Lewis Structures

Tuesday, February 07, 2017 4:51 PM

Consider Hydrogen and Fluorine:

- How many valence electrons does each have?

  H=1 F=7

- How many more electrons are needed to complete the noble gas configuration?

  H=1 F=1

- From a Lewis perspective, why would hydrogen and fluorine make a covalent bond?

  each need to pair up an e⁻

- How is formal charge determined?

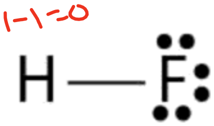
  F.C. = valence e⁻ - owned e⁻

- How is formal charged used to identify a reasonable Lewis structure?

  minimizing F.C. ensures a correct structure

- Here is the Lewis structure of HF, based on formal charge, is this structure reasonable:

  1-1=0

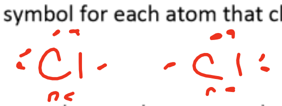


7-7=0

Yep

Consider Cl_2

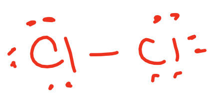
- Draw the Lewis symbol for each atom that clearly shows the number of valence electrons.



- How many electrons does each atom need to make a noble gas configuration?

  1 each

- Using this as a guide, draw the Lewis structure of Cl_2



Consider a molecule made from H and S:

- Draw the Lewis symbols for H and S. 
- How many bonds should S make to create a noble gas electron configuration?

  2

- How many bonds can H make?

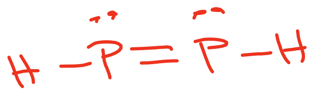
  1

- Using this as a guide, draw the Lewis structure for a compound made from H and S. You can use more than one of each atom if you need to.

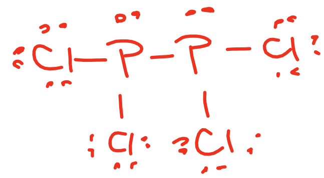
woops

Draw the Lewis structure of these double/triple bond containing molecules:

P_2H_2



P_2Cl_4



Which compound above has:

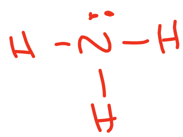
- The longest P-P bond? PCl4
- The weakest P-P bond? PCl4

single bonds are longer and weaker than double

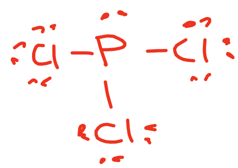
Draw the Lewis structure for each of the following:

H_2O

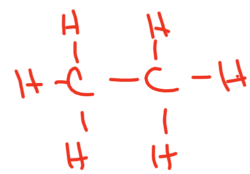
NH_3



PCl_3

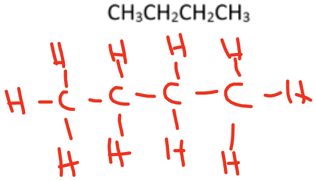


C_2H_6

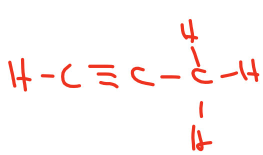


Draw the Lewis structure for each of the following structural formulas:

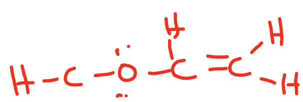
$CH_3CH_2CH_2CH_3$



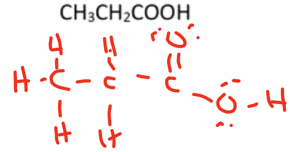
$CHCCH_3$



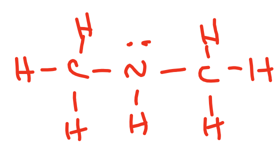
CH_3OCHCH_2



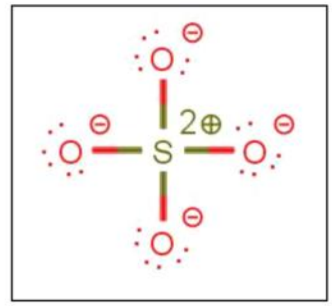
CH_3CH_2COOH



CH_3NHCH_3



Atoms in the 2nd shell are restricted to 8 valence electrons because the 2nd shell has a capacity of 8 electrons. However, atoms that exist in the 3rd shell (or larger) can have more than 8 valence electrons. Consequently, these atoms can expand their octet to make a more stable structure (because they can minimize formal charge).

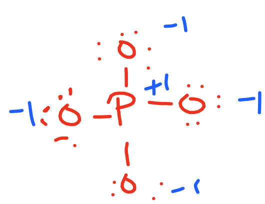


Consider sulfate (SO_4^{2-}). It is shown as we would draw it following the octet rule. Note the amount of formal charge. This makes the molecule very unstable.

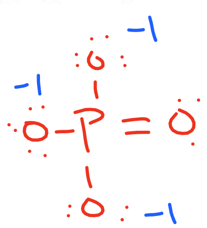
Since sulfur exists in the 3rd shell, it has the ability to expand its octet. By doing so, we can draw sulfate with much less formal charge (do this by taking a lone pair on one of the oxygens and forming a double bond). Note that the stable form of sulfate has two oxygen atoms with a formal charge of -1 instead of drawing the central sulfur with a -2. The latter structure is much less stable because of two reason:

1. Oxygen is more electronegative than sulfur so it is more likely to have a negative charge.
2. Distribution of formal charge is better than having it all placed on one atom.

Now you try. Draw phosphate (PO_4^{3-}) following the octet rule. Assign formal charge to each atom.



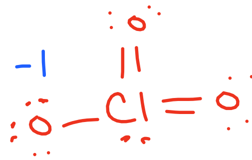
Now minimize formal charge by making double bonds with electrons that come from an oxygen lone pair.

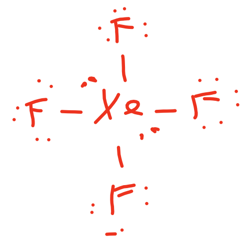


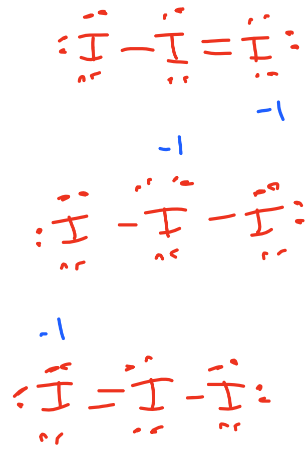
Assign formal charge to your molecule. Is it minimized and distributed? Is negative formal charge on the most electronegative atoms?

Now try drawing these molecules that include expanded octets:

SeF_6 ClO_3^{-1} XeF_4 I_3^{-1}

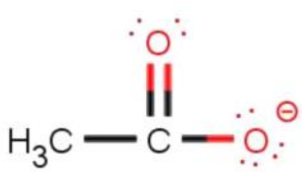


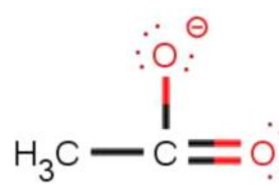




Some molecules can be drawn in multiple ways that are equally stable. These are called resonance structures.

Consider acetate. It is shown two ways that are equally stable: acetate has two resonance forms.

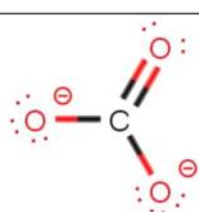


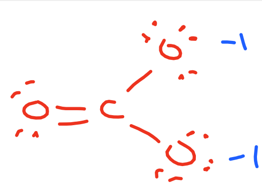


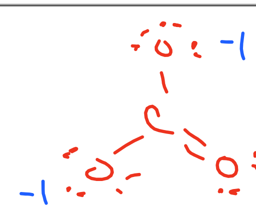
On average:

- Each oxygen has a charge of -½ → (50% * -1 + 50% * 0)
- Each C-O bond is a 1.5 bond (50% single and 50% double).

Look at the structure of carbonate below. Draw the other two resonance forms of carbonate.





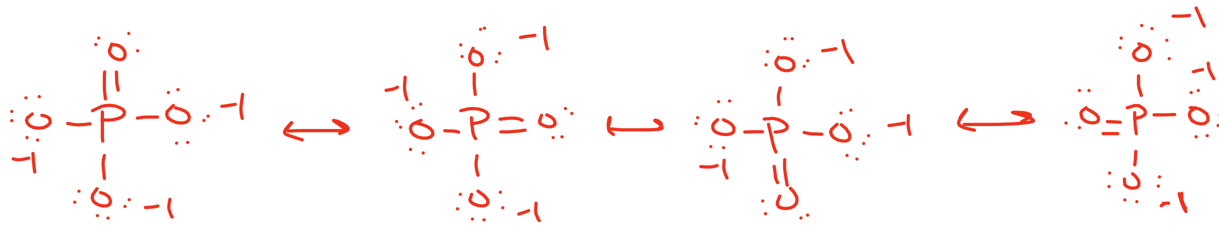


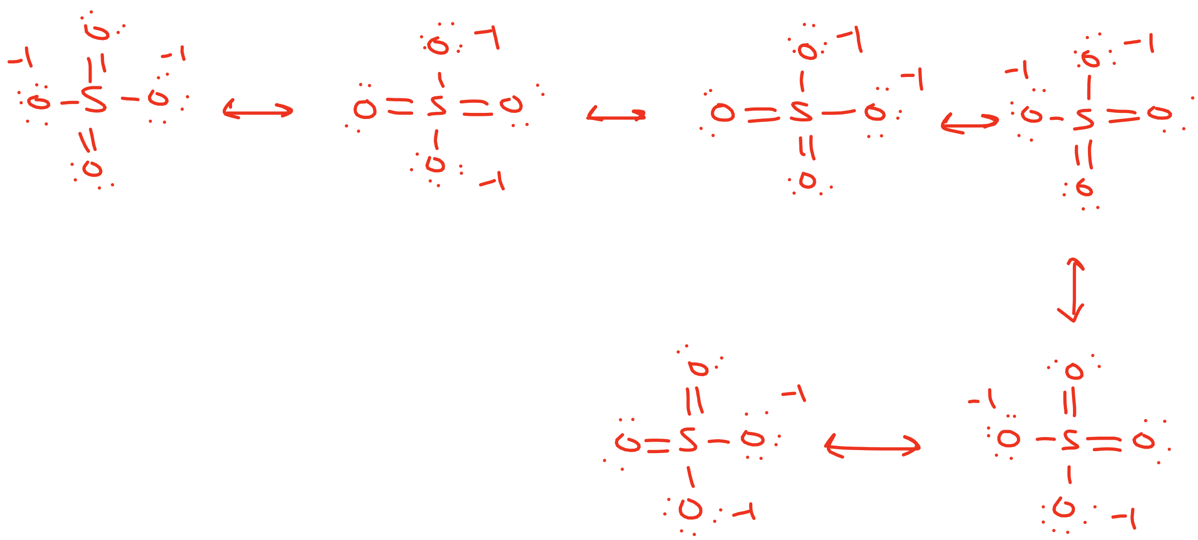
$\frac{1}{3}(0) + \frac{2}{3}(-1) = -\frac{2}{3}$

What is the average charge on each oxygen?

What is the average bond (e.g. single = 1, double = 2, etc.) $\frac{1}{3}(2) + \frac{2}{3}(1) = \frac{4}{3} = 1\frac{1}{3}$

Now try drawing phosphate (PO_4^{3-}) and sulfate (SO_4^{2-}) showing all of the resonance forms.





Still have time? Try these – show all formal charge and resonance forms.

SO_3 SO_3^{2-} HSO_3^{-1} ClO_2^{-1} ClO_4^{-1}