Electronic configurations of Elements.

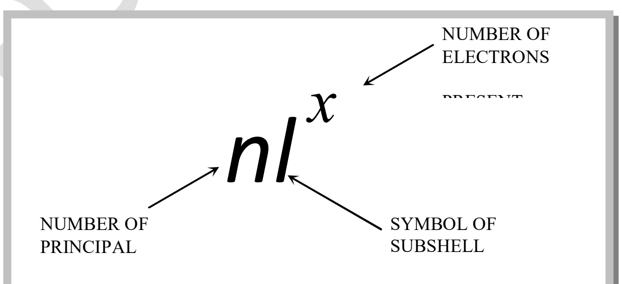
(1) On the basis of the electronic configuration principles the electronic configuration of various elements are given in the following table:

Electronic Configuration (E.C.) of Elements Z=1 to 36

Element	Atomic Number	1s	2s	2p	3s	3p	3d	4s	4p	4d	4f
H	1	1									
He	2	2									
Li	3	2	1								
Be	4	2	2								
B	5	2	2	1							
C	6	2	2	2							
N	7	2	2	3							
O	8	2	2	4							
F	9	2	2	5							
Ne	10	2	2	6							
Na	11	2	2	6	1						
Mg	12				2						
Al	13				2	1					
Si	14	10			2	2					
P	15	electrons			2	3					
S	16				2	4					
Cl	17				2	5					
Ar	18	2	2	6	2	6					
K	19	2	2	6	2	6		1			
Ca	20							2			
Sc	21						1	2			

Ti	22						2	2			
V	23						3	2			
Cr	24						5	1			
Mn	25						5	2			
Fe	26						6	2			
Co	27	18					7	2			
Ni	28	electrons					8	2			
Cu	29						10	1			
Zn	30						10	2			
Ga	31						10	2	1		
Ge	32						10	2	2		
As	33						10	2	3		
Se	34						10	2	4		
Br	35						10	2	5		
Kr	36	2	2	6	2	6	10	2	6		

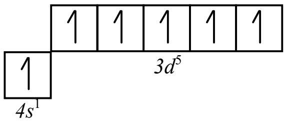
(2) The above method of writing the electronic configurations is quite cumbersome. Hence, usually the electronic configuration of the atom of any element is simply represented by the notation.

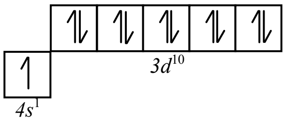


e.g. $1s^2$ means 2 electrons are present in the s- subshell of the 1st main shell.

(3) (I) Elements with atomic number 24(Cr), 42(Mo) and 74(W) have $ns^1(n-1)d^5$ configuration and not $ns^2(n-1)d^4$ due to extra stability of these atoms.

(ii) Elements with atomic number 29(Cu), 47(Ag) and 79(Au) have $ns^1(n-1)d^{10}$ configuration instead of $ns^2(n-1)d^9$ due to extra stability of these atoms.

Cr (24) [Ar] 

Cu (29) [Ar] 

(4) In the formation of ion, electrons of the outer most orbit are lost. Hence, whenever you are required to write electronic configuration of the ion, first write electronic configuration of its atom and take

electron from outermost orbit. If we write electronic configuration of Fe $^{2+}(Z = 26,\ 24\,e^-)$, it will not be similar to Cr (with $24\,e^-$) but quite different.

$$\left.\begin{array}{r} Fe[Ar]\,4s^2\,3d^6 \\ Fe^{2+}[Ar]\,4s^\circ\,3d^6 \end{array}\right\}$$ Outer most orbit is 4th shell hence, electrons from 4s have been removed to make Fe^{2+}.

(5) Ion/atom will be paramagnetic if there are unpaired electrons. Magnetic moment (spin only) is $\mu = \sqrt{n(n+2)}$ BM (Bohr Magneton).$(1BM = 9.27 \times 10^{-24}\ J/T)$ Where n is the number of unpaired electrons.

(6) Ion with unpaired electron in d or f orbital will be colored. Thus, Cu^+ with electronic configuration $[Ar]3d^{10}$ is colorless and Cu^{2+} with electronic configuration $[Ar]3d^9$ (one unpaired electron in 3d) is colored (blue).

(7) Position of the element in periodic table on the basis of electronic configuration can be determined as,

(I) If last electron enters into s-subshell, p-subshell, penultimate d-subshell and anti-penultimate f-subshell then the element belongs to s, p, d and f – block respectively.

(ii) Principle quantum number (n) of outermost shell gives the number of period of the element.

(i) If the last shell contains 1 or 2 electrons (i.e. for s-block elements having the configuration ns^{1-2}), the group number is 1 in the first case and 2 in the second case.

(iv) If the last shell contains 3 or more than 3 electrons (i.e. for p-block elements having the configuration $ns^2 np^{1-6}$), the group number is the total number of electrons in the last shell plus 10.

(v) If the electrons are present in the (n –1) d orbital in addition to those in the ns orbital (i.e. for d-block elements having the configuration (n –1) $d^{1-10} ns^{1-2}$), the group number is equal to the total number of electrons present in the (n –1) d orbital and ns orbital.