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:

\[
\text{H} \cdot + \text{H} \quad \longrightarrow \quad \text{H}:\text{H}
\]

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

\[
\begin{align*}
\cdot\text{F} \cdot + \cdot\text{F} \cdot & \quad \longrightarrow \quad \cdot\text{F}::\text{F} \\
\end{align*}
\]

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:

\[
\begin{align*}
\cdot\text{Cl}::\text{Cl} & \\
\cdot\text{Br}::\text{Br} & \\
\cdot\text{I}::\text{I} &
\end{align*}
\]

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:
Example \(\PageIndex{1}\) : 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.

\[
\begin{align*}
\text{P:} & \quad \text{Cl:} \\
\text{H:} & \quad \text{F:} \\
\text{H:} & \quad \text{Cl:}
\end{align*}
\]

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

\[
\text{Cl:}\text{P:Cl:} \quad \text{or} \quad \text{PCl}_3
\]

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

\[
\text{Se:}\text{H} \quad \text{or} \quad \text{H}_2\text{Se}
\]

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.
<table>
<thead>
<tr>
<th>Full Diagram</th>
<th>With Bond Line</th>
<th>With Lone-Pair Lines</th>
<th>Omitting Lone Pairs</th>
</tr>
</thead>
<tbody>
<tr>
<td>H:H</td>
<td>H—H</td>
<td>H—H</td>
<td>H—H</td>
</tr>
<tr>
<td>:F:F:</td>
<td>F—F</td>
<td>F—F</td>
<td>F—F</td>
</tr>
<tr>
<td>:C:F + P:C</td>
<td>C—P—C</td>
<td>C—P—C</td>
<td>C—P—C</td>
</tr>
<tr>
<td>:C:</td>
<td></td>
<td>C—P</td>
<td>C—P</td>
</tr>
</tbody>
</table>

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