Drawing Lewis Structures

Steps and rules for drawing Lewis structures:

1) Count the total number of valence electrons
2) Put the least electronegative atom in the center. Note: H atoms are always terminal
3) Start by drawing single lines (bonds) between atoms. Each bond represents two valence electrons.
4) Add lone pairs (pairs of nonbonding electrons) to atoms. The goal is to have 8 electrons (an octet) around each atom. Exceptions to the octet rule:
   a) Exception 1: Hydrogen takes 2 electrons
   b) Exception 2: Atoms in row 3 of the periodic table and below can have more than 8 electrons (i.e. SF_6)
   c) Exception 3: Atoms in group II and III can have less than an octet (i.e. $BeCl_2$, BF_3)
   d) Exception 4: Odd-electron molecules will not have octets (i.e. NO, NO_2)
5) When adding lone pairs, the Lewis structure must show the same number of e^- calculated in step 1.
6) If you run out of valence electrons before you achieve octets on each atom, double or triple bonds can be made to satisfy the octet rule.

Example: Draw the Lewis structure for CO_2

Step 1:	Carbon	Group 4	4 valence electrons
	2 Oxygen	Group 6	2×(6 valence electrons)
	Total		16 electrons available for bonding

Steps 2-3:

Carbon is the least electronegative atom, so it goes in the center.

Drawing single bonds between carbon and each oxygen uses 4 electrons.

We have 16-4 = 12 electrons left.

Steps 4-5:

Try to satisfy the octet rule for all atoms using the remaining 12 electrons.
(note: 12 electrons = 6 lone pairs available)

Although this Lewis structure shows all 16 e^- in CO_2, carbon is 4 e^- short of an octet.
So, this is **<u>not</u>** the correct Lewis structure.

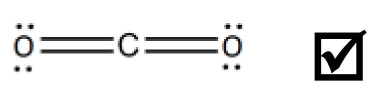
Step 6:

Redraw the Lewis structure with more bonds.
How do you know how many bonds to use?
The previous structure was 4 e^- short of having octets. Since each bond is worth two electrons, you need to draw two additional bonds.

Now you have used 8 electrons.
There are 16-8 = 8 electrons left.

Confirm Lewis structure:

Try to satisfy the octet rule for all atoms using the 8 remaining electrons.
(note: 8 electrons = 4 lone pairs available)



This structure obeys the octet rule and uses all 16 valence electrons in CO_2.
This is the correct Lewis structure.

How did I know to draw two double bonds instead of a single bond and a triple bond? See the back side of this sheet for more information and the answer to this question.

Formal Charge:

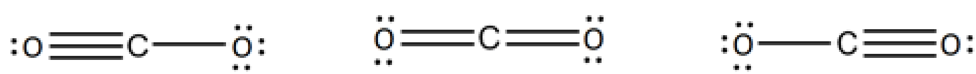
When drawing the Lewis structure for CO_2, we found that a total of four bonds were needed. We could have drawn any of the following Lewis structures:

All of these structures satisfy the octet rule and use 16 electrons. So, which structure is best? To determine this, you need to calculate formal charge (FC). Formal charge equals the number of valence electrons in the free atom minus the number of electrons surrounding the atom in the Lewis structure (note: only one electron from each bond is counted). An easy way to remember this is:

FC = Group number – Number of dots on atom – Number of lines on atom

Consider the CO_2 Lewis structure on the left. The formal charge for carbon = 4 – 0 dots – 4 lines = 0. For the left oxygen, FC = 6 – 2 dots – 3 lines = +1. For the right oxygen, FC = 6 – 6 dots – 1 line = -1.

You can also assign FCs for all atoms in the other two Lewis structures:



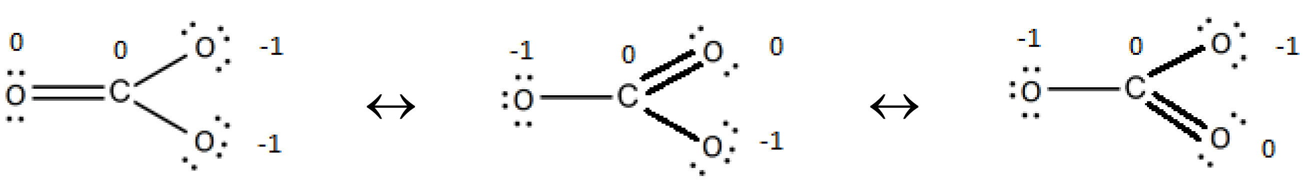
FC: +1 0 -1 0 0 0 -1 0 +1

Note: For neutral molecules, the formal charges on the atoms must add up to zero.

Generally, you want to make formal charge as close to zero as possible. Therefore, the Lewis structure for CO_2 with two double bonds is preferred over the Lewis structures with single and triple bonds.

Resonance structures:

Resonance structures exist when a double or triple bond can be moved around a molecule. For example, consider possible Lewis structures for the carbonate ion (CO_3^{2-}). Note: all formal charges have been assigned next to each atom.



All three of these structures are equivalent, and a hybrid (blend) of these structures is observed. Note: For ions, the formal charges must add up to the charge on the ion. In each structure for CO_3^{2-}, the FC's sum to -2.

Bond order:

Bond order can be calculated by dividing the number of bonds by the number of regions over which the bonds are shared. In the CO_3^{2-} ion, the bond order is 4/3 (four bonds are shared over three CO bonding regions). In CO_2, the bond order is 2 (four bonds are shared over two CO bonding regions). The higher the bond order, the shorter the bond. As such, the CO bonds in CO_2 are shorter than the CO bonds in CO_3^{2-}.