AP Chemistry
Chapter 5: Reactions in Aqueous Solution – Summary Information

3 Major types of reactions in aqueous solutions:

-Acid-Base Reactions: Acid + Base → Salt + Water
-Precipitation Reactions (aka double displacement or exchange reaction):
$AB_{(aq)} + CD_{(aq)} \rightarrow AD_{(s)} + CB_{(aq)}$
-Oxidation-Reduction (not limited to aqueous reactions): Including (but not limited to)
Single Replacement: AB + C → CB + A
Decomposition: AB → A + B
Synthesis: A + B → AB
Combustion: Hydrocarbon + Oxygen → Carbon dioxide + Water

Driving Forces for Product Favored Reactions:

Reaction Type:	Driving Force
Precipitation	Formation of an insoluble compound
Acid-Base Neutralization	Formation of a salt and water (proton transfer)
Gas-Forming	Evolution of water-insoluble gas such as CO_2
Oxidation-Reduction	Electron transfer

Gas-forming Reactions:

Metal carbonate or bicarbonate + acid → metal salt + $CO_{2(g)}$ + $H_2O_{(l)}$
(ex. $Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$)

Metal sulfide + acid → metal salt + $H_2S_{(g)}$
(ex. $Na_2S_{(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2S_{(g)}$)

Metal sulfite + acid → metal salt + $SO_{2(g)}$ + $H_2O_{(l)}$
(ex. $Na_2SO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + SO_{2(g)} + H_2O_{(l)}$)

Ammonium salt + strong base → metal salt + $NH_{3(g)}$ + $H_2O_{(l)}$
(ex. $NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NH_{3(g)} + H_2O_{(l)}$)

Concentration

Molarity = moles solute / Liters of solution   M = mol/L
Molality = moles solute / Kilograms of solvent   m = mol/kg
Normality = mole "equivalents" of a substance per liter of solution   N = mol equiv/L
(Ex. 1M H_2SO_4 is 2N (2 "normal") because each mole of H_2SO_4 can provide two moles of H^+ ions)
For dilute aqueous solutions in which the density is approximately 1.00g/L
1ppm (part per million by mass) = 1mg/L
1ppb (part per billion) = 1ug/L
1ppt (part per trillion) = 1ng/L

Common Dilutions Equation:
$M_1V_1 = M_2V_2$

pH, a Concentration Scale for Acids and Bases

pH (power of hydrogen) = $-\log_{10}[H+]$ hydrogen ion molarity
$[H+] = 10^{-pH}$

Oxidation-Reduction Information:

Common Oxidizing and Reducing Agents			
Oxidizing Agent	**Reaction Product**	**Reducing Agent**	**Reaction Product**
O_2, oxygen	O^{2-}, oxide ion or O combined in H_2O	H_2, hydrogen	$H^+_{(aq)}$, hydrogen ion or H combined in H_2O or other molecule
Halogen, F_2, Cl_2, Br_2 or I_2	Halide ion, F^-, Cl^-, Br^- or I^-	M, metals such as Na, K, Fe and Al	M^{n+}, metal ions such as Na^+, K^+, Fe^{2+} or Fe^{3+} and Al^{3+}
HNO_3, Nitric acid	Nitrogen oxides * such as NO and NO_2	C, carbon (used to reduce metal oxides)	CO and CO_2
$Cr_2O_7^{2-}$, dichromate ion	Cr^{3+}, chromium(III) ion (in acid solution)		
MnO_4^-, permanganate ion	Mn^{2+}, manganese(II) ion (in acid solution)		

*NO is produced with dilute HNO_3, whereas NO_2 is a product of a concentrated acid.

Recognizing Oxidation-Reduction Reactions		
	Oxidation	**Reduction**
In terms of oxidation number	Increase in oxidation number of an atom	Decrease in oxidation number of an atom
In terms of electrons	Loss of electrons by an atom	Gain of electrons by an atom
In terms of oxygen	Gain of one or more O atoms	Loss of One or more O atoms
In terms of hydrogen	Loss of one or more H atoms	Gain of one or more H atoms

Acids and Bases:

Acidic oxides (nonmetallic oxides aka acidic anhydrides) produce acids in solution

Ex. $SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

Basic oxides (metallic oxides aka basic anhydrides) produce bases in solution

Ex. $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$

Common Acids and Bases			
Strong Acids (Strong Electrolytes)		**Strong Bases (Strong Electrolytes)**	
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO_3	Nitric acid		
$HClO_4$	Perchloric acid		
H_2SO_4	Sulfuric acid		
Weak Acids (Weak Electrolytes)		Weak Bases (Weak Electrolytes)	
H_3PO_4	Phosphoric acid	NH_3	Ammonia
H_2CO_3	Carbonic acid		
CH_3CO_2H	Acetic acid		
$H_2C_2O_4$	Oxalic acid		
$H_2C_4H_4O_6$	Tartaric acid		
$H_3C_6H_5O_7$	Citric acid		
$HC_9H_8O_4$	Aspirin (acetylsalicylic acid)		

Redox Equation Balancing

In acid solution:

1. Write two half-reactions, one representing the oxidation, and the other the reduction, that occur in the reaction. It is not necessary to know which is which at this point.

2. In each half-reaction, balance all atoms except hydrogen and oxygen.

3. Balance oxygen atoms in each half-reaction by adding one H_2O molecule for each oxygen atom needed. Never use O_2, OH^-, or any other form of oxygen.

4. Balance the hydrogen atoms by adding hydrogen ions, H^+. Never use H_2, OH^-, or any other form of hydrogen.

5. Balance the charges by adding the proper number of electrons (e^-). If steps 1-4 have been done properly, electrons will be added to the left side of one half-reaction and to the right side of the other.

6. Multiply each half-reaction by the appropriate number so that the two half-reactions have the same number of electrons. Add the half-reactions, and cancel the electrons **(they must cancel)**. Also cancel all common ions and molecules. Simplify the coefficients of the equation if possible.

In basic solution:

7. In addition to steps 1-6, add one OH^- ion for each H^+ ion to both sides of the equation in step 6. Combine the H^+ and the OH^- ions on one side of the reaction into H_2O molecules. Cancel H_2O molecules that appear on both sides of the equation, and simplify if possible.