Drawing Lewis Structures

In the study of covalent compounds, drawing what is known as a Lewis Structure is a method often used to help to understand the bonding within a molecule of the compound. Once a "good" Lewis Structure has been drawn, a number of important things about the molecule can be deduced from this structure. For example, the nature of the individual bonds between atoms can be determined as well as the three-dimensional shape of the molecule. Although Lewis Structures are not exactly "pictures" of how the molecules might look if you could see them, they are very useful in chemistry. Drawing a "good" Lewis Structure is not usually difficult and one only has to follow a few simple rules to accomplish the task. I have summarized the steps in the process below.

Step 1

Before you can draw a Lewis Structure, you must know how the individual atoms are attached to each other. If you do not know or cannot guess which atoms are attached to which other atoms, you cannot draw the structure. Fortunately, it is often easy to guess the connectivity of the atoms of a compound. If that is not possible, you will be told how the atoms are connected. For example, in the compound formaldehyde (CH_2O), the carbon atom is attached to 2 different hydrogen atoms and to the oxygen atom as shown below.

Step 2

Once the connectivity is known, you must determine the total number of electrons in the valence shells of all of the atoms of the molecule. This is done simply by multiplying the Group A number from the Periodic Table by the number of atoms of that element found in the molecule. For formaldehyde the results yield 14 total electrons.

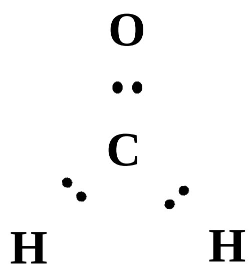
Hydrogen – 2 times 1 = 2
Carbon – 1 times 4 = 4
Oxygen – 1 times 6 = 6

Total Valence Electrons = 12

Step 3

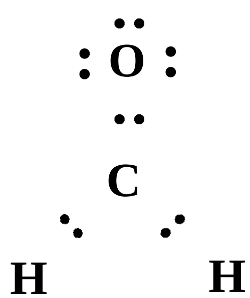
We begin the process of bonding by placing pairs of electrons between all atoms known to be bonded to one another (Step 1). As we use the electrons we keep a running total of how many electrons are remaining of the total we started with in Step 2.

Three pairs of electrons (6) used.
12 – 6 = 6 electrons remaining.



Step 4

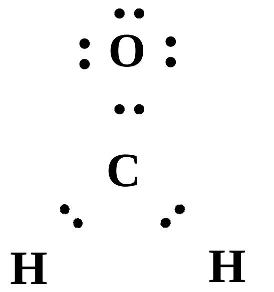
If any electrons remain after Step 3, they should be placed in pairs on the most electro-negative atoms (closest to F on the P.T.) first until each atom has a maximum of 8 electrons or until you run out of electrons. Never place additional electrons on hydrogen as it can accommodate only 2 electrons maximum. If you run out of electrons before you have eight on all atoms (except H), you must stop adding electrons and proceed to Step 5. If you have electrons left over, add them to the most electronegative atom of the compound up to a maximum of 12 total electrons. For our example of formaldehyde, your structure should look like this.



Step 5

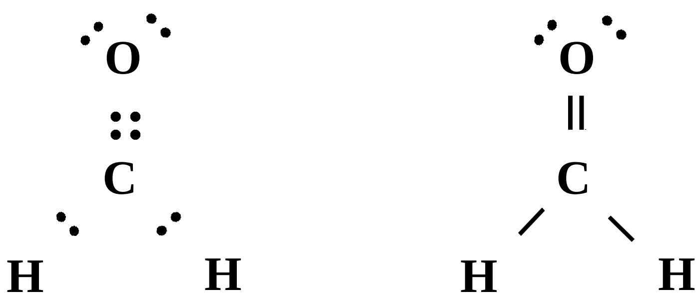
If, after placing all electrons on the atoms of the molecule, any atom other than hydrogen has fewer than eight electrons; you must move unshared pairs from adjacent atoms between the two atoms so that each atom has eight. You may have a maximum of 6 electrons (3 pairs) between any two atoms.

Only 6 electrons on carbon



Move one of the pairs on oxygen

When atoms share one pair of electrons we refer to this as a single bond. When two or three pairs are shared, we call these double and triple bonds respectively. Formaldehyde contains 2 single bonds between the carbon and the hydrogen atoms, and a double bond between the carbon and the oxygen. The oxygen holds 2 unshared pairs of electrons that are often referred to as "lone pairs."

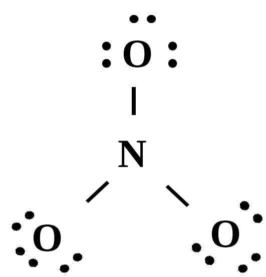


Special Notes

A dash is usually used to indicate shared pair of electrons. Some atoms can form stable molecules with fewer than 8 electrons around a given atom. Exceptions include Be, B, Al, and Ga. Atoms in Period 3 or higher can form compounds with up to 12 electrons around the central atom. Halogens (F, Br, Cl, and I) do not normally form double and triple bonds, only single bonds with 3 lone pairs on the halogen atom. In ions, the octet rule may be violated, even in atoms of Period 2.

Resonance

When one follows the rules for drawing Lewis structures, occasionally an interesting problem arises. An example is the structure for the nitrate ion. When you get to step 5, you are faced with a choice of atoms from which you might obtain unshared pairs of electrons to make a double bond. See the diagram below.



Any one of the three oxygen atoms could be used to supply an electron pair. Which one do you choose? Since all oxygen atoms are equivalent, it really doesn't matter which you choose. It is important to understand that the structure you will draw isn't very descriptive of the nitrate ion anyway. Lewis structures, like all models, have limitations and can only go so far to help us understand the complexity of nature. There is not really a "good" Lewis structure for nitrate as well as for many other compounds and ions. In nitrate actually all three oxygen atoms share electrons with the nitrogen atom and the structure can best be described using orbital theory, not Lewis structures. Another important thing about structures like nitrate is that they are more stable than would be predicted by the bonds described in Lewis theory. This extra stability is called resonance stabilization energy and nitrate is said to have resonance. Whenever you are faced with a choice of atoms from which to get electrons to make multiple bonds, you are likely to have a structure that displays resonance.

Cations, Anions, and Free Radicals

The drawing of Lewis structures for ions is important in organic chemistry as we often encounter ions in the descriptions of the mechanisms of organic reactions. We may actually redraw what may seem like a perfectly good Lewis structure to produce a species with a charge. This is often done to explain why particular parts of molecules are reactive while others are not. Free radicals (structures with a single unshared electron) can also be drawn. You will need to familiar with Lewis structures so that you can understand how to draw these structures and what can and cannot be done with these drawings. Be sure to learn the rules for drawing resonance structures and get comfortable with moving electrons around molecules.