CHAPTER 5. CHEMICAL REACTIVITY AND THE PERIODIC TABLE

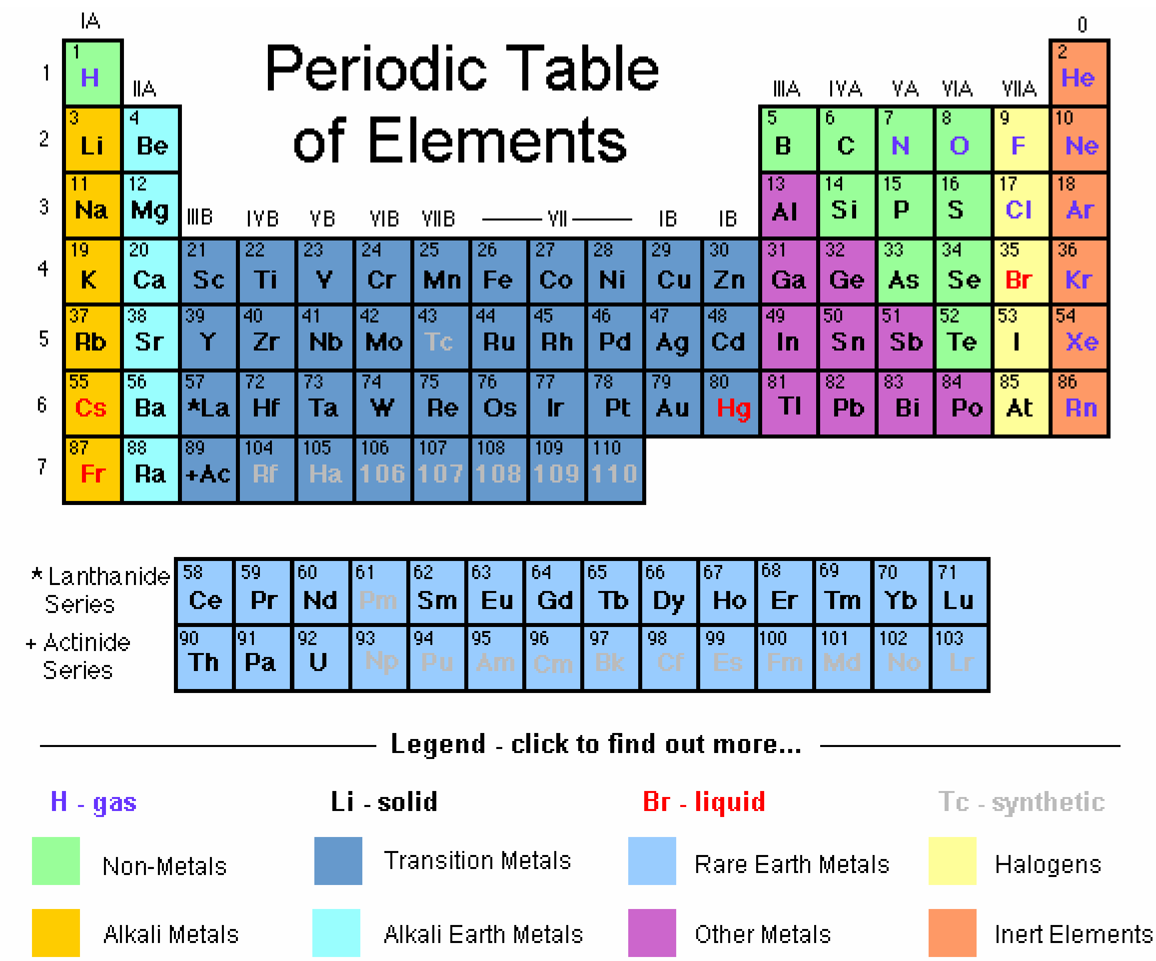
Electron Structure and the Periodic Table

The concept that a few elements combined to form compounds of unlimited number was an important insight in understanding the different forms of matter. Understanding how and why the elements combined, however, posed a difficult problem in the early days of chemistry. Some elements reacted explosively with other substances. Some did not react at all. Some elements were solids, for example the shiny metals like gold and silver that had been purified by the alchemists. Others were gases, like hydrogen and oxygen. How were the chemists to find a unifying concept that would make sense of this multiplicity of properties? Some success was obtained in organizing elements with similar properties into groups or families in the early days of chemistry. Not until the discovery of the atomic structure was a clear explanation given for the ways in which atoms interact with each other.

The nucleus, locked deep in the heart of the atom, does not directly interact with other atoms, although the positive charge of the protons in the nucleus can have important effects on an atom's chemical properties. The electrons form the outer part of the atom, and so the electrons are the part of the atom which interact directly with other atoms. Thus, in order to understand chemical reactivity it is necessary to have some knowledge about electrons. The more thoroughly we understand how electrons behave, the greater will be our ability to understand and predict chemical reactivity.

The electron was not discovered until the twentieth century. Neils Bohr's idea that electrons existed in energy levels explained the colors that had been observed in spectroscopic experiments as atoms absorbed and emitted light. It was also the beginning of the basis of an explanation for the chemical reactivities of elements. Quantum mechanics provided a more sophisticated explanation for the properties of the electrons inside the atom, predicting energy sublevels as well as energy levels and giving an explanation for the numbers of electrons that each energy level and sublevel could hold. Although quantum mechanical calculations are an area of specialty for only a small minority of chemists, the insights they have provided into atomic structure provide a part of the basic working knowledge of chemistry. We will see how the numbers of electrons in energy levels and sublevels are the most important factor in determining patterns of chemical reactivity of the elements.

The **periodic table** is a list of the elements organized in a way that elements which react similarly appear in groups or families. The Russian chemist Mendeleev was the first to succeed in producing a successful periodic table, which he presented to the Russian Chemical Society in 1869 in a paper called *On the Relation of the Properties to the Atomic Weights of the Elements.* In an arrangement which is essentially the same as periodic table we use today, he was able to organize the known elements into horizontal rows called **periods** in such a way that elements with similar properties and reactivities appeared in vertical groups. In order to do so, he had to leave some blank spaces. He predicted that elements would be discovered which would fill in those blank spaces, and he predicted as well the properties that these elements would have. Several of his predictions were soon fulfilled, and his periodic table was adopted by chemists worldwide as the key to understanding the properties of the elements.



Reference: http://61.19.145.8/student/m5year2006-2/502/group11/index.html

We now know that the relationships Mendeleev discovered correspond to configurations of electrons within the atoms. Today we can use his periodic table not only as a key to the properties and reactivities of elements, but as a key to the electron structure which governs these properties and reactivities. From our study of atomic structure we have learned that the maximum number of electrons that can occupy an energy level is given by the formula $2n^2$, where n is the energy level. For the first energy level, $n = 1$. The maximum number of electrons that can occupy this level is $2(1)^2$, or 2. Looking at the first horizontal row of the periodic table, we see only two elements, hydrogen and helium. The maximum number of electrons in this level is given by $2(2)^2$, or 8. Looking at the second row of the periodic table, we see eight elements. For the third electron energy level, the formula for the maximum number of electrons gives $2(3)^2$, or 18. Why does the third row of the periodic table, then, show only eight elements? The answer lies in the existence of energy subshells. One subshell of the third level, the *d* subshell, can contain ten electrons, and the energy level of this subshell is higher than the first, or *n* subshell of the fourth level, which contains two electrons. As a result, predicting the population of energy levels from the periodic table is simple to do only through element 20.

To predict the number of electrons in each energy level for the elements up to atomic number twenty, simply fill the energy levels one at a time, starting with energy level one, the lowest energy level. To find the number of electrons in an atom that will populate the energy levels, recall that the atomic number gives the number of positively charged protons in the nucleus; the number of negatively charged electrons in the atom is the same. For example, hydrogen, the element with atomic number 1, has only one proton and one electron. This electron will occupy the lowest possible energy level, level 1; this is the level closest to the nucleus.

Problem example 5-1: Show the population of electrons in each electron energy level for the element lithium. How many electrons are in the outermost, or highest, energy level? These are the valence electrons.

The periodic table gives the atomic number 3 for the element lithium. The element has therefore 3 protons and 3 electrons. The first two electrons will go into the lowest energy level, level 1, filling that level to capacity. The next electron must go into the next higher energy level, level 2. The resulting electron distribution in the energy levels for lithium is:

Energy level 1: 2 electrons

Energy level 2: 1 electron

The number of electrons in the highest energy level, level 2, of lithium is one. Notice that lithium appears in the second period (horizontal row 2, corresponding to the highest energy level filled, level 2) and that it appears in the first vertical group (Group IA) corresponding to the number of electrons in the highest energy level, 1 electron. So now we have a quick way to find how many valence electrons: it's in **group 1**, so, one **valence electron**!

Problem example 5-2: Show the distribution of electrons in each electron energy level for the element potassium. How many electrons are in the highest energy level? These are the valence electrons.

From the periodic table, the atomic number of potassium (symbol K) is 19. The potassium atom, then, has 19 electrons. The first energy level will fill with two electrons. The second energy level will hold eight electrons, for a total of ten electrons in the first two levels. From the periodic table we see that the next eight electrons go into the third level, for a total of eighteen. The nineteenth electron goes into the fourth energy level.

Energy level 1: 2 electrons

Energy level 2: 8 electrons

Energy level 3: 8 electrons

Energy level 4; 1 electron

The number of electrons in the highest energy level, level 4, of potassium is one. Notice that potassium appears in the fourth period (horizontal row 4, corresponding to the highest energy level filled, level 4) and that it appears in the first vertical group (Group 1A) corresponding to the number of electrons in the highest energy level, 1 electron. So, again, **Group 1, one valence electron**!

Problem example 5-3: Show the distribution of electrons in each electron energy level for the element chlorine. How many electrons are in the highest energy level?

From the periodic table, the atomic number of chlorine (symbol Cl) is 17. The chlorine atom, then has 17 electrons. The first energy level will fill with two electrons. The second energy level will hold eight electrons, for a total of ten electrons in the first two levels. From the periodic table we see that the next seven electrons go into the third level, for a total of seventeen.

Energy level 1: 2 electrons

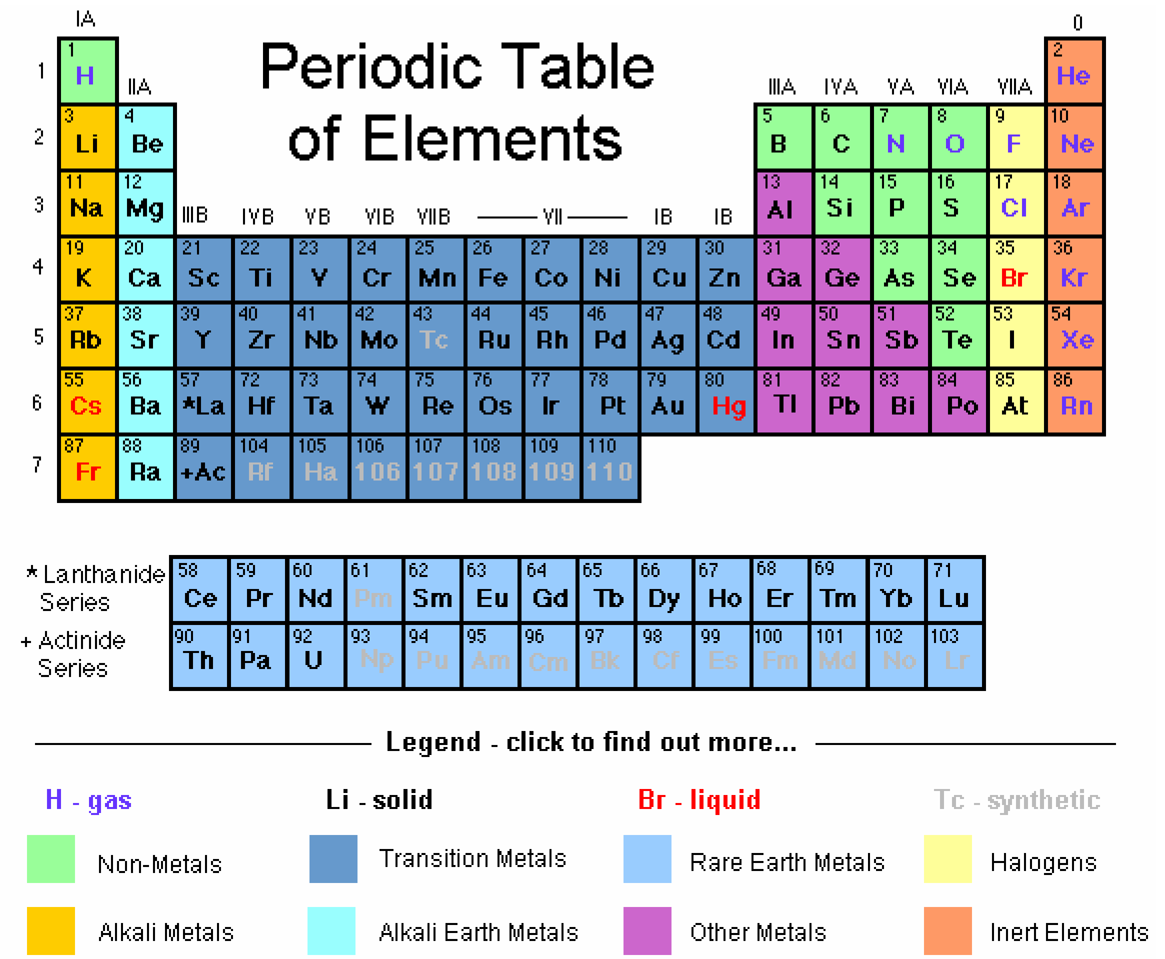
Energy level 2: 8 electrons

Energy level 3: 7 electrons

The number of electrons in the highest energy level, level 3, of chlorine is seven. Notice that chlorine appears in the third period (horizontal row 3, corresponding to the highest energy level filled, level 3) and that it appears in the seventh vertical group (Group 7A) corresponding to the number of electrons in the highest energy level. (Some periodic table are numbered differently, but we will use this easy version.) That way, **group 7, 7 valence electrons.**

Groups of Elements on the Periodic Table

Before we examine in detail the relationship between electron structure and the properties of the elements, it is useful to develop some familiarity with the groups, or families, of elements that appear on the periodic table.



On the far right of the table are a vertical group of elements called the **inert gases**, or the **noble gases**. These elements are all gases, and have the unusual property of being unwilling to react with any element. To the left of the inert gases is another distinctive group of elements, the **halogens**, group 7A. The halogens are gases also, but highly reactive ones. One the far left of the periodic table are group 1A, the **alkali metals**, and group 2A, the **alkaline earth metals**. These are all highly reactive solid substances with the distinctive shiny surface and electrical conductivity which are the properties of all metals. The next ten rows across the middle of the periodic table are called the **transition metals**. In this family of elements are some familiar metallic elements like iron, with the symbol Fe, and copper, with the symbol Cu. The **inner transition elements**, the **lanthanide series** and the **actinide series**, of elements 58 through 103, occupy fourteen vertical rows of the periodic table, and are found placed separately beneath the table so as to make its shape more manageable.

Metals comprise a large portion of the periodic table, from groups 1 through the transition elements, and diagonally in a line through gallium (Ga), tin (Sn), and bismuth (Bi). (See the figure above.) An important exception is hydrogen, in the upper left corner, which is not a metal, but a reactive gas. **Nonmetals** are found on the right of the periodic table beginning with carbon, and between the metals and the nonmetals are the **metalloids** boron, silicon, germanium, arsenic, antimony, and tellurium, in diagonal rows beginning with boron (B).

Electron Structure and Chemical Reactivity

Paradoxically, the first elements to examine in order to understand chemical reactivity are those which do not react at all, the noble gases at the far right of the periodic table; their names are helium, neon, argon, krypton, xenon, and radon. The atomic numbers of the first three of these are 2, 10, and 18, which appear to form a pattern. The inert gas neon appears eight elements after the inert gas helium. Eight elements later, the inert gas argon appears. The Englishman John Newlands noticed this relationship of recurring properties in each succeeding eighth element, which he called the Law of Octaves from its similarity to the musical scale of eight notes. His observations, however, evoked only humorous comments from his colleagues, who could see no benefit to be derived from looking for a musical theme in the elements.

As we have learned in our study of atomic structure, the atomic number gives the number of positively charged protons in the nucleus, and this also is the number of negatively charged electrons in the atom. These inert gas atoms with 2, 10, and 18 electrons, then, appear to have particularly favorable electron configurations. The high reactivities of elements like group 7A, the halogens, can be interpreted as attempts to attain an electron configuration like that of the noble gases.

Electron Dot Formulas and the Octet Rule

Today we know that the patterns of chemical reactivity derive from electron configurations. The device that we often use to predict chemical reactivity, however, is almost as simplistic in appearance as the Law of Octaves. It is often called the **octet rule**: many atoms seek to obtain the favorable electron configuration of eight electrons in the outer electron energy level, or shell. Electrons in the highest occupied energy level, sometimes called the outer shell, are represented by dots, arranged in pairs around the symbol of the element. These are called the **valence electrons**.

Problem example 5-4: Draw an electron dot diagram for the element neon.

The electron dot diagram represents the electrons in the outer energy shell. Before we can draw a dot diagram, we must determine the number of electrons in each energy level as in problem examples 5-1 through 5-3.

Neon has an atomic number of 10, so there are 10 electrons in the neon atom. As in the examples above, the electrons fill the energy levels starting at the lowest level, resulting in electron distributions of:

Energy level 1: 2 electrons
Energy level 2: 8 electrons

Notice that, as in the examples above, the position of neon on the periodic table correlates with its electron configuration. Neon is in horizontal row 2 of the periodic chart, corresponding with the highest energy level filled, level 2. It is in vertical group 8A, corresponding with the number of electrons in this energy level, 8 electrons.

Now that we have determined the number of electrons in the outer energy level, 8, we can draw the electron dot diagram for neon. The element symbol Ne is surrounded by four pairs of electrons arranged to form a sort of square box. Draw them in below:

Ne

Neon, we observe, has a filled outer energy level. The corresponding electron dot diagram shows neon surrounded by four pairs of electrons. If we examined the electron configurations of the other inert gases, we would see that the same is true for all of them. Helium, for example, with two electrons, has a filled outer energy level, level 1. The seemingly "magic" atomic numbers of 2 and 10 for the first two inert gases then, are simply the numbers of electrons in filled energy levels. The nonreactive properties of these inert gases must reflect the fact that their energy levels are filled.

The octet rule is very useful in many cases for indicating electron configurations and predicting chemical reactions. Its limitations, however, should be realized. Though the flat, square shape is useful for arranging an octet, or group of eight, of electrons, atoms are actually neither square nor flat. They are three-dimensional, with electrons whizzing through space in patterns determined by the type of orbital, or energy sublevel, they are in. Moreover, not all elements obey the octet rule. Two conspicuous examples of elements that do not obey the octet rule which we already know about are hydrogen and helium. Because they have only one and two electrons, respectively, they use only the first electron energy level, and hence they require only two electrons, not eight, to have a filled shell.

Problem example 5-5: Draw an electron dot diagram for the element helium.

Helium has an atomic number of 2, so there are 2 electrons in the neon atom. As in the examples above, the electrons fill the energy levels starting at the lowest level, resulting in an electron distribution of:

Energy level 1: 2 electrons

To draw the electron dot formula for helium, we simply place a pair of dots to indicate these two electrons next to the symbol He for helium.

He:

Helium is in the first row of elements in the periodic table because its outer electron energy level is level 1. Therefore it does not obey the octet rule, and no more than two electrons can be put in its electron dot formula.

To draw electron dot formulas for other elements that are not inert gases, the procedure is the same. Determine the number of electrons in the outer energy level, then place a dot for each electron around the symbol for the element. At this point, you have probably realized that there are two possible ways to determine the number of electrons in the outer energy level of the atom, called valence electrons. Looking at the total number of electrons and subtracting the number of electrons in the filled inner shells gives the number of valence electrons. More simply, the vertical group number (given at the top of the group on the periodic table) is the same as the number of valence electrons for all the elements in the group.

Problem example 5-6: Draw an electron dot diagram for the element chlorine.

To determine the number of valence electrons in chlorine, we can start with the total number of 17 electrons and determine the total electron distribution in the energy levels as in Problem Example 5-3. The number of valence electrons, or electrons in the outer energy level, found in this way is seven. Alternatively, as we become familiar with the periodic table, we find that chlorine is in vertical group 7A, corresponding to seven valence electrons in this atom. The electron dot diagram which indicates these seven valence electrons is:

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 . .
: Cl .
 . .
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Problem example 5-7: Draw an electron dot diagram for the element sodium.

Sodium, symbol Na, has atomic number 11. If we find the distribution of all electrons in the atom we find:

Energy level 1: 2 electrons

Energy level 2: 8 electrons

Energy level 3: 1 electron

Alternatively, we can simply find from the fact that sodium is in vertical group 1 that there is one valence electron in the sodium atom. The electron dot formula for sodium is therefore:

Na .

Electron Transfer and the Formation of Ionic Compounds

How can we use electron dot formulas to predict the reactivity of elements? The principle is very simple: If an octet of electrons is an especially favorable configuration for electrons, then elements will seek to fill their outer shells so that they have those eight electrons. An example of a type of chemical reaction that illustrates this principle is given by the reaction of sodium and chlorine. We have already found the electron dot structures for these elements. Sodium has one valence electron and chlorine has seven:

Na .   .

. .
: Cl .
. .

Looking at the dot structures, it is easy to see that the chlorine atom could achieve an octet simply by taking away the electron from sodium. When the elements sodium and chlorine are put together, that is exactly what happens. Transfer of one electron from the sodium atom the chlorine atom occurs. To indicate this process with electron dots, we can move the lone dot from the sodium symbol to the chlorine symbol, thus giving the chlorine eight electrons, a full octet. The chlorine nucleus, however, is unaffected by this electron transfer. It still has only 17 positively charged protons. The extra electron taken from the sodium has brought the electron total for chlorine to 17 + 1 = 18 electrons. The extra electron means that the chlorine now has a negatively charged electron that is not balanced out by a positively charged proton. We need to indicate this negative charge that is now carried by the chlorine by a -1 sign on the chlorine that now bears an octet of electrons. When an atom carries an electrical charge like this, it is called an **ion**. We call the negative ion that has been formed by adding an electron to chlorine a ***chloride ion***. Draw its electron dot formula below by putting 4 pairs of electrons for a total of eight as you did with neon.

: Cl -

Now that the electron has been taken from the sodium atom, how has it been changed? Taking one dot form the electron dot formula leaves no more dots. This, of course, does not mean that sodium is totally devoid of electrons. The electron distribution we found in problem example 5-7 was:

Energy level 1: 2 electrons

Energy level 2: 8 electrons

Energy level 3: 1 electron

With the loss of the valence electron, the electron in the highest energy level, the electron distribution has become:

Energy level 1: 2 electrons

Energy level 2: 8 electrons

Energy level 2 has eight electrons, just like the inert gas neon. The sodium atom has also benefitted, then, from this loss of an electron, attaining a favorable electron energy configuration with a filled outer energy level. Another important change has occurred, however, with the loss of an electron. The negative charge carried by the electron has been lost. Sodium still has 11 positively charged protons in its nucleus, but only 10 negatively charged electrons remaining. Therefore it carries a net charge of +1, and it has now become a *sodium ion*. The electron dot formula for the sodium ion shows that it has no valence electrons left and that it carries a positive charge. It is written simply as :

$$Na^+$$

Show the overall electron transaction between chlorine and sodium using electron dot formulas by writing in the electron dots in the following equation:

$$Na \quad + \quad Cl \quad \text{-->} \quad Na^{+} \quad + \quad Cl^{-}$$

These tiny dots written on the surface of a page hardly do justice to the dramatic reality of the actual chemical reaction. Both elemental sodium and elemental chlorine are highly reactive substances, and therefore can be dangerous to handle unless every precaution is taken in dealing with them. The tendency for the sodium atom to lose its single electron or for the chlorine atom to gain one more electron is a very potent driving force, and these substances will not only react with one another but with many other substances they might come into contact with. If they are carefully brought together, sodium metal and chlorine gas react with violent force. The final product, containing positive sodium ions and negative chloride ions, is called sodium chloride. It is the same as ordinary table salt.
http://jchemed.chem.wisc.edu/jcesoft/cca/CCA3/MAIN/NACL/PAGE1.HTM

Substances like sodium chloride that are composed of positive and negative ions form an important class of chemical compounds known as **ionic compounds**, or **salts**. These compounds are held together by the very strong attraction of positive and negative electrical charges for one another. These strong attractive forces are called **ionic bonds**. Such strongly bonded substances tend not to melt easily, to decompose readily, or to react with other substances. They are held together in orderly arrays called **crystal lattices**, in which the positive ions and negative ions alternate in order to maximize the attractive forces between the ions of opposite charge and minimize the electrostatic repulsion between ions of like charges. The sketch below shows the arrangement of the sodium chloride crystal lattice.

Although the forces holding together the crystal lattice are very strong, the crystal is brittle and can be fractured along the flat planes of the lattice. Table salt, for instance, is made of small grains that were once parts of larger pieces of sodium chloride crystal. The flat planes that form the sides of a large sodium chloride crystal or other large crystal structures you may have seen are typical of crystal structures. On a visible scale they indicate the orderly, flat arrays of the crystal lattice. Similar flat surfaces appear if sodium chloride crystals from a salt shaker are viewed under a microscope.

It is important to recognize that the final product of this reaction, the chemical compound called sodium chloride, is very different from the sodium and chlorine which existed before the reaction. A chemistry student once analyzed the water of the Charles River in Boston for its chloride content. Then to interpret the results the student looked up chlorine in an encyclopedia and was horrified to learn that chlorine is a toxic and highly reactive gas. The conclusion that appeared in the final laboratory report was that an alarming pollution hazard from chlorine existed in the Charles River. In fact, the water had been collected near the mouth of the river, where ocean water mixed with the river and raised its salt content. The student had confused *chlorine*, the reactive element, with *chloride*, its unreactive ionic or salt form. It is probably just as well the student was not aware that the table salt he had at lunch was pure sodium chloride! Some months later, there was a leak in the tank of chlorine gas that was used to disinfect the water in the swimming pool at a local university. The immediate area was evacuated, and rightly so, for the elemental chlorine was so reactive it could have damaged the lungs of anyone who breathed the gas in high concentrations.

The pattern of electron transfer to form ions as sodium and chloride do is followed whenever a metal reacts with a nonmetal. All elements of group 7A easily take on an extra electron as chlorine does to form ionic compounds. Elements of group 6A, which have 6 valence electrons, form ionic compounds by taking on two extra electrons to form an octet. These two extra electrons give ions of group 6A a charge of -2. All elements of group 1A form positive ions like sodium does by losing an electron. Elements of group 2A lose two electrons to form ions with a charge of +2. Other types of ions can be formed as well; the most important ions are listed in Tables 5-1 and 5-2.

Table 5-1. Some Common Cations To Know.

Name of Ion	*Symbol of Ion*
Hydrogen ion	H^+
Lithium ion	Li^+
Sodium ion	Na^+
Potassium ion	K^+
Ammonium ion	NH_4^+
Magnesium ion	Mg^{2+}
Calcium ion	Ca^{2+}
Aluminum ion	Al^{3+}
Copper (I) ion (cuprous ion)	Cu^+
Copper (II) ion (cupric ion)	Cu^{2+}
Silver ion	Ag^+
Zinc ion	Zn^{2+}
Iron (II) ion (ferrous ion)	Fe^{2+}
Iron (III) ion (ferric ion)	Fe^{3+}

Table 5-2. Some Common Anions To Know.

Chloride ion	Cl^-
Bromide ion	Br^-
Iodide ion	I^-
Hydroxide ion	OH^-
Carbonate ion	CO_3^{2-}
Hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
Sulfate ion	SO_4^{2-}
Hydrogen sulfate ion (bisulfate ion)	HSO_4^{2-}
Phosphate ion	PO_4^{3-}
Monohydrogen phosphate ion	HPO_4^{2-}
Dihydrogen phosphate ion	$H_2PO_4^-$
Nitrate ion	NO_3^-
Nitrite ion	NO_2^-
Cyanide ion	CN^-
Acetate ion	$C_2H_3O_2^-$ (or $CH_3CO_2^-$)

Whenever ions are formed, the number of electrons gained in electron transfer to form the negative ion is always the same as the number of ions lost in electron transfer to form the positive ion.

Problem example 5-8: Show with electron dot formulas how sodium sulfide is formed from the elements sodium and sulfur.

First, we need to write electron dot formulas for the elements sodium and sulfur. We have already shown the electron dot formula for sodium with one valence electron:

Na .

Sulfur, in group 6A, has 6 valence electrons. Draw its electron dot formula by putting three pairs of dots around the S.

S

In forming the ionic compound, which is called sodium sulfide, sulfur needs to take on two electrons, but sodium has only one valence electron to give. The compound is formed by reacting two sodium atoms with a sulfur atom, thus providing sulfur with the two electrons it needs to form an octet. The resulting formula for sodium sulfide is Na_2S. To show how Na_2S is formed with electron dot formulas:

$$2Na \quad + \quad S \quad \text{-->} \quad 2Na^{+1} \quad + \quad [\ S\ ]^{-2}$$

Names and Formulas for Ionic Compounds

Since ionic compounds are such an important class of chemical compounds, it is worthwhile to be able to recognize their names. We have already seen in Chapter 2 that not only table salt, but also other ionic compounds like sodium phosphate appear on the list of ingredients on a box of macaroni and cheese mix. As we learn more about the contents of foods, cleaning agents, and other everyday products we will increasingly encounter chemical names of ionic compounds. So, what is a salt? Just like the table salt you are familiar with, most of them appear as white crystals. And all of them, like table salt, are ionic compounds, the combination of a positive ion (cation) and a negative ion (anion).

Learning to decipher both the names and the formulas of ionic compounds is a useful skill which requires learning some names of commonly encountered ions. Table 5-1 lists some common positive ions, called **cations**, and Table 5-2 lists some common negative ions, called **anions**. Many of these names will look quite familiar. As we have seen, the ion formed by sodium is called simply the sodium ion. The other elements of groups 1A and 2A follow this simple pattern as well. Chlorine, we have found, forms an anion called chloride. Other elements in groups 6A and 7A follow this pattern as well, with fluorine forming fluoride ion, for example, and oxygen forming oxide ion. Other ions are less familiar. The polyatomic ions with more than one element, like nitrate, NO_3^{-1}, and phosphate, PO_4^{-3}, are less obvious both in their formulas and their names, but they are commonly encountered in everyday substances.

A common polyatomic ion is the ammonium ion found in household ammonia, with the formula

NH_4^+ N means one nitrogen atom
H_4 means four hydrogen atoms
$^+$ means one plus charge

Notice that a one is often left out of the formula, and understood to be there.

How would you know that this ion is put together this way? Just by becoming familiar with it from Table 5-1.

Problem example 5-9: Give the name of the following compounds:

a. NH_4Cl

b. $NaNO_2$

c. $Ca(NO_3)_2$

Naming ionic compounds like these is simply a matter of knowing the names of the ions or looking them up on a list like Tables 5-1 and 5-2. In both the chemical formula and the chemical name, the cation, or positive ion, always precedes the anion, or negative ion. Notice in (b) and (c) that it is important to distinguish between similar ions like nitrite and nitrate. We will find that their properties can be quite different even though their formulas look similar. In (c), two nitrate ions, which have a charge of -1, must be used in order to balance out the charge on the calcium ion of +2. Parentheses must be used, or the subscript will look like 32. The correct answers are:

a. ammonium chloride

b. sodium nitrite

c. calcium nitrate

Writing the chemical formula if we are given the chemical name if a substance is a slightly more difficult proposition, for the chemical name does not tell us how many of each ion to use. In order to write a chemical formula, first we must know both the formulas for both the cation and the anion and the charges on these ions. If the charge on the cation is the same as the charge on the anion, the positive and negative charges cancel out, and one cation and one anion appear in the chemical formula. As we have seen with sodium sulfide in problem example 5-8, The cation charge is not always the same as the anion charge. In that case different numbers of positive and negative ions must be used in order to balance out the total positive charge with the total negative charge. Another way to understand this principle is to remember that in forming the ions from the elements, electrons lost to form the positive ion in electron transfer must be equal in number to electrons gained by the negative ion.

Sometimes it is possible to determine how many positive and negative ions are present in the formula for an ionic compound by inspection. This is always true when the charges on the cation and the anion are the same, as in the following examples.

Problem example 5-10: Write the chemical formula for sodium nitrite, a food additive found in hot dogs and other processed meats which has been implicated in some forms of cancer.

From Tables 5-1 and 5-2 we find that the formula for sodium ion is Na^{+1} and the formula for nitrite ion is NO_2^{-1}. The charges of +1 on the cation and -1 on the anion balance out. The symbol for the cation always precedes the symbol for the anion. The chemical formula is

$$NaNO_2$$

Problem example 5-11: Calcium carbonate is an ionic compound that is found in the form of marble, blackboard chalk, clam shells, and as the major ingredient in the antacid Tums. What is the chemical formula of calcium carbonate?

From Tables 5-1 and 5-2 we see that the formulas for the calcium ion and carbonate ion are Ca^{2+} and CO_3^{2-}. Since the charges on the cation and anion are the same, the chemical formula for calcium carbonate is

$$CaCO_3$$

When the charges on the cation and anion are different in an ionic compound, the chemical formula must be written subscripts that indicate the correct numbers of each kind of ion. Sometimes this can be done by inspection. A useful device which can be used to help in writing chemical formulas is the "criss-cross" method demonstrated in the examples below.

Problem example 5-12: Write the chemical formula for aluminum hydroxide, the major ingredient in the antacid Amphojel.

From Tables 5-1 and 5-2 the formulas for aluminum ion and hydroxide ion are Al^{3+} and OH^-.

$$Al^{3+} \qquad OH^{1-}$$

Three negative charges are required to balance out the triple positive charge on the aluminum ion. The formula found for aluminum hydroxide is

$$Al(OH)_3$$

Notice the use of the parentheses to indicate that three hydroxide ions and not just three hydrogens are present.

Some students like to do the “criss-cross” method and just put the charges in the superscripts as the subscripts on the other atom or group.

Problem example 5-13: Write the chemical formula for aluminum oxide, the chemical compound found in bauxite, the aluminum ore mined from deposits in the earth.

From Tables 5-1 and 5-2, the formulas for aluminum ion and oxide ion are Al^{3+} and O^{2-}. Use "criss-cross" arrows to show how the numbers which indicate the ion charges become the subscripts on the other ions.

$$Al^{3+} \; O^{2-}$$

The formula found for aluminum oxide is Al_2O_3.

Notice that the charges balance: 2 x (+3) or +6, and 3 x (-2), or -6.

Electron Sharing and the Formation of Covalent Compounds

Electron transfer to form ionic compounds is not the only way that elements can obey the octet rule and attain more desirable electron configurations. Unlike elements in groups 1A and 2A, which can easily lose one or two electrons to form octets, or groups 6A and 7A, which can easily gain one or two electrons to form octets, elements toward the center of the periodic table are more likely to share electrons than to participate in electron transfer. The form of bonding that results from shared electrons is called **covalent bonding**.

Carbon, in the center of the second horizontal row, or period, of the periodic table, is also central to our lives. The chemistry of living things and hence of our bodies is largely that of carbon. Carbon-containing molecules, often called **organic compounds**, are typically held together by the shared electrons called covalent bonds. The covalently bonded compounds formed by electron sharing differ in important ways from ionic compounds. The units which form ionic compounds are the spherical ions arranged in rows. Though they are held together by strong electrostatic forces, the ionic crystals can be cleaved, or split, easily into smaller crystals with these ions still intact. Covalent bonds involve electron sharing to form molecules. Molecules can very in size from very small to very large. They can be rather rigid or very flexible. This variety and complexity found in covalently- bonded molecules makes possible the variety and complexity of the forms of carbon-based life.

As always, let us begin with relatively simple examples in order to understand the basic principles of covalent bonding. Carbon, our most important covalently-bonded element, will be our stating point. First, examine the electron dot structure of carbon. In group 4A, it has four valence electrons. Draw them in around the C by putting one dot above, below, and to each side of the C.

C

In order to achieve an octet, it must gain four more electrons. Pulling away a total of four electrons from other elements to produce an ion with a charge of -4 is an energetically impossible task for carbon; no such ion has ever been observed. Instead, carbon shares electrons with other atoms. The element which most commonly shares electrons in covalent bonds with carbon is hydrogen, which needs only two electrons to complete its outer shell. Its electron dot structure is

H·

How can carbon combine with hydrogen so that carbon is surrounded by eight valence electrons and hydrogen has two? These conditions can be met if four hydrogen atoms surround the carbon atom, sharing electrons. Draw in the resulting electron dot structure by putting a pair of electrons between the C and each H.

H

H C H

H

The resulting molecule, with the formula CH_4, is the substance methane, or natural gas. The electron pairs shared between the carbon and hydrogen atoms constitute **covalent bonds**. Often in structure drawings a line joining the two atoms is substituted for the electron pair,
denoting the covalent bond, so draw a line between the C and each H.

H

H C H

H

The covalently bonded compounds of carbon, many of them featuring long chains of carbon atoms bonded to one another, are so numerous and important that their study constitutes an important branch of chemistry. Chapters 10 and 12 in this text will discuss their structures and reactions.

Carbon and hydrogen are not alone in their ability to share electrons to form covalent bonds. Whenever nonmetals combine with one another, covalent bonds are formed. If, for example, pure chlorine is isolated from other substances in a sealed container, it is unable to find other elements from which it can pull away electrons in order to complete its octet by electron transfer. Frustrated from reaching other sources of electrons, the chlorine atoms can complete their octets by sharing electrons as we have seen carbon and hydrogen do. Using electron dot structures to visualize this process, first we can picture two individual chlorine atoms, each with seven valence electrons:

Cl Cl

It is then easy to see how they could join together, each sharing its unpaired electron. Count the number of electrons around each chlorine atom. Each chlorine atom is now surrounded by eight electrons. The element chlorine, then, in its pure form, does not exist as free atoms, but as chlorine molecules, each with two chlorine atoms. Be warned, however, that this is not the preferred bonding situation for chlorine. As soon as the container of chlorine is opened, the chlorine atoms will seek easier sources of electrons than other electron-hungry chlorine atoms.

Hydrogen is another example of an element which, in its pure form, exists with covalently bonded atoms. Writing first the electron dot structures of two separate hydrogen atoms,

H·   ·H

it is easy to see that they can share electrons, giving each a filled outer shell of two electrons:

H : H

Electron dot structures are useful in predicting the structures of the many possible covalently bonded compounds. The rules for writing such structures are simple.

1. First, write the atomic symbols of the elements involved, as these will serve as the centers around which the electrons are placed. If several atoms are involved, there may be more than one possible arrangement of these symbols representing the atoms. In that case, it is often helpful to examine the structures of similar compounds if they are known. A good general rule is that nature loves symmetry. A symmetrical structure is more likely to be close to the true one than an unsymmetrical structure.

2. Next, determine the number of valence electrons for each element. As in the practice exercises for drawing the electron dot structures for atoms earlier in this chapter, the group numbers of the elements from the periodic table are helpful in determining the number of valence electrons.

3. Last, the electrons are arranged around the atom symbols so that each is surrounded by an octet of paired electrons. Hydrogen is an exception to the octet rule, as it needs only two electrons to complete its outer shell, and thus should have only two adjacent dots in the electron dot diagram. For this reason, hydrogen will never appear between two other elements. (Why not?)

Though the rules are simple, drawing electron dot structures is a skill. <u>Like all skills, it is best learned through practice.</u>

Problem example 5-14: Draw an electron dot diagram for the covalently bonded compound HCl.

First we write the element symbols for the compound:

H Cl

Then we determine the number of valence electrons for each element. Hydrogen, in group 1A of the periodic table, has one valence electron, and chlorine, in group 7A, has seven.

Positioning these eight electrons in pairs so that chlorine is surrounded by eight electrons and hydrogen has eight gives us the electron dot formula. Can you draw in the dots?

H Cl

Problem example 5-15: Draw an electron dot formula for the covalently bonded molecule, carbon tetrachloride, CCl_4.

A symmetrical pattern for the element symbols in a molecule such as this is:

Cl

Cl C Cl

Cl

Each chlorine atom has seven valence electrons, and the carbon atom has four. Arranging these so that each of the atoms is surrounded by an octet of paired electrons gives the electron dot formula, so draw these in now.

Cl

Cl C Cl

Cl

For a molecule with as many valence electrons as this one, it is a good idea to check the number of dots in the formula and compare it with the total number of valence electrons, in this case (4 x 7) + 4 = 32 electrons.

w

Problem example 5-16: Draw an electron dot formula for the covalently bonded molecule of water, H_2O.

Each hydrogen has one valence electron. Oxygen, being in group 6A, has six. Arrange the atom symbols symmetrically and placing the electrons so that oxygen is surrounded by eight electrons and each hydrogen has two

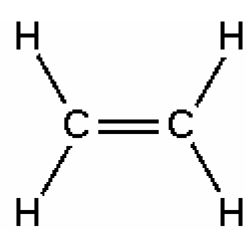
H O H

It is equally possible to place the hydrogen atoms differently and still obey the octet rule for oxygen while placing two electron dots next to each hydrogen:

H O H	O H H

As we shall see, none of these flat diagrams succeeds in describing the three-dimensional shape of the water molecule. The electron dot diagram shows successfully, however, how two hydrogen atoms join with one oxygen atom to form the water molecule.

Sometimes it is not possible to draw an electron dot structure for a molecule with octets around the atoms and one electron pair joining each atom. An electron dot structure with two pairs of electrons between two atoms, however, may be successful in creating octets. This electron dot picture corresponds to a type of covalent bond called a **double bond**. Just as the electron pair or a line between two atoms represents a covalent single bond, a double electron pair or two lines represents a double bond. An example of a molecule with a double bond is the organic molecule ethylene:



Problem example 5-17: Draw an electron dot formula for the covalent molecule carbon dioxide, CO_2. This important molecule is produced whenever fuel is burned and is an important contributor to the greenhouse effect which warms the earth's atmosphere.

A symmetrical arrangement for the carbon atom and the two oxygen atoms is

O C O

Each oxygen atom has six valence electrons, and the carbon atom has four valence electrons. Arranging this total of sixteen electrons using single bonds alone does not work. For example, the structure below has eight electrons around each oxygen atom, but only four around the carbon atom:

:O : C : O: *(incorrect)*

Moving in two of the electron pairs between the carbon and oxygen atoms is the only way to have eight electrons around each atom. The final electron dot structure is

:O::C::O:

Sometimes a **triple bond**, with three pairs of electrons between two atoms, is the only way to write a successful electron dot formula. An alternative way to represent this kind of bonding is with three connecting lines representing the triple bond. An example of an organic compound with a triple bond is acetylene, C_2 H_2:

H:C:::C:H

Problem example 5-18: Write the electron dot structure for the nitrogen molecule, N_2. This molecule comprises 78% of the earth's atmosphere.

Nitrogen is in group 5A, and each nitrogen atom has five valence electrons. Attempting to draw an electron dot structure using a single bond between the two nitrogen atoms is unsuccessful, giving each nitrogen atom only six electrons:

:N : N: *(incorrect)*

Moving one pair of electrons into the position between the two nitrogen atoms to form a double bond gives eight electrons to one atom, but only six to the other:

:N :: N: *(incorrect)*

Only when a second pair of electrons is moved between the two nitrogen atoms to form a triple bond does the electron dot structure show a full octet of electrons for each nitrogen atom:

:N ::: N: *(correct)*

The resulting triply bonded structure predicted by the electron dot formula is in agreement with the experimental properties observed in nitrogen. Held together by the strong forces of the triple bond, this molecule, the predominant one in our atmosphere, is quite unreactive.

Names and Formulas for Covalent Compounds

Predicting the ways that ionic compounds will combine, we have learned, is a matter of knowing the charges on the positive and negative ions and then seeing that they balance out. Predicting the numbers of atoms in a molecular formula is less straightforward, for sometimes the atoms can combine in more than one way to form molecules. Carbon and oxygen, for instance, react together to form carbon dioxide, CO_2. If the quantity of available oxygen is limited, however, they may react to form CO, carbon monoxide.

The names of molecular compounds feature the names of the elements involved, usually in the order in which they appear from left to right on the periodic table, and putting the ending -*ide* after the name of the second element. Prefixes are added where necessary before an element to show the number of atoms. Notice, for example, the use of *mono* to denote one oxygen in carbon monoxide and *di* to denote two oxygens in carbon dioxide. In deciphering the names of such compounds it is useful to know the meaning of these prefixes, which are derived from Greek roots (Table 5-3). Chemistry, of course, has no monopoly on the use of these prefixes, which are used quite generally, especially in reference to geometric forms. (How many sides on the Pentagon building in Washington?)

Problem example 5-19: Name the following molecular compounds:

a. CCl_4

b. SO_2

c. N_2O_4

Answers:

a. carbon tetrachloride

b. sulfur dioxide

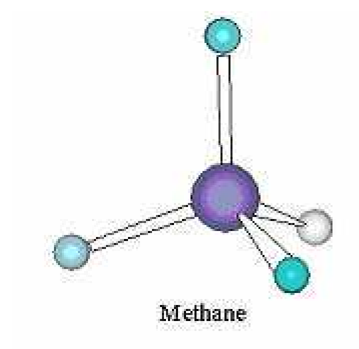
c. dinitrogen tetroxide

Notice in (c) that the *a* in *tetra* is dropped in front of *oxide*, which begins with a vowel.

The Shapes of Molecules

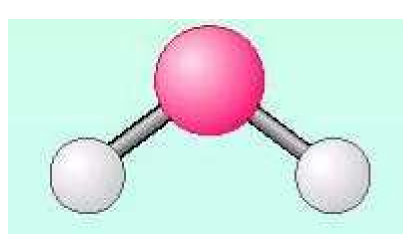
Unlike atoms and ions, which are spherical in shape, molecules can assume a variety of shapes, which depend on the way their atoms are bonded together. Methane, for example, with the formula CH_4, is typical of molecules of single-bonded carbon, having four atoms bonded to a central carbon atom. The four atoms arrange themselves in space in the symmetrical geometric structure called a **tetrahedron**. Four styrofoam balls attached by elastic bands to a central point will naturally assume this tetrahedral shape as well. The balls, like the electron pairs, are equally attracted, or pulled, to the center and naturally assume a symmetrical shape which maximizes that attractive force. In this configuration the electron pairs are also positioned as far away from each other as possible, as their negative charges are mutually repelled (Figure 5-11).

The hydrogen atoms of the methane molecule form a symmetrical tetrahedral structure.



What can we predict about the shape of the water molecule? Our electron dot formulas in problem example 5-16 could be arranged in several possible ways, all of them unsatisfactory because they were flat. The water molecule, H_2O, may seem to have little in common with the tetrahedral molecule methane. Comparing the electron dot structures of these two molecules, however, shows that they have in common the fact that eight electron pairs are arranged around a central atom. If the electrical repulsion between the like charges of electron pairs determines the shapes of electrons, then the four electron pairs around the central oxygen in water, two bonded to hydrogen and two nonbonded, ought to arrange themselves much like the four bonded electron pairs in methane. As a matter of fact, the bent angle of 104.5° that the water molecule assumes is not very different from the angles of 109.5° observed in the methane molecule. This bent shape assumed by water, we will find, is very important in determining the properties of this important molecule.

The water molecule assumes a bent shape.



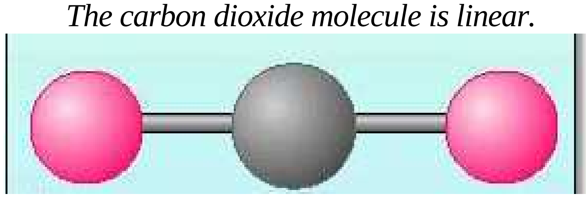
Another important molecule is

carbon dioxide, for which we determined the electron dot structure in problem example 5-17. What shape can we predict for this molecule? To predict its three-dimensional shape, first we draw the electron dot diagram and find how the electron pairs are arranged around the central atom. Unlike methane or water, carbon dioxide has two double bonds to the central atom:

:O::C::O:

The electrons in a double bond, like those in a single bond, are found between the two atoms they join. The central carbon atom, then, is surrounded by only two groups of electrons. The best way for them to get as far away from each other as possible, minimizing the repulsion of the negatively charged electrons, is for the electrons to arrange themselves on either side of the carbon atom in a straight line, much as it appears in the electron dot formula. Unlike the bent water molecule, then, the carbon dioxide molecule is linear, with the atoms in a straight line.

The carbon dioxide molecule is linear.



Problem example 5-20: Predict the shape of carbon tetrachloride, CCl_4.

The electron dot formula is the first step in determining the shape of a molecule. In problem example 5-15 we found the electron dot structure of carbon tetrachloride:

Cl

Cl C Cl

Cl

There are four bonds around the central carbon atom. The structure of carbon tetrachloride, then, is tetrahedral like that of methane.

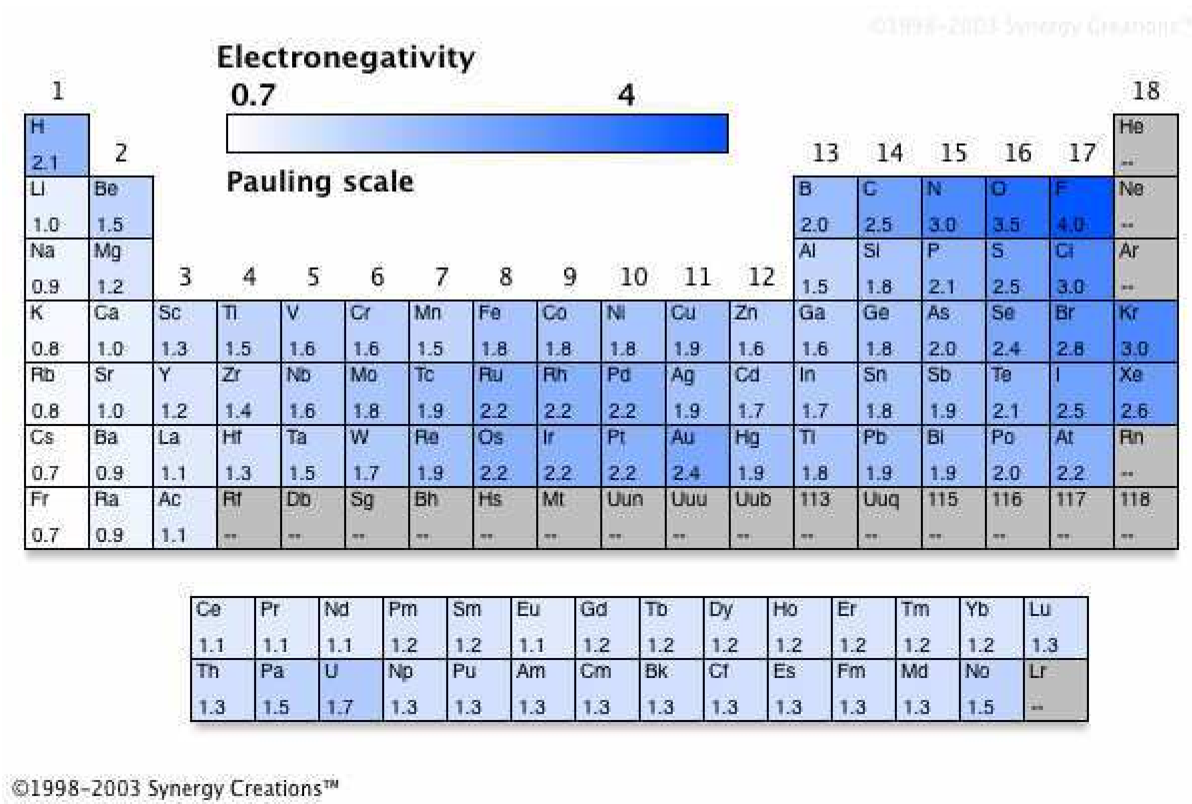
Electronegativity and Bond Polarity

The electron pairs in covalent bonds are not always shared equally between two atoms. In the hydrogen chloride molecule, for example, the bonding electrons, which move rapidly in a pattern between the two atoms, are somewhat more likely to be found in the vicinity of the chlorine atom. In the simplified picture given by the electron dot diagram, we may picture the pair of bonding electrons being pulled in closer to the chlorine atom than the hydrogen atom. The electron pair is rather like a blanket shared on a cold night between two people of unequal pulling power, with more of the shared blanket being found on one side of the bed than the other. This unequal sharing of electrons results in a partial negative charge on the end of the hydrogen chloride molecule with the higher density of electron distribution. A covalent bond in which electrons are unequally shared is called a **polar covalent bond**.

A molecule like hydrogen chloride with a partial positive charge on one end and a partial negative charge on the other end is called a **polar molecule**.

How can we predict whether a covalent bond has a polar character? When the atoms sharing electrons in a covalent molecule are identical, as in H_2, their electron-pulling powers are equal, and the bonding electrons are shared equally. Such a bond is called a **nonpolar covalent bond**. To predict the polar character of bonds between different kinds of atoms, some measure of their relative power to attract electrons is needed. In 1932 the American chemist Linus Pauling developed a relative scale of numbers called the **electronegativity scale** to describe the electron-pulling power of the elements.

The electronegativity scale.



Elements with equal electronegativities share electrons equally in a nonpolar covalent bond. The electronegativity scale can predict the degree of polarity of a covalent bond, since the greater the difference in electronegativities, the more unequally the electrons will be shared. As the diagram above shows, there are some definite trends in electronegativity in the periodic table, with the most electronegative element, fluorine, on the upper left corner of the periodic table, and the least electronegative elements, potassium, rubidium, cesium, and francium, on the lower right corner.

Problem example 5-21: Compare the polarity of the C-H bonds in CH_4 with the polarity of the C-Cl bonds in CCl_4. Which bonds have the greater polarity?

From the electronegativity table we see that the electronegativity of carbon is 2.5, while that of hydrogen is 2.2. The small electronegativity difference of 0.3 means that the bond polarity is slight. Comparing carbon's electronegativity of 2.5 with the value of 3.2 for chlorine, there is greater electronegativity difference of 0.7. The C-Cl bond is more polar than the C-H bond.

Polar and Nonpolar Molecules

As we have seen, the covalent polar bond in HCl resulted in a polar molecule, with a partially positive end and a partially negative end. A polar molecule like HCl will have very different properties from a nonpolar molecule like H_2 because its partially charged ends will be attracted by unlike charges and repelled by like charges on other molecules or ions. Knowing the degree of polarity of a molecule, then, can be very important in predicting its properties. In order to predict whether a molecule is polar, it is necessary to know things: the polarity of the bonds in the molecule, and the molecule's shape.

For example, to predict the polarity of the carbon dioxide molecule, we first look at its bonding as predicted by the electron dot formula. In the CO_2 molecule, carbon forms two double bonds with oxygen:

:O::C::O:

The electronegativity of carbon is 2.5, and that of oxygen is 3.4, a high difference in electronegativity, so these are polar covalent bonds. But the existence of polar covalent bonds does not always result in a polar molecule. The shape of the molecule is important, also. Once more the electron dot formula gives us valuable information about the carbon dioxide molecule. Only two bonds surround the central carbon atom. As we have seen, in order to minimize the repulsion between the negatively charged e lectrons in these bonds, they will spread as far away from each other as possible, on opposite sides of the carbon atom. Carbon dioxide, then, has a carbon atom with a partial positive charge in the center and two oxygen atoms with partial negative charges placed symmetrically on opposite sides of the carbon

Is the carbon dioxide polar? A simpler way to phrase this question is to ask whether it has a negative end and a positive end. Clearly, it does not, for both ends of the molecule have the same charge.

In Chapter 7 we will see how the polar properties of molecules determine the forces that pull molecules together to form solids and liquids.

Problem example 5-22: Is carbon tetrachloride a polar molecule?

We have already determined (problem example 5-21) that the bonds in carbon tetrachloride are polar covalent bonds. Determining the polarity of the molecule, however, we must look at the shape of the molecule and determine whether it has a negative end and a positive end. In problem example 5-20 we determined that carbon tetrachloride is a tetrahedral molecule. In this problem we see the usefulness of molecular models. If a three-dimensional model of carbon tetrachloride is available, it can be seen by looking at the model from all angles that it is perfectly symmetrical, with carbon in the center and the chlorine atoms projecting from the center in such a way that there is no positive or negative end of the molecule. Without an opportunity to view a three-dimensional model, the symmetry of the molecule may be difficult to visualize.

CONCEPTS TO UNDERSTAND FROM CHAPTER 5

The periodic table lists all the elements in an organized form so that elements with similar properties are grouped together.

Electron configurations in the outer energy level of an atom, called the valence electrons, determine the reactivities of an element.

The patterns of reactivity on the periodic table can be related to patterns in electron configuration.

The noble gases, or inert gases, are unreactive because they have filled outer energy levels, or valence shells.

Metals combine with nonmetals to produce ionic compounds, or salts, through the transfer of electrons to produce filled outer shells of electrons in the resulting ions.

Ionic compounds are composed of positive ions (cations) and negative ions (anions) held together by ionic bonding, the mutual attraction between unlike charges.

Nonmetals combine with one another to form molecules, held together by covalent bonding, or shared electrons.

Covalent bonds may be single bonds, consisting of one electron pair, double bonds, with two electron pairs, or triple bonds, with three electrons pairs.

Electron dot formulas for covalent compounds must have an octet of electrons around each atom.

Hydrogen and helium have only two electrons in a filled outer shell, not an octet.

The shape of a covalent molecule is determined by the number of electron pairs, bonded or nonbonded, around a central atom.

The electronegativity scale gives the relative electron-pulling power of the elements in a covalent bond.

Elements in the upper right corner of the periodic table have the highest electronegativities. Elements in the lower left corner have the lowest.

A polar covalent bond has unequally shared electrons between two atoms of different electronegativities. The greater the electronegativity difference, the more polar the bond.

A polar molecule has a partially negative end and a partially positive end.

The polarity of a molecule can be predicted by examining the polarity of its bonds and its shape.

FACTS TO LEARN FROM CHAPTER 5

The periodic table includes groups of elements with similar properties, which you should be able to identify from their position on the periodic table:
The noble gases
The halogens
The alkali metals
The alkaline earth metals
The transition metals
The lanthanide series
The actinide series

The main classes of elements in the periodic table are the **metals, the nonmetals, and the metalloids**. You should be able to place any element in one of these categories by its position on the periodic table.

You should know the names and formulas of the common anions and cations listed ion Table 5-1 and Table 5-2.

SKILLS TO ACQUIRE FROM CHAPTER 5

After finishing this chapter, you should be able to:

Give the number of electrons in each energy level for the first twenty elements.

Draw an electron dot formula for any element in groups 1A, 2A, 3A, 4A, 5A, 6A, or 7A.

Show with electron dot formulas for the groups listed above how metals combine with nonmetals to produce ionic compounds.

Draw electron dot formulas for the covalent compounds formed by elements in these groups.

Given the chemical formula for an ionic compound or a covalent compound, write the name of the compound.

Given the names or the formulas of an anion and a cation, write the chemical formula of the resulting ionic compound.

Chapter 5 Summary: Chemical Reactivity and the Periodic Table

Metals are on the left of the periodic table; nonmetals on the right.

Two types of compounds, characterized by two different kinds of bonds: ionic (charged ions) and covalent (molecules bound together by electron sharing).

Ionic compounds:
When a metal reacts with a nonmetal, an ionic compound is formed.

An ionic compound has a positive ion (cation) and a negative ion (anion).

If you are doing an electron dot formula for an ion, enclose ion formulas with electron dots in a bracket.

Covalent compounds:
The other category of compounds is covalent, in which the type of bonding is called covalent and the atoms share electrons. A pair of shared electrons forms a covalent bond. The pattern of electron sharing is predicted by drawing electron dot formulas so that each atom except H is surrounded by 8 dots. (See text for hints on making dot structures.)
Sometimes two atoms share two pairs of electrons. This is called a double bond. Sometimes they even share three pairs of electrons, forming a triple bond.

Type of compound	**Compound unit**	**Type of bonding that holds the units together**	**Formed by reaction between**
Ionic	Ions + and -	Attraction between + and - ions	Metal and nonmetal
Covalent	Molecule	Shared electrons form covalent bonds	Nonmetal and nonmetal

Polarity of bonds

Covalent nonpolar bonds- if two atoms are identical or close in electronegativity, they share electrons equally and the bond between them is nonpolar.

Covalent polar bonds – if two atoms are different in electronegativity (electon-pulling power), then they share the pair of electrons in the bond between them unequally (example HCl). One atom has more electron density than the other and hence is more negative than the other. This is called a polar bond.

Polarity of molecules

But is the molecule polar? If a molecule has a + end and a – end, the molecule is polar. For a molecule to be polar, it must have polar bonds and have a geometry that produces a + end and a – end. So, HCl has a polar bond, has a + and and a – end, and is a polar molecule.
What about methane? When we did its structure with gumdrops, we found that it was totally symmetrical, with bonds coming out from the center symmetrically. If you looked again at that structure, you would find that even if C has a different electronegativity from H, all the C-H bonds point out from the center symmetrically. You would not be able to find a + end and a – end because of this geometry.

Name ___________________________ **Date** _______________

PROBLEMS TO SOLVE USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 5

5-1. Give the names of:

a. A metal with an atomic weight greater than 50

b. Two elements which are diatomic gases

c. A highly reactive gas

d. A highly reactive metal

e. Four inert gases

f. A metalloid

g. The element with highest electro negativity

h. A halogen

i.An alkaline earth metal

j. A transition metal

k. A lanthanide

5-2. a. Which element would you expect to be totally unreactive, xenon or tellurium?

b. Which element would you expect to be a shiny substance which is a good electrical conductor, vanadium or phosphorous?

c. Which element would you expect to be a gas which is irritating and dangerous to breathe because of its high reactivity, bromine or krypton?

d. Which element would you expect to dissolve in water to produce ions, potassium chloride or carbon tetrachloride?

e. Hydrogen and helium are the lightest gases. Which would you expect to be less reactive with oxygen, and, hence, safer to use in a lighter-than-air balloon like the Hindenburg?
http://www.archive.org/details/hindenberg_explodes

5-3. Give the number of valence electrons for the following atoms:

a. Hydrogen

b. Lithium

c. Fluorine

d. Aluminum

e. Calcium

5-4. Draw electron dot formulas for the following:

a. A chlorine atom Cl

b. A chlorine molecule Cl_2

c. Chloride ion Cl^-

d. Sodium chloride NaCl

e. Hydrogen chloride HCl

5-5. Draw electron dot formulas for the following:

a. A hydrogen atom H

b. A hydrogen molecule H_2

c. Methane, CH_4

d. Carbon tetrachloride, CCl_4

e. Chloroform, $CHCl_3$

5-6. Draw electron dot formulas for the following molecular and ionic compounds. Notice that first you will have to decide whether they are molecular or ionic!

a. Carbon dioxide, CO_2

b. Silicon dioxide, SiO_2

c. Sulfur dioxide, SO_2

d. Magnesium oxide, MgO

e. Calcium oxide, CaO

5-7. Draw electron dot formulas for the following molecules:

a. H_2S

b. NH_3

c. PH_3

d. NF_3

e. HBr

5-8. Match the formula of each compound with the correct name by placing the correct letter in the blank.

_____Na_2CO_3	A. sodium carbonate (found in laundry products)
_____$NaNO_3$	B. sodium chloride (table salt)
_____K_2S	C. sodium hypochlorite (laundry bleach)
_____Na_2NO_2	D. sodium nitrate (found in deli meats)
_____Na_2SO_4	E. sodium sulfate (found in many foods)
_____Na_2SO_3	F. sodium nitrite (found in bacon and ham)
_____SiO_2	G. sodium sulfite (found in salad bars)
_____SO_2	H. potassium sulfide (used in making kraft paper and dyes)
_____NaOCl	I. silicon dioxide (used as a defoamer in foods)
_____NaCl	J. sulfur dioxide (found in raisins)

5-9 . The label of a baking soda box gives its contents as sodium bicarbonate. What is the chemical formula of this compound?

5-10. A package of bacon lists sodium nitrate as part of its contents. What is the chemical formula of this compound?

5-9. Fill in the table below by writing in each blank the ionic substance formed by the combination of the cation in the vertical row with the anion in the horizontal row. One answer, sodium chloride, has been filled in as an example.

	Cl^-	NO_3^-	O^{2-}	PO_4^{3-}
Na^+	*NaCl*			
Ca^{2+}				
Al^{3+}				

5-10. Fill in the table below by writing in each blank the ionic substance formed by the combination of the cation in the vertical row with the anion in the horizontal row.

	Br^-	S^{2-}	CO_3^{2-}	NO_2^-
NH_4^+				
Fe^{3+}				
Fe^{2+}				

5-11. Write chemical formulas for the following molecular and ionic substances:

a. Sodium sulfate

b. Ammonium chloride

c. Iron (II) chloride

d. Calcium carbonate

e. Potassium sulfate

f. Dinitrogen pentoxide

g. Phosphorous pentachloride

h. Sulfur hexafluoride

i. Sulfur trioxide

j. Manganese dioxide

5-12. A box of macaroni and cheese mix lists as an ingredient sodium phosphate. What is the chemical formula of this compound?

5-13. *Decoding chemistry*: Write chemical formulas corresponding to the following chemical names found on a container of a vitamin-mineral supplement:

a. Calcium carbonate

b. Cupric oxide

c. Potassium iodide

d. Magnesium oxide

e. Zinc oxide

f. Potassium chloride

5-14. Another bottle containing a vitamin-mineral supplement simply lists the names of the elements, for example, calcium, iron, and iodine as ingredients. Do you think these three elements are really present in their elemental form? Give your reasoning. You may need to consult references to find the properties of these elements, or you can use your knowledge of the periodic table.

15-15. *Decoding chemistry:* Write chemical names for the following substances (covalent compounds) that are known to be components of polluted air:

a. Nitrogen dioxide

b. Sulfur dioxide

c. Carbon monoxide

d. Sulfur trioxide

15-16. Is methane a polar molecule? Give your reasoning. (What is its shape? Does it have a positive end and a negative end?)

15-17. Is hydrogen bromide a polar molecule? Give your reasoning

.

15- 18. Classify the bonds in the following substances as ionic, polar covalent, or nonpolar covalent.

a. HBr

b. NaBr

c. Br_2

d. O_2

e. H_2O

19. Classify the bonds in the following substances as ionic, polar covalent, or nonpolar covalent.

a. SO_2

b. NO

c. KBr

d. N_2

e. SrO