Drawing Lewis Structures the Jasperse/Edvenson Way Using Formal Charge

The system of drawing Lewis structures described here depends heavily on the recognition, assignment, and application of formal charge. Thus it is vital that you become proficient at handling the concept of formal charge!

Formal charge is the hypothetical charge that an atom in a molecule would have if all of the bonding electrons in the atoms were shared evenly.

The formal charge is the number of electrons in a neutral isolated atom(the group number using the American numbering system) minus the number of electrons assigned to the atom in the Lewis structure, assuming again that the electrons in bonds are shared evenly.

In a covalent bond in which two electrons are shared evenly between two atoms, only one of the two electrons is assigned as "belonging" to each of the two atoms. On the other hand, an unshared pair is assigned as two electrons belonging entirely to the atom on which it resides. Therefore the total number of electrons assigned to any atom in a molecule is the number of bonds plus the number of unshared electrons.

Formal charge = group # – "assigned" electrons
or
Formal charge = group # – (bonds + unshared electrons)
or
Formal charge = group # – unshared electrons – bonds
or
group # – formal charge = unshared electrons + bonds

A practical way to determine the formal charge for an atom is to count the bonds and unshared electrons(assigned electrons) around an atom and compare that number to the group number for the particular element. If the number of assigned electrons is equal to the group number of the element, then the formal charge of that atom is zero. If the count is greater than the group number, then the atom must have a negative formal charge equal to the number of "excess" electrons. If the count is less than the group number, then the atom must have a positive formal charge equal to the number of "missing" electrons.

The formal charges for common bonding situations are worth memorizing. You shouldn't need the formula to recognize that an oxygen with 2 bonds has zero formal charge, and that an oxygen with only 1 bond has a –1 formal charge.

FORMAL CHARGE

# of bonds	C	N	O	F
4	0	+1		
3	–1	0	+1	
2		–1	0	
1			–1	0

Notice that being one bond short normally gives formal charge of -1; having one bond extra normally gives formal charge of +1.

STEPS FOR DRAWING LEWIS STRUCTURES

1. Connect the atoms with single bonds.

2. Give octets to the outside atoms by adding unshared pairs of electrons. (This may be unnecessary, if you remember that oxygens with one bond, for example, will come out to be -1).

3. Assign the formal charges of the outside atoms, then deduce the formal charge necessary on the central atom based on the overall charge of the chemical formula.

> Net charge of formula = the sum of all the individual formal charges
> so:
> F.C. on central atom = net charge– formal charge on outside atoms

4. Add lone pairs to the central atom as needed to give it the formal charge deduced in step 3.

Formal charge = group # – unshared electrons – bonds

A more useful way of thinking about this relationship is is the following.

group # – formal charge = unshared electrons + bonds

Example: If you are trying to make the formal charge of a central S atom equal to zero, then the number of unshared electrons + bonds must equal six(the group number of S). If there are already four bonds to the central sulfur, then it still needs an unshared pair to get up to six assigned electrons.

5. If the central atom is short of an octet, form double (or triple) bonds by changing outside atom lone pairs to bonding pairs. (It may exceed an octet, in which case don't worry).

6. With all electrons now in place, rewrite formal charges.

Notes:
1. The above protocol works for nonmetal atoms, groups 4-8. The octet rule may be broken for metal/metalloid atoms from groups 1-3.
2. For row 2 nonmetals, the octet rule must never be violated (either too many or too few). However, for elements from periods 3-5, more than eight valence electrons is not a problem.