Physical Science
Lecture Notes
Chapters 14, 15, 16-1 & 16-2

Chapter 14

I. Describing Matter: **Matter is anything that has mass and occupies space**

A. Properties of Matter – How is it described: Hot, cold, hard, soft, rough, smooth, shiny, dull, solid, liquid, gas, etc.

B. States of Matter – Solid, Liquid, Gas or Plasma

C. Characteristic Properties – Those properties of a given substance that do not change and therefore can be used to help identify the substance.

1. **Boiling Point** – The temperature at which a liquid changes to a vapor at or below the surface of the liquid
    i. **Evaporation** – liquid changing to a gas only at the surface of the liquid
2. **Melting Point** – Temperature at which a solid changes to the liquid state

D. Changes in Matter

1. **Physical Change** – A change that alters the form of a substance but not the chemical makeup of the substance
    i. Examples of physical changes – water changing from solid to a liquid or a gas, salt dissolving in a liquid, tearing a piece of paper, slicing a metal in half, chopping wood.
2. **Chemical Change** - One or more substances combine or decompose to form a chemically different substance
    i. Examples – Iron rusts, wood burns, H_2 and O_2 forms water, acids react & release Hydrogen

E. Types of Matter

1. **Pure Substances** – Those substances made up of one kind of matter. It has definite characteristic properties
    i. Examples include: Sugar, alcohol, salt, iron, copper, lead, hydrogen, oxygen
    ii. Two types of pure substances
        1. **Elements**
            a. Contain only one type of atom,
            b. identified by a symbol containing a single capital letter or a two letter symbol: H for Hydrogen, Fe for iron
            c. can not be broken down into other substance by chemical means
            d. H, He, Na, Mg, C, N, O, K, P, S, Cl, Cu, etc.
        2. **Compounds**
            a. A pure substance formed by the chemical combination of two or more elements
            b. CO_2 – Carbon dioxide, H_2O – Water, $C_6H_{12}O_6$ – Glucose, NaCl – Table Salt, O_2 – gaseous Oxygen
2. Mixed Substances – two or more substances that are mixed together but not chemically combined.
    i. The chemical properties of each separate substance is maintained.
        1. Homogeneous Mixture: a very well mixed mixture where the substances are difficult to separate, i.e. a solution of sugar water or salt water.
        2. Heterogeneous Mixture: a mixture where the substances are not evenly mixed, salt and pepper mixed together, a handful of dirt, Rocky Road Ice Cream, etc.

II. Measuring Matter

A. SI – **International System of Units** = the metric system

a. **Length** – the one dimentional measurement of distance – SI unit is Meter, Kilometer

b. **Mass** – the amount of matter in a substance – SI unit: gram or kilogram

c. **Weight** – the force of gravity acting on an object – SI unit: Newton

d. **Volume** – how much space an object occupies – SI unit: liter, milliliter, cm^3

i. **Volume = Length x Width x Height**

ii. 1ml = 1cm3

e. **Density** – the amount of mass an object has in a given volume – SI unit: g/ml, g/cm^3

i. **Density= Mass / Volume**

f. **Temperature –** the average kinetic energy of an object.

i. SI unit- degrees Centigrade or degrees Celcius

g. Time: unit of measure: second, minute

III. Particles of Matter

A. **Atoms** – The smallest particle of an Element that retains the chemical properties of that element

B. **Democritus** – 400 BC, a Greek philosopher that coined the term "atomos" which means uncuttable

C. **Dalton** – 1802, Scientist that describes the Atomic Theory

a. **Atomic Theory** made up of 5 important points:

i. Atoms can not be broken into smaller pteces – atoms are like a solid marble (Not entirely accurate)

ii. In an element all atoms are exactly alike (Not entirely accurate)

iii. Atoms of two or more elements can combine to form compounds (this is true)

iv. Atoms of each element have a unique mass (Not entirely accurate)

v. Compounds are always composed of whole number proportions of elements ie CO_2 – Carbon dioxide, H_2O – Water, $C_6H_{12}O_6$ – Glucose, NaCl – Table Salt (this one is true also)

D. **The basic particle of an element is the Atom** – H, He, Fe, etc

E. **The basic particle of a compound is the Molecule** – a group of atoms that are chemically bonded and act as a single unit until the bonds are broken: CO_2, H_2O, $C_6H_{12}O_6$, NaCl

F. How small are atoms?

a. Sheet of paper is approx 10,000 atoms thick

b. 1 drop of water contains 2×10^{21} atoms of oxygen and 4×10^{21} atoms of Hydrogen

Chapter 15 – Changes in Matter

I. Solid, Liquid, Gas or Plasma

a. **Plasma** – state of matter that has had the electrons stripped away, uncommon on the Earth

b. **Solid** – Definite Shape and Volume

i. Particles (atoms or molecules) are packed closely together and stay in a fixed position

ii. Movement consists of vibrating particles staying in place

iii. Two types of Solids

1. Crystalline Solid – particles in a distinct pattern, melt at a specific distinct temperature: examples include sugar, salt, ice, quartz, etc.

2. Amorphous Solid – particles arranged in an irregular pattern, therefore no real distinct melting point. As heat is applied to an amorphous solid, the substance changes from hard to softer and softer until a liquid: examples include rubber, plastic, and Eskimo pies!

c. **Liquid** – No shape of its own, takes on the shape of the container, but it has a definite volume
   i. Particles in a liquid – atoms are loosely bound and are free flow and move, sliding easily over , under, and across each other but remaining in contact w/ one another.
   ii. **Viscosity** – resistance of a liquid to flow:
      1. high viscosity: slow flowing (molasses in January)
      2. low viscosity: fast flowing (pouring water)

d. **Gas** – No shape of its own, takes the shape of the container it is in. No definite volume, easily compressed – More will be discussed on gases later!!

II. Behavior of Gases

a. Measuring Gases – three important measurements taken for gases. All three are closely related!! **Volume Temperature and Pressure**
   i. **Volume** – Gases dissipate to evenly fill the container they occupy
   ii. **Temperature**
      1. The measurement of the average thermal energy of the particles in the gas.
      2. The average speed of a gas molecule at room temperature is fast!! 500 meters per second!
   iii. **Pressure**- gas particles are in constant motion and exert pressure upon the container they occupy. Because the gas particles are in motion, they collide and bounce off each other and the sides of the container. This contact w/ the sides of the container causes and outward push.
      1. $$\textbf{Pressure} = \frac{\textbf{Force}\ \text{(the outward Push)}}{\textbf{Area}\ \text{(of the container)}}$$

b. Relating Temperature and Pressure (at a constant volume)
   i. If the temperature increases, the added thermal energy causes the particles to push harder on the inside surface of the container… this causes the pressure to also go up.
   ii. If the temperature decreases, the pressure decreases.
   iii. Example: leaving a basketball outside on a cold night causes the ball to go flat

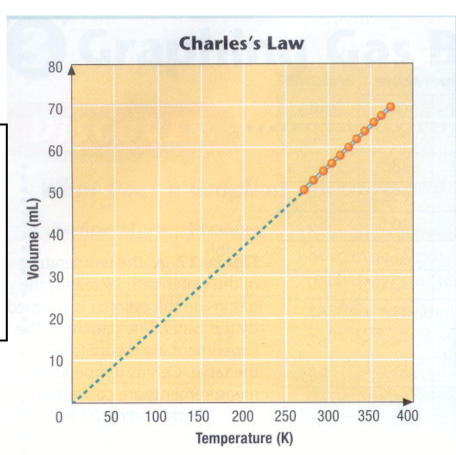
c. Relating Pressure and Volume (at a constant Temperature)
   i. **BOYLES LAW** – As pressure is increased volume will decrease, and conversely; if the pressure is decreased, the volume will increase

d. Relating Volume and Temperature (at a constant Pressure)
   i. **Charles Law**- As the temperature increases the volume will also increase; conversely, as the temperature decreases the volume will also decrease.

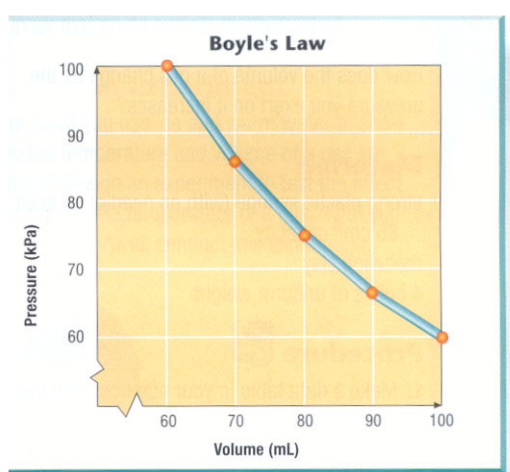
III. Graphing Gas Behavior

a. Graphs
b. **Charles Law** (Temp + → Volume + @ constant pressure)
c. **Boyles Law** (Pressure + → Volume - @ constant temperature)

Responding Variable



Manipulated Variable



IV. Physical vs Chemical Change
  a. Energy and Change
    i. **Physical Change –** alters the form of the substance but not its chemistry
    ii. **Chemical Change –** the substance is changed chemically into an entirely different substance.
    iii. **Thermal Energy** – energy imparted to the moving particles, more energy causes the particles to move faster:
      1. higher temp = higher Thermal Energy
    iv. **Chemical Energy** – The energy stored within the chemical bonds of the substance
    v. **Law of Conservation of Energy**: energy cannot be created or destroyed. It simply changes. The total energy stays the same in all reactions.
  b. Changes between liquid and solids
    i. **Melting:** changing from a solid to a liquid (ice melting in a glass of iced tea)
    ii. **Freezing:** changing from a liquid to a solid (water turning to ice cubes in the freezer)
  c. Changes between Liquid and a gas
    i. **Vaporization:** liquid changes to a gas at or below the surface of the liquid (boiling water)
    ii. **Evaporation:** liquid changing to a gas only at the surface of the liquid (a puddle drying up in the sun)
    iii. C**ondensation**: gas vapor changing to a liquid (rain)
  d. Changes between a solid and a gas
    i. **Sublimation**: Changing from a solid directly to a gas (dry ice turns to carbon dioxide, snow "disappears" w/out melting.

V. 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.

VI. 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.

VII. 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.