

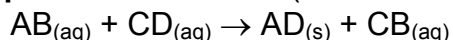
AP Chemistry

Chapter 5: Reactions in Aqueous Solution – Summary Information

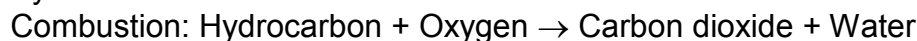
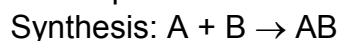
3 Major types of reactions in aqueous solutions:

-**Acid-Base Reactions:** Acid + Base \rightarrow Salt + Water

-**Precipitation Reactions** (aka double displacement or exchange reaction):



-**Oxidation-Reduction** (not limited to aqueous reactions): Including (but not limited to)

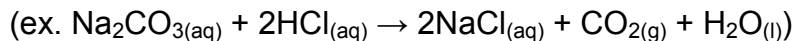


Driving Forces for Product Favored Reactions:

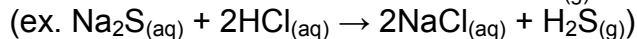
<u>Reaction Type:</u>	<u>Driving Force</u>
Precipitation	Formation of an insoluble compound
Acid-Base Neutralization	Formation of a salt and water (proton transfer)
Gas-Forming	Evolution of water-insoluble gas such as CO ₂
Oxidation-Reduction	Electron transfer

Gas-forming Reactions:

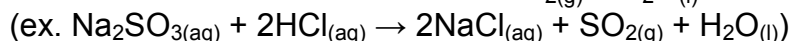
Metal carbonate or bicarbonate + acid \rightarrow metal salt + CO_{2(g)} + H₂O_(l)



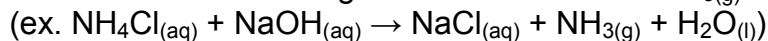
Metal sulfide + acid \rightarrow metal salt + H₂S_(g)



Metal sulfite + acid \rightarrow metal salt + SO_{2(g)} + H₂O_(l)



Ammonium salt + strong base \rightarrow metal salt + NH_{3(g)} + H₂O_(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 H₂SO₄ is 2N (2 “normal”) because each mole of H₂SO₄ 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) = $-\log_{10}[\text{H}^+]$ hydrogen ion molarity
 $[\text{H}^+] = 10^{-\text{pH}}$

Oxidation-Reduction Information:

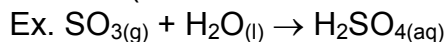
Common Oxidizing and Reducing Agents			
Oxidizing Agent	Reaction Product	Reducing Agent	Reaction Product
O ₂ , oxygen	O ²⁻ , oxide ion or O combined in H ₂ O	H ₂ , hydrogen	H ⁺ _(aq) , hydrogen ion or H combined in H ₂ O or other molecule
Halogen, F ₂ , Cl ₂ , Br ₂ or I ₂	Halide ion, F ⁻ , Cl ⁻ , Br ⁻ 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 ₂ O 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 HNO₃, whereas NO₂ is a product of a concentrated acid.

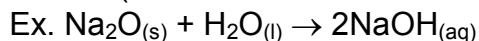
Recognizing Oxidation-Reduction Reactions		
	Oxidation	Reduction
In terms of oxidation number	Increase in oxidation number of an atom	Decrease in oxidation number of an atom
In terms of electrons	Loss of electrons by an atom	Gain of electrons by an atom
In terms of oxygen	Gain of one or more O atoms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



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



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 ₂ SO ₄	Sulfuric acid		
Weak Acids (Weak Electrolytes)		Weak Bases (Weak Electrolytes)	
H ₃ PO ₄	Phosphoric acid	NH ₃	Ammonia
H ₂ CO ₃	Carbonic acid		
CH ₃ CO ₂ H	Acetic acid		
H ₂ C ₂ O ₄	Oxalic acid		
H ₂ C ₄ H ₄ O ₆	Tartaric acid		
H ₃ C ₆ H ₅ O ₇	Citric acid		
HC ₉ H ₈ O ₄	Aspirin (acetylsalicylic acid)		

Redox Equation Balancing

In acid solution:

- 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.
- In each half-reaction, balance all atoms except hydrogen and oxygen.
- Balance oxygen atoms in each half-reaction by adding one H₂O molecule for each oxygen atom needed. Never use O₂, OH⁻, or any other form of oxygen.
- Balance the hydrogen atoms by adding hydrogen ions, H⁺. Never use H₂, OH⁻, or any other form of hydrogen.
- 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.
- 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:

- 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 H₂O molecules. Cancel H₂O molecules that appear on both sides of the equation, and simplify if possible.