

10.5: Writing Lewis Structures for Covalent Compounds

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to **add** or **subtract** the number of electrons necessary to give the total charge on the ion.

For CO_3^{2-} , for example, we **add two electrons** to the total because of the **-2** charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the **least electronegative element** in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-} , which both have C as the central atom), which is another clue to the compound's structure.
- **Hydrogen and the halogens** are almost always connected to only one other atom, so they are usually **terminal** rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O , for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, **use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.**

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

✓ Example 10.5.1: Water

Write the Lewis Structure for H_2O .

Solution

Solutions to Example 10.4.1

Steps for Writing Lewis Structures	Example 10.5.1
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of <u>8 valence electrons</u> .
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be <u>HOH</u> .
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing one bonding pair of electrons between the O atom and each H atom gives
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	$\text{H}-\text{O}-\text{H}$ with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: $\text{H}:\ddot{\text{O}}:\text{H}$
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

✓ Example 10.5.2

Write the Lewis structure for the CH_2O molecule

Solution

Solutions to Example 10.4.2

Steps for Writing Lewis Structures	Example 10.5.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of $[(2)(1) + 4 + 6] = \underline{12 \text{ valence electrons}}$.
2. Arrange the atoms to show specific connections.	$\begin{array}{c} \text{O} \\ \\ \text{H}-\text{C}-\text{H} \end{array}$ <p>Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.</p>

Steps for Writing Lewis Structures

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

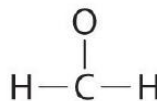
5. If any electrons are left over, place them on the central atom.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

7. Final check

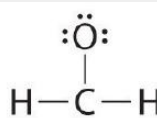
Example 10.5.2

Placing a bonding pair of electrons between each pair of bonded atoms gives the following:



6 electrons are used, and 6 are left over.

Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:

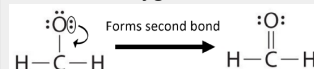


Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.

Not necessary.

There are no electrons left to place on the central atom.

To give carbon an octet of electrons, we use one of the lone pairs on oxygen to form a carbon–oxygen double bond:

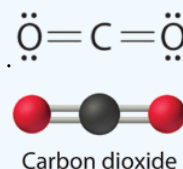


Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

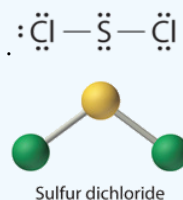
? Exercise 10.5.1

Write Lewis electron structures for CO_2 and SCl_2 , a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO_2



Answer SCl_2



The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions (CK-12)

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+ , is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia (NH_3) molecule in a coordinate covalent bond.

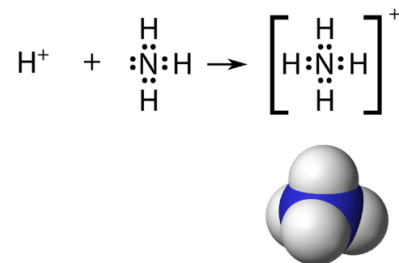


Figure 10.5.3: The ammonium ion. (CK12 License)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

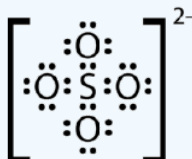
$$\begin{aligned} 1 \text{ N atom} &= 5 \text{ valence electrons} \\ 4 \text{ H atoms} &= 4 \times 1 = 4 \text{ valence electrons} \\ \text{subtract 1 electron for the } 1+ \text{ charge of the ion} \\ \text{total of 8 valence electrons in the ion} \end{aligned}$$

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

✓ Exercise 10.5.2

Draw the Lewis electron dot structure for the sulfate ion.

Answer (CK12 License)



Exceptions to the Octet Rule (BC Campus)

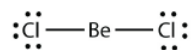
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO₂, and ClO₂. The Lewis electron dot diagram for NO is as follows:

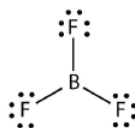


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

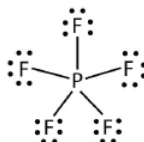
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3 :



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty d orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5 . The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

✓ Example 10.5.3: Octet Violations

Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

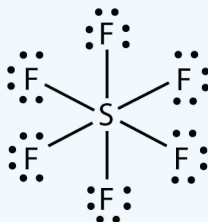
- ClO
- SF_6

Solution

a. With one Cl atom and one O atom, this molecule has $6 + 7 = 13$ valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:



b. In SF_6 , the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



? Exercise 10.5.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer



The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.

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