

IMPORTANT:

Test 3 is postponed to Tuesday May 6 because there is a test (post-test) at the beginning of next lab. The post-test is similar to the pre-test taken at the beginning of the semester. It should be good practice for the test 3 and the final exam.

Stuff on Exam 3: Ch.6.5-6.9, 7.1-7.5, 8.1-8.3, 8.5-8.6 (relevant homework assignments include 10, 11, and 12)

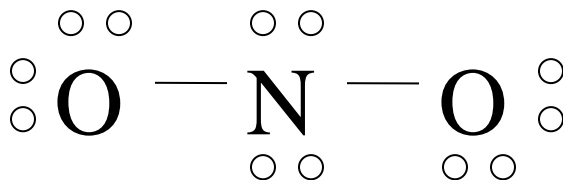
Answers to Group Problem #7

a) Lewis structure for NO_2^+

1) First, you have to figure out which atom in the molecule has the lowest **electron affinity** (this is how much an atom desires an electron). See figure 8.6 on page 312 of the book for electron affinity values of elements. The element with the lowest value goes in the middle when drawing the Lewis structure:

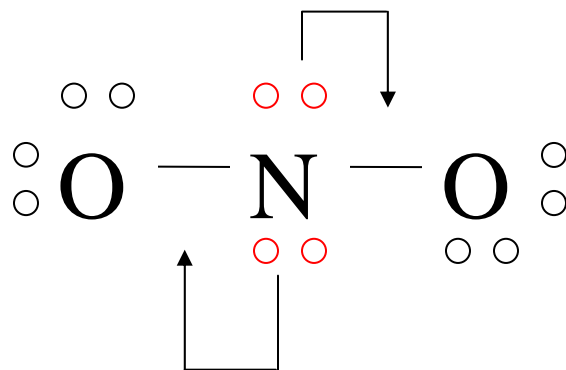


2) Add the rest of the valence electrons as paired dots on each atom. The total number of electrons used in this bond is 5 (from N valence shell) + 12 (from the two O atoms) – 1 (because the whole molecule has a positive charge) = 16 electrons. So:



*Remember that a straight line connecting two atoms equals one shared pair of electrons (2 electrons). So, 2 straight lines and 12 dots = 16 electrons.

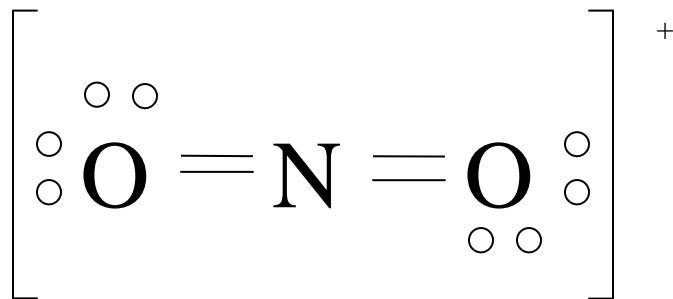
3) However, the octet rule is not satisfied. Each atom must have valence electrons. In the figure above, the left oxygen has 6 electrons, the nitrogen has 8 electrons, and the right oxygen has 6 electrons. For each oxygen atom to have 8 valence electrons each, nitrogen can share its two unshared pairs of electrons; one pair to each oxygen:



OR

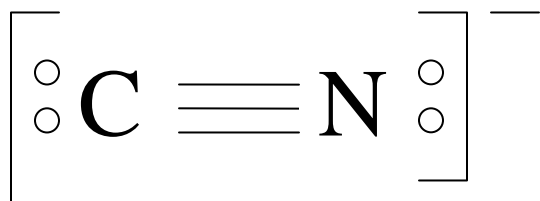


4) All the atoms have 8 electrons for their valence shells but there is one more step before the Lewis structure is complete. Since this is a polyatomic ion with a positive charge, there must be brackets around the whole thing with a plus sign indicating the charge:

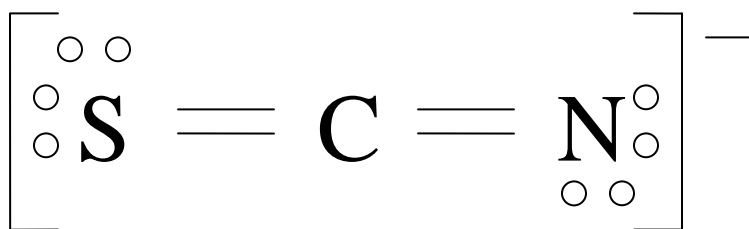


b) Write the Lewis structure for CN^-

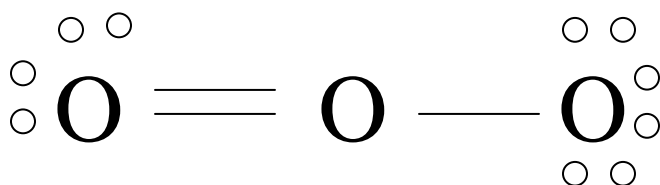
Follow the same rules to get:



c) Lewis structure for SCN^-



d) Lewis structure for O_3

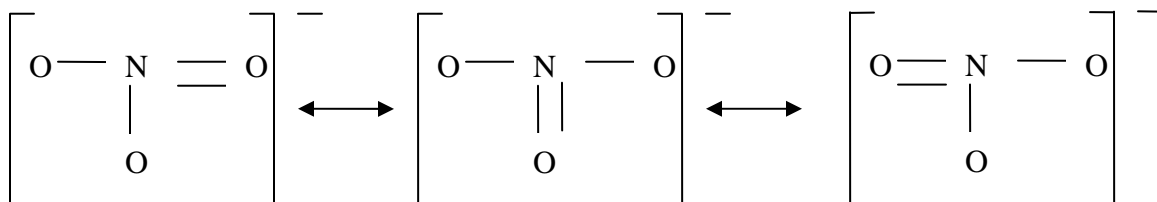


-Lewis structures can be used to predict:

- 1) Resonance
- 2) Bond
- 3) Bond length
- 4) Formal charge

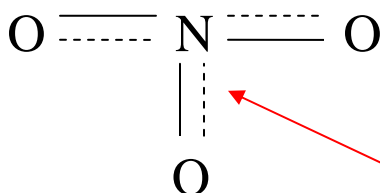
Resonant Structures

Example: (These figures omit the unshared electron pairs because it's too hard to draw them in a typed document, but you should never omit the unshared electron pairs when you draw a structure)



-All of these are nitrate ions but the location of the double bond is different in each example.

-We can picture the nitrate ion as:



-This is not a Lewis structure, but it can be used to emphasize **bond order**. The average bond order can be found by dividing the # of bonds by the # of locations. For the nitrate example above:

$$\frac{\text{\# of bonds}}{\text{\# of locations}} = \frac{4}{3}$$

Bond Strength and Length

-Strength: The greater the bond order, the stronger the bond, or the more energy it requires to break the bond

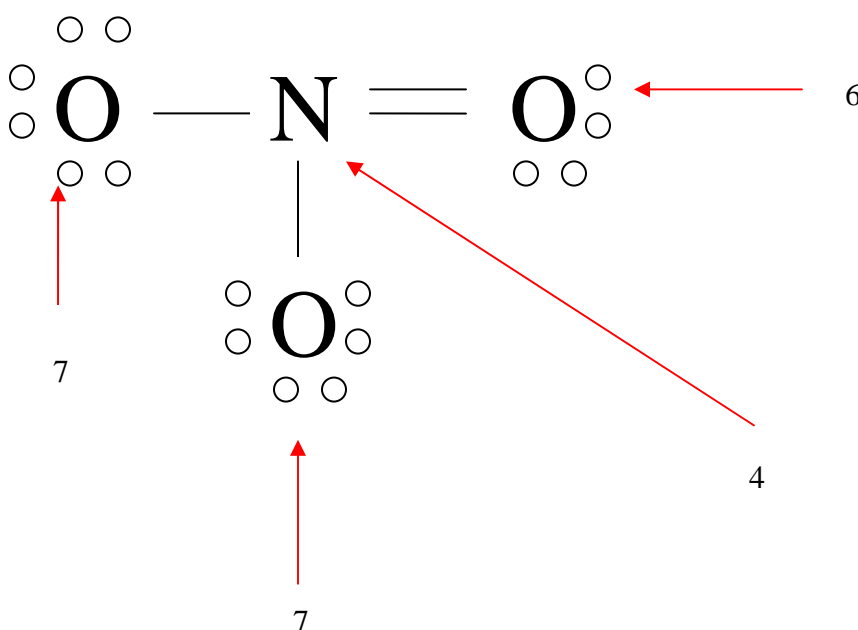
-Length: The greater the bond order, the shorter the bond length
(Lecture 22, page 6 has examples of bond strength and length)

Formal Charges

-**formal charges** are comparisons between the original valence electrons contributed by an atom and the electrons it looks like the atom would have if all bonds were broken and electrons reassigned democratically.

-formal charges are **NOT** oxidation numbers. They are only used to figure out which resonant structure is the most stable.

-the formal charge is found by subtracting the democratically reassigned electrons* of an atom from the original valence electrons of that atom. In the nitrate ion,



oxygen (left) = $6 - 7 = -1$

oxygen (bottom) = $6 - 7 = -1$

oxygen (right) = $6 - 6 = 0$

nitrogen = $5 - 4 = +1$

*democratically reassigned electrons = when the covalent bond splits, each atom receives one electron from a pair

Formal Charges and Alternative Structures

-If there is more than one Lewis structure that exists for a molecule or ion, the most stable form of it is the one where the formal charges make the most sense...following these guidelines:

- 1) negative formal charges are on the atoms with the highest electron affinity
- 2) positive formal charges are on the atoms with the lowest electron affinity