Week 1

The Basics: Bonding and Molecular Structure

1.1 Course Information

9/28:

- No labs this week.
- Virtual lab: Watch a video and record data in your notebook; answer embedded quiz questions.
- Collaborative Learning in Organic Chemistry (CLOC).
  - 2hr Sunday or Monday.
  - Contact Dr. Britni Ratliff (ratliff@uchicago.edu).
  - Pass/Fail grading (based on attendance).
  - You work on problems related to the lecture content under the supervision of someone who's taken the class before.
  - You can opt-in/out on a quarter-by-quarter basis.
- Review syllabus: Download alternate textbooks, put exam dates in the calendar, add office hours to calendar.
- Develop an understanding of how structure affects reactivity — mechanistic principles.
- You don't have to memorize anything, but you have to remember everything.
  - Like learning a language.
    - Vocabulary, grammar (principles), apply to understand and predict.

1.2 Defining Organic Chemistry

- **Organic chemistry**: Traditionally, the chemistry of living organisms. Now, the chemistry of carbon compounds.
  - Carbon is of particular import because it can bond with itself, and it can form strong bonds with other elements (e.g., C, O, H, S, N, and P) as well.
  - Carbon is bound in simple molecules (such as CO_2 and CH_4), and highly complex ones (such as proteins, DNA, and RNA).
- Carbon compounds:

- Natural: Sugars, fats, gasoline, hydrocarbons, hormones, natural drugs, peptides, rubber, silk, starch, cotton, etc.
- Synthetic: Dyes, fragrances, soaps, drugs, medicines, plastics, materials, teflon, nylon, etc.

- OChem is a central science that feeds into fields such as biochemistry, molecular biology, molecular medicine, math/theory (e.g., buckyballs), engineering, and physics.

1.3 Gen Chem Review

- Today:
  1. Intro (done).
  2. Atomic structure and bonding (review from Gen Chem).
  3. Chemical bonds — octet rule.
  4. Writing Lewis structures.
  5. Formal charges.
- Atomic structure and bonding.
  - Atoms → elements → compounds.
  - Nucleus (protons and neutrons) surrounded by electrons.
  - This year, we'll concern ourselves with the main group elements.
  - Electron configuration:
    - Aufbau principle: Electrons fill orbitals from lowest energy to highest energy.
    - Pauli exclusion principle: 2 elections/orbital with opposite spin quantum numbers (must pair $+\frac{1}{2}$ with $-\frac{1}{2}$).
    - Hund's rule: Orbitals with equivalent energy get partially filled first before more electrons are added.
    - Example: $1s^2 2s^2 2p^6 3s^1$ is Na.
  - Valence electrons are key in this class.
- Noble gas configurations and the octet rule.
  - Lewis noticed that there is a special stability associated with a filled outer shell.
  - Thus, we generally have 8 electrons in the filled outer shell.
    - For example, $\text{Cl} \xrightarrow{1\,\text{e}^-} \text{Cl}^-$ and $\text{Na} \xrightarrow[-1\,\text{e}-]{} \text{Na}^+$.
  - Chemical bonds form because they allow the atoms to achieve a filled octet.
  - Two kinds of bonding: Ionic and covalent.
    - Ionic: Not covered much this year. Lose or gain an electron (forming cations and anions, respectively) to for a filled outer shell. Usually involves a metal and a nonmetal.
    - Covalent: Covered a lot this year. Sharing electrons to satisfy the need for an octet.
    - The atoms involved dictate whether bonding will be ionic or covalent.
  - Electronegativity: The ability of an atom to attract its valence shell electrons.
    - Defined by Pauling, who let Li = 1.0 and F = 4.0.
    - This is a very important concept for understanding bonding and reactivity.
    - EN increases across and up on the periodic table: More protons and a shorter distance away from the nucleus both mean a greater pull on the electrons.
    - Mnemonic (highest to lowest electronegativity): F O Cl N Br I S C H P.

    - Non-polar covalent bonds form when $\Delta EN < 0.5$.
    - Polar covalent bonds form when $\Delta EN \approx 0.5 - 1.9$.
  - Exceptions to the octet rule: H wants $2\,e^-$. Be wants $4\,e^-$. B and Al want $6\,e^-$. Molecule has an odd number of electrons (e.g., NO with 11 electrons is stable).

- Lewis structures.
  - General rules/procedure (there are exceptions).
    1. Determine the total number of valence electrons for the molecule. Add electrons for negative charges; remove for positive charges.
    2. Draw a skeleton and join atoms with single bonds. Put the atom that likes to make the most bonds in the center.
    3. Deduct 2 electrons from the count in step 1 for each single bond. Fill outside atoms with lone pair electrons.
    4. The remaining electrons go on the central atom.
    5. If you have too few electrons for every atom to have an octet, use lone pair electrons to convert single bonds to double bonds. We can also use triple bonds.
  - CH_4 and NH_3 presented as a worked example.