Unit 1: Organic Chemistry

Hydrocarbons

Hydrocarbon:

Things to know:

- Number of carbons give base of name
- Follow rules
- C_nH_{2n+2}

Number of carbons:

meth, eth, prop, but, pent, hex, hept, oct, non, dec

Prefix + root + suffix

Numbers location

Di, tri

Alkane: single bonded hydrocarbon

Alkene: double bonded hydrocarbon

Alkyne: triple bonded hydrocarbon

Complete structural diagram

Condensed Structural diagram

Line Structural diagram

Rules:

1. Prefix, branch position(s), name of branches
2. Root number of carbon atoms on parent chain
3. Longest chain of carbons atoms is the parent
4. Parent chain must include double/triple bonds
5. Start the parent chain closest to a branch
6. Name each branch and its location
7. Sort branches in alphabetical order
8. If > 1 of the same branch adds multiplying prefix (di, tri ..)
9. Comma between numbers, hyphen between numbers/letters

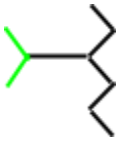
Branches:

-F	fluoro
-Br	bromo
-Cl	chloro
-I	iodo
-OH	hydroxy
$-NO_2$	nitro
$-NH_2$	amino
-COOH	carboxy

Aliphatics, Aromatics, Cyclic Compounds

Carbons join in a ring shape are Cyclic Compounds, prefix (Cyclo)

Alkyl branches:

Isopropyl Y form branch	
N-butyl Butyl branch attached at position 1	
sec-butyl 2nd grade attached at position 2	
tert-butyl 3rd grade + shape	

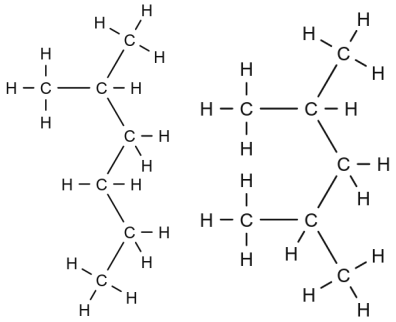
Benzene: C_6H_6, electrons are shared, planar ring

Branch	Location
Para (p)“across”	1,4
Meta (m)“middle”	1,3
Ortho (o)“next door”	1,2

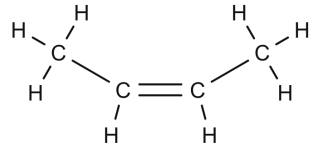
When benzene is a branch it is phenyl

Isomers

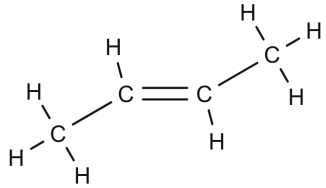
When a molecule has different possible structures it is called an isomer
Constitutional Isomers are molecules that have different bonding with the same formula. Ex. C7H16 can be shown:
2-methylhexane / 2,4-dimethylpentane



Stereoisomers: molecules that have same bonds but different arrangements
Cis-trans are possible for alkenes that posses two branches



Cis: Hydrogens on same sides



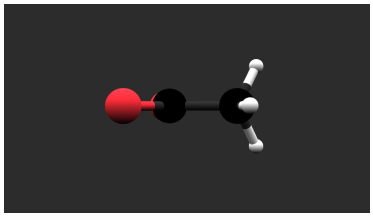
Trans: Hydrogens on opposite sides

Complex Organic Compounds:

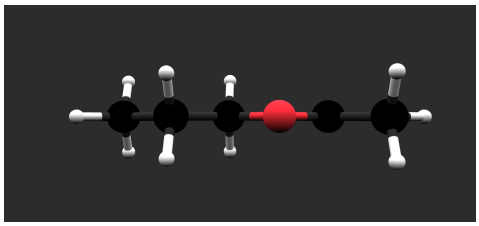
Alcohols, Ethers, Aldehydes, Ketones, Carboxylic Acids, Esters, Amines, Amides

LISTED IN ORDER OF NAMING IMPORTANCE **HIGHEST TO LOWEST:**

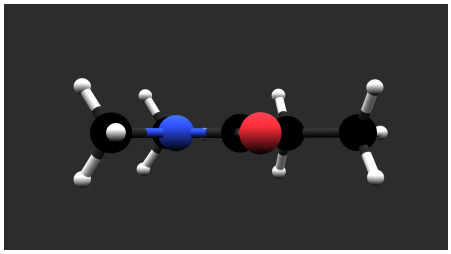
Carboxylic Acids: A hydrocarbon with a carboxyl group (C(=O)OH) attached to an alkyl group
Name parent
Replace -e with -oic acid
The carboxyl group always given position 1



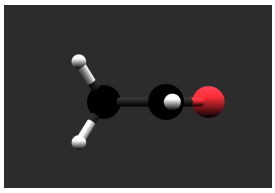
Esters: A hydrocarbon with a double bonded oxygen and a single bonded oxygen (O-C=O) attached to two alkyl groups
1.Longest parent chain with C=O group
2.Replace -e with -oate
3.Put space between alkyl group name attached and the parent chain



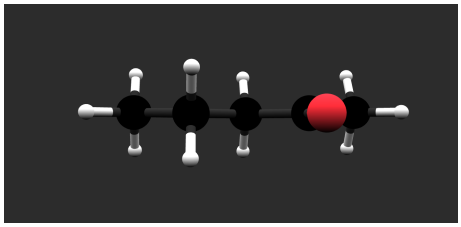
Amides: A hydrocarbon with a bonded nitrogen with a double bonded oxygen (N-C=O) attached to 1 to 3 alkyl groups
Longest that contains C=O is parent
Replace -e with -amide
1 degree, no alkyl groups attached to the nitrogen
2 degree, 1 alkyl groups attached to the nitrogen
3 degree, 2 alkyl groups attached to the nitrogen



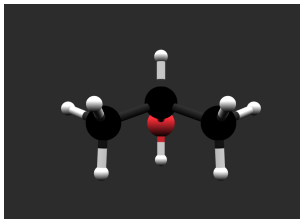
Aldehydes: A hydrocarbon with double bonded oxygen (=O), single bonded hydrogen (H)
Name parent with C=O group, position 1 no numbers
Replace -e with -al



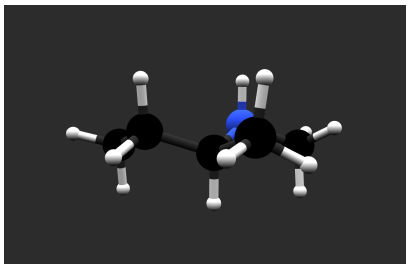
Ketones: A hydrocarbon with a double bonded oxygen (C=O) attached to two alkyl groups
Longest chain with C=O os parent
Replace -e with -one and is needed, di, tri
Number if more than 4 carbons



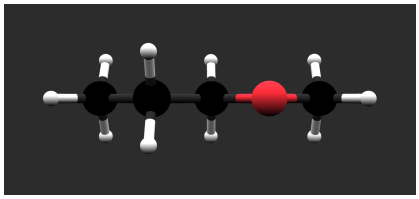
Alcohols: A hydrocarbon with a main hydroxyl (OH) branch
1.Longest chain with OH group
2. Replace -e with -ol
3. Add number if needed and multiplying prefix to -ol



Amines: A hydrocarbon with a bonded nitrogen (N) attached to 1 to 3 alkyl groups
1.Longest attached to nitrogen (-N)
2.Replace -e with -amine
3.N- instead of number locations is gives N to branches on the nitrogen



Ethers: A hydrocarbon with an oxygen (-O-) between two alkyl groups
1.Longest chain is parent
2.Second alkyl group is branch but replace -yl with -oxy, give position number
3.Alkoxy group + parent alkane



Akenes
Alkynes
Alkyl Halides
Alkanes (lowest priority)
BOTTOM OF ORDER OF IMPORTANCE

Types of Reactions:
Combustion, Addition, Substitution, Elimination, Condensation, Hydrolysis, Oxidation, Reduction, Synthesis and Esterification

Markovinkov's Rule: Hydrogen will bond with a carbon with the most hydrogens
Combustion: All hydrocarbons undergo combustion in the presence of oxygen

$$C_nH_{2n+2} + O_2 \rightarrow CO_2 + H_2O$$

Addition: Breaks double/triple bonds and adds atoms to them
$CH_2CHCH_2CH_3 + H_2 \rightarrow CH_3CH_2CH_2CH_3$
Substitution: Hydrogen atom being replaced with a halide atom (needs catalyst)
$CH_3CH_2CH_3 + Cl_2 \rightarrow CH_3CH(Cl)CH_3 + HCl$
Elimination: A double/triple bond formed out of the removal of atoms
$CH_3CH(Cl)CH_3 + NaOH \rightarrow CH_3CHCH_2 + NaCl + H_2O$
Condensation: make water from synthesis of amides
Carboxylic acid + ammonia → amide
$CH_3COOH + NH_3 \rightarrow CH_3ONH_2 + H_2O$
Hydrolysis:
Amines: split water to from carboxylic and amine from an acid/base and amide
$CH_3ONH_2 + H_2O \rightarrow CH_3COOH + NH_3$
Esters: split water with ester into carboxylic acid and alcohol
$CH_3(O)OCH_3 + H_2O \rightarrow HCOOH + CH_3OH$
Oxidation: Carbon forms more bonds with oxygen and less with hydrogen using an oxidizing agent
1 degree Alcohol → aldehyde
$CH_3CH_2OH \ [O]\rightarrow CH_3COH$
2 degree Alcohol → ketone
$CH_3CH(OH)CH_3 \ [O]\rightarrow CH_3COCH_3$
Aldehyde → carboxylic acid
$CH_3COH + H_2O \ [O]\rightarrow CH_3CH_2OH$
Reduction: Aldehyde reduced with hydrogen gas to form 1 degree alcohols
Ketones reduced with hydrogen gas to form 2 degree alcohols
Aldehyde [H]→1 degree Alcohol
$CH_3CH_2OH \ [H]\rightarrow CH_3COH$
Ketone [H]→2 degree Alcohol
$CH_3COCH_3 \ [H]\rightarrow CH_3CH(OH)CH_3$
Synthesis: Halide reacts with ammonia to synthesize an amine
Primary amine
$CH_3Cl + NH_3 \rightarrow CH_3NH_2 + HCl$
Secondary amine
$CH_3Cl + NH_2CH_3 \rightarrow CH_3NHCH_3 + HCl$
Tertiary amine
$CH_3Cl + CH_3NHCH_3 \rightarrow CH_3N(CH_3)CH_3 + HCl$
Esterification: Carboxylic acid and alcohol react to form water and an ester
$HCOOH + CH_3OH \rightarrow CH_3(O)OCH_3 + H_2O$

Properties of Organic Compounds

Solubility: affected by polarity nonpolar molecule will be insoluble in H_2O
Boiling point: affected by types and amount of IMF and shape. L.D. forces are in all atoms. More atoms higher boiling point b/c more IMF. Polar and dipole-dipole have higher boiling points.
Sate: more atoms more solid organic compounds get

Unit 2: Structures and Properties of Matter

Atomic History (Models of the atom)

Aristotle only four elements: earth, air, fire and water
Democritus (300 B.C): atoms are made of tiny, indivisible things
Dalton's Atomic Theory: All matter made up of tiny particles ("**atoms**")
Thomson Model of the Atom: **electrons** could be emitted from matter
Robert Millikan (1909): calculate relative **charge and mass** of the electron
Ernest Rutherford (1911): discovered **atom** is mostly **empty space**
James Chadwick (1932): Proved the existence of the neutron.
Neils Bohr: - The Planetary Model
Max Planck (1900): Quantum Hypothesis made to explain observations.
Einstein (1905): the energy of the photon is transferred to the electron

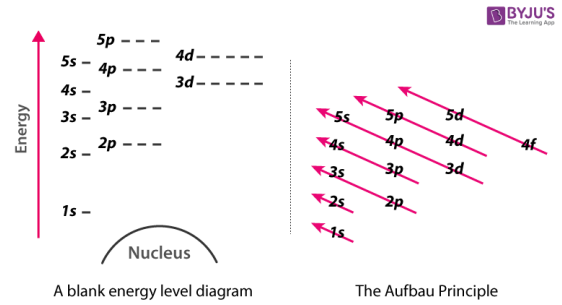
Quantum Model of the Atom (quantum numbers)

Principal (n) - Values from (1,2,3…), size and energy of orbital, n is number of types of orbitals, n^2 number of orbital in ground state

Secondary (*l*) 0 to n-1, shape of orbital, *l*=0 (s) sharp, *l*=1 (p) principal
Magnetic (m_l) *l* to -*l*, orientation of orbital, number of suborbitals
Spin (m_s) direct and magnitude of electrons axial spin, 2 possible spins, 2 electrons per orbital

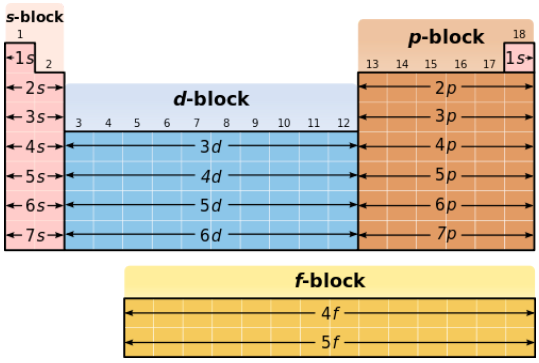
Energy Level Diagrams

n values increase upwards as energy increases, orbitals distance gets smaller further up, overlap at high
Aufbau Principle:
Energy level must be filled before moving up next level
Pauli Exclusion Principle:
No 2 e^- in an atom can have 4 quantum numbers, on 2 e^- can exist per orbital
Hund's Rule:
1 e^- is placed in each suborbital



A blank energy level diagram | The Aufbau Principle

Electron Configuration



n *l* #
n = energy level
l = sublevel
= number of electrons
Iron electron configuration:
1s2 2s2 2p6 3s2 3p6 4s2 3d6
Abbreviated form is from last noble gas
[Ar] 4s2 3d6

Chemical Bonding

Ionic Compounds: made of ionic bonds, usually between metals and nonmetals, bonded in a crystal lattice structure, ΔEN >= 1.7
Molecular Compounds: covalent bonds between 2 nonmetals, ΔEN = 0 to 0.5 is nonpolar covalent / ΔEN = 0.5 to 1.7 is polar covalent
Simple covalent: each atom shares 1 e^- for bonding
Coordinate covalent: 1 atom shares 2 e^- for bonding

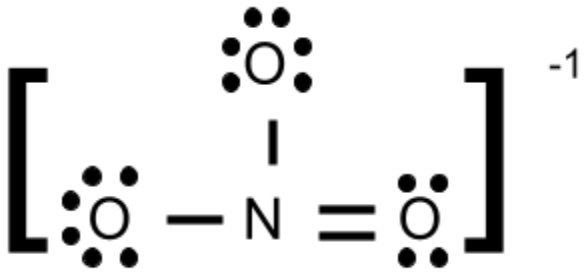
Lewis Structures (including resonance)

Steps:

1. Find number of electrons for every atom and sum them
2. Draw central atom (usually only 1 atom is central)
3. Draw single bonds first (take away 2e^- for every bond)
4. Put remaining e^- on the outer atom to fill their octets
5. Put leftover e^- on middle atom in pairs
6. If central atom does not follow octet rule add double/triple bonds by changing around electrons
7. Check for resonance structures
8. Find vector of molecules by making the atom with the highest ΔEN the negative side and the opposite the positive side
9. Draw vectors if molecule is polar and dipoles
10. Put square brackets around the structure if it is an ion and a charge number

Resonance: when a lewis structure can be changed while not affecting its properties

NO_3^-
N: $5e^-$
O: $6e^-$ x3
= $23e^-$



Formal Charge

F.C. = number of valence e^- - number of unbonded e^- - number of bonds
Structure that has the most zeros is best
Overall number of valence should be equal to zero or if an ion its charge
-O: 6 - 6 - 1 = -1
=O: 6 - 4 - 2 = 0
N: 5 - 0 - 4 = +1

Molecule Polarity and VSEPR Theory

Bond Polarity: disturbing of e^-across single bond
Molecular polarity: is polarity of sum of all bonds in a molecule
Polar molecule: unequal e^- arrangement producing a negative charge at one end and a positive charge at the other
Non-polar molecule: equal e^- arrangement producing no net charge
The type of bond and the location of the bond in the molecule determines if it is polar/non-polar
Get ΔEN
Example:
N: 3.04
O: 3.44
ΔEN = 3.44 - 3.04 = 0.40
∴ bond nonpolar and molecule is nonpolar

VSEPR SUMMARY

Number of electron pairs around central atom		Full description of the molecule					
(BONDING (E))	(LONE (E))	Example	Bond angles °	Geometry of Electron Pairs	Geometry of Atoms	3D Shape	Type
2	0	$BeCl_2$	180	Linear	**Linear**		AB_2
3	0	BF_3	120	Trigonal planar	**Trigonal Planar**		AB_3
2	1	SO_2	Slightly less than 120	Trigonal planar	**Bent or V Shaped**		AB_2E
4	0	CH_4	109.5	Tetrahedral	**Tetrahedral**		AB_4
3	1	NH_3	107.5	Tetrahedral	**Trigonal Pyramidal**		AB_3E
2	2	H_2O	104.5	Tetrahedral	**Bent or V Shaped**		AB_2E_2
5	0	PCl_5	120 in plane, 90 perpendicular to plane	Trigonal bipyramidal	**Trigonal Bipyramidal**		AB_5
4	1	SF_4	Complex	Trigonal bipyramid	**Seesaw**		AB_4E
3	2	ClF_3	Approx. 90	Trigonal bipyramidal	**T-Shaped**		AB_3E_2
2	3	XeF_2	180	Trigonal bipyramid	**Linear**		AB_2E_3
6	0	SF_6	90	Octahedral	**Octahedral**		AB_6
5	1	BrF_5	Approx. 90	Octahedral	**Square Pyramidal**		AB_5E
4	2	XeF_4	90	Octahedral	**Square Planar**		AB_4E_2

∴ trigonal planar for NO_3^-

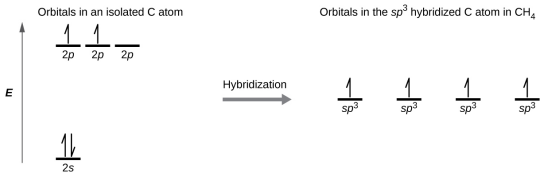
Valence Shell Electron Pair Repulsion
Determines structure around atom by minimizing repulsive force
Bonded and unbonded e^- repel each other far away to minimize repulsive force
Lone pair e^- pushes bonding pairs of e^- together changing angles because they are more repulsive
Treat multiple bonds as single bonds
A = middle atom
X = surrounding atoms
ε = lone pair e^-
∴ AX_3 for NO_3^-

Hybridization

Orbital hybridization can explain the shape of the molecule
Carbon as example can make 4 bonds because it promotes one e^- from 2s to the 2p orbital
Hybrid orbitals only exist when bonding occurs to form a molecule, not in atoms

Sigma bonds: occurs when there is end to end overlap of orbitals - a single covalent bond

Pi bonds: occurs when there is overlap above and below the plane of the bond, “sideways”, overlap of p orbitals - o double or triple bond contains a pi component bond



Unit 3: Energy Changes & Rates of Reactions

Thermochemistry

Thermochemistry means study of change in energy

Energy measured in joules (J) is the ability to do work

Potential Energy is the energy of something because of its position or composition

Kinetic Energy is the energy of an object because of its motion

Thermal energy is the total amount of kinetic and potential energy

Heat is the transfer of thermal energy from object to object

Temperature is a measure of the average kinetic energy

The Law of Conservation of Energy: energy cannot be created or destroyed, it only can be converted

A group of reactants and products being studied is a chemical system

Anything that is not the system is the surroundings

An open system is a system open for energy and matter to move in and out freely

A closed system is a system where energy can enter/leave the system but not matter

An isolated system is an ideal system where matter or energy cannot move in or out

An exothermic chemical reaction is when energy is released

An endothermic chemical reaction is one in which energy is absorbed

Heat Capacity (Q) and Calorimetry

Specific heat capacity (c) is the amount of thermal energy need to increase the temperature of 1g of a chemical by 1 ^{o}C

Specific heat capacity depends on state and substance type

Calorimetry is the process of recording the thermal energy change of a substance

Calorimetry calculations depend on 4 assumptions:

1. Any thermal energy moved from the calorimeter to the outside is not counted
2. Any thermal energy absorbed by the calorimeter is not counted
3. All dilute, aqueous solutions have the same density (1.00g/mL) as water
4. Diluted, aqueous solutions have the same specific heat capacity (4.184 J/(g•^{o}C)) as water

$Q = mc\Delta T$

Q is the total amount of thermal energy/ heat

m is the mass in grams

c is the specific heat capacity in J/(g•^{o}C)

ΔT is the temperature change of substance

$T_f - T_i$

$Q_{system} = - Q_{surroundings}$

An increase in the temperature of the water indicates an exothermic rxn, whereas a decrease in the temperature of the water indicates an endothermic rxn

300g of an unknown metal at 400°C is dropped into 500mL of water at 21.5°C. The water is heated up to 74.6°C, what is the heat capacity of the metal?

$Q=(500g)(4.184\ J/g°C)(53.1°C)$

$Q_{water}=111{,}085.2\ J$

$Q_{metal}=-111{,}085.2\ J$

$c=-111{,}085.2J / (300g)(-840.4°C)$

c=0.44060447 J/g°C
c=0.4406 J/g°C

Enthalpy (ΔH)

Enthalpy (H): is total amount o f thermal energy in a system ro substance
Total kinetic energy + total potential energy = thermal energy
Cannot measure enthalpy of a system directly
Enthalpy change (ΔH) is the energy transferred to or from the surroundings during a chemical or physical change
$\Delta H_{system} = +/-[Q_{surroundings}]$
Endothermic is when ΔH is negative
Exothermic is when ΔH is positive
Signs are always required for enthalpy
Molar enthalpy change (ΔHx) is the enthalpy change of 1 mole of a substance, units kJ/mol
4 way to show enthalpy change:
Thermochemical equations with energy terms: $1/2\ O2_{(g)} + H2_{(g)} \rightarrow H2O_{(l)} + 285.8\ kJ$
Thermochemical equations with ΔH:
$1/2\ O2_{(g)} + H2_{(g)} \rightarrow H2O_{(l)}$
ΔH = -285.8 kJ
Molar Enthalpy (ΔHx):
$1/2\ O2_{(g)} + H2_{(g)} \rightarrow H2O_{(l)}$
ΔHx = -285.8 kJ/mol
Potential energy Diagrams:
ΔH=(n)(ΔHx)
x = type of change

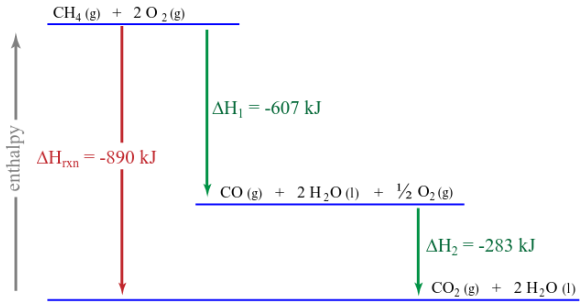
Hess's Law

Hess's La: enthalpy change (ΔH) is determined by initial and final conditions of a system
$\Delta H_{reaction} = \sum \Delta H_{steps}$
Hess's law is used when the reaction target is too slow, too fast, too dangerous, or not possible to do in a calorimeter.
Reverse rxn(flip) must also flip ΔH
Ensure rxn is balanced to mach target (multiply)

Example equation:

$$2C(s,\ \text{graphite}) + \cancel{2O_2(g)} \rightarrow \cancel{2CO_2(g)} \qquad \Delta H = -787.0\ kJ$$
$$H_2(g) + \frac{1}{2}\cancel{O_2(g)} \rightarrow \cancel{H_2O(l)} \qquad \Delta H = -285.8\ kJ$$
$$\cancel{2CO_2(g)} + \cancel{H_2O(l)} \rightarrow C_2H_2(g) + \frac{5}{2}\cancel{O_2(g)} \qquad \Delta H = 1301.1\ kJ$$
$$2C(s,\ \text{graphite}) + H_2(g) \rightarrow C_2H_2(g) \qquad \Delta H = 228.3\ kJ$$

Example chart:



Standard Enthalpy of Formation (ΔH°f)

Standard molar enthalpy: is the change in enthalpy (ΔH°f) when 1 mol is formed directly from its elements
Molar quantity: amount of energy required to form 1 mole of a substance from its elements in their standard states
ΔHf compounds can be found in a table
ΔH = ∑[(n)(ΔHf Products)] - ∑[(n)(ΔHf Reactants)]

Entropy (ΔS) (Spontaneous Change)

Entropy (S) is the amount of chaos of a system (higher entropy means more chaos)
ΔS = ∑[(n)(ΔS Products)] - ∑[(n)(ΔS Reactants)]

1. ΔS > 0
2. Increase in state
3. Number of moles products > reactants

Gibb's Free Energy (ΔG)

ΔG = ΔH -TΔS
If ΔG is + not spontaneous
If ΔG is - is spontaneous
Ideal values for a spontaneous reaction:
ΔH = - (exothermic)
ΔS = + (high entropy)
ΔG = -
Example:
ΔS = +90Kj ΔH = +340Kj T=300K

ΔG = 340-(300)(90)
ΔG = -26660
∴ rxn is spontaneous

Chemical Reaction Rates (Collision Theory and Factors that Affect Reaction Rates)

Reaction Rate: is the speed at which a reaction occurs
Collision theory: chemical reactions can occur only if reactants collide with the right orientation and enough energy to break and form bonds
Activation Energy: the minimum energy for a collision to be effective
5 factors that can cause this increase:

- Chemical nature of the reactants
- Concentration
- Surface Area
- Temperature
- Catalyst

Average Reaction Rate: is the change Δ in reactant or product concentration over a given time interval
Rate determining step: slowest elementary step
The units for average reaction rate are mol/L•s

Reaction Order	K units
0	mol/L x s
1	1/s or s-1
2	L/mol x s
3	L^2/mol^2 x s

Rate Law

Rate – measuring the change in [reactant] or [product] over time.
k is the Rate Constant
x and y are determined experimentally, and do not depend on stoichiometric

rate = $K[A]^x[B]^y[C]^z$
coefficients from balanced equation
total order: sum of the exponents
Average Reaction Rate: is the change Δ in reactant or product concentration over a given time interval

1. Using at least 3 sets of data for the reaction
2. Write out the rate law equation with the reactants used
3. Find 2 experiments where 1 of the reactants is constant but the other changes.
4. Compare the rates and the concentrations to solve for the exponent
5. Find two data points where the 2nd of the reactant remains constant and the other changes
6. Compare the rates and the concentrations to solve for the exponent
7. Using one set of experimental data, and the values for exponent, solve for k
8. Units of k can be found by keeping them in the equation or the table for K units

Unit 4: Chemical Systems and Equilibrium

Dynamic Equilibrium

Most reactions one-way reactions
Most reactions are mixture of both reactants and products
Forward rate: rate at which the products are formed.
Reverse rate: rate at which the reactants are re-formed.
Reaction almost never stops always in a state of equilibrium (back and forth)
Equilibrium: is when $f_{rate} = r_{rate}$
Equilibrium constant is K_{eq} or K_c
K_{eq} (K_c) = $[products]^x / [reactants]^y$

Percent Yield and Reaction Direction

Reactants can become products and products can become reactants
% yield = (Actual / theoretical) x 100%
No reaction (NR) < 1%
Reactants favored >1% but < 50%
Products favored >50% but < 99%
To Completion >99%
Chemical equation
Need to know states
Type of reaction
Balance
Givens
Convert to moles
Find molar ratio of solids
Moles of product solid
% yield = 4.84g / 5.90g x 100%
% yield = 82.033898305%

∴ reaction is product favored

Equilibrium Law (Keq)

$K_{eq} = [products]^{\#mol}/[reactants]^{\#mol}$
[ ] = concentration mol/L
Keq has no units
Keq and Kc are exchangeable
Keq is temperature dependent
Homogeneous equilibria: products and reactants are same phase
All gases
Heterogeneous equilibria: products and reactants are different phase
Liquid → gas
Liquid (*l*): volume does not change, incompressible, constant is concentration
Solids (s): incompressible, constant concentration
Gases (g): compressible, take shape and volume of container, concentration changes
No liquids in Keq formula because [liquid is constant]
Ions in solutions: aqueous solutions separate into ions, cancel out spectator ions
K >>> 1, ($K>=10^3$) forward reaction is preferred, equilibrium is to the right, more product forming
$Keq > 10^{10}$ reaction is complete
K <<< 1 ($K<10^{-3}$) reverse reaction is preferred, equilibrium is to the left, more reactant forming
$Keq > 10^{-10}$ no reaction happening
K close to 1, reactant and products at equilibrium
Keq is equilibrium concentration of chemicals with conditions
Strategy to find k_{eq}:
If question gives concentration at equilibrium and the equilibrium constant then solve for the missing concentration

Initial **C**hange **E**quilibrium Table if question starts with initially or similar

1. Setup ICE table
2. Use E row in K_{eq} equation
3. Solve for unknown variables

Example Approximation rule question:
$K_{eq} = 4.5x10^{-5}$

	$CO_{(g)}$	$Br_{2(g)}$	$COBr_{2(g)}$
I	2.20 mol/L	3.70 mol/L	0
C	-x	-x	+x
E	2.20 -x	3.70 -x	x

3.70 / 6.7x10-5 = 55223.8806 ∴ Yes it is more than 1000x difference, approximation

K_{eq} = [Products] / [Reactants]
$K_{eq} = [COBr_{2(g)})]^1 / [CO_{(g)}]^1[Br_{2(g)}]^1$
$6.7x10^{-5} = [x]^1 / [2.20 -x]^1[3.70 -x]^1$
$6.7x10^{-5} = x / (2.20)(3.70)$
$6.7x10^{-5} = x / 8.14$
$(6.7x10^{-5})(8.14) = x$
$x = 5.45x10^{-4}$ mol/L
$[COBr_{2(g)}] = x$
$[COBr_{2(g)}] = 5.45x10^{-4}$ mol/L
$[Br_{2(g)}] = 3.70 - x$
$[Br_{2(g)}] = 3.70 -5.45x10^{-4}$

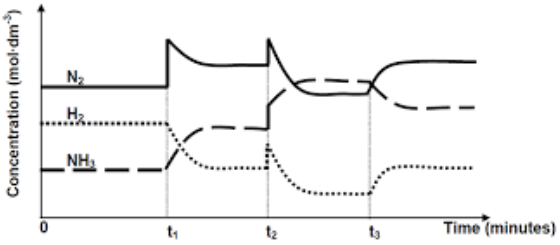
$[Br_{2(g)}] = 3.699455$ mol/L
$[Br_{2(g)}] = 3.70$ mol/L
$[CO_{(g)}] = 2.20 - x$
$[CO_{(g)}] = 2.20 - 5.45 \times 10^{-4}$
$[CO_{(g)}] = 2.199455$ mol/L
$[CO_{(g)}] = 2.20$ mol/L

Reaction Quotient (Q)

Find where the equilibrium of reaction is
Use rate law to determine Q
To find location of the equilibrium, compare Q to K_{eq}
If Q < K, reaction will move forward to get to equilibrium
If Q > K, reaction will reverse to get to equilibrium
If Q = K (within 0.1) the reaction is at equilibrium
Q equation is the same as K_{eq} but with the current concentration of all the reactants and products.
Example:
$K_{eq} = 4.0 \times 10^{-4}$
$[Br_{2(g)}] = 3.70$ mol/L
$[CO_{(g)}] = 2.20$ mol/L
$[COBr_{2(g)}] = 5.45 \times 10^{-4}$ mol/L
$Q = [COBr_{2(g)}]^1 / [CO_{(g)}]^1[Br_{2(g)}]^1$
$Q = [5.45 \times 10^{-4}] / [2.20][3.70]$
$Q = 6.6953317 \times 10^{-5} < K_{eq} = 4.0 \times 10^{-4}$
∴ rxn will proceed to get to equilibrium

Le Châtelier's Principle

A chemical system at equilibrium opposes change when it is disturbed
Concentration, moles, energy changes:
Increase in concentration of reactants will make the equilibrium shift to the right resulting in more products
Increase in concentration of products will make the equilibrium shift to the left resulting in more products
Decrease in energy of system will make the equilibrium shift to the left if endothermic and right if exothermic
Increase in energy of the system will make the equilibrium shift to the right if endothermic and left if exothermic
Volume and Pressure changes:
Boyle's law: P is inversely proportional to V
P↑ V↓ means equilibrium will shift to the side with least total moles
P↓ V↑ means equilibrium will shift to the side with most total moles
Catalyst: No effect on concentration
Inert Gases: No effect on concentration
Example Graph:



t_1:The concentration of nitrogen is increased
t_2: The pressure on the system is increased
t_3: The temperature is increased

Solubility Product Constant (Ksp)

Molar solubility: the # of moles of solute that dissolves in 1L of solvent at a certain temp.
Soluble compounds: completely dissociates into ions; no equilibrium
Insoluble compounds: only a tiny amount dissociates into ions; equilibrium
Solubility Product Constant (K_{sp}): product of molar solubility and more solubility higher k_{so}
General format K_{sp}: Solid (s) ⇌ Aqueous ions (aq)
$K_{sp} = [A_{(aq)}][B_{(aq)}] / ~~[C_{(s)}]~~$ NO (solid)!!
K_{sp} is a measure of the solubility of a chemical
If $K_{sp} << 0$ then low solubility; "insoluble"
If $K_{sp} >> 0$ then high solubility; "soluble"
Solve the same as K_{eq} problems.
$[I] = 4.3 \times 10^{-5}$ mol/L $[Ag] = 7.4 \times 10^{-6}$ mol/L
$K_{sp} = [A][B]$
$K_{sp} = [4.3 \times 10^{-5}][7.4 \times 10^{-6}]$

$K_{sp} = 3.182\times10^{-10}$

Predicting if a precipitate will form:

Use ksp and Q

If $Q = K_{sp}$:

- At equilibrium
- No precipitate
- saturated solution

If $Q > K_{sp}$:

- Right of equilibrium
- Precipitate
- Super saturated
- Many ions

If $Q < K_{sp}$:

- Left of equilibrium
- No precipitate
- Not saturated
- Not enough ions

K_{sp} and Q example:

[I] = 6.3x10-5 mol/L [Ag] = 8.4x10-6 mol/L

$K_{sp} = 3.182\times10^{-10}$

Q = [A][B]

$Q = [6.3x10^{-5}][\ 8.4x10^{-6}]$

$Q = 5.292x10^{-10}$

$Q > K_{sp}$ ∴ a precipitate will form

pH and pOH

Acid: chemical that releases hydrogen ion or hydronium ion when dissolved in water

Base: any chemical that releases hydroxide ions (OH^-) when dissolved in water

Bronsted-Lowery:

Brönsted-Lowry Acid: chemical that releases a hydrogen ion ($H^+_{(aq)}$) or hydronium ion ($H_3O^+_{(aq)}$) when mixed with water

$HF_{(aq)} + H_2O_{(aq)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$

Acid Base Conj. base Conj. Acid

Brönsted-Lowry Base: chemical that accepts a hydrogen ion ($H^+_{(aq)}$) when mixed with water

$CN^-_{(aq)} + H_2O_{(aq)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$

Base Acid Conj. Acid Conj. base

Conjugate acid/base pair: 2 molecules/ions that transfer $[H^+]$ in an acid/base reaction

pH: measure of hydrogen ion concentration

Examples:

pH of HCl when [HCl] = 3.45x10-3 mol/L

$pH = -\log[H^+_{(aq)}]$

$pH = -\log[3.45x10^{-3}]$

pH = 2.45

Find $[H^+]$ of solution when pH = 4.54

$[H^+_{(aq)}] = 10^{-pH}$

$[H^+_{(aq)}] = 10^{-4.54}$

$[H^+_{(aq)}] = 2.88x10^{-5}$

pOH: measure of hydroxide ion concentration

pH of $Ca(OH)_2$ when $[Ca(OH)_2] = 6.25x10^{-2}$ mol/L

$pOH = -\log[OH^-_{(aq)}]$

$[H^+] = -\log[6.25x10^{-2}]$

$[H^+] = 1.20$

$[H^+] = 1.20 \times ½$

$[H^+] = 0.60$

What is the pH of a solution, pOH of 4.5?

pH + pOH = 14

pH + 4.5 = 14

pH = 9.5

Find $[OH^-]$ of solution when pOH = 2.63

$[OH^-_{(aq)}] = 10^{-pOH}$

$[OH^-_{(aq)}] = 10^{-2.63}$

$[OH^-_{(aq)}] = 2.34x10^{-3}$

Sources:

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