**CHE2060 Lecture 1 Examples**

***Not graded***

**1.1 Very brief history of the development of chemistry**

**1.2 What is organic chemistry?**

1. When a solution of sodium chloride in water is mixed with silver nitrate, a white precipitate forms immediately. When tetrachloromethane is mixed with silver nitrate, no precipitate is produced. Explain these events in terms of the types of bonds present in the two compounds.

The ionic bonds of sodium chloride dissociate when water is added because they are relatively weak electrostatic attractions. Once the ions have been separated each interacts with the very polar solvent, water. In a sense, chloride and sodium ions are more likely to bond with water than with one another. In contrast, the bonds of tetrachloromethane are covalent, stronger than ionic bonds, less highly charged and do not dissociate in water. In fact, tetrachlormethane is much less polar than water and the two liquids are immiscible.

2. Here’s a thought question: Why is carbon the basis of the molecules of life? Why not another atom? What makes carbon especially well suited to be the basis of a wide variety of molecules?  
- Carbon, like other molecules in column IVA, is able to form four bonds. That gives it more options for branching than O, N, and the flurorines.

- So why not silicon (Si) also in column IVA? Carbon’s bonds are generally more stable (stronger) than silicon’s. And most silicon-based molecules are solids – less versatile. SO2 is a solid, so respiration with silicon-based organisms wouldn’t involve gases. Silicon couldn’t cycle through the atmosphere, biosphere, geosphere and aquasphere as well as carbon.  
- And carbon is fairly abundant.

**1.3 Atomic models: nuclear to quantum**

3. Think about the nuclear model of the atom, and where each subatomic particle (proton, neutron and electron) is located. Now think about electrostatic force as it applies in everyday life and the science of physics. How does the nuclear structure of the atom appear to contradict everyday electrostatic force?

The nucleus is a very concentrated structure made of protons and neutrons. The packed protons should repel one another, but don’t. A quantum force, the ‘strong nuclear force’ is hypothesized to hold the nucleus together.

*(see ChemiWiki: Components of the nucleus)*

Negatively charged electrons should be sucked down into the positively charged nucleus, collapsing the structure of the atom. But this does not occur.

*(see ChemWiki: Why doen’t atoms collapse?)*

**1.4 All about orbitals**

4. While a number of the atoms involved in organic chemistry include d orbitals, none have variable charges like the transition metals do. Why?

None of the atoms prevalent in organic chemistry have d orbitals in their valence shells. Transition metals have variable charges because their valence shells include d orbitals.

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6. How many electrons can be held in the:

3d orbital 10

2p orbital 6

1s orbital 2

**1.5 How orbitals fill: electron configuration**

7. Using the atom phosphorous and box-arrow diagrams, create electron configuration diagrams that specifically violate each of these rules.

a. Pauli’s exclusion principle

b. the Aufbau principle

c. Hund’s rule

*There are many possible answers; we’ll review some in class.*

8. Write the electron configuration of:

a. atomic chlorine 1s22s2sp63s23p5

b. ionic chloride 1s22s2sp63s23p6

9. All electrons in each atom occupy an assigned (or “ground-state”) energy level.

a. Can that energy level be increased? yes

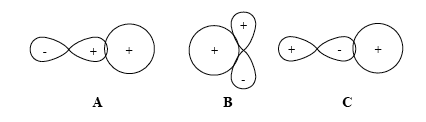
b. If so, how? By an input of energy: light, heat, radiation

c. Is the change permanent or reversible? Reversible – energy is emitted as the electron falls back to ground.

**1.6 Basic bonding: valence electrons & molecular orbitals**

10. Match the orbital overlap shown below with the appropriate type of interaction. Which is:

* bonding C
* anti-bonding A
* non-bonding B



11. How would Figure 1.15 of Chapter 1 change if:

the bond were weaker than that of H2 Higher (shallower) energy well

the bond were stronger than that of H2 Deeper energy well (mimima)

12. If you were able to measure the energy level caused by moving two atoms towards one   
 another, how would you determine the optimal bond length between those two atoms?

The optimal bond length is the inter-nuclear distance at which energy levels are lowest; at the energy minima.

13. How would Figure 1.15 of Chapter 1 change if:

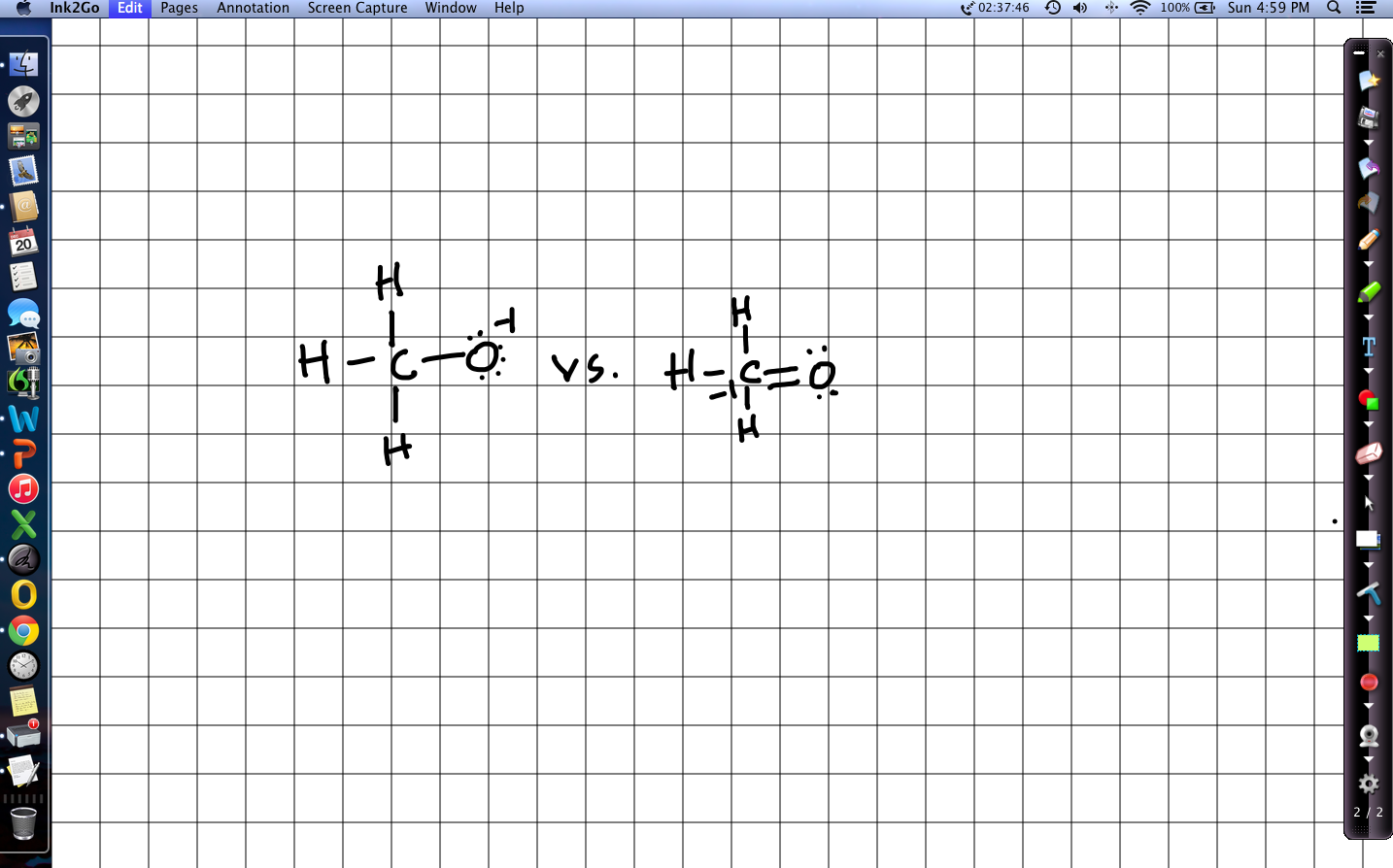
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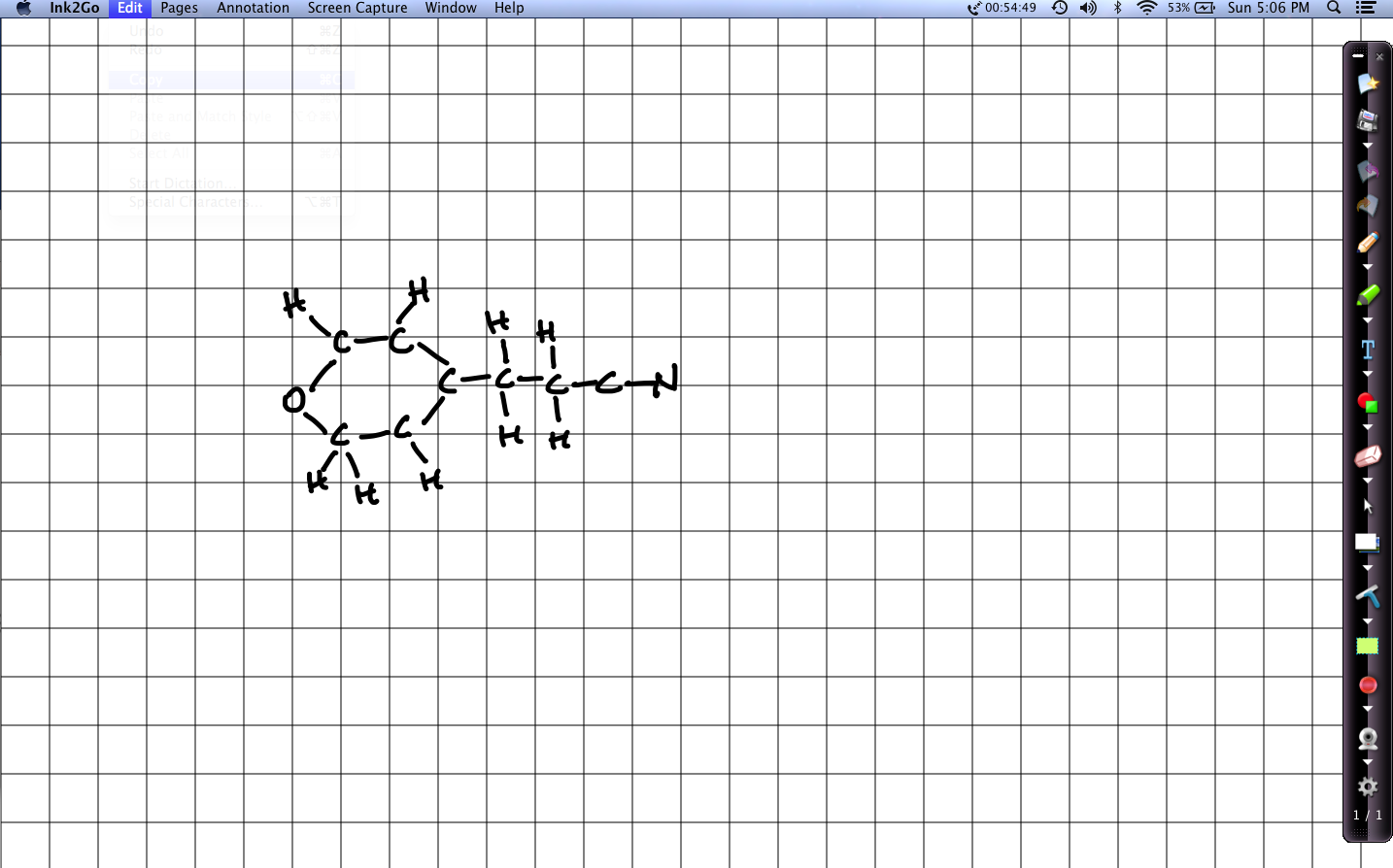
**1.7 Lewis dot structures of molecules**

14. Is it possible to have a double bond between the carbon and the oxygen of the methoxide ion (CH3O-1)? The charge is on the oxygen atom. Draw a Lewis dot structure and explain.

Methoxide has 14 ve-. The correct Lewis structure is shown here on the left. If a double bond were placed between the C and O, the carbon would carry the negative charge as shown in the second structure.



15. Complete this Lewis dot structure by adding bonds and electron pairs where needed.



O gets two free pairs of electrons.

Add another bond between the top two ring carbons.

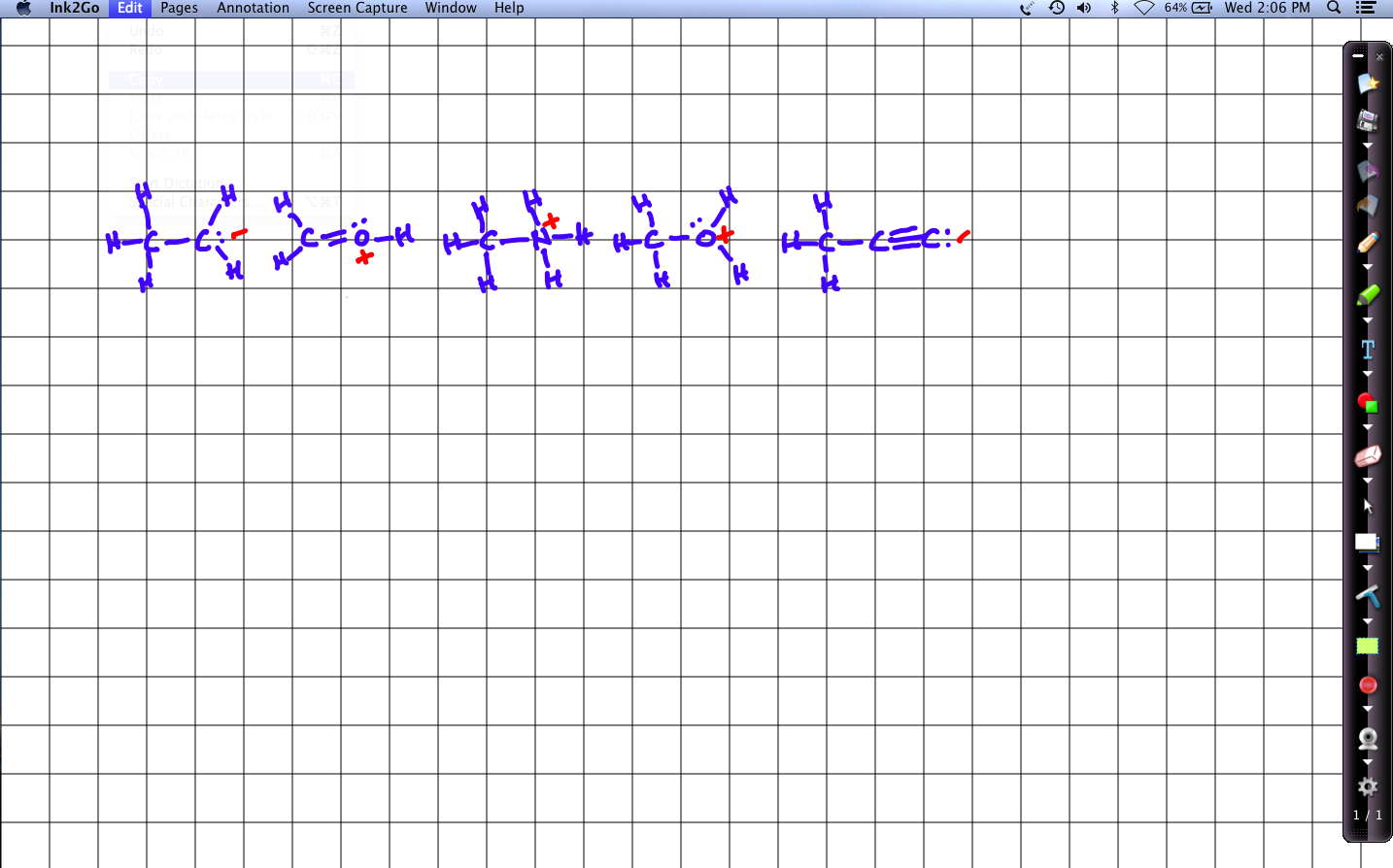
Add another bond between the ring carbon with the acyl chain and the ring carbon below it.

Add two more bonds between C and N.

Add a free pair of electrons to N.

16. Draw Lewis dot structures for these ions. Calculate formal charge to determine which atom carries the charge, and what the charge is.

1. C2H5 anion 14 ve-
2. CH3O cation 12 ve-
3. CH6N cation 14 ve-
4. CH5O cation 14 ve-
5. C3H3 anion (Note that all H are on the same C.) 40 ve-



**1.8 Electronegativity & bond polarity**

17. Consider the N—Br bond. Given the fact that nitrogen and bromine have nearly the same electronegativity (3.04 and 2.96), answer the following questions. Is the N—Br bond covalent, polar covalent, or ionic? Explain your answer. Is the N—Br bond a stable bond?

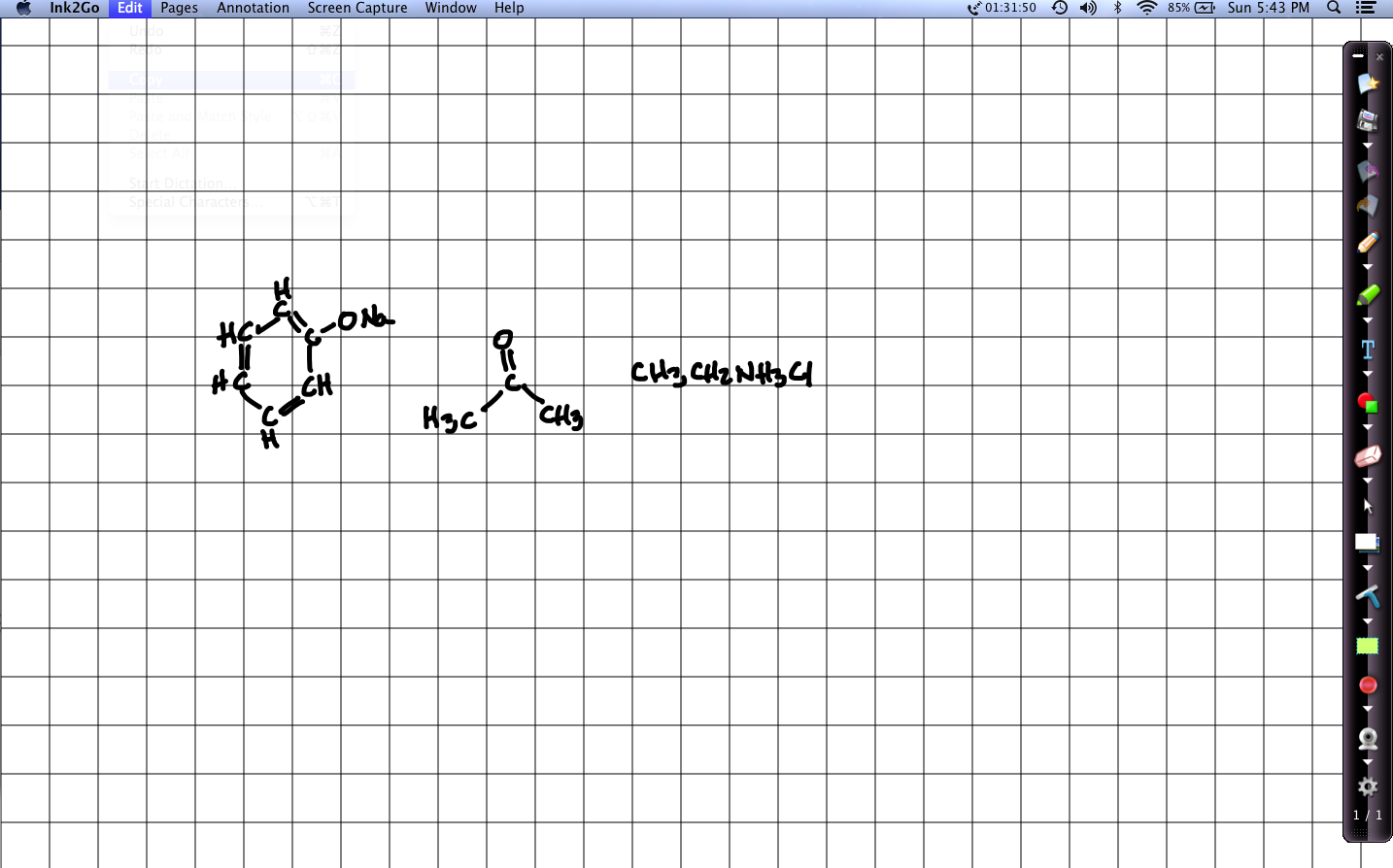
The bond is non-polar as its difference in electronegativity is less than 0.5.  
Yes, non-polar covalent bonds are stable and relatively unreactive.

18. Account for the differences in bond length in the table shown below.

|  |  |
| --- | --- |
| **Molecule** | **C – Cl bond length (pm)** |
| H3C – Cl | 178 |
| ClH2C – Cl | 177 |
| Cl2HC – Cl | 176 |
| FH2C – Cl | 176 |
| Cl3C – Cl | 175 |
| F3C – Cl | 172 |

As you go down the table the chlorine atoms are attached to carbons that are increasingly substituted (bound to) halogens, first chlorines and then fluorines. The halogens have high electronegativity values and pull e-s, including those in the bonds, towards the carbon and halogens and therefore shortening the bond.

19. Which of these are purely covalent and which include ionic bonds?



The ring structure and the third structure are the salts of covalent molecules, so both.

The middle structure is purely covalent.

**1.9 Resonance: a critical concept**

20. A commonly accepted structure of benzene is shown below. Experiments show that all of benzene’s bonds are equal in length.

1. So why is the structure shown below not quite accurate?
2. Draw a more accurate representation of benzene.

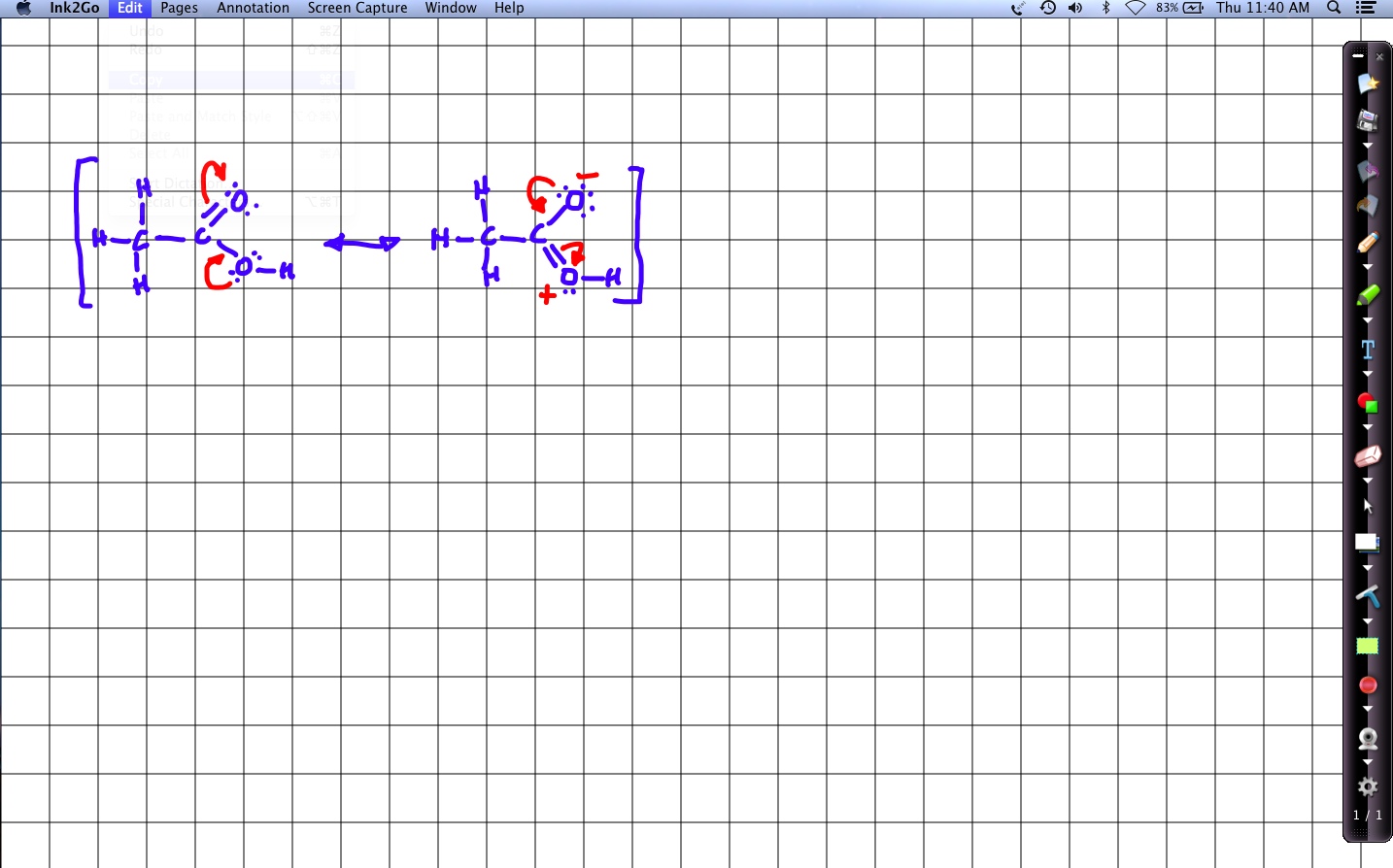
Inaccurate because the structure shown above implies that there are two different bond lengths: shorter double bonds and longer single bonds. A resonance hybrid in which all six bonds have a hybrid length of 1.5 (halfway between single and double) would be more accurate.

21. Acetic acid has the formula,CH3COOH.

a. Draw a Lewis dot structure for acetic acid.

b. Does the molecule have resonance?

c. If so, which resonance structure is the major contributor and why?  
The first structure has fewest formal charges and is therefore the major contributor.



**1.10 Orbital hybridization: key to carbon’s “flexibility”: sp3, sp2 & sp**

22. Label each statement TRUE or FALSE:

a. The number of hybridized orbitals may differ from the number of orbitals that are blended to make them. FALSE

b. Free electron pairs can occupy hybridized orbtials. TRUE

c. Bonding hybridized orbitals exert more repulsive power then hybridized orbitals that hold lone (unbonded) electron pairs. FALSE

23. Rank these bonds in terms of reactivity, from most reactive to least reactive. Explain your answer.

* Double
* Single
* Triple

Triple is most reactive, then double and single is least reactive. Electrons in π bond are further from the nucleus and less tightly held to the nucleus than sigma electrons.

24. Considering the repulsion that exists between electrons in different bonds, give a reason why a planar geometry for methane would be less stable than tetrahedral geometry.

The bonds between carbon and the four hydrogens are electron pairs. The pairs in these bonds are mutually repulsive and seek to get maximum distance from each one another. A tetrahedral geometry allows maximum 3D space between the bonds, therefore the bond electrons.

**1.11 Free electron pairs & radicals**

25. How does a molecules instability (or high energy state) influence its chemical reactivity and why?

Generally, chemical reactivity increases as a molecule’s stability decreases. The increased reactivity is an attempt to increase stability by filling valence shells.

**1.12 VSEPR: classifying molecular geometry & orbital hybridization**

25. Use VSEPR to predict the geometry & orbital hybridization of the following molecules:

PH3 trigonal pyrimidal, sp3

SiH4 tetrahedral, sp3

CCl4 tetrahedral, sp3

:CH3-1 trigonal pyrimidal, sp3

BH3 trigonal planar, sp2

NH4+1 tetrahedral, sp3

BH4-1 tetrahedral, sp3

BeF2 linear, sp

CH3+1 trigonal planar, sp2

26.Indicate which of the following molecules have a dipole moment. Using the geometry of orbital hybridization, draw a three-dimensional representation of the molecule and show the direction of the dipole for the molecule using an arrow.

* CCl4
* CFBr3
* CH3NH2
* CH3Cl
* CH2=CHBr
* CH3OCH3
* CH3CH=CH2
* CH3CHClCH3

*Answers coming in an updated version of this document.*