Mole Notes

1 mole = 6.02×10^{23}

The mole is a conversion to count a substance by weighing it. 1 mole is equal to 6.02×10^{23} representative particles of a substance. Depending on the substance the representative particles may be atoms, molecules, formula unit.

For an Element: 1 mole = 6.02×10^{23}
Ionic Compound: 1 mole = 6.02×10^{23} formula units
Covalent Compound: 1 mole = 6.02×10^{23} molecules

Molar mass

The molar mass of a substance is the mass of one mole of that substance. To calculate the molar mass of a substance you must add the individual atomic mass of all the elements that make up that compound. The unit for a molar mass is in grams/mole.

Example:
$C_6H_{12}O_6$
Carbon: 6 x 12 = 72
Hydrogen: 12 x 1 = 12
Oxygen: 6 x 16 = 96
72+12+96= $\boxed{\text{180 grams/mole}}$

$Fe_3(PO_4)_2$
Iron: 3 x 55.85 = 167.55
Phosphorus: 2 x 31 = 62
Oxygen: 8 x 16 = 128
167.55 + 62 + 128 = $\boxed{\text{357.55 grams/ mole}}$

Stoichiometry Notes (Read Chapter 10 in your text book for more background)

Mole Ratio- whole number relationship of moles of one substance to moles of another substance in a balanced chemical equation

Limiting Reactant- the reactant that is used completely in a chemical reaction.

Stoichiometry- mass and mole relationships between reactants and products in a chemical rxn.

PURPOSES

A + B → C +D

1. To find the amount of substance C a given amount of A will yield.
2. To find the amount of substance B that will react with a given amount of A.

GENERAL FORMS

Grams A → Mol A → Mol B → Grams B

$$\boxed{\text{Grams A}} \rightarrow \frac{\text{1 mol A}}{\text{m.m. A (g)}} \rightarrow \boxed{\text{Mol A}} \rightarrow \frac{\text{\# moles B}}{\text{\# moles A}} \rightarrow \boxed{\text{Mol B}} \rightarrow \frac{\text{m.m. B (g)}}{\text{1 mol B}} \rightarrow \boxed{\text{Grams B}}$$

$$\boxed{\text{grams A} \times \frac{\text{1 mol A}}{\text{m.m. A (g)}} \times \frac{\text{\# moles B}}{\text{\# moles A}} \times \frac{\text{m.m. B (g)}}{\text{1 mol B}} = \text{grams B}}$$

MOLES-MOLES PROBLEMS

1. How many moles of oxygen are necessary in the combustion of 7.50 moles of butane (C_4H_{10})?

$$C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$$

Balance: $\underline{2}\ C_4H_{10} + \underline{13}\ O_2 \rightarrow \underline{8}\ CO_2 + \underline{10}\ H_2O$

Given: 7.50 mol of butane (C_4H_{10}) Asked for: moles of O_2

$$\text{moles A} \times \frac{\text{\# moles B}}{\text{\# moles A}} = \text{moles B}$$

(# = number of moles, coefficients, from balanced equation)

$$7.5 \text{ mol } C_4H_{10} \times \frac{13 \text{ mol } O_2}{2 \text{ mol } C_4H_{10}} = 48.75 \text{ mol } O_2$$

MASS-MASS PROBLEMS

2. How many grams of water would be produced by the combustion of 3.75 g butane?

$$\underline{2}\ C_4H_{10} + \underline{13}\ O_2 \rightarrow \underline{8}\ CO_2 + \underline{10}\ H_2O$$

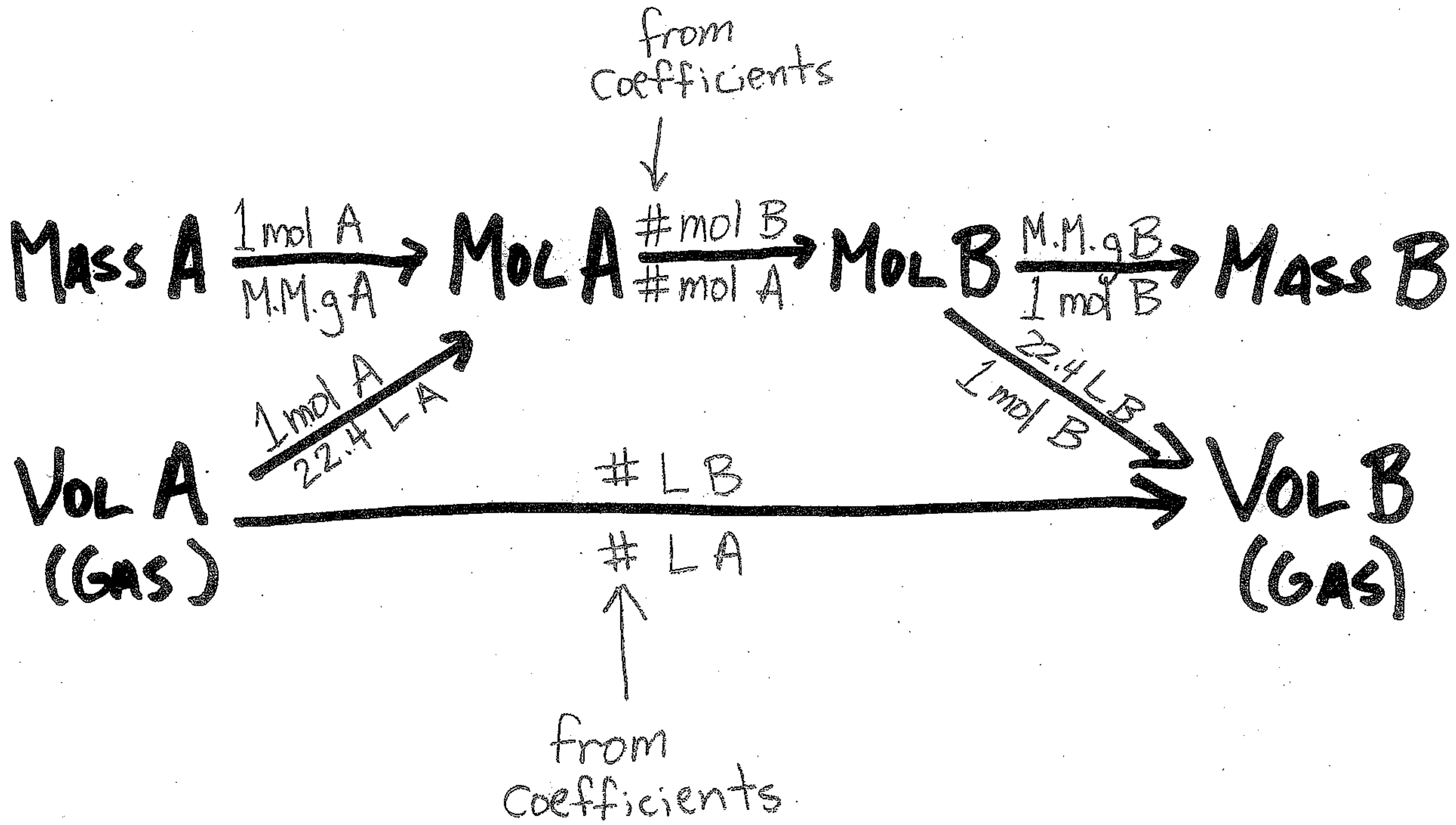
Given: 3.75 g butane (C_4H_{10}) Asked for: grams of H_2O

$$\text{grams A} \times \frac{\text{1 mol A}}{\text{m.m. A (g)}} \times \frac{\text{\# moles B}}{\text{\# moles A}} \times \frac{\text{m.m. B (g)}}{\text{1 mol B}} = \text{grams B}$$

(m.m. = molar mass)

$$3.75 \text{ g } C_4H_{10} \times \frac{1 \text{ mol } C_4H_{10}}{58 \text{ g } C_4H_{10}} \times \frac{10 \text{ moles } H_2O}{2 \text{ moles } C_4H_{10}} \times \frac{18 \text{ g } H_2O}{1 \text{ mol } H_2O} = 5.82 \text{ g } H_2O$$

STOICHIOMETRY



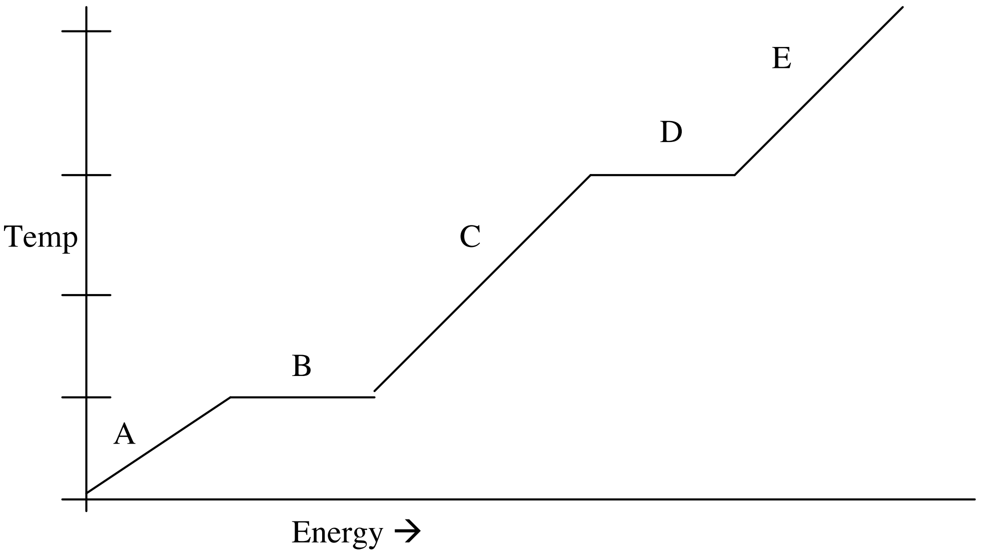
Heat Notes

Heat- the transfer of thermal energy. Dependent on the speed of particles and the amount of particles. Measured in Joules (J)

Temperature (T)- the measure of the average kinetic energy of the particles of matter. Measured with a thermometer in degrees Celsius, Kelvin, or Fahrenheit (°C, °F, K)

Temperature is not a measure of Heat!

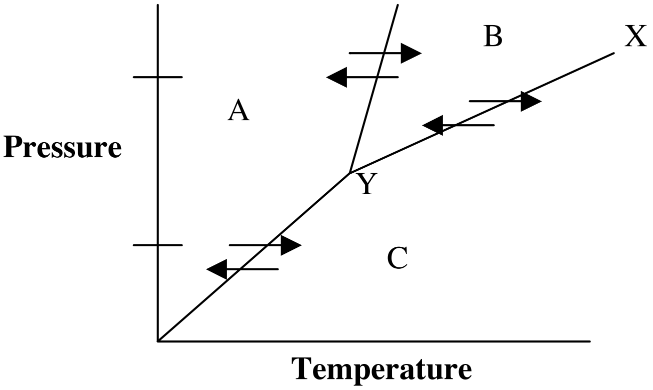
<u>Heating Curve</u>



A- Solid heating up
B- Melting Point- Temperature at which solid changes to liquid
C- Liquid heating up
D- Boiling Point- Temperature at which liquid changes to gas
E- Gas/Vapor heating up
ΔT=0! During a change of state

Temperature *and* Atmospheric Pressure affect changes of state.

Phase Diagram



Y- Triple Point- the specific pressure and temperature for a substance where all three states exist at once. Solid, liquid, and gas in equilibrium.
X- Critical Point-above this temperature the substance is only a gas.

Changes of State
1. Melting- Solid to a liquid
2. Freezing- Liquid to solid
3. Vaporization- Liquid to gas
    -Boiling – gas bubbles all over happens throughout
    -Evaporation- happens at the surface
4. Condensation- Gas to Liquid
5. Sublimation- Solid to Gas
6. Deposition- Gas to Solid

Physical Heat (Q)

When heating a substance without a change in state-
$Q=mC\Delta T$

Specific heat capacity (C or C_P)-the amount of energy required to raise the temperature of 1 gram of a substance 1 degree Celsius. The rate of heat transfer. (J/g˚C)

1 calorie = 4.18 Joules
1000 Joules = 1kJ
1000 calories = 1 kilocalorie = 1 Calorie (food calorie)
C_{H2O}= 4.18 J/g*˚C = 1 cal/g*˚C

The lower the value of the specific heat capacity the faster heat is transferred
The higher the value of the specific heat capacity the slower heat is transferred

At changes of state other equations must be used to determine the amount of energy is used.

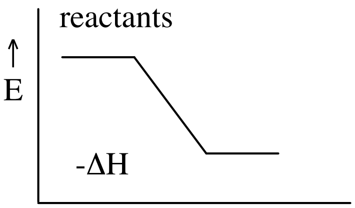
At Melting Point-
$\mathbf{\Delta H = H_f m}$
Heat of Fusion-(H_f)- The amount of energy required to melt 1 gram of a solid (Joules/Cal).

At Boiling Point-
$\mathbf{\Delta H = H_v m}$
Heat of Vaporization-(H_v)- The amount of energy required to boil 1 gram of a solid. (Joules/Cal)

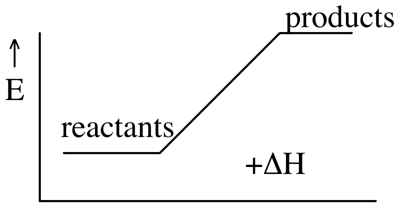
Chemical Heat (ΔH)

	Heat Lost/ Gained	**Heat Released/ Absorbed**	**Heat Reactant/ Product**	**ΔH Value**	**Thermometer Reading**
Endothermic	Gained	Absorbed	Reactant	Positive	Decrease
Exothermic	Lost	Released	Product	Negative	Increase

Energy Profiles



Exothermic



Endothermic

Activation Energy- the energy need to start an exothermic chemical reaction.
For an Exothermic reaction- add heat/spark/

Heat Stoichiometry-
Given an amount of A and the balanced equation with a ΔH value.
$\#A + B \rightarrow C \qquad \Delta H$

To calculate the heat used or released by the reaction.

$$\text{g A} \times \frac{\text{1mol A}}{\text{Molar Mass g A}} \times \frac{\Delta H \ \text{kJ}}{\text{\# mol A}} = \text{kJ}$$

Gas Notes

Kinetic Molecular Theory- all particles have motion.

	Solids	Liquids	Gases
Condensed/Non Condensed	Condensed	Condensed	Non-Condensed
Distance between particles	Closely packed	Closely packed	Large distance between particles, Loosely pack
Compressibility	No	No	Yes
Definite Shape	Yes	No	No
Definite Volume	Yes	Yes	No
Take shape of container	No	Yes	Yes
Will it expand to take the volume of container	No	No	Yes
Energy of Particle (Speed)	Low	Medium	High
Density	High	High	Low

<u>Gas Variables</u>

V = Volume- space occupied by a sample of gas (mL or L)
P = Pressure- force exerted over an area (kPa, atm, mm Hg, torr)
T = Temperature- measure of average kinetic energy of the particles of a gas sample (˚C, K)
n = Moles- the number of moles of a sample (moles)

Variable	Constants	Relationships	Between Variable
P, V	n,T	P↓V↑ P↑V↓	Inverse
T, V	n, P	T↑V↑ T↓V↓	Direct
T, P	n, V	T↑P↑ T↓P↓	Direct
n, P	T, V	n↑P↑ n↓P↓	Direct
n, V	T, P	n↑V↑ n↓V↓	Direct
n, T	P, V	n↓T↑ n↑T↓	Inverse

<u>Conversions</u>

Temperature Units
$T_{Kelvin} = T_{Celsius} + 273$
$T_{Celsius} = T_{Kelvin} - 273$

Pressure Units
1 atm = 760 torr = 760 mm Hg = 101.325 kPa

For all gas law equations, temperatures must be in Kelvin!

<u>STP- Standard Temperature and Pressure</u>

Standard Temperature = 0˚ Celsius = 273 Kelvin
Standard Pressure = 1 atm = 760 mm Hg = 101.325 kPa

Procedure for using gas laws

1. Write down all given values.
2. Determine the correct gas law to use.
3. Make any necessary conversions.
4. Solve the equation for the unknown value.
5. Input the known quantities into the equation.
6. Solve.

Boyle's Law

When T & n are held constant

P & V are inversely proportional

$$\mathbf{P_1V_1=P_2V_2}$$

Charles' Law

n & P are held constant

V & T are directly proportional

$$\mathbf{\frac{V_1}{T_1} = \frac{V_2}{T_2}} \quad \text{or} \quad \mathbf{V_1T_2 = V_2T_1}$$

Gay Lussac's Law

n & V are held constant

P & T are directly proportional

$$\mathbf{\frac{P_1}{T_1} = \frac{P_2}{T_2}} \quad \text{or} \quad \mathbf{P_1T_2 = P_2T_1}$$

Combined Gas Law

n is held constant

$$\mathbf{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}} \quad \text{or} \quad \mathbf{P_1V_1T_2= P_2V_2T_1}$$

Ideal Gas Law

R = 0.0821 atm • L / mol • K

$$\mathbf{PV = nRT}$$

To solve for Molar Mass

$$\mathbf{M.M. = \frac{mRT}{PV}}$$

Dalton's Law of Partial Pressure- total pressure is equal to the sum of the pressure of each individual gas

$P_{Total}= P_1 + P_2 + P_3 \ldots$

Bonding Notes

Valence Electrons- an electron in the outer most energy levels of an atom, where it can participate in bonding

Octet Rule- the tendency for main group elements to gain, lose, or share electrons so that their s and p orbitals are full with 8 electrons

Main Group Elements	Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group18
Electron Ending	s^1	s^2	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6
Valence Electrons	1	2	3	4	5	6	7	8

Ionic Compounds (salts)- compounds that form due to the transfer of electrons, usually bonding a **metal to a non-metal**, smallest unit is called a formula unit.

Covalent (Molecular) Compounds- compounds that form due to the sharing of electrons between elements, usually bonding **non-metals to non-metals**, smallest unit is called a molecule.

Polar Bond- a polar covalent bond arises from the uneven sharing of electrons (bonded atoms that have different electronegativities)

Non-Polar- a non-polar bond has even sharing of electrons in a covalent bond (bonded atoms are the same element or have the same electronegativities)

Polar Molecule- molecules with a slightly positive end and a slightly negative end.

Non-Polar Molecules- molecules with no net charge.

All of the symmetrical molecular shapes (linear, trigonal planar, tetrahedral) yield non-polar molecules as long as all of the outer atoms are identical.

If the outer atoms are different the molecules are usually polar.

The presence of non-bonding electron pairs on the central atom usually lowers the symmetry enough so that a polar molecule results.

Polar molecules mixes with Polar molecules and Non-Polar molecules mixes with Non-Polar molecules.

Intermolecular Forces- the attraction between molecules. The strength of a molecules' intermolecular forces can affect melting point and boiling point.

London Dispersion Forces (LDF, "Big molecule force")
-A weak intermolecular force due to a temporary dipole that is induced by the uneven movement of electrons.
-"The bigger the molecule the stronger the attraction"
-LDF are the only intermolecular forces that can be exhibited on non-polar molecules.

Dipole-Dipole attractions- The attraction of polar molecules to each other due to partial opposite charges.

Hydrogen Bonds- not a true **intramolecular "**bond". A very strong dipole-dipole attraction that occurs between hydrogen of one molecule and a very electronegative atom (ex. N, O, F) of another molecule.

Relative Strength of Intermolecular Forces **for similar sized molecules-**

LDF <<<<<<<<<< Dipole-Dipole <<<<<<<<<< Hydrogen Bonds

Polar Molecules

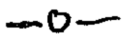
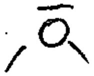
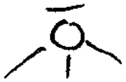
If molecule has:

-one type of atom it is **non-polar.**

-two types of atoms and no unshared pairs of electrons around the central atom, it is **non polar.**

-two types of atoms and has unshared pairs of electrons around the central atom , it is **polar.**

-three types of atoms it is **polar.**

# of e- domains around central atom	Electron Domain Geometry (E.D.G.)	# Bonding Domains	# Non-bonding Domains	3-D Structure	Molecular Geometry (M.G.)	Approx. Bond Angle	Example
2	LINEAR	2	0		LINEAR	180°	
3	TRIGONAL PLANER	3	0		TRIGONAL PLANER	120°	
		2	1		BENT	~120°	
4	TETRAHEDRAL	4	0		TETRAHEDRAL	109.5°	
		3	1		TRIGONAL PYRAMIDAL	107.5°	
		2	2		BENT	104.5°	

ORGANIC CHEMISTRY – HYDROCARBON CHEMISTRY

Carbon will always form ____ bonds in all of these structures. Why??

NAME	BOND TYPE	GENERAL FORMULA	EXAMPLES (note ending of each name)	SATURATED OR UNSATURATED	USE
ALKANES	SINGLES	C_nH_{2n+2}		SATURATED	BURN
CYCLO-ALKANES	SINGLE	C_nH_{2n}		SATURATED	BURN
ALKENES	DOUBLE	C_nH_{2n}		UNSATURATED	BUILD
ALKYNES	TRIPLE	C_nH_{2n-2}		UNSATURATED	BUILD

ABC's of naming organic compounds

Meth- 1

Eth- 2

Prop- 3

But- 4

Pent- 5

Hex- 6

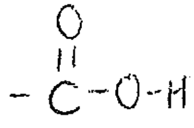
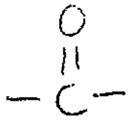
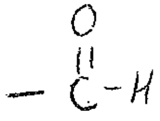
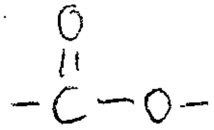
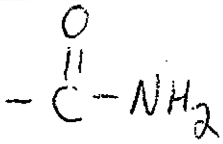
Hept- 7

Oct- 8

Non- 9

Dec- 10

SUBSTITUTED HYDROCARBONS -

NAME OF FUNCTIONAL GROUP	STRUCTURE
HALOGENATED COMPOUNDS	$-X$ $X = F, Cl, Br, I$
ALCOHOLS	$-O-H$
CARBOXYLIC ACIDS	 $-\overset{\overset{O}{\|}}{C}-O-H$
KETONES	 $-\overset{\overset{O}{\|}}{C}-$
ALDEHYDES	 $-\overset{\overset{O}{\|}}{C}-H$
ESTERS	 $-\overset{\overset{O}{\|}}{C}-O-$
ETHERS	$-O-$
AMINES	$-NH_2$
AMIDES	 $-\overset{\overset{O}{\|}}{C}-NH_2$

PART 2 – SUBSTITUTED HYDROCARBONS

HELP WITH NAMING:

FUNCTIONAL GROUP	*HINT*	*EXAMPLES OF NAMES*
Alcohols	"root name of alkane" + "*ol*"	Ethanol, Heptanol
Carboxylic acids	"root name of alkane" + "*oic acid*"	Ethanoic acid, Heptanoic acid
Aldehydes	"root name of alkane" + "*al*"	Ethanal, Heptanal
Ketones	"root name of alkane" + "*one*"	Propanone, Heptanone
Amines	"root name of alkane" + "*amine*"	Ethanamine, Heptanamine
Amides	"root name of alkane" + "*amide*"	Ethanamide, Heptanamide
Halogenated hydrocarbons	# of the C attached to the halogen + "fluoro, chloro, bromo, or iodo" + name of alkane	2-chloro-propane 3-fluoro-heptane
Esters**	End in "oate"	Methyl butanoate
Ethers**	Contain an "oxy" group	1-ethoxypropane

**You will not need to name these. These names will be provided by your instructor. Note the unique names.

pH

$$\begin{tikzcd}[column sep=huge, row sep=huge]
{[H_3O^+]} \arrow[r, shift left, "1x10^{-14}/[H_3O^+]"] \arrow[d, shift right, "- \log [H_3O^+]"'] & {[OH^-]} \arrow[l, shift left, "1x10^{-14}/[OH^-]"] \arrow[d, shift right, "- \log [OH^-]"'] \\
pH \arrow[u, shift right, "10^{-pH}"'] \arrow[r, shift left, "14 - pH"] & pOH \arrow[l, shift left, "14 - pOH"] \arrow[u, shift right, "10^{-pOH}"']
\end{tikzcd}$$

Molarity- Measurement of concentration of a solution

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

Acid/ Base

	Definition	**pH range**	**Typical form**	**Examples**
Acids	Donates H^+	0-6	Starts with H	HCl, H_2SO_4, $HHC_2H_3O_2$
Bases	Accepts H^+	8-14	Ends with OH	$NaOH$, $Ca(OH)_2$, NH_3, NH_4Cl

$pH = -\log [H_3O^+]$ $[H_3O^+] = 10^{-pH}$
$pOH = -\log [OH^-]$ $[OH^-] = 10^{-pOH}$

Strong Acids and Bases completely dissociate in water; Weak Acids and Bases do not.

Equilibrium

LeChatelier's Principle- When a system at equilibrium is disturbed, the system will shift appropriately in order to regain equilibrium.

Chemical systems at equilibrium will:
- shift AWAY from an INCREASE.
- shift TOWARDS a DECREASE.

Kinetics

What factors can increase the rate of a reaction?
1. Temperature
2. Concentration
3. Surface Area
4. Catalyst
5. Nature of Reactants

Electrolytes- ions that can conduct electricity in a solution.
- Strong Electrolytes- Split completely into ions when dissolved in water.
  - Ex: Ionic comounds, stong acids, strong bases
- Weak Electrolytes- Split partially into ions when dissolved in water.
  - Ex: Weak Acids, Weak Bases
- Non- Electrolyte- Does not split into ions when dissolved in water.
  - Ex: Molecular (Covalent) Compounds

Redox Reactions

Oxidation is loss of electrons (charge is increased) ex. $Na^0 \rightarrow Na^+ + e^-$
Reduction is gain of electrons (charge is reduced) ex: $Cl_2 + 2e^- \rightarrow 2Cl^-$

Colligative Properties

Solute particles in a solution can:
- Elevate (raise) the boiling point
- Depress (lower) the freezing point

Nuclear Chemistry

- Nuclear fission is the splitting of the nuclei of atoms.
- Nuclear fusion is the combination of the nuclei of atoms.