Covalent Bonding & Lewis Structures

Unit: Covalent Bonding & Molecular Geometry

NGSS Standards/MA Curriculum Frameworks (2016): HS-PS1-2

Mastery Objective(s): (Students will be able to…)

- Draw Lewis structures for diatomic molecules, showing bonds and lone pairs.

Success Criteria:

- Lewis structures show the correct number of bonds.
- Lewis structures show the correct number of unpaired electrons, drawn in pairs.

Tier 2 Vocabulary: bond, lone pair

Language Objectives:

- Explain how a Lewis structure represents the electrons in a molecular compound.

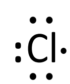
Notes:

covalent bond: a chemical bond consisting of one or more pair(s) of shared electrons.

covalent compound (also known as a molecular compound): a compound made of atoms joined by covalent bonds.

- covalent bonding occurs between non-metals
- electrons are shared in pairs—each pair usually contains one electron that came from each atom.

For example, a chlorine atom has seven valence electrons. The Lewis dot structure for chlorine is:



Note that the chlorine atom has seven valence electrons.

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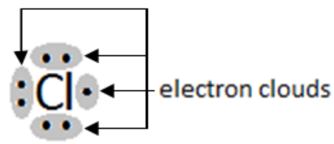
In a Lewis dot structure, we draw the valence electrons in pairs, representing the orbitals that contain them.

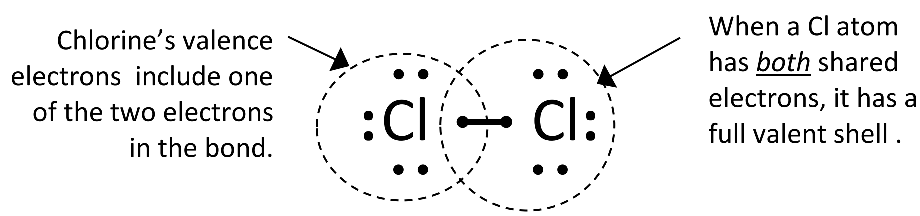
electron cloud: an orbital (a region within the atom) that contains either one or more of an atom's valence electrons or the shared electrons in a covalent bond.



Because the chlorine atom contains seven valence electrons, it needs one more to fill its valent shell. If it is able to receive an electron from another atom (such as sodium), it can from a Cl^{-} (chloride) ion, which can then become part of an inorganic compound, such as NaCl (sodium chloride).

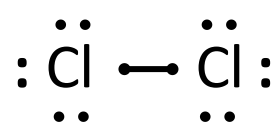
However, if chlorine is only in the presence of other atoms that are also trying to gain electrons, it can form a covalent bond, in which the electrons are shared.

You can think of this as if the chlorine atom were "borrowing" an electron to fill its valent shell. In order to "borrow" an electron, the chlorine atom needs to share an electron of its own. This shared pair of electrons becomes an electron cloud that is shared between the two chlorine atoms, which we call a covalent bond:



When each of the Cl atoms has both of the shared electrons, it has a full valent shell. This means each Cl atom gets a full valent shell some of the time.

This drawing is called a "Lewis structure," named after its inventor, American chemist G.N. Lewis. In a Lewis structure, unshared electrons are shown as dots (just like the Lewis dot diagram), and bonds are shown as line segments:



(Usually, the dots at the ends of the line segments are omitted. They are shown here as a reminder that the line representing the bond contains two electrons.)

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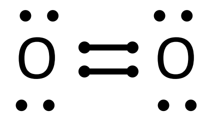
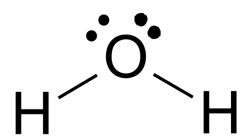
For another example, oxygen has six unpaired electrons, which means an oxygen atom needs two more electrons to fill its valent shell. This means that each oxygen atom in O_2 will share two electrons:



Just like the example with Cl_2, each oxygen atom has a full valent shell whenever it has all four of the shared electrons.

One pair of shared electrons is called a <u>single bond</u>, such as the bond in Cl_2. The O_2 molecule has two pairs of shared electrons, which is called a <u>double bond</u>. Three pairs of shared electrons would be a <u>triple bond</u>. Note that a double bond is a large electron cloud with four electrons, and a triple bond is an even larger electron cloud with six electrons.

An atom can make bonds to more than one other atom. For example, oxygen needs two additional electrons. IT doesn't matter where the oxygen atom gets these electrons from—oxygen can just as easily share one electron with each of two different atoms. An example is the H_2O molecule:



This reasoning can be extended to any non-metal. The rule of thumb is that atoms have to "share one to get one". This means that the number of bonds to any atom will be equal to the number of electrons the atom needs in order to fill its valent shell.

Because of space constraints and the shapes of the *s* and *p* orbitals involved in forming bonds, it is not possible for atoms to make more than a triple bond to a single atom. Also, in a first-year high school chemistry course, we will only consider structures made by electrons in *s* and *p* orbitals. This means the structures we will consider in this course will have no more than four total electron clouds. Structures with expanded octets are studied in AP® Chemistry.

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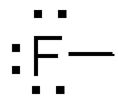
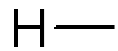
Common Atoms in Lewis Structures

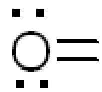
Assuming that each atom (other than hydrogen) needs a full octet (8 electrons), including unshared pairs of electrons ("lone pairs") plus the electrons in bonds. Here are the most common combinations.

The halogens (F, Cl, Br, and I) have 1 bond and 3 lone pairs; hydrogen (H) has 1 bond and no lone pairs:

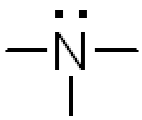


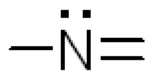


Atoms in the oxygen group (O, S, Se, and Te) always have 2 bonds and 2 lone pairs. The two possible combinations are:

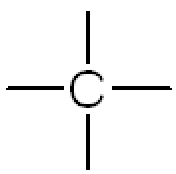


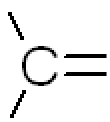
Atoms in the nitrogen group (N, P, and As) always have 3 bonds and one lone pair. There are three possible combinations:

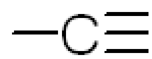




Atoms in the carbon group (C and Si) always have 4 bonds and no lone pairs, There are four possible combinations:







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Rules for Drawing Lewis Structures

1. Neutral atoms will make the same number of bonds as the number of additional electrons they need to fill their valent shell.
2. No pair of atoms can have more than a triple bond between them.
3. Atoms that make only one bond (*i.e.,* that need only one more electron) will almost always be on the outside (*i.e.,* not in the center).
4. The least electronegative atom will almost always be in the center.
5. The total number of bonds in a Lewis structure will usually be:

$$\frac{\text{\# electrons wanted} - \text{actual \# electrons}}{2}$$

If you cannot draw a correct Lewis structure using neutral atoms, try the following. (These are explained in the section titled "Charged Atoms in Lewis Structures" on page 278.)

6. Try moving one electron (at a time) from the central atom to the most electronegative atom that doesn't yet have a negative charge.
7. If the structure is for a charged ion, add or remove the appropriate number of electrons from your structure.
   a. If you need to add electrons, add them to the most electronegative atom first.
   b. If you need to remove electrons, remove them from the least electronegative atom first.
8. If you can draw more than one valid structure, the one with the lowest maximum charge on an atom is preferred. If you can draw more than one such structure, the structure with the fewest number of charged atoms is preferred.

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Homework Problems

Draw a correct Lewis structure for each of the following compounds.

1. PI_3
2. N_2
3. H_2O
4. PBr_3
5. $SiCl_4$
6. HCl
7. FCN
8. HNO
9. PN
10. CH_3NH_2
11. CH_2O
12. CO_2
13. $CBrClFI$
14. NF_3
15. N_2H_2
16. IBr
17. CH_3OH
18. C_2H_4
19. OF_2
20. $COCl_2$

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