

# 5.53 (Molecular Structure and Reactivity I) Notes

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# Weeks

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# Week 1

## Introduction

### 1.1 Introduction

9/5:

- Normally, only about 20 kids enroll in this class per year. This year, there are 40.
  - This is a typical class for the first-year grad students in OChem, but Elkin asks what made advanced undergrads and second-year grad students enroll, as well as just so many of us overall.
  - Radosevich told all the inorganic kiddos to take this class!
  - Bioinorganic and Organometallics also aren't being offered because everyone's on sabbatical.
- Oleta Johnson came to sit in on Masha's class! Oleta is Masha's "best friend."
- The lecture now begins (on MIT Time).
- Masha will teach the first half of the course; Alex will teach the second half.
  - TF is Jonathan Edward, an Elkin kiddo.
    - He will hold weekly OH, study sessions, grades problems and exams, etc.
    - Has a mastery of the subject material (took 5.53 last year), and unrivaled "approachability."
  - Reach out to Masha or Alex if we have issues with the subject material, our own journeys in grad school or undergrad, etc. It's easier to fix problems early in the semester!
- Overview of the course.
  - 1st half.
    - Basically physical organic chemistry.
    - A deep dive on structure and reactivity.
  - 2nd half.
    - Basically reaction mechanisms.
    - Kinetics, rate laws, kinetic isotope effects (KIEs), methodology experiments, etc.
  - The tools presented herein are broadly applicable to various fields of chemistry.
- This course will teach us to...
  - Propose *reasonable* mechanisms for organic reactions;
  - Scrutinize mechanisms in the literature;
    - That is, figure out if a proposed mechanism is reasonable or not, evaluate the authors' evidence, and identify follow-up experiments that can be run.
  - Design experiments to distinguish and test proposed mechanisms;
  - Conduct our own mechanistic study.

- Masha gives the metacognition spiel again.
  - Know our strengths and weaknesses (correct these by reviewing undergrad notes and Googling).
- Course logistics.
  - 2 exams.
    - Fully online; they are trusting us to work alone on the honor system.
  - 4 problem sets.
    - Posted 1 week before they are due.
    - Encouraged to work collaboratively, but submit our own work.
    - Jonathan and Masha will reserve a study room in which we can collaborate.
  - 1 mechanistic proposal.
    - Engage the literature!
  - Textbook: Anslyn and Dougherty (2006).
    - The standard textbook for PhysOrg (do readings and practice problems as needed).
    - Jonathan is working on a correspondence of lectures to chapters.
  - Reach out to Masha, Alex, or Jonathan if we have any questions!
    - If you ever miss class, post a new topic on the Canvas discussion board asking for notes (and be generous in uploading your own).
- We now begin the course content.
- **Mechanism:** An accounting of all bond-making and bond-breaking events in a reasonable sequence.
  - Mechanisms don't exist in the physical sense; it is more of a *model* of how things proceed.
- Mechanisms exist in four levels of depth.
  1. Describe electron movement via arrow pushing.

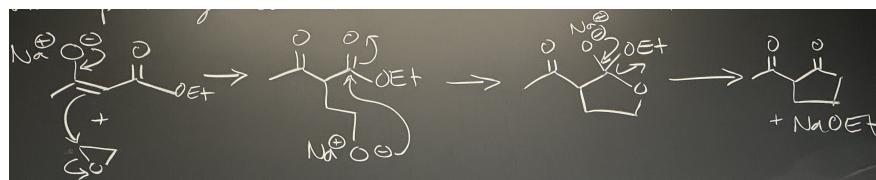


Figure 1.1: Mechanism depth level 1 (arrow pushing).

- Equivalent level: 5.47 & undergrad organic.
- Example: Figure 1.1.
  - In every step that we push arrows, we start at a region of high electron density, we make and break sequential bonds, and we leave the negative charge on an electronegative atom.
  - Once we have completed one step, we can start again from a new region of high electron density, making and breaking bonds, and drawing the product.
  - We repeat this process again and again until we reach the final product.
  - Arrow pushing conserves net neutral charges on molecules.
- Aside: Arrow types.

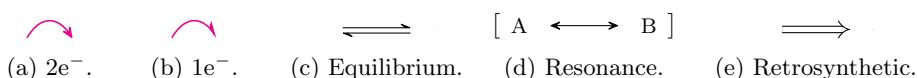


Figure 1.2: Arrow types in arrow pushing.

## 2. Determine the transition-state structures.

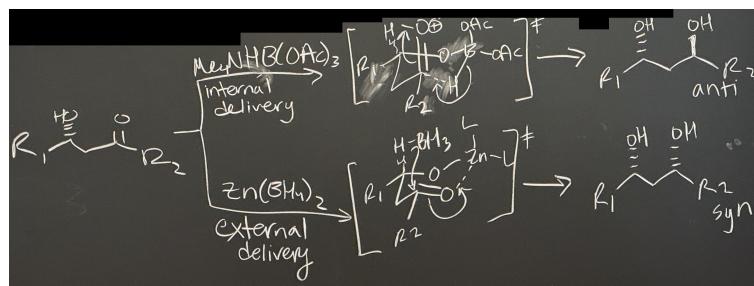
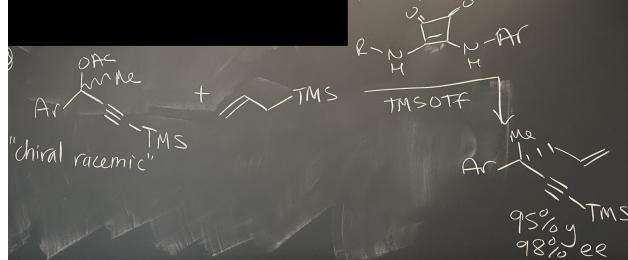


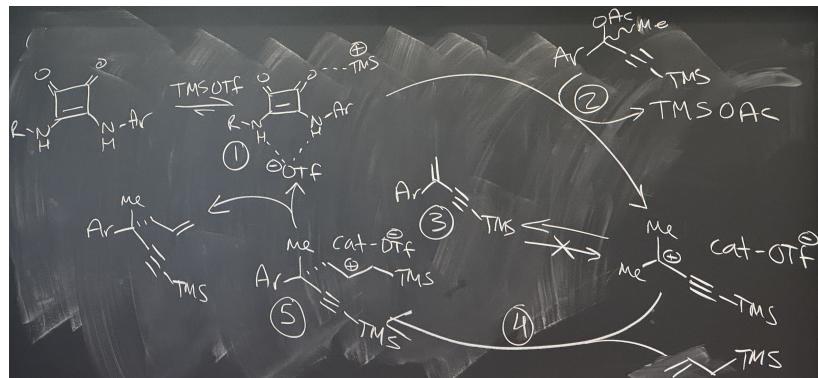
Figure 1.3: Mechanism depth level 2 (transition states).

- Equivalent level: 5.47 & undergrad organic, as well.
- Can't observe these directly — infer from observed selectivities (stereo-, regio-, etc.).
- Example: Figure 1.3.
  - Reacting a  $\beta$ -ketol with two different reducing agents. We can infer the structure of the transition state from the stereochemistry of the product.
  - Internal delivery of tetramethylammonium triacetoxyborohydride yields an *anti*-diol.
  - External delivery of zinc borohydride yields a *syn*-diol.
- Takeaway: We know that in organic chemistry, transition states should have chair-like structures for stability.
- Since we see chair like structures in Figure 1.3, we can infer that these mechanisms are reasonable. Indeed, they have stood for decades!

## 3. Determine the energy landscape and the full reaction coordinate.



(a) A reaction.



(b) The full reaction coordinate.

Figure 1.4: Mechanism depth level 3 (full reaction coordinate).

- Equivalent level: This class!
- This level of analysis enables us to...
  - Rationally design experiments that improve the reaction (i.e., conduct methods development and catalysis);
  - Discover new mechanistic principles.
- Example: Figure 1.4.
  - Figure 1.4a depicts a curious reaction: Propargyl acetate (with racemic chirality) reacts with an allyl silane under a squaramide catalyst and TMSOTf (a Lewis acid).
  - Even though the starting material is racemic, we get an enantioenriched allylated propargyl acetate (95% yield, 98% ee) as a product.
- The mechanism (Figure 1.4b) proceeds in five steps.
  - (1) Activate the catalyst to form an intermediate.
  - (2) Engage the starting material to form a tertiary carbocation.
  - (3) This carbocation can off-cycle to form an elimination product.
  - (4) Preferably, however, we engage our nucleophile (the allyl silane) to get a new cationic adduct, counterbalanced by the catalyst-triflate complex.
  - (5) The adduct goes on to eliminate our product and regenerate the starting intermediate.
- This mechanism originated from a beautiful mechanistic study by this paper's authors. Let's discuss some of their insights.
  - (1) This complex is the **resting state**.
    - Analytical technique(s): Binding experiments between the catalyst and TMSOTf.
    - This is a thermodynamic insight.
  - (2) This step is the **rate-determining step**.
    - Analytical technique(s): The rate law (a kinetic parameter) and **Hammett plots**.
  - (3) This step is an irreversible side reaction.
    - Analytical technique(s): Competition experiments.
    - Since this step is post-RDS in the mechanism, it is quite difficult to study.
    - Takeaway: It is easy to see things between the resting state and RDS, but everything after the RDS is like magic. These steps are very hard — but very important — to probe. Indeed, knowing how and where side reactions originate provides clues on how to stop them!
  - (4) This step is the **stereo-determining step**.
    - Analytical technique(s): The kinetic isotope effect (review this from CHEM 2020!!).<sup>[1]</sup>
      - Revealed that stereoinduction was due to noncovalent interaction (NCIs) between the catalyst and intermediate.
    - Usually, your stereo-determining step is your RDS, but not in this regime. It is very hard to optimize a post-RDS, stereo-determining step.
  - (5) This intermediate is stabilized due to hyperconjugation from silicon.
    - Analytical technique(s):  $\beta$ -silicon effects and  $\alpha$ -silicon effects.
- We will learn all of the techniques mentioned above in this class.
- Impact of this paper.
  - It's one of the first enantioselective S<sub>N</sub>1 reactions.
  - It has a decoupled RDS and stereo-determining step, but gets high ee regardless.
    - This was an unprecedented result, and it changed the way we as chemists think about optimizing entantioselective reactions.
  - Reference: Wendlandt et al. (2018).

<sup>1</sup>KIEs should probably be used in our end-of-class mechanistic proposal, will likely be used in our research, and *can* probe post-RDS steps.

4. Computationally determine the entire multidimensional energy surface.

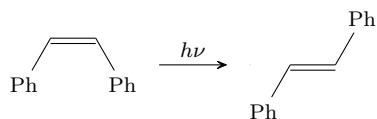


Figure 1.5: Mechanism depth level 4 (full energy manifold).

- Currently only possible for very simple systems.
- Example: Figure 1.5.
  - This is a transformation under light from a *cis*-olefin to a *trans*-olefin.
  - The authors tracked the reaction with femtosecond ( $10^{-15}$ ) Raman spectroscopy.
  - Reference: Takeuchi et al. (2008).
- Full computational modeling is a pipe dream that would hugely enable our work as chemists.
- **Resting state** (of a catalyst): The state of a catalyst such that if you took an NMR of the reaction mixture at any given time, 95% of the sample would look like this.
- **Rate-determining step.** *Also known as rate-limiting step.*
  - Important because if you can speed it up, you can speed up the whole thing!
- **Rate law:** A measure of how the rate of reaction is influenced by the concentration of different components.
- **Hammett plot:** A mechanistic tool to probe what the rate-determining step is.
- **Stereo-determining step:** The step in a reaction mechanism that sets the stereochemistry of the final product; the ee of this step is the ee of the product.
- Takeaway: Keep in mind these various levels when we're trying to work out a reaction!
- Online tool: Reference Resolver!!
  - Give it the journal, year, and page number, and it brings us to the article.
  - There is a website, but also a browser plugin worth getting.
- Now that we've discussed the kinds of mechanisms, let's talk about what a mechanism can and can't do for us.
- A mechanism *can* tell us...
  - Thermodynamics and equilibria: Identity and structure of the ground state species;
  - Kinetics: Identity and structure of the transition state (TS) structures *relative* to the ground state structures;
    - We can't identify anything about the transition state in absolutes, but we can take educated guesses about intermediates and infer their approximate form.
  - Intermediates: Evidence of reaction intermediates;
    - Example: The tetrahedral intermediate.
    - Such intermediates are often called **metastable**.
  - RDS: Insight into selectivity and RDS's.
- **Metastable** (state): An intermediate energetic state within a dynamical system other than the system's state of least energy. *Also known as unstable equilibrium.*
  - A rectangular prism standing on its end under the force of gravity is metastable.

- A mechanism *cannot* be proven.
  - Mechanisms are hypotheses or proposals that can only be *disproven* or *supported*.
  - This is because experimental data often fits several possible mechanisms; there might be a hidden secret mechanism that we never thought of.
  - In sum, a mechanism is an interpretation that is consistent with *all* the data.
    - If a mechanism doesn't fit our data (even a little bit), either our mechanism is missing something (maybe a little something) or our experiment is flawed (and we need to rerun it or run something else).
- Best practices.
  - The best mechanisms provide *testable* predictions.
    - If a mechanism doesn't provide testable predictions, it is not a useful model.
    - If it's not useful, it's not grounded in good scientific practice.
  - The best experiments disprove a mechanistic proposal.
    - In practice, we list all possible mechanisms and try to disprove them with experiments.
    - When we submit to a journal, we do not say that our mechanism is proven, but we state our reasoning and our reviewers try to think of other mechanisms that could fit the data.
- Aside: Both Alex and Masha care about how scientists actually do science and how science can be done ethically.
  - They want this to be a practical class that would enable us to go in the lab and run any of these experiments.
- We study mechanisms to...
  - Ensure a safe, robust (reproducible), and scalable process;
    - This is especially important in process chemistry.
    - Human consequences of failing at safety, scale, and/or robustness:
      - Your ammonia plant could explode; it is essential to watch any runaway exotherms in a mechanism and control them!
      - Your drug might not make it to market if its synthesis can't be scaled up.
  - Improve reaction features such as yield, selectivity, and greenness;
  - Expand scope and enable predictability;
    - Think about reactions we run daily, such as the Suzuki coupling. It always works, and it's easily applicable in a wide range of settings *because* we understand the mechanism.
  - Understand systems on a molecular level.
    - Masha takes 30 seconds to preach about how mechanisms are critically important knowledge that will be passed down the generations.
- Aspects of mechanism: Consider the S<sub>N</sub>2 reaction, Br<sup>-</sup> + Me-I → Br-Me + I<sup>-</sup>.



(a) Orbitals.



(b) Energy surface.

$$\frac{d[\text{MeBr}]}{dt} = k[\text{MeI}][\text{Br}^-]$$

(c) Kinetics.

Figure 1.6: Aspects of mechanism.

- Three things we can consider in this mechanism are the orbital interactions, the potential energy surface along the reaction coordinate, and the kinetics.

# Week 2

## Bonding Models

### 2.1 Bonding Models 1

- 9/10:
- Lecture 1 recap.
    - Aspects of mechanism.
      - Orbitals, energy surface, and kinetics.
      - Masha redraws Figure 1.6.
      - These are the three main pictures that we'll learn about.
    - Today, we'll focus on orbitals.
  - Today: Bonding models.
    - Reading: Anslyn and Dougherty (2006), Chapter 1!!
  - **Bonding:** How electrons are shared between nuclei.
    - This determines all of molecular structure and reactivity (which is the name of this class, and underpins all of organic chemistry!).
    - From bonding, there arise concepts such as nucleophilicity, electrophilicity, etc.
  - There are several levels of bonding theory / models that we'll talk about today.
    - Caveat: *All* of these models are no more than *approximations* of reality that are useful to us.
  - Lecture outline.
    1. Lewis structures.
    2. VSEPR.
    3. Valence Bond Theory (VBT).
    4. Molecular Orbital Theory.
    5. Qualitative Molecular Orbital Theory (QMOT).
  - Lewis structures.



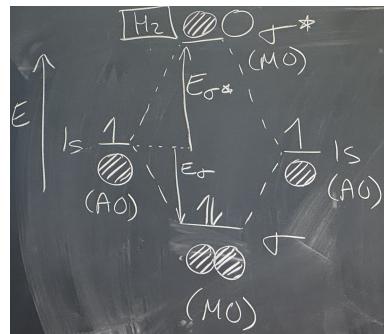
Figure 2.1: Lewis dot structures.

- Developed in 1916 by G. N. Lewis.
  - Chemis-tea: He was nominated 48 times, but never won the Nobel Prize because some people on the review committee didn't like his "interesting personality."
- In this model, we use dots to — on paper — indicate where electrons are in bonds.
- From these **Lewis dot structures**, people developed the "stick structures" that we still use today.
- Lewis structures are very useful in identifying the number of bonds and lone pairs.
- Valence Shell Electron Pair Repulsion (VSEPR).
  - Developed 1939-1957.
  - Key finding: Electrons in bonds repel each other, so you maximize the distance between bonds.
  - This let us go beyond Lewis structures into things like explaining tetrahedral carbon (and its 109.5° bond angles).
  - Issues develop when we try to rationalize other molecules.
    - For example, isobutane has 110.6° Me–C–Me bond angles. The VSEPR purists will cite "sterics."
    - As another example, NH<sub>3</sub> has 107 H–N–H bond angle. The VSEPR purists will cite "lone pair is big."
  - Really, these were just excuses by the VSEPR purists for a bad model, and what we really needed was a new model.
- Valence Bond Theory (VBT).
  - Developed by Linus Pauling, with his seminal paper in 1931.
    - For this work and some other stuff, he won the Nobel Prize in Chemistry in 1954.
    - To be historically accurate, Pauling built off the work of Heitler and London (1926).
    - However, Pauling was the person to both put everybody else's work all together and be visible enough to take the credit.
    - Additional takeaway from Pauling's biography: Don't make your whole life about your work. For example, Pauling was shunned by many of his colleagues after he got into nuclear proliferation, but now we say he was so brave. He even won the Nobel Peace Prize!
    - Takeaway on Pauling vs. Lewis: It pays to not be a jerk. Lewis died via cyanide poisoning (may have been an accident, but was probably suicide).
  - This is a quantum mechanical (QM) description of Lewis structures.
  - Central tenet: Each atom contributes 1 valence electron in a QM-derived atomic orbital (AO).
    - Shows that electrons are delocalized between atoms, and where two electrons overlap and localize is a chemical bond.
    - In other words, electrons are not restricted to tight orbitals.
  - Many concepts arise within VBT until the advent of MO theory.
- VBT was key for many conceptual innovations, such as **hybridization**, **electronegativity**, and **resonance**.
- **Hybridization:** The mixing of orbitals on the same atom to make new orbitals.
  - Specifically, we can take a linear combination of AO waveforms (or AOs).
  - More directional orbitals give you better overlap and therefore stronger bonds.
  - Example: A linear combination  $s + p_y + p_x + p_z$  yields four  $sp^3$ -hybridized orbitals. That's four orbitals with uneven lobes. We can draw all of these on top of each other, and from *there*, we get the tetrahedral carbon.

- We always like new models that agree with old models; this is called a **sanity check**.
- We can also calculate something called the **hybridization index**.
- **Hybridization index:** The number  $i$  in the following formula, expressed as a function of the experimentally determined bond angle  $\theta$ . *Denoted by  $i$ . Given by*

$$1 + i \cos \theta = 0$$
- Example:  $\text{NH}_3$  has a hybridization index of 3.4.
- Example:  $\text{H}_2\text{O}$  has a hybridization index of 4! That's why it has the tiny bond angle. The remaining  $s$ -character is localized on the oxygen, and that's why we say that oxygen is electron dense and nucleophilic.
  - Would this similarly predict that  $\text{H}_2\text{O}$  has longer bonds than  $\text{NH}_3$ ??
- **Electronegativity:** The power of an atom to attract electrons to itself.
  - There are different scales for this. We probably used the **Pauling scale**, but there is also a **Mulliken scale**.
  - More electronegative atoms have lower energy orbitals.
    - This is summarized via the **inductive effect**.
- **Inductive effect:** The withdrawing of electron density through  $\sigma$ -bonds.
  - Example:  $\text{ACN}$ . We think about nitrogen having a partial negative charge and carbon having a partial positive charge. This results in a dipole.
  - Takeaway: Dipoles arise from electronegativity in VBT!
- **Resonance:** The superposition of several Lewis structures. *Antiquated mesomerism*.
  - Example: Consider an  $\alpha, \beta$ -unsaturated ketone. Its resonance structure is a zwitterionic intermediate, and a second resonance structure is a different zwitterion. We have three resonance forms, so that predicts more stable than something with less resonance structures. It also identifies our positive and negative reactive sites.
  - Resonance usually happens through  $\pi$ -networks, but it *can* happen through  $\sigma$ -networks.
  - Takeaway: Delocalization of electron density leads to stability.
  - Know your rules for drawing good resonance structures.
    - We only move bonds, not atoms (no nuclear motion).
    - Prefer to have the least separation of charge.
    - Put the more negative charge on the more electronegative atoms.
- Limitations of VBT.
  - Over time, some key experimental findings emerged that VBT couldn't explain. These results motivated people to develop a new model to explain these rare cases.
    - Nowadays, exceptions to VBT are not so rare.
  - Remember: If a model can't explain certain cases, it's not a useful model.
    - Maxim: Not predictive = not useful.
- Here's a list of the limitations of VBT.
  - Doesn't account for unusual stability/instability (e.g., aromaticity and antiaromaticity).
  - No antibonding orbitals (i.e., no explanation of interactions between molecules).
    - When a nucleophile attacks a ketone, the interaction is with the antibonding orbital of the ketone. Forming a new bond involves populating an antibonding orbital.
  - Thursday is all about aromaticity, and modern ways to conceptualize it.

- This leads to the mother of all bonding models, Molecular Orbital Theory.
  - Central tenet: Molecular orbitals (e.g.,  $\sigma$ ,  $\sigma^*$ ,  $\pi$ ,  $\pi^*$ ) arise from linear combinations of atomic orbitals (in Orgo, this is  $s$  &  $p$ ; we won't consider  $d$ -orbital effects so much).
  - We consider the electronic structure of the whole molecule, not just atoms or bonds.
    - We focus on key molecular orbitals such as the HOMO and LUMO.
  - We also get **group orbitals**: Leads into QMOT, which is MOs for prototypical groups.
- MO theory leads to MO diagrams.

Figure 2.2: MO diagram for  $\text{H}_2$ .

- Two atomic orbitals interact to fill two molecular orbitals.
- We fill the bonding orbital with all the electrons that come in (in this case, 2).
- The energy of stabilization is  $E_\sigma$ .
- The destabilization energy is  $E_{\sigma^*}$ .
- Read Anslyn and Dougherty (2006) for more rules.
- Notes.
  - $|E_{\sigma^*}| > |E_\sigma|$ . Thus, if the antibonding orbitals get populated, the molecule breaks. This is because of nuclear repulsion.
  - The  $\sigma$ -bond is more stable than the  $1s$  orbitals by themselves. This is why the H–H bond forms. This kind of analysis allows us to predict whether or not a bond will form.
- Question for us to consider: Why doesn't He–He form?
  - Because its antibonding MOs would be populated.
- Example MO diagram: Ethylene.

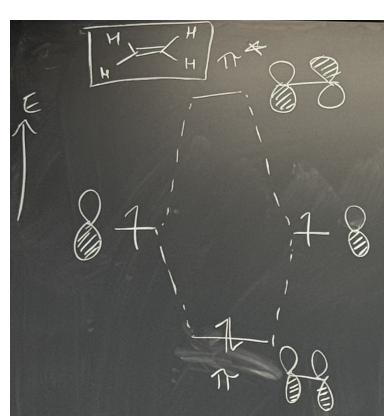
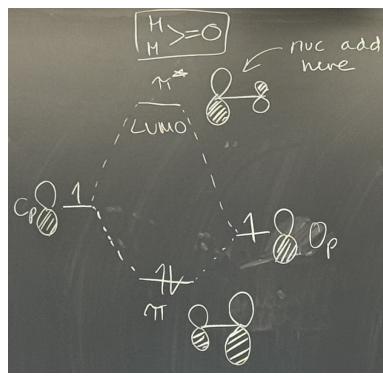


Figure 2.3: MO diagram for ethylene.

- Looking specifically at the  $\pi$ -bond formation.
  - This is why we form a stable  $\pi$ -bond.
  - Example MO diagram: Formaldehyde.
- 
- The diagram illustrates the molecular orbital (MO) formation in formaldehyde (H<sub>2</sub>O). At the top left, the Lewis structure H<sub>2</sub>O is shown with two hydrogen atoms (H) and one oxygen atom (O). A horizontal arrow labeled 'nuc add here' points towards the oxygen atom. Below the molecule, several atomic orbitals (AOs) are depicted as pairs of overlapping circles. One carbon atom (C) has a  $1s$  orbital and a  $2p$  orbital. The oxygen atom (O) has a  $1s$  orbital and a  $2p$  orbital. A dashed line labeled 'LUMO' connects the  $2p$  orbital of the oxygen atom to a pair of overlapping circles below it, which are labeled  $\pi$ . Another dashed line labeled 'H' connects the  $1s$  orbital of the carbon atom to a pair of overlapping circles below it, which are labeled  $\sigma$ . The  $\sigma$  bond is shown as a single horizontal line between the carbon and oxygen atoms.
- Figure 2.4: MO diagram for formaldehyde.
- We mix a C<sub>p</sub> AO and a (lower energy) O<sub>p</sub> AO.
  - These orbitals interact less well than those in ethylene due to their difference in energy.
  - We benefit from constructive phasing, but the lobes are much bigger on oxygen.
  - In the antibonding orbital, the lobes are much bigger on carbon.
  - Principles revealed by this MO diagram.
    - Closer energy AOs give stronger mixing, resulting in lower energy MOs. Lower energy MOs are more stabilizing.
    - More electronegative atoms have lower energy atomic orbitals.
    - The  $\pi$ -orbital is asymmetric because its energetically more similar to O<sub>p</sub> than C<sub>p</sub>.
      - In other words, it's going to look more like the O<sub>p</sub> orbital.
      - One more way of stating this is that the coefficient of oxygen in the LCAO is bigger.
  - We know that the LUMO (frontier orbital) interacts with nucleophiles. The lobe of the LUMO is bigger on carbon, hence why we react there.
  - Qualitative MO theory (QMOT).
    - All about forming group orbitals for common functional groups or motifs.
    - Essentially, we may not need to calculate MOs for the whole molecule to find out how every carbonyl reacts; we can trust that carbonyl group orbitals are decently conserved.
    - There are a bunch of rules for how to form a QMOT diagram.
      - See Table 1.7 in Anslyn and Dougherty (2006) for building QMOT diagrams.
    - This is the basis of **Walsh diagrams**.
    - We can build group MOs from linear combinations of s & p AOs.
  - **Walsh diagram:** A representation of an MO diagram as a function of geometric distortions.
    - This matters because geometry affects orbital overlap, which can be destabilizing or stabilizing.

- Example QMOT diagram:  $\text{CH}_3$ .

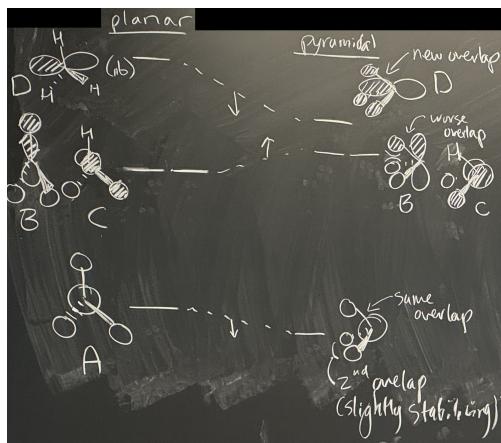


Figure 2.5: QMOT diagram for  $\text{CH}_3$ .

- Key question: What geometry of  $\text{CH}_3$  is favorable?
- Masha defines axes.
- Undetermined yet if this is a radical, cation, or anion. We'll get there!
- We look at a planar set of orbitals first.
  - A. All phases in sync, all  $s$  orbitals.
  - B. Phases align top to bottom with the  $p_x$  orbital of carbon.
  - C. Phases align in and out of the board with the  $p_y$  orbital of carbon.
  - D. Nonbonding; just the  $p_z$  orbital.
- There are also **E**, **F**, and **G** orbitals that are energetically above these, but we won't draw them for now (because we won't fill them with electrons in the carbocation, carbanion, or carbon radical).
  - The **E**, **F**, and **G** orbitals will have the opposite phasing of the lower orbitals!
- We now draw an analogous, pyramidal set of orbitals.
  - A. Overlap is *slightly* more favorable because we have a secondary orbital interaction between the hydrogens now. The C–H overlap stays the same.
  - B. Worse overlap. We're losing a **primary** interaction instead of gaining a **secondary** one, so the energy of **B** actually goes up *more* than **A** went down. We also get some destabilizing secondary interaction between the H orbitals.
  - C. Just like **B**, we get worse primary overlap, and new interfering secondary overlap.
  - D. Gets stabilized the *most* significantly! This is because we've taken something with no bonding interactions and *created* bonding interactions between the  $p$ -orbital and the hydrogens.
- Relationship between QMOT and Walsh diagrams: A Walsh diagram is a QMOT diagram with everything connected.
- Now how do we fill electrons?
  - Consider the  $\text{CH}_3^+$  cation: We have 6 electrons, so we populate the planar orbitals because it's more stable overall.
  - Consider the  $\text{CH}_3^-$  anion: We have 8 electrons, so we populate the pyramidal orbitals because *they're* more stable overall.
- This rigorous prediction of conformation is the benefit of this model.
- We can also use this model for other isostructural molecules.

- Examples.
  - NH<sub>3</sub>: 8 electrons, pyramidal.
  - BH<sub>3</sub>: 6 electrons, planar.
  - ·CH<sub>3</sub>: 7 electrons, *slightly* planar.
    - But this is a special case only for ·CH<sub>3</sub>; any other radical is pyramidal.
- Primary (orbital interaction): An interaction between orbitals on adjacent atoms in a molecule.
- Secondary (orbital interaction): An interaction between orbitals on atoms that are separated by one other atom in a molecule.
- What is quantitative about QMOT?
  - There is a lot more depth in Anslyn and Dougherty (2006). You can calculate the actual potential energy surface and figure out these conformations exactly.
- Example QMOT diagram: CH<sub>2</sub>.

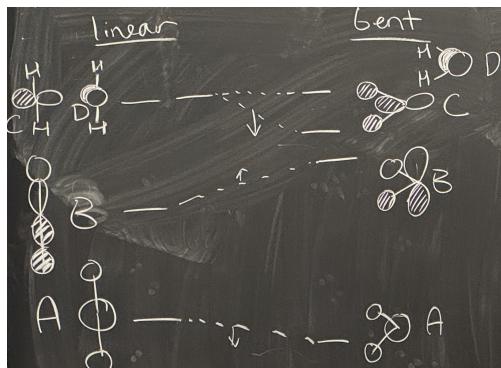


Figure 2.6: QMOT diagram for CH<sub>2</sub>.

- Two geometries: Linear and bent.
- Linear.
  - A. Linear chain of *s*-orbitals with matching phases.
  - B. Linear chain of matching phases orbitals, with *p<sub>x</sub>* on carbon.
  - C. One of the other *p*-orbitals, with no phasing.
  - D. The last remaining *p*-orbital, again with no phasing.
- Bent.
  - A. Goes down slightly. We kept primary, and added secondary.
  - B. Losing primary overlap and gaining a destabilizing secondary interaction; higher *E* like before.
  - C. Adding *significant* constructive interference. Biggest effect again!
  - D. Staying the same; no bonding interactions to begin or end with. We don't consider secondary interactions when there's no density at all there.
- Example species.
- H<sub>2</sub>O: 8 electrons, bent.
  - Note that this model predicts that H<sub>2</sub>O has nondegenerate lone pairs, which has been experimentally verified!
  - Bulk water acts as if it has degenerate lone pairs. We can read Anslyn and Dougherty (2006) about this, but otherwise, it's outside the scope of the class.

- $\text{CH}_2$  (a carbene): 6 electrons, a mix of linear and bent!
  - We'll return to carbenes in a few weeks.
  - We'll define **triplet** (2 electrons in different orbitals) and **singlet** (2 electrons in same orbital) carbenes later.
  - Triplet is  $136^\circ$ , and singlet is  $105^\circ$ , so the triplet is more linear and the singlet is more bent! The triplet has reactivity more characteristic of the linear orbital picture, and the singlet has reactivity more characteristic of the bent orbital picture.
  - The triplet is more favored by 9 kcal/mol
- Example QMOT diagram: Formaldehyde.

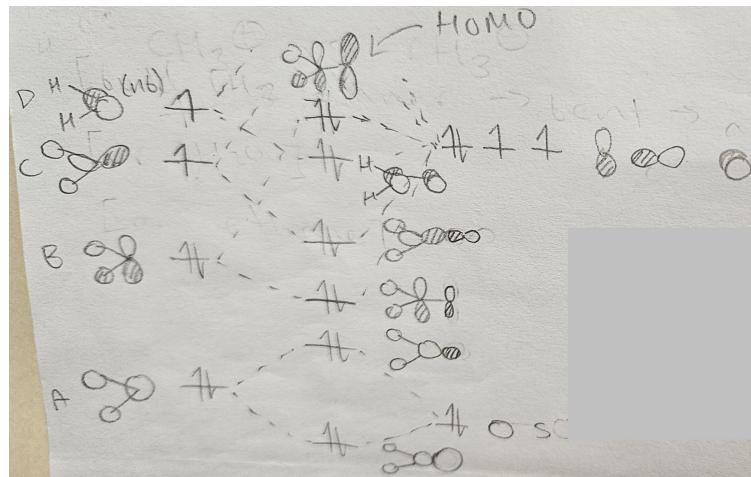


Figure 2.7: QMOT diagram for formaldehyde.

- The HOMO has a larger coefficient on O; this explains why protonation occurs on O and not C!
- Key takeaway: QMOT diagrams and MO diagrams both make the same predictions about the electronic structure and reactivity of formaldehyde (sanity check).
  - Example: They both predict that carbonyls are nucleophilic on oxygen.
  - Example: Orbital mixing is stronger when orbitals are of similar energy.
  - Example: Orbital coefficients are larger on an atom when the MO is closer in energy to the AO that originates with that atom.
  - Example: Orbitals are lower in energy on more electronegative atoms.
  - Etc.

## 2.2 Bonding Models 2

- 9/12:
- Lecture 2 recap.
    - QMOT for formaldehyde (see Figure 2.7).
    - Recall that the HOMO has a larger coefficient on oxygen, which means that protonation occurs on oxygen instead of carbon.
    - No other topics from Lecture 2 are reviewed.
  - Today: Bonding models (continued).

- Lecture outline.
  - Huckel theory.
  - Aromaticity.
  - Banana bonds.
  - Wave functions.
- **Huckel theory:** A quick way to build MOs for conjugated  $\pi$ -systems.
  - Qualitatively great and quantitatively bad.
    - Quick and dirty, but generates useful predictions.
    - Not *accurate*, but definitely *useful*.
  - It is used to analyze the connectivity and topology of the  $\pi$ -system in a planar molecule.
  - Key assumptions.
    - The  $\pi$ -system is independent of the  $\sigma$ -network.
    - You only consider valence electrons.
    - Only neighboring orbitals interact, i.e., only  $\pi$ -orbitals on adjacent atoms.
    - We ignore orbital overlap and electron repulsion.
  - These are some wild simplifications, but it is quick and useful!
  - Rules.
    - The number of  $p$ -AOs you mix equals the number of new MOs you make.
    - The energy of the new MOs is distributed symmetrically around the **nonbonding energy level**.
    - The number of nodes increases by 1 with each energy level.
    - The MOs reflect the symmetry of the molecule.
- **Nonbonding energy level:** The energy of the nonbonding MOs in a Huckel diagram. *Denoted by  $\alpha$ .*
  - This is also the energy of an electron in an empty  $p$ -AO.
- Example Huckel diagram: Ethylene.

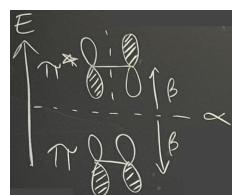
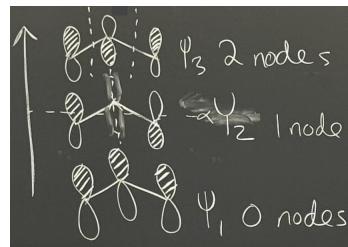


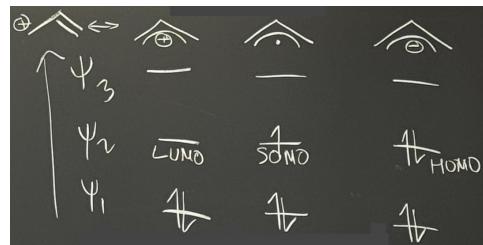
Figure 2.8: Huckel diagram for ethylene.

- Let's first confirm that this diagram meets all four Huckel theory rules.
  - We get two new  $\pi$ -MOs from two  $p$ -AOs.
  - The energy difference from the nonbonding energy level is called  $\beta$ .
  - The number of nodes did increase from 0 to 1.
  - The MOs are symmetric.
- Thus, this is a valid Huckel diagram!
- Note: Do remember that symmetric splitting is *not* accurate!
  - On Tuesday, we (correctly) learned that destabilization energy > stabilization energy.

- Example Huckel diagram: Allyl groups.



(a) Diagram.



(b) Filling orbitals.

Figure 2.9: Huckel diagram for allyl groups.

- The lowest energy orbital is called  $\psi_1$ .
  - It has 0 nodes.
- The middle energy orbital is called  $\psi_2$ .
  - To maintain symmetry, we have to delete the middle orbital and give opposite phases.
- The highest energy orbital is called  $\psi_3$ .
  - It has the 2 nodes we expect.
- We now fill electrons for the allyl cation, radical, and anion (Figure 2.9b).
  - These species have 2, 3, and 4 electrons, respectively.
- Now let's look at where each of these species will react.
  - Nucleophiles will attack the LUMO of the cation.
  - Radicals react with their SOMO (singly occupied molecular orbital).
  - Electrophiles will engage the HOMO of the allyl anion.
- But the LUMO, SOMO, HOMO are all  $\psi_2$ !
  - $\psi_2$  has no density at the middle carbon, so all of these species should only react at the terminal carbons.
  - This prediction of Huckel theory is experimentally confirmed!
  - Intuitively, reacting at the terminals allows you to keep the double bond in play; thermodynamically, you wouldn't want to cleave it by reacting in the middle.

- Example Huckel diagram: Benzene.

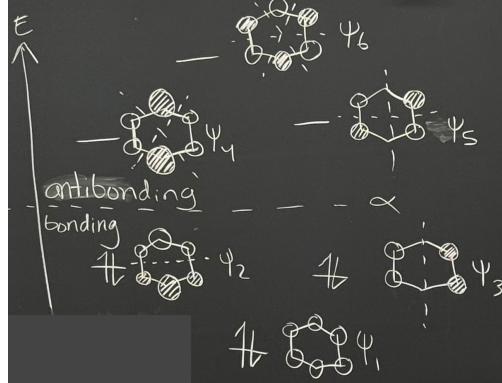


Figure 2.10: Huckel diagram for benzene.

- For cyclic systems, we draw a **Frost circle**.
  - For benzene, the radius of the Frost circle is  $2\beta$ .
- We create  $\psi_1, \dots, \psi_6$ .
  - $\psi_2, \psi_3$  and  $\psi_4, \psi_5$  are degenerate.
  - No electron density on the central  $p$ -orbitals in  $\psi_3$  implies bigger coefficients on the corresponding orbitals in  $\psi_2$ .
    - See Anslyn and Dougherty (2006) for more!!
  - $\psi_4, \psi_5$  have 2 nodes at angles.
  - For  $\psi_6$ , we have 3 nodes through a hexagon, which is alternating shading.
- $\alpha$  is the nonbonding level; higher is antibonding, lower is bonding.
- 6 electrons in benzene's bonding  $\pi$ -system yields stabilization.
  - In particular, we observe stabilization relative to three ethylenes: An extra 36 kcal/mol of stabilization!
  - Huckel theory can't really compare energy between two molecules;  $\beta$  is more a qualitative parameter than a quantitative one.
- **Frost circle:** A circle in which we inscribe a regular  $n$ -gon with one point down — where  $n$  is equal to the number of carbons in the cyclic system — that is used as a guide for drawing Huckel orbitals.
- Example Huckel diagram: Cyclobutadiene.

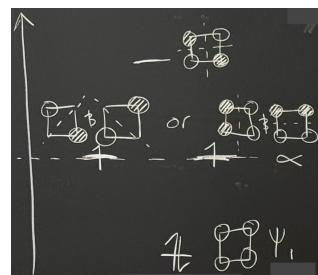


Figure 2.11: Huckel diagram for cyclobutadiene.

- $\psi_1$  has no phase inversion and no nodes.
- There are two different ways to draw the orbitals for  $\psi_2, \psi_3$ .
  - We can dive deeper into this difference in Anslyn and Dougherty (2006).
  - No extra stability relative to three ethylenes!
  - The model also predicts a ground-state triplet diradical.
    - Indeed, this molecule is highly reactive and dimerizes spontaneously at 35 K.
- We now do a deep dive into aromaticity.
- The history of aromaticity.
  - In 1855, Hofmann (not Hoffmann) coins the term “aromatic” because these compounds were smelly.
  - In 1861, we have Kekulé’s dream of a snake eating its tail.<sup>[1]</sup> This inspired a circle of electrons.
  - In 1925, Robinson describes aromaticity as extra stabilization of a molecule.
  - In 1931, Huckel puts forth **Huckel’s rule**.

<sup>[1]</sup>“At least, Kekulé said it was a dream!” - Masha. Good use of reasonable doubt and objectivity in her thinking!

- **Huckel's rule:** Cyclic, planar molecules with  $4n + 2$  continuous  $\pi$ -electrons are aromatic.
  - If you have  $4n$  electrons in a cyclic planar molecule with continuous  $\pi$ -electrons, then you are antiaromatic (extra unstable).
  - Thus, these molecules usually distort out of the plane to break antiaromaticity and become nonaromatic.
    - Both cyclobutadiene and cyclooctatetraene are antiaromatic. Cyclooctatetraene bends into a boat so that its  $\pi$ -orbitals are pointing toward each other.
  - No phase inversions are allowed; we must connect orbitals without crossing the  $\sigma$ -plane.
    - What does this mean??
- Features of aromatic compounds.
  - Aromatic stabilization energy (36 kcal/mol).
  - Equalization of the bond lengths.
    - Essentially, the bond lengths do not alternate but rather share an identical bond order of 1.5.
  - Ring currents and magnetic properties.
    - Those interested in polymer chemistry might be interested in exploiting these properties!
    - Specifically, these are properties that come from a sea of electron density.
  - Benzene vs. hexa-1,3,5-triene.
    - In benzene, all bond lengths are 1.40 Å.
    - In hexa-1,3,5-triene, the single bonds are 1.45 Å, the terminal double bonds are 1.34 Å, and the internal double bond is 1.37 Å.
    - The bond lengths of benzene equalize because benzene has two equally stable major resonance structures.
      - This is why we often draw benzene as a hexagon with a circle in the middle: This is actually the most accurate picture of it!
    - The bond lengths of hexa-1,3,5-triene do *not* equalize because the only resonance structure we can draw of it is a zwitterion, and thus will be a minor contributor.
  - Different kinds of reactivity.
    - Example: Electrophilic aromatic substitution.
    - This is very much distinct from alkene addition chemistry.
- **Möbius aromaticity:** Aromatic rings have one phase inversion (PI), like in a Möbius strip.

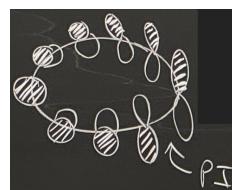


Figure 2.12: Möbius aromaticity.

- This is a different definition of aromaticity.
  - We could research aromaticity for the rest of our lives if we wanted to.
  - There's a whole field of research devoted to it, and we should look into it if we're interested!!
  - A good starting point is Ajami et al. (2003).
- The single phase inversion is called a **Möbius topology**.

- Your PI happens at the sole node.
  - This one node is allowed in Möbius aromaticity, but not in Huckel aromaticity
- The Möbius topology predicts that compounds are aromatic if they have  $4n$  electrons and antiaromatic if they have  $4n + 2$  electrons.
- To be clear, this content is outside the scope of this class, but Masha wants us to know about it and be able to research it if we so choose.
- Ring current.

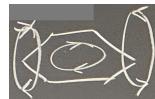


Figure 2.13: Ring current.

- Suppose you have an external magnetic field perpendicular to the  $\sigma$ -plane.
  - This would induce the  $\pi$ -electrons to rotate through their MOs.
  - These rotating electrons would then create an additional magnetic field.
  - This new magnetic field would *reinforce* the external magnetic field outside the aromatic ring and *oppose* the external magnetic field inside the ring.
  - The strength of the induced magnetic field is proportional to the current (i.e., the size of the ring).
- Application (NMR): Ring protons are deshielded (higher  $\delta$ ) outside and shielded (lower  $\delta$ ) inside.
  - Cyclohexene: No ring current, so we get a bit of downfield shift for the vinyl protons ( $\delta$  5.6).
  - Benzene: Has a ring current, so we get a noticeable downfield shift ( $\delta$  7.3).
  - [18]annulene: Has a large ring with many  $\pi$ -electrons, so we get a significant downfield shift for the external protons ( $\delta$  9.3) and a significant *upfield* shift for the internal protons ( $\delta$  -2.9).

- **Quadrupole:** Two dipoles aligned such that there is no net dipole.

- Example: The dipole aligned up and down in benzene — perpendicular to the  $\sigma$ -plane of the molecule — as opposed to (for instance) the linear dipole in fluoromethane.
- Lots of applications beyond the scope of this class, but we can look into it if we want.

- **Banana bond:** A bent chemical bond that contains an unusually high concentration of  $p$ -character.

- The bent  $p$ -lobes of banana bonds look like bananas (see Figure 2.14), hence the name.
- Example banana bonds: Cyclopropane.

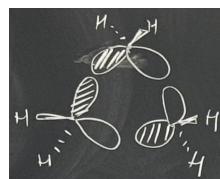


Figure 2.14: Banana bonds in cyclopropane.

- Cyclopropane needs more  $p$ -character because of its  $60^\circ$  bond angles;  $p$ -character helps bonds bend.
  - Specifically, the C–C bonding orbitals in cyclopropane are  $sp^5$ -hybridized.

- The excess of *p*-character in the C–C bonds means that the C–H bonds of cyclopropane have correspondingly more *s*-character.
  - This makes the C–H bonds in cyclopropane shorter than usual!
  - Indeed, there is something of a “conservation” of bonding character: The *s*-character that’s not in the  $\sigma$ -bonds has to go somewhere.
- Group orbitals (HOMO) degenerate.
  - The **Walsh orbitals** have more  $\pi$ -character, so cyclopropane is  $sp^2$ -like.
  - This means it is a good  $\pi$ -donor and a bad  $\pi$ -acceptor.
  - Example of donation: The dimethylcyclopropyl cation is very stable because all of the  $sp^2$ -character is getting donated into the carbocation’s empty orbital. See Figure 4.4c for more.
- Wave functions.
  - Review Anslyn and Dougherty (2006), Chapters 4 & 14!!
    - Also look up your Gen Chem or Quantum notes if it’s been a while.
    - Is there anything relevant to review in Chapter 4??
  - All bonding theories draw upon QM descriptions of electrons as waves existing in **orbitals**.
- **Orbital:** A wave function that is a specific solution to the **Schrödinger equation**.
  - Masha draws the  $1s, 2s, 3s$  orbital penetration graph, as well as what these orbitals look like.
  - Recall that orbitals have **lobes** and **nodes**!
- **Schrödinger equation:** The following equation, where  $E$  is the energy of the electron,  $\psi$  is the wave function describing the position of the electron in space, and  $H$  is the **Hamiltonian operator**. *Given by*

$$H\Psi = E\Psi$$
  - $\psi^2$  is the probability of finding an electron in a specific position (i.e., the electron density!).
  - Big  $\Psi$  is the total molecular wave function, and little  $\psi$  is a molecular orbital.
- **Hamiltonian operator:** A representation of all forces acting on the system, such as the kinetic energy of the electron and nucleus, nuclear-nuclear repulsion, electron-electron repulsion, etc.
- Next week, we’ll talk about DFT and approximating solutions to the Schrödinger equation. It will be like an intro to computational chemistry!
- Example: Electron density in  $H_2$  MOs.
  - Masha redraws Figure 2.2 to start, and Figure 7.3 from Labalme (2023).
  - The point is that...
    - The bonding MO has a lot of electron density between the nuclei, even though you still have some at the atoms;
    - The antibonding MO has minimal to no electron density between the nuclei; the AOs ( $\phi_1^2, \phi_2^2$ ) are very separate.

## 2.3 Chapter 1: Introduction to Structure and Models of Bonding

*From Anslyn and Dougherty (2006).*

9/23: • Good outline of the purpose of Chapter 1.

- Mostly along the lines of what we’ve talked about in class.

- Why bother with simplistic bonding models if we can just compute everything quantum mechanically nowadays?
  - “A string of computer-generated numbers is just no substitute for a well-developed feeling for the nature of bonding in organic molecules” (Anslyn & Dougherty, 2006, p. 3).
  - “It is still true — and will be true for some time — that descriptive models of bonding that are readily applicable to a wide range of situations are the best way to attack complex problems” (Anslyn & Dougherty, 2006, p. 4).

## Section 1.1: A Review of Basic Bonding Concepts

- Vocab for Gen Chem-level quantum mechanics.
  - Most of these I know without review.
- **Spin paired** (electrons): Two electrons in the same orbital with opposite-signed  $m_s$  values.
- **Correlation:** The ability of an electron to feel the trajectory of another electron and therefore alter its own course so as to minimize Coulombic repulsions and keep the energy of the system to a minimum.
- The strengths and weaknesses of Lewis structures.
  - Pros: Predict the number of bonds an atom forms, whether it has lone pairs, and whether any double or triple bonds form.
  - Cons: Does not describe the structure or reactivity of any given species.
- Example: Why formal charge is much more a method of “bookkeeping” nowadays than anything accurate.
  - In the tetramethylammonium cation, we put the formal positive charge on the nitrogen.
  - However, computational studies show that since nitrogen is more electronegative than carbon, a  $\delta^-$  rests on nitrogen in the actual structure, and every carbon shares  $\frac{1}{4}$  of the positive charge.
- VSEPR’s geometrically perfect bond angles are only observed in simple, symmetric molecules.
  - However, words like “tetrahedral” and “trigonal” are broadly used to suggest an idea, even when they’re not *strictly* accurate.
- VSEPR “is not based on any first principles analysis of electronic structure theory. It is a simple way to rationalize observed trends” (Anslyn & Dougherty, 2006, p. 8).
  - Indeed, it is not clear that bonding orbitals or lone pairs really *have* any well-defined notion of size.
  - Note that singly occupied orbitals (e.g., radicals) do not have repulsive effects in VSEPR because they are able to bond to doubly occupied orbitals.
- **Steric repulsion:** The buttressing of filled orbitals that cannot participate in bonding, where the negative electrostatic field of the electrons in the orbitals is repulsive.<sup>[2]</sup>
- **Hybridization:** The method of adding and subtracting atomic orbitals on the same atom.
  - “Remember that orbitals are mathematical solutions to the Schrödinger equation, and that the addition and subtraction of mathematical equations is just an exercise in algebra. It is a perfectly valid operation to add orbitals as long as one also does the corresponding subtraction” (Anslyn & Dougherty, 2006, p. 9).

<sup>[2]</sup>Don’t forget that Alison Wendlandt believes that this definition of sterics is fundamentally flawed, and that overlap is actually beneficial to a point!

- Someday, I should take linear combinations of *s* and *p* wave functions and confirm that they're still orthogonal and in the solution space of the Schrödinger equation!!
  - It's called hybridization because we're literally forming *hybrids* somewhere between the polar extremes of an *s* and *p* orbital!
  - As with VSEPR bond angles, VBT hybridizations deviate from the *sp*, *sp*<sup>2</sup>, or *sp*<sup>3</sup> ideal in most organic molecules, but we loosely retain these terms regardless to convey an idea.
- Anslyn and Dougherty (2006) goes over a cool connection between the hybridization index (as defined in class) and experimentally observed <sup>1</sup>H-<sup>13</sup>C NMR coupling.
- The localization of electrons in a chemical bond per VBT is exactly the impression of bonding that is given by a Lewis structure!
- Some useful content on polar covalent bonding that is slightly beyond the scope of the class.
  - “Introducing polarity into a bond strengthens it” (Anslyn & Dougherty, 2006, p. 12).
  - Pauling electronegativity is based on the BDEs of molecules, while Mulliken electronegativity is based on the IEs of atoms.
    - Additional electronegativity scales by Nagle, Allen, Sanderson, Allred-Rochow, Gordy, Yuan, and Parr.
    - Takeaway: Electronegativity is a hard concept to put your finger on!
  - The real use of electronegativity is in comparing *relative* electronegativities, and all scales more or less agree here.
- **Field effect:** The withdrawing of electron density through space, rather than through  $\sigma$ -bonds.
- Additional irrelevant content.
- Application of quadrupoles: Proving that *sp*<sup>2</sup>-C is more electronegative than H.
- **Resonance energy:** The energy of stabilization imparted by resonance. *Also known as delocalization energy.*
- Why is delocalization stabilizing from the point of view of quantum mechanics?
  - Recall that in the particle in a box, the energy levels are given by
$$E_n = \frac{n^2 h^2}{8mL^2}$$
  - As *L* increases (i.e., as we expand the box/delocalize), the energy goes down.
- Lots more on resonance.

## Section 1.2: A More Modern Theory of Organic Bonding

- Molecular orbital theory has the same predictive power as the models in Section 1.1 (Lewis structures, VSEPR, and VBT), but it can better explain certain structural issues and experimental observations, too.
  - MO theory extracts certain key concepts and trends that result from the output of quantum mechanical calculations to lead to a more rigorous, descriptive model of organic bonding than Lewis structures, VSEPR, and VBT can.
- VBT and MO theory are often fairly interconvertable mathematically!
  - Thus, it is *not* necessarily true that MO theory is “better” than VBT.
  - Rather, both models (and any combination thereof) are approximations of the true answer — a full solution to the Schrödinger equation.
- A more detailed analysis of the QMOT examples from class; definitely worth returning to!!

### Section 1.3: Orbital Mixing — Building Larger Molecules

- Goes over the mixing of fragments/group orbitals in detail. Relevant to PSet 1!!
- The later sections would likely be quite useful for my development as a chemist, but probably aren't immediately relevant to this course.

## 2.4 Chapter 14: Advanced Concepts in Electronic Structure Theory

*From Anslyn and Dougherty (2006).*

### Section 14.1: Introductory Quantum Mechanics

- Some basic and some more advanced quantum mechanics, but all stuff with which I am eminently familiar from my undergrad coursework.

### Section 14.3: A Brief Overview of the Implementation and Results of HMOT

- **HMOT:** Huckel molecular orbital theory.
- A much more mathematical treatment of Huckel theory, similar to what I saw in CHEM 26100. It does rationalize the benzene coefficients, though, so probably worth returning to!!

### Section 14.5: Some Topics in Organic Chemistry for Which Molecular Orbital Theory Lends Important Insights

- More mathematics of aromaticity, cycles, etc. Probably a bit less useful.

# Week 3

## Applications of Bonding Theory

### 3.1 Computational Chemistry

9/17:

- Lecture 3 recap.
  - Huckel theory: A fast way to draw the MOs of conjugated  $\pi$ -systems.
    - If the conjugated  $\pi$ -system in question is cyclic, use a Frost circle.
  - Aromaticity.
    - Huckel's definition:  $4n + 2$ .
    - Möbius's definition:  $4n$ .
    - Leads to properties like stabilization, quadrupoles, and ring current.
  - Cyclopropane:  $sp^2$ -like banana bonds (the only thing we need to remember from that discussion).
  - Wavefunctions: Solutions to the Schrödinger equation.
- Today: Computational chemistry (an overview).
  - Computational chemistry is typically an entire class!
- Lecture outline.
  - Methods of computational chemistry.
  - Molecular mechanics.
  - Semi-empirical methods.
  - Ab initio methods.
  - Hartree-Fock.
  - Density functional theory (DFT).
  - Best practices for calculations.
  - Properties that are especially easy (or hard) to calculate.
- Why do we do computational chemistry?
  - If we could fully solve the Schrödinger equation, we could know the properties of all of our electrons!
  - However, the Schrödinger equation can only be fully solved (practically) for the simplest systems.
    - For now, at least: People are working on this.
  - As such, we *approximate* solutions instead.

- **Computational chemistry:** The science of approximating solutions to the Schrödinger equation.
  - Computational chemistry can be broken up into two general strategies (**ab initio** and **empirical** methods) and one in-between strategy called **semi-empirical** methods.
- **Ab initio** (methods): Make well-defined approximations to the Schrödinger equation, and then solve the approximations mathematically. *Etymology* from Latin “from first principles.”
  - Essentially, make your math simpler.
- **Semi-empirical** (methods): Replace complicated parts of the Schrödinger equation with experimentally derived parameters, such as bond lengths, vibrational frequencies, and more that we can get from spectroscopy.
  - Essentially, shortcut the hardest parts of solving with experimentally derived features.
- **Empirical** (methods): Approximate molecules with force fields that are experimentally derived, and adjust with further experimental parameters.
  - Essentially, start with reality and derive computational things from that.
- We now look at some commonly derived methods. The following list is sorted from methods with high **accuracy** and low **speed** to methods with low accuracy and high speed.
  - Methods at the high end of accuracy and the low end of speed (ab initio).
    - **Coupled cluster**.
    - **Perturbation theory**.
    - **Density functional theory**.
    - **Hartree-Fock**.
  - Methods in the middle (semi-empirical).
    - **Semi-empirical methods**.
  - Methods at the high end of speed and the low end of accuracy (empirical).
    - **Molecular mechanics**.
- **Speed:** Ease of calculations.
- **Accuracy:** Careful and diligent.
- **Coupled cluster:** Useful for approximately 10 **heavy atoms**. *Also known as CC.*
- **Density functional theory:** Useful for approximately 80 heavy atoms, though we can use more (it just gets slower). *Also known as DFT.*
- **Hartree-Fock.** *Also known as HF.*
- **Molecular mechanics:** Useful for hundreds of heavy atoms. *Also known as MM.*
- **Heavy atom:** Any atom that's not hydrogen.
- In this course, we'll discuss further the bottom four methods in the above list of six.
- Molecular mechanics (MM).
  - Atoms are treated as balls and springs (this is a classical analogy and thus much easier to simulate).
  - We use force fields to describe electrons.
    - These force fields are derived from experimental data, i.e., choose a force field that gives us the bonds we calculate from XRD or the vibrations we see in IR.
  - Very fast; often considered “quick and dirty.”

- Gives us a general picture of what we're thinking about.
- Common application: Very large and flexible systems.
  - Think proteins, polymers, etc.
  - Things that have a lot of degrees of freedom.
  - Very useful for chembio, polymer chemistry, etc.
- Subset application: **Molecular dynamics (MD)**.
  - Simulating movement; uses MM as a basis.
- If you're going to use this method, know that it is (in general) only appropriate for approximating the ground states of molecules (not their transition states).
  - However, MM can be a good starting point for higher-level calculations (i.e., more accurate methods).
  - In Orgo, it's mainly used for first approximations to be refined later (and for heavier stuff).
  - All the same, it is a super useful tool with tons of applications, and its simplicity should not lead us to discount it.
- Running MM.
  - If we have a PC, try clicking the MM2 button in Chem3D (which is part of our ChemDraw package).
  - This may not work on Macs; figure this out!!
  - PerkinElmer (who developed ChemDraw) initially developed their stuff for Macs; Masha's not quite sure where they dropped the ball.
- Semi-empirical quantum mechanical (SQM) methods.
  - Use empirical parameters to simplify *ab initio* calculations.
    - Tries to deliver the best of both worlds (speed and accuracy).
  - We can add corrections for missing phenomena and underestimated features.
  - Theoreticians (developers) will draw the line on accuracy somewhere, and then organic chemists will say, "this model fails here."
    - Once that feedback gets into the literature, theoreticians redefine their line.
    - They might need to account for *d*-orbitals, London dispersion forces (LDFs), flexibility, solvent, or more.
      - Methods of accounting for solvent effects are continuously being optimized.
      - It's important to be on top of the literature here, since things are always getting better!
  - Modern implementations (these are getting fast enough to be usable and really good!).
    - Density function based tight binding (DFTB): Approximate DFT.
    - eXtended Tight Binding (xTB)
      - Developed primarily by the Grimme lab.
      - Basically just adding more parameters.
  - LDFs are becoming increasingly important for selective catalysis, so there's a lot of work to approximate them.
    - Catalysis is not about partial positive and negative charges so much as it is about electrons flopping around to achieve incredible selectivities in next-gen catalysts.
  - Very fast (seconds) and pretty accurate. Increasingly used, especially for ML and data science.
    - Nowadays, if you want to do ML, you need these hundreds of experimental data points.
- Ab initio methods.
  - Background theories (neither is technically true, but it is helpful for speed).

- **Born-Oppenheimer approximation.**

- **Independent electron theory.**

- **Born-Oppenheimer approximation:** Nuclei are way bigger than electrons (have over 1000 times more mass), so they are basically fixed in space relative to the electrons.
  - This means that you can treat the nuclei separately; you can use one approach for the nuclei and an entirely different approach for the electrons.
- **Independent electron theory:** Electron movements are not correlated to each other; all electrons whiz around independently.
  - Making this approximation will cause some issues.
- Hartree-Fock (HF).
  - Treat electrons as a delocalized cloud with independent electron movement.
    - Remember the plum pudding model of the atom? This is not that dissimilar from that.
  - This approximation ignores Coulombic interactions (like LDFs).
    - This becomes very problematic for transition states.
  - HF methods are largely historical today.
    - There are applications where they're still used today, but not in Orgo and not without an understanding of their shortcomings.
- We can run any and all of these computations throughout grad school as MIT students, and we should! They're in our toolbelt now, and we should try them out!!
- Density functional theory (DFT).
  - Instead of calculating wavefunctions, we're going to calculate electron density.
    - We're going to do this using **functionals**.
    - We can include functionals for things like Coulombic interactions, etc.<sup>[1]</sup>
  - This is a good workhorse method in organic chemistry.
    - DFT is appropriate for reaction coordinate mapping, transition states, etc.
    - We'll often work with collaborators that can tailor a model to our needs.
  - There are many specific functionals and basis sets.
    - You have to choose the functional (choose what to include), and then choose the basis sets (how much detail do I need for this calculation, e.g., treating polarization, charge, unpaired electrons more accurately).
    - It is best to find a basis set and functional appropriate for our context.
  - Basis sets don't describe all types of elements.
    - Some describe elements 1-30, others do 1-86.
    - Don't be that person who has to redo their entire calculation because they forgot that tin is one of their reagents!
    - We often use **split basis sets** (esp. for transition metals), i.e., certain atoms (i.e., metals) get more functionals.
      - Carbon, hydrogen, and oxygen (CHO) don't need the craziest level of theory to approximate, but that palladium center will!
    - Think about what level of theory you need for each atom.

<sup>[1]</sup>Maybe what I can be known for in research is custom building computational tools for specific organic problems, and turning that into a workflow that people do. Maybe that's what ML already is.

- **Functional:** A function of functions.<sup>[2]</sup> *Also known as higher-order function.*
- Best practices for running calculations for our own things.
  - This part of the lecture is *critical*; it tells us what we need to know to use computational chemistry.
    - If we want to learn the theory for all of these things, we should read a textbook or take Heather Kulik's class.
  - Use the appropriate level of theory for your needs and capabilities.
  - Questions to ask yourself to assess your needs and capabilities.
    - Do you have a supercomputer? How much time on the supercomputer do you have?
    - When do you need this result by? Is your PI breathing down your neck?
    - What am I trying to model?
    - Is this a thought experiment or something serious?
  - Additional things to consider wrt your needs and capabilities.
    - Consider speed vs. accuracy.
      - You can always start at a lower level theory and then ramp it up if you need more accuracy. This is a great general approach.
    - Consider size and flexibility (no HF on proteins, or MM on methane).
    - Consider “weirdness”: If you've got something that's all inverted and Möbius like, you're gonna need something more tailor-made.
    - Find a *reliable* literature precedent for a similar system.
      - If you want to model a cationic cyclization, use a precedent paper's level of theory.
      - How do I model an iridium catalyst? Find an iridium catalyst paper and go from that!
    - Know how your level of theory works.
      - Does it account for polarizability? Charge? Solvent? *d*-orbitals?
      - It is our responsibility as an experimentalist to know this if we're going to publish it; our PI probably won't be as deep into the nitty-gritty as us.
  - Don't blindly trust calculations.
    - Calculations always give you an answer (unless they fail or don't converge). However, just because you get a number doesn't mean that that number is accurate!
    - Benchmark your calculation with experiments whenever we can. Examples: X-ray structure, ratio of products (we can back-calculate from temperature the activation energy barrier, and from the transition states what ratio of products we expect).
    - Redo a couple of calculations at a higher level of theory to see if you get the same answer.
    - Chris Cramer (a founding father of computational chemistry): “There is no particular virtue to the speed at which a wrong answer can be obtained.”
  - Example of doing calculations wrong: Doing an S<sub>N</sub>1 reaction without solvent. These reactions are so solvent-dependent, and there's no gas-phase cation that will replicate this solution-phase reaction.
  - What's easy to calculate?
    - Spectra: IR, Raman, NMR.
      - Masha likes to predict the NMR spectra of wacky intermediates.
      - ChemDraw does this for free.
        - The default solvent is THF; make sure you change it to CDCl<sub>3</sub>!
      - MNova's function is better; it's ML-based, but it also costs money to run?? I think Masha has this wrong for MIT students.

<sup>2</sup>This is the computer science definition; it is largely unrelated to the mathematical definition that is equivalent to linear forms and duals.

- Geometries, conformers, and ground state structures.
  - “Geometry optimization” or “energy minimization” is very common.
  - Draw a 3D structure, give it to our program, move atoms, calculate  $E$ , repeat (let the program perturb the atom’s positions a bit) until we reach a *local* minimum.
  - This is what we’ll do on the problem set.
  - If we want to get the *global* minimum, we have to look for lower energy structures (manually, automatically, or a combination of both).
  - We often start at a low level of theory and then refine. Start with a search of the chemical space to find some stable conformers, and then pop that into DFT.
- Frequencies, well-defined transition states, and **single point calculations**.
  - Important because transition states are saddle points on the potential energy surface with 1 imaginary frequency corresponding to the bond-making or -breaking event.
  - If you have a structure that you think is a ground state, you have to prove this.
- **Single point calculation:** Calculating the energy of a structure without any other atoms around.
- Note: Nucleophiles typically come in at a  $120^\circ$  angle (the **Bürgi-Dunitz angle**), because that’s where it’s easiest to donate into the  $\pi^*$ -lobe.
- What’s “hard” to calculate?
  - Caveat: Do your research!!
    - Many applications require specialized approaches.
    - There’s an army of computational chemists who are trying to develop niche methods for our little problem; find them, connect with them, collaborate with them, etc.
    - It is our responsibility to know what part of a certain experiment is difficult.
      - Example: In photophysics, you need to know the limitations of certain parts of our model.
      - The system will not say, “I’m bad at predicting excited states;” you have to know that.
  - Things that require more finessing to study.
    - Open-shell species, e.g., radicals.
    - Transition metals: Heather researches how to model TMs with SQM, etc. This is really important and really hard.
    - “Unusual structures,” e.g., gas-phase plasma nonsense.
  - Thermochemistry.
    - E.g., thermodynamical parameters, calorimetry, etc.
    - There are specific packages that work for this, if we need to look into them.
    - Masha doesn’t know anything about any of this, but recommends that we can learn them!!
  - Solvent effects, LDFs, etc.
    - For these topics, the methods are getting better all the time (which is code for, “the programs don’t work great yet”).
    - **Implicit solvation vs. explicit solvation.**
- **Implicit solvation:** Treat the solvent as a continuous medium.
- **Explicit solvation:** Draw solvent molecules and add them to the calculation.
  - If you think the solvent is stabilizing the transition state in an  $S_N1$ , you need to draw a little THF donating its lone pair to the carbocation.
  - This gets complicated in proteins, where we need to identify how many waters we need to draw to accurately represent water; recent papers suggest that effects matter up to 44  $H_2O$  molecules away!

## 3.2 Pericyclic Reactions

9/19:

- Lecture 4 recap.
  - Masha redraws the accuracy/speed list of computational techniques.
  - She also rewrites the Cramer quote.
  - Basically, Tuesday was the methods, theory, and applications of computational chemistry.
- Announcements.
  - PSet 1 posted.
  - Conference room booked this afternoon for collaboration on the PSet.
- Today: Pericyclic reactions.
  - Reading: Anslyn and Dougherty (2006), Chapters 15 (pericyclics in general) and 16 (photochemical pericyclics).
  - See Jonathan's reading list!!
- Lecture outline.
  - Pericyclic reaction types and vocabulary.
  - History of pericyclic reactions.
  - Woodward-Hoffmann rules.
  - Dewar-Zimmerman analysis.
  - Frontier molecular orbital theory.
  - Miscellaneous pericyclic reactions.
- **Pericyclic** (reaction): A **concerted** (as opposed to **stepwise**) reaction with a transition state (TS) consisting of a cyclic<sup>[3]</sup> array of atoms and orbitals. *Antiquated thermoreorganization*.
  - Informal definition: “If you can draw circle arrows, it’s a pericyclic reaction.”
  - Can be **synchronous** or **asynchronous**.
    - Indeed, sometimes we see asynchronous concerted Diels-Alders! These make us ask, “is the TS truly symmetric, or are some bonds longer or shorter?”
  - There are 5 (main) types of pericyclic reactions. The first three are “the big three,” and the latter two are less common.
    1. **Electrocyclizations**.
    2. **Cycloadditions**.
    3. **Sigmatropic rearrangements**.
    4. **Group transfers**.
    5. **Chelotropic reactions**.
    6. Etc.
  - This lecture assumes that we’ve seen all of these; if we haven’t seen these reactions before in undergrad or if it’s been a while...
    - Read the textbook;
    - Review your notes from undergrad;
    - Do some Googling (the Wikipedia pages are pretty helpful!!);
    - Ideally, do all of the above!

<sup>3</sup>This stands in contrast to “normal” organic reactions, which prefer to proceed through a *linear* TS.

- **Concerted** (mechanism): A mechanism with no **intermediates**.
- **Stepwise** (mechanism): A mechanism *with* intermediates.
- Concerted and stepwise mechanisms can be differentiated based on their respective energy diagrams.

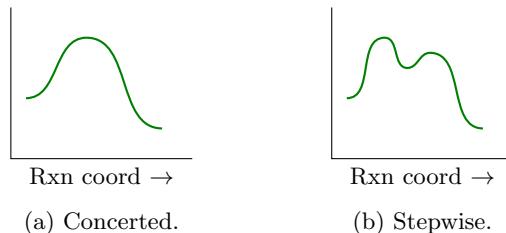
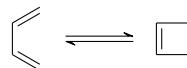


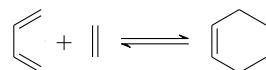
Figure 3.1: Concerted vs. stepwise energy diagrams.

- **Intermediate**: A local ground state structure.
- **Synchronous** (mechanism): All bond-making and -breaking occurs to an equal extent in the TS.
- **Asynchronous** (mechanism): All bond-making and -breaking does *not* occur to an equal extent in the TS.
- David: Would the fact that bond breaking/making happens more sequentially in an asynchronous mechanism imply that these reactions have energy diagrams that differ from the synchronous, concerted ideal of Figure 3.1a?
  - The energy diagram *is* different between synchronous and asynchronous.
  - Look into Dean Tantillo at UC-Davis for more!!
- **Electrocyclization**: A pericyclic reaction in which one  $\pi$ -bond gets converted into one  $\sigma$ -bond or vice versa. *Also known as electrocyclic reaction. Denoted by  $m\pi$ .*



– Example: We can refer to the above reaction as a “ $4\pi$  electrocyclization.”

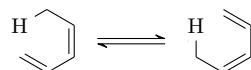
- **Cycloaddition**: A pericyclic reaction in which two or more unsaturated molecules (intermolecular) — or parts of the same molecule (intramolecular) — combine to form a cyclic adduct with a net reduction of bond multiplicity. *Denoted by  $[m + n]$ .*



– Example: We can refer to the above reaction as a “[ $4 + 2$ ] cycloaddition.”

– This specific cycloaddition is also known as a **Diels-Alder reaction!**

- **Sigmatropic rearrangement**: A pericyclic reaction in which a  $\sigma$ -bond migrates along with a corresponding reorganization of the  $\pi$ -electrons. *Also known as sigmatropic reaction. Denoted by  $[m, n]$ .*



– Example: We can refer to the above reaction as a “[ $1, 5$ ]-sigmatropic hydride shift.”

- **Group transfer** (reaction): A reaction that transfers atoms from one molecule to another, but in a concerted pericyclic transition state.
  - **Chelotropic** (reaction): A cycloaddition in which two bonds are made to one atom.
  - **Diels-Alder** (reaction): A  $[4 + 2]$  cycloaddition.
  - Aside (chemis-tea): One of Steve Buchwald's pet peeves.
    - Don't erase the chalkboard with your fingers; use the eraser.
    - If Steve is on your thesis committee, the first thing he'll tell you is to use the eraser.
  - History of pericyclic reactions.

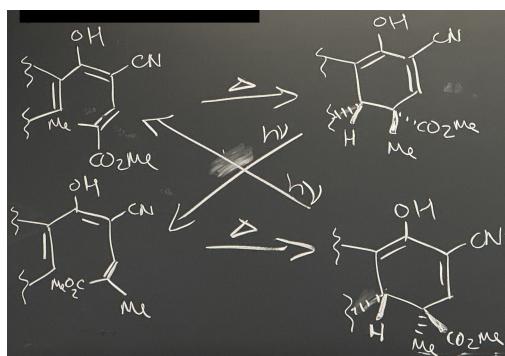


Figure 3.2: First observation of a pericyclic mechanism.

- Long considered to have “no mechanism.”
    - This is because people saw a starting material and a product, with nothing in between.
    - Quote from the ’60s: “No-mechanism is the designation, given half in jest, half in desperation, to thermoreorganization reactions” – von Eggers Doering and Roth (1962).
  - In 1966, Woodward and his army of grad students were synthesizing Vitamin B<sub>12</sub>. During one particular ring-closing ‘thermoreorganization’ reaction, they noticed that using different geometric isomers as starting materials formed different stereoisomers as products (see Figure 3.2).
    - This gave a hint!
    - Shortly after, Woodward and colleagues made a second (even weirder) observation: In the presence of light, the cyclized products would revert to their starting materials, but to the *opposite* geometric isomer (also see Figure 3.2)!
  - These observations kickstarted a series of studies into the mechanism of such reactions.
  - In time, we came to classify the reaction in Figure 3.2 as a “6π electrocyclization,” governed by the mechanism described as follows.



Figure 3.3:  $6\pi$  electrocyclization mechanisms.

- During a **thermal** electrocyclization, the termini of the  $\pi$ -systems rotate in opposite directions.
  - Notice how in Figure 3.3a, both *exo* groups rotate down, but the right one rotates clockwise and the left one rotates clockwise.
- During a **photochemical** electrocyclization, the termini of the  $\pi$ -system rotate in the same direction.
  - Notice how in Figure 3.3b, both axial groups rotate clockwise.
- This “rotation” of the  $\pi$ -systems’ termini is classified as **disrotatory** and **conrotatory**, respectively.
  - The disrotatory/conrotatory phenomenon led us to the **Woodward-Hoffmann rules**.
- The fact that Woodward and colleagues’ forward reaction is thermal but reverse reaction is photochemical is what yields the opposite starting material!
- **Thermal** (reaction): A reaction driven by high temperatures.
- **Photochemical** (reaction): A reaction driven by light.
- **Disrotatory** (electrocyclic reaction): An electrocyclic reaction in which the termini of the  $\pi$ -systems rotate in opposite directions.
- **Conrotatory** (electrocyclic reaction): An electrocyclic reaction in which the termini of the  $\pi$ -systems rotate in the same direction.
- **Woodward-Hoffmann rules:** Pericyclic reactions occur by the conservation of orbital symmetry from starting material to product.

| <b>Activation</b> | <b># e<sup>-</sup></b> | <b>Rotation</b> |
|-------------------|------------------------|-----------------|
| $\Delta$          | $4n$                   | con             |
| $\Delta$          | $4n + 2$               | dis             |
| $h\nu$            | $4n$                   | dis             |
| $h\nu$            | $4n + 2$               | con             |

Table 3.1: Woodward-Hoffmann rules.

- These rules are important because they allows us to predict the stereochemistry of our products.
- Nobel prize (1981) to Hoffmann and Fukui.
  - Fukui was jointly awarded this prize for his work on frontier molecular orbital theory, which we’ll talk about later in this lecture.
  - Woodward didn’t win because he had died. It was ok, though, because he had already won the Nobel once; this would have been his second.
  - Aside (chemis-tea): A spat over who invented the Woodward-Hoffmann rules.
    - E. J. Corey claimed credit for giving Woodward the idea for the Woodward-Hoffmann rules in 2004 — see Corey (2004).
    - Then Hoffmann rebuts Corey with a show-me-the-receipts type article — see Hoffmann (2004).
    - Woodward and Corey were both titans in their field at Harvard, both Nobel laureates, but also both big personalities.
  - Aside: Anyone who believes that science is somehow unbaised and empirical has never worked with a real scientist. Masha: “Scientists are some of the most human, emotional colleagues I’ve ever worked with... and I love them, don’t get me wrong.”
- Historical impact: One of the first successful unions of theory and experiment in chemistry.
  - Credited with leading organic chemists to finally accept MO theory.

- Let's now schematize the orbital machinations underlying the Woodward-Hoffmann rules.
- Correlation diagram:** A method of tracking orbital symmetry from starting materials to products in an electrocyclization.
  - Workflow.
    - Draw MOs.
    - Assign symmetry ( $S$  = symmetric,  $A$  = antisymmetric).
    - Populate with electrons.
    - Correlate orbitals with the same symmetry.
  - We assign symmetry differently depending on whether we're investigating a disrotatory or conrotatory pathway.
    - Disrotatory pathway: Ask yourself, "are the orbitals symmetric with respect to the  $\sigma$ -plane?"
    - Conrotatory pathway: Ask yourself, "are the orbitals symmetric with respect to the  $C_2$  axis perpendicular to the  $\sigma$ -bond that forms in the electrocyclization and lying in the plane of the pericyclic TS?"
  - For clarification on what exactly this all means, we'll look at a few examples. In particular, we'll investigate the favorability of the thermal and photochemical, disrotatory and conrotatory pathways through which the  $4\pi$  electrocyclization of butadiene could proceed.
- $\sigma$ -plane: The mirror plane lying perpendicular to the  $\sigma$ -bond that forms in an electrocyclization.
- Example: The possible thermally activated,  $4\pi$  electrocyclizations of butadiene.

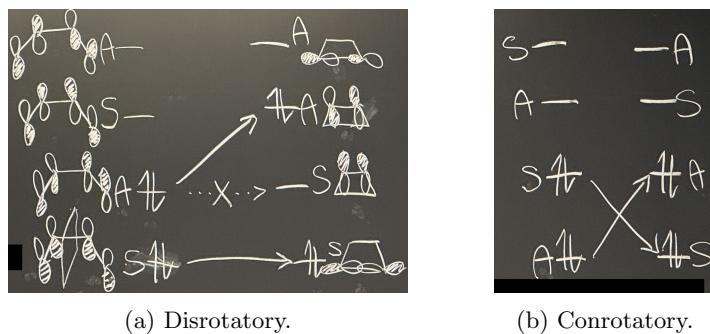


Figure 3.4: Correlation diagrams for butadiene's thermal  $4\pi$  electrocyclizations.

- We will first apply the correlation diagram workflow to the disrotatory case (Figure 3.4a).
  - Draw MOs for both the starting material and product.
    - Let's begin by drawing the MOs for the starting material.
      - We'll draw four  $p$ -orbitals and shade them in to create a conjugated  $\pi$ -system exactly as in Huckel theory.
      - Specifically, notice the no nodes  $\rightarrow$  1 node  $\rightarrow$  2 nodes  $\rightarrow$  3 nodes pattern.
    - Then we draw the product's MOs adjacent.
      - We know that the original four  $p$ -orbitals are transforming into a new  $\sigma$ -bond and a new  $\pi$ -bond, so we draw the bonding and antibonding phases of the new  $\sigma$ -bond as well as the bonding and antibonding phases of the new  $\pi$ -bond.
      - Essentially, we are drawing a  $\sigma$ ,  $\pi$ ,  $\pi^*$ , and  $\sigma^*$  MO.
      - Note that the  $\pi$  and  $\pi^*$  orbitals split less (energetically) than the  $\sigma$  and  $\sigma^*$  orbitals — just like in IChem — because they have less direct overlap; this is why we get the ordering  $\sigma \rightarrow \pi \rightarrow \pi^* \rightarrow \sigma^*$  as opposed to  $\pi \rightarrow \sigma \rightarrow \sigma^* \rightarrow \pi^*$  or something like that.

2. Assign symmetry to each of our drawn MOs.
    - Since we are looking at the *disrotatory* case, we will look at symmetry with respect to the  $\sigma$ -plane.
    - As a guide, we draw in the  $\sigma$ -plane in the bottom-left MO.
    - This particular MO is clearly symmetric with respect to the  $\sigma$ -plane, so we label it “S”.
    - We then perform this analysis for the remaining MOs, noting them as either symmetric or antisymmetric.
  3. Populate the starting MOs with electrons.
    - $4\pi$  electrocyclization, so 4 electrons to fill normally (i.e., per Aufbau, Pauli, and Hund).
  4. Correlate the filled starting MOs to the lowest energy product MOs with matching symmetry.
    - Maxim: An orbital cannot flip its symmetry during an electrocyclization.
    - As such, the lowest energy starting MO (being symmetric) has no problem becoming the lowest energy product MO (which is also symmetric).
    - However, the second-lowest energy starting MO (being antisymmetric) *cannot* become the second-lowest energy product MO because the latter is symmetric. (We say that this transition is formally **forbidden** because symmetry is not conserved.) As such, it must go up in energy to become the third-lowest energy product MO.
    - This population of a higher energy orbital means that the  $4\pi$  electrocyclization of butadiene is *disfavored* to occur through a thermal, disrotatory pathway.
- We now apply the correlation diagram workflow to the conrotatory case (Figure 3.4b).
1. The MOs will be the same as in Figure 3.4a, so we don’t need to redraw them.
  2. The symmetry must be evaluated with respect to that  $C_2$  axis this time, though, so we have to reassign S or A to each MO.
    - For the MOs as drawn in Figure 3.4a, the  $C_2$  axis we need goes into the plane of the page.
  3. We populate the starting MOs as before.
  4. When we correlate, this time we can fill the bottom two product MOs!
    - We didn’t populate electrons directly across, but we *did* populate the lowest energy orbitals again, so the conrotatory pathway is *favored*.
    - Both arrows involve a conservation of orbital symmetry, so (to reiterate) this reaction is **allowed** (thermally).
- David: Why do we only draw some of the molecular orbitals?



Figure 3.5: MOs relevant to butadiene’s  $4\pi$  electrocyclization.

- We only consider the orbitals involved in the reaction; considering the whole  $\sigma$ -network would get more complicated without changing our results.
- This is actually an example of why arrow-pushing is useful! Namely, because it shows that the  $p$ -MOs we consider in the starting material become the  $\sigma$ - and  $\pi$ -MOs we consider in the product.
- **Photochemical reaction:** A reaction driven by the absorption of a photon, leading to an excited state.
  - Later in this course, we’ll go more into detail on photochemical reactions, but this is the only level of detail we need right now.
  - In the reactions we’ll look at today, one electron is kicked up an energy level with no other changes to the structure.

- Example: The possible photochemically activated,  $4\pi$  electrocyclizations of butadiene.

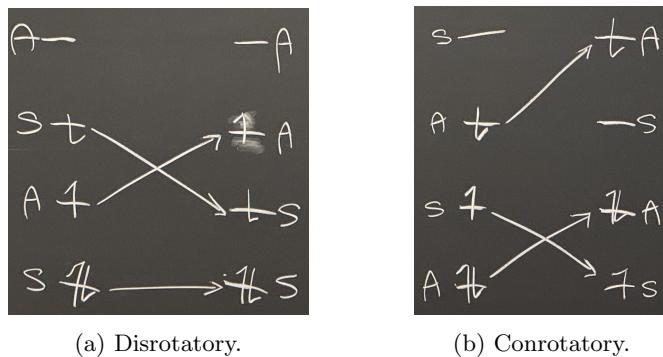


Figure 3.6: Correlation diagrams for butadiene's photochemical  $4\pi$  electrocyclizations.

- We use the same MOs and symmetries as in the corresponding subfigures of Figure 3.4.
- Differences only start to appear when we populate with electrons.
  - In particular, we excite one electron up a level (without altering its spin) in both sets of starting-material MOs.<sup>[4]</sup>
- Disrotatory case (Figure 3.6a).
  - Orbital symmetries are such that we end up with the *same* populations as in the starting material.
  - Therefore, this pathway is allowed/favored.
- Conrotatory case (Figure 3.6b).
  - Orbital symmetries are such that we end up with a *higher-energy* population than in the starting material.
  - Since electrons are in “much higher” energy levels, this pathway is forbidden/disfavored.
- Notice that the photochemical result that disrotatory is favored and conrotatory is disfavored is the opposite of thermal!
  - Thus, we just derived the Woodward-Hoffmann rules (Table 3.1) about what is favored and disfavored!
    - At least we derived the case for 4 electrons.
    - A more general mathematical proof can be done to rigorously verify Table 3.1, but the details are beyond the scope of this class.
  - If you ever forget the WH rules, just rederive them from first principles :)
- A note on the correlation arrows in Figures 3.4 and 3.6.
  - The uppermost correlation arrow in Figure 3.6b corresponds to an *allowed* but *disfavored* electronic transition.
  - The X'ed-out correlation arrow from A to S in Figure 3.4a corresponds to an explicitly *forbidden* electronic transition.
- The Woodward-Hoffmann rules are one way to look at pericyclic reactions, probably the most complex way.
  - We'll now look at two simpler ways.

<sup>4</sup>I.e., without intersystem crossing to a triplet state.

- Dewar-Zimmerman analysis: Aromatic TS theory.
  - Principle: Reactions that go through aromatic transition states are allowed.
  - We already like 6-membered TS's because they're geometrically stable; 6-membered *aromatic* TS's are even lower energy and more favored!
- Examples of aromatic and antiaromatic transition states.



(a) Aromatic TS in a [4 + 2] cycloaddition.



(b) Antiaromatic TS in a [2 + 2] cycloaddition.

Figure 3.7: Aromatic and antiaromatic transition states.

- Figure 3.7a shows that the transition state in a Diels-Alder reaction is aromatic.
  - This is why the forward Diels-Alder reaction is favored (under thermal conditions).
- Figure 3.7b shows that the transition state in a [2 + 2] cycloaddition is antiaromatic.
  - This is why the forward reaction to cyclobutane is disfavored (under thermal conditions).
- Let's get a little more formal now.
- Rules.
  1. Draw orbitals with any phasing, and decide the reaction topology.
    - By “any phasing,” we do mean that you can take *any* of the MOs you would draw and the Dewar-Zimmerman analysis will still work. So “go crazy,” if you want!
      - Masha does not draw out any examples to inductively prove this to us, but it would probably be a good exercise to do this on my own!!
    - By “reaction topology,” we mean how the two reactants approach each other. Is something coming from the top? The bottom?
  2. Connect orbitals through the reaction topology and count the number of phase inversions (PIs).
    - In other words, count how many lines connect lobes of opposite phases.
    - This will tell us whether we should evaluate the transition state for *Huckel* or *Möbius* aromaticity.
    - Essentially...
      - If there are an *even* number of PIs, having  $4n + 2$  electrons will lend *Huckel* aromaticity to the transition state and favor its formation;
      - If there are an *odd* number of PIs, having  $4n$  electrons will lend *Möbius* aromaticity to the transition state and favor its formation.
  3. Count the number of electrons.
    - As mentioned above, this number will tell us (depending on whether we're in Huckel-land or Möbius-land) whether the transition state is aromatic.

- To practice using these rules, let's reevaluate the examples in Figure 3.7.

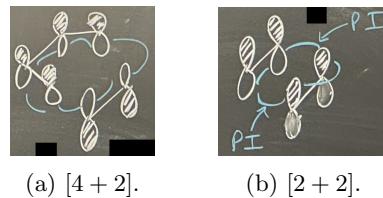
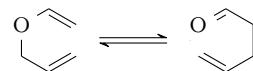


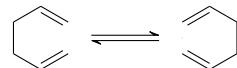
Figure 3.8: Dewar-Zimmerman connections for [4 + 2] and [2 + 2] cycloadditions.

- Example: [4 + 2] cycloaddition (Figure 3.8a).
    1. Draw MOs, and decide the reaction topology.
      - MOs: Let's arbitrarily choose to use the MOs with no nodes for both butadiene and ethene.
      - Topology: Let ethene approach butadiene from the bottom.
    2. Connect orbitals, and count PIs.
      - Connections: We connect all the bottom lobes of butadiene, the two top lobes of ethene, and (since ethene is approaching from the bottom, per the reaction topology) the bottom terminal lobes of butadiene to the top terminal lobes of ethene.
        - The connections are all drawn as blue lines in Figure 3.8.
      - PIs: All of the connected lobes are unshaded, so there are 0 PIs.
      - 0 is an even number, so we are in Huckel-land.
    3. Count the number of electrons.
      - There are  $6 = 4(1) + 2 \pi$ -electrons.
      - Therefore, our TS will be *stabilized* by *aromaticity* of the *Huckel* type.
  - Example: [2 + 2] cycloaddition (Figure 3.8b).
    1. Draw MOs, and decide the reaction topology.
      - MOs: We once again choose (arbitrarily) the MOs with no nodes for both ethenes.
      - Topology: The right ethene approaches the left ethene from the bottom.
    2. Connect orbitals, and count PIs.
      - Connections: We connect the bottom lobes of the left ethene to each other and to the top lobes of the right ethene (which are also connected to each other).
      - PIs: This time — because of the way we have drawn the right ethene — we have 2 PIs.
      - 2 is an even number, so we are in Huckel-land.
    3. Count the number of electrons.
      - There are  $4 = 4(1) \pi$ -electrons.
      - Therefore, our TS will be *destabilized* by *antiaromaticity* of the *Huckel* type.
  - The Dewar-Zimmerman analysis is useful for predicting the feasibility of sigmatropic rearrangements.
    - Recall from above that a *sigmatropic rearrangement* involves the migration of a  $\sigma$ -bond along with a corresponding reorganization of the  $\pi$ -electrons.
    - Specifically, in a sigmatropic rearrangement, the total number of  $\pi$ - and  $\sigma$ -bonds does not change.
  - Example sigmatropic rearrangements.
    1. **Claisen rearrangement**.
    2. **Cope rearrangement**.

- **Claisen rearrangement:** A [3, 3]-sigmatropic rearrangement of allyl vinyl ethers to form corresponding  $\gamma, \delta$ -unsaturated carbonyls.



- **Cope rearrangement:** A [3, 3]-sigmatropic rearrangement of 1,5-dienes to form other 1,5-dienes.



- We should be familiar with both the Claisen and Cope rearrangements; if we're not, Google them!!
- Example: Dewar-Zimmerman analysis of **suprafacial** and **antarafacial** [1,3]-sigmatropic H shifts.

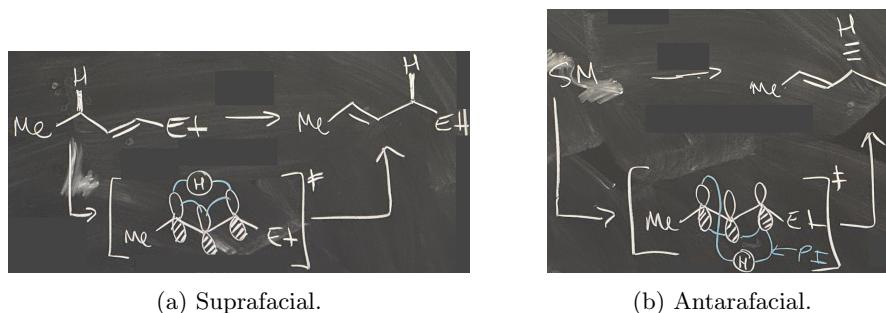


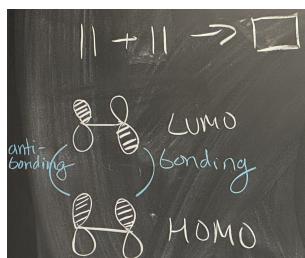
Figure 3.9: Dewar-Zimmerman analysis of [1, 3]-sigmatropic hydride shifts.

– We'll start with the suprafacial case (Figure 3.9a).

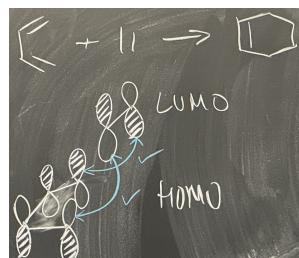
1. Draw MOs, and decide the reaction topology.
  - MOs: We choose the MO with no nodes for the  $\pi$  bonds, and the hydrogen atom's 1s orbital.
    - Notice that we draw a  $p$ -orbital for all three carbons involved in the bond breaking/making process in the transition state, not just the two carbons involved in the initial or final bond!
  - Topology: Draw the H atom in the process of migrating.
2. Connect orbitals, and count PIs.
  - Connections: Connect all the top lobes together and to hydrogen.
  - PIs: 0.
  - Even PIs, hence Huckel.
3. Count the number of electrons.
  - We have two electrons in the C=C  $\pi$ -bond, and two electrons in the C–H  $\sigma$ -bond.
    - We also get a clue that there are 4 electrons present because we drew 4 orbitals!
  - Thus, there are  $4 = 4(1)$  electrons present.
  - Therefore, our TS will be destabilized by Huckel antiaromaticity.
  - It follows that the suprafacial pathway is (thermally) forbidden.
    - This is an interesting result because at first glance, it "looks" like a nice TS (with the H just bouncing over), but nope! It's not allowed.

- We now move onto the antarafacial case (Figure 3.9b).
  1. Draw MOs, and decide the reaction topology.
    - MOs: Same as in Figure 3.9a.
    - Topology: The H atom is switching faces, so it will have to engage with the top lobe on one side and the bottom lobe on the other side.
  2. Connect orbitals, and count PIs.
    - Connections: We connect the three *p*-orbitals as expected, but note the explicit connection of the top-left *p*-lobe and the bottom-right *p*-lobe to the hydrogen, in accordance with the reaction topology.
      - Note that it doesn't really matter which lobes of the *p*-orbitals we connect because we get the same result either way.
    - PIs: 1.
    - This is our first time having an *odd* number of PIs, so we are now in Möbius-land!
  3. Count the number of electrons.
    - As above, there are  $4 = 4(1)$  electrons.
    - However, because we are in Möbius-land, this nevertheless means that our TS will be *stabilized* by *aromaticity* of the *Möbius* type.
    - Thus, “ugly” antarafacial transition states are nevertheless totally allowed!
- Despite the fact that antarafacial [1, 3]-sigmatropic hydride shifts are favored over their suprafacial counterparts, our intuition that the antarafacial transition state would be sterically strained is correct.
  - Indeed, there are examples of [1, 3]-hydride shifts occurring with stereoinversion, but they are rare.
  - Nevertheless, this is a fun and nonintuitive finding!
  - If you are interested, you can look into work on antarafacial [1, 3]-methyl shifts, which are also favored over their suprafacial counterparts!
- Aside: One place where we do see stereoinversions.
  - The keto-enol tautomerization could be thought of as a [1, 3]-sigmatropic hydride shift!
  - Indeed, if it occurs intramolecularly, it would occur with stereoinversion.
  - However, the rate of this intramolecular rearrangement is naturally very slow due to strain, which is why we need a solvent, acid, or base catalyst to do the proton transfer intermolecularly with any appreciable rate.
  - Essentially, the reaction can't really happen intramoleucularly because it'd be forbidden electronically with *cis* hydrogens or very disfavored sterically with *trans* hydrogens.
- **Suprafacial** (sigmatropic rearrangement): A sigmatropic rearrangement in which the bond-breaking and bond-making processes occur on the *same* face of the  $\pi$ -system.
- **Antarafacial** (sigmatropic rearrangement): A sigmatropic rearrangement in which the bond-breaking and bond-making processes occur on *opposite* faces of the  $\pi$ -system.
- Notice how...
  - In Figure 3.9a, the hydride is on the same side of the molecule in both starting material and product, i.e., coming out of the plane of the page;
    - That's why we call this suprafacial!
  - In Figure 3.9b, the hydride is on opposite sides of the molecule in the starting material vs. the product, i.e., coming out of the plane of the page vs. going into the plane of the page.
    - That's why we call this antarafacial!

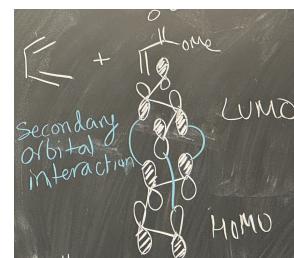
- Frontier molecular orbital theory (FMO).
    - By Fukui, as mentioned above.
    - This is a simplification of some other models in which you only consider the HOMO/LUMO interactions (instead of all MOs).
    - Principle: If the HOMO of the electron-donating species and LUMO of the electron-accepting species mix favorably, then the reaction is allowed.
  - Example: FMO analysis of cycloadditions.



(a) [2 + 2].



(b) [4 + 2].



(c) *endo*-[4 + 2].

Figure 3.10: Frontier molecular orbital analysis of cycloadditions.

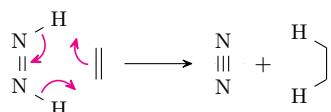
- A [2 + 2] cycloaddition (Figure 3.10a).
    - As in a Dewar-Zimmerman analysis we begin by drawing the HOMO and LUMO of the reactants and deciding the reaction topology.
      - MOs: Recall that for ethene, the HOMO has 0 PIs and the LUMO has 1 PI.
      - Topology: We have decided to have the ‘electron-donating’ ethene attack from the bottom.
    - Then (also as in a Dewar-Zimmerman analysis) we connect lobes and check for bonding and antibonding interactions.
      - Here, we have 1 bonding and 1 antibonding interaction.
    - The presence of an antibonding interaction means that this reaction is forbidden/disfavored.
    - Bonus content: Ketenes engage in [2 + 2] cycloadditions at ambient temperatures!
      - Masha encourages us to look more into this!!
  - A [4 + 2] cycloaddition (Figure 3.10b).
    - Performing an analogous analysis to the above, we observe 2 bonding interactions. This means that this reaction is allowed.
  - A [4 + 2] cycloaddition with a more elaborate dienophile (Figure 3.10c).
    - As in both previous cases, we begin by drawing the HOMO of the ‘electron-donating’ species and the LUMO of the ‘electron-accepting’ species.
      - MOs: Notice that it does not matter that the dienophile’s  $\pi$ -system contains a heteroatom.
      - Topology: This time, we have the dienophile attack from the top.
    - Connecting orbitals.
      - First, observe that we get the same favorable bonding interactions as in Figure 3.10b.
      - In addition, we get a new, secondary orbital interaction if we draw the *endo* transition state.
      - This additional, stabilizing interaction rationalizes why the *endo* transition state is favored in a Diels-Alder reaction!
    - Therefore, this reaction is allowed, and the *endo* product is preferred due to secondary orbital interactions.

- Miscellaneous pericyclics: Group transfer and chelotropic reactions.
- Recall from above that a *group transfer* reaction transfers atoms from one molecule to another, but in a concerted pericyclic transition state.
- Example group transfer reactions.

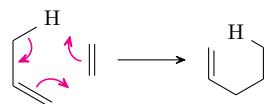
- Diimide reduction.**

- The ene reaction.**

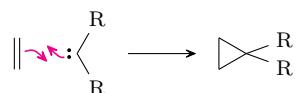
- Diimide reduction:** A group transfer reaction that converts an unsaturated organic compound to a reduced alkane using diimide ( $N_2H_2$ ).



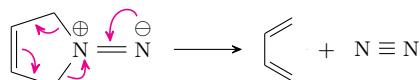
- Remember to draw arrows to the place where you're making the bond, not to an atom unless the electrons are going on that atom!!
- Ene reaction:** A group transfer reaction between an **ene** and an **enophile** that forms a new  $\sigma$ -bond with migration of the ene double bond and a 1,5-hydrogen shift. *Also known as Alder-ene reaction.*



- Ene:** An alkene with an allylic hydrogen.
- Enophile:** A compound containing a multiple bond.
- Moving on, recall from above that a *chelotropic* reaction is a cycloaddition in which two bonds are made to one atom.
- Example chelotropic reactions.
  - Carbene addition.**
  - Certain **cycloreversions**.
- Carbene addition:** The addition of a singlet carbene to an alkene to make a cyclopropane.



- Observe how the two arrows form two  $\sigma$ -bonds to the carbene.
- Cycloreversion:** The reverse of a cycloaddition reaction.



- Since two of the left nitrogen's bonds are being *broken*, this is technically a *retro-chelotropic* reaction.

- Lecture summary: Three models to study pericyclic reactions.
  1. The Woodward-Hoffmann rules.
    - These are all about the conservation of orbital symmetry.
  2. The Dewar-Zimmerman analysis, also known as aromatic TS theory.
    - This is the “I can’t believe it works!” one, where you can draw any phasing and the model still gives you the right answer.
  3. FMO theory.
    - This is where we only look at HOMO/LUMO interactions.
- Matthew: When would you use one model over the others?
  - All three models should always give the same result (otherwise, there’s a problem with the model), but sometimes you care more about one aspect of a reaction or another.
  - For example, if you want to figure out whether you get the conrotatory or disrotatory product, it is easier to use the Woodward-Hoffmann rules.
    - This is because they’re designed specifically for such questions.
  - If you need a quick-and-dirty “is this reaction going to happen,” use FMO.
  - If you want to determine whether a reaction will be antarafacial or suprafacial, use Dewar-Zimmerman.

### 3.3 Chapter 14: Advanced Concepts in Electronic Structure Theory

*From Anslyn and Dougherty (2006).*

#### Section 14.2: Calculational Methods — Solving the Schrödinger Equation for Complex Systems

- 9/23:
- 20 pages on computational methods, but a bit outdated.

### 3.4 Chapter 15: Thermal Pericyclic Reactions

*From Anslyn and Dougherty (2006).*

- This whole chapter is a gold mine. Return to!!

#### Section 15.1: Background

- 

#### Section 15.2: A Detailed Analysis of Two Simple Cycloadditions

- 

#### Section 15.3: Cycloadditions

- 

#### Section 15.4: Electrocyclic Reactions

-

## Section 15.5: Sigmatropic Rearrangements

- 

## 3.5 Chapter 16: Photochemistry

*From Anslyn and Dougherty (2006).*

- Also a treasure trove, though much of this is beyond the scope of the course. However, it is largely within the scope of 5.47, so I should definitely return!!

### Section 16.3: Photochemical Reactions

- Recommended readings for this course: Sections 16.3.4-16.3.5.
- The tail-end of 16.3.5 talks a bit about norbornadiene (PSet 1).

# Week 4

## Ions

### 4.1 Cations

9/24:

- Lecture 5 recap.
  - Pericyclic reactions: Concerted reactions with a TS having a cyclic array of atoms and orbitals.
  - Three models.
    1. Woodward-Hoffmann rules: Conservation of orbital symmetry.
    2. Dewar-Zimmerman analysis: Aromatic TS theory.
    3. Frontier MO theory: HOMO-LUMO interactions.
  - “No mechanism... half in jest, half in desperation... to the thermoreorganization reactions.”
    - Essentially, pericyclic reactions really led to a new blossoming of organic chemistry, and a series of successful mergers between theory and experiment.
- Announcements: PSet 1 due tomorrow; if late, we'll lose a lot of points.
- Today: Cations (mostly carbocations).
  - This is the first in a series of lectures on functional groups: Cations, anions, radicals, and carbenes.
- Lecture outline.
  - Overview of cation structure and reactivity.
  - Measuring a cation's (thermodynamic and kinetic) stability.
  - Stabilizing cations to promote reactivity.
  - Cation reactions.
  - Nonclassical carbocations.
- There are three phases in a cation's lifetime: Synthesis, stability, and reactivity.

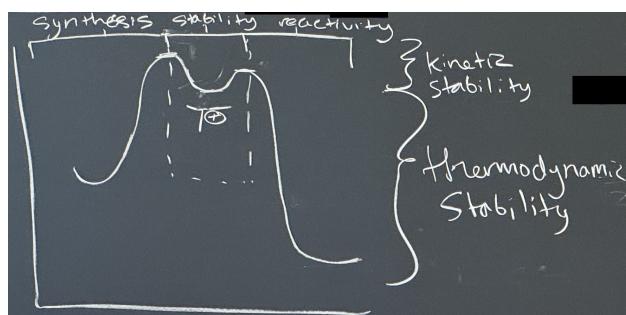


Figure 4.1: Phases in the life of a cation.

- All three phases correspond to specific regions along the reaction coordinate in the energy diagram for a cation-intermediate reaction.
- Stability, in particular, we'll talk about from a kinetic *and* a thermodynamic perspective.
  - Kinetic stability deals with the energy barrier to *form* and to *react* the cation.
  - Thermodynamic stability deals with the energy difference between the cation and the adjacent local ground state structures.
- Cations can have quite “sensitive” energy surfaces, i.e., factors that can stabilize and destabilize cations can have dramatic effects on the synthesis, stability, and reactivity of cations.
- Features that stabilize cations tend to lead to reactions.
  - If you're in the lab, consider stabilizing the cation in order to induce the desired reactivity!
- Cation structure.

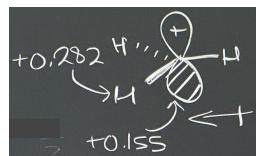


Figure 4.2: Cation structure.

- Figure 4.2 depicts a methyl cation ( $\text{CH}_3^+$ ).
- In general, cations are  $sp^2$ -hybridized, trigonal planar species.
  - Recall that Figure 2.5 explains why cations are trigonal planar instead of pyramidal.
- The cationic charge is delocalized across the entire molecule, not localized on the carbon.
  - Indeed, there is a  $\delta^+$  on the H's, too.
  - In fact, the dipole qualitatively points *toward* the carbon.
  - Quantitatively, the **Mulliken partial charges** are +0.155 on C and +0.282 on each H. Together, these partial charges sum to the total charge of +1:

$$1 \times 0.155 + 3 \times 0.282 \approx 1$$

- Experimental evidence for cation formation.

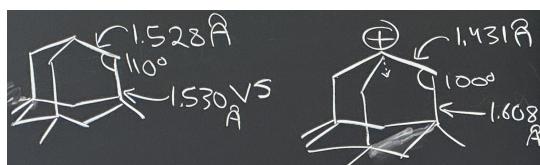


Figure 4.3: Evidence that carbocations exist.

- Experimental evidence primarily comes from some cool adamantane structures.
- For example, consider trimethyl adamantane and its corresponding cation. The cation has structural characteristics indicative of the “flattening” transformation we would expect. Specifically...
  - The adjacent bond angles flatten from 110° to 100°;
  - The bonds immediately surrounding the cation shrink from 1.528 Å to 1.431 Å as the molecular geometry compresses the flattening cation;
  - The bonds  $\alpha, \beta$  to the cation elongate from 1.530 Å to 1.608 Å as electron density is removed from them through hyperconjugation and the no-bond resonance form.
- Reference: Laube (1986).

- Moving on, to measure the thermodynamic stability of a cation, we use the **hydride ion affinity**.
- Hydride ion affinity:** The extent to which cations want to bind a hydride in solution. *Also known as HIA. Given by*



- Always measured in the gas phase.
- Only tells you the *relative* stability.<sup>[1]</sup>
- Example HIAs.

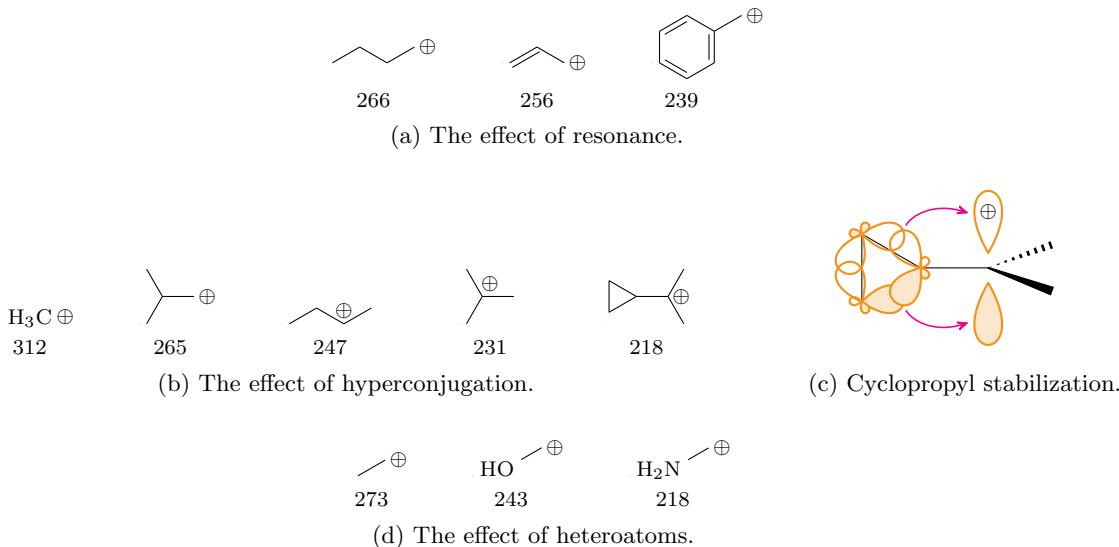


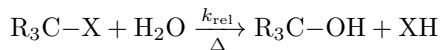
Figure 4.4: Hydride ion affinity examples.

- Alkyl, allylic, and benzylic HIAs (Figure 4.4a).
  - Respectively: 266 kcal/mol, 256 kcal/mol, and 239 kcal/mol.
  - Attributable to resonance delocalization and conjugation.
- Methyl, isobutyl, *sec*-butyl, *tert*-butyl, and dimethylcyclopropyl HIAs (Figure 4.4b).
  - Respectively: 312 kcal/mol, 265 kcal/mol, 247 kcal/mol, 231 kcal/mol, and 218 kcal/mol.
  - Attributable to hyperconjugation.
  - Deeper dive: Hyperconjugation from cyclopropyl rings (Figure 4.4c)
    - This is a follow up to our brief discussion on the same topic in Lecture 3.
    - When this molecule forms, the carbocation's empty *p*-orbital will align with the  $\sigma$ -plane of the cyclopropyl group.
    - With this alignment, *both* adjacent C–C banana bonds can donate into the carbocation through hyperconjugation.
    - The hyperconjugative interaction is so extreme that the barrier to rotation along the bond between the cation and the cyclopropyl group is 13.7 kcal/mol!
    - We can also picture this interaction through no-bond resonance forms that delocalize the positive charge to the back two carbons in the cyclopropyl group.
    - You can look up the crystal structure of this molecule to see the interaction more.
- Ethyl, hydroxymethyl, and aminomethyl HIAs (Figure 4.4d).
  - Respectively: 273 kcal/mol, 243 kcal/mol, and 218 kcal/mol.
  - Attributable to heteroatom stabilization (aka resonance).

<sup>1</sup>Relative to what??

- The stability of the carbocation (as discussed above in terms of HIAs) determines how high the local minimum is in the energy diagram in Figure 4.1.
- We now move onto the kinetic stability/reactivity of cations.
- Two ways of measuring this.
  - Rates of **solvolysis**.
    - Used all the time.
  - Mayr electrophilicity**.
    - More niche, but still good to know.

- Solvolysis:** A type of nucleophilic substitution ( $S_N1$  or  $S_N2$ ) wherein the nucleophile is a solvent molecule. *Given by*



- Rates of solvolysis are reported as a relative rate constant  $k_{\text{rel}}$ .
- Comparing HIAs to rates of solvolysis.

|                  | Bn-Br | All-Br | <i>i</i> Pr-Br |
|------------------|-------|--------|----------------|
| HIA ( $R^+$ )    | 239   | 256    | 249            |
| $k_{\text{rel}}$ | 100   | 52     | 0.7            |

Table 4.1: HIAs and the rate of solvolysis are not correlated.

- To be clear, we are listing the HIA of the benzyl, allyl, and isopropyl cations.
- Benzyl bromide affords a cation that is both the most stable and the most reactive in the set.<sup>[2]</sup>
- Note that in general, solution-phase measures of stability like solvolysis and gas-phase measures of stability like the HIA *don't* correlate. This means that we do have to measure them independently.
- Mayr electrophilicity:** The rate of reaction for various electrophilic and nucleophile pairs. *Given by*



- By Herbert Mayr from 5.47!
  - Mayr defined three parameters ( $S$ ,  $N$ , and  $E$ ) via the equation.
- $$\log k = s(N + E)$$
- $s$  is a nucleophile-specific slope parameter.
  - $N$  is a nucleophile parameter.
  - $E$  is an electrophile parameter.
  - Note that “ $\text{Nuc}^-$ ” indicates a nucleophile, just like the more commonly used  $\text{Nu}^-$ .
  - Mayr has done hundreds of these reactions, measured their rates, had reference nucleophile, etc.
    - His group is still expanding the chart!
    - There's a giant PDF on Mayr's [website](#) that we can download if we want.
  - Reference: Mayr and Patz (1994).

<sup>2</sup>Clarify??

- Example Mayr electrophilicities.

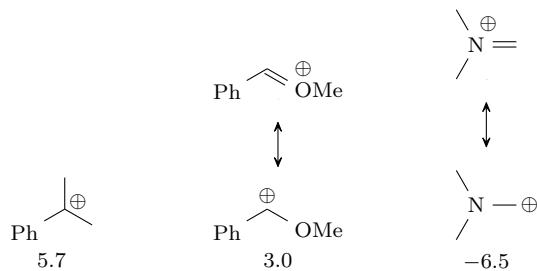


Figure 4.5: Mayr electrophilicity examples.

- To be clear, Figure 4.5 lists the  $E$  value for each species.
- Remember that Mayr electrophilicity is reported on a logarithmic scale, so the difference in  $E$  between the left two species (approximately 3) corresponds to a difference in reactivity of three *orders of magnitude*.
  - Similarly, the difference in reactivity between the right two species is *nine orders of magnitude*!
- Some of these trends should make sense.
  - For example, it stands to reason that the cation with heteroatom stabilization is the least electrophilic.
  - Observe that our most thermodynamically stable carbocation (the 3° one with extensive resonance into the phenyl ring) is also our most Mayr electrophilic one!
    - This is yet another example of thermodynamics being decoupled from the kinetics of reactivity.
- This concludes our discussion of *measuring* kinetic and thermodynamic stability. Let's now talk about *enhancing* carbocation stability.
- Four ways of doing this.
  1. **Hyperconjugation.**
  2. Heteroatom stabilization.
  3. The  $\beta$ -silicon effect.
  4. The neighboring group effect.
- **Hyperconjugation:** The delocalization of electrons through  $\sigma$ -bonds.



(a)  $\sigma_{\text{CH}} \rightarrow p_{\text{C}}$  donation.

(b) No-bond resonance.

Figure 4.6: Stabilizing carbocations: Hyperconjugation.

- Hyperconjugation explains why substituted cations are more stable.
- Recall from 5.13<sup>[3]</sup> that the ethyl cation is stabilized by  $\sigma_{\text{CH}} \rightarrow p_{\text{C}}$  donation.
- Equivalently, we can say that the ethyl cation is stabilized by **no-bond resonance**.
  - What this really tells us is that the C–C bond is shorter than we'd normally expect, and the C–H bond is longer than we'd normally expect.

<sup>3</sup>Figure 4.6a is just Figure 2.3a from Labalme (2024a).

- Example HIA differences caused by hyperconjugation (Figure 4.4b).
  - Increasing from no adjacent C–C bonds to three adjacent C–C bonds decreases the HIA from 312 kcal/mol to 231 kcal/mol.
  - Essentially, as we add more R groups, the cation's empty  $p$ -orbital gets stabilized by additional adjacent  $\sigma$ -orbitals.
- Matthew: Does hyperconjugation induce a barrier to rotation?
  - There's always some barrier.
    - In a normal alkyl molecule, it's approx 3 kcal/mol.
  - In a hyperconjugated cation, we will see bigger differences.
    - In fact, there's a fascinating example somewhere in the literature of the stereochemistry of a product being determined by geometric constraints caused by hyperconjugation!
  - So all this is to say, yes.
- Heteroatom stabilization.

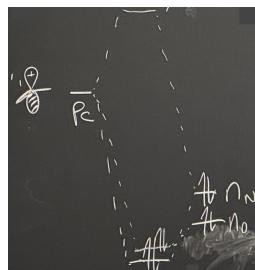


Figure 4.7: Stabilizing carbocations: Adjacent heteroatoms.

- The high-energy empty  $p$ -orbital on carbon and low-energy heteroatom lone pair interact to form new bonding and antibonding MOs.
  - The bonding MO will be lower energy than the lone pair AO, so the electrons in that lone pair will be stabilized.
- Nitrogen vs. oxygen stabilization: Rationalizing why nitrogen is more stabilizing in Figure 4.4d.
  - The  $n_O$  AO is lower in energy than the  $n_N$  AO.
  - This means that there is worse energy overlap between the  $n_O$  AO and the  $p_C$  AO than between the  $n_N$  AO and the  $p_C$  AO.
  - The worse energy overlap with oxygen leads to a resultant decrease in MO splitting, and hence less stabilization for the oxygen lone pair than the nitrogen lone pair receives.
- **$\beta$ -silicon effect:** The stabilization of positive charge at the position  $\beta$  to a silicon atom.
  - Caused by hyperconjugation.
  - Specifically, silicon is a better  $\sigma$ -donor, by which we mean that C–Si bonds are better at sharing their electron density via hyperconjugation than C–C or C–H bonds.<sup>[4]</sup>
  - Silicon is better because...
    - Silicon is less electronegative than other common  $\sigma$ -donors;
      - Indeed,  $EN_C = 2.55$  and  $EN_{\text{C}} = 2.20$ , but  $EN_{\text{Si}} = 1.90$ .
      - Thus, C–Si bonds hold their electrons less tightly and hence are happier to share.
      - C–Si bonds holding their electrons less tightly also implies the following.

<sup>4</sup>Note that we do *not* mean that silicon is a better  $\sigma$ -donor ligand, like in inorganic chemistry.

- C–Si bonds are longer;
  - 1.86 Å vs. the 1.54 Å typical of a C–C bond.
  - This allows for greater overlap with the typically lengthy *p*-orbitals.
- C–Si bonds are more ionic;
  - Polarization toward carbon (more ionicness) means that there's more electron density on the carbon (i.e., near the carbocation).
- The  $\sigma_{\text{CSi}}$  orbital is higher in energy than  $\sigma_{\text{CC}}$  orbital.
  - Thus, like in Figure 4.7, we get closer to the  $p_{\text{C}}$  energy level and have more effective overlap.
- Example HIA differences caused by the  $\beta$ -silicon effect.

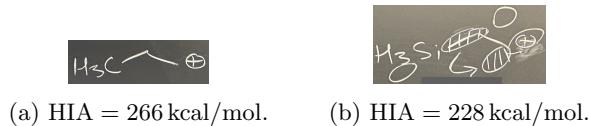


Figure 4.8: Hydride ion affinities subject to the  $\beta$ -silicon effect.

- Changing an alkyl cation to the direct silicon analogue alters the HIA by nearly 40 kcal/mol.
- Examples of how the  $\beta$ -silicon effect alters reactivity.

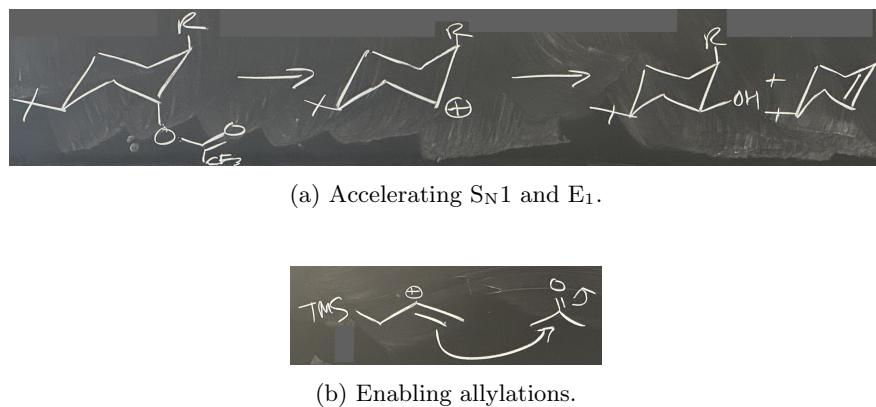


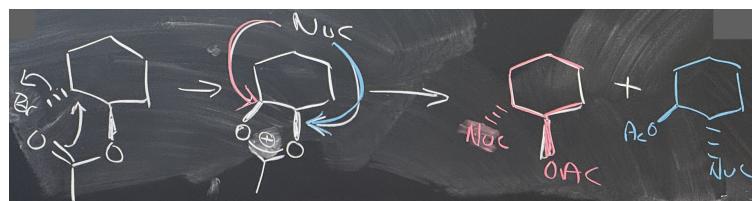
Figure 4.9: The  $\beta$ -silicon effect enables cationic reactivity.

- It can accelerate the departure of a leaving group by orders of magnitude (Figure 4.9a).
  - Suppose we have a trifluoroacetate leaving group on a cyclohexane ring in aqueous solution.
  - Locking  $\text{F}_3\text{CCO}_2^-$  in the axial position with an equatorial *tert*-butyl group aids departure.
    - Specifically, if  $\text{R} = \text{SiMe}_3$ , the anomeric effect will significantly weaken the C–O bond.
  - Once  $\text{F}_3\text{CCO}_2^-$  leaves, the reaction completes through hydration ( $\text{S}_{\text{N}}1$ ) or elimination ( $\text{E}_1$ ).
  - If  $k_{\text{rel}} = 1$  when  $\text{R} = \text{H}$ , then  $k_{\text{rel}} = 2.4 \times 10^{12}$  when  $\text{R} = \text{SiMe}_3$ .
    - There's a reason this effect has a name: It's huge!
- It enables allylations to happen at all (Figure 4.9b).
  - We do allylations with allyl silane because it's the only way this will work.
  - The allyl group attacks the carbonyl as a nucleophile, forming a secondary carbocation that's stabilized by the  $\beta$ -silicon effect at the indicated position.
  - Note that this reaction is *not* an already-formed carbocation somehow engaging in a nucleophilic attack, despite how it's drawn. Here's a helpful [reference](#) on this type of reactivity.

- Motivating the neighboring group effect.



(a) The possible products of a reaction.



(b) The mechanism of the reaction.

Figure 4.10: The neighboring group effect alters cationic reactivity.

- The reaction in Figure 4.10a is a nucleophilic substitution with an enantiopure starting material, and it has four possible product stereoisomers.
- Through which mechanisms could this reaction proceed?
  - If S<sub>N</sub>2: We'll see 100% *syn* and 0% *anti* product because S<sub>N</sub>2 is stereospecific.
    - The *syn* product will be enantiopure due to the stereoinverting nature of the attack.
  - If S<sub>N</sub>1: We'll see 50% *syn* and 50% *anti*, maybe favoring *anti* a bit due to sterics.
    - Both diastereomers will be enantipure (we're not engaging the acetate's chiral carbon).
  - Observed: We get 0% *syn* and 100% *anti*, and it's a racemic mixture of the *anti* diastereomer.
- What's happening here?!
  - The acyl group is not as innocent as it seems.
  - Per Figure 4.10b, the actual mechanism begins with intramolecular displacement of the bromine to form a resonance-stabilized carbocation. This is followed by a backside attack on either carbon, hence selecting the *anti* product and inducing the racemization.
  - Conclusion: The neighboring group effect makes this reaction *trans*-selective and racemizing.
- Neighboring group effect:** The interaction of a reaction center with either an intramolecular lone pair or an intramolecular pair of  $\pi$ -electrons. *Also known as anchimeric assistance.*
  - Note that the intramolecular pair of  $\pi$ -electrons cannot be conjugated with the reaction center; that's just resonance stabilization of the carbocation then.
- Homoconjugation:** A neighboring group effect in which the neighboring group is a  $\pi$ -system.<sup>[5]</sup>
- Example of homoconjugation.



Figure 4.11: Homoconjugation.

<sup>5</sup>This definition is consistent with the definition of homoconjugation as “an overlap of two  $\pi$ -systems separated by a non-conjugating group” because the carbocation counts as a  $\pi$ -system and the carbocation in Figure 4.11 is separated from the  $\pi$ -bond by one methylene group on each side.

- Essentially, the displacement of the tosyl group in Figure 4.11 is much more favorable in the molecule shown than in the saturated analog because a double bond is present nearby (in the unsaturated molecule), and its  $\pi$ -orbitals can donate into the carbocation.
  - Something like 5 orders of magnitude faster.<sup>[6]</sup>
  - We're now done with carbocation stability, and we'll begin discussing their synthesis and reactivity.
  - Acidity.

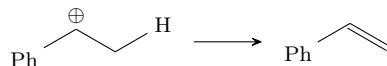
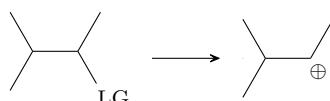
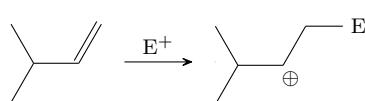


Figure 4.12: Carbocations acidify  $\beta$ -protons.

- Carbocations induce a dramatic acidification of  $\beta$  C–H bonds.
  - Indeed, the  $pK_a$  of the proton drawn in Figure 4.12 is  $-14!$
  - Additionally, this reaction is just the second step in an  $E_1$  mechanism: Adjacent deprotonation is just elimination!
    - The reaction is purely downhill thermodynamically, and adjacent deprotonation is actually a great perspective to take on  $E_1$ .
  - Synthesis of carbocations: Two main ways.



(a) Ionization.



(b)  $\pi$ -activation.

Figure 4.13: Synthesis of carbocations.

### 1. Ionization.

- This is just the departure of a leaving group.

## 2. Activation of a $\pi$ -system.

- Can be done by an electrophile, such as a proton, metal, etc.
  - Gives the Markovnikov adduct.

### • Reactions of cations.



(a) [1, 2]-sigmatropic shift.



(b) General form of a rearrangement.

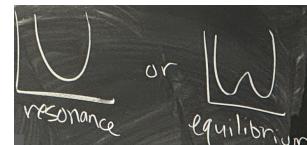
Figure 4.14: Reactions of cations.

<sup>6</sup>Actually 11 orders of magnitude per Wikipedia.

- Cations most typically appear in elimination ( $E_1$ ) and capture/substitution ( $S_N1$ ) mechanisms.
  - Once formed, cations can also do rearrangements, shifts, cyclizations, etc.
  - An important subcategory of cationic shifts is [1, 2]-sigmatropic shifts.
    - These are very common.
    - They are also very fast and very easy to do.
      - The rate of a [1, 2]-sigmatropic hydride shift is  $k_{1,2} = 3 \times 10^7 \text{ s}^{-1}$ , even at  $-139^\circ\text{C}$ .
      - The activation energy  $\Delta G_{1,2}^\ddagger \approx 3 \text{ kcal/mol}$ , which is on the same order of magnitude as bond rotation.
    - If you want these to happen, that's great!
      - If not, you're going to need to think about explicit ways to prevent it by design because [1, 2]-shifts will happen whether or not you want them to — you can't stop it.
    - Migratory aptitude:  $s > sp > sp^2 > sp^3$ .
      - The probability that a substituent will shift depends on the extent to which there is *s*-character in the bonding orbital of the *mobile* group because more *s*-character leads to better orbital overlap in the transition state (Figure 4.14a).
    - Essentially, the mechanism works by taking hyperconjugation “to the extreme” to move the bond (Figure 4.14a).
    - Two final noteworthy things about shifts.
      - We have a 2-electron Huckel aromatic transition state, so it will be allowed/favored by the Dewar-Zimmerman analysis.
      - We retain the stereochemistry of the migrating group (it's a suprafacial shift).
  - There are many named rearrangements.
    - Examples include the **Wagner-Meerwein rearrangement**, **pinacol rearrangement**, and **semipinacol rearrangement**.
      - We are not a named-reactions class, so we will not discuss these much, but you can look them up if you want.
    - These are all variants on a theme, though.
      - They all follow the general form in Figure 4.14b but with different R and LG groups.
      - The naming generally depends on the *identity* of the R and LG, based on whichever chemist discovered and popularized the class.
  - Nonclassical carbocations.



(a) 3c-2e bonds.



### (b) Energy diagrams

Figure 4.15: Nonclassical cations.

- Consider two cations: The  $3^\circ$  *tert*-butyl cation and a  $2^\circ$  cation on norbornane.
    - Interestingly, HIA = 231 kcal/mol for *both* of these cations!
    - How can they both be equally stable?
  - This question led to the discovery of nonclassical 3c-2e bonds (Figure 4.15a).
    - Essentially, we can draw two no-bond resonance forms for this cation. We move one of the  $\sigma$ -bonds in each of these (which we're not usually supposed to do).
    - Thus, we can draw the real structure with two half bonds.

- Aside (chemis-tea): The debate as to whether the true structure of nonclassical cations was barrierless resonance or an equilibrium between two cations raged in the literature for 70 years (Figure 4.15b).
  - On team resonance: Olah (Nobel prize for this cation work), Wintsein, Schleyer, Saunder.
  - On team equilibrium: H. C. Brown (Nobel prize for unrelated work).
  - Brown just thought this was due to poor techniques.
  - They would go to conferences, sit in the front row, yell at each other; publish snarky papers at each other.
  - Debate era: 1940s-2010s.
  - The debate ended at Science, 2013, 62 with an X-ray structure of the nonclassical cation (which really supported the resonance team). Unfortunately, H.C. Brown died in 2004. Anybody who knew Brown said he wouldn't have accepted this either.
  - “One would have thought that the application of careful experiment and intelligent thought would lead to a rapid solution to the [nonclassical carbocation] problem. This has not been the case” - Brown’s book.
  - Until they could prove the structure of one or both, we couldn’t know. This really drove the development of spectroscopy, NMR, low-temperature analysis of exotic species, etc. Essentially, people work hard when their ego is at stake.
- Takeaway from our discussion of nonclassical cations: Cations exist on a spectrum.

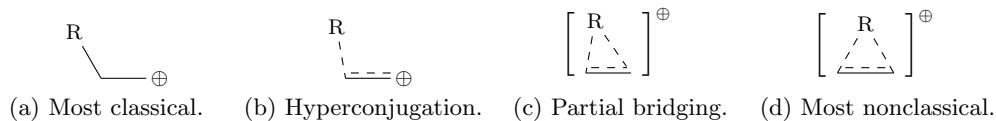


Figure 4.16: A spectrum of cations.

- Most classical: Discrete, trivalent, trigonal carbocations.
  - These rarely exist.
- Next step: Hyperconjugation and resonance.
  - This accounts for most carbocations.
- Next step: Some kind of bridging but asymmetric carbocation.
  - There are some examples.
- Most nonclassical: Bridging, symmetric carbocations (3c-2e).
  - These have to be special cases, such as the norbornane one.

## 4.2 Anions

9/26:

- Lecture 6 recap: Modes of cation stabilization.
  - Heteroatoms: Through lone-pair resonance.
  - Anchimeric assistance: With neighboring lone pairs (Figure 4.10b).
  - Homoconjugation: With nearby  $\pi$ -electrons (Figure 4.11).
  - Hyperconjugation: With adjacent  $\sigma$ -bonding orbitals. (Figure 4.6a).
- Announcements.
  - Don’t cheat on the PSet.
    - You can probably find papers or the key online, but be responsible academics instead.
    - PSets give you a chance to engage with the material; you will not learn if you cheat.
  - Ask for help if you can’t make the deadline.

- Today: Anions.
  - Lecture outline.
    - Acidity.
      - Gas phase.
      - Solution phase.
    - $pK_a$ 's.
      - Common ones.
      - Solvent effects.
      - Misc. influencing factors.
    - Anion structure and inversion.
    - Synthesis of carbanions.
    - Reactions of carbanions.
    - Kinetic vs. thermodynamic acidity.
  - Thermodynamic stability of an anion.
  - **Acidity** (gas phase): The extent to which anions want to bind a proton in the gas phase. *Given by*
- $$\text{R}-\text{H} \rightleftharpoons \text{R}^- + \text{H}^+ \quad \Delta H^\circ = \text{acidity}$$
- Similarities between this definition and that of the HIA!
    - Acidity is also always measured in the gas phase.
    - It is also more useful as a measure of relative stability.
  - Trends in (gas-phase) acidity are not always the same in solution.
  - Example gas-phase acidities.

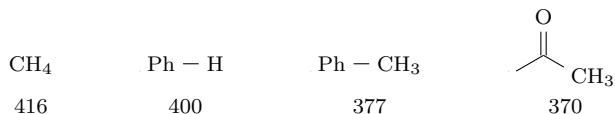


Figure 4.17: Gas phase acidity examples.

- We can intuitively rationalize these with resonance/EWG stabilization of the conjugate base.
  - **Acidity** (solution phase): The extent to which anions want to bind a proton in the solution phase. *Given by*
- $$\text{R}-\text{H}_{(\text{aq})} \rightleftharpoons \text{R}_{(\text{aq})}^- + \text{H}_{(\text{aq})}^+ \quad pK_a = \text{acidity}$$
- A refresher on what exactly “ $pK_a$ ” means.
    - The dissociation of an Arrhenius acid (exemplified by the above chemical equation) obeys the following mass action expression, where  $K_a$  is the **dissociation constant**.

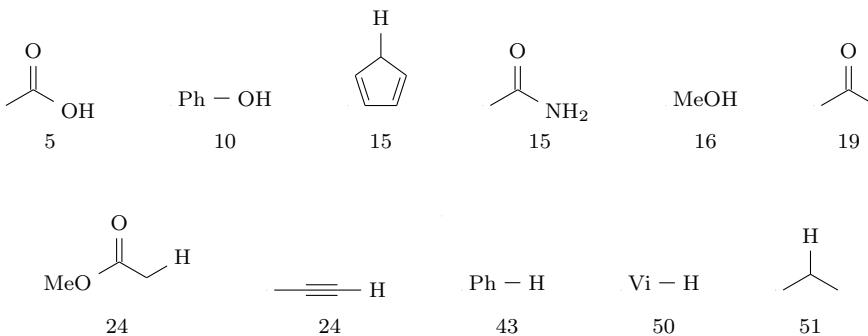
$$K_a = \frac{[\text{R}_{(\text{aq})}^-][\text{H}_{(\text{aq})}^+]}{[\text{RH}_{(\text{aq})}]}$$

- To look at the  $K_a$ 's in more human-readable units, we transform them to a log scale using the **p function**:

$$pK_a = -\log K_a \approx \Delta G$$

- Solution-phase acidity is a quantitative measure of anion stability.

- Some  $pK_a$ 's to know. (Memorize these!! These are the baseline, and they are a common qual question.).



- These are all measured in  $H_2O$ .
- These all come from the [Evans  \$pK\_a\$  table](#).<sup>[7]</sup>
- We should bookmark this page and refer to it regularly when we're trying to work out plausible mechanisms!!
- Solvent effects on the  $pK_a$ .

| <b>R-H</b>                        | <b><math>pK_a</math> (<math>H_2O</math>)</b> | <b><math>pK_a</math> (DMSO)</b> |
|-----------------------------------|--|---------------------------------|
| H-Cl                              | -8   | 1.8                             |
| PhCOOH                            | 4.2  | 11.1                            |
| CH <sub>3</sub> NO <sub>2</sub>   | 10   | 17.2                            |
| H <sub>2</sub> O                  | 15.7   | 32                              |
| H <sub>3</sub> CCN                | 25   | 31.3                            |
| CH <sub>2</sub> (CN) <sub>2</sub> | 11   | 11                              |

Table 4.2:  $pK_a$ 's in  $H_2O$  vs. DMSO.

- The  $pK_a$  of  $H_2O$  is about 15, so it is hard to get accurate measurements in  $H_2O$  for anything less acidic (higher  $pK_a$ ).
- One solution to this problem is to use DMSO as an alternate solvent.
  - The  $pK_a$  of DMSO is  $\approx 35$ .
  - This allows you to characterize a greater range of things.
  - DMSO is also very polar (like water), minimizing conflicting aggregation effects.
- All data on DMSO acidity comes from the [Bordwell  \$pK\_a\$  table](#).
- $pK_a$ 's are typically higher in DMSO than in water, as we can see in Table 4.2.
  - This is because  $H_2O$  is better at anion stabilization than DMSO, so the equilibrium is easier to access.
  - The trends are not always consistent, e.g.,  $CH_2(CN)_2$ .
    - When  $pK_a$ s are similar in different solvents, this tends to be because the anion is being stabilized internally.
    - For example, the  $(CN)_2HC^-$  anion is stabilized by both resonance and  $\sigma$ -EWG inductive effects. Since it is internally stabilized, its stabilization is less dependent on solvent effects.

<sup>7</sup>Not all of Masha's values match the Evans table (e.g., cyclopentadiene is 18.0, not 15). Whose value should we memorize??

- Factors influencing a compound's solution-phase acidity.

1. Electronegativity.

- More electronegative atoms make acids stronger.
  - This is because electronegative atoms inductively (i.e., through the  $\sigma$ -network) withdraw electron density, stabilizing the negative charge through delocalization.
  - Example: HOAc and TFA have  $pK_a = 4.76$  and 0.52, respectively.

2. Hybridization.

- More  $s$ -character leads to a stronger acid.
  - Essentially, orbital electronegativity (and hence stability) decreases  $s > sp > sp^2 > sp^3$ .
  - This is because  $p$ -orbitals "feel" the  $\delta^+$  nuclear charge less, owing to their node at the nucleus. Therefore,  $s$ -orbitals are a better place for  $\delta^-$  charge to reside in.
  - Example:  $-\equiv H$ , Ph-H, and  $^iPr-H$  have  $pK_a = 24$ , 43, and 51, respectively.
  - This effect also extends to nitrogen.
    - Example: Protonated imines are more acidic than protonated amines because their conjugate bases (the neutral imine and amine) have lone pairs in  $sp^2$  vs.  $sp^3$  orbitals.
    - Example: Piperidine is more basic than pyridine because its lone pair is in a relatively destabilized  $sp^3$  orbital.
    - Good qual question: Use a hybridization argument to differentiate basicities/acidities!!

3. Delocalization and aromaticity.

- A more delocalized anion means a stronger acid.
- Example: Cyclopentadiene and cyclopropene have  $pK_a = 15$  and 61 because the former deprotonates to an aromatic anion and the latter deprotonates to an antiaromatic anion.

4. Orbital overlap with adjacent atoms.

- Donation into adjacent  $d$ - or  $\sigma^*$ -orbitals stabilizes anions.
- Example:  $R_3P^+ - ^-CH_2$ .
  - This ylide has a relatively stable anion.
  - Ylides are especially stable when the adjacent atom is S or P; such ylides are synthetically useful (e.g., Swern oxidation and Wittig olefination).

- Anion structure.

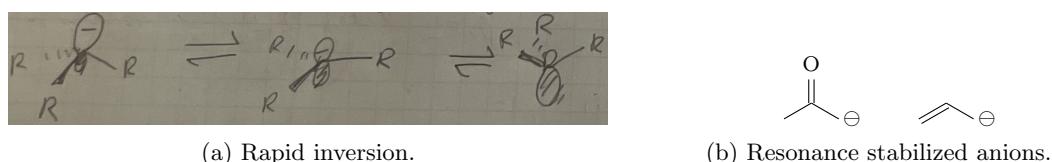


Figure 4.19: Anion structure.

- In general, anions are  $sp^3$ -hybridized, trigonal pyramidal species.
  - Recall that Figure 2.5 explains why anions are trigonal pyramidal instead of planar.
- Inversion is possible through a trigonal planar structure (Figure 4.19a).
  - The inversion barrier for carbanions is an *extremely* low 1-2 kcal/mol.
  - Because of this rapid inversion, anions behave as if they are planar even though they are not!
- Exception: Resonance stabilized anions are *actually* planar (Figure 4.19b).
  - This is because in these cases, the negative charge localizes to a  $p$ -orbital to have better overlap with the  $\pi$ -system with which it resonates.
  - This gives the carbanion atom an orbital structure of  $sp^2 + p$ .

- Some factors can raise the inversion barrier.
  1. Geometric constraints (i.e., incorporation into a small ring) raise the inversion barrier.
    - The planar structure of the ring requires some bond angles (e.g., between the anion lone pair and a substituent) to be larger than in the pyramidal structure.
    - Example: The cyclopropanide anion.
  2. More electronegative substituents raise the inversion barrier.
    - VBT explanation.
      - Electronegative groups prefer to bind to orbitals with more *p*-character (**Bent's rule**) since it's easier to "steal" those electrons because they're further from the nucleus.
    - MO theory explanation.
      - Consider the **D**-MO in Figure 2.5, which is the HOMO for an anion.
      - In the pyramidal structure, the *p*-AO in **D** will hybridize into an *sp*<sup>3</sup>-orbital, shedding some of its *p*-character. But per the "conservation of bonding character" discussed in Figure 2.14, this *p*-character will infuse the bonds to the (now electronegative) substituents.
      - Electronegative substituents then want this influx of *p*-character, lowering the energy.
      - Is this it, or am I missing something else??
- Other XR<sub>3</sub> structures with 8 electrons.
  - Consider H<sub>3</sub>C<sup>-</sup>, H<sub>3</sub>N, F<sub>3</sub>N, and H<sub>3</sub>P.
  - Their respective inversion barriers are 1-2 kcal/mol, 5 kcal/mol, 50 kcal/mol, and 35 kcal/mol.
  - Thus, H<sub>3</sub>C<sup>-</sup> and H<sub>3</sub>N are effectively planar, and F<sub>3</sub>N and H<sub>3</sub>P are pyramidal.
    - F<sub>3</sub>N is pyramidal due to its electronegative substituents.
    - H<sub>3</sub>P is pyramidal due to the HOMO of P being even more stabilized by its larger 3*p*-orbital.
    - Implication: We can have chirality at P, but rarely at N.
- Synthesis of carbanions: Two main ways.

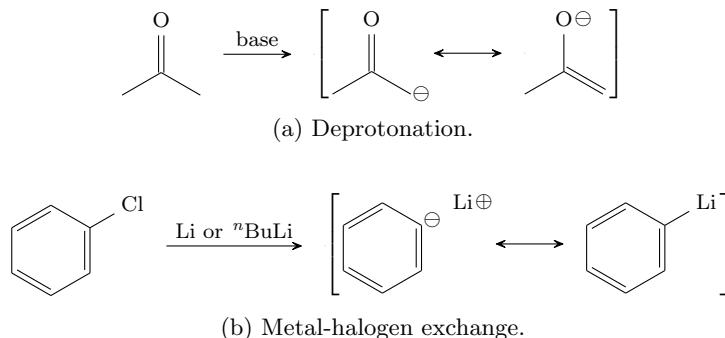


Figure 4.20: Synthesis of carbanions.

- A carbanion created by base deprotonation typically must be stabilized (e.g., by resonance).
- Metal-halogen exchange is typically used to create aryl, vinyl, and primary alkyl anions.
- Proton transfer and lithium-halogen exchange are among the fastest *intermolecular* reactions common in organic chemistry.
  - Some *intramolecular* reactions can be faster, e.g., 1,2-hydride shifts.
  - It is important to know such relative rates for reaction planning.
  - Caveat: Proton transfer from heteroatoms is much faster than proton transfer from carbon ( $k_{\text{rel}} \approx 10^6$ ). This is why we often talk about acidic X-H bonds, e.g., RCOOH.
- Note: B-H bonds are **hydridic** (think inorganic), not protic??

- Reactions of carbanions.

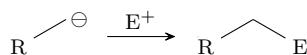


Figure 4.21: Reactions of carbanions.

- Anions are nucleophilic and basic.
- They react with electrophiles, such as protons and metals.
- Kinetic vs. thermodynamic acidity.



(a) Two deprotonation sites. (b) The energy diagram.

Figure 4.22: Energy differences governing kinetic and thermodynamic acidity.

- Rate of deprotonation varies based on whether our base attacks site A or B (Figure 4.22a).
- Site B has less steric clutter, so deprotonation is easier there.
  - This ease of deprotonation manifests as a lower energy  $\text{B}^\ddagger$  relative to  $\text{A}^\ddagger$  (Figure 4.22b).
  - To form this enolate, we should use a base such as LDA that is sterically bulky and has essentially irreversible deprotonation, so we will be under kinetic control.
- On the other hand, **Zaitsev's rule** tells us that the tetrasubstituted enolate we obtain by deprotonating at site A is more thermodynamically stable than the trisubstituted one we obtain by deprotonating at site B.
  - This difference in thermodynamic stability manifests as a lower energy  $\text{A}^\ddagger$  relative to  $\text{B}^\ddagger$  (Figure 4.22b).
  - To form this enolate, we should use a base such as an amine or alkoxide that has reversible deprotonation, so we eventually form the thermodynamic product.
- Great explanation of this phenomenon in Figures 5.3-5.7 (esp. Figures 5.5-5.6) of Labalme (2024c).
- **Zaitsev's rule:** The more substituted alkene is the more stable one.

# Week 5

## Misc. Reactive Intermediates

### 5.1 Radicals

- 10/1:
- Lecture 7 recap.
    - Anion formation:  $\text{R}-\text{H} \rightleftharpoons \text{R}^- + \text{H}^+$ .
    - $pK_a$ 's are a measure of anion stability.
    - Anions are stabilized by...
      - Electronegative substituents (that withdraw electron density);
      - More *s*-character (to hold the negative charge closer to the positive nucleus);
      - Delocalization/resonance (to spread out the negative charge);
      - Orbital overlap with adjacent atoms (along the lines of reverse hyperconjugation, e.g., in the case of ylides).
    - Anions are pyramids with low inversion barriers, and hence are effectively planar.
  - Today: Radicals.
  - Lecture outline.
    - Structure of radicals.
    - Stability of radicals (thermodynamic and kinetic).
    - Bond dissociation energy.
    - Synthesis of radicals.
    - Radical reactions.
    - Probing radical mechanisms: Radical clocks, traps, and cages.
    - Radical ions.
  - Structure of radicals.

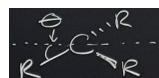
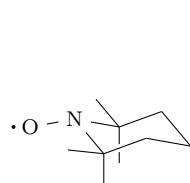


Figure 5.1: Angle of deviation from planarity.

- Most radicals are shallow pyramids with small inversion barriers (< 5 kcal/mol).
  - The methyl radical ( $\cdot\text{CH}_3$ ) is planar by  $\sim 10$  kcal/mol.
  - Recall the discussion of its QMOT diagram (Figure 2.5)!

- As with anions, electronegative substituents raise the inversion barrier.
  - For example, the trifluoromethyl radical ( $\cdot \text{CF}_3$ ) is pyramidal.
- Increasing sterics favor pyramidalization (more  $p$ -character)?? Wouldn't bulky groups push apart?
- Consider the angle of deviation  $\theta$  from planarity.
  - The ethyl radical ( $1^\circ$ ) has  $\theta = 11.9^\circ$ .
  - The isopropyl radical ( $2^\circ$ ) has  $\theta = 18.6^\circ$ .
  - The isobutyl radical ( $3^\circ$ ) has  $\theta = 24.1^\circ$ .
- Thermodynamic stability of radicals.
  - Delocalization stabilizes radicals.
    - Delocalization with neighboring heteroatoms is especially stabilizing!
  - Hyperconjugation stabilizes radicals.
    - Thus, in terms of decreasing stability,  $3^\circ > 2^\circ > 1^\circ$ .
    - This is analogous to cations.
  - More  $p$ -character stabilizes radicals.
    - Thus, in terms of decreasing stability,  $p > sp^3 > sp^2 > sp > s$ .
    - This is because radicals are inherently electron deficient, so they want to be further from the  $\delta^+$  nucleus.
    - This is the opposite of anions!
    - Alternatively: The more  $s$ -character, the stronger the bond, and hence the less stable the radical formed by homolytic bond cleavage.
      - We'll formalize this notion with **BDEs** in just a moment.
  - More electronegative atoms destabilize the radical center.
    - Thus, in terms of decreasing stability,  $\cdot \text{C} > \cdot \text{N} > \cdot \text{O}$ .
  - Larger atomic size stabilizes radicals.
    - Thus, in terms of decreasing stability,  $\cdot \text{S} > \cdot \text{O}$ .
    - This is because larger atoms are more polarizable.
- Kinetic stability of radicals.
  - Consider a relatively stable radical, such as one that has some resonance stabilization. Suppose we add some steric blocking to it. This yields a **persistent radical**.
- **Persistent radical:** A kinetically stable radical that may even be shelf-stable.
  - These are not thermodynamically stable: It's still a radical, so it doesn't want to exist. But it just won't react with anything.
  - Classic example: **TEMPO**.
  - Persistent radicals are useful for radical traps and other experiments discussed later this lecture.
- **(2,2,6,6-Tetramethylpiperidin-1-yl)oxyl:** A common persistent radical. *Also known as TEMPO.*  
*Structure*



(a) Structure.

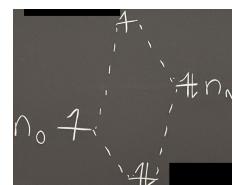
(b)  $\text{N}-\text{O}\cdot$  molecular orbitals.

Figure 5.2: TEMPO.

- Formed by oxidizing 2,2,6,6-tetramethylpiperidine (TMP), a sterically hindered organic base, with  $\text{H}_2\text{O}_2$  in the presence of the tungstate anion ( $\text{WO}_4^{2-}$ ).
- Check: TEMPO *does* have both resonance stabilization (with the adjacent nitrogen heteroatom) and steric blocking (from the adjacent quaternary carbons).
- The MO diagram (Figure 5.2b) reveals that the  $\text{N}-\text{O}\cdot$  bond is an example of a 2c-3e bond.
- The **spin density map** shows that the radical is evenly dispersed on O and N: 50% radical density on O and 50% on N.
- Takeaway: If you want to design your own persistent radical, take something with some resonance, add some steric blockers, and you're good to go!

- **Bond dissociation energy:** The energy it takes to symmetrically break a chemical bond. *Also known as BDE. Given by*



- Observe that this definition is analogous to those of HIA and  $\text{p}K_a$  from the past two lectures!
- X and Y can be organic groups, hydrogen, heteroatoms, etc.
- The above reaction denotes **homolytic** bond cleavage, as opposed to **heterolytic**.
- BDE is an extremely useful measure of “bond strength.” It’s probably one of the top three key concepts we should take away from Phys Orgo to use in the rest of our careers.
  - Guideline: A weak bond yields a more stable radical.
  - BDE is useful for predicting if a reaction is endothermic or exothermic.
  - One big factor that affects BDE is bond polarity.
    - In general, more polar bonds are stronger.
      - This contrasts with heterolytic cleavage, where polar bonds are easier to cleave.
      - Essentially, acidic bonds are “stronger” even if it’s easier to take off the acidic proton with your own “hands” (reagents) in lab.
    - See Table 5.1 for more.
    - Key point: When we talk about “strong bonds,” just remember that we’re talking about the BDE.

- **Homolytic** (bond cleavage): The breaking of a chemical bond in such a way that an *equal* amount of electron density is left on both products.
- **Heterolytic** (bond cleavage): The breaking of a chemical bond in such a way that an *unequal* amount of electron density is left on both products.
- Let’s look at some example comparisons between  $\text{p}K_a$ ’s and BDE’s to see the aforementioned inverse relationship.

|  | $\text{H}_3\text{C}-\text{H}$ | $\text{H}_2\text{N}-\text{H}$ | $\text{HO}-\text{H}$ | $\text{F}-\text{H}$ |
|--|-------------------------------|-------------------------------|----------------------|---------------------|
| $\text{p}K_a$ ( $\text{H}_2\text{O}$ ) | 48                            | 38                            | 15.7                 | 3.2                 |
| BDE (kcal/mol)                         | 105                           | 107                           | 119                  | 135                 |

Table 5.1: BDEs and  $\text{p}K_a$ ’s are inversely related.

- As we go to the right, it becomes easier to remove  $\text{H}^+$ .
- As we go to the left, it becomes easier to remove  $\text{H}\cdot$ .

- Some BDEs to know. (Memorize these!! They will likely come up in your Quals!)
  - The effect of bond polarity.
    - C–C:  $\sim 81$  kcal/mol.
    - C–H:  $\sim 98$  kcal/mol.
    - O–H:  $\sim 105$  kcal/mol.
  - The effect of hyperconjugation.
    - Me–H:  $\sim 105$  kcal/mol.
    - Et–H:  $\sim 100.5$  kcal/mol.
    - $i$ Pr–H:  $\sim 98.1$  kcal/mol.
    - $t$ Bu–H:  $\sim 95.7$  kcal/mol.
  - The effect of hybridization.
    - $\text{RC}\equiv\text{C}-\text{H}$ :  $\sim 132.8$  kcal/mol.
    - $\text{R}_2\text{C}=\text{CH}-\text{H}$ :  $\sim 111.2$  kcal/mol.
    - Ph–H:  $\sim 112.9$  kcal/mol.
  - The effect of resonance.
    - All–H:  $\sim 88.2$  kcal/mol.
    - Bn–H:  $\sim 88.5$  kcal/mol.
    - Note that allyl bonds are broken more easily than benzyl ones because resonance stabilizing the product radical doesn't force you to break aromaticity; indeed, breaking aromaticity is a little less fun than moving a  $\pi$ -bond around.
  - The effect of atomic size and polarizability.
    - Me–I:  $\sim 57.1$  kcal/mol.
    - Me–Br:  $\sim 70.3$  kcal/mol.
    - Me–Cl:  $\sim 83.7$  kcal/mol.
    - Me–F:  $\sim 110.0$  kcal/mol.
  - Peroxides.
    - HO–H:  $\sim 119$  kcal/mol.
    - HO–OH:  $\sim 51$  kcal/mol.
    - $t$ BuOO–H:  $\sim 88$  kcal/mol.
    - $t$ BuO–OH:  $\sim 44$  kcal/mol.
    - $t$ BuO–H:  $\sim 106$  kcal/mol.
- Polar effects on radicals.

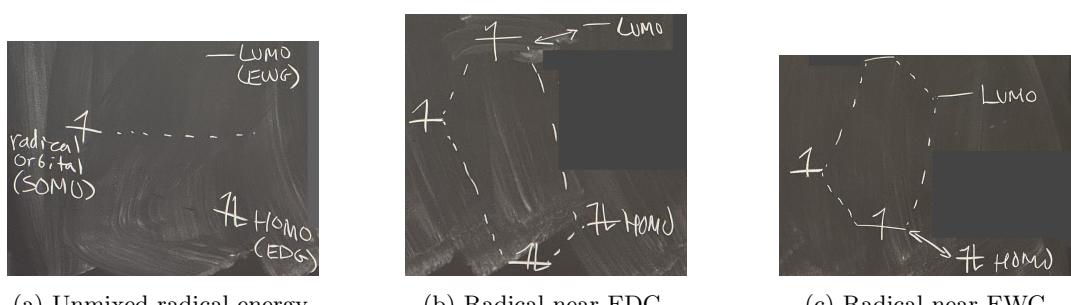


Figure 5.3: Radicals near EDGs and EWGs.

- Both EDGs *and* EWGs can stabilize radicals, despite the fact that radicals are electron deficient.
- Example: A radical  $\alpha$  to a carbonyl (e.g., homolytically cleave one of acetone's C–H bonds).
  - The carbonyl *will* destabilize the radical inductively.
  - However, it *will* also *stabilize* the radical through resonance.
  - The second effect (resonance) is stronger.
- Let's now justify these stabilizing effects using MO theory.
  - Before mixing (Figure 5.3a), a typical radical has energy intermediate between the HOMO of an EDG and the LUMO of an EWG.
  - When a radical's SOMO interacts with the HOMO of an EDG (Figure 5.3b), *two* electrons get stabilized and *one* gets destabilized. It follows that there is a net stabilization of the molecule, as expected.
  - When a radical's SOMO interacts with the LUMO of an EWG (Figure 5.3c), the sole electron present in the system gets stabilized. It follows that there is *still* a net stabilization of the molecule, even here!
- These molecular orbital diagrams reveal two additional attributes of polarized radicals, as well.
  1. EDGs make radicals more nucleophilic.
    - When a radical's SOMO interacts with the HOMO of an EDG (Figure 5.3b), a new radical SOMO (the antibonding orbital) is created.
    - This new SOMO is a better energy match with LUMOs, so the radical electron is more likely to mix with a LUMO since this will lead to greater thermodynamic stabilization of the product than before.
    - In other words, the radical is now more nucleophilic.
  2. EWGs make radicals more electrophilic.
    - When a radical's SOMO interacts with the LUMO of an EWG (Figure 5.3c), a new radical SOMO is once again created, but it is the stabilized bonding orbital this time.
    - This new SOMO is a better energy match with HOMOs, so the radical electron is more likely to mix with a HOMO since this will lead to greater thermodynamic stabilization of the product than before.
    - In other words, the radical is now more electrophilic.
- Synthesis of radicals.
  - Also known as **initiation**, if we're doing a radical chain reaction.
  - Most common way to make a radical: Homolytic cleavage of a weak bond.
    - We'll often use light or heat to give a little burst of energy and break this bond.
  - Commonly used radical initiators.
    - Peroxides are easily broken by light and heat, so we often use them.
      - Example: Organic peroxides react like  $\text{RO}-\text{OR} \xrightarrow[\Delta]{h\nu} \text{RO}\cdot$
    - AIBN.
    - $\text{Br}_2$ .
      - Bromine reacts like  $\text{Br}-\text{Br} \xrightarrow[\Delta]{h\nu} \text{Br}\cdot$
    - Paramagnetic metals, i.e., metals with 1 unpaired electron.
      - Example:  $\mathbf{Cp_2Ti^{III}Cl}$ .
    - Single-electron transfer (SET) or energy transfer (ET).
      - Often done with metals, electrochemistry, or photochemistry.
      - These are increasingly common ways to cycle one-electron oxidation states.
  - Essentially, if we ever need to make a radical, we can choose old-school or new-school based on what we have on hand!

- **Azobisisobutyronitrile:** A common thermal radical initiator. *Also known as AIBN.*



Figure 5.4: AIBN as a thermal radical initiator.

- This radical has a very cool design: Under heat or shock, you cleave the N–C bonds to form tertiary radicals that are additionally stabilized by their proximity to a  $\pi$ -system, and release N<sub>2</sub>.
- You commonly see AIBN used with HSnBu<sub>3</sub> in radical cyclizations.
- **Titanocene monochloride:** An increasingly popular SET agent. *Denoted by Cp<sub>2</sub>Ti<sup>III</sup>Cl.*

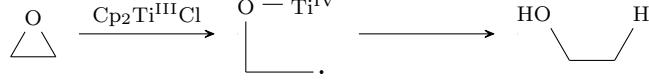


Figure 5.5: Cp<sub>2</sub>Ti<sup>III</sup>Cl as an SET radical initiator.

- For example, Cp<sub>2</sub>Ti<sup>III</sup>Cl can be used for radical-mediated epoxide openings, as shown above.
- Radical reactions.

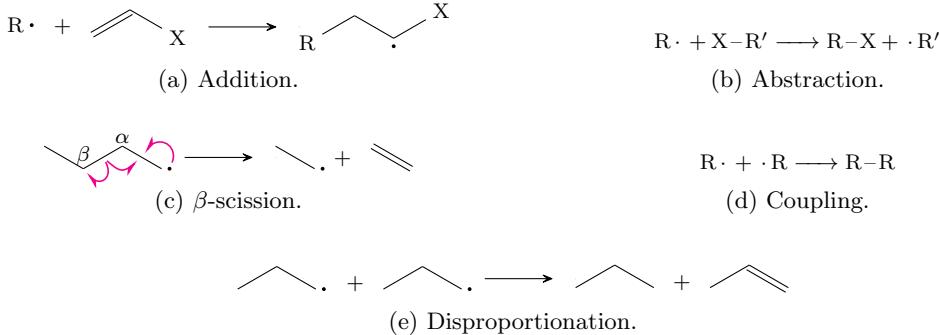


Figure 5.6: Radical reactions.

- Addition to multiple bonds (Figure 5.6a).
  - This can lead to cyclization, quenching, propagation, etc.
  - If it's a cyclization, we follow **Baldwin's rules**.<sup>[1]</sup>
- Abstraction (Figure 5.6b).
  - X is a halogen or hydrogen.
- $\beta$ -scission and fragmentation.
  - In general,  $\beta$ -scission refers to breaking the chemical bond between the carbons  $\alpha$  and  $\beta$  to the radical.
  - Fragmentation can be interpreted more broadly.
  - Example: The second step in the cleavage of benzoyl peroxide would count as fragmentation. Formally, this is called **radical decarboxylation**.
- Radical chain propagation.
  - This occurs by one of the above mechanisms.

<sup>1</sup>These rules are not covered in this course, but basically, they tell us which radical cyclizations are allowed.

- Polymerization.
  - If this is our goal, great! Radical chain reactions are great for making polymers.
  - If this is not our goal, it's a common side reaction for which we need to watch out.
- Radical coupling/dimerization, i.e., a **termination** step.
  - Two radicals form a bond.
  - $\Delta H$  is always negative (from an enthalpic point of view), but sterics can prevent this as with persistent radicals.
- Disproportionation.
  - A reaction in which two radicals form two nonradical products.
  - This is another possible termination step.
- **Barton deoxygenation.**

- **Radical decarboxylation.**

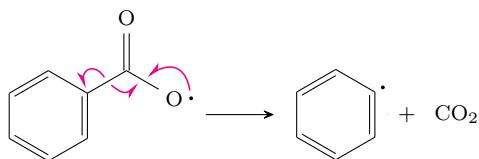


Figure 5.7: Radical decarboxylation.

- Radical decarboxylation can help us generate unstable radicals.
- For example,  $\text{Ph}\cdot$  isn't too stable normally, but we will form it under radical decarboxylation conditions regardless because  $\text{CO}_2$  is a really good leaving group. Basically,  $\text{CO}_2$  helps the thermodynamics work out.
- **Barton deoxygenation.** A method of deleting hydroxyl groups. *Also known as Barton-McCombie deoxygenation.*

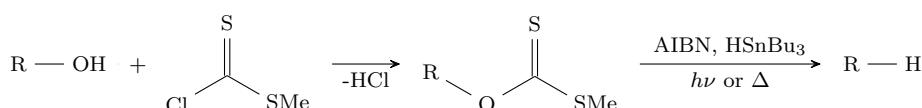


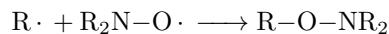
Figure 5.8: Barton-McCombie deoxygenation.

- Masha wanted to include this one named reaction, even though named reactions are not our focus in this class.
- Essentially, we react an alcohol to form a xanthate ester and then cleave it off with radicals.<sup>[2]</sup>
- David: Why would you ever choose to use a thermal initiator over a photochemical one?
  - Practically speaking, chemical initiators can be a bit easier to work with in lab because with photochemical, you have to find the exact right wavelength that will activate our initiator and do nothing else in our reaction.
- Radical **clocks** and **traps**.
  - This is the first of many mechanistic experiments we'll cover in this class, so take note of it in case you want to use it in your final project (the mechanistic proposal)!!
  - Radical clocks and traps both test for the presence of radical intermediates.

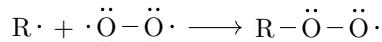
<sup>2</sup>See 5.47 notes for a mechanism.

- **Radical trap:** A species (often a persistent radical) that can quickly sequester a radical intermediate.

– Example: If we add TEMPO ( $\text{R}_2\text{N}-\text{O}\cdot$ ) to our reaction mixture, any radical intermediate  $\text{R}\cdot$  that is formed in solution is likely to react with TEMPO to form a **TEMPO adduct** as follows.



– Example: Per MO theory,  $\text{O}_2$  is a ground-state triplet diradical, so it can interact with  $\text{R}\cdot$  and form the peroxide as follows.



– Interpret the results of a radical trap experiment with caution — they can't be the basis of our whole argument that something is a radical mechanism; they are just a good first piece of evidence.

- **Radical clock:** A reaction with a known (fast) rate that is used to benchmark a radical reaction of interest.

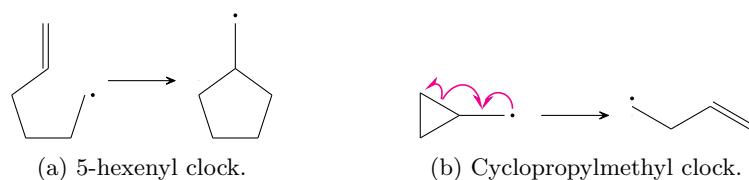


Figure 5.9: Radical clock reactions.

– Method: Synthesize an analogue of your substrate with a certain functional group attached such that if a radical is formed at a certain site, it will react with your new functional group instead of doing the designed reactivity.

– Example: Enable the formation of a 5-hexenyl radical so that it can do a **5-exo-trig cyclization**.<sup>[3]</sup>

■ The rate of this reaction is  $k = 2.3 \times 10^5 \text{ s}^{-1}$ .

– Example: Enable the formation of a cyclopropylmethyl radical so that it can do a radical ring opening and form an olefin.

■ The rate of this reaction is even quicker:  $k = 9.4 \times 10^7 \text{ s}^{-1}$ .

■ This is gold standard for a mechanistic experiment to prove radical mechanisms.

– These are very common kinetic probes for mechanisms!

- Another mechanistic experiment: The **radical cage effect**.

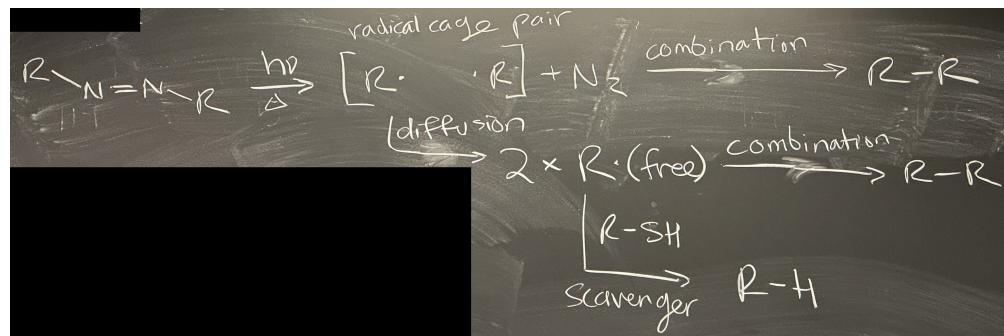


Figure 5.10: Radical cage effect.

<sup>3</sup>This is a type of radical cyclization allowed by Baldwin's rules.

- Consider a radical initiator of the form R-N=N-R.
  - When it decomposes — either thermally or photochemically — it will form two radicals that are in very close proximity to each other in solution.
    - We call these two radicals a **radical cage pair**, where the “cage” is the surrounding solvent molecules.
  - These radicals can easily recombine within the cage in a radical coupling/dimerization reaction to form R-R.
  - However, they can also diffuse out of the cage, drifting apart to yield 2 R· in solution.
    - These “free” radicals can then combine again to form R-R.
    - Or, alternatively, they can interact with a radical scavenger (such as a thiol<sup>[4]</sup>) in solution.
  - Notes on the cage effect.
    - A more viscous solvent makes it harder to escape the cage.
    - A scavenger can differentiate pathways.
    - Stereoretentive radical reactions can occur within a radical cage, because combination within the cage can outcompete stereoinversion.
      - However, if you diffuse out of the cage, forget it.
      - Takeaway: Cages can have stereochemical consequences, such as retention.
  - I need to do more reading on this and figure out exactly what I’m responsible for here!!
- Radical anions and cations.

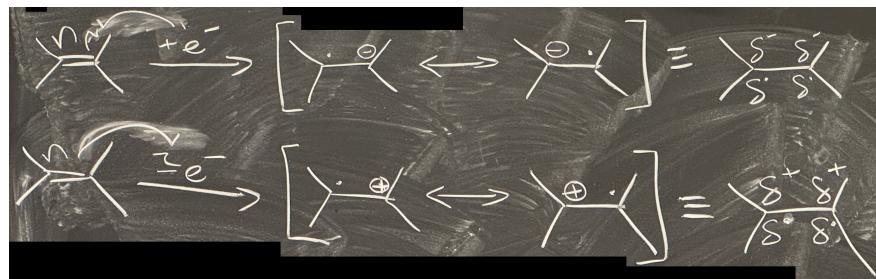


Figure 5.11: Radical ion formation and structure.

- Start with a  $\pi$ -system.
  - When we add an electron to the  $\pi$ -system or subtract one from it, we form a radical ion.
  - These radical ions exist in resonance with each other, giving us partial radical and ion character at both sides of the  $\pi$ -system.
  - Example of radical cations: Mass spec.
- Radical ions are common in aromatic rings.

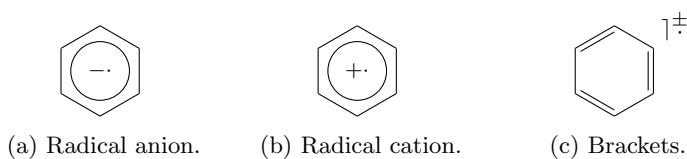
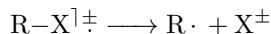


Figure 5.12: Aromatic radical ion notation.

<sup>4</sup> Alison Wendlandt uses thiols (such as adamantane thiol, AdSH) as HAD sources in her research!

- As such, we have a special notation for them.
  - Use a circle for the  $\pi$ -system, and then write “ $\pm$ ” in the center.
  - Alternatively, we can write the “ $\pm$ ” and “ $\cdot$ ” on top of each other outside a bracket surrounding the species.
- Example of aromatic radical anions: The Birch reduction. See Labalme (2024b).
- Radical ions can undergo **mesolytic cleavage** to generate a radical plus an ion.



- This is what happens in mass spec!
- It’s actually a common phenomenon, even though we often don’t think of it in that much detail.
- Example of radical ions: A plug for using these species in catalysis.

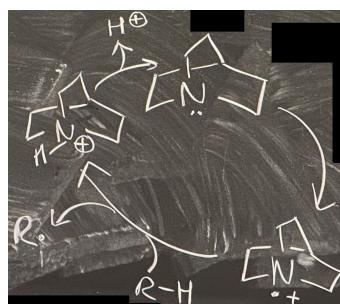


Figure 5.13: Catalysis with radical ions.

- Quinuclidine forms a radical cation, reacts with  $\text{R}-\text{H}$  to form  $\text{R}\cdot$  and quinuclidinium via an H-atom abstraction (HAA) pathway, and then can reform quinuclidine by loss of a proton. The  $\text{R}\cdot$  then goes on to do cool stuff.
- Here, quinuclidine is used as a catalytic initiator!
- Reference: Le et al. (2017).

## 5.2 Carbenes

10/3:

- Lecture 8 recap.
  - Radicals are shallow pyramids with low inversion barriers.
  - They are stabilized by hyperconjugation and resonance.
  - They are destabilized by electronegative atoms and  $s$ -character since they are electron deficient.
  - $\pi$ -EWGs stabilize radicals *and* make them electrophilic.
    - On the other hand,  $\sigma$ -EWGs (e.g., nearby halogens) *only* destabilize radicals.
  - EDGs stabilize radicals and make them nucleophilic.
  - You can make persistent radicals (i.e., radicals that *persist* over long timeframes/are kinetically stable) by using bulky steric blocking groups to prevent dimerization.
    - Recall that persistent radicals are still not thermodynamically stable.
- Today: Carbenes.
  - This is the last lecture in our “reactive intermediates” series.

- Lecture outline.
  - Orbital structure of carbenes (i.e., the origins of singlet vs. triplet states).
  - Substituent effects on carbene stability and state.
  - Synthesis of carbenes.
  - Carbene reactivity.
  - How carbene orbitals determine the transition-state angle of approach.
- Recall the Walsh diagram for  $\text{CH}_2$ .

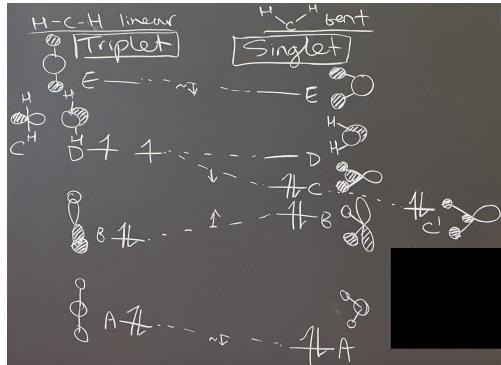


Figure 5.14: Walsh diagram for  $\text{CH}_2$ .

- Masha redraws Figure 2.6, but adds a higher-energy **E** orbital that is antibonding.
    - We will soon see that this **E** orbital is the LUMO of a so-called “triplet carbene.”
    - The **E** orbital also goes down in energy slightly when the carbene is bent.
  - Additional change from Figure 2.6: Per QMOT Rule 7 in Anslyn and Dougherty (2006), **C** also goes down a bit more to **C'** via secondary mixing and hybridization to an “ $sp^2$ -like” orbital.
  - Populating 6 electrons gives us a triplet diradical for linear, and a singlet for bent.
    - The  $sp^2$ -like **C'** ends up being our HOMO for singlet carbenes!
  - Sergei: Why don't secondary bonding interactions stabilize **D**?
    - Because the hydrogen orbitals are in the nodal plane of the constituent  $p$ -orbital, so they don't participate in this MO.
    - It's probably a symmetry thing.
  - **Multiplicity:** The following number, where  $S$  is the sum of the electron spins  $m_s$ . *Given by*
- $$2S + 1$$
- Example: All electrons are paired, save two electrons with parallel spins as in the linear carbene.
    - $S = 1/2 + 1/2 = 1$ , so the multiplicity is  $2(1) + 1 = 3$ . We call this a “triplet.”
  - Example: All spins are paired as in the bent carbene.
    - $S = 0$ , so the multiplicity is  $2(0) + 1 = 1$ . We call this a “singlet.”
  - This proves that our linear form is a **triplet carbene** and our bent form is a **singlet carbene**.
  - **Triplet** (carbene): A carbene with two unpaired electron spins. *Denoted by T.*
  - **Singlet** (carbene): A carbene with all electron spins paired. *Denoted by S.*

- Why would we favor a singlet carbene over a triplet, or vice versa?
  - It comes down to Hund's rule: We pay an energetic penalty for pairing electrons, and we pay an energetic penalty for putting electrons in higher orbitals, so we'll pay whichever is less.
  - Implication: As **C** and **D** get close in energy, we favor the triplet; as **C** and **D** get further apart, the singlet is favored.
- The **C'** and **D** orbitals in singlet and triplet carbenes.

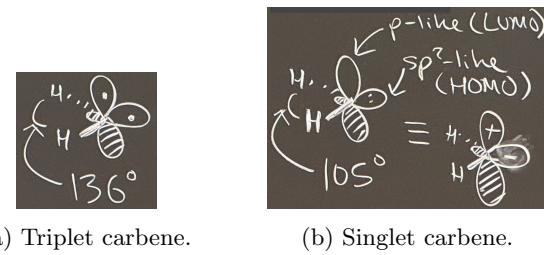
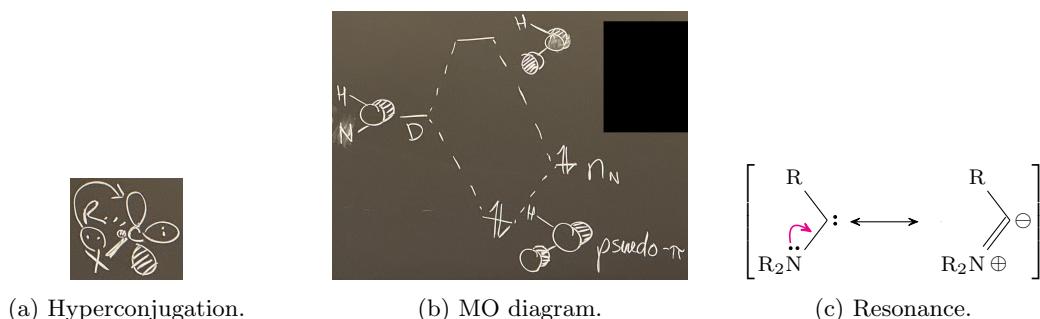


Figure 5.15: Singlet vs. triplet carbene orbitals.

- As we would expect from Figure 5.14, triplet carbenes are more linear and singlet carbenes are more bent.
  - However, the difference isn't as big as we might expect: The actual angles are 136° and 105°.
- We say that triplet carbenes have “diradical character” since we put an unpaired electron in both **C'** and **D**.
- We say that singlet carbenes have “cation + anion character” since we put two paired electrons in the *sp*<sup>2</sup>-like **C'** orbital and none in the *p*-like **D** orbital.
  - Note that this makes **C'** our HOMO and **D** our LUMO.
- Implication: Triplet carbenes react as diradicals, and singlet carbenes react as cations and anions.
- Let's now discuss some ways we can push carbenes to favor the singlet state, or the triplet state.
  - Specifically, we'll discuss the effects of placing  $\pi$ -donors,  $\pi$ -acceptors,  $\sigma$ -EWGs, and multiple bonds (e.g., for conjugation) near the carbene.
  - **Singlet-triplet gap:** The difference in energy between the singlet and triplet states of a carbene.
  - Simple dialkyl carbenes favor the triplet state.
    - Example:  $\text{H}_2\text{C}$ : favors the triplet state by 8.5 kcal/mol.
  - $\pi$ -donor substituents stabilize carbenes and favor the singlet state.

Figure 5.16:  $\pi$ -donor substituents stabilize carbenes and favor the singlet state.

- Example  $\pi$ -donor substituents: Neighboring lone pairs from atoms like N/O/X.
  - Additional example: **NHCs**.
- $\pi$ -donors favor the singlet state because they stabilize cations.
  - Intuitively, we can think of this as the electron density from a  $\pi$ -donor “pushing out” the radical electron from **D** and moving it into **C'**.
- More accurately, molecular orbital theory tells us that  $\pi$ -donors stabilize the high-energy, cationic **D**-orbital through hyperconjugation (Figure 5.16a).
- Formally, this hyperconjugation is manifested as a pseudo- $\pi$ -bonding interaction (Figure 5.16b).
  - **D** mixes with a nonbonding lone pair  $n_X$  to stabilize the lone pair and form a new, higher-energy LUMO.
    - Intuitively, the lone pair is stabilized because it has gotten to delocalize more!
    - This new LUMO is lower than **E** in energy.
  - In the process, **D** is effectively moved way up in energy in the Walsh diagram (Figure 5.14), forcing the carbene’s electrons into **C'**.
- One last way of representing this interaction is via resonance (Figure 5.16c).
  - This resonance structure reveals that the carbene electrons are properly an anion.
- To recap, there are *three* ways we can think about carbene stabilization by  $\pi$ -donors: Hyperconjugation, MO diagrams, and resonance diagrams.

- **N-Heterocyclic carbene: Also known as NHC. Structure**

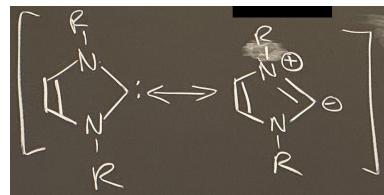


Figure 5.17: *N*-heterocyclic carbene.

- $\sigma$ -EWGs stabilize carbenes and favor the singlet state.
  - A  $\sigma$ -EWG is basically just an electronegative substituent.
  - $\sigma$ -EWGs stabilize carbenes because they stabilize anions.
    - Recall that  $\sigma$ -EWGs also *destabilize* radicals!
  - Example: F<sub>2</sub>C: favors the singlet state by  $\sim 50$  kcal/mol.
- $\pi$ -acceptor substituents stabilize carbenes and favor the triplet state.
  - Examples: Carbonyls, sulfones (SO<sub>2</sub>R), NO<sub>2</sub> groups, boron, etc.
  - These groups mix with **D** just like in Figure 5.16b, except that the new, lower-energy orbital is now empty!
    - This new, lower-energy orbital is our LUMO.
    - Since it has been stabilized, it is now close in energy to the **C'**-orbital.
    - This allows you to release the spin pairing energy by forming the triplet state.
  - These substituents also make carbenes more electrophilic.
- Conjugation stabilizes carbenes and favors the triplet state.
  - Example groups that can conjugate with carbenes: Alkenes, alkynes, and arenes.

- Synthesis of carbenes.

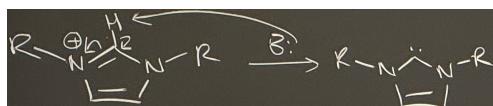


Figure 5.18: Synthesis of carbenes.

- We almost always synthesize carbenes via  $\alpha$ -elimination (Figure 5.18a).
  - If you're forming a carbocation and a carbanion at the same time, you're forming a carbene.
  - We can draw  $\alpha$ -elimination as a concerted or stepwise mechanism.
  - To deprotonate a secondary halide, we need a strong base.
- Aside (practical consideration): Don't run strongly basic reactions in CHCl<sub>3</sub>! Even something like KO<sup>t</sup>Bu can form dichlorocarbenes from chloroform in solution.
- Diazo compounds can also form carbenes pretty easily (Figure 5.18b).
- NHCs can form from an aromatic salt and a strong base (Figure 5.18c).
  - These aren't so much of a recent development anymore; they've definitely left their mark on organic chemistry.
  - These are important catalysts and ligands for transition metal catalysis.
  - If the R-groups are small, then NHCs will dimerize to form a double bond.
- Aside (story): When Masha gave this lecture last year, she talked about how you can put NHCs on quantum dots (to do cool things) in honor of Moungi's Nobel Prize.
- Aside (chemis-tea): On the Nobel Prize.
  - You get a medallion made of solid gold for you to keep and a few (cheaper) replicas to display.
  - You have to pay for the replicas, though. In fact, you have to pay for a lot: They take the expenses for flights, tickets to the gala, etc. out of your Nobel winnings (like 40 grand in total).
  - This was both Moungi's and Schrock's experience, which Masha found out when she had lunch with Schrock last week during his visit.
- Aside (chemis-tea): Get a tungsten wedding ring! Aqua regia will eat a gold or silver wedding band.
- Most synthetically useful reactions involve **carbenoids**, i.e., metal carbenes.
  - We won't talk about carbenoids in here because that's a topic for 5.44 - Organometallics, which we can take next year with Alison if we want.
  - To be clear, we'll only talk about **free** carbenes in this class.
- Free** (carbene): A carbene that is not bound to a metal.
- We will now begin discussing the reactivity of carbenes.
  - Singlet (cation/anion) and triplet (diradical) carbenes naturally react differently.
  - Carbene reactivity is easier to grasp than the carbene stability stuff from earlier :)

- We'll begin by discussing carbenes' reactivity toward [2 + 1] cycloaddition.
- Singlet carbenes add to alkenes in a concerted and stereospecific fashion.

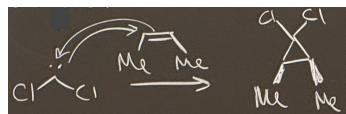
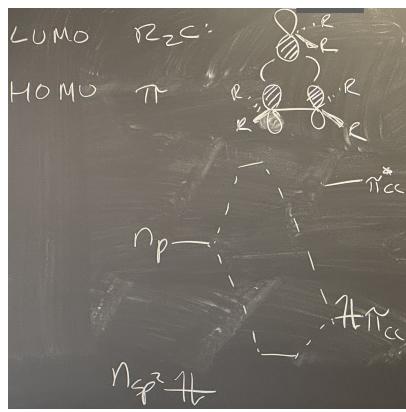
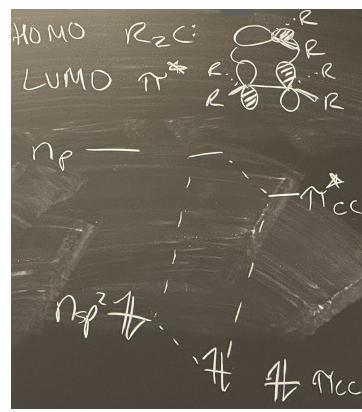


Figure 5.19: Carbene reactions: Chelotropic [2 + 1] cycloaddition.

- Example (Figure 5.19): Dichlorocarbene and 2-butene react to form a 100% *cis*-product.
- This is a chelotropic cycloaddition! Recall that we discussed carbene addition in Lecture 5.
- FMO analysis of this reaction.



(a) Electrophilic carbenes.



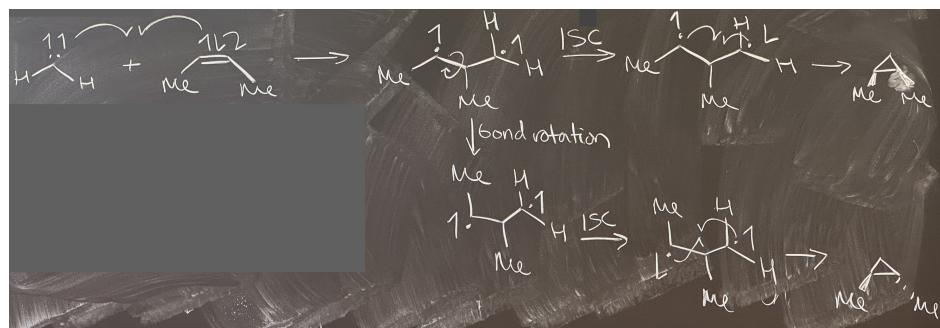
(b) Nucleophilic carbenes.

Figure 5.20: The orbitals behind chelotropic carbene cycloadditions.

- Electrophilic carbenes have low-energy orbitals, so their LUMO engages the alkene's  $\pi$ -HOMO.
- Nucleophilic carbenes have high-energy orbitals, so their HOMO engages the alkene's  $\pi^*$ -LUMO.
- Triplet carbenes also add to alkenes, but in a stepwise and non-stereospecific fashion.



(a) General form.



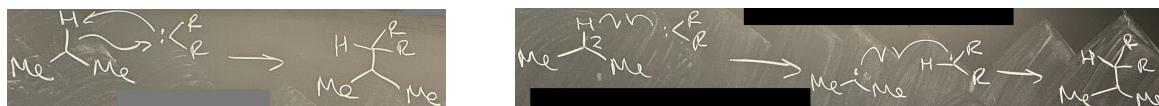
(b) Mechanism.

Figure 5.21: Carbene reactions: Stepwise [2 + 1] cycloaddition.

- Because the mechanism has changed, this [2 + 1] reaction is no longer chelotropic nor pericyclic.
- The rough product distribution is usually about 30% *cis* and 70% *trans* (Figure 5.21a).
- Let's investigate the mechanism (Figure 5.21b).
  - We begin with two unpaired electrons on the carbene (which have the same spin) and two paired electrons in the  $\pi$ -HOMO of the alkene (which, naturally, have opposing spins).
  - One of the carbene electrons will react with the alkene electron having specifically the *opposite* spin.
    - Postulate: Only radicals with paired spins can react.
  - After the first step, two radicals remain (one from the olefin and one from the carbene).
  - Since they both have the same spin, though, we will need **ISC** to the singlet state before we can proceed.
  - Following ISC, the remaining radicals can react to form a second bond.
  - Alternatively, since ISC takes a bit of time, we can have a bond rotation followed by an ISC, followed by bond formation.
- Takeaway: The *cis*-to-*trans* ratio depends on the relative rates of the bond rotation vs. ISC.
- Hint for the mechanistic proposal!!
  - Suppose you're proposing a mechanistic study of a triplet carbene.
  - We can learn from orthogonal experiments how long a bond rotation takes.
  - Thus, the *cis*-to-*trans* ratio can be a good probe for our triplet lifetime!

- **Intersystem crossing:** A spin flip. *Also known as ISC.*

- We will learn more about this later in the semester!
- Steven: Are ISC and pairing always separate, or can they ever happen together?
  - Pairing involves moving electrons into different orbitals, and spin flipping is just spin flipping; essentially, they're two different kinds of processes, and this is why we draw them separately.
  - You can't pair electrons unless they're spin-flipped (except perhaps in some niche application).
- We'll now talk about carbenes' reactivity toward insertions.



(a) Concerted (singlet carbenes).

(b) Stepwise (triplet carbenes).

Figure 5.22: Carbene reactions: Insertions.

- These are reactions with C–H bonds.
- Singlet carbenes reacted through concerted mechanisms, and triplet carbenes react through stepwise mechanisms.
  - This means that only singlet carbenes engage in a *true* insertion into the C–H bond.
  - Triplet carbenes, on the other hand, have an H-atom transfer (HAT) followed by radical recombination.
- Aside: HAT refers to hydrogen ( $H\cdot$ ) transfer.
  - It is distinct from proton ( $H^+$ ) transfer and hydride ( $H^-$ ) transfer.
  - Know which one you're talking about!

- Moving on from insertions, let's talk about ring expansions.



(a) Expanding benzene.



(b) Expanding pyrrole.

Figure 5.23: Carbene reactions: Ring expansions.

- Benzene as a starting material (Figure 5.23a).
  - This reaction involves addition to an olefin followed by a  $6\pi$  electrocyclization.
  - Carbenes are so reactive that they can even break aromaticity!
- Pyrrole as a starting material (Figure 5.23b).
  - This reaction involves the same addition to an olefin as before. However, it is then followed by elimination of a chlorine and proton transfer to afford pyridine.
  - This is an example of single-atom editing because it allows us to insert a specific carbon atom into a heterocycle. Single-atom editing is currently blowing up in the literature.
- Steven: Why do we add to the side double bonds in pyrrole instead of the back one?
  - It will be obvious if you draw the product of such an addition: Said product will be much more charge separated.
    - Specifically, if we're going to engage the back  $\pi$ -bond, we have to do so from the resonance structure with a carbanion on one carbon and a positive charge on the nitrogen.
  - Takeaway: The side bonds in pyrrole tend to react more as separate  $\pi$ -systems.
    - This is because the aromatic stabilization of pyrrole is like 17 kcal/mol, i.e., pyrrole is much less aromatic than benzene (35 kcal/mol).
  - Blocking groups might force the second reaction, but it would def give a less stable product.
  - The second reaction would give 4-chloropyridine instead of 3-chloropyridine.
    - We might get a 99 : 1 ratio naturally, but if we put an EWG at the 2-position, that could stabilize a resonance form, giving us 2-EWG-4-chloropyridine.
- Aside: Learning to draw 7-membered ring.

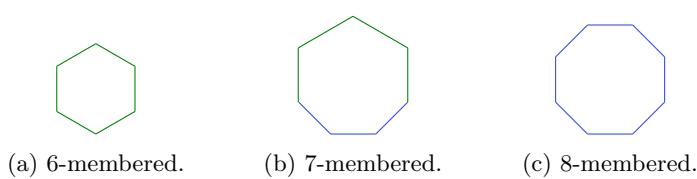


Figure 5.24: Drawing a 7-membered ring.

- On paper, draw half a hexagon and half an octagon. This is easier to read than the “proper” equidistant form, which we should leave that for ChemDraw.

- Masha: “Jonathan [the TF] is like the most brilliant person ever, so whenever I find a chink in the armor, it makes me happy.”
- Lastly, we’ll discuss rearrangements.

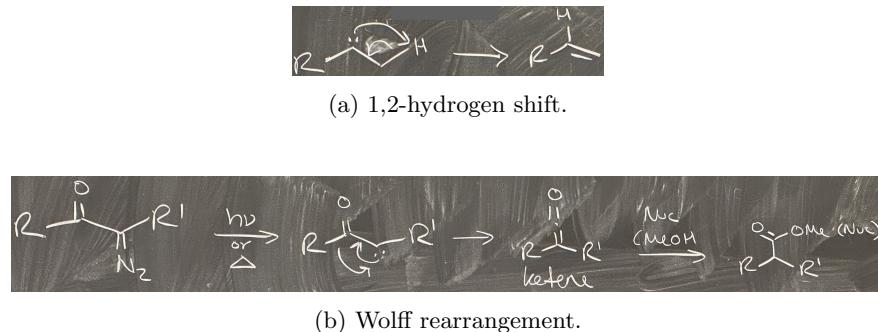


Figure 5.25: Carbene reactions: Rearrangements.

- The most common one is a 1,2-hydrogen shift (Figure 5.25a).
- Then there are a ton of named reactions, e.g., the **Wolff**, **Curtius**, and **Hoffmann rearrangements**.
  - We’ll never run these reactions in our lives, but they teach them all the same regardless... perhaps because they’re mechanistically interesting.
- Example named rearrangement: The Wolff rearrangement (Figure 5.25b).
  - Either light or heat can be used to cleave the diazo functional group to a carbene.
  - Then we get a rearrangement into a ketene.
  - Then a nucleophile attacks the ketene’s central carbon (the electrophilic one) to give a ketone.
  - This rearrangement is useful for lots of things, e.g., a 5.47 problem!
- Angle of approach.

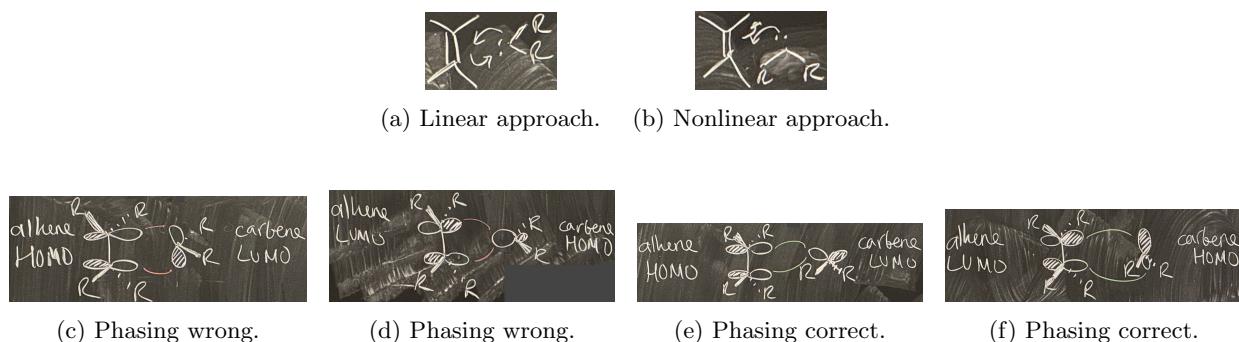


Figure 5.26: Angle of approach in chelotropic [2 + 1] cycloadditions.

- Consider a [2 + 1] with an alkene.
- The mechanism is actually side-to-side per MO theory (Figure 5.26b), not linear and head on as we often draw (Figure 5.26a).
- Linear approach.
  - The alkene and carbene can both be HOMO/LUMO.
  - First, let the alkene be the HOMO and the carbene the LUMO (Figure 5.26c).
    - The phasing is all wrong, because we’ve got a shaded lobe mixing with an unshaded lobe.

- Now consider the alkene LUMO and the carbene HOMO (Figure 5.26d).
  - Here, the phasing is still all wrong, because we're interacting a single carbene lobe with two different shaded  $\pi^*$ -lobes.
- So the phasing is wrong in both cases.
  - Nonlinear approach.
    - Alkene HOMO and carbene LUMO (Figure 5.26e).
      - Good phasing match.
    - Alkene LUMO and carbene HOMO (Figure 5.26f).
      - Good phasing match as well!
  - Conclusion: Nonlinear approach is required!

# Week 6

## Thermodynamics

### 6.1 Selectivity

10/8:

- Lecture 9 recap.
  - Last lecture wrapped up reactive intermediates, focusing specifically on carbenes.
  - Triplet carbenes (Figure 5.15a).
    - More linear.
    - Smaller HOMO-LUMO gap implies 2 SOMOs.
    - React as diradicals.
    - R can be any  $\pi$ -acceptor, such as alkyl, vinyl, aryl, carbonyl,  $\text{SO}_2\text{R}$ ,  $\text{NO}_2$ , B, etc. groups.
  - Singlet carbenes (Figure 5.15b).
    - More bent.
    - Larger HOMO-LUMO gap.
    - React as cations and anions.
    - R can be any  $\pi$ -donor or  $\sigma$ -EWG, such as halogens,  $\text{NR}_2$ , or OR groups.
  - Both types of carbenes...
    - Can be nucleophilic or electrophilic;
    - React by adding into  $\pi$ -systems or inserting into bonds.
      - The mechanisms through which S/T carbenes engage in this reactivity vary slightly.
- Today: Selectivity.
- Lecture outline.
  - Thermodynamic selectivity.
  - Kinetic selectivity.
  - Curtin-Hammett kinetics.
  - Kinetic quench.
  - Principle of microscopic reversibility.
  - Reactivity-selectivity principle.
  - Practical aspects of selectivity (deferred to next time).
- When two products form from a single common intermediate (or starting material), selectivity between these products can arise from **thermodynamic** or **kinetic** factors.

- **Thermodynamic** (selectivity): Selecting for a certain product based on the position of an equilibrium, i.e., the stability of the products.

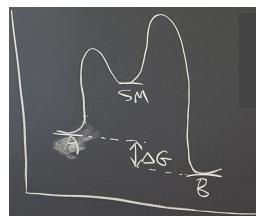


Figure 6.1: Energy variables relevant to thermodynamic selectivity.

- Key words: Thermodynamic = product = equilibrium.
- Relevant reaction coordinate.



- A and B form from a single common starting material (SM).
- The relevant equilibrium constants are  $K_A$  and  $K_B$ .
- $K_A$  and  $K_B$  allow us to define the **selectivity** of this reaction as follows.

$$\text{selectivity} = \frac{[A]}{[B]} = \frac{K_A}{K_B} =: K_{\text{eq}}$$

- Energy diagram of a thermodynamically controlled reaction (Figure 6.1).
  - In order for a reaction to be under thermodynamic control, all steps must be reversible, i.e., all intermediates must interconvert.
  - $\Delta G$  is the difference in energy between the products.
  - Recall from Gen Chem that  $\Delta G = -RT \ln(K_{\text{eq}})$  and hence  $K_{\text{eq}} = e^{-\Delta G/RT}$ .
- Thermodynamic selectivity is very useful if all products are at very different energy levels.
- Example: Olefin isomerization can occur with great selectivity because one product can be much more stable than another.

- **Selectivity** (of a reaction): The preference for one product (A) over another (B), where both A and B originate from a single common intermediate or starting material. *Given by*

$$\text{selectivity} := \frac{[A]}{[B]}$$

- **Kinetic** (selectivity): Selecting for a certain product based on the differences in energies of competing transition states, i.e., by reaction rates.

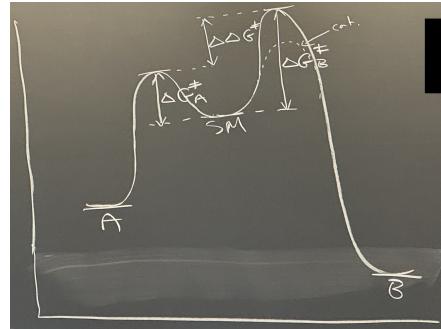


Figure 6.2: Energy variables relevant to kinetic selectivity.

- Key words: Kinetic = transition state = rate.

- Relevant reaction coordinate.



- As before, A and B form from a single common SM.
- The relevant rate constants are  $k_A$  and  $k_B$ .
- $k_A$  and  $k_B$  allow us to define the selectivity of this reaction as follows.

$$\text{selectivity} = \frac{[A]}{[B]} = \frac{k_A}{k_B}$$

- Energy diagram of a kinetically controlled reaction (Figure 6.2).

- $\Delta G_A^\ddagger$  and  $\Delta G_B^\ddagger$  are the activation energies required to form the transition states from the SM to A and B, respectively.
- $\Delta\Delta G^\ddagger$  is then the difference between these transition states' activation energies.
- Recall from Gen Chem that  $\Delta\Delta G^\ddagger = -RT \ln(k_A/k_B)$ .<sup>[1]</sup>
- Often,  $k_A/k_B$  is equal to the relative rate  $k_{\text{rel}}$  of the two reactions ( $\text{SM} \longrightarrow A$  and  $\text{SM} \longrightarrow B$ ).
  - If A and B are enantiomers or diastereomers,  $k_{\text{rel}}$  often equals **er** or **dr**, respectively.
  - Another consequence of the introduction of  $k_{\text{rel}}$  is that  $k_{\text{rel}} = e^{-\Delta\Delta G^\ddagger/RT}$ .
- Note that *catalyzing* a pathway is a kinetic effect, corresponding to a lower activation barrier.
- In contrast to thermodynamic equilibrium, the products formed here are formed irreversibly and do not interconvert.
- Kinetic control is more common than thermodynamic control.
  - Reactions under thermodynamic control have largely been developed and optimized over the last 100 years, so kinetic control gives us a better handle in modern methods development.
  - Everything about a catalytic cycle is based on kinetics! You're not changing the thermodynamics of  $\text{CO}_2$  upcycling; you're making it more energetically feasible.

- **Enantiomeric ratio:** The ratio of the (*S*)-enantiomer to the (*R*)-enantiomer. *Denoted by er.*

- This is more mathematically useful than the enantiomeric excess (ee), so there's currently something of a push to phase out ee in favor of er.
- ee is still used primarily for historical reasons.

- **Diasteriomic ratio:** The ratio of one diastereomer to the other. *Denoted by dr.*

- We now discuss a special type of kinetic control called **Curtin-Hammett kinetics**.

- **Curtin-Hammett (kinetics):** A kinetic regime characterized by two starting materials or intermediates that rapidly interconvert, causing the ratio of products (i.e., the selectivity) to depend only on the transition state energies. *Also known as C/H.*

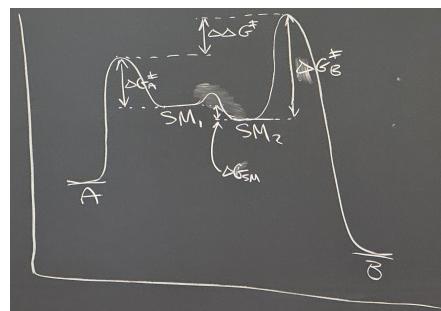
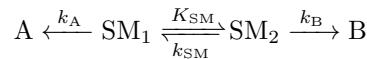


Figure 6.3: Energy variables relevant to Curtin-Hammett kinetics.

<sup>1</sup>This can be derived by dividing the Arrhenius equation for one reaction by the Arrhenius equation for the other reaction and rearranging.

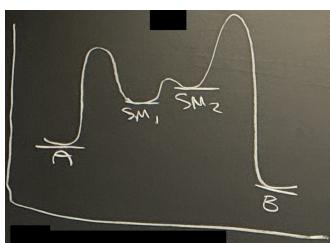
- In particular, the selectivity does *not* depend on the energies of the starting materials.
- Relevant reaction coordinate.



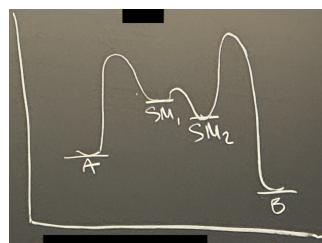
- $k_{SM}$  must be big. Typically, it is approximately ten times faster than  $k_A$  or  $k_B$ .
- Working out the math, we get

$$\text{selectivity} = \frac{[A]}{[B]} = e^{-\Delta\Delta G^\ddagger/RT}$$

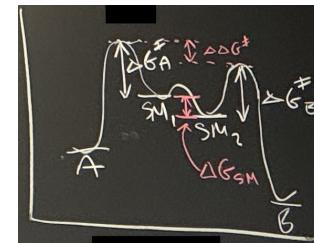
- Indeed, we see that in this regime, the selectivity *mathematically* depends only on the relative energies of the transition states.
- Energy diagram of a reaction under Curtin-Hammett kinetics (Figure 6.3).
  - Note that there is only a small energy barrier between  $SM_1$  and  $SM_2$  because we need fast interconversion.
  - Observe that the products are formed irreversibly and do not interconvert.
  - Indeed, the SMs interconvert freely as long as they stay SMs, but once they go over their barrier to A or B, they do not continue to interconvert.
- Scenarios that manifest Curtin-Hammett kinetics.



(a) More stable reacts more quickly.



(b) Less stable reacts more quickly.



(c) Both react same.

Figure 6.4: Curtin-Hammett scenarios.

1. The more stable starting material reacts more quickly (Figure 6.4a).
  - Let  $SM_1$  be lower energy than  $SM_2$ , and let the  $SM_1 \rightarrow A$  transition state have a lower activation energy than the  $SM_2 \rightarrow B$  transition state.
  - It follows that  $SM_1$  is thermodynamically favored. This means that we'll see more of it in solution:  $[SM_1] > [SM_2]$ .
  - The lower activation energy to form A (i.e.,  $\Delta G_A^\ddagger < \Delta G_B^\ddagger$ ) implies that A is kinetically favored.
  - The product ratio will not be equal to the starting material ratio.
    - You might not even see  $SM_2$  among the starting materials; you might just think that  $SM_1 \rightarrow A + B$ .
    - Takeaway: It isn't always obvious when Curtin-Hammett kinetics are in effect.
2. The less stable starting material reacts more quickly (Figure 6.4b).
  - Let  $SM_1$  be higher energy than  $SM_2$ , and let the  $SM_1 \rightarrow A$  transition state have a lower activation energy than the  $SM_2 \rightarrow B$  transition state.
  - It follows that  $SM_2$  is thermodynamically favored. This means that we'll see more of it in solution:  $[SM_2] > [SM_1]$ .

- The lower activation energy to form A (i.e.,  $\Delta G_A^\ddagger < \Delta G_B^\ddagger$ ) implies that A is kinetically favored.
- The less stable starting material is kinetically favored to react.
- Takeaway: All the reactivity goes through  $SM_1$ , even though we might not even see  $SM_1$ ; you might just think that  $SM_2 \longrightarrow A + B$ .
- This is classic Curtin-Hammett kinetics, wherein the product we observe is from the starting material we don't observe.
  - Results like this can be confusing because the SM we put in the flask doesn't look like it'd give the product we see.
  - This contrasts with Scenario 1, wherein the SM we see logically leads to our product A, and all we miss is that there's a secret equilibrium that helps us get to B.
- 3. Both starting materials react equally quickly (Figure 6.4c).
  - Let  $SM_1$  be higher energy than  $SM_2$ , and let the  $SM_1 \longrightarrow A$  and  $SM_2 \longrightarrow B$  transition states have identical activation energies (i.e.,  $\Delta G_A^\ddagger = \Delta G_B^\ddagger$ ).
  - We call this **ground state control**.
    - Thus,  $\Delta G_{SM}$  suddenly predicts our products; not because it actually does but because  $\Delta\Delta G^\ddagger = \Delta G_{SM}$ .
    - To reiterate:  $\Delta\Delta G^\ddagger$  still controls selectivity; it just happens that it equals  $\Delta G_{SM}$ .
  - Because  $\Delta\Delta G^\ddagger = \Delta G_{SM}$ , we can work out mathematically that the selectivity happens to be the following (even though we still have C/H kinetics).

$$\text{selectivity} = \frac{[A]}{[B]} = \frac{[SM_1]}{[SM_2]}$$

- This regime often arises when A and B are really similar and hence have similar transition states (e.g., if A and B are enantiomers or diastereomers with far apart stereogenic centers).
- It's our job as the responsible scientist to account for the full kinetic picture, even when it may not provide us much additional information!
  - Indeed, the reactions that are the most interesting to develop are the ones that fall in this C/H regime because they have the most subtle reactivity.
- Let's now look at some examples.
  - Pay attention, because this is going to be a super useful skill for grad school and beyond!!
  - Example: Nitrogen rapidly epimerizes while a *tert*-butyl group locks the chair in place.

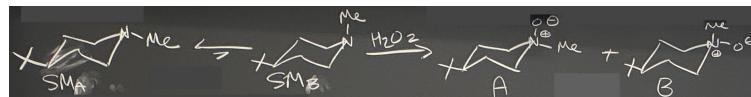


Figure 6.5: Curtin-Hammett kinetics: Kinetically trapping epimers.

- This epimerization (a **nitrogen inversion**) occurs fast relative to product formation.
- It puts  $SM_A$  and  $SM_B$  in a 98 : 2 ratio.
- Either epimer can react with  $H_2O_2$  to form the N-oxo products in a 5 : 95 (A : B) ratio.
- This is an example of Scenario 2 (Figure 6.4b).
  - Your first thought might be that the oxidation occurs with inversion of stereochemistry. This is a great first thought.
  - But then you have to ask about alternate scenarios, and you should think about decoupled Curtin-Hammett steps wherein you're just kinetically trapping the epimers.

- Example: Axial and equatorial tosylates equilibrate before E<sub>2</sub> elimination to form a double bond.



Figure 6.6: Curtin-Hammett kinetics: Elimination.

- Let SM<sub>A</sub> be the axial tosylate (on the left), and let SM<sub>B</sub> be the equatorial tosylate (on the right).
- Because of the large steric bulk of the tosylate group and hence its disfavored 1,3-diaxial interactions, SM<sub>A</sub> and SM<sub>B</sub> occur in a 1 : 14 ratio.
- However, SM<sub>A</sub> has hydrogens antiperiplanar to it, so it reacts faster ( $k_{\text{rel}} = 70$ ).
- So to recap: SM<sub>B</sub> is preferred, but the product comes from SM<sub>A</sub>. Therefore, this must be another example of Scenario 2 (Figure 6.4b).
- Example: *trans* and *cis* alkenes react via bromination to form a *trans*- and *cis*-dibromide.

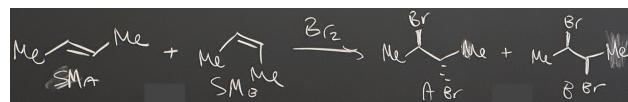


Figure 6.7: Curtin-Hammett kinetics: Bromination of geometric isomers.

- We have a 1 : 1 mixture of SMs, and we form a 1 : 1 mixture of products.
- Thus, based on the selectivity equation, it looks like this could be a candidate for Scenario 3. However, this is not C/H because the SMs do not interconvert! Rather, this is a case of a kinetic quench, which we'll cover next.
- Learn C/H because we will see a lot of it on PSet 2.
- Kinetic quench (not C/H).

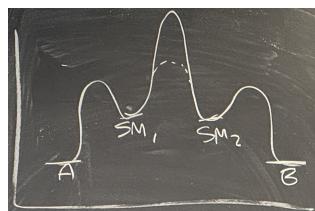


Figure 6.8: Energy variables relevant to a kinetic quench.

- Here, the SM<sub>1</sub> ⇌ SM<sub>2</sub> interconversion is slower than product formation.
- Thus, the ratio of starting materials equals the ratio of products, as follows.

$$\text{product ratio} = \frac{[A]}{[B]} = \frac{[\text{SM}_1]}{[\text{SM}_2]}$$

- This is basically a case of two isolated systems (SM<sub>1</sub> → A and SM<sub>2</sub> → B).<sup>[2]</sup>

<sup>2</sup>Could I come up with one-pot reactions where you have two different starting materials under kinetic quench form two different products and then those products react???

- One tricky thing: When the rate of interconversion approximately equals the rate of product formation (Masha shows this regime with the dotted line in Figure 6.8).
  - In this case, the product ratio is difficult to predict!
  - That's real, messy science.
  - When you encounter such a regime, either you change something to make it simpler, or you do a Wendlandt-style deep dive on the full mechanism where you uncover the secrets of the universe and then publish a bunch of *Science* papers.
  - “Alison’s the master of these really hairy and difficult kinetic pictures and disentangling them and adding to our understanding of chemistry overall.”
- Example of kinetic quench: Protonating two different epimers of an amine.

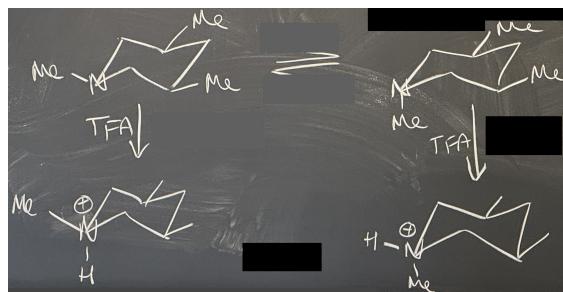


Figure 6.9: Kinetic quench: Protonation.

- The epimer with the equatorial methyl occurs in a  $> 15 : 1$  ratio.
- Epimerization occurs relatively slowly, protonation of the equatorial lone pair occurs fast, and protonation of the axial lone pair is even *faster* than protonation of the equatorial one.
  - What is “fast” and “slow” is all relative! Usually, nitrogen inversion is fast, but proton transfer (PT) to nitrogen is even faster.
- However, the product ratio is also  $> 15 : 1$ , just like the SM ratio. To reiterate, this is because we’re not interconverting between our starting materials.
- Moving on, let’s discuss the **principle of microscopic reversibility**.
- **Principle of microscopic reversibility:** The lowest energy path connecting two intermediates is the same, regardless of the direction in which the reaction proceeds.
  - Basically, if you propose a mechanism from A  $\longrightarrow$  B, the same mechanism (in reverse) has to be true for B  $\longrightarrow$  A.
  - If we proceed through a certain transition state in one direction, we cannot proceed through a different transition state on the way back.
  - Really useful to probe kinetically silent steps.
- A cool example of using the principle of microscopic reversibility to see which mechanism is operative (Figure 6.10).
  - Consider the elimination of a  $\beta$ -hydroxyketone to form an enone (Figure 6.10a).
  - Is the mechanism E<sub>2</sub> (Figure 6.10b) or E<sub>1</sub>CB (Figure 6.10d)?
    - Retro-E<sub>2</sub> (Figure 6.10c): A one-step forward reaction for E<sub>2</sub> means a one-step reverse reaction, wherein HO<sup>-</sup> adds in, the olefin grabs a proton from water, and HO<sup>-</sup> leaves.

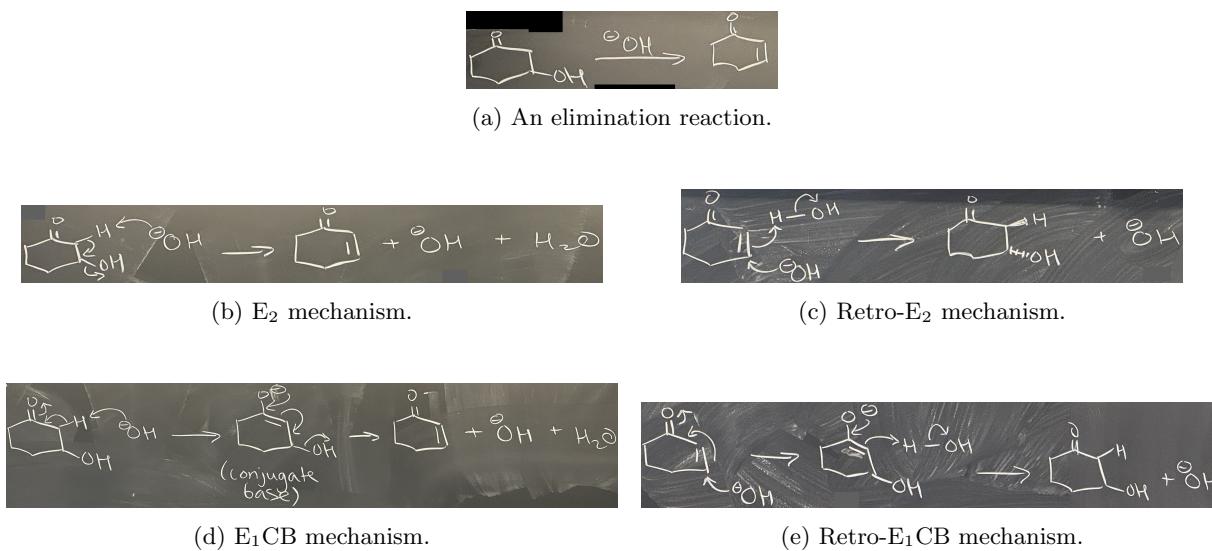


Figure 6.10: Microscopic reversibility to differentiate plausible mechanisms.

- Retro-E<sub>1</sub>CB (Figure 6.10e): This time, a two-step reverse reaction is implied. First, we kick electron density all the way up to oxygen, and second, we kick arrows back down to grab a proton.
  - Which reverse mechanism is more plausible?
    - In Figure 6.10c, we need a termolecular transition state (which is possible, but rare). However, we'd also form only the anti product, and this is flatly inconsistent with experiment.
    - In Figure 6.10e, we have a conjugate addition step followed by an enolate protonation step, both of which are very typical reactions.
      - Molecular orbital theory also implies that the electrons push all the way up through the conjugated system to the oxygen in a concerted step upon nucleophilic addition at the Bürgi-Dunitz angle, like in 5.13!
  - Now remember that the more reasonable mechanism must follow the same steps in the forward and reverse direction.
    - Thus, more reasonable in reverse implies more reasonable in forward!
    - Conclusion: E<sub>1</sub>CB wins!
- **Elimination unimolecular conjugate base:** Just a type of E1 that happens with an acidic proton. Also known as **E<sub>1</sub>CB**.
  - You draw the formation of a conjugate base (i.e., the conjugate base of the SM “acid”) followed by the elimination of something.
  - That wraps it up for microscopic reversibility; let's now move onto another principle.
- **Reactivity-selectivity principle:** It is often observed that a more reactive reactant, intermediate, or reagent corresponds to a less selective reaction.
  - When we say “more reactive,” we typically mean higher energy, more exothermic, etc.
  - This happens because the transition states to different products tend to resemble this higher energy intermediate per the **Hammond postulate**.
  - It follows since the transition state does not resemble the products that it is less sensitive to differences in product energy, so it is harder for the transition state to differentiate between products, so the reaction is less selective.

- **Hammond postulate:** The transition state is most similar in structure to the higher energy intermediate.
- Example of the reactivity-selectivity principle: Radical halogenation.

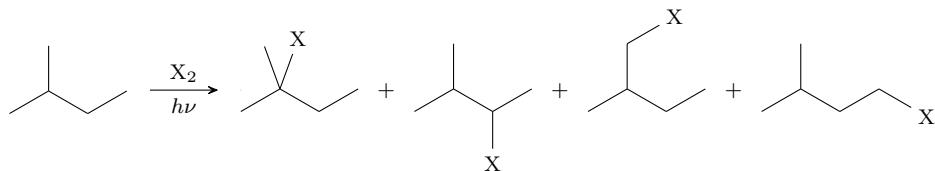


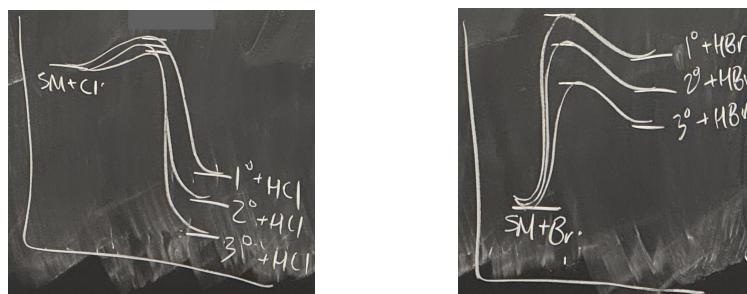
Figure 6.11: Reactivity-selectivity principle in radical halogenation.

- This reaction yields 1 tertiary product, 2 different secondary products, and 1 primary product.
- The reaction in Figure 6.11 forms different product distributions with different halogens.

| $\mathbf{X = Cl \ (\%)}$ | 28 | 35 | 24  | 12  |
|--------------------------|----|----|-----|-----|
| $\mathbf{X = Br \ (\%)}$ | 90 | 9  | < 1 | < 1 |

Table 6.1: Product distribution in radical bromination vs. chlorination.

- Evidently,  $\text{Br}\cdot$  is more selective than  $\text{Cl}\cdot$ .
- Why? Consider BDEs in the selectivity-determining propagation step wherein a halide radical creates an alkyl radical and HX.



(a) Chlorination energy diagram.

(b) Bromination energy diagram.

Figure 6.12: The Hammond postulate explains the reactivity-selectivity principle.

- In radical chlorination: C–H has a BDE of 98 kcal/mol and H–Cl has a BDE of 103 kcal/mol.
  - Thus, the reaction is exothermic with  $\Delta H = -5 \text{ kcal/mol}$ .
  - Then per the reactivity-selectivity principle, we have a high-energy intermediate. This will lead to three energetically close transition states that unselectively determine the product (Figure 6.12a).
- In radical bromination: C–H has a BDE of 98 kcal/mol and H–Br has a BDE of 87 kcal/mol.
  - Thus, the reaction is endothermic with  $\Delta H = 11 \text{ kcal/mol}$ .
  - Then per the reactivity-selectivity principle, we have a low-energy intermediate. This will lead to three energetically distinct transition states that resemble the product more and hence selectively determine it (Figure 6.12b).

- The reactivity-selectivity principle is useful to understand and many-times true, but there are also many exceptions.
  - Example exception: If there are more complicated mechanistic relationships between the SMs and transition states.
  - See Figure 3.4 of Labalme (2024b)!
- Practical aspects of selectivity (will come up on our quals).
  - Numbers worth knowing.
  - We'll go over this in the Lecture 10 recap on Thursday!

## 6.2 Office Hours (Jonathan)

- Would this similarly predict that H<sub>2</sub>O has longer bonds than NH<sub>3</sub>?
  - Perhaps, but other factors make O–H bonds in water shorter than the N–H bonds in ammonia.
- In what way does the HIA only tell us the *relative* stability?
  - The number doesn't tell us anything on its own, and it's not a very useful number.
  - Essentially, all we can learn from these is which cations are more reactive *relative* to other cations.
- How can Bn–Br be the most stable and most reactive species (Table 4.1)?
  - The benzyl *cation* (not the benzyl bromide) is the most stable because it takes the least energy to create it. We had to put more energy into the other two systems to create carbocations, so they are higher energy and hence less stable.
  - The benzyl cation is most reactive toward solvolysis because it has the highest  $k_{\text{rel}}$ .
- Mayr electrophilicity?
  - What I wrote down sounds wrong to Jonathan.
  - It has nothing to do with the thermodynamic stability of anything; it's all about rate constants.
  - I can read the paper if I want, but it's probably not too important.

## 6.3 Linear Free Energy Relationships

10/10:

- Masha's perspective on the Nobel Prize.
  - “Very new, very corporate, very Capitalistic science.”
  - Oleta Johnson: Justice for Bill DeGrado (other *de novo* protein person, along with David Baker).
- Lecture 10 recap.
  - Two types of selectivity: Thermodynamic ( $\Delta G$ ) and kinetic ( $\Delta G^\ddagger$ ).
  - Curtin-Hammett kinetics.
  - Kinetic quench.
  - Principle of microscopic reversibility.
  - Practical aspects of selectivity.
    - If  $\Delta G = 1.4 \text{ kcal/mol}$ , then we get a 10 : 1 ratio at room temperature.
      - This free energy difference can be in the rate ( $\Delta G^\ddagger$ ) or products ( $\Delta G$ ).
    - If there's only one thing you learn in this class, let it be these numbers!!
      - It's a super common qual question.

- Lecture 10 continued: Practical aspects of selectivity.
- The kinetic products are typically favored by short reaction times and low temperature.
  - The thermodynamic products are typically favored by long reaction times and high temperatures.
  - Example: If you want a kinetic enolate vs. a thermodynamic enolate, you'll use different conditions.
- All reactions exist on the spectrum of kinetic control to thermodynamic control.
  - At infinite time, all reactions reach thermodynamic equilibrium.
  - Example: All diamond will eventually convert into graphite because diamond is not the thermodynamically stable form of carbon; it's just kinetically locked.
    - Implication: The “diamonds are forever” jingle is not scientifically true!
- Thermodynamic control.
  - Recall from Gen Chem that  $\Delta G = -RT \ln(K_{\text{eq}})$ .
  - If we plug in the  $K_{\text{eq}}$  for a 10 : 1 ratio (i.e.,  $K_{\text{eq}} = 10$ ), then  $\Delta G = 1.4 \text{ kcal/mol}$  at room temperature.
    - Because of the log scale, if  $K_{\text{eq}} = 100$ , then  $\Delta G = 2.8 \text{ kcal/mol}$ .
    - Implication: Doubling the energy difference doubles the order of magnitude of the selectivity.
  - We rarely think about the energies behind the data we get in the lab. If we get a 10 : 1 selectivity, it feels like that should be because of a big driving force. But it's actually not: It's just a kcal and a half (remember that bond rotation is 3 kcal/mol, for comparison).
- Kinetic control.
  - Recall from Gen Chem that  $\Delta\Delta G^\ddagger = -RT \ln(k_{\text{rel}})$ .
  - Examples of  $k_{\text{rel}}$ : er and dr.
  - To get an ee of 90% (i.e., a 95 : 5 ratio, so er = 19), we only need  $\Delta\Delta G^\ddagger = 1.75 \text{ kcal/mol}$  at room temperature.
  - To get an ee of 99.5% (er = 366), we only need  $\Delta\Delta G^\ddagger = 3.5 \text{ kcal/mol}$  at room temperature.
  - Implication: The energy required for 0-90 ee is the same as for 90-99.5, so it gets progressively harder to get higher ee's.
- Temperature dependence: Lower temperatures mathematically enable higher selectivity, both thermodynamically and kinetically.
  - Example: 1.75 kcal/mol at  $-78^\circ\text{C}$  gives us 98% ee.
  - Example: 1.4 kcal/mol at  $-78^\circ\text{C}$  gives us a 37 : 1 product ratio.
- Rates of completion.
  - A reaction is complete after five half lives (approx 97% yield).
  - A slow reaction (1 day) has a transition state energy of 23 kcal/mol at room temperature.
  - A fast reaction (1 hour) has a transition state energy of 21 kcal/mol at room temperature.
  - Increasing the temperature by  $10^\circ\text{C}$  increases the rate by 2-5 times.
  - Implication: A reaction that finishes in 6 hours at room temperature will finish in 17 minutes at  $50^\circ\text{C}$ .
  - Essentially, high temperatures can put a lot of energy into our system and really accelerate our reactions.
- This concludes the end of last lecture.

- Today: Linear free energy relationships (LFERs).
- Lecture outline.
  - Types of substituent effects.
  - Hammett plots (definition and special cases).
- LFERs are based on **substituent effects**.
- **Substituent effect:** The effect that a new substituent (Y) can have on a reaction rate ( $\Delta G^\ddagger$ ) or equilibrium ( $\Delta G$ ), relative to a reference substituent (X).
- Examples.
  1. **Inductive effects.**
  2. **Field effects.**
  3. **Resonance effects**
  4. **Polarizability effects.**
  5. **Steric effects.**

- **Inductive effect:** The donation or withdrawing of electrons through  $\sigma$ -bonds.
  - Distance dependence: The closer our EWG or EDG is, the bigger effect it has.
- Example of inductive effects' distance dependence.

| Acid            | <chem>CC(=O)O</chem> | <chem>FC(F)C(=O)O</chem> | <chem>CC(F)(F)C(=O)O</chem> | <chem>CC(F)(F)C(F)(F)C(=O)O</chem> |
|-----------------|----------------------|--------------------------|-----------------------------|------------------------------------|
| pK <sub>a</sub> | 4.9                  | 4.2                      | 3.1                         | 0.2                                |

Table 6.2: Inductive effects' distance dependence.

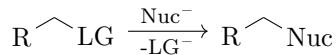
- Let's compare the pK<sub>a</sub>'s of propionic acid, 4,4,4-trifluorobutyric acid, 3,3,3-trifluoropropionic acid, and trifluoroacetic acid.
- The trifluoromethyl EWG stabilizes the anion, resulting in a more acidic proton as the EWG gets closer to the site of deprotonation.
- **Field effect:** The donation or withdrawing of electrons through space.
  - Examples: Dipole moments or charges.
- **Resonance effect:** The donation or withdrawing of electrons through  $\pi$ -bonds.
- Example resonance effect: Deactivating a carbonyl.
  - Consider acetophenone vs. the *para*-methoxy analog.
  - The *para*-methoxy group can donate electron density up through the ring and into the carbonyl, making the carbonyl less electrophilic. We can visualize this donation with resonance structures.
  - This is a further example of the resonance saturation effect.
- **Polarizability effect:** The ability of a substituent to distort an electron cloud.
  - An atom's electron cloud can be **hard** or **soft**.
- **Hard** (atom): An atom that is not polarizable; its electron cloud is difficult to distort.
  - Example: Oxygen.
- **Soft** (atom): An atom that is polarizable; its electron cloud is easy to distort.
  - Example: Sulfur.

- **Steric effect:** The ability of a large group to “deflect” reactants.
- Example steric effect: Changing the rate of an S<sub>N</sub>2 reaction.

|                        |   |                 |                        |
|------------------------|---|-----------------|------------------------|
| <b>R</b>               | H | CH <sub>3</sub> | <sup>t</sup> Bu        |
| <b>k<sub>rel</sub></b> | 1 | 0.33            | 3.3 × 10 <sup>-7</sup> |

Table 6.3: Steric effects on S<sub>N</sub>2.

- Imagine you’re trying to run the following S<sub>N</sub>2 reaction.



- Table 6.3 tells us what happens as we change the R group.
  - In particular, *k<sub>rel</sub>* changes dramatically for bigger groups!
- This concludes our discussion of substituent effects.
  - However, there is still one more major factor that can affect free energy: The solvent.
- **Solvent effect:** The effect on the reaction of changing the solvent.
  - This is *not* a substituent effect, but it can amplify them.
  - You see this a lot, especially in conjunction with field effects and charge.
- What do substituent effects tell us?
  - Identical substituents tend to have similar effects across different reactions and substrates.
    - Examples.
      - NO<sub>2</sub> is almost always an EWG.
      - NR<sub>2</sub> is almost always an EDG.
    - This may be intuitive to us at this point, but it’s not necessarily a given! It’s a blessing that chemistry works out this way.
    - Today, we will discuss a method of quantitatively showing that substituents engender similar effects across reactions and substrates.
  - Substituent effects can tell us a lot about the mechanism and transition states of a reaction.
    - We get mechanistic and transition state information from quantifying how much a substituent “matters,” which we will do with LFERs!
- Let’s now talk about LFERs and the tool through which we visualize them, called a **Hammett plot**.
- Hammett’s program: What did Hammett want to do, and how did he do it?
  - Hammett wanted to study the electronic effects that substituents have on chemical reactions.
  - Initial observation: Substituents thermodynamically favor products with charges that they can help stabilize, and kinetically favor transition states with charges that they can help stabilize.
  - Hammett’s plan: Let’s find a reaction with a product that should obviously be stabilized or destabilized by EWGs and EDGs, let’s vary the EWGs and EDGs on the substrate, and let’s measure the variability in the extent to which the reaction proceeds!
    - The relationship he found happened to be log-linear (hence *linear* free energy relationships), and therefore ended up being very useful.
  - After measuring how each EWG or EDG affected this “reference” reaction, he had a numerical scale on which he could measure EWG/EDG effects on other reactions relative to this reference.
    - Note: Like any relative numerical scale, the origin must be defined arbitrarily. Hammett chose the substituent H as his zero.
  - With this framework, people could measure substituent effects, plot them, and interpret them!

- **Linear free energy relationship:** A correlation of free energy ( $\Delta G$  or  $\Delta G^\ddagger$ ) to parameters that describe substituent effects. *Also known as LFER.*
  - To reiterate: LFERs quantify the effect of substituents on equilibrium or rate.
- Key aspects of LFERs.
  - EDGs accelerate reactions with positive charge buildup in the transition state, and EWGs accelerate reactions with negative charge buildup in the transition state.
    - This is because if you're building up a charge on the transition state, it's more stabilizing to delocalize that charge across the molecule.
  - Later this lecture, we will look at cases where changing the substituents can change the mechanism. In these cases, Hammett plots give us great insight into reaction mechanism!
- One tool in particular helps us study and visualize LFERs: **Hammett plots**.
- **Hammett plot:** A plot of  $\Delta G$  or  $\Delta G^\ddagger$  for a reaction as a function of a **substituent parameter**.
- **Substituent parameter:** A measure of a substituent's ability to stabilize a negative charge. *Denoted by  $\sigma_X$ . Given by*

$$\sigma_X := \log\left(\frac{K_X}{K_H}\right)$$
  - $K_H$  is the equilibrium constant for some reference reaction that creates a negative charge, where the starting material is unsubstituted.
  - $K_X$  is the equilibrium constant for some reference reaction that creates a negative charge, where the starting material has substituent X attached.
  - Higher values of  $\sigma_X$  indicate a stronger ability to stabilize a negative charge.
  - Negative values of  $\sigma_X$  indicate an ability to *destabilize* a negative charge.
    - Alternatively, negative values of  $\sigma_X$  indicate an ability to stabilize a positive charge!
- Hammett quantified how various substituents stabilize the negative charge in benzoate, choosing Figure 6.13 as his reference reaction.

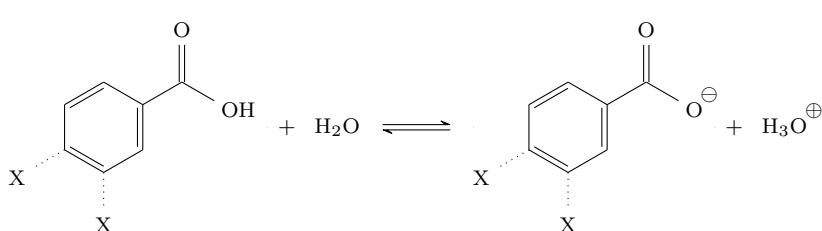


Figure 6.13: Hammett's reference reaction.

- In particular, he looked at the deprotonation of benzoic acid ( $X = H$ ) as a reference reaction, calling its equilibrium constant  $K_H$ .
- Then he looked at the deprotonation of substituted benzoic acids, calling their equilibrium constants  $K_X$ .
- He defined  $\sigma_m$  to measure the substituent's effect when *meta*-positioned on benzoic acid, and  $\sigma_p$  to measure the substituent's effect when *para*-positioned on benzoic acid.

- $\sigma_m$ : A measure of a substituent's ability to stabilize (inductively) the negative charge that builds up when benzoic acid is deprotonated. *Given by*

$$\sigma_m := pK_a(H) - pK_a(X)$$

- Note that the above definition equals  $\log(K_X/K_H)$ , where the equilibrium constants are  $K_a$ 's!
- To measure  $\sigma_m$ , the substituent is *meta*-substituted onto benzoic acid.
  - This way, it *cannot* resonance-delocalize to the *ipso*-position.
- $\sigma_p$ : A measure of a substituent's ability to stabilize (inductively *and* through resonance) the negative charge that builds up when benzoic acid is deprotonated. *Given by*

$$\sigma_p := pK_a(H) - pK_a(X)$$

- Note that the above definition equals  $\log(K_X/K_H)$ , where the equilibrium constants are  $K_a$ 's!
- To measure  $\sigma_p$ , the substituent is *para*-substituted onto benzoic acid.
  - This way, it *can* resonance-delocalize to the *ipso*-position.
- Note that we don't use  $\sigma_o$  (i.e., for *ortho*-substituted substituents) because it incorporates steric effects that are hard to decouple.
  - We will discuss methods of quantifying steric effects next lecture!
- Relating the definitions of substituent parameters to LFERs.
  - The change in free energy  $\Delta G$  of the deprotonation reaction is related to the equilibrium constant  $K_a$ , which is related to  $pK_a$ .
  - Thus, to measure  $\Delta G$ , we can measure the  $pK_a$ !
- Recap: Why benzoate is a great proxy for measuring a substituent's electronic effects.
  - *meta*- and *para*-positioning decouples the substituent's steric effects from its electronic effects.
  - Benzoic acid is aromatic and conjugated, so even though the *para*-position is farther away from the reactive site, there is a minimal difference in distance dependence between the *meta*- and *para*-positions to interfere with comparing inductive effects.
  - Substituted benzoic acids are readily accessible synthetically.
- Example:  $\sigma_p$  and  $\sigma_m$  values for some common substituents.

| X                 | $pK_a$ | $\sigma_p$ | $\sigma_m$ |
|-------------------|--------|------------|------------|
| CH <sub>3</sub> O | 4.5    | -0.27      | 0.10       |
| CH <sub>3</sub>   | 4.3    | -0.14      | -0.06      |
| H                 | 4.2    | 0          | 0          |
| Cl                | 4.0    | 0.24       | 0.37       |
| NO <sub>2</sub>   | 3.4    | 0.81       | 0.71       |

Table 6.4:  $\sigma_p$  and  $\sigma_m$  for common substituents.

- We first measure the  $pK_a$ 's of the *para*-X substituted benzoic acids (see Figure 6.13).
  - These numbers are reported in the  $pK_a$  column in Table 6.4.
  - Plugging them into the definition of  $\sigma_p$  yields the  $\sigma_p$  column in Table 6.4.
  - Examples:

$$\sigma_p(\text{NO}_2) = 4.2 - 3.4 \approx 0.81$$

$$\sigma_p(\text{CH}_3) = 4.2 - 4.3 \approx -0.14$$

- A similar process allows us to measure  $\sigma_m$ .
- Recap: Intuitively interpreting the values of these substituent parameters.
  - When  $\sigma = -$ , we have an EDG (which makes our substrate less acidic than when X = H).
  - When  $\sigma = +$ , we have an EWG (which makes our substrate more acidic than when X = H).
  - When  $\sigma_p$  and  $\sigma_m$  differ, the group is inductively an EWG but by resonance an EDG.
  - Example:  $\text{CH}_3\text{O}$  has  $\sigma_p = -$  (resonance EDG), but  $\sigma_m = +$  (inductive EWG).
- Now that we've established our substrate parameters, let's use them to learn something about the following reaction.

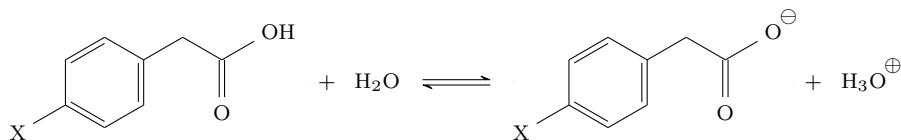


Figure 6.14: The deprotonation of phenylacetic acid.

- In particular, we'll use them to build our first actual Hammett plot.

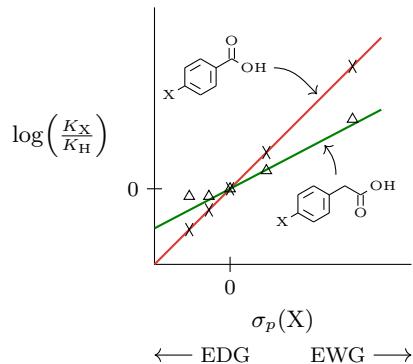


Figure 6.15: Hammett plot for benzoic and phenylacetic acid.

- Measure  $K_X$  for the deprotonation of *para*-substituted phenylacetic acid (see Figure 6.14), as X varies over the substituents in Table 6.4.
- Use these values to calculate a corresponding set of  $\log(K_X/K_H)$  values.
- Plot these values against the  $\sigma_p$  values in Table 6.4 as the triangles in Figure 6.15.
- Perform a regression to fit this data to the appropriate general Hammett equation.

$$\log\left(\frac{K_X}{K_H}\right) = \rho\sigma_X \quad \log\left(\frac{k_X}{k_H}\right) = \rho\sigma_X$$

- We use the left equation above in the case of  $\Delta G$  (e.g., this case).
- We use the right equation above in the case of  $\Delta G^\ddagger$ .
- Recall that  $\sigma_X$  is the substituent parameter.
- $\rho$  is the **sensitivity factor**.
- Performing this analysis, we can determine that  $\rho = 0.56$  for phenylacetic acid (green line in Figure 6.15).
  - Naturally,  $\rho = 1$  for the reference reaction (red line in Figure 6.15).
  - This means that the reaction in Figure 6.14 is about half as sensitive to substituent effects as the reference reaction (Figure 6.13), which makes sense because the carboxylic acid is no longer conjugated to the substituent-bearing aromatic ring.

- **Sensitivity factor:** A measure of how sensitive a chemical reaction is to changes in substituents. Denoted by  $\rho$ .
- Intuitively interpreting the value of  $\rho$ .
  - $\rho > 0$ : The reaction builds up negative charge in the transition state.
    - Such as the *anion-forming* deprotonations in Figures 6.13 & 6.14!
  - $\rho < 0$ : The reaction builds up positive charge in the transition state.
  - $\rho = 0$ : The reaction is not sensitive to substituents.
  - $|\rho| < 1$ : The reaction is less sensitive to substituents than the reference reaction.
    - Such as the deprotonation in Figure 6.14 where, as mentioned, the reactive site is farther from the substituent.
  - $|\rho| > 1$ : The reaction is more sensitive to substituents than the reference reaction.
- Steven: Is the axis labeling correct?
  - Math is definitely not Masha's strong suit.
- To build our intuition and ability to connect  $\rho$  values to mechanistic insights, let's look at some examples.

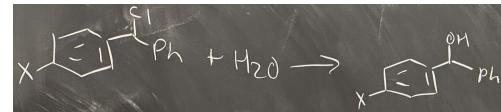
(a) Phenol deprotonation ( $\rho = 2.26$ ).(b) Benzylamine protonation ( $\rho = -1.05$ ).(c) Phenolate attack ( $\rho = -0.95$ ).(d) Nucleophilic substitution ( $\rho = -5.09$ ).

Figure 6.16: Sensitivity factors for simple reactions.

- The deprotonation of *para*-substituted phenols (Figure 6.16a).
  - $\rho = 2.26$ .
  - This means that we're building up negative charge in the transition state, and the reaction is more sensitive to X than the reference reaction.
- The protonation of *para*-substituted benzylamines (Figure 6.16b).
  - $\rho = -1.05$ .
  - This means that we're building up positive charge in the transition state.
- The ring-opening backside attack of *para*-substituted phenolates on epoxides (Figure 6.16c).
  - $\rho = -0.95$ .
  - There's no discrete build up of positive charge in this reaction, but this value indicates that we have a loss of negative charge in the transition state.
- A nucleophilic substitution (Figure 6.16d).
  - $\rho = -5.09$ .
  - Since  $\rho = -$ , we're building up positive charge in the transition state.
  - Since  $|\rho| > 1$ , we're (significantly) more sensitive to substituents than the reference.

- These two facts can actually help us determine the mechanism of this reaction!
  - There are two possible mechanisms by which this reaction can proceed:  $S_N1$  and  $S_N2$ .
  - The RDS of  $S_N1$  is the departure of the leaving group, and  $S_N2$  is concerted. Importantly, this means that  $S_N1$  mechanisms have a significantly greater buildup of positive charge in the “transition state” since they form a true carbocation.
  - So since substituents have a *significant* effect here, the mechanism of this particular nucleophilic substitution must be  $S_N1$ !
  - If it were  $S_N2$ , we’d expect a small negative  $\rho$ .
- Takeaway: Sometimes Hammett plots give us simple insights, and sometimes they are powerful tools to help us probe reaction mechanisms.
  - Usually, we measure  $\rho$  and try to propose mechanisms that would be consistent with that  $\rho$  value; we do not usually draw the mechanism and guess the  $\rho$ .
- While linear Hammett plots can evidently be very helpful, sometimes we get nonlinear relationships. These can also give us important information.
- Example: Consider the following two-step reaction.

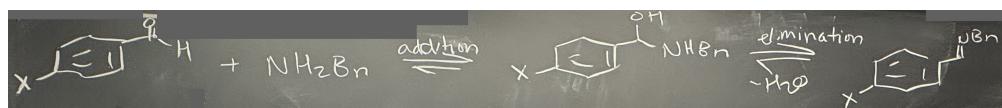


Figure 6.17: Imine formation from a substituted aldehyde.

- This is nucleophilic addition to an aldehyde, forming a hemiaminal, followed by elimination to the imine.
- The Hammett plot for Figure 6.17 is **concave down**.

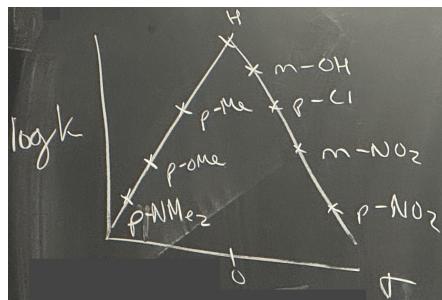


Figure 6.18: Hammett plot for imine formation.

- Notice that this Hammett plot deals with rate ( $\Delta G^\ddagger$ ) because the  $y$ -axis is in  $\log(k)$ , not  $\log(K)$ .
- Hence, this Hammett plot is under the control of two *kinetic* regimes.
- In the left regime, stronger EDGs decrease the rate of reaction.
  - Stronger EDGs will make the initial carbonyl less electrophilic.
  - Thus, with stronger EDGs, addition becomes the rate-limiting step.
- In the right regime, stronger EWGs decrease the rate of reaction.
  - Stronger EWGs will make the initial carbonyl more electrophilic (speeding up addition), and they will destabilize the positive charge that builds up when the hydroxyl group is protonated before elimination.
  - Thus, with stronger EWGs, elimination becomes the rate-limiting step.

- **Concave down** (Hammett plot): A Hammett plot that indicates a change in rate-determining step as X is varied, but the same overall mechanism.

- It should also make intuitive sense that a concave down plot changes the RDS: We have something of an equilibrium at H and all we need is one step slowed down to be the RDS, so pushing one way slows down one step, and pushing the other way slows down the other step!
- Essentially, regardless of which step is accelerated or slowed down by EWGs/EDGs, what matters is that *one* of the steps will be being slowed down, and *that* step will become rate-limiting.

- Further examples of concave down Hammett plots.



Figure 6.19: More concave down Hammett plots.

- Some can have two conjoined downward-sloped lines (Figure 6.19a).
  - This also corresponds to a change in the RDS, but in this case, *both* steps build up positive charge and hence are decelerated by EWGs.
- Some can be curved down (Figure 6.19b).
  - This corresponds to a more gradual change in RDS.
  - We see this when the transition state “moves” with the substituent changes.
- This concludes our discussion of concave down Hammett plots.
- We now look at another example reaction and its Hammett plot.

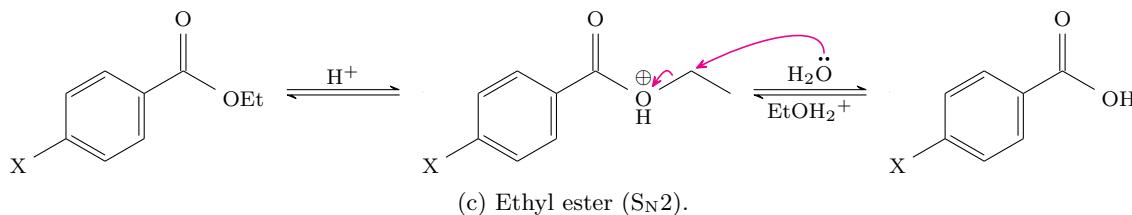
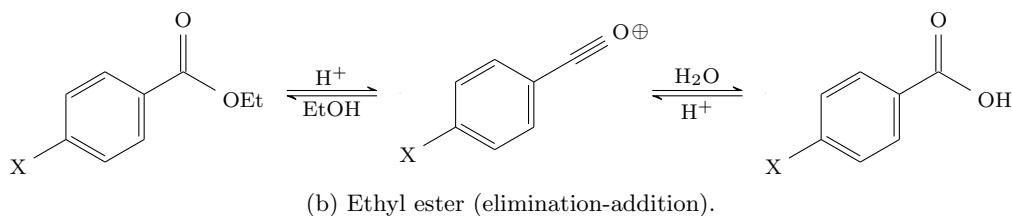
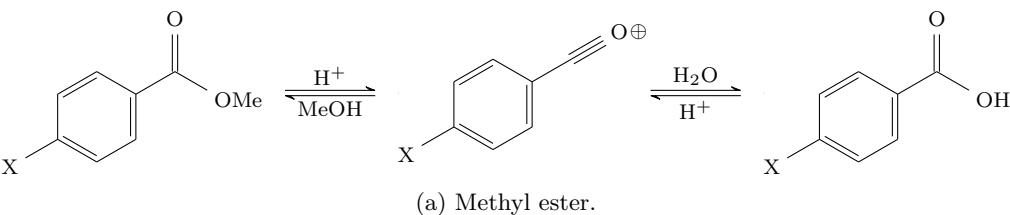


Figure 6.20: Acid-catalyzed ester hydrolysis.

- When a methyl ester hydrolyzes under acidic conditions, there is only one possible mechanism: Protonation of OMe followed elimination of methanol, forming an acylium ion, then addition of water followed by deprotonation to the acid (Figure 6.20a).
  - We call this an “elimination-addition mechanism.”
- However, when an *ethyl* ester hydrolyzes under acidic conditions, it can follow one of two mechanisms.
  1. An analogous elimination-addition mechanism (Figure 6.20b).
  2. Protonation of the ester oxygen followed by an S<sub>N</sub>2-type mechanism (Figure 6.20c).
- The Hammett plot for the hydrolysis of a methyl ester is linear - down (Figure 6.21a), but the Hammett plot for the hydrolysis of an ethyl ester is **concave up** (Figure 6.21b).

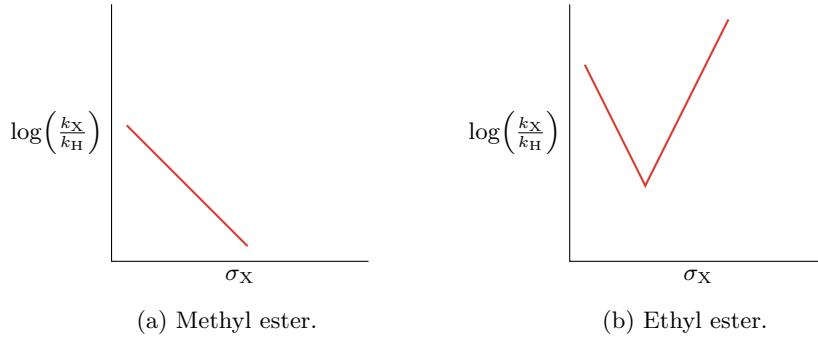


Figure 6.21: Hammett plots for ester hydrolysis.

- The hydrolysis of a methyl ester displays a constant, negative sensitivity factor (Figure 6.21a).
  - This is because the positively charged acylium ion intermediate gets destabilized by stronger EWGs.
- The hydrolysis of an ethyl ester displays a negative sensitivity factor for EDGs, and a positive sensitivity factor for EWGs (Figure 6.21b).
  - When X is an EDG, the acylium ion get stabilized. Weaker EDGs stabilize it less ( $\rho = -$ ), but we still favor the elimination-addition mechanism (Figure 6.20b).
  - As X becomes electron withdrawing, the mechanism in Figure 6.20c becomes operative. This corresponds to a positive Hammett slope.
- **Concave up** (Hammett plot): A Hammett plot that indicates a change in mechanism.
- Why are neither of these mechanisms addition-elimination?
  - That’s better under basic conditions!
- Take-home message: Any deviation from linearity in a Hammett plot indicates a change in the RDS or the mechanism.
- Hammett plots are a very powerful mechanistic tool; think about using them in your final mechanistic proposal!!

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