Chemistry Review for Final Exam

Data Analysis

Absolute error $=$ Experimental Value – Accepted Value   * Absolute error is always positive.

$$\%\ \textbf{Error} = \frac{|\text{ Experimental Value} - \text{Accepted Value}|}{\text{Accepted Value}} \times 100 \quad \textbf{\textit{in other words}} \quad \frac{\text{Big \#} - \text{Small \#}}{\text{True \#}} \times 100$$

$$\frac{\textbf{\textit{Density}}\ (g/ml)}{1} = \frac{mass\ (g)}{volume\ (ml)}$$

Density has units of **g/ml** or **g/cm^3** or **kg/m^3** $\boxed{\textbf{1 ml = 1 cm}^3}$

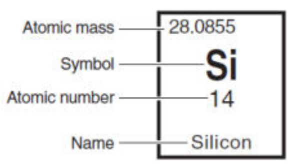
Sig Fig Rules:

Zeros in the FRONT are NEVER significant.	**0.002 = 1 sig figs**
Zeros in the MIDDLE are ALAWAYS.	**2005 = 4 sig figs**
Zeros at the END – ONLY IF A DECIMAL	**5000 = 1 sig fig 5.000 = 4 sig figs**

Base Metric Units: Length (meters)   Volume (liters)   Mass (grams)   Energy (Joules)

1 kilogram = 1000 grams   1 Liter = 1000 milliliters   1 meter = 100 cm

Atomic Structure



Atomic Mass on the periodic table has a decimal because it is based on the average mass of all known **isotopes**. Each isotope has a different number of **neutrons** but the same number of **protons**. Some isotopes are rare and others are very abundant. Abundancy can be given as a **percent** (% Abundance) or as a **decimal** (Relative Abundance). The average mass is based on the weighted abundance of each isotope.

$$\text{Average Atomic Mass} = \frac{(\text{Mass} \times \ \%\ \text{Abundance}) + (\text{Mass} \times \ \%\ \text{Abundance})}{100} \quad \textbf{\textit{or}} \quad =$$
$$\frac{(\text{Mass} \times \text{Relative Abundance}) + (\text{Mass} \times \text{Relative Abundance})}{100}$$

Atomic number = equals the number of protons for given element

Mass # = protons + neutrons for a given isotope

Proton # = atomic number

Neutron # = mass # - atomic number

Electron # = Proton # for neutral elements

48
22 Ti

Mass # = 48
Atomic # = 22
Protons = 22
Neutrons = 48 - 22 = 26

Oxidation # = charge on atom based on # of electrons (**LEO** + goes **GER** -)

Metals lose electrons to become **positive**.

Nonmetals gain electrons to become **negative**.

Electron configuration pattern: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Aufbau Principal – start with lowest orbital

Pauli Principal – only two electrons per orbital

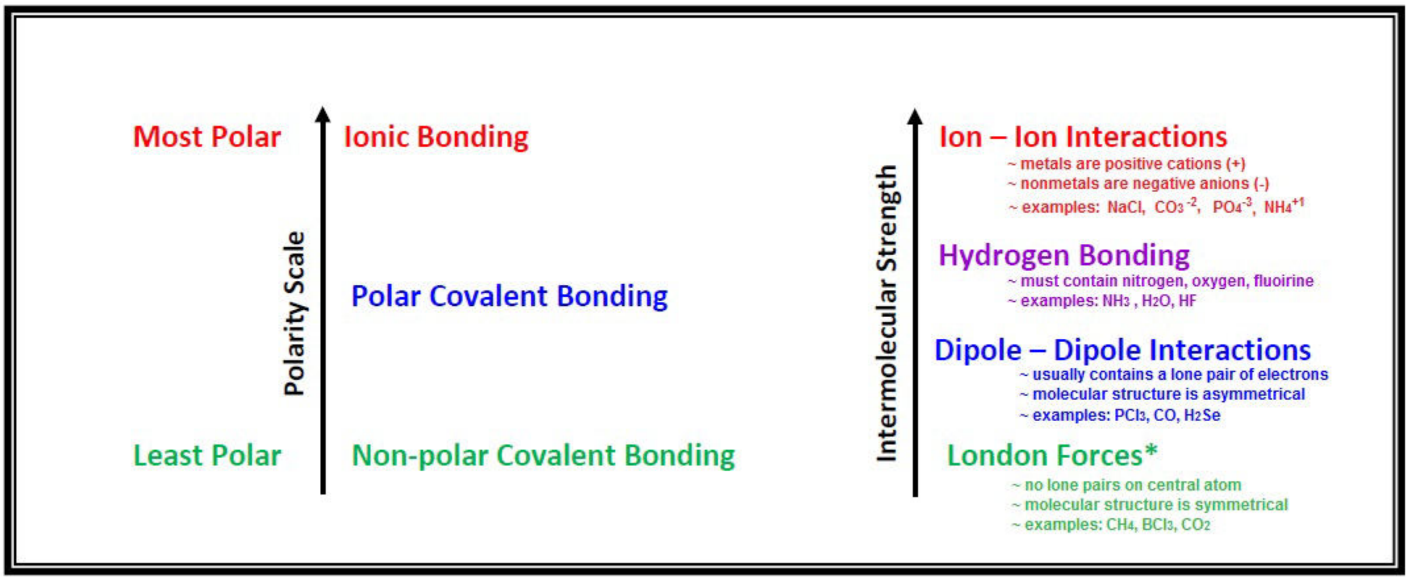
Hund's Rule – fill all up arrows before adding down arrows

Abbreviated Configuration – starts from previous noble gas

			1s	2s	2px	2py	2pz
Lithium,	Li	$1s^2 2s^1$	↑↓	↑			
Beryllium,	Be	$1s^2 2s^2$	↑↓	↑↓			
Boron,	B	$1s^2 2s^2 2p^1$	↑↓	↑↓	↑		
Carbon,	C	$1s^2 2s^2 2p^2$	↑↓	↑↓	↑	↑	
Nitrogen,	N	$1s^2 2s^2 2p^3$	↑↓	↑↓	↑	↑	↑
Oxygen,	O	$1s^2 2s^2 2p^4$	↑↓	↑↓	↑↓	↑	↑
Fluorine,	F	$1s^2 2s^2 2p^5$	↑↓	↑↓	↑↓	↑↓	↑
Neon,	Ne	$1s^2 2s^2 2p^6$	↑↓	↑↓	↑↓	↑↓	↑↓

ex. Strontium = [Kr] $5s^2$

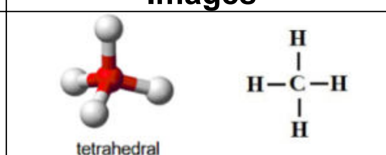
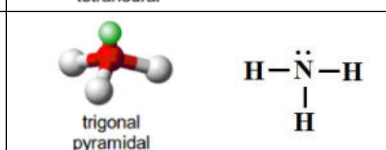
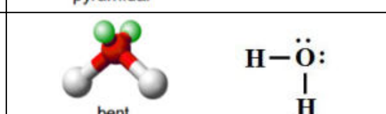
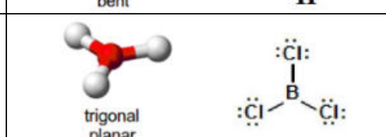
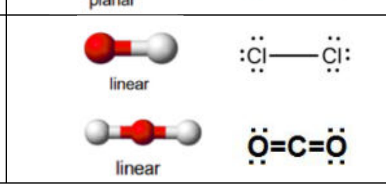
Intermolecular Forces



Intermolecular forces determine the physical properties of a substance. **Stronger forces** result in **higher boiling points**, **higher melting points**, and **higher solubility**, but **lower vapor pressures.**

VSEPR

VSEPR stands for Valence Shell Electron Pair Repulsion. The number and type of bonds on the central atom determine the molecular shape. **Polar** bonds will be asymmetrical, and **nonpolar** bonds will be symmetrical.

VSEPR Shape	Structure	Bond Angle	Images	Examples
Tetrahedral	4 single bonds, no lone pairs	109.5°		CH_4
Trigonal Pyramidal	3 single bonds, 1 lone pair	107°		NH_3, PCl_3
Bent	2 single bonds, 2 lone pairs	105°		H_2O, H_2S
Trigonal Planar	3 single bonds, no lone pairs	120°		BH_3, BCl_3
Linear	1 or 2 bonds, no lone pairs	180°		H_2, Cl_2, O_2 CO_2, HCN

Periodic Trends

		Direction of Increase	Left to Right Across Period	Down Group
A	Atomic Radii		decreases	increases
I	Ionization Energy		increases	decreases
M	Metallic Properties		decreases	increases
E	Electronegativity		increases	decreases
S	Shielding		stays same	increases

As the **atomic number** increases, more **protons** are added to the nucleus. This attracts electrons inward, and makes the **atom smaller** as you go across towards the noble gases.

Metals become **positive ions** by losing electrons and getting **smaller**. **Nonmetals** gain electrons to become **negative ions** that are **larger.**

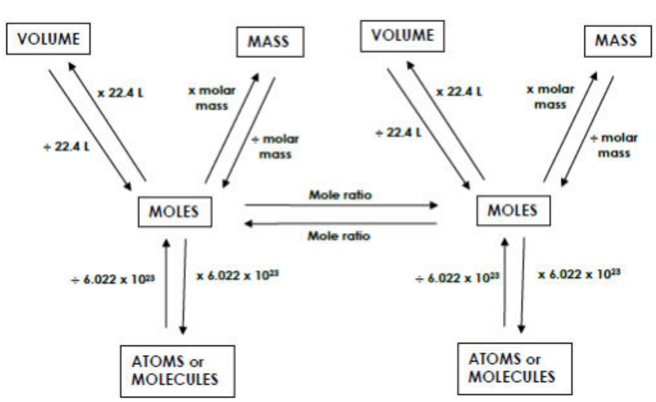
Ionization Energy is a measure of the energy required to **remove** an electron. Therefore, the **noble gases** have **High IE** because their stability makes it difficult to lose electrons.

Electronegativity is a measure of the atoms ability to **attract** electrons. **Fluorine** has the **highest** electronegativity. The **noble gases** have **ZERO** electronegativity because they have no desire to gain electrons because they are already stable.

Shielding refers to the number of orbitals between the valence electrons and the nucleus. The number of shielding orbitals equals the row # minus 1.

Stoichiometry

The Roadmap to Stoichiometry



If you only have 1 chemical, then all conversions are compared to 1 mole. Multiply when you are given moles and want to covert to liters, grams, atoms, or molecules. Divide if you are starting with liters, grams, atoms, or molecules and need to calculate the number of moles.

$$\frac{\textit{Moles}}{1} = \frac{\textit{grams}}{\textit{molar mass}}$$

$$\left|\frac{\blacksquare}{\quad}\right|\frac{\quad}{\blacksquare}\left|\frac{\Delta}{\quad}\right| = \Delta$$

Limiting Reactant is the reactant that will run out first.
Excess Reactant is abundant and will be left over after the reaction is complete.

$$\% \ \textbf{Yield} = \frac{\text{Actual (made in lab)}}{\text{Theoretical (calcualted by math)}} \text{ x } 100$$

Stoichiometry proves the **Law of Mass Conservation** because the total mass of the reactants equals the total mass of the products.

Kelvin (K) = °C + 273 **Celsius °C** = K – 273

Absolute Zero = 0 K or - 273°C
Freezing Point of Water = 0°C or + 273 K
Boiling Point of Water = 100°C or 373 K

Standard Temperature & Pressure (STP)
- Pressure = 1atm = 760 mmHg = 101.3kPa
- Temperature = 0°C = 273K

Boyle's Law $P_1V_1 = P_2V_2$

Guy Lussac's Law $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Charles' Law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Avogadro's Law $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Ideal Law $PV = nRT$

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

Dalton's Laws: $P_1 + P_2 + P_3 = P_{total}$

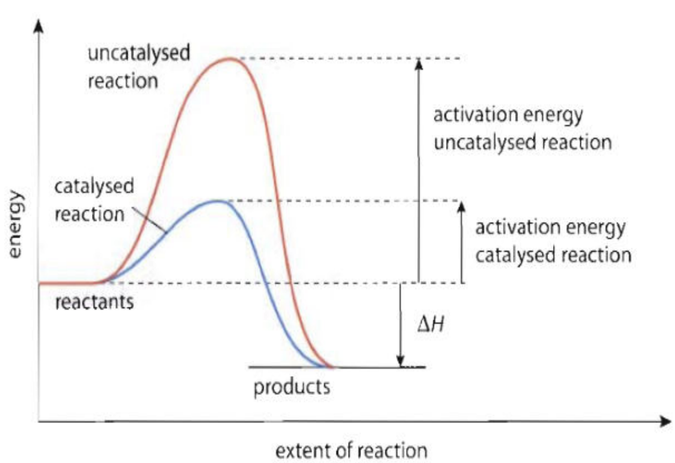
$n_1 + n_2 + n_3 = n_{total}$

$\frac{P_1}{P_{total}} = \frac{n_1}{n_{total}}$

Graham's Law: $\frac{\text{fast rate}}{\text{slow rate}} = \sqrt{\frac{\text{slow molar mass}}{\text{fast molar mass}}}$

<u>Behavior of Gases</u>
Gases move in rapid, random, constant motion.
Gases move in straight line paths.
Gases have negligible volume (basically zero) compared to their containers.
Gases have elastic collisions in which they do not lose kinetic energy when colliding.
Gases prefer high temperature an low pressure.
The average kinetic energy of the molecules is proportional to the temperature.
Gases exert pressure on walls of container and on nearby molecules.
Gases have low intermolecular forces (nonpolar) and are not attracted unless induced.

Causes of Change



A **catalyst** lowers the activation energy and speeds up the reaction.

Activation energy is the energy required to break bonds.

Enthalpy (ΔH) is the energy absorbed or released in a chemical reaction; equals products minus reactants.

Endothermic reactions occur when products are higher than the reactants because the reaction absorbed energy during the reaction; **(+ΔH)**

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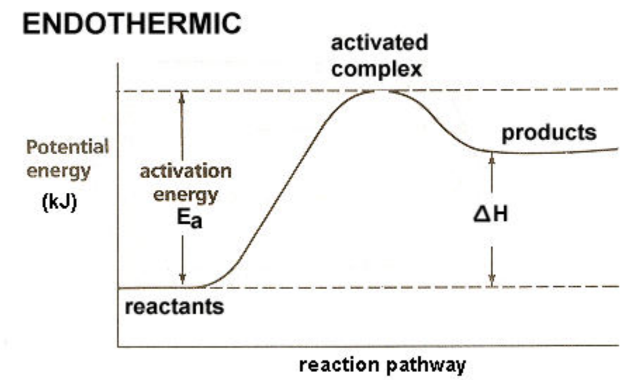
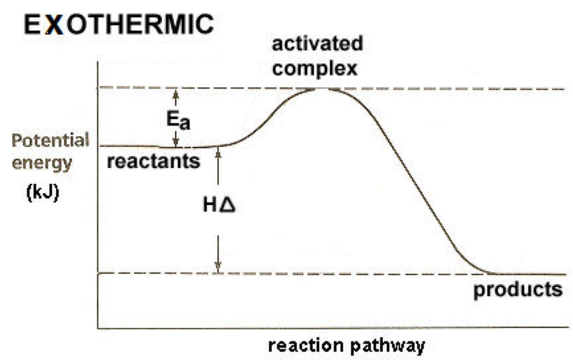
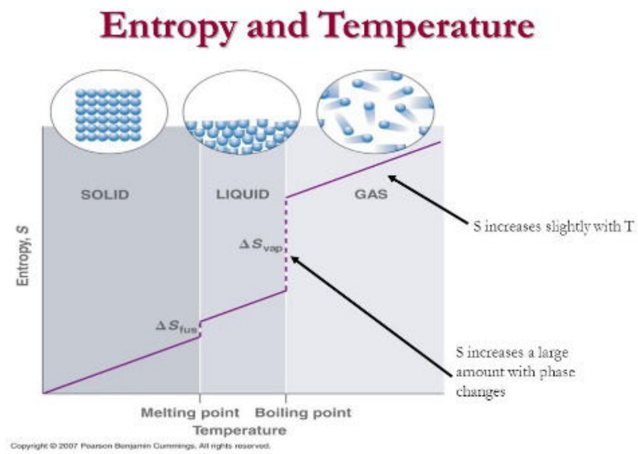




Exothermic reactions occur when products are lower than the reactants because the reaction has released energy during the reaction**; (– ΔH)**



Entropy (ΔS) is a measure of randomness.

Gases (g) and **aqueous ions (aq)** have the most

$$aA + bB \rightleftharpoons cC + dD$$

$$\boxed{K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

Temperature, Pressure, and Concentration can effect Keq.

Keq >1 favors products; forward rxn

Keq = 1 @ equilibrium

Keq < 1 favors reactants; reverse rxn

*Solids and Liquids are NOT included in Keq equation.

$$Q = \text{mass} \cdot \text{specific heat} \cdot (T_{\text{final}} - T_{\text{initial}})$$

$$\Delta G = \Delta H - \left(T \times \frac{\Delta S}{1000}\right)$$

$$\Delta H = \text{Products} - \text{Reactants}$$

Solute, Solvents, and Solutions

$$\frac{\text{Moles}}{1} = \frac{\text{grams}}{\text{molar mass}} \qquad \frac{\text{Liters}}{1} = \frac{\text{milliliters}}{1000} \qquad \frac{\text{Molarity}}{1} = \frac{\text{moles}}{\text{liters}}$$

Preparing dilutions: $M_1 V_1 = M_2 V_2$ **Stock = more concentrated solution**

A **solute** is dissolved by the **solvent** in order to make a **solution.**

Homogeneous – looks the same (Examples: air, bronze, salt water)

Heterogeneous – looks different (Examples: lava lamp; oil and water; ice water, cement)

Acids & Bases

Arrhenius Theory – acids produce $\mathbf{H^+}$ and bases have $\mathbf{OH^-}$

Bronsted – Lowry – acids are **proton donors** and bases are **proton acceptors**

$$\mathbf{pH = -\log [H^+]} \qquad \mathbf{10^{-pH} = [H^+]} \qquad \mathbf{pOH = -\log [OH^-]} \qquad \mathbf{10^{-pOH} = [OH^-]}$$

An acid-base titration is a lab technique that allows you to determine the concentration of an unknown solution. (There are other types of titrations, but this is the most common.) Some terms you need to know:

- **Titrant**: a solution of known concentration (usually); usually the solution in the buret
- **Analyte**: the solution you are trying to determine the concentration of; usually the solution in the beaker or flask
- **Equivalence point:** the volume of titrant added to give **equal moles of acid and base** (in an acid/base titration)
- **End point:** the volume of titrant added to make the color of the indicator change
  **hopefully, the equivalence point and the end point happen at the same time!
- **Indicator:** a solution that changes color in varying pH ranges

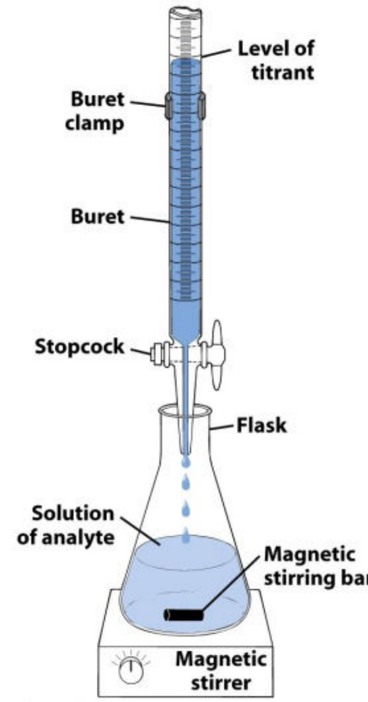


Figure 7-1
Quantitative Chemical Analysis, Seventh Edition
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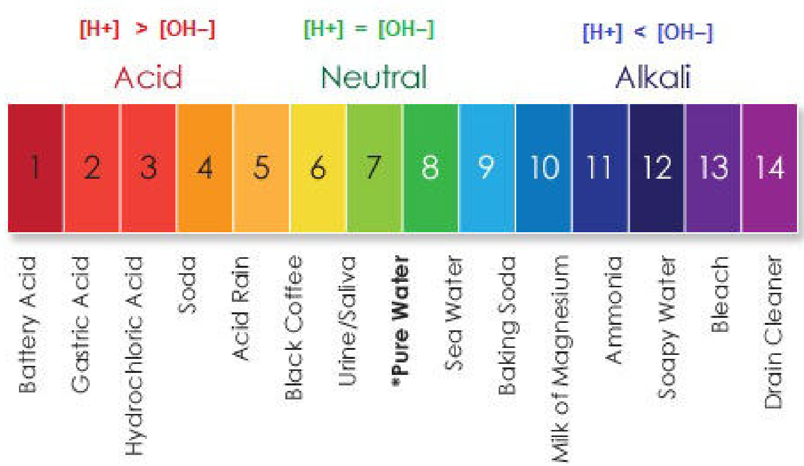
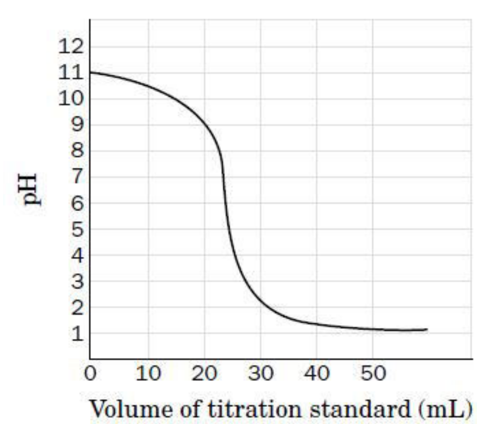
$$\underbrace{M_{acid}}_{\text{MOLARITY OF ACID}}\overbrace{V_{acid}}^{\text{VOLUME OF ACID}} = \underbrace{M_{base}}_{\text{MOLARITY OF BASE}}\overbrace{V_{base}}^{\text{VOLUME OF BASE}}$$

Note: If you are using a diprotic or triprotic acid, then you need to add coefficients to the equation above. (**C_a** = coefficient of acid; **C_b** = coefficient of base. Coefficients are based on the # of moles required for a balanced chemical equation.)

$$\mathbf{M_a\, C_b\, V_a = M_b\, C_a\, V_b}$$

A titration curve shows how the pH changes as you add titrant to the analyte.





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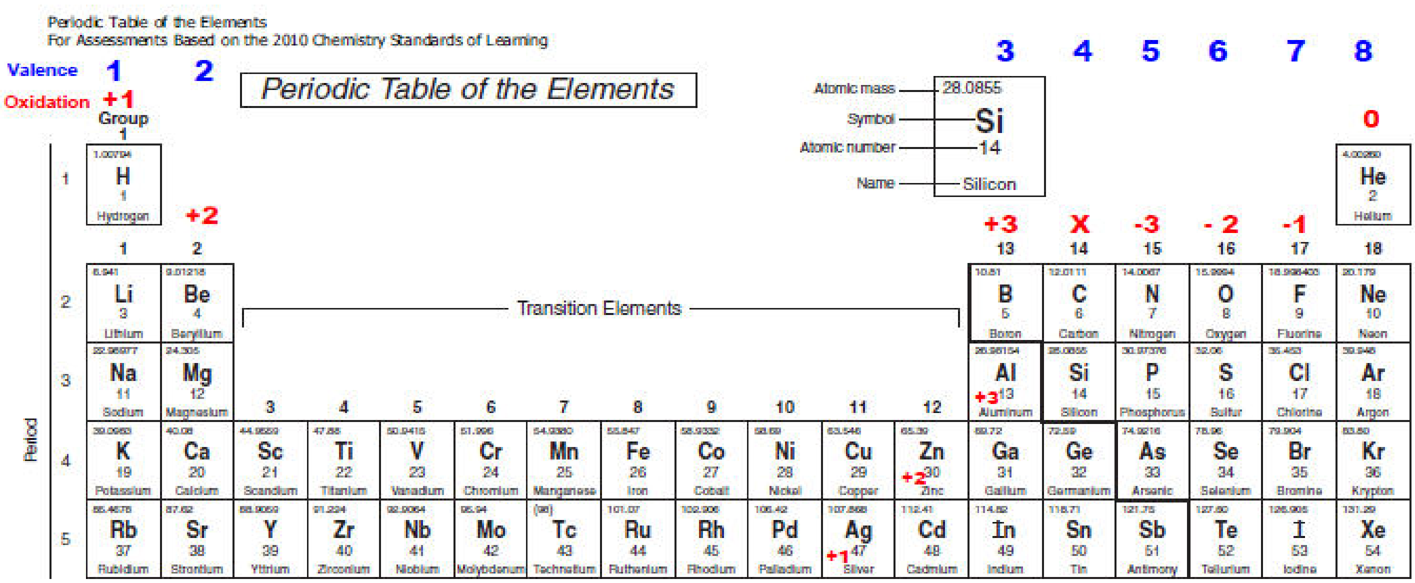
DIATOMIC Molecules

Elements that never exist alone (are always found in pairs) – Start with 7 and make a 7 then top your hat; **N_2, O_2, F_2, Cl_2, Br_2, I_2, H_2**; or Have No Fear Of Ice Cold Beer **H_2, N_2, F_2, O_2, I_2, Cl_2, Br_2**

Polyatomic Ions

Ion Name	Ion Formula	Example
Hydroxide	OH^{-1}	Sodium Hydroxide $NaOH$
Nitrate	NO_3^{-1}	Sodium Nitrate $NaNO_3$
Carbonate	CO_3^{-2}	Potassium Carbonate K_2CO_3
Sulfate	SO_4^{-2}	Magnesium Sulfate $MgSO_4$
Phosphate	PO_4^{-3}	Calcium Phosphate $Ca_3(PO_4)_2$
Ammonium	NH_4^{+1}	Ammonium Hydroxide NH_4OH



Roman numerals are only used for the transition elements in columns 3 – 12, and 14.
Silver (Ag+1) and Zinc (Zn+2) do not change so they do not need Roman numerals.

Greek prefixes (mono, di, tri, tetra) are only used with nonmetal covalent bonds.

Lewis Dots represent **valence electrons** from the **s** and **p** orbitals.

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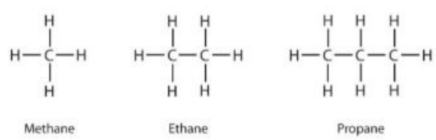
Organic Chemistry

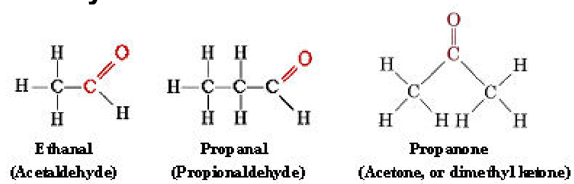
Molecules can only be considered "organic" if they contain CARBON and HYDROGEN.   For example, sugar is organic because it contains $C_{12}H_{22}O_{11}$, but carbon dioxide CO_2 is not organic because it does not contain hydrogen.

PURE hydrocarbons are insoluble in water due to nonpolar London forces. Adding POLAR functional groups, such as alcohols, aldehydes, and carboxylic acids, and increase the solubility in water.

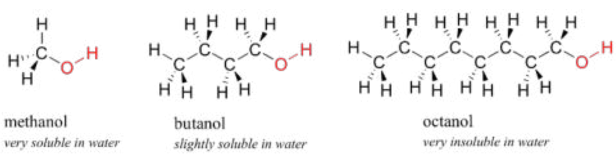
Alkanes



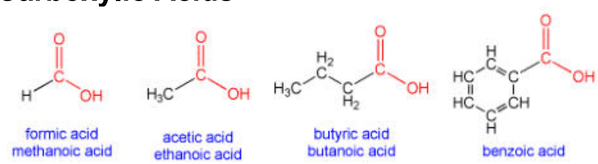
Aldehydes



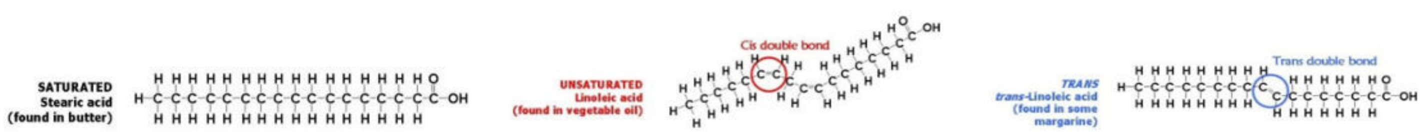
Alcohols



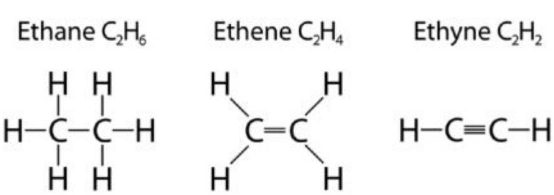
Carboxylic Acids



A hydrocarbon is **SATURATED** when it has the maximum amount of hydrogen bonds possible. Saturated fats are usually solid at room temperature because their linear structures pack tightly together. The bends and kinks in the unsaturated and trans-saturated fats create gaps, causing the intermolecular forces to be weaker and resulting in liquids. The molecules are too large to be gases.



Carbon can have single bonds (**alkanes**), double bonds, (**alkenes**), and triple bonds (**alkynes**).   Single bonds are called **sigma (σ) bonds**, and double and triple bonds are called **pi (π) bonds**. The more bonds, the stronger the **bond energy** between the atoms. The covalent bonds are **attractive forces** connecting the atoms, so the stronger the attractive forces, the closer together the atoms will be, making the **bond length** shorter.



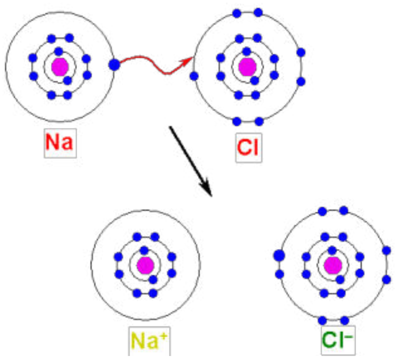
Bond Energy: $C \equiv C$ **>** $C = C$ **>** $C - C$

Bond Length: $C \equiv C$ **<** $C = C$ **<** $C - C$

Molecular Formulas: show actual structure   **Example: $C_{12}H_{24}O_9$**

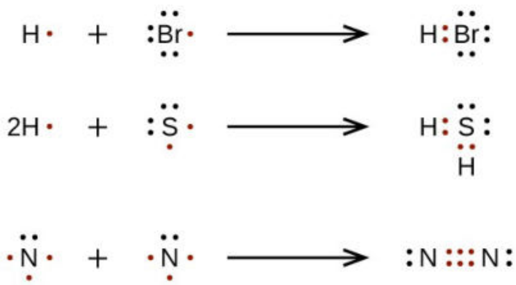
Empirical Formulas: have been reduced   **Example: $C_4H_8O_3$**

Chemical Bonding



Ionic Bonding occurs when metals **TRANSFER** valence electrons to nonmetals. **Metals** become **positive CATIONS** and **nonmetals** become **negative ANIONS**. Ionic bonding is the strongest type of intermolecular forces; therefore, ionic compounds exhibit the **highest melting and boiling points.**

Salts are ionic bonds between a metal and a nonmetal. Salts are usually **soluble** if they contain **Group 1, Group 2, or Group 17** ions. Exceptions include **$BaSO_4$** and **AgCl,** which will always be solid **precipitates**.



Covalent bonds occur when **nonmetals SHARE** valence electrons. As electrons are shared, the atom follows the **OCTET RULE**, pairing electrons until it reaches a maximum of 8 valence electrons.

Exceptions to Octet Rule:
Hydrogen and **helium** can only share up to 2 electrons.
Boron can share up to 6 electrons.

: represents one lone pair of electrons. Lone pairs do not bond with other atoms. A single dot **·** represents an electron wanting to be shared so it is called a bonding electron.

Types of Chemical Reactions

Synthesis	A + B → AB	Only 1 product
Decomposition	AB → A + B	Only 1 reactant
Single Replacement	A + BC → AC + B	Must have an element on both sides
Double Replacement	AB + CD → AD + CB	4 ionic compounds, no water
Neutralization	**H**B + A**OH** → AB + **H(OH)**	acid + base make **salt and water**
Combustion	$\mathbf{C_xH_y} + O_2 \rightarrow CO_2 + H_2O$	makes **carbon dioxide and water**