Lewis Electron Dot Structures

**Rules for writing Lewis Electron Dot Structures.** When applying these rules write each step on paper as you complete it. Do not trust yourself to do this in your head! Little mistakes have disastrous consequences!

1. Count all of the electrons in the valence (outermost) shells of all atoms in the molecule.
   a) For anions add to this number the same number of electrons as the net charge on the anion.
   b) For cations subtract from this number the same number of electrons as the net charge on the cation.
   c) If you find that you have an odd number of electrons be careful, this is not common. However, there are many important such cases. (eg, ClO, NO, CH₃, CH₂O, & CCl₃ These are called free radicals and are very reactive species!)

2. Arrange the atoms (atomic symbols) on the page so that each atom is adjacent to all other atoms to which it is covalently bonded. (If you are not certain you probably do not completely understand the use of chemical formulas or chemical nomenclature. Review this topic in your book.)

3. Place 2 electrons (a pair) between each pair of atoms that has a covalent bond. Then subtract this number of electrons from your initial total from step 1.

4. Determine which atoms are surrounded by 8 electrons (an octet) and those which are not. Determine which of the electron deficient atoms is most electronegative and complete its octet by placing additional electrons, in pairs, around the atom. You may place the electrons into the molecule as lone pairs or in bonds producing double and triple bonds. Proceed until each atom has an octet or there are no more electrons available. (If there are not enough electrons the best Lewis Structure may not give each atom an octet. However unfortunate this may be for the atom(s) in question do not take pity on the poor fellow and add electrons just to “make it look better”.)

5. Some atoms are able to accept 10 or 12 electrons by using d orbitals – those in the 3rd row or below of the periodic table. The most commonly encountered examples are: S, P, and I.

6. If you have placed all the available electrons around atoms in the molecule and do not have octets for each then it is important to check to see if you have maximized the sharing of electrons in multiple bonds. In atoms having C-O bonds, N-N bonds, S-O bonds, P-O bonds, N-O bonds and in similar cases where S substitutes for O, multiple bonds are common. If you have completed the octet of the more electronegative element without forming a multiple bond it may be possible to shift a lone pair of electrons and create a double bond thereby satisfying the octet of a less electronegative element to which it is bonded.

7. Determine the formal charge on each atom by the following method.
   a) Assign to each atom half of all the bonding electrons surrounding it.
   b) Also assign to it any lone pair electrons.
   c) Subtract this number from the balance nuclear charge for the atom (this is the group number); this is the number of electrons the neutral atom should have.
   d) Write the resulting Formal Charge adjacent to the atom.
e) **Check your work** – the sum of the formal charges on the atoms of a molecule must equal the net molecular charge.

8. It is often possible to write more than one Lewis structure according to the rules given above, each giving an octet to all atoms. Such structures are said to be resonance structures. Generally one is more favored than the others. *The most stable resonance form is the one having the smallest number of formal charges, thereby reducing the net separation of charge within the molecule.*

9. Resonance structures - When two different ways of arranging the electrons around a molecule can be drawn while the positions and formal charges of the atoms remain unchanged within the molecule, we say the two structures are related by resonance and are of equal energy. Resonance structures with few formal charges and complete octets are generally more stable than those without octets or those having significant charge separation within the molecule. Further, molecules which have two or more such resonance structures are generally more stable than comparable molecules which lack resonance.

Lewis Structures assignment

Draw complete Lewis structures for each of the following formulas including all resonance structures and formal charges.

**Inorganic** Name each.
- NH₃
- H₂O
- BH₃
- ClO₄⁻¹
- H₂CO₃
- CO₂
- SO₃
- SO₄²⁻
- H₂S
- SO₂
- N₃⁻¹
- HCN
- CO
- SiO₂
- HF
- PH₃
- OCl₂
- ICl
- NO₂
- NO₃⁻¹
- NO
- N₂
- O₂
- O₃
- PCl₅
- SF₆

**Organic** *(Don’t worry about all the names here)*
- CH₄
- C₂H₆
- C₂H₄
- C₂H₂
- \([C₃H₅]¹⁺\)
- \([C₃H₅]⁻¹\)
- CH₃
- CH₃N₂⁺
- CH₃COCH₃
- CH₃CONH₂
- CH₃PO₃(CH₃)₂
- \([CH₃PO₂]³⁻\)
- CH₃O
Draw Resonance Structure w/curved arrows
Two Resonance Structures