## Introduction

In the previous note pack, you learned some about Lewis dot structures, which represent chemical compounds by showing how electrons are distributed between the molecules. We will now learn how to draw Lewis dot structures ("Lewis structures", or "Lewis formulas") for most molecules and ions. The reason we do this is that the Lewis structure gives us key information about the molecule. Given a set of molecules, if we're able to draw the Lewis structure we can tell which molecule should have the higher boiling point, what each molecule might dissolve in, etc. And we get all of this **without** having to handle the chemical in the lab!

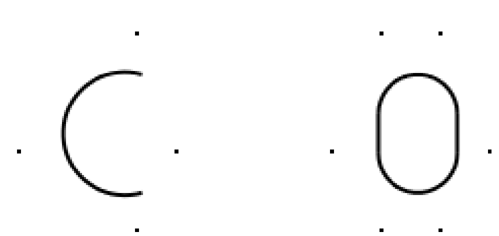
## The basics

In a previous note pack, we discussed the basics of drawing Lewis structures. Dots represent **s** and **p** valence electrons. A single line represents a pair of electrons shared between two atoms, while double and triple lines represent two and three pairs of electrons shared - single, double, and triple **bonds**. Now we will move on to talk about how to draw Lewis structures for complex species - those with more than two atoms.

## Lewis structures - step by step

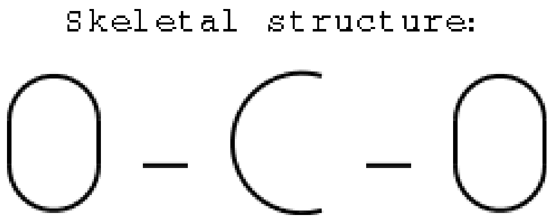
You can use essentially the same set of steps to draw the Lewis structure for almost anything. We'll modify these as we learn a few more things about the structure of atoms, but this discussion will be a good starting point when you need a structure. We will illustrate the process with a simple example - carbon dioxide ($CO_2$).

1) **Count the valence electrons.** You can do this by writing the Lewis dot formula for each atom in the compound, or you can use the periodic table to get the number of valence electrons.

| | |
|---|---|
| Lewis structures for C and O:<br><br>4 electrons 6 electrons | • You don't **have** to draw the Lewis structures - but if you're new to this, it's good practice!<br>• Carbon is in group IVA, which has **4 valence electrons**. Four are unpaired, and carbon is likely to form **four bonds**.<br>• Oxygen is in group VIA, which has **6 valence electrons**. Two are unpaired, and oxygen is likely to form **two bonds**. |
| Illustration 1 - Lewis structures of C and O | |

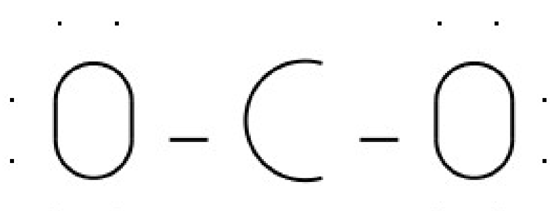
If the species is an **ion**, you need to add electrons (if the ion is negatively charged) or subtract electrons (if the ion is positively charged). Our example molecule has a total of **16 valence electrons**, four from the carbon, and twelve from the two oxygens. This molecule is neutral (uncharged), so we don't need to add or subtract any electrons.

2) **Write a skeletal structure**. You can do this by connecting atoms with a single bond. So how do you arrange the atoms? Sometimes you will be told. Sometimes the central atom is obvious - in $SF_6$, sulfur is the central atom. Complicated molecules may have more than one "center". Unless you're told otherwise, you can draw a structure by arranging the atoms so that the **least electronegative atom is in the center of the structure**. Usually, this means that **the atom that needs to gain the most electrons will be in the center**.

| | |
|---|---|
|   | • The electronegativity of carbon (2.5) is less than that of oxygen (3.5), so put carbon in the center.<br>• Since carbon needs to gain more electrons that oxygen does, it's likely to form more bonds.<br>• Each atom connects to the next with a single bond. |
| Illustration 2 - Skeletal structure of $CO_2$ | |

In our example, carbon is the central atom and we've bonded it using single bonds to each oxygen.

3) **Distribute the electrons**. First, give electrons to the outer atoms. Then, distribute the remaining electrons (if any) to the inner atoms until you run out. Remember that each single bond in your skeletal structure contains **two electrons**, so subtract those out before distributing.

| | |
|---|---|
|  | • We've made two single bonds - this uses **four electrons**.<br>• After we fill the valence shells of each of the oxygen atoms, we've **used all 16 valence electrons**. |
| Illustration 3 - Electrons distributed around $CO_2$ | |

We've used all sixteen electrons, so we now have to evaluate whether this is an acceptable structure for carbon dioxide.

4) **Check to see if each atom in the structure has eight valence electrons - if not, rearrange electrons**. The main reason an atom bonds with another is to get a stable

electron configuration - in other words, they want filled valence shells like the noble gases have. We're only going to concern ourselves with the 's' and 'p' subshells (which can contain a total of eight electrons - two in the 's' subshell and six in the 'p' subshell) when drawing Lewis dot formulas. So, we need to check our structure to make sure that each of our atoms sees **eight electrons**. We call this the **octet rule** and we check to make sure each atom in our structure has an **octet** - eight valence electrons.

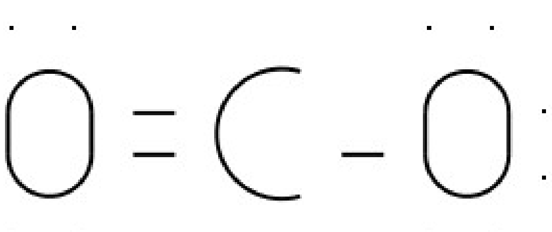
There is a common exception to every atom getting eight valence electrons - hydrogen. Since hydrogen has only an n=1 shell, that shell is full once hydrogen's 's' orbital is full. That means the hydrogen needs only **two** electrons total. When hydrogen forms a molecule, it shares its one electron with another atom, forming only one bond.

Let's count up the electrons that each atom sees in our structure in Illustration 3. Each **single bond** contains two electrons seen by each atom involved in the bond.

| *Atom* | *Electrons from bonds* | *Other electrons* | *Total electrons* |
|---|---|---|---|
| Carbon | 4 (two single bonds) | 0 | **4** |
| Oxygen (left) | 2 (one single bond) | 6 | **8** |
| Oxygen (right) | 2 (one single bond) | 6 | **8** |

We can see a problem immediately with our structure - carbon sees only four electrons, so it doesn't have an incentive to bond. We need to **rearrange the electrons** so that **each** atom has an octet. How do we do this? We need to share more electrons - by creating more bonds.

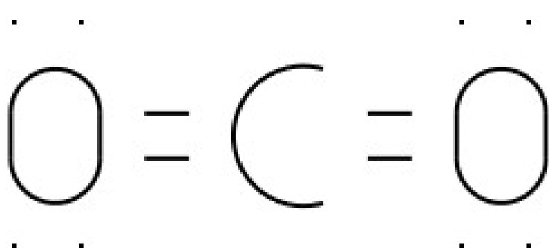
We will add a double bond to the structure and allow oxygen and carbon to share more electrons.

| | |
|---|---|
|  | • We've added a double bond between the oxygen atom on the left and the carbon atom.<br>• The oxygen on the left has four other electrons around it - this gives it eight.<br>• … but the central carbon still does not have enough electrons. |
| Illustration 4 - Proposed structure for $CO_2$ | |

Let's check this structure just like the last one.

| ***Atom*** | ***Electrons from bonds*** | ***Other electrons*** | ***Total electrons*** |
|---|---|---|---|
| Carbon | 6 (one single bond, one double bond) | 0 | **6** |
| Oxygen (left) | 4 (one double bond) | 4 | **8** |
| Oxygen (right) | 2 (one single bond) | 6 | **8** |

Carbon still has only six electrons, but it should now be obvious what to do to give carbon eight - bond the **other** oxygen with a double bond. We could use a triple bond to the first oxygen here, but since there's no reason to suppose one oxygen atom bonds differently from another in this situation, we give the other oxygen a double bond. Using a second double bond gives us the following structure.



- Each oxygen atom has four electrons in addition to the double bond. This gives an octet to each oxygen.
- Carbon has a double bond to each oxygen and no additional electrons.
- … giving it a share in eight electrons.

Illustration 5 - Lewis structure of $CO_2$

Let's check **this** structure.

| ***Atom*** | ***Electrons from bonds*** | ***Other electrons*** | ***Total electrons*** |
|---|---|---|---|
| Carbon | 8 (two double bonds) | 0 | **8** |
| Oxygen (left) | 4 (one double bond) | 4 | **8** |
| Oxygen (right) | 4 (one double bond) | 4 | **8** |

Carbon and both oxygen atoms have an octet. This structure is likely to be the correct one.

<u>Practice</u>

Practice writing Lewis structures for the following simple species. Use the four steps we've outlined in this note pack. Try to do each using only the periodic table for a reference.

| | | | | | | |
|---|---|---|---|---|---|---|
| $CO$ | $N_2$ | $NH_3$ | $NH_4^+$ | $OF_2$ | $CH_2O$ | $CH_2F_2$ |

Answers to these practice structures are in Illustration 6.

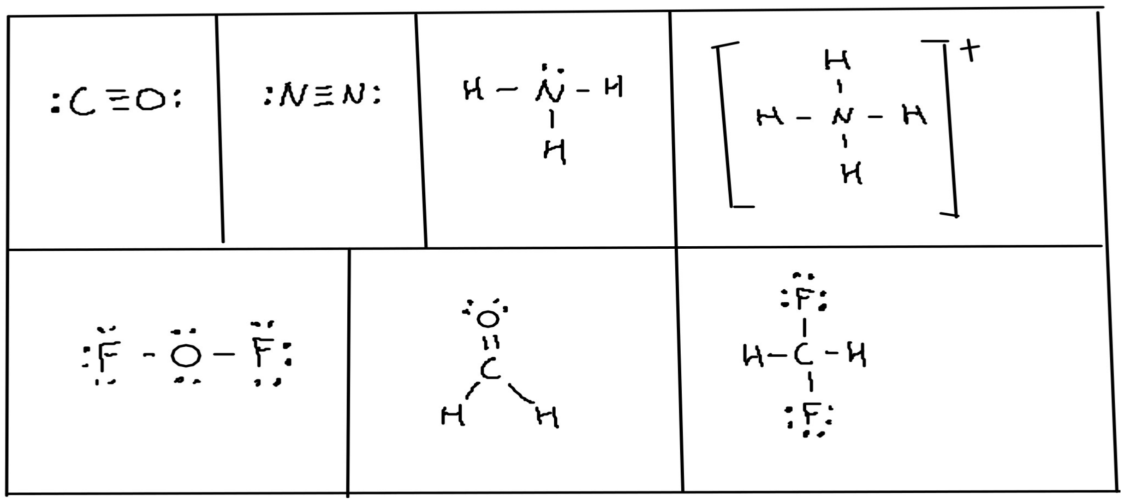


Illustration 6 – Answers to the practice structures

Summary

We've discussed a simple method to write and draw Lewis structures of simple molecules and polyatomic ions. You should now be able to draw simple structures easily and to be able to make an attempt at drawing structures for more complex molecules.