Chemical Bonding
# Chemical Bonding

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1.0 Chemical Bonds

1.1 Types of Chemical Bonds

There are three types of chemical bonds that arise from atomic interactions.

Ionic Bond

An ionic bond is the force of attraction that binds together ions of opposite charges. This occurs by the actual transfer of electrons from one atom to another. Compounds with ionic bonding are usually composed of a metal and a nonmetal, due to the significant differences in their electronegativities. Examples are sodium chloride (NaCl) and potassium iodide (KI).

Covalent Bond

A covalent bond results from the sharing of electrons between two atoms. Molecules with covalent bonding are usually composed of two nonmetals. Examples are carbon dioxide (CO₂) and chlorine gas (Cl₂).

Metallic Bond

Metallic bonds are found in metals such as iron, aluminum and silver. Each atom in metals is bonded to several neighboring atoms. The bonding electrons are relatively free to move throughout the structure of the metal.

1.2 Lewis Symbols

The electrons involved in chemical bonding are the outermost electrons, the valence electrons. The American chemist Gilbert N. Lewis proposed a way to show the valence electrons of atoms. This is known as the Lewis symbol. The Lewis symbol for an element is composed of the chemical symbol for the element plus a dot for each valence electron. The dots can be placed above, below, or on the right and left sides of the symbol for the element. Each side can hold a maximum of two dots. Placing two dots on one side and one dot on another is arbitrary.

For example, the electron configuration of 23Na is 1s²2s²2p⁶3s¹. It has one valence electron. The Lewis symbol for sodium is •Na. The single dot represents the valence electron of sodium. The table below shows more examples of the Lewis symbols of some common elements.
### Table 1.1 Some Elements and their Lewis Symbols

<table>
<thead>
<tr>
<th>Element</th>
<th>Group No.</th>
<th>e- Configuration</th>
<th>Valence e-</th>
<th>Lewis Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1 or IA</td>
<td>1s^1</td>
<td>1</td>
<td>H⁺</td>
</tr>
<tr>
<td>Ca</td>
<td>2 or IIA</td>
<td>1s^22s^22p^63s^23p^64s^2</td>
<td>2</td>
<td>Ca²⁺</td>
</tr>
<tr>
<td>Al</td>
<td>13 or IIIA</td>
<td>1s^22s^22p^63s^23p^1</td>
<td>3</td>
<td>Al⁺</td>
</tr>
<tr>
<td>C</td>
<td>14 or IVA</td>
<td>1s^22s^22p^2</td>
<td>4</td>
<td>C⁺</td>
</tr>
<tr>
<td>N</td>
<td>15 or VA</td>
<td>1s^22s^22p^3</td>
<td>5</td>
<td>N⁺</td>
</tr>
<tr>
<td>O</td>
<td>16 or VIA</td>
<td>1s^22s^22p^4</td>
<td>6</td>
<td>O⁺</td>
</tr>
<tr>
<td>F</td>
<td>17 or VIIA</td>
<td>1s^22s^22p^5</td>
<td>7</td>
<td>F⁺</td>
</tr>
<tr>
<td>Ne</td>
<td>18 or VIII A</td>
<td>1s^22s^22p^6</td>
<td>8</td>
<td>Ne⁺</td>
</tr>
</tbody>
</table>

### 1.3 The Octet Rule

Atoms gain, lose, or share electrons to have the same electron configuration as the noble gas (Group 18) closest to them. The noble gases have very stable electron arrangement. Thus, they have very low electronegativity and electron affinity. They are very unreactive. Because the noble gases (except He) have eight valence electrons, atoms undergoing reaction also end up with eight valence electrons. This is known as the octet rule: atoms tend to gain, lose, or share electrons until they are surrounded with eight valence electrons. An octet of electrons consists of filled s and p subshells on an atom. In the case of hydrogen and lithium, they acquire the configuration of helium to be stable. That is, they tend to have two valence electrons.

Atoms and ions with the same electron configuration are said to be isoelectronic with each other. An example would be the fluoride ion (F⁻) and the neon atom, Ne. A fluorine atom has 9 electrons. A fluoride ion is a fluorine atom that has acquired one more electron, a total of 10 electrons. Its electron configuration is 1s²2s²2p⁶. Neon on the other hand has 10 electrons. Its electron configuration is 1s²2s²2p⁶. Both have the same electron configuration and are isoelectronic with each other.
2.0 Ionic Compounds

2.1 Bond Formation

Ionic compounds are formed by gaining or losing electrons. In general, ionic compounds are formed between a cation and an anion. Metals tend to form cations because of their low electronegativity. They generally have less than four valence electrons and so their tendency is to give up the extra valence electrons to be left with an octet of electrons in their valence shells. Nonmetals on the other hand tend to form anions because of their high electron affinity. They generally contain more than 4 valence electrons and so their tendency is to seek for a few more electrons to have a complete set of eight valence electrons. When sodium and chlorine react together to form sodium chloride, sodium loses one electron to chlorine. Sodium has an electron configuration of $1s^22s^22p^63s^1$ or $[Ne]3s^1$. It will lose the 3s electron to chlorine to end up with a valence configuration of $2s^22p^6$, which is the same with that of neon. It now contains 8 electrons in its valence shell. Chlorine has an electron configuration of $1s^22s^22p^63s^23p^5$ or $[Ne]3s^23p^5$. It has 7 electrons in its valence shell, the 3rd energy level. It will accept the electron from sodium so that it will end up with a configuration of $1s^22s^22p^63s^23p^6$, which is the same with that of argon.

The principal reason for the stability of ionic compounds is the attraction between ions of unlike charges. This attraction draws the ions together, releasing energy and causing the ions to form a solid array or lattice. The sodium chloride lattice is shown in figure 2.2.

---

Figure 2.1 The Bonding of Sodium and Chlorine

![Figure 2.1 The Bonding of Sodium and Chlorine](image1)

Figure 2.2 The Crystal Lattice of Sodium Chloride

![Figure 2.2 The Crystal Lattice of Sodium Chloride](image2)
The measure of the stability of the lattice formed by oppositely charged ions in an ionic solid is given by its lattice energy. The lattice energy is the energy needed to completely separate a mole of a solid ionic compound into its gaseous ions. The higher the lattice energy, the more difficult it is to convert solid ionic compounds into their component ions in the gaseous state. The lattice energy of sodium chloride is +788 kJ/mol. It is the amount of energy needed to break the bond in the solid compound to produce the gaseous chloride and sodium ions.

\[
\text{NaCl} (s) \rightarrow \text{Na}^{+} (g) + \text{Cl}^{-} (g) \quad \Delta H_{\text{lattice}} = +788 \text{ kJ/mol}
\]

The positive sign on the lattice energy indicates that the process is endothermic as energy is absorbed. The opposite process, the combining together of Na\(^+\) (g) and Cl\(^-\) (g) to form NaCl (s), on the other hand, is an exothermic process (\(\Delta H = -788 \text{ kJ/mol}\)). Table 2.1 lists the lattice energies of some ionic compounds. All are large positive values, indicating that the ions are strongly attracted to one another in these solids. The strong attractions also cause most ionic compounds to be strong and brittle, with high melting points (NaCl melts at 801°C). Furthermore, for a given arrangement of ions, the lattice energy increases as the charges on the ions increase and as their radii decrease. Thus, magnesium chloride (MgCl\(_2\)) has higher lattice energy than sodium chloride (NaCl).

<table>
<thead>
<tr>
<th>Compound</th>
<th>Lattice Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaF</td>
<td>910</td>
</tr>
<tr>
<td>NaCl</td>
<td>788</td>
</tr>
<tr>
<td>NaI</td>
<td>682</td>
</tr>
<tr>
<td>MgCl(_2)</td>
<td>2326</td>
</tr>
<tr>
<td>SrCl(_2)</td>
<td>2127</td>
</tr>
<tr>
<td>CaO</td>
<td>3414</td>
</tr>
</tbody>
</table>

2.2 Electron Configurations of the Representative Elements

The representative elements are the elements under the groups 1-3 and 4-8. The energetics of ionic bond formation helps explain why many ions tend to have noble-gas electron configuration. Sodium (Group 1), for example, tends to lose 1 electron to become Na\(^+\), which has the same electron configuration as Ne:

\[
\text{Na} \quad 1s^22s^22p^63s^1 = [\text{Ne}]3s^1
\]

\[
\text{Na}^+ \quad 1s^22s^22p^6 = [\text{Ne}]
\]
The electron lost by sodium comes from the outermost subshell, the 3s. The loss of a second electron would have to come from the inner energy level. This is energetically unfavorable as this would require great amount of energy because it is already closer to the nucleus. The lattice energy is not enough to compensate for the energy required to remove a second electron from an inner shell. This is the reason why we never find compounds that contain Na\(^{+2}\) ions. Thus, sodium and other Group 1 metals are found in ionic substances only as +1 ions.

The addition of electrons to nonmetals is usually exothermic. The electrons are being added to the valence shell. Chlorine, for example, acquires an electron to form Cl\(^-\), which has the same electron configuration as Ar:

\[
\text{Cl} \quad 1s^22s^22p^63s^23p^5 = [\text{Ne}]3s^23p^5 \\
\text{Cl}^- \quad 1s^22s^22p^63s^23p^6 = [\text{Ne}]3s^23p^6 = [\text{Ar}]
\]

The second electron would have to be added to the next higher shell. This is energetically very unfavorable. Therefore we never observe Cl\(^{+2}\) in ionic compounds.

Using this concept, we expect that ionic compounds of the representative metals from groups 1, 2 and 13 will contain cations with charges of +1, +2 and +3, respectively. Likewise, ionic compounds with representative nonmetals from groups 15, 16 and 17 usually contain anions of charges -3, -2 and -1, respectively. Ionic compounds of the nonmetals from group 14 (C, Si and Ge) are rarely found. The heaviest elements in group 14 (Sn and Pb) are metals and are usually found as +2 cations in ionic compounds. This is consistent with the increasing metallic character as one proceeds down a group in the periodic table.

### 2.3 Transition Metal Ions

The transition metals are located on the inner lower box of the periodic table. With the transition metals, the lattice energies are usually large enough to compensate for the removal of up to 3 or 4 electrons. This is despite the fact that ionization energies increase rapidly for each successive electron removed. Thus we find cations with charges of +1, +2, +3 or +4 in ionic compounds. Most transition metals have more than 3 electrons beyond a noble gas core. Silver, for example, has a [Kr]4d\(^{10}\)5s\(^1\) electron configuration. The 5s electron is lost in forming Ag\(^+\). A completely filled 4d subshell results. As in this example, transition metals generally do not form ions with noble gas configurations. The octet rule is limited in scope.

### 2.4 Polyatomic Ions

Many cations and anions are polyatomic. They are composed of more than one atom. Examples are the ammonium cation (NH\(_4^+\)), sulfate anion (SO\(_4^{2-}\)) and the nitrate anion (NO\(_3^-\)). In polyatomic ions, two or more atoms are bonded together by predominantly covalent bonds. They form
grouping that carries a positive or a negative charge. A polyatomic ion, although composed of more than one atom, acts as if it is monatomic. Tables 2.4 and 2.5 list the common polyatomic ions.

2.5 Writing Formulas of Ionic Compounds

2.5.1 Representative Metal Ions

Writing formulas of ionic compounds involving the representative elements is simple. Always remember that the elements under groups 1, 2, 13, 15, 16 and 17 form ions with charges of +1, +2, +3, -2, -3 and -1, respectively. When sodium combines with chlorine to form sodium chloride, the ions would be Na\(^+\) and Cl\(^-\). All we have to do is criss-cross the charges such that the magnitude of the charge of one ion is the subscript of the other. The same is true with polyatomic ions. If the subscript is just 1, however, there is no need to write it anymore. This also means that if an atom in a chemical formula does not contain any subscript, the subscript is just 1. The subscripts are also reduced to their lowest terms except for some, such as hydrogen peroxide, H\(_2\)O\(_2\). The symbol for the less electronegative atom is written first; that is, the symbol for the cation.

When sodium and chlorine combine together to form sodium chloride,

\[
\text{Na}^+ \quad + \quad \text{Cl}^- \rightarrow \text{NaCl}
\]

Sodium acquires a +1 charge since it is under group 1; chlorine has -1, since it belongs under group 17. The magnitude of the charges of sodium ion and chloride ion are both 1. Sodium has a subscript of 1 and so does chlorine when they form a compound. But since the subscripts are just 1, they need not be written anymore. Moreover, the symbol for sodium (the less electronegative element) is written first followed by the symbol for chlorine (the more electronegative element).

When calcium and chlorine combine together to form calcium chloride,

\[
\text{Ca}^{+2} \quad + \quad \text{Cl}^- \rightarrow \text{CaCl}_2
\]

Calcium acquires a +2 charge. It is under group 2 in the periodic table. Chlorine, on the other hand, will have -1 as it is under group 17. Criss-crossing the charges gives calcium a subscript of 1 while chlorine has 2. The symbol for calcium, the cation, is written first, followed by the symbol for chlorine, the anion.

When barium combines with sulfur to form barium sulfide,

\[
\text{Ba}^{+2} \quad + \quad \text{S}^{2-} \rightarrow \text{BaS}
\]
Barium has a charge of +2 while sulfur has -2. When the formula for the compound they form is written, no subscript is seen on either of the ions. This is because they both have a subscript of 2. It is supposed to be $\text{Ba}_2\text{S}_2$. However, as a rule, the subscripts have to be in their lowest terms. In this case, the subscripts can be reduced to 1 by dividing each by 2. And so the formula for barium sulfide is $\text{BaS}$.

2.5.2 Polyatomic Ions

Writing formulas of ionic compounds involving polyatomic ions is very similar to those involving monatomic ions. When ammonium cation and sulfate anion combine together, the resulting ionic compound is ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$. The parentheses are used to separate the subscripts 4 and 2 from each other. Otherwise, the subscript would be 42, as in $\text{NH}_4\text{SO}_4$. It would mean that there are 42 atoms of hydrogen. In the correct one, it means that there are 2 molecules of the ammonium cation. There is no periodic trend with regard to the charges of polyatomic ions. You need to familiarize yourself with the charges of polyatomic ions. The charge on the ammonium cation is +1, while that on the sulfate anion is -2.

$$\text{NH}_4^+ + \text{SO}_4^{2-} \rightarrow (\text{NH}_4)_2\text{SO}_4$$

In the reaction of ammonium cation with a nitrate anion to form ammonium nitrate, notice that the formula for the product does not anymore contain a parenthesis. This is because there are no two succeeding subscripts to be separated. The charge on the ammonium cation is +1, while that on nitrate anion is -1. Again, when the subscripts are just 1, they need not be written out.

$$\text{NH}_4^+ + \text{NO}_3^{-} \rightarrow \text{NH}_4\text{NO}_3$$

Some common examples of formulas of ionic compounds with subscripts of which are not reduced to the lowest terms are hydrogen peroxide ($\text{H}_2\text{O}_2$) and mercury(I) chloride ($\text{Hg}_2\text{Cl}_2$).

2.5.3 Transition Metal Ions

If the ionic compound involves a transition metal, the formula can be correctly written if the charge of the metal is specified. This is because many transition metals occurring in nature have more than 1 oxidation states. Examples are Fe(III) and Fe(II), Pb(IV) and Pb(II), among others. In formula writing, the same rules apply. If iron(III) combines with oxygen in the air to form rust, the reaction would be

$$\text{Fe}^{3+} + \text{O}^{2-} \rightarrow \text{Fe}_2\text{O}_3$$
When copper(II) forms a compound with sulfate anion,

$$\text{Cu}^{+2} + \text{SO}_4^{-2} \rightarrow \text{CuSO}_4$$

Copper can form a +1 and a +2 cation. In this case, it is specified that the copper cation used is in the +2 oxidation state. Sulfate has an oxidation state or a charge of -2. When the magnitude of the charges is criss-crossed, both ions acquire a subscript of 2. They can still be reduced to 1. As this is the case, there are no parentheses written on the formula for the product.

The formula of the ionic compounds can also be used in determining the charge that an ion carries. If the formula is BaF$_2$, it also means that the charge of Ba is +2 and that of F is -1. If we have Al$_2$S$_3$, then the charge of Al is +3 and that of S is -2.

2.6 Nomenclature of Ionic Compounds

In general, the name of the cation is written first followed by the name of the anion.

2.6.1 Representative Metal Cation

If the cation in an ionic compound is a representative metal (group 1, 2, and 3 metal), its name is as how it is called when it is a neutral atom. Or, its name is as how it appears in the periodic table. In NaCl, the cation is Na$^+$. So we write “sodium” first. In BaCl$_2$, the cation is Ba$^{+2}$. We write “barium” first. They are then followed by the name of the corresponding anions.

2.6.2 Monatomic Anion

If the anion is monatomic (one kind of atom), its name is changed to have an –ide ending. In NaCl, the anion is Cl$. Its name is chloride instead of chlorine. The name of the ionic compound is sodium chloride. In CaO, the anion is O$^{2-}$. Its name is oxide instead of oxygen. The name of the compound is calcium oxide. The table below lists some common anions with an –ide ending in their names.

<table>
<thead>
<tr>
<th>Anion</th>
<th>Name with –ide Ending</th>
<th>Anion</th>
<th>Name with –ide Ending</th>
</tr>
</thead>
<tbody>
<tr>
<td>F$^-$</td>
<td>Fluoride</td>
<td>S$^{2-}$</td>
<td>Sulfide</td>
</tr>
<tr>
<td>Cl$^-$</td>
<td>Chloride</td>
<td>Se$^{-2}$</td>
<td>Selenide</td>
</tr>
<tr>
<td>Br$^-$</td>
<td>Bromide</td>
<td>N$^{-3}$</td>
<td>Nitride</td>
</tr>
<tr>
<td>I$^-$</td>
<td>Iodide</td>
<td>P$^{3-}$</td>
<td>Phosphide</td>
</tr>
<tr>
<td>O$^{2-}$</td>
<td>Oxide</td>
<td>C$^{4-}$</td>
<td>Carbide</td>
</tr>
</tbody>
</table>
2.6.3 Transition Metal Cations

Transition metals can have more than one oxidation states. The table below lists some common transition metals and their oxidation states. In naming compounds containing transition metal cations, the oxidation states have to be specified. There are two ways to name ionic compounds containing transition metal cations: the Classical Method and the Stock System.

The Classical Method

The classical method of naming ionic compounds is making use of the Latin names of some transition metals. The Latin name is added with a suffix –ic to indicate the higher charge of the metal and –ous to indicate the lower charge of the metal. This method is however limited almost only to those transition metals with only two different charges or oxidation states. Some metals can have more than two oxidation states. The table below shows the Latin names of some transition metals.

Table 2.3 Some transition metals and their Latin names

<table>
<thead>
<tr>
<th>Transition Metal</th>
<th>Symbol</th>
<th>Latin Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>cuprum</td>
</tr>
<tr>
<td>Gold</td>
<td>Au</td>
<td>aurum</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe</td>
<td>ferrum</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb</td>
<td>plumbum</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg</td>
<td>hydrargyrum</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag</td>
<td>argentum</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn</td>
<td>stannum</td>
</tr>
</tbody>
</table>

Iron can have two oxidation states, +2 and +3. If the cation involved in an ionic bonding is Fe\(^{+2}\), its name is ferrous. If the one involved is Fe\(^{+3}\), its name is ferric. If the compound is the combination of Fe\(^{+2}\) and Cl\(^-\), then its name is ferrous chloride. If the compound is composed of Fe\(^{+3}\) and S\(^{2-}\), its name is ferric sulfide.

The Stock System

As mentioned above, the classical method has limitations. And so another method of naming ionic compounds involving transition metals has been developed. It is the system that is widely used today, the stock system. In this method, the oxidation state of the transition metal involved in the bonding is represented by an equivalent Roman numeral enclosed in parentheses. It is placed immediately after the name of the transition metal. The name of the transition metal used is how it is named in the periodic table. If the metal is Fe\(^{+2}\), its name is iron(II). The name of the metal is
iron, which is how it is named in the periodic table. The oxidation state is +2 and its Roman numeral equivalent is II. The Roman numeral is enclosed in parentheses and is placed right after the name of the metal. If the metal is Fe$^{+3}$, its name is iron(III). If the compound is made by the combination of Fe$^{+2}$ and Cl$^{-}$, then its name is iron(II) chloride. If it is made by combining together Fe$^{+3}$ and O$^{2-}$, then the name is iron(III) oxide.

<table>
<thead>
<tr>
<th>Element</th>
<th>Charge</th>
<th>Name of Ion (Classical System)</th>
<th>Name of Ion (Stock System)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cr</td>
<td>+2</td>
<td>Chromous ion</td>
<td>Chromium(II) ion</td>
</tr>
<tr>
<td></td>
<td>+3</td>
<td>Chromic ion</td>
<td>Chromium(III) ion</td>
</tr>
<tr>
<td>Co</td>
<td>+2</td>
<td>Cobaltous ion</td>
<td>Cobalt(II) ion</td>
</tr>
<tr>
<td></td>
<td>+3</td>
<td>Cobaltic ion</td>
<td>Cobalt(III) ion</td>
</tr>
<tr>
<td>Cu</td>
<td>+1</td>
<td>Cuprous ion</td>
<td>Copper(I) ion</td>
</tr>
<tr>
<td></td>
<td>+2</td>
<td>Cupric ion</td>
<td>Copper(II) ion</td>
</tr>
<tr>
<td>Fe</td>
<td>+2</td>
<td>Ferrous ion</td>
<td>Iron(II) ion</td>
</tr>
<tr>
<td></td>
<td>+3</td>
<td>Ferric ion</td>
<td>Iron(III) ion</td>
</tr>
<tr>
<td>Pb</td>
<td>+2</td>
<td>Plumbous ion</td>
<td>Lead(II) ion</td>
</tr>
<tr>
<td></td>
<td>+4</td>
<td>Plumbic ion</td>
<td>Lead(IV) ion</td>
</tr>
<tr>
<td>Hg</td>
<td>+1</td>
<td>Mercurous ion</td>
<td>Mercury(I) ion</td>
</tr>
<tr>
<td></td>
<td>+2</td>
<td>Mercuric ion</td>
<td>Mercury(II) ion</td>
</tr>
<tr>
<td>Ni</td>
<td>+2</td>
<td>Nickelous ion</td>
<td>Nickel(II) ion</td>
</tr>
<tr>
<td></td>
<td>+3</td>
<td>Nickelic ion</td>
<td>Nickel(III) ion</td>
</tr>
<tr>
<td>Ag</td>
<td>+1</td>
<td>Silver ion</td>
<td>Silver ion</td>
</tr>
<tr>
<td>Sn</td>
<td>+2</td>
<td>Stannous ion</td>
<td>Tin(II) ion</td>
</tr>
<tr>
<td></td>
<td>+4</td>
<td>Stannic ion</td>
<td>Tin(IV) ion</td>
</tr>
</tbody>
</table>

2.6.4 Polyatomic Ions

If the ion involved is a polyatomic ion, its name is unchanged. No prefix or suffix is added. The names of the polyatomic cations and anions have to be familiarized.

<table>
<thead>
<tr>
<th>Element</th>
<th>Charge</th>
<th>Name of Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>+2</td>
<td>Stannous ion</td>
</tr>
<tr>
<td></td>
<td>+4</td>
<td>Stannic ion</td>
</tr>
</tbody>
</table>

Table 2.4 Polyatomic Cations and Their Names
<table>
<thead>
<tr>
<th>Polyatomic cation</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄⁺</td>
<td>Ammonium</td>
</tr>
<tr>
<td>H₃O⁺</td>
<td>Hydronium</td>
</tr>
<tr>
<td>NO⁺</td>
<td>Nitrosyl</td>
</tr>
<tr>
<td>VO⁺²</td>
<td>Vanadyl</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Polyatomic Anion</th>
<th>Name</th>
<th>Polyatomic Anion</th>
<th>Name</th>
<th>Polyatomic Anion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₂H₃O₂⁻</td>
<td>Acetate</td>
<td>CO₃⁻₂</td>
<td>Carbonate</td>
<td>AlO₃⁻³</td>
<td>Aluminate</td>
</tr>
<tr>
<td>HCO₃⁻</td>
<td>Bicarbonate</td>
<td>CrO₄⁻²</td>
<td>Chromate</td>
<td>AsO₄⁻³</td>
<td>Arsenate</td>
</tr>
<tr>
<td>HSO₄⁻</td>
<td>Bisulfate</td>
<td>Cr₂O₇⁻²</td>
<td>Dichromate</td>
<td>AsO₃⁻³</td>
<td>Arsenite</td>
</tr>
<tr>
<td>BrO₃⁻</td>
<td>Bromate</td>
<td>C₂O₄⁻²</td>
<td>Oxalate</td>
<td>BiO₄⁻³</td>
<td>Bismuthate</td>
</tr>
<tr>
<td>ClO₄⁻</td>
<td>Perchlorate</td>
<td>MoO₄⁻²</td>
<td>Molybdate</td>
<td>BO₃⁻³</td>
<td>Borate</td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>Chlorate</td>
<td>O₂⁻²</td>
<td>Peroxide</td>
<td>PO₄⁻³</td>
<td>Phosphate</td>
</tr>
<tr>
<td>ClO₂⁻</td>
<td>Chlorite</td>
<td>MnO₄⁻²</td>
<td>Permanganate</td>
<td>PO₃⁻³</td>
<td>Phosphate</td>
</tr>
<tr>
<td>ClO⁻</td>
<td>Hypochlorite</td>
<td>SO₄⁻²</td>
<td>Sulfate</td>
<td>Fe(CN)₆⁻³</td>
<td>Ferricyanide</td>
</tr>
<tr>
<td>CN⁻</td>
<td>Cyanide</td>
<td>SO₃⁻²</td>
<td>Sulfite</td>
<td>Fe(CN)₆⁻⁴</td>
<td>Ferrocyanide</td>
</tr>
<tr>
<td>BrO⁻</td>
<td>Hypobromite</td>
<td>B₄O₄⁻²</td>
<td>Tetraborate</td>
<td>SiO₄⁻⁴</td>
<td>Silicate</td>
</tr>
<tr>
<td>IO₄⁻</td>
<td>Periodate</td>
<td>S₄O₆⁻²</td>
<td>Tetrathionate</td>
<td>P₂O₇⁻⁴</td>
<td>Pyrophosphate</td>
</tr>
<tr>
<td>IO⁻</td>
<td>Iodate</td>
<td>S₂O₃⁻²</td>
<td>Thiosulfate</td>
<td>As₂O₇⁻⁴</td>
<td>Pyroarsenate</td>
</tr>
<tr>
<td>IO₂⁻</td>
<td>Iodite</td>
<td>ZnO₂⁻²</td>
<td>Zincate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>Nitrate</td>
<td>SnO₃⁻²</td>
<td>Stannate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>Nitrite</td>
<td>PbO₂⁻²</td>
<td>Plumbite</td>
<td></td>
<td></td>
</tr>
<tr>
<td>OH⁻</td>
<td>Hydroxide</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MnO₄⁻</td>
<td>Permanganate</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>SCN⁻</td>
<td>Thiocyanate</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 2.5 Polyatomic Anions and Their Names

If the compound is made by combining together NH₄⁺ and NO₃⁻, the name of the compound is ammonium nitrate.
If the ionic compound is formed by Na\(^+\) and CO\(_3\)^{2-}, the name of the compound is sodium carbonate. If the compound is produced by combining together NH\(_4\)^+ and OH\(^-\), the name of the compound is ammonium hydroxide.

2.7 Writing Formulas of Ionic Compounds Given Their Names

With the above knowledge on how the names of ionic compounds are written, we can easily come up with the formula of an ionic compound. Represent the cation by its symbol. The same is done with the anion. If the name is calcium iodide, we know that the metal is calcium as it is written first. Calcium is under group 2 in the periodic table and so it will have a charge of +2. The cation is Ca\(^{+2}\). The nonmetal is iodine and the anion is I\(^-\) since iodine is under group 17 in the periodic table. We form the compound by criss-crossing the charges to come up with CaI\(_2\). If the name of the compound is lead(II) nitride, the cation is Pb\(^{+2}\). Pb is the symbol for lead and the charge or oxidation state is specified to be +2. The anion is the nitride ion. It is nitrogen. Since nitrogen is under group 15, its charge is -3. So the anion is N\(^-3\). When we criss-cross the charges, we come up with Pb\(_3\)N\(_2\). If the ionic compound has the name ammonium nitrate, we know that both the cation and anion are polyatomic. Going back to the table above, ammonium ion is NH\(_4\)^+ and nitrate ion is NO\(_3\)^-. When the charges are criss-crossed, the formula would be NH\(_4\)NO\(_3\). If the ionic compound is ammonium ferricyanide, the cation is NH\(_4\)^+ and the anion is Fe(CN)\(_6\)^{3-}. When we criss-cross the charges, we come up with the formula (NH\(_4\))\(_3\)Fe(CN)\(_6\).

2.8 Properties of Ionic Compounds

Ionic compounds are usually brittle substances with high melting points. They are usually crystalline, meaning that the solids have flat surfaces that make characteristic angles with one another. Ionic substances can often be cleaved - they break apart along a smooth, flat surface. These characteristics result from electrostatic forces that maintain the ions in a rigid, well-defined three-dimensional arrangement such as that shown in Figure 2.2.

3.0 Covalent Bonding

3.1 Lewis Structures

A chemical bond formed by sharing a pair of electrons is known as a covalent bond. Its formation can be represented by Lewis symbols for the constituent atoms. The formation of H\(_2\) molecules can be represented as shown in Figure 3.1.
Each hydrogen atom acquires a second electron. Each of them acts as if they have sole ownership of the electrons being shared. Both achieve the stable, two-electron, noble gas configuration of helium. The shared electrons are represented by dots. They are valence electrons. In the dash formula, the shared valence electrons are represented by a dash. A single dash indicates that two electrons are being shared. If there are two dashes, then four electrons are being shared. In the dash formula, the valence electrons not involved in the bonding may or may not be shown. They are called lone pairs of electrons or nonbonding electrons. The electrons involved in the bonding are known as the bonding electrons. In the case of the H₂ molecule, there are two bonding electrons and no lone pair of electrons.

The bonding of two fluorine atoms can be represented as shown in the figure below.
Fluorine is under group 17 of the periodic table. Its electron configuration is 1s\(^2\)2s\(^2\)2p\(^5\). It has 7 valence electrons. It lacks one more electron to be isoelectronic with neon. The Lewis symbol for a fluorine atom is

\[ \text{F} \]

When two atoms of fluorine combine together to form a molecule, the resulting Lewis structure is

\[ \text{F} \cdot \text{F} \]

By sharing the bonding pair of electrons, each fluorine atom has eight valence electrons. It thus achieves the noble gas configuration of neon. The two electrons between the two fluorine atoms are the ones being shared. In the dash formula, they are connected with a line to represent the bonding electrons.

3.2 Multiple Bonds

The sharing of a pair of electrons constitutes a single covalent bond. This is generally referred to as a single bond. In many molecules, atoms attain a complete set of eight valence electrons by sharing more than one pair of electrons between them. When two electron pairs are shared, two lines are drawn. These represent a double bond. When three electron pairs are being shared, this is a triple bond, represented by three lines. An oxygen atom is under group 16 in the periodic table. It has six valence electrons. It lacks two more electrons. When two oxygen atoms share electrons, two electron pairs are shared between them for each atom to achieve a complete set of eight valence electrons. Hence, a double bond is formed. A triple bond is exhibited by a nitrogen molecule, N\(_2\). Each nitrogen atom has five valence electrons and so three electron pairs must be shared to achieve the octet configuration.

Figure 3.5 The Lewis Structures of a.) N\(_2\) and b.) O\(_2\)
3.3 Properties of Covalent Molecules
A covalent bond is weaker than an ionic bond. Hence, covalent molecules have generally lower melting and boiling points. They are generally liquids and gases such as water (H₂O) and nitrogen gas (N₂). Many of them are volatile and do not conduct electricity (non-electrolyte).

3.4 Electronegativity
Electronegativity is defined as the ability of an atom in a molecule to attract electrons to itself. The greater the electronegativity of an atom, the greater is its ability to attract electrons to itself. Electronegativity is related to ionization energy and electron affinity. Ionization energy is a measure of how strongly an atom holds its electrons. Electron affinity is a measure of how strongly an atom attracts additional electrons. It has a negative value because it is an exothermic process. An atom with a very negative electron affinity and high ionization energy will strongly attract electrons from other atoms in a molecule. It also resists having its electrons taken away. Both ionization energy and electron affinity are properties of isolated atoms. The first and most widely used electronegativity scale was developed by Linus Pauling. The values are unitless. Fluorine, being the most electronegative atom in the periodic table, has a value of 4.0. The least electronegative, cesium, has a value of 0.7. The periodic trend in electronegativity is that going across a period in the periodic table, electronegativity increases. Nonmetals are more electronegative than metals. Going down a group in the periodic table, electronegativity decreases. Chlorine is more electronegative than sodium. Fluorine is more electronegative than iodine.

3.5 Bond Polarity
If the electrons are shared equally by the atoms in a molecule, the covalent bond is nonpolar. Otherwise, it is polar. Bond polarity arises due to the difference in the electronegativity of atoms in a molecule. Fluorine molecule, F₂, is nonpolar. The electrons are shared equally because both fluorine atoms have equal electronegativity. The force that pulls the electrons toward one fluorine atom is the same as that of the other. A hydrogen fluoride molecule, HF, is polar. Hydrogen is less electronegative than fluorine. Thus, fluorine has a stronger pull of the electrons towards itself. This causes the fluorine atom to have a partial negative charge and hydrogen atom to have a partial positive charge. Lithium fluoride, LiF, is also a polar substance. Lithium is less electronegative than fluorine. The electronegativity difference is greater in lithium fluoride than in hydrogen fluoride. This also means that lithium fluoride is more polar than hydrogen fluoride. The transfer of electronic charge is essentially complete in lithium fluoride. The resulting bond is therefore ionic. In figure 3.6, a region of high electron density is shown in red. It can be noticed that the intensity of the red color around the fluorine atom is greater in lithium fluoride than in hydrogen fluoride. This simply means that the electron cloud is denser around the fluorine atom in lithium fluoride than in hydrogen fluoride. The model for the fluorine molecule shows no red color as there is equal distribution of electrons between the bonded fluorine atoms.
The electronegativity data can be used to predict the bond that exists between two bonded atoms.

Let's have hydrogen molecule, $H_2$, as an example. The electronegativity value of a hydrogen atom is 2.10. The electronegativity difference, $\Delta EN$, of the two hydrogen atoms is 0. Therefore, the bond is a nonpolar covalent bond. In potassium chloride,

\[
\begin{align*}
\text{EN of } K &= 0.9 \\
\text{EN of } Cl &= 3.0 \\
\Delta EN &= (\text{EN of } Cl) - (\text{EN of } K) = 3.0 - 0.9 = 2.1
\end{align*}
\]

Since the electronegativity difference is greater than 1.7, the K-Cl bond is ionic. The strength of the bonds in ascending order is nonpolar covalent bond $> \text{polar covalent bond} > \text{ionic bond}$.

We can indicate the polarity of a molecule in two ways. Using hydrogen fluoride as an example, the polarity of compounds can be represented as

$\delta^+$ $\delta^-$ or $+\quad \rightarrow$

$\text{H} - \text{F}$

Table 3.1 A General Guide on Predicting Bond Type Through $\Delta EN$

<table>
<thead>
<tr>
<th>Electronegativity Difference ($\Delta EN$)</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>$&gt;1.7$</td>
<td>Ionic Bond</td>
</tr>
<tr>
<td>$0.4 - 1.7$</td>
<td>Polar Covalent Bond</td>
</tr>
<tr>
<td>$&lt;0.4$</td>
<td>Nonpolar Covalent Bond</td>
</tr>
</tbody>
</table>

Figure 3.6 Models Showing the Electron Density of a.) $F_2$, b.) HF and c.) LiF
The symbol “\( \delta^+ \)” is read as “delta plus” and “\( \delta^- \)” is “delta minus”. The plus sign indicates the more positive end of the molecule, while the negative sign indicates otherwise. The arrowhead indicates the direction wherein electron density is high. The tail of the arrow is crossed to indicate the more positive end of the molecule.

![Figure 3.7 The Electronegativity Values of Elements](image)

### 3.6 Dipole Moments

The dipole moment of a molecule is the product of the partial charges and the distance between the bonded atoms. Greater distance between charges means greater difference in electronegativities of the two atoms, hence, greater dipole moment. The dipole moment, denoted as \( \mu \), is the quantitative measure of the magnitude of a dipole. Dipole moment of molecules is usually reported in debyes (D), a unit that equals \( 3.34 \times 10^{-30} \) coulomb-meter (C·m). The greater the dipole moment, the greater is the polarity of the two bonded atoms. Polarity helps determine many of the properties of substances that we observe. Polar molecules align themselves with respect to each other and with respect to ions. The negative end of one molecule and the positive end of another attract each other. Likewise, the positive end of a polar molecule is attracted to a negative ion and its negative end is attracted to a positive ion. Not all polar bonds form a polar molecule. Polarity is predicted if the charges in a molecule are separated such that there are some distances between the two charges. Polar covalent molecule indicates a direction along which the charge is separated at one end than the other. An example of a nonpolar molecule consisting of polar bonds is carbon tetrachloride. All the four C-Cl bonds are polar, with carbon being less...
electronegative than chlorine. The whole molecule however is nonpolar because there is no net dipole moment. If we look at each polar bond as a force directed towards chlorine, all the four forces cancel each other. This is because the structure of a $\text{CCl}_4$ is tetrahedral. The force going up is cancelled by the force going down. The force going to the right is cancelled by the force going to the left. There is no net dipole moment. Hence, the molecule is nonpolar covalent. In a water molecule, the dipole on one of the O-H bond is not cancelled by the dipole on the other O-H bond as they are not directed in opposite directions. Water has a bent structure and not linear. Thus, there is a net dipole moment. The molecule is polar covalent. Figure 3.7 presents the molecular structure of $\text{CCl}_4$ and $\text{H}_2\text{O}$.

Figure 3.8  $\text{CCl}_4$ has a Tetrahedral Structure while $\text{H}_2\text{O}$ has a Bent Structure

The bond length of two bonded atoms can also be predicted by the electronegativity difference and hence, the dipole moment. In the table below, data on electronegativity differences, dipole moments and bond lengths of the different hydrogen halides are presented. As we proceed from HF to HI, the electronegativity difference between the bonded atoms decreases. The same is true with the dipole moment. On the other hand, the bond lengths increase. In general, the bond length is inversely proportional to the electronegativity difference and the dipole moment. As the electronegativity difference and the dipole moment decrease, the bond length increases.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Bond Length, Å</th>
<th>Electronegativity Difference ($\Delta$EN)</th>
<th>Dipole Moment, D</th>
</tr>
</thead>
<tbody>
<tr>
<td>HF</td>
<td>0.92</td>
<td>1.9</td>
<td>1.82</td>
</tr>
<tr>
<td>HCl</td>
<td>1.27</td>
<td>0.9</td>
<td>1.08</td>
</tr>
<tr>
<td>HBr</td>
<td>1.41</td>
<td>0.7</td>
<td>0.82</td>
</tr>
<tr>
<td>HI</td>
<td>1.61</td>
<td>0.4</td>
<td>0.44</td>
</tr>
</tbody>
</table>

3.7 Naming Binary Covalent Compounds

Binary covalent compounds are composed of two different nonmetals. $\text{CCl}_4$ is composed of two kinds of nonmetals, carbon and chlorine. Another binary covalent compound is $\text{H}_2\text{O}$. In naming binary covalent compounds, the following rules apply.
1. The element with the lower group number is written first in the name; the element with the higher group number is written second.

   Exception: If the compound contains oxygen and a halogen, the name of the halogen is the first word in the name.

2. If both elements belong to the same group, the name of the element with the higher period number is the first word.

3. The name of the second element in the compound is added with an –ide ending, just like how an anion is named in ionic compounds.

4. Greek prefixes are used to indicate the number of atoms present in the chemical formula.

   Exception: If the compound contains one atom of the element that is written first in the name, the prefix “mono” is no longer used for that element.

   Table 3.3 The Greek Prefixes and Their Corresponding Numbers

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Number Indicated</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Mono</td>
</tr>
<tr>
<td>2</td>
<td>Di</td>
</tr>
<tr>
<td>3</td>
<td>Tri</td>
</tr>
<tr>
<td>4</td>
<td>Tetra</td>
</tr>
<tr>
<td>5</td>
<td>Penta</td>
</tr>
<tr>
<td>6</td>
<td>Hexa</td>
</tr>
<tr>
<td>7</td>
<td>Hepta</td>
</tr>
<tr>
<td>8</td>
<td>Octa</td>
</tr>
<tr>
<td>9</td>
<td>Nona</td>
</tr>
<tr>
<td>10</td>
<td>Deca</td>
</tr>
</tbody>
</table>

   If the prefix ends with a vowel and the first letter of the next word is “oxide”, the last letter of the prefix is usually omitted. An exception to this is dioxide and trioxide. Examples are monoxide instead of monooxide, tetroxide instead of tetraoxide and decoxide instead of docaoxide.

   Let’s name some covalent molecules by having N₂O₃ as an example. Nitrogen is under group 15 while oxygen is under group 16. Nitrogen should be written first in the name because it is of lower group number (Rule 1) and oxygen written second. There are two atoms of nitrogen and so the prefix di- should be attached to its name. There are three atoms of oxygen so we should use tri- (Rule 4). Oxygen, as the second element, should have the –ide ending (Rule 3). So the name of the compound is dinitrogen trioxide. Another example is Cl₂O. Chlorine is under group 17 while oxygen is under group 16. Since oxygen is under a lower group number, its name should be
written first in the name. However, as an exception to Rule 1, the name of the halogen chlorine should be written first. There are two chlorine atoms and there is only one oxygen atom. The name is dichlorine monoxide.

3.8 Drawing Lewis Structures

Lewis structures help us understand the bonding in molecules. The bonding is responsible for the many physical and chemical properties of molecules. The following are the steps to follow in drawing Lewis structures.

1. Sum up the number of valence electrons in all the atoms involved in a molecule.

Let's have ammonia, \( \text{NH}_3 \), as an example. There are three hydrogen atoms and each of them has 1 valence electron. There is one nitrogen atom and it has 5 valence electrons. There are a total of 8 valence electrons. It is most often necessary to consult the periodic table in determining the number of valence electrons.

2. Connect the atoms, using their symbols, by a single dash.

A single dash represents two electrons bonded together. The central atom is always the least electronegative atom. Between nitrogen and hydrogen, nitrogen is less electronegative. It is the central atom in ammonia. Each of the three hydrogen atoms is then connected by a single dash to the nitrogen atom. In doing so, 6 electrons are consumed by the three bonds in ammonia. Since there is a total of eight valence electrons, two electrons remain.

3. Complete the octet of the bonded atoms.

This is done by distributing the remaining electrons to the atoms in the molecule. The distribution of the remaining electrons starts with the central atom. If the central atom attains an octet, the remaining electrons are then distributed to the atoms surrounding the central atom. For ammonia, the remaining two valence electrons are placed on nitrogen. Remember that hydrogen can have only two electrons.
4. If there are not enough electrons to achieve an octet, try multiple bonds.

Let’s have carbon dioxide. The formula is CO₂. Since carbon is less electronegative than oxygen, it is the central atom. Each of the two oxygen atoms is then connected by a single bond to the central atom, carbon. The structure can be written as

\[
\text{O} \quad \text{\_} \quad \text{O} \\
\text{O} \quad \text{\_} \quad \text{O}
\]

Oxygen has 6 valence electrons and carbon has 4. In CO₂, there are a total of 16 valence electrons. Four electrons have been used for the two bonds already constructed. There remain 12 valence electrons left to distribute. Assigning 4 electrons to carbon will give it an octet of electrons. The resulting molecule now has used up 8 valence electrons.

\[
\text{O} \quad \text{\_} \quad \text{O} \\
\text{O} \quad \text{\_} \quad \text{O}
\]

Eight more electrons remain unused. These extra electrons will then be distributed to the oxygen atoms, even if it gives these atoms more or less than an octet of electrons.

\[
\text{O} \quad \text{\_} \quad \text{O} \\
\text{O} \quad \text{\_} \quad \text{O}
\]

In the molecule above, the two oxygen atoms have four nonbonding electrons and a single bond each. They have 6 electrons attached to each of them. Remember that the electrons involved in a bond are assumed to be the sole ownership of the atom to which the bond is attached.

Carbon has 8 electrons surrounding it while each of the oxygen atoms has only 6. Oxygen must achieve an octet of electrons. This is attained by forming multiple bonds. If the oxygen atom on the left side of the molecule is connected by a double bond to the carbon atom using two of its nonbonding electrons, the resulting structure would be

\[
\text{O} = \text{\_} = \text{O} \\
\text{O} = \text{\_} = \text{O}
\]

The oxygen atom on the left has 6 electrons surrounding it. It is not yet an octet. The carbon atom has 10 electrons around it. It is already more than an octet. The oxygen atom on the right also has 6 electrons surrounding it. If the electrons used to form the 2nd bond between oxygen on the left and the carbon atom comes from the carbon atom, the resulting structure would be

\[
\text{O} = \text{\_} = \text{O} \\
\text{O} = \text{\_} = \text{O}
\]
In this molecule, the oxygen atom on the left has 8 electrons surrounding it. Carbon atom also has eight, and the oxygen atom on the right has only 6. If we use the remaining nonbonding electron pair on carbon to come up with a double bond between the carbon atom and the oxygen on the right, the Lewis structure would be

\[
\overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}} = C = \overset{\cdot}{\overset{\cdot}{\overset{\cdot}{O}}}
\]

In the molecule above, all the atoms have 8 electrons surrounding each of them. Then this is the Lewis structure for carbon dioxide. The figures below show the Lewis structures of \( \text{F}_2 \) and \( \text{CONH}_3 \).

Figure 3.9 The Lewis Structures of a.) \( \text{F}_2 \) and b.) \( \text{CONH}_3 \)

If the molecule is an anion, it is enclosed in square brackets and the charge is written on the upper right of the closing bracket.

Figure 3.10 The Lewis Structure of the Hydrogen Sulfate Ion, \( \text{HSO}_4^- \)
3.9 Formal Charge

Many molecules have more than one Lewis structures. To know which one is the most reasonable, we determine the formal charges on the atoms in a molecule or ion. The formal charge is the charge on an atom in a molecule assuming that electrons in a chemical bond are shared equally between the bonded atoms regardless of electronegativity. The formal charge on an atom equals the number of valence electrons in the isolated atom minus the nonbonding electrons in that atom and half of the bonding electrons. In symbols, it can be written as

\[ FC = \text{valence e}^- - (\text{nonbonding e}^- + \frac{1}{2} \text{bonding e}^-) \]

Carbon dioxide has the following Lewis structure.

![Lewis structure of carbon dioxide]

Oxygen has 6 valence electrons. There are 2 lone pairs of electrons surrounding the oxygen atom on the left. Each lone pair is equivalent to 2 nonbonding electrons. Two bonds are attached to it. This means that 2 bonding electrons are assigned to it.

\[ FC = 6 \text{ valence e}^- - (4 \text{ nonbonding e}^- + \frac{1}{2} (4 \text{ bonding electrons})) \]
\[ FC = 6 - (4 + 2) = 0 \]

The oxygen atom on the right side has the same formal charge with that of the oxygen on the left. Carbon has 4 valence electrons. There is no nonbonding electron surrounding it. There are 4 bonds attached to it. This means that 4 bonding electrons are assigned to it.

\[ FC = 4 - (0 + 4) = 0 \]

A formal charge is usually written above the atom. For carbon dioxide,

![Lewis structure with formal charges]
The sum of the formal charges equals the charge on the molecule. Carbon dioxide is a neutral molecule. It has no charge. The sum of the formal charges is 0. But how important is the concept of formal charges on the structure of molecules? When several Lewis structures are possible, the most stable one will be that in which the atoms bear formal charges close to zero and that in which any negative charges reside on the more electronegative atoms. Take thiocyanate ion for example. There are three possible Lewis structures for thiocyanate. The number of valence electrons in nitrogen, carbon and sulfur are 5, 4, and 6, respectively. As they must, the formal charges in each of the structures sum up to -1, the charge of the ion. As is already known, nitrogen is more electronegative than carbon and sulfur. So it is expected that any negative formal charge must reside on nitrogen. Furthermore, we usually choose the Lewis structure containing formal charges of lowest magnitude. For these reasons, the middle structure is the most plausible or the most preferred.

\[ \text{A} \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad \quad

3.10 Resonance Structures

A single Lewis structure often cannot represent the true electronic structure of a molecule. Take ozone, \( \text{O}_3 \), for example. Studies on the molecule have shown that each \( \text{O}-\text{O} \) bonds are of equal length. Oxygen has 6 valence electrons. Drawing its Lewis structure would give one \( \text{O}-\text{O} \) single bond and one \( \text{O}-\text{O} \) double bond. The double bond is shorter than the single bond and this is contrary to the observed bond lengths in ozone. So this structure cannot by itself be correct. Moreover, the double bond can be placed on the other \( \text{O}-\text{O} \) bond such as in

![Lewis structure of ozone](image-url)
The placement of the atoms in the two alternative Lewis structures for ozone is just the same but the placement of electrons is different. Lewis structures of this sort are called resonance structures. A double-headed arrow is used to show that the structures shown are resonance structures. Resonance structures contribute to the true electronic structure of a molecule. As an analogy, a green color is a combination of both blue and yellow colors. A green color cannot be green by just a blue color alone. The same is true with a yellow color alone. It must be a combination of both colors.

Some more examples of resonance structures are the following.

3.11 Exceptions to the Octet Rule
The octet rule also fails in many covalent bonding. These are of three types.

1. Molecules with an odd number of electrons.
2. Molecules in which an atom has less than an octet.
3. Molecules in which an atom has more than an octet.

3.11.1 Odd number of Electrons
In many molecules, the number of electrons is even. This allows the complete pairing of electrons. In a few molecules, the number of molecules is odd. This results to one unpaired electron in the molecule. Examples are ClO$_2^-$, NO, and NO$_2^-$. Molecules with unpaired electrons are known as free radicals. Their high reactivity is due to the presence of this unpaired electron.
3.11.2 Less than an Octet

This occurs when there are fewer than eight electrons around an atom in an ion or molecule. This is most often encountered in the molecules of boron and beryllium.

Boron only has six electrons surrounding it. That is the reason why when an unpaired electron is available nearby, boron trifluoride easily forms a bond with it. An example would be the bond easily formed between boron trifluoride and ammonia, \( \text{NH}_3 \), to form \( \text{NH}_3\text{BF}_3 \).

3.11.3 More than an Octet

The largest class of exceptions consists of ions or molecules in which there are more than an octet of electrons in an atom. An example is \( \text{PCl}_5 \). In this molecule, there are ten electrons surrounding the central atom, phosphorus. This is possible for atoms with unfilled d orbitals. Elements starting on the third row of the periodic table are capable of having more than an octet since they have unfilled d orbitals. An electron from the s and p orbitals can jump and occupy the d orbital. The size of the central atom is also a determining factor whether it can accommodate more than an octet of electrons or not. The bigger the size of the central atom, the more likely atoms can attach to it due to the space available around the central atom. Moreover, the smaller the size of the surrounding atoms, the greater is their number that can form bonds with the central atom. Other molecules with atoms having more than an octet of electrons are \( \text{SF}_4 \) and \( \text{AsF}_6^- \).

3.12 Strengths of Covalent Bonds

The strength of a bond is determined by how easily it can be broken. Breaking a bond requires an input of energy, thus, it is an endothermic process. By convention, the energy values for endothermic processes are positive. Bond strengths are always positive quantities. The bond enthalpy is the enthalpy change, \( \Delta H \), for breaking a particular bond in a mole of a gaseous substance. The bond enthalpy for the bond between two chlorine atoms is 242 kJ (kilojoules). This means that for the bond to break, 242 kJ of energy is needed. Since the breaking of this bond yields two atoms, this is also known as atomization. For molecules, the enthalpy of a single bond is an average value. The bond enthalpy for methane, \( \text{CH}_4 \), is 1660 kJ. Since a mole of methane contains four C-H bonds, the bond enthalpy of a single C-H bond is the average of the total bond enthalpy. That is, 1660 kJ/4 bonds. Each C-H bond has an enthalpy of 415 kJ.
3.13 Bond Length

If a bond is strong, the atoms bound by the bond are strongly attracted to each other. This simply means that they are brought closer to each other. As a consequence, the bond becomes shorter. The reverse is true for weaker bonds. Therefore, the stronger the bond, the shorter it gets. And the weaker the bond, the longer it is. Bond lengths are usually expressed in Angstrom, Å.

Molecules with strong bonds have less tendency to undergo chemical changes than do those with weaker bonds. Moreover, the strength of a bond is inversely proportional to its length.

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Bond Strength</th>
<th>Bond Length</th>
</tr>
</thead>
<tbody>
<tr>
<td>C–C</td>
<td>348 kJ/mol</td>
<td>1.54 Å</td>
</tr>
<tr>
<td>C≡C</td>
<td>614 kJ/mol</td>
<td>1.34 Å</td>
</tr>
<tr>
<td>C≡C</td>
<td>839 kJ/mol</td>
<td>1.20 Å</td>
</tr>
</tbody>
</table>

4.0 Intermolecular Forces of Attraction

4.1 London Dispersion Force

This force of attraction between molecules is temporary. When a nonpolar molecule, such as H₂, gets nearby a polar molecule such as HCl, this can cause an induced dipole on H₂. Imagine having the orientation of HCl such that the Cl atom faces one of the hydrogen atoms of H₂. The more electronegative chlorine atom pulls the electrons towards itself, rendering one of the hydrogen atoms on H₂ to be partially positive. This causes an attraction between that hydrogen atom and the chlorine atom. As the HCl moves away, the attraction is then broken.

4.2 Dipole-dipole Interactions

This force of attraction exists between polar molecules. If two HCl molecules are contained in one container, the electronegative chlorine on one HCl molecule will attract the hydrogen atom of the other such as in the figure below.

![Figure 4.1 The Dipole-Dipole Interaction Between HCl Molecules](image-url)
4.3 Hydrogen Bonding

A hydrogen bond is the strongest intermolecular force of attraction. This exists between a hydrogen atom and a fluorine, oxygen, or nitrogen atom. The high electronegativities of F, O, and N create highly polar bonds with hydrogen. The high boiling point of water is due to this force of attraction.

Figure 4.2 The Hydrogen Bond Between Water Molecules
5.0 Quiz

I. Fill in the Blanks

1. A/an ______________ exists between a hydrogen atom and either F, O or N atoms. This causes the high boiling point of water.
2. A/an ___________ bond exists usually between a metal and nonmetal elements.
3. The ______________ of an atom is the energy needed to remove its outermost electrons.
4. The ______________ of an atom is its ability to attract electrons toward itself.
5. Breaking a bond between two atoms requires energy. This amount of energy is usually expressed in a unit of measure known as ____________________.
6. The molecule NH₄I is named as ________________.
7. CO₂ is called ______________ while HF is ______________.
8. Molecules that are composed of the same number and kinds of atoms but differ on how the atoms are arranged are called ______________ structures.
9. Among the types of bonds, the ______________ is the most polar.
10. The ______________ electrons are the ones involved in chemical bonding.

II. Write > or <.

1. Polarity: HCl ____ NaOH
2. Bond Length: F₂ ____ O₂
3. Electronegativity: N ____ H
4. Electron affinity: Li ____ H
5. Ionization energy: Na ____ I
6. Bond strength: HF-HF ____ F₂-F₂
7. Bond strength: O₂ ____ N₂
8. Formal charge in SCN⁻: S ____ N
9. Bond enthalpy: NaCl ____ KCl
10. Number of oxygen atoms: Ammonium sulfate ____ Magnesium nitrate

III. Name the following.

1. NaNO₂
2. (NH₄)₂SO₄
3. N₂O₃
4. SrCl₂
5. NCl₃

IV. Write the formulas of the following.

1. Dinitrogen trisulfide
2. Lead(IV) iodide
3. Ferric acetate
4. Carbon tetrachloride
5. Silver sulfate