Physical Science
Lecture Notes
Chapters 16, 17 & 18

I. 16-1 Organizing Elements
  a. The periodic table is laid out by increasing **atomic number** as you go across and down the table
  b. Main body of the table is organized into
    i. **18 vertical Groups or Families**
    ii. **7 horizontal Periods**
  c. **Noble Gases - Group 18 –** Non reactive, have a full compliment of valence electrons, 8 and are called the "**Inert Gases**" because they do not react w/ other elements. Examples include **Helium** (very low mass and is used in filling children's balloons and even airships and the "Goodyear Blimp) and **Neon** used in lighted bulbs to make a red glowing light (a neon light).
  d. **Halogens – Group 17 –** Very reactive nonmetals w/ 7 valence electrons. Need only one more electron to fill their outer shell. Will steal an electron from a reactive metal to form ionic bonds. Examples include **Chlorine** (the most abundant halogen), **Iodine** and Bromine (found in Seawater).
  e. **Chalogens** AKA: **Oxygen Family – Group 16 –** nonmetals w/ 6 valence electrons, need 2 electrons to fill the outer shell. Most common oxidation state is -2. Examples are **Oxygen** (ozone is one of its allotropes), **Sulfur** (responsible for that rotten egg smell when it combines w/ oxygen to form sulfur dioxide) **and** **Selenium** (one of the few non metals that are also a good conductor of electricity).
  f. The **Nitrogen Family – Group 15–** Elements w/ 5 valence electrons, need 3 to fill the outer shell. Examples include **Nitrogen** (the most abundant element in air) **and Bismuth** (the only metal in this family. Has a very low melting point so it is used in automatic fire sprinkler systems) and **Phosphorous** (its common allotropes are red and white). **Arsenic** (a sometimes poisonous element that is also used extensively in medicines).
  g. The **Carbon Family – Group 14 –** has 4 valence electrons, needs four more or needs to get rid of the four it has to have none in its outer orbit. Examples include **Carbon** (on which all life is based) and **Silicon**
  h. The **Alkali Metals – Group 1 –** Very reactive metals that have only one valence electron in the outer orbit and will freely give it away to become stable. Very soft metal (you could cut it w/ a plastic knife!). They form ionic bonds w/ Halogens and Chalogens. Examples include **Sodium** and **Potassium**.
  i. The **Alkaline Earth Metals – Group 2 –** not as reactive as Alkali Metals, but still very reactive. They have two valence electrons and generally give them up to nonmetals to form ionic bonds. Examples include **Calcium** and **Magnesium**
  j. **Transition Metals – Groups 3 thru 12 –** vary dramatically in reactivity, Their oxidation states vary. they are a bridge between the very reactive Alkali and Alkaline Earth Metals and the nonmetals.
  k. **Rare Earth Elements - Lanthanides and Actinides** - AKA (**Inner Transition Elements**). These are the bottom two rows of the periodic table set apart from the rest of the table.

II. 16-2 Metals
  a. Most of the elements are metals.
  b. Metals tend to form positive (+) ions.
  c. Physical Properties
    i. Such as hardness, shiny, malleability (pounded into shapes), ductility (stretched or pulled into a wire) electrical conductivity and magnetic. Examples include **Gallium** that has an oxidation state of +3

d. Chemical Properties
   i. Metals show a wide range of chemical properties.

III. 16-3 Non Metals and Metalloids
   a. There are **17 nonmetals**, each are located to the right of the zigzag line in the periodic table.
   b. Non metals tend to steal electrons when they form negative (-) ions.
   c. Physical Properties – in general the physical properties of nonmetals are opposite those of metals. Powdery, gaseous, crumbly, non conductive, dull, not ductile or malleable.
   d. Chemical properties – usually form ionic bonds when combined w/ metals (NaCl, FeO_2, and CaC_{l2}) and usually form covalent bonds when combined w/ other nonmetals (CO2, O_2, $C_6H_{12}O_6$)
      i. Asbestos – substance once used for its fire retardant characteristics but is no longer used because of it's a carcinogen.
   e. Even though **Hydrogen (H)** is located in Group 1, it is still a nonmetal and exhibits oxidation states of +1 and -1.
   f. **Metalloids** – AKA "semi metals" 7 elements on the zigzag border between metals and the non metals. Their properties will sometimes make them act like a metal and then sometimes act like a nonmetal.
      i. Most important characteristic is their varying ability to conduct electricity. Silicon is used to make Semiconductors which are used in making computer chips.

IV. 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 → Product + Product

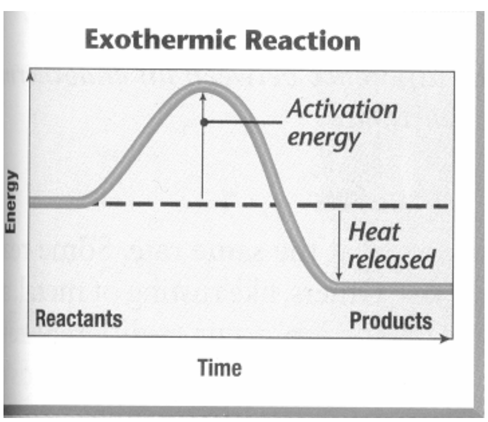
e. **Coefficient**: a whole number in front of an element or molecule in a chemical reaction: Tells how many of each compound or element is present

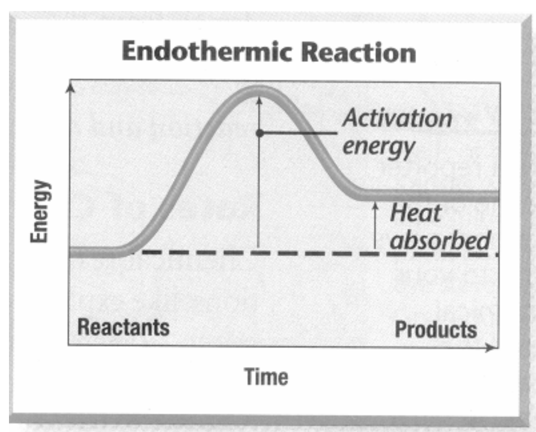
$\mathbf{6}CO_2 + \mathbf{6}H_20 \rightarrow C_6H_{12}O_6 + \mathbf{6}O_2$

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:
      $2CuO + C \rightarrow 2Cu + CO2$

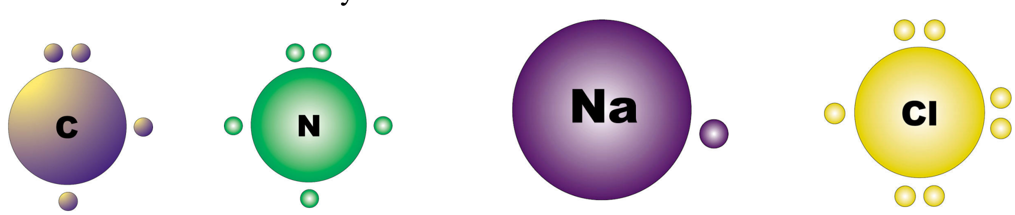
c. 17-3 Controlling Chemical Reactions
   i. Energy in Chemical Reactions
      1. Every chemical reaction involves a change in energy.
         a. Some reactions **release energy** in the form of **heat** (***<u>exothermic</u>***).
         b. Some reactions **absorbs energy** and the container holding 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. $\mathbf{CO_3^{-2}}$ 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 **single bond** forms
    i. H_2O – Oxygen form single bonds with each Hydrogen atom
e. When two pairs of electrons are shared – a **double bond** 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 ***polar***
        4. Covalent bonds that share electrons equally are ***nonpolar***