Physical Science
Lecture Notes
Chapters 17, 18 & 19

I. 17-1: Matter & Its Changes
  a. Changes in matter
    i. **Physical Changes** – Alters form or appearance but doesn't change it into another substance ie. Water evaporates into water vapor, a rock is broken into pieces
    ii. **Chemical change**- changes the material into a new substance i.e. hydrogen and oxygen combine to form water.
      1. Chemical reactions take place when chemical bonds are either formed or broken.
      2. Strong chemical bonds resist change: glass
      3. Weak chemical bonds breakdown easily: wood
  b. 17-3 Describing chemical reactions
    i. Writing Chemical Reactions
      1. Elements are represented by a one or two letter symbol
        a. When symbol is a single letter: always capitalize: Hydrogen=H
        b. When symbol is two letters, capitalize first letter & lower case second letter: Sodium = Na
      2. Chemical formulas show the ratio of elements found in molecules and compounds
        a. **Subscript** numbers designate how many atoms of each element are present: H_2O_2 ; 2 Hydrogen atoms and 2 Oxygen atoms are present in this molecule
        b. When no subscript number is shown: it is understood that there is only one atom present: H_2O = 2 Hydrogen atoms and only one Oxygen atom are present in this molecule
    ii. Structure of an equation: summarizes the changes taking place in a chemical reaction
        a. Beginning materials are **reactants**
        b. Ending materials are **products**
        c. **Conservation of Mass** - Matter cannot be created nor destroyed so there must be the same number of atoms on each side of the equation
        d. Example of Chemical reaction:
           Reactant + Reactant $\rightarrow$ Product + Product
        e. **Coefficient**: a whole number in front of an element or molecule in a chemical reaction: Tells how many of each compound or element is present

$$\mathbf{6}CO_2 + \mathbf{6}H_20 \rightarrow C_6H_{12}O_6 + \mathbf{6}O_2$$

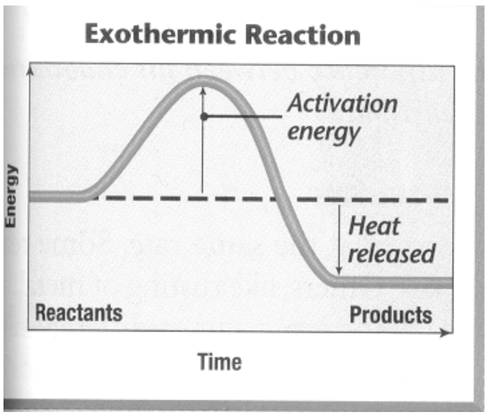
Coefficients

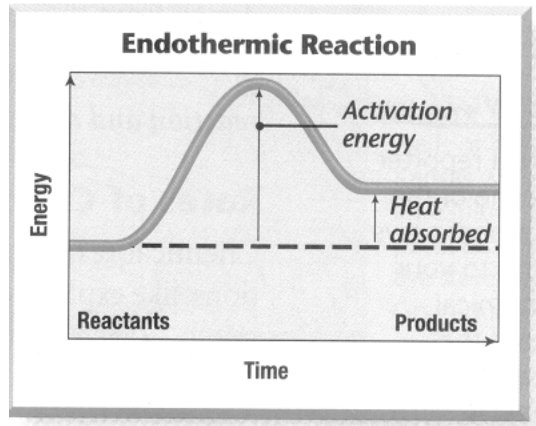
      2. Classifying Chemical Reactions
        a. reactions can be classified into one of three categories depending how the reactants and products change,
          i. **Synthesis**: When two or more substances combine to form a more complex substance: $2H_2 + O_2 \rightarrow 2H_2O$
          ii. **Decomposition**: When a complex substance is broken into two or more simpler substances: $2H_2O \rightarrow 2H_2 + O_2$

iii. **Replacement**: When one element replaces another or when two elements in different compounds change places:

$$2CuO + C \rightarrow 2Cu + CO_2$$

c. 17-3 Controlling Chemical Reactions
   i. Energy in Chemical Reactions
      1. Every chemical reaction involves a change in energy.
         a. Some reactions **release energy** in the form of **heat** (***<u>exothermic</u>***).
         b. Some reactions **absorbs energy** and the container holding the reaction **gets colder** to the touch (***<u>endothermic</u>***)
   ii. Getting Reactions Started





I. 18-1 Inside an Atom
a. Models of Atoms
   i. **Dalton Model**– 1808, atoms are thought to be solid marble like objects
   ii. **Thomas Model** – 1897, atoms thought to be solid positively charged sphere w/ electrons embedded. I.e. a muffin w/ raisins scattered in it.
   iii. **Rutherford** – 1911, first to say positive nucleus w/ electrons in random orbit
   iv. **Bohr** – 1913, agreed w/ Rutherford but said electrons in distinct layers or orbits
   v. **Chadwick** – 1932, discovered neutrons and said they were in the nucleus
   vi. **Modern Model** – 1920's to present, says electrons somewhere in a "cloud" around the nucleus.
b. An atom consists of a nucleus surrounded by one or more electrons
   i. **Nucleus** contains
      1. **Protons – positively** charged (+)– 1 **AMU** (atomic mass unit)
      2. **Neutrons** – neutral charge – 1 **AMU** (atomic mass unit)
   ii. Outer orbits contain electrons w/ a negative charge .0005 AMU
      1. **electrons** (negative charge) (-) travel at extremely high speeds around the nucleus in a "cloud" called an orbit.
   iii. Atoms are **electrically neutral** w/ the **same number of protons as electrons**. The number of positive charges are balanced by the same number of electrons
   iv. Majority of the atom is **empty space**. If nucleus were the size of a pencil eraser, the closest electron would be 100yards away!
c. Electron Orbits and **sub orbits**
   i. Named: 1s,2s,2p,3s,3p,3d,4s,4p,4d,4f,5s,5p,5d,5f,6s,6p,6d,6f,7s
   ii. How many in electrons each sub orbit?
      1. S sub orbits hold 2 electrons
      2. P sub orbits hold 6 electrons

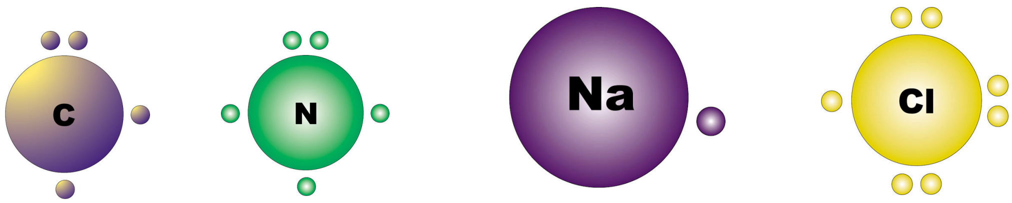
3. D sub orbits hold 10 electrons
4. F sub orbits hold 14 electrons

iii. Elements become **stable** when:

1. **their outer orbit contains 8 electrons** or
2. **their outer orbit becomes empty**

iv. **Valence electrons** are the electrons located in the outermost orbit

1. one way to show the number of valence electrons is w**/ Lewis Dot diagrams**



d. Why atoms form bonds

i. Chemical bonds form between two atoms when valence electrons move between them.

1. Electrons are either shared between them (**covalent bond**)
2. or Electrons are transferred (stolen) from one atom by another **(ionic)**

II. 18-2 Atoms in the Periodic Table

a. **Atomic Number** – the number of protons (+) in an atom

b. Since an atom is electrically neutral (same number of + and – charges), the atomic number also tells us the number of electrons.

c. **Atomic Mass** – the # of AMU's of an atom. An atom's mass. This is simply what the mass of the atom would be if we could "weigh" it. Since a proton has 1 AMU, a Neutron also has 1AMU and an electron is basically 1/2000 of an AMU,

i. The **Atomic Mass is the # Protons plus #Neutrons**

ii. Atoms of an element w/ varying numbers of neutrons are called **Isotopes**.

iii. **Allotrope** - Elements that form different molecular forms (ie oxygen gas O2 and ozone O3)

d. Periodic Table

i. **Columns** are called **families**, or **groups**.

1. Based on the number of **Valence Electrons**
2. Have their own characteristic properties.

ii. **Rows are called Periods** (hence the name "Periodic Table")

III. Chapter 18-3: **Ionic Bonds- Stealing Electrons**

a. Ionic Bonds form when a metal combines with a nonmetal

b. Ionic bonds are generally stronger than Covalent bonds

c. Ion: When an atom gains or looses electrons and becomes electrically charged

i. **Cation**- a positively charged ion

ii. **Anion**- a negatively charged ion

d. Electron Transfer

i. Atoms w/ 1,2 or 3 valence electrons transfer them to other atoms

ii. Atoms w/ 5, 6 or 7 valence electrons "steals" from other atoms

e. Polyatomic Ions

i. Ions made of more than one atom

ii. Stay together when chemically combined w/ other ions

iii. Common Polyatomic Ions: Need-To-Knows:

1. $\mathbf{HCO_3^{-1}}$ Bicarbonate
2. $\mathbf{NO_3^{-1}}$ Nitrate
3. $\mathbf{O^{-2}}$ Oxide
4. $\mathbf{SO_4^{-2}}$ Sulfate

5. **CO_3^{-2}** Carbonate

f. Naming Ionic Compounds – The Rules:
   i. The **cation comes first** and takes the name of the metal or a polyatomic cation
   ii. The **anion comes second**
      1. If it is a single ion, the end of the element's name changes to ***-ide***
      2. If it is a polyatomic ion, the name remains the same

g. Properties of Ionic Compounds
   i. Crystal shape
   ii. High melting points
   iii. Electrical conductivity when in solution or in a liquid state

IV. Chapter 18-4: **Covalent Bonds: Sharing electrons**

a. Covalent bonds form when two or more nonmetals combine

b. Covalent bonds are generally weaker than ionic bonds

c. The number of bonds each element can form equals the number of valence electrons it needs to make a total of 8 valence electrons
   i. Oxygen has 6 valence electrons so it can form 2 bonds
   ii. Caron has 4 valence electrons so it can form 4 bonds
   iii. Chlorine has 7 valence electrons so it can form only 1 bond

d. When only one pair of electrons are shared – a **<u>single bond</u>** forms
   i. H_2O – Oxygen form single bonds with each Hydrogen atom

e. When two pairs of electrons are shared – a **<u>double bond</u>** is formed
   i. O_2 – Oxygen forms a double bond with another Oxygen atom
   ii. CO_2 – Carbon forms double bonds with both of the Oxygen atoms that it is bonded with

f. Properties
   i. Relatively low melting points
   ii. Poor conductors of electricity

g. Unequal Sharing of electrons
   i. Some atoms pull stronger on the shared electrons than other atoms
      1. These electrons move closer to these atoms and they become more negatively charged
      2. The atom that the shared electrons move away from become slightly positively charged
      3. Covalent bonds that do not share electrons equally are ***<u>polar</u>***
      4. Covalent bonds that share electrons equally are ***<u>nonpolar</u>***

V. Chapter 19 Working with Solutions

a. **Solutions and suspensions**
   i. Suspensions are mixtures where particles can be easily separated:
      1. A mixture of water and pepper is a suspension
      2. They are not evenly mixed together
      3. They can be filtered and separated
   ii. Solutions are a well mixed mixtures that can not be easily separated.
      1. Salt and water mixed together
      2. Salt is evenly distributed throughout the solution
      3. Cannot be filtered out by normal means
   iii. Solvents and Solutes
      1. Solvent – the part of the solution that is present in the largest amount
      2. Solute – the part of the solution present in the least amount
   iv. Types of Solutions
      1. Solutions can be made from different states of matter:

Solute	Solvent	Solution
Oxygen –gas	Nitrogen – gas	Air – gas
CO_2 – gas	Water – liquid	Soda Pop
Glycol – liquid	Water – Liquid	Antifreeze – liquid
Salt – solid	Water – liquid	Ocean water - liquid
Zinc – solid	Copper – Solid	Brass - Solid

v. Particles in solution
   1. Solute particles are separated from each other and are surrounded by solvent particles.
      a. Water is polar and easily dissolves ionic compounds i.e. NaCl
      b. Water can also dissolve many "nonpolar" particles because these particles may have a slight polar side of the molecule which allows the polar water to be attracted to these surfaces.
      c. Remember that most molecular bonds are a gradient between pure ionic and pure covalent types of bonds.
   2. Electrical conductivity of a solution depends on the degree of ionic bonding found in the solute. The more ionic the bond the more conductive the solution.
   3. Concentration – defines the amount of solute in solution
      a. Dilute – weak solution "less" solute present
      b. Concentrated – strong solution "more" solute present
   4. Solubility – the amount of solute that will dissolve in a solvent at a given temperature.
      a. Solubility is considered a characteristic property
      b. Saturated – point when no more solute can dissolve into the solvent at the given temperature
      c. Unsaturated – condition before saturation takes place
         i. Generally speaking:
            1. Higher temperatures will allow more of a solid to dissolve into a liquid
            2. Higher temperatures will hold less gas in solution than colder temperatures
   5. Effects of solutes on the solvent:
      a. Increased concentrations of solute in a solution will lower the freezing point and increase the boiling point of the pure solvent
         i. Salt spread over icy roads to melt the ice and turn it into water
         ii. Salt placed into cooking water will increase the temperature of the water before it starts to boil, i.e. decreasing cooking time of pasta as it cooks in hotter water.

VI. Describing Acids and Bases
   a. Properties of Acids – compounds that:
      i. Release free Hydrogen ions into solution (H^+)
      ii. Reacts with metals and carbonates
      iii. Turns blue litmus paper red
      iv. Tastes sour (never taste any solution unless told to do so)
      v. are corrosive, eating away other substances
      vi. have a pH less than 7.0
   b. Important Acids
      i. Hydrochloric HCl
      ii. Nitric Acid HNO_3

        iii. Sulfuric Acid H_2SO_4
        iv. Carbonic Acid H_2CO_3
    c. Acids react with metals like magnesium, zinc and iron and releases Hydrogen gas bubbles into the solution
    d. Acids react with different Carbonates to produce CO_2 gas (carbon dioxide)
        i. example: $2HCl + CaCO_3 \rightarrow CaCl + CO_2 + H_2O$
    e. Acids react with indicator papers like litmus. Turns Blue litmus paper red
        i. Some indicator papers change to a variety of colors depending on the pH of the Solution
    f. Properties of Bases
        i. Bases are compounds that:
            1. Release hydroxide ions (OH^-) into solution
        ii. Has a bitter taste (never taste any solution unless told to do so)
        iii. feels slippery
        iv. Reacts with indicators like litmus by turning red litmus blue
        v. has a pH greater than 7.0
    g. Common Bases
        i. Sodium Hydroxide $NaOH$
        ii. Potassium Hydroxide KOH
        iii. Calcium Hydroxide $Ca(OH)_2$
        iv. Ammonia NH_3

VII. Acids and Bases in Solution
    a. Acids in Solution
        i. Acids are made of a H^+ ion and an Anion (a negatively charged ion)
        ii. In water, acids dissociates (breakdown) into H^+ and anions
            1. $HCl \rightarrow H^+ + Cl^-$
    b. Bases in solution
        i. Most bases release hydroxides ions into the water
            1. $NaOH \rightarrow Na+ + OH^-$
            2. $NH_3 + H_2O \rightarrow NH_4 + OH^-$
    c. Strength of Acids and Bases
        i. Strength is a measure of how well an acid or base dissociates into ions in water
        ii. Strong acids – Hydrochloric and Sulfuric Acids – most of the molecules breakdown into their ionic form
        iii. Weak Acids – Carbonic and Acetic Acids – fewer molecules break into corresponding ions.
        iv. Need – to – know the difference between:
            1. Strong $\rightarrow$ Weak
            2. Concentrated $\rightarrow$ Dilute
            3. A diluted solution of a strong acid can still burn a hole in your skin or clothes!!!
        v. Measuring pH – the unit used to describe the strength of an acid or a base
            1. pH scale is from 0 to 14 and is the measure of the [ ] of hydrogen ions in solution

pH	**Acidic Or Basic**	Example
14	Base	NaOH – Drain cleaner
13	Base	
12	Base	
11	Base	Ammonia
10	Base	
9	Base	
8	Base	Baking Soda
7	Neutral	Pure Water
6	Acid	Milk
5	Acid	Banana
4	Acid	Tomato
3	Acid	
2	Acid	Lemon
1	Acid	
0	Acid	Hydrochloric Acid

d. Acid / Base Reactions
   i. When Acids and Bases are combined a Neutralization reaction produces water and salt
      1. Hydrochloric Acid and Sodium Hydroxide yields water and Sodium Chloride
         a. $HCl + NaOH \rightarrow H_2O + NaCl$
      2. Salt is an ionic compound formed from an acid / base reaction (neutralization)