Basic Principles of Element Bonding

Atoms start with a specific number of valence electrons. They will then form bonds to try to fill up their valence shells. This leads to predictable numbers of bonds and non-bonding electrons because first and second row atoms cannot exceed a full shell.

The number of bonds for a neutral atom is equal to the number of electrons in the full valence shell (2 or 8 electrons) minus the number of valence electrons. This method works because each covalent bond that an atom forms adds another electron to an atom's valence shell without changing its charge.

\[
\text{number of bonds (for a neutral atom)} = (\text{full valence shell}) - (\text{number of valence electrons})
\]

<table>
<thead>
<tr>
<th>Atom</th>
<th>Full valence shell</th>
<th>Number VE</th>
<th>typical # bonds</th>
<th>typical # unbonded e's</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2</td>
<td>1</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>4</td>
<td>4</td>
<td>0</td>
</tr>
<tr>
<td>N</td>
<td>5</td>
<td>5</td>
<td>3</td>
<td>2</td>
</tr>
<tr>
<td>O</td>
<td>6</td>
<td>6</td>
<td>2</td>
<td>4</td>
</tr>
<tr>
<td>F, Cl, Br, I</td>
<td>8</td>
<td>7</td>
<td>1</td>
<td>6</td>
</tr>
<tr>
<td>Si</td>
<td>8</td>
<td>4</td>
<td>4</td>
<td>0</td>
</tr>
<tr>
<td>P</td>
<td>8 (or 10)</td>
<td>5</td>
<td>3 (or 5)</td>
<td>2 (or 0)</td>
</tr>
<tr>
<td>S</td>
<td>6 (or 12)</td>
<td>6</td>
<td>2 (or 6)</td>
<td>4 (or 0)</td>
</tr>
</tbody>
</table>

For example, hydrogen typically makes one bond because its full valence shell is 2 and its valence number is 1. Carbon typically makes four bonds because its full valence shell is 8 and its valence number is 4.

This same method can be used to calculate the number of electrons that are not participating in bonding. The number of non-bonding electrons is equal to the number of electrons in a full valence shell minus the number of electrons which are participating in bonding (which is 2 x the typical number of bonds). The number of lone pairs is the number of non-bonding electrons divided by two.

\[
\text{number of non-bonding electrons (for a neutral atom)} = (\text{full valence shell}) - 2 \times (\text{number of bonds})
\]

For example, hydrogen typically has 0 non-bonding electrons. The full valence shell for hydrogen is 2 and the number of electrons in bonds is also 2. The difference is zero. Oxygen typically has 4 non-bonding electrons (or 2 lone pairs). The full valence shell for oxygen is 8 and the number of electrons in bonds is 4. Therefore, the difference is 4.

Ionic and Covalent Bonds

An ionic bond is where one atom takes one valence electron from the other. This will give one atom a positive formal charge and the other a negative formal charge. The two atoms stay together because of the electrostatic attraction of the plus and minus charges.

A covalent bond is where two atoms share two electrons. They do this to try to fill their valence shells. Covalent bonds have well defined lengths (~1.0 to 2.5 Å) and bond strengths (~100 kcal/mol)

One way to predict whether a bond is ionic or covalent is to look how far apart the two atoms forming the bonds are in the periodic table. If one atom is of the far left (Group 1 or 2) and the other is on the far right (Group 5, 6, or 7), then the atoms will have large differences in EN and will form an ionic bond. Most other pairs of atoms are close enough in EN to form covalent bonds. For example, any two atoms in the main group elements (Groups 3, 4, 5, 6, 7), will usually form a covalent bond. Remember to treat hydrogen like a Group 3 element as its EN is closest to boron.
Covalent and ionic are just the two extremes of the different types of bonds. There is actually a continuum of bond types. Most bonds have some covalent and some ionic character. For example, a C-C covalent bond has little or no ionic character because the two atoms have the same EN's. The Si-F is one of the most polarized covalent bonds and has a lot of ionic character. The N-Li bond is ionic but has significant covalent character.

1. The number of electrons in the outer shell
   A) ion
   B) valence
   C) isotope
   D) atomic mass

2. These elements don't bond with other elements because their outer shell is filled.
   A) metals
   B) none of the answers are correct
   C) noble solids
   D) Inert gases

3. Most atoms adopt one of three simple strategies to achieve a filled shell. Which of the following is NOT one of these strategies.
   A) accept electrons
   B) Keep their own electrons
   C) Give away electrons
   D) Share electrons

4. Which of the following is NOT a type of chemical bond?
   A) Metallic
   B) Ionic
   C) Covalent
   D) All of the answers are correct

5. In ionic bonding
   A) Electrons are accepted
   B) Electrons are shared
   C) Electrons are given away
   D) Two answers are correct

6. In ionic bonding electrical forces between same charged ions holds the atoms together.
   A) True
   B) False

7. In covalent bonding
   A) Some electrons are shared by all the atoms.
   B) Bonding takes place between positively charged areas of one molecule with a negatively charged area of another molecule.
   C) One atom takes the outer shell electrons from another atom.
   D) A couple of atoms share their electrons with each other.

8. Which element forms the basis for organic chemistry
   A) hydrogen
   B) oxygen
   C) all of the above
   D) carbon

9. O₂ is an example of what type of bonding
   A) Metallic
   B) Hydrogen
   C) None of the above
   D) Ionic

10. Which particles play the most active role in chemical bonding?
    A) electrons
    B) neutrons
    C) valence electrons
    D) protons