Lesson 4: Electronic configuration & The periodic Properties of the Elements

Lesson Objective

Dear students!

at the end of the lesson, you will be able to :

- Correlate the electron configurations of elements with their positions inthe periodic table.
- Give a reasonable explanation for the shape of the periodic table.
- Classify elements as representative, transition and inner-transition elements.
- Explain the general trends in atomic radius, ionization energy, electron affinity, electronegativity, and metallic character of elements within a period and group of the periodic table.
- Write the advantages of the periodic classification of elements.
- Demonstrate periodic law and how electronic configurations of atoms are related to the orbital diagrams
- Explain periodic trend describe scientific enquiry skills along this unit: inferring, predicting, classifying, comparing and contrasting, communicating, asking questions and making generalizations.

Brainstorming Question :

why does not an atomic orbital contains more than two electron ?

It contains a single or a pair of electron . Unless it fails the principle.

- What does each box in an orbital diagram represent?
- What are the similarities and differences between a 1s and a 2s orbital?
- What Is the difference between a 2px and a 2py orbital?
- What is the meaning of the symbol 4d6?

key terms / Concepts

- Electron configuration
- Periodic properties
- Ionization energy
- Electron affinities

Electron configuration refers to the arrangement of electrons within an atom or ion, indicating the distribution of electrons among the orbitals and energy levels.

7 Electronic Configurations and Orbitals Diagram

The electronic Configuration for any atom follows the following three principles:

- Aufbau principle
- Pauli’s exclusion Principle
- Hund’s rule

Aufbau principle

- In The Ground State Of The Atoms,the orbitals are filled in order of their increasing energies.
- electrons first occupy the lowest energy orbital available to them and enter into higher energy orbitals only after the lower energy orbitals are filled.
- In General,electrons occupy the lowest-energy orbital available before entering the higher energy orbital.
- The relative energy of an orbital is given by(n+l)rule.
- As(n+l)value increases, the energy of orbital increases.
- The orbital with the lowest (n+l)value is filled firsts
- When two or more orbitals have the same (n+l)value,the one with thelowest‘n’ value (or) highest ‘l ’ value is preferred in filling.

Example 1

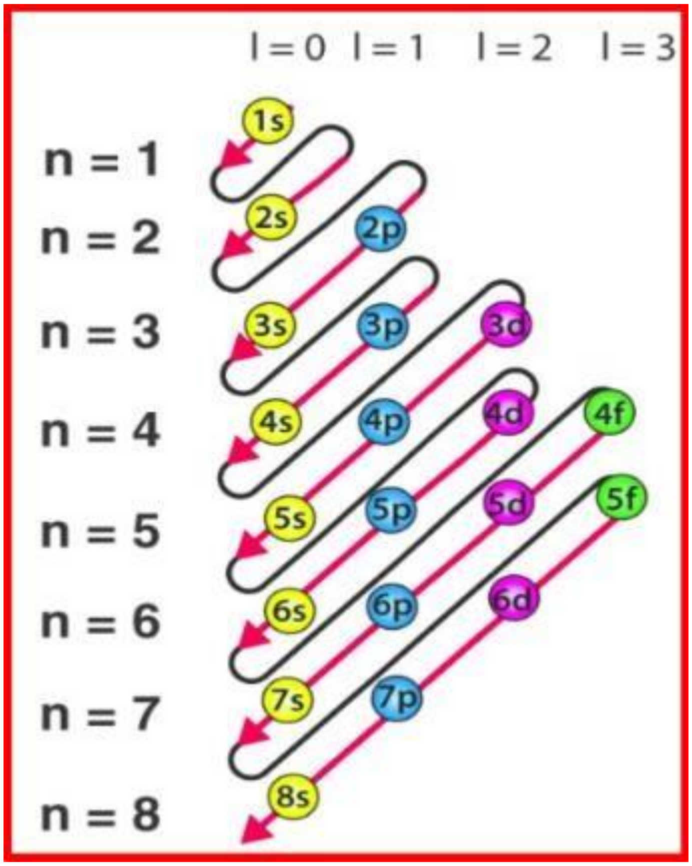
- Consider two orbitals 3d and 4s.n+l value of 3d=3+2=5 and of 4S=4+0=4.Since 4s has lowest (n+l) value, it is filled first before filling taking place in 3d

Example 2, Consider The orbitals 3d,4p and 5,

- The (n+l) value of 3d=3 +2= 5
- The(n+l) value of4p=4+ 1 =5
- The(n+l) value of5s=5+ 0=5

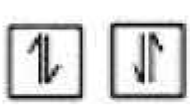
These three values are same. Since the‘n’valueis lower to 3d orbitals, the electrons prefer to entering 3d, then 4p and 5s.

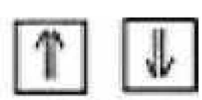
- The Order Of Increasing energy of atomic orbitals Is 1s<2s <2p<3s <3p<4s <3d<4p <5s<4d<5p <6s<4f <5d<6p<7s andso on
- The Sequence In which the Electrons Occupy Various orbitals can be easil remembered with the helpes of Afuba diagrama shown bellow:



2 Paulis Excluision Principle

- stated asked “No two electrons in an atom can have the same set of values for all the four quantum numbers”
- two electrons in an orbital may have the same number ,same three of the quantum number but differing spin quantum number.In an orbital if one electron has clockwise spin, the other has anticlockwise spin.
- It follows that an orbital can hold a maximum of two electrons with opposite spins.
- One orbital cannot have more than two electron ms
- The maximum capacity holding electron in of amain energy shell is $2n^2$
  Example-





lectronic configurationof N(7)is 1s2s22p3
The electrons in 2p subshell are occupied singally.i.e.,1s22s22px1 2py12pz1
3. Hund’sprinciple.?
In degenerate orbital (the orbital in same energy level) electrons electron enter as single before paring up.

Ground State Electronic Configuration Of The Elements

- Two General methods are used to denote electronic cofiguration

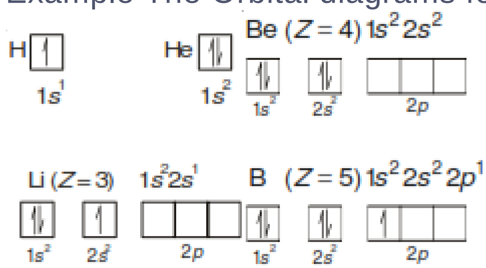
- 1.The Subshell (sublevel) method-notation uses numbers to designate the principal energy levels or principal quantum number and the letters s, p, d and f to identify the sublevels or subshell

- Example , hydrogen(H;Z=1)–1s1
- helium(He;Z=2)–1s2
- Lithium(Li; Z=3)–1s22s1

2. Orbtal diagram method -,

which consists of a box (or circle, or just aline)for each orbital available in a given energy level, grouped by sublevel, with an arrow indicating the electron's presence and its direction of spin.

- Example The Orbital diagrams for the first three elements are



Note:The expected configuration of,chromium and copper based on theaufbau principle,is not the ones observed through the emission spectra and magnetic properties of the elements.

	Expected	Observed
Cr (Z = 24)	[Ar] $4s^2 3d^4$	[Ar] $4s^1 3d^5$
Cu (Z = 29)	[Ar] $4s^2 3d^9$	[Ar] $4s^1 3d^{10}$

ason:The presence of half-filled and completely filled degenerate orbitals gives greater stability to atoms.

Electronic Configuration some Elements

atoms.

B	C	D
H	$1s^1$	$1s^1$
He	$1s^2$	$1s^2$
Be	1s22s2	[He]2s2
B	1s22s22p1	[He]2s22p1
C	1s22s22p2	[He]2s22p2
N	1s22s22p3	[He]2s22p3
O	1s22s22p4	[He]2s22p4
F	1s22s22p5	[He]2s22p5
Ne	1s22s22p6	[He]2s22p6
Na	1s22s22p63s1	[Ne]3s1
Mg	1s22s22p63s2	[Ne]3s2
Al	1s22s22p63s23p1	[Ne]3s23p1
Si	1s22s22p63s23p2	[Ne]3s23p2
P	1s22s22p63s23p3	[Ne]3s23p3
S	1s22s22p63s23p4	[Ne]3s23p4
Cl	1s22s22p63s23p5	[Ne]3s23p5
Ar	1s22s22p63s23p6	[Ne]3s23p6
K	1s22s22p63s23p64s1	[Ar]4s1
Ca	1s22s22p63s23p64s2	[Ar]4s^{2}
Sc	1s22s22p63s23p63d14s2	[Ar]3d^{1}4s2
Ti	1s22s22p63s23p63d24s2	[Ar]3d^{2}4s2
V	1s22s22p63s23p63d34s2	[Ar]3d^{3}4s2
Cr	1s22s22p63s23p63d54s1	[Ar]3d^{5}4s1
Mn	1s22s22p63s23p63d54s2	[Ar]3d^{5}4s2
Fe	1s22s22p63s23p63d64s2	[Ar]3d^{6}4s2
Co	1s22s22p63s23p63d74s2	[Ar]3d^{7}4s2
Ni	1s22s22p63s23p63d84s2	[Ar]3d^{8}4s2
Cu	1s22s22p63s23p63d104s1	[Ar] 3d^{10}4s1
Zn	1s22s22p63s23p63d104s2	[Ar] 3d^{10}4s2

1.8 Electronic Configurations AND THE PERIODIC Table Of The Element

- periodic law states that certain sets of physical and chemical properties recur at regular intervals(periodically)when the elements are arranged according to increasing atomic number.
  1.8.1 Classification Of The Elements

- Elements are placed in the periodic table in accordance to valence electron entering the orbital of lowest energy.
- There are 18 groups and 7 periods in the modern periodic table.
  The Table Is thus divided into 4 blocks namely–s,p,d,f blocks.

- depending on the occupation of the respective orbitals by the valence electrons of an element.
  Representative or main group elements:

- These consistof all s-and p-blockel ements.
- The chemical properties of the representative elements are determined by the number of valence electrons in their atoms.
- s-blockelements:
- The Valence Electrons In The These elements are only in the s-orbital.
- There can bemaximum 2 electrons in the valence shell as s-orbital can only accommodate 2 electrons in it.
  Example:Li(Z=3)1s22s1.
  Na(Z=11)1s22s22p63s1.
  K (Z=19) 1s2 2s2 2p6 3s2 3p6 4s1. Ca(Z=20)1s22s22p63s23p64s2 Mg (Z= 12) 1s22s22p63s2
  Alkali(Group1) and Alkaline Earth Metals(Group2) and Helium are s-block elements all s-block elements.

p-block elements:

The valence electrons occupy p-orbitals
i.e the lastelectron enters the p-orbital.

- Valance electrons in P -Orbital

Example: Carbon 1s22s22p2.
Silicon 1s22s22p63s23p2. Chlorine1s22s22p63s23p5.
Argon 1s22s22p63s23p6.

Transition elements:

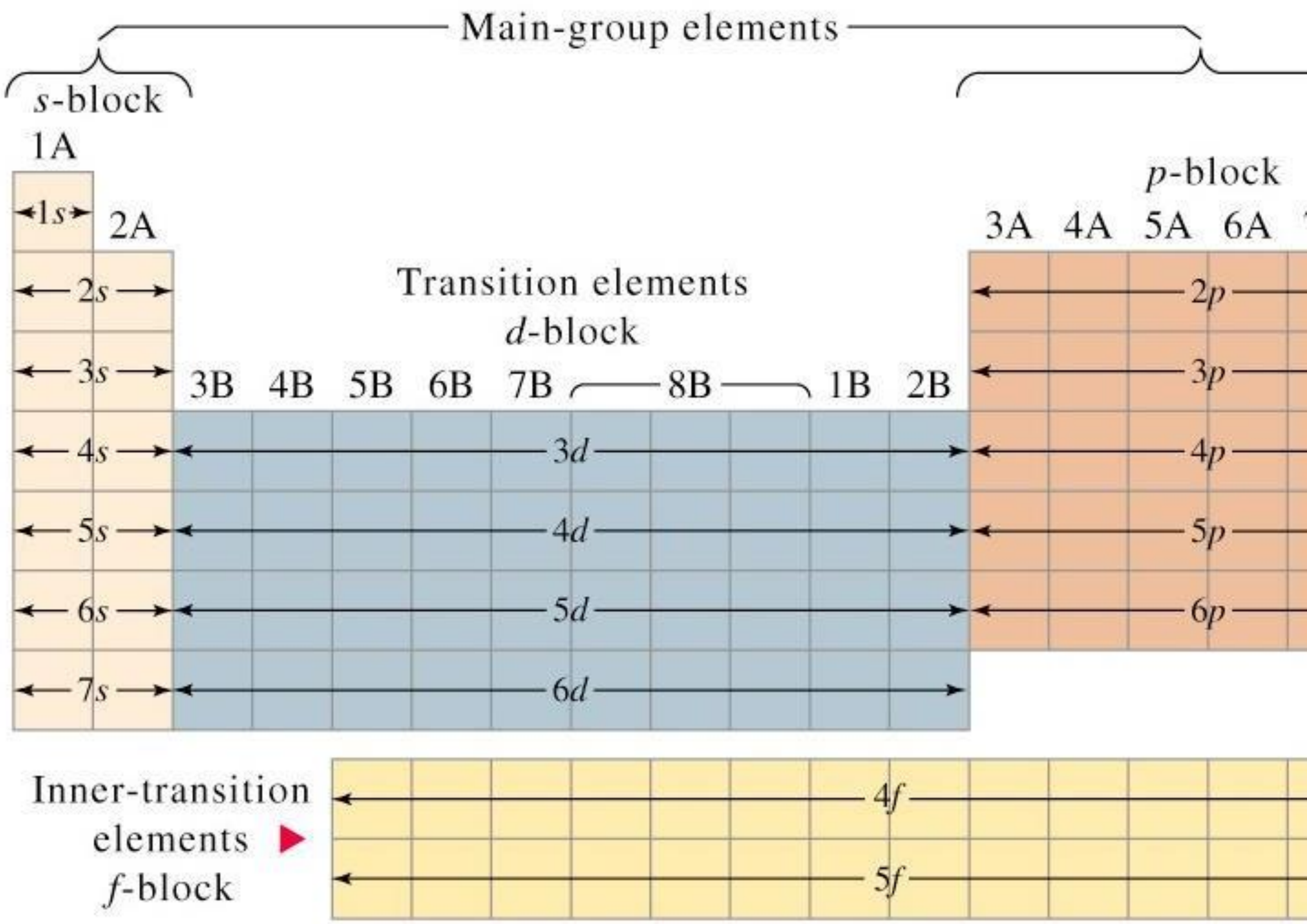
These are d-block elements.

- There are four series of transitional elements, 3d, 4d, 5d and 6d depending on the energylevels of d-orbitals.

- Intheseelements,thepenultimate(n-1)d-orbitalisfilledwithelectrons.
- the energy of the nsshell is always less than (n-1) d shell. So the electrons fill the nsshell first and later they enter the penultimate (n-1)d orbital.
- Example:energyof4sorbital<energyof3dorbital.So,electronsenter4sorbitalfirstand3d orbital later.
- Theorderoffillingofelectronsinorbitalsis1s2s2p3s3p4s3d4p5s4d5p6s4f5d6p 7s 5f 6d
- Theseelementshavethushavevariableoxidationstates.
- e.g.Ironioncanexistintwostates,Fe+2andFe3+
- The general outer electronic configuration of transition elements is (n-1)d1-10ns1-2.
- Transition Metals Have Very Strong Metallic Bonds And Thus Have High Melting And Boiling points(except Zn,Cd and Hg – on account of completely filled orbitals).
- There are4 sets of transition series–
- First Transition Series–3d Orbital Get sfilled–Scandium(Z=21)toCopper(Z=29).
- Second Transition Series–4d orbital gets filled–Yttrium(Z=)toSilver(Z=)
  Third transition series–5d orbital gets filled – Lanthanum (Z=) to Hafnium (Z=) except Cerium to Lutetium.
  Fourth Transition Series–6d orbital gets filled–Actinum(Z=)to(Z=)–Thisisan incomplete period.

Inner Transition elements:
are f-block elements

- These are called the inner transition elements and they are placed separately at the bottom of the main periodic table.
- There are two series of f-block elements,4f and 5f series called lanthanides and actinides, respectively.
- The periodic table is unable to include the inner transition elements inits mainframe.
- The two Inner transition Series are–
- Lanthanide series–4f orbital gets filled–Cerium(Z=58)toLutetium (Z=71)
- Actinide series–5f orbital gets filled–Thorium(Z=90)to Lawrencium(Z=103)-after Uranium(Z=92), the other elements in this group are synthesized synthetically.
- All actinides are radioactive.
- The general outer electronic configuration of Actinidesis5f1-146d0-17s2



1.8.2Periodic Properties

Atomic Size (AtomicRadii)-

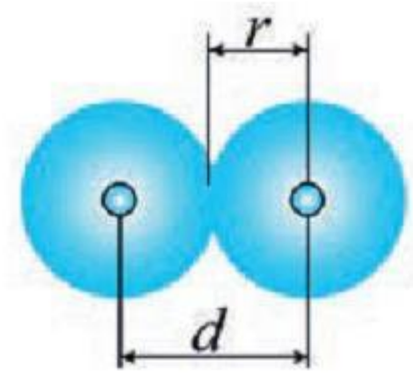


Figure 1.19: Atomic radius for a diatomic molecu

$$r = \frac{d}{2}$$

- is the distance between the nuclei of two adjacent atoms,
- atomic radii increases from top to bottom within a group of the periodic table(each succeeding member has one more principal shell occupied by electrons)
- **Atomic radii of the in main group element atomic size tend to decrease from left to right in a periodof the periodic table (The effectivenuclearcharge increases from left to right which cause decrease in size).**

Example From left to right,size of atoms decrease in the periodic table.

Ionization Energy

$M(g) + IE \rightarrow M^{+}(g) + e^{-}$ $Atom(g) \rightarrow Ion^{+}(g) + e^{-} \quad \Delta E = IE_1 > 0$

$Ion^{+}(g) \rightarrow Ion^{2+}(g) + e^{-} \quad \Delta E = IE_2 \quad (always > IE_1)$

- The Ionization Energy(IE)is the amount of energy required to remove the outermost electron in an isolated gaseous atom or ion.
- Itis expressed in kJmol–1(kilojulespermole)
- first ionization energy(IE1) – is the energy needed to remove anelectron from the highest occupied sublevel of the gaseous atom.
- -The first ionization energy is a key factor in an elements chemical reactivity because, atoms with a low IE1 tend to form cations during reactions, whereas those with a high IE1 (except the noble gases) often form anions.
  second ionization energy (IE2) removes the second electron.IE2is always larger than IE1

$B\ (g)$	$\rightarrow$	B^{+}	$(g) + e^{-}$	IE_1
$B^{+}\ (g)$	$\rightarrow$	B^{2+}	$(g) + e^{-}$	IE_2
$B^{2+}\ (g)$	$\rightarrow$	B^{3+}	$(g) + e^{-}$	IE_3
$B^{3+}\ (g)$	$\rightarrow$	B^{4+}	$(g) + e^{-}$	IE_4
$B^{4+}\ (g)$	$\rightarrow$	B^{5+}	$(g) + e^{-}$	IE_5

The Variation In the magnitude of Ionization energy of elements in the periodic table is mainly dependent on the following factors:

- The size of the atom (inverse relation IE and atomic size.)
- The magnitude of the nuclear Charge on the atom,
- The extent of screening the type of orbital involved(s,p,d, or f), s>p >d>f
- In general, ionization energy increases across a period; it is easier to remove an electron from an alkali metal than from a noble gas.
  Except:

- The first ionization energy of B(boron) is lower than that of Be(beryllium), Boron,Z =5(1s22s22p) has smaller IE1
  Reason: The 2p electron of boron is at a higher energy than the 2s electron of beryllium and is therefore easier to remove. Thiskind of discontinuity occurs generally in proceeding from a GroupIIA or Group IIB elements to a Group IIIA element.
  The Ionization Enthalpy of Al (aluminium)is lower than that of Mg(magnesium); the first ionization enthalpy of O(oxygen)is lower than that of N(nitrogen).

Electron Affinity (EA)

$$X_{(g)} + e^{-} \longrightarrow X^{-}_{(g)}$$

iexample.

$$He(g) + e^{-} \not\longrightarrow He^{-}(g) \qquad EA = \quad 0\ kJ/mol$$

$$Cl(g) + e^{-} \longrightarrow Cl^{-}(g) + 349\ kJ \qquad EA = -349\ kJ/mol$$

Electron affinity is the energy released for one mole of neutral atoms in a gaseous state when electron is accepted by each atom
Electro affinity increase from left to right and decrease down a group.

ELECTRONEGATIVITY

Table 1.4: (b) Electronegativity values (on Pauling scale) down a grou

Atom (Group IA)	Electronegativity value	Atom (Group VIIA)	Elect
Li	1.0	F	
Na	0.9	Cl	
K	0.8	Br	
Rb	0.8	I	
Cs	0.7	At	

Is the ability of an atom to atract the shared electron towards itself.

FON, The highest electron negative element, F, N, O in decreasing order. electronegativity decreases downing any group and increases from left to right in the period.

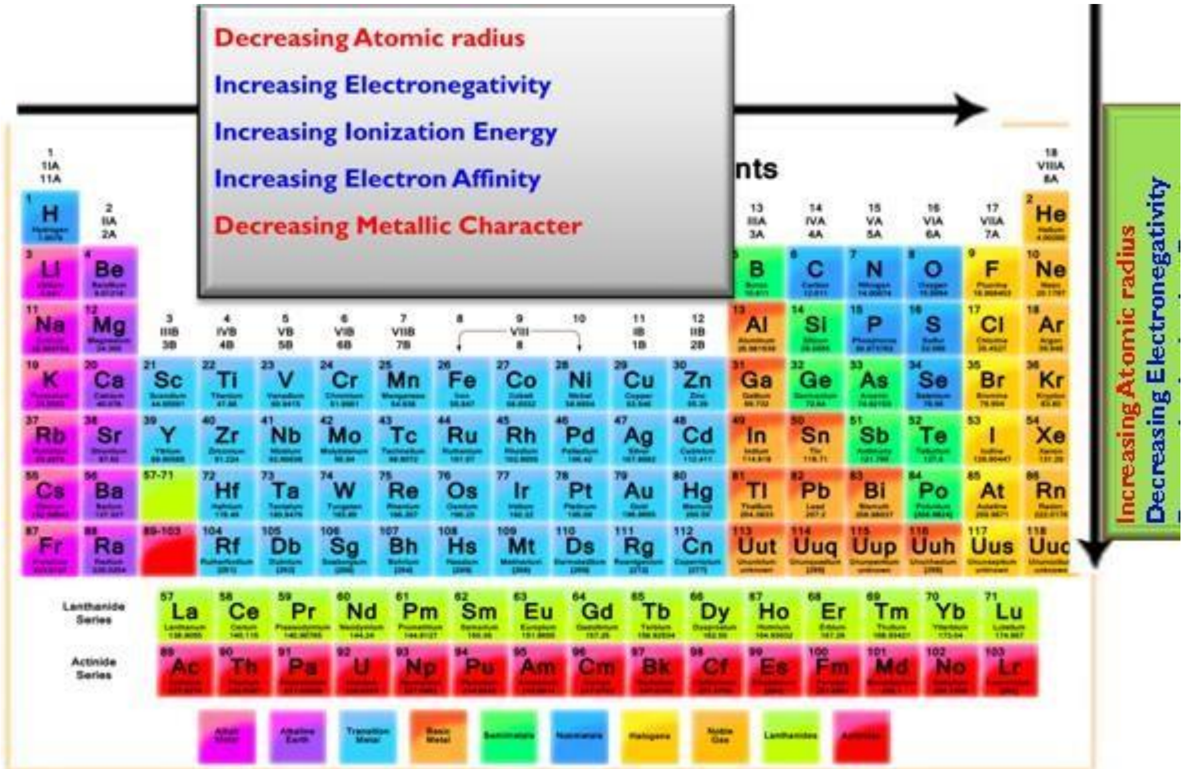
Metallic Character

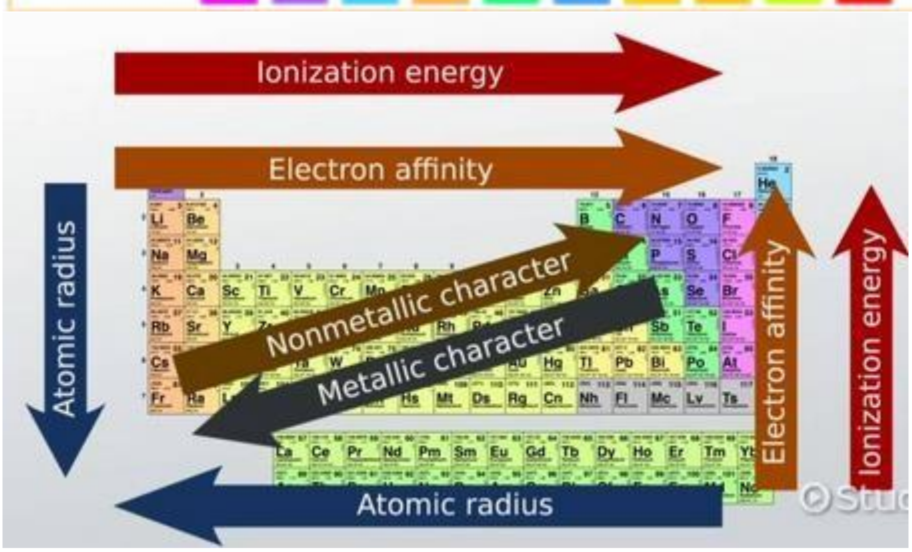
- Metallic Character refers to the Chemical Properties Associated With elements classified as metals.
- These properties arise from the elements ability to lose electrons
- the metallic character decreases from left to right and increase down agroup

Advantages of Periodic Classification of the Elements

- The classification of elements is based on the atomic number,which is a fundamental property of an element.
- The reason for placing isotopes at one place is justified as the classification is onthebasis of atomic number
- It explains the periodicity of the properties of the elements and relates them to their electronic configurations.
- The position of the elements that were misfits on the basis of mass number (anomalous pairs like argon and potassium) could be justified on the basis of atomic number.
- The lanthanides and actinides are placed separately at the bottom of the periodic table.

Summary





Lesson 4: Summary

- **Electronic configuration** refers to the arrangement of electrons in atoms or ions. It specifies the distribution of electrons among energy levels (shells) and sub-levels (orbitals) based on quantum mechanical principles.
  Quantum Numbers Involved:
- **Principal Quantum Number ((n))**: Specifies the main energy level or shell (1, 2, 3, …).

- **Angular Momentum Quantum Number ((l))**: Determines the shape of the orbital (s, p, d, f).
- **Magnetic Quantum Number ((m_l))**: Specifies the orientation of the orbital in space.
- **Spin Quantum Number ((m_s))**: Describes the spin state of electrons (+1/2 or -1/2).

Aufbau Principle:

- Electrons fill orbitals in order of increasing energy.
- Follows the sequence: (1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, …).

Pauli Exclusion Principle

- No two electrons in an atom can have the same set of four quantum numbers (n, l, m_l, m_s).

Hund's Rule:

- Electrons occupy orbitals singly before pairing up, and all singly occupied orbitals have the same spin.

Examples

- **Carbon (C)**: ($1s^2 2s^2 2p^2$)
- **Neon (Ne)**: ($1s^2 2s^2 2p^6$)
- **Chlorine (Cl)**: ($1s^2 2s^2 2p^6 3s^2 3p^5$)

Periodic Properties of Elements:

Atomic Radius:

Electron Affinity

- **Definition**: Energy change when an electron is added to a neutral atom to form a negative ion.

Electronegativity

- **Definition**: Measure of the ability of an atom to attract electrons in a chemical bond

Metallic Character

- Property of elements associated with the ability to lose electrons and form positive ions (cations)

Chemical Reactivity:

- **Metals**: Tend to lose electrons and form positive ions (cations).
- **Non-Metals**: Tend to gain electrons and form negative ions (anions)
- **Electronic configuration** defines how electrons are distributed in atoms based on quantum mechanical principles (Aufbau, Pauli exclusion, Hund's rules).
- **Periodic properties** describe trends across the periodic table related to atomic size (radius), tendency to gain or lose electrons (ionization energy, electron affinity), ability to attract electrons (electronegativity), and behavior in chemical reactions (metallic character, reactivity).