A Lewis Structure is a very simplified representation of the valence shell electrons in a molecule. It is used to show how the electrons are arranged around individual atoms in a molecule. Electrons are shown as "dots" or for bonding electrons as a line between the two atoms. The goal is to obtain the "best" electron configuration, i.e. the octet rule and formal charges need to be satisfied.

Note

Lewis structure does NOT attempt to explain the geometry of molecules, how the bonds form, or how the electrons are shared between the atoms. It is the simplest and most limited theory on electronic structure.

How to draw Lewis Diagrams

The following is an example of how to draw the "best" Lewis structure for NO$_3^-$ (learning by example).

1. First determine the total number of valence electrons in the molecule. This will be the sum of the group number a of all atoms plus the charge.

   \[
   \begin{array}{c|c}
   \text{N} & 5 \\
   \text{O (x 3)} & 18 \\
   \text{charge} & 1 \\
   \hline
   & 24
   \end{array}
   \]

2. Draw a skeletal structure for the molecule which connects all atoms using only single bonds. The central atom will be the one that can form the greatest number of bonds and/or expand its octet. This usually means the atom lower and/or to the right in the Periodic Table, N in this case.

   \[
   O-N-O \\
   \text{O}
   \]

3. Now we need to add lone pairs of electrons. Of the 24 valence electrons available in NO$_3^-$, 6 were used to make the skeletal structure. Add lone pairs of electrons on the terminal atoms until their octet is complete or you run out of electrons.

   \[
   :O-N-O:: \\
   :O: \\
   \]

4. If there are remaining electrons they can be used to complete the octet of the central atom. If you have run out of electrons you are required to use lone pairs of electrons from a terminal atom to complete the octet on the central atom by forming multiple bond(s). In this case the N is short 2 electrons so we can use a lone pair from the left most O atom to form a double bond and complete the octet on the N atom.
5. Now you need to determine the **FORMAL CHARGES** for all of the atoms. The formal charge is calculated by:

\[(\text{group number of atom}) - (\frac{1}{2} \text{ number of bonding electrons}) - (\text{number of lone pair electrons})\], i.e. see the figure below.

\[
\begin{align*}
\text{O} & : & 6 - 1 - 6 & = -1 \\
\text{N} & = & 6 - 2 - 4 & = 0 \\
\text{O} & : & 5 - 4 - 0 & = +1
\end{align*}
\]

No Lewis structure is complete without the formal charges. In general you want:

- the fewest number of formal charges possible, i.e. formal charges of 0 for as many of the atoms in a structure as possible.
- the formal charges should match the electronegativity of the atom, that is negative charges should be on the more electronegative atoms and positive charges on the least electronegative atoms if possible.
- Charges of -1 and +1 on adjacent atoms can usually be removed by using a lone pair of electrons from the -1 atom to form a double (or triple) bond to the atom with the +1 charge. **Note:** the octet can be expanded beyond 8 electrons but only for atoms in period 3 or below in the periodic table. In our present example N can not expand beyond 8 electrons so retains a formal charge of +1, but the S atom below can expand its octet.

6. You have determined the "best" Lewis structure (octets completed and lowest formal charges) for NO$_3^-$, but there are a number of ways to show this structure. Although it is most common to use a line to indicate a bonding pair of electrons they can be shown as electrons, see the left most image below. It is also common to show only the net charge on the ion rather than all of the formal charges, i.e. see the right most figure below.
Why are there different ways for the "same" Lewis structure? It depends what you want to show. While the most complete structure is more useful for the novice chemist, the simplest is quicker to draw and still conveys the same information for the experienced chemist. You should learn to recognize any of the possible Lewis structures.

Contributors and Attributions

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