Lewis Structure Practice

ORGANIZATION

- Mode: Practice in pairs as assigned by the instructor. Demonstrate determining a structure on the board in class.
- Grading: Lab Notebook and Post-Lab Questions
- Safety: None

GOAL:

To be able to draw Lewis structures, calculate formal charges, identify resonance structures and predict the "best" Lewis structures for a molecule or an ion.

I: BACKGROUND

In this lab you will practice the skills that you developed in the pre-lab exercise. Your instructor will assign you and your partner structures to draw, and groups will take turns showing their work on the whiteboard.

Rules for Determining Lewis Structures

1. Add all of the valence electrons for all of the atoms—call this number "available" for the total *available* electrons for the molecule.

2. If the molecule is a cation, subtract one or more electrons from the available number. If the molecule is an anion, add one or more electrons to the available number.

3. Add up how many electrons are needed for all atoms to have an octet/duet—call this number "needed", because it is how many electrons are *needed* in the entire molecule.

4. Subtract: *Needed* − *Available* = *Shared*; **Shared** is the number of electrons that will be shared between atoms. In other words, this is the number of electrons that will be in bonds.

5. After you have drawn the skeleton with the number of bonds between the atoms, fill in the other "available" electrons as lone pairs, starting with the outer atoms first. Each bond (line between atoms) represents 2 shared electrons. Every atom should have an octet/duet after all electrons are filled in. (Sometimes the central atom will not have an octet—see exceptions below). You must use all "available" electrons, so check back on the final Lewis structure to make sure the number of electrons represents the number of "available" electrons.

6. When deciding between two possible Lewis structures, examine the formal charges of the atoms in each structure (see below).

Exceptions to the Octet Rule

1. *Octet-Deficient Molecules*—Covalent compounds of Group 3A atoms (B, Al, Ga) often have 6 electrons, rather than 8, in stable molecules. Example—BF_3
2. *Octet-Expanded Molecules*—When atoms from Period 3 or below are central atoms in a molecule, they can accommodate more than 8 electrons. Example—SF_6
3. *Odd-Electron Molecules*—Some molecules are stable with one lone electron (known as a free radical). Example—NO_2

Assigning Formal Charge

$$\begin{aligned} \text{formal charge } (q) = & \text{ valence electrons } (VE) \\ & - \text{non-bonding electrons } (NBE) \\ & - \frac{1}{2} \times \text{ bonding pair electrons } (BE) \end{aligned} \tag{1}$$

Assign the formal charge of each atom in a molecule to determine the feasibility of that Lewis structure. The molecule with the lowest formal charges is the best structure (ideally 0 on all atoms). If there must be a non-zero formal charge, negative formal charges go on the most electronegative elements.

II: EXERCISES

You will be assigned several Lewis structures to draw. At least one will be an ion, and at least one will require resonance structures. You will be given the name of one compound, and you will then have to determine its chemical formula and structure. You will be asked to demonstrate how to draw one of these structures.

For each molecule you must perform the following steps and be able to explain them.

1. Determine the total number of electrons in the molecule.
2. Identify the central atom.
3. Identify whether the central atom can have an expanded octet.
4. Go through the steps outlined in the pre-lab to construct a Lewis structure. Be sure to include all lone pairs.
5. Draw any additional resonance structures.
6. Label all atoms with their formal charges and demonstrate how they were calculated.
7. Explain why you believe this to be the best Lewis structure for this molecule.
8. Determine the bond order for all bonds in the molecule.