**Review of Solutions Unit**

**Solutions** are homogeneous mixtures made of a **solvent** and one or more **solutes.**

**Solvents** are the medium in which thing get dissolved or mixed in.

**Solutes** are the substances that dissolve within the solvent**.**

**Aqueous solutions** are homogeneous mixtures where the solvent is water.

**Concentration** is the ratio of the amount of solute contained within a solution.

Solutions come in varying degrees of concentration. They can be dilute, concentrated, saturated or unsaturated.

We will start by breaking solutions into three categories:

**Strong Electrolytes, Weak Electrolytes** and **Nonelectrolytes.**

A **Strong Electrolyte** is a substance that when dissolved in an aqueous solution will conduct electricity well because it will create mobile ions in the water. Many salt solutions are electrolytes, A salt breaking into ions is called **dissociation.**

Example: NaCl*(s)* Na+*(aq)* + Cl-*(aq)*

Strong electrolytes generally dissociate almost completely, with only a few of the molecules remaining undissociated.

Other examples of strong electrolytes include strong acids and strong bases.

Some examples of strong acids include: HCl, H2SO4, HNO3, HBr, HI, HClO4

Some examples of strong bases include: NaOH, KOH, LiOH, Ba(OH)2

Acids don’t dissociate in water, they get ***ionized***in water. However, the result is similar because mobile ions still are formed.

Example: HCl*(aq)* + H2O*(l)*  H3O +*(aq)* + Cl-*(aq)*

**Precipitation Reactions**

A **precipitation reaction** refers to the formation of an insoluble salt when two solutions containing soluble salts are combined. The insoluble salt that falls out of solution is known as the **precipitate**, hence the **reaction's** name. **Precipitation reactions** can help determine the presence of various ions in solution.

Example: **AgNO3*(aq)* + KCl*(aq)* KNO3*(aq)* + AgCl*(s)***

A **molecular equation** is sometimes simply called a balanced **equation**. In a **molecular equation**, any ionic compounds or acids are represented as neutral compounds using their chemical **formulas**. The state of each substance is indicated in parentheses after the formula.

Example: 2NaOH*(aq)* + MgCl2*(aq)* → 2NaCl*(aq)* + Mg(OH)2*(s)*

Having learned that some substances are strong electrolytes and exist entirely as ions in solution, we can write these chemical equations in a way that better represents what actually happens in solution.

The equation above is a [molecular equation](file:///C%3A%5CUsers%5CCentralScienceLive%5CChapter04%5CCH04_4.2_Main.html##) in which none of the species is represented as ionized. A molecular equation shows the complete chemical formulas for the reactants and products. We know, however, that several of the species in the equation dissociate completely in solution. We convert the equation to a [**complete ionic equation**](file:///C%3A%5CUsers%5CCentralScienceLive%5CChapter04%5CCH04_4.2_Main.html##) by identifying the strong electrolytes and representing them as separated ions. The equation above becomes

The ionic equation reveals that two of the species in solution (sodium ion and hydroxide ion) do not undergo any change in the course of the reaction. Ions that are present but play no role in the reaction are called [**spectator ions**](file:///C%3A%5CUsers%5CCentralScienceLive%5CChapter04%5CCH04_4.2_Main.html##). Eliminating the spectator ions from both sides of the equation gives the [**net ionic reaction**](file:///C%3A%5CUsers%5CCentralScienceLive%5CChapter04%5CCH04_4.2_Main.html##).

The net ionic equation for the combination of aqueous sodium sulfate and aqueous barium hydroxide is the same as the net ionic equation for any combination of a soluble sulfate and a soluble barium compound.

**How to Write a Net Ionic Equation**

1. Write a balanced molecular equation.
2. Rewrite the equation to show the ions that form in solution when each soluble strong electrolyte dissociates or ionizes into its component ions. Only *dissolved strong electrolytes* are written in ionic form.
3. Identify and cancel spectator ions that occur on both sides of the equation.

There are many different types of reactions in chemistry. In this class we will characterize reactions into five simple topics. In order to figure out the product of a reaction, you must first learn to identify which **type** of reaction you are looking at, even if you are only given the reactants. In other words, you will learn to figure out the products of a reaction by learning the correct type of reaction.

There are also many guidelines for a reaction. These are special rules that you will learn to help you figure out reactivity patterns. It is important to realize at this point that you are not required to know every single reaction that has ever existed. Simply learn the basics and as you practice doing reactions; learn to develop an instinct on what type of reaction you are looking at, and what products will form from that reaction.

Try one on your own:

Write the net Ionic equation for the reaction of

Fe(NO3)3*(aq)* + 3NaOH*(aq)* → 3NaNO3*(aq)* + Fe(OH)3*(s)*

Fe+3 + 3NO3-+ 3Na+ + 3OH- → 3Na+ + 3NO3- + Fe(OH)3*(s)*

Fe+3 + 3OH- → Fe(OH)3*(s)*

How can we determine whether a product will precipitate out?

Below is a solubility guideline chart. It shows which ions are soluble in water and which are not, and their exceptions.

**Solubility Rules for Common Ionic Compounds in Water**

**Soluble Compounds Exceptions**

Most common acids

Group 1 metals (Li+, Na+, K+, Rb+, Cs+) None

Nitrates (NO3- ) None

Chlorates (ClO3- ) None

Perchlorates (ClO4- ) None

Hydrogen carbonate (HCO3- ) None

Acetates (C2H3O2- ) (silver acetate only slightly soluble)

Ammonium (NH4+ ) (ammonium hydroxide breaks up)

Halides (F-, Cl-, Br-, I-) (Ag+, Pb2+, Hg22+)

Sulfates (SO42- ) (Ag+, Pb+2, Hg22+, Ca2+, Ba2+, Sr2+)

**Insoluble Compounds Exceptions**

Carbonates (CO32- ) Group 1 metals, ammonium, dilute acids

Oxides (O2-) Group 1 metals, ammonium, dilute acids

Phosphates (PO43- ) Group 1 metals, ammonium, dilute acids

Sulfides (S2-) Group 1 metals, Ca2+, Sr2+, Ba2+, ammonium

Sulfites (SO32-) Group 1 metals, ammonium, dilute acids

Hydroxides (OH- ) Group 1 metals, Ca2+, Sr2+, Ba2+, dilute acids

Chromates (CrO42- ) Group 1 metals, Ca2+, Mg2+, dilute acids

Solid precipitates and molecular compounds do not dissociate or ionize. Therefore, they are considered **nonelectrolytes**. Weak acids and bases only break up in small percentages and are therefore considered **weak electrolytes**.

**Metathesis Reactions (double replacement)**

When you see two binary ionic compounds (including acids), the compounds switch partners to form two new compounds. The driving force and product is either a gas, a precipitate, or a weak electrolyte.

|  |  |
| --- | --- |
| a gas | **Bubbles will be produced in this reaction** |
| a precipitate | **Solids will be produced in this reaction** |
| a weak electrolytes | **Conductivity will decrease in this reaction** |

**Molarity**

**Molarity (M)** is defined as the moles of a solute per liters of a solution. Molarity is also known as the molar concentration of a solution.

Molarity = moles of solute / Liters of solution

**M = moles Liters**

**Problem #1:** Sea water contains roughly 28.0 g of NaCl per liter. What is the molarity of sodium chloride in sea water?

**Solution:**

28.0 g of NaCl x 1mol / 58.443 g = 0.4790993 moles

M = 0.4790993 moles / 1 Liter = **0.479 M**

**Problem #2:** How many grams of Na2CO3 are there in 10.0 L of 2.00 M solution?

**Solution:**

M = moles of solute / liters of solution

2.00 M = x / 10.0 L x = 20.0 mol

20.0 mol of Na2CO3 x 105.965 g / mol = 2119.3 grams = **2120 grams**

Often, a chemist will need to change the concentration of a solution by changing the amount of solvent. **Dilution** is the addition of solvent, which decreases the concentration of the solute in the solution. To determine how much volume is needed to dilute a known concentration of a solution, we can use an equation called the **Dilution formula:**

***M*1*V*1 = *M*2*V*2**

The volumes must be expressed in the same units.

**Problem #3**

If 25.0 mL of a 2.19 M solution are diluted to 72.8 mL, what is the final Molarity?

**Solution**

(2.19 M)(25.0 mL) = *M*2(72.8 mL) ***M*2 = 0.752 M**

**Problem #4**

If a chemist wants to dilute a 1.00 M KCl solution in order to make100. mL of a 0.0500 M solution of KCl, what volume does he need to take out of the original solution?

(1.00 M) V1 = (0.0500 M)(100. mL) **V1 = 5.00 mL**

Practically. The chemist would use a 5.00 mL pipette to draw out the 5 mL of KCl solution and deposit it into a 100-mL volumetric flask. Them he/she would dilute the solution up to the meniscus with distilled water until it reached 100 mL.

**Molarity of ions**

A solution is prepared by dissolving 9.82 grams of copper (II) chloride (CuCl2) in enough water to make 600. milliliters of solution. What is the molarity of the Cl- ions in the solution?

### Solution

[Molar mass](https://www.thoughtco.com/atomic-mass-and-atomic-mass-number-606079) of CuCl2 = 134.45 g/mol

[Number of moles](https://www.thoughtco.com/molecules-and-moles-603801) of CuCl2 = 9.82 g x 1 mol/134.45 g = 0.0730383 mol CuCl2

M = 0.0730383 mol/(600. mL x 1 L/1000 mL) = 0.122 mol/L

CuCl2 dissociates by the reaction

CuCl2 → Cu2+ + 2Cl -

Note that for every 1 mole of CuCl2 there are two moles of Cl- produced.

**Therefore, the** M of Cl- = **0.244 M**

**Arrhenius Acid** – A species that gives off H+ ions in solution.

HCl(aq) H+ + Cl-

H2SO4(aq) 2H+ + SO4-2

**Arrhenius Base** – A species that gives off OH- ions in solution.

NaOH(aq) Na+  + OH-

Ca(OH)2(aq) Ca+2  + 2OH-

**Neutralization** – When equal amounts of **H+** **ions** and **OH-** **ions** are added together to completely react an acid with a base. In other words, you “neutralize” an acid with a base or vice versa.

Example: HCl + NaOH NaCl + H2O

Total Ionic: H+ + Cl- + Na+ + OH-  Na+ + Cl- + H2O

Net Ionic: H+ + OH-  H2O

**Titration of Acids with Bases**

In a **titration**, a solution of known concentration (the titrant) is added to a solution of the substance being studied (the analyte). In an **acid**-**base titration**, the titrant is a strong **base** or a strong **acid**, and the analyte is an **acid** or a **base**, respectively.

**The Titration Equation**: **MAVA = MBVB**

Where **MA**  = the molarity of the H+ ions

 **VA** = the volume of the acid

 **MB**  = the molarity of the OH- ions

 **VB** = the volume of the base

**Example:** If 25.0 mL of 0.100 M KOH is needed to neutralize 100. mL of HNO3, calculate the concentration of the acid?

 **MA**  = ???M HNO3 **VA** = 100. mL of HNO3,

 **MB**  = 0.100 M KOH **VB** = 25.0 mL of KOH

 **MAVA = MBVB**

 (X)(100. mL) = (0.100 M)(25.0 mL)

 **X= 0.0250 M HNO3**

But be careful!!! If your acid has more than one **H+** or your base has more than one **OH-**, then you need to multiply by the number of those ions.

Lets try the same example again with one major change!

**Example:** If 25.0 mL of 0.100 M KOH is needed to neutralize 100. mL of **H2SO4**, calculate the concentration of the acid?

 **MA**  = ???M **H2SO4** **VA** = 100. mL of HNO3,

 **MB**  = 0.100 M KOH **VB** = 25.0 mL of KOH

 **MAVA = MBVB**

 **2**(X)(100. mL) = (0.100 M)(25.0 mL)

 **X= 0.0125 M H2SO4**

The reason for the **2** is because **H2SO4** has **2** **H+** ions not one.

So, a good thing to do is to alter the formula like this:

 (# of H+ ions) x **MAVA = MBVB** x (# of OH- ions)

**Now you try one:** How many milliliters of a 0.200 M solution of H3PO4 is needed to neutralize 50.0 mL of 0.150 M Ca(OH)2?

 (# of H+ ions) x **MAVA = MBVB** x (# of OH- ions)

 (3 H+ ions) x (**0.200** **M)( VA ) =**  (**0.150 M**)(**50.0 mL**)x (2 OH- ions)

**VA = 25.0 mL**