# ATOMICSTRUCTURE

1st year Chemistry
New Syllabus
Punjab Boards

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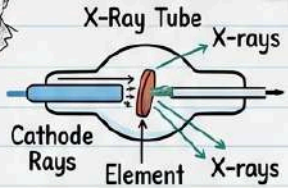
# 2.1 ATOMIC NUMBER, PROTON NUMBER & NUCLEON NUMBER, IDENTITY OF AN ELEMENT

## ★ Moseley's Law (1913)

1913: Moseley observed X-rays of characteristic frequencies when different elements are bombarded with cathode rays.

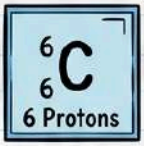



**STATEMENT**

The square root of the X-ray frequency ($\sqrt{\nu}$) is $\sqrt{\nu} \propto Z$ *directly proportional to the ATOMIC NUMBER (Z) of the element.

## ★ Proton Number (Z) ★

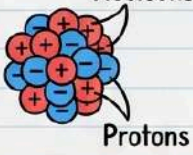
Moseley concluded Z is a FUNDAMENTAL PROPERTY of an element.
Also called PROTON NUMBER.




Z = 6

## ★ Nucleon Number (A) ★

The number of Protons and Neutrons in the nucleus of an atom is collectively called Nucleon Number (A), or Mass Number.




$A = Z + N$ Nucleon No. = Proton No. + Neutron No.

## ★ Example ★

$^{27}_{13}Al_{13}$ → A (Mass Number/Nucleon No.), X (Element Symbol), Z (Atomic Number/Proton No.)

For $^{27}Al_{13}$ :
Atomic Number (Z) = 13 (Protons)
Mass Number (A) = 27

$N = A - Z$

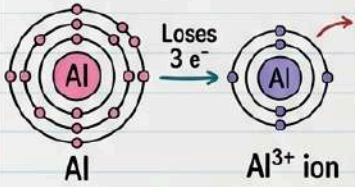
N = 27 - 13 = 14 (Neutrons)

# CALCULATION OF NO. OF ELECTRONS, PROTONS, & NEUTRONS

## ★ Calculation for Ions (Cations) ★

A neutral atom LOSES ELECTRONS to form a POSITIVE ION (CATION).

Example: A neutral Aluminium atom loses three electrons to form $Al^{3+}$ ion.




**Protons - No. of electrons lost**

No. of Protons = 13

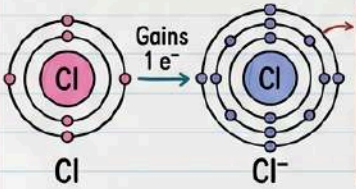
No. of Neutrons = 14

No. of Electrons = 13 - 3 = 10

## ★ Calculation for Ions (Anions) ★

A neutral atom GAINS ELECTRONS to form a NEGATIVE ION (ANION).

Similarly, a neutral Chlorine atom gains an electron to form $Cl^-$ ion.




**Protons + No. of electrons gained**

No. of Protons = 17

No. of Neutrons = 18

No. of Electrons = 17 + 1 = 18

## ★ Other Examples (Electron Gain) ★

Table 2.1: Number of protons, electrons and neutrons in different ions

| Species | Neutrons | Protons | Electrons |
|---|---|---|---|
| $O^{2-}$ | 8 | 8 | 10 |
| $S^{2-}$ | 16 | 16 | 18 |
| $P^{3-}$ | 16 | 15 | 18 |

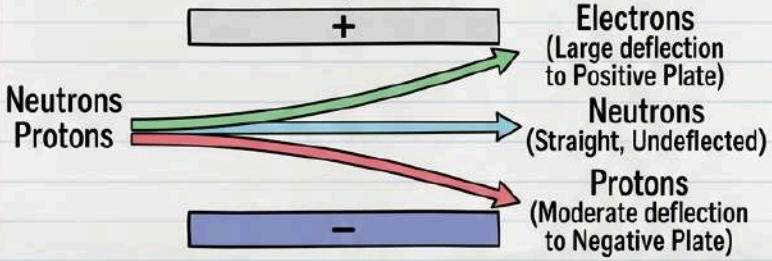
Thus, the atomic number and proton number represent the same concept.

# 2.2 EFFECT OF ELECTRIC FIELD ON FUNDAMENTAL PARTICLES

## ★ General Principle ★

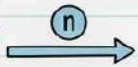
Behaviour in an electric field depends on:
1. MASS and 2. CHARGE.

Assuming beams of electrons, protons, and neutrons pass one-by-one at the same speed through an electric field:



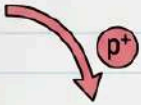

## 1. ★ Effect on Neutrons ★

Neutrons are NEUTRAL. They are **not deflected** but travel in a straight path perpendicular to the electric field direction.



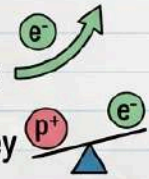

## 2. ★ Effect on Protons ★

Protons are POSITIVELY CHARGED. They are **deflected towards the negative plate.**




## 3. ★ Effect on Electrons ★

Electrons are NEGATIVELY CHARGED. They are **deflected towards the positive plate.**
Deflection is TO GREATER EXTENT because they are *thousands of times lighter* than protons.




## * Measurement of Amount of Deviation ★

The deviation from the original path is measured in two ways:
(i) Angle of deflection ∝ charge / mass
(ii) Radius of deflection ∝ mass / charge

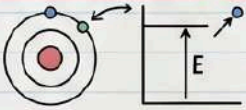

Possible by imagining the particle moves in a circular path. Radius of deflection factors RECIPROCAL to Angle of deflection factors.

# 2.3 EXPERIMENTAL EVIDENCES FOR THE ELECTRONIC CONFIGURATION

★ Modern theory of electronic structure originates from Bohr Model of Atom

Evidence derives principally from two sources:

★ Atomic Spectra ★ ★ Ionization Energies ★




## 2.3.1 Atomic Spectra

### ✷ Two Types of Atomic Spectra ★

★ Atomic Emission Spectrum ★

Element heated to high temp or subjected to electrical discharge in gaseous state emits radiation.
Spectrum contains COLOUR COLOUR COLOURED LINES.

★ Atomic Absorption Spectrum ★

White light passes through gaseous sample of element in COLD STATE, absorbing certain wavelengths. Spectrum shows up as DARK LINES on a continuous spectrum background.

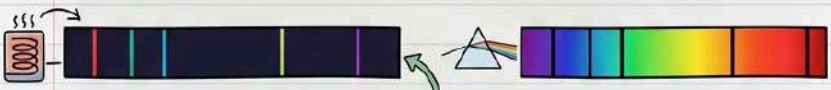


Figure 2.2(a): Hydrogen Emission Spectrum

Wavelengths of dark lines match exactly with coloured lines!

Figure 2.2(b): Hydrogen Absorption Spectrum

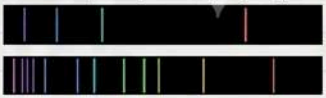
### Importance of Atomic Spectra ★

Each element has a UNIQUE range of fixed ENERGY LEVELS.
Wavelengths absorbed or emitted are also UNIQUE.
Characteristics of electron arrangement.

**Conclusion:** Every element has its own characteristic spectrum for identification.

Atomic spectra are the "FINGERPRINTS" of the elements!

Example Spectra (Na, Hg)



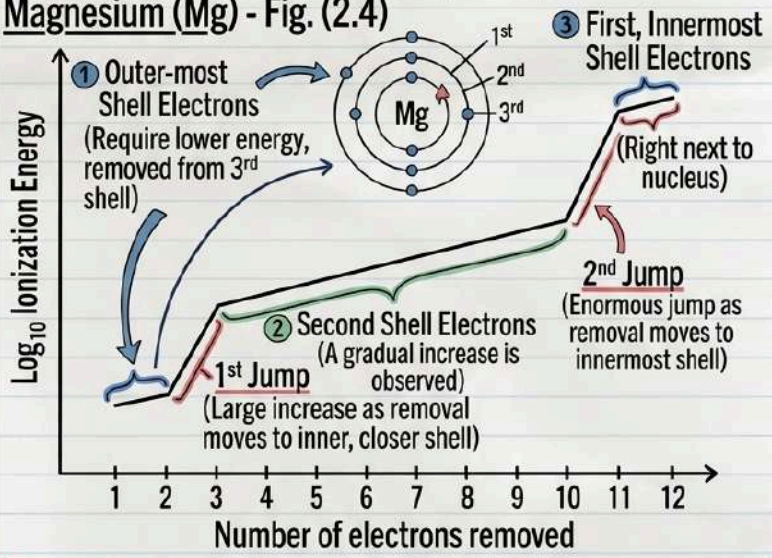
# IONIZATION ENERGY AND ENERGY LEVELS (ELECTRONIC SHELLS)

## ★ MEASURING ELECTRON ENERGIES INVESTIGATES ELECTRONIC CONFIGURATION

### (i) Successive Ionization Energies of the same element

We can look at an atom of a particular element and measure the energy required to remove each of its electrons, one by one. This sequence is called successive ionization energies and clearly shows the arrangement of electrons in shells around the nucleus.

### Successive Ionization Energies plot for Magnesium (Mg) - Fig. (2.4)




**Conclusion:** OVERALL, WE OBSERVE TWO LARGE JUMPS in the series of successive ionization energies. These two large jumps provide very good evidence that the electrons in the magnesium atoms exist in three different shells.

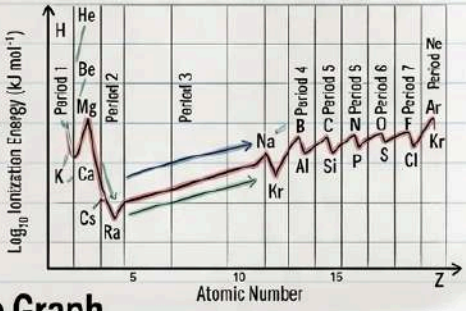
Ionization Energies are a crucial investigative tool!

# FIRST IONIZATION ENERGIES OF DIFFERENT ATOMS (VARIATION WITH ATOMIC NUMBER)

## ★ Relation between I.E and atomic number

The second way ionization energies show details of electronic configuration is by observing how first ionization energies (IE) vary with atomic numbers (Z).

**First Ionization Energy Plot for the first 88 Elements (Fig. 2.5)**



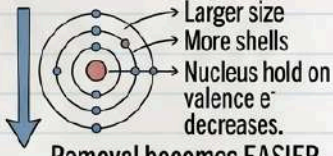

## ★ Conclusions from the Graph

### (a) Endothermic nature

- All first ionization energies are STRONGLY ENDOTHERMIC.


### (b) Variation of I.E down the Group

Going down a Group (e.g., He to Ne to Ar, or Li to Na to K), ionization energies DECREASE.



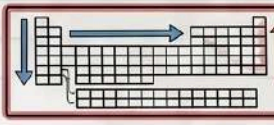

Removal becomes EASIER.

### (c) Variation of I.E across the Period

Ionization energies GENERALLY INCREASE going across a Period.


Group 1 elements (alkali metals) have LOWEST IE in each period.

- Across a Period, shell number remains same.
- As proton number increases, valence e⁻ are added to same shell.
- Nucleus attracts valence e⁻ more STRONGLY.



First Ionization Energy Trends are a key indicator of atomic structure!

# 2.4.3 MAGNETIC QUANTUM NUMBER ($m$)

(Angular Momentum Quantum Number)

## ★ DESCRIBES ORBITAL ORIENTATION ★

Within a subshell, values of $m$ depend on $\ell$.

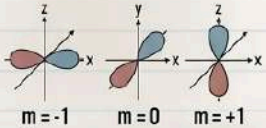
For a given $\ell$, there are $(2\ell + 1)$ integral values of as follows:

... -1, 0, +1 ...

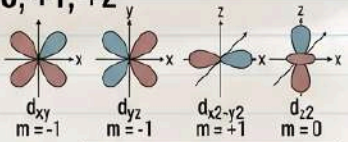
### Relationship between $\ell$ and $m$ ( orbitals in subshells )

- s-subshell ($\ell = 0$) ⇒ m = 0 (only one value).
  ⇒ s-subshell has 1 orbital.

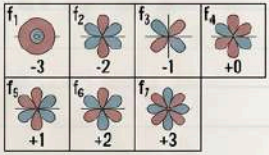

- p-subshell ($\ell = 1$) ⇒ m = -1, 0, +1
  ⇒ p-subshell has 3 orbitals.
  (three values).




- d-subshell ($\ell = 2$) ⇒ m = -2, -1, 0, +1, +2
  ⇒ five values).
  ⇒ d-subshell has 5 orbitals.

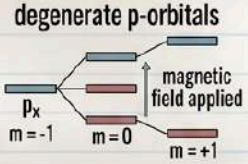



- f-subshell ($\ell = 3$) ⇒
  ⇒ m = -3, -2, -1, 0, +1, +2, +3
  ⇒ f-subshell has 7 orbitals.




## ★ DEGENERATE ORBITALS ★

Orbitals of the same subshell have the same energy (without a magnetic field) and are called degenerate orbitals.




### Table 2.5 Relationship between $\ell$ and m

- for $\ell = 0$ (s-subshell) ⇒ m = 0 ⇒ One orbital.
- for $\ell = 1$ (p-subshell) ⇒ m = -1, 0, +1 ⇒ Three degenerate p-orbitals.
- for $\ell = 2$ (d-subshell) ⇒ m = +2, +1, 0, -1, -2 ⇒ Five degenerate d-orbitals.
- for $\ell = 3$ (f-subshell) ⇒ m = +3, +2, +1, 0, -1, -2, -3
  0, -1, -2, -3 ⇒ Seven degenerate f-orbitals.

**Conclusion: The three quantum numbers (n, $\ell$, m) describe the size, shape, and orientation of an orbital! Fourth (spin) is next...**

# 2.4.4 SPIN QUANTUM NUMBER (s)

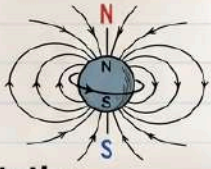
## ★ DESCRIBES ELECTRON SPIN ★

Electrons are thought of spinning on their own axes, like the Earth.



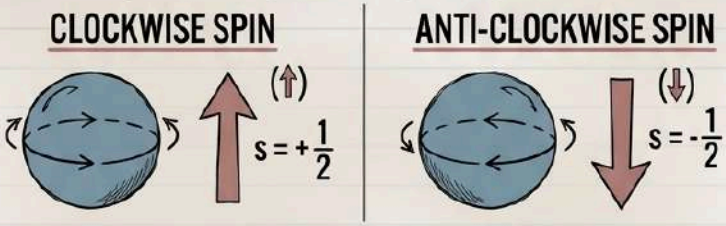
## Relationship between spinning motion and behavior

- According to electromagnetic theory, a spinning charge generates a magnetic field.
- This motion causes an electron to behave like a tiny magnet.




## Direction of electron spin & representation

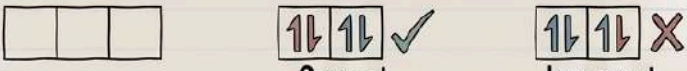
Fig. 2.4: Two possible spinning motions




- Clockwise is represented by an upwards-pointing arrow (↑), while anti-clockwise is by a downwards-pointing arrow (↓).

## ★ Orbital capacity & requirement ★

- Each orbital can accommodate at the most two electrons, *provided* the two electrons have **opposite spins.**




**Conclusion:** It takes three quantum numbers ($n$, $\ell$, $m$) to describe an orbital, but a fourth (s) to differentiate between the two electrons within it!

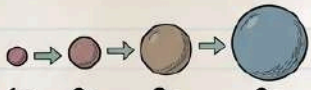
# 2.5 SHAPES OF ATOMIC ORBITALS

## ★ ORBITAL: MAXIMUM PROBABILITY REGION ★

Atomic orbital defines the 3D region in space around the nucleus where the probability of finding an electron is maximum.


## 2.5.1 s-Orbital ★ SPHERICAL SHAPE ★

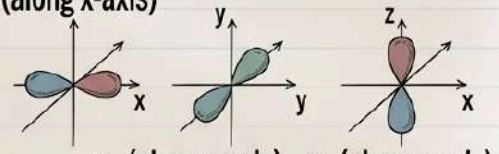
- Electronic density uniformly distributed in all directions.
- Size becomes larger with an increase in principas in Principal Quantum Number (n).




## 2.5.2 p-Orbitals ★ DUMBBELL SHAPE ★

- NOT spherically symmetric; distribution for 2p: Two lobes on any axis. 2p is shown.

- NAMED $p_x$, $p_y$, $p_z$ according to their axes.

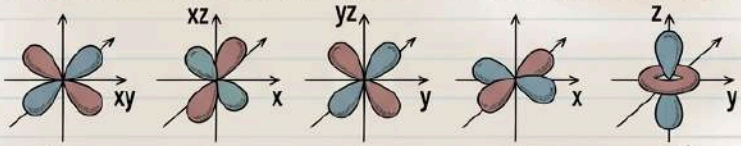



$p_y$ (along y-axis) $p_z$ (along z-axis)

## 2.5.3 d and f-Orbitals

## ★ d-ORBITALS (COMPLICATED ORIENTATIONS) ★

- In a given shell, 'd' orbitals ($\ell$=2) have different shapes & orientations.




$d_{xy}$ $d_{xz}$ $d_{yz}$ $d_{x^2-y^2}$ $d_{z^2}$ (doughnut" in xy plane)

Fig. 2.9: Shapes of the five d orbitals.

- f-subshell has seven orientations ($\ell$=3) → Seven complicated f-orbitals.

**Conclusion:** Principal quantum number (n) sets the shell size, while azimuthal ($\ell$) and magnetic (m) determine the orbital type and 3D shape!

# 2.6 ELECTRONIC CONFIGURATION

## Definition

"Electronic configuration is the distribution of electrons among available shells, subshells, or orbitals of an atom or ion."

- Each group of orbitals in a subshell is labeled by its subshell notation. 1s 2p 3d
- An electron in an orbital is shown by an arrow.
- The arrow points upward, when $s = +\frac{1}{2}$ and downward when $s = -\frac{1}{2}$. $s=+\frac{1}{2}$ $s=-\frac{1}{2}$ N

## Example

The orbital diagram of boron ($_{11}$5B) is as follows

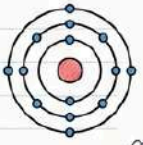
$1s^2$ $2s^2$ $2p^1$

Total electrons = 5 (2+2+1)

## 2.6.1 Distribution of Electrons in Shells

### Definition

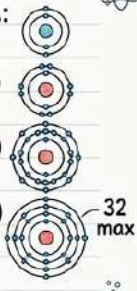
The electronic configuration of an atom describes the distribution of electrons in its atomic shells.



### Shell capacities

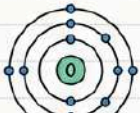
Each shell has a specific capacity for electrons:

- K shell (n=1) : 2 electrons maximum (1s)
- L shell (n=2) : 8 electrons maximum (2s, 2p)
  $n=1 \rightarrow 2n^2 = 2(3)^2 = 8$
- M shell (n=3) : 18 electrons maximum (3s, 3p)
  $n=3 \rightarrow 2n^2 = 2(3)^2 = 18$
- N shell (n=4) : 32 electrons maximum (4s, 4p)
  $n=4 \rightarrow 4n^2 = 4(3)^2 = 32$ max

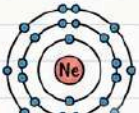



### Examples


C ($6e^-$) : K=2, L=4
or $1s^2\ 2s^2\ 2p^2$



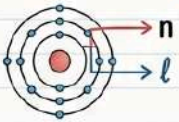

O ($8e^-$) : K=2, L=6
or $1s^2\ 2s^2\ 2p^4$




Ne ($10e^-$) : K=2, L=8
or $1s^2\ 2s^2\ 2p^6$

# Arrangement of subshells

- **How to arrange the subshells energy wise?**
- Since the energy of a subshell, in the absence of any magnetic field, depends upon the principal quantum number (n) and the azimuthal quantum number ($\ell$), hence the order of filling subshells with electrons may be obtained from the **summation ($n + \ell$)**.




**$(n + \ell)$**

### (a) Orbitals of low $(n+\ell)$ values

The subshell having lower $(n + \ell)$ value has lower energy and is filled first.

4s orbital

$n = 4, \ell = 0 \rightarrow (n+\ell) = 4 + 0 = 4$

3d orbital

$n = 3, \ell = 2 \rightarrow (n+\ell) = 3 + 2 = 5$


### (b) Orbitals of equal $(n + \ell)$ values

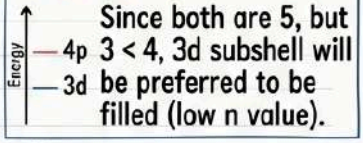
In case two subshells have equal $(n + \ell)$ values, then the subshell with lower 'n' value will be filled first.

4p subshell

$n = 4, \ell = 1 \rightarrow (n+\ell) = 4 + 1 = 5$

3d subshell

$n = 3, \ell = 2 \rightarrow (n+\ell) = 3 + 2 = 5$




- **Arrangement of orbitals**

According to this rule the energy-wise arrangement of orbitals should be:

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p <

< 4d < 5p < 6s < 4f < 5d < 6p < 7s

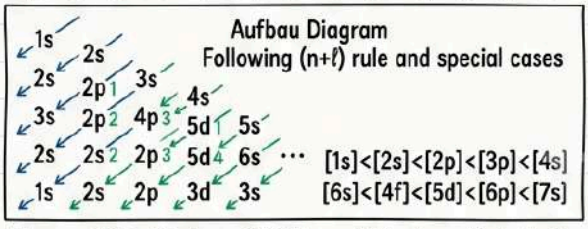
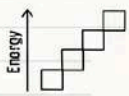



Figure 2.10: Order of Filling of Various Subshells

# 2.6.3 Distribution of Electrons in Orbitals

A useful way of representing electronic configurations is a diagram that places electrons in boxes Figure 2.11.

- Each box represents an atomic orbital.
- Each orbital can occupy maximum of two electrons.
- An electron is represented by an arrow. →
- The boxes (orbitals) can be arranged in order of increasing energy from bottom to top.




## (i) Pauli's exclusion principle Statement

According to this principle: 'No two electrons in an atom can have the same values for all the four quantum numbers' or, 'Two electrons in an orbital will always have opposite spins'.

### Example

In the first shell of helium (He) atom, there are two electrons. They are present in 1s orbital. According to the concept of quantum numbers and Pauli's exclusion principal, the values of their quantum numbers are:

Table 2.7 Values of quantum numbers of two electrons in the same orbital

| Electron | n | $\ell$ | m | s |
|---|---|---|---|---|
| Electron 1 | 1 | 0 | 0 | +½ ↑ |
| Electron 2 | 1 | 0 | 0 | −½ ↓ |

The two electrons having the same values of 'n', '$\ell$' and 'm' can have different values of 's'. It means that their spins are in the opposite directions.


Figure 2.11

# 2.6.4 (ii) Hund's rule

This rule gives an idea for filling electrons into the orbitals having equal energies.

## Statement

degenerate orbitals. ☐ ☐ ☐

"When degenerate orbitals are available and more than two electrons are to be placed in them, they should be placed in separate orbitals with the same spin rather than in the same orbital with opposite spins."

**Example:** Three p-orbitals, i.e., $p_x$, $p_y$ and $p_z$ have equal energy. To understand it, let us take an example in which three electrons are to be filled into three p-orbitals. There are two different ways to do this as shown below:

✗ ↑↓ | | |  (Wrong)

↑ | ↑ | ↑  ✓ (Correct)

Which of the above two is correct?

According to the Hund's rule, the correct way of filling three electrons in three p orbitals is that in which each orbital is singly occupied.

Electronic configuration in terms of orbitals in carbon, nitrogen and oxygen is provided in Figure 2.12

| Symbol | Atomic Number | Orbital Diagram | Electron Configuration |
|---|---|---|---|
| C | 6 | ↑↓ ↑↓ \| \| \| | $1s^2\ 2s^2\ 2p^2$ |
| N | 7 | ↑↓ ↑↓ \| ↑ \| ↑ \| | $1s^2\ 2s^2\ 2p^3$ |
| O | 8 | ↑↓ \| ↑ \| ↑ \| ↑ \| | $1s^2\ 2s^2\ 2p^4$ |

## Electronic configuration of some elements

The electronic configurations of the periodic tabll in the ligh light of the abovementioned principles are given in Table 2.6

Table 2.6 Configuration of some elements

| Element | Z | Principal Quantum Number shells K | L | ss | N | Full Subshell Configuration |
|---|---|---|---|---|---|---|
| H | 0 | K | | | | $1s^0$ |
| He | 10 | K | K | | | $1s^2$ |
| Li | 11 | 4 | L | L | | $1s^2 2s^1$ |
| Be | 6 | 4 | 5 | M | | $1s^2 3s^2\ 2p^0$ |
| B | 6 | 4 | 3 | M | N | $1s^2 2s^2\ 2p^0\ 2p^1$ |
| N | 7 | 4 | 3 | M | N | $1s^2 2s^2\ 2p^0\ 3p^1$ |
| O | 8 | 4 | 4 | N | N | $1s^2 2s^2\ 2p^0\ 3p^1$ |

Figure 2.11

# 2.7 ELECTRONIC CONFIGURATION AND THE PERIODIC TABLE

We have seen that the electronic configurations of elements are related to their position in the periodic table. The periodic table is structured so that elements with the same pattern of outer-shell (valence) electron configuration are arranged in the same groups.

You can easily write the electron configuration of an element based on its location in the periodic table. The pattern is summarized below. Notice that elements are grouped in terms of the type of orbital into which theelectrons are placed.

Table 2.8: Outer-Shell Configuration Patterns

| Periodic Table Block | Groups | Outer-Shell Configuration | Element Type | |
|---|---|---|---|---|
| s-block | Groups 1 & 2 | $ns^1$ (Alkali Metals), $ns^2$ (Alkaline Earth Metals) | Representative Elements | |
| p-block | Groups 13 to 18 | $ns^2np^{1-6}$ (e.g., $ns^2np^1$ for Gp 13) | Representative Elements | |
| d-block | Transition Metals | $(n-1)d^{1-10}ns^{1-2}$ | Transition Elements | |
| f-block | Lanthanides & Actinides | $(n-2)f^{1-14}$ | Inner Transition Elements | |

## Position of s-block element

On the left are two columns of elements, alkali metals and alkaline earth metals (groups 1 and 2), in which outer-shell s orbitals are being filled. Group 1 and 2 elements have $ns^1$ and $ns^2$ 2 elements have $ns^1$ and $ns^2$ outer configurations respectively.

## Position of p-block elements

Jump to group 13 elements, they have $ns^2np^1$ configuration. On the right is a block of six columns where outermost 'p' orbitals are being filled.

## Position of d-block elements

In the middle of the table is a block of ten columns containing the transition metals, where d orbitals are being filled.

## Position of f-block elements

Below the main portion of the table are two rows containing fourteen columns. These are f-block elements, where 'f' orbitals are being filled.

NOTE: Recall that the numbers 2, 6, 10, and 14 are precisely the number of electrons that can fill the s, p, d, and f subshells, respectively.

Figure 2.13

## 2.7.1 Valence Electrons

- ✓ The electrons in an atom in the outermost shell are called **valence electrons.**
- ✓ These are such electrons that are primarily involved in chemical reactions.

### Similarities of chemical properties

- ✓ The similarities among the configurations of valence electrons account for similarities of the chemical properties among groups of elements.
- ✓ Table 2.8 shows a periodic table which include the valence shell configurations included. Note the similarity in electron configuration within any group (column) of elements.

Table 2.8: Outer-Shell Configuration Patterns

| Group 1 | Group 2 | Group 13 | Group 14 | |
|---|---|---|---|---|
| Li [$2s^1$] | Be [$2s^2$] | B [$2s^2 2p^1$] | C [$2s^2 2p^2$] | Similar valence configuration |
| Na [$3s^1$] | Mg [$3s^2$] | Al [$3s^2 3p^1$] | Si [$3s^2 3p^2$] | |

## 2.7.2 Classification of Elements of Periodic Table

- ✓ The main group or representative elements all have valence-shell configurations $ns^a\ np^b$. They have some choice of 'a' and 'b'.
- ✓ In other words, the outer 's' or 'p' subshell is being filled. Similarly, in the d-block transition elements, '(n-1)d' subshell is being filled.
- ✓ In the f-block transition elements or inner-transition elements, an '(n-2)f' subshell is being filled.

### Building up order

There is a definite pattern to the order of filling of the subshells as we go through the elements in the periodic table. From this we can write down the **building-up order.** The overall sketrich of periodic table is given in Table 2.9. This shows the blocks of elements.

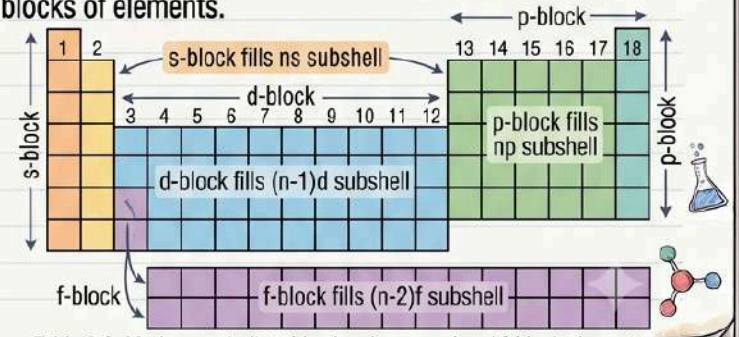
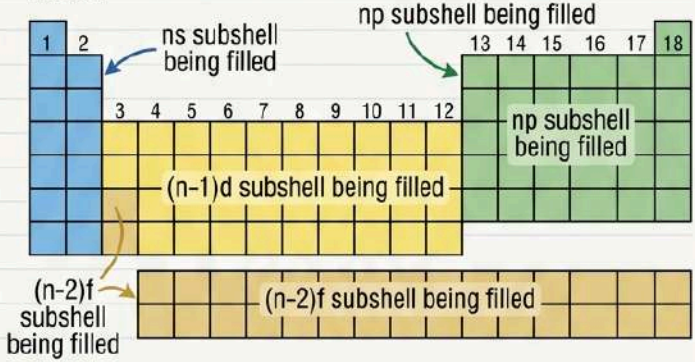



Table 2.9: Modern periodic table showing s, p, d and f-block elements

## Example

- ✓ The blue colored area, an 'ns' subshell is being filled.
- ✓ In the green colored area, an "np" subshell is being filled.
- ✓ In the yellow area, an (n – 1) d subshell is being filled.
- ✓ In the light golden area and (n–2) f subshell is being filled.




We start building-up order by starting with the first period, in which the '1s' subshell is being filled.

- ✓ In the second period, we have '2s'; then staying in the same period but jumping across, we have '2p'.
- ✓ In the third period, we have '3s' and '3p'; in the fourth period, '4s', '3d and thin '4p. in the fourth period, '4s', '3d' and then '4p'. This pattern should become clear enough to visualize with a periodic table.

To illustrate this,


## Classification of elements

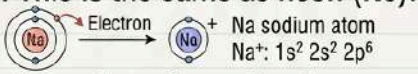
- ✓ The classification of elements of periodic table has been done into metals, non-metals, representative, transition elements, periods, groups and blocks (s, p, d, f).
- ✓ This division is done on the basis of electronic configuration and their valence electrons.
- ✓ The chemical properties of various categories of elements discussed above can be easily assessed.

# 2.8 ELECTRONIC CONFIGURATION OF IONS AND FREE RADICALS

## 2.8.1 Ions

Electronic configuration of positive ions

- ✓ Positive ions are formed when electrons are removed from atoms.
- ✓ Example: The sodium ion, $Na^+$ (proton number = 11), has 10 electrons. Its electronic configuration is $1s^2\ 2s^2\ 2p^6$. This is the same as neon (Ne).

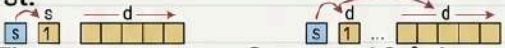



Electronic configuration of negative ions

- ✓ Negative ions are formed when atoms gain electrons.
- ✓ Example: The sulfide ion, $S^{2-}$ (proton number = 16), has 18 electrons. Its electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$, the same as argon (Ar).


**NOTE:**

- ✓ In general, electrons in the outer subshell are removed when metal atoms form the positive ions.
- ✓ **However, the d-block elements behave slightly differently.** Reading across from potassium to zinc, the 4s fills before the 3d. But when atoms of a d-block element lose electrons to form ions, the 4s electrons are lost first.




Ti: atom/ Ti atom
Cr: atom/ $Ti^{2+}$ ion

Cr atom/ $Cr^{3+}$ ion
Cr atom/ $Cr^{3+}$ ion

## 2.8.2 Free Radicals

- ✓ "A free radical is a species that has one or more unpaired electrons".
- ✓ Example: A simple free radical is a free chlorine atom. The configuration of this radical is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$. In the 3p subshell, two orbitals have paired electrons, and the third contains a single unpaired electron.

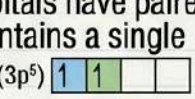



Figure 2.14

- ✓ NOTE: Apart from single atoms, groups of atoms can also be free radicals.
- ✓ Examples: groups of atoms like hydroxyl (·OH), or methyl ($\cdot CH_3$).

# 2.9 ELECTRONIC CONFIGURATION AND THE FORMATION OF SEMICONDUCTORS

Figure 2.15

## Semiconductors

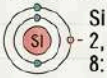
"Semiconductors are materials that can conduct electricity under some conditions."

- ✓ Used in many electronic devices, including smartphones, laptops, and cars.
- ✓ **Example:** Elements like silicon, germanium, and arsenic etc. can act as semiconductors.

## Formation of semiconductors

Formation is possible because of a unique electronic configuration. The etectrance of a unique electronic configuration. Let us consider Si and explore p-type and n-type transformation.

- ✓ Electron configuration of $_{14}Si = 2, 8, 4$; meaning 4 valence electrons.

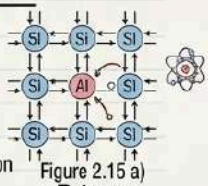



- ✓ Pure Si crystal: Each atom atom bonded to four others, no conduction.


- ✓ **P-type** and **N-type** are formed by '**doping**' with impurity atoms. Doping Process ⟶
  - **P-type**: Created by adding trivalent impurities (B or Al), creating '**holes**' (positive charge carriers).
  - **N-type**: Created by adding pentavalent impurities (P or As), creating '**extra electrons**' (negative charge carriers).

## 2.9.1 P-type Semiconductor Formation

- ✓ Trivalent impurity atoms (like Al) are added to the pure semiconductor.
- ✓ Some take the place of Si atoms. Lack of electrons creates **holes** (as positive charge carriers).

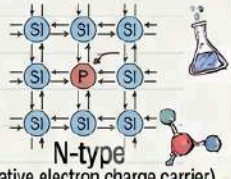

- ✓ Moving electrons from an external source can travel through the material.

## 2.9.2 N-type Semi-conductor Formation

- ✓ Pentavalent impurity atoms (like Phosphorus) are added. Some replace Si atoms.
- ✓ Some replace Si atoms. Vicinity can make, bonds, and the fifth is an **extra electron**.








- ✓ Impurity atoms contribute extra electrons that act as free negative charge carriers.