**5. Chemical Bonding: The Covalent Bond Model**  
**5.1 The Covalent Bond Model**

Almost all chemical substances are found as aggregates of atoms in the form of molecules and ions produced through the reactions of various atoms of elements except the noble-gas elements which are stable mono-atomic gases.

**Chemical bond** is a term that describes the attractive force that is holding the atoms of the same or different kind of atoms in forming a molecule or ionic solid that has more stability than the individual atoms. Depending on the kinds of atoms participating in the interaction there seem to be three types of bonding:

**Gaining or Losing Electrons**:

**Ionic bonding**: Formed between many ions formed by **metal** and **nonmetallic** elements.

**Sharing Electrons:**

**Covalent bonding**: sharing of electrons between **two atoms** of **non-metals**.

Metallic Bonding: sharing of electrons between **many atoms** of **metals**.

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Ionic Compounds** | **Covalent Compounds** | **Metallic Compounds** |
|  | **Metal** and **non-meal** element combinations. | **Non-metal** and **non-meal** elements combinations. | **Metal** of one type or, combinations of two or metal elements combinations. |
|  | **High melting brittle** crystalline solids. | **Gases, liquids, or waxy, low melting soft** solids. | **Conducting, high melting, malleable, ductile** crystalline solids. |
|  | Do not conduct as a solid but conducts **electricity** when **molten**. | **Do not** conduct electricity at **any state**. | Conduct electricity at **solid and molten** states. |
|  | Dissolved in water produce conducting solutions (**electrolytes**) and few are soluble in non-polar solvents. | Most are soluble in non-polar solvents and few in water. These solutions are non-conducting (**non-electrolytes**). | Insoluble in any type of solvents. |

The differences in these three bonding types are mainly due to the number of valence election of the interacting atoms compared to noble gas elements. Noble gases show no inherent tendency to form any type of bonding apparently due to their closed valence shell electron configurations. Non-metals need only few electrons to achieve a closed shell. We focus mainly on covalent bonding in this chapter.

**Covalent bond**: bond in which one or more pairs of valence electrons are shared by two atoms in achieving a noble gas in contrast to complete transfer of electrons in forming ionic bonds.  
 Covalent bonding can be visualized with the aid of the **octet rule** and the **Lewis structure** we learned previously in chapter 4.

A molecular or covalent compound is made up of two or more **atoms** which are held together by a **covalent bond**. Even though covalent bonds are mainly formed between two atoms of non-metals, there could be many metal and nonmetal combination that are held by covalent bonds.

**5.2 Lewis Structures for Molecular Compounds**

Two atoms in molecule or ion in which each atom **shares one** valence electron with the other atom to from a single bond that keeps the two atoms together in forming a covalent bond.



**Lewis Theory of Covalent Bonding:**

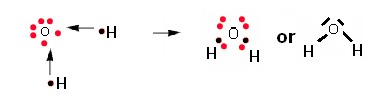
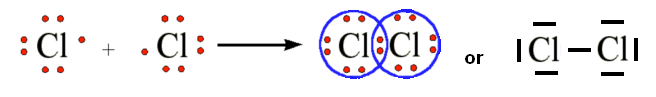
The idea that the stability of noble gas electron configurations and the realization of the connection of reactivity of elements to achieve **octet of valance electrons** through formation of electron-pair bonds with other atoms of main group elements except duet for hydrogen.

**E.g.** In fluorine molecule, F2 each atom F atom has 7 valence electrons and shares one valence electrons to form a covalent bonding pair. Bonding pair is represented by a line. or .



**Lewis Electron-Dot (Line) Formulas**: A simple way of writing out a formula that shows the disposition of the shared valance electron pairs between the different atoms in a molecule.

In chlorine molecule, Cl2 each atom Cl atom has 7 valence electrons and shares one valence electrons to form a covalent bonding pair. Bonding pair is represented by a line.



In water, H2O molecule, one oxygen atom with 6 valence electrons shares one valence electrons with two hydrogen atom each with one electron in forming two covalent bonding pair. Bonding pair is normally represented by a line.

**Bond pairs**: An electron pair **()**shared by two atoms in a bond.



**Lone pair**: An electron pair () found solely on a single atom.



**5.3 Single, Double, and Triple Covalent Bonds**

**Single covalent bond:** An electron pair **()** shared by two atoms in a bond. Always there has to be a single covalent before forming multiple covalent bonds.



**Multiple Covalent Bonds in Lewis Structures:**  For every pair of electrons shared between two atoms, a single covalent bond is formed. Some atoms can share multiple pairs of electrons, forming multiple covalent bonds.

**Double covalent bond:** Where two electron pairs are shared between two atoms. **For example**, oxygen, O2 (which has six valence electrons) needs two electrons to complete its valence shell. In O2, two pairs of electrons are shared, forming two covalent bonds called a **double covalent bond**.

|  |  |
| --- | --- |
| Six valence electrons from each O  Total of 12electrons or 6 electrons pairs. |  |

**Triple covalent bond:** Found in unsaturated hydrocarbons with formula

**For example**, nitrogen, N2 (which has 5 valence electrons) needs three electrons to complete its valence shell. In N2, three pairs of electrons are shared, forming three covalent bonds called a **Triple double covalent bond**.

|  |  |
| --- | --- |
| Five valence electrons from each N  Total of 10 electrons or 5 electrons pairs. |  |

**Bond Length and Bond Energy**

**Bond Energy order**: single=1 < double=2 < triple=3  
**Bond length**: single (1pair) **>** double (2 pairs) **>** triple (3 pairs)

**Bond lengths from periodic trends in atomic radii**

Bond length is proportional to the sun of atomic radii forming the bond.

**Atomic radii trend**: Li> Be> B> C> N> O> F **decrease across a period**

**Bond Length Order:** Li-H> Be-H> B-H> C-H> N-H> O-H> F-H

**5.4 Valence Electrons and Number of Covalent Bonds Formed**

**Predicting number of covalent bonds formed by representative elements:**

Lewis dot symbols are useful in showing the ways in which non-noble gas electron configurations could be achieved by sharing electrons in forming covalent bonds. Number of covalent bonds formed is equal to the number of single electrons in the Lewis symbol of the representative element.

**Group I A:** alkali metals (Li, Na, K, Rb, Cs)

Common Lewis symbol of Group I A could only form a single covalent bond as found in NaH.



**Group II A:** alkali earth metals (Be, Mg, Ca, Ba)

Common Lewis symbol of Group II A : alkali earth metals (Be, Mg, Ca, Ba) could form two covalent bond as found in BeCl2.

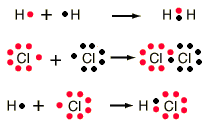


**Group III A**: (B, Al, Ga, In)

Common Lewis symbol of Group III A : (B, Al, Ga, In) could form two covalent bond as found in NH3.



Hydrogen, H2 gas forms the simplest covalent bond in the **homo-nuclear diatomic molecule**, H2. The halogens such as chlorine also exist as **homo-nuclear diatomic** gases by forming covalent bonds. The nitrogen and oxygen which makes up the bulk of the atmosphere also exhibits covalent bonding in forming diatomic molecules. Hydrogen and chlorine forms the simplest covalent bon d in the **hetero-nuclear diatomic molecule**, HCl.

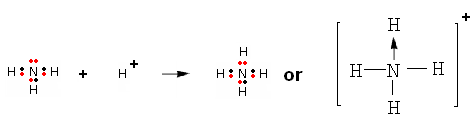


|  |  |
| --- | --- |
|  | **Predicting charge of Ions of Representative Elements** |
| **Periodic Table Group Number** |  |
| **Common Lewis symbols** |
| **Lewis symbols**  **2nd period elements** |
| **Number of covalent bonds formed** |

**5.5 Coordinate Covalent Bonds**

**Coordinate Covalent bonds:** Two atoms in molecule in which one atom **shares two** valence electrons with the other atom to from a bond that keeps the two atoms together.

**E.g.** In ammonia ion fluorine molecule, NH4+, N atom has a lone pair of electrons and shares with the H+ ion to from a single bond that keeps the two atoms together. Coordinate covalent bonding pair is represented by an arrow pointing to atom receiving the electron. Chemistry dealing with coordinative covalent bonds is called **Coordination Chemistry**.



**Covalent Bonds in Lewis Structure of Molecules and Polyatomic Ions:** Lewis electron dot structures are representations of the distribution of electrons in molecules and polyatomic ions.

**5.6 Systematic Procedures for Drawing Lewis Structures**

A Lewis structure can be drawn for any molecule or ion could be obtained by following six steps:

**Steps for Writing Electron-Dot Lewis structures for Molecules and Polyatomic Ions**

1. **Calculate the number of valence electron pairs** (including charges, if any). The total number of valence electrons for the molecule or ion is calculated by adding up the valence electrons for each of the atoms in the species (molecule or ion). The number of valence electrons contributed by an atom equals the group number of the element. If the species has a **charge**, you need to **add** (anions) or **subtract** (cations) electrons from the neutral atoms to give that charge. Add one electron for each negative charge and subtract one electron for each positive charge.
2. **Pick the central atom**: Lowest EN (**electronegativity**) atom, **largest atom**, and/or atom forming most bonds is usually the central atom.
3. **Connect central atom to all terminal atoms.** Connect atoms by single bonds, using either a pair of **dots (** or a dash ( ). You may decide later that one or more of these bonds will add up to a double or triple bond.



1. **Fill octet to terminal atoms.** Distribute electron **dots ()** or a dash ( ) to the atoms surrounding the central atom to satisfy the octet rule in them. Calculate the number of remaining valence electrons pairs so far used.



1. **Fill octet to central atom**: Distribute the remaining valence electrons to the central atom. If there arc fewer than eight electrons around the central atom, this suggests that you need to make double triple bonds taking **lone-electron pairs** ()from terminal atoms. When you have a choice of atoms from which to obtain such a lone-pair, note that C. N, 0, and S atoms tend to form multiple bonds, so you take the lone pair from such an atom.



1. **Check that total number of valance electron pairs** in the structure match with the **dots (** or a **dash** ( ) or lines.



**Draw Lewis Structure of NH3**

|  |  |  |  |
| --- | --- | --- | --- |
| 1) **Count** **Valence electrons**:  5 + 3 x 1 = 8 = 8/2 **4 electron pairs** | **4 electron pairs** | **4) Fill octet to terminal atoms.** Exception duet to H atoms. | or |
| 2) **N is the central atom**: Forms most covalent bonds |  | **5) Fill octet to terminal atoms.** | or  **Lewis Structure** |
| **3) Connect central atom to terminal** Add terminal H atoms |  | **6) Check that total number in 5.** | 3 bond pairs = 3 pairs  1 lone pairs = 1 pairs  **4 electron pairs** |

**Problem**: Draw the Lewis structures for the following molecules: H2O

|  |  |
| --- | --- |
| **a) H2O** sum of valence electrons: 2 + 6 =8 ; 4 pairs  Pick the central atom: usually the biggest atom and the atom forming most bonds.  **O is the central atom**  Connect central atom to other terminal atoms with a line representing a covalent bond (two electrons). Fill octet to central atom:  A dot represents one electron and  line represents two electrons.  Check to see whether duet on H | Count electrons and check valence electrons on C.  2 bond pairs = 2 x 2 = 4  2 lone pairs = 2 x 2 = 4  Total 8 = 4 electron pairs (an Octet)  **Bond pairs**: an electron pair shared by two atom in a bond. E.g. two pairs between O--H in water.  **Lone pair** : an electron pair found solely on a single atom.  E.g. two pairs found on the O atom at the top and the bottom. |
| H2S has the same Lewis Structure as H2O  since S is in the same group as O.  As a rule if an elements from the period 2 is replaced by another element from the period 3 and same group in the periodic table  they will have same Lewis structure. |  |
| **Draw Lewis Structure of NH3**  i) Valence electrons: 5 + 3 x 1 = 8 = **4 electron pairs**  ii) N is the central atom:  iii) Add terminal Hs  vi) Give octet to N  v)Check duet on Hs  iv) Count electrons:  3 bond pairs = 3 pairs  1 lone pairs = 1 pairs  **4 electron pairs** | Lewis Structures of NH3, PH3 has the same Lewis Structure  3 bond pairs and 1 lone pair are found on central N. |
| Draw Lewis structure of Nitrogen (N2) | **:**N**:::**N**:** triple bond between N and N |
| Draw Lewis structure of CO2  i) Valence electrons: 4 + 2 x 6 = 16 ( 8 pairs)  ii) Central atom C; O -- C -- O  iii) Give octet to carbon  --  O -- C -- O  -- | Try to fill octet to O  iv) Count electrons:  4 bond pairs = 4 pairs  4 lone pairs = 4 pairs  8 electron pairs |
| Draw Lewis structure of **CH4**  i) Sum of valence electrons:  4 ( from one C) + 4 (from four H) = 8 electrons = **4 electron pairs**  ii) Central atom is C  iii) Octet on C atom and duet on H are already complete.  iv) Count valence electrons on the Lewis structure: | bond pairs = 2 x 4 = 8 = **4 electron pairs**  lone pairs = 0 |
| Draw Lewis structure of Ethane( CH3CH3 )  Draw two carbons and attach six hydrogen.  Count valence electrons:  2C = 8  6H = 6  14 (7 electron pairs) | Octet on each carbon and duet on each hydrogen |
| HNO3  H = 1 valence electron  N = 5 valence electrons  O x 3 = 6 x 3 = 18 valence electrons  Total of 24 valence electrons |  |

**5.7 Bonding in Compounds with Polyatomic Ions Present**

term that  
**5.8 Molecular Geometry**

**Lewis Structures and Molecular Geometry; VSEPR Theory**

The valence shell electron pair repulsion (VSEPR) theory helps to explain (or predict) molecular geometry, including linear, trigonal planar, and tetrahedral arrangements of attached atoms. Exceptions to the octet rule exist. **Valence-Shell Electron-Pair**

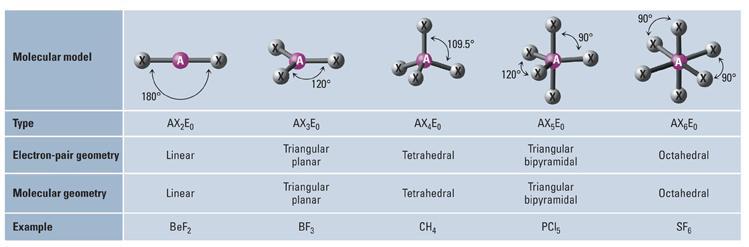
**Repulsion Theory (VSEPR Theory)**

In the theory, valence shell electron pairs are assumed to repel each other, assuming orientations to minimize repulsions and establish certain groups of molecular shapes.

**Basic Geometries created by molecules with 2,3,4,5 and 6 electron pairs**.



A = central atom; X= terminal atoms E= lone pair of electrons



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**Molecular Structures with Lone Pairs**

The molecular configuration is not exactly the same as the electron configuration. The reason for this is that lone pairs of electrons do not show up in the molecular configuration, although their effects are still seen (such as pushing bonding pair electrons closer together). This gives rise to 6 more configurations (in addition to the 5 basic electron configurations).

|  |  |  |
| --- | --- | --- |
| Angular or Bent: | AX2E2 | H2O, OF2 |
| Triangular Pyramidal or Trigonal Bipyamid | AX3E1 | NH3, NCl3 |
| Seasaw | AX4E1 | SF4 |
| T-shape | AX4E2 | CF3 |
| Square pyramidal | AX5E1 | BrF5 |
| Square Planar | AX4E2 | XeF4 |

First case of angular structures has three electron pairs surrounding the central atom. The possible *molecular* configurations are either trigonal planar (if all of the electron pairs are bonding pairs) or angular Angular configuration could be based on trigonal planar or triangular planar electron geometry or bent (If one is a lone pair).

The second case of angular structure has four electron pairs surrounding the central atom. The water molecule (H:O:H) has two lone pairs and two bond pairs (4 total pairs on O). The shape is angular or bent with a bond angle which is less than that in NH3. This is due to the greater repulsions with two lone pairs.

that  
Chemistry at a Glance: The Geometry of Molecules  
**5.9 Electronegativity**

term that  
**5.10 Bond Polarity**

**5.11 Molecular Polarity**

**Lewis Structures and Polarity**

Concepts of polarity are better understood when viewed from the perspective of a Lewis structure. A polar covalent molecule has at least one polar covalent bond. An understanding of the concept of electronegativity helps to assess the polarity of a bond. A molecule containing only non-polar bonds must be non-polar. a molecule containing polar bonds may be polar or non-polar, depending on the relative position of the bonds.

**Molecular Polarity**

Molecules composed of covalently bonded atoms may also be polar or non-polar. For the molecule to be polar, it must, of course, have polar bonds. But the key factor for determining the polarity of a molecule is its shape. If the polar bonds (dipoles) are symmetrical around the central atom, they offset each other and the resulting molecule is non-polar. However, if the dipoles are not symmetrical around the central atom, the electrons will be pulled to one end of the molecule. The resulting molecule is polar. Ball and stick models are often used to demonstrate molecular shape. In this exercise you will build several covalent molecules and predict each molecule's polarity on the basis of its molecular shape.

The geometry of the molecule affects its dipole moment or polarity - the asymmetric distribution of positive and negative charges.

Polar Molecules will orient in a magnetic field

When two elements with different electronegativities are bonded together, one obtains a polar bond.

Example: HF

Depending on the molecular shape, polar bonds can give rise to a polar molecule, one with an overall uneven distribution of electron charge. We can see this charge imbalance when the molecules are placed in an electric field. Polarity is a key feature of a molecule because it can influence physical, chemical, and even biological properties.  
  
 For a diatomic molecule, a polar bond *must* lead to a polar molecule. Consider hydrogen fluoride, shown here as the Lewis structure changes to a ball-and-stick model enclosed within the space-filling shape. Note the polar arrows and the colors. If red indicates high electron density and blue indicates low, you can see that the F end of the molecule is much more negative than the H end, and thus HF is highly polar. Between two electric plates with the field off, the molecules lie every which way. With the field on, however, they become oriented with their negative ends facing the positive plate and their positive ends facing the negative plate.

If a molecule has more than two atoms, its shape can affect the polarity in a crucial way. For example, in carbon dioxide, since oxygen is more electronegative than carbon, each bond is highly polar. But the linear molecular shape makes the bond polarities cancel each other, so the CO2 molecule is nonpolar. Notice that the orientation of the molecules is random, whether the field is off or on.

The situation is very different for water. As in CO2, the highly electronegative oxygen pulls electron density toward itself, so each bond is polar. But the V-shape of the molecule allows the bond polarities to reinforce each other, so water is highly polar, as the large polar arrow shows. With the field off, the molecules are oriented randomly, but with it on, their poles become oriented toward the oppositely charged plates.

The case of boron trifluoride is similar to that of CO2. Because the three highly polar bonds point to the corners of an equilateral triangle, the bond polarities cancel each other, and BF3 is nonpolar. The molecules are oriented randomly with the field off or on.

Like BF3, ammonia has four atoms and three polar bonds, but the trigonal pyramidal shape means that the bond polarities reinforce each other. Thus, ammonia is highly polar for the same reason that water is. Sometimes two molecules have similar overall shapes, but their slightly different compositions lead to very different polarities. Carbon tetrachloride has four polar carbon-chlorine bonds that point to the corners of a tetrahedron, so they cancel each other and give a nonpolar molecule. Substituting a hydrogen for one of the chlorines gives chloroform, another tetrahedral molecule, but now the polar bonds reinforce each other. As a result, chloroform is highly polar.

**Using VSEPR molecular structures predict the polarity (net dipole moment or zero dipole moment) of the following molecules: H2O, NH3, CO2, SO3, SF6, PCl5, XeF4**

|  |  |  |  |
| --- | --- | --- | --- |
| **Molecule** | **Shape** | **Electron**  **pair distribution** | **Polarity** |
| **H2O** | bent | asymmetric | polar |
| **NH3** | Trigonal pyramid | asymmetric | polar |
| **CO2** | liner | symmetric | Non-polar |
| **SO3** | Trigonal planer | symmetric | Non-polar |
| **SF6** | octahedral | symmetric | Non-polar |
| **PCl5** | Trigonal bipyramid | symmetric | Non-polar |
| **XeF4** | Square planar | symmetric | Non-polar |

Chemistry at a Glance: Covalent Bonds and Molecular Compounds  
**5.12 Naming Binary Molecular Compounds**

**Covalent Compounds**

Most covalent compounds are formed by the reactions of non-metals with another non-metal. Covalent compounds exist as molecules and are named using prefixes that denote the number of each element present in the compound.

**Prefixes used:**

|  |  |  |  |
| --- | --- | --- | --- |
| 1 | mono | 6 | hexa |
| 2 | di | 7 | hepta |
| 3 | tri | 8 | octa |
| 4 | tetra | 9 | nona |
| 5 | penta | 10 | deca |

Many familiar covalent compounds have common names. It is useful to correlate both systematic and common names with the corresponding molecular formula.

**Problem**: Give the names of following formulas from the names of following covalent compunds names:

1. H2S
2. CS2
3. PCl5
4. P2O5

**Answer**:

a. hydrogen sulfide

b. carbon disulfide

c. phosphorus pentachloride

d. diphosphorus pentoxide

**Naming Binary Molecular Compounds**

Binary molecular compounds are composed of only two elements. Examples are H2O, NO, SF6 etc. .Sometimes these compounds have generic or common names (e.g., H2O is "water") and they also have systematic names (e.g., H2O, dihydrogenmonoxide). The common name must be memorized. The systematic nameis more complicated but it has the advantage that the formula of the compound can be deduced from the name.

|  |  |  |
| --- | --- | --- |
| **Compound** | **Systematic name** | **Common name**  **(if it has one)** |
| NF3 | nitrogen trifluoride |  |
| NO | nitrogen monoxide  **note: for first element we don't use mono- prefix** | nitr**ic** oxide  higher oxidation # |
| NO2 | nitrogen dioxide |  |
| N2O | dinitrogen monoxide | laughing gas  nitr**ous** oxide  lower oxidaton # |
| N2O4 | dinitrogen tetraoxide |  |
| PCl5 | phosphorous pentachloride |  |
| SF6 | sulfur hexafluoride |  |
| S2F10 | disulfur decafluoride |  |
| H2O | dihydrogen monoxide | water |
| H2S | dihydrogen monosulfide | hydrogen sulfide |
| NH3 | nitrogen trihydride | ammonia |
| N2H4 | dinitrogen tetrahydride | hydrazine |
| PH3 | phosphorous trihydride | phosphine |

**Problem**: Give the formula of following covalent compounds:

1. Nitrogen trifluoride
2. Carbon monoxide

**Answer**:

a. NF3

b. CO

**Names of Acids and Bases**

Binary acids: made up of only two elements - hydrogen and one other element.

Naming binary acids:

Begin with the prefix hydro.

Determine the "stem" - part of the name of the element that combines with hydrogen.

Add the suffix ic.

Examples:

HF -hydro fluor ic - hydrofluor ic acid

HCl - hydro chlor ic - hydrochloric acid

HBr - hydro brom ic - hydrobromic acid

HI -hydro iod ic - hydroiodic acid

**Ternary acids**: made up of three elements - hydrogen, oxygen, and another element.

**Naming ternary acids:**

Acids made up of three elements including hydrogen

Determine the "stem" - part of the name of the third element.

The most common acid is given the suffix ic.

Add the prefix per for the acid with one more oxygen.

The suffix ous is given to the acid with one less oxygen.

Add the prefix hypo for the acid with two less oxygen atoms.

**Examples**:

HClO4 - per chlor ic - perchloric acid - one more oxygen atom.

HClO3 - chlor ic - chloric acid - the most common form of the acid.

HClO2 - chlor ous - chlorous acid - one less oxygen atom.

HClO - hypo chlor ous - hypochlorous acid - two less oxygen atoms.

HNO3 -nitric acid

HNO2 -nitrous acid

H2SO4 - sulfuric acid

H2SO3 -sulfurous acid

H3PO4 -phosphoric acid

H3PO3 -phosphorous acid

H3BO3 -boric acid

term that Covalent Compounds produced by nonmetal + nonmetal sharing of electrons.   
Exercise 08.   Use electron-dot symbols to show the sharing of electrons between nonmetal atoms to form covalent compounds.  Then name the compound.

|  |  |  |  |
| --- | --- | --- | --- |
| Phosphorus + hydrogen | silicon + hydrogen | carbon + flourine | nitrogen + chlorine |
| [**Answer**](http://chem.latech.edu/~snowchem/chem120/C5HW05BB.jpg) | [**Answer**](http://chem.latech.edu/~snowchem/chem120/C5HW06BB.jpg) | [**Answer**](http://chem.latech.edu/~snowchem/chem120/C5HW07BB.jpg) | [**Answer**](http://chem.latech.edu/~snowchem/chem120/C5HW08BB.jpg) |

Covalent compounds are named in different ways than are ionic compounds (although there is some overlap).

Many covalent compounds have common names such as "methane", "ammonia" and "water".

Simple covalent compounds are generally named by using prefixes to indicate how many atoms of each element are shown in the formula.   
Also, the ending of the last (most negative) element is changed to -ide.

|  |  |  |
| --- | --- | --- |
| The prefixes used are mono-, di-, tri-, tetra-, penta-, hexa-, and so forth. The mono- prefix is usually not used for the first element in the formula. The "o" and "a" endings of these prefixes are dropped when they are attached to "oxide." | 1 | mono- |
| 2 | di- |
| 3 | tri- |
| 4 | tetra- |
| 5 | penta- |
| 6 | hexa- |
| 7 | hepta- |
| 8 | octa- |
| 9 | nona- |
| 10 | deca- |

Examples:   
 P2O5 - this is named diphosphorus pentoxide, because there are two phosphorus atoms and five oxygen   
 atoms.   
 CO - this is carbon monoxide (you need the "mono-" because there's only one oxygen atom).   
 CF4 - this is carbon tetrafluoride, because there's one carbon atom and four fluorine atoms.

Exercise 09.  Name the following covalent compounds.   
CS2     N2S4     PF5     S2F10     CBr4    Cl2O7     P4S10     I2O5   
[**Answers**](http://chem.latech.edu/~snowchem/chem120/C5HWNAM1.jpg)  
  
Chemical Connections: Nitric Oxide: A Molecule Whose Bonding Does Not Follow "The Rules"; Molecular Geometry and Odor

**Drawing Lewis Structures of Molecules and Polyatomic Ions**

The electronic structure of atoms, ions, and molecules, which is closely allied to the properties of these substances, can conveniently be represented using Lewis structures, or electron‑dot diagrams based on the octet rule.

**Covalent bond** - a bond between two atoms, in which the two atoms share a pair of electrons. Covalent bonds can join atoms together to form molecules.

**Lewis-electron-dot or Lewis structure of a molecule or ion**

Lewis electron dot structures are representations of the distribution of electrons in molecules and polyatomic ions. They are useful in determining the three-dimensional shape of a molecule or ion. A Lewis structure can be drawn for a molecule or ion by following three steps:

1. **Calculate the number of valence electrons (including charges, if any)**

2. **Write skeleton structure**

* Lowest EN (electronegative) atom, largest atom, and/or atom forming most bonds is usually central atom
* Hydrogen cannot be central
* Connect all atoms with a single bond. Single bond is represented by line or two dots.
* In oxoacids the oxygen bonds to central, the H to oxygen. Like in sulfuric acid, H2SO4.
* Compounds are usually compact and symmetrical structures

3**. Count up used electrons**

Distribute remaining electrons to terminal atoms filling their eight electrons(octet rule)

4. **Central atom octet is filled last**

Remaining electrons become lone pairs on central.

If central atoms do not have an octet, move terminal a pair to form multiple bonds on at a time forming double and triple bonds.

**Terminology used in describing Lewis structures of molecules**

**Bond pairs**: An electron pair shared by two atoms in a bond.

**Lone pair**: An electron pair found solely on a single atom.

**Single covalent bond** - Bond between two atoms when they shared 1 pair

**Double covalent bond** - Bond between two atoms when they shared 2 pairs.

**Triple covalent bond** - Bond between two atoms when they shared 3 pairs.