

TutorTube: Lewis Dot Structures

Summer 2021

Introduction

Hello and welcome to TutorTube, where The Learning Center's Lead Tutors help you understand challenging course concepts with easy-to-understand videos. My name is Rebecca Yohannes Lead Tutor for Biology and Chemistry. In today's video, we will explore Lewis Dot Structures. Let's get started!

Definitions

Let's first understand what Lewis structures are. They are a simplified representations of elements' valence shells using dots, lines and element symbols.

The dots represent the valence electrons. The numbers of the main groups in the Periodic Table are the same as the valence electrons of the elements in the respective group. This rule, however, does not apply to the Transition metals.

For example, fluorine which is in group 7A has 7 valence electrons, oxygen which is in main group 6 has 6 valence electrons, and so on.

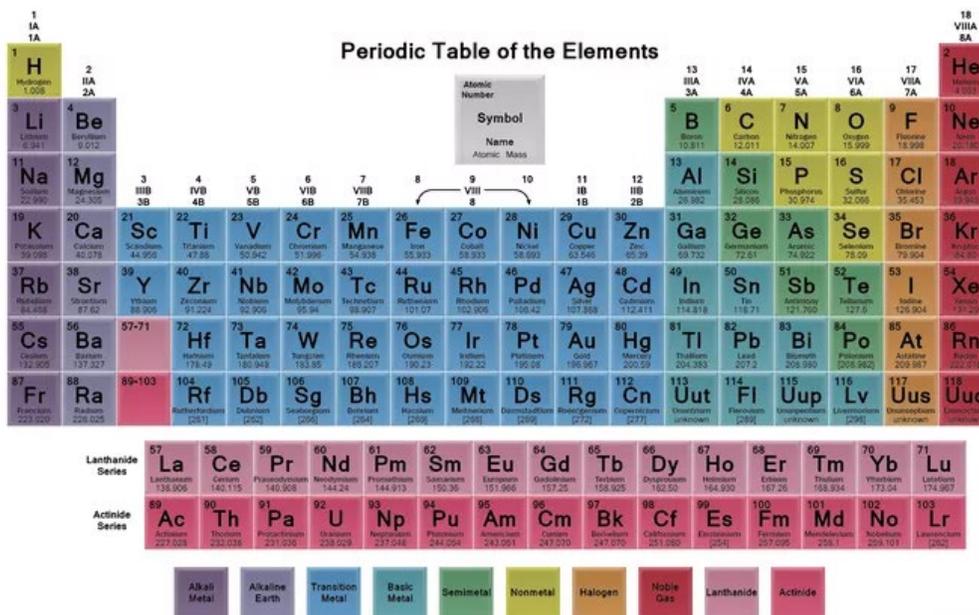


Figure 1: Periodic Table

Valence electrons aren't just drawn anywhere. Imagining a square around an element's symbol, there can only be a maximum of 2 dots on each side of the element's symbol. Chlorine has 7 valence electrons and its Lewis structure is drawn this way:



Figure 2: Lewis Structure of chlorine

When drawing the Lewis structures of compounds and molecules, bonds are illustrated by lines. Each line involves 2 electrons. Single bonds are depicted with one line, double bonds are 2 lines and triple bonds are 3 lines.

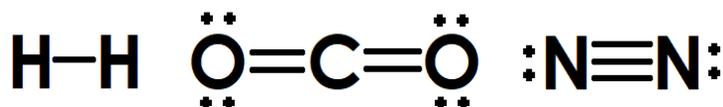


Figure 3: Lewis Structures of hydrogen molecule, carbon dioxide and nitrogen molecule

Octet Rule

The octet rule states that elements lose or gain electrons to have a total of 8 valence electrons and become stable.

Elements that tend to attract electrons to fulfill the octet rule have 4 or more valence electrons and these elements are typically found on the left side of the Periodic Table from group IVA to VIIA.

Elements on the second period of the Periodic Table lack d sublevels and only have s and p sublevels that comprise the octet or the 8 valence electrons. Thus, C, N, O, and F on the second period follow the Octet Rule without any exceptions (Lumen Learning, n.d.).

Let's take a look at some examples:

C has 4 valence electrons, so must form 4 bonds to have 8 valence electrons to fulfill the Octet Rule.

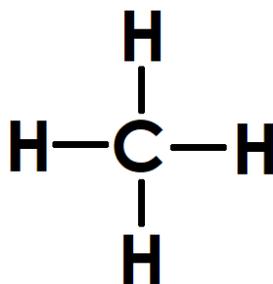


Figure 4: Lewis structure of methane

O has 6 valence electrons so must gain 2 more electrons to fulfill the Octet Rule. It does this by forming 2 covalent bonds.

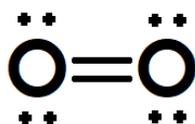


Figure 5: Lewis structure of molecular oxygen

F has 7 valence electrons and only needs one more electron to have 8 valence electrons. Thus, in a molecule or compound, it is likely to form only one bond.

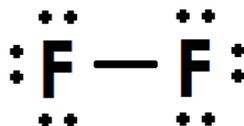


Figure 6: Lewis structure of fluorine molecule

Exceptions to the Octet Rule

Certain elements can have a stable electron configuration while having less than 8 valence electrons after forming bonds. These elements are said to have *incomplete octets*. Some common examples of elements with incomplete octets are H, Li, Be, B, and Al.

In a hydrogen molecule, each hydrogen has 2 valence electrons in its shell after sharing a bond with the other hydrogen.

Beryllium in beryllium fluoride has 4 valence electrons after bonding with two fluorine atoms.

Boron trichloride is stable with only 6 valence electrons on the boron.

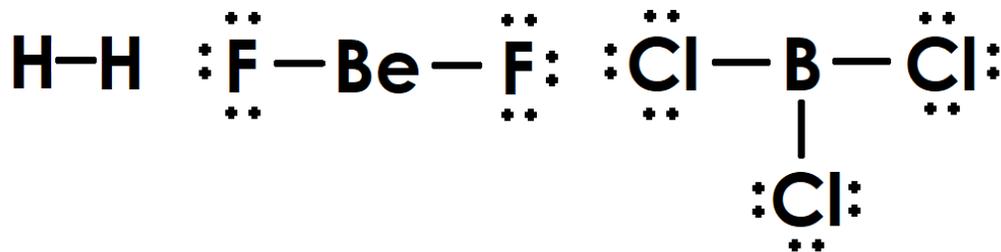


Figure 7: Lewis structures of molecular hydrogen, beryllium fluoride, and boron trichloride as examples of incomplete octets

Elements that can potentially have more than 8 valence electrons after forming bonds are said to have *expanded octets*. For elements after Al, a d-sublevel exists that can accommodate more electrons. This allows elements, such as Si, P, S, and Cl to have expanded octets.

Periodic Table of the Elements

1

2

13

14

15

16

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18

Periodic Table of the Elements																	
1 H Hydrogen 1.008																	2 He Helium 4.002
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.064	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 84.798
37 Rb Rubidium 84.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.711	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.29
55 Cs Cesium 132.905	56 Ba Barium 137.328	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.227	78 Pt Platinum 195.085	79 Au Gold 196.967	80 Hg Mercury 200.592	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [209]	85 At Astatine 209	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [269]	111 Rg Roentgenium [272]	112 Cn Copernicium [277]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [293]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown
57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.242	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.055	71 Lu Lutetium 174.967			
89 Ac Actinium 227.028	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium 237.048	94 Pu Plutonium 244.064	95 Am Americium 243.061	96 Cm Curium 247.070	97 Bk Berkelium 247.070	98 Cf Californium 251.080	99 Es Einsteinium [254]	100 Fm Fermium 257.095	101 Md Mendelevium 258.1	102 No Nobelium 259.101	103 Lr Lawrencium [262]			

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Figure 8: Elements that obey the Octet Rule and elements that can have an expanded octet on the Periodic Table

For example, phosphorus in phosphorus pentachloride has 10 valence electrons after forming single bonds with 5 chlorine atoms.

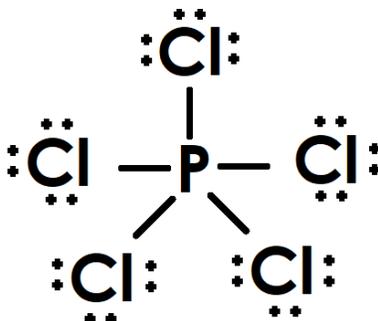


Figure 9: Lewis structure of phosphorus pentachloride

Drawing Lewis structures

Let's go over the steps of drawing the Lewis structure of a fluorine molecule (F_2)

First, count all the valence electrons involved in the molecule. You can always refer to the main group number to find out what the valence electron of an element is. Since fluorine is found in the 7th group, it has 7 valence electrons. There are 2 fluorine atoms in this molecule, so there are a total of 14 valence electrons.

Next, draw the skeletal structure of the molecule. Don't worry about the lone electrons just yet, simply draw single bonds between the atoms involved:



Figure 10: Skeletal formula of molecular F

Now, calculate how many valence electrons are left after drawing the bonds. Since a single bond contains 2 electrons and there is only one single bond in the skeletal structure, subtract 2 electrons from 14 which tells us that 12 valence electrons are left. Draw these electrons around each fluorine atom as lone electrons or dots. This will be the resulting structure:

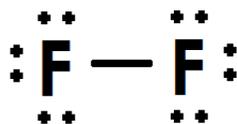


Figure 11: Lewis structure of molecular fluorine

You can now check if both atoms in this diatomic molecule have fulfilled the octet rule. Since both fluorine atoms have 8 electrons around them, this is the correct Lewis structure for molecular fluorine.

Let's work on the next example, phosphorus trichloride (PCl_3).

First, count all the valence electrons involved in this compound. Phosphorus has 5 valence electrons and chlorine has 7 valence electrons. Since there is only one phosphorus atom and three chlorine atoms in this compound, there will be a total of 24 valence electrons in this compound.

$$\text{PCl}_3: 5 + (3 \times 7) = 24 \text{ valence electrons}$$

Let's now determine which element is the central element and which are the terminal atoms. Generally, the least electronegative element is the central atom, except hydrogen which cannot be the central atom because hydrogen can't form more than one bond. In phosphorus trichloride, phosphorus is the least electronegative element; therefore it is the central atom and the chlorine atoms are the terminal atoms. We can now draw the skeletal structure:

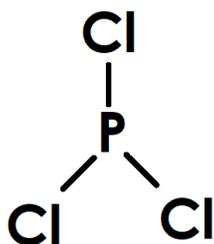


Figure 12: Skeletal structure of phosphorus trichloride

You can try to predict how many bonds each element must form to fulfill the octet rule at this point. Since chlorine has 7 valence electrons, it only needs to form one bond to have a total of 8 valence electrons. Phosphorus has 5 valence electrons, thus is likely to form 3 bonds to have a total of 8 valence electrons around it.

Since 3 bonds take up 6 electrons, then we need to subtract that from the total valence electrons: $24 - 6 = 18$ valence electrons remaining

Place the remaining electrons around the terminal atoms first. Then if there are valence electrons left, add them on to the central atom. Since there would still be 2 electrons left after filling the chlorine atoms' valence shells with lone pairs, draw those 2 electrons on the central atom phosphorus:

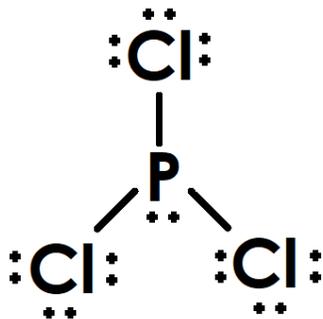


Figure 13: Lewis structure of phosphorus trichloride

Finally, count the electrons around each atom to make sure that all the elements fulfill the Octet Rule. Since each element has 8 valence electrons now, this is the correct Lewis structure for phosphorus trichloride.

Another example we can work on is Nitrosyl chloride, NOCl

First, calculate the total number of valence electrons present in the compound; nitrogen has five valence electrons, oxygen has 6 and chlorine has 7 valence electrons, which gives a total of 18 valence electrons in this compound.

$$\text{NOCl} = 5 + 6 + 7 = 18 \text{ valence electrons}$$

Next, identify which element goes in the middle and which are terminal. Since nitrogen is the least electronegative element in this compound compared to oxygen and chlorine, it is the central atom and oxygen and chlorine are the terminal atoms bonded to nitrogen. Now, we can draw the skeletal structure:



Figure 14: Skeletal formula of nitrosyl chloride

We can try to predict the number of bonds each element is likely to form as practice. Chlorine is only going to form one bond since it has 7 valence electrons, nitrogen will probably need to form 3 bonds to fulfill the Octet Rule since it only has 5 valence electrons, and oxygen needs to form 2 bonds since it has 6 valence electrons. So, we can expect to see multiple bonds in this compound, specifically between nitrogen and oxygen. Before we draw multiple bonds, however, let's first place the remaining valence electrons as lone electrons around this skeletal structure. Remember to first draw the lone electrons on the terminal atoms then move on to the central atom if there are still more electrons left.

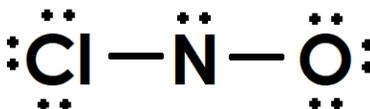


Figure 15: Incorrect Lewis structure of nitrosyl chloride

Check if all the elements fulfill the octet rule. Nitrogen only has 6 valence electrons around it, so it doesn't satisfy the Octet Rule. This is when we can try to use multiple bonds to ensure that nitrogen has 8 valence electrons. Based on our initial prediction, nitrogen needs 3 bonds and oxygen typically forms 2 bonds to have a total of 8 valence electrons around each of them. Since chlorine already meets our expectation, we can leave it as it is. So, let's use one of oxygen's lone pair to form a second bond between N and O. Remember that we can't simply add an extra pair of electrons into the compound out of nowhere because we already calculated that this compound only has a total of 18 valence electrons. That's why we have to remove a lone pair from O in order to have a second bond between N and O.

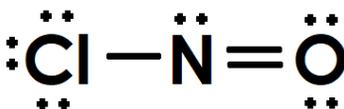


Figure 16: Lewis structure of nitrosyl chloride

Now, all of the elements have 8 valence electrons around them, so this is the final Lewis dot structure of NOCl.

Summarized Steps of Drawing Lewis Dot Structures

1. Add up all the valence electrons of the elements in the compound
2. Determine the central element and terminal atoms according to their electronegativities
 - a. Remember that least electronegative element is the central atom
 - b. Hydrogen can never be the central element since it can't form more than one bond
3. Draw skeletal structure:
 - a. At this point, try to predict which atoms form bonds and if single or multiple bonds are needed to fulfill the Octet Rule, but don't draw the multiple bonds just yet. Only draw the single bonds for now
 - b. Notice if there are elements that can have incomplete or expanded octets

4. Compute and place the remaining valence electrons around the elements as lone pairs of electrons
 - a. Remember to place the lone electrons on the terminal atoms first, then on the central atom if there are still more electrons remaining
5. If there are elements that are not fulfilling the Octet Rule, draw multiple bonds
 - a. Think back to the predictions made earlier when drawing the skeletal structure of the compound (in step 3)

Formal Charges

Sometimes, you may find multiple possible ways to draw the Lewis structure of a compound. In these cases, you must calculate the formal charge to determine the preferred Lewis structure.

FC of an element = valence electron of the element – number of bonds it formed – lone electrons of that element

Choosing among multiple Lewis Structures

1. Lewis structures with more zero formal charges are generally preferred.
2. Negative formal charges should be on the more electronegative atom and positive formal charges should be on the less electronegative atom.
3. Lewis structures with formal charges closest to zero are more stable, thus preferred.
4. The further away atoms with opposite formal charges are, the more stable the Lewis structure is.

Let's look at sulfite ion (SO_3^{2-}) as an example. There is one sulfur atom in this ion contributing six valence electrons, three oxygen atoms give us a total of 18 valence electrons, and there are two extra electrons in this ion represented by the negative two charge on the compound, therefore there will be a total of 26 valence electrons in this ion.

$\text{SO}_3^{2-} = 6 + (3 \times 6) + 2 = 26$ total valence electrons.

Sulfur is the less electronegative element, so it is the central atom. The skeletal structure of the compound looks like this:

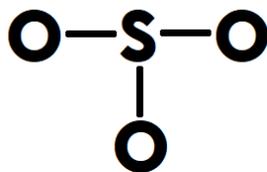


Figure 17: Skeletal formula of sulfite

Let's predict if there would be multiple bonds in this compound. Oxygen has 6 valence electrons and needs to form 2 bonds to fulfill the octet rule, so we might see some multiple bonds here. Sulfur also has 6 valence electrons and remember that sulfur is one of the elements that can have an expanded octet, so it can form more than 2 bonds and carry more than 8 valence electrons on it.

Now, let's calculate the remaining valence electrons.

$$26 - (2 \times 3) = 20 \text{ valence electrons left.}$$

When we draw the 20 valence electrons on the compound, this is what it looks like:

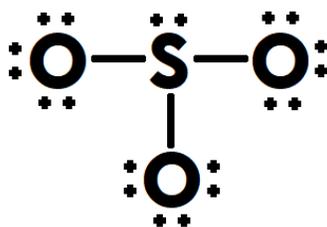


Figure 18: Possible Lewis structure of sulfite

This structure is one possible Lewis structure of sulfite ion because all the elements have satisfied the Octet Rule. However, remember we predicted that oxygen usually likes to form 2 bonds and sulfur can have an expanded octet, so let's take a lone pair from one of the oxygen atoms and use it to make a second bond with sulfur.

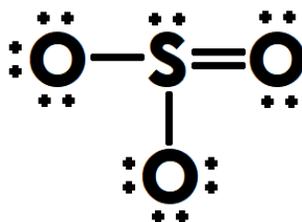


Figure 19: Second possible Lewis structure of sulfite

This is another way to draw the Lewis structure of sulfite.

Let's calculate the formal charge for each structure to determine the preferred one.

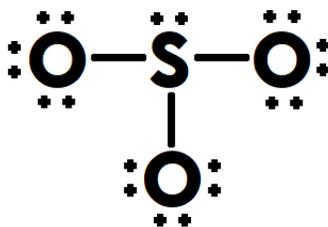


Figure 20: First possible Lewis structure of sulfite

The formal charge of sulfur on the first Lewis structure (figure 20) that we drew is calculated by subtracting its 6 valence electrons with the number of bonds its forming (3) and the lone electrons (2) on it. This results in a positive one charge on sulfur.

$$FC_1 \text{ of S} = 6 - 3 - 2 = +1$$

$FC_1 \text{ of O} = 6 - 1 - 6 = -1$, since all three oxygen atoms have the same number of bonds and lone electrons, the formal charge is the same for all 3 of them.

This structure results in multiple formal charges on the atoms. The overall FC is $(+1) + (-1) + (-1) + (-1) = -2$, which corresponds with the ion charge of the compound (SO_3^{2-})

Let's do the same thing for the second structure (Figure 21)

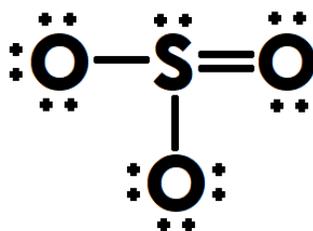


Figure 21: Second possible structure of sulfite

$$FC_2 \text{ of S} = 6 - 4 - 2 = 0$$

$FC_2 \text{ of O} = 6 - 1 - 6 = -1$, for the 2 oxygen atoms that are forming single bonds with S

$FC_2 \text{ of O} = 6 - 2 - 4 = 0$, for the oxygen atom that is forming a double bond with S

The overall FC of this structure is also -2 which makes sense because it is the same as the ion charge of the compound (SO_3^{2-}).

Comparing the formal charges on this structure to the first (Figure 20) structure, this second structure (Figure 21) has a greater number of atoms/elements that have a zero formal charge which makes it the preferred Lewis dot structure.

Outro

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References

Figure 1: Helmenstine, A. M., Ph.D. (2020, August 25). What Is the Importance of Color on the Periodic Table? Retrieved from <https://www.thoughtco.com/color-on-the-periodic-table-608827>

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