VALENCIACOLLEGE

Experiment 5:
Observing and Classifying Reactions

Version 5

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Careful observations of eleven potential chemical reactions will allow determination of products and reaction classification for those that reacted.

Objectives

- Observe the characteristics of reactant substances, the characteristics of product substances, and energy effects to establish whether a chemical reaction has occurred.
- Classify reactions as combination (synthesis), decomposition, single replacement, or double replacement (metathesis).
- Write molecular equations for different types of reactions.
- Write complete ionic equations and net ionic equations.

Learning Outcomes

- Understand the nature of matter and its physical and chemical characteristics.
- Understand the nature and characteristics of chemical bonds.
- Understand and apply the rules of inorganic chemical nomenclature.
- Understand and apply the rules to write chemical formulas of ionic compounds.
- Perform essential lab techniques in laboratory setting.
- Write molecular, ionic, and net ionic equations.
- Balance chemical equations.

Definitions

- **Anion** – a particle with a net negative charge
- **Aqueous** – dissolved in water
- **Cation** – a particle with a net positive charge
- **Combination reaction** – see synthesis reaction
- **Complete ionic equation** – a chemical equation that lists all of the ions present individually
- **Compound** – a substance composed of two or more elements chemically bonded in a fixed, definite proportion
- **Decomposition reaction** – chemical reaction in which a single reactant breaks down into two or more products
- **Double replacement reaction** – chemical reaction where two compounds react, and the positive ions (cations) and the negative ions (anions) of the two reactants switch places, forming two new compounds or products
- **Electrolytes** – substances that conduct electricity when dissolved in water
- **Half-reactions** – chemical equations that show a portion of an oxidation-reduction reaction to include the electrons as either reactants or products

- **Ionic substance** – compound composed of cations and anions chemically bonded through electrostatic attraction
- **Ion** – atom or a group of bonded atoms with a net charge
- **Ionization** – process of gaining or losing electrons to become an ion
- **Law of conservation of mass** – a law that states that in a chemical reaction matter is neither created nor destroyed
- **Metathesis reaction** – see double replacement reaction
- **Molarity, M** – unit of concentration; expressed as moles of solute per liters of solution, mol/L
- **Molecular equation** – a balanced equation that shows complete neutral formulas for all compounds and their phases
- **Molecules** – a group of atoms bonded together, representing the smallest fundamental unit of a molecular compound (also applies to multi-atomic elements such as N_2 or S_8)
- **Moles** – a counting unit (a word that substitutes a counted amount), specifically the amount of material containing 6.022×10^{23} (Avogadro's number) particles
- **Oxidation** – process in which a reactant loses one or more electrons
- **Products** – chemicals formed or produced from a chemical reaction; found on the right side of the arrow in a chemical equation
- **Reactants** – starting chemicals involved in a chemical reaction; found on the left side of the arrow in a chemical equation
- **Reduction** – process in which a reactant gains one or more electrons
- **Single replacement reaction** – chemical reaction in which an element reacts with a compound and takes the place of another element in that compound
- **Solute** – the component in lesser amount in a solution
- **Solution** – a homogeneous mixture of a solute(s) and solvent
- **Solvent** – the major component of a solution
- **Stoichiometry** – the quantitative relationship between substances in a chemical equation, based upon the law of definite proportions and the law of conservation of mass
- **Spectator ions** – ions that do not participate in the reaction and therefore remain in solution as free ions
- **Strong electrolytes** – substances that dissociate/ionize in water completely or almost completely, forming ions; all ionic compounds and strong acids are strong electrolytes
- **Synthesis or combination reaction** – chemical reaction in which two or more reactants combine to form one single chemical compound
- **Weak electrolytes** – substances that dissolve in water but do not completely ionize and therefore only weakly conduct electricity in solution; weak acids and weak bases are weak electrolytes

Introduction

You have been awarded a conservation scientist internship at Laboratorio Scientifico dell'Opificio, in Florence, Italy. Conservation scientists are chemists that "help conservators understand materials and methods artists used to make works of art and, in some instances, they help determine what is original to the work and what was added in previous restorations. They also develop and test new materials - often from the chemical industry - to help conservators preserve artworks."[1]

Figure 1. Fresco "La Leggenda della Vera Croce", painted in the 15th century by Piero della Francesca before and after restoration using the Ferroni-Dini method. [2]

Molecular, Complete Ionic, and Net Ionic Equations

When an artist or restorer prepares art materials, such as grounds (a preparatory layer put on the support before a paint is applied), and paints, which consist of binders and pigments, they become immediately involved in chemical change. Chemical changes deal with changes in the structure of substances, the **reactants**. Bonds between atoms in reactants are broken and atoms are rearranged and bonded into new substances called **products**. In the process of chemical change, atoms are conserved. That is, all the atoms present in the reactants are present in the products, a requirement of the law of conservation of matter. A chemical change can be written as a chemical equation.

$$Na_3PO_4(aq) + 3AgNO_3(aq) \rightarrow Ag_3PO_4(s) + 3NaNO_3(aq) \qquad \textit{Equation 1}$$

A **chemical equation** is simply a shorthand way of describing a chemical reaction. Equation 1 could be written out in a sentence that would read: "1 mole of aqueous sodium phosphate reacts with 3 moles of aqueous silver nitrate, to produce one mole of solid silver phosphate and 3 moles of aqueous sodium nitrate." The chemical equation says the same thing in a lot less space! If you examine this equation closely and count the number of atoms on each side of the arrow, you should see that this is also a **balanced chemical equation.**

Equation:	*Na*	*P*	*O*	*Ag*	*N*
Left side	3	1	13	3	3
Right side	3	1	13	3	3

A balanced chemical equation obeys the law of conservation of mass. It includes several important things:

1. The correct formulas for **reactants** and correct formulas for **products.**
2. An arrow to mean produces or yields.
3. The correct ratio (stoichiometry) in which the atoms, molecules, or moles of starting materials react with each other.
4. The correct stoichiometry of atoms, molecules, or moles of products formed.
5. Additional information might be included such as the substance's state written in parenthesis to the right of the substance (g - gas, l - liquid, s – solid, aq - aqueous) or special conditions placed above the arrow (Δ - heat, c - catalyst).

Equation 1 is also called a **molecular equation** because it shows complete neutral formulas for all the compounds.

Since aqueous ionic compounds and aqueous strong acids (Appendix 4) are strong electrolytes, they exist as virtually 100% dissociated in water. Therefore, a **complete ionic equation**, an equation that shows all the ions present, is a truer representation of the reaction. Equation 2 is the complete ionic equation for Equation 1.

$$3\,Na^{+}(aq) + PO_4^{3-}(aq) + 3\,Ag^{+}(aq) + 3\,NO_3^{-}(aq) \rightarrow Ag_3PO_4(s) + 3\,Na^{+}(aq) + 3\,NO_3^{-}(aq) \qquad \textit{Equation 2}$$

Notice that the three sodium ions and the three nitrate ions appear the same on both sides of the equation. Since they have not actually reacted, they are spectator ions.

A **net ionic equation** only includes the ions that actually react and the product(s) they form. Equation *3* is the net ionic equation for Equation 1.

$$PO_4^{3-}(aq) + 3Ag^{+}(aq) \rightarrow Ag_3PO_4(s) \qquad \textit{Equation 3}$$

To prepare for this internship, you need to recognize some general types of reactions, and how to write molecular, ionic, and net ionic equations. You also need to learn Italian, so get busy!

General Types of Reactions

We often classify reactions. Though there are many types of reactions, the most common types of reactions include: (1) synthesis or combination reactions, (2) decomposition reactions, (3) single replacement reactions, and (4) double replacement or metathesis reactions.

1. **Synthesis or Combination Reactions.** A synthesis reaction occurs when two or more chemicals react to form one single chemical compound.

   Metals often combine with non-metals to form compounds. Heating may be necessary to start reaction. Paleolithic artists used the product of Equation 4, MnO_2, as a black pigment[3]; it is also used as a glass additive to impart pink color to the glass.[4]

   $$2\,\mathrm{Mn(s)} + \mathrm{O_2(g)} \rightarrow 2\,\mathrm{MnO_2(s)} \qquad \textit{Equation 4}$$

   Nonmetals may combine with oxygen to form binary compounds called nonmetal oxides:

   $$S_2(s) + 3\,O_2(g) \rightarrow 2\,SO_3(g) \qquad \textit{Equation 5}$$

   Synthesis reactions are not limited to elements; compounds also participate in synthesis reactions. Nonmetal oxides (such as CO_2, SO_2, SO_3) are a serious hazard to artwork because they react with water to form acids. This combination reaction is illustrated in the formation of sulfuric acid in *Equation 6* below.

   $$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq) \qquad \textit{Equation 6}$$

2. **Decomposition Reactions.** A decomposition reaction is a reaction in which a single reactant breaks down into two or more products.

   For example, a number of solid substances will decompose to form simpler substances when heated, as seen in Equation 7. Many Renaissance painters used $Cu(OH)_2$ as the pigment in pale blue paint. Many of these paintings now have black spots because the decomposition product of $Cu(OH)_2$, CuO, is black.

$$Cu(OH)_2 \rightarrow CuO(s) + H_2O(g) \qquad \textit{Equation 7}$$

3. **Single Replacement Reactions.** A single replacement reaction is a reaction in which an element reacts with a compound and takes the place of another element in that compound, as depicted in Figure 2.

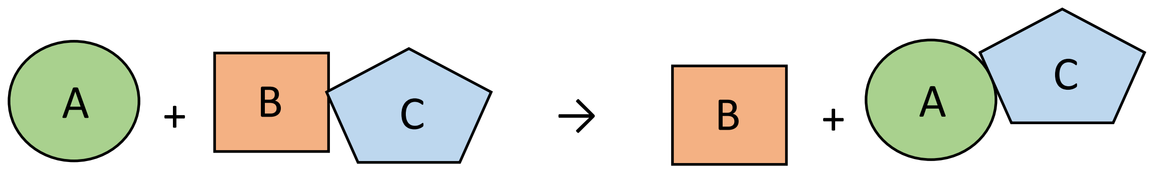

Figure 2. Representation of a Single Replacement Reaction

   When a metal is placed in a solution containing ions of a second metal that is less reactive than it, the first metal donates electrons to the ions of the second metal in the solution and ions of the second metal will be converted to the free metal.

   Equation 8a is an example of a single replacement reaction. The change in charge, which is easily observed in the net ionic equation (Equation 8b), is due to a transfer of electrons. Reactions in which electrons are transferred from one substance to another are called oxidation-reduction reactions or redox reactions.

$$Zn(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Zn(NO_3)_2\,(aq) \qquad \textit{Equation 8a}$$

$$\mathrm{Zn(s)} + 2\,\mathrm{Ag^+(aq)} \rightarrow 2\,\mathrm{Ag(s)} + \mathrm{Zn^{2+}(aq)} \qquad \textit{Equation 8b}$$

   To be able to "see" the electrons transferred, let's split Equation 8b into two half reactions, the oxidation half-reaction and the reduction half-reaction:

oxidation half-reaction $\quad Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$ $\quad$ *Equation 8c*

reduction half-reaction $\quad 2\,Ag^+ + 2\,e^- \rightarrow 2\,Ag(s)$ $\quad$ *Equation 8d*

   Notice that Equation 8a and Equation 8b do not show the electrons because the number of electrons lost has to be the same as the number of electrons gained - in other words, electrons are transferred. It is necessary to have a coefficient of 2 in front of the silver metal and silver ion in Equation 8a and Equation 8b to allow the two electrons lost by zinc to be completely transferred since one silver cation only accepts one electron.

   Metals more reactive than hydrogen will react with hydrogen ions in solution to produce hydrogen gas.

$$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g) \qquad \textit{Equation 9}$$

4. **Double Replacement or Metathesis Reactions.** Many chemicals, when dissolved in water, form solutions containing positive and negative ions. If two such solutions are mixed, a reaction will occur if ions from the two solutions will combine with each other.

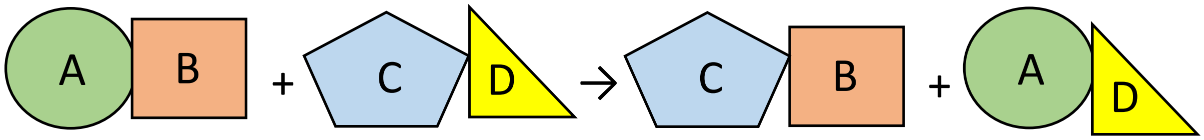

Figure 3. Representation of a Double Replacement Reaction

You will observe three types of double replacement reactions today:

Precipitation reactions: a solid or solids form. Equation 1 is an example of a precipitation reaction. The solubility rules for common ionic compounds in Appendix 3 can help you determine which ions will combine to form solids.

a. Formation of a gas, as seen in the example below.

$$Na_2S\,(aq) + 2HCl(aq) \rightarrow H_2S(g) + 2\,NaCl(aq) \qquad \textit{Equation 10}$$

Another example involves the reaction between sulfuric acid, H_2SO_4 (formed in Equation *6*) and the surface of frescoes. (A fresco is a painting done on the surface of wet, fresh plaster of calcium carbonate. Though origins date back to the 1st and 2nd centuries BCE in Greece, it was developed into a fine art during the Renaissance - 15th and 16th centuries CE. Perhaps the best-known example is the ceiling of the Sistine Chapel, painted by Michelangelo.) This reaction produces calcium sulfate, $CaSO_4$ and carbon dioxide. The calcium sulfate often crystallizes just beneath the surface and throughout the paint layer of frescoes causing severe blistering and loss of the fresco surface. The gas produced also affects the fresco surface. This reaction is shown in Equation 11.

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g) \qquad \textit{Equation 11}$$

Examples of some common gases:

H_2S HCN NH_3 CO_2 SO_2 NO_2 NO

b. Neutralization reactions between an acid and a base produce water and an ionic compound. Acids cause changes in the color of many paints and degrade marble statues.

$$HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) \qquad \textit{Equation 12}$$

Many weak electrolytes are acids and bases.

Examples of some weak electrolytes:

H_2O $HC_2H_3O_2$ H_3PO_4 HCN HF H_3BO_3 H_2SO_3 HNO_2 H_2S NH_4OH

When writing complete ionic and net ionic equations, keep weak electrolytes in a molecular form, as is observed in the exercise below, because weak electrolytes dissociate very little.

Exercise 1. Write the complete ionic and net ionic equations for the reaction between acetic acid and aqueous sodium hydroxide.

Equation 12 is the molecular equation for the reaction between acetic acid and sodium hydroxide. Since acetic acid, is a weak electrolyte, it should not be dissociated in the complete ionic or net ionic equations.

Ionic equation:

$$HC_2H_3O_2(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + C_2H_3O_2^-(aq) + H_2O(l)$$

Net ionic equation:

$$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$$

Techniques

The following techniques are used in the experimental procedure:

- Technique 1: Cleaning glassware
- Technique 8: Decanting
- Technique 9, Video Tech. 9: Using a Bunsen burner
- Technique 10: Using a centrifuge
- Technique 11: Disposing chemical waste
- Technique 12: Using pH paper
- Technique 15: Testing for Odors
- Splint test – explained in experiment

Technique

List of Chemicals

- *0.1 M ammonium chloride
- *0.2 M ammonium oxalate
- *0.1 M barium chloride
- *0.1 M calcium chloride
- solid copper
- *0.1 M copper (II) nitrate
- deionized (DI) water
- *1.0M hydrochloric acid
- **6.0 M hydrochloric acid
- *0.5 M silver nitrate
- *1.0 M sodium hydroxide
- **6 M sodium hydroxide
- *0.1 M sodium sulfate
- solid zinc

*Irritant to the skin, eyes, and mucous membranes. Avoid inhalation and other contact.

Safety

**Severe irritant to the skin, eyes, and mucous membranes. Avoid inhalation and other contact.

List of Equipment and Glassware

- one glass stirring rod
- one large beaker labeled as waste
- one test tube rack (if plastic, do not place a very hot test tube in it)
- twelve 13 mm x 100 mm test tubes (Once you use these, follow the steps listed in the Clean-up/Disposal section, and reuse them. They do not need to be dry inside.)
- one Bunsen burner
- one splint (a thin piece of wood)
- one test tube holder
- one medium size beaker (100-mL or 250-mL)
- centrifuge

Experimental Procedure

Part A General Instructions

1. Be sure to carefully read the label on each chemical bottle. For example, Na_2SO_4 will not work if Na_2SO_3 is required. Check the concentration as well as the name. For example, the properties of concentrated (18 M) sulfuric acid are quite different from 3 M sulfuric acid.
2. Use 13 mm x 100 mm test tubes unless stated otherwise. Wash and rinse each test tube with deionized water. Dry the outside. These can be used wet, just shake to remove excess water.
3. When approximate volumes are needed, use the dropper on the bottles: 20 drops are approximately 1 mL.
4. Pour one reactant into the other - the order does not matter unless stated otherwise.
5. Mix all reactions thoroughly with a glass rod. After using the glass rod, rinse it with deionized water. Collect rinses in your waste beaker. No need to dry it before reusing.
6. Record all observations. Some observations to include (if applicable) are:
   - physical description: color, phase, hardness, crystallinity, shininess
   - temperature change (touch the test tubes before mixing and again after mixing; DO NOT TOUCH A TEST TUBE AFTER HEATING IT IN THE BUNSEN BURNER
   - gas (bubbling or effervescence is evidence of gas formation)
   - odor
   - solid (observed as a precipitate – bottom of test tube, or cloudiness – suspended in the solution)
   - any other observations you notice

Technique 15

7. Did a reaction occur? If so, write yes in the table, and classify the reaction based on the Reaction Types listed earlier in this experiment. If there is no reaction, then simply write no, and write a dash in the reaction type. Then move on.
8. If a chemical reaction occurs, write the molecular, ionic, and net ionic equations below as shown in Figure 4. Write these before moving on to the next reaction, unless stated otherwise. There is no need to try to write an equation if a reaction did not occur.

25) drops 0.1 M silver nitrate + 2 mL 0.1 M sodium chloride

Observations	
Reactants before combining	After combining and mixing
$AgNO_3$ – clear, colorless solution NaCl – clear, colorless solution	White precipitate, clear solution above. Test tube felt cool to the touch.
Reacted? Yes	Reaction type: Double replacement

Molecular Equation:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Complete Ionic Equation:

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + NO_3^-(aq) + NO_3^-(aq)$$

Net Ionic Equation:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Figure 4. Recording Data and Equations – an Example

9. Some reactions ask for your instructor's signature. Have your instructor sign as soon as you complete the table and have written the equations, if applicable.
10. The following materials may be useful when writing the chemical equations:
    - Appendix 3: Solubility Rules
    - Appendix 4: Common Strong Acids and Strong Bases
    - Common gases listed after Equation 11
    - Some weak electrolytes listed after Equation 12

Part B. Reactions and No Reactions

Perform each reaction enumerated below. Record observations for each reactant in the corresponding Experimental Data Table. Then mix them together. Record observations and complete the table.

1. Reaction 1: Add 2 mL of deionized water into a 13 mm x 100 mm test tube. Then add a small amount (about half the size of a small pea) of zinc granules.
2. Reaction 2: Add 2 mL of 6.0 M hydrochloric acid into another 13 mm x 100 mm test tube. Place the test tube in a test tube rack. Since the reaction is fast, good coordination and team work is needed between you and your lab partner in the next steps involving the splint test:

Technique 9
Video Tech. 9

   a. Turn on the Bunsen burner and ignite a splint (a thin piece of wood), blow out most of the flame (leave some embers that glow).
   b. Add a small amount (about half the size of a small pea) of zinc granules into the test tube containing the hydrochloric acid, and immediately place the tip of the splint with glowing embers, just inside the opening of the test tube. (If you do not detect a change, relight the splint, do not blow out the flame, and hold over the test tube. If you still do not detect a change, repeat the test with fresh reactants.)

c. Interpret the results of the splint test:
- glowing splint will burst into flame when oxygen is present;
- glowing splint will cause a small explosion (pop) when hydrogen is present; or
- glowing splint will extinguish when carbon dioxide is present.

d. Dispose of the splint in the trash can (wet it to ensure the flame is extinguished).

e. Mix with glass rod. Observe closely: has the amount of solid zinc changed?
<u>Obtain instructor's signature after you write the reactions, if applicable</u>.

3. Reaction 3. Add a few drops of 0.2 M ammonium oxalate (the oxalate ion is $C_2O_4^{2-}$) to 1 mL of 0.1 M calcium chloride.
<u>Obtain instructor's signature after you write the reactions, if applicable</u>.
4. Reaction 4. Place a small piece of copper metal (turnings or wire) in 1 mL of 6.0 M hydrochloric acid. Observe the reaction when the wire is first placed in the solution, and then observe again several times for the next 10 to 15 minutes. (Continue with the next reactions; remember to come back and observe this reaction every so often).
<u>Obtain instructor's signature after you write the reactions, if applicable</u>.
5. Reaction 5. Place a small piece of copper wire in 2 mL of 0.5 M silver nitrate solution. Observe the reaction when the wire is first placed in the solution, and then observe again several times for the next 10 to 15 minutes. (Continue with the next reactions; do not forget to observe this one every so often). If the solution turns blue, this is evidence of the presence of Cu^{2+}.
<u>Obtain instructor's signature after you write the reactions, if applicable</u>.
6. Reaction 6. Add a few drops of 0.1 M barium chloride to 1 mL of 0.1 M sodium sulfate solution.
7. Reaction 7. Using graduated cylinders, place 2.0 mL of 1.0 M sodium hydroxide in one test tube and 2.0 mL of 1.0 M hydrochloric acid in another. Record how the temperature of the lower part of the test tubes feels to the touch, and measure and record the pH of each solution using pH paper. Pour the hydrochloric acid solution into the sodium hydroxide solution and swirl to mix. Record how the temperature of test tube feels to the touch and measure the pH of the mixture.
A change in temperature or a change in pH of two or more pH units is evidence of a chemical reaction.
(Note: it is sometimes difficult to deliver exactly the exact same amount; an extra drop of acid or base can affect the final pH observed.)
<u>Obtain instructor's signature after you write the reactions, if applicable</u>.

Technique 4
Technique 12

8. Reaction 8. Add 2 mL of 0.1 M ammonium chloride to 2 mL of 0.1 M copper (II) nitrate.

Part C. A Sequence of Reactions

Fill out the data table for each enumerated reaction below. Write equations if a reaction occurs.

9. Reaction 9. Add 6.0 M sodium hydroxide dropwise to 2 mL of 0.1 M copper(II) nitrate until no further reaction is observed. Save the products for the next step.
10. Reaction 10. Centrifuge the product from the step above. Decant the liquid. CAREFULLY heat the solid remaining in the test tube with a Bunsen burner. Keep the test tube moving in the flame and keep the test tube out of the flame most of the time. Be sure that the tube is not pointing at anyone. Continue moving it in and out of the flame until the solid changes color. Save the product. (If the test tube rack is plastic or insulated metal wire, place the hot test tube in

Technique 10
Technique 8

a medium size beaker while it cools.) In this reaction, one reactant makes two products: water in the gas phase and a solid.

11. Reaction 11. Add 6.0 M hydrochloric acid dropwise to the solid produced in the step above. Stir after the addition of each drop. Add only enough hydrochloric acid to have a complete reaction.

Clean-up/Disposal

- Pour all solutions and solids into the inorganic waste container.
- Rinse each piece of glassware with deionized water, and collect the rinses into the inorganic waste container.
- Wash all glassware with soap and water, and then rinse with DI water. Dry the outsides of the glassware. Return all glassware to its place.

Technique 1
Technique 11

Pre-lab

1. The chemistry of fresco painting is the chemistry of limestone and lime plaster. Limestone is calcium carbonate ($CaCO_3$), an abundant, naturally occurring mineral. Two key reactions, shown below, are involved in the process of converting limestone to lime plaster (calcium hydroxide). Classify each of these reactions as one of the four **types of reactions** listed in this experiment (decomposition, synthesis, single replacement, or double replacement).

   a. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

   b. $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

2. The Ferroni-Dini method was used to repair the damaged fresco shown in Figure 1. This method involves production of barium sulfate, a stable white solid, using the two-step process shown below. This solid was then dispersed using nanotechnology[5] to restore the fresco surface. Write **the complete ionic equation** for:

   a. $CaSO_4 \cdot 2H_2O(s) + (NH_4)_2CO_3(aq) \rightarrow (NH_4)_2SO_4(aq) + CaCO_3(s) + 2H_2O(l)$

   b. $Ba(OH)_2(aq) + (NH_4)_2SO_4(aq) \rightarrow BaSO_4(s) + 2NH_3(g) + 2H_2O(l)$

3. Balance the following equation:

   $\_\,NaHCO_3(s) + \_\,H_2SO_4(aq) \rightarrow \_\,Na_2SO_4(aq) + \_\,H_2O(l) + \_\,CO_2(g)$

4. A black precipitate was formed when 1 mL of 0.1 M $Pb(NO_3)_2$ was mixed with 1 mL of 0.1 M Na_2S.

   a. Write the chemical formula of the solid formed.

   b. Write the chemical formula of the spectator ion(s).

5. Two elements in their pure metallic form are used in this experiment. What are they?

6. Some solutions used in today's experiment are listed below. Write their chemical formula:

   a. ammonium oxalate
   b. calcium chloride
   c. copper(II) nitrate
   d. hydrochloric acid
   e. sodium hydroxide

Post-lab

- You do not need to write a conclusion for this experiment.
- The laboratory report should include the following items:
  a. Purpose
  b. Data tables for the 11 experiments.
  c. Molecular, complete ionic, and net ionic equations when applicable.

References

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Experiment 5: Observing and Classifying Reactions

Experimental Data and Equations

Name: ____________________ **Date:** __________

Lab Partner: ____________________ **Section:** ______

Purpose

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____

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Reactions and No Reactions

1. Deionized water + of zinc granules.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

Molecular Equation (ME):

Ionic Equation (IE):

Net Ionic Equation (NIE):

Name: ____

2. 6 M hydrochloric acid + zinc granules. Perform splint test.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

Instructor's Signature ______________________________

3. 0.2 M ammonium oxalate (the oxalate ion is $C_2O_4^{2-}$) + 0.1 M calcium chloride.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

Instructor's Signature ______________________________

Name: ______________________________

4. Copper metal + 6 M hydrochloric acid. Observe reaction several times for 10 - 15 minutes.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

Instructor's Signature ______________________

5. Copper metal + 0.5 M silver nitrate. Observe reaction several times for 10 - 15 minutes.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

Instructor's Signature ______________________

Name: ____

6. 0.1 M barium chloride + 0.1 M sodium sulfate.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

7. 1.0 M sodium hydroxide + hydrochloric acid. Measure temperature and pH before and after.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

Instructor's Signature ____________________________

Name: ______________________________

8. 0.1 M ammonium chloride + 0.1 M copper (II) nitrate.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

A Sequence of Reactions

9. 6.0 M sodium hydroxide dropwise + 0.1 M copper (II) nitrate.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

Name: ______________________________________

10. Centrifuge the product from experiment 9. Decant the liquid. Heat the solid.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:

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11. 6 M hydrochloric acid + solid produced in experiment 10.

Observations	
Reactants before combining	After combining and mixing
Reacted?	Reaction type:

ME:

IE:

NIE:
