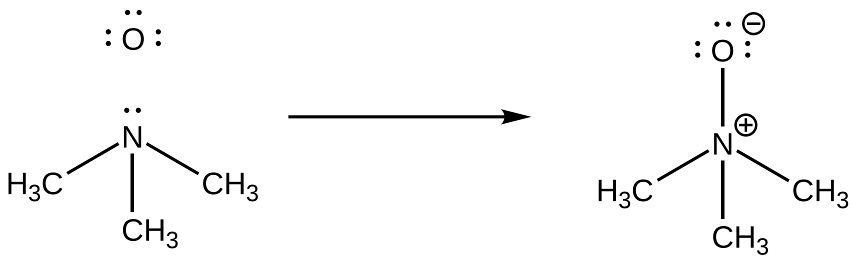
Formal Charges in Lewis Structures

Skills to Develop

- Determine and illustrate formal charges for Lewis structures

When you draw Lewis structures, sometimes the electrons are shared in a way which seems "unfair." For instance, in $(CH_3)_3NO$, to give N 8 electrons (and not more, since N can't have more than 8), you have to draw a single bond to oxygen. If you imagine a reaction between $(CH_3)_3N$ and an oxygen atom, both electrons that form the bond to O come from N (the former lone pair). (This is a rare example of a reaction that is both a Lewis acid-base reaction and a redox reaction.) However, once the bond is made, these electrons are shared. If you imagine that they are shared equally, then there is a single positive charge on N and negative charge on O. These are called **formal charges**.



The Lewis acid-base reaction to form trimethylamine oxide, a molecule with formal charges.

How are Formal Charges Different from Oxidation Numbers?

You might remember oxidation numbers from the discussion of redox chemistry earlier. Oxidation numbers are found by assuming that the bonding electrons are entirely owned by the atom that pulls hardest. Formal charges, in contrast, are calculated by assuming that the bonding electrons are shared evenly between the two atoms. The truth is usually somewhere in between. Both formal charges and oxidation numbers are used for "bookkeeping" or counting purposes. They don't tell you much about the real position of the electrons in the bond.

How Should You Find Formal Charges?

To find formal charges in a Lewis structure, for each atom, you should count how many electrons it "owns". Count all of its lone pair electrons, and half of its bonding electrons. The difference between the atom's number of valence electrons and the number it owns is the formal charge. For example, in NH_3, N has 1 lone pair (2 electrons) and 3 bonds (6 electrons total, so count 6/2 =3), so it owns 5 electrons, which is the same as the number of valence electrons. The formal charge is 0. For each H atom, it has 1 bond and thus 1 electron, so its formal charge is also 0. This is good, because all the formal charges of each atom must add up to the total charge on the molecule or ion. For the ammonium ion, NH_4^+, each H is still 0. Now N has 4 bonds and no lone pairs, so it owns 4 electrons. 5 - 4 = +1, so N has a +1 charge. This matches the +1 charge of the whole ion.

What Do You Need to Do with Formal Charges?

When you write Lewis structures, you should include formal charges next to each atom with a formal charge that isn't 0. (Usually, you circle the charge so it's clear.) This can also help you tell which Lewis structures are good. Usually negative formal charges should be on atoms that pull electrons strongly (like O or F, elements from the top right of the periodic table that have high ionization energies and high electron affinities). Positive formal charges should be on elements that pull electrons less. Big formal charges (more than 2) are usually bad. And it's better to have opposite formal charges right next to each other (so you get a formal "ionic bond"), and like formal charges farther from each other.

Outside Links

- CrashCourse Chemistry: Atomic Hook-ups (10 min)
- Khan Academy: Formal charges and dot structures (12 min)

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