Chapter 5: Electron Configuration, Lewis Dot Structure, and Molecular Shape

Electron configuration.

The outermost electrons surrounding an atom (the ***valence electrons***) are responsible for the number and type of bonds that a given atom can form with other atoms, and are responsible for the chemistry of the atom. The shape of the modern periodic table reflects the arrangement of electrons by grouping elements together into *s, p, d*, and *f* blocks. If we compare the arrangement of electrons, the logical grouping of elements becomes apparent. The ***electron configuration*** (the arrangement of electrons in an element) is a direct result of the work of Bohr, Heisenberg, Schroedinger, and other physicists in the early 20th century.

Elements are arranged in horizontal rows, called ***periods***. Each period is a ***principal energy level***, numbered 1, 2, 3, 4, etc. The first principal energy level starts with hydrogen, the second level starts with lithium, the third with sodium, etc. Each principal energy level holds a maximum number of electrons equal to $2n^2$, where n is the principal energy level number. The second energy level starting with lithium holds $2(2^2) = 8$ electrons maximum; the third level starting with sodium holds $2(3^2) = 18$; the fourth can hold $2(4^2) = 32$, and so on.

Principal energy levels are divided into ***sublevels*** following a distinctive pattern, shown in Table 5.1 below.

Principal level	Sublevels
1	*s*
2	*s, p*
3	*s, p, d*
4	*s, p, d, f*
5	*s, p, d, f, g*
6	*s, p, d, f, g, h*
7	*s, p, d, f, g, h, i*

Table 5.1. Principal energy levels and their sublevels.

The principal energy levels and sublevels build on top of each other. The sodium atom in the third principal energy level has the first principal level composed of an *s*-sublevel, and the second principal level composed of *s*- and *p*-sublevels, underneath the third principal level. Probably the simplest mental image would be an onion. Each principal level corresponds to a layer of the onion. As you move farther from the center of the onion, the layers get larger. Larger layers can be sub-divided into more pieces (sublevels) than can smaller layers.

Each sublevel is in turn divided into ***orbitals***, specific locations for the electrons. The number of orbitals for each sublevel also follows a distinctive pattern, shown in Table 5.2 below.

Sublevel	Total orbitals
s	1
p	3
d	5
f	7
g	9
h	11
i	13

Table 5.2. Total orbitals for each type of sublevel.

All *s*-sublevels have 1 orbital, regardless of whether they are 1*s*, or 2*s*, or 3*s*, etc. All *p*-sublevels contain 3 orbitals, and *d*-sublevels contain 5 orbitals, and so on. Every orbital, regardless of the sublevel, holds a maximum of 2 electrons.

In order to compare the arrangement of electrons around atoms, we need some way of predicting the filling order of the principal levels and sublevels. Figure 5.1 below guides us in arranging the electrons.

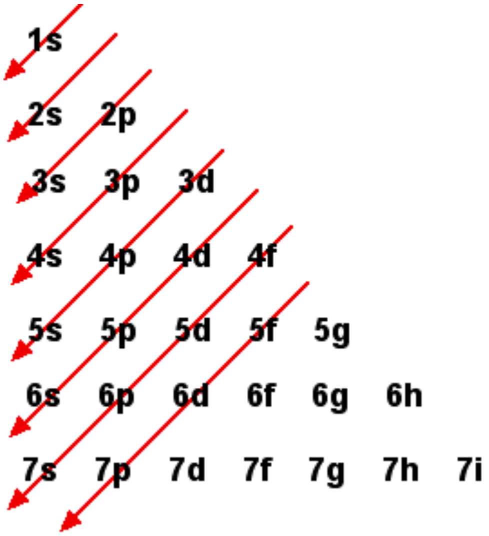


Figure 5.1. Filling order of principal levels and sublevels.

Using Figure 5.1 is as simple as following the arrows. Start with the top arrow, and follow its direction from tail to head; you see that the first arrow passes through $1s$, and the $1s$ sublevel is the first to fill with electrons. When the $1s$ sublevel is full, continue to the second arrow which passes through the $2s$ sublevel. The third arrow passes through the $2p$ and $3s$ sublevels (in this order), while the fourth arrow passes through the $3p$ and $4s$ sublevels, in order. The fifth arrow passes through the $3d$, $4p$, and $5s$ sublevels in order.

Using the order provided from Figure 5.1, and remembering the total number of orbitals in each sublevel, we can write electronic configurations for the elements. The ***electronic configuration*** is the electronic structure of the atoms; the specific levels, sublevels, and number of electrons occupying orbitals for a given atom.

Let's look at group 1A elements. Hydrogen has 1 electron, and the electronic configuration is $1s^1$. This tells us that the first principal energy level is being filled, the *s*-sublevel is the specific location within the principal energy level, and that there is 1 electron in the *s*-orbital (the superscript "1" is the number of electrons). Electronic configurations for all group 1A elements are given below.

H (1 electron): $1s^1$

Li (3 electrons): $1s^2, 2s^1$

Na (11 electrons): $1s^2, 2s^2, 2p^6, 3s^1$

K (19 electrons): $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

Rb (37 electrons): $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^1$

Cs (55 electrons): $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^1$

Fr (87 electrons): $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^1$

Notice that in all cases the outermost energy level contains a single sublevel: an *s* type sublevel, and this sublevel contains 1 electron. The outside layers of these elements are very similar. The group 2A elements contain one additional electron, and in all cases this additional electron is in the *s* type sublevel. Beryllium will be $1s^2, 2s^2$. Magnesium will resemble sodium with an additional electron ($1s^2, 2s^2, 2p^6, 3s^2$), calcium will resemble potassium with an additional electron, and so on.

This pattern follows throughout the periodic table. The group 3A elements all have an outer layer following the pattern $\mathrm{n}s^2, \mathrm{n}p^1$ (where n is the principal energy level number). Boron's outside layer is $2s^2, 2p^1$, aluminum's is $3s^2, 3p^1$, and so on. Group 4A elements follow the pattern $\mathrm{n}s^2, \mathrm{n}p^2$; Group 5A's pattern is $\mathrm{n}s^2, \mathrm{n}p^3$, and so on.

Group 8A, the Noble gases, have a very interesting electronic structure.

He (2 electron): $1s^2$
Ne (10 electrons): $1s^2$, **$2s^2$, $2p^6$**
Ar (18 electrons): $1s^2$, $2s^2$, $2p^6$, **$3s^2$, $3p^6$**
Kr (36 electrons): $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, **$4s^2$**, $3d^{10}$, **$4p^6$**
Xe (54 electrons): $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, **$5s^2$**, $4d^{10}$, **$5p^6$**
Rn (86 electrons): $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^6$, **$6s^2$**, $4f^{14}$, $5d^{10}$, **$6p^6$**

The outside layer of electrons is full! With the exception of helium, which is too small to hold 8 electrons, all other elements in this group follow the pattern ns^2, np^6. This arrangement is called a ***Noble gas configuration***, 8 electrons completely filling the n*s* and n*p* sublevels, and this arrangement confers extraordinary stability on these compounds. The Noble gases don't participate in very many chemical reactions, and then only under extreme conditions.

When we get into the *d* block (transition elements) and the *f* block (lanthanides/actinides) the same general pattern holds. Group 3B elements have similar electron configurations, as do group 4B, 5B, and so on.

Lewis Dot Structures.

Metallic elements, to the left of the staircase dividing line, tend to loose one or more electrons and form ions. Hydrogen is the single general exception to this trend; when combined with non-metals, hydrogen tends to form covalent compounds.

All of the non-metals have filled *s*-sublevels, and partially filled *p*-sublevels. There are a fixed number of electrons on the outside layer of these elements, and this number is equal to the group number. Boron is the only group 3A non-metal, and has 3 electrons in its outside layer. Carbon and silicon are group 4A non-metals and have 4 electrons in their outside layers. Groups 5, 6, 7, and 8A follow similarly.

In 1916, American physical chemist Gilbert Newton Lewis (1875 – 1946) discovered the covalent bond and developed a method of representing bonding between non-metals using simple diagrams called Lewis dot structures. Lewis dot structures are the elements chemical symbol, with the valence electrons arranged uniformly around the four sides of the symbol. Lewis dot structures for the first few non-metals are shown in Figure 5.2.

Figure 5.2. Lewis dot structures for the first few non-metals.

Silicon is similar to carbon, with “Si” replacing “C”, phosphorous and arsenic are similar to nitrogen; sulfur, selenium, tellurium are similar to oxygen; chlorine, bromine, and iodine are similar to fluorine; and the Noble gases are similar to neon.

In Lewis dot structures, some dots are single, representing single electrons, while others are double representing pairs of electrons. We can predict the number of bonds each element forms by counting the number of single electrons in the Lewis dot structure. Boron can form 3 bonds, carbon can form 4 bonds, nitrogen can form 3, oxygen can form 2 bonds, fluorine forms 1 bond, and neon doesn’t form any bonds because it has no single electrons, only pairs.

When one element shares a single electron with another element having a single electron, the shared pair of electrons is a chemical bond. For example, hydrogen (having only 1 electron) can share this electron with fluorine to make hydrogen fluoride (HF). The Lewis dot structure of this compound is:

Figure 5.3. Lewis dot structure of hydrogen fluoride.

For clarity, the electron from hydrogen is colored red. The chemical bond is composed of one (black) electron from fluorine and one (red) electron from hydrogen.

In water, two hydrogen atoms are connected to an oxygen atom, and the Lewis dot structure of this compound is:

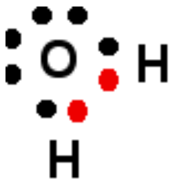


Figure 5.4. Lewis dot structure of water.

Before anyone gets wrong ideas; the electrons aren't permanently fixed to one particular side of the chemical symbol. We can always distribute the electrons in any fashion that is convenient or pleasing to us, provided that we keep them either as single electrons or as pairs of electrons (no groups of 3 or more electrons).

We could always draw the water molecule as:

Figure 5.5. Alternate Lewis dot structure of water.

The shared pair will always be somewhere between the two bonded atoms. Often, we show a shared pair as a single line, representing the chemical bond. Using this method we can draw our water molecule as:

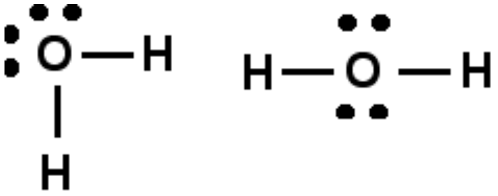


Figure 5.6. Lewis structures of water, showing two equivalent structures.

Strictly speaking, using a combination of lines for bonds and dots for unshared pairs of electrons are *Lewis structures*, not Lewis dot structures. The distinction is not particularly important for this course and you can use either. However, don't try to use both! Some students will draw a single line, and then include two dots as electrons (above the line, below the line, or straddling the line). This is wrong and confusing.

With two exceptions (hydrogen and boron) all non-metals must have a total of 8 electrons. These electrons can be shared pairs (bonds) or unshared pairs. Hydrogen forms one bond, and is satisfied having 2 electrons. Boron has 3 electrons, and can make 3 bonds for a total of 6 electrons. All other non-metals need 8 electrons.

The structures of many simple molecules can be determined by comparing the number of bonds that each element can form. However, not all molecules are equally simple, and small molecules can sometimes be deceptive. Consider carbon monoxide (CO). Carbon has 4 unshared electrons, and can make 4 bonds. Oxygen has 2 unshared electrons, and can make 2 bonds. What does carbon monoxide look like?

When the simple inspection method doesn't work, we use the "pooled electron" method. We add together the total number of valence electrons for our elements, and then distribute them so that each element has 8 electrons (either as shared pairs or as unshared pairs). For carbon monoxide, carbon contributes 4 electrons, oxygen contributes 6, and the total valence electrons are 10. These 10 electrons have to be distributed so that bonds are formed, and so that both carbon and oxygen have 8 electrons. These two requirements result in the following Lewis dot structure:

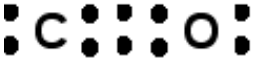


Figure 5.7. Lewis dot structure of carbon monoxide.

Or if we prefer to use lines for bonds:

Figure 5.8. Lewis structure of carbon monoxide.

This method works especially well for polyatomic ions. In sulfate (SO_4^{-2}), sulfur contributes 6 electrons, and each oxygen atom contributes 6 electrons. There are two additional electrons indicated by the -2 charge, for a total of 32 electrons. One possible structure would be a central sulfur atom surrounded by the oxygen atoms:

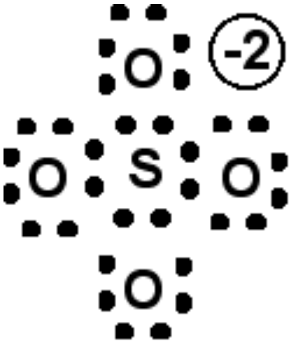


Figure 5.9. One possible Lewis dot structure for sulfate.

Other arrangements meeting the requirements that all electrons are used and each atom has 8 electrons are possible.

For positively charged polyatomic ions like ammonium and hydronium, we must subtract the positive charge from the valence electrons, since loss of electron(s) results in positive charge(s). Ammonium has a total of 8 valence electrons, as does hydronium, and their Lewis dot structures are:

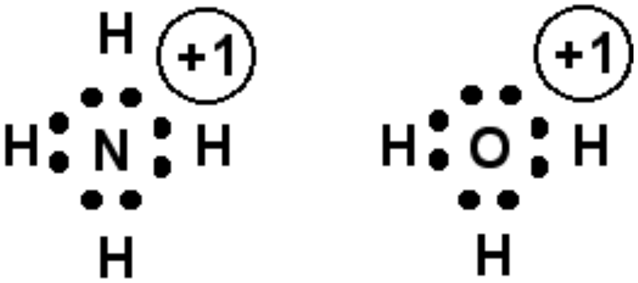


Figure 5.10. Lewis dot structures of ammonium (left) and hydronium (right).

Sometimes, a compound or ion can have more than one valid structure. An example is carbonate, which has three equivalent Lewis dot structures:

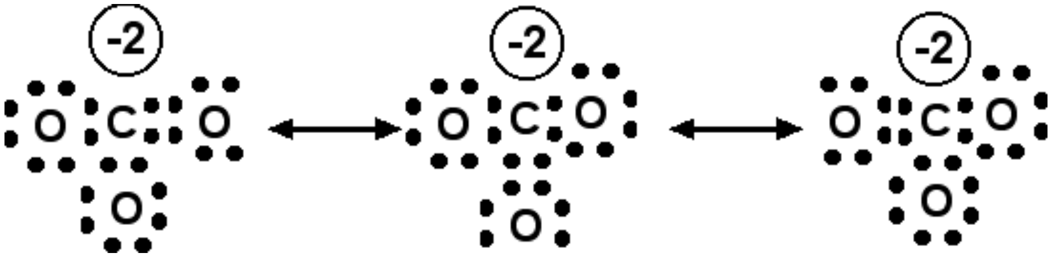


Figure 5.11. Three Lewis dot structures for the carbonate anion. The three structures are equivalent.

Notice that one of the oxygen atoms shares 4 electrons with carbon (it is double bonded to carbon). In each structure a different oxygen atom is double bonded to carbon. This is generally explained by the idea of ***resonance***. Resonance does **NOT** mean that the ion is rapidly interchanging between the three different forms, nor does it mean that there is an equal mixture of the three forms. Instead, the extra carbon-oxygen bond is somehow “spread out” over the entire molecule. Every carbon-oxygen bond is actually 1.3333 bonds.

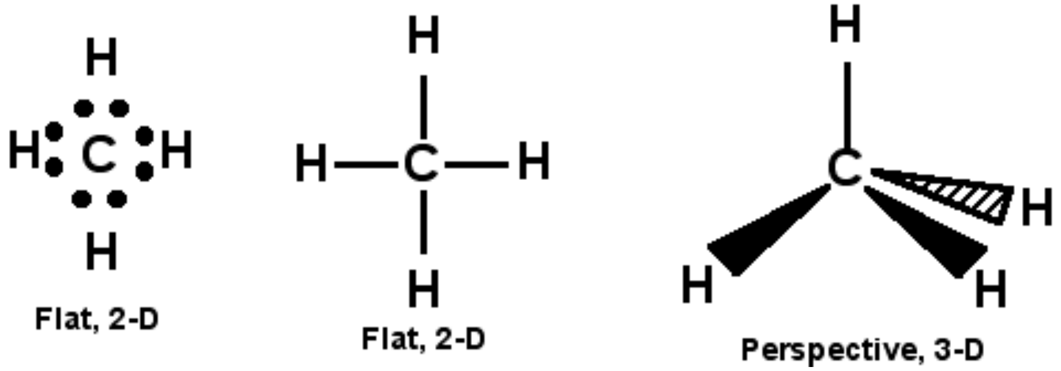
Now, this idea that we can have “fractional” bonds doesn’t make very much sense if we are talking about bonds being equivalent to two electrons, because what sense is there is talking about a fraction of an electron? However, you need to remember that the Bohr model of the atom, with electrons as small negatively charged particles orbiting the nucleus, is **WRONG** (although very convenient for many purposes). In reality, the electron (when “orbiting” the nucleus) is an electromagnetic wave. Just like two sound waves can combine to form a new sound wave, or two water waves can combine to form a larger water wave, two electromagnetic waves can combine to form a new electromagnetic wave. When two “electron-waves” combine, they form a new “electron-wave” that we call a "chemical bond". Having 1/3 of an electron is a difficult notion, but having one wave 1/3 as large as another wave isn't nearly as difficult to picture. Imagining how two electrons, as negatively charged particles, can hold two atoms together is difficult. Imagining two electrons combining to form an "electron-wave" may not help very much, but it is a better description of the chemical bond.

I'm going to continue talking about pairs of electrons as if they are actual small negatively charged bits of matter, but when they are IN the atom, they aren't.

VSEPR theory.

Valence shell electron pair repulsion theory (VSEPR) allows us to predict the geometry of a molecule, based on the arrangement of atoms and unshared electron pairs around a central atom. As we have seen from our Lewis dot structures, each atom tends to have four pair of electrons around it. If a molecule is sufficiently large (5 atoms total), then we have 4 atoms arranged around fifth atom at the center. Methane (CH_4) is a good example of this arrangement.

When we draw a methane molecule on a flat surface, the flat surface forces 2-dimensional geometry (unless we try to make some sort of perspective drawing). Below are some common representations of a methane molecule.



Flat, 2-D Flat, 2-D Perspective, 3-D

Figure 5.12. Three different representations of the methane molecule.

The perspective drawing tries to emphasize the 3-dimensional nature of molecules. By convention, the dark triangular lines are bonds projecting out of the surface of the paper, while the hashed triangle is a bond projecting into and through the paper. Methane is a tetrahedral molecule – a pyramid made of 4 equal triangles (as opposed to a normal pyramid made of 4 equal triangles and a square base). Tetrahedral geometry is the only 3-D geometry allowing four objects to be equally distributed around a central fifth object.

VSEPR theory states that a pair of electrons occupies almost the same volume of space as a hydrogen atom. Hydrogen fluoride, water, sulfate, ammonia, and hydronium ion have shapes similar to methane, but with unshared electron pairs replacing hydrogen atoms in some cases. These substances are shown below.

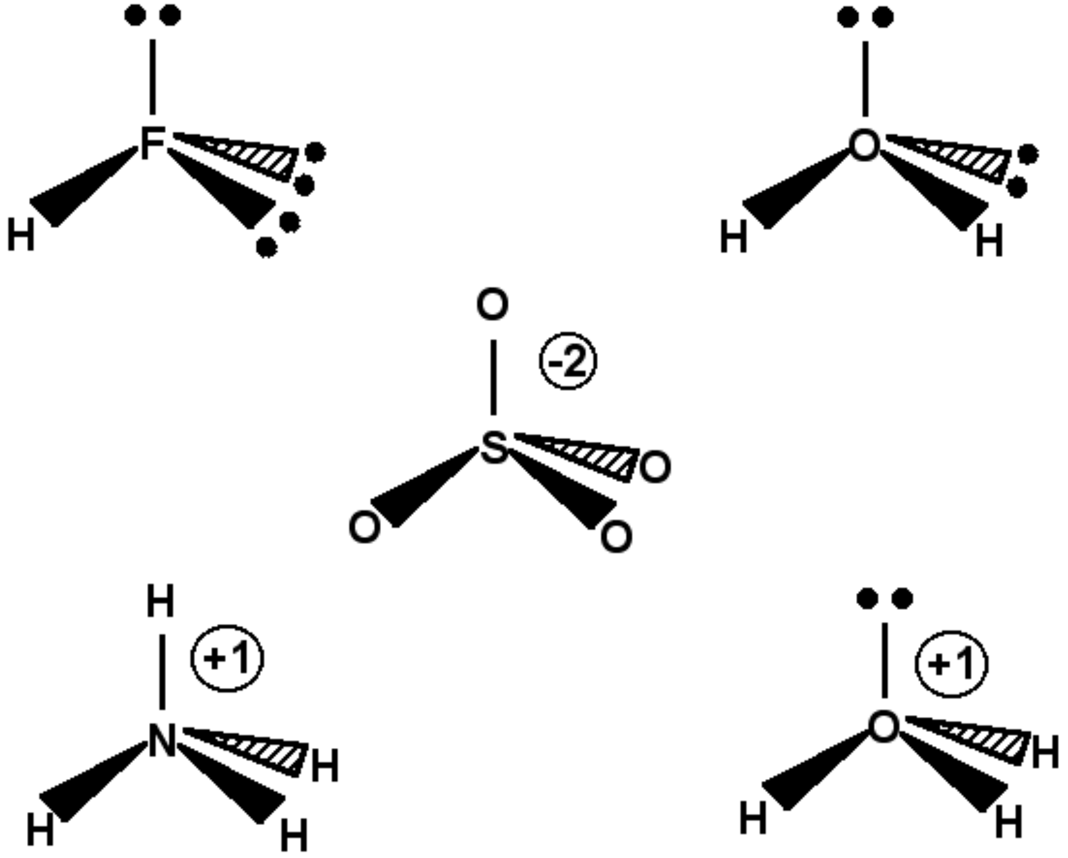


Figure 5.13. Geometry of hydrogen fluoride, water, sulfate, ammonium, and hydronium ions.

If there are fewer than 4 atoms bonded around a central atom, or if there are no unshared electron pairs, the molecular geometry changes. Table 5.3 summarizes various possible combinations of atoms and unshared pairs, and the resulting geometry. These “rules” apply to central atoms obeying the octet rule. Some elements, such as sulfur, have the ability for form “expanded octets”, resulting in more than 8 electrons around sulfur. These are exceptions, and aren’t particularly useful in learning the general pattern exhibited by most compounds, so we won’t worry about these examples here.

# of atoms bonded to central atom	# of unshared electron pairs	Molecular geometry	Examples
2	0	Linear	Carbon dioxide
2	1	Bent	Sulfur dioxide
2	2	Bent	Water
3	0	Trigonal planar	Boron trifluoride
3	1	Trigonal pyramidal	Ammonia
4	0	Tetrahedral	Methane

Table 5.3. Molecular geometry based on bonded atoms and lone pairs.

Bond polarity and molecular polarity.

When identical atoms share electrons in a chemical bond, the sharing must be exactly equal. Consider two hydrogen atoms bonded to form a molecule. Neither hydrogen atom can exert greater control of the electrons than the other. The electron pair is shared equally between the two atoms, because the two atoms are identical.

When dissimilar atoms share electrons, the situation is different. Each element has a characteristic ***electronegativity***, a chemical property describing the tendency of an atom to attract electrons towards itself. Table 5.4 shows Pauling electronegativity values for selected non-metal compounds. “Pauling electronegativities” are named for American chemist Linus Carl Pauling (1901 – 1994), who developed the concept of electronegativity. Pauling was one of the founders of quantum chemistry and of molecular biology. He is the only person to win two unshared Nobel Prizes, the first for Chemistry and the second for Peace.

We can estimate how equally electrons are shared between elements by comparing electronegativities for pairs of elements – simply subtract the smaller value from the larger. This result, the difference in electronegativity (DEN) is used to determine the ***bond polarity***.

If DEN < 0.6, then the bond is considered to be “nonpolar covalent”. It is essentially the same as a bond between identical atoms. Carbon-hydrogen bonds have DEN = 0.4, so they are nonpolar covalent.

H 2.1	B 2.0	C 2.5	N 3.0	O 3.4	F 4.0
		Si 1.9	P 2.2	S 2.6	Cl 3.2
				Se 2.5	Br 3.0
					I 2.7

Table 5.4. Pauling electronegativity values for selected non-metals.

If DEN < 1.6, then the bond is considered to be "polar covalent". In this bond, the electrons are shared, but they are strongly attracted to the element having the higher electronegativity value. Hydrogen-oxygen bonds have DEN = 1.3, and the shared electrons are attracted towards oxygen and away from hydrogen.

If DEN > 2.0, the bond is considered ionic. The only example we have of this with the non-metals is the silicon-fluorine bond (DEN = 2.1). However, many metals (not shown) have very low electronegativity values and readily form ionic bonds. Sodium or calcium's electronegativity is 1.0. Sodium-oxygen bonds have DEN = 2.4, sodium-chlorine bonds have DEN = 2.2. Generally, we classify any compound between metals and non-metals as ionic as a matter of course.

If DEN is between 1.6 and 2.0, then the bond classification depends on the type of elements combined. If one of the elements is a metal, then the bond is considered ionic, while if both elements are non-metals, then the bond is polar covalent. Hydrogen fluoride has DEN = 1.9, but since both elements are non-metals the bond is polar covalent. Scandium's electronegativity is 1.3, and scandium chloride has DEN = 1.9. Since scandium is a metal, this bond is ionic.

We can readily show the attraction of shared pairs towards one element by using arrows instead of straight lines, with the arrowhead pointing towards the element that attracts the electrons (Figure 5.14). If the bond is nonpolar covalent, we can use either a line or a double-headed arrow. Notice the **"δ+"** and **"δ-"** symbols – these represent partial + and – charges, due to the imbalance in electron sharing. In water for example, the electrons are strongly attracted to oxygen, resulting in oxygen being a little bit negative. The hydrogen atoms are slightly deprived of electrons and become a little bit positive.

Molecular polarity is the result of unshared electron pairs, bond polarity, and geometry. A polar molecule will have one side or end of the molecule slightly positive and the opposite end or side slightly negative. If you use the following guidelines, you will rarely go wrong in assigning molecular polarity. Once again, these guidelines apply only to central atoms obeying the octet rule. Atoms with expanded octets have modified guidelines.

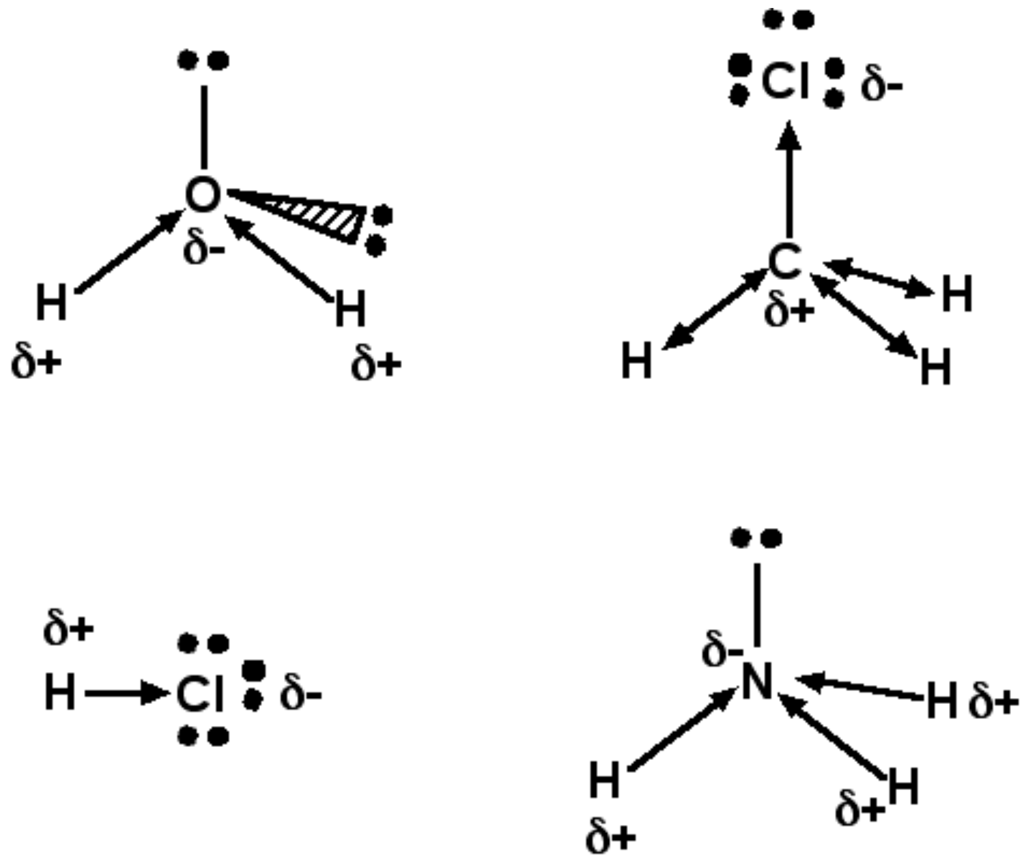


Figure 5.14. Bond polarities are shown using arrows. Shared electrons are drawn towards the element having higher electronegativity.

1. Does the central atom have unshared electron pairs? If "yes", then the molecule is polar. Water and ammonia are classic examples; unshared electrons on the central atom guarantee that the molecule is polar.

2. If there are no unshared electron pairs on the central atom, then does the atom have polar covalent or ionic bonds? If all of the bonds are nonpolar covalent, then the molecule is nonpolar. Methane (CH_4) is the classic example of this guideline in action. The Lewis dot structure shows that the central carbon atom does not have any unshared electron pairs. The carbon-hydrogen bonds have DEN = 0.4, clearly nonpolar covalent bonds. The methane molecule is therefore nonpolar.

3. If there are no unshared electron pairs on the central atom, but there are polar covalent bonds, then the molecule's geometry determines molecular polarity. Geometry can cancel the effects of bond polarity. The simplest way to see this effect is to consider a series of compounds, CH_4, CH_3F, CH_2F_2, CHF_3, and CF_4 (Figure 5.15). Methane is nonpolar as described above. Carbon-fluorine bonds are polar covalent, shared electrons are attracted to the fluorine atom(s), and the molecule has slightly positive/negative sides. In

CF_4, the “outside” of the molecule is uniformly negative, while the “inside” (the carbon atom) is slightly positive. However, nothing can get near the slightly positive carbon atom, so the molecule is effectively non-polar.

4. If the substance is made from ions, then it is polar.

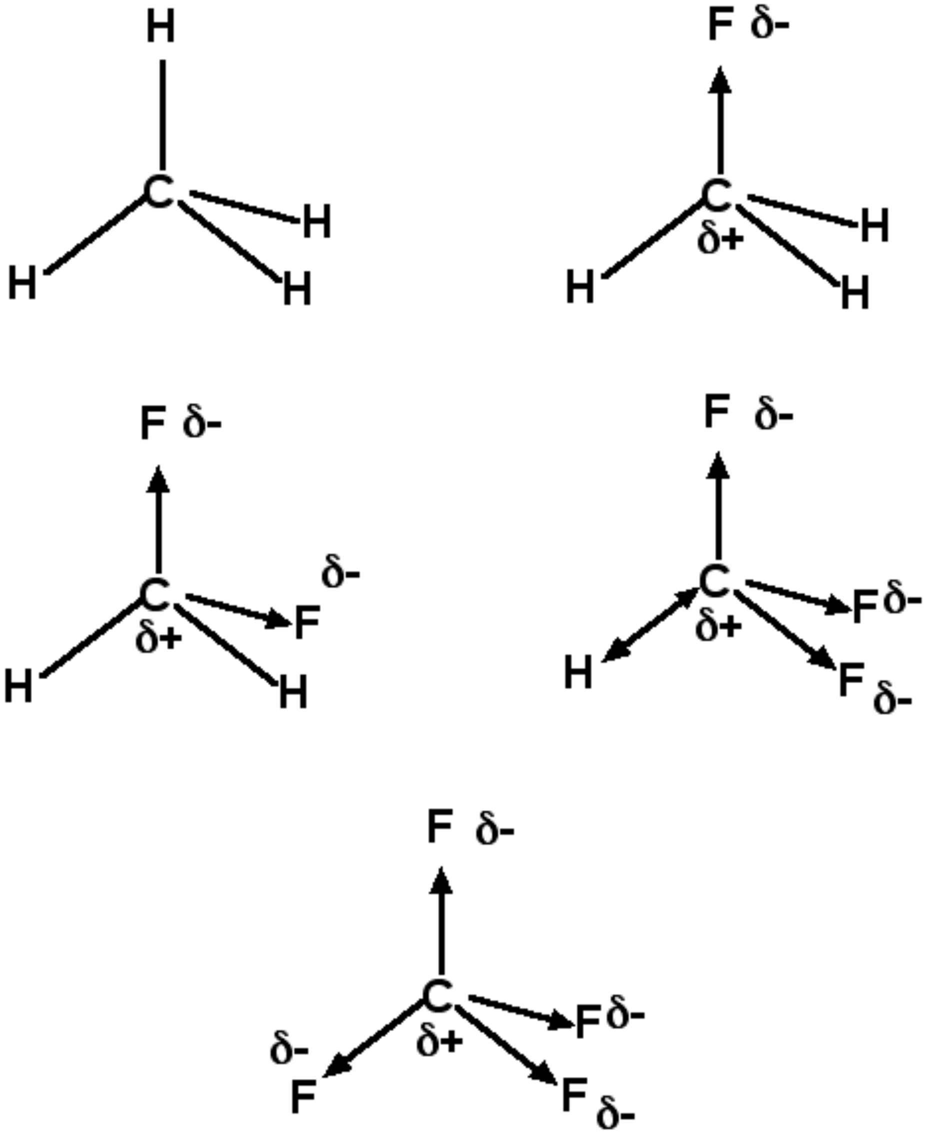
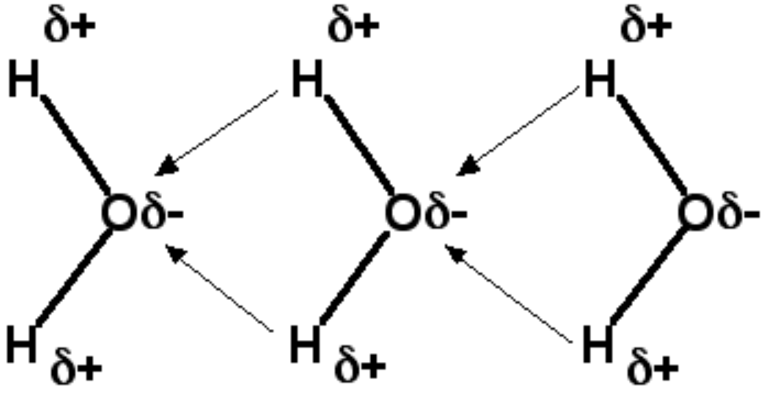


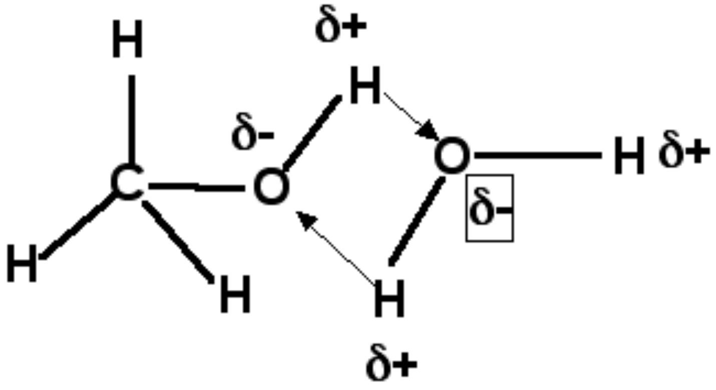
Figure 5.15. Bond polarity effects can be cancelled out by the molecules geometry.

Molecular polarity explains a variety of chemical and physical properties. One example is the expression “like dissolves like”. Polar and ionic substances readily dissolve in polar materials, and are insoluble in nonpolar substances. Sodium chloride (ionic) or ammonia (polar) dissolves readily in water (polar).

Hexane (C_6H_{14}, no unshared pairs and nonpolar bonds) is nonpolar and doesn't dissolve particularly well in water. ***Hydrogen bonding*** is an intermolecular force in which the hydrogen atom attached to an oxygen, nitrogen, or fluorine atom is attracted to the oxygen, nitrogen, or fluorine atom of an adjacent molecule. Water is the classic example of hydrogen bonding, either with another water molecule or with any molecule containing oxygen, nitrogen, or fluorine (Figure 5.16).



Hydrogen bonding in water.



Hydrogen bonding between water and alcohol.

Hydrogen bonds are shown as arrows.

Figure 5.16. Hydrogen bonding examples.

Chapter 5 Homework:

Vocabulary. The following terms are defined and explained in the text. Make sure that you are familiar with the meanings of the terms as used in chemistry. Understand that you may have been given incomplete or mistaken meanings for these terms in earlier courses. The meanings given in the text are correct and proper.

Valence electrons

Principal energy level

Noble gas configuration

Bond polarity

Electron(ic) configuration

Sublevels

Resonance

Hydrogen bonding

Periods

Orbitals

Electronegativity

1. For the following elements, write out proper electronic configurations.
   a. He
   b. N
   c. Al
   d. Fe
   e. Br
   f. Y
   g. Ag
   h. I
   i. U

2. For the following elements, draw proper Lewis dot structures.
   a. C
   b. O
   c. Cl
   d. S
   e. P
   f. B

3. For the following compounds and ions, draw proper Lewis dot structures.
   a. BH_3
   b. CH_2Cl_2
   c. OCl^-
   d. SO_2
   e. SO_3^{-2}
   f. NO_3^-
   g. CH_2O

4. For the substances in question 3, use VSEPR theory to determine the shape of the molecule or ion.

5. For the molecular substances in question 3, describe whether or not the molecule is polar or nonpolar. Use the electronegativity values and table in the text to assign bond polarities.

6. For the following compounds, indicate which ones can form hydrogen bonds with water molecules.

   a. BH_3
   b. CH_2Cl_2
   c. H_2S
   d. HCl
   e. CO_2
   f. C_6H_6
   g. SO_2
   h. CH_3CH_2SH
   i. NH_3
   j. CH_3COCH_3
   k. CH_3NH_2

Answers:

1. For the following elements, write out proper electronic configurations.

   a. He 2 electrons; $1s^2$

   b. N 7 electrons; $1s^2$, $2s^2$, $2p^3$

   c. Al 13 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^1$

   d. Fe 26 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^6$

   e. Br 35 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^5$

   f. Y 39 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^1$

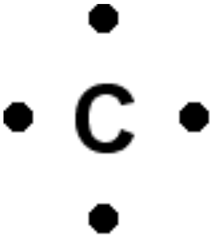
   g. Ag 47 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^9$

   h. I 53 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^5$

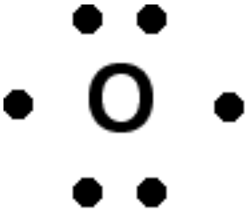
   i. U 92 electrons; $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^6$ $6s^2$, $4f^{14}$, $5d^{10}$, $6p^6$, $7s^2$, $5f^4$

2. For the following elements, draw proper Lewis dot structures.

   a. C

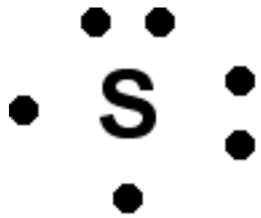


   b. O

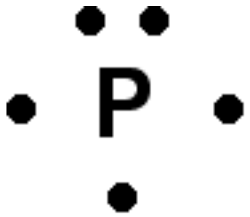


c. Cl

d. S



e. P

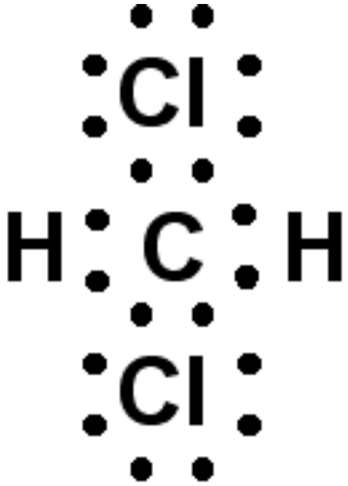


f. B

3. For the following compounds and ions, draw proper Lewis dot structures.

a. BH_3

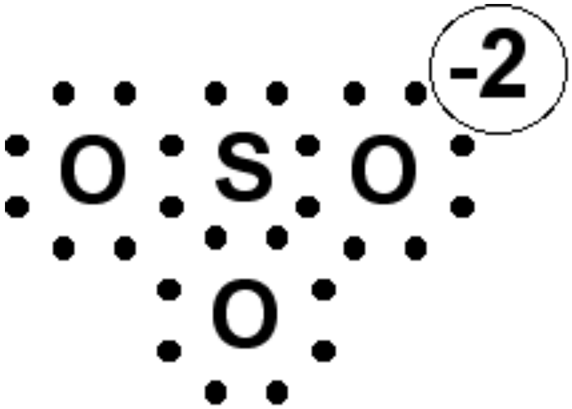
b. CH_2Cl_2



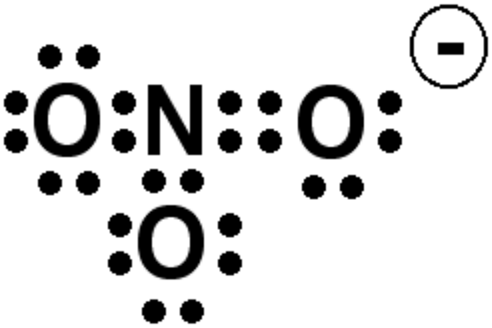
c. OCl^-

d. SO_2

e. SO_3^{-2}



f. NO_3^-



g. CH_2O

4. For the substances in question 3, use VSEPR theory to determine the shape of the molecule or ion.

a. BH_3 3 atoms bonded to the central B atom, no unshared pairs on B, shape is trigonal planar.

b. CH_2Cl_2 4 atoms bonded to the central C atom, no unshared pairs on C, shape is tetrahedral.

c. OCl^- No central atom, just two atoms bonded together. Linear.

d. SO_2 2 atoms bonded to central S atom, one unshared pair on S, shape is bent.

e. SO_3^{-2} 3 atoms bonded to central S atom, one unshared pair on S, shape is trigonal pyramidal.

f. NO_3^- 3 atoms bonded to central N atom, one unshared pair on N, shape is trigonal pyramidal.

g. CH_2O 3 atoms bonded to central C atom, no unshared pairs on C, shape is trigonal planar.

5. For the molecular substances in question 3, describe whether or not the molecule is polar or nonpolar. Use the electronegativity values and table in the text to assign bond polarities.

a.	BH_3	The DEN for B-H bonds is 0.1, which makes the bond non-polar. And there are no unshared electrons on B, so the molecule is non-polar.
b.	CH_2Cl_2	The C-H bonds are non-polar (DEN = 0.4), but C-Cl bonds are polar covalent (DEN = 0.7). The chlorine side of the molecule has a partial negative charge (δ-) while the hydrogen end has a partial positive charge (δ+). The molecule is polar.
c.	OCl^-	This is an anion, so it has a full negative charge. Polar/nonpolar doesn't apply.
d.	SO_2	The sulfur – oxygen bond is polar covalent (DEN = 0.8), and the Central sulfur atom has an unshared pair of electrons. The molecule is polar.
e.	SO_3^{-2}	This is an anion, and polar/nonpolar doesn't apply.
f.	NO_3^-	This is an anion, and polar/nonpolar doesn't apply.
g.	CH_2O	The C-H bonds are nonpolar (DEN = 0.4), but C=O is polar covalent (DEN = 0.9). There is a partial negative charge on oxygen (δ-) and a partial positive charge on hydrogen (δ+). The molecule is polar.

6. For the following compounds, indicate which ones can form hydrogen bonds with water molecules.

a.	BH_3	No hydrogen bonding
b.	CH_2Cl_2	No hydrogen bonding
c.	H_2S	Hydrogen bonding possible
d.	HCl	No hydrogen bonding
e.	CO_2	Hydrogen bonding possible
f.	C_6H_6	No hydrogen bonding
g.	SO_2	Hydrogen bonding possible
h.	CH_3CH_2SH	Hydrogen bonding possible
i.	NH_3	Hydrogen bonding possible
j.	CH_3COCH_3	Hydrogen bonding possible
k.	CH_3NH_2	Hydrogen bonding possible