In Today's Class:

- ★ Perfect for Beginners
- ★ Good for Boards, JEE Main
- ★ Will Build a Good Level for JEE Advanced too
- ★ All Important Concepts Covered
- ★ Lots of Learning via Questions (Like an IITian)
- ★ High Focus on Understanding
- ★ Lots of Question Practice
- ★ HW Questions will be given

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PAHULPSG

Class 9th

NCERT

⇒ Dalton → atom

⇒ electron, proton, neutron

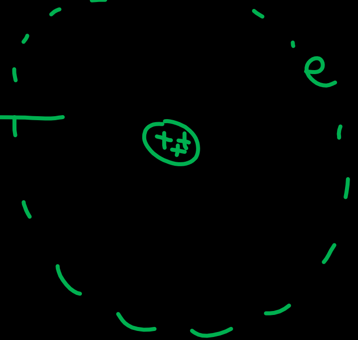
⇒ J J Thompson

⇒ Geiger - Marsden

Gold foil experiment by Rutherford

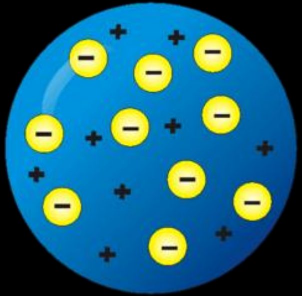
★ atom has a nucleus

★ most of the atom is empty space.

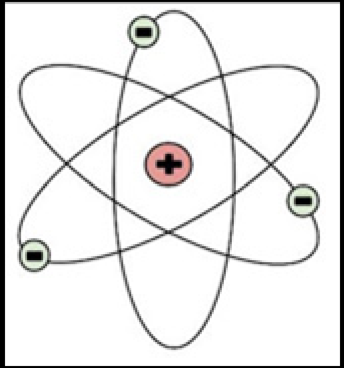


Old Models of the Atom

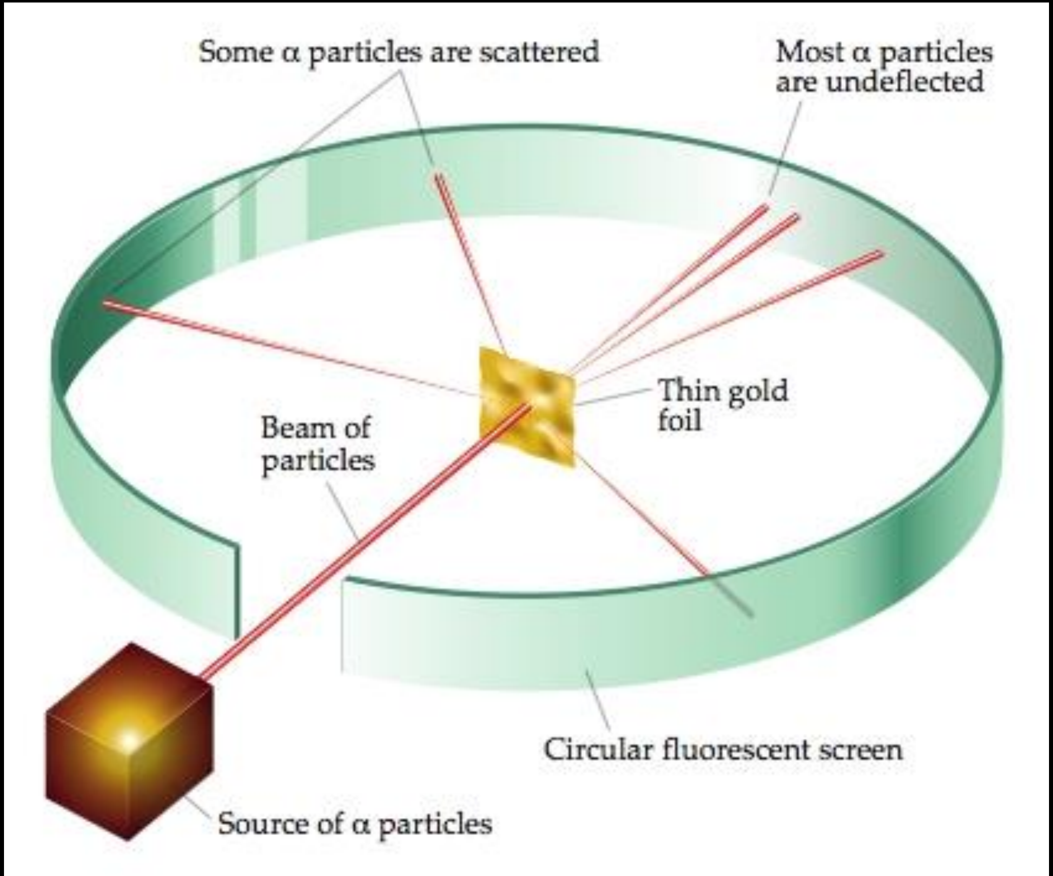
1. **JJ Thomson's Model**



2. **Ernest Rutherford's Model**



Rutherford's Experiments

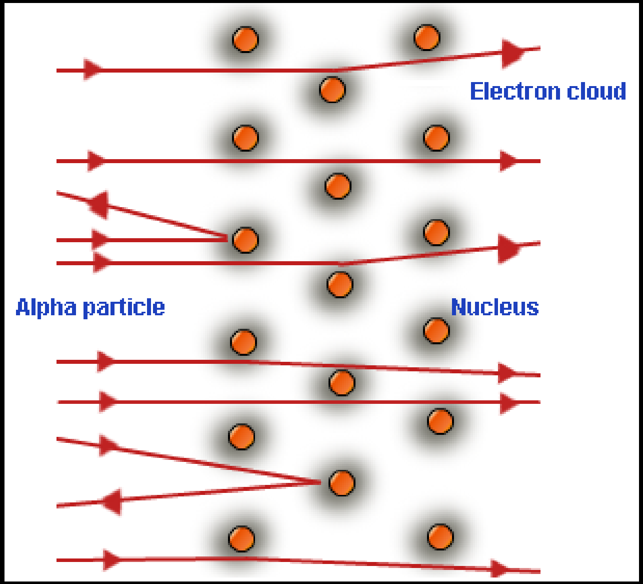


- The **Geiger–Marsden experiment(s)** (also called the **Rutherford gold foil experiment**) were a landmark series of experiments by which scientists discovered that **every atom contains a nucleus where all of its positive charge and most of its mass are concentrated**. They deduced this by measuring how an alpha particle beam is scattered when it strikes a thin metal foil. The experiments were performed between 1908 and 1913 by Hans Geiger and Ernest Marsden **under the direction of Ernest Rutherford** at the Physical Laboratories of the University of Manchester.

- To check the conclusions of Thomson's model, they bombarded a **thin sheet of Gold by fast moving alpha particles** coming from a radioactive source and observed their deviations after passing through the foil.

- Alpha-particles are positively charged helium nuclei with atomic mass 4 a.m.u. Gold foils are of thickness around 400 nm.

Rutherford's Experiments



1. Most of the space inside an atom is empty.

1. Some positively charged region in the atom.

1. Positively charged region was very small as compared to the atom.

The biggest achievement of the model was the **discovery of the Nucleus**. The order of diameter of a nucleus is 10^{-15} m and that of an atom is about 10^{-10}m.

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg	Mass/u	Approx. mass/u
Electron	e	-1.6022×10^{-19}	-1	9.10939×10^{-31}	0.00054	0
Proton	p	$+1.6022\times10^{-19}$	+1	1.67262×10^{-27}	1.00727	1
Neutron	n	0	0	1.67493×10^{-27}	1.00867	1

$$\boxed{e = 1.6\times10^{-19} \text{ Coulombs}}$$

$$m_e = 9.1\times10^{-31}\text{ kg}$$

$$m_p \approx m_n = 1.67\times10^{-27}\text{ kg.}$$

Isobars

Elements with **same mass number(A)** but with **different atomic number (Z)** are called isobars.

$\rightarrow \boxed{p+n}$ $\rightarrow$ p

* p & n are also called nucleons

$^{40}_{18}Ar$ *and* $^{40}_{20}Ca$

$\rightarrow$ 18 p
22 n

$\rightarrow$ p = 20
n = 20

Example of a series of isobars would be ^{40}S, ^{40}Cl, ^{40}Ar, ^{40}K, and ^{40}Ca. The nuclei of these nuclides all contain 40 nucleons; however, they contain varying numbers of protons and neutrons

Isotones

Elements with **same number of neutrons** are called isotones.

$^{40}_{20}Ca$ and $^{37}_{17}Cl$

$40-20$

$= 20$ neutrons

$^{23}_{11}Na$ and $^{24}_{12}Mg$

Isoelectronic Species

Ions which have **same number of electrons** are called isoelectronic species.

N^{3-} O^{2-} F^{1-} Ne Na^{+} Mg^{2+}

$\boxed{e^- = 10 \text{ each}}$

Electromagnetic Radiations (waves)

Electric field
wave

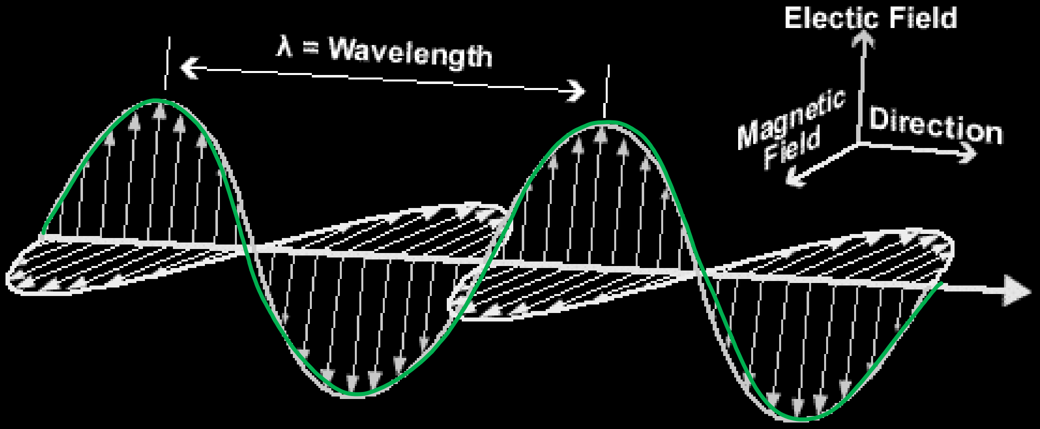
Magnetic field
wave

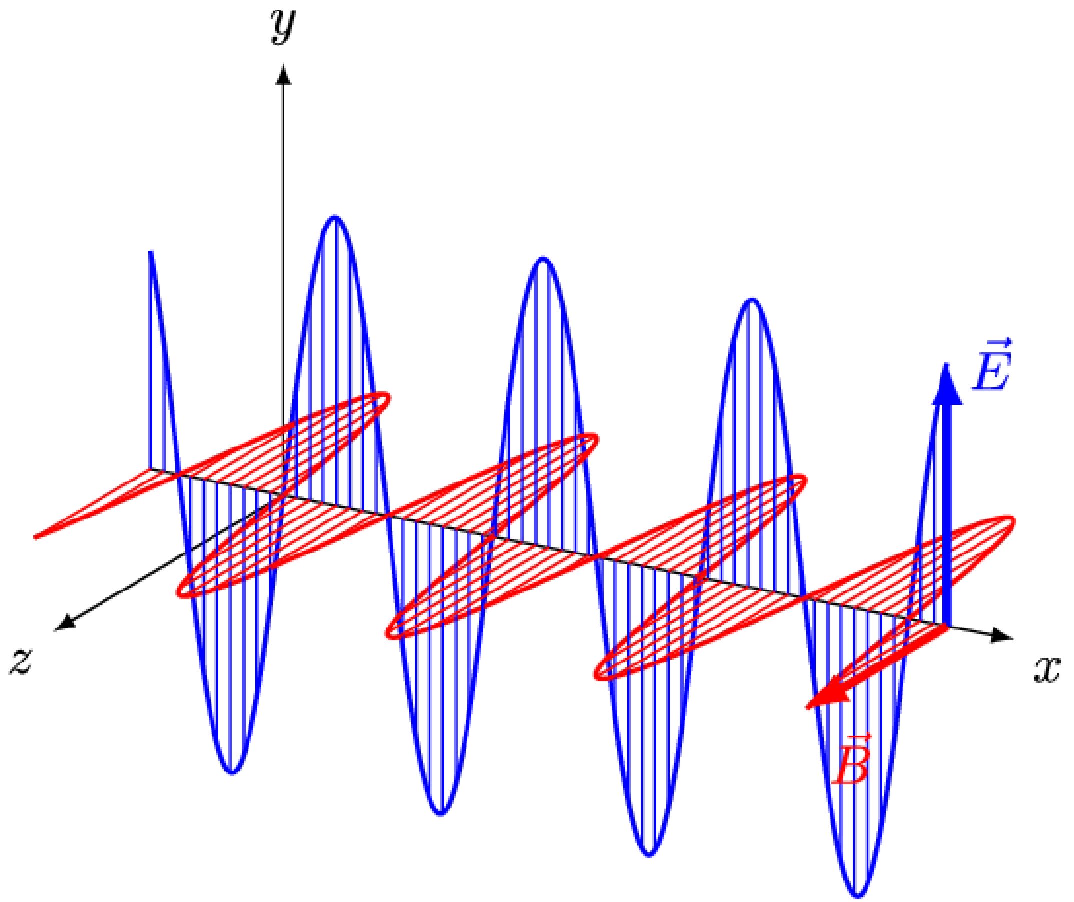
✓ Radio waves
✓ Microwaves
✓ Infrared rays
✓ visible light
✓ UV light
✓ X-Rays

high frequency

Electromagnetic Radiations

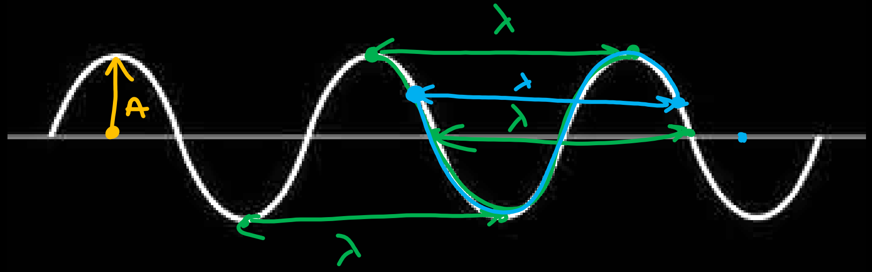
Radiations which consist of an **electric field (E)** and a **magnetic field (B)** oscillating perpendicular to each other and both perpendicular to the direction of propagation.



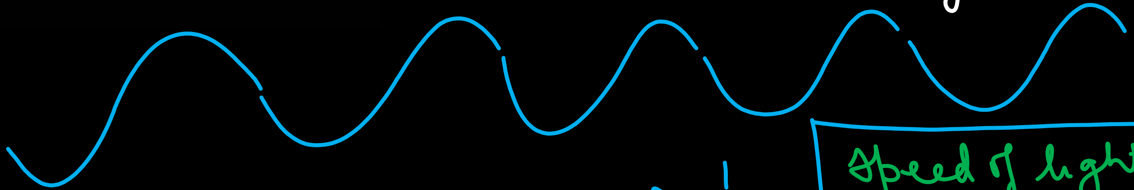


y
$\vec{E}$
z
x
$\vec{B}$

Characteristics of a Wave



$\lambda \Rightarrow$ is the shortest part of the wave that can be repeated to create the entire wave again



$$\text{speed} = \underbrace{\lambda}_{\text{length}} \times \nu$$

$$\nu \rightarrow \frac{1}{\text{time}}$$

$$\boxed{\text{Speed of light (em radiations)} = 3 \times 10^{8}\ \text{m/s}}$$

ν nu

λ lambda

$\bar{\nu}$ nu bar

Main Characteristics of a Wave

- Wavelength (λ)
- Frequency (ν)
- Velocity (v)
- Wavenumber ($\bar{\nu}$) $= \frac{1}{\lambda}$
- Amplitude (A)

no of waves per unit length

Let's Solve.

$\star$ $1 nm = 10^{-9} m$

The wavelength of 400 nm of an EM radiation does not correspond to:

(A) Frequency = 7.5×10^{14} Hz ✓

(B) Wave number = $2.5 \times 10^{6}\ m^{-1}$ ✓

(C) Velocity = $3 \times 10^{8}\ ms^{-1}$ ✓

(D) $\lambda = 40$ Å ✗

$c = \nu \lambda$

$3 \times 10^8\ m/s = \nu \cdot 400 \times 10^{-9}\ m$

$\nu = \dfrac{3 \times 10^8}{400 \times 10^{-9}}$ $\dfrac{1}{s}$ = Hertz

1 Angstrom (Å) = 10^{-10} m

400×10^{-9} m

4000×10^{-10} m = 4000 Å

$\bar{\nu} = \dfrac{1}{\lambda} = \dfrac{1}{400 \times 10^{-9}\ m} =$ m^{-1}

$= \dfrac{10^{7}}{400} = \dfrac{25 \times 10^6}{4}$

Which of the following radiations have lesser wavelength than the visible light?

- **A** Infrared
- **B** UV
- **C** Microwave
- **D** Radio wave

The frequency of yellow light having wavelength 600 nm is

(A) 5.0×10^{14} Hz

(B) 2.5×10^{7} Hz

(C) 5.0×10^{7} Hz

(D) 2.5×10^{14} Hz

$c = \nu \lambda$

$\nu = \frac{c}{\lambda}$

$= \frac{3 \times 10^{8} \text{ m/s}}{600 \times 10^{-9} \text{ m}}$

The mass to charge ratio for A^{+} is 1.97×10^{-7} kg /C. Find the mass of A^{+}.

→ charge from 1 proton

↳ 1.6×10^{-19} C

charge

given ⇒ $\dfrac{\text{Mass of } A^{+}}{\text{Charge of } A^{+}} = 1.97 \times 10^{-7} \dfrac{\text{kg}}{\cancel{C}} = \dfrac{?}{1.6 \times 10^{-19} \cancel{C}}$

$? = 1.97 \times 10^{-7} \times 1.6 \times 10^{-19}$ kg

$= 3.15 \times 10^{-26}$ kg

Find the ratio of e/m value for electron to proton.

$\downarrow$
charge

$$\frac{\left(\frac{\cancel{\text{charge}}}{\text{mass}}\right)_e}{\left(\frac{\cancel{\text{charge}}}{\text{mass}}\right)_p} = -\frac{(\text{mass})_p}{(\text{mass})_e}$$

$$\approx -1840$$

a) Calculate the number of electrons, protons and neutrons in $^{80}_{35}Br$.

$80 \rightarrow p+n \rightarrow 45\ n$

$35\ p$

$35 \rightarrow p$ $\quad 35\ e$

b) The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign with proper Symbol.

e $\quad$ p $\quad$ n

$$\left.{}^{32}_{16}X^{2-}\right\} \qquad {}^{32}_{16}S^{2-}$$

H.W

1. Vividh Bharti station of All India Radio, Delhi broadcasts on a frequency of 1368 kHz. What is the wavelength (λ) of the radiation ? $c = 3 \times 10^8$ m/s

a. 300 m b. 220 m c. 500m d. 800m

2. Calculate wave number of a radiation having wavelength 5800 Å in m^{-1}

a. 1.7×10^6 b. 21.7×10^8 c. 9×10^6 d. 5.8×10^8

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JEE 2025
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$\boxed{1999 \leftrightarrow 4999}$

range

Links in Description

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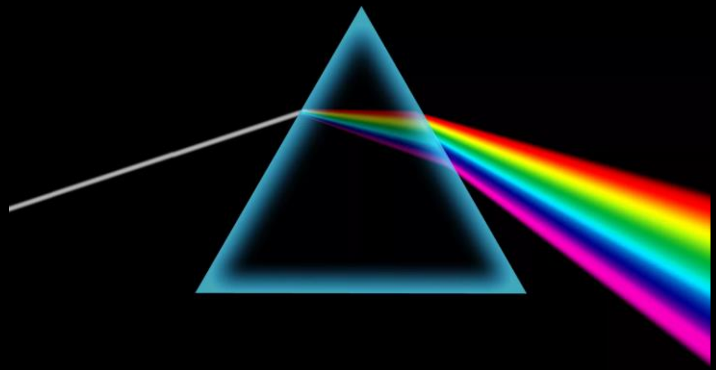
What's a spectrum?

$c = \nu \lambda$

$\nu = \frac{c}{\lambda}$

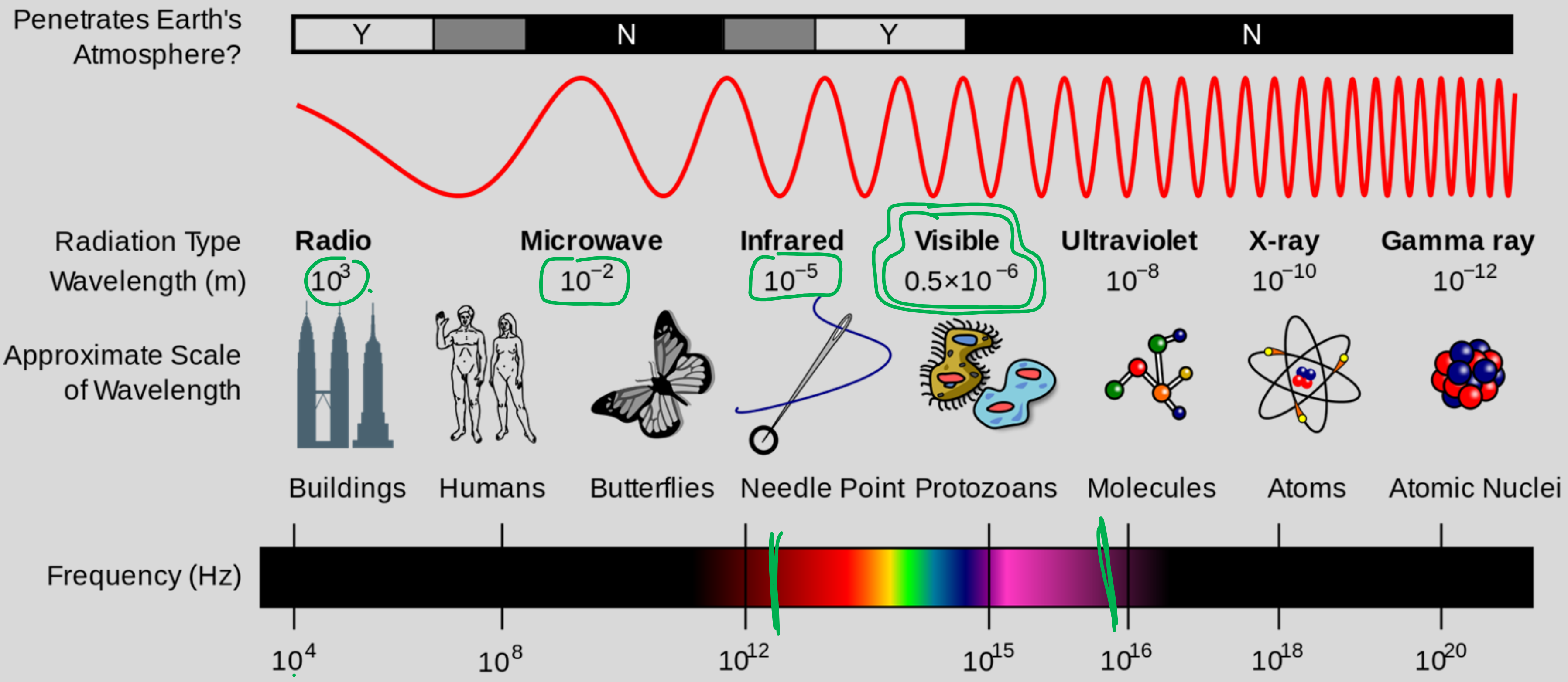


visible light ≈

$\underbrace{400}_{\text{violet}} \sim \underbrace{750}_{\text{red light}}$ nm

- A spectrum is an array of entities ordered in accordance with the magnitudes of a common physical property
- **VIBGYOR**

Electromagnetic Spectrum



Light is a wave ✓

Light is a $\boxed{\text{particle}}$!!!

$\Rightarrow$ Planck's Quantum Theory

Black Body Radiations

- a theoretical concept
- perfect energy absorber & emitter
- block of $\boxed{\text{IRON}}$

heat

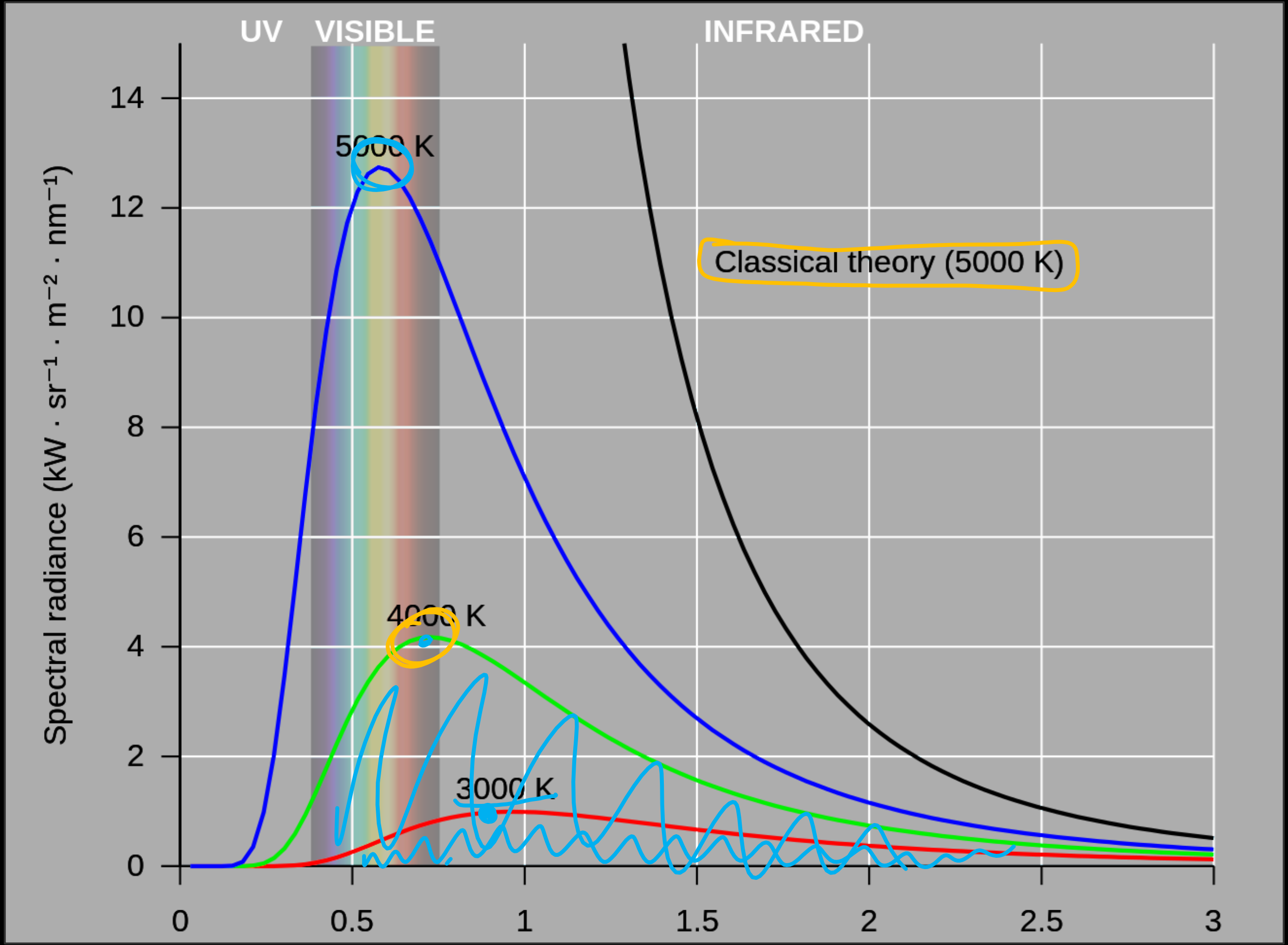
frequency

$\boxed{\text{color}}$ of light is Related to Energy

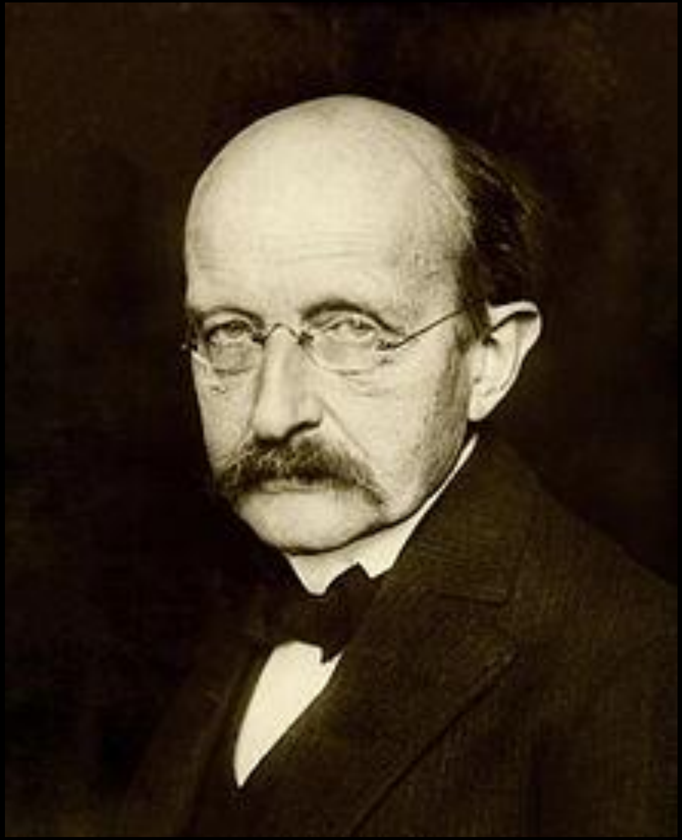
frequency of light is related to the energy of light.

low E low ν

high E high ν



UV
VISIBLE
INFRARED
5000 K
4000 K
3000 K
Classical theory (5000 K)
Spectral radiance (kW · sr⁻¹ · m⁻² · nm⁻¹)
0
2
4
6
8
10
12
14
0
0.5
1
1.5
2
2.5
3



light contains $\boxed{\text{small small PACKETS of Energy}}$

$E_{\text{of each packet}} \propto \nu_{\text{light}}$

$\rightarrow$ Quantization

What is Planck's Quantum Theory?

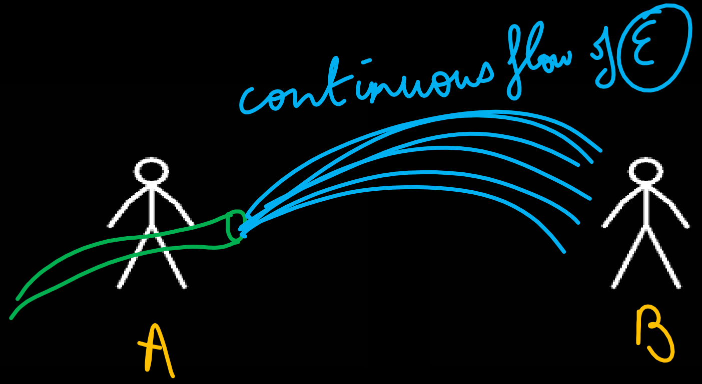
1. Energy is emitted in form of radiations from a source in a discontinuous manner that is in form of packets of energy.

1. And energy of each packet depends on the frequency of radiation. These packets are called "Quantum of energy".

Fight

Round 1

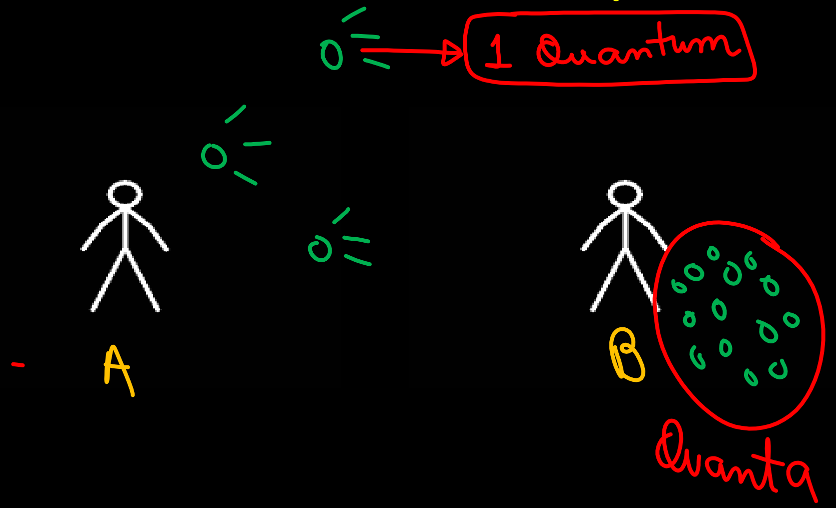
A attacks B



Round 2

B attacks A

small, but discontinuous packets of E.



- Energy of Quantum

$$E \propto \nu$$

- The constant of proportionality is called $\boxed{\text{Planck's constant}}$ represented by "h"

$$\underset{J}{E} = h\underset{\frac{1}{s}}{\nu} \qquad \boxed{E = h.\nu} \rightarrow \text{Energy of 1 Quantum}$$

- E is in J and the units of h are Js (Joule-seconds).

$$h = \underline{6.626} \times \underline{10}^{-34}\ \underline{Js}.$$ Remember

- If it contains "n" Quanta of the same frequency then the total energy will become:

$$\boxed{E = nh\nu} \qquad E = nh\nu$$

☆ for visible light

1 Quantum is also called a Photon

only 1 wavelength

$\rightarrow \nu = \frac{c}{\lambda}$

A 100 watt bulb emits $\boxed{\text{monochromatic}}$ light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb

a. 2×10^{20}

b. 2×10^{10}

c. 2×10^{30}

d. 2×10^{50}

1 watt = 1 Joule/second

in 1 sec, $\varepsilon = 100 J$

ε of 1 quantum

$\varepsilon = h\nu = \frac{hc}{\lambda}$

$\varepsilon = n \frac{hc}{\lambda}$

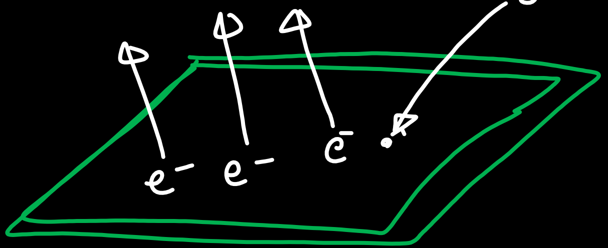
$n = \frac{\varepsilon \lambda}{hc}$

$= \frac{\cancel{100} J \times \cancel{400} \times 10^{-9} m}{\cancel{6.626} \times 10^{-34} J s \times \cancel{3} \times 10^{8} m}$

$= 200 \times 10 \times 10^{-9} \times 10^{34} \times 10^{-8}$

Photoelectric Effect

light (e-m-radiation)



✓ light's energy is being transferred to electrons

1e^- can absorb the energy of only 1 photon!!!

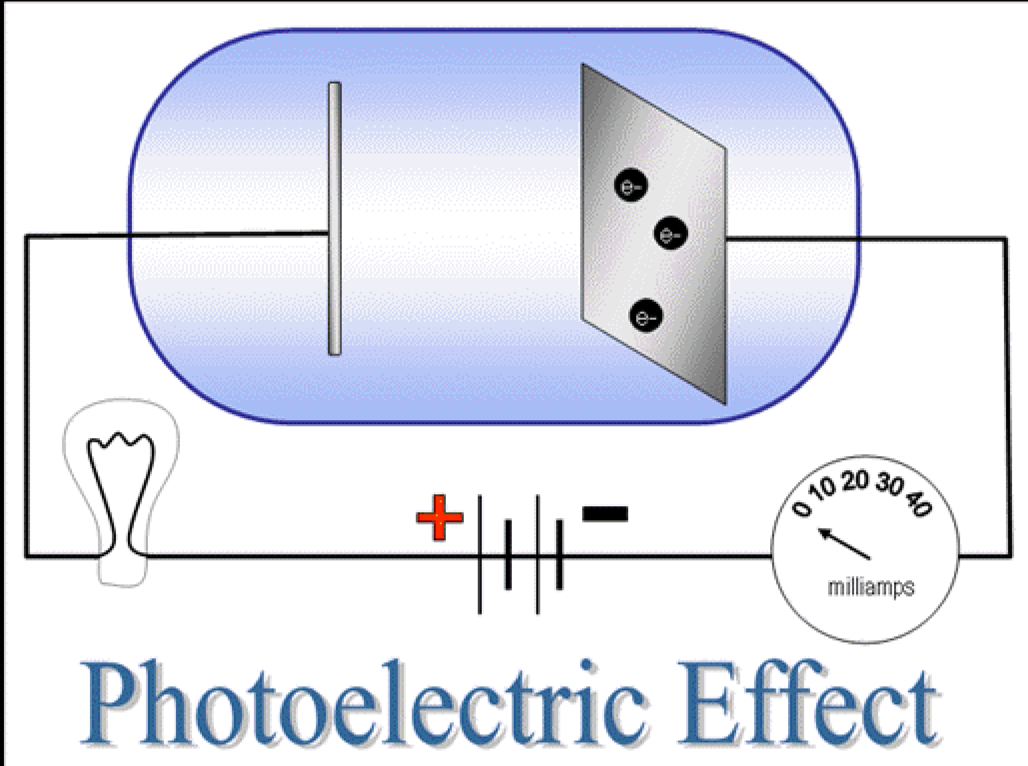
★ observation:-

low ν light $\rightarrow$ no e^- coming out

after a certain ν $\Rightarrow$ threshold frequency

high ν light $\rightarrow$ e^- leaves the metal

ϕ phi



say, for a metal.

min E required to remove an e^- = ϕ

incident photon's Energy =

$E = h\nu$

$h\nu - \phi$ = Kinetic Energy of e^-

$$\underbrace{h\nu}_{\text{incident energy}} - \underbrace{\phi}_{\boxed{\text{min E required}}} = \frac{1}{2} m_e v^2$$

min E required = work function

Photoelectric Effect

$$E - \phi = KE$$

$$h\nu - \phi = \frac{1}{2} m_e v^2$$

$$\phi = h\nu_0$$

threshold frequency

$$\boxed{h\nu - h\nu_0 = \frac{1}{2} m_e v^2 \qquad \nu > \nu_0}$$

The minimum energy that must be possessed by photons in order to produce the photoelectric effect with platinum metal is:

[Given: The threshold frequency of platinum is $1.3 \times 10^{15}\ s^{-1}$ and $h = 6.6 \times 10^{-34}$ J s.]

(A) 3.21×10^{-14} J

(B) 6.24×10^{-16} J

(C) 8.58×10^{-19} J

(D) 9.76×10^{-20} J

$$h\nu - \boxed{h\nu_0} = \frac{1}{2} m_e v^2$$

min E

$$\parallel$$

$$6.6 \times 10^{-34}\ J\ \cancel{s} \times 1.3 \times 10^{15}\ 1/\cancel{s}$$

A light source of wavelength λ illuminates a metal and ejects photo-electrons with $(K.E.)_{max}$ = 1 eV

Another light source of wavelength $\frac{\lambda}{3}$, ejects photo-electrons from same metal with $(K.E.)_{max}$ = 4 eV

Find the value of work function ?

Options:

(a) 1 eV

(b) 2 eV

$\boxed{\text{(c) 0.5 eV}}$ ✓

(d) None of these

Case - 1

$$\varepsilon = \frac{hc}{\lambda}$$

$$\frac{hc}{\lambda} - \phi = 1\,eV \Rightarrow (i)$$

Case - 2

$$\varepsilon = \frac{hc}{\lambda/3} = \frac{3hc}{\lambda}$$

$$\frac{3hc}{\lambda} - \phi = 4\,eV \Rightarrow (ii)$$

$$1\,eV = 1.6 \times 10^{-19}\,J$$

KE_{max} ?

3 (i) − (ii)

$$\cancel{\frac{3hc}{\lambda}} - 3\phi = 3\,eV$$

$$-\cancel{\frac{3hc}{\lambda}} + \phi = -4\,eV$$

$$-2\phi = -1\,eV$$

If photon of wavelength 150pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of 1.5×10^7 m/s. Calculate the energy with which it is bound to the nucleus.

Calc ϕ

$\lambda = 150 \times 10^{-12}$ m

$v = 1.5 \times 10^{7}$ m/s

$KE = \frac{1}{2} m_e v^2$

$\frac{hc}{\lambda} - \phi = KE$

$\phi = \frac{hc}{\lambda} - KE$

$$\frac{6.626 \times 10^{-34}\,Js \times 3 \times 10^{8}\,m/s}{150 \times 10^{-12}\,m} - \frac{1}{2} \times 9.1 \times 10^{-31}\,kg \times \left(1.5 \times 10^{7}\,\frac{m}{s}\right)^2$$

$\phi = 12.22 \times 10^{-16}$ J

If a certain metal was irradiated by using two different light radiations of frequency 'x' and '2x', the maximum kinetic energy of the ejected electrons are 'y' and '3y' respectively. The threshold frequency of the metal will be :

(a) x/3 (b) x/2

(c) 3x/2 (d) 2x/3

$① \rightarrow hx - h\nu_0 = y$

$② \rightarrow h2x - h\nu_0 = 3y$

$3① - ② \Rightarrow$

$3hx - 3h\nu_0 = 3y$

$- 2hx + h\nu_0 = -3y$

$1\cancel{h}x - 2\cancel{h}\nu_0 = 0$

4000 Å photon is used to break the iodine molecule, then the % of energy converted to the K.E. of iodine atoms if bond dissociation energy of I_2 molecule is 246.5 kJ/mol

(a) 8%

(b) 12%

(c) 17%

(d) 25%

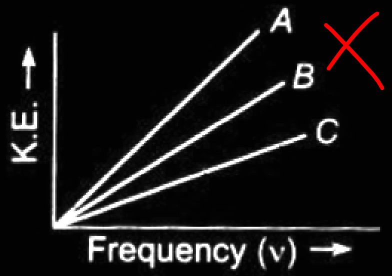
Hw

try

photon

I — I

Photoelectron emission is observed for three different metals A, B and C. The kinetic energy of the fastest photoelectrons versus frequency 'ν' is plotted for each metal. Which of the following graph shows the phenomenon correctly ?

(a)


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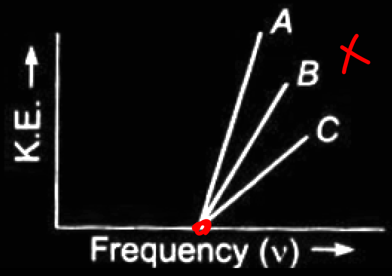
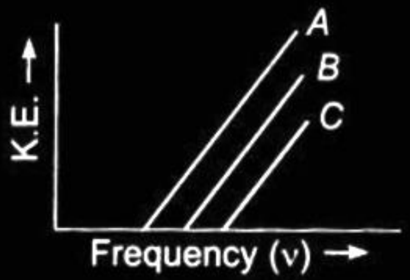


(c)


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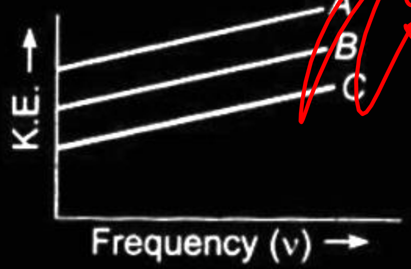
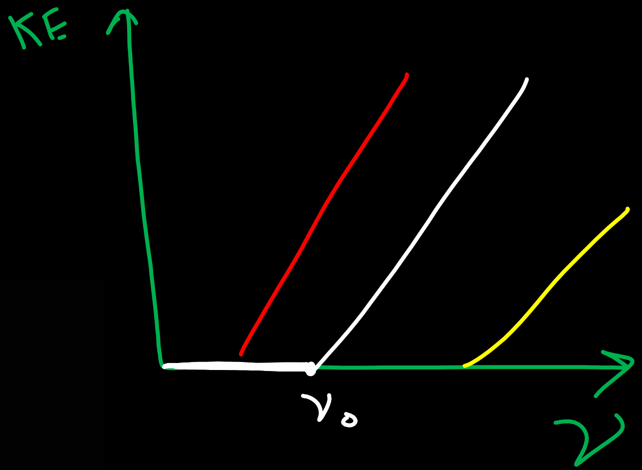
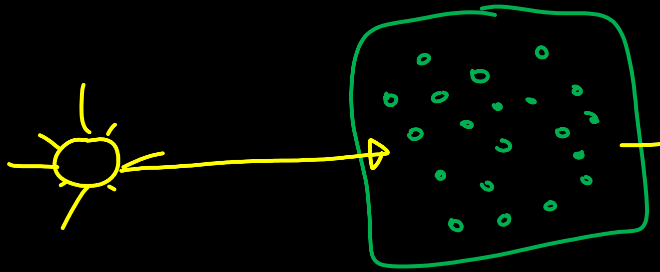




Atomic spectrum

Fingerprint of an atom.



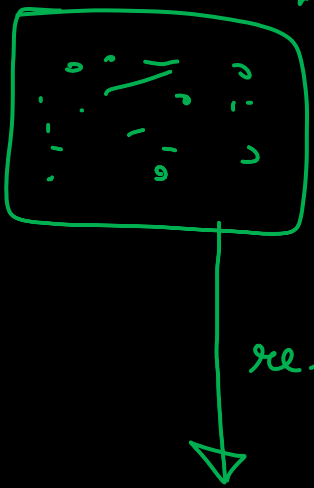
atoms absorb

some energy

(some photons)

↓

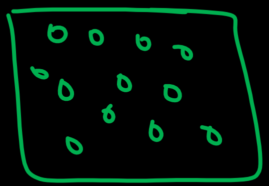
some frequencies



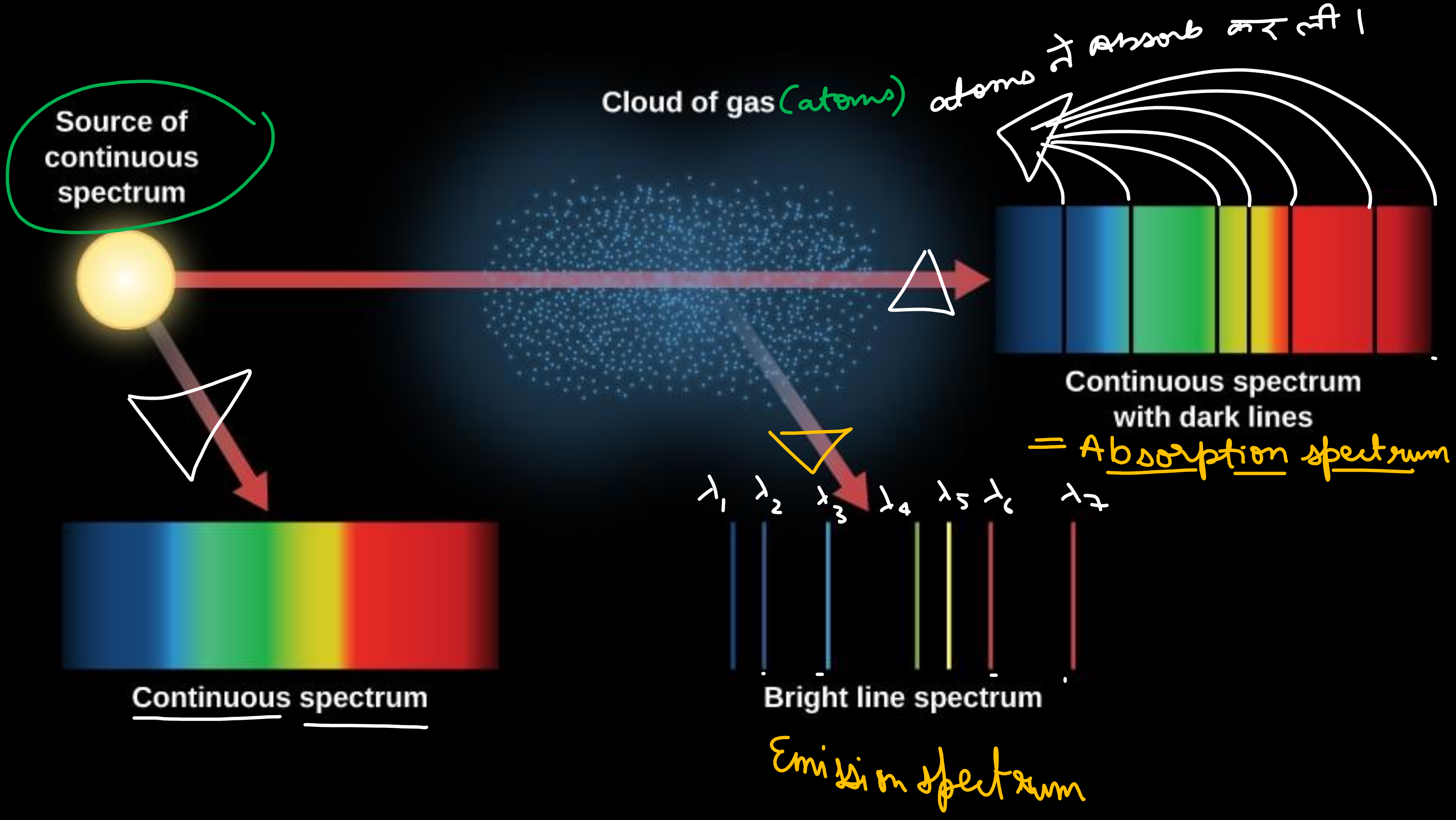
higher energy

Excited

release photons.



★★ only certain frequencies of light were were absorbed


Source of continuous spectrum
Cloud of gas (atoms)
atoms ने absorb कर ली।
Continuous spectrum with dark lines
= Absorption spectrum
λ1 λ2 λ3 λ4 λ5 λ6 λ7
Continuous spectrum
Bright line spectrum
Emission spectrum

Hydrogen Spectrum

Balmer in 1885 - observed the lines of H Spectrum in visible range

Rydberg generalised the Formula:

$$\frac{hc}{\lambda} = E$$

$$\boxed{\frac{1}{\lambda}} = 109677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \boxed{cm^{-1}}$$

$\boxed{n_1, n_2}$ → natural numbers ✓

$\boxed{n_1 < n_2}$

$$\frac{1}{\lambda} = 109677\left(\frac{1}{1} - \frac{1}{n_2^2}\right) cm^{-1}$$

Say.

$n_1 = 1$	$n_2 = 2, 3, 4, 5, 6$
$n_1 = 2$	$n_2 = 3, 4, 5, 6 \ldots \ldots$
$n_1 = 3$	$n_2 = 4, 5, 6, 7, 8 \ldots$
$n_1 = 4$	$n_2 = 5, 6, 7, 8, 9 \ldots$

α alpha

β beta

γ gamma

δ delta

$n_1 = 1$

	α	β	γ	δ		
$n_2 =$	2	3	4	5	6	.

first line of Lyman series

2nd line

3rd line

When $n_1 = 1$ lines of Lyman Series

When $n_1 = 2$ lines of Balmer series

When $n_1 = 3$ lines of Paschenn series.

3rd line

$n_1 = 2$

	α	β	γ	δ		
$n_2 =$	3,	4,	5,	6,	7	...

1st line of Balmer

2nd line.

Series	n_1	n_2	Spectral Region
Lyman	1	2, 3...	Ultraviolet
Balmer	2	3, 4...	Visible
Paschen	3	4, 5...	near infrared
Brackett	4	5, 6...	mid infrared
Pfund	5	6, 7...	far infrared

Estimate λ of 2nd line of Balmer series for H atom

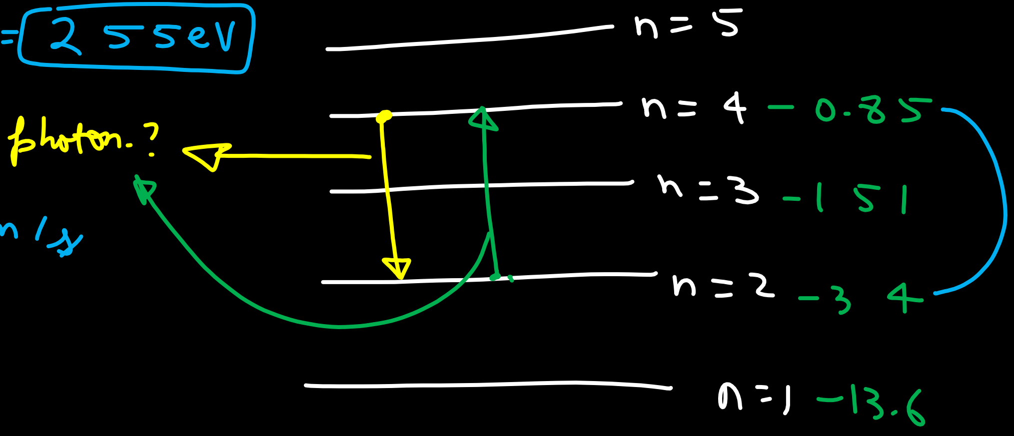
$n_1 = 2$

$n_2 = 4$

$E = \frac{hc}{\lambda}$

$\lambda = \frac{hc}{E}$

$\Delta E = 3.4 - 0.85 = \boxed{2.55\,eV}$



$$\lambda = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}\ m/s}{2.55 \times 1.6 \times 10^{-19}\ J}$$

$= 487\,nm$

Estimate the λ of second line (also called β -line) in Balmer series of H-atom.

$\boxed{n_1 = 2}$

$\boxed{n_2} = 3,\ \boxed{4},\ 5,\ 6.$

1st line (3), 2nd line (4), 3rd line (5)

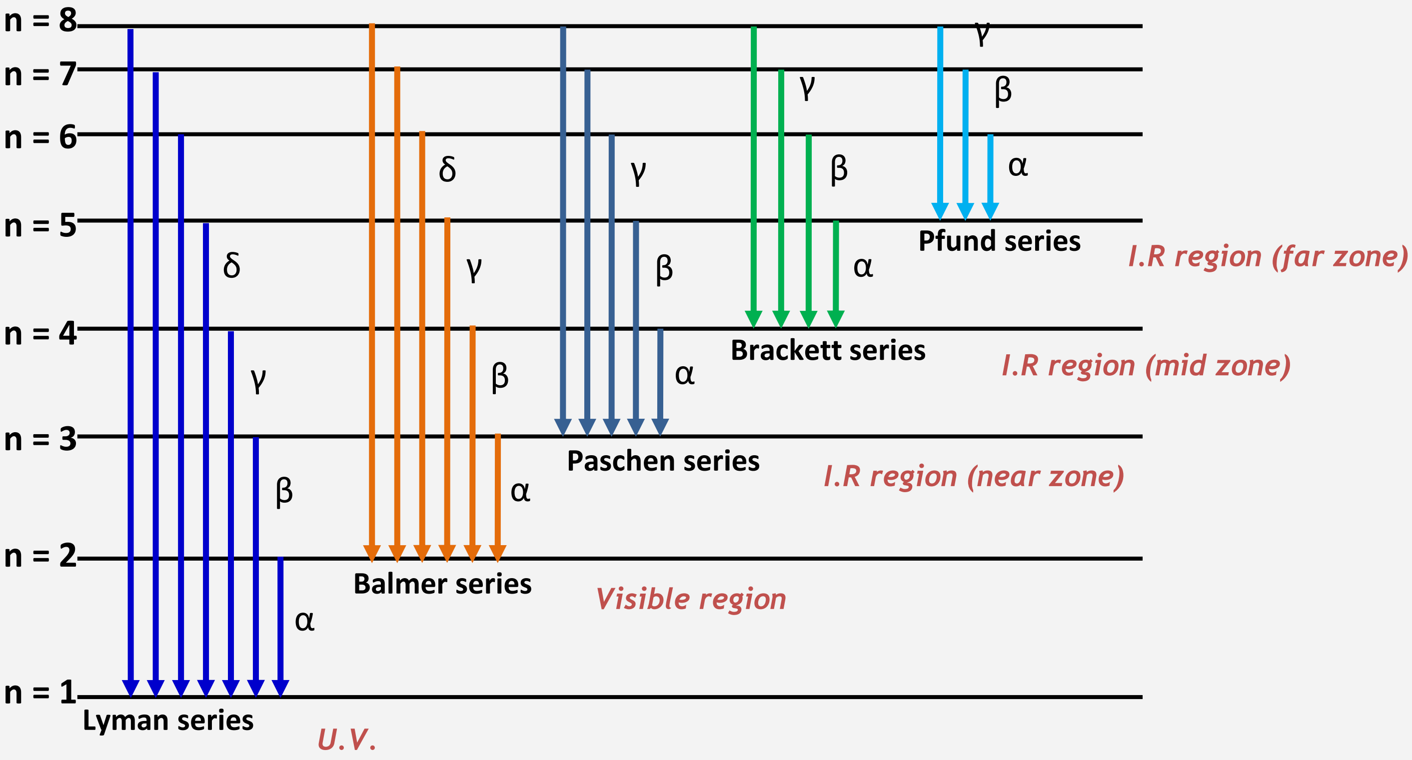
$$\frac{1}{\lambda} = 109677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) cm^{-1}$$

$$= 109677\left(\frac{1}{4} - \frac{1}{16}\right) cm^{-1}$$

$$\frac{1}{\lambda} = 109677\left(\frac{3}{16}\right) cm^{-1}$$

$$\lambda = \frac{16}{3 \times 109677}\ cm \Rightarrow \boxed{487\ nm}$$

Hydrogen



n = 8
n = 7
n = 6
n = 5
n = 4
n = 3
n = 2
n = 1
Lyman series
U.V.
Balmer series
Visible region
Paschen series
I.R region (near zone)
Brackett series
I.R region (mid zone)
Pfund series
I.R region (far zone)
α
β
γ
δ

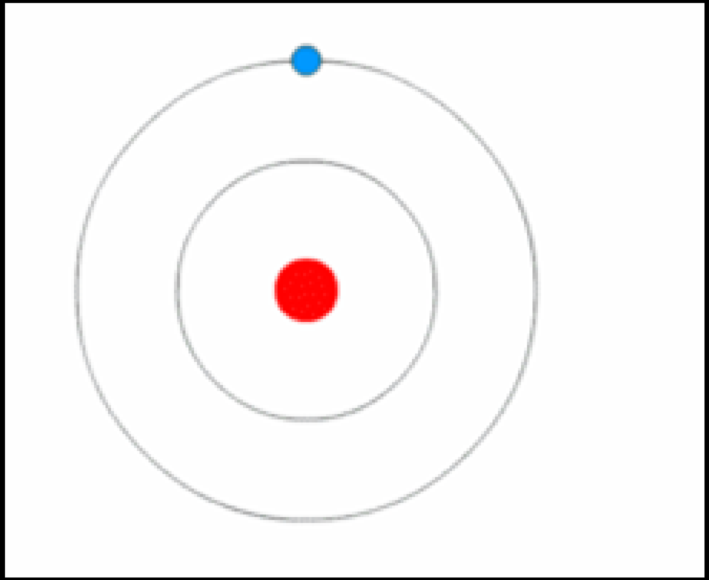
Understand this first

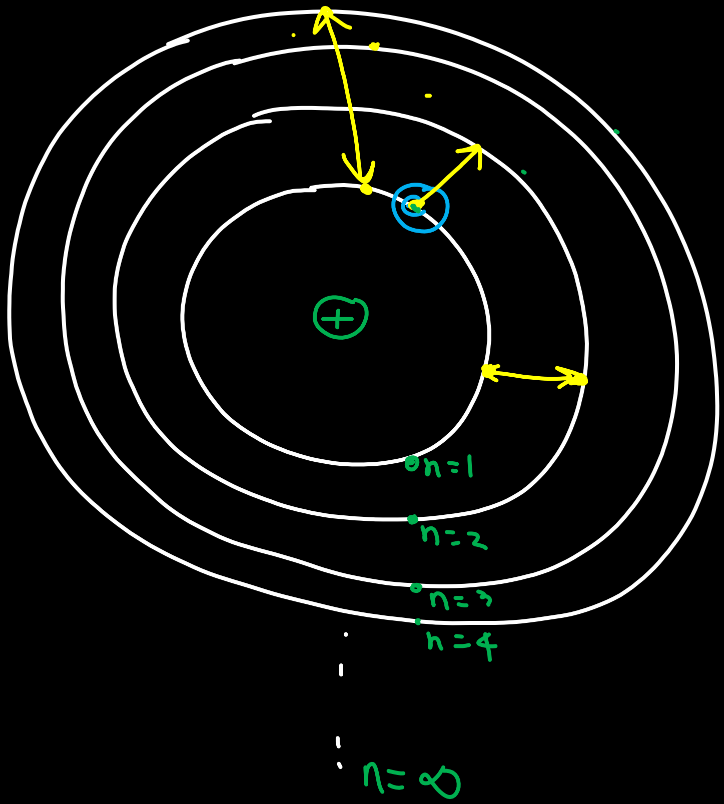
1. Bohr Model stems from the H Spectra
2. It is a THEORY given by Niels Bohr
3. Only for Single Electron Systems

$\rightarrow$ H, He^{+}, Li^{2+}, Be^{3+}, B^{4+}, C^{5+}

1. Electron in the H atom can revolve around the nucleus in **ORBITS** - places with fixed radius and energy
2. Energy of the ORBIT doesn’t change
3. The electron can gain energy and JUMP! Between orbitals

from a photon





n=1
n=2
n=3
n=4
n=∞

Bohr's ideas –

→ $f = \dfrac{mv^2}{r}$ (Centripetal force)

→ Electrostatic force $\left(\dfrac{K q_1 q_2}{r^2}\right)$

$$\boxed{\frac{mv^2}{r} = \frac{KZe^2}{r^2}}$$

$p \rightarrow +e$

$e \rightarrow -e$

★★ Bohr's postulate

angular momentum is quantized

$$mvr = n\frac{h}{2\pi}$$

n → no of orbit

n → natural number

Total

Energy of e^- in an orbit

$$E = -13.6\ Z^2/n^2\ \ \textit{ev/atom}$$

$$\boxed{E = -13.6 \times \frac{Z^2}{n^2}}\ \ eV/atom$$

Z → atomic number

n → number of the orbit

$$E = -2.18\times10^{-18}\ Z^2/n^2\ \ \textit{J/atom}$$

Radius

$$r = \boxed{A_o}\, n^2/Z$$

$$\boxed{r = 52.9 \cdot \frac{n^2}{Z}\ pm}$$

$$A_o = 52.9\ \textit{pm}$$

$$Mvr = nh/2\pi$$

Velocity

$$v = 2.18 \times 10^6\ Z/n\ \ \textit{m/s}$$

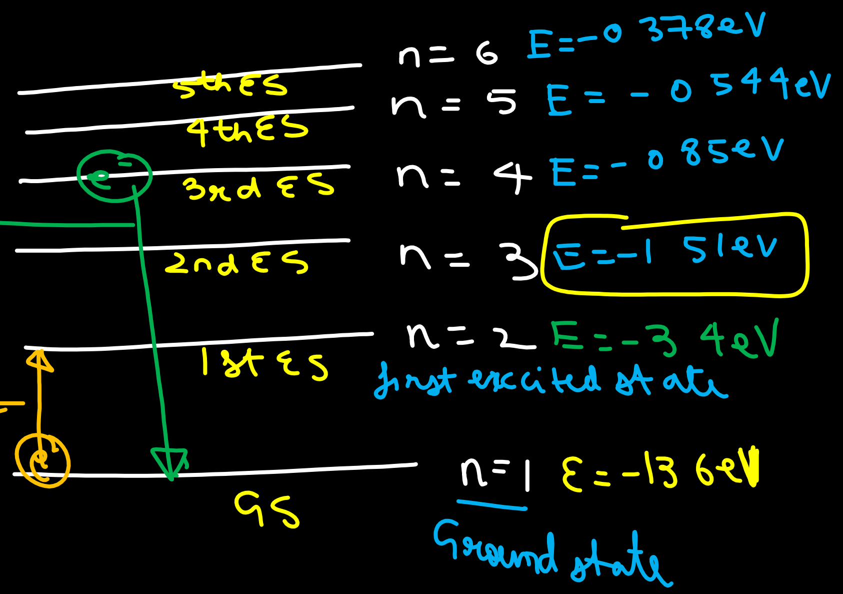
$$\boxed{v = 2.18\times10^6\ \frac{Z}{n}\ m/s}$$

$$E = -13.6\frac{Z^2}{n^2}\,eV$$

$Z^2 \to 1 \leftarrow$ for H atom $\Rightarrow$

E of photon released $= 12.75\,eV$

$\Delta E = +10.2\,eV$



JEE Level Practice

In a hydrogen atom, if the energy of an electron in the ground state is – 13.6 eV, then that in the 2nd excited state is?

(a) – 1.51 eV

(b) – 3.4 eV

(c) – 6.04 eV

(d) – 13.6 eV

JEE Level Practice

The energy of an electron in the first Bohr orbit of H-atom is –13.6 eV. The possible energy value(s) of the excited state(s) for electrons in Bohr orbits of hydrogen is (are)

(a) –3.4 eV

(b) – 4.2 eV

(c) – 6.8 eV

(d) + 6.8 eV

JEE Level Practice

Which hydrogen like species will have same radius as that of Bohr orbit of hydrogen atom ?

$r = 52.9 \frac{n^2}{Z}$ pm

$H_{n=1}$ $r = 52.9$ pm $n=1$ $Z=1$

(a) n = 2, Li2+ $\rightarrow \frac{n^2}{Z} = \frac{4}{3}$

(b) n = 2, Be3+ $\rightarrow \frac{n^2}{Z} = \frac{4}{4} = 1$

(c) n = 2, He+ $\rightarrow \frac{n^2}{Z} = \frac{4}{2}$

(d) n = 3, Li2+

$\rightarrow \frac{n^2}{Z} = \frac{9}{3}$

Why is E (-ve) convention

$$\boxed{E = -13.6\,\frac{Z^2}{n^2}\,eV}$$

total Energy

$$PE = 2\,\text{Total Energy}$$

$$= -2 \times 13.6 \times \frac{Z^2}{n^2}\ eV/atom$$

$$KE = TotE - PE$$

$$= -TotE = +13.6\,\frac{Z^2}{n^2}\ eV/atom$$

dist = ∞

$\oplus$ nu

no attraction

$\boxed{PE = 0}$ ✓

e^-

$\oplus \leftarrow r \rightarrow e^-$

$\boxed{PE \text{ is} < 0}$

For He^+, a transition takes place from the orbit of radius 105.8 pm to the orbit of radius 26.45 pm. The wavelength (in nm) of the emitted photon during the transition is ___.

JEE adv PYQ

[Use:

Bohr radius, a = 52.9 pm

Rydberg constant, $R_H = \boxed{2.2 \times 10^{-18}\ J}$

Planck's constant, $h = 6.6 \times 10^{-34}$ J s

Speed of light, $c = 3 \times 10^8\ m\ s^{-1}$]

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

$$= 4 \times 2.2 \times 10^{-18}\left(\frac{1}{1} - \frac{1}{4}\right) J = \frac{hc}{\lambda}$$

$$\lambda \times \cancel{8.8} \times 10^{-18}\left(\frac{\cancel{3}}{\cancel{4}}\right) = \cancel{6.6} \times 10^{-34} \times 3 \times 10^8$$

$$\boxed{\lambda = 30\,nm}$$

n=3

n=2 $r = 52.9 \times \frac{4}{2}$ = 105.8 pm

-13.6 × ... eV

n=1 $r = 52.9 \times \frac{1}{2}$ = 26.45 pm

$-13.6 \times \frac{Z^2}{n^2}$

$= -13.6 \times \frac{4}{1} = -54.4\ eV$

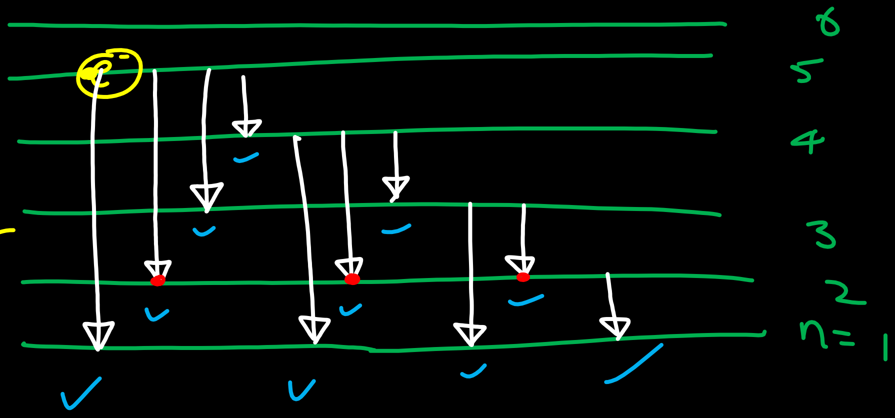
$\Delta E = \boxed{40.8\ eV}$

use $\frac{hc}{\lambda} = E$

A certain transition in H-spectrum from an excited state to ground state in one or more steps gives rise to a total of ten lines. How many of these belong to visible spectrum ?

(a) 3 (b) 4

(c) 5 (d) 6



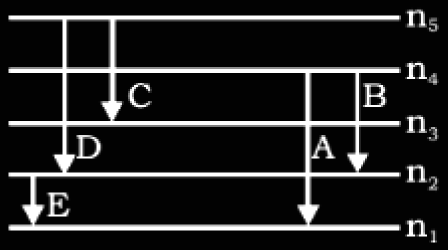
no of spectral lines:-

$$\frac{(\Delta n)(\Delta n+1)}{2}$$

when a group of H atoms (all e^- in 5^{th} orbit) come back to ground state, 10 spectral lines are obtained

HW

For a hypothetical H like atom which follows Bohr's model, some spectral lines were observed as shown. If it is known that line 'E' belongs to the visible region, then the lines possibly belonging to ultra violet region will be (n_1 is necessarily ground state)



[Assume for this atom, no spectral series shows overlaps with other series in the emmission spectrum]

(a) B and D
(b) D only
(c) C only
(d) A only

Cuvette

INR ~ 15K- 20K

NMR spectrum

दुनिया का हर Particle
एक wave भी है

Any moving particle has a wavelength associated with it

$$E = mc^{\cancel{2}} = \frac{h\cancel{c}}{\lambda}$$

$$\lambda = \frac{h}{mc} \Rightarrow \lambda = \frac{h}{mv}$$

momentum

light → wave
light → particle

$\Big\}$ dual nature of light

wave- particle duality

De-Broglie's Wave-Particle Duality

$$\lambda_{db} = \frac{h}{p} = \frac{h}{mv}$$

b

60 kg , 5 m/s

$$\lambda = \frac{\cancel{6.6}^{\,11} \times 10^{-35} \text{ Js}}{\cancel{60}_{10} \text{ kg} \times \cancel{5}\text{ m/s}}$$

$$\approx \ldots \times 10^{-34.36} \text{ m}$$

insignificant for macroscopic particles

Heisenberg's Uncertainty ?

"It is impossible to determine the position and momentum of a particle simultaneously and accurately."

$$\Delta x \times \Delta p \geq \boxed{h/4\pi}$$

uncertainty in position

uncertainty in momentum

$\sim 10^{-34}$

Find the De Broglie wavelength associated with the electron in Bohr's first orbit

$$v = \boxed{2.18 \times 10^6}\,\frac{Z}{n} \quad m/s$$

$$\lambda = \frac{h}{mv}$$

$$\lambda = \frac{6.626 \times 10^{-34}}{9.1 \times 10^{-31} \times 2.18 \times 10^{6}}\ m$$

$$\approx \boxed{334\ pm}$$

↳ significant for e^- !!

electron's wave nature is important !!!

charge = $\boxed{e}$ = -1.6×10^{-19} C

An electron is accelerated through a potential difference of V_o volts.

Find the De Broglie wavelength associated with the electron.

$$\frac{1}{2} m_e v^2 = e V_0$$

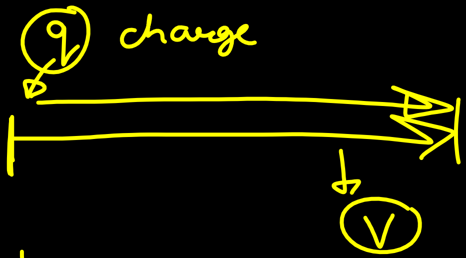
$$\lambda = \frac{h}{mv}$$

$$m_e \left(m_e v^2 = 2 e V_0 \right)$$

$$\sqrt{m_e^2 v^2} = \sqrt{2 m_e e V_0} = m_e v$$

$$\lambda = \frac{h}{\sqrt{2 \cdot m_e e V_0}}$$

fact -



Kinetic Energy gained by this charge = qV

$$\frac{1}{2} m v^2 = qV$$

Calculate the uncertainty in position assuming uncertainty in momentum within 0.1% for:

(a) a tennis ball weighing 0.2 kg and moving with a velocity of 10 m/s.
(b) a electron moving in an atom with a velocity of $\boxed{2 \times 10^6}$ m/s.

Tennis ball -

$$p = mv = 0.2 \times 10 = 2 \, kg \frac{m}{s}$$

$$\Delta p = \boxed{0.1\% \text{ of } p} = \frac{0.1}{100} \times 2 = \boxed{2 \times 10^{-3}}$$

$$\Delta x = \frac{h}{4\pi \Delta p}$$

$$= \frac{6.626 \times 10^{-34}}{4 \times 3.14 \times 2 \times 10^{-3}} = 2.6 \times 10^{-32} \, m$$

insignificant

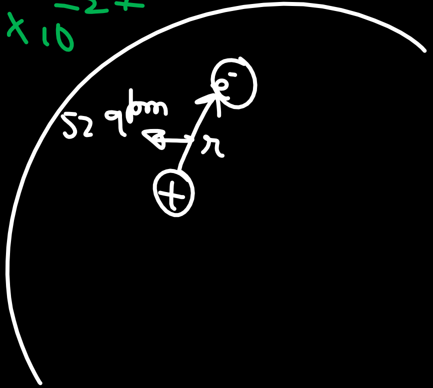
$$\Delta mv = \Delta p = 9.1 \times 10^{-31} \times 2 \times 10^{6} \times \frac{0.1}{100}$$

$$\Delta p = 1.8 \times 10^{-27}$$

$$\Delta x = \frac{6.626 \times 10^{-34}}{4 \times 3.14 \times 1.8 \times 10^{-27}}$$

$$= 2.9 \times 10^{-8} \, m$$

$$\approx \boxed{29 \, nm}$$



$A \sin Bx$

define the character of the wave

wave-function

$$\frac{d}{dx} e^{x} = \boxed{1} \cdot e^{x}$$

operator

eigenvalue

ψ ⇒ we can use multiple operators with e^- wave function

$\hat{H}$ ⇒ hamiltonian operator

$$\hat{H}\psi = E \cdot \psi$$

energy

Erwin Schrödinger

Gave the notorious Wave Equation. Let us study that now

→ e^- की wave nature

→ Heisenberg's unc principle

e^- is a wave → study

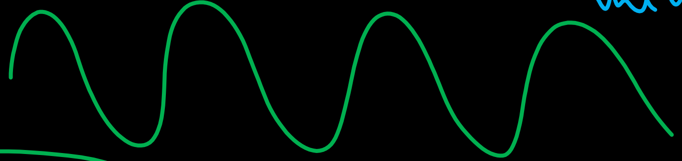
e^- का जो एक wavefunction होगा।

wave → wave function

↳ has all details of the wave.

* electronic wave function -

ψ ψ ψ (psi)



sine wave

Time-independent Schrödinger equation (*general*)

$$\hat{\mathrm{H}}\,|\Psi\rangle = E|\Psi\rangle$$

Time-independent Schrödinger equation (*single nonrelativistic particle*)

$$\left[\frac{-\hbar^2}{2\mu}\nabla^2 + V(\mathbf{r})\right]\Psi(\mathbf{r}) = E\Psi(\mathbf{r})$$

We will now think about the atom in terms of Wavefunctions.

These Wavefunctions are Characterized by $\boxed{\text{Quantum Numbers!}}$ ✓✓ ⇒ Give us the address of an e^-

n l m s

A certain set of Quantum Numbers defines an Atomic Orbital \m/

→ is the house of an electron

→ a volume region in space where there is maximum (95%) probability of finding an electron

Writing down the orbitals

$n=1$

$l=0$

$\hookrightarrow m=0$

1s orbital

can fill → $2e^-$

$n=2$

$l=0$

$m=0$

2s orbital

$2e^-$

$l=1$

$m: -1, 0, 1$

(2p orbitals)3

$6e^-$

$n=3$

$l=0$

m 0

3s orbital

$2e^-$

1

-1

0

1

3p orbitals

(3)

$6e^-$

2

-2

-1

0

1

2

3d orbitals

(5)

$10e^-$

$n \Rightarrow$ principal quantum number

city

shell

1, 2, 3, 4, 5 ∞

area

$l \Rightarrow$ azimuthal quantum number

subshells

0 $n-1$

$m \Rightarrow$ magnetic quantum number

house number

orientation of orbital

$-l \;. \;\; . +l$

$s \Rightarrow$ spin quantum number

room number

spin

$-\frac{1}{2}, +\frac{1}{2}$

$n = 5 \Longrightarrow$ we have $\boxed{25 \text{ orbitals}}$

→ can house a max of $50\ e^-$ in it

	s	p	d	f	g
$l \Rightarrow$	0	1	2	3	4
m	0	−1	−2	−3	−4
		0	−1	−2	−3
		1	0	−1	−2
			1	0	−1
			2	1	0
				2	1
				3	2
					3
					4

The correct designation of an electron with n = 4, l = 3, m = 2, and s = 1/2 is :

(a) 3d (b) 4f

(c) 5p (d) 6s

0 1 2 3
s p d f

4f

Which of the following sets of quantum number is correct for an electron in $\boxed{4f}$ orbital ?

(a) ~~n = 3, ℓ = 2, m = –2, s = +1/2~~

(b) ~~n = 4, ℓ = 4, m = –4, s = –1/2~~

(c) n = 4, ℓ = 3, m = +1, s = +1/2 ✓

(d) ~~n = 4, ℓ = 3, m = +4, s = +1/2~~

n = 4

l = 3

↳ m ⇒ -3, -2, -1, 0, 1, 2, 3

s ⇒ $+\frac{1}{2}, -\frac{1}{2}$

For each value of ℓ, the number for m_s values are

(a) 2ℓ (b) $n\ell$

(c) $2\ell + 1$ (d) $n - \ell$

$l = 2$

$-2 \quad -1 \quad 0 \quad 1 \quad 2$

5 orbitals

A subshell with $n = 6$, $\ell = 2$ can accommodate a maximum of

(a) 10 electrons (b) 12 electrons

(c) 36 electrons (d) 72 electrons

0 1 2 3

s p d

Let us Understand Electronic Configuration

Aufbau Rule ✓

Pauli's Exclusion Principle ✓

Hund's Rule of Maximum Multiplicity ✓

Aufbau Rule

In the ground state of atoms, the orbitals are filled in order of their increasing energies. (n+l) rule

for atomic orbitals, their energy $\propto$ $n+l$ value

	n	l	$n+l$
1s	1	0	1
2s	2	0	2
2p	2	1	3
3s	3	0	3
3p	3	1	4
3d	3	2	5
4s	4	0	4
4p	4	1	5

★ if $n+l$ value is same, fill lower n first

fill 4s before 3d orbitals

Pauli's Exclusion Principle

No two electrons in an atom can have the same set of four quantum numbers

→ only max 2 e^- can stay in an orbital

Hund's rule of maximum multiplicity

Pairing of electrons belonging to the same subshell (s,p,d,f) does not take place until each orbital belonging to that subshell has got one electron each - ie. it is singly occupied

$O \rightarrow 8e^- \rightarrow 1s^2\, 2s^2\, 2p^4$

$2p^4$: $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$

$N \Rightarrow 7e^-$

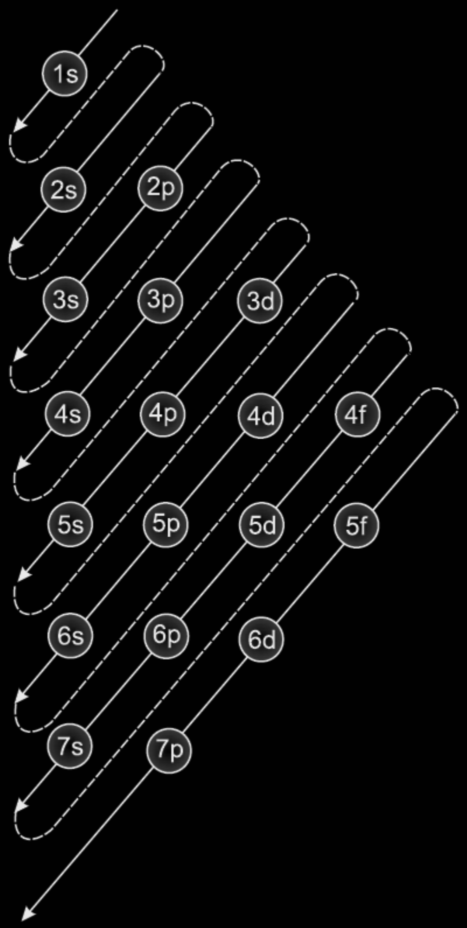
$1s^2$: $\boxed{\uparrow\downarrow}$ $2s^2$: $\boxed{\uparrow\downarrow}$

~~$2p$ opt A~~: $\cancel{\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{}}$

opt B ✓: $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

$F \rightarrow 9e^- \rightarrow 1s^2\, 2s^2\, 2p^5$

$2p^5$: $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$


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

Let us write some configurations!

$Cr \rightarrow 24$

$1s^2\ 2s^2\ 2p^6\ 3s^2$

$3p^6\ 4s^1\ 3d^5$

$\overset{29}{Cu}$

$1s^2\ 2s^2\ 2p^6\ 3s^2$

$3p^6\ 4s^1\ 3d^{10}$

$C \rightarrow 6e^-\quad 1s^2\ 2s^2\ 2p^2$

$O\quad 8e^-\quad 1s^2\ 2s^2\ 2p^4$

$Ne\quad 10e^-\quad 1s^2\ 2s^2\ 2p^6$

$Mg\quad 12e^-\quad 1s^2\ 2s^2\ 2p^6\ 3s^2$

$Ca\quad 20e^-\quad 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \cancel{3d}\ 4s^2$

$Ti\quad 22e^-\quad 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

$Fe\quad 26e^-\quad 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

In manganese atom, Mn (Z = 25), the total number of orbitals populated by one or more electrons (in ground state) is

(a) 15 (b) 14

(c) 12 (d) 10

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

1 1 3 1 3 1 5

The correct set of quantum numbers for the unpaired electron of chlorine atom is

try

	n	ℓ	m
(a)	2	1	0
(b)	2	1	1
(c)	3	1	1
(d)	3	0	0

The quantum number of four electrons are given below : JEE 2019

I. $n = 4, l = 2, ml = -2, ms = -1/2$ $n+l = 6$

II. $n = 3, l = 2, ml = 1, ms = +1/2$ $n+l = 5$

III. $n = 4, l = 1, ml = 0, ms = +1/2$ $n+l = 5$

IV. $n = 3, l = 1, ml = 1, ms = -1/2$ $n+l = 4$

The correct order of their increasing energies will be -

(1) ~~I < III < II < IV~~

(2) IV < II < III < I

(3) ~~I < II < III < IV~~

(4) ~~IV < III < II < I~~

3d orbitals

$n = 3, \ l = 2$

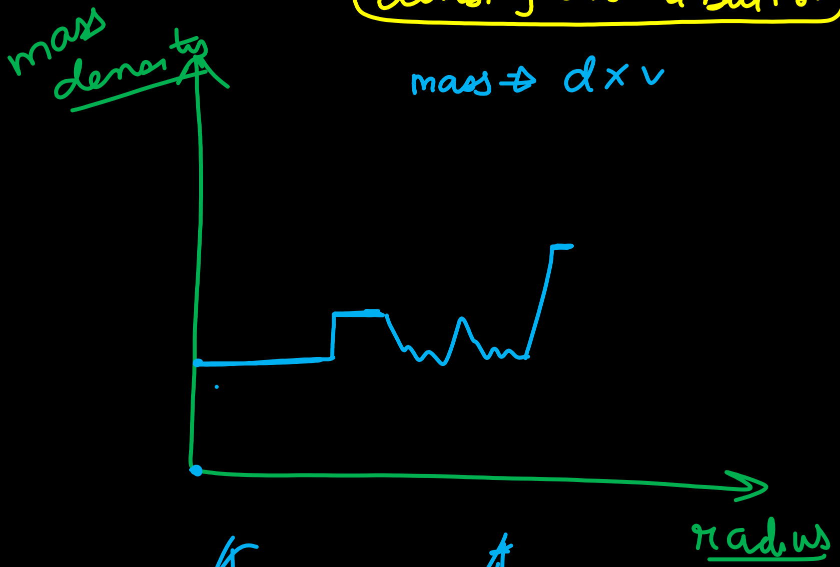
$m \rightarrow$ $\boxed{\begin{array}{c} -2 \\ -1 \\ 0 \\ 1 \\ 2 \end{array}}$

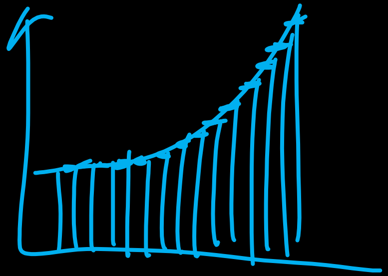
Same energy

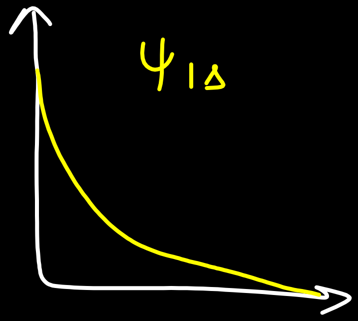
degenerate orbitals

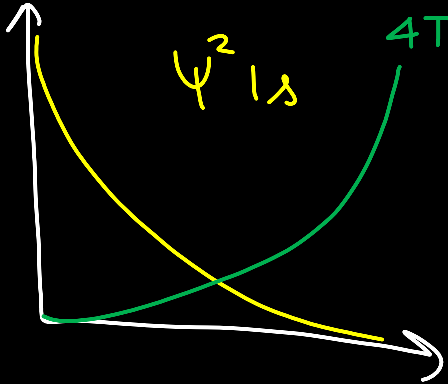
$\boxed{\text{density distribution}}$ function w r t radius

mass ⇒ $d \times v$





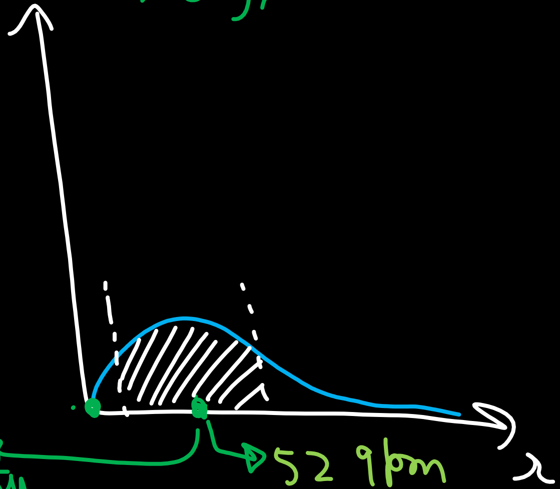




$\Longrightarrow$

e^- in a $\boxed{1s}$ orbital

↳ say, this is H atom

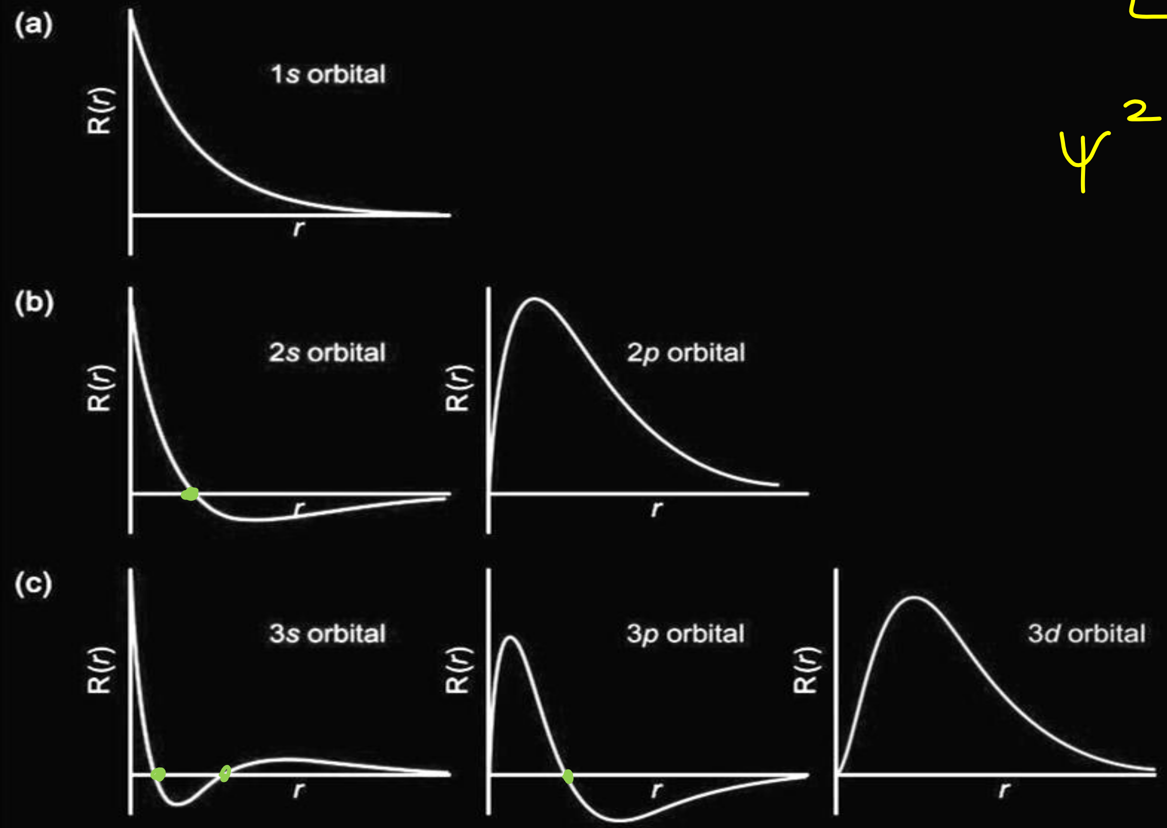


radius of max probability (RoMP) → 52.9 pm

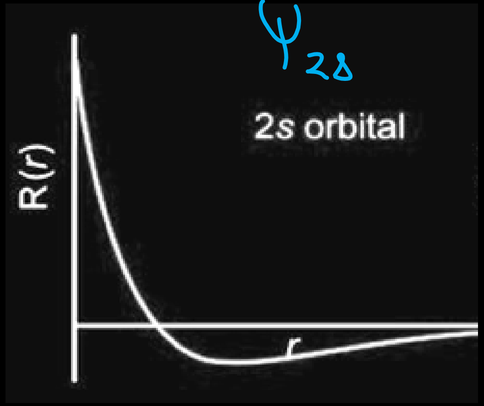
Plot of Radial Wave Function = f(r) = $\boxed{\psi}(r)$

$\rightarrow$ useless

ψ^2 } Probability density distribution (P/V)

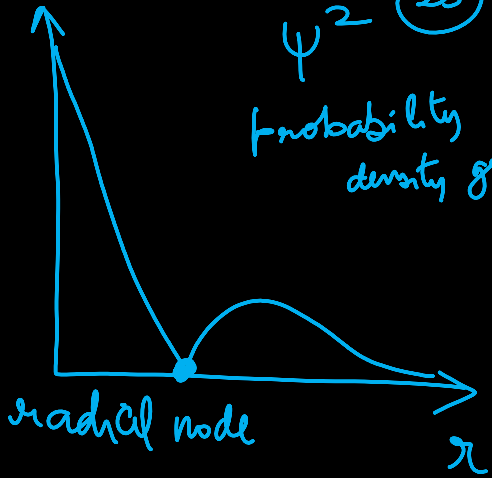


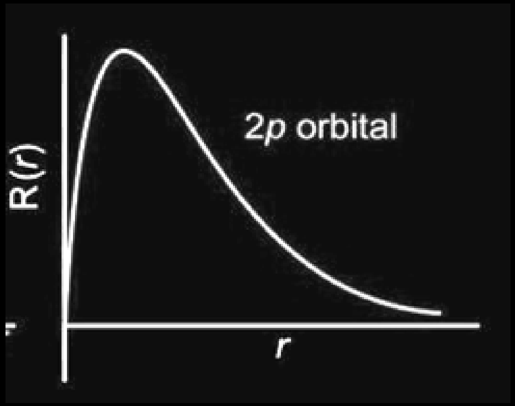
Ψ_{2s}

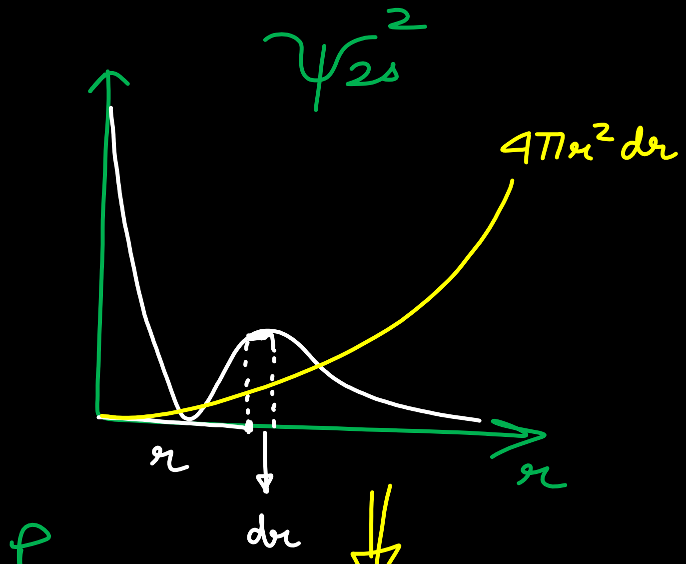


ψ^2 2s

probability density graph







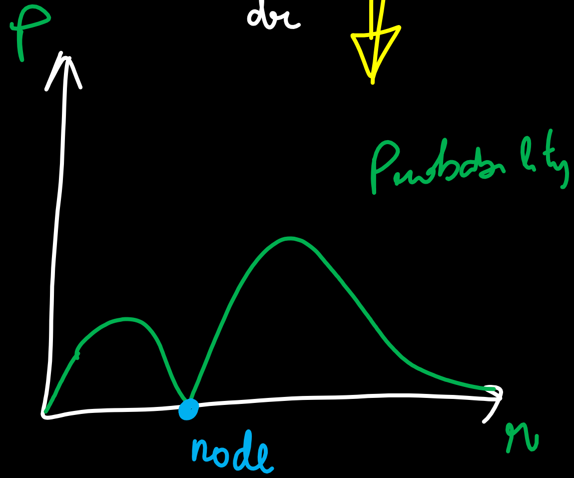
at $r \Rightarrow$ $\psi^2 \times$ vol of small shell of thickness dr

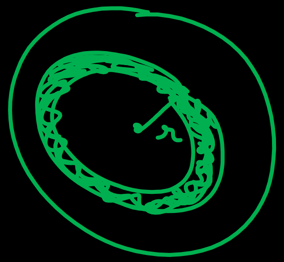
Prob

vol $= 4\pi r^2 \times h$

$\psi^2 \times \underline{4\pi r^2 \, dr}$?

Probability vs radius





JEE 2020

ψ^2

The correct statement about probability density (except at infinite distance from nucleus) is:

(1) ~~It can be negative for 2p orbital~~
(2) It can be zero for 3p orbital →
(3) ~~It can be zero for 1s orbital~~ →
(4) ~~It can never be zero for 2s orbital~~ →

3p →

rad	ang
1	1

The electrons are more likely to be found :

(1) in the region a and c

(2) only in the region c

(3) in the region a and b

(4) only in the region a

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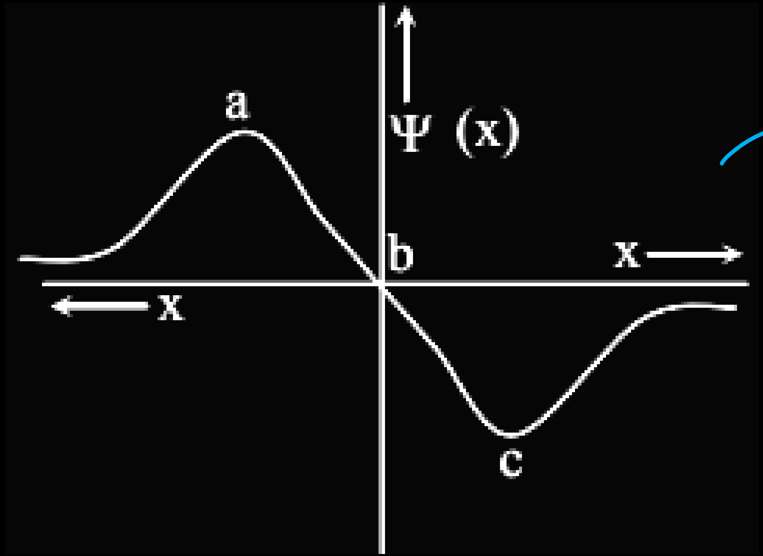
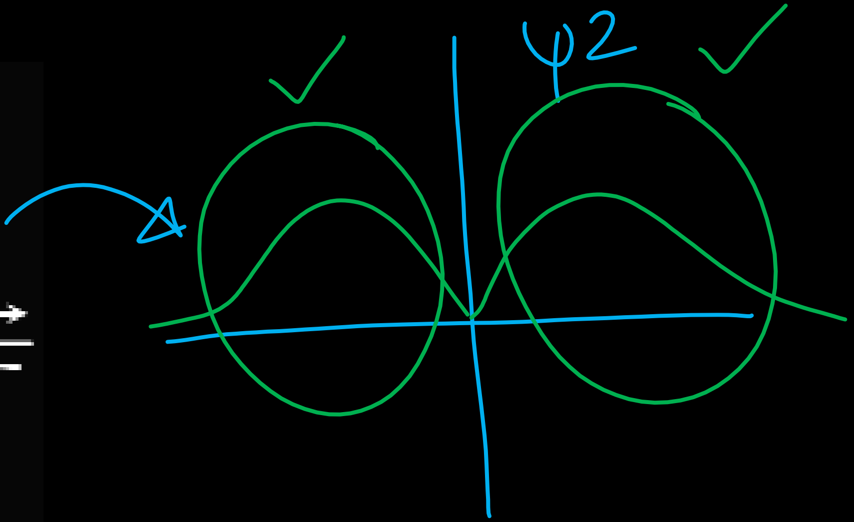




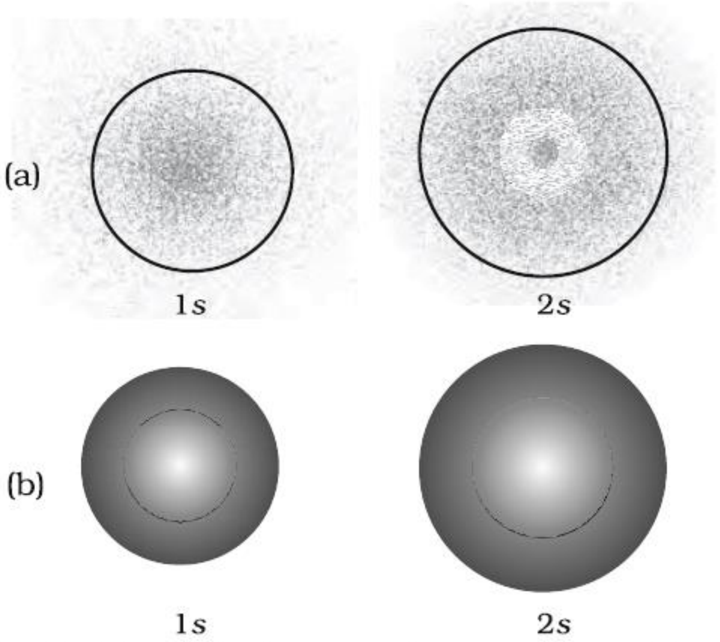


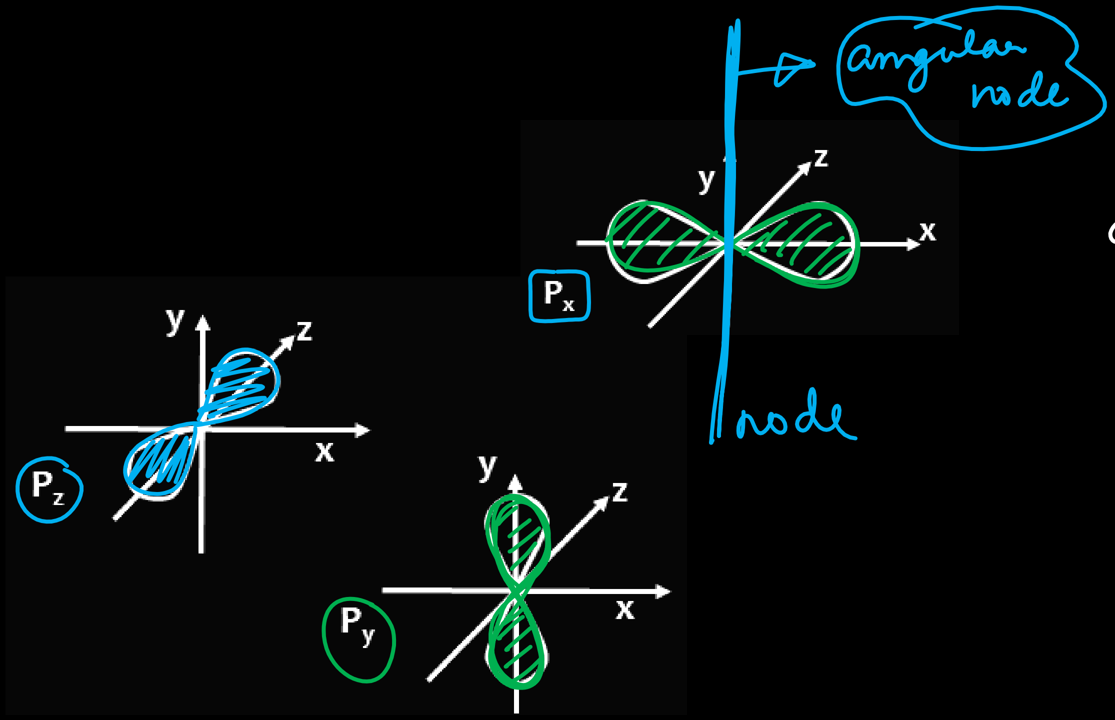
Fig. 2.13 *(a) Probability density plots of 1s and 2s atomic orbitals. The density of the dots represents the probability density of finding the electron in that region. (b) Boundary surface diagram for 1s and 2s orbitals.*

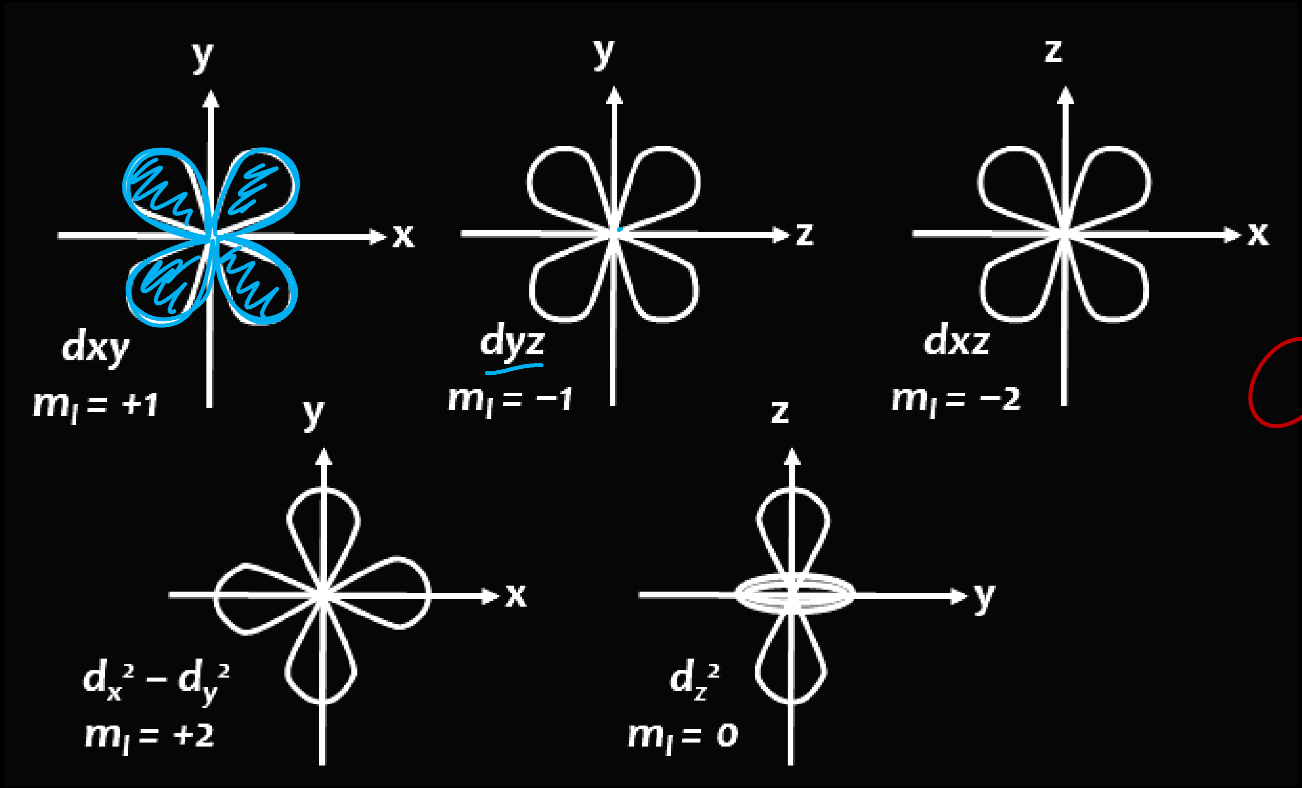
$p \rightarrow l = 1$

$m \rightarrow -1, 0, 1$

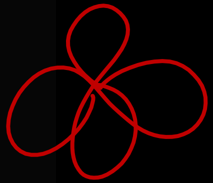
$p_y \quad p_z \quad p_x$

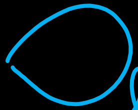
dumbell shaped



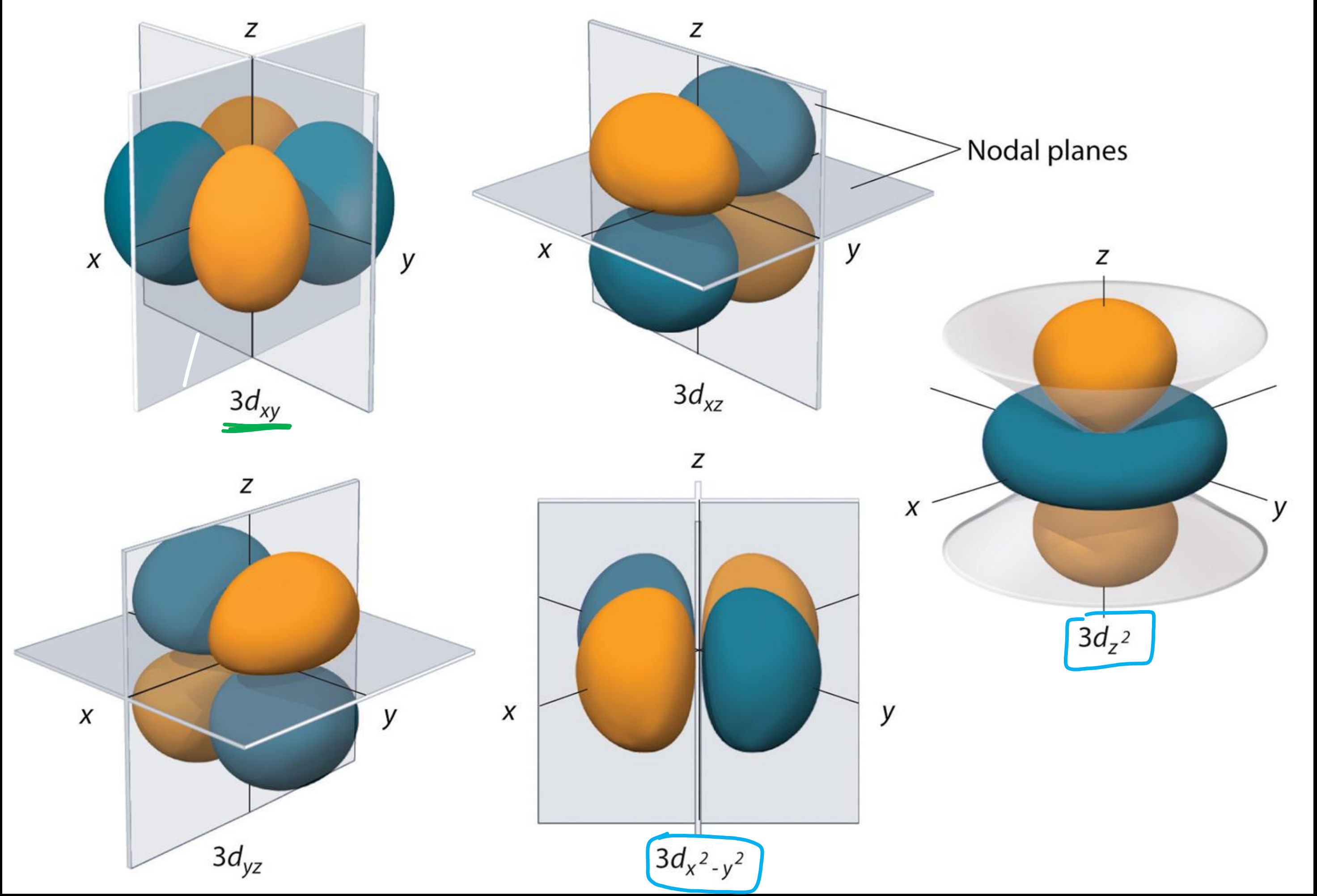


double dumbell





lobe

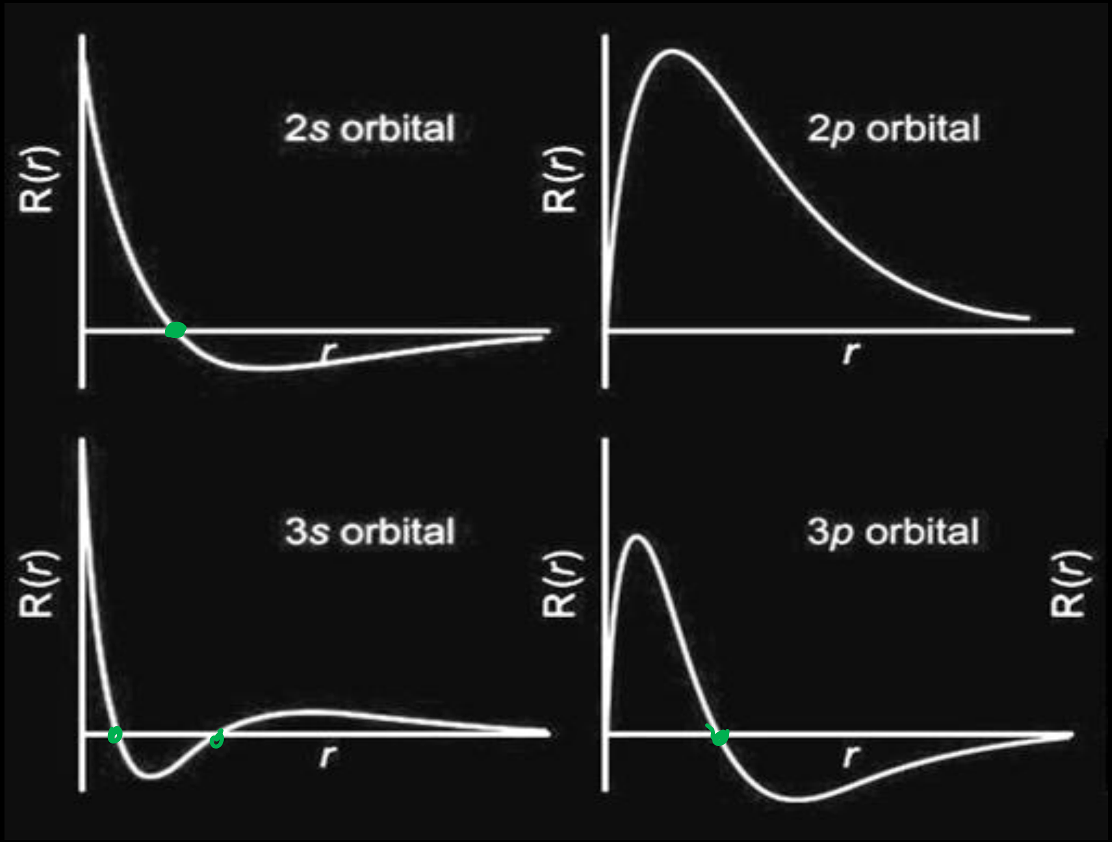

z
x
y
$3d_{xy}$
z
x
y
Nodal planes
$3d_{xz}$
z
x
y
$3d_{z^2}$
z
x
y
$3d_{yz}$
z
x
y
$3d_{x^2-y^2}$

$\boxed{\text{node}} \Rightarrow$ the region where probability of finding e^- is $\boxed{0}$.

(nucleus & ∞ are not counted as nodes)

$$\psi = 0$$

What are Nodes?



The Schrodinger equation for the hydrogen atom is:

$$\Psi_{2s} = \frac{1}{4\sqrt[2]{2\Pi}} \left(\frac{1}{a_0}\right)^{3/2} \left[2 - \frac{r_0}{a_0}\right] e^{-r/a_0} = 0$$

Where a_0 is Bohr radius. If the radial node in 2s is at r_0, then find r in terms of a_0?

$\to 52.9\ pm$

$2 - \frac{r}{a_0} = 0$

$2 = \frac{r}{a_0}$

$r = 2a_0$

Cartesian
coordinates

x, y, z

not good for

e^- wavefunc

polar
coordinates

θ ϕ

theta phi

r

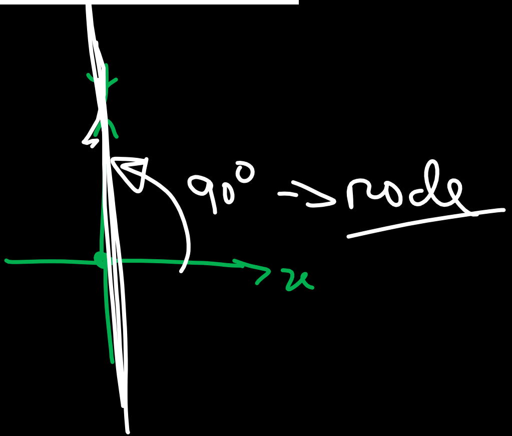
radius

radial
component

angular component

$$\psi_{210}(r,\theta,\phi) = \frac{1}{\sqrt{32\pi a_0^5}}\, r e^{-\frac{r}{2a_0}} \boxed{\cos\theta} \rightarrow \cos 90^\circ = 0$$

2p



no of radial nodes = $n - l - 1$

no of angular nodes = l

total nodes = $\boxed{n-1}$

	rad	ang
3s	2	0
4p	2	1
5d	2	2

0	1	2	3
s	p	d	f

spin only magnetic moment → formula (find)

→ bec of unpaired electron