REGENTS CHEMISTRY: THE YEAR IN REVIEW

Unit 1: Atomic Concepts

Rutherford Gold-foil Experiment

In this experiment alpha particles were shot at a very thin piece of gold foil. A fluorescent screen was used to determine where the particles went after contact with the foil.

➢ **Observations of the experiment**
1. Most of the alpha particles went straight through the gold foil.
2. A small number of alpha particles were deflected slightly from the nucleus. An even smaller number of particles were greatly deflected.

➢ **Conclusions of the experiment**
1. The atom is made up of mostly empty space.
2. **The atom contains a small, dense, positively charged center called the nucleus**. Most of the mass of the atom is found in the nucleus.
3. This experiment disproved the idea that atoms were solid, dense spheres.

Atomic number
- The **atomic number is the number of protons in the nucleus** of an atom.
- It is also equal to the number of electrons for neutral atoms.
- Elements are arranged according to their atomic number on the periodic table.
- It does not change for a specific element.

Mass number
- **The mass number is the number of protons and neutrons in the nucleus of an atom.**
- Protons and neutrons are found in the nucleus of an atom. They can be referred to as nucleons.

To solve for the number of neutrons: **Neutrons = Mass Number - Atomic Number**

Characteristics of Subatomic Particles

Mass of a proton = 1 amu	charge of a proton = +1
Mass of a neutron = 1 amu	charge of a neutron = 0
Mass of an electron = 1/1836 amu	charge of an electron = -1

Isotopes
- Isotopes are atoms with the same atomic number but different mass number.
- Only the number of neutrons is different. (The protons and electrons are the same.)

Ex. Carbon-12 and Carbon-14 The number after the name of the element is the mass number.

The number of neutrons is different, but everything else is the same.

Ions
- Ions are atoms or groups of atoms that have a charge. (Polyatomic ions are listed on table E.)
- Atoms form ions by losing or gaining electrons.

Example: Compare the number of electrons for;

Be and Be^{2+}
Be loses 2 electrons to become Be^{2+}

F and F^{-1}
F gains one electron to become F^{-1}

Atomic Mass

The atomic mass is the weighted average mass of the naturally occurring isotopes of an element.

Example: The element copper contains the naturally occurring isotopes ^{63}Cu and ^{65}Cu. The relative abundances and atomic masses are 69.2% (mass = 62.93amu) and 30.8% (mass = 64.93amu), respectively. Calculate the average atomic mass of copper.

Answer:

Percentage	x	mass		
.692	x	62.93	=	43.55
.308	x	64.93	=	+ 19.99
		Total	=	**63.54 amu This is the atomic mass of copper.**

Wave-Mechanical Model (Electron Cloud Model) **This is the current model of the atom.**

- Electrons of an atom are found in regions of space around the nucleus called **orbitals.**
- **An orbital is a region of space where electrons are likely to be found.**
- Each electron in an atom has its own distinct amount of energy.

Ground State

The ground state is when all electrons of an atom are occupying the lowest possible energy levels (ex. 2-8-2). The electron configurations on the periodic table represent the ground state for that atom.

Excited State

The excited state is when electrons are in higher energy levels than the lowest possible energy levels (higher than the ground state).

For example:

1. Lower energy levels are not completely full before filling outer energy levels.
   ex. 2-7-3 The second energy level can hold up to eight electrons.
   Since it has only 7 it is an excited state.

Electrons **absorb energy** and the move into **higher energy levels.** This is how they form the excited state. Electrons **release light energy as spectral lines when they drop back to lower energy levels.** This emitted energy can be used to identify an element. **Each element has a unique set of spectral lines.**

Valence Electrons

- **Valence electrons are the outermost electrons of an atom.**
- Valence electrons affect the chemical properties of an element.
- All atoms in the same group have the same number of valence electrons.

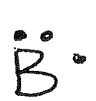
Ex. Nitrogen atoms have 5 valence electrons.

Lewis Electron - Dot Structures

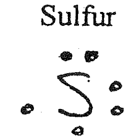
In this type of diagram, each dot represents one valence electron of an atom.
The element symbol represents the kernel (the nucleus and the inner electrons).

Sample Lewis dot diagrams

1. Boron



2. Sulfur



3. Nitrogen

Unit 2: Periodic Table

The elements on the periodic table are arranged in order of increasing atomic number.

Metals

- Most elements are metals (2/3 of the chart).
- They are found to the left of the staircase (except Hydrogen which is a nonmetal).

Properties of Metals

- Metallic luster
- Good conductors of electricity in all phases (due to mobile electrons)
- Malleable (gold - Au)
- Ductile (copper - Cu)
- Low electronegativity
- Low ionization energy
- Tend to lose electrons and form positive ions

Nonmetals

Nonmetals are found to the right of the staircase.

Properties of Nonmetals

- Poor conductors of electricity in all phases
- Brittle
- Low melting and boiling points
- High electronegativity
- High ionization energy
- Tend to gain electrons and form negative ions

Metalloids (semi-metals)

- Metalloids have properties of metals and nonmetals.
- They are found on or below the staircase.

Examples: (B, Si, As, Te, Ge, Sb)

For Groups 1, 2, and 13-18 on the Periodic Table, elements within the same group have the same number of valence electrons (He is the exception), therefore, similar chemical properties.

Alkali metals (Group 1) - (Remember the Rich and Holly story - Sodium in water explosion)

- These are the most reactive metals in any period.
- They exist as compounds in nature, never found uncombined.
- They react violently with water to produce hydrogen gas.

Alkaline Earth metals (Group 2)

- They exist as compounds in nature, never found uncombined

Halogens (Group 17)

- These are the most reactive nonmetals in any period.
- They contain elements in all 3 phases of matter (at STP, F_2 and Cl_2 are gases, Br_2 is a liquid, I_2 is a solid).
- They exist as compounds in nature, never found uncombined.

Noble Gases (Group 18) (What does a King give off when he has indigestion? - Noble gas)

- They have filled valence shells (complete outer shells - 8 valence electrons)
- They are unreactive but some can bond to fluorine (Ex. XeF_6)

Trends

Compare and contrast properties of elements within a group or a period for Groups 1, 2, 13-18 on the Periodic Table. Many properties are listed on Table S.

➢ **Atomic radius**
- listed on Reference Table S
- generally decreases from left to right across a period due to an increase in nuclear charge (the number of protons in the nucleus)
- generally increases as you go down a group due to an increase in the number of principal energy levels

➢ **Ionic radius** (Not listed on Table S)
- positive ions are smaller than their parent atoms.
  (Ex. Na^{+1} is smaller than Na because it lost one electron.)
- negative ions are larger than their parent atoms.
  (Ex. F^{-1} is larger than F because it gained one electron.)

➢ **Electronegativity**
- listed on Reference Table S
- **Electronegativity is the attraction for electrons in a chemical bond**
- generally increases from left to right across a period
- generally decreases as you go down a group

➢ **First Ionization Energy**
- listed on Reference Table S
- **amount of energy needed to remove the most loosely bound electron** from a neutral gaseous atom
- generally increases from left to right across a period
- generally decreases as you go down a group

➢ **Metallic Character**
- decreases from left to right across a period
- increases as you go down a group (Remember the Rich and Holly story - explosion is bigger as we move down the group)

➢ **Nonmetallic Character**
- increases from left to right across a period
- decreases as you go down a group

Diatomic molecules - H_2, O_2, F_2, Br_2, I_2, N_2, Cl_2 (Remember my teacher from high school: Mr. HOFBrINCl)

At STP (Standard Temperature and Pressure)

- Most elements are solids
- Mercury (Hg) and bromine (Br_2) are the only liquid elements
- **F_2 and Cl_2 are gases, Br_2 is liquid, and I_2 is a solid** due to increasing strength of Van der Waals forces for larger molecules

Allotropes

- Allotropes are elements that can exist in two or more forms in the same phase.
- Allotropes have different chemical and physical properties

(Ex. Carbon - graphite, diamond, buckminsterfullerene.)
Oxygen: O_2 - diatomic oxygen and O_3 - ozone)

Unit 3: Moles/Stoichiometry

Formula Writing -find the oxidation numbers of the elements and/or polyatomic ions (Reference Table E) and criss-cross-reduce

Ex. 1) Al^{+3} Cl^{-1} is $AlCl_3$ 2) Al^{+3} SO_4^{-2} is $Al_2(SO_4)_3$

Naming Compounds

➢ **Binary Ionic Compounds**
Name the metal first and non-metal second. Change the non-metal name ending to *-ide.*

Example: 1) MgS is magnesium sulfide 2) $CaCl_2$ is calcium chloride

➢ **Ternary Compounds**
These compounds contain polyatomic ions. Use reference table E.

Example: 1) $CaSO_4$ is calcium sulfate 2) NH_4Br is ammonium bromide

➢ **Stock System**
A roman numeral is used after the name of the metal when the metal has more than one oxidation state. (Remember the teacher that wanted to order iron chloride - without the roman numeral I wouldn't know which compound he wanted.)

Example: 1) $FeCl_2$ is iron (II) chloride 2) $FeCl_3$ is iron (III) chloride

➢ **Binary Covalent Compounds**

- These involve two nonmetals with the lower electronegative element named first.
- Change the second non-metal name ending to - *ide*
- Prefixes are used to tell the number of atoms of each element (1 is mono-, 2 is di-, 3 is tri-, 4 is tetra-, 5 is penta)
- If only one atom of the first element is present, the prefix *mono-* is not used

Example: 1) CO is carbon monoxide 2) N_2O_4 is dinitrogen tetroxide
3) CCl_4 is carbon tetrachloride 4) SO_3 is sulfur trioxide

Molar Mass (gram formula mass)

- The gram-formula mass is the formula mass expressed in grams.
- It equals one mole of the substance.

Example: What is the formula mass of K_3PO_4?

Element	Number of atoms	x Atomic mass	
K	3	39.1	117.3
P	1	31.0	31.0
O	4	16.0	+ 64.0
		Total =	212.3 g/mole

Molecular Formula
The molecular formula is the actual amount of atoms in a molecule.

Empirical Formula
The empirical formula is the simplest possible ratio of elements.

Example: Molecular formula is C_2H_6, the empirical formula is CH_3.
Molecular formula is N_2O_4, the empirical formula is NO_2.

Sample empirical formula problem
Example: A compound has a molecular mass of 180 g/mole and an empirical formula of CH_2O. What is its molecular formula?

Answer:
Find the mass of the empirical formula : CH_2O is 30g

Divide the molecular mass by the empirical mass: 180 / 30 = 6

Multiply the answer by the empirical formula: 6 x (CH_2O) = $\mathbf{C_6H_{12}O_6}$

Structural Formula
- This shows all the bonds that connect the atoms and the approximate molecular shape.
- You must be able to draw structural formulas of organic molecules.

Balancing Chemical Equations

Equations must have a conservation of mass, energy and charge. Only coefficients can be used to balance equations. You can not change or add subscripts to a formula to balance an equation.

Endothermic Processes
An endothermic reaction requires energy in order to occur. (absorbs energy)

Ex. H_2O (s) + energy → H_2O (l)
$6CO_2$ + $6H_2O$ + energy → $C_6H_{12}O_6$ + $6O_2$

*in these reactions the energy terms is on the reactant side. (left side)

Exothermic Processes
An exothermic reaction releases energy when it takes place.

Ex. H_2O (l) → H_2O (s) + energy

* in these reactions the energy term is on the product side

CH_4 + $3O_2$ → CO_2 + $2H_2O$ + energy (This reaction is complete combustion.)
Water and carbon dioxide are the products.

Percent Composition - Part/Whole x 100 (percent by mass)

Example: What is the percent composition of oxygen in $Ca(OH)_2$?

Mass of oxygen atoms is 32.0 g
Formula mass of $Ca(OH)_2$ is 74.1 g/mole.

Percent of oxygen is = 32.0 / 74.1 x 100 = 43.2%

Example: What is the percent of water in $CuSO_4 \cdot 5H_2O$?

Mass of $5H_2O$ is 90.0 grams.
Formula mass of $CuSO_4 \cdot 5H_2O$ is 249.3 g/mole.
Percent of water = 90.0 / 249.3 x 100 = 36.1%

Mole
The mass of one mole of a substance is its formula mass (molecular mass).

Conversions of moles to grams and grams to moles can be made using the equation found on reference table T.

$$\textbf{Number of moles} = \frac{\textbf{given mass (g)}}{\textbf{Gram-formula mass}}$$

Example: How many moles are equivalent to 14.1 g of CaS?

$$\text{Moles} = \frac{14.1\text{g}}{72.2 \text{ g/mole}}$$
$$\text{Moles} = 0.195$$

Example: How many grams are present in 2.50 moles of CO_2?

$$2.50 \text{ moles} = \frac{X \text{ grams}}{44.0 \text{ g/mole}} \qquad X = 110. \text{ grams}$$

Types of Reactions - Identify the four major types.

1. **Synthesis** - two or more reactants yield one product

   Ex: $P_4 + 6\,Cl_2 \rightarrow 4\,PCl_3$

2. **Decomposition** - one reactant breaks down into two or more products

   Ex: $2\,NO_2 \rightarrow 2\,NO + O_2$

3. **Single replacement** - an element reacts with a compound to form a new element and a new compound. Table J can be used to determine if the reaction will take place.

   Ex: $Zn + 2\,HCl \rightarrow ZnCl_2 + H_2$

4. **Double replacement** - two compounds react to form two new compounds.
Table F can be used to determine if the reaction will take place by looking up the solubility of the compounds.

Ex: $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaNO_3(aq)$

Since one of the products is not soluble ($BaSO_4(s)$), this double replacement reaction will take place. The $BaSO_4(s)$ is a precipitate.

Unit 4: Chemical Bonding

- When a bond is formed, energy is released.
- When a bond is broken, energy is absorbed.
- Atoms form bonds to become more stable (get 8 electrons in their outer shell).

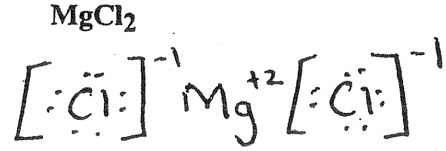
There are three types of bonds: Ionic, covalent, and metallic

Ionic Bonds
- Ionic bonds exist between a metal and nonmetal.
- The valence electrons are transferred from the metal to the nonmetal so atoms attain a stable noble gas electron configuration

Example: Lewis dot structures of KF and $MgCl_2$. **Only ionic compounds use the brackets in the dot structure.**

KF

$MgCl_2$



metal charge only

nonmetal get 8 dots in brackets.

charge outside brackets

Covalent Bonds
- Covalent bonds form between two nonmetals.
- The valence electrons are shared between the two nonmetals to attain a more stable electron configuration.
- Each covalent bond is made up of two electrons.

Nonpolar covalent bonds
- exist between two of the same nonmetal atoms
- electronegativity difference between the two bonded atoms is equal to zero
- equal sharing of valence electrons

Ex. All diatomic elements are held together by nonpolar covalent bonds.

Polar covalent bonds
- exist between two different nonmetal atoms (the elements have different electronegativities)
- unequal sharing of electrons
- the shared electrons are closest to the atom with the higher electronegativity

Coordinate covalent bonds

- exist in polyatomic ions
- any substance with a free, unshared pair of electrons could form this bond

Molecular Polarity (Remember " It's a SNAP"

S- symmetrical
N- nonpolar

A- asymmetric
P- polar

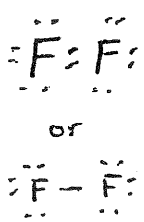
- Determined by the shape and distribution of charge in the molecule.
- Molecules can be polar or nonpolar.

➢ **Nonpolar molecules**

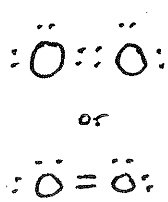
- Nonpolar molecules are symmetrical.
- A molecule that has only nonpolar bonds is nonpolar (Ex. Diatomic molecules)
- A molecule that has polar bonds but are arranged symmetrically is nonpolar (Ex. CO_2, CCl_4, CH_4)

Ex. Lewis dot structures for F_2, O_2, N_2, CO_2 and CCl_4. The shape of each molecule is listed.

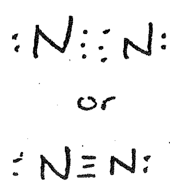
F_2
linear



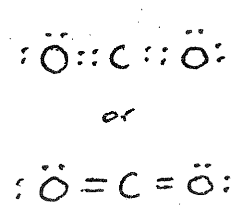
O_2
linear



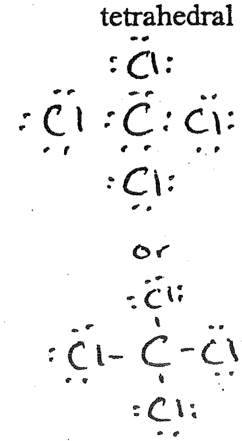
N_2
linear



CO_2
linear



CCl_4
tetrahedral

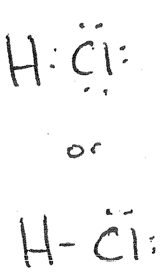


➢ **Polar molecules**

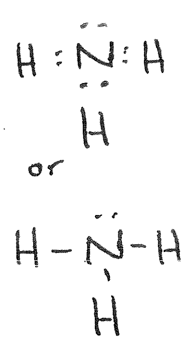
- asymmetrical molecules
- must have at least one polar covalent bond
- all molecules that have the bent or pyramidal shape are polar

Ex. Lewis dot structures for HCl, NH_3, H_2O, CH_3Cl. The shape of each molecule is listed.

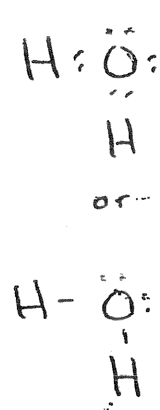
HCl
Linear



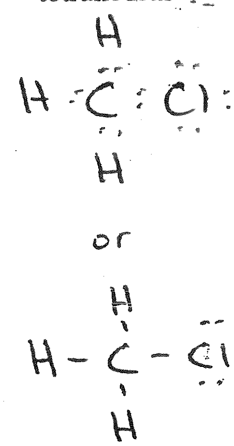
NH_3
trigonal pyramid



H_2O
angular(bent)



CH_3Cl
tetrahedral



Metallic Bonds

- metal by itself (Ex. Cu, Pb, Na, Fe, Hg)
- **positive ions immersed in a sea of mobile electrons**
- solid metals can conduct electricity due to the mobile electrons

Substances can be classified as Ionic or Molecular

Ionic Solids (Remember the gum drops and toothpicks)

- contain ionic bonds in a crystalline structure (crystal lattice)
- do not conduct electricity unless melted (liquid phase) or dissolved in liquid (aqueous solution if water is the solvent)
- are very hard
- high melting points
- ionic compounds containing polyatomic ions have both ionic and covalent bonding (NH_4Cl)

Molecular Substances

- contain covalent bonds
- do not conduct electricity
- are soft
- low melting points

Intermolecular Forces (Attraction Between Molecules)

The stronger the force of attraction between molecules, the stronger the molecules are held together. Molecules with strong forces of attraction have high boiling points.

Remember: Ionic compounds have higher melting than molecular compounds. Thus NaCl has a much higher melting point than H_2O.

van der Waals' Forces

These are very weak forces that exist between nonpolar molecules.

Ex. Ne - Ne or I_2 - I_2 or CH_4 - CH_4

The attraction is stronger for larger molecules and increases as molecules are placed closer together.

Dipole-Dipole Forces

- This is the attraction between polar molecules such as HCl.

Hydrogen Bonding (chemistry is "FON")

- much stronger intermolecular forces than dipole-dipole attractions
- exist when hydrogen is bonded to fluorine, oxygen, or nitrogen (these atoms are small in size and have high electronegativities)
- responsible for the high boiling point of H_2O as compared to H_2S

Know these three examples for compounds with hydrogen bonding:

Ex. HF - HF , H_2O - H_2O , NH_3 - NH_3

Molecule - Ion Attraction: (present in solutions of ionic compounds)

- results when an ionic compound dissociates in water
- the positive ions are surrounded by the negative end of water (oxygen end)
- the negative ions are surrounded by the positive end of water (hydrogen end)

Unit 5: Physical Behavior of Matter

Matter
Matter consists of pure substance or a mixture of substances.

Substance
- A substance is an element or a compound.
- Substances are homogeneous which means they have constant properties throughout a sample.

Phases of Matter

Solid (s) - definite shape and volume, molecules are close and in a regular, repeating, geometric pattern
Liquid (l) - definite volume but no definite shape
Gas (g) - no definite shape or volume, molecules are far apart and randomly arranged

Element
- Elements cannot be decomposed by chemical change.
- Elements are composed of atoms that have the same atomic number.

Compound
- Compounds are two or more elements that are chemically combined.
- Ex. NH_3 (ammonia), H_2SO_4 (sulfuric acid), Glucose ($C_6H_{12}O_6$)
- Compounds can be decomposed by chemical change and have a definite proportion by mass.

Mixture
- Mixtures are two or more substances that are physically combined.
- usually heterogeneous **(solutions are homogeneous mixtures)**
- each component in the mixture retains its own properties
- proportions of the components in a mixture can vary (no definite proportion by mass)
- components of a mixture can be separated based on differences in the properties

Techniques for the separation of mixtures include:

1. **filtration (particle size)**
   Filtration will separate insoluble solids from liquids. A mixture of salt and water can not be separated by this technique (the salt and water pass through the filter paper). However, sand and water can be separated by this technique (the sand gets caught in the filter paper).
2. **distillation (boiling point)**
   A mixture of two different liquids or a mixture of salt and water can be separated by this technique.
3. **chromatography (solubility and/or polarity and/or boiling point)**

Solution
- A solution is a homogeneous mixture.
- The solute is dissolved in a solvent.
- Remember the "like dissolves like" rule. Polar substances will dissolve in polar solvents.

Example: ammonia dissolves in water to form $NH_3(aq)$
(Both ammonia and water are polar substances)

Factors that affect solubility
Gases - most soluble at low temperature and high pressure
Solids - generally more soluble at high temperature, unaffected by pressure

Solubility Curves

- Reference Table G
- know the terms saturated, unsaturated and supersaturated solutions
- a point on the line, the solution is saturated
- a point below the line, the solution is unsaturated
- a point above the line, the solution is supersaturated

Solubility Guidelines

- Reference Table F
- Used to determine if an ionic compound is soluble or insoluble in water
- can be used to determine whether or not a double replacement reaction will work

Solution Concentration

Qualitative descriptions include dilute and concentrated.
Quantitative description used to express the specific amount of solute in a solution

Molarity (M) $M = \frac{\text{moles of solute}}{\text{liters of solution}}$ on Reference Table T

Example: How many grams of NaOH are needed to prepare 250. mL of 2.0 M NaOH(aq)?

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

2.0M = x/.250 liters x= 0.50 moles moles = grams/ formula mass

0.50 = X / 40.0 g/mole x = 20.0 grams

Parts per Million -used for extremely dilute solutions

$\text{ppm} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1{,}000{,}000$ on Reference Table T

1 ppm = 1mg of solute per liter of solution

Percent by Mass -commonly used to express concentration for many household items

$\text{percent mass} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100$ NOT on Reference Table T

Percent by Volume - commonly used to express concentration when 2 liquids are in a mixture

$\text{percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$

Note: Any percent will be part/whole x 100.

Colligative Properties

Any nonvolatile solute (salts and sugars) added to water **lowers the freezing point and raises the boiling point.**

Kinetic Molecular Theory

This theory is used to explain the behavior of a gas with these assumptions about the gas particles:

- they are in random, constant, straight-line motion
- have no volume
- do not attract each other
- are separated by great distances
- have collisions that may result in a transfer of energy between particles, but the total energy of the system remains constant
- a gas that follows these assumptions is called an ideal gas

Ideal Gases

- Assumed to have no volume and no force of attraction between molecules.
- **The most ideal gas is the gas with the lowest formula mass.**

  Ex. H_2 is more ideal than Cl_2 because it is lighter and smaller.
- **Gases act most like ideal under conditions of high temperature and low pressure.**

Real Gases

These gases deviate from ideal behavior for two reasons:

- Real gases have forces of attraction between molecules.
- Real gases have a small, but significant volume.

STP (standard temperature and pressure)

- Values are listed on Reference Table A.
- To change Celsius to Kelvin: $K = {}^{\circ}C + 273$ (Reference Table T)

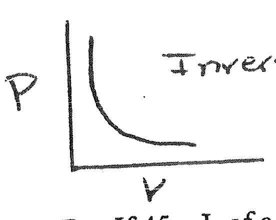
Combined Gas Law **(found on reference table T)**

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

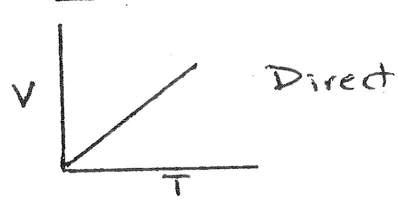
-temperature must be in Kelvin

Know the relationships and graphs for the following combinations: (Remember P T V with your pen - think of the ruler hanging in front of the room)

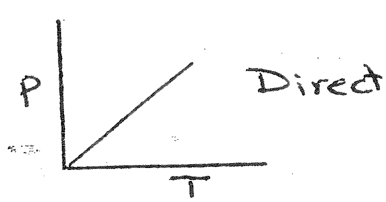
Pressure and Volume



Volume and Temperature



Pressure and Temperature



Ex. If 45 mL of a gas is at STP, what volume will the gas occupy if the temperature is increased to 85°C and the pressure is increase to 2.5 atm?

convert to Kelvin

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{(1\,atm)(45\,mL)}{273\,K} = \frac{(2.5\,atm)(x)}{358\,K}$$

$x = 24\ mL$

Temperature

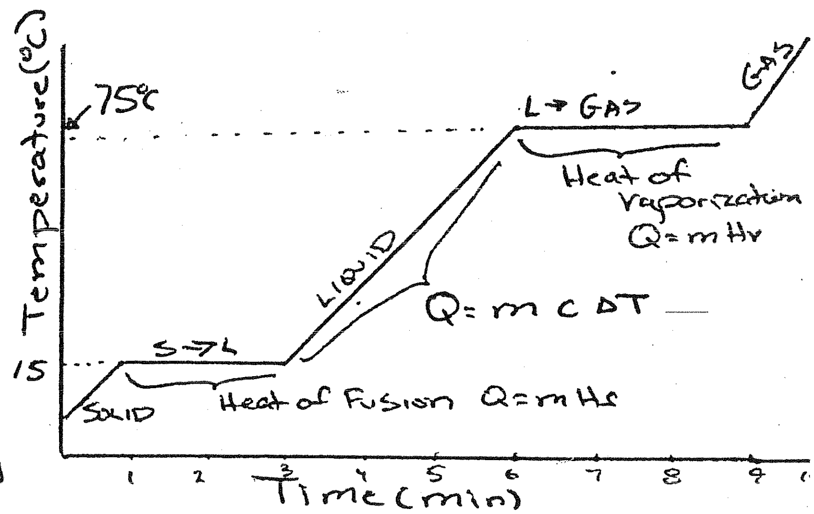
Temperature measures the **average kinetic energy** of a substance's particles.

Heating and Cooling Curve

The higher the temperature, the greater the kinetic energy of the molecules. During a change of phase, the potential energy is increasing and the temperature remains constant.

Sample data Table for a Heating curve

Time (minutes)	Temperature (°C)	
0	10	
1	15	Freezing Part
2	15	
3	15	
4	35	
5	55	
6	75	Boiling Pt
7	75	
8	75	
9	75	
10	100	



Celcius

Sublimation

- This is a phase change from a solid directly to a gas.

  Example: Iodine (I_2) and solid carbon dioxide (CO_2) also called dry ice
- Molecules with weak intermolecular forces will sublime.

Deposition

- This is a phase change from a gas directly to a solid (opposite of sublimation).

Vapor Pressure

- The weaker the forces of attraction between molecules the higher the vapor pressure of a substance
- As temperature increases, the vapor pressure of a liquid increases. (Reference Table H)

Forces of attraction also explain the rate at which a substance evaporates as well as how much is energy necessary to change the phase of a substance. Molecules with weak forces of attraction will evaporate quickly.

Heat Calculations

The equations are listed on Reference Table T and the constants on Table B.

Example: How many joules are absorbed when 150. g of water are heated from 25.0 °C to 52.0 °C?

$Q = m\,c\,\Delta T$

$Q = (150\,g)(4.18\,J/g\cdot{}^{\circ}C)(27.0)$ $\quad Q = 16{,}900\,J$

Heat of Fusion - heat energy needed to change a gram of solid to liquid
Heat of Vaporization - heat energy needed to change a gram of liquid to gas

Sample Problems:

1) How many joules of energy are required to melt 35.0 g of ice at $0^{o}C$?

$$Q = m H_f$$
$$Q = (35.0g)(334 J/g)$$
$$Q = 11{,}700 J$$

2) How many joules of energy are required to vaporize 35.0g water at $100^{o}C$?

$$Q = m H_v$$
$$Q = (35.0g)(2260 J/g)$$
$$Q = 79{,}100 J$$

Boiling Point

Boiling occurs when the vapor pressure equals the atmospheric pressure.
The lower the air pressure the lower the normal boiling point. Think about when I said I would move to Denver, Colorado. Water boils at temperature below $100^{o}C$ due to the elevation (higher elevation - lower air pressure)

Unit 6: Kinetics/Equilibrium

Collision Theory

- Reactant particles must collide for a chemical reaction to occur.
- Only "**effective collisions**", those with proper energy and orientation, result in a reaction.

Reaction Rate

The rate of a reaction increases with an increase in effective collisions .

Ex. List the factors that would increase the rate of a reaction.

1. Increase the temperature (Temperature increases the number of collisions and effectiveness)
2. Increase the concentration (increases the number of collisions)
3. Increase the surface area (increases the number of collisions)
4. Increase the pressure (works for gases only)
5. **Add** a catalyst

Catalyst

A catalyst increases the reaction rate by lowering the activation energy.

Equilibrium **(Equal rates, constant amounts)**

- Equilibrium can exist for reversible chemical and physical changes.
- The **rate** of the forward reaction **equals the rate** of the reverse reaction.
- The **concentrations** of reactants and products remain **constant.**

Le Chatelier's Principle

When a stress is applied to a system at equilibrium, the equilibrium will shift toward the direction that relieves the stress.

Example: Given the following equation: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) + \text{heat}$

Stress
Change in Concentration

Direction of Equilibrium Shift
Shifts away from the side of the equation increased or toward the side decreased

Examples for concentration:

Change in concentration	Direction of Shift
Increase Nitrogen gas (N_2)	Right
Increase hydrogen gas (H_2)	Right
Increase ammonia gas (NH_3)	Left
Decrease nitrogen gas	Left
Decrease hydrogen gas	Left
Decrease ammonia gas	Right

Stress
Change in Temperature

Direction of Equilibrium Shift
An increase in temperature shifts the reaction in the endothermic direction. A decrease in temperature shifts the reaction in the exothermic reaction.

Change in temperature	Direction of Shift
Increase temperature	Left
Decrease temperature	Right

Stress
Change in Pressure

Direction of Equilibrium Shift
An increase in pressure shifts the reaction toward the side of the equation with fewer moles of gas. A decrease in pressure shifts the reaction toward the side of the equation with more moles of gas.

Change in pressure	Direction of Shift
Increase pressure	Right
Decrease pressure	Left

Adding a catalyst will increase the rates of both the forward and reverse reactions equally but have no effect on equilibrium. There is no shift when a catalyst is applied.

Enthalpy (H)

Enthalpy is the heat content of a substance. (It is measured in joules or kilojoules.)

The change in enthalpy is called **heat of reaction (ΔH).**

Look on Reference Table I for the ΔH of specific reactions.

A minus sign indicates an exothermic reaction: $\Delta H = (-)$

A plus sign indicates an endothermic reaction: $\Delta H = (+)$

Ex. Combustion reactions are exothermic; they release energy to the surroundings.

Entropy (S)

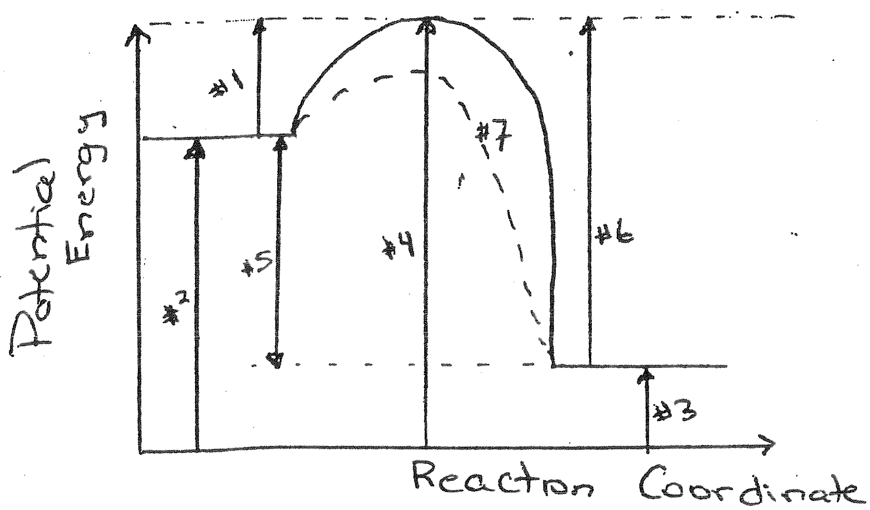
- Entropy is a measure of the randomness or disorder of a system (greater disorder, greater entropy).
- **Entropy increases as you go from solid → liquid → gas. (ΔS is positive)**
- **Entropy increases for decomposition reactions. (ΔS is positive)**

Note: Systems in nature tend to undergo changes toward <u>lower enthalpy and higher entropy</u>.

Potential Energy Diagrams

These diagrams show the changes in potential energy as a reaction proceeds.

Example: Draw a potential energy diagram for an exothermic reaction. Indicate the activation energy(#1), P.E. of the reactants(#2), P.E. of the products(#3), P.E. of the activated complex(#4), heat of reaction (ΔH)(#5), activation of the reverse reaction(#6). Also indicate what happen if a catalyst is added to the reaction(#7).



High to low.... exo.

Unit 7: Organic Chemistry

Properties of Organic Compounds

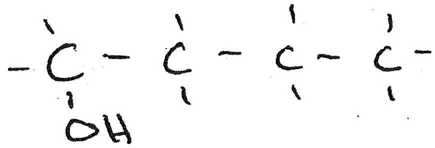
- have low melting points
- generally nonpolar
- generally insoluble in water
- generally nonelectrolytes (only organic acids can conduct electricity in solution)
- react slowly (molecular compounds react slower than ionic compounds)

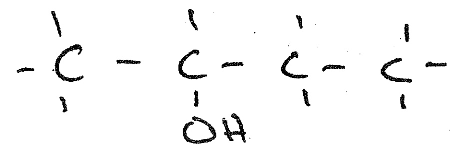
Isomers

- Isomers have the same molecular formula but different structures and properties.
- The greater the number of carbons the greater the number of possible isomers.
- Hydrocarbons must have at least 4 carbon atoms in order to exist as isomers.

Examples:

1) 1-butanol and 2-butanol

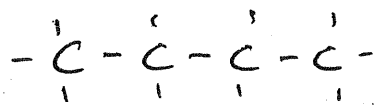




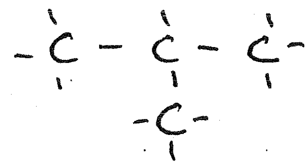
2) Butane and methyl propane

important to know

Butane



methylpropane



Homologous Series (Families) of Hydrocarbons

Hydrocarbons contain hydrogen and carbon only.
Each member differs from the preceding member by one carbon atom and two hydrogen atoms.

Alkanes
- -general formula is C_nH_{2n+2}
- -saturated hydrocarbons (contains all single bonds between C atoms)
- -methane (CH_4) (first member of the alkane series)
- -ethane (C_2H_6) (second member)
- -propane (C_3H_8) (third member)
- -butane (C_4H_{10}) (two isomers - see above)
- -pentane (C_5H_{12}) (three isomers)

Alkenes
- -general formula is C_nH_{2n}
- -unsaturated hydrocarbons (contains one double bond [two pairs of electrons are shared] between C atoms)
- -ethene (C_2H_4) (first member of the alkene series)
- -butene(C_4H_8) (two isomers) 1-butene and 2-butene

Alkynes
- -general formula is C_nH_{2n-2}
- -unsaturated hydrocarbons (contains one triple bond [three pairs of electrons are shared] between C atoms)
- -ethyne (C_2H_2) (first member of the alkyne series)
- -butyne(C_4H_6) (two isomers) 1-butyne and 2-butyne

Remember when naming or drawing organic compounds

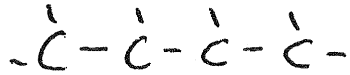
-refer to Reference Tables P and Q

-side group names

CH_3- is methyl

C_2H_5- is ethyl

- prefixes are used to indicate the number of side groups
- di is 2, tri is 3, tetra is 4, and penta is 5
- example: 2,3-dichlorobutane



- **Functional Groups**

A functional group is an atom or groups of atoms that replace hydrogen in a hydrocarbon. Functional groups give the compound distinctive physical and chemical properties.

Remember when classifying, naming or drawing organic compounds with functional groups refer to Reference Table R. Listed in this table are functional groups, general formulas, and examples.

Halides

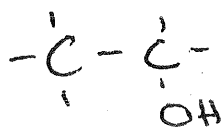
-functional groups are
- -F (fluoro)
- -Cl (chloro)
- -Br (bromo)
- -I (iodo)

-the alkane chain is numbered to show the location of the halogen

Alcohol

-functional group is -OH (OH is called hydroxyl)
-nonelectrolytes (CH_3OH looks like a base but it is an alcohol)
-polar molecules
-soluble in water
-name compound by replacing the final *-e* from the corresponding alkane name to *-ol*
-the alkane chain is numbered to show the location of the alcohol group

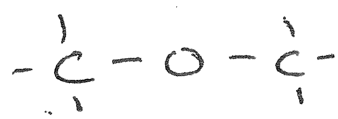
Examples: Ethanol



Ether

-functional group is -O-
-name compound by naming groups on either side of oxygen and ending with *ether*

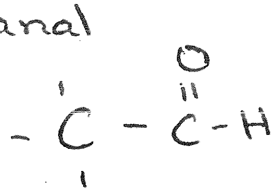
Examples: dimethy ether



Aldehyde

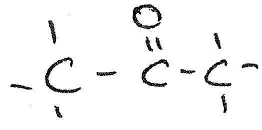
-functional group is -CHO
-name compound by replacing the final *-e* from the corresponding alkane name to *-al*

Examples: Ethanal



Ketone

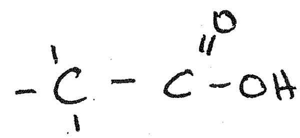
-functional group is -CO-
-name compound by replacing the final *-e* from the corresponding alkane name to *-one*
-the alkane chain is numbered to show the location of the ketone group
Examples: Propanone



Organic Acid

-functional group is -COOH (carboxyl)
-weak electrolytes (acetic acid is CH_3COOH- listed on table K)
-soluble in water
-turn litmus red and phenolphthalein colorless
-name compound by replacing the final *-e* from the corresponding alkane name to *-oic acid*

Examples: Ethanoic acid

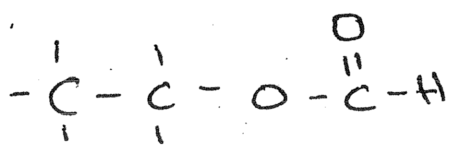


Ester

-functional group is -COO-
- Name the carbon chain without the carbonyl (C=O) first. This part of the name uses the appropriate prefix from table P and ends in the letters "yl".

Name the carbon chain with the carbonyl (C=O) second. This part of the name will end in the letters "oate".

Examples: ethyl methanoate

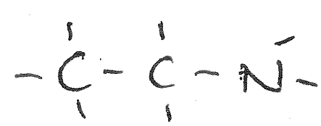


Amine

-functional group is -N-
-formed when one or more hydrogen atoms of ammonia are replaced by an alkyl group such as methyl or ethyl
-name the compound by replacing the final *-e* from the corresponding alkane name to *-amine*
-the alkane chain is numbered to show the location of the amine group

Examples: Ethyl amine (Ethanamine)

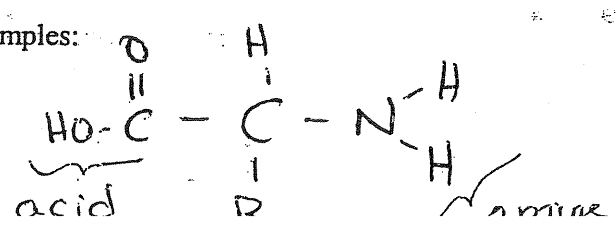
C



Amino Acid

-contains both a carboxylic group (-COOH) and an amine group (-N-)
-the amine group is attached to the C atom that is adjacent to the COOH
-amino acids are the building blocks of proteins

Examples:

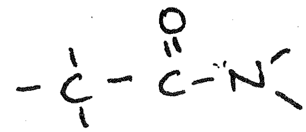


Amide -functional group is -CO-NH

-also called a peptide

-formed by the combination of two amino acids by a condensation reaction

Examples: Ethanamide



Organic Reactions

-generally occur more slowly than inorganic reactions

-be able to identify reaction type and/or determine a missing reactant or product in a balanced equation

Combustion

-almost all organic compounds will burn

-when sufficient oxygen is present products will be carbon dioxide and water

Ex. $CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$ (complete combustion)

-when the supply of oxygen is limited, carbon monoxide may be produced instead of carbon dioxide

Ex. $2CH_4 + 3O_2 \rightarrow 4H_2O + 2CO$ (incomplete combustion)

Substitution

-occurs with saturated hydrocarbons (alkanes)

-two reactants yield two products

Ex. $C_2H_6 + Br_2 \rightarrow C_2H_5Br + HBr$

Addition

-occurs with unsaturated hydrocarbons (alkenes, alkynes)

-double or triple bonds become a single bond

-two reactants yield one product

Ex. $C_2H_4 + Cl_2 \rightarrow C_2H_4Cl_2$

$C_2H_2Cl_2 + 2\,Cl_2 \rightarrow C_2H_2Cl_4$

$C_2H_4 + H_2 \rightarrow C_2H_6$

Esterification

-the reaction is; alcohol + acid → ester + water

-name ester by using the alkyl name of the alcohol followed by the acid group modified to end in -oate

Ex. $C_2H_5OH + HCOOH \rightarrow HCOOC_2H_5 + H_2O$

ethanol + methanoic acid → ethyl methanoate + water

Saponification

- in general, the reaction is ester + base → alcohol + soap

-one of the most common reactions is

Ex. $C_3H_5(C_{17}H_{35}COO)_3 + 3NaOH \rightarrow C_3H_5(OH)_3 + 3C_{17}H_{35}COONa$

fat (ester) + base → glycerol (alcohol) + soap

Fermentation

sugar → alcohol + CO_2

yeast is needed to secrete the enzyme zymase

Ex. $C_6H_{12}O_6 \rightarrow 2\,C_2H_5OH + 2\,CO_2$

Polymerization
This is the formation of large molecules from small molecules (monomers).
Naturally occurring polymers include proteins, starches, cellulose.
Synthetic polymers include plastics such as nylon, rayon, polyethylene.

Ex. $n(C_2H_4) \rightarrow (-C_2H_4-)n$

Unit 8: Oxidation - Reduction

Review handout on assigning oxidation numbers.

REDOX (Oxidation and Reduction) "Leo the Germ" or "Oil Rig"

- The loss of electrons is oxidation.
- The oxidation number increases for an atom that gets oxidized.
- The gain of electrons is reduction.
- The oxidation number goes down for an atom that gets reduced.

Oxidizing agent
This is the species that gets reduced.

Reducing Agent
This is the species that gets oxidized.

Remember
The answer to "What is reduced, oxidized, the oxidizing agent, or the reducing agent?" will always be found on the left side of the equation.

Ex. Write the half reactions for the reaction below. What is reduced, oxidized, the oxidizing agent, and the reducing agent?

$$2Ag^{1+}(aq) + Ni(s) \rightarrow 2Ag + Ni^{2+}(aq)$$

Reference Table J

Use Reference Table J to predict whether or not a single replacement reaction will take place
On the left side of table J:

- A metal listed on the table will react with the compound of a metal that is below it.
- All metals listed above H_2 will react with acids to release H_2 gas and product a salt.

On the right side of table J:

- A nonmetal will replace a less active nonmetal in a compound that is below it.

Example: Will the following reactions work?

The element without the charge must be higher than the ion in order for a react to occur.

$Ag + NaCl \rightarrow$ No reaction

$Zn + 2HBr \rightarrow ZnBr_2 + H_2(g)$

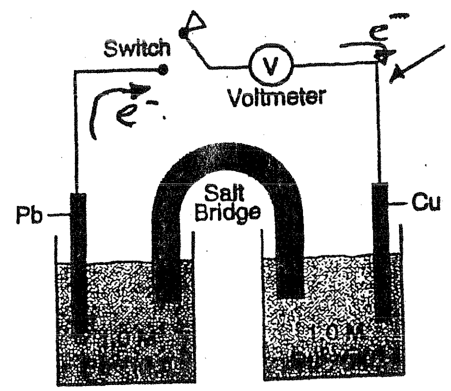
$Br_2 + 2KI \rightarrow 2KBr + I_2$

Electrochemical Cells

- There are two types; voltaic and electrolytic.
- Oxidation occurs at the anode. (An ox)
- Reduction occurs at the cathode. (Red cat)
- Electrons flow through the wire from anode to cathode

Voltaic Cell

- This is a spontaneous redox reaction that produces an electric current.
- There is a conversion of chemical energy to electrical energy.
- Oxidation occurs at the anode, reduction at the cathode. (An ox, red cat)
- Ions migrate across the salt bridge.
- The anode decreases in mass while the cathode increases in mass.



Pb(s) is the anode.

The anode will decrease in mass.

Cu(s) is the cathode

The cathode will increase in mass.

Pb is above Cu on Table J
Pb is the anode.

Electrolytic Cell

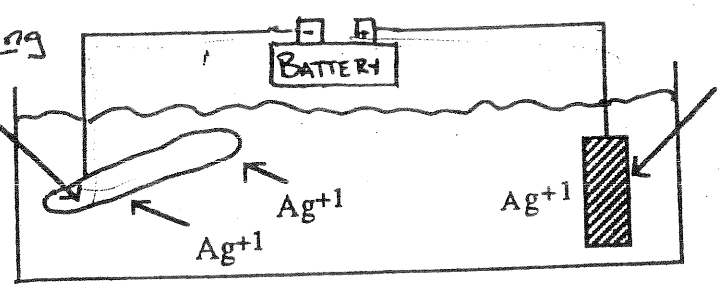
- Involves electrolysis (the use of electricity to cause a chemical reaction.)
- Non-spontaneous redox reactions.
- These cells require electrical energy to produce chemical change.
- The ions become neutral atoms.

* The cathode is negative in an electrolytic cell.

Electroplating

Object to be plated

Cathode
negative charge
$Ag^{+1} + 1e^- \rightarrow Ag$



Metal to be plated on the object.
Anode: Ag (s)
positive electrode

$Ag \rightarrow Ag^{+1} + 1e$

The positive ions are attracted to the negative electrode where they are reduced to a solid on the object to be plated. (The object to be plated is the cathode).

Unit 9: Acids, Bases, and Salts

Common acids and bases can be found on Reference Tables K and L.
Review notes on characteristic properties of acids and bases.

Arrhenius Acid/Base theory

- Acids yield H^+ or H_3O^+ (hydronium) as the only positive ion in aqueous solution.
- Bases yield OH^- (hydroxide) as the only negative ion in aqueous solution.

Bronsted-Lowry (also referred to as an "alternative theory")

- An acid is a proton donor (loses H^+).
- A base is a proton acceptor (gains H^+).

Neutralization

Acid + Base → Salt + Water

Titration is performed to find unknown concentrations of either an acid or base. In this method, a volume of solution of known concentration is used to determine the concentration of another solution.

For calculations: $M_A \times V_A = M_B \times V_B$

Example: 50.0 milliliters of hydrochloric acid is titrated with 25.0 mL of 0.40 M sodium hydroxide. Write the neutralization reaction and also calculate the concentration of hydrochloric acid.

Answer: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

$M_A \times V_A = M_B \times V_B$

$(50.0ml) \times V_A = (25.0ml) \times (0.40M)$ $V_A = 0.20$ M

Electrolytes

- substances that when dissolved in water forms a solution capable of conducting an electric current (ions are required)
- the higher the concentration of ions, the better the solution conducts electricity
- strong acids, strong bases, and soluble salts are all strong electrolytes

pH

- pH is a measure of the hydrogen ion concentration in an aqueous solution.
- each decrease of one unit of pH represents a tenfold increase in $[H^+]$
- acids have a pH less than 7, $[H^+]$ is greater than $[OH^-]$
- neutral substances have a pH of 7, $[H^+] = [OH^-]$
- bases have a pH greater than 7, $[H^+]$ is less than $[OH^-]$

Acid-Base Indicators

Indicators are substances that undergo a color change to show the relative level of acidity or alkalinity of a solution.
Some listed on Reference Table M

Example: What is known about the pH of a solution if an aqueous solution of the substance is yellow after bromcresol green is added?

The pH must be less than 3.8.

Unit 10: Nuclear Chemistry

Transmutation

- the changing of a nucleus of one element into that of a different element
- can occur naturally (alpha decay, beta decay, and positron emission) or be induced by the bombardment of the nucleus by high energy particles (artificial transmutation)
- stability of an isotope based on a neutron:proton ratio of about 1:1

Radioactive Isotope

Radioactive isotopes have unstable and spontaneously decay emitting radiation.
types of radiation include (see Reference Tables N and O):

	Ionizing Power	Penetrating Power
alpha decay	Best	Low
beta decay	Moderate	Moderate
gamma rays	Low	High

Sample alpha decay: $^{226}_{88}Ra \rightarrow \, ^{4}_{2}He + \, ^{222}_{86}Rn$

Sample beta emission: $^{60}_{27}Co \rightarrow \, ^{0}_{-1}e^{-} + \, ^{60}_{28}Ni$

Sample positron emission: $^{53}_{26}Fe \rightarrow \, ^{0}_{+1}e + \, ^{53}_{25}Mn$

Artificial Transmutation

- These reactions are induced by the bombardment of the nucleus of a stable atom by high-energy particles (protons, neutrons or alpha particles).
- These reactions have 2 reactants and 2 products.

Ex. Complete the following nuclear equation:

$^{27}_{13}Al + \, ^{4}_{2}He \rightarrow \, ^{1}_{0}n + \, ?$ Ans: $^{30}_{15}P$

Half life

- The half life is the time it takes for a radioactive sample to decay to one half its original mass.

  After one half life, 1/2 of the original sample remains.
  After two half lives, 1/4 of the original sample remains.
  After three half lives, 1/8 of the original sample remains.

- All half life values are listed on Reference Table N.
- The half life is not affected by temperature or pressure, it is constant.

Nuclear Energy

- In a nuclear reaction there is a conversion of mass to energy.
- The energy released in nuclear reactions is much greater than the energy released during chemical reactions

Risks with using radioactivity and radioactive isotopes

- biological exposure
- long-term storage and disposal
- nuclear accidents, storage of waste materials is a major problem

Beneficial uses of radioactivity and radioactive isotopes

- radioactive dating
- tracing chemical and biological processes
- industrial measurement
- nuclear power
- detection and treatment of diseases

Specific uses by some common radioisotopes

- C-14 to C-12 ratio in dating living organisms
- I-131 for diagnosing and treating thyroid disorders (ThIroid)
- U-238 to Pb-206 ratio in dating geological formations (U-238 is used to date)
- Co-60 in treating cancer

Fission

- Fission is the splitting a heavy nucleus into two lighter nuclei. (U-235, will divide)

Ex. $^{235}_{92}U + ^{1}_{0}n \rightarrow ^{141}_{56}Ba + ^{91}_{36}Kr + 3\,^{1}_{0}n + \text{energy}$

Fusion

- Fusion is the combining two small nuclei to produce one larger nucleus.
- Involves combining isotopes of hydrogen to form helium.

Ex. $^{2}_{1}H + ^{2}_{1}H \rightarrow ^{4}_{2}He$