CHEM 123: Summary of Important Concepts
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Thermodynamics

Equation/Concept	Info About Equation/Concept
$PV = nRT$	This is the ideal gas law. It is the *equation of state* for a gas of point particles that do not interact with each other. P is the pressure of the system, V is the volume of the system, n is the number of moles of gas, R is the gas constant, and T is the absolute temperature.
$\Delta E = q + w$	This is the equation for the first law of thermodynamics. It states that the change in energy *of the system* is equal to the heat flow q into the system (positive q is heat flow *into* the system) plus the work done on the system w. Work done *on the system* corresponds to positive w.
$đw = -P_{\text{ext}}dV$	The most general definition of the increment of work. This equation is for the *improper* increment of work done *on* the system. It is improper because work is a path variable, so the increment of work done is not an exact differential.
$w = -\int_{V_i}^{V_f} P_{\text{ext}}dV$	The definition of total work done on the system, for a process. It is the integral from the initial volume to the final volume of the incremental amounts of work done along that path. This is the most general expression for work done.
$w = -nRT \ln\left(\frac{V_f}{V_i}\right)$	The is the work done on an *ideal gas* expanding/contracting *reversibly and isothermally*. Isothermal expansion means that the temperature is constant throughout the process.
$W = -P_{\text{ext}}\Delta V$	This is the work done on a system with constant external pressure
$\Delta E = q_V$	Change in internal energy when the volume of the system remains constant. The subscript just indicates that the heat is exchanged at constant volume.
$q = C\Delta T$	The heat flow into the system is the product of the heat capacity C and the change in temperature $\Delta T = T_f - T_i$. Note that this heat capacity is an extensive variable. This is only truly applicable for small changes in temperature, as C a temperature dependent variable.
$H = E + PV$	Definition of enthalpy.
$\Delta H = \Delta E + P\Delta V$	The change in enthalpy for a constant pressure process.
$\Delta H_{\text{rxn}}^{\circ} = \sum_{\text{P}} \Delta H_{\text{f,P}}^{\circ} - \sum_{\text{R}} \Delta H_{\text{f,R}}^{\circ}$	Hess's Law. This law applies equally well to any state variable, such as total energy, entropy, etc. It is a reflection of the fact that the change in a state variable does not depend on the path taken.
$\Delta S_{\text{Universe}} \geq 0$	The second law of thermodynamics. The equality is only satisfied for a reversible process, for all non-reversible processes the inequality is true.
$\Delta S = \frac{q_{\text{rev}}}{T}$	True for reversible heat exchange only. Often how changes in entropy are measured.
$G = H - TS$	Definition of Gibbs free energy.
$\Delta G^{\circ} = -RT \ln(K)$	The standard Gibbs free energy. ΔG° is the Gibbs free energy change per mole of reaction of pure reactants to pure products at standard conditions (298 K, 1 atm, solutions at 1 M concentration, pure liquids and solids). K is the equilibrium constant for that reaction. $\Delta G^{\circ} < 0$ indicates products are favoured; $\Delta G^{\circ} > 0$ indicates reactants are favoured.
$S = k_B \ln(\Omega)$	The definition of entropy for an isolated system. Ω is the number of possible microstates and k_B is the Boltzmann constant, which is included to convert the units.
$\Delta G = \Delta H - T\Delta S$	The change in Gibbs free energy for a process done at constant temperature and pressure.
$\Delta G = \Delta G^{\circ} + RT \ln Q$	This equation can be used to predict the direction of a reaction. Q is called the "reaction quotient", and is the same as the equilibrium constant K, but calculated when a reaction has not yet reached equilibrium. $\Delta G < 0$ indicates that there are too many reactants and the reaction will favour products. $\Delta G > 0$ indicates that the reaction will proceed to the reactants side.
$K = \frac{[\text{Products}]}{[\text{Reactants}]}$	The definition of the equilibrium constant. Square brackets indicate concentrations, and the concentrations should be raised to the power of their stoichiometric coefficients.
$\Delta G = -T\Delta S_{\text{Universe}}$	This equation applies at constant temperature and pressure for a system. It is the reason that the change in Gibbs free energy is a measure of spontaneity under those conditions. $\Delta G < 0$ corresponds to an increase in the entropy of the universe when T and P are held fixed.

Equation	Info About Equation/Concept
$\ln K = \frac{-\Delta H^\circ}{R}\frac{1}{T} + \frac{\Delta S^\circ}{R}$	Van't Hoff plot: plot $\ln K$ vs $\frac{1}{T}$. The slope $m = \frac{-\Delta H^\circ}{R}$ and the y-intercept $b = \frac{\Delta S^\circ}{R}$. Using these plots assumes that the standard change in enthalpy and standard change in entropy are independent of T. This is only true for small changes in T.
S°	The standard molar entropy. This is the absolute entropy of 1 mole of a substance in its standard state, usually measured at $298\,\text{K}$. By the third law, this is non-zero.
$S_{0\,\text{K,perfect crystal}} = 0$	The third law of thermodynamics. Here $S_{0\,\text{K,perfect crystal}}$ refers to the entropy of a pure, perfectly crystalline solid at $0\,\text{K}$. The third law of thermodynamics gives us a reference point to define the zero of entropy. Basically, it states that there's only one possible microstate for a perfect crystal with no internal energy. The entropy for a system with only one microstate is $S = k_B \ln(1) = 0$.

Thermodynamics of Acids and Bases

Equation/Concept	Info About Equation/Concept
$K_a = \frac{[H_3O^+][A^-]}{[AH]}$	Definition of the acid dissociation constant. This is simply the equilibrium constant for acid reactions. AH is an arbitrary acid and A^- is its conjugate base. Concentrations should all be raised to the power of their stoichiometric coefficients.
$K_b = \frac{[OH^-][BH^+]}{[B]}$	Definition of the base dissociation constant. This is simply the equilibrium constant for base reactions. B is an arbitrary base and BH^+ is its conjugate acid. Concentrations should all be raised to the power of their stoichiometric coefficients.
$K_W = K_a K_b = [H_3O^+][OH^-]$	Water self-dissociation constant. This is equal to 10^{-14} at $25\,°\text{C}$. It is the equilibrium constant for pure water undergoing dissociation.
$pH = -\log_{10}[H_3O^+]$	Definition of pH, pOH is defined similarly, where $[H_3O^+]$ is replaced by $[OH^-]$.
$pK_W = pK_a + pK_b = pH + pOH$	Equal to 14 at $25\,°\text{C}$. The water self-dissociation constants in terms of negative logs.
$[AH]_{\text{Buffer}} \approx 0.1[A^-]$ to $10[A^-]$	For effective buffers, the acid and conjugate base should be within 10 times concentration of each other.
$pH_{\text{Buffer}} \approx pK_a \pm 1$	A more succinct way of stating that buffers must be within 10 times concentration of each other.
$pH = pK_a + \log_{10}\left(\frac{[A^-]}{[HA]}\right)$	The Henderson-Hasselbalch equation. Here AH and A^- are an acid base conjugate pair. This equation is useful in determining how the pH changes as a result of addition of an acid or base to a *buffer solution*. The equation implicitly assumes that the change in concentrations of the acid and conjugate base is negligible. This is only appropriate for solutions that begin as buffers and stay as buffers upon addition of some small amount of another acid or base.

Kinetics

Equation/Concept	Info About Equation/Concept
Factors That Increase Reaction Rates	1. Increase the temperature to increase the average kinetic energy per particle to overcome the transition state energy barrier. 2. Add a catalyst which lowers the transition state energy barrier of the reaction. 3. Increase the concentration of reactants (except for zero order reactions) allowing more collisions to occur per unit time.
$\Delta G^{\ddagger}$	The Gibbs free energy of activation. This is the free energy change required to reach the transition state, measured as the difference in Gibbs free energy between the transition state and the reactants.
Rate Determining Step	For a multi-step reaction, the rate determining step is the step with the highest Gibbs free energy barrier, $\Delta G^{\ddagger}$.
Elementary Reaction	A reaction that proceeds through only one step, has only one energy barrier, and one transition state. For a general elementary reaction $aA + bB \to cC$ the rate law is always $Rate = k[A]^a[B]^b$ where k is the rate constant.

Equation/Concept	Info About Equation/Concept
Zero Order Reactions	These types of elementary reactions do not depend on the concentration of the reactants. Such reactions usually depend on some catalyst, which is limiting the rate of the reaction. Zero order kinetics are an artifact of the reaction conditions, and will collapse to higher order kinetics when the concentration of reactants becomes low enough to limit the rate of reaction (instead of the catalyst).
$\text{Rate} = -\frac{b}{a}\frac{d[A]}{dt} = k$	The zero order differential rate law for a general reaction $a[A] \to b[B]$. This is used to calculate the rate constant k. $[A]$ is the concentration of reactants and a is its stoichiometric coefficient. b is the stoichiometric coefficient of the products. The "Rate" here is always the rate of product formation. Often times this equation will be written without the factor of $\frac{b}{a}$, with $k' = \frac{a}{b}k$.
$[A](t) = -kt + [A]_0$	This is the integrated rate law for a zero order reaction, note that the rate constant here is really $\frac{a}{b}k$. This equation will provide the concentration of the reactants as a function of time. Plots of $[A]$ vs t will be linear for zero order reactions, with a slope of $-k$ and a y-intercept of $[A]_0$.
$t_{\frac{1}{2}} = \frac{[A]_0}{2k}$	The half-life of a zero order reaction. This depends on the initial concentration of reactants.
First Order Reactions	These types of reactions depend on the concentration of only one reactant, which a stoichiometric coefficient of 1. For example, the elementary reaction $A \to bB$.
$\text{Rate} = -b\frac{[A]}{dt} = k[A]$	The differential rate law for a first order reaction $A \to bB$. Often the rate constant k will include the factor of $\frac{1}{b}$.
$[A](t) = [A]_0 e^{-kt}$	The integrated rate law for first order reactions. This can also be written as $\ln[A] = -kt + \ln[A]_0$. This equation can be used to determine the concentration of reactants as a function of time. Plots of $\ln[A]$ vs t will be linear for first order reactions, with a slope of $-k$ and a y-intercept of $\ln[A]_0$.
$t_{\frac{1}{2}} = \frac{\ln 2}{k}$	The half-life for a first order reaction. This does not depend on the initial concentration of the reactants.
Second Order Reactions	These types of reactions depend on the concentration of one reactant with a stoichiometric coefficient of 2, or on the concentration of two reactants. For example, the elementary reaction $2A \to bB$.
$\text{Rate} = -\frac{b}{2}\frac{[A]}{dt} = k[A]^2$	The differential rate law for a first order reaction $A \to bB$. Often the rate constant k will include the factor of $\frac{2}{b}$.
$\frac{1}{[A](t)} = kt + \frac{1}{[A]_0}$	The integrated rate law for second order reactions. This equation can be used to determine the concentration of reactants as a function of time. Plots of $\frac{1}{[A]}$ vs t will be linear for second order reactions, with a slope of k and a y-intercept of $\frac{1}{[A]_0}$.
$t_{\frac{1}{2}} = \frac{1}{[A]_0 k}$	The half-life for a second order reaction. This depends on the initial concentration of the reactants.
$k = Ae^{-\frac{E_a}{RT}}$	The Arrhenius equation. This applies to any reaction and describes the effect of temperature on the rate of reaction. A is the pre-exponential factor, and is usually assumed to be temperature independent over small changes in temperature; it describes the frequency of collisions that occur with the correct orientation (different for each reaction). E_a is the activation energy per mole, and is the minimum energy required to overcome the energy barrier for reaction. RT is the average thermal energy per mole. This equation is sometimes written $\ln k = -\frac{E_a}{R}\frac{1}{T} + \ln A$ so that plots of $\ln k$ vs $\frac{1}{T}$ are linear with a slope of $-\frac{E_a}{R}$ and a y-intercept of $\ln A$.
$k = \frac{k_B T}{h}e^{-\frac{\Delta G^{\ddagger}}{RT}}$	The Eyring equation. This applies only to elementary reactions, and can be used to determine the Gibbs free energy of activation. k_B is the Boltzmann constant, which is just $\frac{R}{N_A}$ where N_A is Avagadro's number. This reaction comes about by assuming that formation of the transition state is kinetically controlled (i.e. is in equilibrium) but the formation of products is thermodynamically controlled (i.e. is irreversible).
$\ln\left(\frac{k}{T}\right) = -\frac{\Delta H^{\ddagger}}{R}\frac{1}{T} + \frac{\Delta S^{\ddagger}}{R} + \ln\left(\frac{k_B}{h}\right)$	The linear form of the Eyring equation. Plots of $\ln\left(\frac{k}{T}\right)$ vs $\frac{1}{T}$ will be linear. The slope will be $-\frac{\Delta H^{\ddagger}}{R}$ and the y-intercept will be $\frac{\Delta S^{\ddagger}}{R} + \ln\left(\frac{k_B}{h}\right)$.
Reaction Intermediates	These are species that are generated in one step of a reaction and consumed in a subsequent step.
$\frac{d[\text{intermediates}]}{dt} = 0$	This is the steady-state approximation. Useful for reactions where the rate determining step is the production of an intermediate. This often greatly simplifies the rate laws for multi-step reactions.

Equation/Concept	Info About Equation/Concept
$V_0 = \frac{V_{\max}[S]}{K_m + [S]}$	The Michaelis-Menton equation. Describes the kinetics of the enzyme equilibrium reaction $E + S \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} ES \xrightarrow{k_{cat}} E + P$. Here, $K_M = \frac{[E][S]}{[ES]} = \frac{k_{-1} + k_{cat}}{k_1}$ is the equilibrium constant for the enzyme-substrate binding; V_0 is the rate of product formation; $V_{\max} = k_{cat}[E]_{total}$ is the reaction rate when all available enzymes are saturated. Derivation of this equation involves the steady state approximation $\frac{d[ES]}{dt} = 0$. Notice that in the limit where the concentration of substrate is large compared to that of the enzyme, $K_M \ll [S]$ such that $K_M + [S] \approx [S]$ and $V_0 \approx V_{\max}$. In the limit of very low substrate concentration, $K_M \gg [S]$ such that $K_M + [S] \approx K_M$ and the kinetics become first order with respect to $[S]$.

IUPAC Organic Nomenclature Notes

- The highest priority functional group is given the suffix, all other lower priority functional groups are named with prefixes.

Naming esters with IUPAC:

- If the ester is the highest priority functional group, assign the alpha carbon (the carboxyl carbon) to number 1.

- Name the other side of the ester as if it is a substituent group. Place the name of this substituent group in front. Name the primary chain as normal, but add "-oate" to the end (eg: ethyl methanoate).

Naming ethers with IUPAC:

- Determine which side of the ether is the higher priority side: the shorter carbon chain becomes the "substituent" chain.

- Name the shorter chain as normal, and then the "-oxy" suffix to the end. Place this name in front followed by a space.

- The higher priority chain is then named as normal, with no suffix.

Naming sulfides with IUPAC:

- Determine which side of the sulfide is the higher priority side: the shorter carbon chain becomes the "substituent" chain.

- Name the shorter chain as normal, and then the "-thio" suffix to the end. Place this name in front followed by a space.

- The higher priority chain is then named as normal, with no suffix.
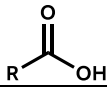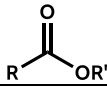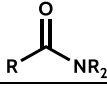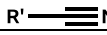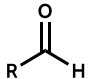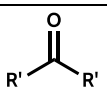
Naming acyl halides (acid halides) with IUPAC:

- These functional groups have higher priority than amides, but lower than carboxylic esters.

- Name it as if its a carboxylic acid, but remove the "-oic acid" suffix.

- Add the suffix "-oyl [hal]-ide" where "[hal]-ide" is replaced by the halogen in question (fluoride, chloride, bromide, iodide).

Naming Carboxylic anhydrides with IUPAC:

- These functional groups have higher priority than esters, but lower priority than carboxylic acids.

- Name each component carboxylic acid separately.

- Alphabetically arrange them, separated by spaces followed by the word "anhydride" (eg ethanoic propanoic anhydride).

Organic Nomenclature Guide

Functional Groups (in order of decreasing priority)

Functional group	Formula	Shorthand	Principle name	Substituent name
Carboxylic acids	(structure: R–C(=O)–OH)	R-COOH	-oic acid	carboxy
Esters	(structure: R–C(=O)–OR')	R-COOR'	-oate	(R)oxycarbonyl
Amides	(structure: R–C(=O)–NR$_2$)	R-CONR$_2$	-amide	amido
Nitriles	R'–≡N	R'-CN	-nitrile	cyano
Aldehydes	(structure: R–C(=O)–H)	R-CHO	-al	oxo
Ketones	(structure: R'–C(=O)–R')	R'-CO-R'	-one	oxo
Alcohols	R'-OH		-ol	hydroxy
Thiols	R'-SH		-thiol	sulfanyl
Amines	R'-NR$_2$		-amine	amino
Hydrocarbons	R'-H		-e	(R)yl
Ethers	R'-O-R'		-ether	(R)oxy
Sulfides	R'-S-R'		-sulfide	(R)ylsulfanyl
Halides	R'-X (X=F, Cl, Br, I)		-(hal)ide (e.g. bromide)	(hal)o (e.g. fluoro)

Note: R=saturated C/alkyl or H; R'=saturated C/alkyl

Prefixes 1 (parent/substituent length)

R (# of carbons)	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-
9	non-
10	dec-

Infixes

Nature of C-C bonds	Infix
Single bond(s)	-an-
Double bond(s)	-en-
Triple bond(s)	-yn-

Prefixes 2 (multiple substituents/infixes)

# of repeated elements	Infix
2	di-
3	tri-
4	tetra-
5	penta-

IUPAC Nomenclature Rules

1. Identify the highest priority functional group and corresponding suffix.
2. Identify the parent chain (longest carbon chain that includes (if C-containing) or is directly attached to (if non-C-containing) the highest priority functional group) and corresponding prefix.
3. Identify the infix based on the nature of the C-C bonds in the parent chain (single/double/triple)
4. Number the parent chain from end to end, giving the lowest # to (1) the highest priority functional group; (2) multiple bonds; (3) substituents closest to the end; (4) groups that come first alphabetically
5. Identify the position and name of each substituent.
6. If two or more substituents are present on the same carbon atom, use the same number twice.
7. If a substituent group or infix appears more than once, use the prefixes: di, tri, tetra, penta, hexa etc.
8. List the substituents alphabetically at the beginning of the name (use the root names from rule 5, not the prefixes from rule 7 to alphabetize).
9. Numbers are separated from letters by dashes (-). Multiple numbers are separated by commas (,).
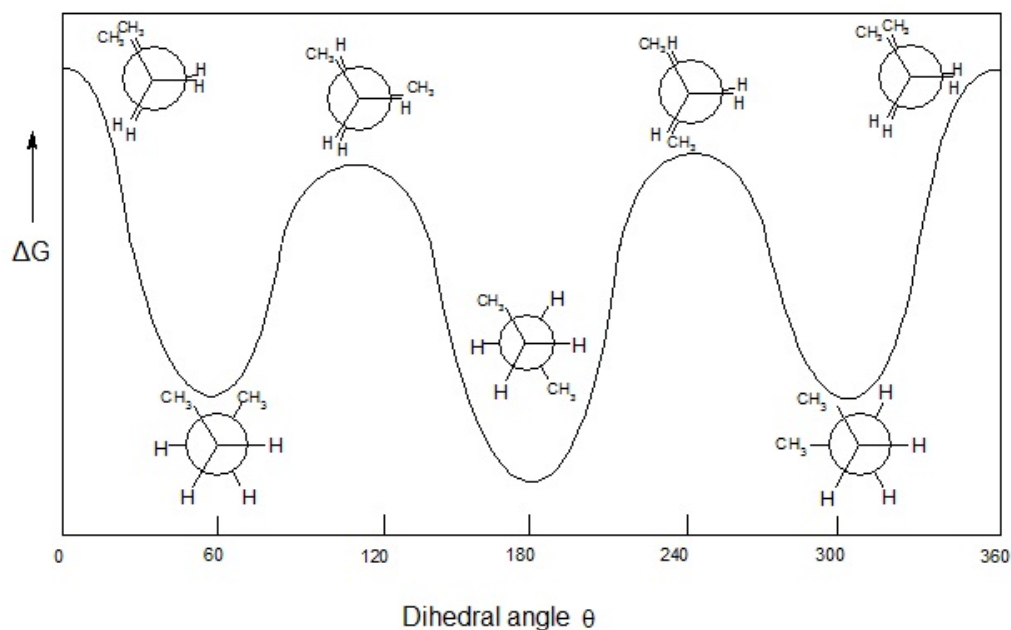
Organic Chemistry: Bonding Theories

- **Formal Charge:** (Valence electrons) - (Lone Pair electrons) - $\frac{1}{2}$(Bonding electrons)

- **Electronegativity:** The strength that an atom has to attract electrons. More electronegative atoms have lower energy orbitals. Electronegativity increases $\rightarrow$ and $\uparrow$ accross the periodic table.

- **Lewis Dot Structures:** The simplest bonding theory. Non-bonding valence electrons are represented by dots (eg H·) and bonding electrons are represented by lines (eg H—H). Good for connectivity, but does not show molecular geometry.

- **VSEPR Theory:** Valence Shell Electron Pair Repulsion theory. Is able to predict molecular geometry by assuming substituents take on the lowest possible energy arrangement around a central atom. Does not predict molecular orbitals, and so cannot be used for reactivity.

- **Valence Bond Theory (VBT):** Simplified version of molecular orbital theory, only involving the valence electrons. Predicts hybridization and good approximations for orbitals.

- **Molecular Orbital (MO) Theory:** Generates fully accurate molecular orbitals which accurately predict reactivity. Becomes extremely complicated for large molecules.

- **Frontier MO theory:** Remixes the hybrid orbitals generated in VBT using the same method as MO theory.

Organic Chemistry: Reactivity and Acid-Base Chemistry

- **Resonance:** The concept in which two or more dot formulae for the same arrangement of atoms are necessary to fully describe the bonding in a molecule or atom. Basically an artifact of the inaccuracies inherent in our line-bond structures, which do not allow for partial bonding or electron delocalization to be shown from a single structure. Resonance allows movement of electrons through the π-system only, and always lowers the energy of the structure. Electrons flow form source (δ^-, -) to sink (δ^+, +).

- **Nucleophile:** An electron-rich species that seeks positive charge to react with. Bases are often but not always good nucleophiles (counter example: extremely bulky bases are not good nucleophiles due to steric hindrance). Anions are usually good nucleophiles

- **Electrophile:** An electron-deficient species that seeks negative charge to react with. Positively or partially positively charged species make good electrophiles.

- **pK$_a$:** Lower pK$_a$ means stronger acid. Higher pK$_a$ means stronger base. A better acid is better at donating protons, a better base is better at accepting protons.

- **Strong Acids and Bases:** These are unstable species, their conjugates will be much more stable and hence highly favoured at equilibrium.

- **Factors That Effect Stability:**

  1. **Resonance:** Allows delocalization of charge, leading to stabilization. Resonance motifs: $\pi - \sigma - \pi$; $\pi - \sigma-$charge; $\pi - \sigma-$lone pair.

  2. **Electronegativity:** Negative charge can be better stabilized by more electronegative species. Positive charge can be better stabilized by less electronegative (more electropositive) species. Comparison of electronegativity for stability is most important across a row on the periodic table.

  3. **Polarizability:** More polarizable atoms can spread out charges better, and hence stabilize both negative and positive charges. Comparison of polarizability is most important down a column on the periodic table.

  4. **Inductive Effects:** Nearby substituents can withdraw or donate electrons to help stabilize charge. This effect reduces with distance from the charge. More electronegative substituents are good electron withdrawing groups, helping to stabilize negative charge. Electron donating groups (alkyl groups, electron-rich groups) help stabilize positive charge.

- **Electronegativity of Identical Atoms:** Electronegativity increases with increasing s character, as s-orbitals are lower in energy. Electronegativity: $sp > sp^2 > sp^3 > sp^3d >$ etc.
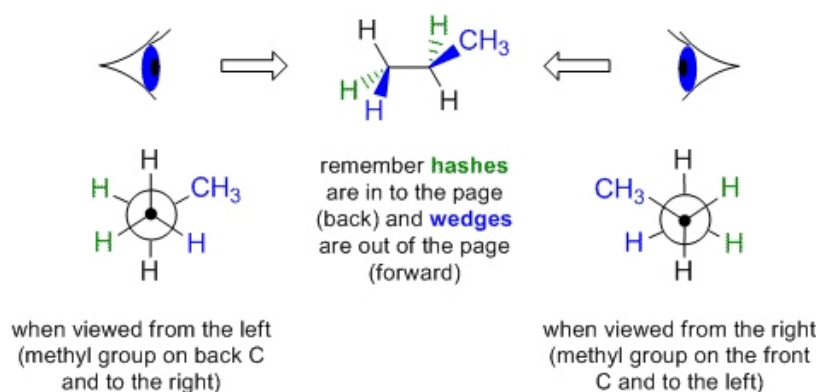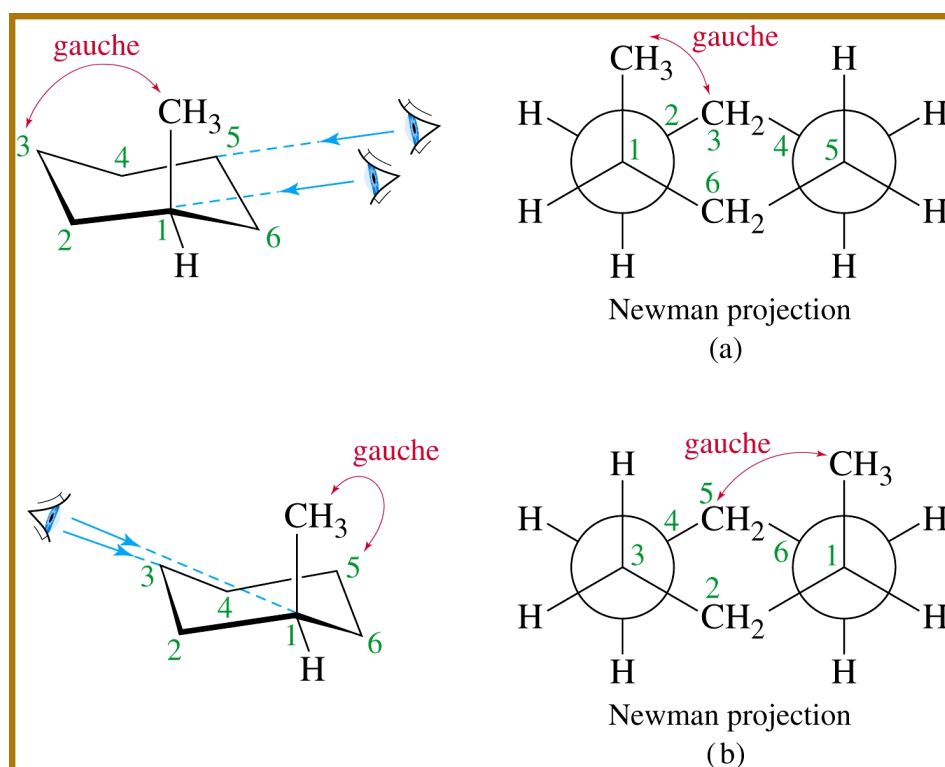
Organic Chemistry: Geometry and Stereochemistry

Dihedral angle θ

- **Butane Conformations:** For the dihedral angle between terminal methyl groups: Syn $= 0°$; Gauche $= 60°$; Eclipsed $= 120°$; Anti $= 180°$. Notice eclipsed is always higher energy than staggered. Energy is raised as the bulky substituents are forced closer together.

- **Angle Strain:** Always raises energy. If an sp^3 hybridized carbon is force to take on a geometry with 90° bond angles as in cyclobutane, the energy of the system is raised.

- **Cyclohexane Conformations:** Two possible conformation types: chair or boat. Chair is always lower in energy. Chair flipping must pass through the boat conformation, and changes all equatorial substituents to axial and vice versa. The lowest energy chair conformation is the one that most reduces 1,3-diaxial interactions.

- **1,3-Diaxial Interactions:** These are interactions of the three axial substituents in the chair conformation of cyclohexane on both faces. 1,3-Diaxial interactions are reduced when as few bulky substituents as possible occupy axial positions. Longer bonded groups like Br do not have strong 1,3-diaxial interactions.

- **Cahn-Ingold-Prelog Priority Rules (CIP Rules):** Assign highest priority to the highest atomic number atom at the first point of difference along substituents.

- **E/Z System:** Rank priority of substituents on both sides of the double bond separately using the CIP rules. If the highest priority substituents are cis (same side): Z configuration. If the highest priority substituents are trans (opposite side): E configuration.

- **R/S System:** Rank priority of substituents around an asymmetric carbon using the CIP rules. Point the lowest priority substituent backward and determine the direction from highest priority group to lowest $(1 \rightarrow 2 \rightarrow 3)$. If clockwise, then the group has R configuration. If anticlockwise, the group has S configuration.

- **E/Z and R/S Nomenclature:** Determine the underlying IUPAC root name as normal, then in brackets out in front of the name, list the positions of each stereocenter followed by its configuration in numerical order, separated by commas. (eg: (3S,5S)-5-bromo-3-chloro-5-fluoro-3-hexanamine).

- **Stereoisomers:** Isomers of identical constitution (connectivity) but differing in the arrangement of their atoms in space. Can refer to enantiomers, diastereomers, or conformers.

- **Chirality:** The property of any molecule or object of being non-superimposable with its own mirror image (e.g. hands are chiral). Chiral molecules cannot have a plane of symmetry or inversion center. Chiral molecules rotate plane-polarized light.

- **Enantiomers:** Two structures that that are non-superimposable mirror images of one another. Enantiomers must be chiral. Enantiomers have the (R/S) configuration at all asymmetric carbons flipped, but have the same (E/Z) configuations. Enantiomers have identical physical and chemical properties, except when interacting with other chiral compounds and in their interaction with plane-polarized light. Enantiomers rotate plane polarized light in opposite directions by the same amount.

- **Diastereomers:** Two structures that have the same connectivity (are stereoisomers) but have at least one difference in the (R/S) configurations at asymmetric carbons, but not at all asymmetric carbons.

- **Meso Compounds:** Compounds that would otherwise be chiral, but contain a plane of symmetry or inversion center which make them achiral. Meso compounds have asymmetric carbons arranged symmetrically, such that when the (R/S) configurations of all asymmetric carbons are flipped, an identical molecule is produced.
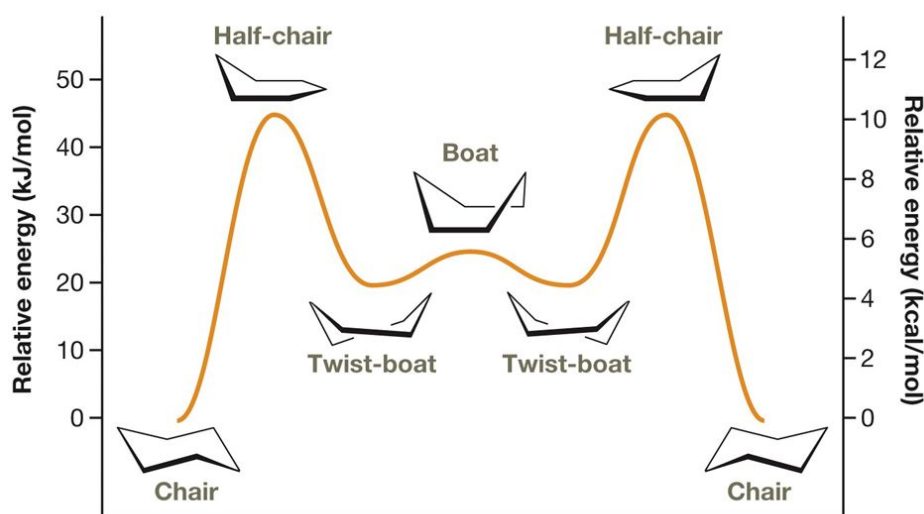
Newman Projections

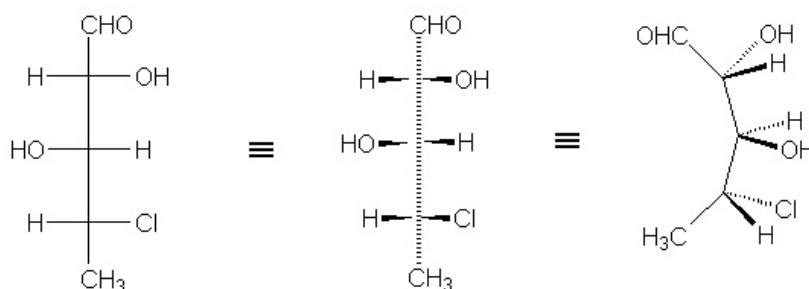## Cyclohexane and Double Newman Projections

8

More Notes About Cyclohexane

- **Conformations:** There are a maximum of only two possible chair conformations for any substituted cyclohexane, and they can be switched through a chair flip. Chair flipping requires energy as it must pass through the boat conformation (see below). The boat conformation is always higher in energy than either of the chair conformations.

- **1,3-Diaxial Interactions:** One of the two possible chair conformations of a substituted cyclohexane may be lower in energy, and therefore the preferred orientation. To determine which, look at the strength of the 1,3-diaxial interactions in each conformation.

- **Factors That Affect 1,3-Diaxial Interactions:**

  - **Bond length.** Larger atoms like bromine or iodine have long bond lengths, and thus do not interact much through 1,3-diaxial interactions with hydrogen or methyl groups, but will interact with other groups that have long bonds!

  - **Degree of substitution.** When comparing the strength of 1,3-diaxial interactions for alkyl groups, look at the number of non-hydrogen groups bound to carbons that are in the axial orientation. Alkyl groups that are highly substituted *at the first carbon* interact more strongly through 1,3-diaxial interactions and are thus less stable. Substitution at carbons beyond the first carbon in a chain are less important, but still may play a small role in determining relative energy.

Cyclohexane Chair Flip

Fischer Projections

Summary of S_N1 vs S_N2 Reactions

		SN2	SN1
characteristics	**mechanism**	simultaneous bond making and breaking	bond breaking first then bond making
	key intermed or TS	crowded TS	carbocation intermed
	stereochemistry	inversion of stereochem	both inversion and retention of stereochem
	rate	$= k\,[\text{R-X}]\,[\text{Nuc}^-]$	$= k\,[\text{R-X}]$
factors that effect the rate of reactions	R-X	**Methyl > 1° > 2°**	**3° > 2°**
	Nuc	**must be good Nuc$^-$**	can be weak (is often the solvent)
	LG	good LG is better	**must be good LG**
	solvent	more polar is better **(aprotic is better)**	more polar is better **(protic is better)**

- More substituted electrophiles form more stable carbocations, and hence react more quickly through S_N1 reactions. More substituted electrophiles are less likely to react through S_N2 due to steric hindrance.

- In S_N1 reactions, the rate limiting step is the carbocation formation. Hence, the strength of the nucleophile is not very important.

- In S_N2 reactions, the rate limiting step is the formation of a concerted transition state. Stronger nucleophiles lower the energy of the transition state and hence improve reaction rates.

- Weak bases make good leaving groups for both types of substitution reaction. Weak bases are more stable as free species than strong bases, and hence do not raise the energy of the system too much upon dissociation.

- S_N1 can be catalyzed by protic solvents or acids in solution. A proton can be added to a bad leaving group, turning it into a better leaving group. Polar solvents also help to stabilize the carbocation.

- S_N2 reactions favor polar aprotic solvents, as these allow for stronger (charged) nucleophiles. Protic solvents will lose their proton to strong nucleophiles, making them weak. Polar protic solvents also tend to hydrogen bond with the nucleophile, creating a solvent shell around it and hindering its reactivity.

- Aprotic solvents do not solvate negatively charged nucleophiles that well. Therefore, negatively charged nucleophiles are more reactive in aprotic solvents.

- More polarizable nucleophiles react faster because they are not solvated as readily. Their charge is spread out more, so that Van der Waals bonds to solvent molecules are weaker. Small nucleophiles are highly solvated, and tend to react slowly in S_N2 reactions. Relative polarizability for nucleophiles is most important along the same columns of the periodic table.

- Adjacent π-systems help stabilize electrophiles in both S_N1 and S_N2 reactions through the possibility of resonance. In S_N1 reactions the carbocation is stabilized by resonance, whereas in S_N2 the concerted transition state is stabilized by resonance.

- S_N2 reactions only occur on sp^3 hybridized carbons.

- The rate determining step of S_N2 reactions is a bi-molecular step. Hence it has second order kinetics and the rate of reaction depends on the concentrations of both the nucleophile and electrophile.

- The rate determining step of S_N1 reactions is unimolecular. Hence it has first order kinetics and the rate of reaction only depends on the concentration of the electrophile.

Physical Constants and Periodic Trends (for Reference)

Constant	Value and Units
Atomic mass unit (u)	$1\,u = 1.660\,538\,9 \times 10^{-24}\,g$ $1\,g = 6.022\,142 \times 10^{23}\,u$
Avogadro's number (N_A)	$N_A = 6.022\,142 \times 10^{23}\,mol^{-1}$
Boltzmann's constant (k_B)	$k_B = 1.380\,648\,52 \times 10^{-23}\,J\,K^{-1}$ $k_B = R/N_A$
Gas constant (R)	$R = 8.314\,459\,8\,J\,K^{-1}\,mol^{-1}$ $= 8.314\,459\,8 \times 10^{-3}\,kJ\,K^{-1}\,mol^{-1}$ $= 8.314\,459\,8\,kg\,m^2\,s^{-2}\,K^{-1}\,mol^{-1}$ $= 8.314\,459\,8\,m^3\,Pa\,K^{-1}\,mol^{-1}$ $= 0.082\,057\,338\,L\,atm\,K^{-1}\,mol^{-1}$ $= 1.987\,203\,6 \times 10^{-3}\,kcal\,K^{-1}\,mol^{-1}$ $= 8.205\,733\,8 \times 10^{-5}\,m^3\,atm\,K^{-1}\,mol^{-1}$ $= k_B N_A$
Mass of electron (m_e)	$m_e = 5.485\,799 \times 10^{-4}\,u$ $= 9.109\,383 \times 10^{-28}\,g$
Mass of proton (m_p)	$m_p = 1.007\,276\,5\,u$ $= 1.672\,621\,7 \times 10^{-24}\,g$
Mass of neutron (m_n)	$m_n = 1.008\,664\,9\,u$ $= 1.674\,927\,3 \times 10^{-24}\,g$
Planck constant (h)	$h = 6.626\,069 \times 10^{-34}\,J\,s$
Reduced Planck constant ($\hbar$)	$\hbar = h/2\pi = 1.054\,572\,66 \times 10^{-34}\,J\,s$
Speed of light (c)	$c = 2.997\,924\,58 \times 10^{8}\,m\,s^{-1}$