Lesson #3 Covalent Bonds

WHY DO ATOMS FORM BONDS?

- Bond formation begins with atoms colliding.
- For example, as two hydrogen atoms approach each other, their kinetic energy increases as each electron cloud is attracted to the other’s approaching positive nucleus.
- As the atoms continue moving together, they overlap enough to cause attractive forces to exceed the repulsive ones.
- The two valence electrons will move into the region of space between the atoms nuclei and convert their kinetic energy into potential energy (bond energy).

**Shared electron pair = covalent bond**

- As the atoms get close to each other, their e- clouds overlap enough to cause attractive forces.
- The two valence electrons will move into the region of space between the atoms nuclei and convert their kinetic energy into potential energy (bond energy).

**Single covalent bond**

- The atoms of noble gases have completely filled outer shells and so are stable.
- This makes the noble gases very unreactive and so they do not usually form bonds.

**A COVALENT BOND** is formed when two atoms complete their octets by sharing one or more pairs of electrons.

- Usually involves a non-metal and a non-metal.
- Both nuclei get to be attracted to more electrons.
- Both atoms complete their valence shells.
- Covalent bonds are very strong.
I. Single (Covalent) Bonds

- Formed when two atoms share a single pair of electrons.
- Simplest examples are the homonuclear diatomic molecules.

Many non-metal elements, such as hydrogen, exist as simple diatomic molecules that contain covalent bonds.

How is a covalent bond formed in diatomic molecules?

Example 1

\[ \text{H} + \text{H} \rightarrow \text{H}_2 \]

\( \text{H}^0 \rightarrow \text{H}^- + \text{H}^+ \)

Overlap of the s orbitals of each atom.

Example 2

\[ \text{F} + \text{F} \rightarrow \text{F}_2 \]

As they bond the s and p orbitals “blend” to form sp-orbitals (2e–)

\( \rightarrow 4 \text{ sp-orbitals.} \)

The number of electrons that an atom can share is usually the same as its valence.

# of unpaired electrons.

Example 3  How many single bonds would you think Carbon could form? Oxygen? Bromine?

\[
\begin{align*}
\text{C} & : [\text{He}] 2s^2 2p^2 \\
\text{O} & : [\text{He}] 2s^2 2p^4 \\
\text{Br} & : [\text{Ar}] 4s^2 4p^5
\end{align*}
\]

Valence electrons:

- C: 4
- O: 6
- Br: 7

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- C: 4
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VSEPR indicates 4 sp-orbitals.

\( \rightarrow \text{4 bonds possible} \)

\( \rightarrow \text{1 bond possible} \)

II. Double & Triple (Covalent) Bonds:

Non-metal atoms with less than 7 valence e–s are able to share more than single pair of e–s.

Examples:

- Oxygen
- Nitrogen
- Carbon (Dioxide)

\( \text{double bond:} \quad 2 \text{e}^- \text{pairs} = 4 \text{e}^+ \text{total} \)

\( \text{triple bond:} \quad 3 \text{e}^- \text{pairs} = 6 \text{e}^+ \text{total} \)
III. Coordinate Covalent Bonds – atoms that have a non-bonding e- pair in their octet can share to allow an "electron-deficient oxygen" to complete its octet thereby.

Examples:

Ammonium

\[
\text{NH}_4^+ \quad \text{NH}_3 + H^+ \quad \text{lost its only } e^- \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H}
\end{array}
\]

Hydronium

\[
\text{H}_3O^+ \quad \text{H}_2O + H^+ \quad \text{only has a 1s orbital.. only holds } 2e^- \quad \begin{array}{c}
\text{H} \\
\text{H} \\
\text{H}
\end{array}
\]
Orbital “blending” during bonding is complex, and certainly extension. It can be helpful in determining the bond type: single, double or triple bond.

<table>
<thead>
<tr>
<th>Bond type</th>
<th>No. of σ bond</th>
<th>No. of π bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Single (C-H)</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>Double (C=C)</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>Triple (C≡C)</td>
<td>1</td>
<td>2</td>
</tr>
</tbody>
</table>

How is the ratio of atoms calculated?

To calculate the ratio of atoms in a stable covalent compound:

1. Work out how many electrons are needed by each non-metal element to complete its outer electron shell.
2. Work out the ratio of atoms that will provide enough shared electrons to fill all the outer shells.

EXAMPLE:

How do carbon and hydrogen atoms form covalent bonds in a molecule of methane?

<table>
<thead>
<tr>
<th>element</th>
<th>C</th>
<th>H</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron configuration</td>
<td></td>
<td></td>
</tr>
<tr>
<td>electrons needed</td>
<td></td>
<td></td>
</tr>
<tr>
<td>ratio of atoms</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
EXERCISE:

68. Which of the following atom pairs would you expect to form covalent bonds when they join?
   (a) S and O
   (b) Ba and O
   (c) Fe and Cl
   (d) N and O
   (e) H and S
   (f) C and H

69. (a) When the distance between two covalently-bonded atoms increases, what happens to the electrostatic attraction of their nuclei to the shared electrons in a covalent bond?
    (b) What would you expect to happen to the strength of the covalent bond between two identical halogen atoms when going down the halogen family from F₂ to I₂?

70. What would you expect to happen to the strength of a covalent bond when the number of shared electrons increases?

71. The distance between the nuclei of two atoms involved in a bond is called the BOND LENGTH. What should happen to the bond length as the number of shared electrons in the bond increases? Why will this happen?

72. Predict the formula of the compound formed by bonding together the following.
   (a) P and Cl
   (b) B and O
   (c) C and S
   (d) P and O
   (e) H and Se
   (f) F and O
   (g) H and O
   (h) N and I
   (i) B and C
   (j) C and Cl
   (k) Si and P
   (l) Si and S