Chapter 5
How Are the Chemical Elements Organized?

As more and more chemical elements were discovered, scientists wondered whether there was any way of organizing them in categories. It was in the early 20th century that the answer was to be found in the number and arrangement of the electrons.

5.1 Background

There had been so many elements discovered by the 1820s that scientists started looking for some way of organizing them. A German chemist by the name of Johann Döbereiner (1780-1849) (Figure 5.1) noticed that there were sometimes similarities amongst groups of three elements. For example, the group of three elements: chlorine, bromine, and iodine, had similarities. Another group of three with similar properties were lithium, sodium, and potassium. Such groups of three which resembled each other were called Döbereiner's Triads.

Figure 5.1 Similarities in groups of three elements was noticed by Johann Döbereiner

Some years later, a British chemist, John Newlands (1837-1898), realized that, if the elements were placed in order of increasing atomic mass, each element resembled the eighth one after it (Figure 5.2). For example, the two elements lithium, and sodium, the eighth element after lithium, share many similarities in their chemistry. Likewise, the other pairs, such as carbon and silicon resemble each other in their chemistry. This discovery was called Newland's Law of Octaves.

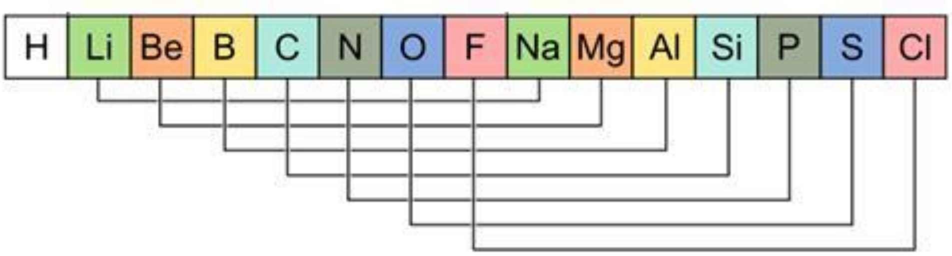


Figure 5.2 The pattern of repeating properties of elements recognized by John Newlands.

It was a Russian chemist, Dmitri Mendeléev (1834-1907), who realized that Newland's Law was not totally valid. First, he found that, to fit the elements then known, he had to put some of the elements in an eighth column. Second, in order to fit the pattern of repeating properties, he needed to leave some gaps in the table. One of several early versions of the Mendeléev Periodic Table is shown below (Figure 5.3).

Group / Period	I	II	III	IV	V	VI	VII	VIII
1	H=1							
2	Li=7	Be=9.4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27.3	Si=28	P=31	S=32	Cl=35.5	
4	K=39	Ca=40	?=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56,Co=59 Ni=59
5	Cu=63	Zn=65	?=68	?=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	?=100	Ru=104,Rh=104 Pd=106
7	Ag=108	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140				
9								
10			?Er=178	?La=180	Ta=182	W=184		Os=195,Ir=197 Pt=198
11	Au=199	Hg=200	Tl=204	Pb=207	Bi=208			
12				Th=231		U=240		

Figure 5.3 One of the early versions of Mendeléev's Periodic Table.

It was the realization that many elements remained to be discovered that made his Table so much more significant. Up until then, geochemists roamed the planet looking for unusual minerals which might contain previously unknown elements. It was all very much guesswork. The Periodic Table changed this random hunt to a more focused one. For example, at the time, the element with the next higher mass to calcium was titanium. However, the next "box" to calcium should have been an element similar to aluminum. Titanium was similar to silicon, the next "box" along. So Mendeléev proposed that there was a then-unknown element with properties similar to that of aluminum. Chemists could then look for ores that contained aluminum in the hope that the unknown element would have similar properties. Using the Mendeléev Periodic Table, chemists were able to discover elements to fill seversl empty spaces.

However, Mendeléev's later versions of the Periodic Table, when atomic masses were known more precisely, had several flaws:

1. The Group VIII elements contained three elements per box.

   For example, to keep the repeating eight columns, iron, cobalt, and nickel, had to be fitted in one 'box.'

2. Some element orders had to be reversed to match properties.

   For example, tellurium (Te) has an average atomic mass of 127.6 u while iodine has an average atomic mass of 126.9 u, so I should precede Te. However, tellurium is an element like selenium (Se) while iodine is like bromine (Br). So to match up the chemical properties, Mendeléev chose to invert the order.

3. When the 14 ‘rare earth’ elements were subsequently discovered, the only way to fit them in the Table was to squish them all into one box.
   The discovery of these elements proved totally perplexing as they were all chemically similar and there was no space for them.
4. The Table could not anticipate a whole missing column.
   When the noble gases began to be discovered, it was realized a whole new column had to be added.

It was the British physicist, Henry Moseley, who, in 1913, realized that each element had a unique integer associated with it – what we now know to be the number of protons. From then on there was no uncertainty as to the order of the chemical elements or the possibility of missing elements or columns of elements.

5.2 The Modern Periodic Table

Since Moseley’s time, chemists have agreed on a standard structure of the Periodic Table. The long form of the Periodic Table is shown in Figure 5.4. The vertical columns are called *groups* while each row is called *periods*. The elements in the first two and last six groups are called the *main group* elements; the two rows from La to Yb and Ac to No are called the *lanthanoids* and *actinoids* respectively. The rectangle of elements across from Sc to Zn and down to Lr and across to Uub are called the *transition metals*.

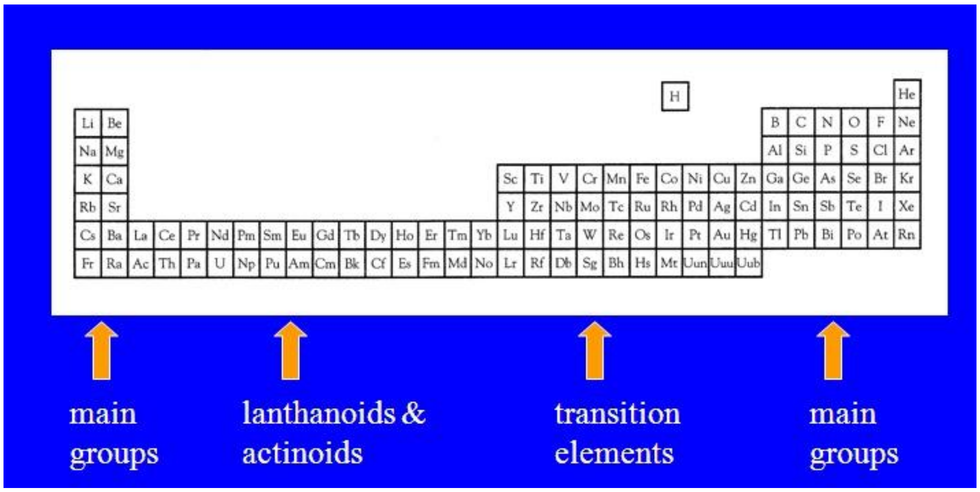


Figure 5.4 The long form of the Periodic Table with the different segments identified.

THE SHORT FORM OF THE PERIODIC TABLE

The long form of the Periodic Table shows the elements in order, but it is an awkward shape. Also, the lanthanoid and actinoid elements are rarely discussed. So a more compact form of the Periodic Table, the short form, was developed in which those two sets of elements are placed underneath and the gap between the left-hand pair of main groups and the transition elements is closed up (Figure 5.5). It is the short form that is used for most purposes.

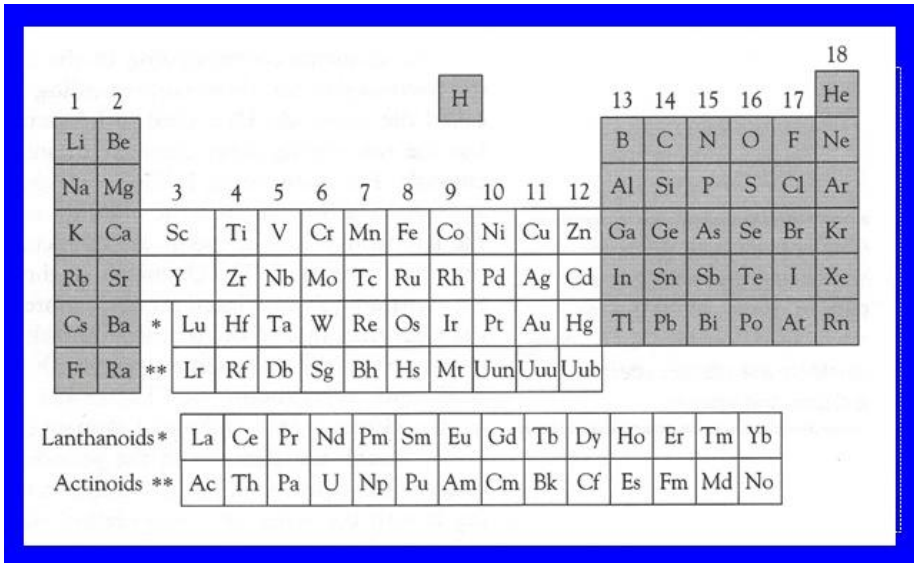


Figure 5.5 The short form of the Periodic Table

In addition to giving names to blocks of elements, some of the groups are given individual names. These are shown in Figure 5.6.

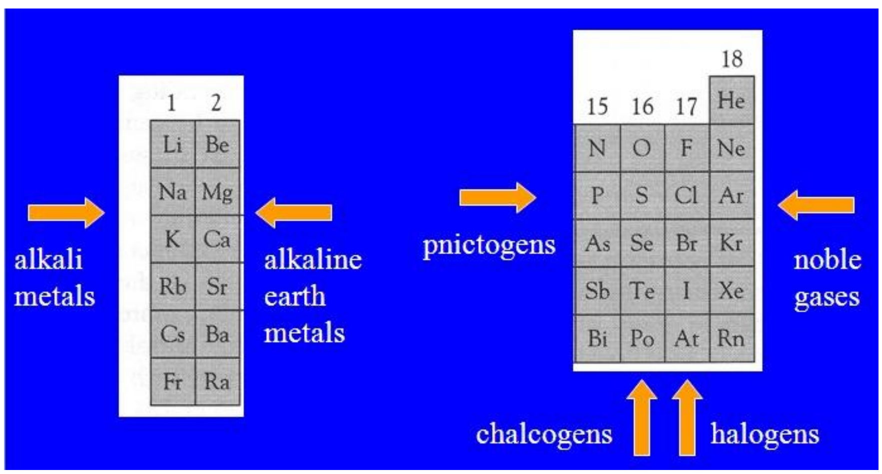


Figure 5.6 The specific names given to six of the groups in the Periodic Table.

THE DISCOVERY OF NEW ELEMENTS

Chemists are sure that there are no more stable elements to be discovered. However, groups of nuclear scientists are constantly trying to synthesize new chemical elements. These elements have to be synthesized atom-by-atom. As all of these new elements are radioactive, there is only a short time between the instant of synthesis and the time of decay into other elements.

When there is convincing evidence that atoms of a new element have been synthesized, then a tentative name and symbol based upon a combination of Latin and Greek prefixes is used (Figure 5.7) followed by the ending "-ium." The temporary chemical symbol is a three-letter combination of these components. For example, if atoms of the element with 125 protons was synthesized, it would be provisionally named unbipentium and it would have the provisional symbol of Ubp.

0	nil	5	pent
1	un	6	hex
2	bi	7	sept
3	tri	8	oct
4	quad	9	enn

Figure 5.7 The hybrid Latin-Greek prefixes used for the preliminary naming of new chemical elements.

It is the group of scientists whose discovery is first confirmed who choose the definitive name, which then has to be approved by an international board of scientists.

5.3 Classifications of the Chemical Elements

There are several ways of categorizing the chemical elements. One way is to divide them into metals (shiny and conduct electricity) and non-metals (all others). When chemists attempt to do so, there are a few elements with sort-of-in-between properties. These elements, called semimetals, are often shiny but do not conduct electricity well. The three categories are shown in Figure 5.7. Notice that all the nonmetals are clustered in the upper right quadrant of the Periodic Table with the semimetals along the boundary between the metals and nonmetals. Most of the chemical elements, then, are metals. However, in this text we will focus on the main group elements, the majority of which are nonmetals or semimetals.

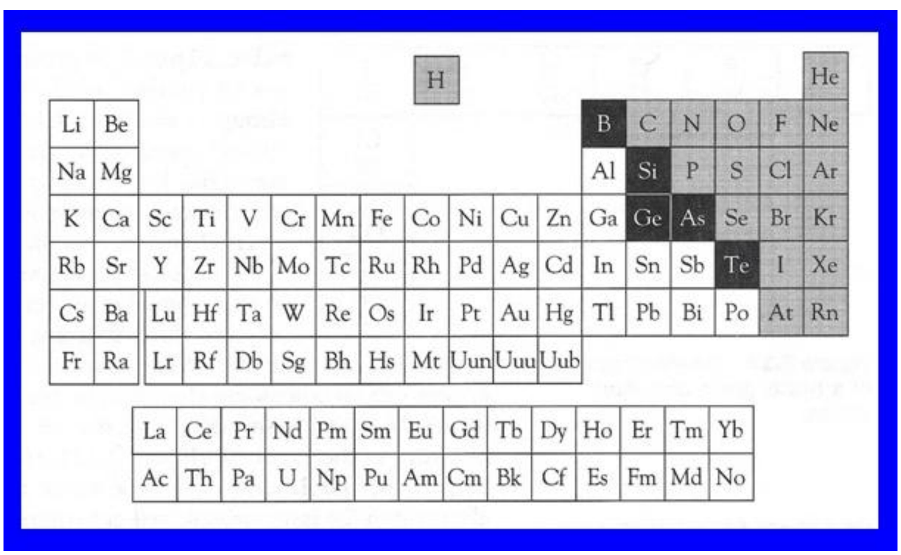


Figure 5.7 Classification of the elements as metals (white), semimetals (black), and nonmetals (grey)

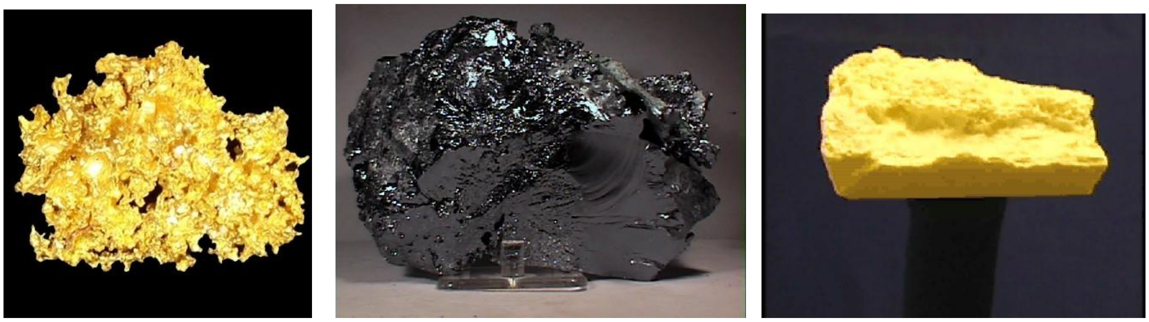


Figure 5.8 Left – gold (a metal); centre – silicon (a semimetal); right – sulfur (a nonmetal)

Another way to classify elements is by their phase at room temperature: gas, liquid, or solid. This categorization is shown in Figure 5.9. Notice that the gaseous elements are clustered in the upper-right quadrant of the Periodic Table and that they are all nonmetals. There are only two liquid elements at room temperature: mercury (a metal) and bromine (a nonmetal).

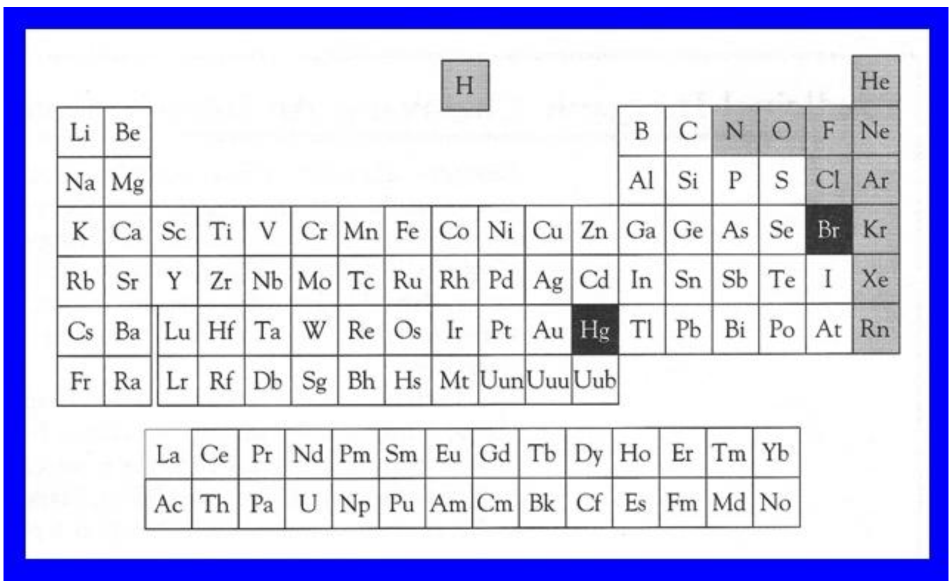


Figure 5.9 Classification of the elements as solids (white), liquids (black), and gases (grey).

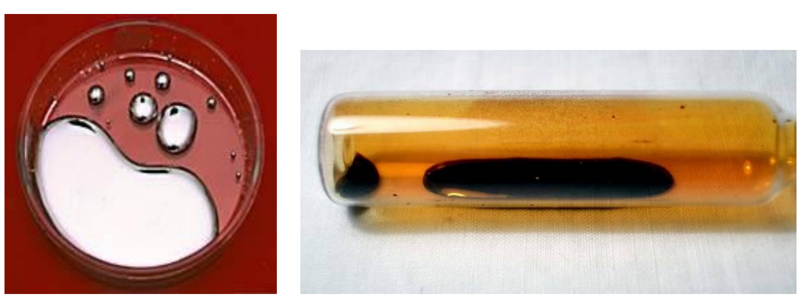


Figure 5.10 The two elements which are liquids at room temperature: left – silvery mercury (in a glass dish); right – dark red-brown bromine (in a glass ampoule)

5.4 Why are Elements Arranged in This Way?

Up to now, you have seen that, based on chemical similarities, the elements can be organized in a repeating pattern of Periods and Groups. But what is the underlying cause of this format? The answer is found with the electrons.

When electrons were initially discovered, it was thought that they could be anywhere around the nucleus. However, a simple experiment was to prove that was not true. If a solid is heated until it glows white-hot, and light is passed through a prism, then all the colours of the rainbow are seen (Figure 5.11).

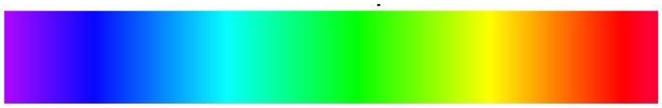


Figure 5.11 The continuous spectrum emitted by a white-hot solid.

If the experiment is repeated with heating a gas by passing a high voltage electric current through it, it is not a continuous spectrum but a line spectrum, one in which only a few very specific wavelengths of light are seen (Figure 5.12). Each chemical element gives its own unique line spectrum.

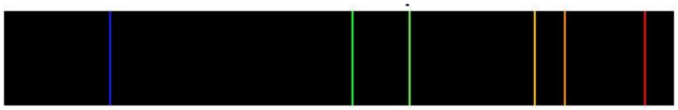


Figure 5.12 The line spectrum emitted by hydrogen gas when heated using high-voltage electricity.

So, how is it possible to explain these line spectra? It was the Danish scientist, Niels Bohr, who, in 1913, contended the electrons were in layers around the nucleus. Each layer has a specific energy, the closer to the nucleus, the lower in energy. Light is emitted when electrons drop from one energy level to one closer to the nucleus. These energy levels are given the symbol, n. The closest energy level to the nucleus was assigned as $n = 1$, and theoretically, there are an infinite number of energy levels. The levels become closer and closer together as the distance from the nucleus increases.

When an electron drops from any level to the level closest to the nucleus, $n = 1$, the light released is of such high energy, it is in the ultraviolet part of the spectrum. When an electron drops to the $n = 2$ level, the light emitted is in the visible region. When an electron drops to any of the higher levels, the light emitted is in the infrared part of the spectrum. The possible transitions for energy levels from $n = 1$ to $n = 6$ is shown in Figure 5.13.

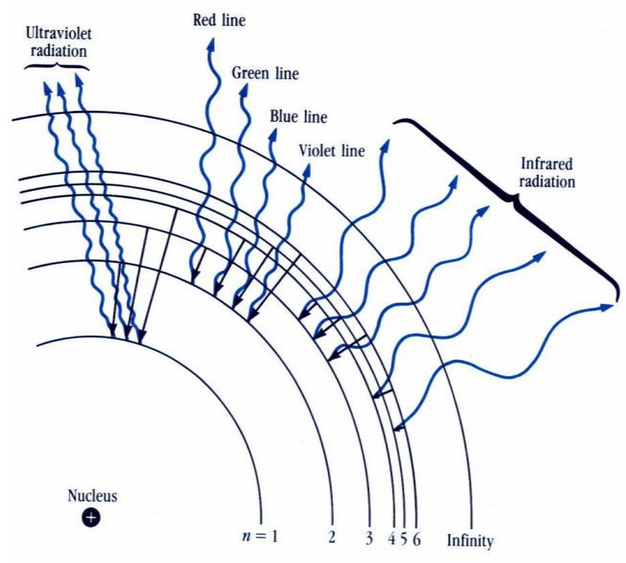


Figure 5.13 Some of the possible energy level transitions for an atom, showing the levels of n = 1 to n = 6.

In atoms, the electrons are normally in the lowest possible energy level (called the ***ground state***) and there is a maximum number for each level. The maximum number of electrons permissible in each level is shown in Figure 5.14.

Level	Number of electrons
1	2
2	8
3	8
4	18
5	18
6	32
7	32

Figure 5.14 Electron energy levels and the corresponding numbers of electrons permissible in each level.

5.5 Energy Levels and the Periodic Table

Having established the maximum number of electrons in each energy level, the next step is to see how the electron layers match with the organization of the Periodic Table. As the hydrogen atom has one proton, so in the neutral atom, it must have one electron. This electron will go in the n = 1 level. With helium, there are two protons, so there are two electrons, and both of these can go in the n = 1 level. The electron structures of these two elements are shown in Figure 5.15.

Element	protons	electrons $n = 1$
Hydrogen	1	1
Helium	2	2

Figure 5.15 The filling of the n = 1 energy level

The next element is lithium, and this has three protons, and so three electrons. Only two of the electrons can fit in the n = 1 level, so the third electron must go in the n = 2 level. With each successive element, the nucleus contains one more proton and so has one additional electron filling the n = 2 level. The filling of the n = 2 level is complete with the element neon as can be seen from Figure 5.16.

Element	protons	electrons $n = 1$	electrons $n = 2$
Lithium	3	2	1
Beryllium	4	2	2
Boron	5	2	3
Carbon	6	2	4
Nitrogen	7	2	5
Oxygen	8	2	6
Fluorine	9	2	7
Neon	10	2	8

Figure 5.16 The filling of the n = 2 energy level

The n = 3 level can be filled up similarly and this comprises the elements from sodium to argon. The filling of the electron levels then enables chemists to make sense of how the elements fall into patterns in the Periodic Table. The locations of the first 18 elements are shown in Figure 5.17.

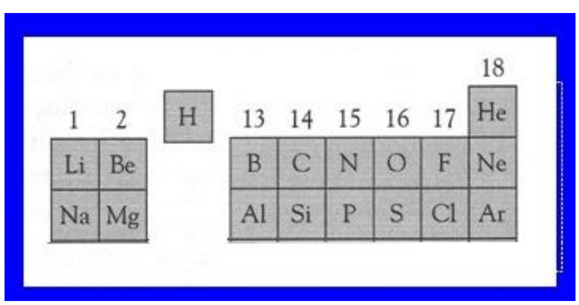


Figure 5.17 The filling of the first three electron levels corresponds to the first three periods of the Periodic Table.

As can be seen, the first Period consists of the two elements where the *n* = 1 level is filled, then the eight elements corresponding to the filling of the *n* = 2 level, then the eight elements filling the *n* = 3 level. As mentioned before, the elements are organized in Groups (columns) in which the elements have similar properties. Hydrogen is left on its own as its properties are totally unique. The numbering of the groups is also a helpful guide in immediately identifying the number of electrons in the outermost level containing electrons. Thus magnesium, like beryllium, is in Group 2 and has two electrons in its outer level (two in *n* =1; eight in *n* = 2; and two in *n* = 3; making the total of 12). Sulfur, like oxygen is in Group 16, and has six electrons (the second digit of the group number) in its outer level (two in *n* =1; eight in *n* = 2; and six in *n* = 3; making the total of 16).

The next two periods after argon contain a maximum of 18 electrons as was shown in Figure 5.14. The additional ten electrons reside in a separate sub-level which starts to fill after two electrons fill the 'normal' *n* = 4 level. When the next ten electrons are 'sandwiched' in between for the first row of the transition metals, then the final six electrons of the 'regular' *n* = 4 level fill up. The same happens with the *n* = 5 level. These first five periods of the Periodic Table are shown in Figure 5.18. The reason why these intervening electrons occur is better left to a more advanced course in chemistry.

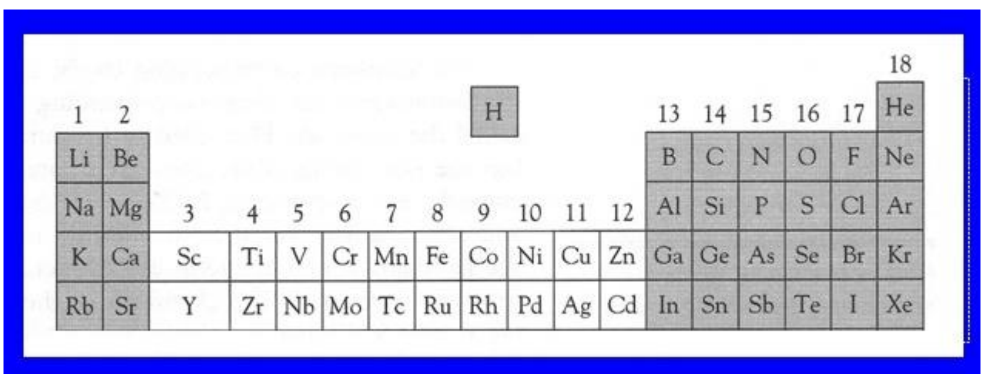


Figure 5.18 The filling of the first five electron levels corresponds to the first five periods of the Periodic Table.

After xenon, element 54, the next level contains 32 electrons. Just as the previous two levels had 10 electrons 'sandwiched' in between the 'regular' electrons, so for the next two periods, there are an additional 14 electrons that fill another sub-level, before the sub-level of the transition metals fill, and only then does the final six electrons of the 'regular' level fill. This filling order then accounts for the arrangement of the Periodic Table (Figure 5.19).

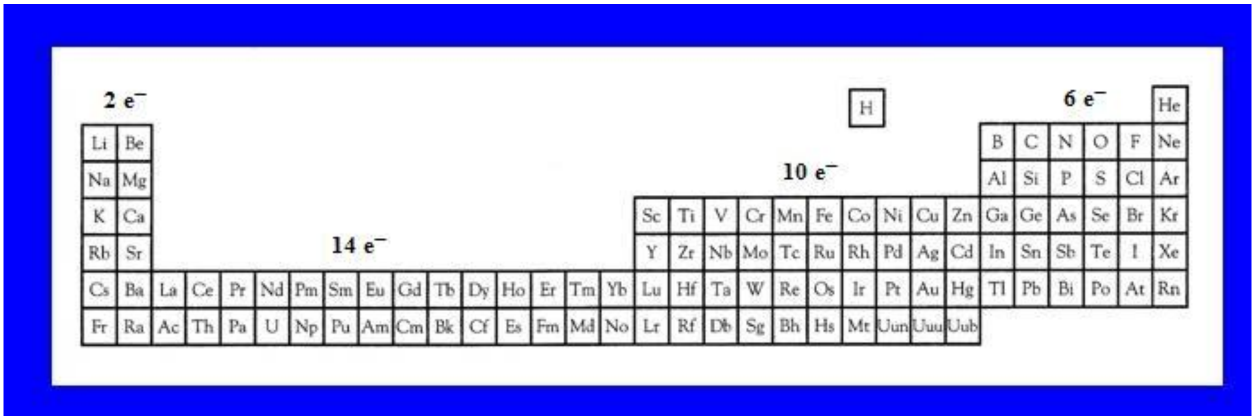


Figure 5.19 The long form of the Periodic Table showing the electron filling pattern.

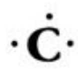
5.6 Writing Electron-Dot Symbols

It is the outermost partially-filled level of electrons that determines the chemistry of an element. As they are so important, chemists have developed a symbolic way of depicting the information. These are called *electron-dot symbols*.

In this method, the element symbol itself is used to depict the nucleus plus the filled inner levels of electrons. Dots are then used to display each of the outer electrons (also called *valence electrons*). For example, sodium, in Group 1, has one electron in its outer level (n = 3). Its electron-dot symbol will therefore be:

Na·

Carbon, in Group 14, has four electrons in its outer level. Each successive electron is placed at right, left, top, and bottom. So the four electrons of carbon in the n = 2 level are placed as follows:



Beyond four electrons, any additional electrons are "paired up" on each side in turn. For example, fluorine, in Group 17, has seven electrons in its outer layer (n = 2). These seven electrons are placed as follows:

Thus each element in Group 1 will have one electron-dot; each element in Group 2, two dots; each element in Group 13, three dots; each element in Group 14, four dots; each element in Group 15, five dots; each element in Group 16, six dots; each element in Group 17, seven dots; and each element in Group 18, eight dots – except helium which will have only two dots (as there are a maximum of two electrons in the n = 1 level).

5.7 Where Next?

There are now known to be over 100 chemical elements. One of the most amazing things is that these elements can combine together to form millions of different compounds. In the next chapter, you will see that there are two very different ways in which atoms can combine. And that the key to understanding the combination of atoms is the number of electrons in the outer filled level of each atom.