Chemical Reactions: Introduction to Reaction Types

Introduction:
In this experiment, you will observe (on video and in lab) examples of the six basic types of chemical reactions which are described below. You will write your observations of the reactions. Based on those observations and the given word equations, you will write complete, balanced chemical equations to effectively communicate the chemistry of the reactions.

Six Reaction Types:

1. Combination: A + B \rightarrow AB
In a combination reaction, a new compound can be formed in one of three ways, by combining: a) 2 elements, b) 1 element and 1 binary compound (consisting of 2 elements), or c) 2 binary compounds. The following are examples of combination reactions:

- The rusting of iron: 4Fe (s) + 3O_2 (g) \rightarrow 2Fe_2O_3 (s)
- The formation of one kind of acid rain: SO_3 (g) + H_2O (l) \rightarrow H_2SO_4 (aq)

2. Decomposition: AB \rightarrow A + B
In a decomposition reaction, a compound is broken 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. The following are examples of decomposition reactions:

- Heating mercury (II) oxide produces oxygen gas: 2HgO (s) \rightarrow 2Hg (l) + O_2 (g).
- Leaving the cap off the carbonated soft drink bottle allows the carbonic acid to release carbon dioxide: H_2CO_3 (aq) \rightarrow H_2O (l) + CO_2 (g).

3. Single Replacement: A + BC \rightarrow AC + B
In this type of reaction, a more “active metal” displaces another element in solution. These reactions can be further classified as a solid metal reacting with a) a metal ion solution, b) an acid solution, or c) water. An example of each is provided below:

- a) When a solid metal reacts with a metal solution, the solid metal’s ions go into solution while the metal ions originally in solution plate out onto the surface of solid metal—e.g., Zn (s) + CuSO_4 (aq) \rightarrow ZnSO_4 (aq) + Cu (s).
- b) When a solid metal reacts with an acid, the metal replaces hydrogen in the acid to produce hydrogen gas while the metal ion goes into solution with the anion from the acid—e.g., Mg (s) + 2HCl (aq) \rightarrow H_2 (g) + MgCl_2 (aq).
- c) When a solid metal reacts with water, the metal replaces hydrogen in the water to produce hydrogen gas while the metal ion goes into solution with hydroxide ion—e.g., Ca (s) + 2 H_2O (aq) \rightarrow H_2 (g) + Ca(OH)_2 (aq).

To predict whether or not a single-replacement reaction will occur, we refer to the Activity Series for Metals (shown on the next page).
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

If a solid metal is “more active” – i.e., higher on the Activity Series – than the metal ion in solution or hydrogen for acids, the more active metal will displace the less active ion, so a reaction occurs. The more active metal goes into solution while the less active ion either plates out for a metal ion or bubbles out as a gas for hydrogen ion. If the solid metal is less active than the metal ion or hydrogen in the compound, then no reaction occurs. Only six metals (Li, K, Ba, Sr, Ca, and Na) – called “active metals” – react directly with water to produce hydrogen gas and a metal hydroxide solution. These active metals are the first six metals in the Activity Series.

4. Acid-Base Neutralization: \( \text{HX (aq)} + \text{YOH (aq)} \rightarrow \text{H}_2\text{O (l)} + \text{YX (aq)} \)

These reactions occur between an acid and a base. In general, *acids* are compounds that **produce 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, \( \text{HCl(aq)} \), nitric acid, \( \text{HNO}_3(aq) \), and sulfuric acid, \( \text{H}_2\text{SO}_4(aq) \).

*Bases* are compounds that **produce hydroxide ions** \( (\text{OH}^-) \) when dissolved in water. A few common bases are sodium hydroxide, \( \text{NaOH} \), potassium hydroxide, \( \text{KOH} \), calcium hydroxide, \( \text{Ca(OH)}_2 \), and barium hydroxide, \( \text{Ba(OH)}_2 \). Other types of bases contain carbonate ion, \( \text{CO}_3^{2-} \), and hydrogen carbonate (or bicarbonate) ion, \( \text{HCO}_3^- \).

The two types of acid-base neutralization reactions involve a) an acid reacting with a base (containing the hydroxide ion \( (\text{OH}^-) \)) to produce water and a salt (an ionic compound) or b) an acid reacting with a base containing carbonate \( (\text{CO}_3^{2-}) \) or hydrogen carbonate ion \( (\text{HCO}_3^-) \) to produce water, carbon dioxide gas, and a salt. An example of each is provided below:

a) When **an acid reacts with a base containing hydroxide ion** \( (\text{OH}^-) \) **to produce water and a salt**, the hydrogens from the acid combine with the hydroxide from the base to form water while the salt is formed by combining the cation from the base with the anion from the acid. The following is an example of this type of reaction:

\[
\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{H}_2\text{O (l)} + \text{NaCl (aq)}.
\]

b) When **an acid reacts with a base containing carbonate** \( (\text{CO}_3^{2-}) \) **or hydrogen carbonate ion** \( (\text{HCO}_3^-) \) **to produce water, carbon dioxide gas, and a salt**, the hydrogens from the acid combine with the carbonate or hydrogen carbonate from the base to form water and carbon dioxide gas while the salt is formed by combining the cation from the base combining with the anion from the acid:

\[
\text{HCl (aq)} + \text{NaHCO}_3 (aq) \rightarrow \text{H}_2\text{O (l)} + \text{CO}_2 (g) + \text{NaCl (aq)}.
\]

5. Combustion Reactions: Hydrocarbon \( \text{C}_x\text{H}_y + \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + \text{H}_2\text{O (g)} \)

In a combustion reaction, a hydrocarbon (composed of C and H) or a hydrocarbon derivative (composed of C, H, and O) is burned in oxygen to produce carbon dioxide gas and steam. One example is the combustion of methane (natural gas): \( \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O (g)} \)
6. Double Replacement/Precipitation:  $AB + CD \rightarrow AD + CB$

These reactions involve the mixing of two aqueous ionic compounds to produce a **precipitate**, an insoluble ionic compound. The products of a double-replacement/precipitation reaction can be predicted by switching the cations of the two compounds and using the Solubility Rules (see below) to determine if the compounds produced are soluble or insoluble.

### Solubility Rules for Ionic Compounds in Water

<table>
<thead>
<tr>
<th>The compound is SOLUBLE if it has:</th>
<th>The compound is INSOLUBLE if it has:</th>
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<tbody>
<tr>
<td>1. Li⁺, Na⁺, K⁺, or NH₄⁺ ion (ALWAYS!)</td>
<td>5. $CO_3^{2-}$, $CrO_4^{2-}$, $PO_4^{3-}$, except compounds with Li⁺, Na⁺, K⁺, NH₄⁺ are soluble</td>
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<tr>
<td>2. $C_2H_3O_2^-$, NO₃⁻, ClO₄⁻</td>
<td>6. $S^{2-}$, except compounds with Li⁺, Na⁺, K⁺, NH₄⁺, Ca²⁺, Sr²⁺, Ba²⁺ are soluble</td>
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<tr>
<td>3. $Cl^-$, $Br^-$, or $I^-$, except compounds with Ag⁺, Pb⁺², Cu²⁺, and Hg₂⁺² are insoluble</td>
<td>7. Hydroxide ion, OH⁻, except compounds with Li⁺, Na⁺, K⁺, NH₄⁺ are soluble</td>
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<tr>
<td>4. $SO_4^{2-}$, except compounds with Ag₂SO₄, CaSO₄, SrSO₄, BaSO₄, PbSO₄, and Hg₂SO₄ are insoluble</td>
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</tr>
</tbody>
</table>

Soluble ionic compounds will dissolve in water, so their physical states are indicated as aqueous, (aq), while insoluble ionic compounds will not dissolve in water, so their physical states are indicated as solid, (s). For a precipitation reaction to occur, at least one of the products must be insoluble; if both products are soluble, then no reaction occurs. The presence of a precipitate is observed in the lab as a cloudy mixture that results when two solutions are mixed.

**Molecular Level Drawing of Solutions**

We can represent ionic solids in water as ions if they are soluble and as solids if they are insoluble. Refer to the solubility rules when looking at the examples below. Notice that the solvent water molecules are not shown.
When we balance chemical reactions, we use coefficients so that the “Law of conservation of matter” is not violated: Matter cannot be created or destroyed. The following example shows how we draw one HCl\((aq)\) versus three HCl\((aq)\). The coefficient tells us how many of the molecules are present, not that the molecules are bonded together.

Example of a double replacement reaction: the reaction between aqueous lead(II) nitrate with aqueous potassium bromide, as shown on the next page:

\[
\text{Pb(NO}_3\text{)}_2(aq) + \text{KBr}(aq) \rightarrow ?
\]

When we pour the two solutions together we have four types of ions in solution: Pb\(^{2+}\)(aq), NO\(^3\text{-}\)(aq), K\(^+\)(aq), and Br\(^-\)(aq). (Remember (aq) means dissolved in water) Consider the possible combinations of ions hitting in the solution. Pb\(^{2+}\) and NO\(^3\text{-}\) ions can hit each other but will not bond because otherwise they would not have dissociated in the first place. The same goes for K\(^+\) and Br\(^-\) ions. Pb\(^{2+}\) and K\(^+\) ions can hit each other, but positive charges repel. Similarly, the negatively charged NO\(^3\text{-}\) and Br\(^-\) ions will not bond because negative charges repel. The only possible combinations that could form products are Pb\(^{2+}\) with Br\(^-\), to form lead(II) bromide, PbBr\(_2\), and K\(^+\) with NO\(^3\text{-}\), to form potassium nitrate, KNO\(_3\), as shown below:

\[
\text{Pb(NO}_3\text{)}_2(aq) + \text{KBr}(aq) \rightarrow \text{PbBr}_2 + \text{KNO}_3
\]

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 soluble or insoluble. Based on Rule #3, 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:

\[
\text{Pb(NO}_3\text{)}_2(aq) + \text{KBr}(aq) \rightarrow \text{PbBr}_2(s) + \text{KNO}_3(aq)
\]

Because a precipitate formed—PbBr\(_2\)(s)—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:

\[ \text{Pb(NO}_3\text{)}_2(aq) + 2 \text{ KBr}(aq) \rightarrow \text{PbBr}_2(s) + 2 \text{ KNO}_3(aq) \]

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.

If you carry out this reaction in a laboratory, you will observe that the lead(II) nitrate and potassium bromide solutions are both clear and colorless. When you mix them, the resulting solution 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, two clear solutions becoming cloudy when combined is often interpreted as experimental evidence of a precipitate forming.

The following animation will help you understand precipitation reactions:
http://www.mpcfaculty.net/mark_bishop/precipitation.htm