Introduction

We will now describe several different classes of chemical reactions, first briefly and then in detail.

1. **Precipitation reactions** involve the formation of a **precipitate** (a solid substance that falls out of solution). You will do many of these in your labs this semester.
2. **Neutralization** reactions involve the transfer of *protons* (hydrogen ions) between reagents.
3. **"Redox"**, or **oxidation-reduction** reactions involve the transfer of *electrons* between reagents. There are several subclasses of these reactions.

Precipitation reactions

A simple-to-understand type of chemical reaction is the precipitation reaction. At this point, you have done some of these in lab. To perform a precipitation reaction, you mix solutions of the reactants and observe the formation of a solid material, the **precipitate**, from the two clear solutions. A good example would be the reaction between silver nitrate and sodium chloride.

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

The net result of this reaction is the formation of solid **silver chloride**, so this is a precipitation reaction. The sodium nitrate exists as free ions in solution.

Silver nitrate and sodium chloride are ionic species - and ionic theory tells us that they exist in solution as free ions. You can think of the mixture of silver nitrate and sodium chloride as an "ion soup" that lets all ions interact with each other. **A precipitation will occur if there is a pair of ions that can join together to form an insoluble compound!**

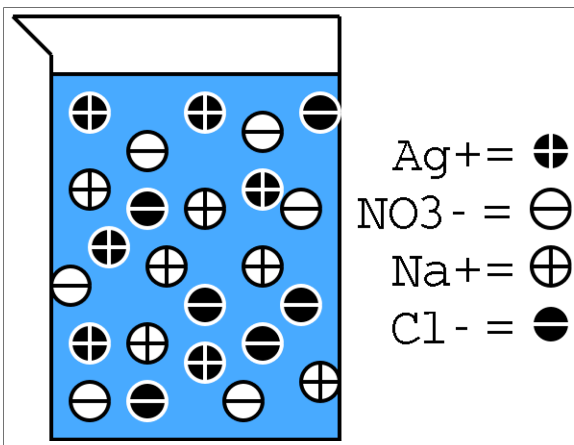
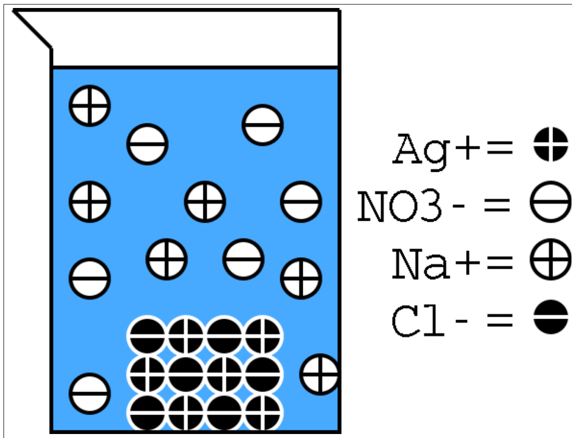


Illustration 1 - "Ion soup" formed from mixing $AgNO_3$ and NaCl solutions

- At the instant the silver nitrate and sodium chloride solutions are mixed, you have a mixture of all four ions in solution.
- Oppositely charged ions will tend to attract each other, leading to four possible pairings: $AgNO_3$, AgCl, $NaNO_3$, and NaCl.
- If any of the possibilities are **insoluble**, they will fall out of solution as a precipitate.

How can you tell if a mixture of two ionic substances will result in a precipitate (i.e. a chemical reaction)? Use the solubility rules to determine if a combination of ions will be insoluble. In our example, there are four possible combinations of ions: **$AgNO_3$, AgCl, $NaNO_3$, and NaCl**. Using the solubility rules we've discussed previously, all of the possibilities **except AgCl** are soluble, so a reaction does take place. **AgCl is the precipitate**.



- AgCl is insoluble, so it will crystallize and fall out of solution.
- Sodium and nitrate ions (Na^+ and NO_3^-) remain in solution as free ions.

Illustration 2 - AgCl precipitate forming from mixing $AgNO_3$ and NaCl solutions

Let's look at another mixture: NaCl is mixed with $CaBr_2$. The possible compounds formed from this mixture are NaCl, NaBr, $CaCl_2$, and $CaBr_2$. Since all four of these compounds are **soluble**, **no precipitate forms** (and there is **no reaction**).

But there's an easier way to do this. Note what happens in the overall reaction between silver nitrate and sodium chloride:

$$\begin{array}{ccccccc} AgNO_3(aq) & + & NaCl(aq) & \rightarrow & AgCl(s) & + & NaNO_3(aq) \\ \text{A-B} & + & \text{C-D} & \rightarrow & \text{A-D} & + & \text{C-B} \end{array}$$

The letters below the equation are there to make it a little easier to see what happens. The ions have simply switched partners, and it is for this reason that these reactions are often called **exchange reactions**. You can write possible products for a precipitation reaction by simply making the ions change partners. Let's try another.

$$Na_2CO_3(aq) + AgNO_3(aq) \rightarrow \text{what?}$$

Exchange the ions. Pair up Na^+ with NO_3^- and Ag^+ with CO_3^{2-}. When you're pairing you should ignore subscripts from the original compounds unless they're **part of the ion**, so the products are $NaNO_3$ and Ag_2CO_3. The charges on each product compound **must sum to zero**.

$$Na_2CO_3(aq) + AgNO_3(aq) \rightarrow NaNO_3 + Ag_2CO_3$$

This equation is not balanced (and we need to balance it), but first, let's check to make sure there is a reaction. There's no sense balancing a reaction that doesn't occur!

Sodium nitrate is **soluble** - all sodium compounds **and** all nitrates are soluble. Silver carbonate is insoluble (it's a carbonate not hooked to a "rule 1" cation), so it is the precipitate and a reaction **does occur**. Let's balance and write the equation in the proper form:

$$Na_2CO_3(aq) + \underline{2}AgNO_3(aq) \rightarrow \underline{2}NaNO_3(aq) + Ag_2CO_3(s)$$

And let's try *one more* example.

$$CuCl_2(aq) + Na_2SO_4(aq) \rightarrow \text{what?}$$

Again, we'll let our ions exchange partners. The possible products are $CuSO_4$ and NaCl. The sodium chloride is soluble. The copper(II) sulfate is also soluble, as most sulfates are soluble and copper's not listed as an exception. So we don't need to go any further:

$$CuCl_2(aq) + Na_2SO_4(aq) \rightarrow NR \text{ (no reaction)}$$

<u>Neutralization reactions (also called acid-base reactions)</u>

Arrhenius (who gave us the ionic theory of solutions) observed that certain substances behaved in particular ways when dissolved in aqueous solution. He called these substances acids and bases, and noted that acids would react with bases in a predictable way.

Arrhenius said that **acids** were materials that releases H^+ (hydrogen ions, or protons) when dissolved in water. **Bases,** on the other hand, release OH^- when dissolved in water. The classic examples of these types of substance are HCl (hydrochloric acid) and NaOH (sodium hydroxide, a base).

Arrhenius's definition involves H^+ and OH^-. If you put those together, you get the following reaction:

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

This equation is the net ionic equation of a **neutralization reaction** between an "Arrhenius" acid and an "Arrhenius" base - like the reaction between hydrochloric acid and sodium hydroxide.

$$\mathbf{H}Cl(aq) + Na\mathbf{OH}(aq) \rightarrow \mathbf{H_2O(}\boldsymbol{l}\mathbf{)} + NaCl(aq)$$

This is the classic example of a neutralization reaction between an acid and a base.

*Helpful hint: This reaction is simply another **exchange reaction**, where water molecules form instead of precipitates.*

If you look closely at the neutralization reaction, you'll see what really happens in the reaction is the transfer of a **proton** (here, that's another name for H^+ ion … not a part of an atomic nucleus) from the acid to the base. Bronsted and Lowry noticed this and applied the acid-base nomenclature to a wider variety of substances. They said that an **acid** was any species that could **donate a proton** to another species, and a **base** was any species that could **accept a proton** from another species. A **neutralization reaction**, then, was a **proton-transfer reaction**.

For example, let's look at the reaction between hydrochloric acid and ammonia through Bronsted-Lowry's viewpoint.

$$\mathbf{H}Cl(aq) + NH_3(aq) \rightarrow N\mathbf{H}_4Cl(aq)$$

What's happened here? The HCl has **donated a proton** to the NH_3, leaving Cl^-. The NH_3 becomes NH_4^+ when it **accepts the proton** and the salt ammonium chloride is formed. In the Bronsted-Lowry view, the HCl is an acid and the NH_3 is a base and this transfer of a proton from the acid to the base is a neutralization reaction.

<u>Strong vs. weak</u>

Acids and bases can be classified as "strong" or "weak". So, what makes a strong acid/base different from a weak acid/base?

Strong acids are **completely ionized in water**. In other words, strong acids are strong electrolytes. A common **strong acid** is **hydrochloric acid**, which does this in water:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Hydrochloric acid is a strong acid. This means that when HCl dissolved, no molecular HCl exists in solution - it's all H^+ and Cl^- ions.

Acetic acid (the main ingredient in vinegar) is a common **weak acid**. It does this in water:

$$HC_2H_3O_2(aq) \rightarrow H^+(aq) + C_2H_3O_2^-(aq)$$

When acetic acid dissolves, the above reaction takes place. However, the reaction does not proceed very far. In fact, most of the acetic acid remains as the molecule $HC_2H_3O_2$. Relatively little (compared to a strong acid) H^+ is released by a weak acid.

Since they are so common, you should know a few of the strong acids by name and formula.

Strong acid's name	***Chemical formula***
sulfuric acid	H_2SO_4
hydrochloric acid	HCl
nitric acid	HNO_3

Strong bases (like strong acids) are also completely ionized in water, while weak bases only ionize to a small degree. You should be able to recognize some common strong bases. Common strong bases are the **soluble metal hydroxides**. The two most common strong bases we use in lab are **NaOH** (sodium hydroxide) and **KOH** (potassium hydroxide).

Both strong and weak acids/bases will react in acid-base reactions. Whether the acid or base is strong or weak **does not change** the form of the acid-base reaction.

Summary

In this note pack, we have introduced two of the three common reaction types that you'll need to understand: precipitation reactions (where a solid product is formed from mixing two ionic compounds in aqueous solution) and neutralization reactions (where an acid and a base react together and neutralize each other). Both of these kinds of reaction can be treated as exchange reactions, where ions effectively "switch partners" to form new compounds. Next, we will discuss **oxidation-reduction reactions**.