**HW Set 1: Atoms, orbitals & bonding topics**

Problems must be solved, or written out, in their entirety with all work shown. If you need more space, you must use engineering graph paper. You must label each set in the upper left hand corner with your name, the date and the chapter. Problems must be identified by number and all work must be shown. Be sure your handwriting is legible.

**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?

**1.2 What is organic chemistry?**

2. Using the molecules shown in the first chapter and lecture as examples, what are the most common atoms in organic compounds; name five.

**1.3 Atomic models: nuclear to quantum**

3. Bohr’s planetary model of the atom is a more sophisticated version of the nuclear model of Rutherford.

a. Describe both models of the atoms. Use diagrams if you like.

b. What are the *most critical* differences between these two models of the atom?

**1.4 All about orbitals**

4. 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.

5. The s orbital:

a. Draw the shape of an s orbital.

b. Draw the electron density plot of the 1s orbital. The x-axis is distance from the nucleus & y-axis is energy level. ^ energy

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 distance from nucleus

c. Do nodes ever appear in s orbitals? If so, draw a picture or diagram to show where.

6. 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.

**1.5 How orbitals fill: electron configuration**

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

a. Can that energy level be increased?

b. If so, how?

c. Is the change permanent or reversible?

8. Find calcium in the periodic table.

a. Write the electron configurations for the calcium atom.

b. How many valence electrons does calcium have?

c. How many core electrons does calcium have?

d. Write the electron configuration of the calcium ion.

e. Is calcium likely to be found in an organic molecule?

9. Describe how you understand the term “spin” used by Pauli.

10. Write the complete abbreviated electron configurations of the third row elements, Na through Ar.

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

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

Atom Ion

Ca

N

P

Mg

C

O

12. Methane is formed by the covalent bonding of one carbon and four hydrogen atoms.

a. How many molecular orbitals are formed?

b. How many of those are molecular bonding orbitals?

c. How many of those are molecular anti-bonding orbitals?

d. Which have higher energy levels, bonding or anti-bonding?

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?

**1.7 Lewis dot structures of molecules**

14. Draw the Lewis dot structures of the molecules listed below and show all valence electrons (all necessary dots).

a. H2S

b. CH2NH

c. SOF2

15. Draw the Lewis dot structure of hydroxylamine, H3NO, and calculate the formal charge of each atom.

16. 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!



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

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

**1.8 Electronegativity & bond polarity**

18. These two bonds have polarities of opposite direction. Show the direction of each and explain.

ICl

FCl

19. For each of these molecules:

1. draw the Lewis structure showing all valence electrons;
2. draw a polarity arrow to show direction of polarity; and
3. classify the molecule as nonpolar covalent, polar covalent or ionic.

HBr

H2O

LiI

BrCl

NH3

KF

20. Working with the Lewis dot structure shown here,

a. Add free electron pairs to complete the structure; and

b. Calculate formal charge for each atom in the structure.



**1.9 Resonance: a critical concept**

21. Nitromethane, CH3NO2, is a molecule with resonance.

a. Draw the two Lewis dot resonance structures of this uncharged molecule.

b. Calculate the formal charges on all atoms of one resonance structure.

c. Draw the resonance hybrid.

22. Draw all resonance structures ***and*** the resonance hybrid of the carbonate ion, CO3-2.

23. Add arrows to the first resonance structure to create the second resonance structure.

24. This enamine is a resonance structure.

a. Draw the second resonance structure by moving the π bond and free electron pair.

b. Add arrows to move π bonds and free electron pairs to convert the first structure to the second and vice versa.

c. Which resonance structure is more stable; the major contributor?



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

25. Draw slot-dot diagrams to show the hybridization of the valence electrons of N in ammonia.

26. Norethidrone is used in oral contraceptives. On the structure shown here, locate an example of each of these features:

* A nonpolar covalent bond
* A highly polar covalent bond
* An sp hybridized carbon
* An sp2 hybridized atom
* An sp3 hybridized atom
* A bond between atoms of different hybridization



27. Allene has the structure CH2 = C = CH2.

1. What is the hybridization of each carbon?
2. What orbitals are involved in each bond?
3. What is the molecule’s shape?

**1.11 Free electron pairs & radicals**

28. Think about atoms as they are presented in the periodic table. Some are radicals and some are not. Look at row two and classify each of its elements as a radical or not.

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

29. Create a diagram that shows a molecule of two carbon atoms joined by a triple bond and bound to two hydrogen atoms (acetylene or ethyne). Show atoms, all orbitals (hybridized and not hybridized), and all bond angles.

30. Using the VESPR table and Lewis dot structures for these three molecules, determine the:

i. central atom hybridization; and

ii. molecular geometry

a. H2S

b. CH2NH (C & N are both central)

c. SOF2