmus paper turns red in acidic solution, and red litmus paper turns blue in basic solution.

Combustion Reactions

Combustion Reactions: C_xH_y + $O_2(g)$ → $CO_2(g)$ + $H_2O(g)$

In a combustion reaction, a **hydrocarbon,** C_xH_y (a compound containing only carbon and hydrogen atoms), or a **hydrocarbon derivative,** $C_xH_yO_z$ (a compound containing carbon, hydrogen, and oxygen atoms), burns in oxygen to produce carbon dioxide gas, $CO_2(g)$, and steam, $H_2O(g)$. (Steam is formed rather than water because the reactions occur at high temperatures.)

For example, when butanol, $C_4H_9OH(l)$, burns in oxygen present in air, the products are carbon dioxide gas and steam:

$$C_4H_9OH(l) \quad + \quad O_2(g) \xrightarrow{\Delta} \quad CO_2(g) \quad + \quad H_2O(g)$$

To balance the equation, first balance carbons, then hydrogens, and finally the oxygens. Thus, the balanced equation is:

$$C_4H_9OH(l) \quad + \quad 6\ O_2(g) \xrightarrow{\Delta} \quad 4\ CO_2(g) \quad + \quad 5\ H_2O(g)$$

If we carry out this reaction in the laboratory, we see when lit with a flaming splint, liquid butanol burns and appears to completely disappear as the products are both colorless gases.

Balancing Chemical Equations

Chemical equations are written to represent chemical reactions, and these equations must be balanced to show that atoms are never created or destroyed in a chemical reaction but simply rearranged. Balancing a chemical equation means reconciling the number of atoms of each element on both the reactant and product sides of the equation. In general, using the following order when balancing equations will make the process as easy as possible:

- Metals
- Polyatomic ions (balance as a whole unit, not as individual atoms)
- Carbon
- Hydrogen
- Oxygen
- All other elements

For most of the reactions observed in this experiment, a chemical equation will be balanced. Additional chemical equations will also be balanced and identified as one of the following types of reactions:

- Combination reaction (C)
- Decomposition reaction (D)
- Single-Replacement reaction (SR)
- Double-Replacement/Precipitation reaction (DR)
- Acid-Base Neutralization reaction (N)
- Combustion reaction (B)

Procedure

NOTE: All waste for this experiment should go into the appropriate waste containers in the hood. *Do not dispose of any solids down the drain.* Carefully extinguish all burning or smoldering splints in a beaker of water, then place in the waste container in the hood.

I. Combination Reactions

a. Obtain a a piece of copper wire, Cu*(s)*. Holding the copper metal with your crucible tongs, heat it in a Bunsen burner flame. Record your observations on your Lab Report form.

CAUTION: The burning of Magnesium metal results in an extremely bright light. Hold the magnesium metal away from you, and do not look directly at the piece when heating the magnesium.

b. Repeat step a with a small ribbon of magnesium metal, Mg*(s)*. Record your observations on your Lab Report form.

Waste disposal: Dispose of used copper and magnesium into the waste container in the hood, NOT in the trash.

II. Decomposition Reactions

a. Place approximately a pea-size amount of copper(II) carbonate, $CuCO_3$*(s)*, in a dry test tube. Observe the color of the sample. Using a test tube clamp, heat the test tube over a Bunsen burner until you notice a color change. **DO NOT** *return the hot test tube to the plastic test tube rack, or it will melt the plastic.* Instead, place the test tube in an empty beaker to cool. Record the color of the solid sample after heating.

Waste disposal: When cool, dispose of used copper(II) carbonate into the waste container in the hood, NOT in the trash.

III. Single-Replacement Reactions

a. Place about 10 drops of 0.1 M copper (II) nitrate, $Cu(NO_3)_2$*(aq)*, in a clean small test tube. Note the color of the solution on your Lab Report. Now add a few pieces of zinc mesh (or granules), and let the solution sit for a few minutes. Note any changes in the color of the solution and the formation of any solid. Record your observations.

CAUTION: Silver nitrate solution, $AgNO_3(aq)$, stains skin and clothing. Wash any spilled silver nitrate immediately with plenty of water.

b. Clean a short (4-5 cm) piece of Cu wire with sandpaper. Coil the Cu wire and place it into a clean small test tube. Add enough 0.1 M silver nitrate, $AgNO_3$*(aq)*, to completely cover the copper wire, and allow the mixture to sit for 5-10 minutes. Record the appearance of the copper wire and the silver nitrate solution *before* any reaction. Record the appearance of the copper wire after sitting in the silver nitrate solution for 5-10 minutes, and record the appearance of the solution.

CAUTION: Hydrochloric acid, HCl(aq), is toxic and corrosive, so it can cause chemical burns on contact. Any HCl(aq) spilled on skin must be washed immediately with plenty of water for at least 10 minutes.

c. Drop 2-3 pieces of magnesium turnings in a test tube. Add enough 0.1 M hydrochloric acid, HCl*(aq)* to completely cover the magnesium. Record your observations on your Data Sheet.

IV. Double-Replacment/Precipitation Reactions

Caution: Lead (II) nitrate solutions are toxic. Avoid contact with skin, eyes, and clothing. If the solution contacts your skin, rinse with water immediately.

a. Add ~10 drops of 0.1 M lead (II) nitrate, $Pb(NO_3)_2$*(aq)*, to a clean small test tube. Add an equal amount of 0.1 M potassium iodide, KI*(aq)*. Mix the solutions, and record your observation.

b. Silver nitrate, $AgNO_3$, is very sensitive to the presence of chloride ions and is used to detect chloride ions, Cl^-, in aqueous solution. Add ~10 drops of 0.1 M sodium chloride, NaCl*(aq)*, in a clean small test tube. Next add a few drops of 0.1 M $AgNO_3$, and record your observation on your Data Sheet.

V. Acid-Base Neutralization Reactions

Caution: Hydrochloric acid is corrosive and can cause chemical burns as well as damage clothing. Any hydrochloric acid spilled on skin must be rinsed immediately with water for 15 minutes. Any acid spilled on your work area must be neutralized, then the entire area should be washed and dried.

a. Blue litmus paper turns red in an acidic solution, so it is used to identify acidic substances. Place a piece of blue litmus paper on a watch glass. Add 1 drop of 0.1 M hydrochloric acid, HCl*(aq)*, and record your observations.

Caution: Sodium hydroxide (NaOH) can easily damage eyes. It is corrosive and can cause chemical burns and damage clothing. Any NaOH splashed into eyes or spilled on skin must be rinsed immediately with water for 15 minutes. Any base spilled on your work area must be neutralized, then the entire area should be washed and dried.

b. Red litmus paper turns blue in a basic solution, so it is used to identify basic substances. Place a piece of red litmus paper on a watch glass. Add 1 drop of 0.1 M sodium hydroxide, NaOH*(aq)*, and record your observations.

c. Phenolphthalein is an acid-base indicator that is colorless in acidic and neutral solutions and pink in basic (or alkaline) solution. Fill a small clean test tube about 1/5 full with distilled water. Add a drop of phenolphthalein indicator. Next add 1-2 drops of 0.1 M hydrochloric acid, HCl*(aq)*, stir with a glass stirring rod, and record your observations.

d. Fill a small clean test tube about 1/5 full with distilled water. Add a drop of phenolphthalein indicator. Next add 1-2 drops of 0.1 M sodium hydroxide, NaOH*(aq)*, stir with a stirring rod, and record your observations.

e. Fill a small clean test tube about 1/5 full with distilled water. Add 5 drops of 0.1 M hydrochloric acid, HCl*(aq)*. Add 1 drop of phenolphthalein, and stir with a stirring rod. Add 0.1 M sodium hydroxide, NaOH*(aq)*, to this mixture one drop at a time, *mixing the solution with your glass stirring rod between additions*. Count how many drops of sodium hydroxide are needed for a permanent color change. Record your observations.

Caution: ***Sulfuric acid is corrosive and can cause chemical burns and damage clothing. Any sulfuric acid spilled on skin must be rinsed immediately with water for 15 minutes. Any acid spilled on your work area must be neutralized, then the entire area should be washed and dried.***

f. Repeat step e, substituting 0.1 M sulfuric acid, $H_2SO_{4(aq)}$ for 0.1 M hydrochloric acid, HCl*(aq)*. Again, count how many drops of sodium hydroxide are needed for a permanent color change. How many drops of NaOH*(aq)* were required to neutralize the sulfuric acid? Record your observations.

VI. Combustion Reactions

a. Add 15 drops of ethanol into a clean ceramic evaporating dish. Use your Bunsen burner to light a wood splint, then light the ethanol with the splint. Allow the ethanol to burn completely. Record your observations, including how the evaporating dish appears after the ethanol burns. Wash and dry the evaporating dish.

b. Add 15 drops of isopropyl alcohol into the clean evaporating dish. Light the isopropyl alcohol with the splint, and allow it to burn completely. Record your observations, including how the evaporating dish appears after the isopropyl alcohol burns.

NOTE: All waste for this experiment should go into the waste containers in the hood. Extinguish burning or smoldering splints in a beaker of water, then dispose of splints in the waste container in the hood. Wash your entire work area with a wet paper towel, then dry. Wash your hands completely with soap and water before you leave.

VII. Balance and identify the reaction type for the equations on page 13.

CHM 130LL: Chemical Reactions

Name: ______________________

Partner: ______________________

Section Number: ______________

LAB REPORT

I. Combination Reactions

	Observations
a. Copper metal, Cu(s), is heated.	
b. Magnesium metal, Mg(s), is heated.	

Balance the following equations representing the reactions observed:

a. _____ Cu*(s)* + _____ O_2*(g)* $\xrightarrow{\Delta}$ _____ CuO*(s)*

b. _____ Mg*(s)* + _____ O_2*(g)* $\xrightarrow{\Delta}$ _____ MgO*(s)*

Note: The reaction arrow with a triangle ($\xrightarrow{\Delta}$) indicates reactants were heated.

II. Decomposition Reactions

	Observations
a. Copper (II) Carbonate, $CuCO_3$(s), is heated.	

Balance the following equation representing the reaction observed:

a. _____ $CuCO_3$*(s)* $\xrightarrow{\Delta}$ _____ CuO*(s)* + _____ CO_2*(g)*

III. Single-Replacement Reactions

	Observations
a. Zinc (Zn) metal in $Cu(NO_3)_2$(aq)	Original color of **$Cu(NO_3)_2$(aq).**: ____________________ Zinc metal before reaction: __________________________ Color of **solution** after reaction: ____________________ Zinc metal after reaction: ___________________________
b. Copper (Cu) metal in $AgNO_3$(aq)	Original color of **$AgNO_3$(aq)**: ______________________ Copper wire before reaction: _________________________ Color of **solution** after reaction: ____________________ Copper wire after reaction: __________________________
c. Magnesium (Mg) metal in HCl*(aq)*	

Balance the following equations representing the reactions observed:

a. _____ $Zn(s)$ + _____ $Cu(NO_3)_2(aq)$ → _____ $Cu(s)$ + _____ $Zn(NO_3)_2(aq)$

b. _____ $Cu(s)$ + _____ $AgNO_3(aq)$ → _____ $Ag(s)$ + _____ $Cu(NO_3)_2(aq)$

c. _____ $Mg(s)$ + _____ $HCl(aq)$ → _____ $H_2(g)$ + _____ $MgCl_2(aq)$

IV. Double-Replacement/Precipitation Reactions

	Observations
a. $Pb(NO_3)_2(aq) + KI(aq)$	
b. $NaCl(aq) + AgNO_3(aq)$	

Balance the following equations representing the reactions observed:

a. _____ $Pb(NO_3)_{2(aq)}$ + _____ $KI(aq)$ → _____ $PbI_2(s)$ + _____ $KNO_3(aq)$

b. _____ $NaCl(aq)$ + _____ $AgNO_3(aq)$ → _____ $AgCl(s)$ + _____ $NaNO_3(aq)$

V. Acid-Base Neutralization Reactions

	Observations
a. Blue litmus paper + $HCl(aq)$	
b. Red litmus paper + $NaOH(aq)$	
c. $HCl(aq)$ + phenolphthalein	Color of the solution: ________________________
d. $NaOH(aq)$ + phenolphthalein	Color of the solution: ________________________
e. $HCl(aq)$ + phenolphthalein + $NaOH(aq)$	Number of drops of NaOH required for permanent color change: ____________
f. $H_2SO_4(aq)$ + phenolphthalein + $NaOH(aq)$	Number of drops of NaOH required for permanent color change: ____________

Balance the following equations representing the reactions observed:

e. _____ $HCl(aq)$ + _____ $NaOH(aq)$ → _____ $H_2O(l)$ + _____ $NaCl(aq)$

f. _____ $H_2SO_4(aq)$ + _____ $NaOH(aq)$ → _____ $H_2O(l)$ + _____ $Na_2SO_4(aq)$

VI. Combustion Reactions

	Observations
a. Burning ethanol, C_2H_5OH*(l)*	
b. Burning isopropyl alcohol, C_3H_7OH*(l)*	

Balance the following equation representing the burning ethanol reaction:

a. _____ $C_2H_5OH(l)$ + _____ $O_2(g)$ $\xrightarrow{\Delta}$ _____ $CO_2(g)$ + _____ $H_2O(g)$

VII. Balancing Equations and Classifying Chemical Reactions

Balance each of the chemical equations given, and identify each as one of the following:

- Combination reaction (C)
- Decomposition reaction (D)
- Single-Replacement reaction (SR)
- Double-Replacement/Precipitation reaction (DR)
- Acid-Base Neutralization reaction (N)
- Combustion reaction (B)

TYPE

_______ 1. _____ $Al(s)$ + _____ $O_2(g)$ $\xrightarrow{\Delta}$ _____ $Al_2O_3(s)$

_______ 2. _____ $HNO_3(aq)$ + _____ $Ba(OH)_2(aq)$ $\rightarrow$ _____ $H_2O(l)$ + _____ $Ba(NO_3)_2(aq)$

_______ 3. _____ $Mg(s)$ + _____ $Al(NO_3)_3(aq)$ $\rightarrow$ _____ $Al(s)$ + _____ $Mg(NO_3)_2(aq)$

_______ 4. _____ $C_3H_6O(l)$ + _____ $O_2(g)$ $\xrightarrow{\Delta}$ _____ $CO_2(g)$ + _____ $H_2O(g)$

_______ 5. _____ $H_3PO_4(aq)$ + _____ $KOH(aq)$ $\rightarrow$ _____ $H_2O(l)$ + _____ $K_3PO_4(aq)$

_______ 6. _____ $Al(s)$ + _____ $HBr(aq)$ $\rightarrow$ _____ $H_2(g)$ + _____ $AlBr_3(aq)$

_______ 7. _____ $CaCl_2(aq)$ + _____ $K_3PO_4(aq)$ $\rightarrow$ _____ $Ca_3(PO_4)_2(s)$ + _____ $KCl(aq)$

_______ 8. _____ $KHCO_3(s)$ $\xrightarrow{\Delta}$ _____ $K_2CO_3(s)$ + _____ $H_2O(g)$ + _____ $CO_2(g)$

_______ 9. _____ $C_3H_8(g)$ + _____ $O_2(g)$ $\xrightarrow{\Delta}$ _____ $CO_2(g)$ + _____ $H_2O(g)$

_______ 10. _____ $NaOH(aq)$ + _____ $AlCl_3(aq)$ $\rightarrow$ _____ $NaCl(aq)$ + _____ $Al(OH)_3(s)$