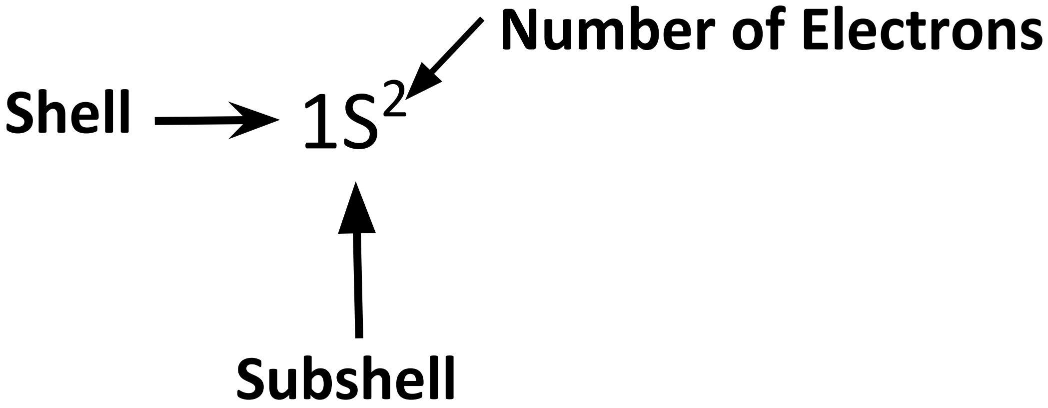
CHAPTER V
ELECTRONIC CONFIGURATIONS AND PERIODIC PROPERTIES OF ELEMENTS

I/ Electronic configuration of the elements of the periodic table

I.1/ The Electronic Configuration of Representative Elements

Electronic configuration shows how electrons are arranged in an atom. It tells us where the electrons are — in which energy levels and orbitals.



Example

Oxygen (Z = 8) → $1s^2\ 2s^2\ 2p^4$

I.2/The Main Principles of Electronic Configuration

The electronic configuration follows three main principles:

✔ Energy Levels of Subshells and the Filling of Orbitals

According to this principle, electrons fill orbitals starting with the lowest energy and moving to higher ones. The same rule applies inside each subshell. This way, the atom keeps its lowest total energy.

Experiments show that orbitals are filled in the following order of increasing energy (with some small exceptions):

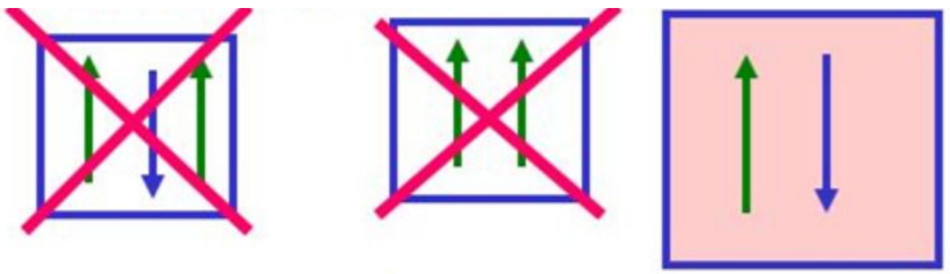
1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f → 5d → 6p → 7s → 5f → 6d → 7p

✔ The Pauli Exclusion Principle

The Pauli Exclusion Principle states that no two electrons in the same atom can have the same set of four quantum numbers.

It follows that an atomic orbital can contain no more than two electrons, and these two electrons must have opposite spins.

It is customary to represent orbitals by quantum boxes, which are squares in which electrons are symbolized by arrows (↑ or ↓). The upward arrow (↑) arbitrarily represents the spin value +1/2.



✔ Hund's Rule

Hund's rule states that electrons occupy degenerate orbitals singly first and with parallel spins before pairing up.

I.3/ Definition of the Valence Shell

The valence shell is the outermost electron shell of an atom.

It contains the valence electrons, which are the electrons involved in chemical bonding and reactions.

Example Oxygen (Z = 8) → $1s^2\ 2s^2\ 2p^4$

In the atom of oxygen (Z = 8), the valence shell is the second shell (n = 2), which contains six electrons.

I.4/ Electronic configurations of periodic table elements

Each element in the periodic table has a unique electronic configuration. It shows how the electrons are arranged in the different shells and subshells around the nucleus.

The electronic configuration helps us understand the position of the element in the periodic table and explains its chemical properties.

Elements in the same group have similar outer electron configurations, so they show similar chemical behavior.

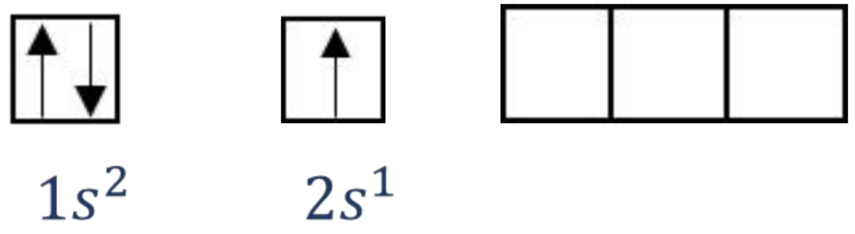
I.4.a/Electronic configuration of s- and p-block elements

✔ Group IA Elements (1)

Example : Li (Z = 3) : $1s^2\ 2s^1$

Group IA elements : ns^1 (Paramagnetic element).

The Group IA belongs to the s-block.

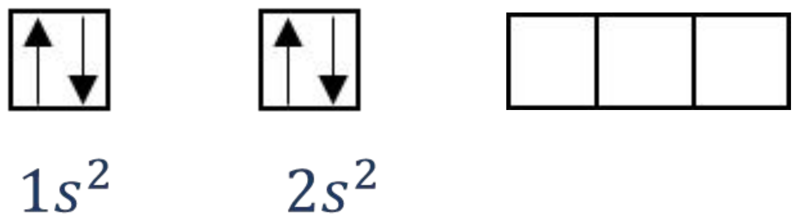


✔ Group IIA Elements (2)

Example: Be (Z = 4) : $1s^2\ 2s^2$

Group IIA elements : ns^2 (Diamagnetic element).

The Group IIA belongs to the s-block.



✔ Group IIIA Elements (13)

Example : B (Z = 5) : $1s^2\ 2s^2\ 2p^1$

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

Group IIIA elements : $ns^2\ np^1$(Paramagnetic element).

The Group IIIA belongs to the p-block.

✔ Group VIA Elements (14)

Example : C (Z = 6) : $1s^2\ 2s^2\ 2p^2$

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

Group IVA elements : $ns^2\ np^2$(Paramagnetic element).

The Group IVA belongs to the p-block.

✔ The Group VA Elements (15)

Example : N (Z = 7) : $1s^2\ 2s^2\ 2p^3$

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

Group VA elements : $ns^2\ np^3$(Paramagnetic element).

The Group VA belongs to the p-block.

✔ The Group VIA Elements (16)

Example : O (Z = 8) : $1s^2\ 2s^2\ 2p^6$

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

Group VA elements : $ns^2\ np^3$(Paramagnetic element).

The Group VIA belongs to the p-block.

✔ The Group VIIA Elements (17)

Example: F (Z = 9) : $1s^2\ 2s^2\ 2p^5$

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

Group VIIA elements : $\mathrm{ns}^2\ np^5$(Paramagnetic element).

The Group VIIA belongs to the p-block.

✔ The Group VIIIA Elements (18)

Exemple : Ne (Z = 10) : $1s^2\ 2s^2\ 2p^6$

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

Group VIIIA elements : $\mathrm{ns}^2\ np^6$(Diamagnetic element).

The Group VIIIA belongs to the p-block.

I.4.b/ The electronic configuration of transition elements

✔ Group IIIB

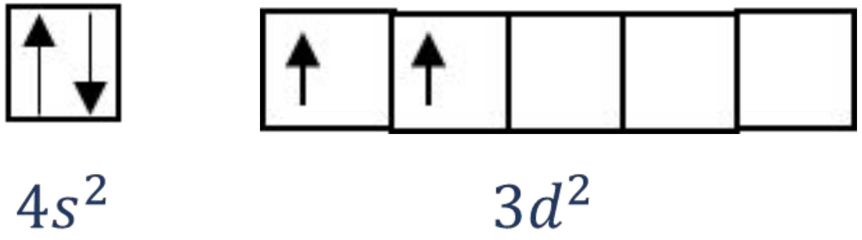
Example : Sc (Z = 21) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^1$

The elements of group IIIB: $ns^2\ (n-1)d^1$ (Paramagnetic element).

Group IIIB belongs to the D block.

✔ Group IVB

Example : Ti (Z = 22) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^2$



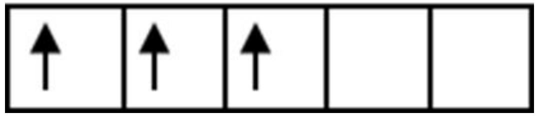
The elements of group IVB: $ns^2\ (n-1)d^2$ (Paramagnetic element).

Group IVB belongs to the D block.

Group VB

✓ Example : V (Z = 23) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^3$

$4s^2$



$3d^3$

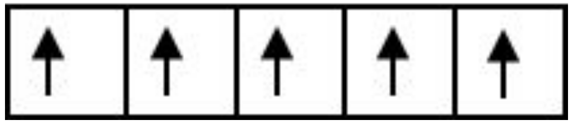
The elements of group VB: $ns^2\ (n-1)d^3$ (Paramagnetic element).

Group VB belongs to the D block.

✔ Group VIB

Example : Cr (Z = 24) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^1\ 3d^5$

$4s^1$



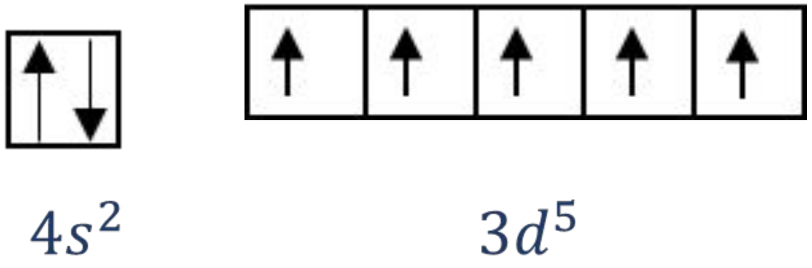
$3d^5$

The elements of group VIB: $ns1\ (n-1)d^5$ (Paramagnetic element).

Group VIB belongs to the D block.

✔ Group VIIB

Example : Mn (Z = 25) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^5$



The elements of group VIIB: $ns^2\ (n-1)d^5$ (Paramagnetic element).

Group VIIB belongs to the D block.

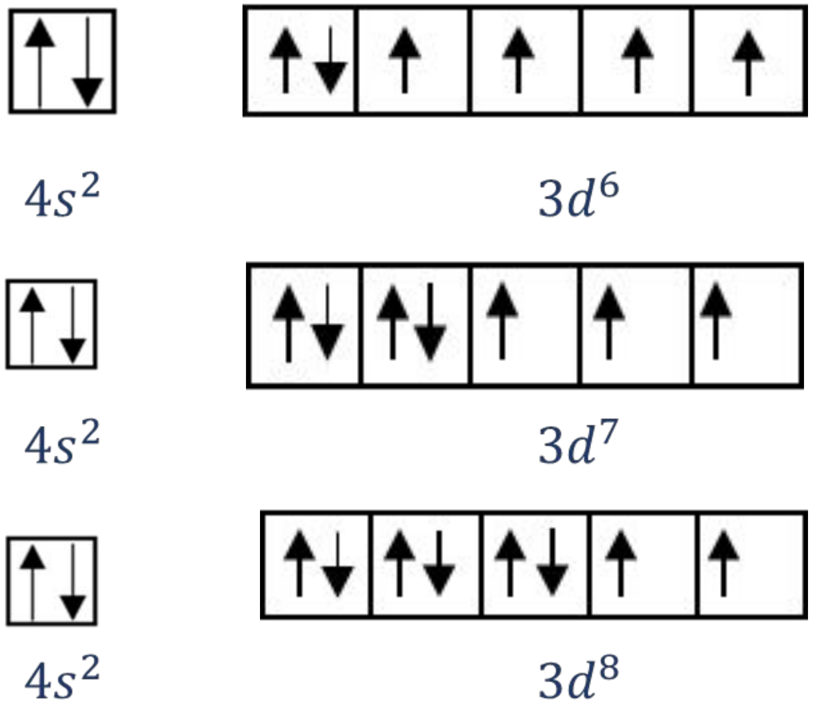
✔ Group VIIIB

Example :

Fe (Z = 26) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^6$

Co (Z = 27) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^7$

Ni(Z = 28) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^8$



The elements of group VIIB*:

$ns^2\ (n-1)d^6$ (Paramagnetic element).

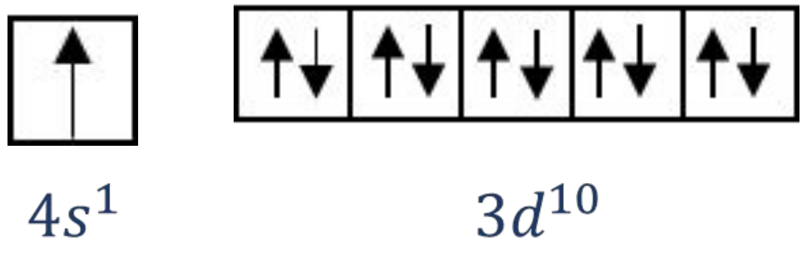
$ns^2\ (n-1)d^7$ (Paramagnetic element).

$ns^2\ (n-1)d^8$ (Paramagnetic element).

Group VIIIB belongs to the D block.

✔ Group IB

Example : Cu (Z =29) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^1\ 3d^{10}$

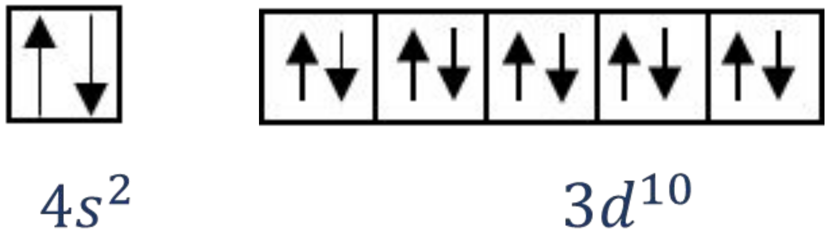


The elements of group IB: $ns^2\ (n-1)d^{10}$ (Paramagnetic element).

Group IB belongs to the D block.

✔ Group IIB

Example : Zn (Z =30) : $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6 4s^2\ 3d^{10}$



The elements of group IIB: $ns^2\ (n-1)d^{10}$ (Diamagnetic element).

Group IIB belongs to the D block.

II/ The Periodic Table and the Properties of Elements

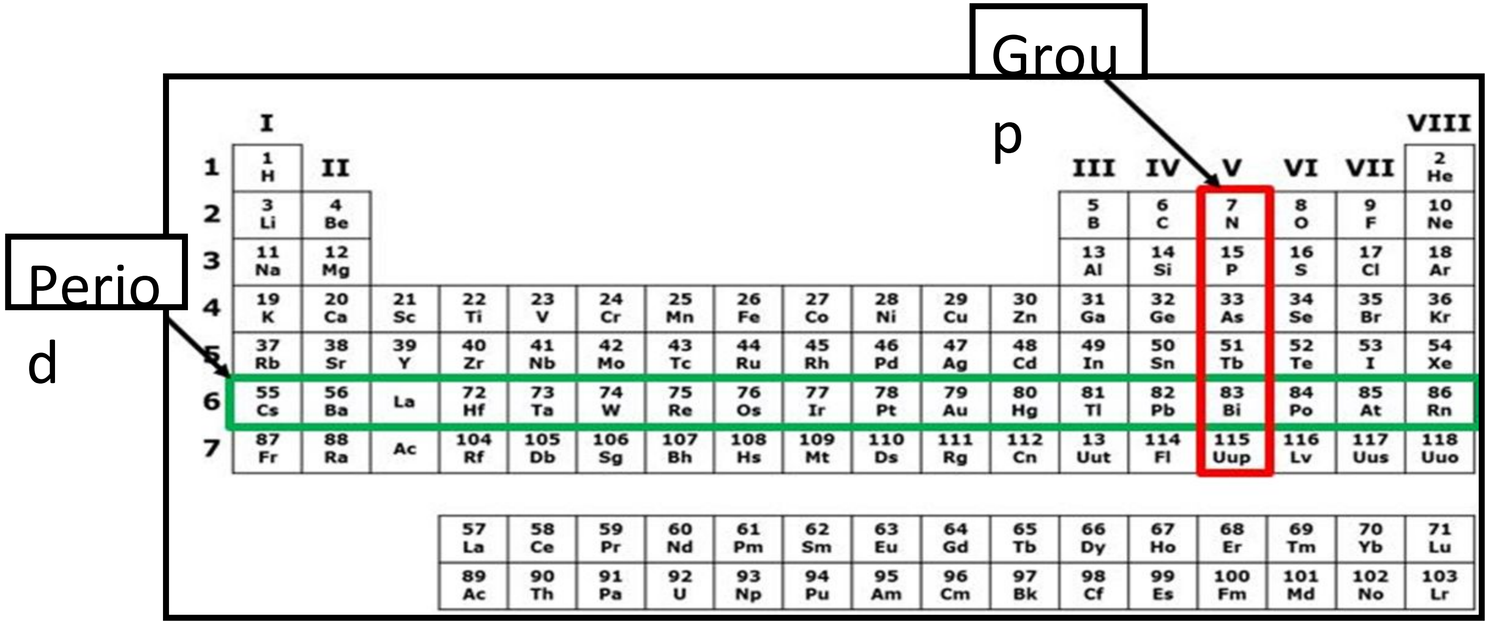
II.1/The Periodic Table

The classification proposed by Dmitri Ivanovich Mendeleev

- ✔ In a period, the elements have the same value of n (the principal quantum number) for their valence shell.
- ✔ In a group, the elements have the same valence configuration.

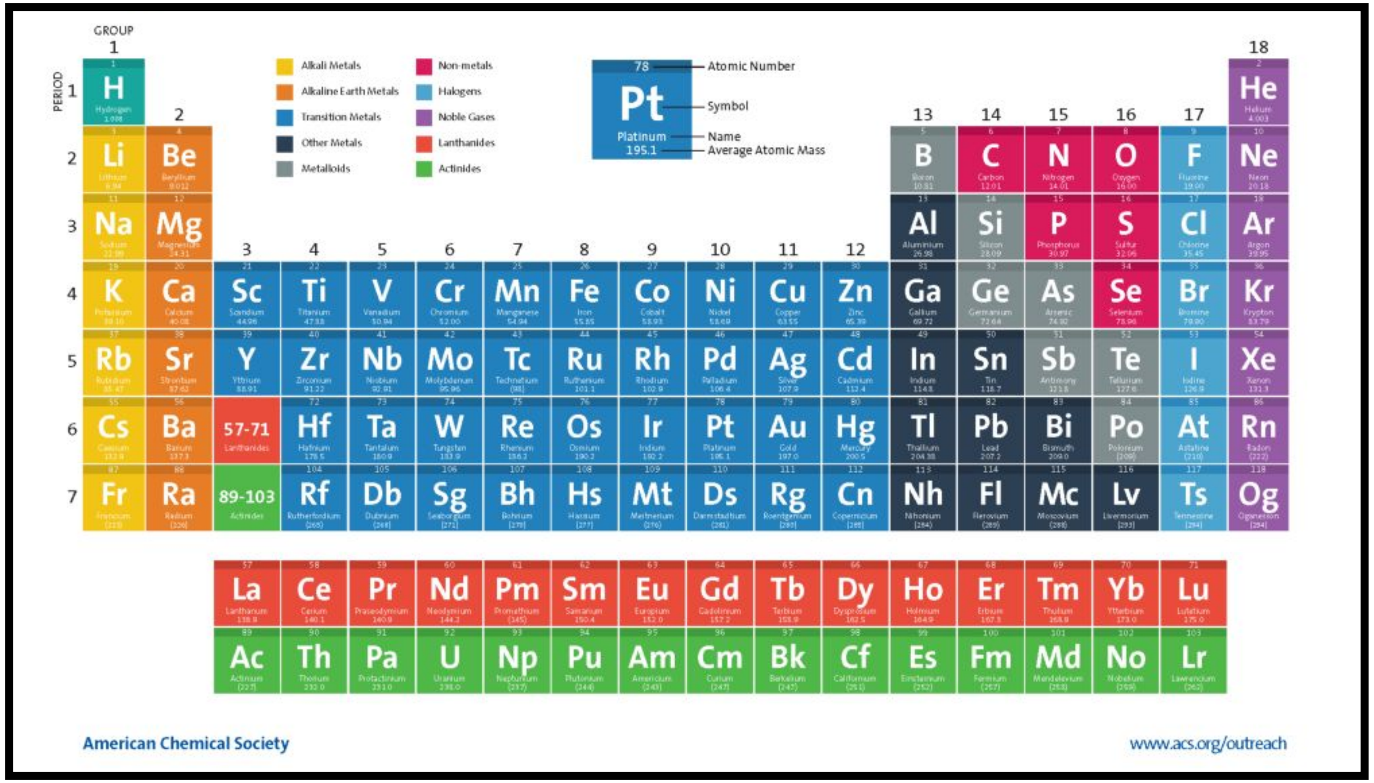
Note

Helium belongs to the S block but is classified in the P block because it has the same chemical properties as the elements in the last column.



The periodic table contains 7 periods and 18 columns (the f block is set apart).

- ✔ **IA and IIA** → S block: ns^1 and ns^2.
- ✔ **IIIA, IVA, VA, VIA, VIIA, and VIIIA** → P block: $ns^2\ np^x$.
- ✔ **IB, IIB, IIIB, IVB, VB, VIB, VIIB, and VIIIB** → D block: $(n–1)d^x\ ns^2$.
- ✔ **4f: Lanthanides** (14 elements).5f: **Actinides** (14 elements). → F block: $(n–2)f^x\ (n–1)d^y\ ns^2$.



II.2/ The families of elements

- **Alkali metals** (Group IA or 1): Very reactive, especially with water. Belong to the **s** block.
  Examples: Lithium (Li), Sodium (Na), Potassium (K)

- **Alkaline earth metals** (Group IIA or 2) : Less reactive than alkali metals. Belong to the **s** block
  Examples: Magnesium (Mg), Calcium (Ca), Barium (Ba).

- **Halogens** (Group VIIA or 17): Very reactive, form salts with metals. Belong to the **p** block
  Examples: Fluorine (F), Chlorine (Cl), Iodine (I)

- **Noble gases** (Group VIIIA or 18): Chemically inert (do not react easily).Belong to the **p** block.
  Examples: Helium (He), Neon (Ne), Argon (Ar)

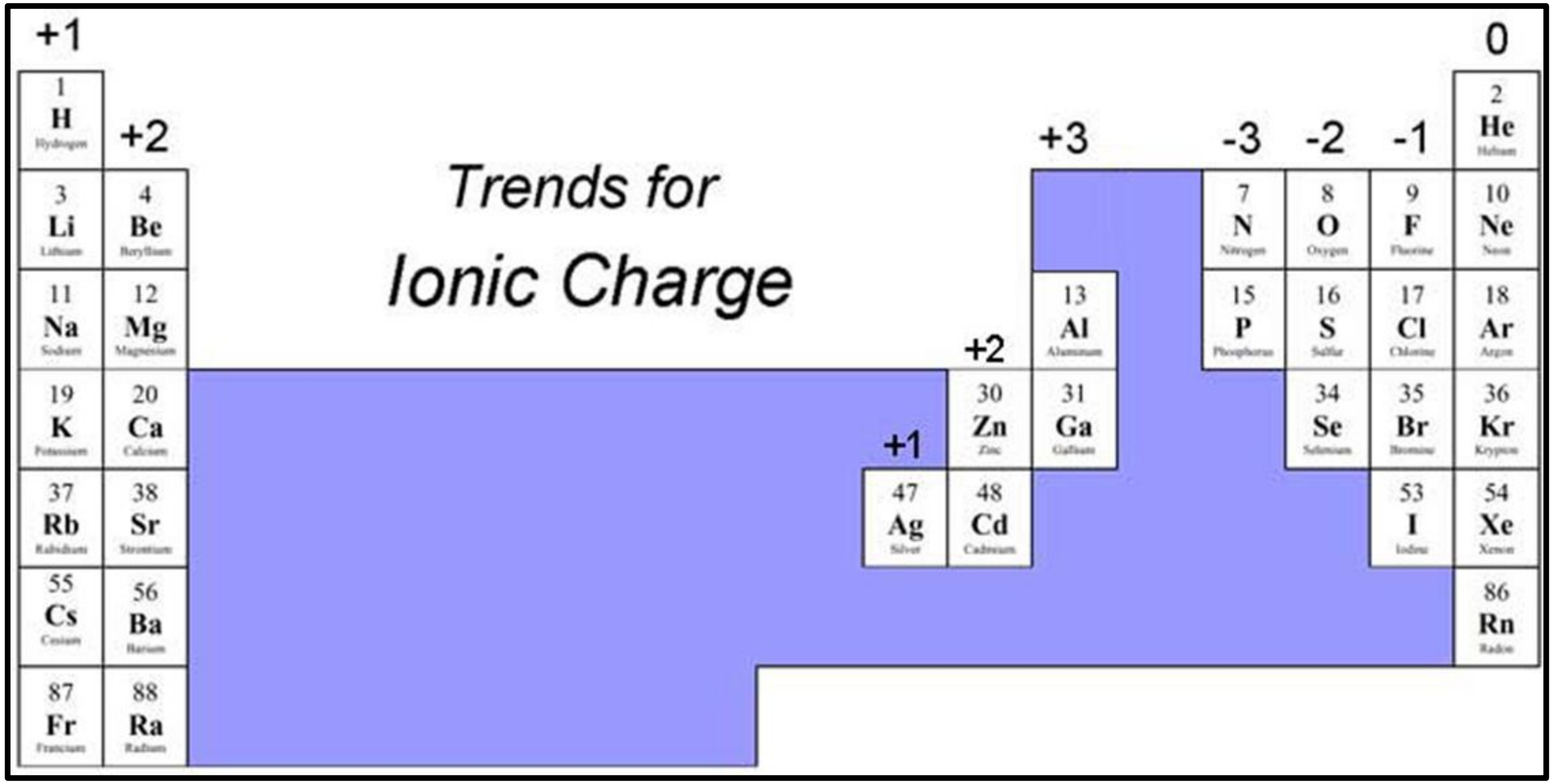
- ✔ **Transition metals** (Groups IIIB to IIB or 3 to 12): Good conductors of heat and electricity. Belong to the **d** block.
  Examples: Iron (Fe), Copper (Cu), Zinc (Zn)

- ✔ **Lanthanides**: from Lanthanum (La) to Lutetium (Lu) → **4f**
- ✔ **Actinides**: from Actinium (Ac) to Lawrencium (Lr) → **5f**
  Metals often used in industry and technology.

II.3/ The charge of monoatomic ions

All the elements in the same family form ions with the same charge.

Elements gain or lose electrons to obtain the (stable) electronic configuration of the nearest noble gas in the periodic table.



II.4/ Trends of covalent radius across the periodic table

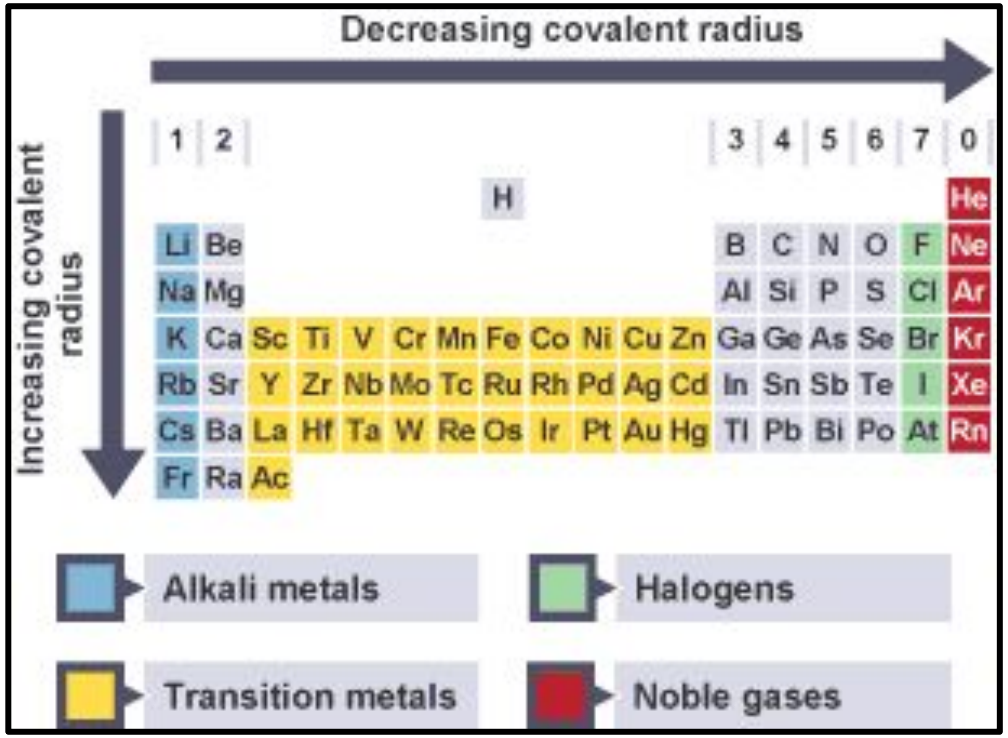
The covalent radius is half the distance between the nuclei of two identical atoms joined by a covalent bond.

The radius decreases when we move from left to right in a period. It decreases because the nuclear charge increases, and the electrons are more attracted to the nucleus.

The radius increases when we move from top to bottom in a group because when we go down the table, extra electron shells are added.

Note

Cations are always smaller than their parent (neutral) atoms because they have one less electron, while anions are larger because one electron is added to them.

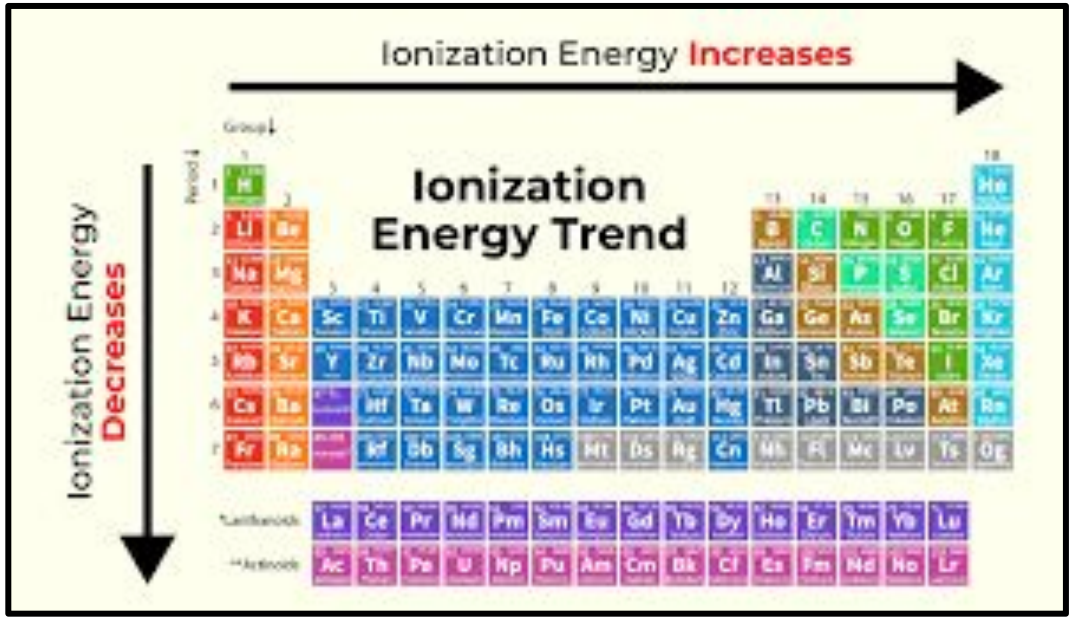


II.4/Trends of ionization energy across the periodic table

Ionization energy is the energy needed to remove one electron from an atom in its gaseous state.

The ionization energy increases from left to right across the table because the atomic radius are smaller, so the electrons are more strongly attracted to the nucleus.

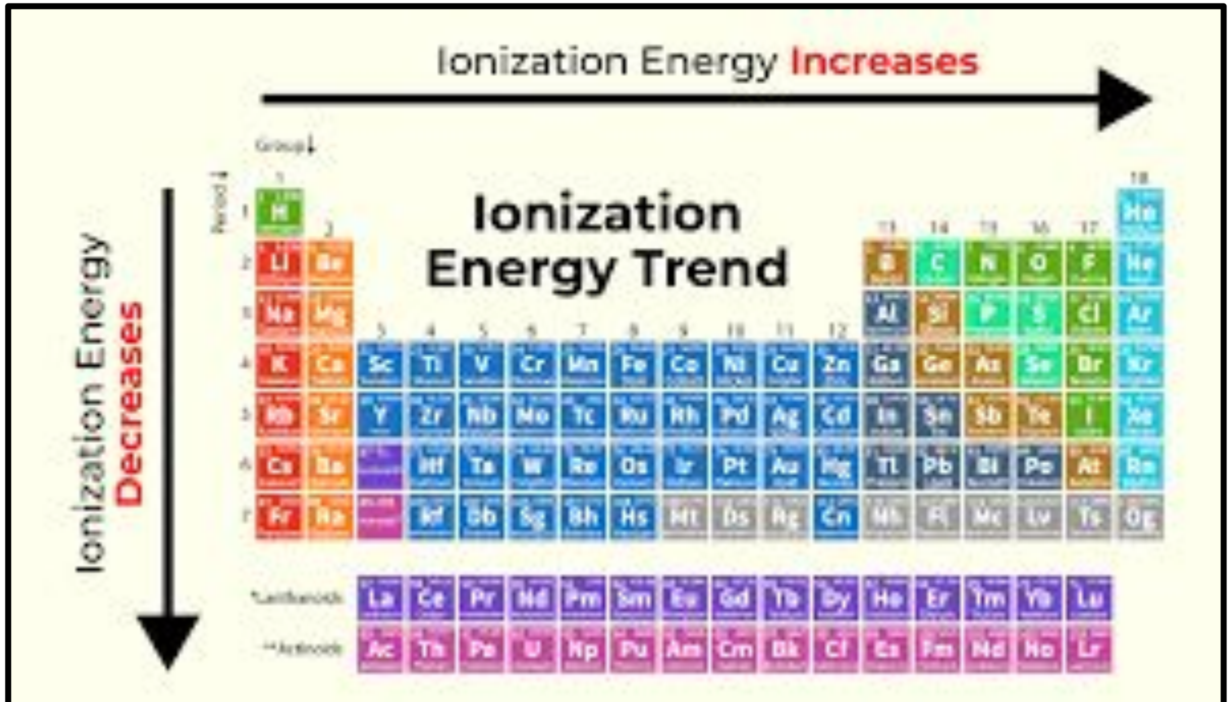
From top to bottom in a group, the ionization energy decreases because the electrons are in shells farther and farther from the nucleus.



The lowest ionization energy is that of the elements in group IA, because when these elements lose one electron, they reach the stable configuration of a noble gas.

The noble gases (ns^2 np^6) (He, Ne, Ar, Kr, Xe, Rn) have high ionization energies. This is normal because the process involves breaking a pair of electrons.

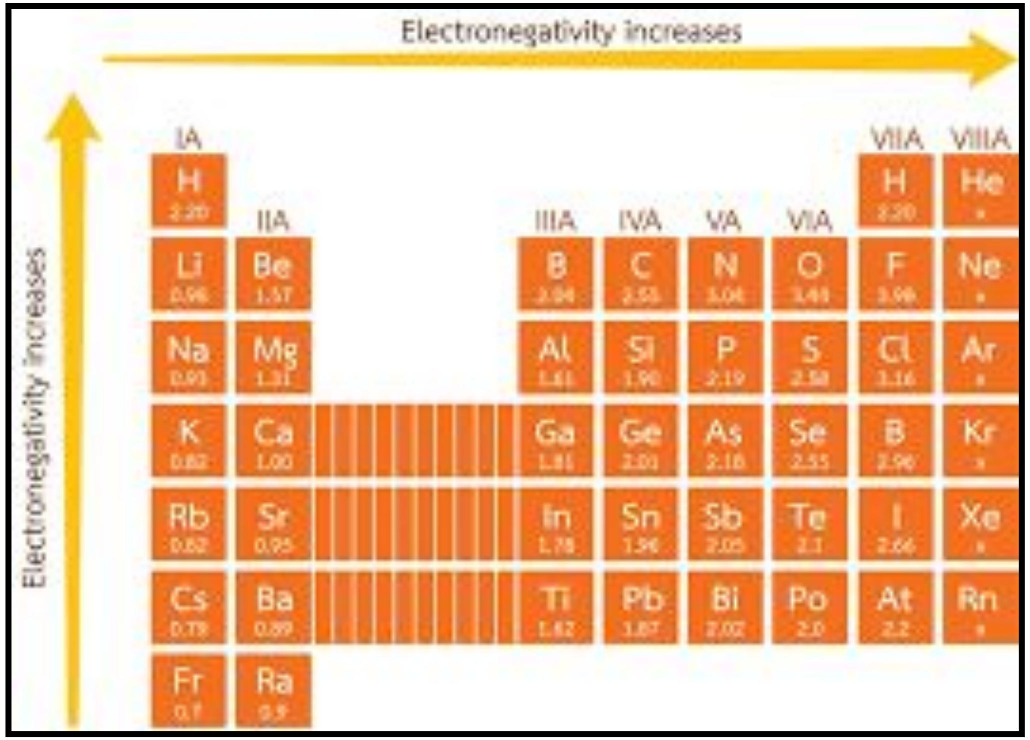
The ionization energy increases when more electrons are removed. This is because the attraction force of the nucleus is stronger in cations.



II.5/ / Trends of electronegativity (χ) across the periodic table

Electronegativity is the ability of an atom to attract the electrons in a chemical bond.

Electronegativity increases from left to right and from bottom to top in the periodic table.



In summary

- ✔ The covalent radius **decreases** from left to right and **increases** from top to bottom.
- ✔ The ionization energy **increases** from left to right and **decreases** from top to bottom.
- ✔ **Cations** are smaller than their parent atoms, while **anions** are larger.
- ✔ **Electronegativity** increases from left to right and from bottom to top.
- ✔ **Noble gases** have very high ionization energies because their electron shells are full.