Chemistry Help Guide: **Bonding Diagrams**

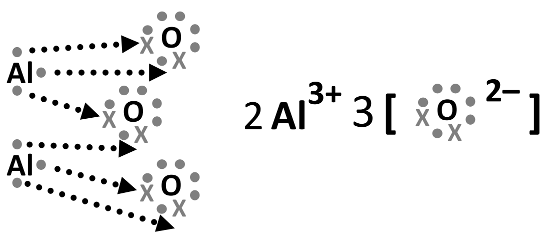
What are Bonding Diagrams?

The way atoms are bonded together determines the physical and chemical properties of the compounds they form. This includes the way a substance looks and especially the way it will react with other substances.

Scientists use a set of diagrams known as "Lewis diagrams" to visually represent how a substance's atoms are bonded together. There is a different way to draw Lewis diagrams for each type of bonding; always start by determining whether you are working with an ionic, metallic, or covalent compound.

Ionic compounds are made of metals bonding with nonmetals; they form **ionic bonds**.

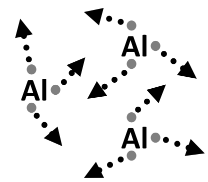
aluminum oxide, Al_2O_3



In an ionic compound, metal atoms give their electrons to nonmetal atoms, creating ions with opposite charges.

Metallic compounds are made of metals bonding with other metal atoms; they form **metallic bonds**.

aluminum metal, Al



$3\ \mathbf{Al}^{3+}$

In a metallic compound, all of the metal atoms lose their electrons, creating the metallic "sea" of mobile electrons.

Covalent compounds are made of nonmetals bonding with other nonmetals; they form **covalent bonds**.

Steps for drawing Lewis structures for covalent molecules:

Step 1: Count the valence electrons. Use the chemical formula to determine how many atoms of each element are present, and add up the total number of available valence electrons based on the elements' locations on the Periodic Table. This is your "electron budget."

Step 2: Draw the skeleton structure. Draw single bonds connecting each of the atoms around the central atom. The **central atom** is always the **first element** in the formula. **Hydrogen** can **never** be the central atom.

Step 3: Draw an octet for each of the outside atoms. Remember that each bond represents 2 electrons; both atoms get to count these 2 electrons as their own. There are 2 **exceptions** to the octet rule: **hydrogen can only have 2 electrons** (1 bond) and **boron can only have 6 electrons** (3 bonds).

Step 4: Count the electrons you've already drawn. Have you already reached your budget? If not, **leftover electrons always go to the central atom**. If you've already used all of your available electrons, you'll need to make the atoms share by adding more **bonds**. If you are AT your budget, erase ONE pair of electrons from an outside atom and draw an extra bond so it can share this pair with the central atom. If you are OVER your budget, erase TWO pairs of electrons (one from each atom) and add an extra bond between the two atoms.

Step 5: Check that all atoms have access to a full octet. Make sure you have used all of your available electrons and that you have not gone over your budget.

oxygen dichloride, OCl_2

Step 1:

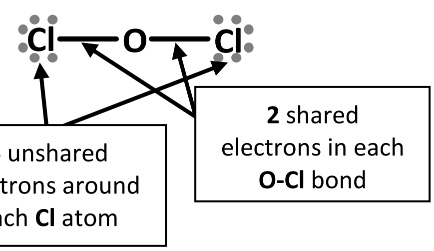
According to the Periodic Table, **O** has **6** valence Electrons and **Cl** has **7** valence electrons.

(1 x **O**) + (2 x **Cl**)
(1 x **6**) + (2 x **7**)
6 + 14
= **20 valence electrons** available

Step 2: Cl—O—Cl

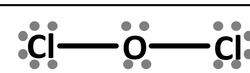
Step 3:

Step 4:



4 electrons in bonds **+ 12** electrons around outside atoms = **16** already used electrons

20 available electrons – 16 used electrons = 4 leftover electrons



Step 5: All atoms have an octet and all 20 available electrons have been used.

In a covalent compound, all nonmetal atoms want to keep their own electrons and share with other atoms in order to gain their full octet. Each nonmetal atom will form a bond for each electron it still needs.

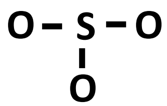
Need to see a few more examples of Lewis structures for covalent molecules?

sulfur trioxide, SO_3

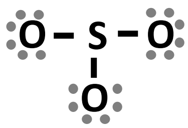
Step 1:
S and **O** both carry **6** valence electrons.

(1 x **S**) + (3 x **O**)
(1 x **6**) + (3 x **6**)
6 + 18
= **24 valence electrons** available

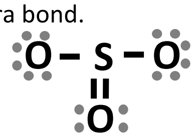
Step 2:



Step 3:



Step 4: All 24 of the available valence electrons have been used. Each O atom has a full octet, but the S atom is missing 2 electrons. Since we are AT our budget, we need to erase ONE pair from 1 O and draw an extra bond.



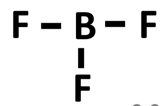
Step 5: All atoms have an octet and all 24 available electrons have been used.

boron trifluoride, BF_3

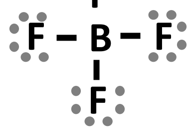
Step 1:
B has **3** valence electrons and **F** has **7** valence electrons.

(1 x **B**) + (3 x **F**)
(1 x **3**) + (3 x **7**)
3 + 21
= **24 valence electrons** available

Step 2:



Step 3:



Step 4: All 24 of the available valence electrons have been used. Each F atom has a full octet. B can only have 6 electrons; since B is an exception to the octet rule, it is also full. No more electrons are needed.

Step 5: All atoms have a full set of electrons (8 for each O atom and 6 for the B atom) and all 24 available electrons have been used.

nitrogen, N_2

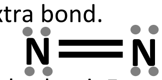
Step 1:
N has **5** valence electrons.

(2 x **N**)
(2 x **5**)
= **10 valence electrons** available

Step 2: There is no central atom in diatomic molecules; the two atoms will "split" the center across the bond.

Step 3:

Step 4: We have used 14 electrons, which is 4 more than we have available. Since we are OVER budget, we need to erase TWO pairs (1 from each N atom) and draw an extra bond.



We're still over budget! Erase TWO more pairs and draw a 3rd bond.

Step 5: All atoms have an octet and all 10 available electrons have been used.