Organic Chemistry Laboratory 1 – Writing Lewis Dot Structures

Purpose

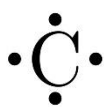
1. To learn how to draw Lewis Dot Structures for atoms and molecules.
2. To develop skills for drawing Lewis Dot Structures for more complex molecules quickly.

Background

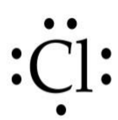
To understand a molecule's chemical behavior, it is necessary to know its connectivity – that is, which atoms are bonded together, and by what types of bonds (single, double, triple). It is also useful to know which valence electrons participate in bonding and which do not. Lewis dot structures (or, Lewis structures) are a convenient way to convey this information.

Before we begin some of the exercises, let's review some basic conventions of Lewis structures.

● ***Lewis structures take into account only valence electrons.*** In a complete Lewis structure, such as that for C by itself, or Cl, the chemical symbol for the element is drawn, and all valence electrons are drawn around the symbol, represented by dots:



4 valence electrons



7 valence electrons

Lewis symbols do not take into account *core* electrons (those occupying lower energy shells). Carbon has two core, and four valence electrons.

● ***Bonding and nonbonding electrons are clearly shown.***

□ Single, double, and triple bonds are indicated by one, two, or three lines (i.e., −, =, or ≡), respectively, which represent the sharing of two, four, or six electrons. Each line represents a shared pair of electrons.

□ Nonbonding electrons are indicated by dots, and are usually paired (:). These are called lone pairs of electrons. In some species (free radicals), nonbonding electrons are unpaired, and are represented by single dots.

● ***Atoms in Lewis structures obey the duet rule and the octet rule.*** Atoms are especially stable when they have complete valence shells: two atoms (a duet) for hydrogen and helium, and eight electrons (an octet) for atoms in the second row of the periodic table. These more stable duets and octets can be achieved through the formation of covalent bonds.

Steps for drawing Lewis Structures

Molecules form, when two atoms form a covalent bond. Whether we are considering a simple molecule such as O_2, or a more complex one such as caffeine, the steps for writing the Lewis structure for each, are the same.

1. Count the total number of valence electrons in the molecule.

   a. The number of electrons contributed by each atom is the same as its group number (H = 1, C = 4, N = 5, O = 6, F = 7).
   b. Each negative charge increases the number of electrons by one; each positive charge decreases the number of electrons by one.

2. Write the skeleton of the molecule, showing only the atoms and the single bonds required to hold them together. If molecule connectivity is not given to you, the central atom (the one with the greatest number of bonds) is usually the one with the smallest electronegativity.

3. Subtract two electrons from the total in Step 1 for each single covalent bond drawn in Step 2.

4. Distribute the remaining electrons as lone pairs.

   a. Start with the outer atoms and work inward.
   b. Try to achieve a filled valence shell on each atom – namely, an octet on each atom other than hydrogen, and a duet on each hydrogen atom.

5. If there is an atom with an incomplete valence shell, convert lone pairs from neighboring atoms into bonding pairs – creating double and triple bonds.

Problem 1: Draw a Lewis structure of CCl_4, where carbon is the central atom.

Step 1. Count total valence electrons.

quantity of given atom	atomic species (H,C,O,Cl)	# valence electrons in atom	total valence electrons
Total # valence electrons: ___________			

Step 2. Write the skeleton of the molecule, showing only atoms and single bonds required to hold them together.

Step 3. Subtract two electrons from total in Step 1 for each covalent bond drawn in above step, and write that amount here: _________

Step 4. Rewrite the structure from above, and distribute the remaining electrons as lone pairs.

Step 5. For atoms with incomplete valence shell, convert lone pairs from neighboring atoms into bonding pairs, making double or triple bonds.

Problem 2: Draw a Lewis structure of HCO_2^-, where carbon is the central atom.

Step 1. Count total valence electrons.

quantity of given atom	atomic species (H,C,O,N)	# valence electrons in atom	total valence electrons
Total # valence electrons: ___________			

Step 2. Write the skeleton of the molecule, showing only atoms and single bonds required to hold them together.

Step 3. Subtract two electrons from total in Step 1 for each covalent bond drawn in above step, and write that amount here: _________

Step 4. Rewrite the structure from above, and distribute the remaining electrons as lone pairs.

Step 5. For atoms with incomplete valence shell, convert lone pairs from neighboring atoms into bonding pairs, making double or triple bonds.

Problem 3: Draw a Lewis structure for C_2H_3N. A carbon that is bonded to three hydrogen atoms is bonded to the second carbon, which is bonded, in turn, to the nitrogen.

Step 1. Count total valence electrons.

quantity of given atom	atomic species (H,C,O,N)	# valence electrons in atom	total valence electrons
Total # valence electrons: __________			

Step 2. Write the skeleton of the molecule, showing only atoms and single bonds required to hold them together.

Step 3. Subtract two electrons from total in Step 1 for each covalent bond drawn in above step, and write that amount here: _________

Step 4. Rewrite the structure from above, and distribute the remaining electrons as lone pairs.

Step 5. For atoms with incomplete valence shell, convert lone pairs from neighboring atoms into bonding pairs, making double or triple bonds.

Strategies for Success

In solving Problems 1 and 2, you may have noticed that each type of atom tends to form a specific number of bonds and to have a specific number of lone pairs of electrons. In the laboratory for Writing Formal Charges you will learn that *those are the number of bonds and lone pairs for atoms that bear no formal charge*.

Refer to the table below to complete the following Lewis structures.

Common Number of Covalent Bonds and Lone Pairs for Selected Uncharged Atoms		
Atom	Number of Bonds	Number of Lone Pairs
H	1	0
C	4	0
N	3	1
O	2	2
F, Cl, Br, I	1	3
Ne	0	4

Problem 4(a – c): Draw Lewis structures for each of the following molecules, applying the rules from the table above.

(a) CH_5N (contains a bond between C and N)

(b) CH_2Cl_2

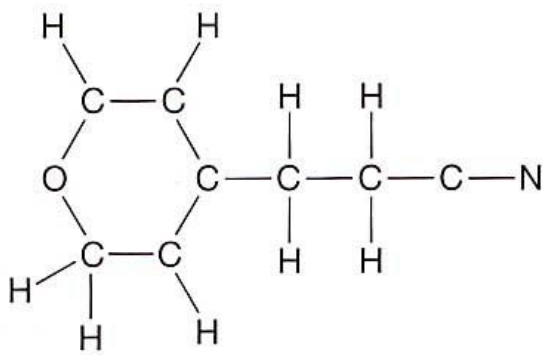
(c) BrCN

Problem 5 (a, b): Complete the Lewis structures, using table above only as a guide – all species will have a *formal charge*.

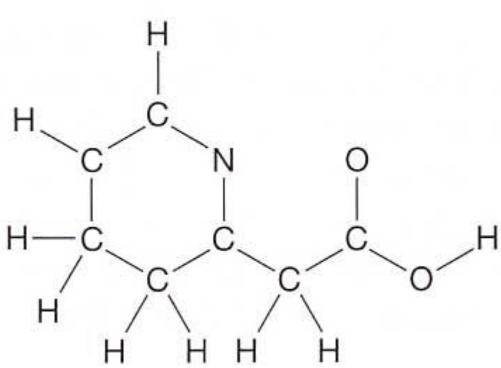
(a) CH_3NO_2 (contains a bond between C and N, but not between any C and O)

(b) CH_2O

Problem 6: Complete the Lewis structure for the compound whose skeleton is shown below. Assume that all atoms are uncharged.

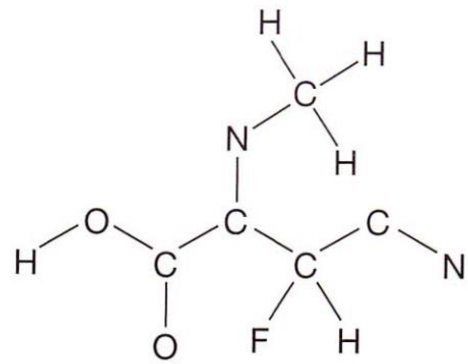


Problem 7: Complete the Lewis structure for the molecule with the following connectivity. You may assume all atoms have the number of bonds and lone pairs as indicated in table above.

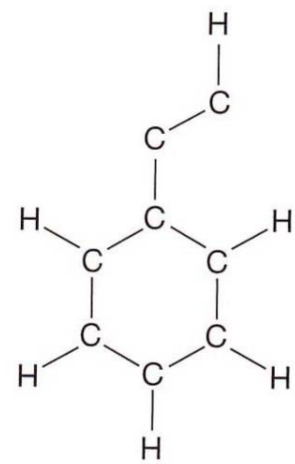


Problem 8(a – c): Complete the Lewis structures for each of the following molecules, using the information provided in table above. You may assume that no formal charges exist, and all H atoms are shown. Add only bonding pairs and lone pairs of electrons.

(a)



(b)



(c)