Chapter 4 -- Aqueous Reactions and Solution Stoichiometry

## Key Vocabulary

**aqueous solutions** - solutions in which water is the dissolving medium
**solvent** - the component in a solution that is of the greatest quantity.
**solute(s)** - the component(s) dissolved in the solvent.
**electrolyte** - a substance whose aqueous solutions contain ions and therefore conduct electricity.
**non-electrolyte** - a substance that does not form ions in solution.
**strong electrolytes** - compounds which exist in solution completely or nearly completely as ions. Nearly all ionic compounds and a few covalent compounds are strong electrolytes.
**weak electrolytes** - molecular compounds which produce a small concentration of ions when they dissolve.
**precipitation reactions** - reactions which result in the formation of an insoluble **precipitate**.
**solubility** - the amount of a substance which can be dissolved in a given quantity of solvent.
**exchange/metathesis reactions** - reactions in which positive ions and negative ions appear to exchange partners.
**molecular equation** - an equation which shows the complete chemical formulas of all reactants and products.
**complete ionic equation** - an equation in which all soluble strong electrolytes are written as ions.
**spectator ions** - ions which appear in identical forms among both reactants and products of a complete ionic equation.
**net ionic equation** - equations in which spectator ions are omitted.
**acids** - substances that ionize in aqueous solutions to form a hydrogen ion, increasing the H+ concentration. They are also called proton donors.

- monoprotic - acids which can ionize to form one H+ ion.
- diprotic - acids which can ionize to form two H+ ions. Ionization occurs in two steps, and only the first ionization is complete.

**bases** - substances that readily accept H+ ions. They also increase the concentration of OH- ions in solution.
**strong acids and bases** - acids and bases that are strong electrolytes.
**weak acids and bases** - acids and bases that are weak electrolytes.
**neutralization reaction** - the reaction that occurs when an acidic and a basic solution are mixed.
**salt** - an ionic compound whose cation comes from a base.
**oxidation** - loss of electrons, or an increased oxidation number.
**reduction** - gain of electrons, or a reduced oxidation number.
**oxidation numbers** - the actual charge of a monatomic ion. In other ions, the hypothetical charge assigned to the atom.
**displacement reactions** - reactions in which an ion in solution is replaced through the oxidation of an element.
**activity series** - a list of metals arranged in order of decreasing ease of oxidation.
**concentration** - the amount of solute dissolved in a given quantity of solvent or solution.
**dilution** - a process by which solutions of lower concentration can be obtained by adding water.
**standard solution** - a solution of known concentration
**titration** - a process used to determine an unknown concentration.
**equivalence point** - the point at which stoichiometrically equivalent quantities are brought together.

## Tips

- Don't confuse the extent to which electrolytes dissolve with whether they are strong or weak. Double arrows indicate weak electrolytes ionizing. Single arrow indicates strong electrolytes ionizing.
- If solubility < $1.0*10^{-2}$ mol/L, the substance is considered insoluble.
- Concentration may change when an ionic compound dissolves, depending on the formula.

## Stuff to Memorize

### Solubility Rules

**Soluble Compounds are those containing**:
NO3-
C2H3O2-
Cl- (Exceptions: compounds of Ag+, Hg2(2+), and Pb2+)
Br- (Exceptions: compounds of Ag+, Hg2(2+), and Pb2+)
I- (Exceptions: compounds of Ag+, Hg2(2+), and Pb2+)
SO4(2-) (Exceptions: compounds of Sr2+, Ba2+, Hg2(2+), and Pb2+)

**Insoluble Compounds are those containing:**
S2- (Exceptions: compounds of NH4+, alkali metal cations, Ca2+, Sr2+, and Ba2+)
CO3(2-) (Exceptions: Compounds of NH4+ and the alkali metal cations)
PO4(3-) (Exceptions: Compounds of NH4+ and the alkali metal cations)
OH- (Exceptions: compounds of the alkali metal cations, Ca2+, Sr2+, and Ba2+)

| Metal | Oxidation Reaction | | | | |
|---|---|---|---|---|---|
| Lithium | Li | $\rightleftharpoons$ | $Li^+$ | + | $e^-$ |
| Rubidium | Rb | $\rightleftharpoons$ | $Rb^+$ | + | $e^-$ |
| Potassium | K | $\rightleftharpoons$ | $K^+$ | + | $e^-$ |
| Barium | Ba | $\rightleftharpoons$ | $Ba^{2+}$ | + | $2e^-$ |
| Calcium | Ca | $\rightleftharpoons$ | $Ca^{2+}$ | + | $2e^-$ |
| Sodium | Na | $\rightleftharpoons$ | $Na^+$ | + | $e^-$ |
| Magnesium | Mg | $\rightleftharpoons$ | $Mg^{2+}$ | + | $2e^-$ |
| Aluminum | Al | $\rightleftharpoons$ | $Al^{3+}$ | + | $3e^-$ |
| Manganese | Mn | $\rightleftharpoons$ | $Mn^{2+}$ | + | $2e^-$ |
| Zinc | Zn | $\rightleftharpoons$ | $Zn^{2+}$ | + | $2e^-$ |
| Chromium | Cr | $\rightleftharpoons$ | $Cr^{3+}$ | + | $3e^-$ |
| Iron | Fe | $\rightleftharpoons$ | $Fe^{2+}$ | + | $2e^-$ |
| Cobalt | Co | $\rightleftharpoons$ | $Co^{2+}$ | + | $2e^-$ |
| Nickel | Ni | $\rightleftharpoons$ | $Ni^{2+}$ | + | $2e^-$ |
| Tin | Sn | $\rightleftharpoons$ | $Sn^{2+}$ | + | $2e^-$ |
| Lead | Pb | $\rightleftharpoons$ | $Pb^{2+}$ | + | $2e^-$ |
| Hydrogen | $H_2$ | $\rightleftharpoons$ | $2\,H^+$ | + | $2e^-$ |
| Copper | Cu | $\rightleftharpoons$ | $Cu^{2+}$ | + | $2e^-$ |
| Silver | Ag | $\rightleftharpoons$ | $Ag^+$ | + | $e^-$ |
| Mercury | Hg | $\rightleftharpoons$ | $Hg^{2+}$ | + | $2e^-$ |
| Platinum | Pt | $\rightleftharpoons$ | $Pt^{2+}$ | + | $2e^-$ |
| Gold | Au | $\rightleftharpoons$ | $Au^{3+}$ | + | $3e^-$ |

## Activity Series

Metals at the top of the table are most easily oxidized.

### Rules for Determining Oxidation Numbers

1. For an atom in its elemental form, the oxidation number is always zero.
2. For any monatomic ion, the oxidation number equals the charge on the ion.
3. Nonmetals usually have negative oxidation numbers, although they can sometimes be positive.
   a. The oxidation number of oxygen is usually -2.
   b. The oxidation of hydrogen is usually +1 when bonded to nonmetals and -1 when bonded to metals.
   c. The oxidation number of fluorine is always -1. The other halogens have an oxidation number of -1 in most binary compounds. They may have positive oxidation states when combined with oxygen.
4. The sum of the oxidation numbers of all atoms in a neutral compound is zero. The sum of the oxidation numbers in a polyatomic ion equals the charge of the ion.

### Writing Net Ionic Equations

1. Write a balanced molecular equation.
2. Rewrite the equation, dissociating soluble strong electrolytes into their component ions.
3. Cancel spectator ions.

## Reactions and Equations

### Double Displacement

Of the form AX + BY → AY + BX

### Single Displacement

A + BX → AX + B

### Neutralization

acid + base → water + salt

### Equation for Concentration

molarity M = moles solute / volume of solution (L)

## Example Problems

### Example 1

Write the net ionic equation for the precipitation reaction that occurs when solutions of calcium chloride and sodium carbonate are mixed.

Step 1. *Write a balanced molecular equation*. CaCl2 (aq) + Na2CO3 (aq) → CaCo3 (s) + 2NaCl (aq)

Step 2. *Rewrite the equation, dissociating soluble strong electrolytes into their component ions.*

Ca2+ (aq) + 2Cl- (aq) + 2Na+ (aq) + CO3(2-) (aq) → CaCO3 (s) + 2Na+ (aq) + 2Cl- (aq)

Step 3. *Cancel the spectator ions.* Ca2+ (aq) + CO3(2-) (aq) → CaCO3 (s)

### Example 2

How many grams of Na2SO4 are required to make 0.350 L of 0.500 M Na2SO4?

M Na2SO4 = (moles NA2SO4)/liters solution

Moles Na2SO4 = liters solution * M Na2SO4

= (0.350 L solution) * ((0.500 mol Na2SO4)/(1 L solution)) = 0.175 mol Na2SO4

(0.175 mol Na2SO4) * ((142 g Na2SO4)/(1 mol Na2SO4)) = 24.9 g Na2SO4

### Example 3

How many moles of H2O form when 25.0 mL of 0.100 M HNO3 solution is completely neutralized by NaOH?

Moles HNO3 = M * L = (0.100 (mol HNO3/L)) * (0.0250 L) = 2.50 * 10^-3 mol H2O

The equation for this reaction is:

HNO3 (aq) + NaOH (aq) → H2O (l) + NaNO3 (aq)

So for every mole of HNO3, there is 1 mole of H2O.

Moles H2O = (2.50 * 10^-3 mol H2O) * ((1 mol H2O)/(1 mol HNO3)) = 2.50 * 10^-3 mol H2O