
GENERAL CHEMISTRY
3
Dr M.F Dr A.E

Chapter Two

Electronic Structure

Part (1)

Electronic Structure of Atoms

Define

Electronic Structure

The way that electrons arranged in an atoms

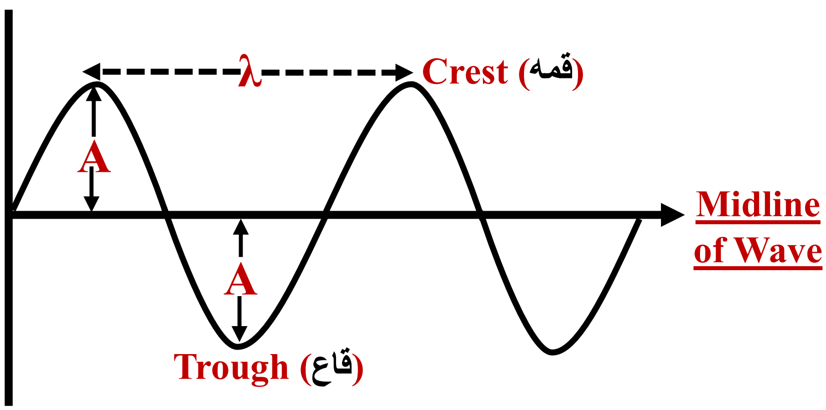
Notes

1. The importance of electronic structure to Know the **Chemical Properties** of any element.

2. The theory of electronic structure is derived from the Experiments that involve Electromagnetic radiation (EMR).

Radiant Energy (Light) (EMR)

Light travel through space in Wave motion (حركة موجيه)



Note

Speed of light in vacuum:

$= 2.997 \times 10^8$ m/s

$\approx 3\times10^8$ m/s

$= 3\times10^{10}$ cm/s

Characterization of Wave

1 Amplitude (A) (السعة)

- The height of the crest (ارتفاع القمة)

OR - The depth of the Trough (انخفاض القاع)

OR - The distance between the top or the bottom of the wave and its midline

Note

- Amplitude depends on the **intensity** of wave (**I**)

2 Wavelength (λ) (الطول الموجي)

- The distance between two **successive** identical points (2Crests or 2 troughs) (متتالية)

Unit

- m, cm, nm or Å

Unit	meter
1 cm (centimeter)	10^{-2} m
1mm (millimeter)	10^{-3} m
1µm (micrometer)	10^{-6} m
1nm (nanometer)	10^{-9} m
1 Å (Angstrom)	10^{-10} m

3 Frequency (υ) (التردد)

- **<u>The number of cycles or waves passing through given point in one second</u>**

Unit

- Sec^{-1} (S^{-1}) ≡ Hz (Hertz)

4 Wavenumber ($\bar{\upsilon}$) (العدد الموجى)

- **<u>The number of or waves per unit distance</u>**

$$\bar{\upsilon} = \frac{1}{\lambda}$$

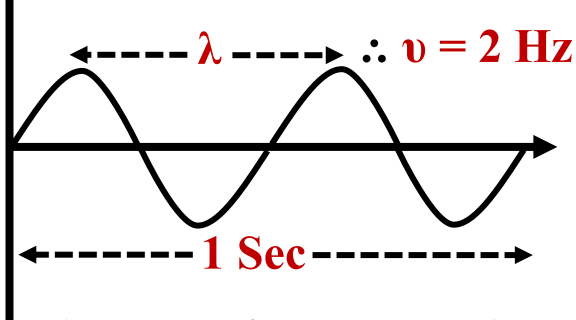
<u>If</u> λ:
- m ∴ $\bar{\upsilon}$ (m^{-1})
- nm ∴ $\bar{\upsilon}$ (nm^{-1})

Unit

- m^{-1}, cm^{-1} or nm^{-1}

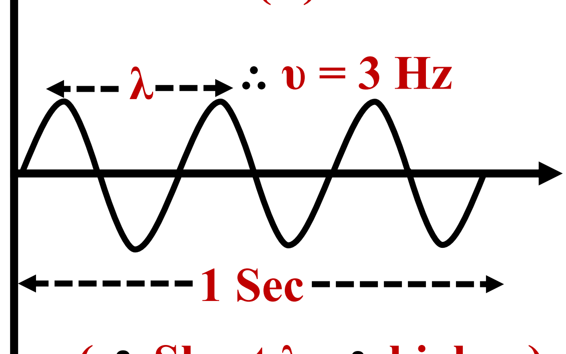
Relation between Frequency and wavelength

(A)



(♣ Long λ ♣ Low υ)

(B)



(♣ Short λ ♣ high υ)

$\because$ $\boxed{\upsilon \,\alpha\, \frac{1}{\lambda}}$

$\therefore$ $\boxed{\upsilon = \frac{C}{\lambda}}$ **OR** $\boxed{C = \upsilon\, \lambda}$ (حفظ)

(C: is the speed of light = 3×10^{8} m/s)

Note

When λ increases:

2- υ decrease   1- $\bar{\upsilon}$ decrease   3- E decrease

(Energy (E) $\propto \upsilon \propto \bar{\upsilon} \propto \frac{1}{\lambda}$)

Types of Electromagnetic radiation (EMR)

E ↓ υ ↓ λ ↑ →

γ-ray	x-ray	UV	Visible	IR	microwave	Radio TV wave

Visible ↓

Violet	Blue	Cyan	Green	Yellow	Orange	Red

- UV: ultraviolet (الاشعه فوق البنفسجية)
- IR: infrared (الاشعه تحت الحمراء)

Problems

1 **The Yellow light given off by sodium lamp has a wavelength equal 589 nm, what is the frequency of the radiation:**

Answer

$$\because\ \upsilon = \frac{C}{\lambda}$$

$$\because\ C = 3 \times 10^{8}\ \text{m/s}$$

$$\because\ \lambda = 589\ \text{nm} = 589 \times 10^{-9}\ \text{m}$$

$$\therefore\ \boxed{\upsilon = \frac{3 \times 10^{8}}{589 \times 10^{-9}} = 5.09 \times 10^{14}\ \text{Hz}}$$

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2 **The light produced from xenon lamp with a wavenumber equal 4.115x 10^{6} m^{-1}, what is the frequency of the radiation:**

Answer

$$\because\ \upsilon = \frac{C}{\lambda} = C\ \bar{\upsilon}$$

$$\because\ C = 3 \times 10^{8}\ \text{m/s}$$

$$\because\ \bar{\upsilon} = 4.115 \times 10^{6}\ \text{m}^{-1}$$

$$\therefore\ \boxed{\upsilon = 3 \times 10^{8} \times 4.115 \times 10^{6} = 1.23 \times 10^{15}\ \text{Hz}}$$

Planck quantum theory :

1 Atoms or molecules can be absorb (gain) or emitt (lose) Energy.

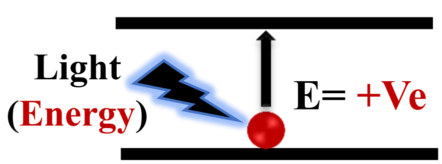
2 The amount of energy absorbed or evolved is quantized.

Define

Quantum

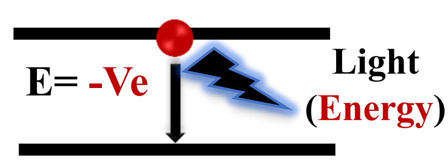
The smallest amount of energy could be absorbed or emitted

Adsorption



(Lower→ Higher)

Emission



(Higher→ Lower)

3 Energy of light (E) :

$\because E \propto \upsilon \quad \therefore \boxed{E = h\,\upsilon}$

Unit

- Energy (E): Joule
- Frequency (υ) : Hz or S^{-1}
- Plank constant (h)= 6.63×10^{-34} J.S

Planck Law

$$E = h\upsilon = h\frac{C}{\lambda} = hC\bar{\upsilon}$$

❖ The smallest amount of energy (quantum)

Note

According to Planck : energy is always emitted or absorbed in multiples of hυ (hυ, 2 hυ, ...) not 1.67 hυ or 3.98 hυ .

Problems

1 **Calculate the smallest amount of energy that an object can absorb from yellow light, whose wavelength is 589 nm**

Answer

$$E = h\upsilon = h\frac{c}{\lambda}$$

$$E = \frac{6.63x10^{-34} x\ 3x10^{8}}{589x10^{-9}}$$

$$\boxed{E = 3.37x10^{-19}\ J}$$

Note

- 1 Joule = 10^7 erg

∴ Planck's constant = $6.63x10^{-34}$ J.s = $6.63x\ 10^{-27}$ erg.s

Dual Nature of Light by Einstein

1 **Besides the wave Nature of Light, light has a particle nature (photons)**

Law

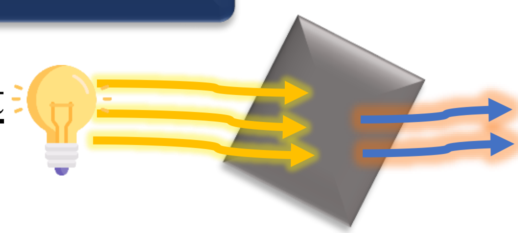
$$E = m \times C^2$$

(m: mass of photon)

∴ Light has a dual nature (Wave Nature + Particle Nature)

2 Photoelectric effect

Light



e,s **(Used in photocells)**

Metal Surface (Cs)

A **When photon strikes the metal, its energy is transferred to electron.**

B **Number of e,s ejected ∝ Intesnity of light**

C **Energy of photons ∝ the Frequency of light**

Define

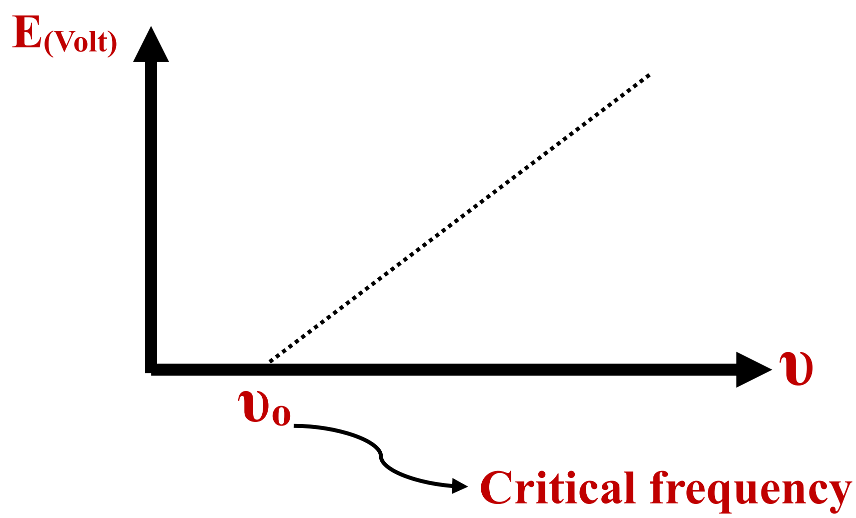
Threshold (critical) Frequency (التردد الحرج)

. The minimum frequency needed to ejected e,s from metal.

OR

. It is frequency above which e,s will be ejected.

Draw the relation between the energy and frequency in photoelectric effect :



∵ E = Ionization Energy + Kinetic Energy

∴ $\boxed{E = h\upsilon_o + \frac{1}{2} mV^2}$

Light and Continuous Spectra

Light

Polychromatic

. Light of different λ,s

. Ex: White light (Sun light)

Monochromatic

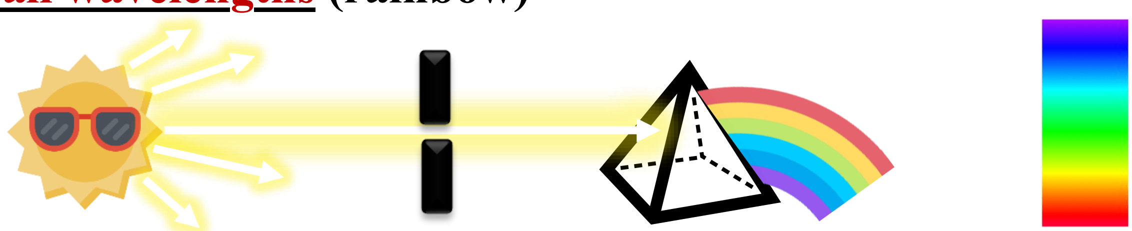
. Light of single λ

. Ex: Laser

Spectra

A Continuous

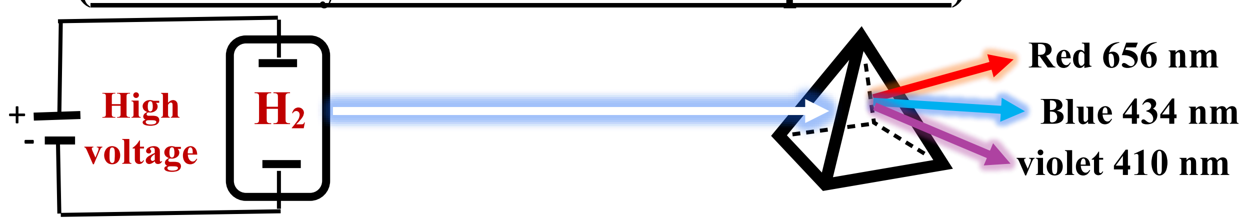
- When white light (sunlight) is passed through Prism it is containing light of all wavelengths (rainbow)



Continuous Spectra

B Line Spectra

- Result from hydrogen lamp due to emission of an excited H-atoms (Exhibit only certain color with a specific λ)



Line spectrum of Hydrogen

- **The emission of line spectra of H-atoms were detected by (Rydberg Equation) :**

Rydberg Equation

$$\bar{\upsilon} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

OR

Rydberg Equation

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

(To determine Wavenumber of Spectra)

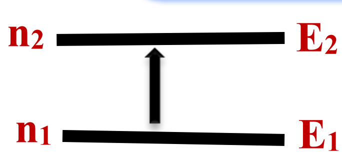
Where

- **n_1 and n_2 : number of energy Level and should be integer number (1,2,3,…..) and should be $n_2 > n_1$**
- **R (R_H) : Rydberg Constant = $1.097x10^5$ cm^{-1} = $1.097x0^7$ m^{-1}**

Important Notes

1. **$n_2 > n_1$ (to give +Ve), due to $\bar{\upsilon}$ and λ always positive values**
2. **$cm^{-1} \xrightarrow{X\ 10^2} m^{-1}$ while $cm \xrightarrow{X\ 10^{-2}} m$**
3. **Energy difference (ΔE)**

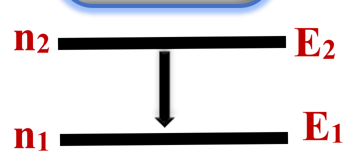
Absorption



∴ Lower (n_1)→ Higher(n_2)

∴ $\Delta E = E_2 - E_1$ = +Ve

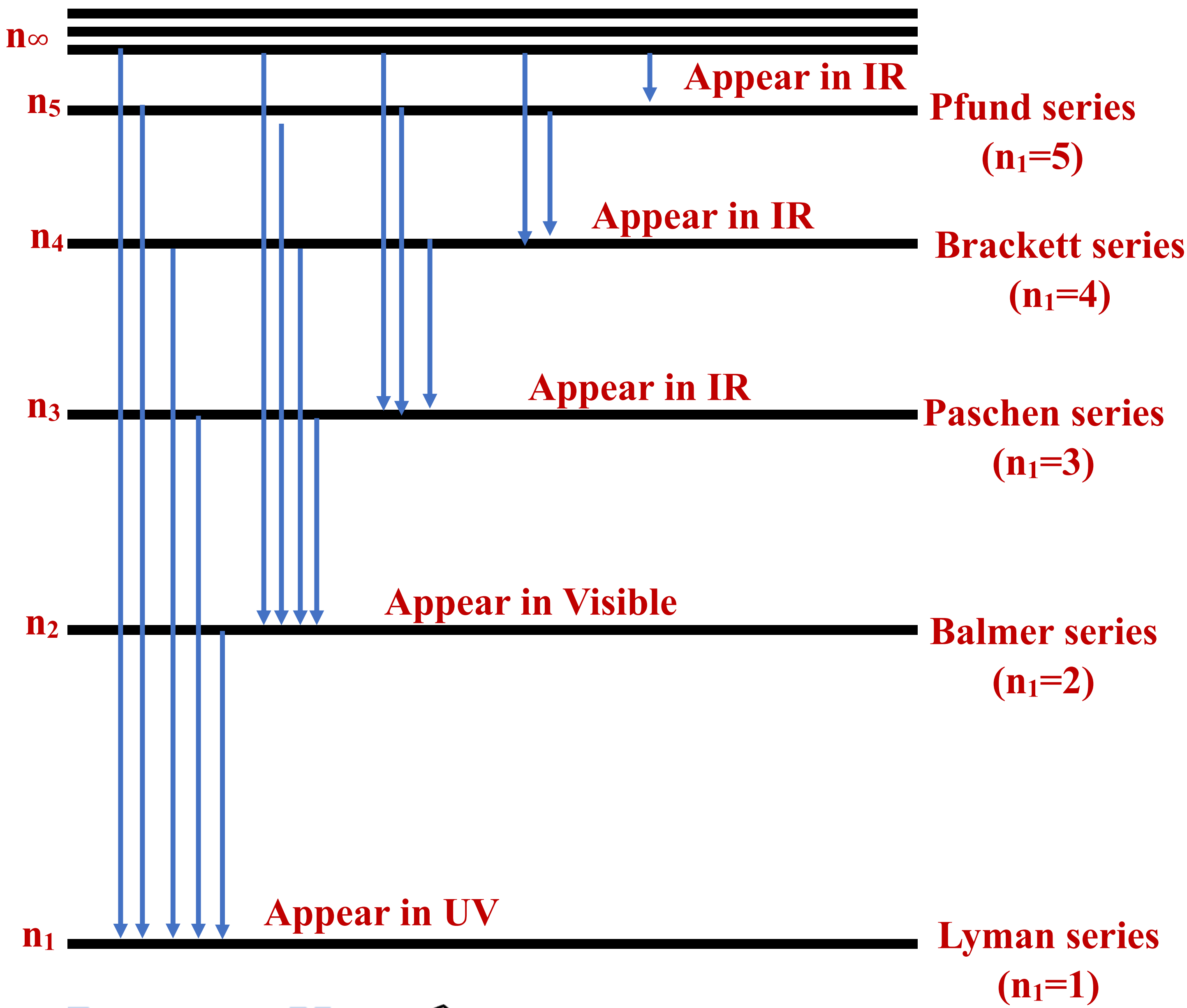
Emission



∴ Higher (n_2)→ Lower (n_1)

∴ $\Delta E = E_1 - E_2$ = -Ve

Emission of Line spectrum of H-atom



Important Notes

- **Energy of Lyman > Balmer > Paschen > Brackett > Pfund**
- **λ of Lyman < Balmer < Paschen < Brackett < Pfund**

Series of line Spectra of H-atom

Series	n_1	n_2
		1st line, 2nd line, 3rd line,..
Lyman	1	2 , 3 , 4
Balmer	2	3 , 4 , 5
Paschen	3	4 , 5 , 6
Brackett	4	5 , 6 , 7
Pfund	5	6 , 7 , 8

Problems

1 **Calculate the wavelength in nm in line spectrum of hydrogen atom corresponding to transition of an electron from n=4 to n=2:** (R_H=1.097 x10^7 m^{-1})

Answer

- $n_1 = 2$ and $n_2 = 4$

$$\because \frac{1}{\lambda} = \overline{\upsilon} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\because \frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{4^2}\right) = 20.57\times10^5\ m^{-1}$$

$$\therefore \lambda = 1/20.57\times10^5\ m^{-1} = 4.86\times10^{-7}\ m$$

$$\therefore \boxed{\lambda = 4.86\times10^{-7}\ m \xrightarrow{\times10^9} 486\ nm}$$

2 **Calculate the wavelength in nm in third line emitted in Paschen series:** ($R_H = 1.097 \times 10^7\ m^{-1}$)

Answer

∵ Paschen ∴ $n_1 = 3$ and ∵ third line ∴ $n_2 = 6$

$$\because \frac{1}{\lambda} = \bar{\upsilon} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\because \frac{1}{\lambda} = 1.097 \times 10^7\left(\frac{1}{3^2} - \frac{1}{6^2}\right) = 9.14 \times 10^5\ m^{-1}$$

$$\therefore \boxed{\lambda = 1.09 \times 10^{-6}\ m = 1090\ nm}$$

3 **Calculate the $\bar{\upsilon}$, λ and υ in nm of line which corresponding to the transition from n=6 to n=1:** ($R_H = 1.097 \times 10^7\ m^{-1}$)

Answer

A

$$\because \bar{\upsilon} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\therefore \bar{\upsilon} = 1.097 \times 10^7\left(\frac{1}{1^2} - \frac{1}{6^2}\right)$$

$$\therefore \boxed{\bar{\upsilon} = 1.07 \times 10^7\ m^{-1}}$$

B

$$\because \frac{1}{\lambda} = \bar{\upsilon}$$

$$\therefore \lambda = \frac{1}{\bar{\upsilon}} = \frac{1}{1.07X10^7}$$

$$\therefore \boxed{\lambda = 94X10^{-9}\text{ m} = 94\text{ nm}}$$

C

$$\because \upsilon = \frac{C}{\lambda}$$

$$\because C = 3\text{ x}10^8\text{ m/s}$$

$$\therefore \boxed{\upsilon = \frac{3x10^8}{94x10^{-9}} = 3.19x10^{15}S^{-1}}$$

4

The frequency of ultraviolet light is $2.37x\ 10^{-16}$ and yellow light is $5.26x10^{14}\ s^{-1}$

(Calculate the energy (J) of each photon and compare these photons by calculating the energy ratio)

Answer

For Ultraviolet

$\because$ E=hυ

$\because E_{UV} = 6.63x10^{-34}\text{ x }2.73x10^{16}$

$\therefore E_{UV} = 1.18\text{ x }10^{-17}$ J

For Yellow Light

$\because E_Y = 6.63x10^{-34}\text{ x }5.26x10^{14}$

$\therefore E_Y = 3.49\text{ x }10^{-19}$ J

Ratio between 2 light

$$\because \frac{E_{UV}}{E_Y} = \frac{1.81X10^{-17}}{3.49X10^{-19}}$$

$$\therefore \frac{E_{UV}}{E_Y} = 51.9$$

$\therefore$ Ultraviolet light is larger than yellow light by 51.9 time

Chapter Two

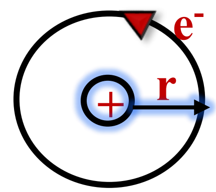
Electronic Structure
Part (2)

Bohr's Theory

Bohr explained H-atom and H-like atoms (having only one electron) → Ex: He^{+}, Li^{2+}, Be^{3+}, Na^{10+}, …..

Bohr's Postulates

1 The electron moves around the nucleus in a circular orbit of constant radius (r).



(مدار دائرى)

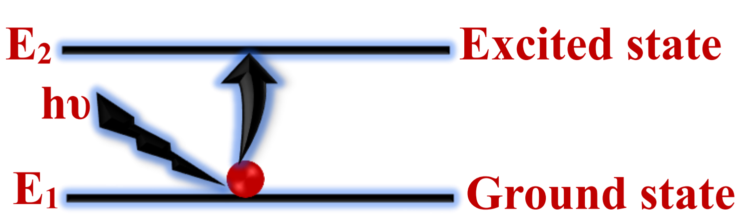
2 The electrons of the same orbital have the same energy (r).

$\therefore$ Energy is quantized (محدده)

3 The electrons undergo the transition from a lower energy state (ground state) to a higher energy state (Excited state) when the atom absorbs energy.

Note

The amount of energy absorbed or emitted is equal energy difference (ΔE) between 2 levels.



$\boxed{\Delta E = h\upsilon = E_2 - E_1}$

(Jump to a higher level)

From Bohr's theory can be calculated

A. Radius (r) B. Energy of orbit (E)

C. Enrgy difference emitted or absorbed (ΔE)

A Radius (r) of an orbit (n)

(حفظ)

$r = a_o n^2$ H-atom (كلما زاد رقم مستوى n زاد نص القطر r)

$r = \frac{a_o\, n^2}{Z}$ H-like atom (H^+, Li^{2+}, Be^{3+}, …..)

Where

- Z: Atomic number
- a_o: radius constant (a_o= 0.529 Å = 52.9 Pm)

(Pm : Picometer = 10^{-12} m)

B Energy (E) of an orbit (n)

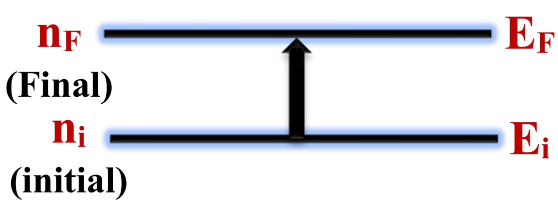
(حفظ)

$$\boxed{E = \frac{-K}{n^2}}$$ **H-atom**

$$\boxed{E = \frac{-KZ^2}{n^2}}$$ **H-like atom**

K: Bohr's constant (K = 2.18X10^{-18} J)

C Energy difference (ΔE) (Absorbed or emitted)



$\because \Delta E = E_F - E_i$ And $E = \frac{-K}{n^2}$

$$\therefore \Delta E = \frac{-K}{n_F^2} - \frac{-K}{n_i^2}$$

$$\boxed{\Delta E = K\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)}$$ **H-atom**

$$\boxed{\Delta E = KZ^2\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)}$$ **H-like atom**

(ΔE= +Ve)

$n_F > n_i$

∴ **Absorption** امتصاص

(ΔE= -Ve)

$n_i > n_F$

∴ **Emission** انبعاث

Important Notes

A. As n of an orbit increases, its size increase

B. As electrons far from the nucleus, their energy becomes larger

Rydberg equation from Bohr theory

From Bohr

$$\boxed{\Delta E = K\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)}$$

$$\because \Delta E = h\upsilon = \frac{h\,C}{\lambda} = h\,C\,\bar{\upsilon}$$

$$\therefore \Delta E = h\,C\,\bar{\upsilon} = K\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)$$

$$\therefore \bar{\upsilon} = \frac{K}{h\,C}\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)$$

$$\because \frac{K}{h\,C} = R_H$$

$$\therefore \boxed{\bar{\upsilon} = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)}$$ (Rydberg Equation)

- $C = 3\text{x}10^{8}$ m/s
- $K = 2.18\text{x}10^{-18}$ J
- $R = 1.097\text{x}10^{7}$ m^{-1}
- $h = 6.63\text{x}10^{-34}$ J.S

Ionization energy (Ionization potential)

- The energy required to remove an electron from atom. (التآين يعني الخروج النهائي للالكترون)
- $n_i = 1, n_f = \infty$ (لحساب طاقة التأين نضع)

Disadvantages (Failure) of the Bohr model

1. Bohr can't explain the atomic spectra of species containing more than one electron (Only fit 1 electron system)

2. Bohr predicted: electron is Particle that has fixed distance From the nucleus. so It's Possible to locate position of electron (which accually not Possible)

Important Notes للمسائل

A. ΔE from Bohr's theory can be (+Ve) or (-Ve) according to absorbance or emitted.

B. $\overline{\upsilon}$ from the Rydberg equation is always (+Ve) value

C. In case of ionization of electron ⟹ Put $n_f = \infty$ and $n_i = 1$ ✔

Problems

1 **Calculate the radius (r) for the first Bohr orbit of hydrogen atom**

Answer

For H-atom : $\because \; r = a_o \, n^2$

$$\therefore \; r = 52.9 \times 1^2 = 52.9 \text{ Pm}$$

OR $\therefore \; r = 0.529 \times 1^2 = 0.529 \text{ Å}$

2 **Calculate the Energy corresponding to the transition of the electron in H-atom from n=4 to n=2 and calculate frequency**

($h = 6.63 \times 10^{-34}$ J.S and $K = 2.18 \times 10^{-18}$ J)

Answer

$\because \; n_i = 4$ and $n_F = 2$

A- $\because \; \Delta E = K \left(\frac{1}{n_i^2} - \frac{1}{n_F^2} \right)$

$$\therefore \; \Delta E = 2.18 \times 10^{-18} \left(\frac{1}{4^2} - \frac{1}{2^2} \right)$$

$$\therefore \; \boxed{\Delta E = -4.0875 \times 10^{-19} \text{ J}} \rightarrow \text{Emission}$$

B- $\because \Delta E = h\upsilon$

$\therefore -4.0875\times10^{-19} = 6.63\times10^{-34}\ \upsilon$

$\therefore \boxed{\upsilon = 6.17\times10^{14}\ Hz}$

C- **For wavelength**

$\because \lambda = \frac{C}{\upsilon}$

$\because \boxed{\lambda = \frac{3\times10^{8}}{6.17\times10^{14}} = 486\times10^{-9}\ nm}$

3 **Calculate the Energy required for the ionization of the electron from the ground state of the H-atom**

Answer ($K = 2.18\times10^{-18} J$)

$\because H \rightarrow H^{+} + e^{-}$ (ionization)

$\because n_i \longrightarrow$ ground state ($n_i = 1$)

$\because n_F \longrightarrow$ ionization ($n_f = \infty$)

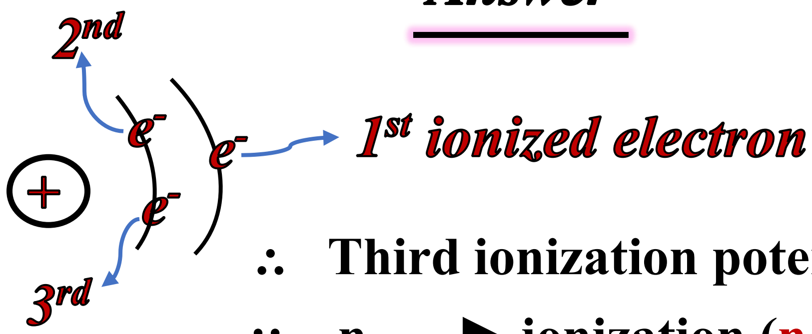
$\because \Delta E = I.E = K\left(\frac{1}{n_i^2} - \frac{1}{n_F^2}\right)$

$\therefore \boxed{\Delta E = 2.18\times10^{-18}\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right) = 2.18\times10^{-18}\ J}$

4 Calculate the third ionization Potential of the lithium atom

(K = 2.18x10^{-18} J and Z=3)

Answer



∴ Third ionization potential n_i ⟶ ($n_i = 1$)

∵ n_F ⟶ ionization ($n_f = \infty$)

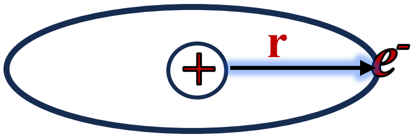
∵ $\Delta E = Z^2 K \left(\frac{1}{n_i^2} - \frac{1}{n_F^2} \right)$

∴ $\Delta E = 3^2 \times 2.18 \times 10^{-18} \left(\frac{1}{1^2} - \frac{1}{\infty^2} \right) = 1.96 \times 10^{-17}$ J

Sommerfield model

The electron moves in an elliptical orbit with a variable radius

(مدار بيضاوى)



De-Broglie model

Show the wave nature of an electron (Particle)

The electron has Dual Nature

A- Wave Nature (Plank)

$\because \boxed{E = h\,\upsilon = h\,\frac{c}{\lambda}}$

B- Particle Nature (Einstein)

$\because \boxed{E = m\,C^2}$

$\therefore$ **From A and B** : $E = h\,\frac{c}{\lambda} = m\,C^2$

$$\therefore \boxed{\lambda = \frac{h}{m\,C}}$$

(**For Photon**)

m: mass of photon

OR For any Particle like (electron, Proton,……..)

$$\therefore \boxed{\lambda = \frac{h}{m\,V}}$$

(**For Particle**)

m: mass of particle

V: velocity of particle

Important Notes

A. mC → momentum of Photon (كميه التحرك للفتون)

B. mV → momentum of Particle (كميه التحرك لجسم)

C. Velocity is inversely proportional with mass ($V \propto \frac{1}{m}$)

Ex : mass of Baseball > Neutron > Proton > electron
Velocity of Baseball < Neutron < Proton < electron

Problems

1 **In a baseball, a fast pitch has time at 44.1 m/s calculate (λ) associated with a baseball** (h = 6.63×10^{-34} J.S and m=146g)

Answer

$$\because \quad \lambda = \frac{h}{m\,V}$$

$\because$ mass = 146 g = 146×10^{-3} kg

$$\therefore \quad \boxed{\lambda = \frac{6.63x10^{-34}}{44.1x146x10^{-3}} = 1.03 \times 10^{-34} \text{ m} = 1.03 \times 10^{-25} \text{ nm}}$$

mass should be by Kg (due to J= $Kg.m^2/S^2$)

2 **The velocity of an electron is 2.19×10^6 m/s calculated (λ) associated with an Electron** ($m_e = 9.11 \times 10^{-28}$ g)

Answer

$$\because \quad \lambda = \frac{h}{m\,V}$$

$\because$ mass = 9.11×10^{-28} g = 9.11×10^{-31} kg

$$\therefore \quad \lambda = \frac{6.63x10^{-34}}{2.19x10^{6}x9.11x10^{-31}} = 3.32 \times 10^{-10}\ \text{m} = 0.332\ \text{nm}$$

Uncertainty Principle by Heisenberg (مبدأ عدم التاكد لهيزنبرج)

It is impossible to determine the position and the momentum (velocity) of an electron at the same time

$$\Delta X\ m\ \Delta V \geq \frac{h}{4\,\pi}$$

Where

- **ΔX:** uncertainty in position of electron
- **ΔV:** uncertainty in the Velocity of electron
- **m:** mass of electron
- **π:** 22/7

Problems

What is uncertainty in the velocity of:

A- For Baseball (m=0.416 Kg)

B- For electron (m=$9.11 x10^{-34}$ Kg)

(If the position of each object is located in $1x10^{-11}$ m)

Answer

A-

$$\because \ \Delta X \ m \ \Delta V \geq \frac{h}{4\pi}$$

$$\therefore \ 1x10^{-11} \ x \ 0.416 \ \Delta V \geq \frac{6.63x10^{-34}}{4 \ x\frac{22}{7}}$$

$$\therefore \ \boxed{\Delta V = 3.6 \ X10^{-23} \ m/s}$$

B-

$$\because \ \Delta X \ m \ \Delta V \geq \frac{h}{4\pi}$$

$$\therefore \ 1x10^{-11} \ x \ 9.11x10^{-31} \ \Delta V \geq \frac{6.63x10^{-34}}{4 \ x\frac{22}{7}}$$

$$\therefore \ \boxed{\Delta V = 5.8 \ X10^{6} \ m/s}$$

Schrodinger Equation (Quantum mechanics)

Schrodinger Wave Equation

- Describe Particle-Wave Nature of electron (Describe **Behaviour** and **Energy of electron**)
- Electrons do not exist in circular orbits but <u>Orbitals</u> (**Region in space where electrons are most likely to be found**).
- As the electron become Closer to nuclues, the higher probability to find electron
- Electrons found in **clouds** around nucleus

Schrodinger wave equation

$$\nabla^2\psi + \frac{8\pi^2 m}{h^2}(E-V)\,\psi = 0$$

Where

- ✓ **∂ :** Partial differentiation
- ✓ **m:** mass of electron
- ✓ **E:** Total energy
- ✓ **ψ:** Wavefunction (Function in X, Y and Z)
- ✓ **V:** Potential Energy $= \frac{e^2}{4\pi\varepsilon_o r}$
- ✓ **h:** Planck's Constant
- ✓ $\nabla^2 = \frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2}$

Wave function Separation

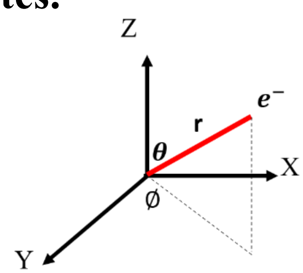
In Cartesian coordinates

$$\therefore \quad \psi_{(X,Y,Z)} = f(x) \cdot f(y) \cdot f(z)$$

In Spherical polar coordinates

$$\therefore \quad \psi(r,\theta,\emptyset) = R(r) \cdot f(\theta) \cdot g(\emptyset)$$

- The location of an electron in atom is determined by spherical coordinates.



r : distance from the origin.
θ : polar angle.
Ø : Azimuthal angle.

- These are related to Cartesian coordinates by following relations:

$X = r \sin\theta \cos\emptyset$

$Y = r \sin\theta \sin\emptyset$

$Z = r \cos\theta$

$r = \sqrt{x^2 + y^2 + z^2}$ $\qquad \theta = \cos^{-1}\frac{z}{r}$ $\qquad \emptyset = \tan^{-1}\frac{Y}{X}$

ψ: Wavefunction describes:

✔ (1) Probability of the position of an electron in space (Orbitals).

✔ (2) The energy of electron in atom.

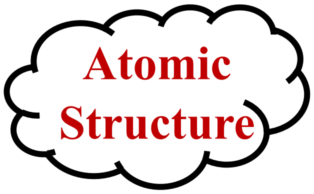
Important Notes

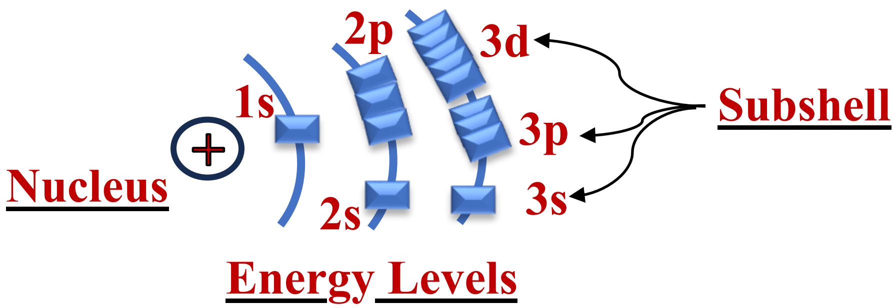
A- This equation was solved only for H-atom

B- Solution Schrodinger equation gives 3 Quantum No :

1. Princpal quantum number (n)
2. Angular quantum number (ℓ)
3. Magnetic quantum number (m_L)

✔ (Spin quantum number not determined by Schrodinger)





Quantum number

A Principal quantum number (n) (عدد الكم الرئيسي)

Number of energy Levels (Shells)

Where

n = 1, 2, 3, 4, 5, ……. (Integer Number) عدد صحيح

1st Level

Shell	K	L	M	N	O	P	Q
n	1	2	3	4	5	6	7

Important Notes

A- Determine the value of energy level from Known (n), where the small value of (n) will have a low Energy.

B- As the value of (n) becomes larger (farther from the nucleus) As the size increase of the orbit.

B Angular (Azimuthal) quantum number (ℓ) (عدد الكم الثانوى)

- Number of Subshells (sublevels)
- Describe the shape of (volume) space that e's occupied

Where

ℓ = 0,1, 2, 3, ……. (n-1)

Subshell	s	p	d	f
ℓ	0	1	2	3

Important Note

ns np

Principle quantum No — Angular quantum No

Example

n	ℓ	subshell
1	0	1s
2	0	2s
	1	2p
3	0	3s
	1	3p
	2	3d

C Magnetic Quantum Number (m_ℓ) (عدد الكم المغناطيسي)

- **Define number of orbitals in each subshells**
- **Tell us the orientation (اتجاه) of orbitals in space**

Where

$$m_\ell = -\ell, \ldots\ldots, 0, \ldots\ldots, +\ell$$

Example

n	ℓ	m_ℓ	Subshell
1	0	0	1s $m_\ell=0$
2	0	0	2s $m_\ell=0$
	1	-1,0,+1	2P -1 0 +1

D Spin Quantum number (m_s or S) (عدد الكم المغزلي)

- **Define the direction (spin) of Electrons**

Clockwise (↑)

$\therefore m_s = +½$

Anti Clockwise (↓)

$\therefore m_s = -½$

(m_s not obtain from Schrödinger equation)

Write The four quantum Numbers

A- $1s^1$

Answer

$n = 1 \leftarrow 1s^1 \rightarrow \ell = 0 \rightarrow m_\ell = 0$

$\therefore m_s = +\frac{1}{2}$ $\boxed{\uparrow}$

يتم توزيع الالكترونات فى الاوربيتالات فرادى اولا و اتجاها لفوق ($S = \frac{1}{2}$)

و عند الازدواج يكون الاتجاه لاسفل ($S = -\frac{1}{2}$)

B- $2p^2$

Answer

$n = 2 \leftarrow 2p^2 \rightarrow \ell = 1 \rightarrow m_\ell = -1, 0, +1$

m_ℓ: -1, 0

$\boxed{\uparrow}\boxed{\uparrow}\boxed{\ }$

m_s: $+\frac{1}{2}$ $+\frac{1}{2}$

C- $3d^4$ ⟶ $\therefore n = 3$ $\therefore \ell = 2 \rightarrow m_\ell = -2, -1, 0, +1, +2$

m_ℓ: -2, -1, 0, 1

$\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\ }$

m_s: $+\frac{1}{2}$ $+\frac{1}{2}$ $+\frac{1}{2}$ $+\frac{1}{2}$

Important Notes

1 **Number of orbitals in each subshells= $2\ell+1$**

Each orbital filled with = 2e's

Subshell	ℓ	No of Orbitals	No of electrons
s	0	1	2e's ⥮
P	1	3	6 e's ⥮ ⥮ ⥮
d	2	5	10 e's ⥮ ⥮ ⥮ ⥮ ⥮
f	3	7	14 e's ⥮ ⥮ ⥮ ⥮ ⥮ ⥮ ⥮

2 **Number of orbitals in each shell = n^2**

Shell (n)	No of Orbitals	Subshell
K (1)	1	1s
L (2)	4	2s, 2p
M (3)	9	3s, 3p, 3d
N (4)	16	4s, 4p, 4d, 4f

3 **Number of electrons in each shell = $2n^2$**

Ex : in second shell (n=2) → ∴ it contains ($2x2^2$) ∴ 8 e,s

shell	n	ℓ	m_ℓ	subshell	No of orbitals
K	1	0	0	1s	1
L	2	0	0	2s	1
		1	-1, 0, +1	2p	3
M	3	0	0	3s	1
		1	-1, 0, +1	3p	3
		2	-2, -1, 0, +1, +2	3d	5
N	4	0	0	4s	1
		1	-1, 0, +1	4p	3
		2	-2, -1, 0, +1, +2	4d	5
		3	-3, -2, -1, 0, +1, +2,+3	4f	7

حفظ

m_ℓ= 0
s

m_ℓ= -1 0 +1
p

m_ℓ= -2 -1 0 +1 +2
d

m_ℓ= -3 -2 -1 0 +1+2 +3
f

Write The four quantum Numbers

A- $3s^2$

Answer

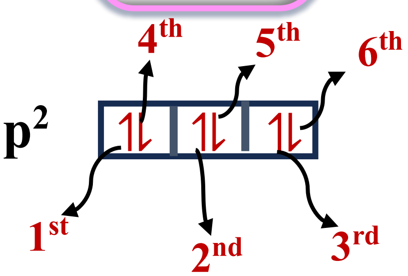
$3s^2$ $\boxed{\uparrow\downarrow}$

Electron	n	ℓ	m_ℓ	m_s
1st	3	0	0	+1/2
2nd	3	0	0	-1/2

B- $2p^2$

Answer

$2p^2$



Electron	n	ℓ	m_ℓ	m_s
1st	2	1	-1	+1/2
2nd	2	1	0	+1/2
3rd	2	1	+1	+1/2
4th	2	1	-1	-1/2
5th	2	1	0	-1/2
6th	2	1	+1	-1/2

Quiz

❖ **Write 4 quantum number for each following :**

A- $2s^2$ B- $3p^2$ C- $4p^4$

D- $4d^4$ E- $3d^7$ F- $5f^1$

Pauli Exclusion Principle

There are no two electrons in same atom having same four quantum number

(at least different in one quantum No)

Example

$2p^3$ $\boxed{\uparrow \mid \uparrow \mid \uparrow}$

$\therefore$ **All three electrons have n=2, $\ell = 1$ and $m_s = +\frac{1}{2}$**

But different in m_ℓ= -1, 0, +1

Summary (ملخص)

1

Shells (n=1, 2,...) $\xrightarrow{\text{Divided into}}$ Subshells $\xrightarrow{\text{Divided into}}$ orbitals $\xrightarrow{\text{contained}}$ e,s

Subshells	orbitals	e,s
$\ell = 0 \rightarrow s$	$m_\ell = 0$	2
$\ell = 1 \rightarrow p$	$m_\ell = -1 \quad 0 \quad +1$	6
$\ell = 2 \rightarrow d$	$m_\ell = -2 \; -1 \quad 0 \quad +1 \; +2$	10
$\ell = 3 \rightarrow f$	$m_\ell = -3 \quad -2 \quad -1 \quad 0 \quad +1 +2 \; +3$	14

2

No of orbitals in each Subshell	$2\ell + 1$
No of orbitals in any shell	n^2
No of electrons in any shell	$2n^2$

3

Electrons in same shell → have same (n)

Electrons in same Subshell → have same (n and ℓ)

Electrons in same orbitals → have same (n, ℓ and m_ℓ)

Atomic Orbitals

The region in space at which electron is most likely to be found

Shape of orbitals

1. s-orbitals
2. p-orbitals
3. d-orbitals

1 s-orbitals

$l = 0$ (starts at n= 1)

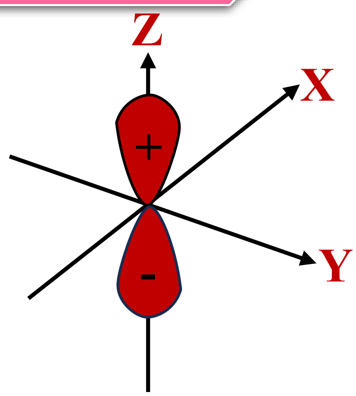
- **Spherical and symmetrical**
- **The radius of the sphere increases (size) with a value of (n)**

2 P-orbitals

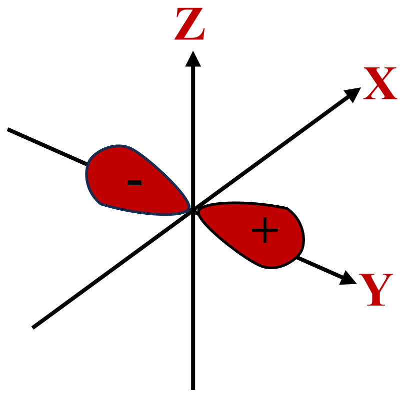
$l = 1$ (starts at n= 2)

- **Dumbbell-shaped (2 lobes) orbitals of electron density with a node at the nucleus that doesn't contain electrons.**
- **Three dengenrate orbitals** (degenrate = same Energy)

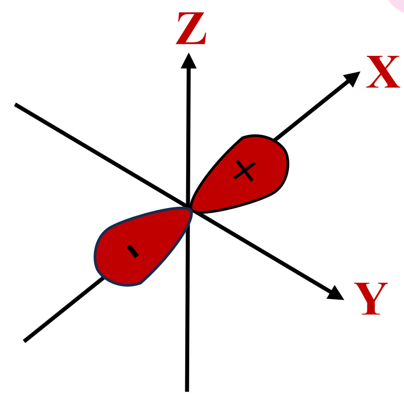
P_X P_Y P_Z → **Degenerate** (same Energy) **but differ in orientation**



P_Z



P_Y

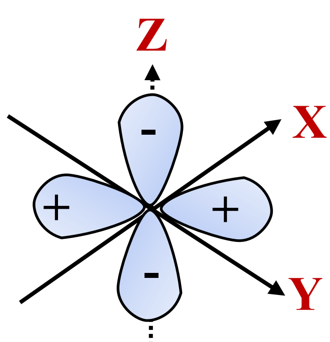


P_X

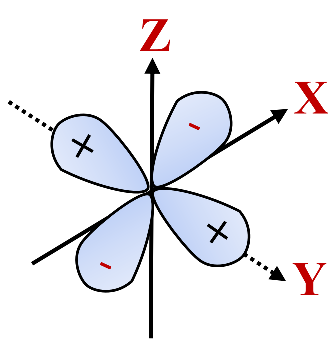
3 d-orbitals $l = 2$ (starts at n= 3)

Five degenerate orbitals of different orientations in space.

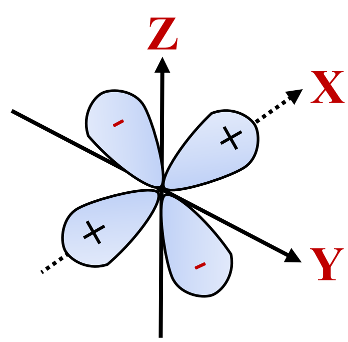
- d_{XY} , d_{XZ} , d_{YZ}, $d_{X^2Y^2} \longrightarrow$ **4 Lobes**
  - $d_{Z^2} \longrightarrow$ **Orinted along Z-axis (Like P-Orbital)**



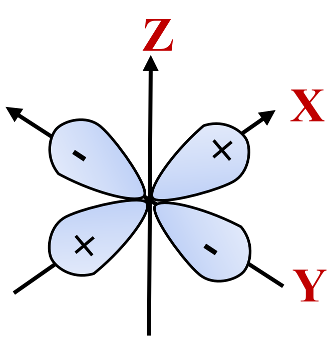
d_{XY}



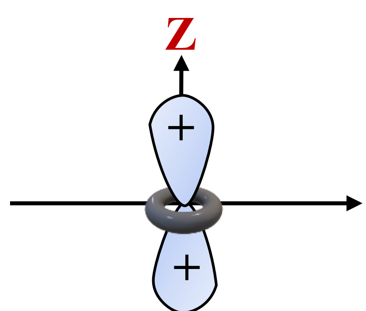
d_{XZ}



d_{YZ}



$d_{X^2-Y^2}$



d_{Z^2}

Quiz 3

I. Choose the correct answer:

1. **(Midterm 2022)** What are the correct four quantum numbers for a 3d orbital? The answers are expressed as (n, l, m_l, m_s)

A) 3, 0, 0, -½	B) 3, 2, 2, -½	C) 3, 2, 2, 1	D) 3, 1, -1, -½

2. **(Midterm 2022)** Consider the ground state of Cr atom (Z = 24). The numbers of electrons with the azimuthal quantum numbers, l = 1 and 2 are, respectively:

A) 12 and 4	B) 12 and 5	C) 16 and 4	D) 16 and 5

3. **(Midterm 2022)** The formalism describes the dual nature of electron is:

A) $E = mc^2$	B) $2\pi r = n\lambda$	C) $E = hc/\lambda$	D) $\lambda = h/(mv)$

4. **(Midterm 2022)** Which of the following statements is correct for an electron that has the quantum numbers n =4 and m_l = -3?

A) The electron may be in a d orbital
B) the electron is in the fourth principal
C) the electron may be in a p orbital
D) none of the above applies to this electron

5. **(Midterm 2022)** The correct set of four quantum numbers for the valence electron of rubidium atom (Z = 37) is

A) 5, 0, 0, +½	B) 5, 1, 0, + ½	C) 5, 1, 1, +½	D) 6, 0, 0, +½

6. **(Midterm 2020)** Which one of the quantum numbers does not result from the solution of the Schrodinger equation?

A) magnetic	B) spin	C) principal	D) azimuthal

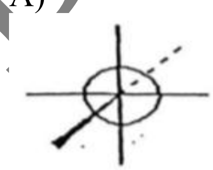
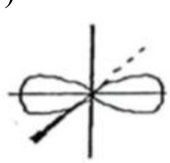
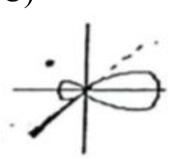
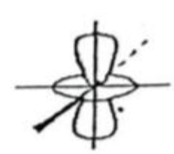
7. **(Midterm 2020)** In a Px orbital, the subscript x denotes the........ of the electron cloud.

A) orientation	B) shape
C) energy	D) size of the orbital

8. **(Midterm 2020)** For the atom S; the number of orbitals in the last subshell.

A) 6	B) 4	C) 3	D) 10

9. **(Midterm 2019)** The probability area that best represents the shape of one of the atomic 2p orbitals is...

A) 	B) 	C) 	D) 

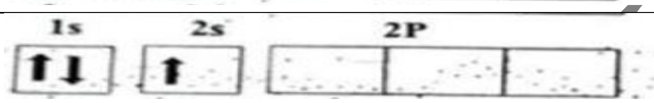
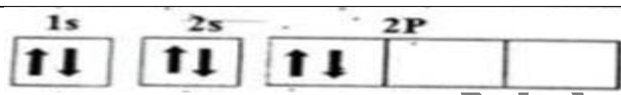
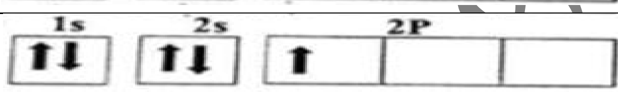
10. **(Midterm 2019)** What is the maximum number of electrons in an atom that can have the following set of quantum numbers? ($n = 4$, $l = 3$, $m_l = -2$, $m_s = +{}^{1}/_{2}$)

A) 0	B) 1	C) 10	D) 6	E) 2

11. **(Midterm 2019)** Which of the following statements is correct for an electron that has the quantum numbers $n = 4$ and $m = -2$?

A) the electron may be in a p orbital
B) the electron may be in a d orbital
C) the electron is in the second principal shell
D) the electron must have a spin quantum number $m_s = +{}^{1}/_{2}$

12. **(Midterm 2019)** Which electron configuration does not agree with Pauli Exclusion Principle?

A) 
B) 
C) 
D) 

13. **(Midterm 2018)** The set of quantum numbers, $n = 4$, $l = 3$, $m_l = 2$

A) describes one of seven orbitals of a similar type
B) is not allowed
C) describes an electron in a 4d orbital
D) describes an electron in a 3p orbital

14. **(Midterm 2018)** Which one of the following sets of quantum numbers is not possible?

	n	l	m_l	m_s
A)	4	3	-2	$+{}^{1}/_{2}$
B)	3	2	-3	$-{}^{1}/_{2}$
C)	3	0	0	$+{}^{1}/_{2}$
D)	4	1	1	$-{}^{1}/_{2}$

15. **(Midterm 2017)** What is the principle quantum number (n) of the first shell to have F-orbitals

A) 1	B) 2	C) 3	D) 4

16. Which of the following statements is correct for an electron that has the quantum numbers n = 4 and m_l = -3?

A) The electron may be in a d orbital
B) the electron is in the fourth principal shell
C) the electron may be in a p orbital
D) none of the above applies to this electron

17. Consider the sulfur atom. How many electrons does this atom have l = 0?

A) 6	B) 3	C) 16	D) 2

18. Which of the following sets of quantum numbers describes the most easily removed electron in an potassium atom in its ground state?

A) n = 4, l = 0, m_l = 0, m_s = +1/2	B) n = 3, l = 1, m_l = 1, m_s = -1/2
C) n = 2, l = 1, m_l = 0, m_s = -1/2	D) n = 4, l = 0, m_l = 1, m_s = +1/2

19. How many orbitals are allowed in a subshell if the angular quantum number (l) for electrons in that subshell is 3?

A) 1	B) 3	C) 5	D) 7	E) 9

20. For all atoms of the same element, the 2s orbital is larger than the 1s orbital.

A) True	B) False

21. An electron in a 3p orbital could have a value of 2 for its angular momentum quantum number (l)

A) True	B) False

22. Each shell (principal energy level) of quantum number n contains n subshells

A) True	B) False

23. The maximum number of electrons permitted in a $3d_{xy}$ orbital are 10.

A) True	B) False

24. The ratio of Oxygen in NO_2 & NO_3 with fixed mass of Nitrogen is ………

A) 3 : 1	B) 2 : 1	C) 1 : 3	D) 1 : 1

25. **(Important)** Which of the following pairs of compounds can be used to illustrate the law of multiple proportions?

A) CH_4 and CO_2	B) NH_4 and NH_4Cl	C) ZnO_2 and $ZnCl_2$
D) H_2O and HCl	E) NO and NO_2	

26. **(Important)** Which compound contains the highest percent by mass of hydrogen?

A) H_2O	B) HF	C) H_2S	D) HCl	E) H_2SO_4

27. **(Important)** In an investigation of the electronic absorption spectrum of a particular element, it is found that a photon having λ = 500 nm provides just enough energy to

promote an electron from the second quantum level to the third. From this information, we can deduce ……….

A) the difference in energies between n = 2 and n = 3.
B) the energy of the n = 2 level.
C) the energy of the n = 3 level .
D) the sum of the energies of n = 2 and n = 3.
E) all of these.

28. **(Important)** How many electrons in an atom can have the quantum numbers n = 4, l = 2?

A) 12	B) 10	C) 6	D) 5	E) 14

29. **(Important)** How many electrons can be contained in all of the orbitals with n = 4?

A) 10	B) 2	C) 18	D) 32	E) 8

30. How many orbitals are allowed in a subshell if *l* = 3

A) 1	B) 3	C) 5	D) 7	E) 9

Answers:

Q:	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A:	B	B	D	B	A	B	A	C	B	B	B	A	A	B	D
Q:	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
A:	B	A	A	D	A	B	A	B	C	E	A	A	B	D	D

II. What is the maximum number of electrons in an atom that can have the following quantum numbers :

(1) n = 2, $m_s = +½$

(2) n= 4, $m_l = +1$

(3) n= 4, l = 3, $m_l = –2$

(4) n= 2, l = 0, $m_s = –½$.

Answer

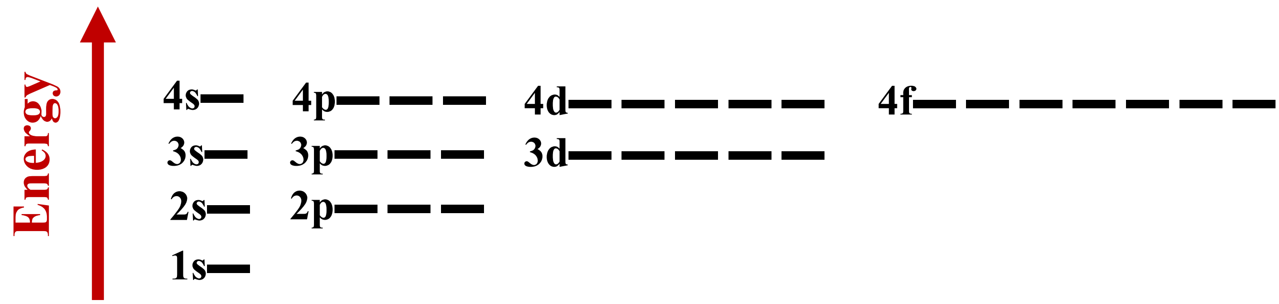
Number	1	2	3	4
No. of electrons	4 es	6 es	2 es	1 es

Energy of orbitals

1 one (single) electron atom

(Ex: H-atom)

Depend only on principal quantum No (n)

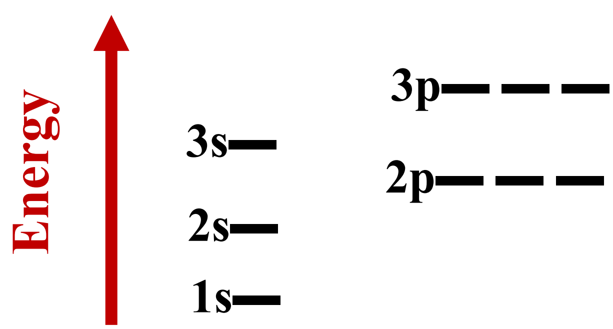


∴ Orbitals have Same (n) → Same energy (degenerate)

Ex : 2s = 2p in energy

2 multielectron atom

Depend on (n) + (ℓ) → **larger** ($n + \ell$) **have higher energy.**



∴ Orbitals have Same (n) → different energy (not degenerate)

EXAMPLE

a) 2s , 2p

❖ 2s $\longrightarrow$ $(n + \ell) = 2 + 0 = 2 \longrightarrow$ lower
❖ 2p $\longrightarrow$ $(n + \ell) = 2 + 1 = 3 \longrightarrow$ higher

$\therefore$ 2s filled at first

b) 3s , 2p

❖ 3s $\longrightarrow$ $(n + \ell) = 3 + 0 = 3$
❖ 2p $\longrightarrow$ $(n + \ell) = 2 + 1 = 3$

$\left.\begin{matrix} 3s \\ 2p \end{matrix}\right\}$ Same $(n+ \ell)$

$\therefore$ Orbital has low (n) will be filled first $\therefore$ 2p filled 1^{st}

c) 4s , 3d

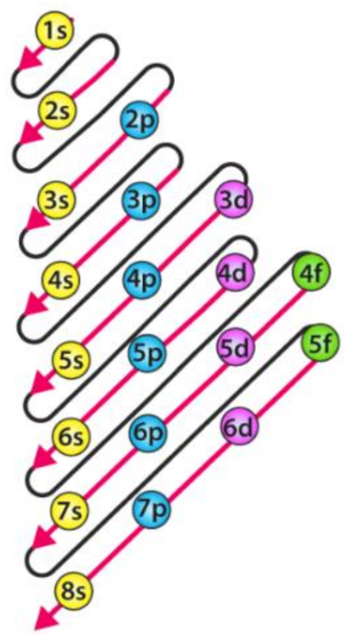
❖ 4s $\longrightarrow$ $(n + \ell) = 4 + 0 = 4 \longrightarrow$ lower
❖ 3d $\longrightarrow$ $(n + \ell) = 3 + 2 = 5 \longrightarrow$ higher

$\therefore$ 4s filled at first

The electronic configuration of elements

1 Aufbau Principle (building up principle)

The electrons occupy the subshell of lower energy and then occupy the subshell of higher energy.



Energy of 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d< 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p

∴ 1s 2s 2p 3s 3p 4s 3d 4p 5s

(سس بس بس دبس دبس فدبس)

2 Hund's Rule

The most stable arrangement of electrons in subshells in the one with the greatest number of parallel spins.

Ex : P^2

$\boxed{\uparrow\downarrow}\boxed{}\boxed{}$ ✗ $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$ ✓

Electronic Configuration

2) Long notation : $1s^2\ 2s^2\ 2p^6$

1) Orbital notation : $\underset{1s}{\boxed{\uparrow\downarrow}}\quad \underset{2s}{\boxed{\uparrow\downarrow}}$

3) Shorthand notation (Noble gas configuration)

$$\left.\begin{array}{ll} \text{A-} & {}_{10}[Ne]\ 3s^2\ 3p^6 \longrightarrow 18 > Z > 10 \\ \text{C-} & {}_{18}[Ar]\ 4s^2\ 3d^{10}\ 4p^6 \longrightarrow 36 > Z > 18 \\ \text{B-} & {}_{36}[Kr]\ 5s^2\ 4d^{10}\ 5p^6 \longrightarrow 54 > Z > 36 \end{array}\right\}$$

حفظ

Important Notes

1

Shells (n=1, 2,...) $\xrightarrow[\text{into}]{\text{Divided}}$ Subshells $\xrightarrow[\text{into}]{\text{Divided}}$ orbitals $\xrightarrow{\text{contained}}$ e,s

$\ell = 0 \rightarrow s$	$m_\ell = 0$	2
$\ell = 1 \rightarrow p$	$m_\ell = -1 \quad 0 \quad +1$	6
$\ell = 2 \rightarrow d$	$m_\ell = -2\ -1\ \ 0\ +1\ +2$	10
$\ell = 3 \rightarrow f$	$m_\ell = -3\ \ -2\ \ -1\ \ 0\ \ +1\ +2\ +3$	14

$\therefore$ $s^2\ p^6\ d^{10}\ f^{14}$ (Max number of e's to be completely filled)

2

Order of orbitals stability

Completely filled > half-filled > partially filled

Completely filled → More stable

partially filled → less stable

Example

1) $_{24}$Cr (Chromium) مهم جداً

- $_{18}$[Ar] $4S^2$, $3d^4$ → (unstable) (Not prefer) ✗
- $_{18}$[Ar] $4S^1$, $3d^5$ → (Extra stability of half-filled $3d^5$) ✓

2) $_{29}$Cu (copper) مهم جداً

- $_{18}$[Ar] $4S^2$, $3d^9$ → (unstable) (Not prefer) ✗
- $_{18}$[Ar] $4S^1$, $3d^{10}$ → (Extra stability of complete filled $3d^{10}$) ✔

3 Exception in configuration (الحالات الشاذه فى التوزيع)

Half-filled d - orbital		Completely filled d-orbital	
$_{24}$Cr	$_{18}$[Ar] $4s^1$ $3d^5$	$_{29}$Cu	$_{18}$[Ar] $4s^1$ $3d^{10}$
$_{42}$Mo	$_{36}$[Kr] $5s^1$ $4d^5$	$_{47}$Ag	$_{36}$[Kr] $5s^1$ $4d^{10}$
$_{74}$W	$_{54}$[Xe] $4f^{14}$ $6s^1$ $5d^5$	$_{79}$Au	$_{54}$[Xe] $4f^{14}$ $6s^1$ $5d^{10}$

♣ Mo →Molybdenum ♣ W →Tungsten

♣ Ag →Silver ♣ Au →Gold

Quiz

Which one of the following pairs is more stable

❖ $_{28}$Ni or $_{25}$Mn

Answer

✔ $_{28}$Ni : $_{18}$[Ar] $4S^2$ $3d^8$ → partially filled

✔ $_{25}$Mn: $_{18}$[Ar] $4S^2$ $3d^5$ → Half-filled

∴ $_{25}$Mn is more stable than $_{28}$Ni

Show electronic configuration

5) $_{7}N$ ➔ $1S^2\ 2S^2\ 2P^3$ or $_{2}[He]\ 2S^2\ 2P^3$

3) $_{19}K$ ➔ $_{18}[Ar]\ 4S^1$

2) $_{22}Ti$ ➔ $_{18}[Ar]\ 4S^2\ 3d^2$

4) $_{40}Zr$ ➔ $_{36}[Kr]\ 5S^2\ 4d^2$

5) $_{12}Mg$ ➔ $1S^2\ 2S^2\ 2P^6\ 3S^2$ or $_{10}[Ne]\ 3S^2$

Electronic configuration of ions

2) **Atom** → Number of electrons = Atomic number (Z)

1) **+Ve ion** → Number of electrons < Atomic number (Z)

Ex : Na^+ (Z = 11 , No of e`s = 10)

3) **-Ve ion** → Number of electrons > Atomic number (Z)

Ex : Cl^- (Z = 17 , No of e`s = 18)

Example

A) Na^+ (Z = 11)

Na ➔ $1S^2\ 2S^2\ 2P^6\ 3S^1$ → loss 1 electron

∴ Na^+ ➔ $1S^2\ 2S^2\ 2P^6\ 3S^0$

B) Ca^{+} (Z = 20)

Ca → $_{18}[Ar]\ 4S^2$ → loss 2 electron

∴ Ca^{2+} → $_{18}[Ar]\ 4S^0$

C) Al^{3+} (Z = 13)

Al → $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^1$

∴ Al^{3+} → $1S^2\ 2S^2\ 2P^6$

D) O^{2-} (Z = 8)

O → $1S^2\ 2S^2\ 2P^4$

∴ O^{2-} → $1S^2\ 2S^2\ 2P^6$

Important Notes

❖ **For elements having Z ≥ 21 :** (Z = 21 : 30)

Loss of e`s from 4s 1st then 3d 2nd.

EX Cu^{+} (Z = 29)

Cu → $_{18}[Ar]\ 4S^1\ 3d^{10}$

Cu^{+} → $_{18}[Ar]\ 3d^{10}$

الفقد اولا من 4s

ثم بعد ذلك من 3d

Example

A) $_{47}Ag \rightarrow {}_{36}[Kr]\ 5S^1\ 4d^{10}$ loss 1st from 5s

$\therefore Ag^+ \rightarrow {}_{36}[Kr]\ 4d^{10}$

B) $_{21}Sc \rightarrow {}_{18}[Ar]\ 4S^2\ 3d^1$ (valence e\`s = 3)

$Sc^{3+} \rightarrow {}_{18}[Ar]$

C) $_{27}Co \rightarrow {}_{18}[Ar]\ 4S^2\ 3d^7$

$Co^{2+} \rightarrow {}_{18}[Ar]\ 4S^0\ 3d^7$

$Co^{3+} \rightarrow {}_{18}[Ar]\ 4S^0\ 3d^6$

D) $_{24}Cr \rightarrow {}_{18}[Ar]\ 4S^1\ 3d^5$

$Cr^{+} \rightarrow {}_{18}[Ar]\ 4S^0\ 3d^5$

$Cr^{2+} \rightarrow {}_{18}[Ar]\ 4S^0\ 3d^4$

Quiz

✔ <u>Write electronic configuration</u> : $_{24}Cr$, Cr^{3+} and Cr^{6+}

Which one of the following pairs is more stable

1) O^- OR O^{2-} (Z=8)

✔ $_8O \rightarrow 1S^2\ 2S^2\ 2P^4$

✔ $_8O^- \rightarrow 1S^2\ 2S^2\ 2P^5$

Partially filled

✔ $_8O^{-2} \rightarrow 1S^2\ 2S^2\ 2P^6$

Completely filled

∴ O^{2-} more stable than O^-

---

2) Fe^{2+} OR Fe^{3+} (Z=26)

✔ $_{26}Fe \rightarrow {}_{18}[Ar]\ 4S^2\ 3d^6$

✔ $Fe^{2+} \rightarrow {}_{18}[Ar]\ 4S^0\ 3d^6$

Partially filled

✔ $Fe^{3+} \rightarrow {}_{18}[Ar]\ 4S^0\ 3d^5$

Half -filled

∴ Fe^{3+} more stable than Fe^{2+}

Quiz

❖ **Which of the following is more stable Cu^+ or Cu^{2+}**

Valence electrons (الكترونات التكافؤ)

Electrons in the outer shell

Ex (1): $_{11}Na \rightarrow 1S^2\ 2S^2\ 2P^6\ 3S^1$ → Valence electrons = 1 e^-

Outer shell

Ex (2): $_{6}C \rightarrow 1S^2\ \underbrace{2S^2\ 2P^2}_{\text{Outer shell}}$ ∴ Valence electrons = 4 e^-

Ex (3): $_{16}S \rightarrow 1S^2\ 2S^2\ 2P^6\ \underbrace{3S^2\ 3P^4}_{\text{Outer shell}}$ ∴ Valence electrons = 6 e^-

Important Notes

For transition metals (العناصر الإنتقالية), valence e`s are equal to the sum of electrons in (d) and (S) subshells

Ex (1): $_{26}Fe \rightarrow {}_{18}[Ar]\ \underbrace{4S^2\ 3d^6}_{\text{Outer shell}}$ ∴ Valence electrons = 8 e^-

Ex (2): $_{29}Cu \rightarrow {}_{18}[Ar]\ \underbrace{4S^1\ 3d^{10}}_{\text{Outer shell}}$ ∴ Valence electrons = 11 e^-

Magnetic Properties (Magnetism)

Paramagnetic

Have unpaired e`s $\boxed{\uparrow}$

Diamagnetic

All e`s are paired $\boxed{\uparrow\downarrow}$

EXAMPLE

Ex (1): $_{11}Na \rightarrow 1S^2 \quad 2S^2 \quad 2P^6 \quad 3S^1$

$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow} \quad \boxed{\uparrow}$

∴ Number of unpaired electrons = 1 e^-

∴ Na is Paramagnetic

Ex (2): $_{26}Fe \rightarrow {}_{18}[Ar]\ 4S^2 \quad 3d^6$

$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow|\uparrow|\uparrow|\uparrow|\uparrow}$

∴ Number of unpaired electrons = 4 e^-

∴ Fe is Paramagnetic

Ex (3): $_{10}Ne \rightarrow 1S^2 \quad 2S^2 \quad 2P^6$

$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$

∴ All electrons are paired

∴ Ne is Diamagnetic

Isoelectronic

Atoms or ions having same no. of electrons(same electronic configuration)

EXAMPLE

A) Na^+ (Z = 11)

$_{11}Na \rightarrow {}_{10}[Ne]\ 3S^1$

$\therefore\ Na^+ \rightarrow {}_{10}[Ne]$

$\therefore$ **Na^+ has same No. of e\`s of Ne (same e – configuration)**

✔ $\therefore$ **Na^+ is isoelectronic with Ne**

B) F^- (Z = 9)

$_9F \rightarrow 1S^2\ 2S^2\ 2P^5$

$\therefore\ F^- \rightarrow 1S^2\ 2S^2\ 2P^6 = {}_{10}[Ne]$

$\therefore$ **F^- has same No. of e\`s of Ne (same e – configuration)**

✔ $\therefore$ **F^- is isoelectronic with Ne**

Periodic table

1) Mendeleev

Arrangement the elements according to increasing their atomic weight.

2) Mosely (Modern Periodic table)

arrangement of elements according to increasing atomic number (z).

Moder Periodic table

groups

18 Vertical columns

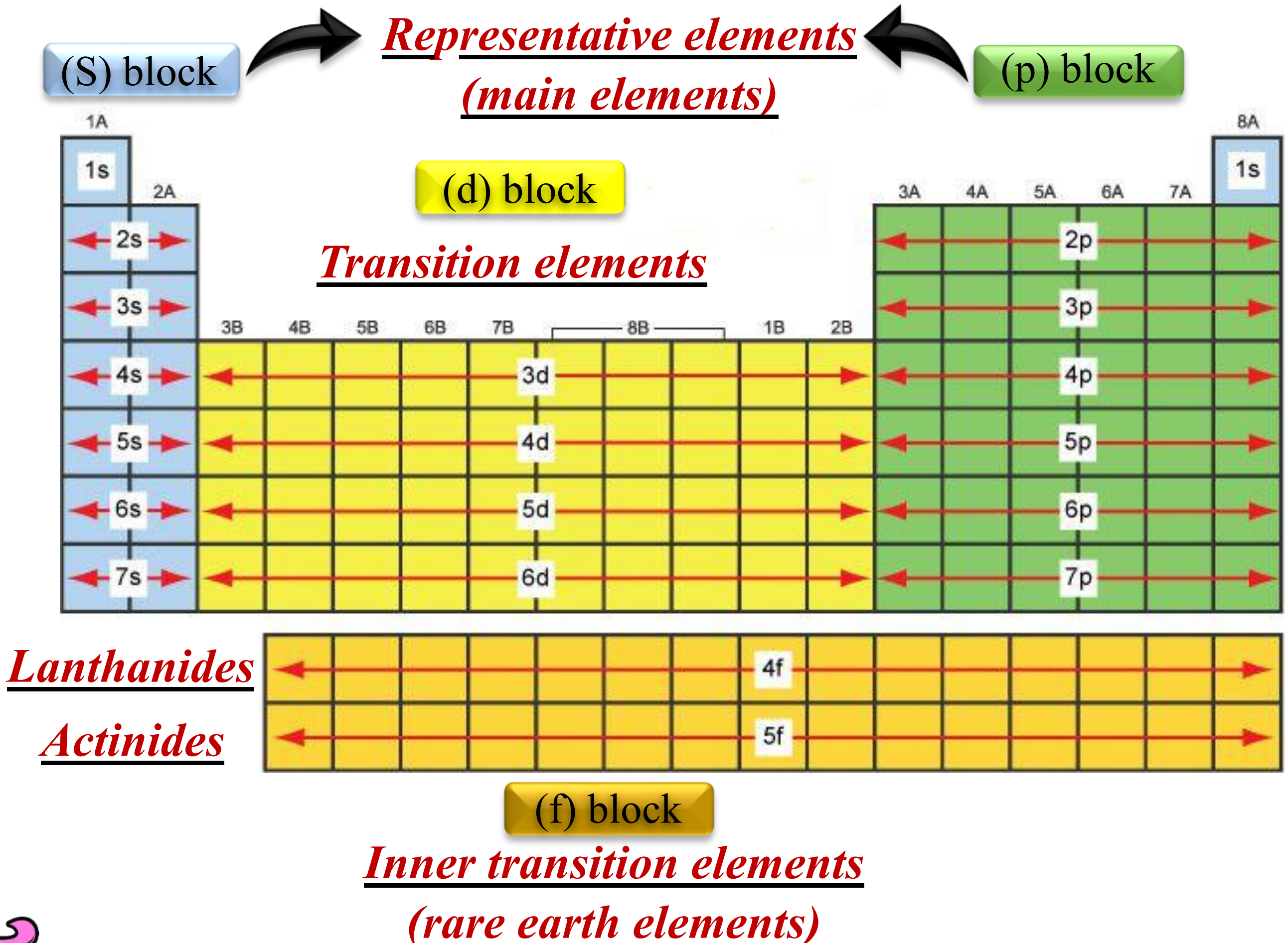
8 Gropus (A)

10 Gropus (B)

Periods

7 Horizental raws

Periodic Table



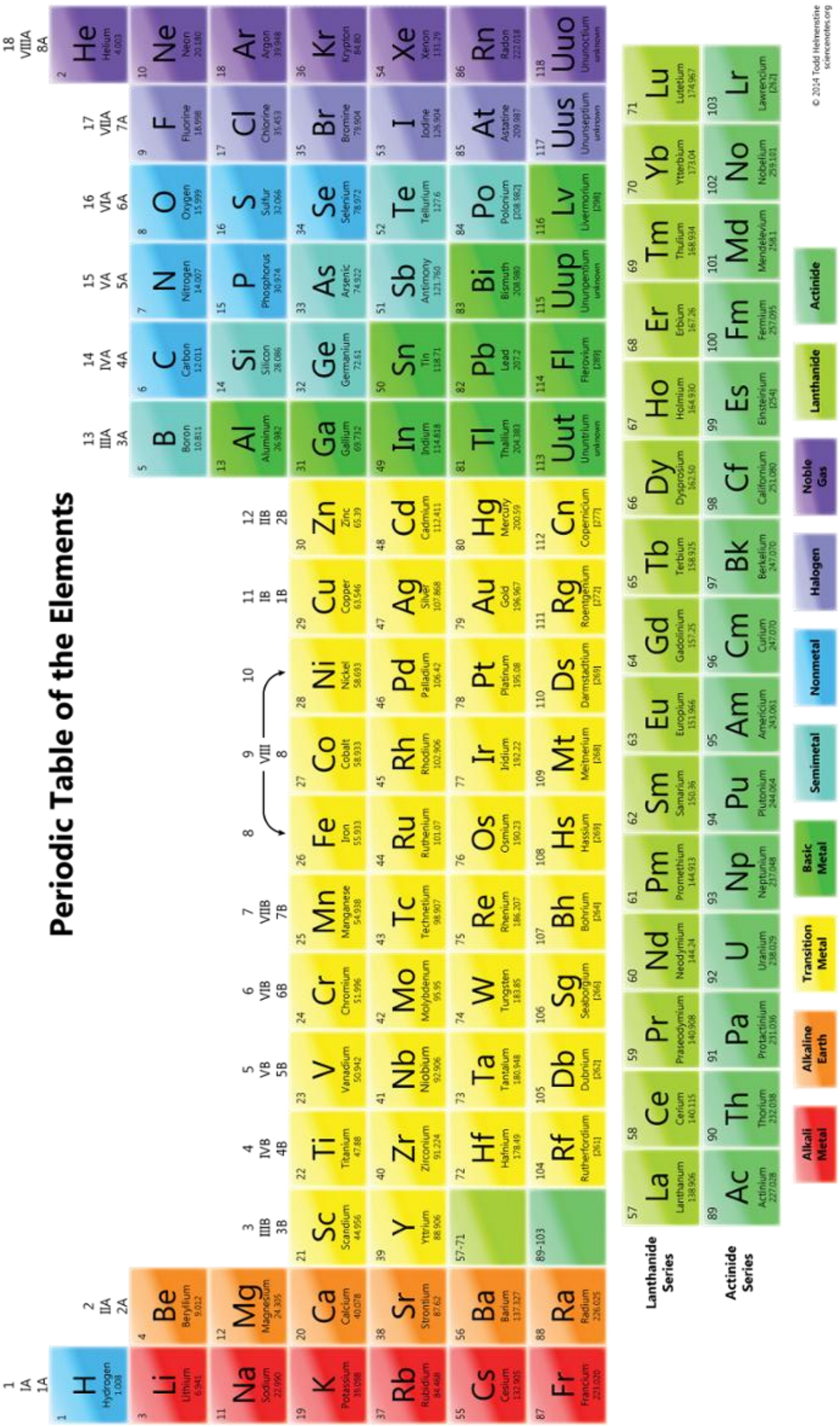
Name of groups

gp 1A (IA) ⟶ Alkali metal

gp 2A (IIA) ⟶ Alkaline earth metal

gp 7A (VIIA) ⟶ Halogens

gp 8A (VIIIA) ⟶ Nobel (inert) gases


Periodic Table of the Elements
1 H Hydrogen
2 He Helium
3 Li Lithium
4 Be Beryllium
5 B Boron
6 C Carbon
7 N Nitrogen
8 O Oxygen
9 F Fluorine
10 Ne Neon
11 Na Sodium
12 Mg Magnesium
13 Al Aluminum
14 Si Silicon
15 P Phosphorus
16 S Sulfur
17 Cl Chlorine
18 Ar Argon
19 K Potassium
20 Ca Calcium
21 Sc Scandium
22 Ti Titanium
23 V Vanadium
24 Cr Chromium
25 Mn Manganese
26 Fe Iron
27 Co Cobalt
28 Ni Nickel
29 Cu Copper
30 Zn Zinc
31 Ga Gallium
32 Ge Germanium
33 As Arsenic
34 Se Selenium
35 Br Bromine
36 Kr Krypton
37 Rb Rubidium
38 Sr Strontium
39 Y Yttrium
40 Zr Zirconium
41 Nb Niobium
42 Mo Molybdenum
43 Tc Technetium
44 Ru Ruthenium
45 Rh Rhodium
46 Pd Palladium
47 Ag Silver
48 Cd Cadmium
49 In Indium
50 Sn Tin
51 Sb Antimony
52 Te Tellurium
53 I Iodine
54 Xe Xenon
55 Cs Cesium
56 Ba Barium
57-71
72 Hf Hafnium
73 Ta Tantalum
74 W Tungsten
75 Re Rhenium
76 Os Osmium
77 Ir Iridium
78 Pt Platinum
79 Au Gold
80 Hg Mercury
81 Tl Thallium
82 Pb Lead
83 Bi Bismuth
84 Po Polonium
85 At Astatine
86 Rn Radon
87 Fr Francium
88 Ra Radium
89-103
104 Rf Rutherfordium
105 Db Dubnium
106 Sg Seaborgium
107 Bh Bohrium
108 Hs Hassium
109 Mt Meitnerium
110 Ds Darmstadtium
111 Rg Roentgenium
112 Cn Copernicium
113 Uut Ununtrium
114 Fl Flerovium
115 Uup Ununpentium
116 Lv Livermorium
117 Uus Ununseptium
118 Uuo Ununoctium
Lanthanide Series
57 La Lanthanum
58 Ce Cerium
59 Pr Praseodymium
60 Nd Neodymium
61 Pm Promethium
62 Sm Samarium
63 Eu Europium
64 Gd Gadolinium
65 Tb Terbium
66 Dy Dysprosium
67 Ho Holmium
68 Er Erbium
69 Tm Thulium
70 Yb Ytterbium
71 Lu Lutetium
Actinide Series
89 Ac Actinium
90 Th Thorium
91 Pa Protactinium
92 U Uranium
93 Np Neptunium
94 Pu Plutonium
95 Am Americium
96 Cm Curium
97 Bk Berkelium
98 Cf Californium
99 Es Einsteinium
100 Fm Fermium
101 Md Mendelevium
102 No Nobelium
103 Lr Lawrencium
Alkali Metal
Alkaline Earth
Transition Metal
Basic Metal
Semimetal
Nonmetal
Halogen
Noble Gas
Lanthanide
Actinide
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Groups

- Divided into 4 blocks where each block contains number of groups equal number of electrons in outer subshell

1) s-Block

- Contain two groups⟶ IA (Alkali metal) and IIA (Alkaline earth metal) & H and He

2) p-Block

- Last six groups from IIIA to VIII A

3) d-Block

- Consist of 10 groups (IB-VIIIB) and Known as transition metals

4) f-Block

- Consist of 14 groups Known as inner transition elements (Lanthanides and actinides)

Periods

1) Period 1 — 2 Elements (H and He)

2) Period 2 and 3 — Have 8 elements

3) Period 4 and 5 — Have 18 elements

4) Period 6 — Have 32 elements

5) Period 7 — Not completed yet due to new exotic or manmade elements

Position of element in periodic table (مهم)

A gp No of element — No of electrons in outer level (valence e's) ✓

B Period No of element — Principal quantum number (n) ✓

Example

1- $_{11}Na \rightarrow 1s^2, 2s^2, 2p^6, 3s^1$

Valence electron = 1

∴ gp No = 1

n = 3 of outer shell

∴ period No = 3

∴ (Na in gp 1A (IA) and in Period 3) ✓

Important Notes

✔ **Determine the Position of Transition element :**

Group Number in transition element = No of electrons exist in s & d due to outer level is (n-1s, nd) :

Except :

✔ No of e's in (s+d) = 8, 9 or 10 → ∴ gp VIIIB

✔ No of e's in (s+d) = 12 → ∴ gp IIB

✔ No of e's in (s+d) = 11 → ∴ gp IB

Families in Periodic table

A) Hydrogen

2) Colorless, diatomic gas H_2

3) React with nonmetals to form molecular compounds

Ex: $H_2 + Cl_2 \rightarrow 2HCl_{(g)}$

(HCl dissolves in H_2O to form acid [also other hydrogen halides]

Ex: $H_2 + O_2 \rightarrow H_2O_{(L)}$

1) React with metals to form metal hydrides (Ex: NaH)

B) Alkali metals

1) **Soft, Low melting points and low density**
2) **Very reactive elements so never exist alone in nature (often exothermic reaction)**
3) **React with water to form Basic (alkaline solution) and H_2 (g)**

Ex: $Na + H_2O \rightarrow NaOH + H_{2\,(g)}$

C) Alkali earth metals

3) **harder, higher melting points and denser than alkali metal**
2) **Less reactive than alkali metal**
1) **Form stable insoluble basic oxide (naturally extracted as basic oxide so-called alkaline earth metals)**

D) Halogens

1) **Reactive nonmetal diatomic molecules**
2) **F_2, Cl_2 → gases & Br_2→liquid & I_2→Solid**
3) **With metals form ionic compounds**
4) **With hydrogen forms acids→ size: HI > HBr > HCl > HF**

E) Nobel gases

1) Chemically inert **monoatomic** gases
3) **Very low** melting point and boiling point
2) **All gases** at room temperature

Atom loss or gain to be stable (nearest noble gas)

Atom

Loss e's → +Ve ion (Cation)

gain e's → -Ve ion (Anion)

Ex: Na^{+}, Mg^{2+}, Al^{3+}

$\therefore$ ($P > e^{-}$)

Ex: F^{-}, Cl^{-}, O^{2-}, S^{2-}

$\therefore$ ($P < e^{-}$)

H^{+}																	
li^{+}														N^{3-}	O^{2-}	F^{-}	
Na^{+}	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^{-}	
K^{+}	Ca^{2+}	SC^{3+}	Ti^{4+}	V^{5+} V^{4+}	Mn^{2+} Mn^{4+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^{+} Cu^{2+}	Zn^{2+}					Se^{2-}	Br^{-}	
Rb^{+}	Sr^{2+}							Pd^{2+}	Ag^{+}	Cd^{2+}			Sn^{2+} Sn^{4+}	Sb^{3+} Sb^{5+}	Te^{2-}	I^{-}	
Cs^{+}	Ba^{2+}							Pt^{2+}	Au^{3+}	Hg^{2+}			Pb^{2+} Pb^{4+}	Bi^{3+} Bi^{5+}			

Molecule

Aggregate between two atoms in definite arrangement by chemical force (chemical bond)

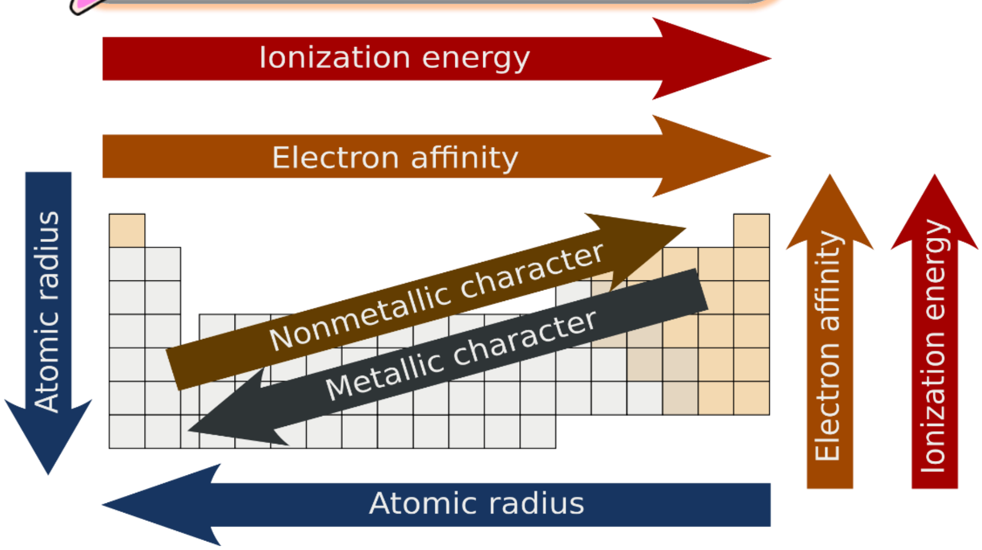
(Two atoms) **Diatomic** — Ex: N_2, H_2, O_2

Polyatomic (>Two atoms) — Ex: NO_2, H_2O, CO_2

Notes

- **Zn, Cu, Fe, I and Co exist in the human body in trace amount (0.1%) but are very important in:**
  1) Growth 2) Transport of oxygen 3) metabolism
  4) Defense against disease

Properties of Periodic table



Effective Nuclear Charge (Z_{eff})

- Nuclear Charge ⟶ Atomic Number (Z)
- Valence electrons ⟶ Number of e's in valence shell (outer shell)
- Inner electrons ⟶ Shielding e's (σ) or (s) (Number of electrons in inner shells)

Effective Nuclear Charge (Z_{eff})= Z-S = (Valence electrons)

Important Notes

1) Inner e's (shielding e's) ⟶ <u>shield (تحجب) the effect of nuclear charge on valence e's</u>.

2) As (Z_{eff}) decreasing, the effect of nuclear charge on valence e's decreasing also:

   ∴ <u>Ionization (loss e's) become easily</u>.

3) Shielding effect through period remain constant while in group increase.

EXAMPLE

Ex (1): $_{6}C \rightarrow \underbrace{1S^2}\ 2S^2\ 2P^2$ ∴ **Valence electrons = 4 e⁻'s**

∴ **Shielded electrons = 2 e⁻'s**

(Z_{eff})= Z-S = 6-2 =4 (Valence electrons)

Ex (2): $_{11}Na \rightarrow \underbrace{1S^2\ 2S^2\ 2P^6}\ 3S^1$ ∴ **Valence electrons = 1 e⁻**

∴ **Shielded electrons = 10 e⁻'s**

(Z_{eff})= Z-S = 11-10 =1 (Valence electrons)

Ex (3): $_{12}Mg \rightarrow \underbrace{1S^2\ 2S^2\ 2P^6}\ 3S^2$ ∴ **Valence electrons = 2 e⁻'s**

∴ **Shielded electrons = 10 e⁻'s**

(Z_{eff})= Z-S = 12-10 = 2 (Valence electrons)

∴ **gp (IIA) and Period = 3**

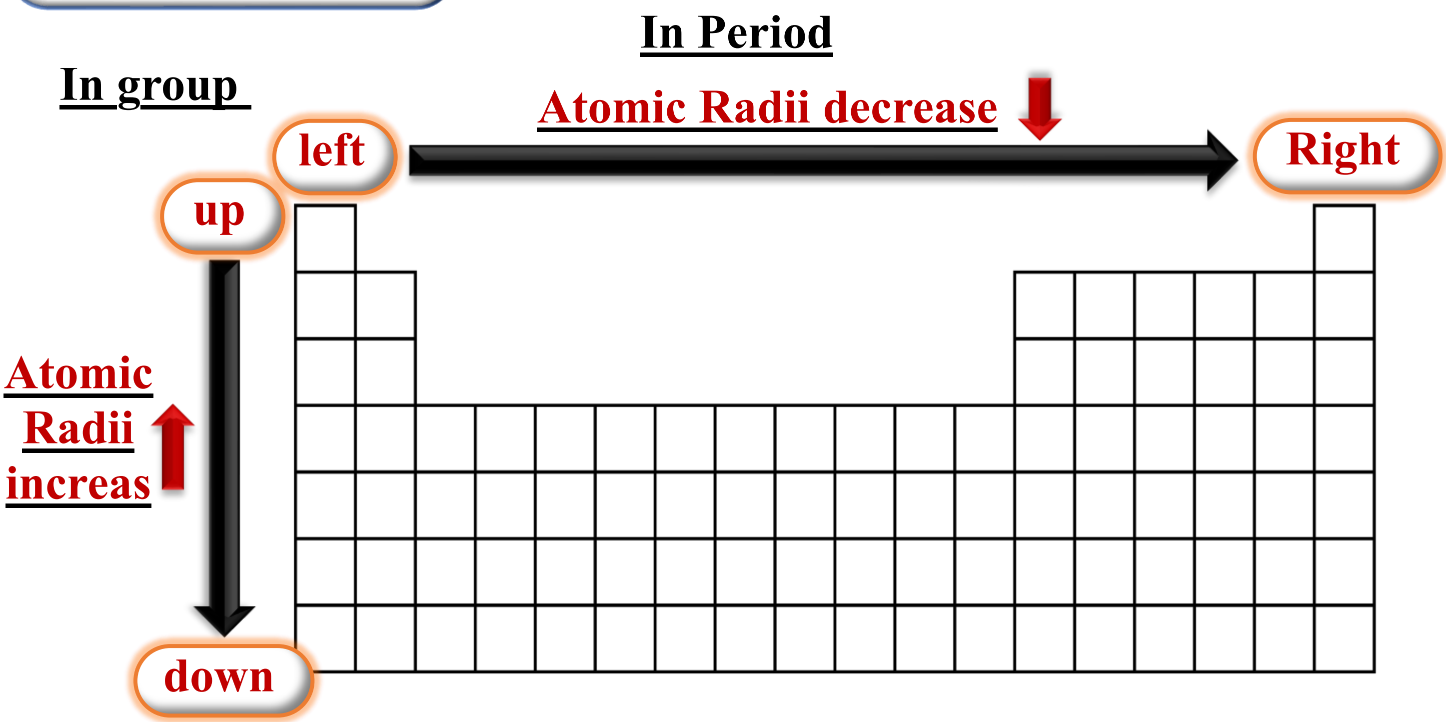
Ex (4): $_{10}Ne \rightarrow \underbrace{1S^2}\ 2S^2\ 2P^6$ ∴ **Valence electrons = 8 e⁻'s**

∴ **Shielded electrons = 2 e⁻'s**

(Z_{eff})= Z-S = 10-2 = 8 (Valence electrons)

Properties of elements in Periodic table

1) Atomic radius



In **Period** (From Left→ Right)

Atomic radii **decrease** due to Z_{eff} **increase**

Ex: Radius of Li > Be > C > N > O > F

Z_{eff}: +1 +2 +3 +4 +5 +6

In **group** (From Up→ Down)

Atomic radii **Increase** due to Number of shell **increase** and Z_{eff} **remain constant**

Ex: Radius of Li < Na < K < Rb < Cs

Z_{eff}: +1 +1 +1 +1 +1

Important Notes

Radius of -Ve ion	>	Radius of neutral atom	>	Radius of +Ve ion
Repulsion between e's increase				Repulsion between e's decrease

Ex (1) : Radius of Na > Na^+

Ex (2): Radius of F^- > F

Isoelectronic Series (Same number of electrons)

Higher +Ve ion → Atomic Radii decrease → Higher -Ve ion

Ex (1) : Arrangement radius of Cl^- , Ca^{2+}, K^+, S^{2-}

∴ $S^{2-} > Cl^- > K^+ > Ca^{2+}$

Ex (2) : Arrangement radius of Cl^- , Ar, Ca^{2+}

∴ $Cl^- > Ar > Ca^{2+}$

Ex (3) : Arrangement Na^+, F^- , Al^{3+}, Mg^{2+},O^{2-}, Ne

∴ $O^{2-} > F^- > Ne > Na^+ > Mg^{2+}$

2) Ionization Energy (IE)

The energy required to remove an electron from an atom in a gaseous state.

Ex: $H_{(g)} \xrightarrow{IE} H^+ + e^-$

In Period (From Left→ Right)

IE increase due to Z_{eff} increase and atomic radius decrease

In group (From Up→ Down)

IE decrease due to Z_{eff} constant , number of shells increase and atomic radius increase

$$IE = \frac{1}{Size\ (radius)}$$

Important Notes

- **Nobel gases (gp VIII A) ns^2, np^6 → having highest IE :**
  - ✔ **Smallest radius** ✔ **Highest Z_{eff}** ✔ **Completely filled shell**

Exception

1 **IE of gp (IIA) > gp (IIIA)**

Be, Mg — B, Al

Reason

gp IIA

Ex: Be → $1S^2$ $2S^2$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

∴ Completely filled subshell

∴ More Stable

∴ Needed higher IE to remove an electron

gp IIIA

Ex: B → $1S^2$ $2S^2$ $2p^1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\,|\;\;|\;\;}$

∴ Partially filled P-subshell

∴ Less Stable

∴ Needed lower IE to remove an electron

2 IE of gp (VA) > gp (VIA)

gp (VA): N, P

gp (VIA): O, S

Reason

gp VA

Ex: $_7N \rightarrow 1S^2 \; 2S^2 \; 2p^3$

[⇅] [⇅] [↿ | ↿ | ↿]

∴ Half-filled subshell

∴ More Stable

∴ Needed higher IE to remove an electron

gp VIA

Ex: $_8O \rightarrow 1S^2 \; 2S^2 \; 2p^4$

[⇅] [⇅] [⇅ | ↿ | ↿]

∴ Partially filled P-subshell

∴ Less Stable

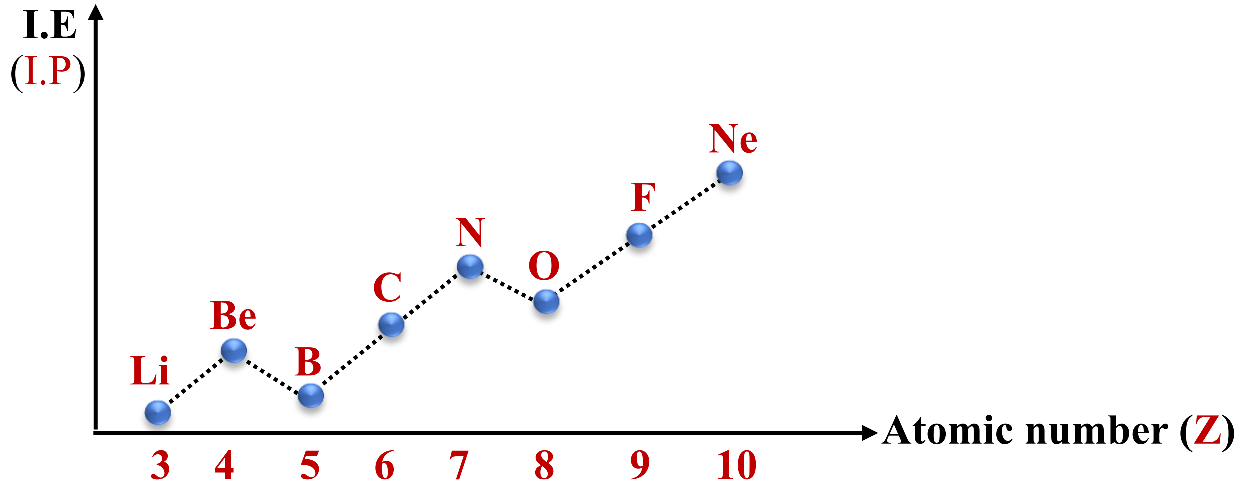
∴ Needed lower IE to remove an electron

Important Notes

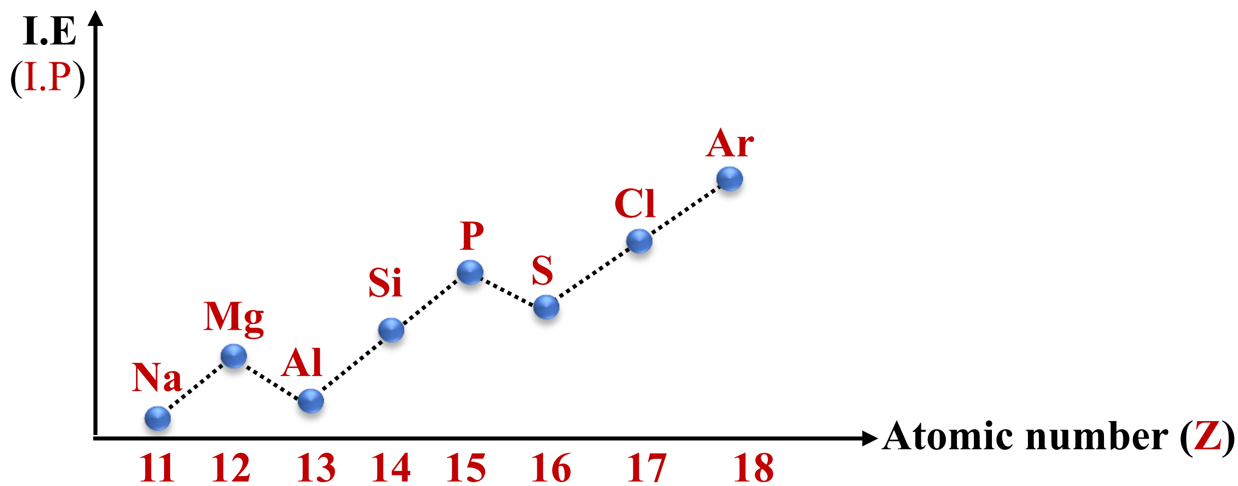
- ❖ $Li \longrightarrow Li^+ + e^-$ ∴ (1st IE)
- ❖ $Li^+ \longrightarrow Li^{2+} + e^-$ ∴ (2nd IE)
- ❖ $Li^{2+} \longrightarrow Li^{3+} + e^-$ ∴ (3rd IE)

∴ 3rd IE > 2nd IE > 1st IE

∴ due to Z_{eff} of $Li^{2+} > Li^+ > Li$



Draw the relation between Ionization Rnergy and the atomic number of 2nd Period:



Draw the relation between Ionization Energy and the atomic number of 3rd Period:

3) Electron Affinity (EA)

The charge in energy when an election is added to an atom in gaseous state.

When e^- is easily added → (**High EA**)

∴ EA = more -ve value (released → Exothermic)

Ex: $F_{(g)} + e^- \longrightarrow F_{(g)}^-$ EA = -328 KJ/mol

When e^- is difficulty added → (**Low EA**)

∴ EA = less -ve value (or zero)

Ex: $B_{(g)} + e^- \longrightarrow B_{(g)}^-$ EA = -23 KJ/mol

$N_{(g)} + e^- \longrightarrow N_{(g)}^-$ EA = Zero

When e^- is high difficult added → (**Energy need**)

∴ EA = + ve value (Endothermic → very low EA)

Ex: $Be_{(g)} + e^- \longrightarrow Be_{(g)}^-$ EA = +241 KJ/mol

Important Notes

EA
- -Ve Exothermic → ∴ e^-s easily added
- +Ve Endothermic → ∴ e^-s difficult added

In Periodic Table

In period : EA ↑ increase (more -ve)

In Group : EA ↓ decrease (less -ve)

If e^- is added

Easily	Difficulty
∴ Higher EA	∴ Lower EA

A) In period (from left →to right): (EA↑)

Radius (size) ↓ decrease and Z_{eff} ↑ increase

∴ e is easily added ∴ EA ↑ increase (more ↑ -ve)

B) In Group (from up →to down): (EA↓)

Radius (size) ↑ increase

∴e is difficulty added ∴ EA ↓ decrease (less -ve)

Important Notes

In Group :

Atomic No (Z) ↑
But Z_{eff} Const

∴Radius↑ ∴IE↓

∴EA↓

In period

Atomic No (Z) ↑ → Z_{eff} ↑

∴ Radius↓ ∴ IE↑ ∴ EA↑

Exception

1 EA of gp (IA) > gp (IIA)

gp (IA)	gp (IIA)
Li	Be
Na	Mg

Reason

gp IA

Ex: Li → $1S^2$ $2S^1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ ← e-

∴ Half-filled subshell

∴ less Stable

∴ e^- is easily added to $2s^1$ to be $2s^2$ (completely filled) (stable)

∴ Higher EA

gp IIA

Ex: Be → $1S^2$ $2S^2$ $2p^0$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{}\boxed{}\boxed{}$ ← e-

∴ Completely filled s-subshell

∴ More Stable

∴ If e- added $2p^0$ to be $2p^1$ (partially filled) less stable

∴ Difficult to add e^-

∴ Low EA

2 EA of gp (IVA) > gp (VA)

gp (IVA): C, Si

gp (VA): N, P

Reason

gp IVA

Ex: ${}_6C \rightarrow 1S^2 \; 2S^2 \; 2p^2$

$[\uparrow\downarrow] \; [\uparrow\downarrow] \; [\uparrow \mid \uparrow \mid \;]$ ← e-

∴ Partially -filled subshell

∴ Less Stable

∴ e^- is added to $2p^2$ to be $2p^3$ (half filled, more stable)

∴ e^- is easily added

∴ **Higher EA**

gp VA

Ex: ${}_7N \rightarrow 1S^2 \; 2S^2 \; 2p^3$

$[\uparrow\downarrow] \; [\uparrow\downarrow] \; [\uparrow \mid \uparrow \mid \uparrow]$ ← e-

∴ Half-filled P-subshell

∴ More Stable

∴ e^- is added to $2p^3$ to be $2p^4$ (partially filled, less stable)

∴ e^- is hardly added

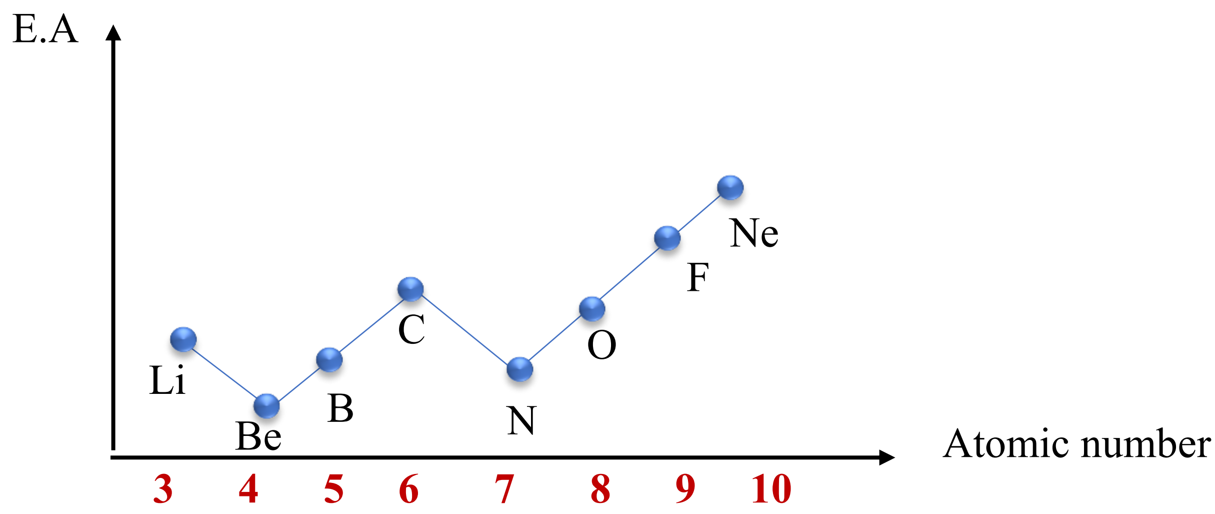
∴ **Lower EA**

3 EA of 2nd period < EA of 3rd period

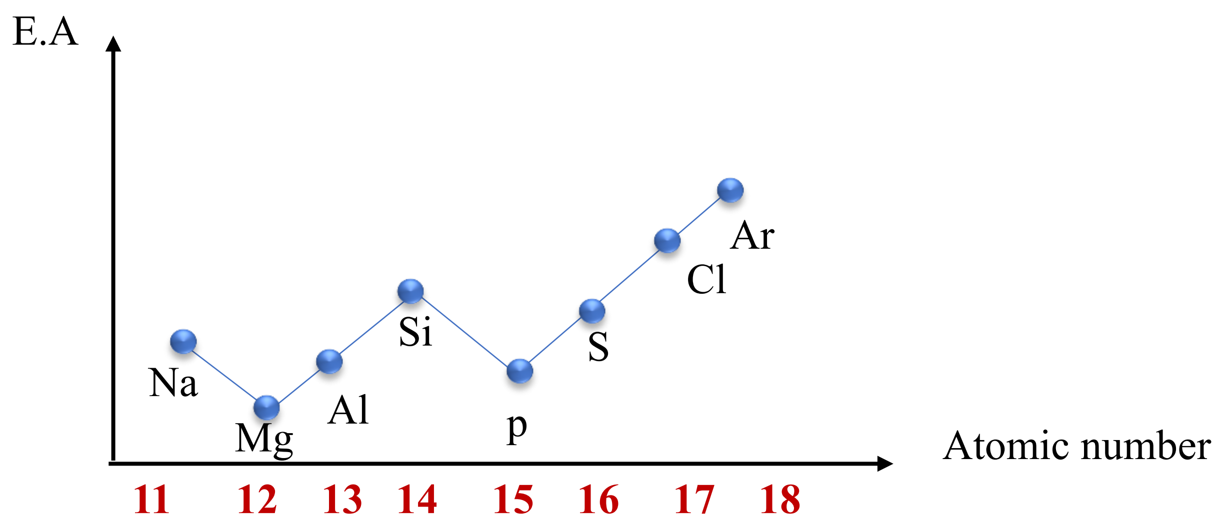
(Li → F) (Na → Cl)

Due to the **smaller size** of elements in 2nd period so a strong **repulsion** between the e`s of atom and in Coming e^-.

∴ e^- is added difficulty. ∴ Lower EA.



Draw the relation between Electron affinity and the atomic number of 2nd Period:



Draw the relation between Electron affinity and the atomic number of 3rd Period:

Quiz

1) EA of F < Cl why?

Answer

Due to the smaller size of the F-atom and Strong repulsion between the incoming e^- and e^-'s of F-atom.

∴ e^- is added difficulty ∴ Low EA

2) Arrange the Following atoms according to descending in their radii:

$_{11}Na$, $_{13}Al$, $_{17}Cl$, $_{39}Rb$

Answer

∵ From periodic table

Na, Al, Cl (same period) Na > Al > Cl

Na, Rb (same group) Rb > Na

∴ $_{39}Rb > {}_{11}Na > {}_{13}Al > {}_{17}Cl$

4) Electron Negativity (EN)

Ability of an atom in molecule to withdraw (attract bonded electrons to its).

$$EN = \frac{1}{Radius\ (Size)}$$

In Group :

EN ↓ (decrease)

radius increase ↑

EN of F > Cl > Br > I

In period

EN ↑ (increase)

because radius decrease ↓

EN of F > O > N > C > H

EN Bond polarity and type of Bond:

Non-polar covalent

Sharing of e^-s between similar atoms.

♦ (Difference in EN = 0)

Polar covalent

Partial transfer of e^-s

♦ Difference in EN ($0 < \Delta EN < 2$)

Ionic

completely transfer of e^-s.

♦ Difference in EN >2

Example

Compound	F_2	HF	LiF
EN-difference	$4 - 4 = 0$	$4 - 2.1 = 1.9$	$4 - 1 = 3$
Type of bond	Non-polar covalent F — F (same atoms)	Polar covalent $H \rightarrow F$ (EN of F >H)	Ionic Bond Li^+ F^-

Important Notes

1] **Polarity of bond α EN**

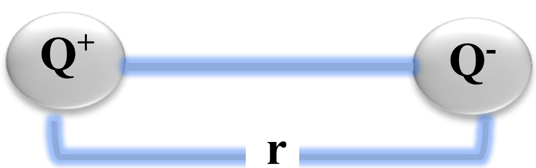
Ex: Polarity of $C^{+\delta} \rightarrow F^{-\delta} > C^{+\delta} \rightarrow Cl^{-\delta} > C^{+\delta} \rightarrow Br^{-\delta}$

2] **Polarity of bonds depend on dipole moment: (μ)**

$$\mu = Q\ r$$

Charge

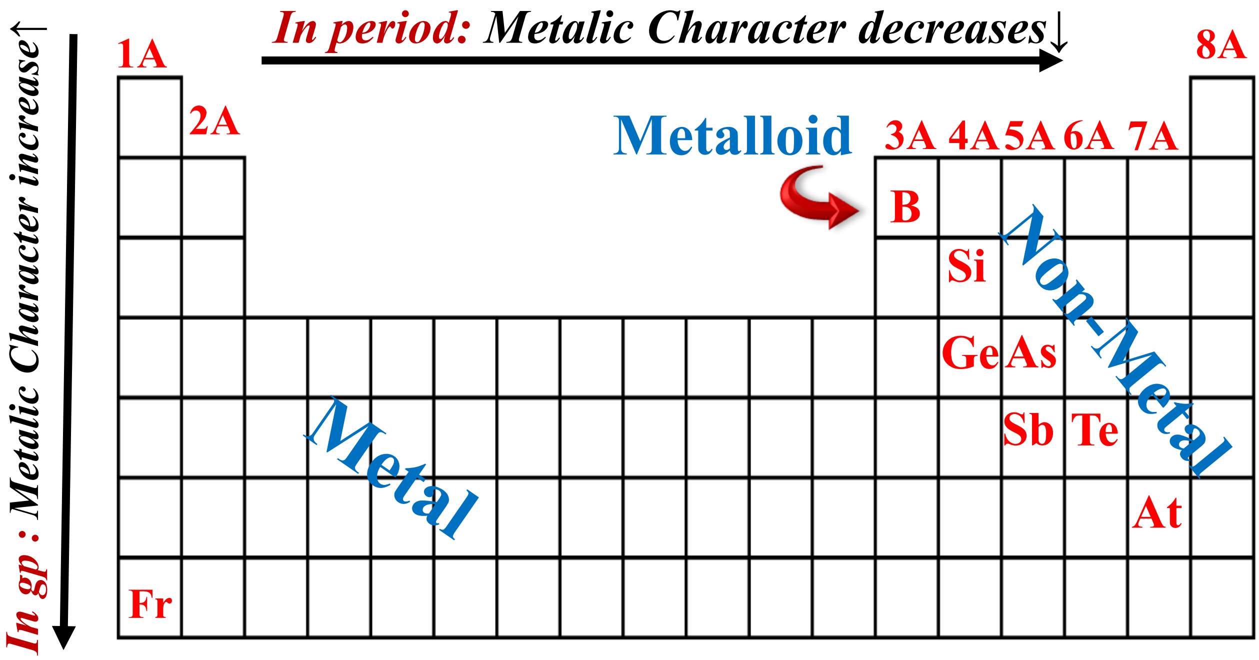
Distance between two atoms



$$\therefore \mu \begin{cases} = 0 \text{ (Non-polar)} \\ \neq 0 \text{ (Polar)} \end{cases}$$

unit: Debye

5) Metallic Character:



Metals	Non- Metals	Semi-metal (Metalloid)
♣ left of the table ♣ Lustrous (shine) لامع ♣Malleable قابل للطرق ♣Ductile قابل للسحب ♣Conduct electricity ♣Conduct heat ♣Solid Except (Hg) is a liquid ♣Form ionic and basic oxide ♣Low EN, IE, EA (lose e⁻s → +ve ions) Fr is the most metallic	♣ Right of table ♣ Non- Lustrous ♣ Brittle, Hard or soft. ♣Non-Conductors ♣Poor Conductors ♣Solid, liquid or gas ♣Form covalent and acidic oxides ♣High IE & EA (gains e⁻s → -ve ions) F is most Non-metallic	♣Borderline Intermediate properties Semi-conductors B, Si, Ge, As, Sb, Te, At.

6) Inert pair effect: (inert-s-pair effect)

- **Elements (In, Sn, Sb, TL, Pb, Bi and Po) are less reactive than expected.**
- **The maximum oxidation states of these elements are less than expected by 2.**
  - **Because of $5s^2$ or $6s^2$ 2 are inert. (not lost e^-s)**

III A	IV A	V A	
In	Sn	Sb	
Tl	Pb	Bi	Po

Tl: $ns^2\ np^1$

Pb: $ns^2\ np^2$

Po: $ns^2\ np^3$

Tl
- +3 expected (not Found) X
- +1 more stable √

Pb
- +4 expected (not Found) X
- +2 more stable √

Bi
- +5 expected (not Found) X
- +3 more stable √

6) Diagonal Relationship

Refers to similarities between elements of different groups and periods.

Li	Be	B	C
Na	Mg	Al	Si

The Chemistry of:

$$Li \xrightarrow{\text{similar to}} Mg$$

$$Be \longrightarrow Al$$

$$B \longrightarrow Si$$

❖ **These Similarities are due to the similarity in their:**

A- EN

B- Polarizing power (Charge density)

$$\text{Polarizing Power} = \frac{\text{Charge (Z)}}{\text{Radius (r)}}$$

Summary

In Group :

Atomic No (Z)

(increasing No of shell)

Z_{eff} Const

∴Radius (size)↑

∴IE↓, EA↓, EN↓

In period

Atomic No (Z) ↑ (same No of shell)

Z_{eff} ↑

∴Radius (size)↓ ∴IE↑, EA↑, EN↑

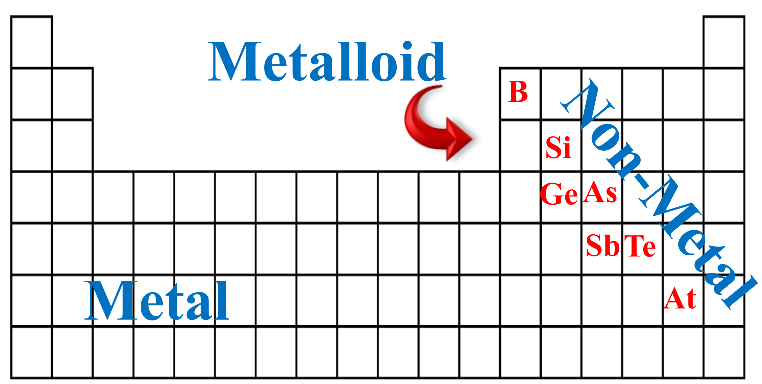
∴ Metallic character ↓

Exception in IE:

1) IE of gp IIA > IIIA
2) IE of gp VA > VI A (N > O)

Exception in EA:

1) EA of gp IA > IIA (Li > Be)
2) EA of gp IVA > V A (C > N)
3) EA of period (2) < 3rd (F < Cl)



Inert Pair Effect:

$5s^2$, $6s^2$ are inert

(In, Sn, Sb, Tl, Pb, Bi, Po)

Pb	+4 X	Bi	+5 X
	+2 √		+3 √

Diagonal Relationship:

Li ↔ Mg, Be ↔ Al, B ↔ Si

Li	Be	B
Mg	Al	Si

Quiz

1. **How many f orbitals have the value n = 3?**

   A) 3 B) 7 C) 0 D) 1 E) 5

2. **How many electrons can be contained in all of the orbitals with n = 4?**

   A) 10 B) 2 C) 18 D) 32 E) 8

3. **Which of the following electron configurations are different from those expected?**

   A) V B) Ca C) Cr D) Ti E) Sc

4. **How many electrons in an atom can have the quantum numbers n = 4, l= 2?**

   A) 12 B) 10 C) 6 D) 5 E) 14

5. **Which of the following atomic symbols is incorrect?**

   A) $_{19}K$ B) $_{8}N$ C) $_{17}Cl$ D) $_{15}P$ E) $_{6}C$

6. **Which metal ion has a d^5 electron configuration?**

   A) Os^{2+} B) Co^{2+} C) Pd^{2+} D) Ag^{+} E) Fe^{3+}

7. **Fe has ………….that is (are) unpaired in its d orbitals.**

   A) 4 electrons B) 1 electron C) 2 electrons
   D) 3 electrons E) none of these

8. **Which of the following sets has elements with the most nearly identical atomic radii?**

   A) Mg, Ca, Sr, Ba B) C, P, Se, I C) Cr, Mn, Fe, Co
   D) Ne, Ar, Kr, Xe E) Be, B, C, N

9. **Which of the following atomic symbols is incorrect?**

   A) $^{39}_{19}K$ B) $^{14}_{8}N$ C) $^{36}_{17}Cl$ D) $^{32}_{15}P$ E) $^{14}_{6}C$

10. Which metal ion has a d^1 electron configuration?
A) Os^{2+} B) Sc^{2+} C) Pd^{2+} D) Ag^{+} E) Fe^{3+}

11. For the elements Rb, F, and O, the order of increasing electronegativity is:

A) O < F < Rb B) F < Rb < O C) Rb < O < F D) Rb < F < O

13. Which of the following is a d^7 ion

A) Cu(II) B) At least two of the above (a-d) are d^7 ions.

C) Mn(IV) D) Co(II) E) Mn(II)

14. In which groups do all the elements have the same number of valence electrons?

A) None B) P, S, Cl C) Ag, Cd, Ar D) Na, Ca, Ba E) P, As, Se

15. For which of the following elements does the e's configuration for the lowest energy state show a partially filled d orbital?

A) Rb B) Ga C) Kr D) Cu E) Ti

16. Which of the following are incorrectly paired?

A) Ni, transition metal B) O, halogen C) K, alkali metal

D) Ba, alkaline earth metal E) Ne, noble gas

17. Which of these is an isoelectronic series?

A) Na^+, Mg^{2+}, S^{2-}, Cl^- B) Li, Be, B, C C) K^+, Ca^{2+}, Ar, S^{2-}

D) Na^+, K^+, Rb^+, Cs^+ E) none of these (a-d)

18. How many electrons can be described by the quantum numbers n = 3, l = 3, ml = 1?

A) 14 B) 10 C) 2 D) 0 E) 6