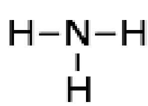
DRAWING LEWIS DOT STRUCTURES

Drawing a simple electron-dot structure

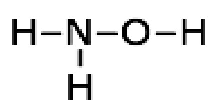
Drawing electron-dot structures is easiest if you follow the simple steps that are outlined below, using hydroxylamine, NH2OH, as an example.

1. Write out a simple structural diagram of the molecule in order to clearly show which atom is connected to which.

- If you are not sure of the bond connectivity, the structures of other similar molecules can often provide a useful clue. Thus if you know that ammonia, NH3, has nitrogen as its central atom, then you might recognize that hydroxylamine is just a derivative of ammonia in which one of the hydrogen's has been replaced with a hydroxyl group:



ammonia

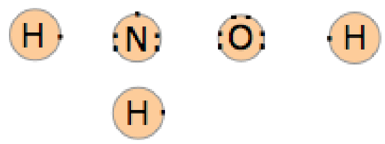


hydroxylamine

- It's also useful to know that the atoms hydrogen, nitrogen and oxygen are commonly connected to 1, 3, and 2 other atoms, respectively.
- The simple molecules for which electron-dot structures are drawn can often be thought of as consisting of one or more "central" atoms to which other atoms are attached.

The points outlined above will become part of your "chemical intuition" as you develop more experience in writing out structures from molecular formulas.

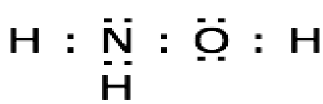
2. Draw electron-dot structures of the individual atoms in the molecule.



What you are doing here is showing how many valence electrons each atom contributes to the structure. We see that the total number of valence electrons is 14.

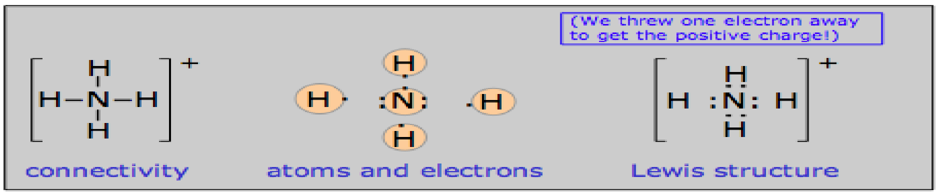
3. Bring the atoms together in a way that places eight electrons around each atom wherever possible.

Of the 14 valence electrons, 8 are needed to form the four covalent bonds in hydroxylamine, leaving six to be distributed as lone pairs. Placing these on the nitrogen and oxygen atoms as shown yields a structure conforming to the octet rule.

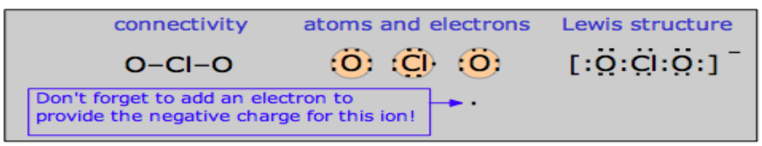


Electron-dot structures of ions

Ions are treated in basically the same way as neutral species, the only complication being that the number of electrons must be adjusted to account for the net electric charge of the ion. In other words, a negative ion contains more electrons than are provided by the valence shells of the constituent atoms, and a positive ion has fewer electrons than do the combining atoms. Here we show examples of the ammonium ionNH4+ which contains 9–1 = 8 electrons

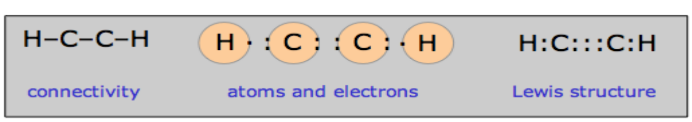


... and of the chlorite ion ClO2–, which contains 19+1 = 20 electrons:



Electron structures involving multiple bonds

The need for double or triple bonds usually becomes apparent when you have more electrons than can be accommodated in single bonds and lone pairs. Thus in the molecule acetylene (ethyne) C2H2, the single electrons contributed by the hydrogen atoms can be thought of as being paired up as one component of the carbon-carbon triple bond:



Source: http://www.chem1.com/acad/webtext/chembond/cb03.html