Lewis Structures: The Clark Method	Name:	
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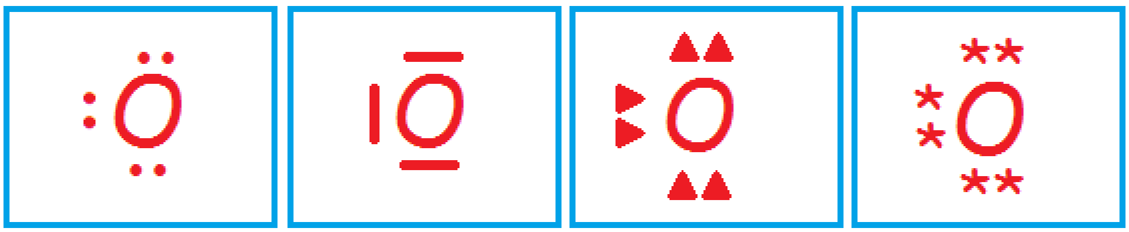
Lewis Structures

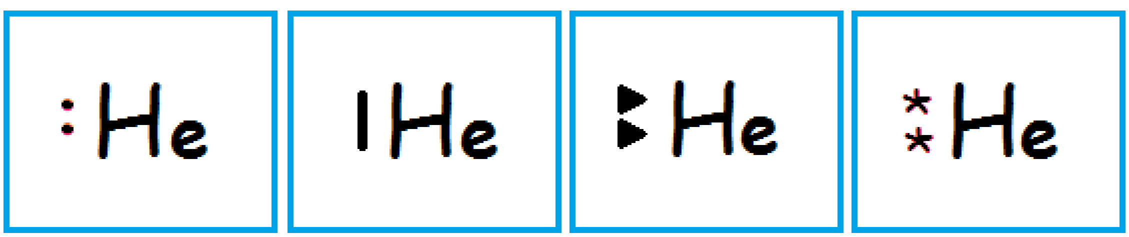
Lewis Structures, for all intents and purposes, utilize different symbols, e.g., dots, squares, "x's", triangles, ad nauseum, to represent the electrons in the outer valence shell of the atoms:



In the graphic above, the 6 valence electrons of oxygen are represented by dots, lines, triangles and asterisks. Note that the elemental symbol and the electron representatives color-match. This is crucial when determining electron alignment as you're drawing Lewis structures of more complex compounds or molecules.

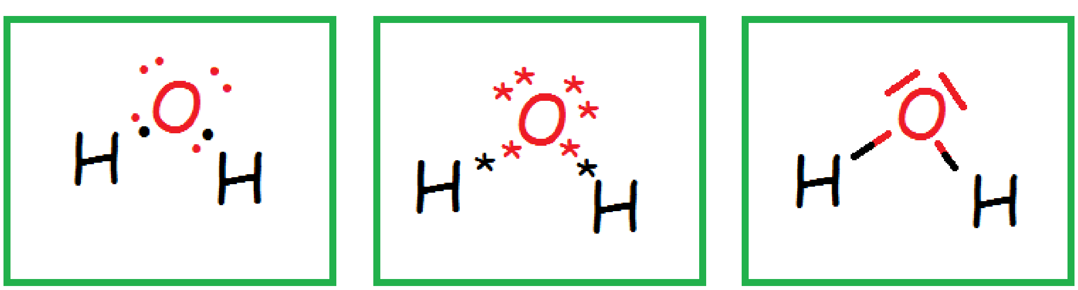
A note: there are those who use lines to represent 2 electrons as in the second images, above and below. **I discourage using this to mark valance electrons because students, often times, lose track of the electrons in the learning stages of Lewis structures**.

Helium's Lewis structures are represented, below:



Remember that **it takes 2 electrons to make a single bond**. Whenever possible, you need to place 8 electrons around your atoms (**octet rule**). At times you will only get 2 electrons around some atoms, e.g. H, Li, Na, K and that fulfills the octet (**duet**) rule, too, e.g., water, below:

Do you see how the oxygen in the water molecule, below, now has 8 electrons around it? That's the octet rule in action. Do you see how hydrogen has 2 electrons around it? That's the duet rule in action.



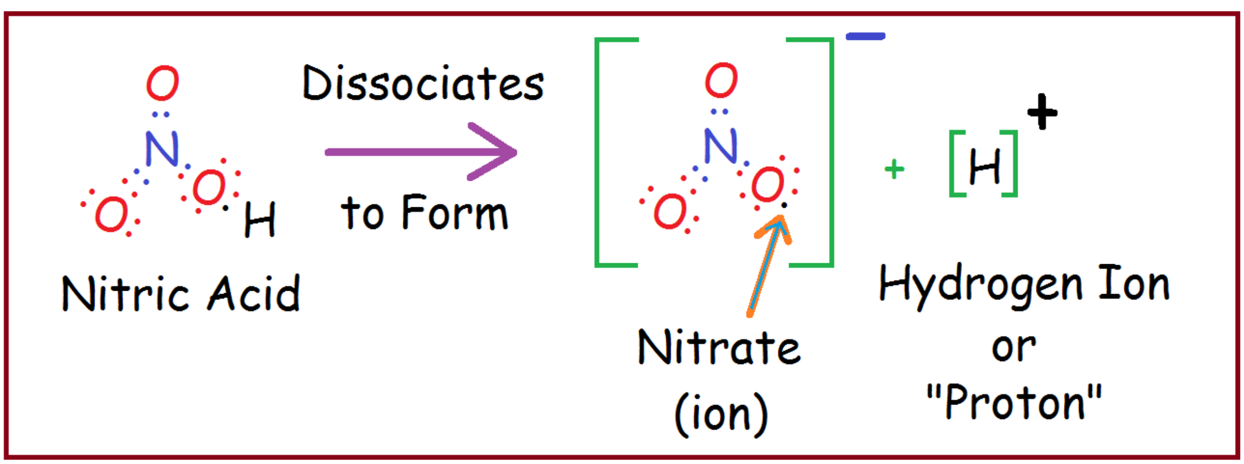
Fortunately for us, the first 20 of the representative elements are the easiest and the ones we'll focus on. Remember that these groups are in numerical order from I to VIII. **The number of the group tells you exactly how many electrons are in the outer valence shell of the atom with which you are working.**

Not only are we able to draw Lewis Structures of neutral compounds and elements, we are also able to draw Lewis Structures of polyatomic ions (be they **anions OR cations**). When it comes to the polyatomic ions (we're ignoring the simple mono-atomic ions, e.g., Cl^{-1} as they're easy to pick up after drawing the polyatomic ions), it's helpful to know their origin at the basic sense. The table, below summarizes some of these polyatomic ions and their compound of origin:

Polyatomic Ion	Polyatomic Ion Name	Compound of Origin	Name of Compound of Origin
NO_3^-	Nitrate	HNO_3	Nitric Acid
NO_2^-	Nitrite	HNO_2	Nitrous Acid
SO_4^{2-}	Sulfate	H_2SO_4	Sulfuric Acid
SO_3^{2-}	Sulfite	H_2SO_3	Sulfurous Acid
PO_4^{3-}	Phosphate	H_3PO_4	Phosphoric Acid
NH_4^+	Ammonium Ion	NH_3	Ammonia

Did you notice that the ammonium ion is a polyatomic cation? Good!

Let's use Nitric acid and the nitrate ion to illustrate the origin of the nitrate ion and to demonstrate each Lewis structure:

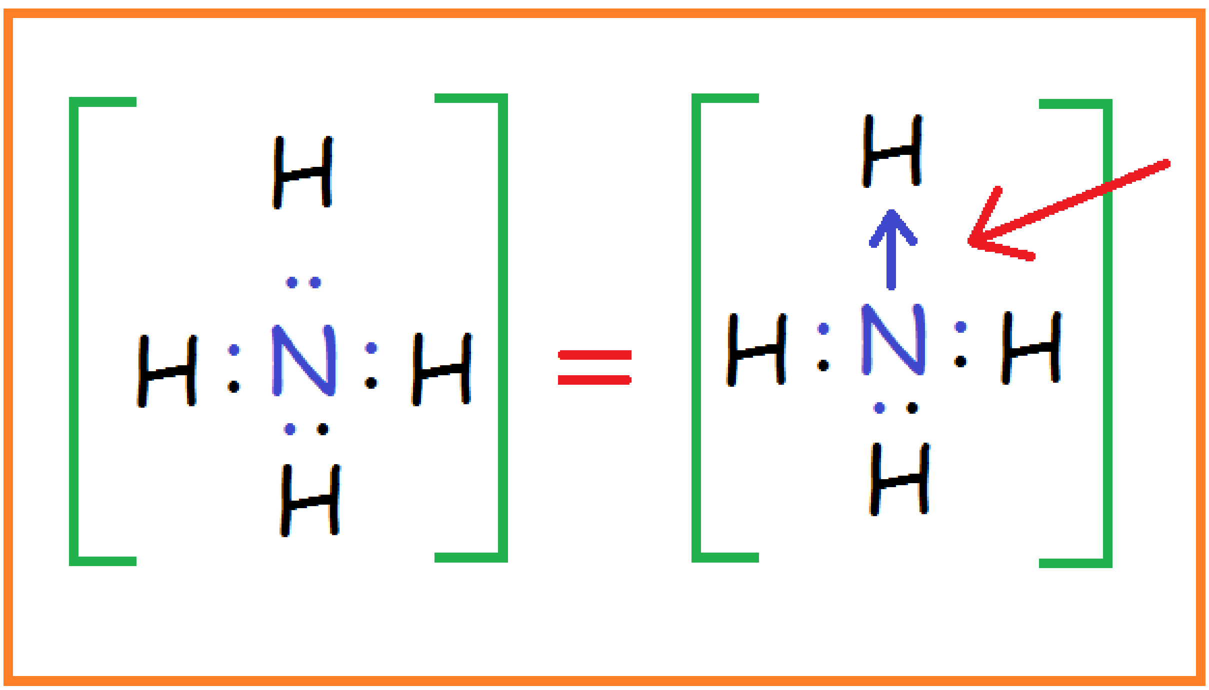


Note the adherence to both the octet and the duet rule in the nitric acid. Do you see the double bond between the N and the O in nitric acid? That's a double bond (remember from above that it takes 2 electrons to make one bond; in this case we have 4 electrons which make 2 bonds).

In the nitrate ion, do you see the orange and blue arrow? It's pointing to the electron that the H left behind when the acid dissociated (separated) in aqueous solution (a water based solution). That extra electron gives the nitrate ion its -1 charge (excess of 1 electron).

The hydrogen ion is often-times referred to as a "proton": that's all that's left from the H dissociation, i.e., **the H electron is "left behind" with the NO_3 to give the ion its charge (NO_3^{-1})**.

We can do the same with the other polyatomic anions, as well, and will go over them in class. What about the ammonium ion, though? It's a bit different:



The ammonium ion is a protonated ammonia molecule, i.e., a proton (hydrogen ion) reacts with ammonia in such a way that the paired electrons in the top of the N in each graphic, above, are donated to the proton. This sort of bond, **where one entity donates both electrons to make the bond is called a coordinate covalent bond**. We can represent it with an arrow drawn from the electron donor to the electron recipient (at the end of the red arrow).

How, then, do we know how to set up Lewis Structures to draw them correctly? It takes lots of practice. In the "old days" it took lots of experience, too. In about 1984, though, a fellow named Clark [1] came up with a quick and dirty method for determining how to draw Lewis Structures.

Let's start with a periodic table of the elements.

Periodic Table of the Elements

	IA	IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA
H																		He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg	IIIB	IVB	VB	VI	VIIB	VIIIB			IB	IIB	Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn

[2]

"*A Periodic Table is provided with Hi (red), Lo (gray) and Intermediate (green) Electronegativities indicated to aid in predicting ionic vs. covalent compounds.* [cf also [3]]

Recall that positively charged ions retain the element name in compounds, while negatively charged elemental ions are given the -ide ending. For covalent compounds the more positive (less electronegative) element is usually given the element name, while the more electronegative element is given the -ide ending.

In general we can predict how electrons distribute in compounds using the Hi-Lo-Intermediate electronegativity rules:

1. *Compounds composed of a "Hi" and a "Lo" element will be ionic with the respective ions combining to give a neutral compound.*
2. *All other combinations will give covalent compounds following the "octet rule" whenever possible."* [2]

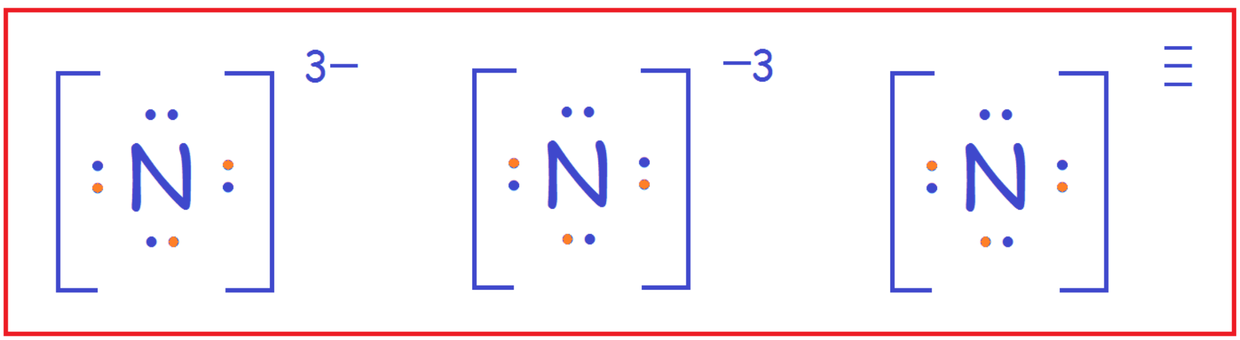
Clark's Method for determining bonding in covalent Lewis Structures in succinct form is as follows: Add up all of the valence electrons in the structure.

1. If $\Sigma e^- = 6y + 2$ (where y = # atoms other than H), then the octet rule is followed with single bonds only.
2. If $\Sigma e^- < 6y + 2$ (where y = # atoms other than H) then the ion or molecule probably has multiple bonding.
3. The number of multiple bonds = the difference between step 2 and the total number of electrons divided by 2 (remember a triple bond is 2 multiple bonds!).

4. If you can draw more than one structure, then choose the most symmetrical.
5. **Remember to add one electron for each negative charge in anions, or subtract one electron for each positive charge in cations**.

Let’s try this out with the nitride ion:

- The **-ide** ending indicates nitride is the ***elemental anion*** *(negatively charged ion)* for nitrogen, so we will have to ***add*** electrons.
- Looking at the Periodic Table, N is in Group 5, so valence electrons on N = 5, so will add 3 electrons to get the nitride ion with a -3 charge (alternately, we can determine the charge by subtracting 8 from the group number, 5 - 8 = -3).
- Arranging the resulting eight outer (valence) electrons into four pairs, and placing them on the four sides of a square around the symbol for nitrogen, N and "decorating" it, we get:



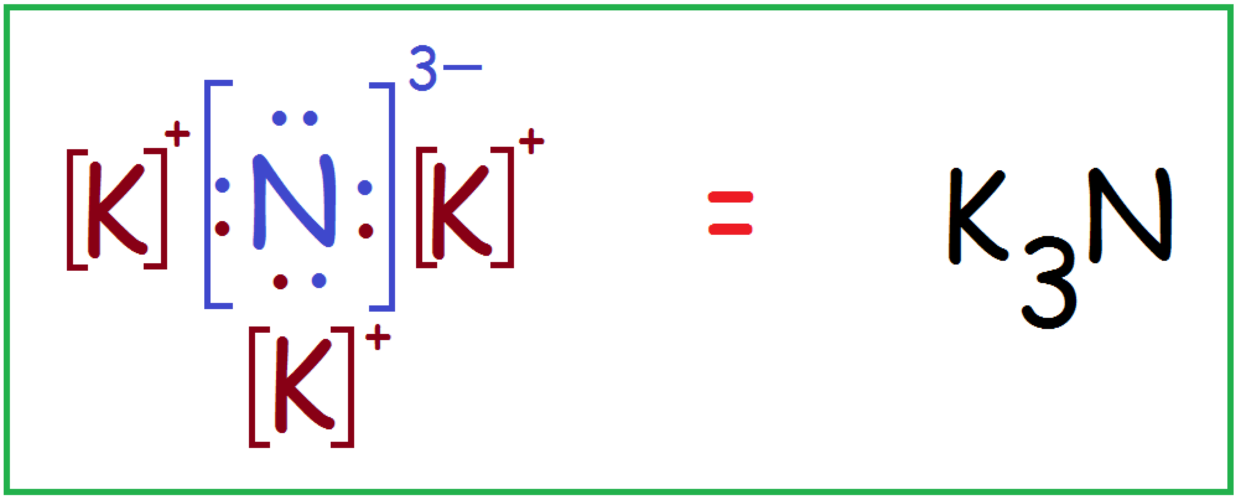
- Note that the ion has a full octet of electrons (four pairs) and that **the structure is enclosed in brackets to demonstrate that the anion "owns" all of the electrons**, and that the charge is distributed over the entire ion.

Note the orange electrons. These are accepted by nitrogen to fill its outer-most shell to generate the anion. Also note the 3 different ways in which to denote the negative three charge – each way is equally as acceptable.

Now, let’s try this with potassium nitride. Looking at the Periodic Table, K is Lo and N is Hi, therefore potassium nitride is an ionic compound;

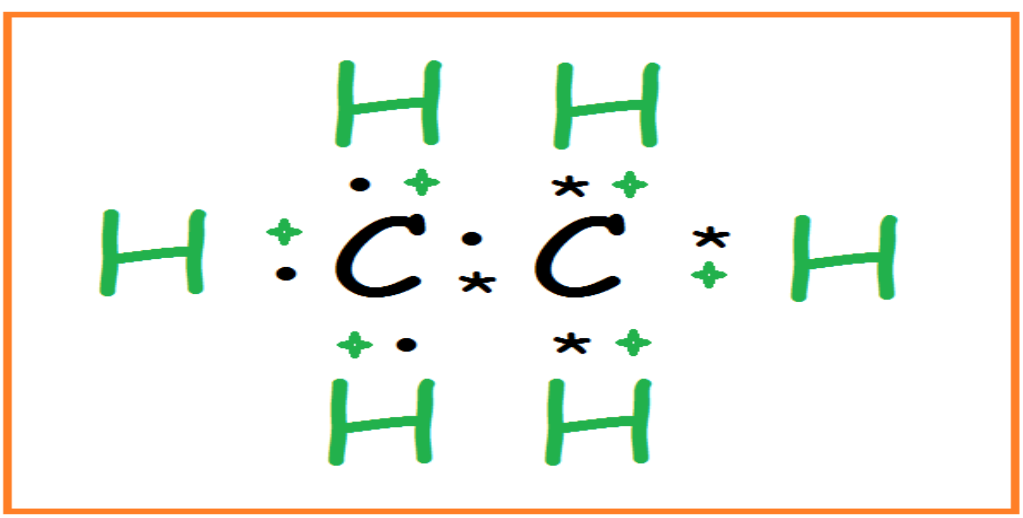
1. The valence electrons on K = 1 (in Group I), so K will lose an electron to give K^+ ion.
2. Nitrogen is in group V as we saw above, so N will gain 3 electrons to give us the N^{-3} ion.

3. Note that to get a balanced compound that the charges do not equal zero: +1 -3 = -2. To get a neutral compound, positive and negative charges have to balance out to zero, hence, we need 3 K^+ ions (+3-3 = 0).



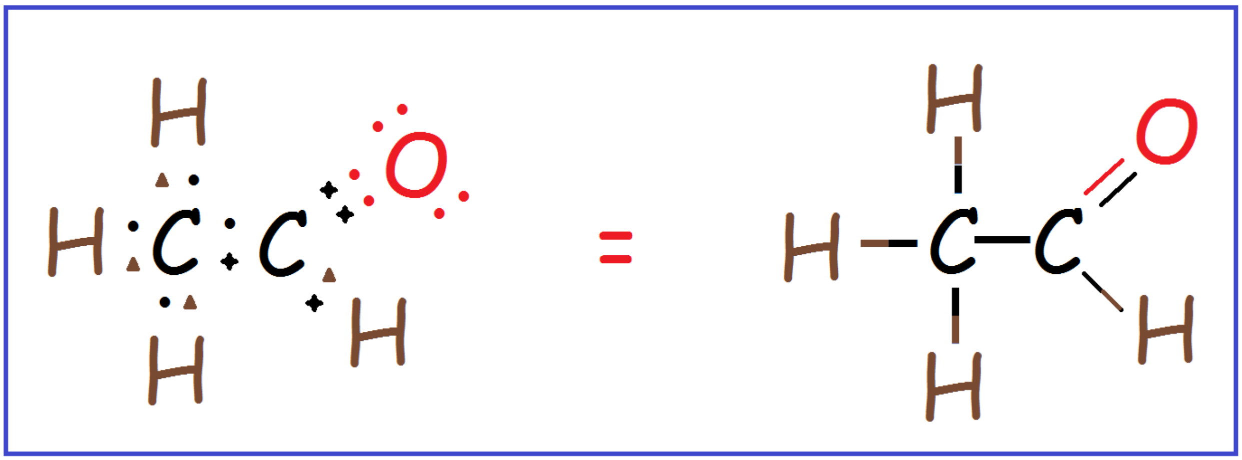
Now, let's try this with ethane (C_2H_6). By looking at the periodic table (or the table of electronegativity [3] values), we can see that this is a covalent molecule.

- Valence electrons = 2x4 (C) + 6x1 (H) = 14
- From Clark's Method need 6y + 2 = (6 x 2) + 2 = 14 electrons to fulfill the octet rule with only single bonds (**2 because there are 2C's**);
- Comparing with the valence electrons, the molecule has the identical number of electrons required for all single bonds, hence, (14 – 14)/2 = 0 multiple bonds:



Do you see how having different colors and symbols help with electron "flow" in these diagrams? Thus far, we have dealt with simple ions or compounds. What about compounds of a covalent (electron sharing) nature with multiple bonds? Let's try this method with acetaldehyde (ethanal). Looking at the Periodic Table, C and H are Intermediate while O is Hi, therefore acetaldehyde is covalent.

- Valence electrons = 2x4 (C) + 4x1 (H) + 1x6 (O) = 18
- From Clark's Method need 6y + 2 = (6 x 3) + 2 = 20 electrons to fulfill the octet rule with only single bonds (**3 because there are 2C's and 1 O**);
- Comparing with the valence electrons, the molecule has 2 less electrons than required for all single bonds, hence, (20 – 18)/2 = 1 double bond:



Experimental

Using these rules and examples, as well as information gleaned from lecture and any outside sources (including, but not limited to, books in the library), **during your lab period**, you are to name the compound or ion, draw and explain how you derived the Lewis structures in the space provided (front and back if needed) for the following compounds (**you may work in 2's or 3's but no larger groups**):

H_2SO_4	HNO_2	H_3PO_4	PO_4^{3-}	NO_2^{-1}
NH_3	H_3AsO_4	HCl	HCN	H_2SO_3
Ca_3N_2	H_2O	CO_2	CO	H_2CO_3
CO_3^{2-}	HCO_3^{-1}	Li_2S	Na_3P	Mg_3N_2

Remember to have your completed work signed off before you leave the lab for the day.

References

1. Clark, Thomas J.: *J. Chem. Educ.*, 1984, *61* (2), p 100

2. http://users.humboldt.edu/rpaselk/C107.F09/Chem_Discuss/LewisStruct.html, accessed 19 August 2014

3. http://www.drcarman.info/kem121lx/mekanx.pdf, p. 54, *et seq,* accessed 19 August 2014

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