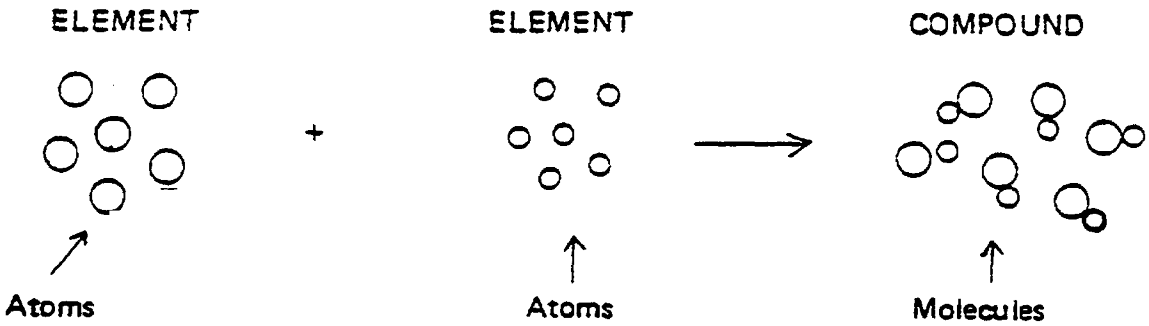


d. Classification of Mixed Matter. Matter that can be separated by physical means is called mixed matter and may be homogeneous or heterogeneous.

(1) *Homogeneous mixtures.* Mixtures which are uniform throughout are called homogeneous. An example would be a solution of sugar in water. Any small part of this solution would exhibit the same properties as any other small part, thus it would be uniform throughout the mixture.

(2) *Heterogeneous mixtures.* Mixtures which are not uniform are called heterogeneous. A good example is a mixture of water and oil. If a small sample is taken, it may not be the same as another small sample taken from elsewhere in the mixture. This is because oil and water do not mix well and give a nonuniform mixture.

4. ENERGY

There are many things in our surroundings which we know exist, yet which are not matter. They are forms of energy. Heat, light, and electricity are examples of energy. Energy may be simply defined as the ability to do work or overcome resistance.

5. ATOMIC STRUCTURE

Early scientists felt that all matter must be built from some basic unit, just as a wall may be constructed from a basic unit, the brick. In trying to find this basic unit, they separated matter by all the methods (chemical and physical) available to them until they could not separate it any further. They felt this separation must result in the building block of matter, which they called the atom (from the Greek word for indivisible). They also observed that the basic units or atoms for various elements differed in their properties, as iron was certainly different from carbon. This led them to try to find the structure of the atom. The difficulty of this problem can be seen when you consider that one cubic centimeter of gold contains as many as 59,000,000,000,000,000,000,000 atoms. The atom is so small that it defies conception. Through ingenious methods, particularly in the last 100 years, we have discovered many facts about this tiny particle which enables us to understand many of the changes that occur around us.

a. **Atomic Model**. In order for us to picture what an atom looks like, we can use a description with which most people are familiar, the solar system model. In this model, the atom is thought of as a tiny solar system in which there is a central core (like the sun) with other particles traveling in circular paths or orbits (like the planets). While more complex and exact models have been developed, this is the best approximation for general use.

b. **The Nucleus**. The central core from the solar system model is called the nucleus (which is derived from the Latin word *nucis* meaning nut or kernel). The nucleus contains two types of particles, the proton and the neutron.

(1) *The proton*. The proton is a particle which has a mass (or weight) of one amu (atomic mass unit) and a positive one (+1) electrical charge. The symbol for the proton is *p*, *p+*, or H^+.

(2) *The neutron*. The neutron has a mass of one amu (atomic mass unit) but has no electrical charge; that is, it is a neutral particle. In an atom which has more than one proton, the positive charges tend to repel each other. The neutrons serve to bind the protons so that this electrical repulsion does not cause them to fly off into space. The symbol for the neutron is *n*.

(3) Atomic number and atomic weight. Two important figures commonly used when discussing an atom are its atomic number and its atomic weight.

(a) Atomic number. The atomic number of an atom is equal to the number of protons in the nucleus of the atom. For example, a carbon atom has six protons in its nucleus; therefore, the atomic number of carbon is 6.

(b) Atomic weight. The atomic weight of an atom is equal to the number of protons in the nucleus of the atom (one amu each) plus the number of neutrons in the nucleus of the atom (one amu each). Therefore a carbon atom with six protons and six neutrons has an atomic weight of 12.

c. **The Outer Structure**. The particles which orbit the nucleus (as the planets orbit the sun) are called electrons. These particles have an electrical charge of negative one (−1), but their mass is so small that it is considered to be zero. Actually, the mass of the electron is 1/1837 of the mass of a proton, but the mass which it contributes to the atom is so small that it is not important. The symbol for the electron is e^- or $^-$.

(1) *Electron configuration*. Since we may have many electrons going around the nucleus, it might appear that there could be a collision of electrons. Collisions do not occur because the electrons are located in orbits which are different distances from the nucleus and because of the repulsion between like charges. The number of electrons and their locations are called the electron configuration. This electron configuration is different for each element.

(2) *Electron shell.* The term electron shell (or energy level) describes where electrons are located; i.e., a specific region around the nucleus. Since electrons can be forced to leave their atoms, the term energy level indicates the amount of energy required to remove the electrons from the various levels or shells. A nucleus can have seven shells, but most chemicals of pharmaceutical importance contain electrons in the first four, which are labeled the K, L, M, and N shells. The K shell is the closest to the nucleus and the N shell is the farthest from the nucleus (fig. 1). These shells contain different numbers of electrons. The maximum number each shell can hold is equal to $2n^2$, where n is the number of the shell (K = 1, L = 2, M = 3, etc.). Thus the maximum number of electrons that each of the first four shells can hold are:

$K = 2(1)^2 = 2$
$L = 2(2)^2 = 8$
$M = 2(3)^2 = 18$
$N = 2(4)^2 = 32$

Since, for example, the M shell can contain as many as 18 electrons, the possibility for collision might still appear to exist. The reason collisions do not occur is that a shell is subdivided into smaller energy levels, called subshells and orbitals, which we will not need to consider.

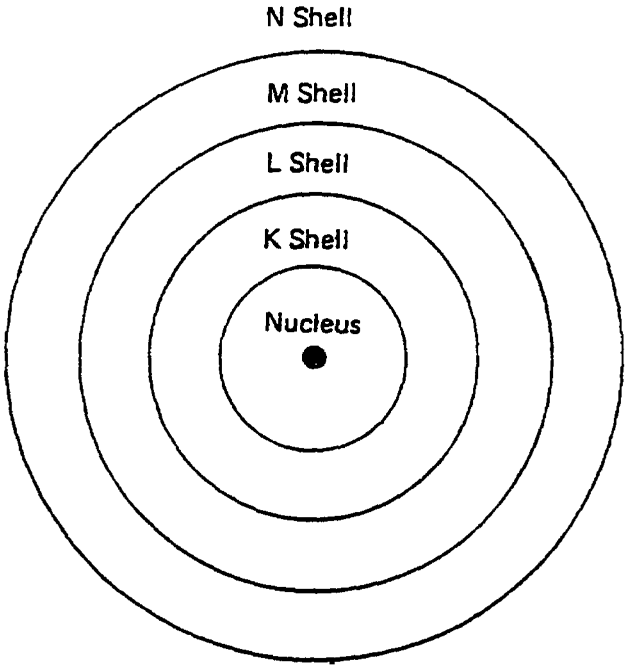


Figure 1. First four electron shells.

(3) *Number of electrons.* What determines the number of electrons an atom will contain? For an atom to exist freely in nature, it must be electrically neutral (without a charge). There are two particles in an atom which have charges, the proton which is positive, and the electron which is negative. For electrical neutrality, the sum of the charges must be zero. In other words, the number of electrons (negative charges) must equal the number of protons (positive charges).

d. **Atomic Structure of Elements.** As previously stated, each element consists of a single type of atom. Since all atoms consist of the three basic particles we have just discussed (except hydrogen, which usually has no neutrons), the only ways in which elements can differ are atomic number (the number of protons) and atomic weight, (the number of protons and neutrons). There are 106 different elements which scientists know to have a different atomic number and atomic weight. These elements have a large assortment of properties. Two elements are liquids at room temperature, eleven are gases, and all others are solids.

e. **Periodic Law.** While investigating the properties of the elements, scientists discovered an interesting fact which is now called the periodic law. This law states that the properties of the elements are periodic functions of the atomic number. As the atomic number increases, the properties of the elements repeat themselves at regular intervals.

f. **Periodic Table.** The periodic law allowed the scientists to group together the elements that had similar properties and form a systematic table of the elements. This table is the periodic table (table 2). The vertical columns are called groups, and the horizontal rows are called periods. This table contains a lot of information which we will not generally use; however, we are concerned with the basic information we can obtain about the elements. Figure 2 includes four blocks for elements from the periodic table showing the information which can be obtained from it. You should note that the number of neutrons is not given in the periodic table. This can be determined by subtracting the atomic number from the atomic weight.

g. **Isotopes.** All the atoms of a particular element are not identical. Slight variations in the number of neutrons are found to occur naturally and can also be produced in reactors. Atoms that have the same number of protons but a different number of neutrons (same atomic number, but different atomic weights) are called isotopes. Sometimes, isotopes are referred to by their mass numbers, H^2, H^3, U^{239}, etc. All of the isotopes of a particular element have identical electronic configurations; and since electronic configurations determine chemical properties, isotopes of an element exhibit identical chemical behavior. Induced nuclear reactions can produce both stable and radioactive nuclei. If the nucleus of the atom is unbalanced during the bombardment reaction, the atom is called a radioisotope. Radioisotopes, such as cobalt^{60} for treatment of cancer and iodine^{131} for diagnosing of thyroid tumors, are of vital importance in the medical field. The presence of isotopes helps to explain why many atomic weights in the periodic table are not whole numbers since all of the isotopes must be considered when computing the average atomic weight of the element.

8. VALENCE AND CHEMICAL BONDING

We have now developed the concept that matter was built from a basic unit called the atom and discussed the nature of the atom. We know, however, that very little matter exists as free elements. Most of the things around us are combinations of elements. Logically, the next step is to consider how things combine.

1 H 1.00797	2 He 4.0026

LIGHT METALS — TRANSITION HEAVY METALS — NON METALS — INERT GASES

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
																	2 He 4.0026
3 Li 6.939	4 Be 9.0122											5 B 10.811	6 C 12.01115	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.183
11 Na 22.9898	12 Mg 24.312											13 Al 26.9815	14 Si 28.086	15 P 30.9738	16 S 32.064	17 Cl 35.453	18 A 39.948
19 K 39.102	20 Ca 40.08	21 Sc 44.956	22 Ti 47.90	23 V 50.942	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.71	29 Cu 63.54	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.909	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.905	40 Zr 91.22	41 Cb 92.906	42 Mo 95.94	43 Tc 99	44 Ru 101.07	45 Rh 102.905	46 Pd 106.4	47 Ag 107.870	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.904	54 Xe 131.30
55 Cs 132.905	56 Ba 137.34	57-71 Lanthanide series	72 Hf 178.49	73 Ta 180.948	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 196.967	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.980	84 Po 210	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226	89- Actinide series															

Lanthanide series	57 La 138.91	58 Ce 140.12	59 Pr 140.907	60 Nd 144.24	61 Pm 147	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.924	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.97
Actinide series	89 Ac 227	90 Th 232.038	91 Pa 231	92 U 238.03	93 Np 237	94 Pu 242	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 249	99 Es 254	100 Fm 253	101 Mv 256	102 No 254	103 Lw 257

Table 2. Periodic table of the elements.

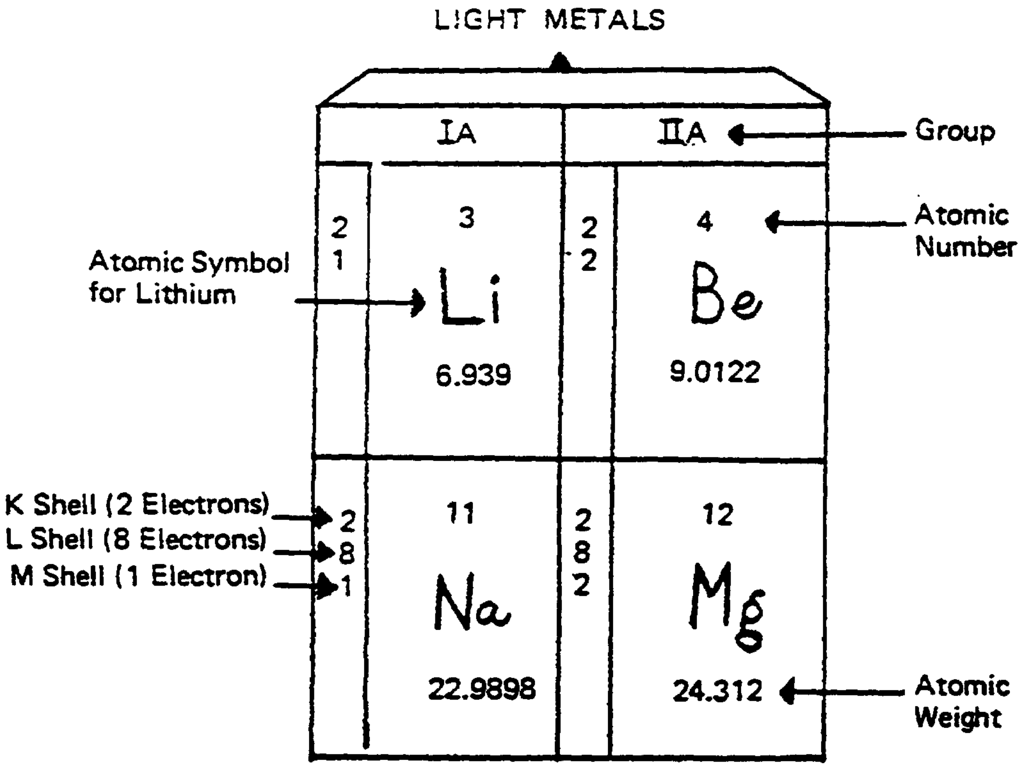


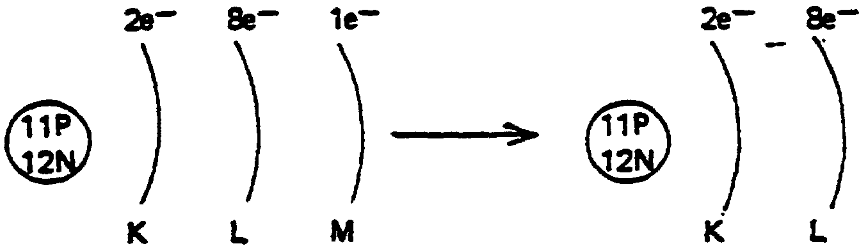
Figure 2. Identifying the components of the periodic table.

a. **Valence.** The valence of an element can be defined as a measure of its combining power or the number of electrons an atom must gain, lose, or share to have a full or stable outer electron shell. The reason atoms combine is contained in this definition. There are certain electron configurations in nature that are unusually stable (unreactive). The elements that have these configurations are in Group VIII A of the periodic table. They are sometimes referred to as the inert or noble gases because they are found in very few combinations in nature. Other elements, by gaining, losing, or sharing electrons, can try to make their outer electron shells resemble the shells of the noble gases and hence become very stable. We can see how this works by considering the two simplest elements, hydrogen and helium. Hydrogen has one electron in the K shell since it has only one proton; therefore, hydrogen is a very reactive element, occurring naturally in many compounds. Helium, a noble gas, has two electrons in the K shell since it has two protons. Helium is very unreactive. Note that helium, by having two electrons, has a completed outer shell, since the K shell can hold only two electrons. Hydrogen would like to be as stable as helium and could be if it could gain or share one more electron to give it a completed outer shell. Hydrogen seeks this electron in nature by combining with other elements.

b. Octet Rule. If you examine the noble gases, you will see that not all have a completed (full) electron shell like helium. Except for helium, the noble gases have eight electrons in their outer shell, yet they are still very stable. Chemists have observed that other elements sometimes gain, lose, or share electrons in order to have eight electrons in their outer shell. This observation led to the development of the octet rule, which states that outer electron shells prefer to have eight electrons even though the shell may not be full. (Octet means a group of eight.) Below are some examples of the electron configurations for various elements which indicate to us how many electrons they can gain, lose, or share to fit the octet rule or have a completed outer shell.

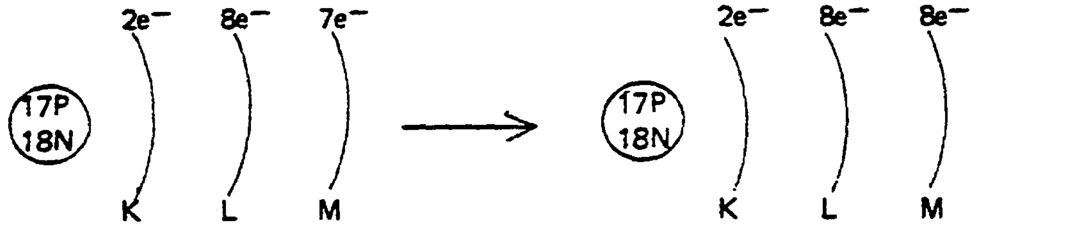
ELEMENT	ATOMIC NUMBER	SHELLS K(2)	L(8)	M(18)	N(32)
H	1	1			
He	2	2			
Li	3	2	1		
Be	4	2	2		
Na	11	2	8	1	
K	19	2	8	8	1

c. **Positive Valence.** An atom that must give up electrons to become stable will have more protons than electrons in its stable configuration and will not be electrically neutral. It will be positive since there are more positive charges than negative charges. This is indicated by a + sign. The number of electrons it gives up is the numerical value of its valence. Consider, for example, the element sodium which has 11 protons and 11 electrons in its free state. It has one electron in the M shell which it loses easily to become stable. After it loses the electron, it will have a positive one charge and its valence will be +1.



Sodium (free state) Sodium (+1 valence state)

d. **Negative Valence.** An atom that must gain electrons to become stable will have more electrons than protons in its stable configuration and will not be electrically neutral. It will be negative since there are more negative than positive charges. This is indicated by a − sign. The number of electrons it gains is the numerical value of its valence. Consider, for example, the element chlorine which has 17 protons and 17 electrons in its free state. It is one electron short of fitting the octet rule in the M shell as that shell contains 7 electrons. After it gains the electron, it will have a negative one charge and its valence will be −1.



Chlorine (free state) Chlorine (−1 valence state)

e. **Important Symbols and Valences.** Since it is very tedious to continually write complete names for elements, chemists developed the symbols for the elements which you observed on the periodic table. It will not be necessary for you to know all the symbols for your work but a number of them appear frequently enough that they should be memorized. Table 3 lists important elements with their symbols and valences. These should be committed to memory. (Please note that most, but not all, valences conform either to the completed shell or octet rules.)

f. **Ions.** Any atom that gains or loses electrons becomes charged (electrical charge) and is called an ion. An *ion* can be defined as any charged atom or group of atoms. If the ion is positively charged, it is called a *cation*. If it is negatively charged, it is called an *anion*. A group of atoms that has a charge and goes through a reaction unchanged is called a *radical*. Whenever we write the symbol for an element and wish to indicate it is an ion, we write the charge as a superscript to the symbol, e.g., Cl^{-1} or Na^{+1}.

g. **Chemical Bonding.** When elements combine to form chemical compounds, the electrons in the outer shell may be transferred from one atom to another or there may be a mutual sharing of the electrons. In either case, a chemical bond is produced. This means the two atoms do not travel or react independently of one another but are held together by the exchange or sharing of the electrons. Both atoms involved in the reaction attain a completed outer orbit, and stability results. There are three types of chemical bonds—electrovalent, covalent, and coordinate covalent.

(1) *Electrovalent (ionic) bonding.* A transfer of one electron from one atom to another resulting in opposite charges on the two atoms which holds them together by electrostatic (opposite charges attract) attraction is called an electrovalent or ionic bond. A good example of this is the bond formed between a Na (sodium) and a Cl (chlorine) atom.

1e⁻ in M shell 7e⁻ in M shell

Sodium has a complete outer shell and a charge of +1. Chlorine has met the octet rule in the M shell and has a charge of −1.

NAME	SYMBOL	VALENCE
Acetate	$C_2H_3O_2$	-1
Aluminum	Al	+3
Ammonium	NH_4	+1
Antimony	Sb	-3, $\underline{+3}$, +5
Arsenic	As	-3, $\underline{+3}$, +5
Barium	Ba	+2
Bicarbonate	HCO_3	-1
Bismuth	Bi	$\underline{+3}$, +5
Bromine	Br	$\underline{-1}$, +1, +3, +5, +7
Calcium	Ca	+2
Carbon	C	+2, $\underline{+4}$, $\underline{-4}$
Carbonate	CO_3	-2
Chlorine	Cl	$\underline{-1}$, +1, +3, +5, +7
Copper	Cu	$\underline{+1}$, $\underline{+2}$,
Fluorine	F	-1
Gold	Au	$\underline{+1}$, $\underline{+3}$
Hydrogen	H	+1
Hydroxide (Hydroxyl)	OH	-1
Iodine	I	$\underline{-1}$, +1, +3, +5, +7
Iron	Fe	$\underline{+2}$, $\underline{+3}$
Lead	Pb	$\underline{+2}$, $\underline{+4}$
Lithium	Li	+1
Magnesium	Mg	+2
Manganese	Mn	$\underline{+2}$, +3, +4, +6, +7
Mercury	Hg	$\underline{+1}$, $\underline{+2}$
Nickel	Ni	$\underline{+2}$, $\underline{+3}$
Nitrate	NO_3	-1
Nitrogen	N	+1, $\underline{-3}$, $\underline{+3}$, $\underline{+5}$
Oxygen	O	-2
Permanganate	MnO_4	-1
Phosphate	PO_4	-3
Phosphorus	P	-3, $\underline{+3}$, $\underline{+5}$
Potassium	K	+1
Silver	Ag	+1
Sodium	Na	+1
Strontium	Sr	+2
Sulfate	SO_4	-2
Sulfur	S	$\underline{-2}$, +2, +4. +6
Zinc	Zn	+2

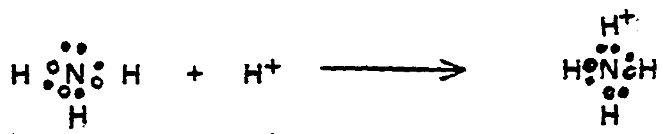
NOTE: The most common valences are underlined where there may be more than one valence.

Table 3. Valences.

(2) *Covalent bond.* If two atoms each donate an electron which is shared with the other atom, the bond is a covalent bond. An example of this is the bond between two H (hydrogen) atoms. Double and triple covalent bonds are also possible.

H• + °H ⟶	H°•H
Both atoms have $1e^-$ in the K shell.	By sharing their electrons each hydrogen has $2e^-$ in the K shell and both are stable because of the completed shell.

(3) *Coordinate covalent bond.* If one atom donates two electrons for sharing with another atom (which donates no electrons), it is called a coordinate covalent bond. An example of this type of bond is the bond between N (nitrogen) in ammonia and a hydrogen ion (proton).



Ammonia —N has a complete outer shell	H^+ has an empty K shell	By sharing the two e^-, H^+ has a completed K shell and N still has a completed outer shell.

7. FORMULA WRITING

a. **Formulas.** Formulas are combinations of symbols that represent a compound. A formula indicates which elements are involved and the number of atoms of each element contained in the compound. In writing formulas, we use subscripts, coefficients, and parentheses in addition to the symbols of the elements. *Subscripts* indicate the number of atoms of an element, as in H_2 where 2 is the subscript meaning two hydrogen atoms. If there is no subscript with a symbol, it is assumed there is only one atom of that element. *Coefficients*, numbers in front of the formula, indicate the number of molecules of compound, as in 4HCl where 4 is the coefficient indicating 4 molecules of HCl. *Parentheses* are used to separate a radical from the rest of the formula when it would be confusing not to do so. In HNO_3 it is not necessary to include parentheses for the NO_3^- radical since there is little chance for confusion. However, we use parentheses for the same radical if it appears in a compound such as $Hg(NO_3)_2$ where the 2 indicates that we have two NO_3^- radicals.

b. **Steps in Formula Writing.** In writing formulas for compounds, there are four steps which should be followed.

(1) Determine the symbols for the elements in a compound.

(2) Determine the valence of each of the atoms or radicals.

(3) Write the positive element's symbol first, followed by that of the negative element.

(4) Make the compound electrically neutral by placing subscripts.

c. **Example.** Write the formula for calcium chloride.

(1) Calcium = Ca, chloride = Cl.

(2) Ca valence is +2, Cl valence is −1.

(3) $Ca^{+2}Cl^{-1}$. If we add the charges, we find that this compound is not neutral (+2 − 1 = +1). Therefore, we must proceed to step (4).

(4) To have two negative charges to balance the two positive charges, we must have two Cl^{-1} ions (−1 x 2 = −2). Thus, the formula would be $CaCl_2$.

d. **Rule of Crossing Valences.** A convenient rule for determining what subscripts are necessary in writing formulas is the rule of crossing valences. This rule states that one can take the valence of the element at the left and make it the subscript of the element at the right, and in like manner take the valence of the element at the right and make it the subscript of the element at the left. For example:

$Fe^{+3}\ SO_4^{-2}$ becomes $Fe_2(SO_4)_3$