Topic 1: The Atom

J.J. Thomson; **Cathode Ray Tube** was used to discover the electron, negatively charged molecules

Plum Pudding Model: Atom made up of negatively charged particles embedded in a positively charged region

Rutherford Gold Foil Experiment: Rutherford **observed** a lot of glow directly behind the gold foil on the fluorescent screen and few to the side

Conclusion: 99.999% of atom empty space, Very small + region in space. Center of atom, nucleus

Bohr: 7 Energy levels. Energy levels increases, stability decreases as force of attraction weaken

Excited State: Conditions that exist when the electrons of an atom occupy higher energy levels while lower energy levels are vacant. Electrons absorb exact amount of energy to move up from ground to excited state, then move back down to ground state. Emit same amount of energy they absorb up to excited state

Wave Mechanical and Planetary Model: Both have a nucleus and electrons are on energy levels

Differences

Planetary Model: Electrons are on orbits, Energy level is more definite

Wave Mechanical Model: Electrons in orbitals, Energy level is much wider

Isotopes: Atoms of the same element that have the same atomic number but different atomic mass

Electromagnetic spectrum: Electrons emit energy as light when they jump from excited state to ground state. Bright line spectra are produced

Types of Matter

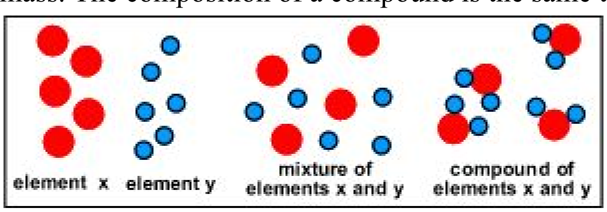
Mixture: 2 or more substances physically combined. Components can separate by physical methods.

Homogeneous: Particles evenly distributed; all parts are the same. 1 part; uniform in composition

Heterogeneous: Sample whose particles are not evenly distributed. Different parts

Pure Compound: Homogeneous and can be decomposed into simpler substances

Compounds: Composed of 2 or more elements that are chemically combined in definite proportions by mass. The composition of a compound is the same throughout



Topic 2: Formulas and Equations

Diatomic molecule: Elements exist in nature as 2 identical atoms. (they are all nonmetals)
Empirical Formula: Represents the simplest integer ratio in which atoms combine to form compounds.
Molecular Formula: Covalently bonded substances form discrete units called **molecules.** The molecular formula may be a multiple of the empirical formula

<u>Naming Compounds</u>

Binary compounds containing a metal and a nonmetal: Contain a metallic cation and a nonmetallic anion.Alkali metals form only ions with a +1 charge, the alkaline earth metals form only ions with a +2 charge, and aluminum forms only the ion Al^{3+}. For these ions, the name of the element followed by the term ion is an unambiguous name.

- $NiCl_2$ is nickel(II) chloride
- K_2S is potassium sulfide
- $CaBr_2$ is calcium bromide
- ZnO is zinc(II) oxide

Binary compounds containing two nonmetals but not hydrogen: Are molecular rather than ionic. They do not contain cations and anions. Carbon dioxide (CO_2) and phosphorus trichloride (PCl_3) are examples of such compounds. They are named using prefixes to state how many atoms of an element are in one molecule of the compound.
The name of the second element is modified to the root of its name followed by the ending ide. In both the formula and the name of these compounds, the most nonmetallic element comes first . The prefix mono is often omitted for the first element but never omitted for the second. Thus,

- CO is carbon monoxide
- SF_6 is sulfur hexafluoride
- N_2O is dinitrogen monoxide

Endothermic Process: Ice melting. Energy required is absorbed from surroundings ,lowing surrounding temperature. Reactions absorb energy as they become products, more potential energy.
Exothermic Processes: Release thermal energy. Burning of carbon in oxygen. Freezing of water. Energy released is given to surroundings. Products have less energy than the reactants.
Types of Reactions
Synthesis: When two or more simple compounds combine to form a more complicated one. General form of: A + B ---> AB **Example**: 8 Fe + S_8 ---> 8 FeS
Decomposition: Complex molecule breaks down to make simpler ones. General form: AB ---> A + B
Example: 2 H_2O ---> 2 H_2 + O_2
Single replacement : When one element trades places with another element in a compound.
A + BC ---> AC + B **Example**: Mg + 2 H_2O ---> $Mg(OH)_2$ + H_2
Double displacement: When anions and cations of 2 different molecules switch places, forming 2 entirely different compounds. These reactions are in the general form: AB + CD ---> AD + CB
Example: $Pb(NO_3)_2$ + 2 KI ---> PbI_2 + 2 KNO_3

Topic 3: The Mathematics of Formulas and Equations

Formula Mass: The mass of the smallest unit of the compound is the formula mass, which is the sum of the atomic masses of all the atoms present.

- Molecular mass is used to represent the mass of a unit of a compound, formula mass is preferred because ionic and network solids do not form discrete molecules

Gram Formula Mass: The formula mass expressed in grams instead of amu.
Percentage Composition: The percentage composition of a substance represents the composition as a percentage of each element compared with the total mass of the compound
Mole: Defined as the number of atoms of carbon present in 12 grams of C-12
Avogadro's Number: 6.02×10^{23} : Number of particles in a mole of a substance
Moles to Gram: # of moles * G.F.M./1 mol
Gram to mole: grams * 1 mol/G.F.M.
Mole Relations in Balanced Equations: Coefficients represent both the basic unit and mole ratios in balanced equations
Avogadro's Hypothesis: At STP, 22.4 liters of any gas contains 6.02×10^{23} ***molecules.***
Mol: 6.02×10^{23} molecules of any substance

- **Ionic Bonds use formula units**
- **Gram Molecular Mass:** Can not use for ionic and network formula
- **Formula Mass:** The sum of the masses of all atoms in a formula. IN AMU

Topic 4: Physical Behavior of Matter

Calorimetry: Measurement of heat q= mc(Δ)t (Remember this as q= mcAt)
Fusion: Constant temperature process in which particles in the solid phase gain enough energy to break away into the liquid phase; also known as melting
Sublimation: Phase change from solid to gas
Deposition: Phase change from gas to solid
Charles' Law: At constant pressure, temperature (IN KELVIN) and volume are directly proportional.
Pressure Increases: (T=K) P= 1/V

1. Distance between particles decreases, volume of the system decreases
2. Collision increase, pressure of gas particles on the walls of the state increases
3. Concentration increases- number of gas particles per unit area

Boyle's Law: Given mass of gas at constant temperature, volume of gas varies inversely with pressure
Kinetic Theory: If temperature is constant, as pressure of a gas increases, volume decreases. In turn, as pressure decreases, volume increases.

$$P_1V_1 = P_2V_2$$

P = Pressure of the gas
V = Volume of the gas

Temperature must be constant

Kinetic Molecular theory(KMT): Compare behavior of a real gas to an ideal gas
Perfect Gas: Gas particles move in straight line motion, change direction when they collide with wall of their system or with each other. Must have same amount of energy to run in constant motion.

1. Collisions, elastic, no energy lost, all energy transferred. Does not use heat

Gay Lussac's Law: States that the pressure of a gas is directly proportional to the Kelvin temperature if the volume remains constant

- As temperature of an enclosed gas increases, the pressure increases if the volume is constant

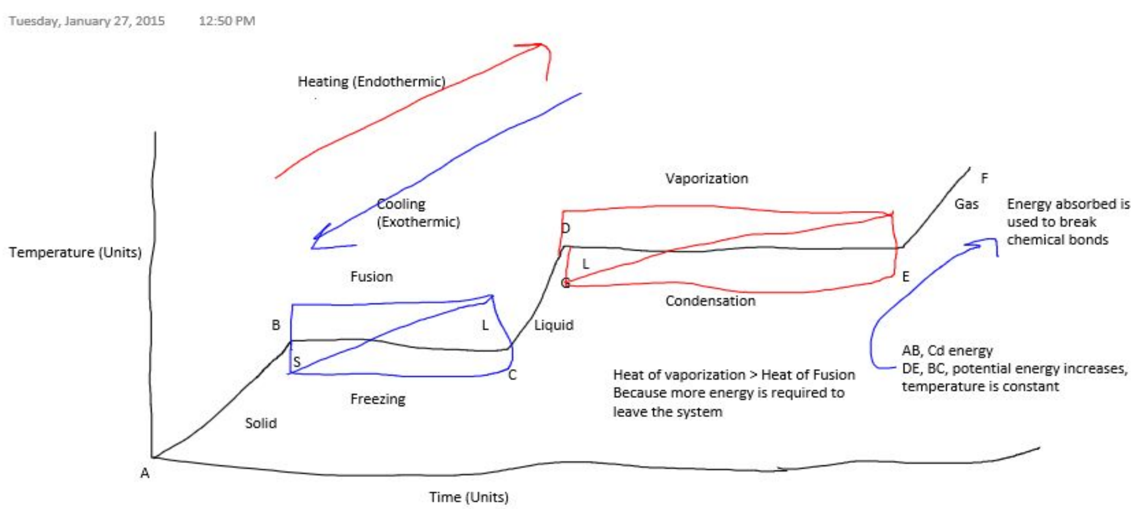
Combined Gas Law:If volume is constant, # of gas particles increase, collisions increase, pressure on the walls increases

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Dalton's Law: Contribution of each gas in a mixture makes to total pressure, partial pressure exerted by gas. In a mixture of gases, total pressure is the sum of the partial pressure of the gases.
Dalton's Law of partial pressure: At constant volume and temperature, total pressure exerted by a mixture of gases is equal to sum of partial pressure of the component gases.
If percent composition of a mixture of gases does not change, fraction of pressure exerted by a gas does not change as total pressure changes.

- **Diffusion:** Tendency of molecules to move toward areas of lower concentration until the concentration is uniform throughout
- During **effusion**, a gas escapes through a tiny hole in its container. **Gases of lower molar mass diffuse and effuse faster than gases of higher molar mass.**
- **If 2 objects with different masses have same kinetic energy, lighter object moves faster**

Phase-change diagram



Vapor pressure: **Why liquid boils:** Atmospheric pressure is greater than the vapor pressure
*The first liquid to boil has weakest force of attraction: propanol has the strongest vapor pressure

<u>Methods for Separating Mixtures</u>

1. **Filtration**: separates insoluble components from a solution
2. **Boiling or evaporation**: can reclaim a dissolved salt from salt water
3. **Distillation**: Separates components of a mixture based upon different boiling points.
4. **Decantation**: To pour off the top layer without disturbing the bottom layer
5. **Separatory Funnel**: can be used to separate two immiscible liquids based upon density.

Topic 5: The Periodic Table: When 2 atoms have same # of electrons means its isoelectronic.

Periodic Table (History):

- **Mendeleev organized elements in the periodic table in order of increasing atomic mass.**
- However, with this, pairs of elements that seemed out of order (I.g Iodine and Te)
- Henry Moseley created current periodic table. Determined atomic number based on X-rays

Periodic Law states: **Properties of all elements are periodic functions of their atomic numbers.**
Groups: Down, metallic character increases as e- farther away from nucleus, easier to remove them.
Properties of groups: Chemical reaction hydrogen often loses valence electron,form metallic hydrides.

- **Group 1 (alkali metals) and group 2 (alkaline earth metals) easily lose their electrons and are never found in nature in their atomic state. Francium is most reactive metal**
- Members of group 17 are called halogens when they are in the free state. When they form 1- ions the salts formed are called halides.

Metals: All metals are solid with the exception of mercury

- Most active metals are located in groups 1 and 2
- Metallic properties of elements increase from the top to the bottom of the group
- Elements of group 3 to 12 are called transition elements, or sometimes transition metals
- Characterized by multiple oxidation states
- Another unique property of such elements are that when they form ions they gain color
- **When metals become positive ions they decrease in size as they lose an energy level.**

Metalloids: Right of metals are metalloids (only 6 metalloids), B Si, Ge, As, Sb,and Te

- Nonmetals include solids, a liquid (bromine) and gases.
- Many nonmetals are gases or molecular or network solids at room temperature. Exception is bromine at room temperature.
- Not malleable or ductile, they tend to be brittle, solid nonmetals lack luster, high ionization energy and high electronegativity values, poor conductors of heat and electricity.

Noble gases: Only a few stable compounds containing noble gases have been formed.
Allotropes: Some non metals can exist in two or more forms in the same phase, known as allotropes. (Ex: Oxygen and Ozone)

- Allotropes have different physical properties and they also differ chemically.

Atomic Radii: Defined as half the distance between two adjacent atoms in a crystal or half the distance between the nuclei of identical atoms that are bonded together.

- **Down a period (from left to right) pattern is that atomic radii decreases**. Metals have larger radii than nonmetals. Down a period number of valence electrons are on the same energy level, number of protons attracting electrons increases, causing radius to decrease.
- **Down a group with more energy levels and the same amount of valence electrons, the atomic radius increases because the amount of energy levels between the valence electrons and the nucleus increases, decreasing attractive force.**
- **As atomic radius' increases they favor losing electrons. lower ionization energy**

Atomic radius decreases as number of protons increases, forces of attraction increase, p/e- ratio
Ionic Radii: Distance from the nucleus to the outer energy level of the ion.
When atom loses v.e and becomes positively charged it loses an energy level and radius decreases.
Radius of metallic ion smaller than radius of atom.When nonmetals form ions the radius increases
Electronegativity: Can be used to predict the type of bond that will form between two atoms.
Reactivity of elements: Some elements can be found uncombined in nature, the atomic state.

- Elements in Groups 1, 2, and 17 are rarely found to be uncombined.

Properties of Groups

Groups 1 and 2: Elements of Group 1, alkali metals and group 2 elements, alkaline earth metals, show typical metallic characteristics.Never found in nature in their atomic state, always found in compounds. Can be reduced to their free states by electrolysis of their compounds. From top to bottom reactivity increases. Group 1 elements are more reactive in than Group 2 elements.

- **Group 15:** Nitrogen, stable gas at room temperature as it contains a triple bond
- **Group 16:** Elements show progression from nonmetal to metal with increasing atomic number. Selenium and tellurium are metalloids. Oxygen is diatomic at room temperature. Oxygen is a reactive elements, easily forming compounds with other elements.
- **Group 17:** Are halogens when they are in the free state. When atoms of elements in this group gain an electron, the salts formed are halides. All elements are nonmetals,
- **Group 18:** Nonpolar and therefore held in the liquid and solid phase by van der Waals forces.

Topic 6: Bonding

Chemical Bond: Breaking of a chemical bond, absorbs energy. Formation of a bond, releases energy

- Chemical bond is formed, resulting compound has less potential energy than the substances from which it was formed. Energy is always released when a bond is formed. Greater the energy released during formation of a bond, greater its stability.

Lewis dot/Electron dot diagram: Kernel of every atom, positively charged. v.e surround kernel and usually represented by small dots, x's. Dots go clockwise formation after first 2 dots are placed on top. *****Atoms** that lose electrons don't have dot in electron dot diagram of ion. Square bracket around kernel and ionic charge written as superscript outside brackets indicate diagram is an ion

- Nonmetals v.e gained shown in its electron dot diagram. Common to show added valence electrons with a different symbol.

Lewis Electron- Dot Diagrams of Compound: Used to show how atoms combine to form molecules. When atoms combine to form diatomic molecules, single, double and triple covalent bonds can show.

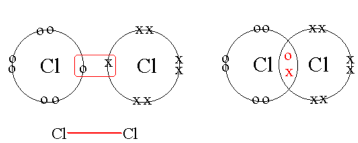
- Instead of dots to show pair of electrons, single dash may be used to show covalent bond. H-H

Members of group 17, halogens, form single bonds between them in their diatomic molecules. In addition to single covalent bond, each atoms has 6 additional valence electrons for a complete octet of 8

Nitrogen, 5 valence electrons and forms diatomic N2 molecule by forming a triple covalent bond

When drawing Lewis diagrams for compounds, pair of electrons forming a covalent bond usually represented by a dashed line, any unpaired electrons are represented by a pair of dots

Metallic Bonds: Bonds in solid and liquid phases are strong. Atom's valence electrons surround kernel. Kernels of metallic atoms making a metallic solid are arranged in fixed positions of a crystalline lattice. V.E move freely throughout crystal and do not belong to any given atom. Freely moving v.e give metals properties of good conductivity.



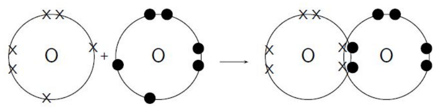
Metallic Bond: Results from force of attraction of mobile valence electrons for an atom's kernel.

- Freedom of valence electrons often leads to term "sea of mobile electrons" to describe metallic bonding.

Covalent Bonds: When 2 atoms approach 1 other, electrons and positive nuclei repel each other, tending to push atoms apart. Attractive force between atom comes from attraction of positively charged nucleus of 1 atom for negatively charged electrons of other atom. Chemical bonds occur when attractive forces between atoms are greater than repulsions forces

- Formed when 2 nuclei share electrons. QCovalent bonds often between nonmetal atoms of same element. Atoms of different elements may combine to form covalent bonds.

Nonpolar Covalent Bonds: Equal sharing of electrons. Formed between atoms having equal or close electronegativity values.



Multiple Covalent Bonds: Atoms sharing more than 1 pair of electrons.
Polar Covalent Bonds: Unequal sharing of electrons. Unequal electronegativity values.

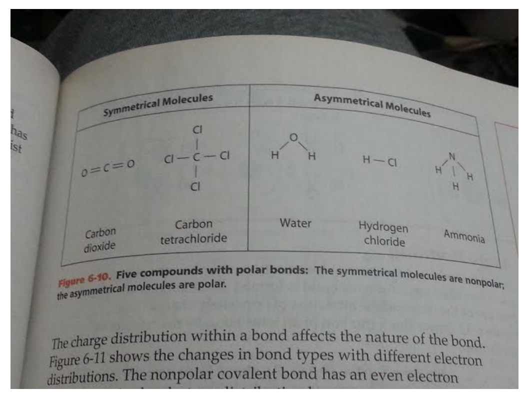
Molecular Substances

Molecule: Smallest particle of an element or compound formed by covalently bonded atoms. Each atom in a molecule usually has electron configuration of a noble gas. Molecular substances may exist as solids, liquids and gases, depends on strength of forces of attraction between molecules.

- Molecules generally have properties associated with covalent bonding. Molecules are generally soft, are poor conductors, and have low melting and boiling points.

Polar Molecules: Molecule may contain polar bonds, not necessarily that molecule itself is polar.

- Water, hydrogen chloride and ammonia molecules are asymmetric. Carbon dioxide, and carbon tetrachloride molecules are symmetrical. The symmetrical shape of the carbon dioxide and carbon tetrachloride molecules causes the pull of the various polar bonds to be offset by other bonds- the net results is a nonpolar molecule
- Charge distribution within a bond affects nature of bond, Nonpolar covalent bond have an even electron distribution, As the electron distribution becomes unequal, a polar covalent bond is formed. When the electron is actually transferred , an ionic bond is formed.



Electronegativity: As electronegativity between 2 atoms in a bond increases, bond more polar in nature. As electronegativity difference increases, bond ionic in character.

- If electronegativity difference between bonding atoms is 1.7 or greater, bond is generally ionic. Ionic bonds are generally formed between metals and nonmetals.

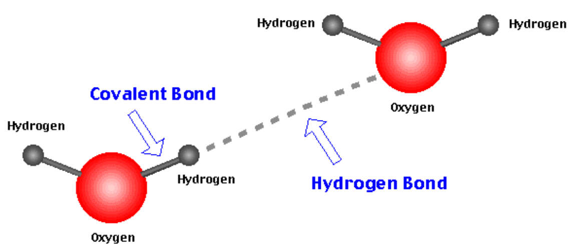
Distinguishing Bond Types: Metallic, covalent and ionic bonds different properties that can distinguish among them. Mercury, liquid standard conditions and an exceptions along with metals of Group 1. Ionic substances, conductors when they are melted (fused) or dissolved in aqueous solutions. Molecular substances held by covalent bonds. No charged particles to conduct an electric current. Molecular substances poor conductors, whether they are solid state, liquid state or aqueous solution

Properties of metallic, ionic and covalent bonds

Bond Type	Melting and Boiling Points	Hardness	Conductivity: Solid	Conductivity: Liquid	Conductivity: Aqueous
Metallic	High	Hard	Yes	Yes	Yes
Covalent	Low	Soft	No	No	No
Ionic	High	Hard	No	Yes	Yes

Intermolecular Forces: Molecules have attractive forces acting on them in solid and liquid states. As polar molecules have positive and negative ends, polar molecules are also dipoles. Positive area of 1 dipole molecule is attracted to the negative portion of an adjacent dipole molecule. These attractive forces are dipole-dipole forces.

Hydrogen Bonds: Is an intermolecular bond between a hydrogen atom in one molecule and a nitrogen, oxygen or fluorine atom in another molecule. Perhaps the most important example is water. Water is a polar molecule, so you can expect it to be held by other water molecule by dipole-dipole attractions.

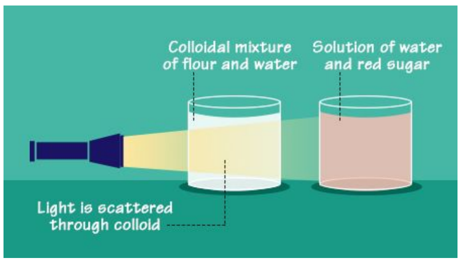


- Hydrogen bonding forces are much stronger than dipole-dipole attractions. Hydrogen bonding is responsible for the relatively high boiling point of water.
- Hydrogen bonding only occurs between hydrogen and nitrogen, oxygen and fluorine. N,O,F are highly electronegative and have small radii and only 2 energy levels.

Solutions : Homogeneous mixture of substances in the same physical state. Solutions contain atoms, ions or molecules of one substance spread uniformly throughout a second substance. When salt is stirred into water, the individual ions of the salt separate and uniformly spread throughout the water, forming a solution

Types of Solutions: A solid may be dissolved in another solid.When metals are mixed to form a solution, the results is called an alloy. Air is an examples of a mixture of gases forming a solution

- Most common type of solution is one in which a solid or a liquid is dissolved in a liquid



- **Solute**: Substance being dissolved, present in the smaller amount.
- **Solvent**: Substance that dissolves solute, present in greater amounts.
- Once the salt and water are stirred and the mixture becomes homogeneous, the dissolved particles will not settle. Liquid solutions are clear, and light will pass through a solution without being dispersed.
- Solutions may or may not have color.

Characteristics of liquid solutions

- Solutions are homogeneous mixtures
- Solutions are clear and do not disperse light
- Solutions can have color
- Solutions will not settle on standing
- Solutions will pass through a filter

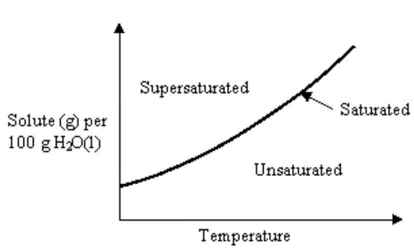
Nature of Solute and Solvent.

- When sodium chloride dissolves in water it does so becauses it's positively and negatively charged ions are attracted to the oppositely charged ends of the polar water molecule. The positively charged sodium ions are attracted to the negative pole of the water molecules. The attractive forces between the water molecules and sodium ions are greater than the attractive forces between sodium and chloride ions. In like manner, the negatively charged ions are attracted to the positively end of the water dipole and are dissolved Ionic and polar substances dissolve in polar solvents.

Solubility Summary

Solute Type	Nonpolar Solvent	Polar Solvent
Nonpolar	Soluble	Insoluble
Polar	Insoluble	Soluble
Ionic	Insoluble	Soluble

Temperature: As temperature increases, most solids become more soluble in water. Gases react in the opposite manner. As temperature rises, solubility of all gases in liquids decreases



Pressure: Pressure has little or no effect on the solubility of solid or liquid solutes. Pressure does affect the solubility of gases in liquids. As pressure increases, the solubility of gases in liquids increases.
Concentration of Solutions: Because solutions are homogeneous mixtures, their compositions can vary. Dilute and concentrated are relative terms and are not precise regarding amount of solute involved
Molarity: The molarity of a solution is the number of moles of solute in 1L of a solution
Molarity = Moles of solute/ liters of solution
Percent By Mass: The mass of an ingredient divided by the total mass, expressed as a percent
Percent Mass = Mass of part/ Mass of whole * 100%
Percent By Volume: The ratio of the volume of an ingredient divided by the total volume and expressed as a percent
Percent By Volume = Volume of solute/ Volume of solution = 100%
Parts per Million: Similar to percent composition as it compares masses. Parts per million (ppm) is the ratio between the mass of a solute and the total mass of the solution.
ppm = grams of solute/ grams of solution * 1,000,000 ppm

Boiling Point
- The heat required to chang 1 mole of a qa from a liquid at its boiling point to 1 mole of a vapor is the **heat of vaporization**

Topic 8: Kinetics and Equilibrium

Kinetics
- In order for a reaction to occur, reactant aww particles must collide- Collision Theory
- Collisions between particles can produce a reaction if both the spatial orientation and energy of the colliding particles are conductive to a reaction

<u>Factors Affecting Rate of Reaction</u>

1Nature of Reactants: Covalent bonds are slower to react than ionic substances due to the greater number of bonds that must be broken before a reaction can occur
- Breaking more bonds requires that the particles must have more energy when they acollide

Presence of a Catalyst: Catalysts are substances that increase the rate of reaction by providing a different and easier pathway for a reaction
- Catalysts take part in a reaction, but are unchanged when the reaction is complete

Factors	**Increases Rate**
Nature of Reactants	Ionic more than covalent
Concentration	Increased Concentration
Pressure	Increased pressure for gases
	Increased Temperature
Surface Area	Increased Surface Area
Catalyst	Presence of a Catalyst

Equilibrium

- Each potential energy diagram show a forward reaction and reverse reaction and even at the same time. When both the aQforward and reverse reaction occur at the same rate, the condition is called equilibrium. An equation representing equilibrium uses a double arrow instead of a single arrow to show that reaction ar proceeding in both directions
- **Equilibrium** is a state of balance between the rates of 2 opposite processes that are taking place at the same rate. Equilibrium can **only occur** in a system in which neither reactants nor products can leave the system
- **Equilibrium is a dynamic process**, it implies motion in which the interactions of reactant particles are balanced by the interactions of product particles. Equilibrium is an important concept because many chemical reactions and physical processes are reversible, they are able to proceed in both directions
- Quantities of reactants and products are not necessarily equal at equilibrium.

Physical Equilibrium: Equilibrium occurs during _physical processes such as change of state or dissolving

Phase Equilibrium: Can exist between the solid and liquid phases of a substance, the melting point of the solid phase or the freezing point of the liquid phase.

Solution Equilibrium: Solids in liquids exist in equilibrium in a saturated solution. Equilibrium may also be attained in a closed system between a gas dissolved in a liquid and the undissolved gas.

- The equilibrium can be disturbed by a change in temperature. If the temperature is raised, a solid generally becomes more soluble in a liquid.
- As temperature increases, rate of the gas escaping from liquid increases, while rate at which gas particles dissolve decreases. This decreases solubility of the gas in the liquid. As the temperature rises, solubility of all gases decreases in a liquid

Chemical Equilibrium: Condition which occurs when the concentration of reactants and products participating in a chemical reaction exhibit no net change over time. Chemical equilibrium may also be called a "steady state reaction." This does not mean the chemical reaction has necessarily stopped occurring, but that the consumption and formation of substances has reached a balanced condition. The quantities of reactants and products have achieved a constant ratio, but they are almost never equal. There may be much more product or much more reactant.

Le Chatelier's Principle: Any change in temperature, concentration or pressure or an equilibrium system is called a stress. Le Chatelier's principle explains how a system at equilibrium responds to relieve any stress on the system

- Increasing the temperature of a system in dynamic equilibrium favours the endothermic reaction. The system counteracts the change you have made by absorbing the extra heat.
- Decreasing the temperature of a system in dynamic equilibrium favours the exothermic reaction. The system counteracts the change you have made by producing more heat.

Enthalpy: Tendency in nature to change to a state of lower energy. Exothermic reactions move toward a lower energy state as some of the energy contained in the reactants is released. THe products have less potential energy than the reactants Drive toward lower energy is also the drive toward lower enthalpy

Entropy: Measure of disorder or randomness of a system. Greater the disorder, higher the entropy.

- **Examples** of entropy change are physical changes from the solid, crystalline phase (great order, low entropy), to the liquid phase (more randomness, higher entropy), to the gaseous phase (maximum randomness, highest entropy). FOr chemical changes, compounds represent a state of greater order and lower entropy than the free elements of which they are composed

1. Generally, the side of the equation with the greater number of molecules has the greater amount of entropy

Topic 10: Acids, Bases and Salts

Acids:

Arrhenius acid: Substance whose water solution contains the hydrogen ion or hydronium as the only positive ion.

Bronsted-Lowry acid: Any substance that donates a proton.

Properties:

- Sour taste
- Aqueous Solution conduct an electrical current. Substances that conduct an electrical current are called electrolytes. The greater the concentration of ions in a solution, the greater the electrical conductivity. If the solution of an acid is a good conductor of electricity, it is a strong acid.
- Acids react with bases to form salt and water, a neutralization reaction. The salt that is formed as a product is an ionic substance composed of a positively charged metallic or polyatomic ion and a negative ion other than the hydroxide ion
- Acids react with certain metals to produce hydrogen gas. Metals in Table J that are above hydrogen will react
- Acids cause acid-base indicators to change colors

Bases:

Arrhenius base: a substance whose water solution containing the hydroxide ion as the only negative ion.

Bronsted-Lowry base: any substance that receives a proton.

Properties:

- Bitter Taste
- Slippery or soapy feel
- Conduct an electric current. Higher concentration of ions that conduct electricity are a strong base
- Bases react with acids to produce water and a salt
- Bases cause acid-base indicators to change color.

Neutralization:

- Acid + Base = H_2O (water) + Salt
- Is Basic if Base is strong and Acid is weak
- Is Acidic if Acid is strong and Base is weak
- Is neutral if both Acid and Base is strong

Hydrolysis:

- H_2O + Salt = Acid + Base
- The reverse of neutralization

Bronsted-Lowry Acid

- A substance that donates a hydrogen ion (H^+)
- All Arrhenius Acids are Bronsted-Lowry Acids
- Includes acids that aren't aqueous

Bronsted-Lowry Base

- Any substance that accepts a proton (H^+)
- All Arrhenius Bases are Bronsted-Lowry Bases and more

Conjugated Acid-Base Pairs

- An Acid-Base pair are two chemical formulas that are different by a hydrogen ion.
- Example
    - $HNO_3 + H_2O = H_3O^+ + NO_3^-$
    - HNO_3 is an acid because it is donating a proton to make H_3O^+
    - H_2O is a base because it is accepting the proton given by HNO_3
    - HNO_3 and NO_3 are Conjugated Acid-Base Pairs because they only difference is an electron

Moles= molarity x volume
Molarity= moles/volume
$M_aV_a=M_bV_b$ ← Titration formula/ concentration formula. Concentration of H+ in acid= Concentration OH^- in a base.

Topic 12: Nuclear Chemistry

Nuclear Radiation: Unlike chemical, nuclear reactions not affected by changes in temperature, pressure. Nuclear reactions of a given radioisotope cannot be slowed, speeded up or stopped

Radioactivity: The spontaneous emission of rays or particles from certain elements, like uranium

Type	Consists of	Charge	Mass (amu)	Common Source	Penetrating Power
Alpha Radiation	Alpha Particles (Helium Nuclei)	2+	4	Radium-226	Low
Beta Radiation	Beta Particles (electrons)	1-	1/1837	Carbon 14	Moderate
Gamma Radiation	Gamma Ray	0	0	Cobalt-60	High

Nuclear Transformations: Neutron-to-proton ratio in a radioisotope determines type of decay

Nuclear Force: An attractive force that acts between all nuclear particles that are extremely close together, such as protons and neutrons in a nucleus

Positron: A particle with the mass of an electron but with a positive charge

Transmutation:In a nuclear reaction the nucleus of one element becomes the nucleus of another element. Identity changes, number of protons changes

- The nucleus of an atom changes if it is unstable
- The neutron to proton ratio determines the stability of the nucleus
- The greater the difference between the neutrons and protons, the more radioactive a nucleus
- Atoms in periodic table from 1-82 are considered stable nuclei. 83 and up are all radioactive
- An isotope that emits radiation is called radioisotope

Natural Transmutation: Result of unstable neutron-to-proton ratio,or by radioactive decay.

Artificial Transmutation: Bombarding nuclei with high-energy particles brings about the change

Transuranium Elements: Elements with atomic numbers above 92, atomic number of uranium

Fission and Fusion: In a chain reaction, some of the emitted neutrons react with other fissionable atoms, which emit neutrons that react with still more fissionable atoms

Fission: Split a heavy nucleus into lighter nuclei**.** Nuclei of certain isotopes bombarded with neutrons.

Fusion:Nuclei combine to produce a nucleus of a greater mass, release much more energy than fission reactions, in which large nuclei split apart to form smaller nuclei

- Most common ex. sun where hydrogen nuclei react in a series to produce helium molecules
- Extremely high temperatures and pressures needed
- 1 advantage as an energy source: Product not highly radioactive, like fission reactions
- In a nuclear reaction, mass of the products is less than the mass of the reactants. Mass lost is converted into energy E= mc^2
- Mass defect: Mass unaccounted for when protons and neutrons are put together in a nucleus.

Radioactive decay at a constant rate is not dependent on factors such as temperature. Random event

Nuclei Stability: When molecules in a cell are ionized, they may no longer carry on their normal functions and this may cause the death of the cell

Positron Emission : Atomic number decreases by 1: Number of protons decreases by 1

- Number of neutrons increases by 1: Mass number remains the same

Uses and Dangers of Radioisotopes

- Scientists can use the ratio of U-238 : P-206 to date rocks and other geological formations.

Chemical Tracers: Ability to detect radioactive materials and their decay products makes it possible to determine their presence/absence in a substance. Any radioisotope used to follow path of a material in a system is a tracer. (C-14 is used to map path of carbon in metallic processes.)

- I-131 haas uses both in detection and treatment of thyroid conditions.
- Co-60 (cobalt) emits gamma radiation as it decays. Can be used at cancerous tumors.
- Co-60 and Ce-138 are 2 of the elements used to produce gamma rays to treat anthrax bacilli.
- When Te-99 is given to patients with cancerous tumors it accumulates in the tumors.

Industrial Applications: Radiation products can be used to measure the thickness

Electrons are located in the orbital: Imaginary location of highest probability where electrons can be located : electrons have properties of both wave and particle