**CHE-2060 Lab: Comprehensive Lewis structure review**

**Covalent bonding**

In inorganic chemistry you learned a bit about covalent bonding: bonds formed when two atoms each contribute an electron to form a shared pair of electrons that holds the two atoms together. Atoms form covalent bonds in order to fill their valence shells. Remember that a full valence shell is eight atoms in the outmost electron shell for all atoms except hydrogen and helium, satisfied with just two valence electrons.

* Valence electron number is an atom’s column (or group) number.
* Lewis structures: Free valence electrons are represented as dots, while covalent bonds (a pair of shared valence electrons) are represented by lines between the atoms contributing to the pair.

Example: Neither hydrogen (I) nor carbon (IV) have full valence shells. However when they share their valence electrons all atoms achieve a full valence shell. The molecule shown here is methane, CH4.

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***Problem 1:*** Prove that all atoms in the structure of methane have a full valence shell by counting the number of valence electrons shared by each atom in the molecule’s structure.

Each hydrogen atom has a single bond or two valence electrons, satisfying its desire for a ‘duet’ of two valence electrons.

The carbon has four single bonds for a total of eight valence electrons, satisfying its desire for an octet of eight valence electrons.

**Common bonding patterns**

It’s easier to draw Lewis structures, and to begin to understand and visualize atomic and molecular geometry, if you can recognize common bonding patters of the atoms found in organic molecules.

Common bonding patterns for **uncharged atoms** are show here. Note these are the ‘biological’ bonding patterns of S and P.



Common bonding patterns for **charged atoms**:

* Note that S and P can have more than an octet; S, P and other large atoms, are exceptions to the octet rule. In these exceptional cases, S and P will be charged.
* The formal charge of atoms is calculated by subtracting the sum of ‘dots + sticks’ from the atoms valence number: FC = ve- - (dots + sticks).



**Steps for writing Lewis dot structures:**

1. Sum all valence electrons for all atoms in the structure, including charge.
2. Write atomic symbols and connect each with one line. Single atoms are central and repeated atoms surround the center.
3. Complete the octets of peripheral atoms (those connected to the central atom).
4. Place any leftover valence electrons on the central atom.
5. If there are not enough valence electrons to complete the central octet, use some electron pairs to create multiple bonds to the central atom.



**Isomers** are two different molecules that share a molecular formula. Their structures, physical and chemical properties differ.



**Calculating formal charge:**

An atom’s formal charge depends on its bonding pattern. For each atom in a Lewis dot structure calculate formal charge using this ‘formula’:

**formal charge = (# valence electrons) – (dots + sticks)**



**Determining bond polarity:**

For each bonded pair of atoms in the Lewis dot structure, calculate the difference in electronegativity values for those two atoms:

**difference in electronegativity = |atom 1 EN – atom 2 EN|**



Notice that the difference is an absolute value, so sign is not an issue. To determine which bond is the most polar compare differences. The largest difference indicates the most polar bond.

**Resonance:**

A molecule has resonance if its bonding pattern (aka arrangement of electrons) can be drawn in two or more equivalent ways. This generally occurs when there are two or more repeated or equivalent atoms and when a double bond can be placed in more than one position without changing the atom-to-atom bonding pattern. All possible resonance structures are shown connected by a double-headed arrow that indicates that they are resonance structures. Some resonance structures may be more stable than others and these are said to be ‘preferred’, ‘predominant’ or ‘major’ while less stable structures are said to be ‘minor’ contributors. Stability is often determined by lower or fewer formal charge(s). Resonance hybrids are a more realistic depiction of the phenomenon of resonance and spread double bonds across all possible locations.

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***Problem 2:*** Calculate the formal charges of the common bonding patterns or charged atoms shown on page 1.

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| --- | --- | --- |
| **atom** | **bonding pattern** | **formal charge** |
| H | No bonds, no free pairs | 1 – (0+0) = +1 |
| H | No bonds, one free pair | 1 – (2+0) = -1 |
| C | Three single bonds, no free pairs | 4 – (0+3) = +1 |
| C | Three single bonds, one free pair | 4 – (2+3) = -1 |
| N | Four single bonds, no free pairs | 5 – (0+4) = +1 |
| N | Two double bonds, no free pairs | 5 – (0+4) = +1 |
| N | One single, one triple bond, no free pairs | 5 – (0+4) = +1 |
| O | Three single bonds, one free pair | 6 – (2+3) = +1 |
| O | One single, one double bond, one free pair | 6 – (2+3) = +1 |
| O | One double bond, three free pairs | 6 – (6+2) = -2 |
| S | Three single bonds, one free pair | 6 – (2+3) = +1 |
| S | One single, one double bond, one free pair | 6 – (2+3) = +1 |
| S | One double bond, three free pairs | 6 – (6+2) = -2 |
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***Problem 3:*** For each of these formulas, CH4O and CH5N:

1. Draw the Lewis dot structure.
2. Calculate formal charge for each atom.
3. Find the most polar bond.

***Problem 4:*** For each of these formulas, C2H6, C3H8 & C2H7N:

1. Draw the Lewis dot structure.
2. Calculate formal charge for each atom.
3. Find the most polar bond.

***Problem 5:*** For each of these formulas, N2, C2H4 & C2H2:

1. Draw the Lewis dot structure.
2. Calculate formal charge for each atom.
3. Find the most polar bond.

***Problem 6:*** Make models of each of these and submit photographs of them.

1. H2 & N2
2. C2H6 vs. C2H4 vs. C2H2

***Problem 7:*** Consider C2H6O.

1. Draw two different Lewis dot structures for this molecule: isomers.
2. Calculate formal charge for each atom in each structure.
3. Find the most polar bond in each structure. Does polarity change?
4. Build a model of each molecule. Can you change the shape of the either molecule without breaking or making new bonds? If so, describe what happens.

***Problem 8:*** Consider NO2-1, CO, CO2, NO+1, NO3-1, and C6H6 (the carbons form a ring).

1. Draw the Lewis dot structure.
2. Calculate formal charge for each atom.
3. Find the most polar bond.
4. If the molecule has resonance, draw all possible resonance structures and indicate if any are major or minor.
5. If the molecule has resonance, draw the resonance hybrid.

