Chem 1010/1800 Tip Sheet

Lewis Structures

Covalently bonded molecules are often represented as Lewis dot structures, where electrons (e^-) are denoted by lines (for bonds) and dots (for lone pairs). Knowing how to properly draw these structures lets us predict many things, such as molecular geometry and bond strength.

> **Octet rule:** most elements like to be surrounded by 8 electrons (an octet) so that they have a complete valence shell. These electrons can either be in lone pairs or in bonds. Carbon, nitrogen, oxygen, and fluorine must always follow the octet rule.

How to draw Lewis dot structures:

1) Determine the central atom. It is usually the least electronegative element that is not hydrogen.

    e.g. HCN: carbon is less electronegative than nitrogen, so carbon is the central atom. Hydrogen can never be the central atom because it can only have two electrons in its valence shell, and can therefore only form one bond.

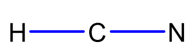
2) Count the total number of valence electrons in the molecule. Tip: do this by looking at each element's group number on the periodic table (e.g., C is in group 14, has 4 valence electrons; Cl is in group 17, has 7 valence electrons).

    e.g. HCN: 1 e^- from H + 4 e^- from C + 5 e^- from N = 10 e^- total

Note: if the molecule is charged, this will change the electron count. A -1 charge will add an electron, and a +1 charge will remove an electron.

3) Draw a line (single bond) from the central atom to each surrounding atom. Count the number of electrons that must still be accounted for.

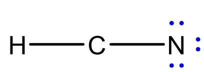
    e.g. HCN:



4/10 electrons have been used – 6 electrons still needed

4) Complete the octet of any outer atoms using lone pairs. Count the number of electrons that must still be accounted for. If there are remaining electrons, put them on the central atom as lone pairs.

    e.g. HCN:



10/10 electrons have been used – no more needed

For more information or to book an appointment

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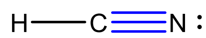
5) If the central atom does not have a complete octet, use electrons from the outer atoms to form double or triple bonds until it does. Recall that every line represents 2 electrons in a bond, so a double bond will contain 4 electrons and a triple bond will contain 6 electrons.

e.g. HCN:



two of nitrogen's lone pairs were moved to become bonding e⁻

6) You're finished! Double check that all elements have a full valence shell (generally a complete octet) and that the correct number of electrons was accounted for.

Useful Tools

1) Recognize patterns. Many <u>neutral</u> elements tend to have a pattern when it comes to predicting how many bonds and lone pairs they will have. If you remember the following pattern, you can often draw structures more quickly by doing steps 3 to 5 at the same time:

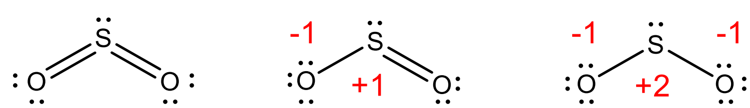
Preferred number of...	C	N	O	F
Lone pairs	0	1	2	3
Bonds	4	3	2	1

Note: this pattern is <u>only</u> followed when the above elements have a formal charge of zero.

2) Consider formal charges. Atoms in a covalent molecule will not always be neutral (i.e., their formal charge might not be zero). You will often find that there is more than one way to configure the electrons in a molecule, and need to decide which structure is the most stable. If you are unsure, calculate the formal charge of each atom with the following formula:

$$\text{formal charge} = (\#\ \text{valence electrons}) - (\#\ \text{lone pair electrons}) - (\#\ \text{of bonds})$$

To decide which structure is best, always avoid having non-zero charges when possible. If you need to have a non-zero charge, avoid charges that are bigger than ±1. If you need to have a negative formal charge, this should generally be assigned to the most electronegative element in the molecule. Example:



best worst

3) Not all elements follow the octet rule! Sometimes this is obvious, such as with hydrogen and helium, which can only hold 2 electrons in their valence shell. However, there are other notable exceptions: boron will often only have 6 electrons, and elements in the 3rd row or lower (e.g., P, Si, S, Br, etc.) can support more than 8 electrons – this is called an expanded octet.

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