

of the elements from sodium to argon can be written as [Na, [Ne] $3s^1$] to [Ar, [Ne] $3s^2 3p^6$]. The electrons in the completely filled shells are known as core electrons and the electrons that are added to the electronic shell with the highest principal quantum number are called **valence electrons**. For example, the electrons in Ne are the core electrons and the electrons from Na to Ar are the valence electrons. In potassium (K) and calcium (Ca), the $4s$ orbital, being lower in energy than the $3d$ orbitals, is occupied by one and two electrons respectively.

A new pattern is followed beginning with scandium (Sc). The $3d$ orbital, being lower in energy than the $4p$ orbital, is filled first. Consequently, in the next ten elements, scandium (Sc), titanium (Ti), vanadium (V), chromium (Cr), manganese (Mn), iron (Fe), cobalt (Co), nickel (Ni), copper (Cu) and zinc (Zn), the five $3d$ orbitals are progressively occupied. We may be puzzled by the fact that chromium and copper have five and ten electrons in $3d$ orbitals rather than four and nine as their position would have indicated with two-electrons in the $4s$ orbital. The reason is that fully filled orbitals and half-filled orbitals have extra stability (that is, lower energy). Thus p^6 , p^3 , d^5 , d^{10} , f^7 , f^5 etc. configurations, which are either half-filled or fully filled, are more stable. Chromium and copper therefore adopt the d^5 and d^{10} configuration (Section 2.6.7) [caution: exceptions do exist].

With the saturation of the $3d$ orbital, the filling of the $4p$ orbital starts with gallium (Ga) and is complete at krypton (Kr). In the next eighteen elements from rubidium (Rb) to xenon (Xe), the pattern of filling the $5s$, $4d$ and $5p$ orbitals are similar to that of $4s$, $3d$ and $4p$ orbitals as discussed above. Then comes the turn of the $6s$ orbital. In caesium (Cs) and the barium (Ba), the s orbital contains one and two electrons, respectively. Then from lanthanum (La) to mercury (Hg), the filling up of electrons takes place in $4f$ and $5d$ orbitals. After this,

filling of $6p$, then $7s$ and finally $5f$ and $6d$ orbitals takes place. The elements after uranium (U) are all short-lived and all of them are produced artificially. The electronic configurations of the known elements (as determined by spectroscopic methods) are tabulated in Table 2.6 (page 66).

One may ask what is the utility of knowing the electron configuration? The modern approach to the chemistry, in fact, depends almost entirely on electronic distribution to understand and explain chemical behaviour. For example, questions like why two or more atoms combine to form molecules, why some elements are metals while others are non-metals, why elements like helium and argon are not reactive but elements like the halogens are reactive, find simple explanation from the electronic configuration. These questions have an answer in the MO or an model of atom. A detailed understanding of the electronic structure of atoms is, therefore, very essential for getting an insight into the various aspects of modern chemical knowledge.

2.6.6 Stability of Completely Filled and Half Filled Subshells

The ground state electronic configuration of the atom of an element always corresponds to the state of the lowest total electronic energy. The electronic configurations of most of the atoms follow the basic rules given in Section 2.6.5. However, in certain elements such as Cu, or Cr, where the two subshells ($4s$ and $3d$) differ slightly in their energies, an electron shifts from a subshell of lower energy ($4s$) to a subshell of higher energy ($3d$), provided such a shift results in all orbitals of the subshell of higher energy getting either completely filled or half filled. The valence electronic configurations of Cr and Cu, therefore, are $3d^5 4s^1$ and $3d^{10} 4s^1$ respectively and not $3d^4 4s^2$ and $3d^9 4s^2$. It has been found that there is extra stability associated with these electronic configurations.

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