One Page Lesson: Drawing Lewis Structures

The atoms that need the *most* electrons to achieve a noble gas electron configuration (and thus form the *most* covalent bonds) should be used as *central* atoms. The atoms that need very *few* electrons to achieve a noble gas electron configuration (and thus form very *few* bonds) should be used as *peripheral* atoms.

Examples of central atoms:

- Carbon has *FOUR* valence electrons. It needs *FOUR more* electrons to achieve an octet. It gets those four electrons by *sharing*: by forming *FOUR covalent bonds. CARBON'S octet: FOUR BONDS and NO LONE PAIR.*
- Nitrogen has *FIVE* valence electrons. It needs *THREE more* electrons to achieve an octet. It gets those three electrons by *sharing*: by forming *THREE covalent bonds. NITROGEN'S octet: THREE BONDS and ONE LONE PAIR.*
- Oxygen has *SIX* valence electrons. It needs *TWO more* electrons to achieve an octet. It gets those two electrons by *sharing*: by forming *TWO covalent bonds. OXYGEN'S octet: TWO BONDS and TWO LONE PAIRS.*

Examples of peripheral atoms:

- Fluorine has *SEVEN* valence electrons. It needs *ONE more* electron to achieve an octet. It gets that one electron by *sharing*: by forming *ONE covalent bond. FLUORINE'S octet* (and that of the other halogens): *ONE BOND and THREE LONE PAIRS.*
- Hydrogen has *ONE* valence electron. It needs *ONE more* electron to achieve a duet. It gets that one electron by *sharing*: by forming *ONE covalent bond. HYDROGEN'S duet: ONE BOND.*

Indicate the *covalent bonds* between atoms with *lines*. Indicate *lone pairs* of electrons (unshared electrons) with *dots*. Arrange the bonds and lone pairs so that *each* atom achieves its noble gas electron configuration (octet, or duet for hydrogen). Be sure to draw *NO MORE* and *NO FEWER* electrons than the *TOTAL* number of valence electrons that the atoms bring with them to the molecule.

- A "*single bond*" is *one pair* of electrons shared between two atoms, and is represented in the Lewis structure by a *single line.*
- A "*double bond*" is *two pairs* of electrons shared between two atoms, and is represented by drawing *two lines* between the bonding atoms. Double bonds are most common between atoms of carbon, nitrogen, oxygen, and sulfur. (In organic and biological molecules, there are many examples of carbon-carbon, carbon-nitrogen, and carbon-oxygen double bonds.)
- A "*triple bond*" is *three pairs* of electrons shared between two atoms, and is represented by drawing *three lines* between the bonding atoms. Triple bonds are most common between carbon and nitrogen atoms.

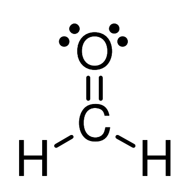
Example Lewis Structure: Formaldehyde, CH_2O

The *TOTAL* number of valence electrons the atoms bring with them to the molecule:

1 carbon atom with 4 valence electrons	4
2 hydrogen atoms, each with 1 valence electron	2(1)
1 oxygen atom with 6 valence electrons	6
Total available electrons	12

Consider the bonding patterns of our component atoms: Carbon is always a central atom because it forms *four bonds.* Oxygen tends to form *two bonds* and *two lone pairs.* Hydrogen atoms form *one bond.*

Since carbon forms the most bonds, we'll place it at the center and attach the oxygen and hydrogen atoms to it. Since the carbon needs *four* bonds, and the oxygen *two* bonds, we'll place a double bond between them. Two lone pairs of electrons are added to the oxygen to complete its octet. The total number of valence electrons the atoms brought into the molecule is 12. This Lewis structure accounts for *ALL* 12 of those electrons: two electrons in *each* of the four bonds (eight bonding electrons), plus two lone pairs (four unshared/non-bonding electrons).



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