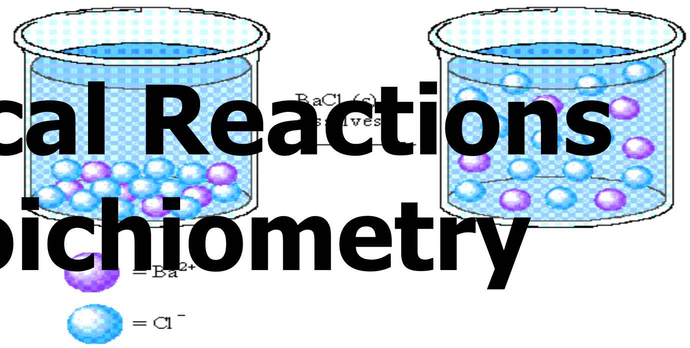
Honors Chemistry 2
Unit 3

Types of Chemical Reactions & Solution Stoichiometry



Students should be able to:

- ✓ Predict to some extent whether a substance will be a strong electrolyte, weak electrolyte, or nonelectrolyte.
- ✓ Predict the ions that an electrolyte dissociates into.
- ✓ Identify substances as acids, bases, and salts.
- ✓ Using solubility rules, predict if a precipitate forms in a metathesis reaction. Next, predict its products and write a balanced equation.
- ✓ Predict the products and write a balanced chemical equation for neutralization reactions.
- ✓ After constructing molecular reactions for metathesis reactions, be able to identify spectator ions and write the net ionic equations.
- ✓ Assign oxidation numbers to atoms.
- ✓ Determine whether a reaction is Redox (single replacement) or not.
- ✓ Use the activity series to predict whether a Redox reaction will occur and be able to write the molecular and net ionic equations if it does.
- ✓ Calculate moles of solute, volume of solution, or Molarity of the solution from the other two.
- ✓ Recognize and work dilution problems.
- ✓ Calculate the volume of a certain molarity solution required to react with another solution of known molarity.
- ✓ Calculate the mass of a substance that would be required to react with a given volume of a solution of known molarity.
- ✓ Calculate mass of solute or concentration of an unknown solution from titration data.

Keywords:

- concentration
- titration
- indicators
- solvent
- strong electrolyte
- precipitate
- molecular equation
- spectator ions
- salts
- reduction
- molarity
- standard solution
- aqueous
- electrolyte
- weak electrolyte
- solubility
- (complete) ionic equation
- acids
- neutralization
- redox reaction
- dilution
- equivalence point
- solute
- nonelectrolyte
- activity series
- metathesis
- net ionic equation
- bases
- oxidation
- oxidation number

I. Aqueous Solutions

A. What is an aqueous solution?

B. What is the solvent?

C. What is the solute?

D. What does concentration mean?

E. How do you measure molarity (M)?

F. Sample Exercise 14.1 – Calculate the molarity of a solution made by dissolving 23.4 g of sodium sulfate, Na_2SO_4, in enough water to form 125 mL.

G. Sample Exercise 14.2 – How many grams of Na_2SO_4 are required to make 0.350 L of 0.500 M Na_2SO_4?

H. How do you make a dilution?

1. What is the formula that you can use?

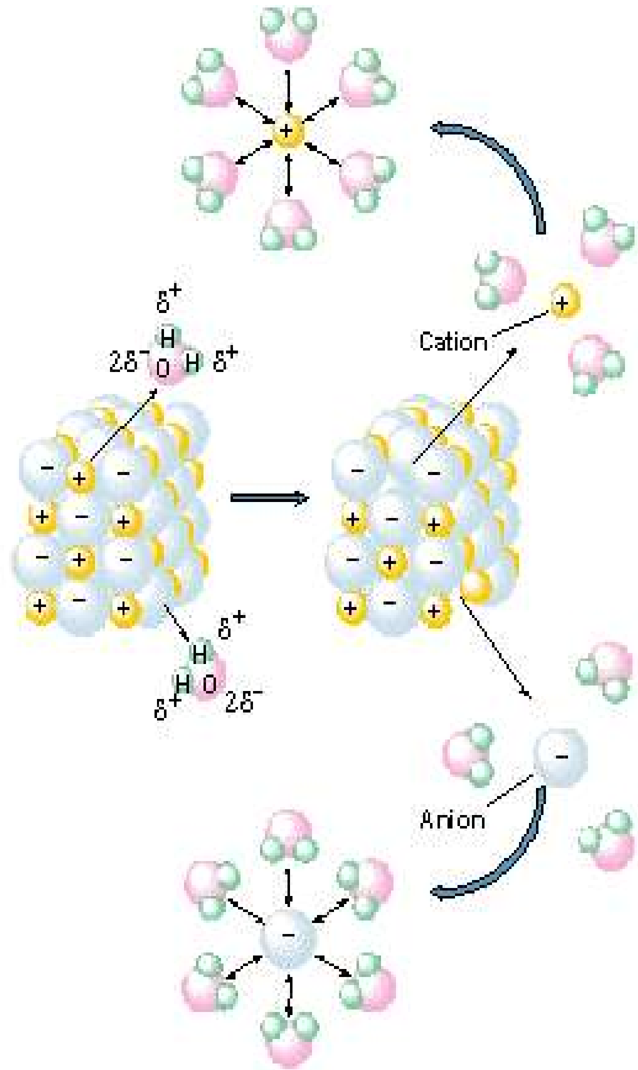
I. Sample Exercise 14.3 – How much 3.0 M H_2SO_4 would be required to make 500 mL of 0.10 M H_2SO_4?

II. Electrolytes

A. Electrolyte vs. Nonelectrolyte

B. Strong vs. Weak Electrolytes

1. What happens when an ionic substance dissolves?



C. Sample Exercise 4.4 – What are the molar concentrations of all ions present in a 0.025 M aqueous solution of calcium nitrate?

III. Acids, Bases, and Salts

A. What is an Acid?

1. What is the difference between a Monoprotic Acid and a Diprotic Acid?

2. What is the difference between a strong acid and a weak acid?

3. What are the strong acids?

B. What is a <u>Base</u>?

1. What types of compounds make strong bases?

C. What are <u>Salts</u>?

D. What is a <u>neutralization reaction</u>?

E. Sample Exercise 4.6 – Write a balanced equation for the reaction of hydrobromic acid with barium hydroxide in aqueous solution.

IV. Ionic Equations

A. Spectator Ions –

B. Net Ionic Equations –

C. Sample Exercise 4.7 – Write the net ionic equation for the neutralization of two of the acidic hydrogens of phosphoric acid by sodium hydroxide in aqueous solution.

V. Metathesis Reactions

A. What is a Metathesis Reaction?

B. What are the *driving forces* for a metathesis reaction?

C. Precipitation Reactions

1. Precipitate –

2. Solubility –

D. Solubility Rules

SOLUBLE SALTS	INSOLUBLE SALTS
Group I compounds and ammonium compounds	Hydroxides (**EXCEPT** Group I and ammonium, hydroxides of Ca^{2+}, Sr^{2+} and Ba^{2+} are slightly soluble)
Nitrates, hydrogen carbonates and chlorates	
Chlorides, bromides and iodides (**EXCEPT** those of Pb^{2+}, Ag^{+} and Hg_2^{2+})	Carbonates, phosphates, chromates and sulfides (**EXCEPT** group I and ammonium salts, sulfides of group II are soluble)
Sulfates (**EXCEPT** Ag^{+}, Sr^{2+}, Ba^{2+}, Pb^{2+} and Ca^{2+})	

E. Sample Exercise 4.8 – Write balanced molecular, ionic, and net ionic equations for the precipitation reactions (if any) that occur when solutions of the following compounds are mixed: (a) $BaCl_2$ and Na_2SO_4 (b) KCl and Na_2SO_4.

F. Reactions in which a weak electrolyte of nonelectrolyte forms:

G. Reactions in which a Gas forms:

H. Sample Exercise 4.9 – Write balanced complete ionic and net ionic equations for any reactions that occur when the following compounds are mixed: (a) $Cr(C_2H_3O_2)_2$ (aq) and HNO_3 (aq) (b) $FeCO_3$ (s) and HCl (aq) (c) PbS (s) and H_2SO_4 (aq).

VI. Reactions of Metals

A. Oxidation and Reduction

B. Oxidation of Metals by Acids and Salts

C. Sample Exercise 4.10 – Write the balanced molecular and net ionic equations for the reaction of aluminum with hydrobromic acid.

D. The Activity Series

Metal	*React with Acid?*	*React with Steam*	*React with Cold Water?*
Li	***YES***	***YES***	***YES***
K	***YES***	***YES***	***YES***
Ca	***YES***	***YES***	***YES***
Na	***YES***	***YES***	***YES***
Mg	***YES***	***YES***	*NO*
Al	***YES***	***YES***	*NO*
Zn	***YES***	***YES***	*NO*
Fe	***YES***	***YES***	*NO*
Sn	***YES***	*NO*	*NO*
Pb	***YES***	*NO*	*NO*
H	**-**	*NO*	*NO*
Cu	*NO*	*NO*	*NO*
Ag	*NO*	*NO*	*NO*
Pt	*NO*	*NO*	*NO*
Au	*NO*	*NO*	*NO*

E. Sample Exercise 4.11 – Will an aqueous solution of iron (II) chloride oxidize magnesium metal? If so, write the balanced molecular and net ionic equations for the reaction.

VII. Solution Stoichiometry

A. How to Solve Solution Stoichiometry Problems:

B. Sample Exercise 4.12 – How many moles of H_2O form when 25.0 mL of 0.100 M HNO_3 solution is completely neutralized by NaOH?

C. <u>Titrations</u> –

1. <u>Standard Solutions</u> –

2. <u>Equivalence Point</u> –

3. <u>Indicators</u> –

D. Sample Exercise 4.13 – One method used commercially to peel potatoes is to soak them in a solution of NaOH for a short time, remove them from the NaOH, and spray off the peel. The concentration of NaOH is normally in the range 3 to 6 M. The NaOH is analyzed periodically. In one such analysis, 45.7 mL of 0.500 M H_2SO_4 is required to react completely with a 20.0 mL sample of NaOH solution:

H_2SO_4 (aq) + 2 NaOH (aq) → 2 H_2O (l) + Na_2SO_4 (aq)

What is the concentration of the NaOH solution?

E. Sample Exercise 4.13 – The quantity of Cl^- in a water supply is determined by titrating the sample with Ag^+ :

Ag^+ (aq) + Cl^- (aq) → AgCl (s)

What mass of chloride ion is present in 10.0 g sample of the water if 20.2 mL of 0.100 M Ag^+ is required to react with all the chloride in the sample?

VIII. Rules for assigning OXIDATION STATES (numbers):

A. UNCOMBINED ELEMENTS (ELEMENTS NOT BONDED TO ANY OTHER TYPE OF ELEMENT) have an oxidation number of ZERO. This includes any formula that has *only* one chemical symbol in it (single elements & diatomic elements).

Examples:

B. In COMPOUNDS (remember, they are neutral and have 2+ different elements bonded together), the sum of the CHARGES must ADD UP TO ZERO so the ions within a compound have oxidation numbers equal to their OXIDATION # FOUND ON PERIODIC TABLE/INDIVIDUAL CHARGES.

Ex: $NaCl$ Ex: Mg_3N_2

Ex: HNO_3

* The OXIDATION NUMBER is the number INSIDE the PARENTHESES. It is the charge on JUST ONE atom of that element!

** Remember that we almost always write the + ION FIRST and the - ION LAST in a compound formula.

EXAMPLE:

EXCEPTION to this rule:

C. GROUP 1 METALS always have an oxidation number of +1 when in a compound (bonded to another species). Likewise, combined GROUP 2 METALS always therefore have a +2 oxidation number when located within a compound.

Ex:

D. FLUORINE is always a -1 in compounds. The other HALOGENS (ex: Cl, Br) are also -1 as long as they are the most electronegative element in the compound.

Ex:

E. HYDROGEN is a +1 in compounds unless it is combined with a metal (and is at the back of the formula), then it is -1.

Ex:

F. OXYGEN is USUALLY -2 in compounds.

Ex:

When combined with fluorine (F), which is more electronegative, oxygen is +2.

Ex:

When in a PEROXIDE oxygen is -1. A peroxide is a compound that has a formula of X_2O_2.

Ex:

G. The sum of the oxidation numbers in polyatomic ions must equal the CHARGE ON THE ION (SEE TABLE E).

Ex: $Cr_2O_7^{2-}$