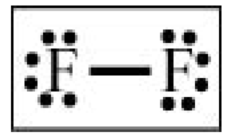
Name ____________________ Honors Chemistry ___/___/___

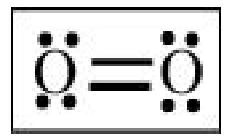
Lewis Structures & Resonance Structures

Last chapter we studied ionic compounds. In ionic compounds electrons are gained or lost. In this chapter we are going to study covalent compounds. In a covalent compound, atoms **share** electrons. Just like ionic compounds, though, all atoms need eight electrons to fill their outer energy level. But, whereas elements would gain or lose electrons in an ionic compound, in covalent compounds atoms will **share** between two and six electrons in order to reach their goal of eight.

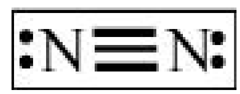
- One exception is **hydrogen**, which only needs **two** electrons to fill its outer energy level.
- **Beryllium** will form compound where it only has **four** electrons.
- **Boron** will form compounds where it only has **six** electrons in its structural formulas.
- These electron deficient compounds are very reactive and will react with other substances to achieve an octet.
- Second-row elements **carbon, nitrogen, oxygen** and **fluorine always obey the octet rule** and must have 8 electrons in their structural formulas. **Second row elements never exceed the octet rule**, since their valence orbitals (2s and 2p) can accommodate only 8 electrons.
- Third row and heavier elements often satisfy the octet rule but can exceed the octet rule by using their empty valence d orbitals. (ex: PCl_5, SF_6)
- When writing the Lewis structure for a molecule, satisfy the octet rule for the atoms first. If electrons remain after the octet rule has been satisfied, then place them on the central atom.



The most common covalent bond is a **single covalent bond**, the bond that is formed when **one pair of electrons is shared between two atoms**. Fluorine, F_2, forms a single covalent bond. This bond is represented by a single dash drawn between the two element symbols. The dots drawn on each symbol represent unshared electrons. They are almost always written in pairs.

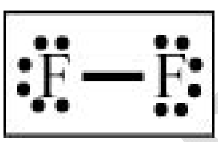


When **4 electrons are shared**, a **double bond** is formed. The double bond is represented by a pair of dashes drawn between the element symbols. Oxygen, O_2, forms a double bond.

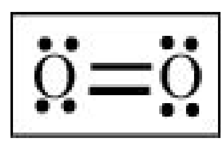


When **6 electrons are shared**, a **triple bond** is formed. A triple bond is represented by three dashes drawn between the element symbols. Nitrogen, N_2, forms a triple bond.

Observe that in each of the above examples each element has 8 electrons. Remember, a dash means that both elements connected by that dash are sharing those two electrons. So, when counting electrons for each element you must remember to count those electrons for both elements. Each dot represents one electron and they are possessed only by the element on which they are drawn.



.For example, look at F_2 from above. Each fluorine atom shares 2 electrons, represented by the single dash between them. Since each element has two electrons, each needs 6 more electrons to reach 8 and have a full outer energy level. Each atom gets the needed six from the six dots drawn around its symbol.



In O_2 there is a double bond. That means that each of the oxygen atoms shares 4 electrons. Thus, each atom still needs 4 more electrons to get to 8 and have a full outer shell. The last four electrons four each element are in the form of unshared pairs of electron, drawn as 4 dots on each element.

Nitrogen, N_2, follows the same pattern. Six electrons are shared between the two elements and the last two are drawn as dots representing unshared pairs of electrons.

Now, there is a specific procedure for writing structural formulas. It is not just a matter of drawing dashes and dots. The procedure for writing structural formulas is outlined below.

Part 1: How to Write Lewis Structures

Write the structural formula for diatomic nitrogen, N_2

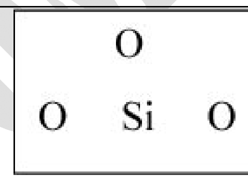
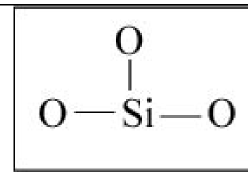
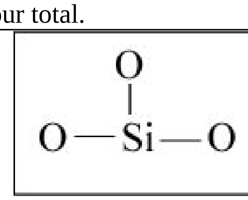
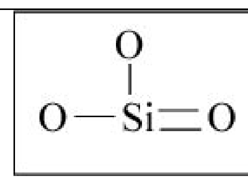
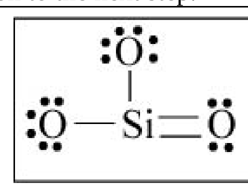
Description of Action	Action
1. Determine how many electrons you have to work with; this is done by multiplying the number of valence electrons for that element by the number of those atoms. Label your total as "HAVE".	**1.** N_2 nitrogen's valence # of nitrogen **HAVE** 5 x 2 = **10** So, we have 10 electrons with which to work.
2. If the given molecule is an ion, add or subtract electrons from your total, depending on the charge of the molecule. If the molecule has a positive charge, subtract electrons from your total. If the molecule has a negative charge, add electrons to your total.	**2.** HAVE N_2 5 x 2 = 10 N_2 is not an ion so we do not have to add or subtract electrons to or from our total.
3. Write the symbols for the given formula with some space between them. If there are more than two atoms, put the one with the **lowest electronegativity** in the center and the others on each of the four sides. **Exception:** Hydrogen can NEVER be in the center. An element's electronegativity can be found on the front of your periodic table.	**3.** HAVE N_2 5 x 2 = 10 N N
4. Draw one dash between the symbols to represent a bond. Each dash drawn represents 2 electrons. Subtract 2 from your "HAVE" amount for each bond you draw.	**4.** HAVE N_2 5 x 2 = 10 N—N **- 2** **8**
5. Each element in a covalent compound wants to share 8 electrons. We have to check to see how many electrons each element needs to reach 8. Every bond that is drawn to a symbol should be counted as 2 electrons for that element. Add these numbers together and label this number as "NEED".	**5.** HAVE N_2 5 x 2 = 10 N—N **NEED** **- 2** **8** 6 6 = **12** Each nitrogen atom has 2 electrons because of the bond between them. That means each nitrogen still needs 6 more electrons. 6 + 6 = 12, so we NEED 12 electrons.
6. Compare the amount of electrons you HAVE with the number of electrons you NEED. If the values match, draw the correct amount of dots on each symbol. Dots must be drawn in pairs. If the number of electrons you HAVE and the number of electrons you NEED do not match, draw another bond between the symbols and subtract 2 more electrons from your total.	**6.** HAVE N_2 5 x 2 = 10 N—N **NEED** **- 2** **8** 6 6 = **12** **- 2** N=N **6** 4 4 = **8** **-2** **4** N≡N **2** **2** = **4**
7. Once the numbers match, draw in your dots. The number of dots to be drawn on each symbol is indicated by the number below it once the number of electrons you NEED and the number of electrons you HAVE match. All dots must be written in pairs.	**7.** Because we have a 2 under each nitrogen, each nitrogen gets 2 dots. Write them as a pair. :N≡N:
8. If the molecule is an ion and has an overall charge, put brackets around the structural formula and write the charge outside the brackets in the upper right corner. (See the next example to see what I mean.)	**8.** Nitrogen is not an ion. Thus, our structural formula does not need brackets. We are done.

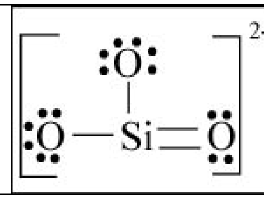
Write the structural formula for each of the following compounds.

1. F_2
2. Cl_2
3. I_2
4. N_2
5. O_2
6. H_2
7. ClF
8. NH_3
9. CH_4
10. H_2S
11. PH_3
12. BrI
13. CO
14. CO_2
15. SO_2
16. CCl_4
17. SiF_4
18. CCl_2F_2
19. SiO_2
20. H_2O
21. CH_2ClF
22. HCN
23. AsF_3
24. CH_2O

Part 2: How to Write a Structural Formula for an Ion

Write the structural formula for SiO_3^{2-}

Description of Action	Action
1. Determine how many electrons you have to work with; this is done by multiplying the number of valence electrons for that element by the number of those atoms. Label your total as "HAVE".	**1.** SiO_3^{2-} element's valence # of that element Si 4 x 1 = **4** O 6 x 3 = **18** **22 HAVE**
2. If the given molecule is an ion, add or subtract electrons from your total, depending on the charge of the molecule. If the molecule has a positive charge, subtract electrons from your total. If the molecule has a negative charge, add electrons to your total.	**2.** SiO_3^{2-} is an ion. It has a –2 charge, which means it has gained two electrons. Thus, we must add two electrons to our total. 22 + 2 = **24**. We **HAVE 24** electrons to work with.
3. Write the symbols for the given formula with some space between them. If there are more than two atoms, put the one with the **lowest electronegativity** in the center and the others on each of the four sides. **Exception:** Hydrogen can NEVER be in the center. An element's electronegativity can be found on the front of your periodic table.	**3.** Silicon has a lower electronegativity than oxygen. So silicon will go in the middle and the oxygen will be placed around it as shown to the right. 
4. Draw one dash between the symbols to represent a bond. Each dash drawn represents 2 electrons. Subtract 2 from your "HAVE" amount for each bond you draw.	**4.** HAVE 24 -6 18 We must draw bonds between silicon and each of the oxygens.  Since we must draw three bonds and each bond is worth 2 electrons, we must subtract 6 from our total.
5. Each element in a covalent compound wants to share 8 electrons. We have to check to see how many electrons each element needs to reach 8. Every bond that is drawn to a symbol should be counted as 2 electrons for that element. Add these numbers together and label this number as "NEED".	5. HAVE NEED 24 O: 6 -6 O: 6 **18** O: 6 Si: 2 **20**  Since each oxygen has 2 electrons, they each need 6 more to reach 8. 6 x 3 = 18. Plus, silicon has 6 which means it still needs 2. 18 + 2 = **20.**
6. Compare the amount of electrons HAVE with the number of electrons you NEED. If the values match, draw the correct amount of dots on each symbol. Dots must be drawn in pairs. If the number of electrons you need and the number of electrons you have do not match, draw another bond between the symbols and subtract 2 more electrons from your total.	**6.** . HAVE NEED 24 O: 6 -6 O: 6 18 O: 4 -2 Si: 0 **16** **16**  Our NEED and HAVE values do not match, so we must add another bond and subtract two. Once we do that, our NEED and HAVE values do match. Move on to the next step.
7. Once the numbers match, draw in your dots. The number of dots to be drawn on each symbol is indicated by the number below it once the number of electrons you NEED and the number of electrons you HAVE match. All dots must be written in pairs.	**7.** Since our HAVE and NEED values matched, I drew in the correct amount of dots for each element. Silicon, since it has 4 dashes (8 electrons) does not need any dots. The oxygen with the double bond needs only 4. The other two oxygens with the single bonds each need 6 dots. 

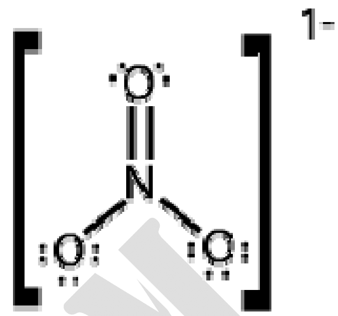
8. If the molecule is an ion and has an overall charge, put brackets around the structural formula and write the charge outside the brackets in the upper right corner.	8.	

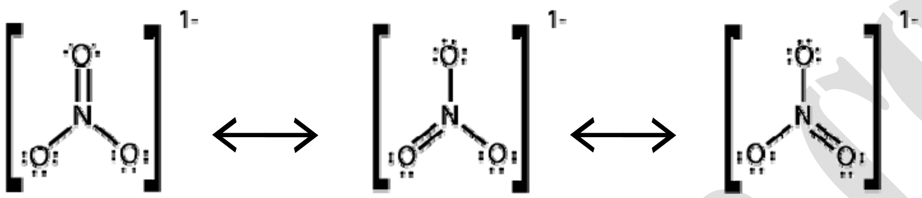
Write the structural formula for each of the following charged covalent compounds.

1. PO_4^{3-}
2. SeO_4^{2-}
3. SiO_3^{2-}
4. NH_4^{1+}
5. NO_3^{1-}
6. ClO_3^{1-}
7. SO_4^{2-}
8. NO_2^{1+}
9. H_3O^{1+}
10. SO_3^{2-}
11. NO_2^{1-}
12. CO_3^{2-}

Part 3: Resonance

- Sometimes more than one valid Lewis structure is possible for a given molecule.
- Observe the Lewis structure for nitrate to the right. It shows one double bond and two single bonds. But, experiments clearly show that only one type of N – O bond occurs with length and strength between those expected for a single and double bond.
- The structure to the right is a valid Lewis Structure but it does not correctly represent the bonding in NO_3^-.
- Resonance occurs when more than one valid Lewis structure can be written for a particular molecule.
- The resulting electron structure of the molecule is the average of the resonance structures. The three resonance structures for nitrate are shown below.





- Resonance shows that electrons are not localized to one atom but instead travel throughout the molecule.
- Resonance structures got their name because scientists originally thought that the bonds would switch or resonate between the different positions. Further research revealed that compounds that can be written with resonance structures actually have bonds that are hybrids of the other bonds in the compound

Write all resonance structures for each of the following.

1. SO_2

2. $(NO_3)^{1-}$

3. $(CO_3)^{2-}$

4. $(NO_2)^{1-}$