

Chapter 4 Notes - Types of Chemical Reactions and Solution Chemistry

4.1 Water, the Common Solvent

A. Structure of water

1. Oxygen's electronegativity is high (3.5) and hydrogen's is low (2.1)
2. Water is a bent molecule
3. Water is a polar molecule

B. Hydration of Ionic Solute Molecules

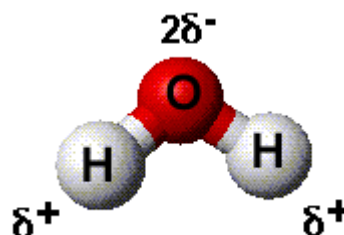
1. Positive ions attracted to the oxygen end of water
2. Negative ions attracted to the hydrogen end of water

C. Hydration of Polar Solute Molecules

1. Negative end of polar solute molecules are attracted to water's hydrogen
2. Positive end of polar solute molecules are attracted to water's oxygen

D. "Like Dissolves Like"

1. Polar and ionic compounds dissolve in polar solvents like water
2. Nonpolar compounds like fats dissolve in nonpolar solvents like _____?



4.2 The Nature of Aqueous Solutions: Strong and Weak Electrolytes

A. Definition of Electrolytes

1. A substance that when dissolved in water produces a solution that can conduct an electric current

B. Strong electrolytes conduct current very efficiently

1. Completely ionized when dissolved in water
 - a. Ionic compounds
 - b. Strong acids ($\text{HNO}_3(\text{aq})$, $\text{H}_2\text{SO}_4(\text{aq})$, $\text{HCl}(\text{aq})$)
 - c. Strong bases (KOH , NaOH)

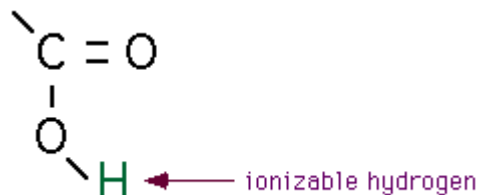
C. Weak electrolytes conduct only a small current

1. Slightly ionized in solution
 - a. Weak acids (organic acids - acetic, citric, butyric, malic)

- b. Weak bases (ammonia)

D. Nonelectrolytes conduct no current

1. No ions present in solution
 - a. alcohols, sugars



4.3 The Composition of Solutions

A. Molarity

1. Moles of solute per liter of solution
 $M = \text{molarity} = \text{moles of solute/liters of solution}$

B. Concentration of Ions in Solution

1. Ionic compounds dissociate in solution, multiplying the molarity by the number of ions present

C. Moles from Concentration

1. Liters of solution x molarity = moles of solute

D. Solutions of Known Concentration

1. Standard solution - a solution whose concentration is accurately known
2. Preparation of Standard solutions

How much x How strong x What does it weigh?

$$L \times \text{mol/L} \times \text{g/mol} = \text{grams required to prepare the standard}$$

E. Dilution

1. Dilution of a volume of solution with water does not change the number of moles present
2. Solving dilution problems

$$M_1V_1 = M_2V_2$$

4.4 Types of Chemical Reactions

A. Precipitation reactions

1. When two solutions are mixed, an insoluble solid forms

B. Acid-Base reactions

1. A soluble hydroxide and a soluble acid react to form water and a salt

C. Oxidation-Reduction reactions (redox rxns)

1. Reactions in which one or more electrons are transferred

4.5 Precipitation Reactions

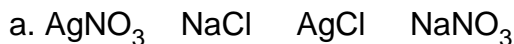
A. Dissociation

1. Ionic compounds dissolve in water and the ions separate and move independently



B. Determination of Products

1. Recombination of ions



2. Elimination of reactants as products

a. AgNO_3 and NaCl are reactants and can't be products

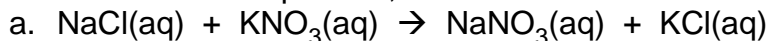
3. Identifying the precipitate

a. "Switch Partners" of reactant pairs to determine the names of the products.

b. AgCl and NaNO_3 are the products

c. AgCl is insoluble, so it is the white precipitate

d. If there is no insoluble product, the reaction does not occur



Both products are soluble and all ions remain independent in solution; no reaction occurs

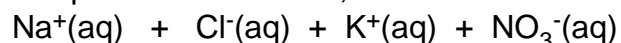


Table 4.1 Simple Rules for the Solubility of Salts in Water

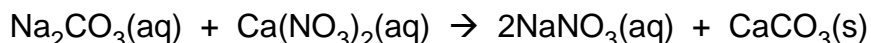
1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts containing the alkali metal ions (Li^+ , Na^+ , K^+ , Cs^+ , Rb^+) and the ammonium (NH_4^+) ion are soluble.
3. Most chloride, bromide and iodide salts are soluble. Exceptions are salts containing the ions Ag^+ , Pb^{2+} , and Hg^{2+} .
4. Most sulfate salts are soluble. Notable exceptions are BaSO_4 , PbSO_4 , HgSO_4 and CaSO_4 .
5. Most hydroxide salts are only slightly soluble. The important soluble hydroxides are NaOH and KOH . The hydroxides of barium, strontium and calcium are marginally soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}) and phosphate (PO_4^{3-}) salts are only slightly soluble.

AP Standard – Students must know that salts containing ammonium, sodium, potassium, and nitrate ions are ALWAYS soluble. Memorization of other solubility rules is outside the scope of the course and will not be tested.

4.6 Describing Reactions in Solution

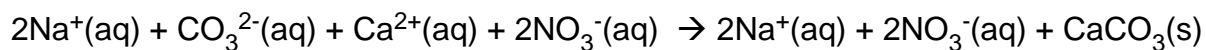
A. The Molecular Equation

1. Gives the overall reaction stoichiometry, not necessarily the actual forms of reactants and products in solution



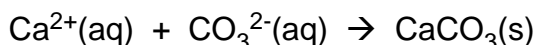
B. The Complete Ionic Equation

1. Represents as ions all reactants and products that are strong electrolytes



C. The Net Ionic Equation

1. Includes only those components that take part in the chemical change
2. Spectators are eliminated



4.7 Stoichiometry of Precipitation Reactions

- A. Determine what reaction takes place
- B. Write the balanced net ionic equation for the reaction
- C. Calculate the moles of reactants
- D. Determine which reactant is limiting
- E. Calculate the moles of product or products
- F. Convert to grams or other units, as required

4.8 Acid-Base Reactions (Neutralization Reactions)

A. Definitions

1. Brønsted: Acids are proton donors, bases are proton acceptors
2. Arrhenius: Acids produce H^+ ions in water, bases produce OH^- ions in solution

B. Net ionic equation for acid-base reactions

1. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
2. The hydroxide ion can be assumed to completely react with even a weak acid in solution

C. Stoichiometry Calculations for Acid-Base Reactions

1. List the species present in the combined solution before any reaction occurs; decide what reaction will occur
2. Write the balanced net ionic equation for this reaction
3. Calculate the moles or reactants
 - a. For reactions in solution, use volumes of the original solutions and their molarities
4. Determine the limiting reactant where appropriate
5. Calculate the moles of the required reactant or product
6. Convert to grams or volume of solution as required

D. Acid-Base Titrations

1. Vocabulary
 - a. Titrant - Solution of known concentration
 - b. Analyte - Solution of unknown concentration
 - c. Equivalence point - Point at which the amount of titrant added to analyte results in perfect neutralization
 - d. Indicator - a substance added at the beginning of the titration that changes color at the equivalence point
 - e. Endpoint - the point at which the indicator changes color
2. Requirements for a successful titration
 - a. the exact reaction between titrant and analyte must be known
 - b. the reaction must proceed rapidly
 - c. the equivalence point must be marked accurately (select the appropriate indicator)
 - d. the volume of titrant required to reach the equivalence point must be known accurately
 - e. for acid-base titrations, the titrant should be a strong acid or a strong base

4.9 Oxidation-Reduction Reactions (redox)

A. Electron transfer (LEO says GER)

1. Gain electrons = reduction
2. Lose electrons = oxidation

B. Examples of redox rxns

1. Photosynthesis
2. Combustion of fuels
3. Oxidation of sugars, fats, proteins for energy

Rules for Assigning Oxidation Numbers	Summary
1. the oxidation number of the atom of a free element is zero	Element = 0
2. the oxidation number of a monatomic ion equals its charge	
3. In compounds, oxygen has an oxidation number of -2, except in peroxides, where it is -1	Oxygen = -2
4. In compounds containing hydrogen, hydrogen has an oxidation number of +1	Hydrogen = +1
5. In compounds, fluorine is ALWAYS assigned an oxidation number of -1	Fluorine = -1
6. The sum of the oxidation states for an electrically neutral compound must be zero	

C. Noninteger Oxidation states

1. Fe₃O₄ - Magnetite

- Oxidation number for each iron averages to +8/3
- Magnetite contains two Fe³⁺ ions and one Fe²⁺

D. Characteristics of Oxidation-Reduction Reactions

- the oxidized substance:
 - loses electrons
 - increases oxidation state
 - is the reducing agent
- the reduced substance
 - gains electrons
 - decreases oxidation state
 - is the oxidizing agent

4.10 Balancing Oxidation-Reduction Equations

There are no notes for this section. The only way to master the balancing of redox equations is to actually balance them. While there are some minor variations in the processes used for acidic and basic solutions, the skills involved are identical. We will practice balancing numerous redox equations as a class.

There are two methods for balancing redox equations covered in this text:

Section 4.10 – Oxidation States Method

Section 18.1 – The Half-reaction Method

We will practice both methods in Unit 2.