

Chapter V :

Periodic table

1- Introduction :

Did you know that although electrons are minuscule compared to other parts of an atom, the way they are arranged around the nucleus is the biggest factor in determining the chemical properties of an element? The periodic chart is ordered by atomic number, but drastic shifts in chemical properties can occur from one element to the next. These shifts are explained by how the elements are displayed on the periodic table.

2- Electronic composition of atom:

a- Orbital Diagrams :

- **S** orbital consists of one quantum quase



$$l = 0$$

(S)

- **P** orbital consists of three quantum quases



$$l = 1$$

(p)

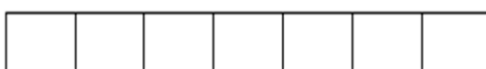
- **d** orbital consists of five quantum quases



$$l = 2$$

(d)

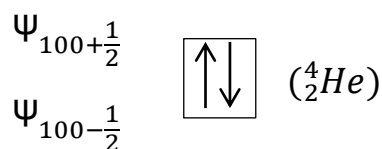
- **f** orbital consists of seven quantum quases



$$l = 3$$

(f)

There can be no two electrons with the same four quantum numbers. The first three quantum numbers (n , l , and m_l) can be the same, but the fourth must be different. A single orbital can only house two electrons with opposing spins. One electron spins up ($m_s = +1/2$), while the other spins down ($m_s = -1/2$).



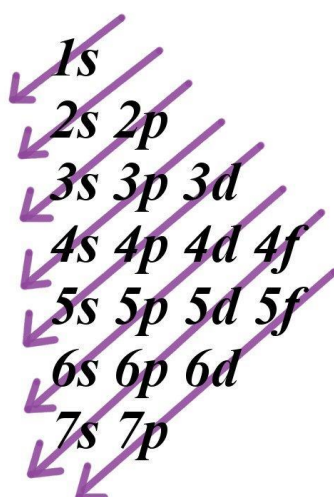
b- Rules for Assigning Electron Orbitals :

- **Occupation of Orbitals :**

Electrons fill orbitals in a way to minimize the energy of the atom. Therefore, the electrons in an atom fill the principal energy levels in order of increasing energy (the electrons are getting farther from the nucleus). The order of levels filled looks like this:

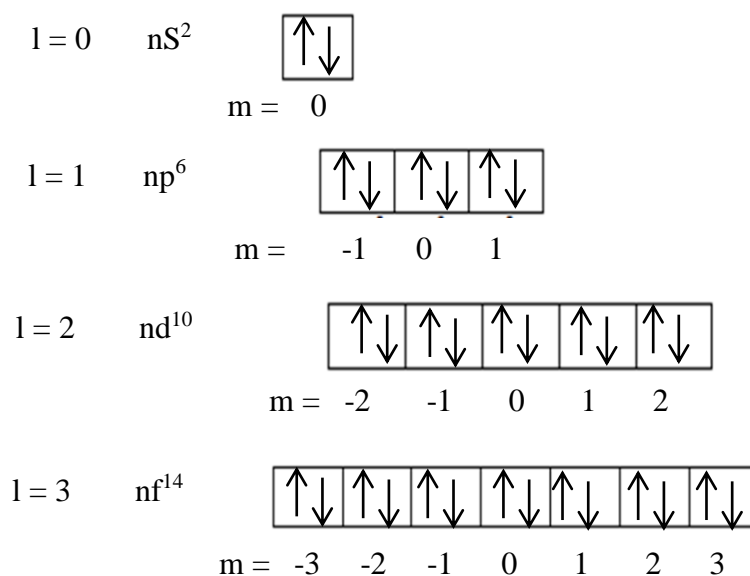
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, and 7p

One way to remember this pattern, probably the easiest, is to refer to the periodic table and remember where each orbital block falls to logically deduce this pattern. Another way is to make a table like the one below and use vertical lines to determine which subshells correspond with each other.



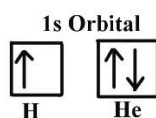
- **Pauli Exclusion Principle :**

The Pauli exclusion principle states that no two electrons can have the same four quantum numbers. The first three (n , l , and m_l) may be the same, but the fourth quantum number must be different. A single orbital can hold a maximum of two electrons, which **must** have opposing spins; otherwise they would have the same four quantum numbers, which is forbidden. One electron is spin up ($m_s = +1/2$) and the other would spin down ($m_s = -1/2$). This tells us that each subshell has double the electrons per orbital. The s subshell has 1 orbital that can hold up to 2 electrons, the p subshell has 3 orbitals that can hold up to 6 electrons, the d subshell has 5 orbitals that hold up to 10 electrons, and the f subshell has 7 orbitals with 14 electrons.



example 1: Hydrogen and Helium

The first three quantum numbers of an electron are $n=1$, $l=0$, $m_l=0$. Only two electrons can correspond to these, which would be either $m_s = -1/2$ or $m_s = +1/2$. As we already know from our studies of quantum numbers and electron orbitals, we can conclude that these four quantum numbers refer to the 1s subshell. If only one of the m_s values are given then we would have $1s^1$ (denoting hydrogen) if both are given we would have $1s^2$ (denoting helium). Visually, this is be represented as:



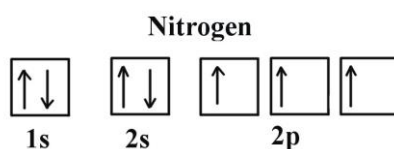
As shown, the 1s subshell can hold only two electrons and, when filled, the electrons have opposite spins.

- **Hund's Rule :**

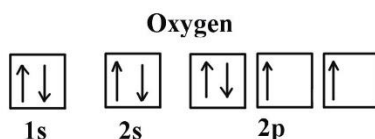
When assigning electrons in orbitals, each electron will first fill all the orbitals with similar energy (also referred to as degenerate) before pairing with another electron in a half-filled orbital. Atoms at ground states tend to have as many unpaired electrons as possible. When visualizing this processes, think about how electrons are exhibiting the same behavior as the same poles on a magnet would if they came into contact; as the negatively charged electrons fill orbitals they first try to get as far as possible from each other before having to pair up.

Example 2: Oxygen and Nitrogen

If we look at the correct electron configuration of the Nitrogen ($Z = 7$) atom, a very important element in the biology of plants: $1s^2 2s^2 2p^3$



We can clearly see that p orbitals are half-filled as there are three electrons and three p orbitals. This is because Hund's Rule states that the three electrons in the 2p subshell will fill all the empty orbitals first before filling orbitals with electrons in them. If we look at the element after Nitrogen in the same period, Oxygen ($Z = 8$) its electron configuration is: $1s^2 2s^2 2p^4$ (for an atom).



Oxygen has one more electron than Nitrogen and as the orbitals are all half filled the electron must pair up.

c- Writing Electron Configurations :

When writing an electron configuration, first write the **energy level** (the period), then the **subshell** to be filled and the **superscript**, which is the number of electrons in that subshell. The total number of electrons is the atomic number, Z . The rules above allow one to write the electron configurations for all the elements in the periodic table. Three methods are used to write electron configurations:

1. orbital diagrams
2. spdf notation
3. noble gas notation

Each method has its own purpose and each has its own drawbacks.

✓ Orbital Diagrams :

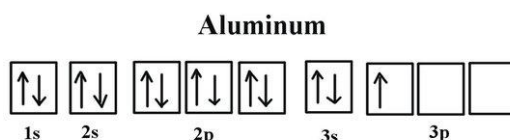
An orbital diagram, like those shown above, is a visual way to reconstruct the electron configuration by showing each of the separate orbitals and the spins on the electrons. This is done by first determining the subshell (s,p,d, or f) then drawing in each electron according to the stated rules above.

Example 4: Aluminum and Iridium

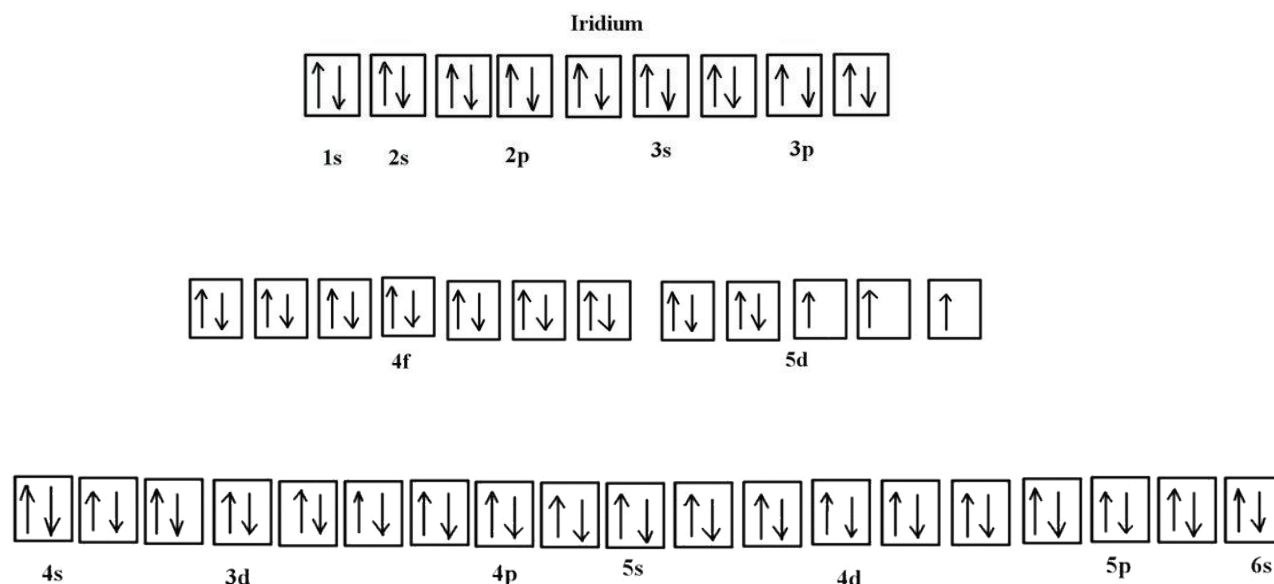
Write the electron configuration for aluminum and iridium.

SOLUTION

Aluminum is in the 3rd period and it has an atomic number of $Z=13$. If we look at the periodic table we can see that its in the p-block as it is in group 13. Now we shall look at the orbitals it will fill: 1s, 2s, 2p, 3s, 3p. We know that aluminum completely fills the 1s, 2s, 2p, and 3s orbitals because mathematically this would be $2+2+6+2=12$. The last electron is in the 3p orbital. Also another way of thinking about it is that as you move from each orbital block, the subshells become filled as you complete each section of the orbital in the period. The block that the atom is in (in the case for aluminum: 3p) is where we will count to get the number of electrons in the last subshell (for aluminum this would be one electron because its the first element in the period 3 p-block). This gives the following:



Note that in the orbital diagram, the two opposing spins of the electron can be visualized. This is why it is sometimes useful to think about electron configuration in terms of the diagram. However, because it is the most time consuming method, it is more common to write or see electron configurations in spdf notation and noble gas notation. Another example is the electron configuration of iridium:



The electron configuration of iridium is much longer than aluminum. Although drawing out each orbital may prove to be helpful in determining unpaired electrons, it is very time consuming and often not as practical as the spdf notation, especially for atoms with much longer configurations. Hund's rule is also followed, as each electron fills up each 5d orbital before being forced to pair with another electron.

✓ spdf Notation :

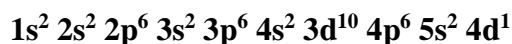
The most common way to describe electron configurations is to write distributions in the spdf notation. Although the distributions of electrons in each orbital are not as apparent as in the diagram, the total number of electrons in each energy level is described by a superscript that follows the relating energy level. To write the electron configuration of an atom, identify the energy level of interest and write the number of electrons in the energy level as its superscript as follows: $1s^2$. This is the electron configuration of helium; it denotes a full s orbital. The periodic table is used as a reference to accurately write the electron configurations of all atoms.

Example 5: Yttrium

Write the electronic configuration of Yttrium.

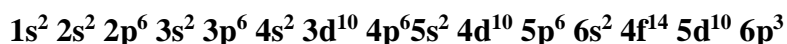
SOLUTION

Start with the straightforward problem of finding the electron configuration of the element yttrium. As always, refer to the periodic table. The element yttrium (symbolized Y) is a transition metal, found in the fifth period and in Group 3. In total it has thirty-nine electrons. Its electron configuration is as follows:



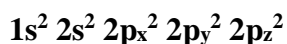
This is a much simpler and more efficient way to portray electron configuration of an atom. A logical way of thinking about it is that all that is required is to fill orbitals across a period and through orbital blocks. The number of elements in each block is the same as in the energy level it corresponds. For example, there are 2 elements in the s-block, and 10 elements in the d-block. Moving across, simply count how many elements fall in each block. Yttrium is the first element in the fourth period d-block; thus there is one electron in that energy level. To check the answer, verify that the subscripts add up to the atomic number. In this case, $2+2+6+2+6+2+10+6+2+1=39$ and $Z=39$, so the answer is correct.

A slightly more complicated example is the electron configuration of bismuth (symbolized Bi, with $Z = 83$). The periodic table gives the following electron configuration:

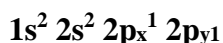


The reason why this electron configuration seems more complex is that the f-block, the Lanthanide series, is involved. Most students who first learn electron configurations often have trouble with configurations that must pass through the f-block because they often overlook this break in the table and skip that energy level. It's important to remember that when passing the 5d and 6d energy levels that one must pass through the f-block lanthanoid and actinoid series. Keeping this in mind, this "complex" problem is greatly simplified.

Another method (but less commonly used) of writing the spdf notation is the expanded notation format. This is the same concept as before, except that each individual orbital is represented with a subscript. The p, d, and f orbitals have different sublevels. The p orbitals are p_x , p_y , and p_z , and if represented on the 2p energy with full orbitals would look like: $2p_x^2 2p_y^2 2p_z^2$. The expanded notation for neon (Ne, $Z=10$) is written as follows:



The individual orbitals are represented, but the spins on the electrons are not; opposite spins are assumed. When representing the configuration of an atom with half filled orbitals, indicate the two half filled orbitals. The expanded notation for carbon is written as follows:



Because this form of the spdf notation is not typically used, it is not as important to dwell on this detail as it is to understand how to use the general spdf notation.

✓ Noble Gas Notation :

This brings up an interesting point about elements and electron configurations. As the p subshell is filled in the above example about the Aufbau principle (the trend from boron to neon), it reaches the group commonly known as the noble gases. The noble gases have the most stable electron configurations, and are known for being relatively inert. All noble gases have their subshells filled and can be used them as a shorthand way of writing electron configurations for subsequent atoms. This method of writing configurations is called the noble gas notation, in which the noble gas in the period above the element that is being analyzed is used to denote the subshells that element has filled and after which the valence electrons (electrons filling orbitals in the outer most shells) are written. This looks slightly different from spdf notation, as the reference noble gas must be indicated.

Example 6: Vanadium

What is the electronic configuration of vanadium (V, Z=23)?

SOLUTION

Vanadium is the transition metal in the fourth period and the fifth group. The noble gas preceding it is argon (Ar, Z=18), and knowing that vanadium has filled those orbitals before it, argon is used as the reference noble gas. The noble gas in the configuration is denoted E, in brackets: [E]. To find the valance electrons that follow, subtract the atomic numbers: $23 - 18 = 5$. Instead of 23 electrons to distribute in orbitals, there are 5. Now there is enough information to write the electron configuration:



This method streamlines the process of distributing electrons by showing the valence electrons, which determine the chemical properties of atoms. In addition, when determining the number of unpaired electrons in an atom, this method allows quick visualization of the configurations of the valance electrons. In the example above, there are a full s orbital and three half filled d orbitals.

3- Modern Periodic Table :

In 1869, the Russian chemist Dmitri Mendeleev first proposed that the chemical elements exhibited a "periodicity of properties." Mendeleev had tried to organize the chemical elements according to their atomic weights, assuming that the properties of the elements would gradually change as atomic weight increased. What he found, however, was that the chemical and physical properties of the elements increased gradually and then suddenly changed at distinct steps, or periods. To account for these repeating trends, Mendeleev grouped the elements in a table that had both rows and columns.

The modern periodic table of elements is based on Mendeleev's observations; however, instead of being organized by atomic weight, the modern table is arranged by atomic number (z).

Modern Periodic Law:

- Properties of elements are the periodic function to their atomic numbers.
- The periodicity in properties is due to repetition of similar outer shell electronic configuration at a certain regular intervals.
- In modern periodic table is based on modern periodic law in which elements are arranged in increasing order of their atomic numbers.
- In the modern periodic table, the elements are arranged in rows and columns. These rows and columns are known as periods and groups respectively.
- The table consists of 7 periods and 18 groups
- Period indicates the value of 'n' (principal quantum number) for the outermost or valence shell.
- Same number of electrons is present in the outer orbitals (that is, similar valence shell electronic configuration

H 1 (1.008)																	He 2 (4.003)
Li 3 (6.941)	Be 4 (9.012)											B 5 (10.811)	C 6 (12.011)	N 7 (14.007)	O 8 (15.999)	F 9 (18.998)	Ne 10 (20.180)
Na 11 (22.990)	Mg 12 (24.305)											Al 13 (26.982)	Si 14 (28.086)	P 15 (30.974)	S 16 (32.065)	Cl 17 (35.453)	Ar 18 (39.948)
K 19 (39.098)	Ca 20 (40.078)	Sc 21 (44.956)	Ti 22 (47.883)	V 23 (50.942)	Cr 24 (51.996)	Mn 25 (54.938)	Fe 26 (55.845)	Co 27 (58.933)	Ni 28 (58.693)	Cu 29 (63.546)	Zn 30 (65.38)	Ga 31 (69.723)	Ge 32 (72.630)	As 33 (74.922)	Se 34 (78.96)	Br 35 (79.904)	Kr 36 (83.80)
Rb 37 (85.468)	Sr 38 (87.62)	Y 39 (88.906)	Zr 40 (91.224)	Nb 41 (92.906)	Mo 42 (95.94)	Tc 43 (98.906)	Ru 44 (101.07)	Rh 45 (102.91)	Pd 46 (106.42)	Ag 47 (107.87)	Cd 48 (112.41)	In 49 (114.82)	Sn 50 (118.71)	Sb 51 (121.76)	Te 52 (127.6)	I 53 (126.91)	Xe 54 (131.29)
Cs 55 (132.91)	Ba 56 (137.33)	La 57 (138.905)	Hf 58 (178.49)	Ta 59 (180.948)	W 60 (183.84)	Re 61 (186.207)	Os 62 (190.23)	Ir 63 (192.22)	Pt 64 (195.084)	Au 65 (196.967)	Hg 66 (200.59)	Tl 67 (204.38)	Pb 68 (207.2)	Bi 69 (208.98)	Po 70 (209)	At 71 (210)	Rn 72 (222)
Fr 87 (223)	Ra 88 (226)	Ac 89 (227)	Rf 104 (261)	Db 105 (262)	Sg 106 (263)	Bh 107 (264)	Hs 108 (265)	Mt 109 (266)	Ds 110 (267)	Rg 111 (268)	Cn 112 (269)			Fl 114 (289)			Lv 116 (289)
Ce 58 (140.12)	Pr 59 (140.91)	Nd 60 (144.24)	Pm 61 (145)	Sm 62 (150.36)	Eu 63 (151.96)	Gd 64 (157.25)	Tb 65 (158.93)	Dy 66 (162.50)	Ho 67 (164.93)	Er 68 (167.26)	Tm 69 (168.93)	Yb 70 (173.05)	Lu 71 (174.967)				
Th 90 (232.04)	Pa 91 (231.04)	U 92 (238.03)	Np 93 (237)	Pu 94 (244)	Am 95 (243)	Cm 96 (247)	Bk 97 (247)	Cf 98 (251)	Es 99 (252)	Fm 100 (257)	Md 101 (258)	No 102 (259)	Lr 103 (260)				

✓ **Characteristics of Periods :**

- First period is called shortest period and contains only two elements. Second and third periods are called short periods containing eight elements each. Fourth and fifth periods are long periods containing eighteen elements each. Sixth period is the longest period with thirty-two elements. Seventh period is an incomplete period containing nineteen elements. Numbers 2, 8, 8, 18, 18, 32 are called magic numbers.
- Lanthanide and actinide series containing 14 elements each are placed separately under the main periodic table. These are related to sixth and seventh periods of III group respectively.
- Elements of third period from sodium (Na) to Chlorine (Cl) are called representative or typical elements.
- Valency of an element in a period increases from 1 to 7 with respect to oxygen.

Na₂O	MgO	Al₂O₃	SiO₂	P₂O₅	SO₃	Cl₂O₇
1	2	3	4	5	6	7

- From left to right in a period generally
 - Atomic weight, effective nuclear charge, ionisation potential, electronegativity and electron affinity of an element increases.
 - Atomic radius, electropositive character and metallic character of an element decreases.
- Diagonal relationship — Elements of second period Li, Be and B resemble closely with the elements Mg, Al and Si of third period in the next higher group.

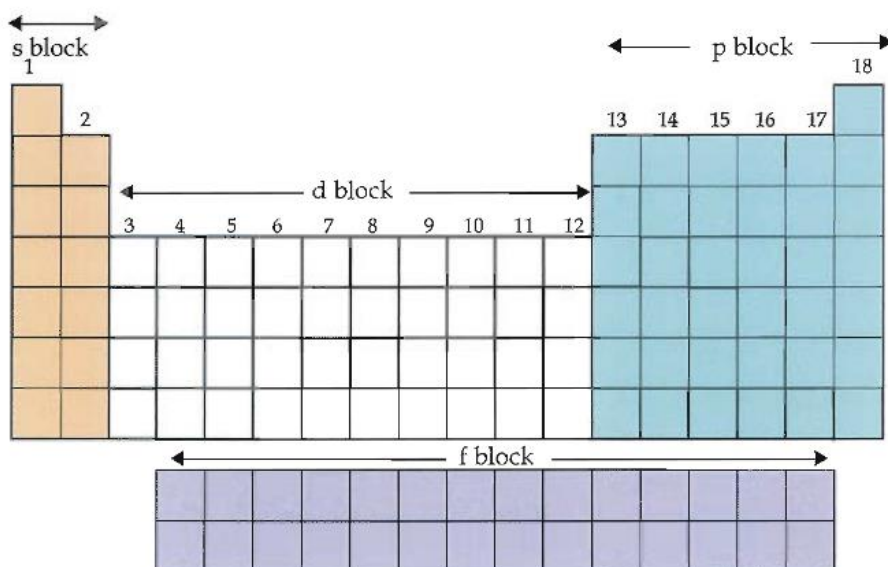
Second Period	Li	Be	B	C
Third Period	Na	Mg	Al	Si
- Elements of second period are called bridge elements.

✓ **Characteristic of Groups :**

- There are nine groups in modern periodic table and they are represented by roman numerals as I, II, III, IV, V, VI, VII, VIII and zero.
- Groups I to VII are further divided into two subgroups A and B, Group VIII consists of three sets, each one containing three elements.
- Inert gases or noble gases are placed in zero group.
- The valency of an element in a group is equal to the group number.

- The elements of the groups which resemble the typical elements are called normal elements. For example IA, IIA, IIIA, IVA, VA, VIA, VIIA group elements are normal elements.
- Those elements of the groups which do not resemble the typical elements are called transition elements. For example- IB, IIB, IIIB, IVB, VB, VIB, VIIB, and VIII group elements are transition elements. Hydrogen is placed in both IA and VIIA groups.
- Atomic weight, atomic size, electropositive character and metallic character of elements increases down the group.
- Ionisation potential, electron affinity and electronegativity of elements decreases down the group.

✓ **Classification of Elements :**



On the basis of electronic configuration, the elements may be divided into four groups.

s-block elements

- These are the elements in which last electron enters s subshell.
- These are present in the left part of the periodic table.
- These include 1 and 2 group elements.
- All the s- block elements are metals.
- The general electronic configuration of valence shell for s block elements is ns^{1-2} (n = 1 to 7).

p – block elements

- These are the elements in which last electron enters p subshell.

- These are present in the right part of the periodic table.
- These include 13 to 18 of the periodic table.
- Most of the p block elements are metalloids and non metals but some of them are metals also.
- The general electronic configuration of the valence shell is ns^2np^{1-6} ($n = 2$ to 7).
- ns^2np^6 is stable noble gas configuration. The electronic configuration of He is $1s^2$.

d-Block Elements

- These are the elements in which last electron enters d-subshell.
- These are present in the middle part of the periodic table (between s & p block elements)
- d block elements include group 3 to 10 groups of the periodic table.
- All d block elements are metals.
- The last electron fills in $(n - 1)d$ orbital.
- The outermost electronic configuration of d block elements is $(n-1)d^{1-10} ns^{1-2}$ ($n = 4$ to 7).
- There are three series of d-block elements as under
 - 3d series – Sc(21) to Zn (30)
 - 4d series – Y (39) to Cd (48)
 - 5d series – La (57), Hf (72) to Hg (80)

f -Block Elements

- These are placed separately below the main periodic table.
- These are mainly related to IIIB i.e. group 3 of the periodic table
- There are two series of f-block elements as under
 - 4f series – Lanthanides – 14 Elements i.e. Ce (58) to Lu (71)
 - 5f series – Actinides – 14 Elements i.e. Th (90) to Lw (103)
- the last electron fills in $(n - 2)f$ -orbital
- Their general outermost electronic configuration is $(n-2)f^{1-14} (n-1)s^2 (n-1)p^6 (n-1)d^{0-1} ns^2$ ($n = 6$ and 7).

4- Chemical families :

In the periodic table of elements, there are seven horizontal rows of elements. Each of these rows are called *periods*. The vertical columns of elements are called groups, or *families*. The most common way the periodic table is classified is by metals, nonmetals, and metalloids.

- **Periods in the periodic table :**

In each period (horizontal row), the atomic numbers increase from left to right. The periods are numbered 1 through 7 on the left-hand side of the table.

Elements that are in the same period have chemical properties that are not all that similar. Consider the first two members of period 3: sodium (Na) and magnesium (Mg). In reactions, they both tend to lose electrons (after all, they are metals), but sodium loses one electron, while magnesium loses two. Chlorine (Cl), down near the end of the period, tends to gain an electron (it's a nonmetal).

- **Families in the periodic table**

Just like human families, members of the families (vertical columns) in the periodic table have similarities or similar properties. The families are labeled at the top of the columns in one of two ways:

- The older method uses Roman numerals and letters. Many chemists prefer and still use this method.
- The newer method uses the numbers 1 through 18.

So why do the elements in the same family have similar properties? You can examine four families on the periodic table and look at the electron configurations for a few elements in each family.

Alkali Metals/Alkali Earth Metals :

The Alkali metals are comprised of group 1 of the periodic table and consist of Lithium, Sodium, Rubidium, Cesium, and Francium. These metals are highly reactive and form ionic compounds (when a nonmetal and a metal come together) as well as many other compounds. Alkali metals all have a charge of +1 and have the largest atom sizes than any of the other elements on each of their respective periods.

Alkali Earth Metals are located in group 2 and consist of Beryllium, Magnesium, Calcium, Strontium, Barium, and Radium. Unlike the Alkali metals, the earth metals have a smaller atom size and are not as reactive. These metals may also form ionic and other compounds and have a charge of +2.

Transition Metals :

The transition metals range from groups IIIB to XIIB on the periodic table. These metals form positively charged ions, are very hard, and have very high melting and boiling points. Transition metals are also good conductors of electricity and are malleable.

Lanthanides and Actinides :

Lanthanides and Actinides, form the block of two rows that are placed at the bottom of the periodic table for space issues. These are also considered to be transition metals. Lanthanides are form the top row of this block and are very soft metals with high boiling and melting points. Actinides form the bottom row and are radioactive. They also form compounds with most nonmetals.

Metalloids :

As mentioned in the introduction, metalloids are located along the staircase separating the metals from the nonmetals on the periodic table. Boron, silicon, germanium, arsenic, antimony, and tellurium all have metal and nonmetal properties. For example, Silicon has a metallic luster but is brittle and is an inefficient conductor of electricity like a nonmetal. As the metalloids have a combination of both metallic and nonmetal characteristics, they are intermediate conductors of electricity or "semiconductors".

Halogens :

Halogens are comprised of the five nonmetal elements Fluorine, Chlorine, Bromine, Iodine, and Astatine. They are located on group 17 of the periodic table and have a charge of -1. The term "halogen" means "salt-former" and compounds that contain one of the halogens are salts. The physical properties of halogens vary significantly as they can exist as solids, liquids, and gases at room temperature. However in general, halogens are very reactive, especially with the alkali metals and earth metals of groups 1 and 2 with which they form ionic compounds.

Noble Gases :

The noble gases consist of group 18 (sometimes referred to as group O) of the periodic table of elements. The noble gases have very low boiling and melting points and are all gases at room

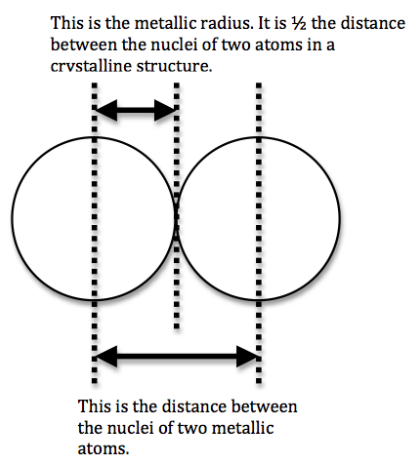
temperature. They are also very nonreactive as they already have a full valence shell with 8 electrons. Therefore, the noble gases have little tendency to lose or gain electrons.

5- Periodic Properties of the Elements :

The elements in the periodic table are arranged in order of increasing atomic number. All of these elements display several other trends and we can use the periodic law and table formation to predict their chemical, physical, and atomic properties. Understanding these trends is done by analyzing the elements electron configuration; all elements prefer an octet formation and will gain or lose electrons to form that stable configuration.

➤ Atomic Radius

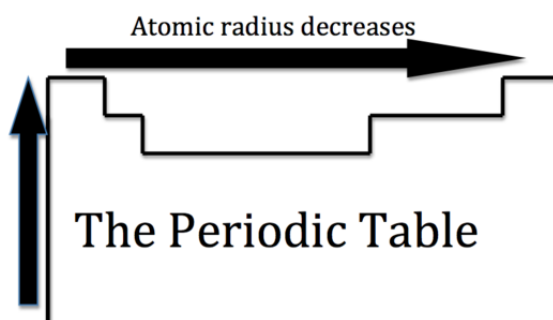
We can never determine the atomic radius of an atom because there is never a zero probability of finding an electron, and thus never a distinct boundary to the atom. All that we can measure is the distance between two nuclei (internuclear distance). A covalent radius is one-half the distance between the nuclei of two identical atoms. An ionic radius is one-half the distance between the nuclei of two ions in an ionic bond. A metallic radius is one-half the distance between the nuclei of two adjacent atoms in a crystalline structure. The SI units for measuring atomic radii are the nanometer (nm) and the picometer (pm). $1\text{nm}=1\times 10^{-9}\text{m}$ and $1\text{pm}=1\times 10^{-12}\text{m}$.



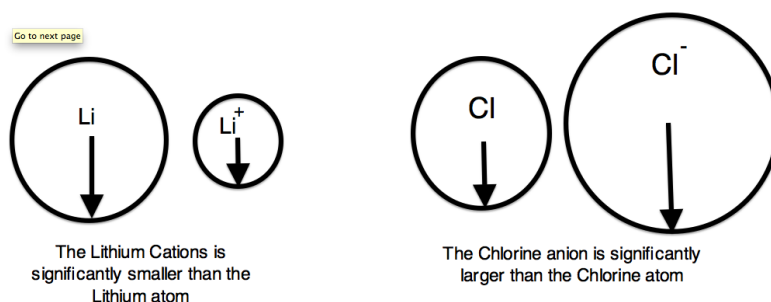
Metallic Radii.

description the atomic radius trend in the periodic table:

The atomic number increases moving left to right across a period and subsequently so does the effective nuclear charge. Therefore, moving left to right across a period the nucleus has a greater pull on the outer electrons and the atomic radii decreases. Moving down a group in the periodic table, the number of filled electron shells increases. In a group, the valence electrons keep the same effective nuclear charge, but now the orbitals are farther from the nucleus. Therefore, the nucleus has less of a pull on the outer electrons and the atomic radii are larger.



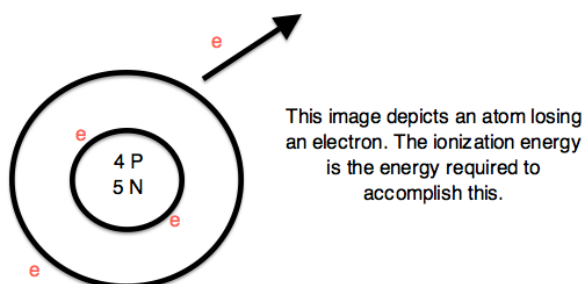
We can now use these concept to explain the atomic radius differences of cations and anions. A cation is an atom that has lost one of its outer electrons. Cations have a smaller radius than the atom that they were formed from. With the loss of an electron, the positive nuclear charge out powers the negative charge that the electrons exert. Therefore, the positive nucleus pulls the electrons tighter and the radius is smaller. An anion is an atom that has gained an outer electron. Anions have a greater radius than the atom that they were formed from. The gain of an electron does not alter the nuclear charge, but the addition of an electron causes a decrease in the effective nuclear charge. Therefore, the electrons are held more loosely and the atomic radius is increased.



➤ Ionization Energy (ionization potential)

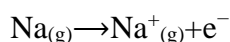
Expelling an electron from an atom requires enough energy to overcome the magnetic pull of the positive charge of the nucleus. Therefore, ionization energy (I.E. or I) is the energy required

to completely remove an electron from a gaseous atom or ion. The Ionization Energy is always positive.

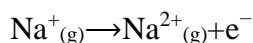


The energy required to remove one valence electron is the first ionization energy, the second ionization energy is the energy required to remove a second valence electron, and so on.

- 1st ionization energy



- 2nd ionization energy



Ionization energies increase relative to high effective charge. The highest ionization energies are the noble gases because they all have high effective charge due to their octet formation and require a high amount of energy to destroy that stable configuration. The highest amount of energy required occurs with the elements in the upper right hand corner. Additionally, elements in the left corner have a low ionization energy because losing an electron allows them to have the noble gas configuration. Therefore, it requires less energy to remove one of their valence electrons

Table 1: Ionization Energies of certain elements (1st IE, 2nd IE, etc)

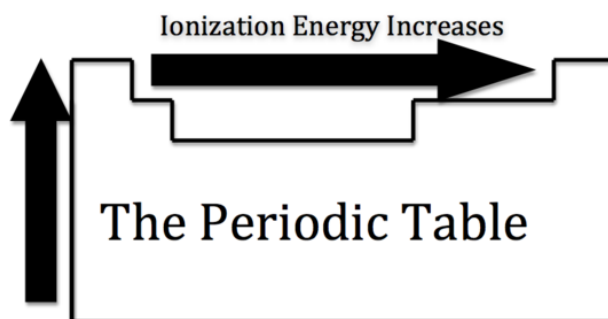
Element	1st	2nd	3rd	4th	5th	6th	7th
Na	496	4562					
Mg	738	1451	7733				
Al	577	1817	2745	11580			
Si	786	1577	3232	4356	16090		
P	1060	1903	2912	4957	6274	21270	
S	999.6	2251	3361	4564	7013	8496	27110
Cl	1256	2297	3822	5158	6542	9362	11020

Table 1: Ionization Energies of certain elements (1st IE, 2nd IE, etc)

Element	1st	2nd	3rd	4th	5th	6th	7th
Ar	1520	2666	3931	5771	7238	8781	12000

These are the ionization energies for the period three elements. Notice how Na after in the second I.E, Mg in the third I.E., Al in the fourth I.E., and so on, all have a huge increase in energy compared to the proceeding one. This occurs because the proceeding configuration was in a stable octet formation; therefore it requires a much larger amount of energy to ionize.

Ionization Energies increase going left to right across a period and increase going up a group. As you go up a group, the ionization energy increases, because there are less electron shielding the outer electrons from the pull of the nucleus. Therefore, it requires more energy to out power the nucleus and remove an electron. As we move across the periodic table from left to right, the ionization energy increases , due to the effective nuclear charge increasing. This is because the larger the effective nuclear charge, the stronger the nucleus is holding onto the electron and the more energy it takes to release an electron.

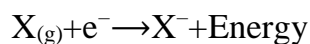


The ionization energy is only a general rule. There are some instances when this trend does not prove to be correct. These can typically be explained by their electron configuration. For example, Magnesium has a higher ionization energy than Aluminum. Magnesium has an electron configuration of $[\text{Ne}]3s^2$. Magnesium has a high ionization energy because it has a filled 3s orbital and it requires a higher amount of energy to take an electron from the filled orbital.

➤ **Electron Affinity :**

Electron affinity (E.A.) is the energy change that occurs when an electron is added to a gaseous atom. Electron affinity can further be defined as the enthalpy change that results from the addition of an electron to a gaseous atom. It can be either positive or negative value. The greater the negative value, the more stable the anion is.

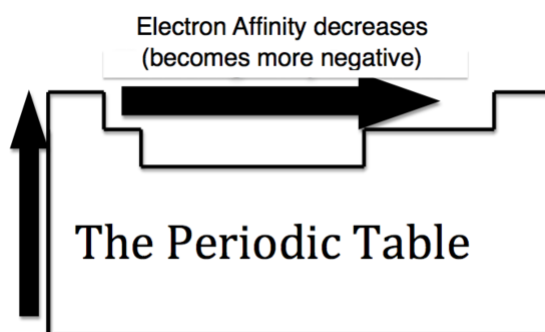
- (Exothermic) The electron affinity is positive



- (Endothermic) The electron affinity is negative

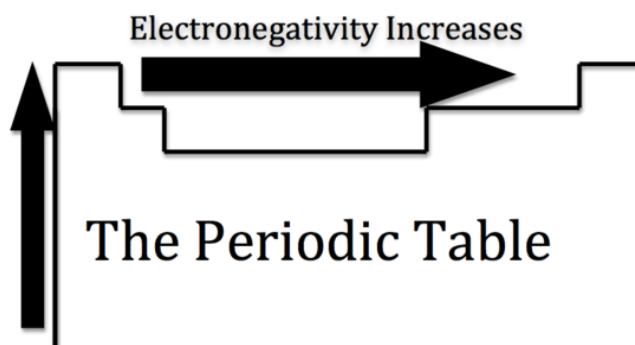


It is more difficult to come up with trends that describe the electron affinity. Generally, the elements on the right side of the periodic table will have large negative electron affinity. The electron affinities will become less negative as you go from the top to the bottom of the periodic table. However, Nitrogen, Oxygen, and Fluorine do not follow this trend. The noble gas electron configuration will be close to zero because they will not easily gain electrons.



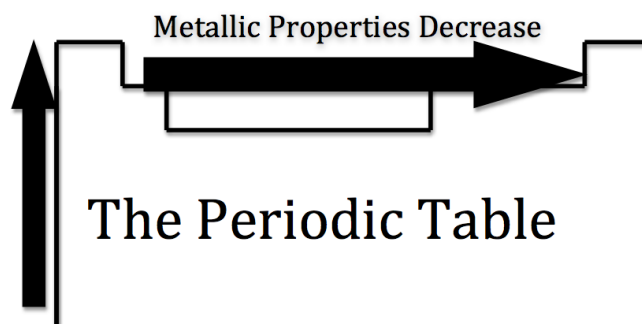
➤ **Electronegativity :**

Electronegativity is the measurement of an atom to compete for electrons in a bond. The higher the electronegativity, the greater its ability to gain electrons in a bond. Electronegativity will be important when we later determine polar and nonpolar molecules. Electronegativity is related with ionization energy and electron affinity. Elements with low ionization energies have low electronegativities because their nuclei do not exert a strong attractive force on electrons. Elements with high ionization energies have high electronegativities due to the strong pull exerted by the positive nucleus on the negative electrons. Therefore the electronegativity increases from bottom to top and from left to right.



➤ Metallic Character

The metallic character is used to define the chemical properties that metallic elements present. Generally, metals tend to lose electrons to form cations. Nonmetals tend to gain electrons to form anions. They also have a high oxidation potential therefore they are easily oxidized and are strong reducing agents. Metals also form basic oxides; the more basic the oxide, the higher the metallic character.



As you move across the table from left to right, the metallic character decreases, because the elements easily accept electrons to fill their valence shells. Therefore, these elements take on the nonmetallic character of forming anions. As you move up the table, the metallic character decreases, due to the greater pull that the nucleus has on the outer electrons. This greater pull makes it harder for the atoms to lose electrons and form cations.

Summary of Periodic Trends :

