**Lecture 1 Practice Quiz Key**

Practice quizzes are a resource you may choose to use to test you understanding of a topic, or to prepare for graded quizzes or exams. The answer keys to practice quizzes are available to you at all times.

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

1. In the lecture materials, I’ve very broadly sketched out four periods of the development of human knowledge:

1. Divine authority
2. Institutional authority that interprets divine authority
3. The pre-enlightenment period that developed alchemy
4. The enlightenment forward

In which of these periods was knowledge ‘received’ and in which was knowledge actively sought after?

In the first two periods, divine authority and institutional authority, man **received** knowledge from the divine, either directly or through a religious / governmental institution.  
During the dawn of the enlightenment, and in the enlightenment itself, man **actively** sought after knowledge through observation, experimentation and logic.

**1.2 What is organic chemistry?**

2. What was vitalism and How did chemists disprove ‘vitalism’ in the early 1800s?

Vitalism was the belief that biomolecules (organic molecules) contained a divine spark and could only be produced by the divine and the natural world created by the divine. In the early 1800s, chemists like Wöhler were able to synthesize small organic molecules, showing that man could produce the molecules of life.

**1.3 Atomic models: nuclear to quantum**

3. Draw some diagrams to represent orbitals:

a. Draw one atom with these orbitals: 1s, 2s, 2p (all three ps)

b. Draw a nucleus with these orbitals: 2s, 3s, one p orbital in principle energy level 2 and one p orbital in principle energy level 3.

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4. What is the major difference between the Rutherford model of the nuclear atom and the Bohr model of the quantum atom?

The Rutherford model merely holds that the atom has a tiny, dense, central nucleus of protons and neutrons and that the vast majority of the atom’s volume consists of a diffuse cloud of electrons. The Bohr model structures the electrons into orbits around the nucleus at specific and fixed quantum energy levels.

5. What’s the relationship between an atom’s valence electron number and the charge it takes as an ion? Why?

With exception of atoms in column IVA, an atoms charge is determined by how many ve- the atom must lose (if less than four ve-) or gain (if greater than four ve-) in order to fill the atom’s valence shell. For most atoms a full valence shell requires 8 electrons; for H and He only 2 electrons are needed to fill the valence shell.   
If an atom’s number of ve- is less then 4, then the atom will take a positive charge equal to the number of ve-.

If an atom’s number of ve- is greater than 4, then the atom will take a negative charge equal to 8 – ve- number.

6. What does the term “ground state” mean?

An atom or molecule’s ground state is it’s lowest energy state.

**1.4 All about orbitals**

7. How do orbitals change with principle quantum numbers?

a. Draw p orbitals at three different principle quantum numbers (or shells): 1, 2 and 3.

b. Add an arrow showing in which direction energy levels of these orbitals increases.

c. Add an arrow showing in which direction orbital size increases.

d. Add an arrow showing in which direction distance from the nucleus increases.

a. All p orbitals are ‘dumbell’ shaped.

b. Energy level increases 1 🡪 3.

c. Orbital size increases 1 🡪 3.

d. Distance from the nucleus increases 1 🡪 3.

8. In which of these orbitals does an electron possess the most potential energy?

1. 4s
2. 5s
3. 5p
4. 5d 🡨
5. 4d

**1.5 How orbitals fill: electron configuration**

9. Three rules govern the way in which electrons fill atomic orbitals. State each of these ‘rules’.

a. the Aufbau principle

b. Pauli’s exclusion principle

c. Hund’s rule

a. Aufbau: Electrons must fill orbitals in order from low to high and cannot skip an orbital.

b. Pauli’s: One orbital can contain no more than 2 electrons (0, 1 or 2).

c. Hund’s: When 2 electrons are present in one orbital, they must have opposite spins.

10. Draw the Lewis dot structures of these atoms and their ions.

Atom Ion

Ca 2 dots no dots, +2 charge

N 5 dots 8 dots, -3 charge

P 5 dots 8 dots, -3 charge

Mg 2 dots no dots, +2 charge

C 4 dots 4 dots, no charge NO ION!

O 6 dots 8 dots, -2 charge

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

11. What molecular measurement is a result of the energy minimum between two bonded atoms?

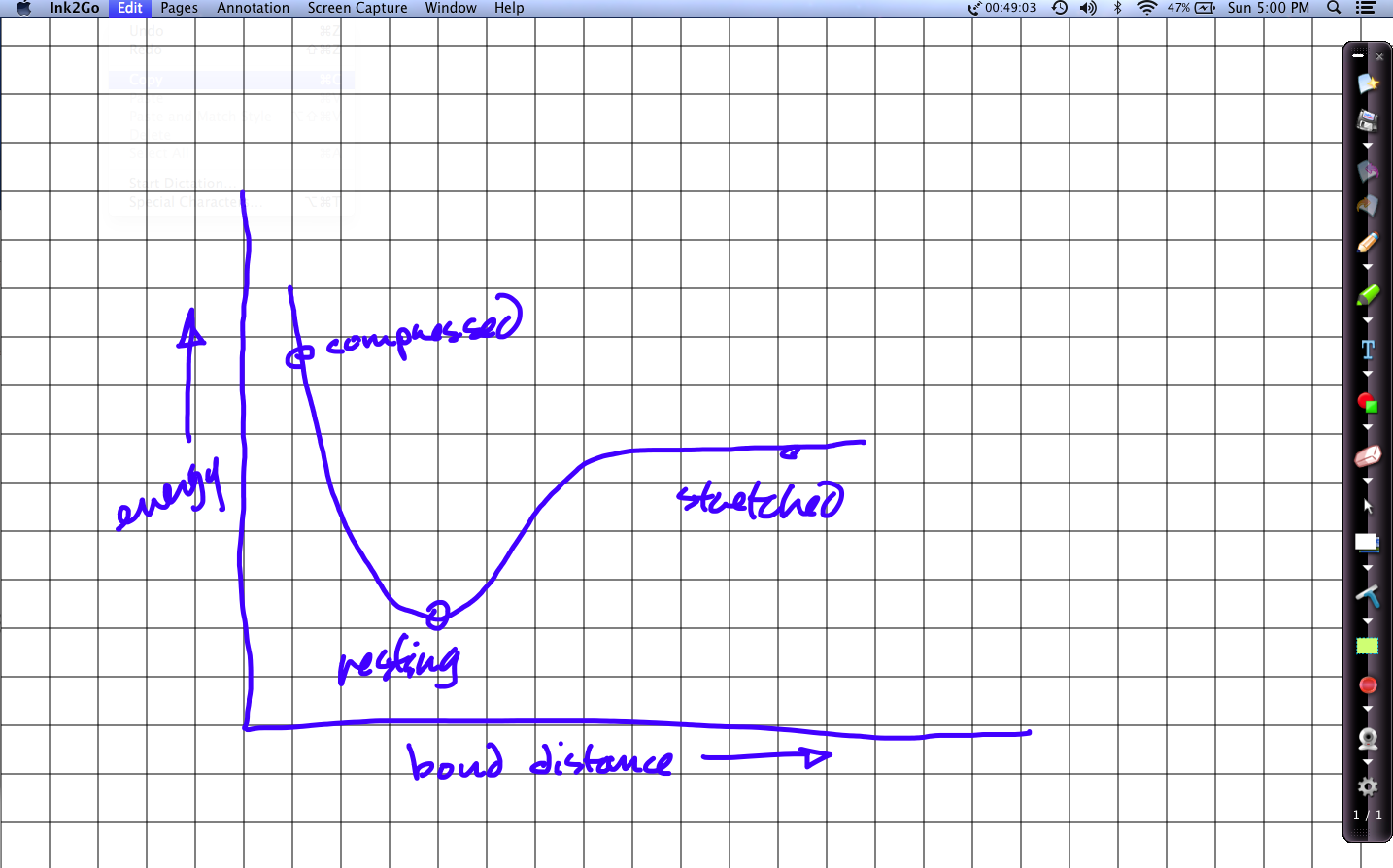
The nucleus-to-nucleus distance; bond length

12. For each of the molecules listed below, answer these questions:

1. How many valence orbitals are there in the molecule?
2. How many are bonding
3. How many are non-bonding?
4. How many are anti-bonding?
   * H2S 6, 2, 2, 2
   * HCl 5, 1, 3, 1
   * CO2 12, 4, 4, 4
   * CH3OH 12, 5, 2, 5

13. How can a spring be used to represent how bond energies change as the distance between two nuclei travels from too close, through optimal bond length, and then past that bond length?

When the spring is at rest (normal length, untouched) it represents the optimal bonding length and optimal bonding energy (lowest energy point) of a chemical bond.  
Pull on either end of the spring and make it longer. This represents the input of energy required to pull two atoms apart and break their bond.  
Push the ends of the spring in to compress it, shortening its length. This represent the input of energy required to push two atoms together closer than there are at their optimal boding distance.



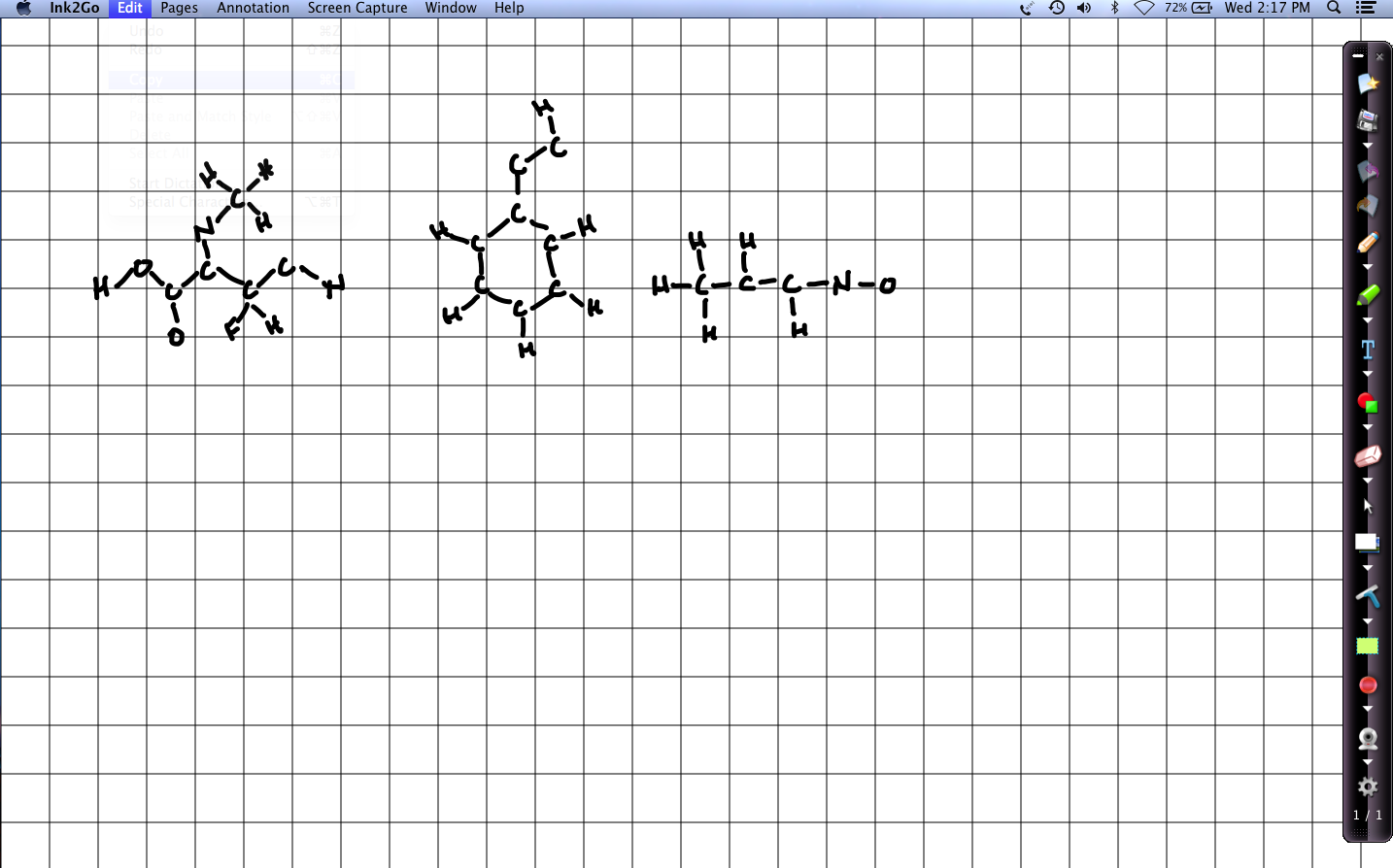
**1.7 Lewis dot structures of molecules**

14. Draw Lewis dot structures, showing all valence electrons, for these molecules:

* CO2
* HCN
* N2
* H2S
* SiO2
* CH2NH

*Answers attached at the end of this document.*

15. Complete these Lewis dot structures by adding π bonds and free electron pairs as needed. None of the molecules have formal charges. Don't add any atoms!



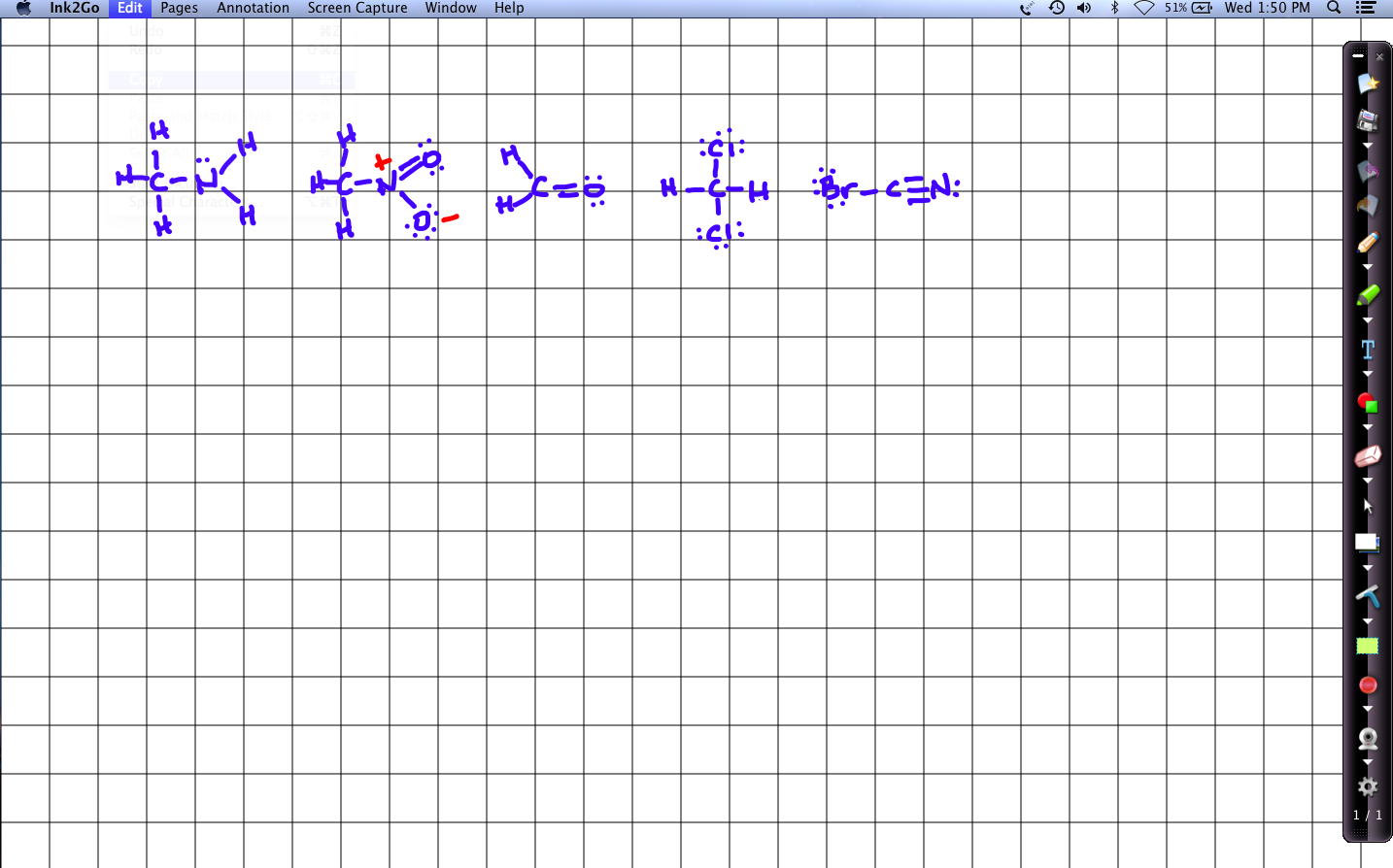
**Left**: 2 free pairs to first oxygen; double bond and 2 free pairs to second O; double bond C-N and free pair for N; 3 free pairs for F; two more bonds C – N and free pair for N.

**Center**: Add 3 pi bonds in the ring, alternating with single bonds. Create triple bond between the two carbons outside the ring.

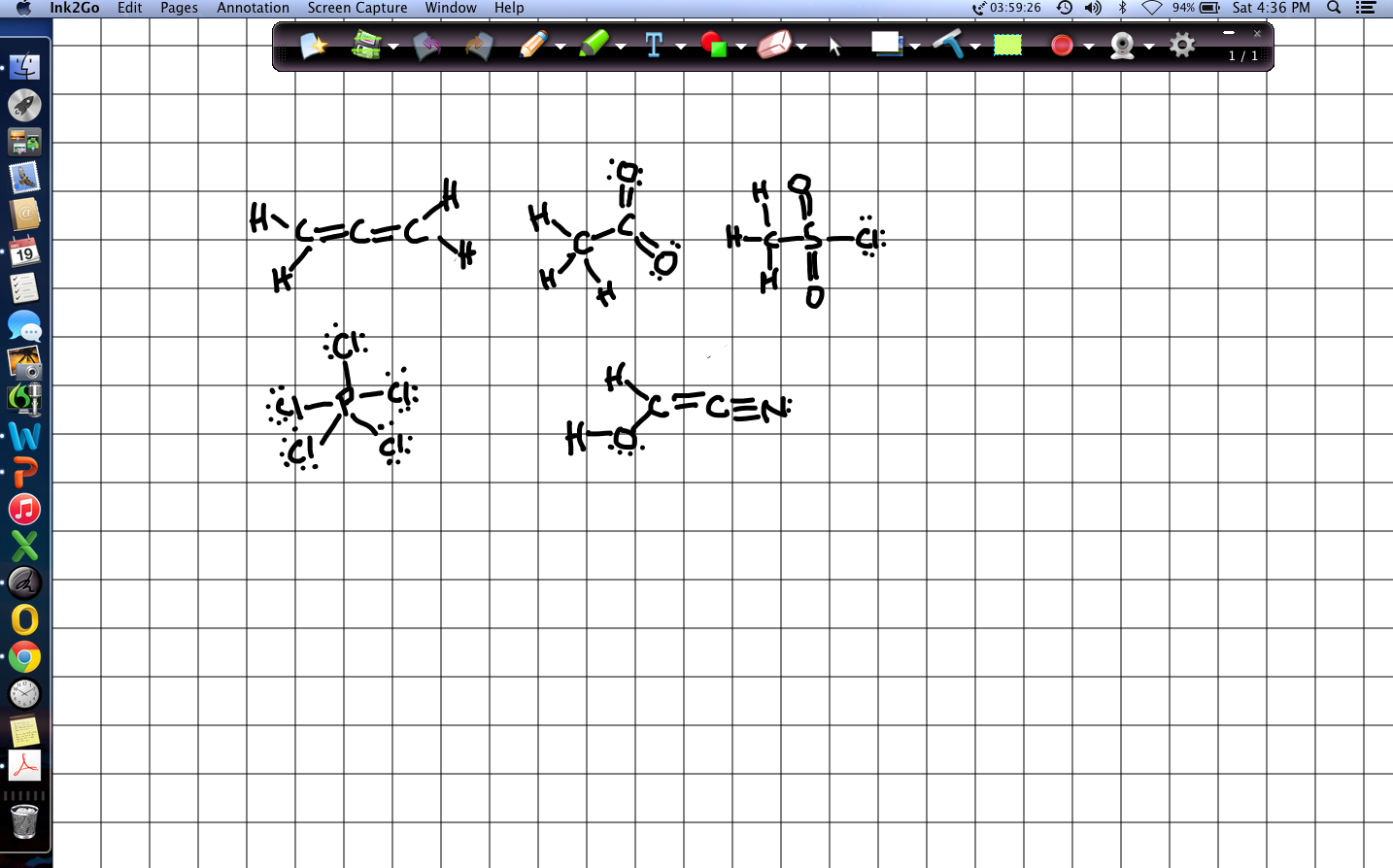
**Right**: Add pi bond between the carbons two and three. Add pi bond between N and O. Add 1 free pair to N and two free pairs to O.

16. Draw Lewis dot structures for each of these molecules:

1. CH5N (There is a bond between C and N.) 14 ve-
2. CH3NO2 (There is a bond between C and N, but not between C and O.) 24 ve-
3. CH2O 12 ve-
4. CH2Cl2 20 ve-
5. BrCN 14 ve-

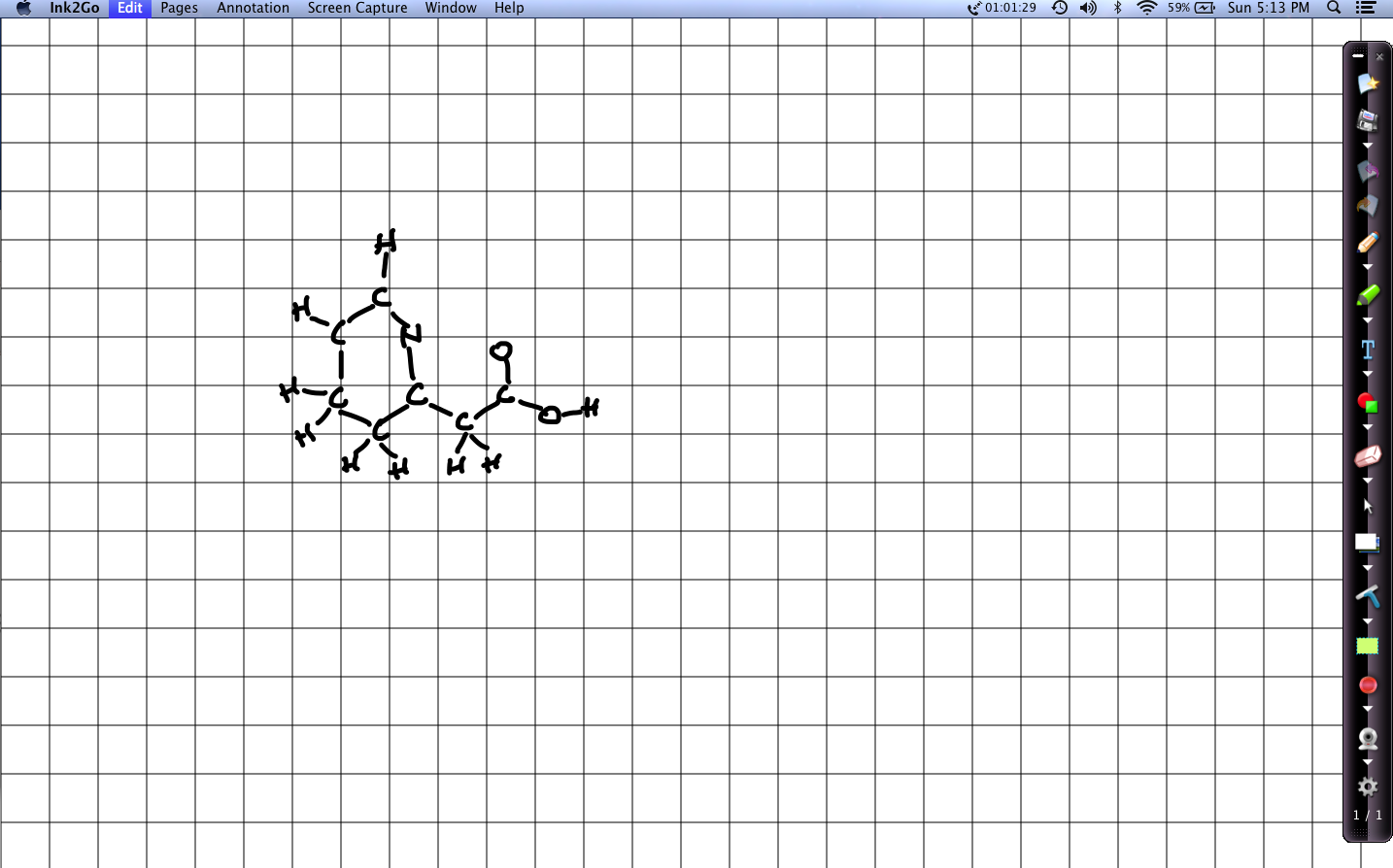


17. Some of the Lewis structures show here are correct and some are not. For those that are not correct, draw a modified structure that is correct.

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Upper & lower left correct. Upper middle: one O should be single bonded and negatively charged. Lower right, central carbon has too many bonds. Upper right, Oxygens need 2 free pairs each.

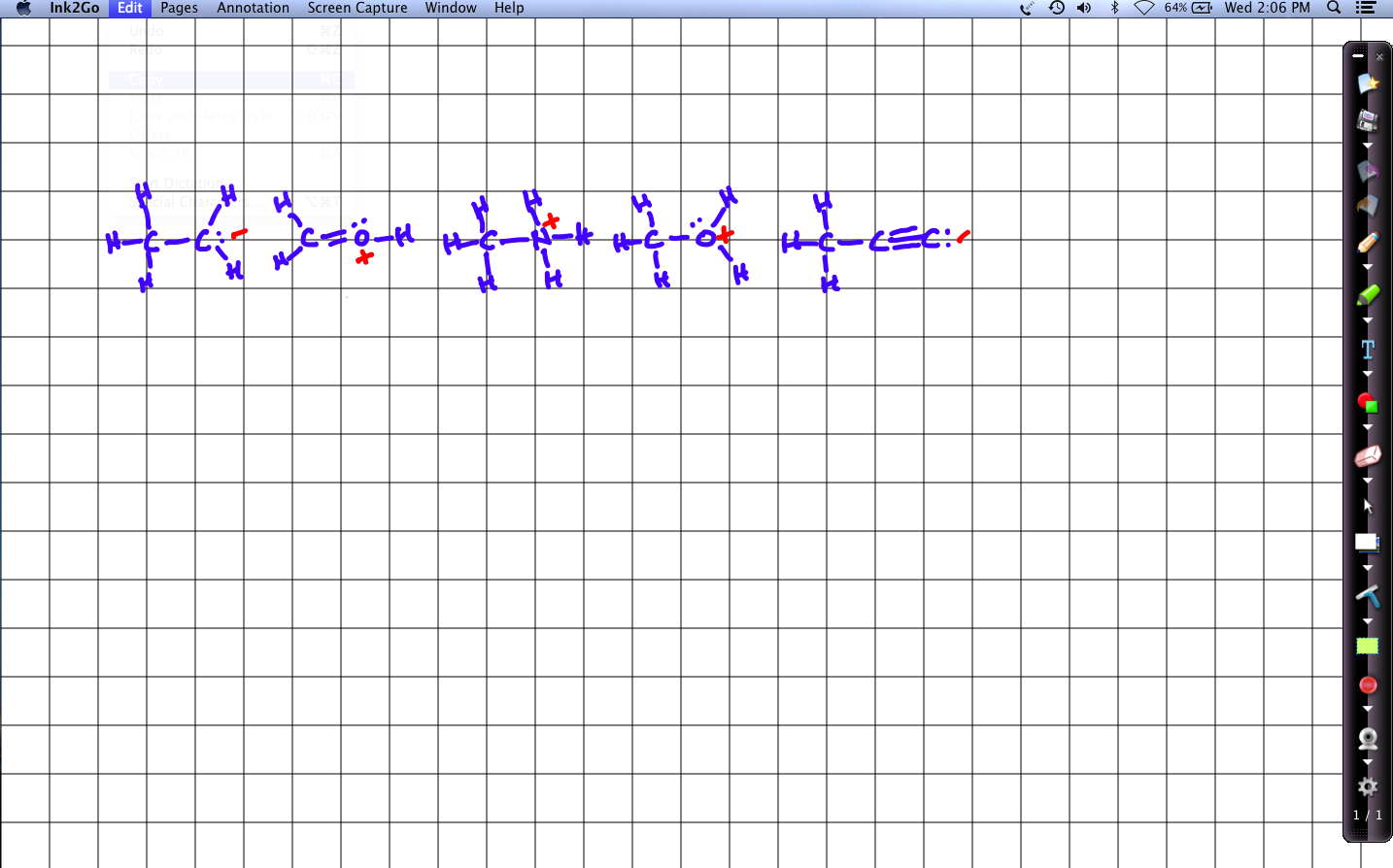
18. Complete this Lewis dot structure by adding bonds and electron pairs where needed.



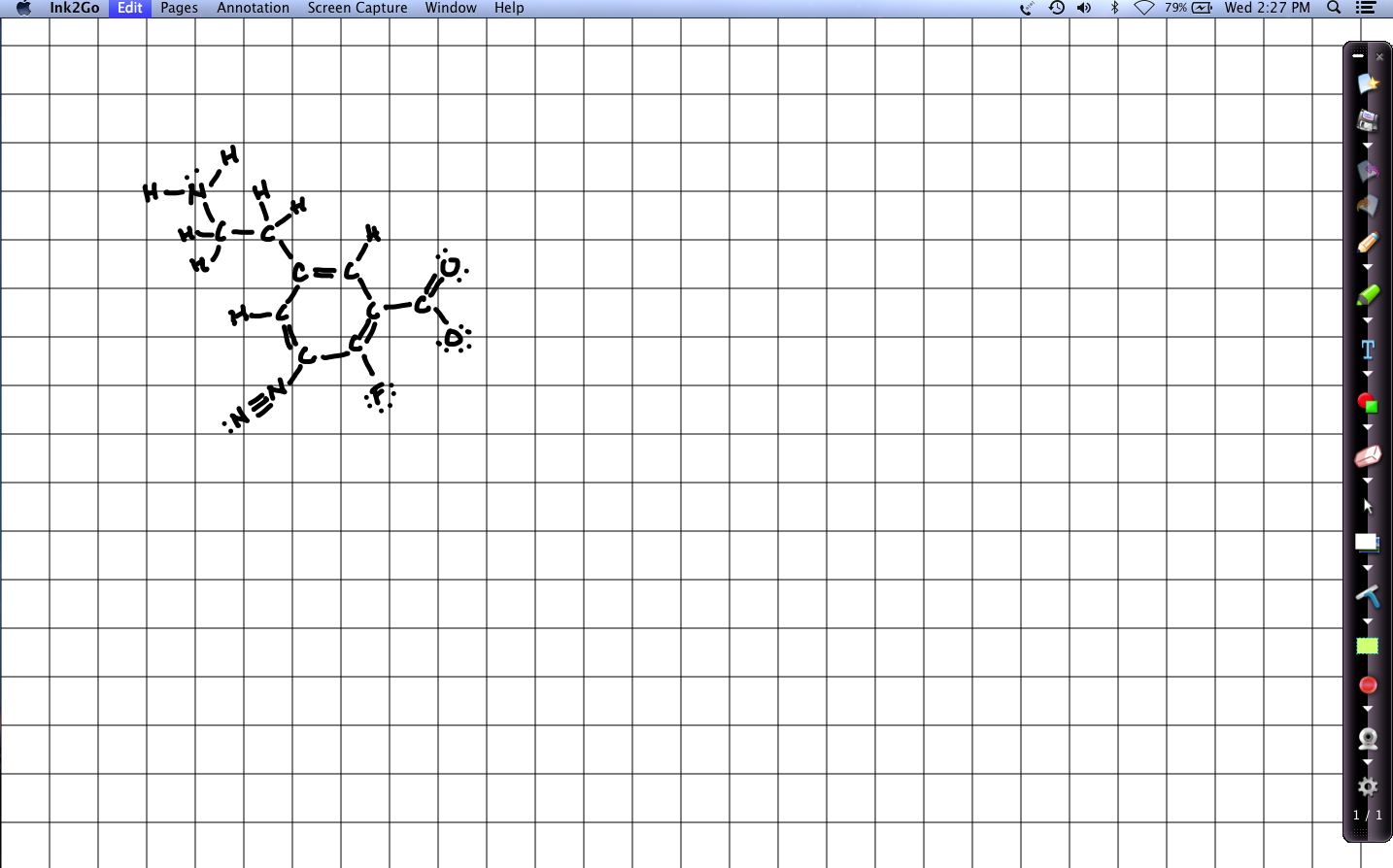
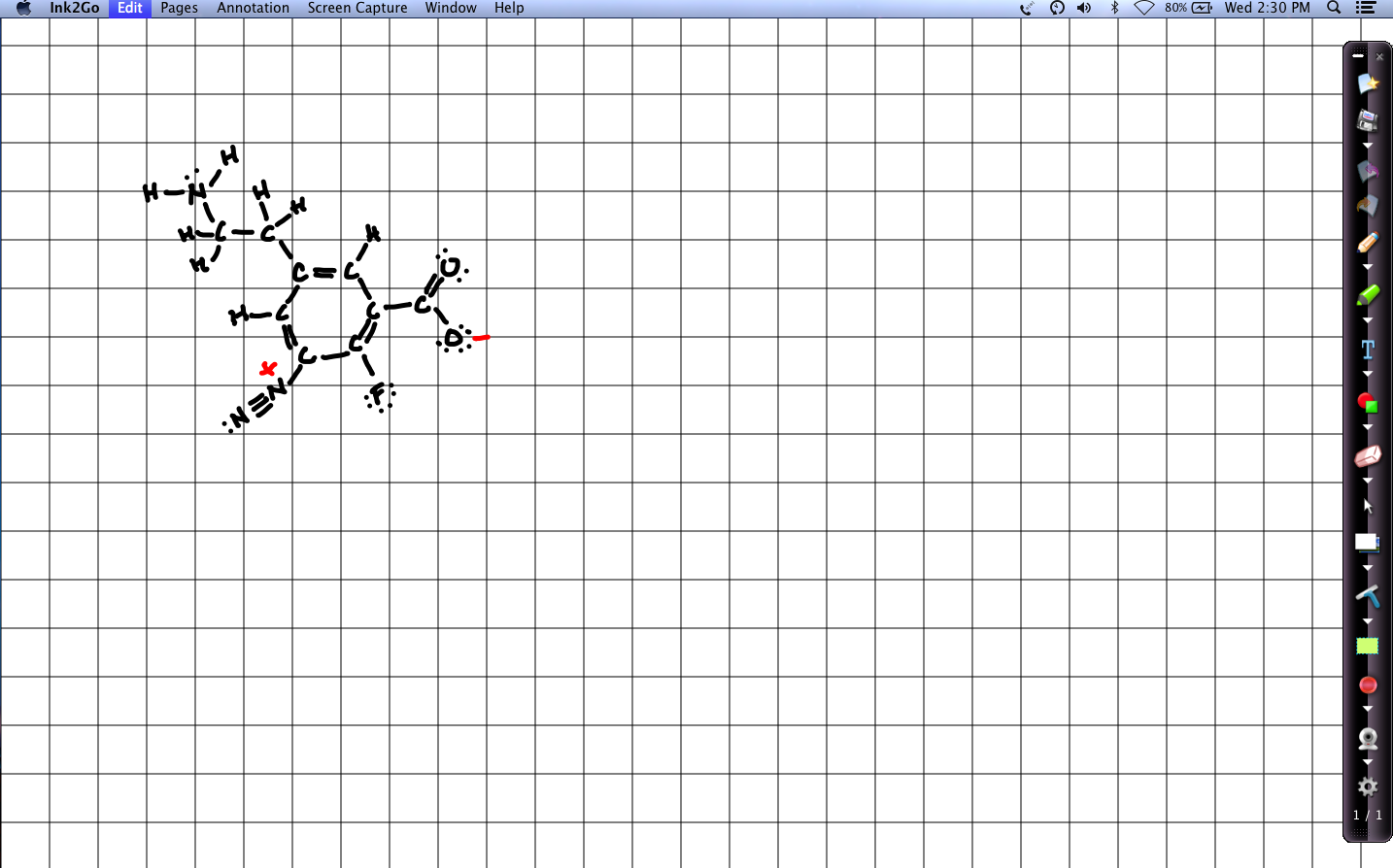
Add pi bond between the two upper ring carbons. Add pi bond between N and the carbon below it. Add 1 free pair to the N. Add pi bond between C and O above it. Add 2 free pairs to each O.

19. 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-



20. Calculate the formal charge on each atom of this molecule. Show any charges at their atom.

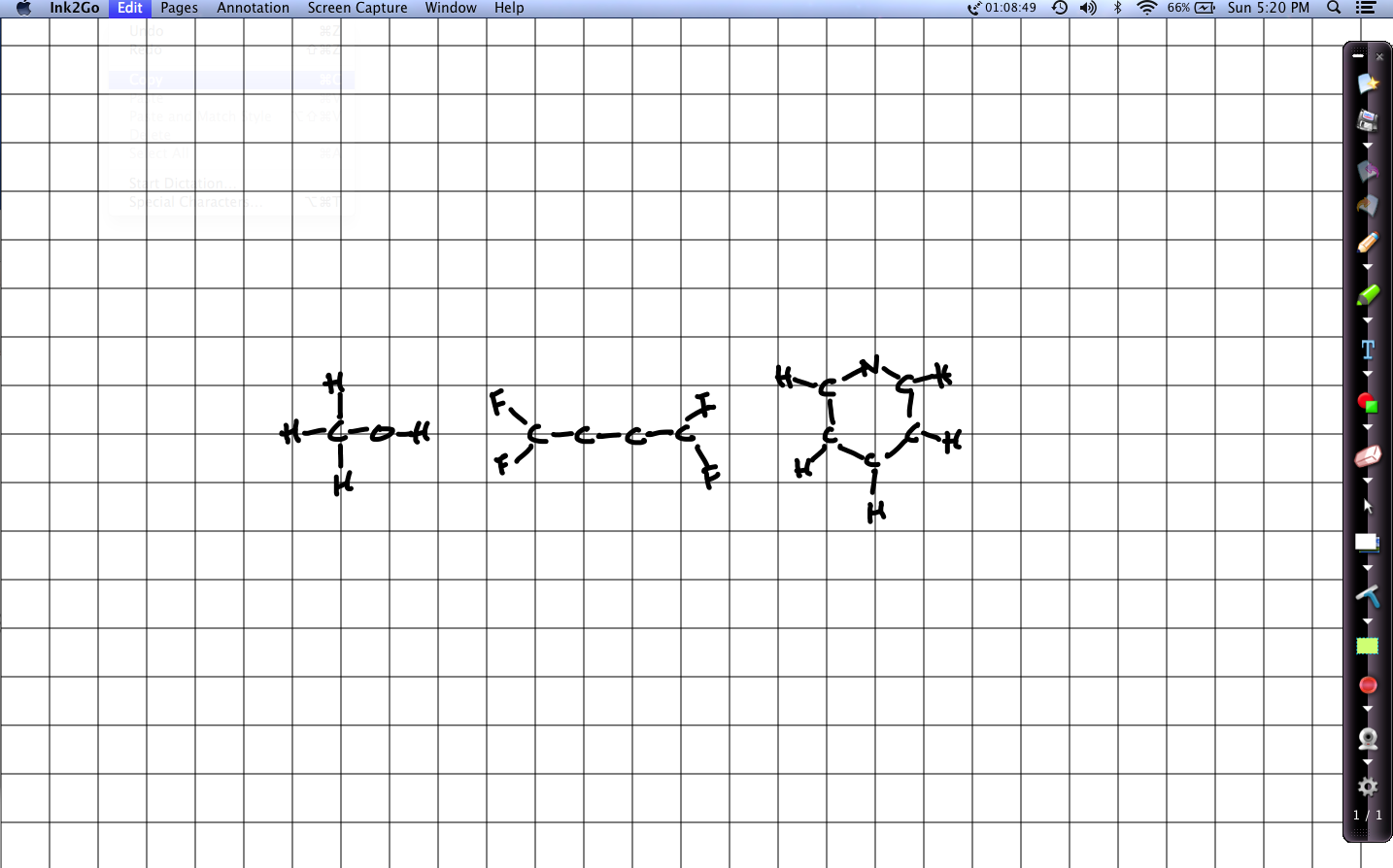
 

**1.8 Electronegativity & bond polarity**

21. For each of the structures shown below,

a. Complete the Lewis dot structure by adding bonds and free electron pairs where needed, and;

b. Draw dipole arrows along each polar covalent bond. Use thicker or longer arrows to show the degree of bond polarity.



22.For each of the following compounds, tell whether the bonds are ionic or covalent. If covalent, show the polarity by drawing a polarity arrow (cross at the non-arrow end) indicating the direction of the dipole.

CaF2 ionic

NaH ionic

FCl covalent: toward F

SiH4 covalent: toward Si

BeF2 ionic

NaI polar covalent towards I

BH3 covalent: towards H

PH3 covalent: towards P

**1.9 Resonance: a critical concept**

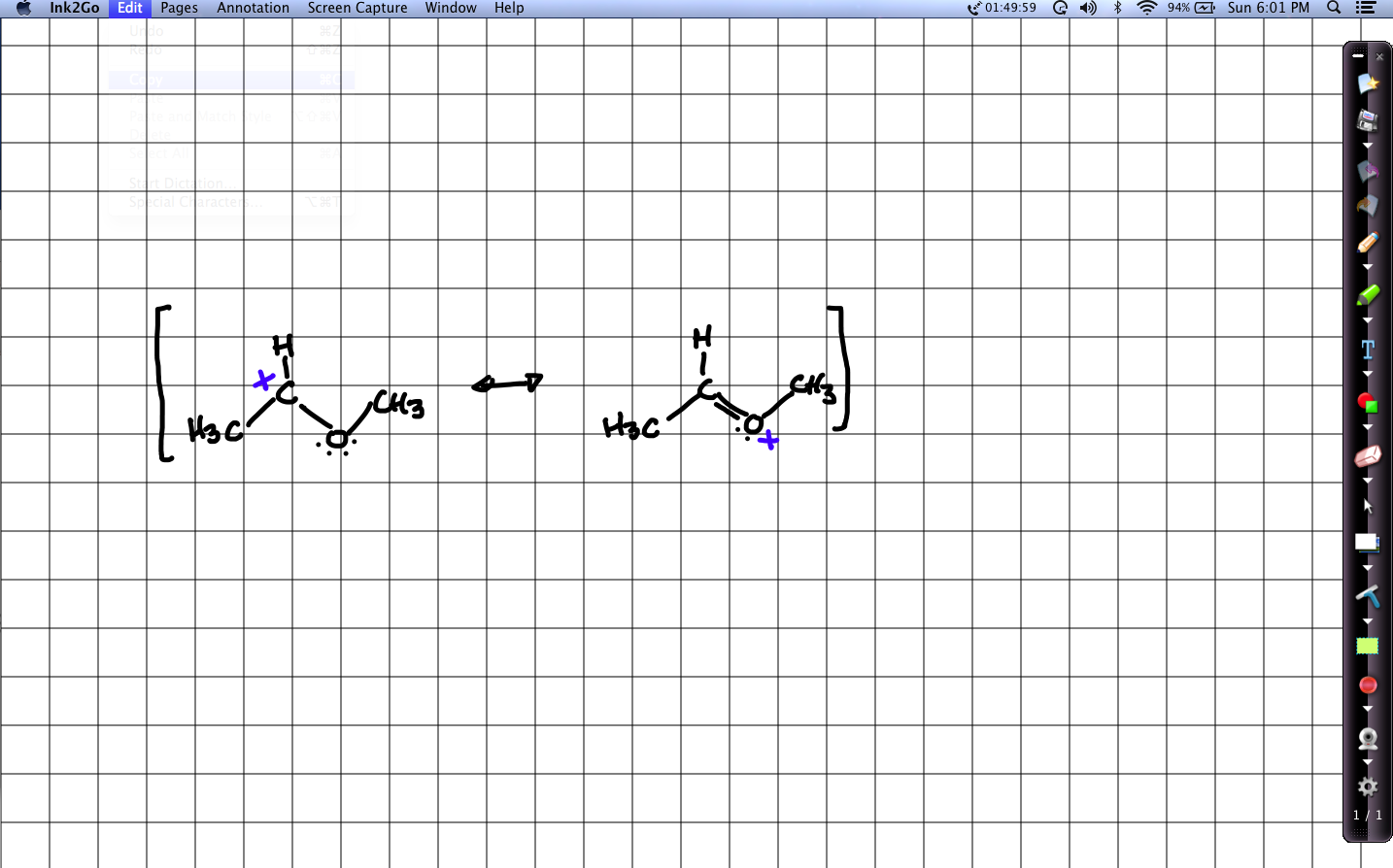
23. Acetic acid has the formula,CH3COOH.

a. Draw a Lewis dot structure for acetic acid.

b. Does the molecule have resonance?

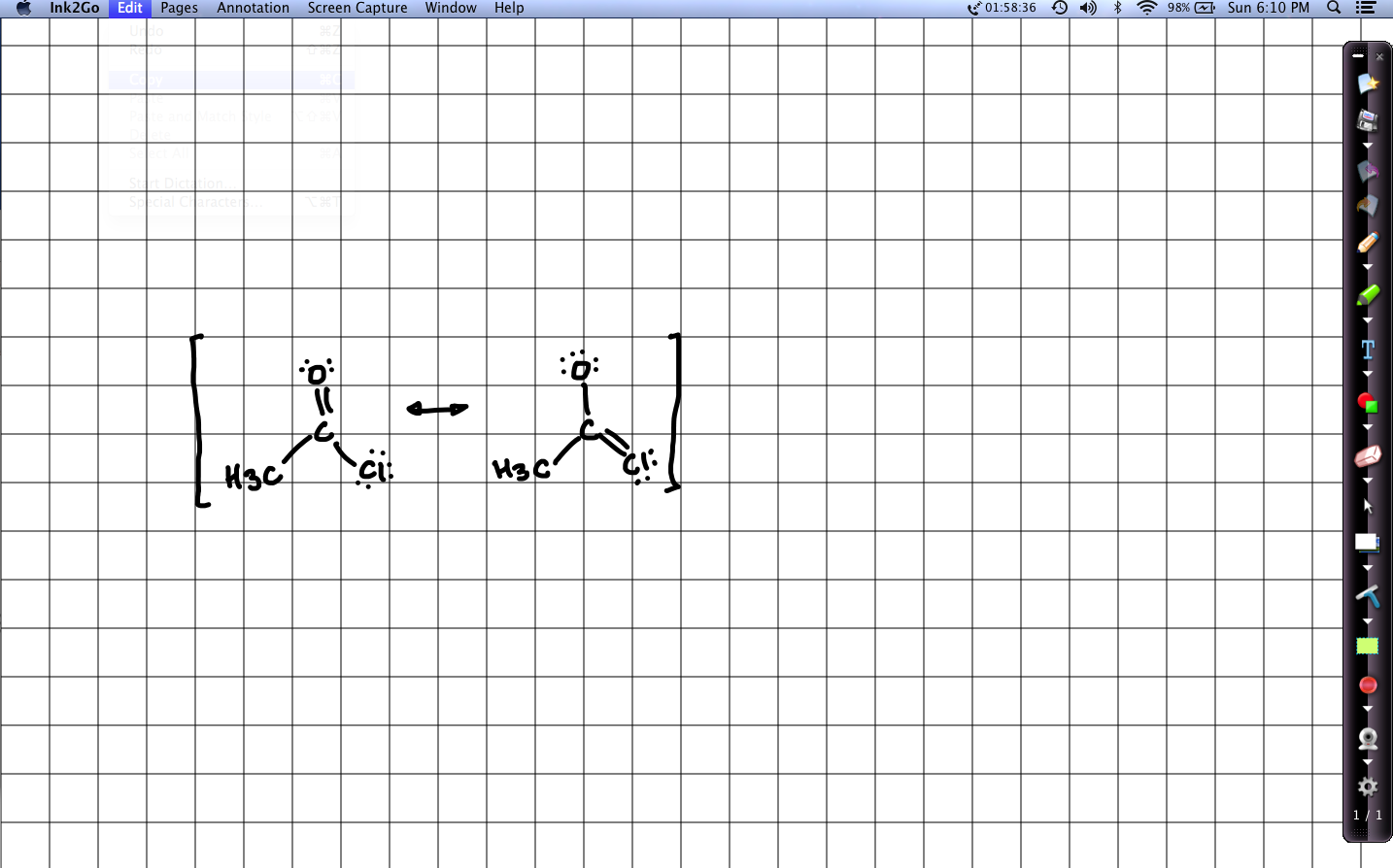
c. If so, which resonance structure is the major contributor and why?

24. Which of the two resonance structures shown below makes a greater (the major) contribution to the resonance hybrid?



In the left-hand structure, the second C does not have an octet and has a + charge.  
In the right-hand structure, the oxygen has a + charge but also has an octet.   
The right hand structure makes the major contribution because all atoms have an octet and because oxygen (with a higher EN value) is better able to hold a charge.

25. Which of the two resonance structures shown below makes a greater (the major) contribution to the resonance hybrid? [NB: Formal charges are not shown below!]



The left-hand structure has no formal charges and all atoms have full valence shell.  
In the right-hand structure, oxygen has an FC of -1 and the Cl has an FC of +1.  
Therefore, the left-hand structure is more stable (lower energy) and is the major resonance contributor.

26. Draw the two resonance structures shown in the problem above. Add arrows to move π bonds or free electron pairs to covert the first resonance structure into the second, and vice versa.

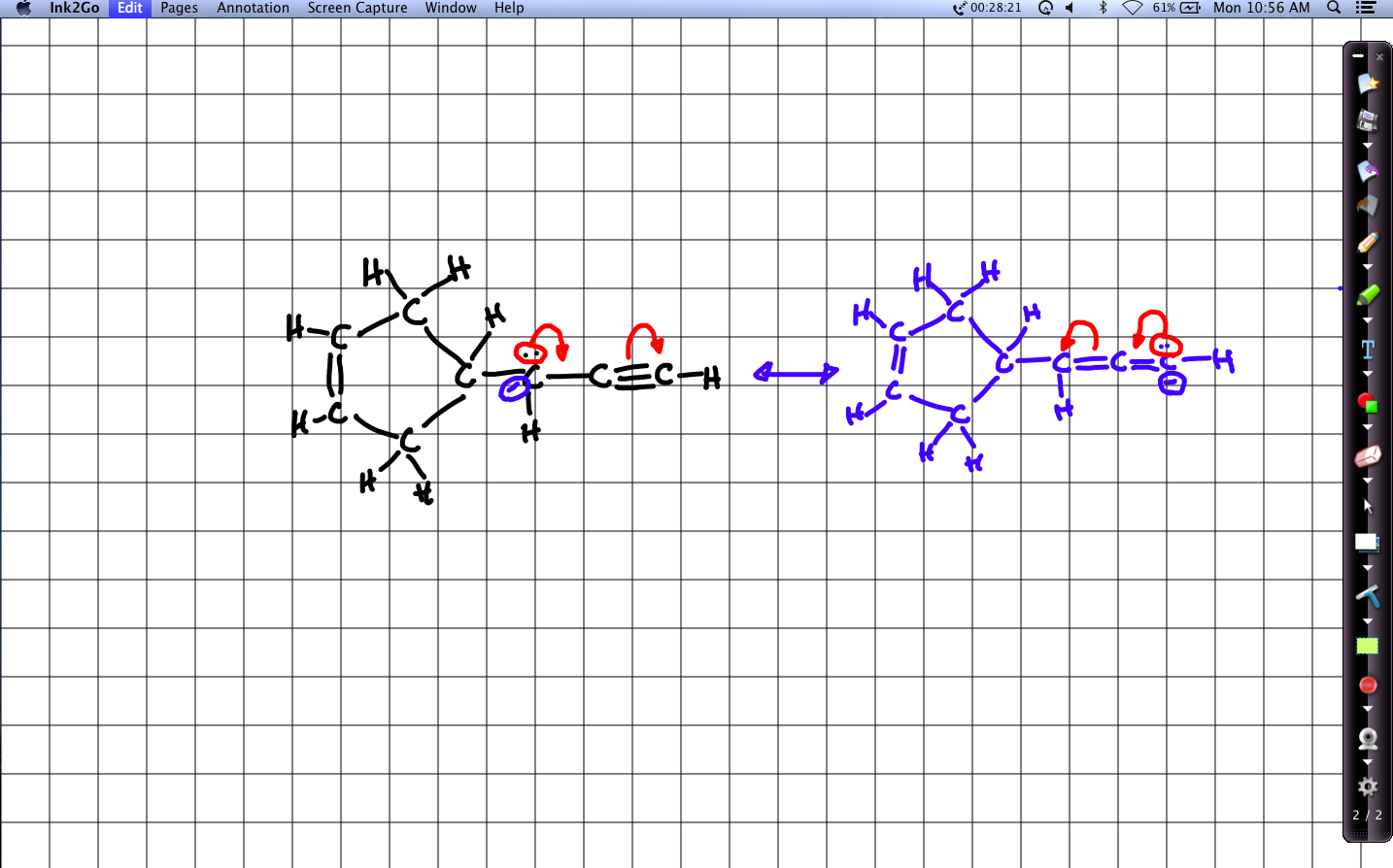
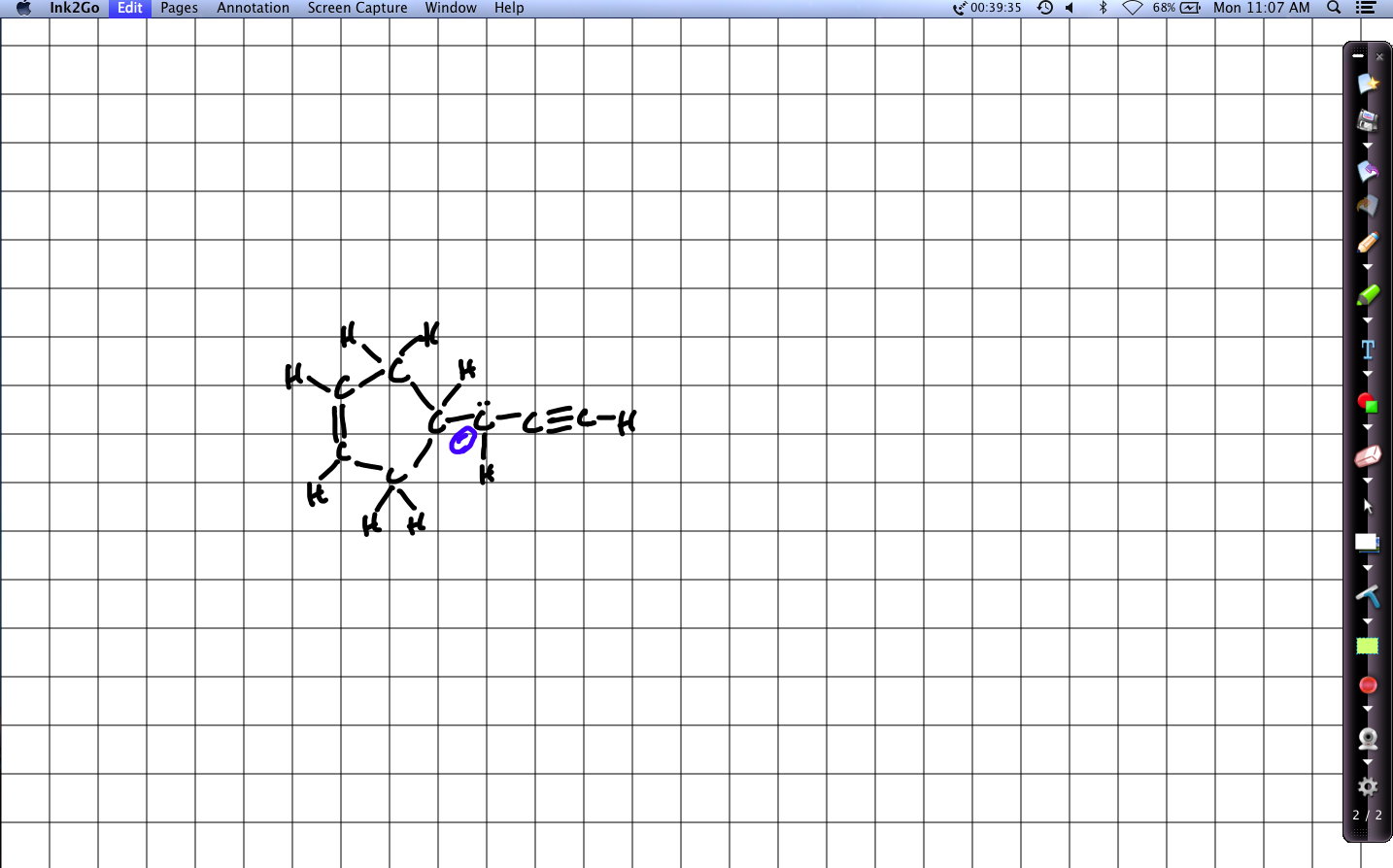
27. The structure shown below has resonance and a negatively charged carbon.

a. Draw the second resonance structure for this molecule by moving the positions of the π bond and negative charge (free electron pair). Be sure to show the new

location of the charge.

b. Use arrows to move π bonds and free electron pairs to convert the first structure

to the second and vice versa.



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

28. Draw the two resonance structures in problem 25 here and Identify the orbital hybridization in each of atom in each structure. There many not be hybridization of all atoms.)

29. Draw the two resonance structures in problem 25 here and Identify the orbital hybridization in each of atom in each structure. There many not be hybridization of all atoms.)

**1.11 Free electron pairs & radicals**

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

30.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 attached at the end of this document.*

31. What does VSEPR stand for? And how does VSEPR help predict & explain molecular geometry?

VSER stands for valence shell electron pair repulsion. Orbitals containing pairs of electrons (whether bonded or unbonded) have a negative charge and repel each other. Repulsive force ensures that orbitals will seek as much space for themselves as possible and will seek to be equally far from one another. So four hybridized orbitals will achieve a tetrahedral geometry, 3 hybridized orbitals will be trigonal planar, and 2 hybridized orbitals will be linear.