5.13 (Organic Chemistry II) Notes

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Unit 1

Structure Determination

1.1 Intro + Elemental Analysis

9/4:

- Teaching team.
  - Prof. Masha Elkin.
  - Prof. Steve Buchwald.
  - 8 Teaching Fellows (TFs).
- Masha Elkin begins. Steve Buchwald and all TFs introduce themselves. Special roles:
  - Head TF: Minh Le.
  - Electronic TF (contact with questions on Canvas, Piazza, BACON): Angel Garcia-Ramirez.
- In this class, you will learn...
  - New things in organic chemistry;
  - Old things at a deeper level;
  - Real-world applications of chemistry.
- Why study organic chemistry?
  - Chemists manipulate matter, and that's awesome!
  - By "manipulate matter," we mean making molecules, breaking molecules, making polymers, making detergents, and making sure that all of these things break down in the environment :)
- Core questions.
  - *How* do we make molecules?
  - What molecules *should* we make?
- Course logistics.
  - Seven (7) units total (2 big units before the halfway mark & 5 smaller units after).
  - The units.
    - Unit 1: How do we know what molecule(s) we have?
    - Unit 2: How do electrons move?
    - Units 3-7: How do we make molecules? How do reactions work?
  - Exams after units 1, 2, 4, and 6; final exam after unit 7.
  - Questions? Ask your TF first, then the Head TF, then the profs (Masha & Steve).

- Prerequisites.
  - Official prerequisites: 5.12 (equivalent to Orgo I, in case you took it elsewhere) & Gen Chem.
  - Recommended reading for review: Chapters 1-2 of the main textbook, referred to in these notes as Clayden et al. (2012).
- Grading.
  - Your grade will (hopefully) be a reflection of your learning.
  - There are no curves in this class or at MIT, so *everyone can get an A!!!*
  - How to improve your grade: Do problems!
    - Problem sets (PSets) and recitation worksheets will be provided.
    - You may also do as many textbook problems as you want. Feel free to buy the solutions manual, or borrow a copy from the ChemEd office[1] to check your answers.
- How to learn organic chemistry.
  - Analogy: Learning Orgo is like learning a language.
    - Basic vocab and grammar that must be memorized. Examples: Drawing structures, curved arrow formalism, etc.
    - Recognizing patterns and trends. Examples: Nucleophiles tend to have lone pairs (or be other regions of high electron density).
    - Developing intuition.
    - Practice, practice, practice! (Focus on drawing structures.)
  - Tips for success.
    - Be active and participate in lecture, recitation, etc. Take notes while you're here!
    - Practice **metacognition**, i.e., learn how you learn.
      - Do you learn best in a crowded coffee shop, or in your own room? Would you rather recopy your notes, or read the textbook?
      - Note that what works for somebody else may not work for you, and vice versa!
      - Invest the time and effort that *you* need to succeed. This may be more (or less) than other students, and that's ok!
    - Communicate with *the whole* teaching team. They're here to help!!!
      - Seek out accommodations as needed: It's the student's responsibility to ask.
- **Metacognition**: Being aware of your own understanding.
- We now begin the content for Unit 1.
- Goal: Learn how to determine the chemical structure of a given organic compound.
- Why do we need to determine structures?

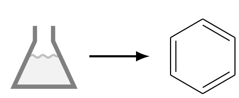


Figure 1.1: Why we study structure determination.

  - With the naked eye, organic chemists see a flask with a colorless liquid. But we draw the skeletal diagram for benzene (which is a colorless liquid). What tools enable us to convert from the flask to the structure?

[1] Located in 6-203.

- Here's another reason: Suppose we run a brand new chemical reaction. Organic chemists do this all the time in research! How do we now what the product is? How do we know which atoms it contains, and in what arrangement?

- Structure determination workflow.
  1. Identify the atoms present.
     - Questions to answer: What is the molecular formula?
     - Relevant tools: Elemental analysis (EA) and mass spectrometry ("mass spec" or MS).
  2. Identify the functional groups and substructures present.
     - Questions to answer: Do we have ketones? Esters? Alcohols? Rings?
     - Relevant tools: MS, infrared spectroscopy (IR), and nuclear magnetic resonance (NMR).[2]
  3. Identify how all the functional groups fit together.
     - Questions to answer: Are they close? Far apart? Ortho/meta/para? What stereochemistry?
     - Relevant tools: NMR and X-ray diffraction.
- We now begin talking about EA.
  - History: Began development in the 1820s.
  - Purpose: Determine which elements are present, and in what quantities (in a given sample).
- In this course, we will apply EA to compounds containing carbon, hydrogen, and oxygen *exclusively*.
  - To reiterate, in an EA problem for this course, we will *not* have to worry about any other elements.
  - The typical EA technique for such compounds is **combustion analysis**.
- **Combustion analysis**: Burn the sample and measure the products.
  - All C in the sample becomes CO_2.
  - All H in the sample becomes H_2O.
  - O is then determined via process of elimination, explained as follows.
- Advanced techniques (beyond the scope of this class): Nitrogen to NO or NO_2, sulfur to SO_2, etc.
- A schematic of combustion analysis.

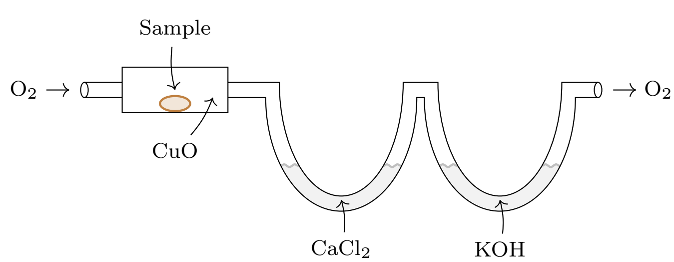


Figure 1.2: Combustion analysis schematic.

  - Burn the sample in the presence of an oxidant such as cupric oxide (CuO).
  - Flow O_2 into the combustion chamber to facilitate burning as well.
  - The combusted gas then flows through a series of reaction containers.
    - The first one contains a desiccant (like $CaCl_2$) that absorbs the water.

[2] NMR is an organic chemist's best friend!

  - The second one contains a base (like KOH) that absorbs the CO_2.
- The remaining oxygen flows out the end.

- The *analysis* part of combustion analysis.
  - The amount of H is equal to the change in mass of the $CaCl_2$.

$$\Delta\text{mass}(CaCl_2) = \text{mass}(H_2O) \to \text{ratio(H)}$$

  - The amount of C is equal to the change in mass of the KOH.

$$\Delta\text{mass}(KOH) = \text{mass}(CO_2) \to \text{ratio(C)}$$

  - The amount of O is equal to the change in mass of the sample.

$$\text{mass(sample)} - \text{mass(H)} - \text{mass(C)} = \text{mass(O)} \to \text{ratio(O)}$$

  - Result: We get an **empirical formula** of the form $C_xH_yO_z$. Remember that this is *not* (necessarily) the **molecular formula**; it is *only* a ratio of elements.

- EA example: Let's burn 0.5 g of propanol (C_3H_8O).
  - Suppose we obtain 0.600 g H_2O and 1.09 g CO_2.
  - This means that there was 0.067 g (H) and 0.300 g (C) in the sample. The remaining 0.133 g must then be due to O.
  - Therefore, the elements exist in a 3:8:1 (C:H:O) ratio.
  - Bonus: Convert the masses to a ratio via stoichiometry.
    - $0.600\,\text{g } H_2O \times \frac{1\,\text{mol } H_2O}{18.02\,\text{g } H_2O} \times \frac{2\,\text{mol H}}{1\,\text{mol } H_2O} \times \frac{1.01\,\text{g H}}{1\,\text{mol H}} = 0.067\,\text{g (H)}$
    - $1.09\,\text{g } CO_2 \times \frac{1\,\text{mol } CO_2}{44.01\,\text{g } CO_2} \times \frac{1\,\text{mol C}}{1\,\text{mol } CO_2} \times \frac{12.01\,\text{g C}}{1\,\text{mol C}} = 0.300\,\text{g (C)}$
    - $0.5\,\text{g propanol} - 0.067\,\text{g (H)} - 0.300\,\text{g (C)} = 0.133\,\text{g (O)}$

- A note on the previous example.

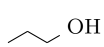
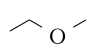
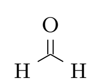
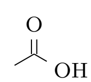
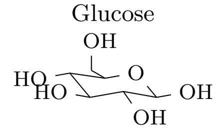
Name	Propanol	Methyl ethyl ether	Formaldehyde	Acetic acid	Glucose
Structure					
Emp. formula	C_3H_8O	C_3H_8O	CH_2O	CH_2O	CH_2O
Mol. formula	C_3H_8O	C_3H_8O	CH_2O	$C_2H_4O_2$	$C_6H_{12}O_6$

Table 1.1: Questions that EA can't answer.

  - EA has given us the empirical formula, but it has *not* confirmed that the sample is propanol. For example, methyl ethyl ether has the same empirical formula!
  - Additionally, we don't yet have the molecular formula. Consider, for instance, the breadth of compounds with empirical formula CH_2O!
  - Takeaway: EA gives you the empirical formula; we need MS to get the molecular formula (we'll see this on Friday), and we may need even more to get the atomic connectivity.

- Application of EA to real-world chemistry.
  - A home furnace burns natural gas — which is mostly methane (CH_4) — for heat.

- **Ideal combustion**[3] corresponds to the reaction

$$\mathrm{CH_4} + \mathrm{O_2} \longrightarrow \mathrm{CO_2} + \mathrm{H_2O}$$

- Real-world combustion is incomplete; you make

$$\mathrm{CH_4} + \text{air} \longrightarrow \mathrm{CO_2} + \mathrm{H_2O} + \mathrm{NO_2} + \mathrm{CO}$$

- When a technician comes to your home, they analyze the flue gas (i.e., your furnace exhaust).
  - Their analysis could determine that our combustion has too much O_2, which is called "air rich." This is inefficient and doesn't yield enough heat.
  - They could also determine that you have too much CO_2 and CO, which is called "fuel rich." This yields too much soot and CO. CO can be dangerous and lead to carbon monoxide poisoning, which makes you sleepy before it kills you.
- To measure this flue gas, though, they have a little handheld elemental analysis device!

- Note that there is a relation between ideal/real-world combustion and the CuO oxidant in Figure 1.2: The CuO ensures that when we combust our EA sample, all the carbon is fully oxidized to CO_2! Without it, some CO would be formed, and our stoichiometry would be thrown off.

1.2 Mass Spectrometry

9/6: • Lecture 1 recap.

- Elemental analysis (EA).

$$\mathrm{SM} + \mathrm{O_2} \xrightarrow{\Delta} \mathrm{CO_2} + \mathrm{H_2O}$$

  - SM means "starting material."
  - SM's we will focus on: Compounds of the form $C_xH_yO_z$.
- Empirical formula vs. molecular formula (see Table 1.1).

- Today: Mass spectrometry (MS).
  - Purpose: Convert empirical formulas to molecular formulas (and more!).
  - Reading: Clayden et al. (2012), Chapter 3.

- Lecture outline.
  - Mass spectrometer schematic.
  - Mass spectrum elements.
  - Fragmentation, and common types.
  - Isotope effects in MS.
  - Ionization methods.

- **Mass spectrometry**: A structure determination technique that tells us the exact mass of molecules and their "fragments." *Also known as* **MS**, "**mass spec**."

- Overview.

$$[\mathrm{M}] \xrightarrow{e^- \quad 2e^-} [\mathrm{M}]^{+\cdot} \begin{array}{l} \nearrow\ [\mathrm{M}{-}a]^+ \ + \ [a]^\cdot \\ \longrightarrow\ [\mathrm{M}{-}b]^+ \ + \ [b]^\cdot \\ \searrow\ [\mathrm{M}{-}c]^+ \ + \ [c]^\cdot \end{array}$$

[3]You can learn more about in a chemical engineering/ChemE course.

- You have a sample — denoted by [M] — that you bombard with electrons (e^-). When an electron hits a molecule of your sample, it knocks off one of the molecule's electrons (and flies off itself). This ionizes your molecule to a **radical cation**, denoted by $[\mathrm{M}]^{+\cdot}$ and called the **molecular ion**.
- This radical cation is unstable and fragments into a proper cation and a proper radical. The radical is usually not detected, but any cationic fragment produced — the $[\mathrm{M}-a]^+$, $[\mathrm{M}-b]^+$, and $[\mathrm{M}-c]^+$ above — usually *is* detected.

- A (stepwise) schematic of a mass spectrometer.

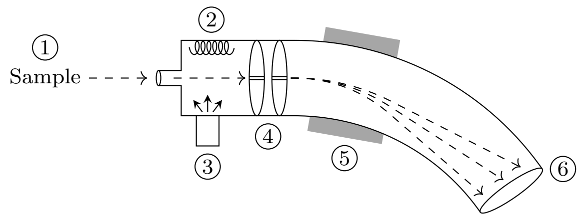
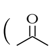


Figure 1.3: Mass spectrometer schematic.

1. The sample is injected into a curved tube.
2. A heater vaporizes the sample.
3. An electron source (also known as an electron gun) shoots electrons at the vaporized sample, ionizing it. The ionized sample starts fragmenting.
4. The fragments encounter a series of negatively charged plates with slits in the middle. These negatively charged plates accelerate the positively charged cations.
5. A magnet deflects the accelerated, positively charged ions. The magnet deflects them based on their **mass-to-charge ratio**. Because of physics, the lightest ions are deflected the most, and the heaviest ions are deflected the least.
6. A detector records where the ions hit. This data is converted into a mass-to-charge ratio for each ion. This yields a spectrum of all the fragments' masses.

- **Mass-to-charge ratio** (of a cation): The cation's mass divided by its net charge. *Denoted by* $\boldsymbol{m/z}$.
  - For the purposes of this class, $z = 1$.
- Example mass spectrum: Acetone ().

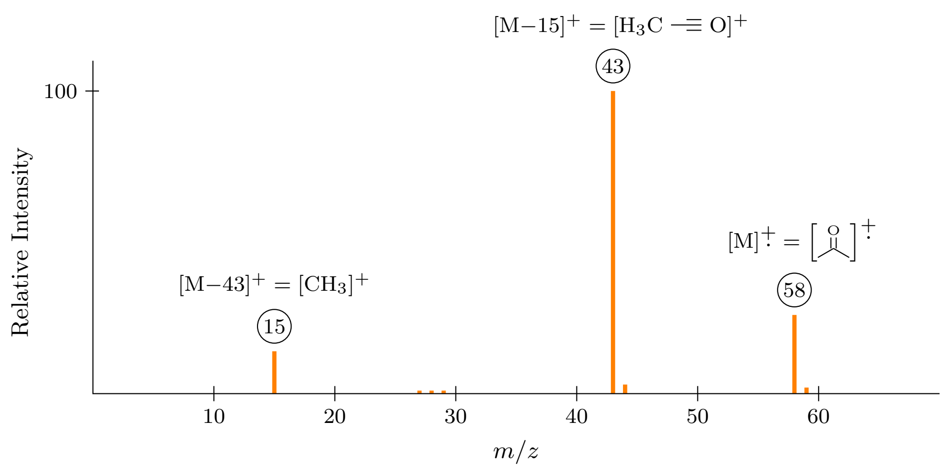


Figure 1.4: Mass spectrum of acetone.

- The x-axis is the mass-to-charge ratio, and the y-axis is the "relative intensity" of each peak.
  - If a certain fragment gets produced more than another (and hence recorded more than it), we say it has a "higher relative intensity."
- We identify two special types of peaks in a mass spectrum: The **parent peak** and the **base peak**. In the case of acetone...
  - The parent peak lies at 58;
  - The base peak lies at 43.
- The peak at 15 also has a relatively large magnitude, and from the fact that the mass of a methyl cation is approximately 15, we can infer that this peak corresponds to the methyl cation fragment.
  - Notice that its intensity is significantly lower than the intensity of the base peak because we may recall from Orgo I that the methyl cation is a far less stable cation than the resonance-stabilized, secondary acylium ion at 43.
- There are a number of smaller peaks, too, but they give less information.
- Note that the major peaks may be appropriately referred to by *any* of the three nomenclature methods in Figure 1.4: By exact mass, by $[\mathrm{M}-a]^+$, and/or by structure.

• **Parent peak**: The peak in a mass spectrum corresponding to the molecular ion.
  - The parent peak is always the rightmost peak in the spectrum.[4] This is because it is created by the heaviest ion, and you can't have more mass than your initial molecule!
  - It is typically *not* the tallest peak in the spectrum.
  - Useful information: It gives the molecular weight of the molecule.

• **Base peak**: The tallest peak in a mass spectrum.
  - The base peak corresponds to the fragment that the molecule forms most preferentially, which is usually also the most stable fragment.

• **Fragmentation peak**: Any peak to the left of the parent peak.

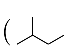
• Maxim: Molecules fragment in predictable ways to form stable cations.

• At this point, let's formally define **fragmentation**.

• **Fragmentation**: The formation of stable(-ish) cations.
  - Recall from Orgo I (review your notes on cation stability!!) that stable cations tend to be more substituted, delocalized, atom-stabilized (e.g., close to a heteroatom), etc.

• Let's now discuss some common species that we analyze via MS — and how they fragment.

• Alkane fragmentation: Preferentially break bonds to get more substituted (e.g., 2° & 3°) carbocations.

• Example: 2-methylbutane ().

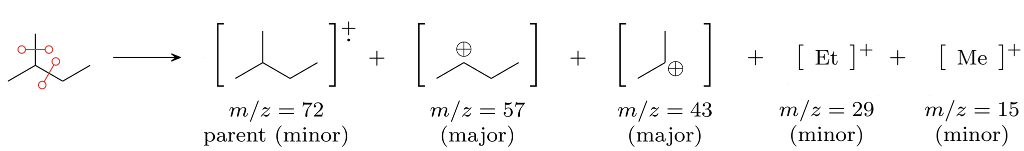
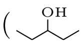
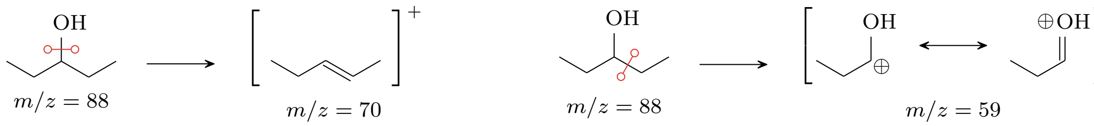


Figure 1.5: Fragmentation of alkanes.

  - All these peaks will appear, but the tallest will correspond to the species labeled "major" above.

[4]Excepting isotope effects; discussed later in this lecture.

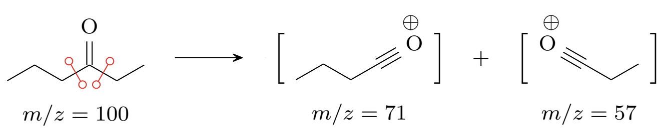
- Alcohol fragmentation.
  - Dehydration: Yields an $[\mathrm{M}-18]^+$ peak, corresponding to the loss of water.
  - α-cleavage: Leads to a resonance-stabilized product.
- Example: Pentan-3-ol ().



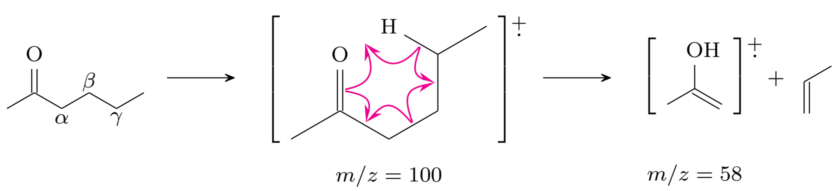
(a) Dehydration. (b) α-cleavage.

Figure 1.6: Fragmentation of alcohols.

- Ketone fragmentation.
  - α-cleavage: Leads to a resonance-stabilized product, once again.
  - McLafferty rearrangement: Only happens for ketones with a γ-proton.
    - We select for this type of ketone because in this case, we can form a six-membered transition state. Recall that six-membered transition states are super stable in organic chemistry!
    - This fragmentation leads to a charged enol (that we see in the spectrum) and an uncharged olefin (that we don't see in the spectrum).
- Example: Hexanones.



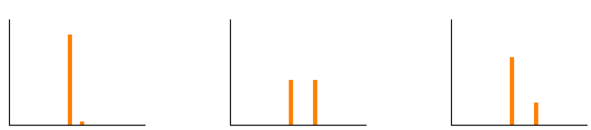
(a) α-cleavage.



(b) McLafferty rearrangement.

Figure 1.7: Fragmentation of ketones.

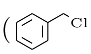
- Isotope effects.
- Principle: Mass spectrometry weighs individual molecules, so molecules containing a heavier (or lighter) isotope will appear separate from other molecules in the mass spectrum.
- Atoms with notable isotope effects.



(a) Carbon. (b) Bromine. (c) Chlorine.

Figure 1.8: Isotope effects in MS.

- Carbon: The $^{12}\mathrm{C} : {}^{13}\mathrm{C}$ ratio is $99 : 1$.
  - Implication: For every $[\mathrm{M}]^{+\cdot}$, we see 1% $[\mathrm{M+1}]^{+\cdot}$.
  - This is why we see tiny "shadow" peaks to the right of the parent peak and base peak in Figure 1.4!
    - Note that the "shadow" of the parent peak is 3% its height (not 1%) because there are *three* carbons in the acetone molecular ion.
    - Similarly, the "shadow" of the base peak is 2% its height because there are *two* carbons in the acylium ion.
- Bromine: The $^{79}\mathrm{Br} : {}^{81}\mathrm{Br}$ ratio is $1 : 1$.
  - Implication: The $[\mathrm{M}]^{+\cdot}$ and $[\mathrm{M+2}]^{+\cdot}$ peaks exist in a $1 : 1$ ratio, i.e., have the same height/relative intensity.
  - The splitting of the molecular ion peak into two such peaks is a super recognizable, distinct, and useful fingerprint of bromine-containing compounds!
- Chlorine: The $^{35}\mathrm{Cl} : {}^{37}\mathrm{Cl}$ ratio is $3 : 1$.
  - Implication: The $[\mathrm{M}]^{+\cdot}$ and $[\mathrm{M+2}]^{+\cdot}$ peaks exist in a $3 : 1$ ratio.
  - Similar to bromine, this peak splitting is a fingerprint of chlorine-containing compounds.

- Combining everything we've learned up to this point, let's do another example.
- Example: Benzyl chloride (

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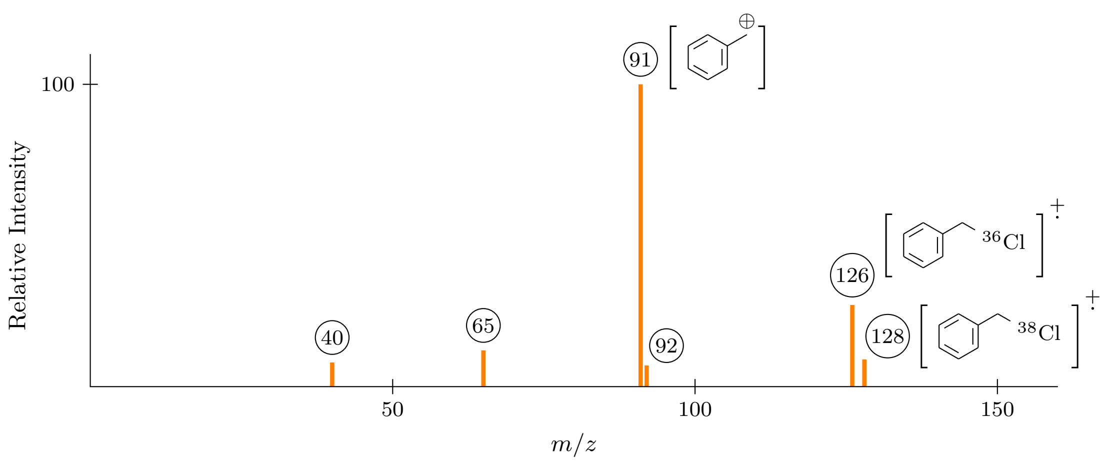


Figure 1.9: Mass spectrum of benzyl chloride.

- The parent peak will lie at 126, and the corresponding chlorine isotope peak will lie at 128 and be one-third the height.
- The base peak will lie at 91, and the corresponding carbon isotope peak will lie at 92 and be 7% the height (to account for the 7 carbons in the benzylic cation that may be heavy).
  - It will correspond to the most stable fragment, which in this case is the benzylic cation.
  - The benzylic cation is super stable because its positive charge can be resonance delocalized to four different atoms!
  - A large peak at $m/z = 91$ strongly suggests the presence of an aromatic system.
- This example focused on predicting the peaks in a mass spectrum based on reasonable fragmentation patterns. But what if we are given the mass spectrum? What data can we pull out then?

- To answer this question, here are some guidelines for the interpretation of mass spectra.

- Guidelines for interpretation.
  - The parent peak provides the molecular weight of the molecule.
    - This allows you to convert an empirical formula obtained from EA to the molecular formula.
  - The parent peak also reveals key atoms via distinct isotopic fingerprints.
    - Examples include bromine and chlorine.
    - An additional one is the **nitrogen rule**.
  - Fragmentation patterns can identify substructures.
    - Recall from Lecture 1 (9/4) that identifying substructures is part of the second step of the structure determination workflow!
    - Common fragments:
      - Loss of a methyl group is -15.
      - Loss of an OH group is -17.
      - Loss of H_2O is -18.
      - Loss of CO_2 is -44.
      - Loss of a ^{t}Bu group is -57.
    - Look at the m/z of the fragments *and* the difference in m/z between certain fragments.
      - Example: Maybe a certain fragment is formed by losing both a methyl group *and* water.
  - Important note: These guidelines are just a guide; we will need multiple forms of evidence to support an assignment.
- **Nitrogen rule**: If you have an odd number of nitrogen in a molecule, you will get an odd molecular weight.
  - The basis for this rule lies in the fact that nitrogen is trivalent but has an even mass.
    - This means that nitrogen tends to bond an odd number of groups (specifically, 3), making the overall mass odd.
  - Examples: Ammonia has an odd mass of $17 = 14 + 1 + 1 + 1$ and methylamine has an odd mass of $31 = (14 + 1 + 1) + (12 + 1 + 1 + 1)$, while methane has an even mass of $16 = 12 + 1 + 1 + 1 + 1$ and ethane has an even mass of $30 = (12 + 1 + 1 + 1) + (12 + 1 + 1 + 1)$.
  - You can read more about the nitrogen rule here.
  - Implication: If you see an odd molecular weight, you *might* have a nitrogen present!
- Types of ionization.
- **Electron ionization**: A beam of electrons. *Denoted by* **EI**. *Also known as* **hard ionization**.
  - This is the method we are using in this class.
- **Electrospray ionization**: Forms charged droplets. *Denoted by* **ESI**. *Also known as* **soft ionization**.
  - ESI causes less fragmentation.
  - One implication of this is that you observe a larger parent peak.
  - Another consequence is that ESI can analyze a broader range of compounds via mass spectrometry than EI can, since some sensitive compounds (like proteins) would never survive an electron beam.
    - Nobel Prize in Chemistry (2002) for this application of MS to biology!
- **High resolution mass spectrometry**. *Denoted by* **HRMS**.
  - In “normal” low-resolution mass spectrometry (LRMS), both N_2 and C_2H_4 have $m/z = 28$.
  - In HRMS, N_2 has $m/z = 28.0061$ and C_2H_4 has $m/z = 28.0314$.

- HRMS leads nicely into our application for today!
- Application of MS to real-world chemistry: Isotopic signatures.
  - Today, you learned that the ^{12}C : ^{13}C ratio is 99 : 1.
    - In reality, this is an *average* value.
    - The actual ratio of isotopes is globally uneven, and we as humans have mapped it.
  - Indeed, isotope abundances vary by time and location due to air patterns, etc.
  - For example, Montana is home to 2% more ^{13}C than Florida!
  - Implication: We can tell if a narcotic is made in the US (and where) or another country based on the isotopic abundance in it.
  - We can also track where a person, drug, or uranium sample is from.
    - Naturally, the government is very interested in this technology :)
  - You can also tell if a person eats corn or rice because this leads to different ratios of nitrogen isotopes in our bodies.

1.3 Infrared Spectroscopy

9/9:
- Lecture 2 recap.
  - In mass spectrometry, you ionize your sample [M] to the molecular ion $[M]^{+\cdot}$.
  - $[M]^{+\cdot}$ is detected as the parent peak.
    - The parent peak provides the molecular weight (MW) of the molecule.
    - The parent peak also reveals any isotopic signatures.
  - Many molecular ions — once formed — will fragment into cations $[M-a]^+$, $[M-b]^+$, $[M-c]^+$, etc.
    - More stable cations are formed more often, resulting in higher relative intensities.
    - The *most* stable fragment gives rise to the base peak.
  - Common fragments include those resulting from...
    - The loss of a methyl group;
    - The loss of a water molecule;
    - α-cleavage;
    - The McLafferty rearrangement (for ketones).
- Today: Infrared Spectroscopy (IR).
  - Reading: Clayden et al. (2012), Chapter 3.
  - Prof. Elkin highly recommends the section on IR; be sure to read this!!
- Lecture outline.
  - IR spectrometer schematic.
  - IR theory.
  - IR spectrum elements.
  - Key regions of an IR spectrum.
- Principle: Irradiate a sample with infrared waves and detect where the sample absorbs these waves.
  - This technique is useful for identifying certain functional groups, namely those that absorb IR waves well.

- A schematic of an infrared spectrometer.

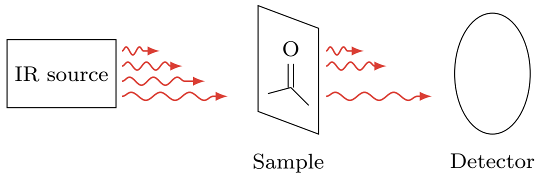
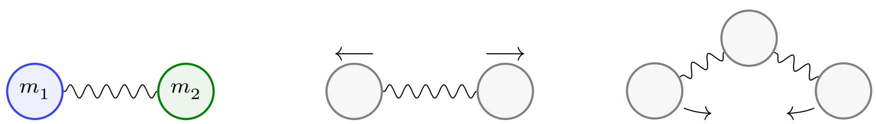


Figure 1.10: Infrared spectrometer schematic.

  - We begin with a source of infrared radiation. This source shoots waves at our sample, which could be a molecule like acetone. The IR waves that the source emits have a range of frequencies.
  - The sample will absorb certain frequencies, and the frequencies that are not absorbed are detected by a detector. In other words, the director detects the **transmittance** of the sample.
- **Transmittance**: How much of each frequency of radiation passes through the sample.
- IR theory.



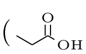
(a) Springy bonds. (b) Stretch. (c) Bend.

Figure 1.11: Infrared spectroscopy theory.

  - Fundamental assumption: A chemical bond is like a spring between atoms.
    - Recall from Gen Chem that in science, we often call a spring a **harmonic oscillator**. If you don't quite remember this term, review your Gen Chem notes or Google it!!
  - Let's dive a bit deeper into this analogy: Imagine we have two different atoms of masses m_1, m_2 joined by a "spring," as in Figure 1.11a.
    - Just like a real spring, chemical bonds can vibrate in different ways: They can stretch and contract (as in Figure 1.11b), bend (as in Figure 1.11c), etc.
    - All of these different motions are called the **vibrational modes** of the chemical bonds.
  - Bonds absorb energy from IR waves when the frequency (ν) of the IR wave matches the frequency of the stretching/bending motion.
    - In other words, when you hit the resonance frequency, you absorb energy.
    - This absorption of energy is detected as the loss of transmittance.
  - The change in energy between vibrational modes is related to characteristics of the bond as follows.

$$\Delta E \approx \sqrt{\frac{k(m_1 + m_2)}{m_1 m_2}}$$

    - k is the force constant (proportional to the bond strength).
    - m is the mass of atom 1 or 2.
    - Implication: Stronger bonds (i.e., those with larger values of k) require more energy (i.e., higher ν IR waves) to absorb.
    - Implication: Lighter atoms (i.e., those with lower values of m) require more energy (i.e., higher ν IR waves) to absorb.
  - One additional requirement: The chemical bond must have a dipole in order to absorb IR waves.
    - Example: C≡O absorbs because O is more electronegative than C, but N≡N does not.

- Questions on IR theory.
  - Why do bonds absorb energy *only* when the frequency of the IR waves matches the frequency of the bond's vibration?
    - The answer to this question is beyond the scope of the class, but Prof. Elkin gives the quantum mechanical explanation.
    - Essentially, when a chemical bond absorbs energy, it gets excited to a higher-energy vibrational mode, which we may think of as a more intense vibration.
    - However, because vibrational modes are separated by a set amount of energy, lower energy photons won't have enough energy to make it to the next vibrational mode while higher energy photons will provide too much energy to reach anything stable.
  - Why don't bonds without dipoles absorb IR waves?
    - The explanation is also quantum mechanical, and hence also beyond the scope of this class.
    - Essentially, symmetric bonds and molecules lack something called a dipole moment, and zero dipole moment zeroes out the absorption in the math of quantum mechanics.
    - Note that there is some really cool math and physics underlying the answer to this question, and Prof. Elkin recommends you look it up if you're interested!!
    - In organic chemistry, however, we're more interested in what we can do with IR spectroscopy than in *exactly* how it works. Essentially, for this class, you should learn how it works well enough to make sense of the trends in spectrum interpretation presented in this lecture, but you don't need to go deeper than that for now.
  - Why do lighter atoms require more energy? It seems like it would take more energy to push around a heavier atom.
    - Check out the explanation in Clayden et al. (2012); it's pretty comprehensive and understandable.
- Example IR spectrum: Propionic acid (
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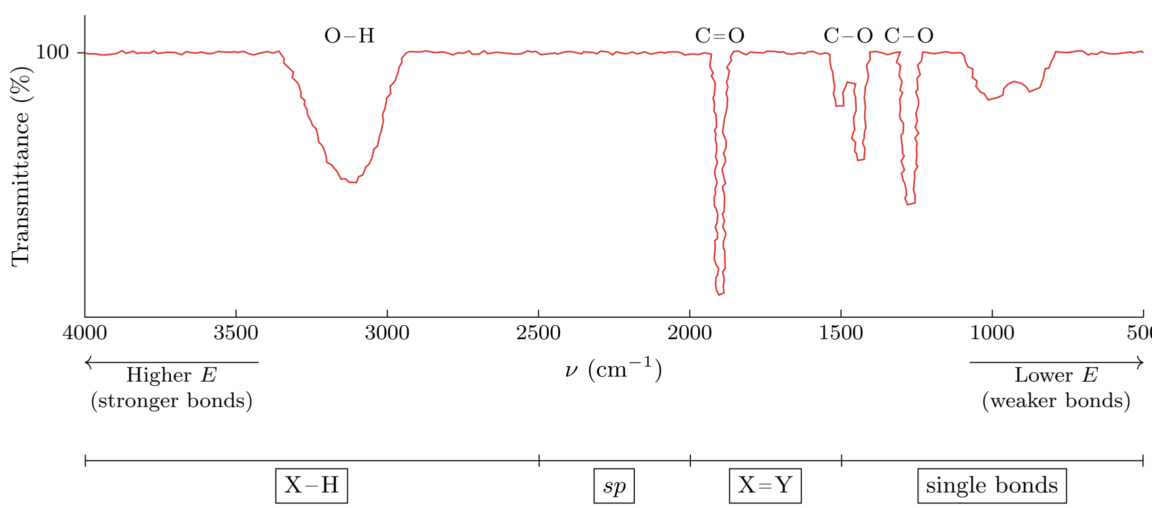


Figure 1.12: Infrared spectrum of propionic acid.

  - The x-axis is the frequency of the IR waves, measured in wavenumbers (cm^{-1}).
    - We typically are interested in the region from 4000–1500 cm^{-1}.
  - The y-axis is the percent transmittance.
  - The **baseline** is 100%, which means pure transmittance (aka, no **absorption**).
    - Then we have **absorbance peaks**, each of which corresponds to a different chemical bond.

- We can further break the spectrum down into regions.
  - X−H bonds occur in the 4000–2500 cm^{-1} region.
    - Peaks in this region are often broad. Specifically, a peak will be broad if the corresponding protons are **exchangeable**.
    - Hydrogen bonding can also lead to broadening.
    - We see this effect in both IR and NMR, so we'll talk about it more later this week!
    - For example, the O−H peak is broad because this acidic proton is exchangeable.
  - *sp*-hybridized atoms occur in the 2500–2000 cm^{-1} region.
    - In other words, polar triple bonds show up here.
    - Examples: C≡N and C≡C′.
  - X=Y bonds occur in the 2000–1500 cm^{-1} region.
    - This is for polar double bonds.
    - Examples: C=O, C=C′,[5] and C=N.
  - Single bonds occur in the 1500–500 cm^{-1} region.
    - Examples: C−C′, C−O, and C−F.
  - Some of these regions are useful, and some less so.
- As you can infer from Figure 1.12, IR spectra look a bit like icicles.
- Note that C−O has two peaks because there are multiple bonding modes per bond.

- **Absorption**: The loss of tranmittance.
  - We typically plot transmittance in a spectrum, but the two measures are inversely proportional.
- **Exchangeable** (proton): A hydrogen atom that is liable to break off of the rest of the molecule and be replaced by another hydrogen atom in solution.
  - This is very much related to acidic protons! Recall that a Brønsted acid will donate its proton and then the conjugate base will pick up a new (possibly new) proton all the time.
- **Diagnostic regions**: 4000–1500 cm^{-1} (useful) and 1500–500 cm^{-1} (useless).
- **Fingerprint region**: The region of an IR spectrum from 1500–500 cm^{-1}.
  - Within the fingerprint region, we have so many overlapping peaks that the spectrum becomes difficult to interpret.
  - However, its shape is characteristic of a molecule, even if it doesn't tell you anything specifically. This is just like a real fingerprint! Your fingerprint doesn't tell anyone else your name, age, date of birth, etc. — but it does tell people that you're you!
- Key regions.

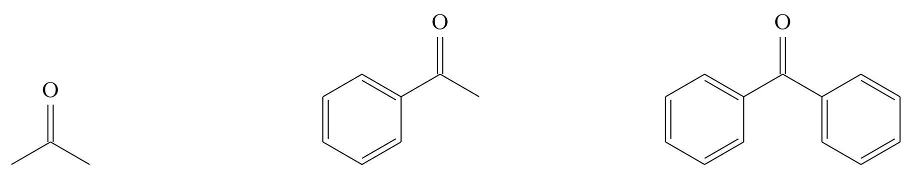
X−H		***sp***		**X=Y**	
FG	**ν (cm^{-1})**	**FG**	**ν (cm^{-1})**	**FG**	**ν (cm^{-1})**
O−H	3600–3200	C≡N	2200	C=O	1840–1630
N−H	3100–2700	C≡C′	2100	C=N	1700–1600
C−H	3000–2850	C=C=C′	1950	C=C′	1670–1600

Table 1.2: Key regions of an infrared spectrum.

[5]The prime on the second carbon indicates that the carbons have different substituents. This is necessary if we are to have a dipole (symmetric C=C bonds are nonpolar).

  - Note that functional groups listed higher up in each column of Table 1.2 have stronger bonds, and thus absorb higher energy/higher ν photons.
  - Both O−H and N−H peaks are broad *if* the proton is exchangeable.
    - There is an example in Clayden et al. (2012) of an O−H that is so sterically encumbered that you don't get proton exchange!!
  - Note also that C−H peaks are often weak, and may not show up at all in some spectra.
  - Should this information be memorized, or will it be provided in a reference chart?
    - Memorize the general regions and trends (as presented in the discussion following Figure 1.12), but not the explicit data in Table 1.2.

- **Broad** (peak): An absorbance peak that stretches over a wide range of wavenumbers.
- **Sharp** (peak): An absorbance peak that is restricted to a narrow range of wavenumbers.
- What determines the *exact* absorption frequency of a chemical bond?



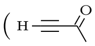
(a) Acetone. (b) Acetophenone. (c) Benzophenone.

Figure 1.13: Related molecules with slightly different infrared absorption peaks.

  - The exact frequency is determined by the atoms and functional groups surrounding the bond.
  - For example, consider acetone, acetophenone, and benzophenone.
    - These three molecules all have C=O bonds, but their C=O bonds absorb IR waves at 1715, 1692, and 1664 cm^{-1}, respectively.
    - This effect can be attributed to increasing conjugation with the π-systems of the aromatic rings.
  - Indeed, the more conjugated the C=O bond, the weaker it is. Conjugation takes off 20–30 cm^{-1} per conjugation!
  - Conjugation is just one example, however; many other group of atoms can affect the absorption frequency.

- Guidelines for interpretation.
  - Look for the presence or absence of key functional groups.
    - This is really good for O−H, N−H, C≡N, C=O, C=N, C=C′, etc.
  - We'll also rationalize trends.
    - Stronger bonds have higher frequencies, and hence get shifted to the left.
    - Weaker bonds have lower frequencies, and hence get shifted to the right.
    - Etc.

- Why do we use wavenumbers instead of per second for frequency?
  - Historical reasons; this is just the way chemists have always done it.

- Example spectrum: But-3-yn-2-one ().

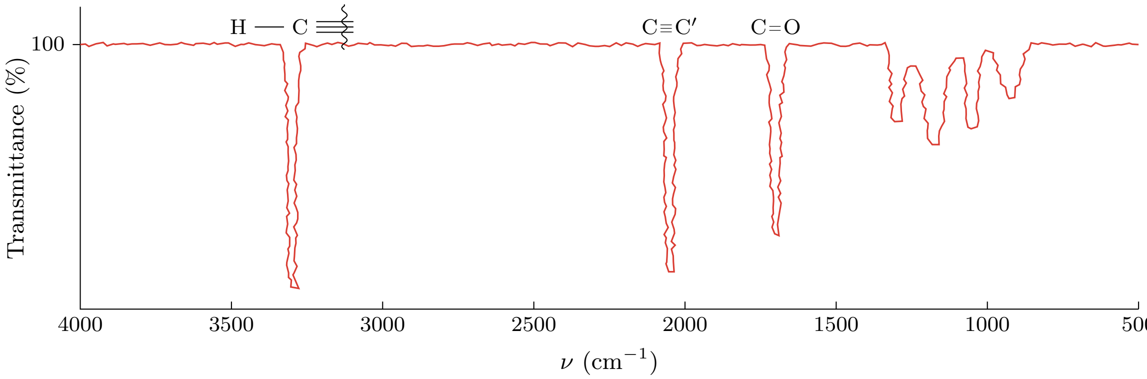


Figure 1.14: Infrared spectrum of but-3-yn-2-one.

  - This spectrum is composed of four major elements: A sharp peak at 3300 cm^{-1}, a sharp peak just to the left of 2000 cm^{-1}, a sharp peak at 1700 cm^{-1}, and the fingerprint region.
  - The sharp peak at 3300 cm^{-1} can be attributed to the propionic C−H bond.
  - But wait: We said in Table 1.2 that C−H bonds lay between 3000–2850 cm^{-1}. What gives?
    - The leftward shift is due to the unique chemical environment of this specific C−H.
    - In particular, the carbon in this bond is *sp*-hybridized. It follows that this C−H bond is more polarized. Thus, the bond is stronger than usual, and we need higher frequency IR waves.
  - Evidence that propionic C−H bonds are stronger: Bond dissociation energies (BDEs).[6]
    - The BDE for a propionic C−H is about 125 kcal/mol, while the BDE for an alkane C−H is about 98 kcal/mol.
    - This difference is also reflected in the relative pK_a's of the two hydrogens: Alkane C−H's have pK_a's in the 50s, while propionic C−H's have pK_a's in the 20s.
  - Note that in this molecule, the sp^3 C−H stretch only absorbs weakly, hence why we don't see a peak around 3000 cm^{-1}.
- There is some theory on how much a certain vibration will absorb, but for our purposes, we'll assume that all stretches absorb a good healthy amount of radiation.
- Application: IR is nondestructive.

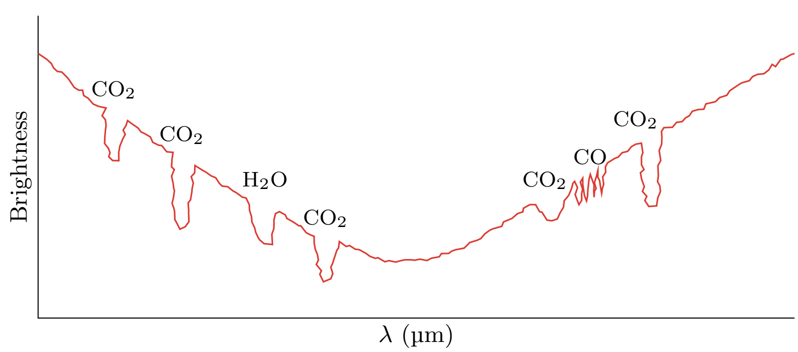


Figure 1.15: Infrared spectrum of the atmosphere of Mars.

[6] Look up BDEs in your Orgo I and Gen Chem notes if you don't remember them. These are important to know!!

  - EA and MS are destructive analytical techniques, meaning that the sample gets destroyed (e.g., by burning or fragmentation) in the process. This requires sample in hand, some of which we can destroy.
  - IR is nondestructive. This means that we can recover our sample after the experiment! In other words, IR spectroscopy can be run from afar.
  - For example, consider the spectrum in Figure 1.15.
    - This is still an IR spectrum, even though the x-axis is in wavelength (λ) — measured in µm — and y-axis is in brightness.
    - The spectrum has a bad baseline, but we'll just forgive this.
    - A number of vibrational modes of CO_2, H_2O, and CO are recorded.
  - What is this spectrum?
    - It is an IR spectrum of the atmosphere of Mars!
    - It was taken by the James Webb Telescope two years ago, in 2022.
    - We've had an IR spectrum of the moon since the 1940s, but this is new and cool!
  - To generalize, here are some major applications of IR spectroscopy.
    - Space.
      - Just like the example in Figure 1.15, IR spectroscopy can be used to find new molecules in celestial bodies.
      - If you ever see a news story along the lines of "Amino acids found on an asteroid," the amino acids in question were probably detected using IR spectroscopy.
    - Climate science.
      - Example: Measuring the concentrations of methane (a potent greenhouse gas) over the arctic.
    - Art.
      - Example: Authenticating old paintings.
      - Indeed, we can use IR to look for diagnostic pigments.
      - A nondestructive method like IR is better in this context than a destructive method like EA or MS because you obviously don't want to chip off a bit of the paint just for an analysis!

- Why is CO_2 (a nonpolar molecule) IR active?
  - The stretching modes are IR silent.
  - However, some of the bending modes induce a dipole, and these are the IR active modes.
- Could we use IR to detect the presence of oxygen on Mars?
  - Oxygen is probably not IR active, so we could not use IR to detect its presence on Mars. There is probably another way, though!

References

Clayden, J., Greeves, N., & Warren, S. (2012). *Organic chemistry* (Second). Oxford University Press.