# 6.14: Covalent Molecules and the Octet Rule

The idea that a molecule could be held together by a shared pair of electrons was first suggested by Lewis in 1916. Although Lewis never won the Nobel prize for this or his many other theories, the shared pair of electrons is nevertheless one of the most significant contributions to chemistry of all time. Wave mechanics was still 10 years in the future, and so Lewis was unable to give any mathematical description of exactly how electron sharing was possible. Instead of the detailed picture presented in the previous section, Lewis indicated the formation of a hydrogen molecule from two hydrogen atoms with the aid of his electron-dot diagrams as follows:

$$H\cdot + \cdot H \longrightarrow H{:}H$$

One H with a small dot plus another H with a small dot forms covalently bonded H that are bonded together through 2 shared dots in the center of the 2 H.

Lewis also suggested that the tendency to acquire a noble-gas structure is not confined to ionic compounds but occurs among covalent compounds as well. In the hydrogen molecule, for example, each hydrogen atom acquires some control over *two* electrons, thus achieving something resembling the helium structure. Similarly the formation of a fluorine molecule from its atoms can be represented by

There are two F symbols shown with seven dots drawn around it. They react to form a covalently bonded Fluorine gas molecule with a complete octet set of electrons around each of the bonded F.

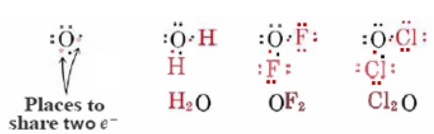
Again a pair of electrons is shared, enabling each atom to attain a neon structure with eight electrons (i.e., an octet) in its valence shell. Similar diagrams can be used to describe the other halogen molecules:

Lewis diagrams showing a complete octet among each of the bonded atoms in a diatomic molecule such as chlorine, fluorine, and iodine.

In each case a shared pair of electrons contributes to a noble-gas electron configuration on *both* atoms. Since only the valence electrons are shown in these diagrams, the attainment of a noble-gas structure is easily recognized as the attainment of a full complement of eight electron dots (an *octet*) around each symbol. In other words *covalent* as well as ionic compounds obey the octet rule.

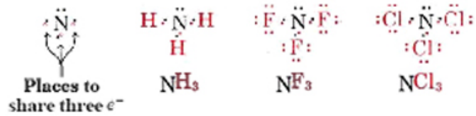
The octet rule is very useful, though by no means infallible, for predicting the formulas of many covalent compounds, and it enables us to explain the usual valence exhibited by many of the representative elements. According to Lewis' theory, hydrogen and the halogens each exhibit a valence of 1 because the atoms of hydrogen and the halogens each contain one less electron than a noble-gas atom. In order to attain a noble-gas structure, therefore, they need only to participate in the sharing of *one* pair of electrons. If we identify a shared pair of electrons with a chemical bond, these elements can only form *one* bond.

A similar argument can be extended to oxygen and the group VI elements to explain their valence of 2. Here *two* electrons are needed to complete a noble-gas configuration. By sharing two pairs of electrons, i.e., by forming two bonds, an octet is attained:



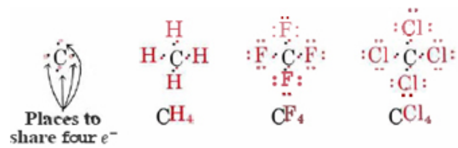

Lewis diagram of O is shown with an arrow pointing to two vacant spots around the atom which are places to share two electrons. The sharing of the electrons to fill the vacant spots are highlighted with examples showing lewis diagrams of H 2 O, O F subscript 2, and C l subscript 2 O.

Nitrogen and the group V elements likewise require three electrons to complete their octets, and so can participate in three shared pairs:




Lewis diagram of N is shown with an arrow pointing to three vacant spots around the atom which are places to share three electrons. The sharing of the electrons to fill the vacant spots are highlighted with examples showing lewis diagrams of N H subscript 3, N F subscript 3, and N C l subscript 3.

Finally, since carbon and the group IV elements have four vacancies in their valence shells, they are able to form four bonds:




Lewis diagram of C is shown with an arrow pointing to four vacant spots around the atom which are places to share four electrons. The sharing of the electrons to fill the vacant spots are highlighted with examples showing lewis diagrams of C H subscript 4, C F subscript 4, and "C" "C l" subscript 4.

### Example 6.14.1 : Lewis Structures

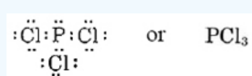
Draw Lewis structures and predict the formulas of compounds containing (a) P and Cl; (b) Se and H.

**Solution:**

**a)** Draw Lewis diagrams for each atom.

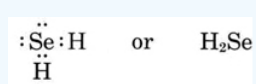
P symbol with two dots on the top and one dot on the left, right, and bottom respectively. C l symbol with two dots on the top, right, and bottom respectively as well as one dot on the left.

Since the P atom can share three electrons and the Cl atom only one, three Cl atoms will be required, and the formula is



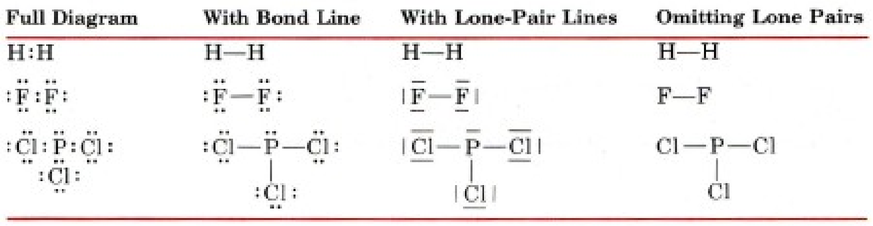

Lewis diagram of P Cl subscript 3. Each of the atom is drawn with a complete octet, meaning 8 dots around each atom.

**b)** Since Se is in periodic group VI, it lacks two electrons of a noble-gas configuration and thus has a valence of 2. The formula is




Lewis diagram of H subscript 2 S e. The s e symbol has eight electrons around it while the two H bonded to S e has two electrons each, completing the duplet rule for hydrogen.

In drawing Lewis structures, the bonding pairs of electrons are often indicated by a **bond line** connecting the atoms they hold together. Electrons which are not involved in bonding are usually referred to as **lone pairs** or **unshared pairs**. Lone pairs are often omitted from Lewis diagrams, or they may also be indicated by lines. Here are some of the alternative ways in which $H_2$, $F_2$, and $PCl_3$ can be written.



| Full Diagram | With Bond Line | With Lone-Pair Lines | Omitting Lone Pairs |
|---|---|---|---|
| H:H | H—H | H—H | H—H |
| :F:F: | :F—F: | \|F—F\| | F—F |
| :Cl:P:Cl: / :Cl: | :Cl—P—Cl: / :Cl: | \|Cl—P—Cl\| / \|Cl\| | Cl—P—Cl / Cl |

A table shows the four different ways a lewis structure can be drawn. There is a full diagram with all electrons shown as dots. There is also with bond line which represents shared electrons as a line, while the unbonded electrons are still drawn out as dots. There is with lone pair lines which shows unpaired electrons as shorts lines around each atom. Finally there is omitting lone pairs which only shows a single line between two atoms that are bonded. Examples are shown for H 2, F 2, and P C l 3.