Chemistry Regents Review: Matter and Energy

Chemistry = The study of **matter**, its structure and composition and the changes matter can undergo, as well as the **energy** changes that accompany a change in matter.

Matter = anything that has volume (takes up space) and has mass.

Volume = length x width x height or water displacement (drop object into graduated cylinder half filled with water, see how much additional space the object takes up).
Units: cm^3, milliliters, liters etc. Make sure you can convert ml-->liters etc.

Mass = weight (without the force of gravity, so it is a constant). Determined on triple-beam-balance (scale).
Units: grams, kilograms etc. Make sure you can convert between units.

All measurements must be made with precision and accuracy, following the rules of **Significant Figures**.

All matter can be divided into three phases: Solid, Liquid and Gas.
Solid -definite volume and shape, crystalline structure. The molecules are arranged in and organized pattern and although vibrating, remain in fixed positions.
Liquid - definite volume but takes the shape of its container. Intermolecular bonds are weaker than those of a solid.
Gas - no definite shape or volume. Will fill any container it's placed into because intermolecular bonds are so weak.

The movement of molecules, intermolecular attraction and therefore the phase of matter is determined by the temperature (Kinetic Energy -energy of motion) of the substance.
Units: degrees Celsius
Kelvin = $^0C + 273$
Heat, measured in Joules, is the measure of energy change in a chemical reaction. Heat flows from areas of high temperatures to areas of low temperatures.

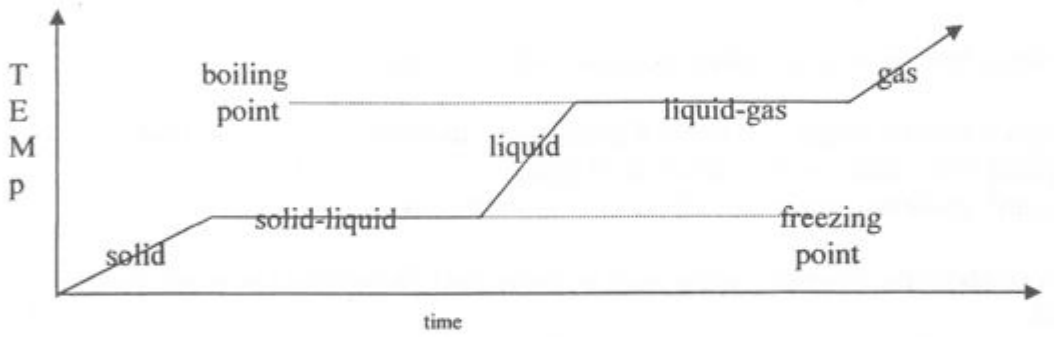
$$q = mc\Delta T$$

Joules = mass of water X specific heat of water (4.18 J/gram) X Δ Temperature ($temp_{final} - temp_{initial}$)

example: How many joules of heat are released when 25grams of water is cooled from 70^0C to 35^0C?

Heat of fusion (H_{FUS}) - amount of heat needed to change 1 gram of a substance from *solid phase to liquid phase*. **$q = mH_{FUS}$ (H_{FUS} & H_{VAP} can be found on Table B)**
Heat of vaporization (H_{VAP}) - amount of heat needed to change 1 gram of a substance from *liquid phase to gas phase*. **$q = mH_{VAP}$**
Reactions that absorb or use heat are ENDOTHERMIC.
Reactions that release or produce heat are EXOTHERMIC.

The Heating Curve of a substance is a graphic representation of the heat changes and accompanying phase changes of a substance as heat is added at a constant rate.



During phase changes there is no change in temperature or KE. The PE increases.

Phase changes

← Freezing ← ← Condensation ←

Solid → Fusion/ Melting → **Liquid** → Boiling/Vaporization → **Gas**

Sublimation

Gas - pressure, volume and temperature all have affects on the behavior of gases. Gases are measured at STP, standard temperature (0^0C or 273K) and pressure (1 atmosphere or 101.3Kpa).

The relationship between pressure and volume is inverse **PV=K** (Boyle's law)
The relationship between volume and Kelvin temperature is direct **V/T=K** (Charles' law)
The combined gas law is $\frac{P_1xV_1}{T_1} = \frac{P_2xV_2}{T_2}$

Ideal gas vs. Real gas - the gases we study are real gases, ideal gases do not actually exist. To be classified an ideal gas the molecules of gas must have no volume and no forces of attraction for each other. This does not actually occur. Real gases that are most like ideal gases are those that have small mass and volume (ex: Hydrogen, Helium). Conditions that would make a real gas behave more like an ideal gas are high temperature and low pressure (under those conditions the molecules will have the least intermolecular attraction like an ideal gas).

Liquids - The pressure of liquid molecules trying to escape the container in the form of a gas is called Vapor Pressure. The higher the temperature of the liquid, the greater the vapor pressure. When the vapor pressure equals atmospheric pressure, the liquid boils. The substance with strongest intermolecular bonds needs the most energy to escape as a gas and therefore have the highest boiling point. **Reference Table H** converts liquid temperatures to vapor pressures.

Properties of water

Normal freezing/melting point = 0°C
Normal boiling/condensation point = 100°C

[normal = at 101.3 kPa atmospheric pressure]

Chemistry Regents Review: Atomic Structure / Periodic Table

Atom - the smallest unit of matter

The atom, it was determined after many experiments (most notably Rutherford's gold foil experiment), is made up **of mostly empty space**. There is a small dense **nucleus** containing **protons** and **neutrons**, surrounded by rings or shells of **electrons**.

Protons - are positively charged particles found in the nucleus (nucleons) that have a mass of 1 amu (atomic mass unit).

Neutrons -are neutral particles (no charge) found in the nucleus (nucleons) that have a mass of 1 amu.

Electrons - are negatively charged particles found in rings around the nucleus. They have almost no mass. They are held on the atom by the + (nucleus) / - (electron) attraction. The exact arrangement of these electrons around the nucleus is called the **electron configuration**.

The number of protons, neutrons and electrons of any atom can be determined from the Periodic Table. The atomic number represents the number of protons. The mass number is equal to the number of protons plus neutrons. (The number of neutrons is = mass number - atomic number) The number of electrons is equal to the number of protons in an unreacted (neutral) atom.

The electron configuration of an atom tells how many electrons are in each shell (principal energy level) of the atom. This information is found at the bottom of each element's box on the periodic table and can be used to draw a Bohr Model.

When electrons are placed in the proper configuration, in the lowest energy levels possible, the atom is in the ground state. If the atom is heated the electrons can jump to higher energy levels than they should normally be in. The atom is then in the excited state. When the electron loses its energy, or relaxes, it gives off this energy in the form of a light spectrum (colored light).

The **valence electrons** are those electrons in the outermost ring. An electron dot diagram can be constructed to show only the valence electrons.

The organization of the **Periodic Table** is based on the reactivity of the different elements. Each element differs from another in the number of protons and electrons found in the atoms of the element. The number of neutrons for any element is not constant as there are heavier and lighter forms of the same element called **isotopes**.

The periodic table is arranged by organizing elements into groups and periods.
Groups - vertical (up/down) families of elements that have similar properties.
Periods - horizontal rows of elements that show trends as you move across the table.

The periodic table is divided into two types of elements (substances that cannot be decomposed) **metals** (left side of staircase) and **non-metals** (right side of stairs). Those elements sitting on the staircase are called **metalloids** or semi-metals. Those elements in groups 3→12 are called **transition metals**. Elements in group 1 all have 1 valence electron and are called the **alkali metals**. These elements are the most reactive metals and react by losing this electron and forming a +1 ion. Elements in group 2 are called **alkaline earth metals**. All have 2 valence electrons which are lost when the elements react forming +2 ions. The last group of atoms, group 18, is called the **noble gases**. They are generally unreactive since their valence shell of electrons is full. Group 17, the **halogens**, each have 7 valence electrons. Each needs one electron to completely fill their outermost ring (octet) and will therefore gain an electron when it reacts, forming a -1 ion. They are the most reactive nonmetals. The halogens are the only group with solids (I, At), liquids (Br), and gases (F, Cl).

Electronegativity - how strong an element attracts the electrons in a bond
scale: 0 → 4 0=weakest 4=strongest

Ionization Energy (IE) - How much energy is necessary to pull off a valence electron?
If the electron is closer to the (positive) nucleus it will be held tighter than an electron further away from the nucleus. Loosely = low IE / Tightly = high IE

Atomic Radius - how big is the atom?
As you go down a group or family, the number of shells of electrons increases - increasing the radius. As you move right across a period the number of shells is the same but the number of protons in the nucleus increases so the radius decreases.

Ionic Radius - the size of an atom once it has become an ion (gained or lost electrons).
An ion is bigger than the original atom if it is a negative ion (anion). The ionic radius is smaller than the original atom if it is a positive ion (cation).

The trends of the periodic table allow opposite corners to have opposite properties. The **lower left** corner of the periodic table (**Fr**) has the lowest electronegativity, lowest ionization energy, biggest atomic radius, most metallic, and the most reactive metal. The **upper right** corner (**He**) has the highest ionization energy, smallest radius, and least metallic. **Fluorine** (**F**) has the highest electronegativity and is the most reactive nonmetal since helium does not react.

Chemistry Regents Review: Bonding

Intramolecular Bonds

These are bonds between 2 or more atoms forming a molecule. Molecules, unlike elements, can be decomposed.

Ionic bonds- bonds between 2 ions (charged atoms) that have transferred electrons. These atoms have charges as a result of losing electrons or gaining electrons. Metals tend to lose electrons and form positive ions, while non-metals tend to gain electrons and form negative ions (see sheet 2).

Positive and negative ions stick together forming an ionic bond- a very strong bond due to the +/- interaction.

Elements forming ionic bonds generally have an electronegativity difference (subtract) greater than 1.7. **A metal and a nonmetal always form an ionic bond.**

example: Na (.9) and Cl (3.2) form NaCl with a difference of 2.3.

Compound containing ionic bonds do not generally exist as discreet molecules, rather as **ionic crystals**. Ionic crystals have high melting points, dissolve to some extent in water (see table F) and are electrolytes (conduct electricity in water).

Covalent Bonds- unlike ionic bonds that require the gain or loss of electrons, covalent bonds are those bonds between atoms that *share a pair of electrons*. This occurs because one atom does not have the strength to remove a tightly held electron from another atom- instead they share.

A) **Polar covalent bonds**- bonds that involve sharing electrons but the sharing is **unequal.** The electron pair spends more time around one atom than the other. To determine which atom holds the electron pair more than the other, compare their electronegativities. The one with the higher electronegativity holds the electron pair more tightly than the other atom, causing the atom to be δ- (slightly negative), while the other atom is δ+ (slightly positive).

Polar covalent bonds have an electronegativity difference less than 1.7 but more than 0.

example: H--------Cl Electronegativity difference = 1.1 polar covalent bond

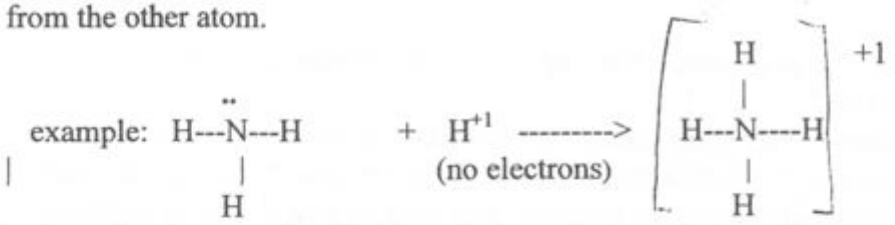
2.1 δ+ 3.2 δ-

B) **Non-polar covalent bond**- bonds between two atoms that share their electrons **evenly**. No atom has a higher electronegativity than another in the bond so no one atom holds the electrons tighter than the other. This type of bond exists in the diatomic elements- those molecules made up of 2 atoms *of the same element* bonded together. The only elements that can exist in the diatomic form are H, O, N, Cl, Br, I, F. These elements are found in nature as H_2, O_2, N_2, Cl_2, Br_2, I_2, F_2. Each molecule is held together with a non-polar covalent bond.

Non-polar covalent bonds have an electronegativity difference of 0 (the two atoms have the same electronegativity because they are identical).

C) **Coordinate covalent bond** - Usually a covalent bond represents a shared pair of electrons, one electron from each of the elements involved in the bond. In this type of

bond, both electrons in the bond come from one atom and none of the electrons come from the other atom.



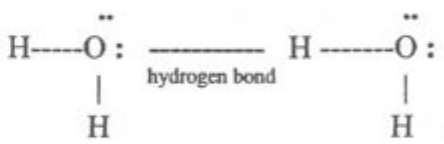
Intermolecular bonds- attractions between two or more molecules. The strength of the attraction will determine the physical properties of the substance. The stronger the attractions, the more energy needed to break them. The result is a higher melting point and boiling point.

Dipole - dipole interaction- If two molecules are dipoles (one side of the molecule is slightly positive and the other side is slightly negative) the +/- attraction between the sides of the molecules will hold the molecules together. This is a relatively strong bond.

example: H-----Cl ---- H-------Cl
δ+ -δ +δ –δ

Hydrogen bonds- An attraction in which the Hydrogen from one molecule attaches to the Fluorine, Oxygen or Nitrogen of another molecule.

Example:



Van der Waals forces- this weakest type of intermolecular attraction exists between 2 or more molecules that have no reason to stick together (no +/- interaction) but do so anyway.

a) This type of bond exists between molecules with non-polar covalent bonds (no +/- side).

example: F_2--------F_2 Van der Waals forces

b) This type of bond also exists between two or more molecules that each contain polar covalent bonds but the overall molecule is non-polar because *it does not have a + and - side to the whole molecule*- all of the bonds are arranged **symmetrically** as in $CH_{4(g)}$ and $CO_{2(g)}$.

Example:

When comparing the strengths of Van der Waals forces, and the resulting boiling points of the substances, **the larger the molecules involved, the stronger the attractions and the higher the boiling point.**

Molecular substances tend to have low melting points and be poor conductors of electricity.

Midwood High School Mrs. Goldberg

Chemistry Regents Review: Forms of Matter

All matter can be divided into two categories: substances and mixtures.
Substances are defined by their inability to be decomposed physically and by their uniform composition and properties (homogeneous). Substances can be divided into two categories: elements and compounds.

Elements -substances that cannot be decomposed into simpler substances by chemical change. The simplest form of matter, they are listed on the Periodic Table.

Compounds-substances that can be decomposed by chemical change because they are composed of two or more different elements chemically combined in a fixed ratio. As a result of the chemical combination of the elements, the properties of a compound are different from the properties of the elements that compose it.

Mixtures-composed of two or more substances (elements or compounds) physically combined. As a result of the physical combination of the substances, the properties of the mixture are intermediate (in the middle) of its components (some of each).
Mixtures can be heterogeneous (non uniformly combined) such as in soil or they can be homogeneous (uniformly combined) such as in a solution (i.e. salt water). Water solutions have the abbreviation (aq) for aqueous.

Chemical Formulas of Compounds (from their names)
Determining the formula (ratio of atoms in a molecule) from the name of a compound can be done by using the oxidation states (charges) of the elements in the 'criss-cross' method. -
examples: Mg^{+2} and Cl^{-1} Ca^{+2} and NO_3^{-1} (polyatomic ion)
$MgCl_2$ $Ca(NO_3)_2$
Reduce, if possible, to lowest terms to determine the empirical formula (simplest ratio) for all ionic compounds.

Naming Chemical Compounds (from their formulas)
Ionic compounds- **The Stock System**
a) Name the positive ion (generally the metal).
b) Name the negative ion (the non metal).
c) If the compound is a binary compound (composed of only two different elements), the suffix of the negative ion is changed to -ide.
If the compound contains three or more elements, use Table E for the suffix.
d) When naming compounds in which the metal ion has multiple oxidation states (more than one possible charge),the specific charge affects the formula of the compound and must be specified in the name as a Roman Numeral.

Examples: $Cu(NO_3)_2$ = copper (II) nitrate
$Fe_2(SO_4)_3$ - iron (III) sulfate

Molecular Substances (covalently bonded)- **The Prefix System**
a) Name the first element. If two or more atoms of the element are present in the compound, use a prefix (di-, tri-, tetra-, penta- etc.) to indicate the number.
b) Name the second non metal. Always use a prefix (including mono-) to indicate the number of atoms
c) If the compound is a binary compound (composed of only two different elements), the suffix of the negative ion is changed to -ide.

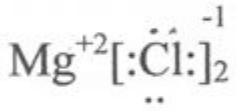
Examples: N_2O = dinitrogen monoxide
SF_6 = sulfur hexafluoride

Electron Dot Structures of Compounds

Ionic compounds

The structure must represent the *transfer* of electrons between the metal and non-metal.
The metal is written without its valence electrons and with a positive charge.
The non-metal is written with its valence electrons and those it has gained giving it an octet (generally 8).
The non-metal with its electrons is placed inside of brackets and its oxidation state must be written.
The two ions are drawn near each other, with subscripts to indicate the ratio of ions in the compound.

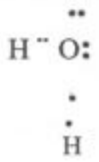
Example: Magnesium chloride



Molecular Substances

The structure must represent the *sharing* of electrons between the two non-metals.
Each atom must be drawn with all of its valence electrons (draw one on each side before doubling up). Any unpaired electrons can participate in a bond. A covalent bond between two atoms will be represented as a pair of electrons, one belonging to each atom, drawn sandwiched between the two atoms.

Example: water (dihydrogen monoxide)



Chemistry Regents Review: Mole Relationships

The Concept of a Mole

A MOLE is a quantity of items. This quantity is 6.02×10^{23}. (It may be helpful to substitute the number 12 and the word dozen into your questions until you are confident in this topic). The number 6.02×10^{23} only represents the number of molecules in a mole, it says nothing about the weight, since a molecule of one type has a different mass than one molecule of another type.

The MASS of one mole of any atom is the mass of one atom from the periodic table, but with the units of GRAMS instead of atomic mass units (amus). This mass of one mole is known as the gram formula mass.

The mass of one mole of a molecule is the total mass of all atoms in that molecule with the units in grams.

example: mass of one atom of sulfur [S] = 32 amu
mass of one mole of S = 32 grams
number of atoms in one mole of S $= 6.02 \times 10^{23}$
Mass of one mole of H_2SO_4 = 98 grams
1x2+ 32+16x4=98grams per mole

You can use the following formula to convert easily between grams and moles of any molecule.

$$\#\ \text{MOLES} = \frac{\text{GRAMS}}{\text{Gram formula mass (from periodic table)}}$$

Examples: Find the mass of .35 moles of H_2SO_4
How many moles of H_2SO_4 are contained in 250grams?
How many molecules of H_2SO_4 are present in .75 moles?
How many atoms re contains in 1.5 moles of H_2SO_4?

Percent Composition

If you can determine the molecular weight of any compound, then you can determine the mass contributed by each of the different atoms in the compound using the formula:

$$\frac{\text{Mass of part}}{\text{mass of compound}} \times 100$$

example: find the percent by mass of Oxygen in H_2SO_4.

Using the concept of percent composition, we can use the percentage mass of each element to determine the ratio of atoms (same as moles) in the formula of an unknown compound.

example: Find the empirical formula (ratio of smallest whole numbers) of a compound that consists of 58.80% Barium, 13.75% Sulfur, and 27.45% Oxygen by mass.

1) Assume you have 100 grams of the substance. This eliminates the % and replaces it with grams.

2) Convert the grams to moles using the formula above (divide each by its molecular weight).

3) Divide each of those moles by the smallest number to get a ratio of the smallest whole numbers. (If you get a number that cannot be rounded off to a whole number, multiply all numbers by 2 or 3 to get whole numbers).
4) To determine the molecular formula of the compound (actual formula) divide the gram formula mass by the mass of the empirical formula to determine the number of times to multiply the empirical formula.

Moles of Gas

Equal volumes of gas at the same temperature and pressure contain equal numbers of moles and equal numbers of molecules.

Balancing equations

When balancing equations, the number of each type of atom must be equal on both sides of the equation.

Rules: 1) List types and numbers of each atom on each side.
2) You may never add subscripts.
3) You may never add numbers in the middle of a molecule.
4) You may only add coefficients. These coefficients apply to the whole molecule.

Suggestions: 5) Balance Hydrogen last.
6) Balance Oxygen second to last. All other atoms in any order.

When an equation is balanced you can use its **coefficients** as a **ratio of molecules** or of **moles**, or in the case of gases, **volumes** as well. This will be helpful in predicting the moles or grams of product if given information about the reactants.

Example: Given the following balanced equation.....

$$2C_2H_{6\,(g)} + 7O_{2(g)} \text{ ----------> } 4CO_{2(g)} + 6H_2O_{(g)}$$

If 21 moles of oxygen react completely, how many moles of water will be formed?
If 18grams of ethane (C_2H_6) react completely, how many grams of carbon dioxide will be produced? REMEMBER TO CONVERT TO MOLES FIRST.
If 35 liters of ethane react completely, how many liters of carbon dioxide will be formed? YOU DO **NOT** NEED TO CONVERT TO MOLES FIRST.

Chemistry Regents Review: Solutions

Solution- a homogenous mixture of two substances, a solute (material being dissolved, can be solid, liquid or gas) and a solvent (substance in excess, solute dissolves in this). Solutions in which water is the solvent have the abbreviation (aq) for aqueous.

Solutions can be classified as *unsaturated* (the solvent can still dissolve more solute at a specific temperature), *saturated* (the solvent is exactly filled to capacity with the solute at a specific temperature) or *supersaturated* (the solvent is dissolving more solute than is actually expected at a given temperature). Using Reference Table G- Solubility Curves, any solution in water can be classified as saturated, unsaturated or supersaturated for any given temperature. Generally the solubility of solids and liquids increases with increasing temperature while the solubility of a gas will decrease.
Solutes dissolve well in like (similar) solvents. Ionic and polar molecules generally dissolve well in polar solvents (like water). Non-polar solutes do not generally dissolve well in water.
Use Reference Table F- Solubility Guidelines in Aqueous Solutions, to compare the solubilities of ionic compounds.

The exact concentration (strength) of any solution is a measure of the moles of solute dissolved in a liter of the solution.

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Liters of solution}}$$

(convert ml to L by moving the decimal 3 places left)
The moles of solute can easily be determined from the grams of solute and vice versa (table T).

The concentration of a dilute (weak) solution can be determined in parts per million (ppm).

$$\text{ppm} = \frac{\text{grams solute}}{\text{grams solution}} \times 1{,}000{,}000$$

[grams solute + grams solvent]

Once the water is part of a mixture it no longer has a normal boiling point of 100^0C and a normal freezing point of 0^0C. The boiling point will elevate (increase) for each mole of solute dissolved, while the freezing point will depress (decrease) for each mole of solute. A more concentrated solution will have a higher boiling point and lower freezing point than a more dilute solution.
Ionic solutes such as salts ionize (break apart into ions) when dissolved in water. To determine the BP and FP of these solutions take into account the number of ions as this will double or triple the concentration of particles thus having a greater affect on the boiling point and freezing point.

Chemistry Regents Review: Acid - Base

Acid

Arrhenius- Any substance that when added to water gives off a hydrogen ion (H^+).
Example: $HCl \rightarrow H^+ + Cl^-$

Base

Arrhenius- Any substance that when added to water gives off an hydroxide ion (OH^-)
Example: $NaOH \rightarrow Na^+ + OH^-$

Bronsted-Lowry - Proton (H^+) donor

Bronsted-Lowry - Proton (H^+) acceptor

example- $HCl + H_2O \rightarrow H_3O^+ + Cl^-$
acid base
(accepts H^+)

Acids and bases are both **electrolytes** since they are found as ions in water and as a result can conduct electricity.

Amphoteric or amphiprotic substances are those substances that can behave as either an acid or a base (according to Bronsted-Lowry) depending on the reaction.

Salts are ionic solids composed of a metal and a non-metal that are neither acids nor bases. They ionize in water to conduct electricity (electrolytes).
Neutral solutions are those solutions that contain equal amounts of H^+ (from acid) and OH^- (from base).

Neutralization Reaction
When an acid and base are mixed in equal amounts the results are always a salt and water.

example- $HCl + NaOH \rightarrow NaCl + H_2O$
acid + base → salt + water

Titration
Physically performing a neutralization reaction to determine the molarity or volume of an unknown solution from the molarity or volume of a known solution is called a titration. The calculations can always be done with the same formula:

$$Molarity_{H+\ ions} \times Volume_{acid} = Molarity_{OH-\ ions} \times Volume_{base}$$

Example: How many milliliters of 3M NaOH will exactly neutralize 75 milliliters of 6M HCl?

Note: Be careful of diprotic acids and dihydroxy bases. (1M H_2SO_4 = 2M H^+ ions)

Tables K and L can be used to find names and examples of some acids and bases.

The pH Scale

We can determine the pH of a solution if we know the hydrogen ion (H^+) concentration or the hydroxide ion (OH^-) concentration. The pH is the exponent or number of decimal places moved in the H^+ concentration.

Example: $[H^+] = 0.0001M$ pH = 4 (moved four decimal places)

$[H^+] = 1 \times 10^{-9}M$ pH = 9 (the exponent)

If we are given the OH^- concentration we can use the same method but we must subtract from 14 because we are really finding the pOH.

Example: $[OH^-] = 1 \times 10^{-12}$ pOH = 12 so pH = 14-12 = 2

Note: $[H^+] = [H_3O^+]$ ←→ hydronium ion = hydrogen ion

A pH of 7 indicates a neutral solution. Above 7 is a basic solution. Below 7 is an acidic solution.

Each number change on the pH scale represents a 10x change in concentration.

If a solution with a pH of 4 becomes 100 times more acidic the pH will be 2.

If a solution with a pH of 4 becomes 100 times more basic the pH will be 6.

Acid / Base Indicators

Examples of acid-base indicators are given on **Table M**. The color change is indicated showing the color the substance is when the sample is on either end of the pH range for color change.

Example: bromcresol green 3.8 - 5.4 yellow to blue

At a pH less than 3.8 bromcresol green will be yellow. At a pH greater than 5.4 bromcresol green will be blue. We can only say it changes at some point between 3.8 and 5.4 but not what color it would be at any specific pH. (Using more than one indicator we can narrow down ranges of possible pHs.)

Chemistry Regents Review: Kinetics

Kinetics- the branch of chemistry dealing with the *rate* (speed) and *mechanism* (step by step processes including energy changes) of chemical reactions.

Factors that Affect the Rate (speed) of a Chemical Reaction:
1) **Temperature** - increasing the temperature of any chemical reaction will speed up the reaction (increase the rate, decrease the time) by increasing the number of effective collisions between reacting molecules.

2) **Concentration** - increasing the concentration of reactants will speed up the reaction by increasing the number of effective collisions between reacting particles.

3) **Pressure/Volume** - changing the pressure (or volume) of the reaction chamber will affect the rate of reacting GASES. An increase in pressure or decrease in volume will cause reactions involving gas molecules to speed up by increasing the number of collisions between molecules. No effect on solids and liquids.

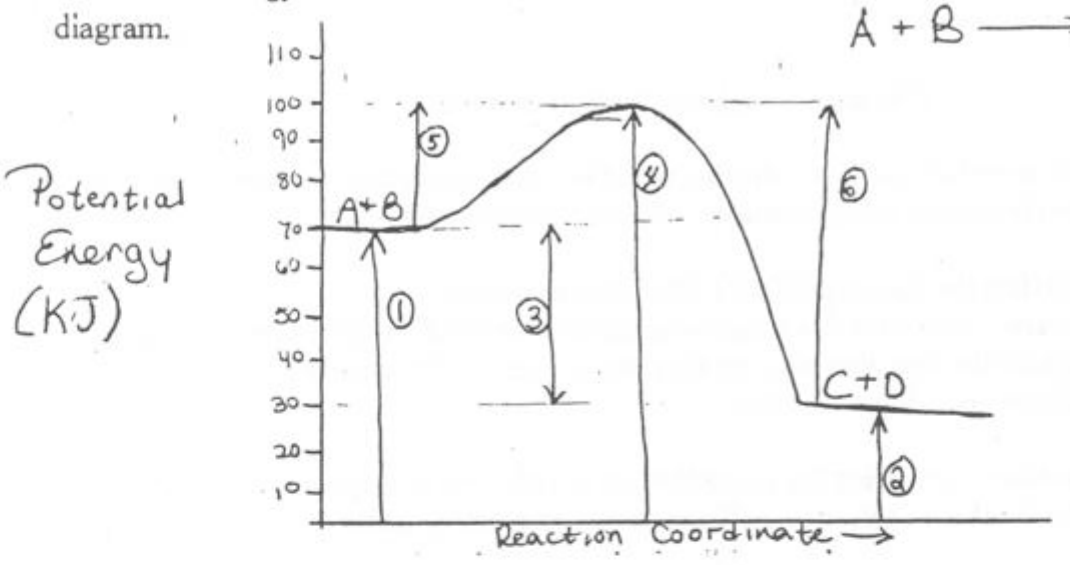
4) **Catalyst** - a substance, that when added to a chemical reaction, *speeds it up by lowering the activation energy* of that reaction. If a smaller 'energy boost' is needed the reaction will occur more quickly.

5) **Surface area** - the reactions that include reactants with a large surface area will react more quickly than those with a small surface area since there will be more collisions between molecules. example: dissolving a sugar cube vs. a teaspoon of sugar

6) **Nature of the reactants** - Some reactions proceed more quickly than others simply due to the chemical composition of the reactants. The number and strength of bonds broken and formed during the course of the reaction will determine the speed of the reaction. The reaction can only proceed as quickly as the slowest step of the reaction *(the rate determining step).*

The role of energy in a chemical reaction is best illustrated with a potential energy diagram.

$A + B \longrightarrow C + D + 40kJ$



Numbers below correspond to the arrows in the diagram above.

1- Potential energy (heat) of reactants
2- Potential energy (heat) of products
3- Enthalpy = Heat of reaction = ΔH

$Heat_{products} - Heat_{reactants}$

+ ΔH indicates an endothermic reaction
- ΔH indicates an exothermic reaction

4- Energy of the activated complex
5- Activation energy
6- Activation energy of the reverse reaction

Reference Table I lists several chemical reactions and their ΔH values. You can interpret which reactions are endothermic and which are exothermic.
Heat, measured in joules or kilojoules is included (as an absolute value) in a chemical reaction as a reactant in an endothermic reaction and as a product of an exothermic reaction.

Many chemical reactions are reversible. If the forward and reverse reactions are proceeding at the same rate (speed) the reaction system is said to be at **equilibrium**.

Entropy is a measure of the chaos or randomness of the molecules of a substance.
Solids have the least entropy followed by liquids and gases have the most entropy. By examining the phases of the products compared to the reactants in any reaction, you can determine if there was an increase or a decrease in entropy in the reaction.

A reaction will occur **spontaneously** (without outside help) if it is *both* exothermic (decreasing enthalpy) and increasing in entropy (solid→liquid→gas).
A reaction will not occur spontaneously if it has increasing enthalpy (endothermic) and decreasing entropy (gas→ liquid→ solid)

Chemistry Regents Review: Equilibrium

Equilibrium - the condition in which the forward and reverse reactions proceed at the same rate producing a constant ratio of concentrations of all products and reactants. The most important prerequisite for chemical equilibrium is a **closed system**, a system to which nothing is added (examples- heat, reactants etc.) and nothing is removed (examples- escaping gas, precipitating solid etc.). A saturated solution is said to be at equilibrium, the dissolving and crystallization of molecules (the forward and reverse reactions) are proceeding at the same rate.

Shifting Equilibrium

Since we know that temperature and pressure changes, catalysts and concentration of reactants will affect the rate of a chemical reaction, we can determine their affects on a system at equilibrium. Once we open the system and alter the conditions, the equilibrium system will no longer exist- one reaction (the forward or reverse) will proceed at a faster rate than the other. We will attempt to predict which way the equilibrium will shift (which reaction will go faster) as well as how this shift will affect the concentrations of reactants and products.

example: $N_{2\,(g)} + 3H_{2\,(g)} \text{<----------> } 2NH_{3\,(g)} + \text{Heat}$

a) Increasing the **temperature** will speed up any chemical reaction. However, the increase in temperature will have a *greater impact on an endothermic reaction* (one that absorbs heat) than on an exothermic reaction (one that releases heat). Decreasing the temperature has an opposite affect. Since any reaction that is endothermic in one direction is exothermic in the reverse (and vice versa), we can determine which reaction will be proceeding at a faster rate, the one proceeding to the right (shift right) or the one proceeding to the left (shift left). Whichever reaction proceeds faster will produce more products and as a result, fewer reactants for that reaction will remain.

example- in the reaction above, the equilibrium will shift to the left to speed up the endothermic reaction (the one that absorbs heat as a reactant) thereby increasing the concentrations of N_2 and H_2 and decreasing the concentration of NH_3.

b) Increasing the **concentration** of a reactant in a chemical reaction will speed up the reaction. However, adding a reactant to a system at equilibrium is also adding a product to one of the reactions- which has the opposite affect. To determine the direction of equilibrium shift, determine to which reaction (forward or reverse) this substance added is a reactant. That reaction will proceed at a faster rate. The concentration of the products of that reaction will increase while the concentration of the reactants of that reaction will decrease. Decreasing the concentration of a reactant has the opposite affect as the reaction attempts to compensate by increasing the rate of the reaction that produces more of the missing reactant.

example- in the system above, if the concentration of N_2 is increased, the equilibrium will shift to the right to use up the extra N_2. This will increase the concentration of NH_3 and decrease the concentration of H_2 as it is used up by this reaction.

example- decrease the $[H_2]$ and the equilibrium shifts left to make more.

c) Increasing the **pressure** (decreasing the volume) will speed up reactions of gas molecules. However, since there are two reactions proceeding at the same time, we must predict which of the two reactions will proceed at a faster rate relative to the other. Increasing the pressure will speed up (favor) the reaction that produces *fewer moles of gas*. Decreasing the pressure will favor the reaction producing more moles of gas. If the two reactions produce equal volumes of gas (count up coefficients of all gas reactants) or no moles of gas, then a change in pressure will have no affect on the system.

example- increasing the pressure on the system above will the equilibrium to the right as the reaction proceeding right produces 2 moles of gas while the reaction proceeding left produces 4 moles of gas.

d) Adding a **catalyst** will speed up a reaction by *lowering the activation energy* of the reaction. However, since in a system at equilibrium there are two reactions proceeding at the same time, we must determine which of the two reactions will proceed faster than the other as a result of the catalyst. Adding a catalyst to the system will lower the activation energy of both reactions, *speeding up both reactions by the same amount* (see potential energy diagram). As a result, both reactions are going faster, but at the same speed relative to each other. *There is no shift in equilibrium.*

Chemistry Regents Review: Redox

REDOX

REDOX is an acronym for *reduction* and *oxidation*. **Reduction** is the *gaining of electrons*, negative charges. This has the net effect of causing the ion to become *more negative*. **Oxidation** is *the loss of electrons* which has the net effect of causing the ion to become *more positive*.

Helpful reminder: LEO the lion says GER- losing electrons oxidation
gaining electrons reduction

In order to determine if a chemical reaction is a redox reaction we must determine the oxidation numbers (charges) of each ion on the reactant side of the equation and each ion on the product side. If any oxidation state has changed, then reduction and oxidation have occurred. Single replacement reactions are **always** REDOX.

Rules for determining oxidation states

1. **The total charge on any molecule** (when adding all positive and negative charges) must **be equal to 0** (unless the charge is given, polyatomic ions, in which case the total charge of all elements must equal the given charge.)
2. Any element in its **natural state**- alone, uncombined, has a charge of **0**. (The number of protons and electrons are equal). Exceptions: Diatomic elements H, O, N, Cl, Br, I and F are found in pairs in their natural state. Therefore in the diatomic form they have no charge but individually or combined as part of a molecule they do have charges.
3. All **alkali metals** (group IA) lose 1 electron when they react to form part of a molecule. As a result, alkali metal ions have an oxidation state of **+1**.
4. All **alkaline earth metals** (group IIA) lose 2 electrons when they react to form part of a molecule. As a result, alkaline earth metal ions have an oxidation state of **+2**.
5. **Fluorine**, when mono-atomic or when combined as part of a molecule has an oxidation state of **-1** due to gaining 1 electron.
6. **Oxygen**, when mono-atomic or combined as part of a molecule has an oxidation state of **-2** due to gaining 2 electrons. Exceptions: when combined with fluorine in $\mathbf{OF_2}$ oxygen has a charge of **+2,** when in peroxide such as $\mathbf{H_2O_2}$ oxygen has a charge of **-1.**

Example: $Zn(s) + CuS(aq) \dashrightarrow ZnS(aq) + Cu(s)$

oxidation states: 0 +2 -2 +2 -2 0

In this case $Zn(s)$ is oxidized to Zn^{+2}

Cu^{+2} is reduced to $Cu(s)$

Oxidizing and Reducing Agents

An **oxidizing agent** is something that causes oxidation, the loss of electrons to occur. **The species that causes oxidation is reduced- the one that gains electrons.**

A **reducing agent** is something that causes reduction, the gaining of electrons. **The species that causes reduction is oxidized- the one that loses electrons.**

Helpful hint: *The species reduced is the same as the oxidizing agent. The species oxidized is the same as the reducing agent.* **All must be *reactants* in the chemical equation.**

Half Reactions

You should be able to identify the oxidizing and reducing half-reactions for any given REDOX reaction. There are two half-reactions one for the oxidation that occurs and the other for the reduction. Half-reactions must be balanced with respect to mass and charge and can be used to balance a simple net REDOX reaction.

Example: $Li + NaCl \rightarrow Na + LiCl$

Oxidation half reaction is $Li \rightarrow Li^{+} + 1e^{-}$

Reduction half reaction is $Na^{+} + 1e^{-} \rightarrow Na$

$\_\_2\_\_ Fe^{+3} + \_\_1\_\_ Ca \rightarrow \_\_2\_\_ Fe^{+2} + \_\_1\_\_ Ca^{+2}$

Half-reactions	Balanced
$Fe^{+3} + 1e^{-} \rightarrow Fe^{+2}$	$2Fe^{+3} + 2e^{-} \rightarrow 2Fe^{+2}$
$Ca \rightarrow Ca^{+2} + 2e^{-}$	$1Ca \rightarrow 1Ca^{+2} + 2e^{-}$

Table J - Activity Series

The activity series can be used to determine if a single replacement reaction will occur spontaneously. The metals that are higher up on the table are more active meaning they are more likely to be oxidized. This implies that they will replace any metal ion below them from a solution. Any metal above H_2 will react with acids to produce hydrogen gas.

Examples:

$K + NaCl \rightarrow$ will occur spontaneously because K is above Na on Table J.

$Zn + CaO \rightarrow$ will not occur spontaneously because Zn is below Ca on Table J.

Al will react with HCl to produce hydrogen gas whereas Ag will not.

Chemistry Regents Review: Electrochemistry

The use of a chemical reaction, specifically a redox reaction, to harness electrical energy (electricity) is called electrochemistry.

Any spontaneous REDOX reaction will have a loss of electrons (oxidation) and the gaining of electrons (reduction) without any external help. If these transferred electrons, from the two half reactions, travel through wires we can harness their energy. The cell (battery- power source) produced is called an electrochemical cell.

Electrochemical Cell (a.k.a. voltaic cell or battery)
This cell is based on having two separate half reactions with the transferred electrons flowing through wires between the two half reactions.
The parts of the electrochemical cell are as follows:
1) The reduction side containing a) the solid metal cathode which is positively charged [RED CAT - reduction at cathode] [posi-cat = positive cathode] b) the metal's ions in solution
2) The oxidation side containing a) the solid metal anode which is negatively charged [AN OX - anode oxidation] b) the metal's ions in solution
3) A wire (external conductor) to connect the two electrodes, the anode and the cathode. A voltmeter can be inserted to measure the voltage.
4) The salt bridge which connects the two metal ion solutions. Its function is to balance the ionic charges of the two solutions. As one solution becomes more + (the oxidation side) - ions in the salt bridge migrate toward that side. As the other solution becomes more - (reduction side) the + ions migrate in that direction.
The electrons migrate through the wires from the oxidation side to the reduction side.

In an electrochemical cell, as the reaction proceeds, the solid metal *anode* (higher metal on chart J) *gets smaller* as it breaks into + ions and electrons, making the solution more +. These electrons flow through the wire and the cathode to the + ions in solution. These ions are reduced (ions of lower metal on table J) by gaining these electrons and forming more of the solid metal *cathode* which *gets larger*. The solution, by losing its + ions becomes more negative forcing the flow of + ions in the salt bridge.

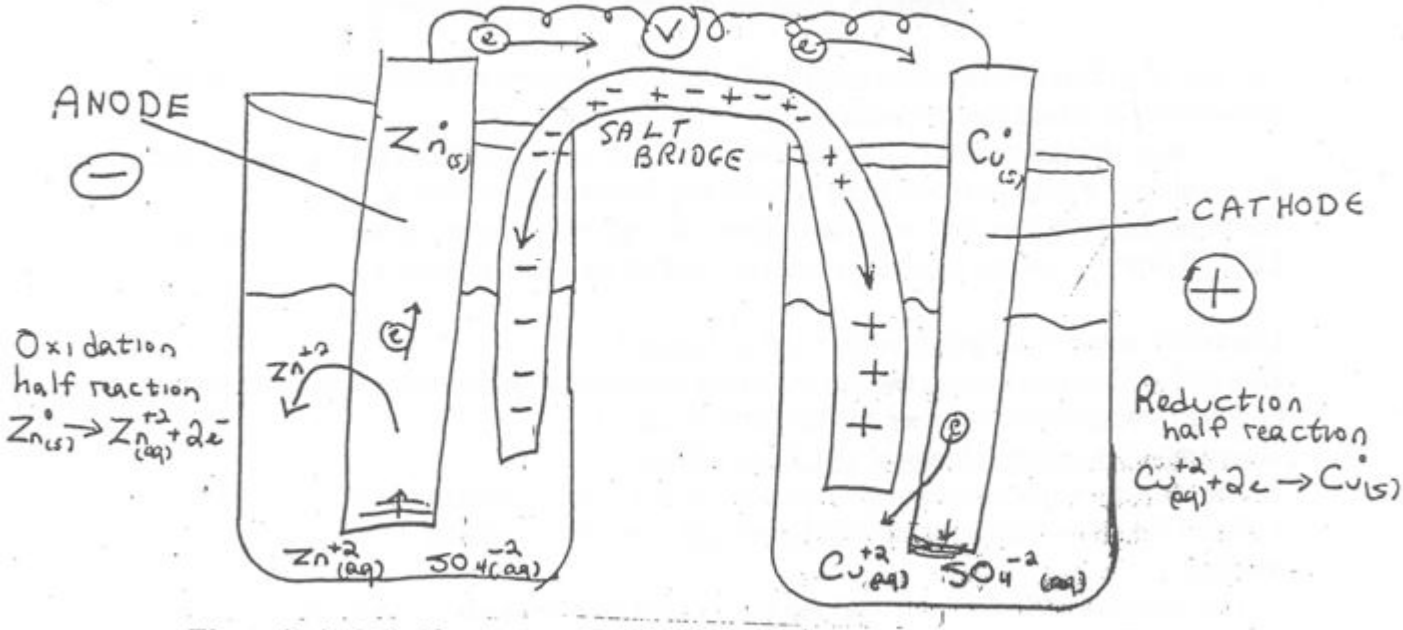
Example: Form an electrochemical using Cu / Cu^{+2} and Zn / Zn^{+2}
[Cu reaction is reduction because it is higher on table N}

$$Zn_{(s)} + CuSO_{4(aq)} \text{ ----------> } ZnSO_{4(aq)} + Cu_{(s)}$$
$$0 \quad +2\ -2 \quad +2\ -2 \quad 0$$

reduction

oxidation

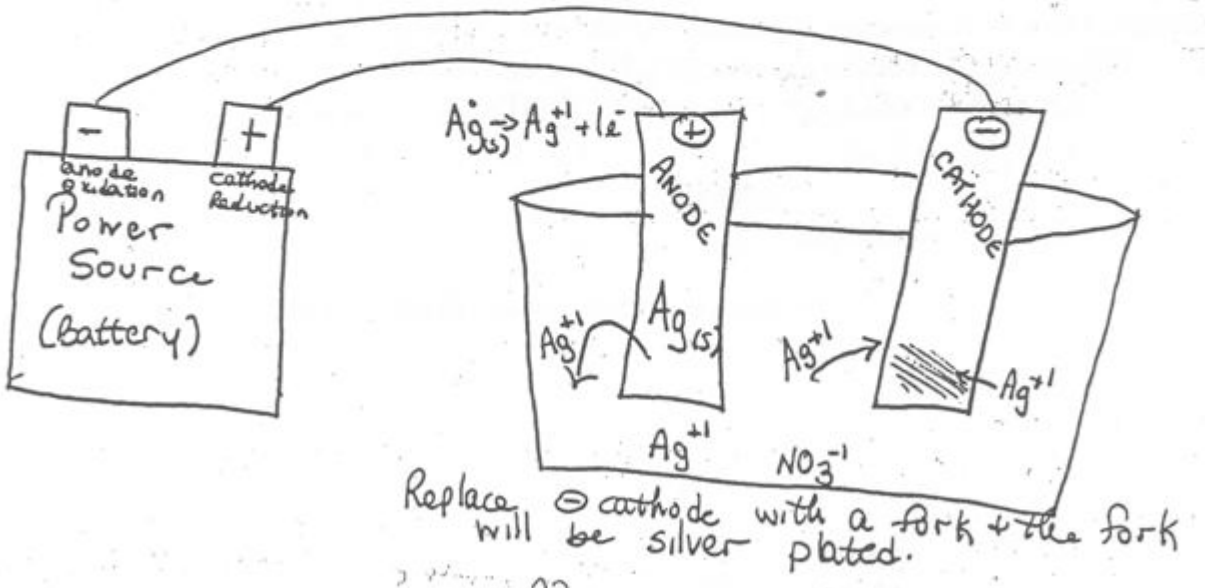
see electrochemical cell on reverse side of this paper



Electrolytic Cell *(for non spontaneous reactions)*
The application of an electrical current (external power source) to a non-spontaneous redox reaction will cause the reaction to take place and the flow of electrons to occur.

As in an electrochemical cell, the anode is the solid metal from the oxidation reaction and the cathode is the solid metal from the reduction reaction.
However, unlike in an electrochemical cell, in the electrolytic cell

1) the **anode is positive** as a result of being hooked up to the + side of the power source and the **cathode is negative** as a result of being hooked up the - side of the power source.
Helpful hint: **A POX in ELECTROLYSIS**
Anode Positive for OXidation in ELECTROLYSIS
2) there is no salt bridge since both half reactions occur in the same container
3) the *negative cathode* can be replaced with a fork or other conductor and can be **electroplated** (coated) with the positive metal ions from the solution.



Chemistry Regents Review: Organic Chemistry

Organic chemistry is the study of organic molecules- those **molecules containing carbon** (with hydrogen). First we will concentrate on *hydrocarbons*, those organic molecules *containing only carbon and hydrogen*. Then we will focus on organic molecules containing oxygen in addition to carbon and hydrogen.

Carbon contains four valence electrons (group 14) and can therefore form four covalent bonds with hydrogen or other carbon atoms. The four bonds are arranged in a shape called a **tetrahedron** so that the bonds maintain the maximum possible distance between them.

Characteristics of Organic Molecules

1) They are generally **non-polar** (no + or - side)
2) They are generally **insoluble in water** (because water *is* polar and organic compounds are not).
3) They are generally **non-electrolytes** (due to characteristics 1 and 2)
4) As non-polar compounds the molecules are held together by weak Van der Waals forces. As a result they generally have **low melting points** and **low boiling points**. However, when comparing characteristics of organic molecules, *the larger the molecule the higher the boiling point and freezing point* (larger molecules have stronger Van der Waals attractions).
5) Because organic reactions require high **activation energy**, organic reactions generally proceed *more slowly* than inorganic reactions.

Hydrocarbons- molecules containing carbon and hydrogen only

Chains of Hydrocarbons

All hydrocarbon chains can be divided into three homologous series based on the bonding of the carbons.

1) **Alkanes** - all the bonds are single bonds
2) **Alkenes** - the molecules contain a carbon-to-carbon double bond
3) **Alkynes** - the molecules contain a carbon-to-carbon triple bond

All hydrocarbons are named for the number of carbon atoms they contain (prefix) and for the homologous series to which they belong (suffix: -ane, -ene or -yne).

The prefixes used in naming hydrocarbons are listed on table P.

General Structural Formulas

A hydrocarbon can be identified as an alkane, alkene or alkyne by its molecular formula. All hydrocarbons in the same homologous series follow the same general formula:

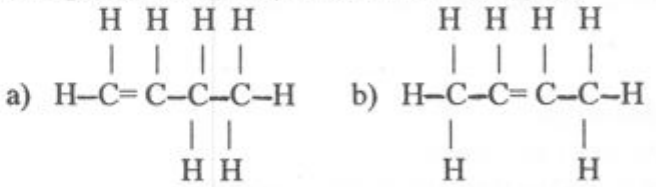
Alkane - C_nH_{2n+2}
Alkene - C_nH_{2n} [n represents the number of carbon atoms]
Alkyne - C_nH_{2n-2}

Alkanes are said to be **saturated hydrocarbons** because they contain the maximum possible number of hydrogen atoms (2n + 2). Since each carbon bond is a single bond, every possible bond is occupied by a hydrogen atom. In alkenes two hydrogen atoms are replaced by a second carbon-to-carbon bond and in alkynes four hydrogen atoms are replaced by a second and third carbon-to-carbon bond. **Alkenes** and **alkynes** are said to be **unsaturated hydrocarbons**.

Isomers

Two molecules with the same molecular formula but different structural formulas (different arrangement of atoms) are said to be isomers.

example:



molecule a) 1 butene molecule b) 2 butene

When **naming hydrocarbons** follow these simple rules:

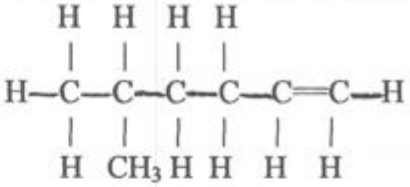
1) Find the longest continuous chain of carbon atoms containing the double or triple bond if present.

2) Name the hydrocarbon for that chain (pick prefix and suffix from tables P and Q).

3) Number the carbon atoms in the chain (from both ends) giving the double or triple bond the lowest possible number. The number is written before the name.

4) If any extra groups are attached to the chain, name the group based on the number of carbons (as above) with the suffix -yl. State the carbon number the group is attached to.

example:



5-methyl 1-hexene

Chemistry Regents Review: Organic Reactions

All hydrocarbons (molecules containing only carbon and hydrogen) are divided up into three categories.

1) chained/ branched hydrocarbons - alkanes, alkenes and alkynes
2) cyclic hydrocarbons
3) aromatic hydrocarbons

Cyclic Hydrocarbons

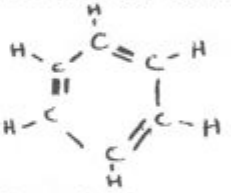
A saturated (all single bonds, alkane) hydrocarbon arranged in a ring. A minimum of three carbon atoms is required and it is named like an alkane, based on the number of carbon atoms with the prefix cyclo-.

examples: cyclopropane, cyclohexane

Aromatic Hydrocarbons

Aromatic hydrocarbons are called aromatic because of their characteristic pleasant smells. Aromatic hydrocarbons are arranged in a ring in which every other bond is a double bond. You should be familiar with the most common aromatic hydrocarbon:

benzene



or

or

Chemical Reactions of Organic Compounds

1) **Combustion** Reaction (burning)

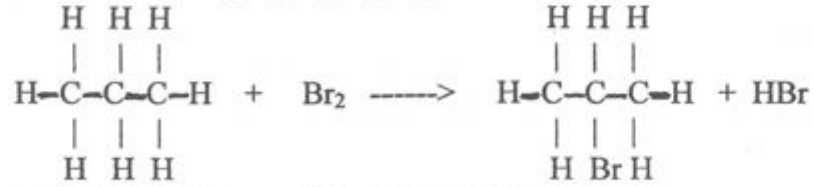
any hydrocarbon + O_2 ----------> CO_2 + H_2O

example: $2C_2H_6 + 7O_2$ -------> $4CO_2 + 6H_2O$

2) **Substitution** Reaction

any alkane (saturated) + any diatomic halogen (Cl_2, Br_2, I_2, F_2) ----> substituted halocarbon + acid

example:



Notice the hydrogen was replaced with one of the bromine atoms.

The hydrogen on the center carbon was replaced rather than hydrogen from one of the extreme carbons because of the 'degree' of the carbon.

Primary (1^0) carbon- a carbon attached to only 1 other carbon

Secondary (2^0) carbon- a carbon attached to 2 other carbons

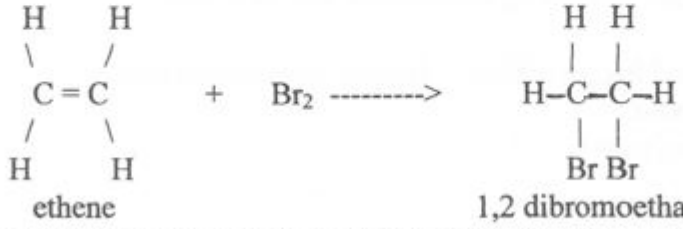
Tertiary (3^0) carbon- a carbon attached to 3 other carbons

The higher the degree of carbon the more likely that its hydrogen will be substituted.

3) **Addition** Reaction

Any unsaturated hydrocarbon (alkene or alkyne) + diatomic molecule ----> a more saturated dihalocarbon

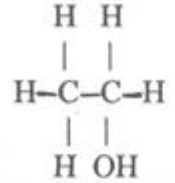
example:



ethene 1,2 dibromoethane

Notice the double bond was opened and 1 bromine attached to each carbon.

If water, H_2O (H-OH) had been added to ethene the product would be:

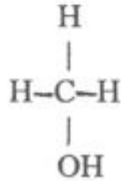


ethyl alcohol

an **alcohol** (any hydrocarbon attached to -OH)

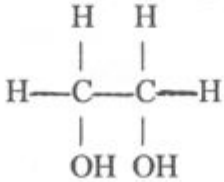
Primary Alcohol- hydroxy group (OH) is on a primary carbon
Secondary Alcohol- OH is on a secondary carbon
Tertiary Alcohol- OH is on a tertiary carbon

Monohydroxy alcohol
one OH group



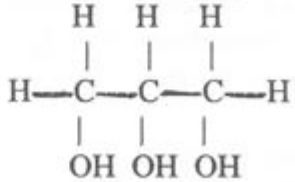
methanol

Dihydroxy alcohol
two OH groups



ethylene glycol

Trihydroxy alcohol
three OH groups



glycerol

4) **Fermentation Reaction**

sugar + enzyme-------> alcohol + CO_2

$C_6H_{12}O_6$ ----------> $2C_2H_5OH + CO_2$

5) **Polymerization** reaction

Forming a polymer (chain of repeating units) from monomers (small, basic units) is called polymerization. There are two types of polymerization:

a) **Condensation Polymerization** - forming a polymer with the removal of water

$HO\text{-}CH_2\text{-}CH_2\text{-}OH + HO\text{-}CH_2\text{-}CH_2\text{-}OH + HO\text{-}CH_2\text{-}CH_2\text{-}OH \rightarrow$ (a polymer)

$HO\text{-}CH_2\text{-}CH_2\text{-}O\text{-}CH_2\text{-}CH_2\text{-}O\text{-}CH_2\text{-}CH_2\text{-}OH + H_20$

b) **Addition Polymerization** - forming a polymer from two or more unsaturated molecules with an addition reaction that makes a more saturated polymer.

$H_2C{=}CH_2 + H_2C{=}CH_2 \longrightarrow —[-CH_2—CH_2—CH_2—CH_2-]—$

The double bond opened so another molecule can be added on each side.

The reaction can be represented as

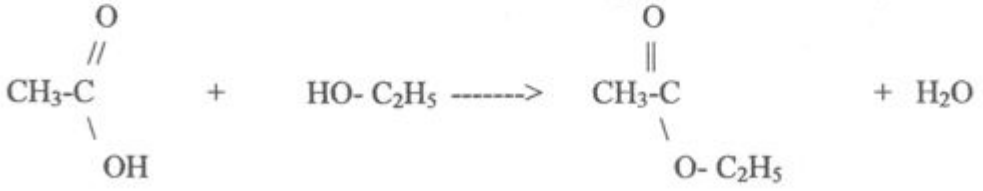
$$n\ CH_2{=}CH_2 \rightarrow -[CH_2\text{-}CH_2]\text{-}_n$$

Naturally occurring polymers include: Starch and cellulose (made of simple sugars) and proteins (made of amino acids). Synthetic polymers include polyester, nylon and plastic.

6) **Esterification** Reaction

organic acid + alcohol -------> ester + water

When water is removed to attach 2 molecules the process is called dehydration synthesis.



7) **Saponification** reaction

The reverse of an esterification reaction - the hydrolysis (break down with the addition of water) of an ester in the presence of a base. Soap and glycerol are always the products of saponification.

The **Structures of Organic Compounds** and examples of how to name them can be found on table R.

Chemistry Regents Review: Nuclear

Nuclear chemistry is the study of the **nucleus** of the atom. The nucleus contains protons and neutrons and has a net positive charge. During the entire chemistry course you studied chemical reactions which were dependent upon changes in the electrons of the atom. Only in nuclear chemistry can a nucleus (protons and neutrons) be altered.

Radioactivity - the release of energy (and particles) from an unstable nucleus as it decays.
There are several types of emissions (particles and rays) that are given off by a radioactive atom. Their names, masses, charges and symbols are listed on Reference Table O.
The radioactive emissions can be separated by passing them through a magnetic field (with a + and - side). **Positive alpha particles** ($^{4}_{2}He$) and **positrons** ($^{0}_{+1}e$) are attracted to the negative side of the field, **Negative beta particles** ($^{0}_{-1}e$) are attracted to the positive side and **gamma rays** ($^{0}_{0}y$) which are uncharged, pass right between the positive and negative sides of the field without being deflected.

Natural vs. Artificial Radioactivity
The **stability** of an isotope is based on the ratio of neutrons to protons in its nucleus.
Natural radioactivity is the spontaneous decay of a nucleus (the nucleus emits a radioactive particle). No element with an atomic number above 83 (Bi) has any stable isotopes. Elements below could have stable and/or unstable isotopes.
Artificial radioactivity is the man-made decay of a nucleus for the purpose of energy production. In order to cause nuclear decay, the nucleus must be made unstable by bombarding it with nuclear particles (example: alpha, beta or neutrons).

Natural Radioactivity
We are able to determine the *rate of decay* of a radioactive substance by a constant called its **half life**. A half life is the amount of time it takes for *one-half* of the nuclei in a sample to decay. Reference Table N lists various radioactive isotopes, their half lives and their mode of decay (particle they emit).
Using this table we can determine what mass of a sample will remain after a given amount of time by *halving* (dividing by 2) *our sample* with the passage of each half life. We can also determine the mass of the original sample or the length of a half life given enough information. We can determine the fraction of the original sample remaining by assuming the original sample was 1 (whole) and dividing by 2 with each half life (1/2, 1/4, 1/8 etc.).

The theory of half lives is used in radio-**carbon dating**. The constant rate of decay (half life) of an isotope can be used to determine the age of fossils and trees. The ratio of Carbon-12 to Carbon-14 (the unstable isotope) is constant when an organism is alive. When the organism dies, the Carbon-14 continues to decay, changing the ratio. We can thus determine the age of the tree. Uranium-238 is used for determining the ages of rocks.

Radioactive isotopes are also used in **diagnostic procedures** such as **iodine-131** used to determine thyroid abnormalities. The isotope may not be harmful to the body and must have a short half life so it is quickly eliminated from the body. Radioactive isotopes are also used in **cancer therapies**.

Transmutation - We can determine which new element will be formed as the atoms of elements decay (different # of protons is a different element). This process of forming a new element is called transmutation. The sum of the masses and the sum of the charges (protons) must remain constant throughout the change.

example: $^{42}_{19}K \rightarrow {}^{0}_{-1}e + X$ X must have a mass of 42 and an atomic number of 20. Atomic # 20 = Ca

(always look at atomic # to determine the element formed)

Artificial Radioactivity

We can cause the decay of an otherwise stable nucleus by making it an unstable isotope. Bombarding a stable nucleus with nuclear particles (protons, neutrons, alpha, beta etc.) will cause a transmutation to an unstable nucleus (radioactive isotope). This unstable nucleus will then decay. The decay of a nucleus would be caused for the purpose of harnessing the energy released.

The energy released during a nuclear reaction is the result of its **mass defect**. The mass of the protons and the neutrons individually is greater than the mass of the protons and neutrons when together in the nucleus. This mass or mass defect, is not lost, rather it is converted to energy. The amount of energy produced is related to the mass defect based on Einstein's formula: $E = mc^2$ where E is the energy produced, c is the speed of light and m is the mass defect.

Nuclear Reactions

The two types of nuclear reactions are fusion and fission.

Fusion is the formation of a heavy nucleus from lighter ones.

(example: in the sun, hydrogen atoms are converted to helium atoms)

Fission is the formation of lighter nuclei from a heavier one. Fission is usually a chain reaction in which the products also undergo fission and so on.

(example: conversion of uranium to krypton)

Both fusion and fission reactions produce great amounts of energy which can be controlled as in a *nuclear reactor* or released in a single burst of energy as in a *nuclear bomb.*

Radiation Risks

High doses of radiation can cause serious illness and death. Radiation is also known to cause mutations in cells. Nuclear power plants produce radioactive wastes with long half-lives that are difficult to store and dispose of properly. There are isotopes of radon that are very dangerous gases that are harmful in high levels.