# Unit 8

## Chemical Reactions

### Textbook Chapter 8!!!

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<th>Monday</th>
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<th>Friday</th>
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<td>No School</td>
<td>STUDY GUIDE DUE (pg 30-33)</td>
<td></td>
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</tr>
</tbody>
</table>

Name: [Key]  
Period: ___
Day 1: Nomenclature Review!!!!!!!!!

Before naming/writing any formula, identify the type of compound.

- Ionic → Metal and a Nonmetal
  - If Ionic: Does the metal need roman numerals in the name?
  - If Ionic: Does the compound have a Polyatomic Ion?
- Molecular → Only Nonmetals (Use Prefixes in the name)

Ionic Compounds

Ionic bonds are formed between a metal and a nonmetal OR a cation + anion.

Naming ionic compounds

* First name the cation and then the anion.
  - (metal)         (nonmetal)

* Change the ending of the anion to -ide.

Example:

\[
\text{MgCl}_2 \quad \text{magnesium chloride}
\]

Prefixes used for naming

Covalent Compounds

<table>
<thead>
<tr>
<th>Prefix</th>
<th>number indicated</th>
</tr>
</thead>
<tbody>
<tr>
<td>mono-</td>
<td>1</td>
</tr>
<tr>
<td>di-</td>
<td>2</td>
</tr>
<tr>
<td>tri-</td>
<td>3</td>
</tr>
<tr>
<td>tetra-</td>
<td>4</td>
</tr>
<tr>
<td>penta-</td>
<td>5</td>
</tr>
<tr>
<td>hexa-</td>
<td>6</td>
</tr>
<tr>
<td>hepta-</td>
<td>7</td>
</tr>
<tr>
<td>octa-</td>
<td>8</td>
</tr>
<tr>
<td>nona-</td>
<td>9</td>
</tr>
<tr>
<td>deca-</td>
<td>10</td>
</tr>
</tbody>
</table>

Covalent Compounds

To name a covalent compound, name the first element, then name the second one and change its ending to -ide. Use prefixes to show how many atoms of each element you have.

- \( \text{P}_2\text{O}_5 \) diposphorus pentoxide
- \( \text{CCl}_4 \) carbon tetrachloride

To write the formula of a covalent compound, simply translate the prefixes and element names.

- Dinitrogen dioxide \( \text{N}_2\text{O}_2 \)
- Diphosphorus heptoxide \( \text{P}_2\text{O}_7 \)
Let's Practice Nomenclature

Name the following compounds.

1. Fe$^{3+}$ and NO$_3^-$  
   
   Fe(NO$_3$)$_3$  

2. sodium oxide  
   
   Na$_2$O  

3. dinitrogen trioxide  
   
   N$_2$O$_3$  

4. iron(III) chloride  
   
   Fe$_2$Cl$_3$  

5. calcium nitrate  
   
   Ca(NO$_3$)$_2$  

6. HgO  
   
   Mercury II oxide  

7. LiBr  
   
   Lithium Bromide  

8. SO$_2$  
   
   Sulfur Dioxide  

9. Zn$_3$(PO$_4$)$_2$  
   
   Zinc Phosphite  

10. CoCO$_3$  
    
    Cobalt II Carbonate  

Law of Conservation of Mass Experiment

Problems:

1. How can the law of conservation of mass be verified?
2. How can we determine if a chemical change occurred?

Background:

Antoine Lavoisier formulated the Law of Conservation of Mass, which states that matter cannot be created or destroyed. During a chemical reaction, the bonds of the reactants are broken, and the atoms are rearranged to form new substances. Because matter must be conserved, the mass of the products, must be the same as the mass of the reactants. In this lab you will verify the law of conservation of mass using the following reaction.

$$2\text{NaOH} \text{(aq)} + \text{CuSO}_4 \rightarrow \text{Na}_2\text{SO}_4 \text{(aq)} + \text{Cu(OH)}_2 \text{(s)}$$

Pre-Lab Questions:

1. 2S (s) + 3O$_2$ (g) $\rightarrow$ 2SO$_3$
   
   In the above reaction, 2.6g of S reacts with 2.0g O$_2$. How many grams of SO$_3$ are recovered.

   $$2.6 + 2.0 = 4.6\text{ g}$$

2. 2HCl + Mg $\rightarrow$ MgCl$_2$ + H$_2$
   
   In the above reaction, 5.0g of HCl react with 2.0g Mg. 3.0g of MgCl$_2$ are recovered, how many grams of H$_2$ were lost.

   $$7 = 3 + x$$
   $$x = 4\text{ g of H}_2$$

3. What are the four indications of a chemical change?
   
   a. Production of gas  
   b. Formation of Precipitate  
   c. Color change  
   d. Formation of heat or light
Procedure:

1. Pour 10.0 ml of CuSO₄ into a plastic measuring cup and label it.
2. Pour 10.0 ml of NaOH into a second plastic cup and label it. Carefully put each solution on the balance and record the mass in the data table.
3. Pour the NaOH into the CuSO₄. Use the stirring rod to mix the solutions. Describe what happens in your qualitative observation section.
4. Put both containers (including the empty NaOH cup!) back on the balance and record the new mass. Did the mass change? Record your answer in the observation section. Dispose of chemicals as teacher directs.

Data Table 1:

<table>
<thead>
<tr>
<th>Procedure</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of cup 1</td>
<td></td>
</tr>
<tr>
<td>Mass of cup 1 and NaOH</td>
<td></td>
</tr>
<tr>
<td>Mass of NaOH</td>
<td></td>
</tr>
<tr>
<td>Mass of cup 2</td>
<td></td>
</tr>
<tr>
<td>Mass of cup 2 and CuSO₄</td>
<td></td>
</tr>
<tr>
<td>Mass of CuSO₄</td>
<td></td>
</tr>
<tr>
<td>Mass of cup 1 and 2 after mixing</td>
<td></td>
</tr>
</tbody>
</table>

Calculations:

Calculate the total mass of the reactants. (Add 3. and 6. from data table) __________ 9

Post-Lab Questions:

1. The mass before the reaction is calculated above. The mass after the reaction is in your data table. Do your values for the total mass before and after each reaction verify the Law of Conservation of Mass? Explain in complete sentences.

2. What indicators of a chemical change were observed? Which indicators were not observed?
Day 1: Homework

\[ \text{Na} + \text{CuS} \rightarrow \text{Na}_2\text{S} + 2 \text{Cu} \]

1. In the above chemical reaction, you reacted 25.0g Na with 15.5g CuS. What is the total mass of your products?

\[ 25 + 15.5 = 40.5 \text{g} \]

2. The reaction for the burning of 15.0 g of ethanol \((\text{CH}_3\text{CH}_2\text{OH})\) looks like this:

\[ 2 \text{CH}_3\text{CH}_2\text{OH} (l) + 7 \text{O}_2 (g) \rightarrow 4 \text{CO}_2 (g) + 6 \text{H}_2\text{O} (l) \]

The mass of the products is 30.0g. How much oxygen was consumed in the reaction?

\[ 15 \text{g} \]

3. A solid has a mass of 35 g. When it is mixed with a solution, a chemical reaction occurs. If the final total mass of products is 85 g, what was the mass of the solution that the solid was mixed with?

\[ 35y + x = 85y \quad x = 50 \]

4. The word equation for the following reaction is as follows:

v vinegar + baking soda \rightarrow \text{sodium acetate} + \text{water} + \text{carbon dioxide}

The chemical equation for the reaction is:

\[ \text{C}_2\text{H}_3\text{O}_2\text{H} + \text{NaHCO}_3 \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} + \text{CO}_2 \]

Use the data table below to find the answers to the three ?.

<table>
<thead>
<tr>
<th>Mass of Weigh Boat</th>
<th>2.0 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of Weigh Boat and baking soda ((\text{NaHCO}_3))</td>
<td>17.0g</td>
</tr>
<tr>
<td>Mass of Baking soda</td>
<td>(10 g)</td>
</tr>
<tr>
<td>Mass of graduated cycliner</td>
<td>50.0g</td>
</tr>
<tr>
<td>Mass of Weigh Boat and vinegar ((\text{C}_2\text{H}_3\text{O}_2\text{H}))</td>
<td>100.0g</td>
</tr>
<tr>
<td>Mass of vinegar ((\text{C}_2\text{H}_3\text{O}_2\text{H}))</td>
<td>(10 g)</td>
</tr>
<tr>
<td>Mass of (\text{NaC}_2\text{H}_3\text{O}_2)</td>
<td>40.0g</td>
</tr>
<tr>
<td>Mass of (\text{H}_2\text{O})</td>
<td>10.0g</td>
</tr>
<tr>
<td>Mass of (\text{CO}_2)</td>
<td>? (10 g)</td>
</tr>
</tbody>
</table>

5. A full, unopened bottle of Dr. Pepper® has a mass of 1,250.0 grams. After opening the bottle (but before taking a sip!)- the beverage now weighs 1050.0 grams. In your own words, explain why the soda seemed to have lost mass.

- Gas bubbles have mass and they were lost to the surroundings.
Day 2:
Nomenclature Review: Write the formulas for the compounds below.

- aluminum oxide: $\text{Al}_2\text{O}_3$
- sodium carbonate: $\text{Na}_2\text{CO}_3$
- nitrogen trioxide: $\text{N}_2\text{O}_3$
- lead(IV) oxide: $\text{PbO}_2$
- copper(II) nitrate: $\text{Cu(NO}_3\text{)}_2$

So now let’s write some chemical equations!

**Summary:**
- When 1 or more **compounds** are reacted to produce totally new compounds that have different **chemical** and physical **properties** than they did before.
- A chemical reaction is represented by writing a **chemical equation**.
  - Using chemical formulas, symbols, and **coefficients**.
- An equation represents the identities and relative amounts of what are called **reactants** and **products** in a chemical reaction.
  - Reactants are the substances you start with in the reaction.
  - Products are the results of the reaction.

**EVIDENCE OF CHEMICAL CHANGE**

1. **Production of gas** This is usually indicated by bubbles. Bubbles don’t always mean a gas is being produced (Air bubbles, bubbles when something boils).

2. **Formation of precipitate** This is called a **precipitate** (ppt). If 2 liquids are combined and a solid substance suddenly appears, this means that a new substance has been formed.

3. **Color change** When a different color or odor appears suddenly, this means a new substance with new properties has formed.

4. **Formation of heat** If a substance becomes warmer in temperature even though no heat is being added, that is a clue that a chemical change is occurring. Substances can also get lower in temperature due to a chemical change.
Writing and Balancing Chemical Reactions: Common Symbols in Reactions

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>+</td>
<td>Separates Reactant from Reactant or Separates Product from Product</td>
</tr>
<tr>
<td>→</td>
<td>“Yields” (indicates result of reaction) Separates Reactants from Products</td>
</tr>
<tr>
<td>(s)</td>
<td>solid</td>
</tr>
<tr>
<td>(l)</td>
<td>pure liquid</td>
</tr>
<tr>
<td>(g)</td>
<td>gas</td>
</tr>
<tr>
<td>(aq)</td>
<td>dissolved in water or in solution</td>
</tr>
<tr>
<td>Δ</td>
<td>heat is added to the reaction</td>
</tr>
<tr>
<td>cat.</td>
<td>A catalyst is added to the reaction.</td>
</tr>
<tr>
<td>↑</td>
<td>A gas is produced.</td>
</tr>
<tr>
<td>↓</td>
<td>A solid is produced.</td>
</tr>
</tbody>
</table>

Common Substances in Chemical Reactions

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula(s) to memorize</th>
</tr>
</thead>
<tbody>
<tr>
<td>Diatomic Elements</td>
<td>H₂, N₂, O₂, F₂, Cl₂, Br₂, I₂</td>
</tr>
<tr>
<td>Water</td>
<td>H₂O</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH₃</td>
</tr>
</tbody>
</table>

Key Words Used in Describing Chemical Reactions

<table>
<thead>
<tr>
<th>Words used to separate reactants from other reactants.</th>
<th>Reacts with</th>
<th>Bubbled through</th>
<th>Mixed together</th>
</tr>
</thead>
<tbody>
<tr>
<td>Words used to separate reactants from products.</td>
<td>Yield</td>
<td>Composted</td>
<td>Burned</td>
</tr>
</tbody>
</table>

Class Practice

For each of the following examples write the correct chemical formulas and symbols.

1. When heated solid tungsten metal reacts with oxygen gas to produce solid tungsten (VI) oxide.
   \[ W(s) + O₂(g) \xrightarrow{\Delta} WO₂(s) \]

2. Solutions of sodium iodide and lead (II) nitrate are mixed and form the precipitate, lead (II) iodide and aqueous sodium nitrate.
   \[ NaI(aq) + Pb(NO₃)₂(aq) \rightarrow PbI₂(s) + NaNO₃(aq) \]

3. When heated, solid aluminum oxide decomposes to form aluminum metal and oxygen gas.
   \[ Al₂O₃(s) \rightarrow Al(s) + O₂(g) \]
**Day 2: Homework**

**Directions:** For each item in Column A, write the letter of the matching item in the word bank. Each choice in Column B may be used once, more than once, not at all.

**Column A**

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>A</strong></td>
<td>The new substances that are formed in a chemical reaction</td>
</tr>
<tr>
<td><strong>B</strong></td>
<td>A chemical reaction that involves one or more substances change into new substances</td>
</tr>
<tr>
<td><strong>C</strong></td>
<td>Shows the relationship between the reactants and products in a chemical reaction</td>
</tr>
<tr>
<td><strong>D</strong></td>
<td>States the mass is neither created nor destroyed during a chemical reaction</td>
</tr>
<tr>
<td><strong>E</strong></td>
<td>The starting substances in a chemical reaction</td>
</tr>
<tr>
<td><strong>F</strong></td>
<td>The substances appearing on the left side of the arrow</td>
</tr>
<tr>
<td><strong>G</strong></td>
<td>The substances appearing on the right side of the arrow</td>
</tr>
</tbody>
</table>

**Word Bank**

- a. Chemical Change
- b. Reactants
- c. Products
- d. Chemical Equation
- e. Law of Conservation of Mass

**Directions:** For the following problems, write the chemical equation for each reaction. Answer the question using the equation and the Law of Conservation of Mass. Show all work on your paper.

8. In a laboratory, 178.8g of water is separated in hydrogen gas and oxygen gas. The hydrogen gas has a mass of 20.0g. What is the mass of oxygen gas produced?

\[
\text{H}_2\text{O} \rightarrow \text{H}_2(\text{g}) + \text{O}_2(\text{g})
\]

178.8 g = 20.0 g + ___

9. When calcium is placed in water, it reacts to form calcium hydroxide and hydrogen gas. A student reacts 2.3g of calcium reacts with 13.8g of water. After the reaction is completed, the student determines that a total of 1.2g of hydrogen gas was produced. How much calcium hydroxide was produced in this reaction?

\[
\text{Ca} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2(\text{g})
\]

14.9 g Ca(OH)_2 2.3 g + 13.8 g = ___ + 1.2 g

**Directions:** Answer questions #10-14 based on the reaction below.

\[
\text{H}_2\text{SO}_4(\text{aq}) + \text{Na}_2\text{CO}_3(\text{s}) \rightarrow \text{CO}_2(\text{g}) + \text{Na}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})
\]

10. What are the reactants in this chemical reaction? \(\text{H}_2\text{SO}_4(\text{aq}) + \text{Na}_2\text{CO}_3(\text{s})\)

11. What is the physical state of water in this chemical reaction? **Liquid**

12. What is the physical state of sodium carbonate in this chemical reaction? **Solid**

13. What sign of a chemical change is demonstrated by this reaction? **Produced new substances**

14. What does (aq) mean? **Aqueous** dissolved in water
Day 3: Review / BALANCING CHEMICAL EQUATIONS

Antoine Lavoisier: He is the father of modern chemistry, and first described the "Law of Conservation of Mass."

Antoine Lavoisier found that the mass of the reactants and the products are equal, even when the states of matter change. Matter is neither created nor destroyed.

**Conservation of atoms**: The number of each type of atom on the reactants side of the chemical equation MUST be equal to the number of each type of atom on the products side of the equation.

**Coefficient**: Represents the number of units of each substance taking part in the reaction.

**Balanced chemical equation**: The same number of atoms of each element on both sides of the equation.

Four Steps to Balance Equations:

1. Set up your equation
2. Count the number of atoms you have on both sides
3. Balance by changing the coefficients and recounting
4. Start the process again if it still doesn't balance

Is this chemical equation balanced?

\[
\begin{array}{ccc}
\text{N}_2 & \text{+} & 3\text{H}_2 \rightarrow \text{2NH}_3 \\
\underline{2} & \underline{6} & \underline{2} \\
\end{array}
\]

Guided Practice: Balance these equations: **How many of each element do we have? Are both sides equal?**

1. \(2\text{H}_2 + 1\text{O}_2 \rightarrow \text{2H}_2\text{O}\)

\[
\begin{array}{ccc}
\text{2H}_2 & \text{+} & 1\text{O}_2 \rightarrow \text{2H}_2\text{O} \\
4 & 2 & 2 \\
\end{array}
\]

2. \(1\text{Mg} + 2\text{HCl} \rightarrow 1\text{H}_2 + \text{1MgCl}_2\)

\[
\begin{array}{ccc}
1\text{Mg} & \text{+} & 2\text{HCl} \rightarrow 1\text{H}_2 + 1\text{MgCl}_2 \\
2+1 & 2 \\
\end{array}
\]

3. \(2\text{Na} + 2\text{HCl} \rightarrow 2\text{NaCl} + 1\text{H}_2\)

\[
\begin{array}{ccc}
2\text{Na} & \text{+} & 2\text{HCl} \rightarrow 2\text{NaCl} + 1\text{H}_2 \\
2+1 & 2 \\
\end{array}
\]
Day 3 Homework

Writing and Balancing Chemical Equations

1. \( \text{N}_2 + \text{O}_2 \rightarrow \text{NO}_2 \)

2. \( \text{Mg}^{2+} \text{NO}_3^- \)

2. Magnesium Nitrate + Calcium Hydroxide → Calcium Nitrate + Magnesium Hydroxide **use polyatomic Ion list.

\[ \text{Mg(NO}_3\text{)}_2 + \text{Ca(OH)}_2 \rightarrow \text{Ca(NO}_3\text{)}_2 + \text{Mg(OH)}_2 \]

3. \( \text{H}_2 + \text{Cl}_2 \rightarrow \text{HCl} \)

4. Magnesium Chloride + Sodium Iodide → Magnesium Iodide + Sodium Chloride

\[ \text{MgCl}_2 + 2 \text{NaI} \rightarrow \text{MgI}_2 + 2 \text{NaCl} \]

5. \( \text{HgO} \rightarrow \text{Hg} + \text{O}_2 \)

6. \( \text{Ca} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2 \)

7. \( \text{C} + \text{O}_2 \rightarrow \text{CO} \)

8. \( \text{Na}_2\text{O}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}_2 \)

9. Sulfuric Acid(\(\text{H}_2\text{SO}_4\)) → Sulfur Trioxide + Water

\[ \text{H}_2\text{SO}_4 \rightarrow \text{SO}_3 + \text{H}_2\text{O} \]
10. \[ \text{2} \text{HCl} + \text{1} \text{Mg} \to \text{1} \text{MgCl}_2 + \text{1} \text{H}_2 \]

11. \[ \text{2} \text{Al(s)} + \text{3} \text{O}_2(g) \to \text{2} \text{Al}_2\text{O}_3 \]

12. \[ \text{2} \text{Ag}_2\text{O(s)} \xrightarrow{\Delta} \text{4} \text{Ag} + \text{1} \text{O}_2 \]

13. \[ \text{1} \text{S(s)} + \text{3} \text{O}_2(g) \to \text{2} \text{SO}_3 \]

14. \[ \text{1} \text{N}_2\text{O}_5(g) + \text{1} \text{H}_2\text{O(l)} \to \text{2} \text{HNO}_3 \]

15. \[ \text{1} \text{CaO(s)} + \text{1} \text{H}_2\text{O(l)} \to \text{1} \text{Ca(OH)}_2 \]

16. \[ \text{1} \text{H}_2\text{O}_2(l) \xrightarrow{\Delta} \text{1} \text{H}_2 + \text{1} \text{O}_2 \]

17. \[ \text{1} \text{CH}_4 + \text{2} \text{O}_2 \to \text{2} \text{H}_2\text{O} + \text{1} \text{CO}_2 \]

18. \[ \text{1} \text{SrF}_2 + \text{1} \text{Na}_2\text{SO}_4 \to \text{1} \text{SrSO}_4 + \text{2} \text{NaF} \]

19. \[ \text{1} \text{Pb(NO}_3)_2 + \text{1} \text{K}_2\text{CrO}_4 \to \text{1} \text{PbCrO}_4 + \text{2} \text{KNO}_3 \]

20. \[ \text{1} \text{K}_2\text{SO}_4 + \text{1} \text{BaCl}_2 \to \text{2} \text{KCl} + \text{1} \text{BaSO}_4 \]

21. \[ \text{2} \text{KOH} + \text{1} \text{H}_2\text{SO}_4 \to \text{1} \text{K}_2\text{SO}_4 + \text{2} \text{H}_2\text{O} \]

22. \[ \text{2} \text{Fe} + \text{3} \text{H}_2\text{SO}_4 \to \text{1} \text{Fe}_2\text{SO}_4\text{O}_3 + \text{3} \text{H}_2 \]
Day 4: Oxidation-Reduction Reactions (Redox Reactions)

Reaction Types are put into three categories:

All of the reaction types we’ve discussed can be placed into one of 3 categories:

* Acid-Base (Neutralization)
* Precipitate
* Oxidation-reduction (Redox)

In this unit we will be looking at Precipitate and Redox.

A. Oxidation-Reduction (Synthesis, Decomposition and Single Replacement reactions)

In redox reactions the oxidation number for an element changes in a chemical reaction.

Oxidation Number: A number assigned to an element, based on the distribution of electrons. The same element can have very different properties in different oxidation states.

<table>
<thead>
<tr>
<th>Rules for Assigning Oxidation #s</th>
<th>Examples</th>
<th>Oxidation #</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 The oxidation number of any uncombined ELEMENT is 0</td>
<td>Na, O₂</td>
<td>Na = 0, O₂ = 0</td>
</tr>
<tr>
<td>2 The ox. # of an ION equals the charge of the ion</td>
<td>Cl⁻</td>
<td>Cl⁻ = -1</td>
</tr>
<tr>
<td>3 The ox. # of elements in COMPOUNDS typically, but not always (unless noted otherwise), follow a trend on the periodic table: Group 1 = +1 ALWAYs</td>
<td>LiF</td>
<td>Li = +1</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Ca = +2</td>
</tr>
<tr>
<td></td>
<td></td>
<td>HF</td>
</tr>
<tr>
<td></td>
<td></td>
<td>H₂O</td>
</tr>
<tr>
<td>Transition metals AND Group 14 = variable charges Note: H has exceptions but we will use +1 in this course</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4 The sum of ox. #s of all atoms in a NEUTRAL COMPOUND is 0</td>
<td>LiF</td>
<td>Li = +1+</td>
</tr>
<tr>
<td></td>
<td></td>
<td>+F = 1-</td>
</tr>
<tr>
<td></td>
<td></td>
<td>0 = 0</td>
</tr>
<tr>
<td></td>
<td></td>
<td>CaCO₃</td>
</tr>
<tr>
<td></td>
<td></td>
<td>+C⁺⁺(4) = +4</td>
</tr>
<tr>
<td></td>
<td></td>
<td>0 - 2 (x3) = -8</td>
</tr>
<tr>
<td></td>
<td></td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td></td>
<td></td>
<td>+O⁻⁻⁻⁻ (x4) = -8</td>
</tr>
<tr>
<td></td>
<td></td>
<td>0 = -2</td>
</tr>
</tbody>
</table>

Note: Additional rules/exceptions to these rules do exist, but are beyond the scope of this course.
**Oxidation** is a reaction in which there is the loss of electrons.

\[ \text{Ex: } \text{Na} \rightarrow \text{Na}^{+1} + e^- \]

**Reduction** is a reaction in which there is the gain of electrons.

\[ \text{Ex: } \text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^- \]

"LEO the lion says GER"

Losing Electrons is Oxidation

Gaining Electrons is Reduction

"OIL RIG"

Oxidation Is Loss (of electrons)

Reduction Is Gain (of electrons)

Since oxidation is the **loss** of electrons and reduction is the **gain** of electrons, they must occur simultaneously.

**Any chemical process in which elements undergo changes in oxidation number is an oxidation - reduction reaction, or redox reaction for short.**

Practice determining whether the following elements have been oxidized or reduced and label the reaction.

**Example 1:**

\[ 4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 \]

**TYPE of Reaction: S/D/SR/DR**

<table>
<thead>
<tr>
<th>Element</th>
<th>Ox.# Reactants side</th>
<th>Ox.# Products side</th>
<th>Lose/Gain e⁻</th>
<th>Oxidized/Reduced</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe</td>
<td>0</td>
<td>+3</td>
<td>Lose 3</td>
<td>Oxidized</td>
</tr>
<tr>
<td>O</td>
<td>0</td>
<td>-2</td>
<td>Gain 2</td>
<td>Reduced</td>
</tr>
</tbody>
</table>

**Example 2:**

\[ \text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \]

**TYPE of Reaction: S/D/SR/DR**

<table>
<thead>
<tr>
<th>Element</th>
<th>Ox.# Reactants side</th>
<th>Ox.# Products side</th>
<th>Lose/Gain e⁻</th>
<th>Oxidized/Reduced</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>+1</td>
<td>0</td>
<td>Gain 1</td>
<td>Reduced</td>
</tr>
<tr>
<td>Mg</td>
<td>0</td>
<td>+2</td>
<td>Lose 2</td>
<td>Oxidized</td>
</tr>
<tr>
<td>Cl</td>
<td>-1</td>
<td>-1</td>
<td>neither</td>
<td>neither</td>
</tr>
</tbody>
</table>

**Example 3:**

\[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]

**TYPE of Reaction: S/D/SR/DR**

Note: **ALL** of the reaction types we have learned are redox reactions **EXCEPT** for **double replacement** reactions.
Day 4 Homework

1. Label the following as an oxidation or reduction reaction.
   
   \[
   \text{Ag} \rightarrow \text{Ag}^+ + e^- \quad \text{oxidized}
   \]
   
   \[
   \text{F}_2 + 2e^- \rightarrow 2\text{F}^- \quad \text{reduction}
   \]
   
   \[
   \text{Cu} \rightarrow \text{Cu}^{2+} + 2e^- \quad \text{oxidation}
   \]

2. Complete the tables below.

   1. \(2\text{Ag} + \text{S} \rightarrow \text{Ag}_2\text{S}\)
   
   \[
   \begin{array}{|c|c|c|c|c|}
   \hline
   \text{Element} & \text{Ox.# Reactants side} & \text{Ox.# Products side} & \text{Lose/Gain } e^- & \text{Oxidized/Reduced} \\
   \hline
   \text{Ag} & 0 & 1+ & \text{Lose 1} & \text{oxidized} \\
   \text{S} & 0 & 2- & \text{Gain 2} & \text{reduced} \\
   \hline
   \end{array}
   \]

   2. \(2\text{Na} + \text{FeCl}_2 \rightarrow 2\text{NaCl} + \text{Fe}\)
   
   \[
   \begin{array}{|c|c|c|c|c|}
   \hline
   \text{Element} & \text{Ox.# Reactants side} & \text{Ox.# Products side} & \text{Lose/Gain } e^- & \text{Oxidized/Reduced} \\
   \hline
   \text{Na} & 0 & 1+ & \text{Lose 1} & \text{oxidized} \\
   \text{Fe} & 2+ & 0 & \text{Gain 2} & \text{reduced} \\
   \text{Cl} & 1- & 1- & \text{Neither} & \text{Neither} \\
   \hline
   \end{array}
   \]

   3. \(2\text{AlCl}_3 \rightarrow 2\text{Al} + 3\text{Cl}_2\)
   
   \[
   \begin{array}{|c|c|c|c|c|}
   \hline
   \text{Element} & \text{Ox.# Reactants side} & \text{Ox.# Products side} & \text{Lose/Gain } e^- & \text{Oxidized/Reduced} \\
   \hline
   \text{Al} & 3+ & 0 & \text{Gain 3} & \text{reduced} \\
   \text{Cl} & 1- & 0 & \text{Lose 1} & \text{oxidized} \\
   \hline
   \end{array}
   \]
Day 5: Types of Redox Reactions

Synthesis (also called Combination or Composition)

Synthesis means “put together”

2 or more elements/simple compounds combine to form 1 compound.

General form: \( A + B \rightarrow AB \)

Identifying feature: only one \( \text{product} \)

Analogy: A boy and a girl come to the dance separately, but end up dancing together.

\[ \text{EX: } 2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO} \]

EX: The combination of iron and sulfur to form iron (II) sulfide: \( \text{Fe} + \text{S} \rightarrow \text{FeS} \)

EX: Burning Charcoal: \( \text{C} (s) + \text{O}_2 (g) \rightarrow \text{CO}_2 (g) \)

2. Decomposition

Decomposition means “break apart”

1 compound is broken down into 2 or more simpler elements/compounds.

General form: \( AB \rightarrow A + B \)

Identifying feature: only one \( \text{reactant} \) (opposite of Synthesis!)

Analogy: A boy and a girl are a couple, but they had an argument. This resulted in the boy and the girl staying apart for the rest of the night.

\[ \text{Examples: } 2 \text{NaCl} \rightarrow 2 \text{Na} + \text{Cl}_2 \]
\[ 2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2 \]

EX: The electrolysis of water to make oxygen and hydrogen gas: \( 2\text{H}_2\text{O} (l) \rightarrow 2\text{H}_2 (g) + \text{O}_2 (g) \)

3. Combustion

Also known as \( \text{burning} \)

Always follows the same form: \( \text{compound containing } C \ + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)

Compound containing \( C \) and \( H \) (and sometimes \( O \)) \( + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)
Note: In a combustion reaction, the compound always burns in oxygen gas and always releases carbon dioxide and water.

During incomplete combustion (a limited amount of \( O_2 \)), carbon monoxide (CO) is also produced.

**Examples:**

\[
\text{CH}_4 \ + \ 2\ O_2 \rightarrow \ CO_2 \ + \ 2\ H_2O
\]

\[
2\ C_3H_6 \ + \ 9\ O_2 \rightarrow \ 6\ CO_2 \ + \ 6\ H_2O
\]

**Trick for balancing tough combustion reactions:**

**Balance from right to left.**

1. Balance the H first by placing a coefficient in front of the water.
2. If that coefficient is **ODD**, **double it** and proceed with balancing C, then O.
3. If even, leave it alone and proceed with C, O.
4. **Always balance oxygen last!!!**

**Example:**

\[
\begin{array}{c}
2 \text{C}_7\text{H}_{14} + 10.5 \text{O}_2 \rightarrow 14 \text{CO}_2 + 7 \text{H}_2\text{O} \\
\end{array}
\]

In class practice for balancing combustion.

1. \[
\begin{array}{c}
\underline{1} \text{C}_4\text{H}_8 + \underline{6} \text{O}_2 \rightarrow \underline{4} \text{CO}_2 + \underline{4} \text{H}_2\text{O} \\
\end{array}
\]

2. \[
\begin{array}{c}
\underline{1} \text{C}_5\text{H}_{12} + \underline{8} \text{O}_2 \rightarrow \underline{5} \text{CO}_2 + \underline{6} \text{H}_2\text{O} \\
\end{array}
\]

3. \[
\begin{array}{c}
\underline{2} \text{C}_9\text{H}_{18} + \underline{13.5} \text{O}_2 \rightarrow \underline{9} \text{CO}_2 + \underline{9} \text{H}_2\text{O} \\
\end{array}
\]

4. \[
\begin{array}{c}
\underline{2} \text{C}_{11}\text{H}_{22} + \underline{33} \text{O}_2 \rightarrow \underline{22} \text{CO}_2 + \underline{22} \text{H}_2\text{O} \\
\end{array}
\]
Day 5: Types of Reactions Homework
Balance and Identify the following: synthesis = S
decomposition = D
combustion = C

Find Oxidized and Reduced Elements in 1., 3, and 5.

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>4Al(s) + 3O₂(g) → 2Al₂O₃</th>
<th>S</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>Al</td>
<td>0</td>
<td>3+</td>
</tr>
<tr>
<td>O₂</td>
<td>0</td>
<td>2-</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>2S(s) + 3O₂(g) → 2SO₃</th>
<th>S</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>S</td>
<td>0</td>
<td>2-</td>
</tr>
</tbody>
</table>

Use Cu¹⁺

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>2Cu(s) + S(s) → Cu₂S⁺</th>
<th>S</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>Cu</td>
<td>0</td>
<td>1+</td>
</tr>
<tr>
<td>S</td>
<td>0</td>
<td>2-</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>1SO₃(g) + 1H₂O(l) → 1H₂SO₄</th>
<th>S</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>Ag</td>
<td>1+</td>
<td>0</td>
</tr>
<tr>
<td>O₂</td>
<td>2-</td>
<td>0</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>2Ag₂O(s) → 4Ag + 1O₂</th>
<th>D</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>Ag</td>
<td>1+</td>
<td>0</td>
</tr>
<tr>
<td>O</td>
<td>2-</td>
<td>0</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>2C₆H₆ + 15O₂ → 12CO₂ + 6H₂O</th>
<th>C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>Ag</td>
<td>1+</td>
<td>0</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>2C₆H₁₈ + 25O₂ → 8CO₂ + 18H₂O</th>
<th>C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Ox# Reactants</td>
<td>Ox# Products</td>
</tr>
<tr>
<td>Ag</td>
<td>1+</td>
<td>0</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>C₆H₁₂O₆ + 6O₂ → 6CO₂ + 6H₂O</th>
<th>C</th>
</tr>
</thead>
</table>

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>C₆H₁₂O₆ + 6O₂ → 6CO₂ + 6H₂O</th>
<th>C</th>
</tr>
</thead>
</table>
Day 6: Single Replacement

Single Replacement

1 element takes the place of another in a compound ("like replaces like")

General forms: $A + BC \rightarrow B + AC$ (metal replacement)
$D + BC \rightarrow C + BD$ (halogen replacement)

Identifying feature: 1 element + 1 compound on each side of the arrow

Analogy: A boy and a girl are dancing, but then another boy "cuts in" and dances with the girl, leaving the first boy alone. Or a boy and a girl are dancing, but then another girl "cuts in" and dances with the boy, leaving the first girl alone. "Like" must replace "like"

Examples:

Metal replacement: $2 \text{Na} + \text{CuCl}_2 \rightarrow 2 \text{NaCl} + \text{Cu}$

Halogen replacement: $\text{F}_2 + 2 \text{KCl} \rightarrow 2 \text{KF} + \text{Cl}_2$

BUT the boy/girl will not always be able to "cut in." Sometimes the other boy/girl will not let them!

**Rules for Reactions Including: Metals with Metals / Metals with Acids**

We must use the **Activity Series** to predict whether or not the replacement will occur.

$A = $ Metal
$B = $ Cation
$C = $ Anion
This is what it looks like! You can find this on the BACK of the Periodic Table of the Elements.

Metals found higher up in the activity series means the metal is MORE REACTIVE.

You can use the activity series to predict whether or not certain reactions will occur. A specific metal can replace any metal listed below it that is in a compound. It cannot replace any metal listed above it. For example, copper atoms replace silver atoms in a solution of silver nitrate. However, if you place a silver wire in aqueous copper (II) nitrate, the silver atoms will not replace the copper. Silver is listed below copper in the activity series, so no reaction occurs. The letters NR (no reaction) are commonly used to indicate that a reaction will not occur.

**This applies to the nonmetal halogens as well. They increase as you move up the family.

Predicting Single Replacement Reactions

**Important Reminder: A metal can only replace another metal or hydrogen.**

<table>
<thead>
<tr>
<th>Rule #1: Metals will replace metals that are lower than themselves on the Activity Series.</th>
<th>(The substances will switch partners)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A and B are metals, and X represents a negative ion.</td>
<td></td>
</tr>
<tr>
<td>-Will the reaction proceed? If yes, what are the products?</td>
<td></td>
</tr>
<tr>
<td>$A + BX \rightarrow B + AX$</td>
<td></td>
</tr>
<tr>
<td>(Higher) (Lower)</td>
<td></td>
</tr>
<tr>
<td>$A + BX \rightarrow$</td>
<td></td>
</tr>
<tr>
<td>(Lower) (Higher)</td>
<td></td>
</tr>
<tr>
<td><strong>no reaction</strong></td>
<td></td>
</tr>
</tbody>
</table>
Using your Activity Series, predict which reactions will occur:

1. $3\text{Mg(s)} + 2\text{AlCl}_3(\text{aq}) \rightarrow 3\text{MgCl}_2 + 2\text{Al}$

2. $\text{FeCl}_2(\text{aq}) + \text{Cu(s)} \rightarrow \text{No reaction}$

3. $\text{Sodium + Lithium Chloride} \rightarrow \text{No reaction}$

4. $\text{Silver Nitrate + Zinc} \rightarrow \text{Zinc Nitrate + Silver}$

Predicting the Products (metals and metals):

1. $3\text{Mg(s)} + 2\text{AlCl}_3(\text{aq}) \rightarrow 2\text{Al} + 3\text{MgCl}_2$

2. $\text{Silver Nitrate + Zinc} \rightarrow 2\text{AgNO}_3 + \text{Zn} \rightarrow 2\text{Ag} + \text{Zn(NO}_3)_2$

3. $\text{Fe(s)} + \text{Pb(NO}_3)_2(\text{aq}) \rightarrow \text{Pb} + \text{Fe(NO}_3)_2$

Rule #2: Any metal above $\text{H}_2$ will react with an acid to produce a compound + $\text{H}_2$. Put $\text{H}_2$ at the end of your arrow.

Example:

$\text{M} =$ metal

$\text{X} =$ negative ion

Using your Activity Series, predict which reactions will occur:

1. $\text{Zn} + \text{HCl} \rightarrow \text{YES!}$

2. $\text{Cu} + \text{H}_2\text{SO}_4 \rightarrow \text{NO!}$

3. $\text{Calcium + Phosphoric Acid} \rightarrow \text{YES!}$

Predicting the Products (metals and acids):

1. $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$

2. $\text{Calcium + Phosphoric Acid} \rightarrow 3\text{Ca} + 2\text{H}_3\text{PO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + 3\text{H}_2$
Day 6: Homework

1. Circle any reactions that will NOT occur.

   \[
   \begin{align*}
   &\text{Ba + NaNO}_3 & \text{Fe + Pb(C}_2\text{H}_3\text{O}_2)_2 & \text{KCl + Ni} \\
   &\text{Mg + HCl} & \text{HBr + Au} & \text{Cu + AgNO}_3
   \end{align*}
   \]

2. Using your activity series, determine if the reactions below will occur. If the reaction will not occur, write no reaction. If the reaction will occur, predict the products of the reaction and write a balanced chemical reaction.

   a. \[\text{Ag + Cu(NO}_3)_2 \rightarrow \text{NO Reaction}\]

   b. \[\text{MgCl}_2 + \text{Ba} \rightarrow \text{BaCl}_2 + \text{Mg}\]

   c. \[\text{Zn + 2HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2\]

   d. \[\text{2Al + 3Cu(NO}_3)_2 \rightarrow 2\text{Al(NO}_3)_3 + 3\text{Cu}\]

   e. \[\text{Au + HBF} \rightarrow \text{NO Reaction}\]
Day 7: Redox Reactions continued...

Rule #3: Any metal that is sodium or above will react with water.

Helpful Hint: Think of water as $\text{H} - \text{O} - \text{H}$.

Using your Activity Series, predict which reactions will occur:
1. $\text{Na} + \text{H}_2\text{O} \rightarrow \text{yes!}$

2. Calcium metal is placed in water. $\text{yes!}$

3. A sample of cadmium is placed in a test tube of water. $\text{No!}$

Predicting the Products (metals and water):

1. $\frac{2}{\text{Na}} + \frac{2}{\text{H}_2\text{O}} \rightarrow 2\text{NaOH} + \text{H}_2$

2. Calcium metal is placed in water.
   \[
   \text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2
   \]

Rule #4: Halogens will only replace other halogens that are higher in the Activity Series. Halogens will NOT replace metals! Use the family on the periodic table.

Using your Halogen Activity Series, predict if the following reactions will occur:

1. $\text{F}_2 + \text{AlCl}_3 \rightarrow \text{yes!}$

2. Chlorine gas is bubbled through a solution of sodium iodide. $\text{yes!}$

3. Liquid bromine is mixed with a solution of potassium chloride. $\text{No!}$

Predicting the products (halogens):

1. $\frac{3}{\text{F}_2} + \frac{2}{\text{AlCl}_3} \rightarrow 2\text{AlF}_3 + 3\text{Cl}_2$

2. Chlorine gas is bubble through a solution of sodium iodide.
   \[
   \text{Cl}_2 + 2\text{NaI} \rightarrow 2\text{NaCl} + \text{I}_2
   \]
Experiment on the Types of Reactions (Chalk Lab)

Purpose

How do you identify the various types of reactions?

Procedure

1. Obtain a small piece of white chalk, which is basically calcium carbonate.
2. Holding the chalk with a set of tongs, heat a piece of chalk in the hottest part of the Bunsen burner for three to five minutes.
3. While heating, add 100mL of water to a 150mL beaker.
4. Add 2-3 drops of phenolphthalein to the water. Record the initial color of the solution.
5. After heating the chalk, place the chalk into the water solution.
6. Using a stirring rod, stir the mixture in the beaker. Note any change in color of the solution.

<table>
<thead>
<tr>
<th>Initial Color of the Solution</th>
<th>Final Color of the Solution</th>
</tr>
</thead>
</table>

Analysis Questions

1. When the chalk or calcium carbonate is heated, it forms solid calcium oxide and carbon dioxide gas. Write the balanced chemical equation. Be sure to indicate the physical state of each substance and show how you can represent the heat of the Bunsen burner.

2. What type of reaction is being described in question #1?

3. In order to decompose the chalk, you relied on the combustion of methane as your source of heat. Write the balanced chemical equation for the combustion of methane, CH₄.

4. The solid calcium oxide, which is produced while the chalk is being heated, undergoes another reaction when it is placed in the water. The solid calcium oxide reacts with water to produce aqueous calcium hydroxide. Write the balanced chemical equation for this reaction. Be sure to include the physical states of all reactants and products.

5. What type of reaction is being described in question #4?  

**ANSWER IN COMPLETE SENTENCES!!**
Day 7: Homework

1. Circle any reactions that will NOT occur.
   
   \[ \text{LiCl} + \text{F}_2 \quad \text{NaF} + \text{I}_2 \]
   
   \[ \text{Na} + \text{H}_2\text{O} \quad \text{H}_2\text{O} + \text{Fe} \]

3. Using your activity series, determine if the reactions below will occur. If the reaction will not occur, write no reaction. If the reaction will occur, predict the products of the reaction and write a balanced chemical reaction.

   a. \[ \text{H}_2\text{O} + \text{Li} \rightarrow \text{LiOH} + \text{H}_2 \]
   
   b. \[ \text{H}_2\text{O} + \text{Pb} \rightarrow \text{No reaction} \]
   
   c. \[ \text{H}_2\text{O} + \text{Ca} \rightarrow \text{Ca(OH)} + \text{H}_2 \]
   
   d. \[ \text{KI} + \text{Br}_2 \rightarrow \text{KBr} + \text{I}_2 \]
   
   e. \[ \text{I}_2 + \text{NaF} \rightarrow \text{No reaction}! \]
   
   f. \[ \text{LiI} + \text{Br}_2 \rightarrow \text{LiBr} + \text{I}_2 \]
Day 8: Precipitate or Double Replacement Reactions

Ions from 2 ionic compounds switch places.

Cations (+) exchange anions (-)

\[ AB + CD \rightarrow CB + AD \]

* A and C (the cations from each compound) switch places *
* Combine inner ions with outer ions *

Identifying feature: 2 compounds on each side of the equation

Analogy: 2 boy-girl pairs are dancing, and they switch partners.

Example: \( \text{FeCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow 2 \text{NaCl} + \text{FeCO}_3 \)

---

<table>
<thead>
<tr>
<th>A soluble compound</th>
<th>An insoluble compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>- will dissolve</td>
<td>- will not dissolve</td>
</tr>
<tr>
<td>- will be labeled aqueous or (aq) in a chemical reaction.</td>
<td>- will be labeled as a solid or(s) in a chemical reaction.</td>
</tr>
</tbody>
</table>

How to determine if a substance is soluble?
You will use this chart

<table>
<thead>
<tr>
<th>Soluble compounds contain</th>
<th>Common exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>C(_2\text{H}_3\text{O}_2)^-</td>
<td>None</td>
</tr>
<tr>
<td>CH(_3\text{COO})^-</td>
<td>None</td>
</tr>
<tr>
<td>NH(^+)</td>
<td>None</td>
</tr>
<tr>
<td>NO(_3^-)</td>
<td>None</td>
</tr>
<tr>
<td>CN(^-)</td>
<td>None</td>
</tr>
<tr>
<td>Cl(^-)</td>
<td>None</td>
</tr>
<tr>
<td>ClO(_2^-)</td>
<td>None</td>
</tr>
<tr>
<td>ClO(_3^-)</td>
<td>None</td>
</tr>
<tr>
<td>ClO(_4^-)</td>
<td>None</td>
</tr>
<tr>
<td>Br(^-)</td>
<td>Compounds of Ag(^+), Pb(^{2+}), and Hg(^{2+})</td>
</tr>
<tr>
<td>I(^-)</td>
<td>Compounds of Ag(^+), Pb(^{2+}), and Hg(^{2+})</td>
</tr>
<tr>
<td>SO(_4^{2-})</td>
<td>Compounds of Sr(^{2+}), Ba(^{2+}), Pb(^{2+}), and Hg(^{2+})</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Insoluble compounds contain</th>
<th>Common exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO(_3^{2-})</td>
<td>Compounds of NH(_4^+) and the alkali metal cations</td>
</tr>
<tr>
<td>PO(_4^{3-})</td>
<td>Compounds of NH(_4^+) and the alkali metal cations</td>
</tr>
<tr>
<td>CrO(_2^{2-})</td>
<td>Compounds of NH(_4^+) and the alkali metal cations</td>
</tr>
<tr>
<td>CrO(_4^{2-})</td>
<td>Compounds of NH(_4^+) and the alkali metal cations</td>
</tr>
<tr>
<td>OH(^-)</td>
<td>Compounds of NH(_4^+), the alkali metal cations, Ca(^{2+}), Sr(^{2+}), and Ba(^{2+})</td>
</tr>
<tr>
<td>S(_2^-)</td>
<td>Compounds of NH(_4^+), the alkali metal cations, Ca(^{2+}), Sr(^{2+}), and Ba(^{2+})</td>
</tr>
</tbody>
</table>
Determine if the following compounds are soluble or insoluble by using the solubility table.

A) Sr(NO$_3$)$_2$ (aq) → **Soluble**
B) AgOH (s) → **Insoluble**
C) PbCl$_2$ (s) → **Insoluble**
D) Na$_3$PO$_4$ (aq) → **Soluble**

Predicting the Products of Double Replacement Reactions

• A precipitate is a solid that forms when two solutions are mixed.
  \[
  \text{NaCl(aq)} + \text{AgNO}_3 (aq) \rightarrow \text{AgCl(s)} + \text{NaNO}_3(aq) \quad \text{precipitate}
  \]

**TIPS FOR DOUBLE REPLACEMENT REACTIONS**

• Don't use the activity series.
• Determine if a precipitate will form.
  o Precipitate will be an insoluble product!
  o It should be marked as a solid (s).
• If a precipitate forms, a reaction occurs.
• If a precipitate does not form, then no visible reaction will occur.
• General Formula \[ AX + BY \rightarrow BY + AX \]
  The **positive** ions should switch places!

**Examples**

• \( \text{CuCl}_2 (aq) + \text{Na}_3\text{PO}_4 (aq) \rightarrow \text{NaCl(aq)} + \text{Cu}_3(\text{PO}_4)_2 (s) \)
• \( \text{NaNO}_3(aq) + \text{KCl(aq)} \rightarrow \text{KNO}_3(aq) + \text{NaCl(aq)} \)
• \( 3\text{KOH(aq)} + \text{Fe(NO}_3)_3(aq) \rightarrow \text{Fe(OH)}_3(s) + 3\text{KNO}_3(aq) \)
• \( \text{CaCl}_2(aq) + (\text{NH}_4)_2\text{SO}_4(aq) \rightarrow \text{NH}_4\text{Cl(aq)} + \text{CaSO}_4(aq) \)
  = No Reaction
Day 8: Precipitate Homework

1. Predict if the following compounds are soluble (aq) or insoluble (s). Label with (s) or (aq).
   a. $\text{KCl (aq) Soluble}$
   b. $\text{K}_2\text{S (aq) Soluble}$
   c. $\text{CaS (aq) Soluble}$
   d. $\text{BaSO}_4 (s)$ Insoluble
   e. $\text{NH}_4\text{OH (aq) Soluble}$
   f. $\text{Ba(OH)}_2 (aq)$ Soluble
   g. $\text{PbS (s)}$ Insoluble
   h. $\text{FeBr}_3 (aq)$ Soluble

2. Determine if the following pairs of solutions will form a precipitate.
   - If they do form a precipitate, write the balanced reaction.
   - If they do not form a precipitate, write NVR for no visible reaction.
   a. $2\text{AgNO}_3 (aq) + \text{CaCl}_2 (aq) \rightarrow \text{Ca(NO}_3)_{2(aq)} + 2\text{AgCl (s)}$
   b. $\text{Pb(NO}_3)_{2(aq)} + \text{K}_2\text{H}_3\text{O}_2 (aq) \rightarrow \text{KNO}_3 (aq) + \text{PbC}_2\text{H}_3\text{O}_2 \Rightarrow \text{NO REACTION}$
   c. $\text{CuSO}_4 (aq) + 2\text{KOH (aq)} \rightarrow \text{K}_2\text{SO}_4 (aq) + \text{Cu(OH)}_2 (s)$
   d. $\text{Na}_3\text{PO}_4 (aq) + \text{FeCl}_3 (aq) \rightarrow \text{FePO}_4 (s) + 3\text{NaCl (aq)}$
   e. $\text{AlCl}_3 + 3\text{NaOH} \rightarrow 3\text{NaCl (aq)} + \text{Al(OH)}_3 (s)$
Day 9: Compound Formation Lab

Objectives

- Predict the products of the double replacement reactions.
- Identify the formation of a precipitate experimentally.

Background

In chemistry there are four signs that indicate that a chemical change has most likely taken place.

1. A Change in Color
2. Formation of a Gas (bubbles)
3. Release or Absorption of Energy (temperature change)
4. Odor
5. Formation of a Precipitate (solid)

A precipitate is a solid that is formed when two aqueous solutions are mixed. It usually separates from the mixture by settling to the bottom of the container. Precipitates often appear to be cloudy, milky, gelatinous or grainy.

Pre-Lab Questions

1. In the following equation, circle the precipitate.

   \[ \text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow 2 \text{NaCl}(\text{aq}) + \text{BaSO}_4(\text{s}) \]

2. Write the chemical equation for the following double replacement reactions. Predict the products of the reactions, and using your solubility chart identify which product is the precipitate (insoluble), if there is one. If no precipitate is formed, then write NO VISIBLE REACTION (NVR) as your product.

   a. A solution of ammonium hydroxide is mixed with aqueous iron (III) chloride.

   b. Ammonium hydroxide dissolved in water reacts with a solution of zinc (II) chloride.

   c. Aqueous lead (II) chloride is mixed with aqueous ammonium hydroxide.

   d. A solution of sodium hydroxide is mixed with aqueous zinc (II) chloride.

   e. Aqueous sodium hydroxide reacts with iron (III) chloride.
f. A solution of lead (II) chloride is mixed with aqueous sodium hydroxide.

g. A solution of sodium iodide reacts with aqueous copper (II) chloride.

**Procedure**

1. Follow the directions on the Reaction Surface Chart.
2. Be sure to place all pipettes tip up to prevent contamination!
3. Record the appropriate data in your data table below.
4. To clean up, soak up your solutions using a paper towel. Dispose of the paper towel in the trash can.
5. Wash your overhead with water.
6. Dry your overhead completely!

**Data**

In each box, record the precipitate (insoluble, or solid, compound) produced by the reaction, and the color of the precipitate if one is formed. Record "NVR" for no visible reaction if no precipitate is formed.

<table>
<thead>
<tr>
<th></th>
<th>ZnCl₂</th>
<th>FeCl₃</th>
<th>PbCl₂</th>
<th>NaI</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄OH</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaOH</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CuCl₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Analysis** **WRITE IN COMPLETE SENTENCES**

1. If a product is soluble, why do you think we don’t see a precipitate?

2. If a product is insoluble, why do you think that a precipitate is visible?

3. Predict the products (and circle the precipitate) if a solution of sodium sulfate is mixed with calcium nitrate. (Be sure to write the BALANCED chemical equation!)
Nomenclature Practice:
Write the correct formula
1. Aluminum chloride $\text{AlCl}_3$
2. Lead (II) nitrate $\text{Pb(NO}_3\text{)}_2$
3. Sulfur dioxide $\text{SO}_2$
4. Sulfurous acid $\text{H}_2\text{SO}_3$
5. Copper (II) oxide $\text{CuO}$
6. Dinitrogen tetroxide $\text{N}_2\text{O}_4$
7. Calcium sulfate $\text{CaSO}_4$
8. Hydrofluoric acid $\text{HF}$
9. Barium phosphate $\text{Ba}_3(\text{PO}_4)_2$
10. Nitric acid $\text{HNO}_3$

What symbols represent the following terms?
A reaction is heated: $\Delta$
A catalyst is used: $\text{cat}\rightarrow$
A solution of: $(\text{aq})$
A precipitate: $(\text{s})\text{ or }\downarrow$
Gaseous: $(\text{g})\text{ or }\uparrow$
Vapor: $(\text{g})$

11. Write a balanced chemical equation for each word and appropriate symbols, formulas, and balancing.
   a. Sodium reacts with water to form hydrogen gas and sodium hydroxide.

$$2\text{Na} + 2\text{H}_2\text{O} \rightarrow \text{H}_2(\text{g}) + 2\text{NaO}_2\text{H}$$

   b. Solid lead (IV) oxide is heated, decomposing to lead metal and oxygen gas.

$$\text{PbO}_2(\text{s}) \xrightarrow{\Delta} \text{Pb}(\text{s}) + \text{O}_2(\text{g})$$

   c. In the presence of a catalyst, nitrogen and hydrogen gases react to form gaseous ammonia.

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \xrightarrow{\text{cat}} 2\text{NH}_3(\text{g})$$

12. Complete the Single Replacement reactions. Then use the activity series of metals and Halogens to determine which of the below reactions will occur. If a reaction will occur, write the balanced equation.
   a. $\text{Au}(\text{s}) + \text{KNO}_3(\text{aq}) \rightarrow \text{NR}$
   b. $\text{Zn}(\text{s}) + 2\text{AgNO}_3(\text{aq}) \rightarrow \text{Zn(NO}_3\text{)}_2 + 2\text{Ag}$
c. \( 2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3 + 3H_2 \)

d. \( MgBr_2 + F_2 \rightarrow Br_2 + MgF_2 \)

e. \( Cu(s) + H_2O(l) \rightarrow NR \)

f. \( 2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3 + 3Cu \)

g. \( NaCl + Br_2 \rightarrow NR \)

e. \( 2H_2O + 2Na \rightarrow 2NaOH + H_2 \)

13. Balance the following combustion reactions.

a. \( C_4H_8 + 6O_2 \rightarrow 4CO_2 + 4H_2O \)

b. \( 2C_6H_{18} + 25O_2 \rightarrow 10CO_2 + 18H_2O \)

14. Predict the products in the following double replacement reactions. Identify the precipitate formed when solutions of these ionic compounds are mixed. If there is NO PRECIPITATE (no insoluble solid), write NVR!

a. \( [CaSO_4] + [BaCl_2] \rightarrow CaCl_2(aq) + BaSO_4(s) \)

b. \( [Al_2(SO_4)_3] + [NH_4Cl] \rightarrow (NH_4)_2SO_4(aq) + AlCl_3(aq) \)

c. \( 2AgNO_3 + K_2S \rightarrow Ag_2S(s) + 2KNO_3(aq) \)

d. \( CaCl_2 + Pb(NO_3)_2 \rightarrow Ca(NO_3)_2(aq) + PbCl_2(s) \)

e. \( Ca(NO_3)_2 + NaBr \rightarrow NaN_3(aq) + CaBr_2(aq) \)

<table>
<thead>
<tr>
<th>Balance the following</th>
<th>C,S,D, SR,DR</th>
<th>ppt, acid-base, redox</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. $2\text{Hf} + 2\text{N}_2 \rightarrow \text{Hf}_3\text{N}_4$</td>
<td>S</td>
<td>Redox</td>
</tr>
<tr>
<td>b. $\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2$</td>
<td>SR</td>
<td>Redox</td>
</tr>
<tr>
<td>c. $\text{C}_2\text{H}_6 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$</td>
<td>C</td>
<td>Redox</td>
</tr>
<tr>
<td>d. $\text{Hg(NO}_3)_2 + 2\text{NH}_4\text{OH} \rightarrow \text{Hg(OH)}_2 + 2\text{NH}_4\text{NO}_3 + \text{H}_2\text{O}$</td>
<td>DR</td>
<td>PPT</td>
</tr>
<tr>
<td>e. $\text{H}_2\text{O}_2 \rightarrow \text{H}_2 + \text{O}_2$</td>
<td>D</td>
<td>Redox</td>
</tr>
<tr>
<td>f. $2\text{KBr} + \text{Cl}_2 \rightarrow 2\text{KCl} + \text{Br}_2$</td>
<td>SR</td>
<td>Redox</td>
</tr>
</tbody>
</table>

16. Based on the oxidation numbers known for the other element(s), determine the oxidation number of the underlined element.

| a) Na$^+$ SO$_4^{2-}$ | O - 4(-2) = -8 |
| b) K$_2$CrO$_4$ | O - 4(-2) = -8 |
| c) KCl | O - 4(-2) = -8 |
| d) Ag | O - 4(-2) = -8 |

18. In an oxidation-reduction reaction, reduction is defined as the —

A. gain of protons.  
B. gain of electrons.  
C. loss of electrons.  
D. loss of protons.
19. Which changes occur when Pt²⁺ is reduced?

A. The Pt²⁺ loses electrons and its oxidation number decreases.
B. The Pt²⁺ gains electrons and its oxidation number decreases.
C. The Pt²⁺ gains electrons and its oxidation number increases.
D. The Pt²⁺ loses electrons and its oxidation number increases.

20. The ionic chemical equation below is balanced.

\[ \text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s) \]

Which equation represents the oxidation half-reaction?

A. \( \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2 \text{e}^- \)
B. \( \text{Cu}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Cu}(s) \)
C. \( \text{Zn}^{2+}(aq) \rightarrow \text{Zn}(s) + 2 \text{e}^- \)
D. \( \text{Zn}(s) + 2 \text{e}^- \rightarrow \text{Zn}^{2+}(aq) \)

21. The chemical equation below represents a redox reaction.

\[ 4 \text{HCl} + \text{MnO}_2 \rightarrow \text{MnCl}_2 + 2 \text{H}_2\text{O} + \text{Cl}_2 \]

What occurs during this reaction?

A. The manganese is oxidized, and its oxidation number changes from +2 to +4.
B. The manganese is reduced, and its oxidation number changes from +4 to +2.
C. The manganese is oxidized, and its oxidation number changes from +4 to +2.
D. The manganese is reduced, and its oxidation number changes from +2 to +4.

22. Which half-reaction correctly represents oxidation?

A. \( \text{Fe}(s) + 2 \text{e}^- \rightarrow \text{Fe}^{2+}(aq) \)
B. \( \text{Fe}(s) \rightarrow \text{Fe}^{2+}(aq) + 2 \text{e}^- \)
C. \( \text{Fe}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Fe}(s) \)
D. \( \text{Fe}^{2+}(aq) \rightarrow \text{Fe}(s) + 2 \text{e}^- \)

Quick Nomenclature Review:
• Ionic Compounds: Name the metal unchanged, Name the nonmetal with “ide” on the end.
• Ionic with multivalent metals: Name the metal unchanged and write the roman numeral for the oxidation of the metal, Name the nonmetal with “ide” on the end.

<table>
<thead>
<tr>
<th>Roman Numeral</th>
<th>I</th>
<th>II</th>
<th>III</th>
<th>IV</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
<th>VIII</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxidation #</td>
<td>1+</td>
<td>2+</td>
<td>3+</td>
<td>4+</td>
<td>5+</td>
<td>6+</td>
<td>7+</td>
<td>8+</td>
</tr>
</tbody>
</table>

• Covalent/Molecular Compounds: Use the prefixes listed below to name the compounds

<table>
<thead>
<tr>
<th></th>
<th>1*</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mono-</td>
<td>Di-</td>
<td>Tri-</td>
<td>Tetra-</td>
<td>Penta-</td>
<td>Hexa-</td>
<td>Hepta-</td>
<td>Octa-</td>
<td>Nona-</td>
<td>Deca-</td>
</tr>
</tbody>
</table>

• Ionic Compounds with polyatomic ions: Write both the metal and the polyatomic ions AS IS.
• Acids: Use the chart below

<table>
<thead>
<tr>
<th>ION TYPE</th>
<th>ION ENDING</th>
<th>ACID NAME BEGINNING</th>
<th>ACID ENDING</th>
</tr>
</thead>
<tbody>
<tr>
<td>Polyatomic</td>
<td>-ate</td>
<td>NO hydro- beginning</td>
<td>-ic</td>
</tr>
<tr>
<td>Polyatomic</td>
<td>-ide</td>
<td>NO hydro- beginning</td>
<td>-ic</td>
</tr>
<tr>
<td>Monatomic</td>
<td>-ide</td>
<td>hydro- beginning</td>
<td>-ic</td>
</tr>
</tbody>
</table>
### Periodic Table of the Elements

<table>
<thead>
<tr>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>He</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
</tr>
<tr>
<td>1</td>
<td>4</td>
<td>3</td>
<td>4</td>
<td>13</td>
<td>14</td>
<td>15</td>
<td>16</td>
</tr>
<tr>
<td>He</td>
<td>2</td>
<td>8</td>
<td>8</td>
<td>8</td>
<td>16</td>
<td>17</td>
<td>18</td>
</tr>
</tbody>
</table>

#### References Materials

**Polyatomic Ions**

- Acetate: $\text{C}_2\text{H}_3\text{O}_2^-$, $\text{CH}_3\text{COO}^-$
- Ammonium: $\text{NH}_4^+$
- Carbonate: $\text{CO}_3^{2-}$
- Chlorate: $\text{ClO}_3^-$
- Cyanide: $\text{CN}^-$
- Dichromate: $\text{Cr}_2\text{O}_7^{2-}$
- Dichromate: $\text{Cr}_2\text{O}_7^{2-}$
- Hydrogen carbonate: $\text{HCO}_3^-$
- Hydroxide: $\text{OH}^-$
- Hypochlorite: $\text{ClO}^-$
- Nitrate: $\text{NO}_3^-$
- Nitric: $\text{N}_2\text{O}_5^-$
- Perchlorate: $\text{ClO}_4^-$
- Permanganate: $\text{MnO}_4^-$
- Phosphate: $\text{PO}_4^{3-}$
- Sulfate: $\text{SO}_4^{2-}$
- Sulfite: $\text{SO}_3^-$

**Solubility of Common Ionic Compounds in Water**

- **Soluble compounds contain**
  - $\text{C}_2\text{H}_3\text{O}_2^-$, $\text{CH}_3\text{COO}^-$
  - $\text{NH}_4^+$
  - $\text{ClO}_3^-$
  - $\text{CN}^-$
  - $\text{ClO}_4^-$
  - $\text{Br}^-$
  - $\text{Cl}^-$
  - $\text{SO}_4^{2-}$

- **Common exceptions**
  - None

- **Insoluble compounds contain**
  - $\text{CO}_2^-$
  - $\text{PO}_4^{3-}$
  - $\text{Cr}_2\text{O}_7^{2-}$
  - $\text{OH}^-$

- **Common exceptions**
  - Compounds of $\text{NH}_4^+$ and the alkali metal cations
  - Compounds of $\text{NH}_4^+$ and the alkali metal cations
  - Compounds of $\text{NH}_4^+$ and the alkali metal cations
  - Compounds of $\text{NH}_4^+$ and the alkali metal cations

**Activity Series**

- Metal:
  - Lithium
  - Potassium
  - Sodium
  - Calcium
  - Barium
  - Strontium
  - Magnesium
  - Aluminum
  - Titanium
  - Vanadium
  - Chromium
  - Iron
  - Cobalt
  - Nickel
  - Tin
  - Lead
  - Mercury
  - Silver
  - Gold