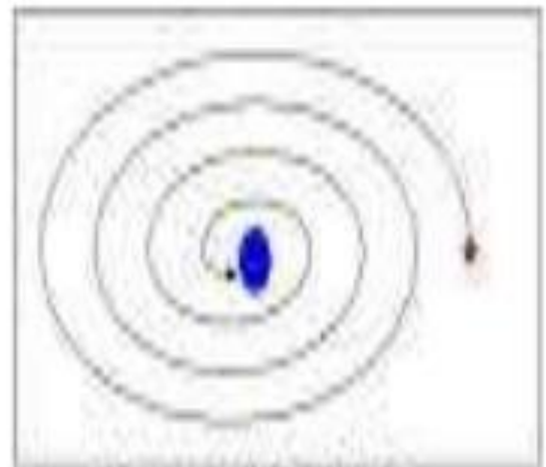
DEFECT OF RUTHERFORD'S THEORY

There were two fundamental defects in Rutherford's atomic model:

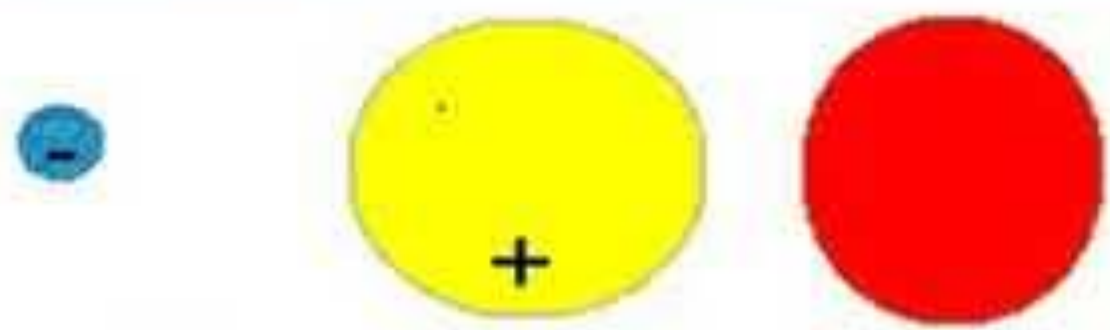
- According to classical electromagnetic theory, being a charge particle electron when accelerated must emit energy. We know that the motion of electron around the nucleus is an accelerated motion, therefore, it must radiate energy. But in actual practice this does not happen. Suppose if it happens then due to continuous loss of energy orbit of electron must decrease continuously. Consequently electron will fall into the nucleus. But this is against the actual situation and this shows that atom is unstable.



- If the electrons emit energy continuously, they should form continuous spectrum. But actually line spectrum is obtained

Properties of electrons, Protons and Neutrons

	Electron	Proton	Neutron
Mass	The electron carries a negligible mass. Its mass is 1 ÷ 1836 part of that of proton.	A proton is 1836 times heavier than electron.	Its mass is almost equal to that of proton.
	The actual mass of an electron is 9.109×10^{-31} Kg.	The actual mass of a proton is 1.672×10^{-27} Kg.	The actual mass of a neutron is 1.675×10^{-27} Kg.



9.109×10^{-31} kg $\quad$ 1.672×10^{-27} kg $\quad$ 1.675×10^{-27} kg

Proton	Neutron	Electron
i) It is positively charged. ii) Its mass is equivalent to a hydrogen atom i.e. 1 a.m.u . iii) It is present inside the nucleus of the atom.	i) It is neutral ii) It is equal in mass to a proton. iii) It is also found inside the atomic nucleus.	i) it is negatively charged. ii) Its mass is 1/1838 of the mass of a proton. iii) It is found outside the nucleus of the atom.

Write the properties of fundamental particles ?

PROPERTIES OF THREE FUNDAMENTAL PARTICLES

Particle	Charge (coloumb)	Relative charge	Mass (kg)	Mass (amu)
Proton	$+1.6022X10^{-19}$	+1	$1.6727X10^{-27}$	1.0073
Neutron	0	0	$1.6750X10^{-27}$	1.0087
ELECTRON	$-1.6022X10^{-19}$	-1	$9.1095X10^{-31}$	$5.4858X10^{-4}$

Today's Learning

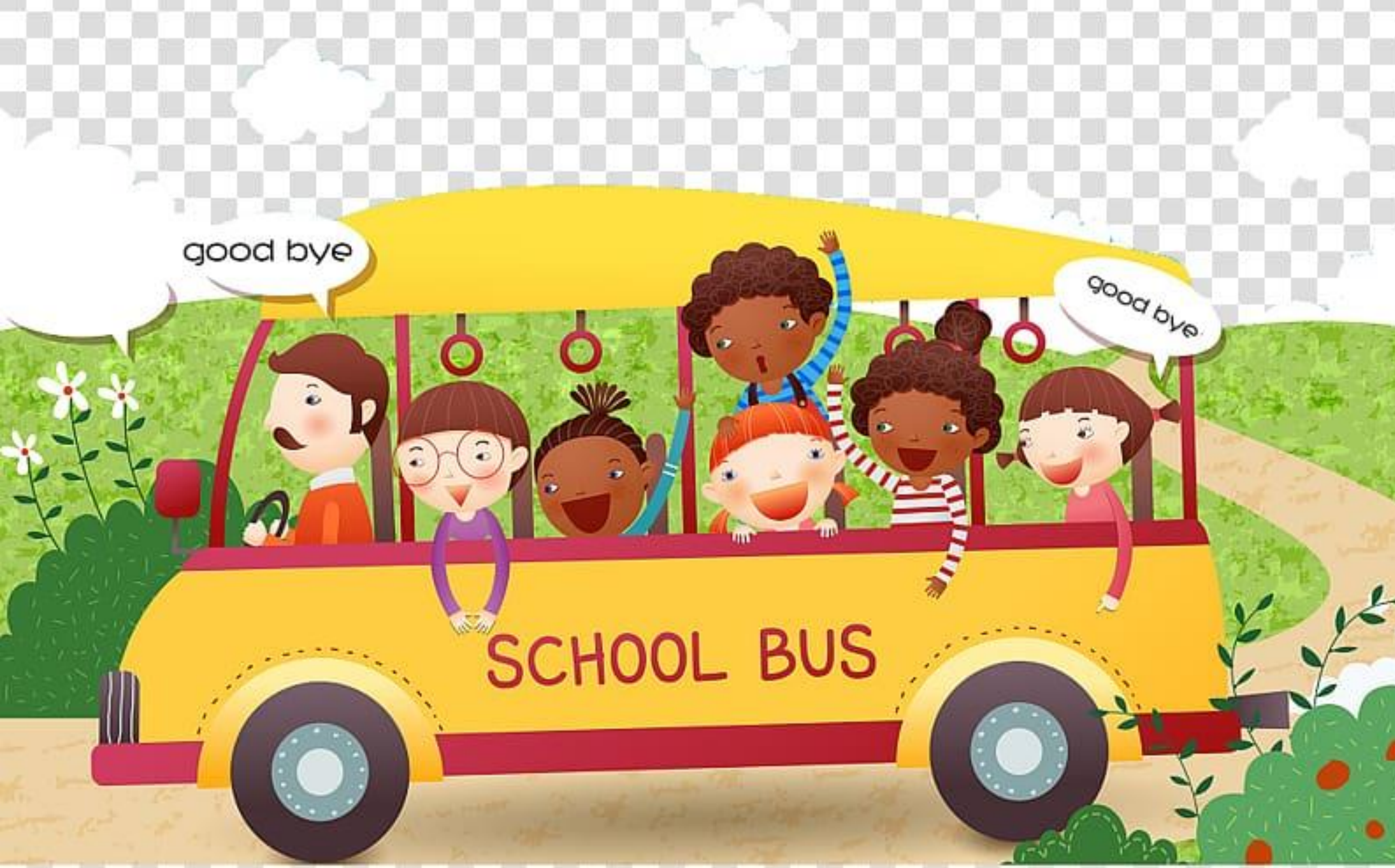
Don't forget to put your name on!

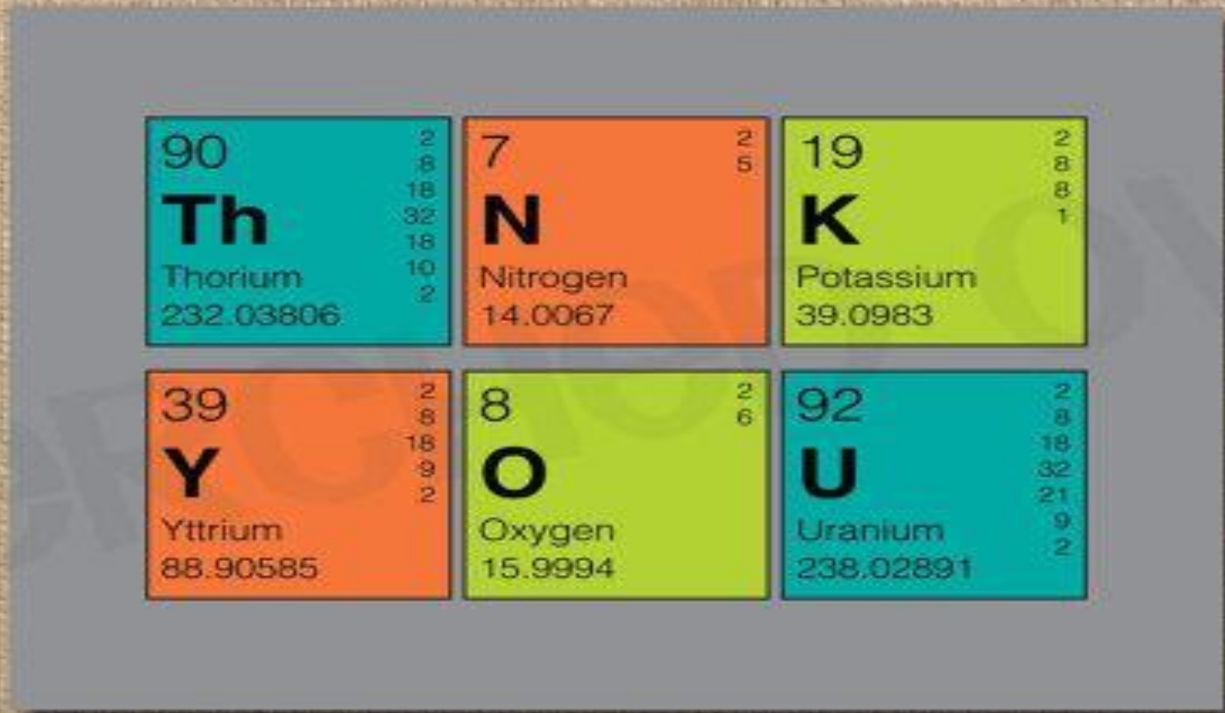
Something I can do now that I couldn't do before the lesson is…	A question I would like to know the answer to is…	I need to improve on…

HOME-WORK:

- NOTE:ATTEMPT ANY ONE QUESTION:
- Write the limitations of Rutherford’s atomic model?
- What were the defects of Rutherford’s atomic model?

LESSON CLOSURE:


good bye
good bye
SCHOOL BUS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2 8 18 9 2
Y
Yttrium
88.90585
8
2 6
O
Oxygen
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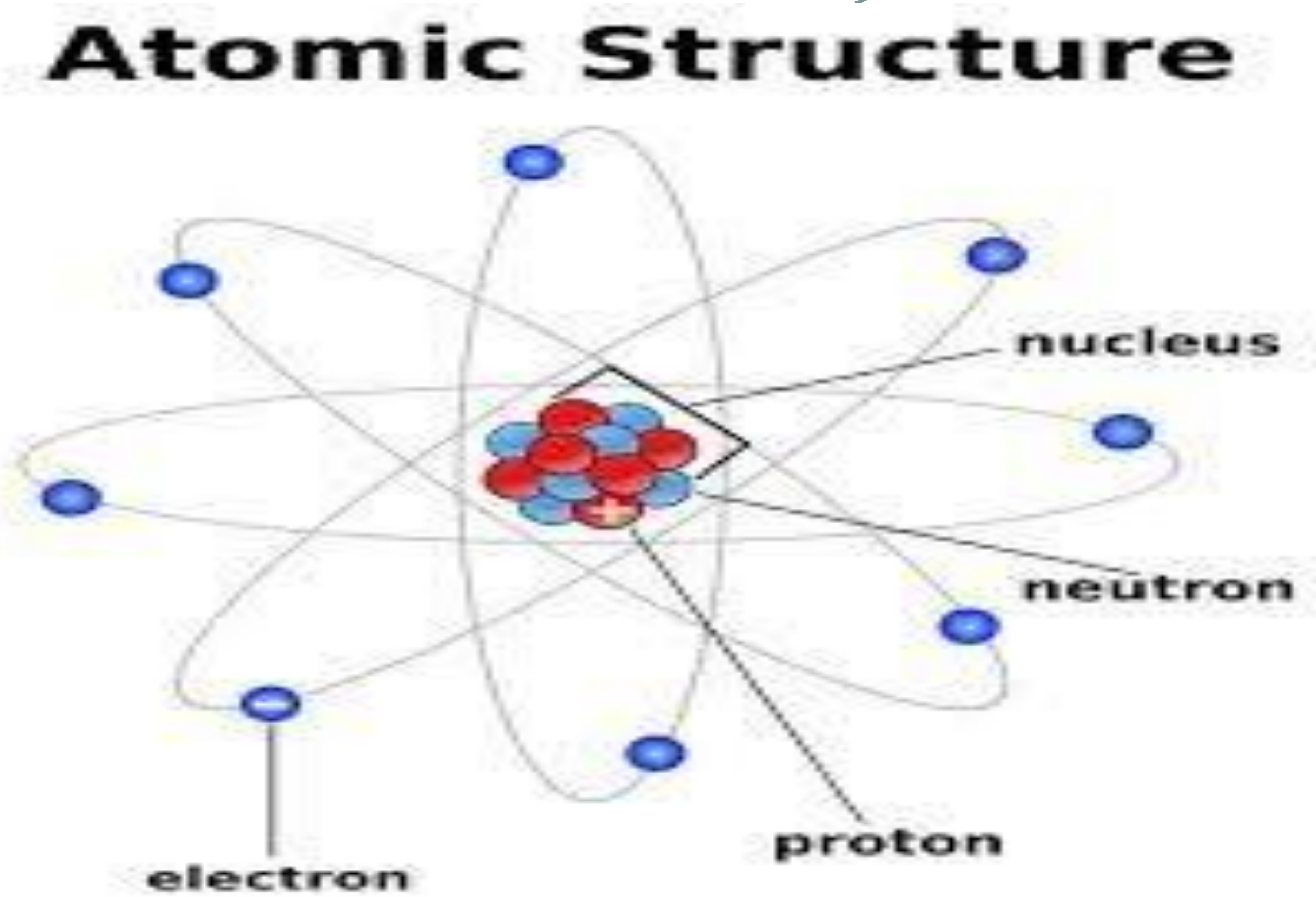
2 8 18 32 21 9 2
U
Uranium
238.02891

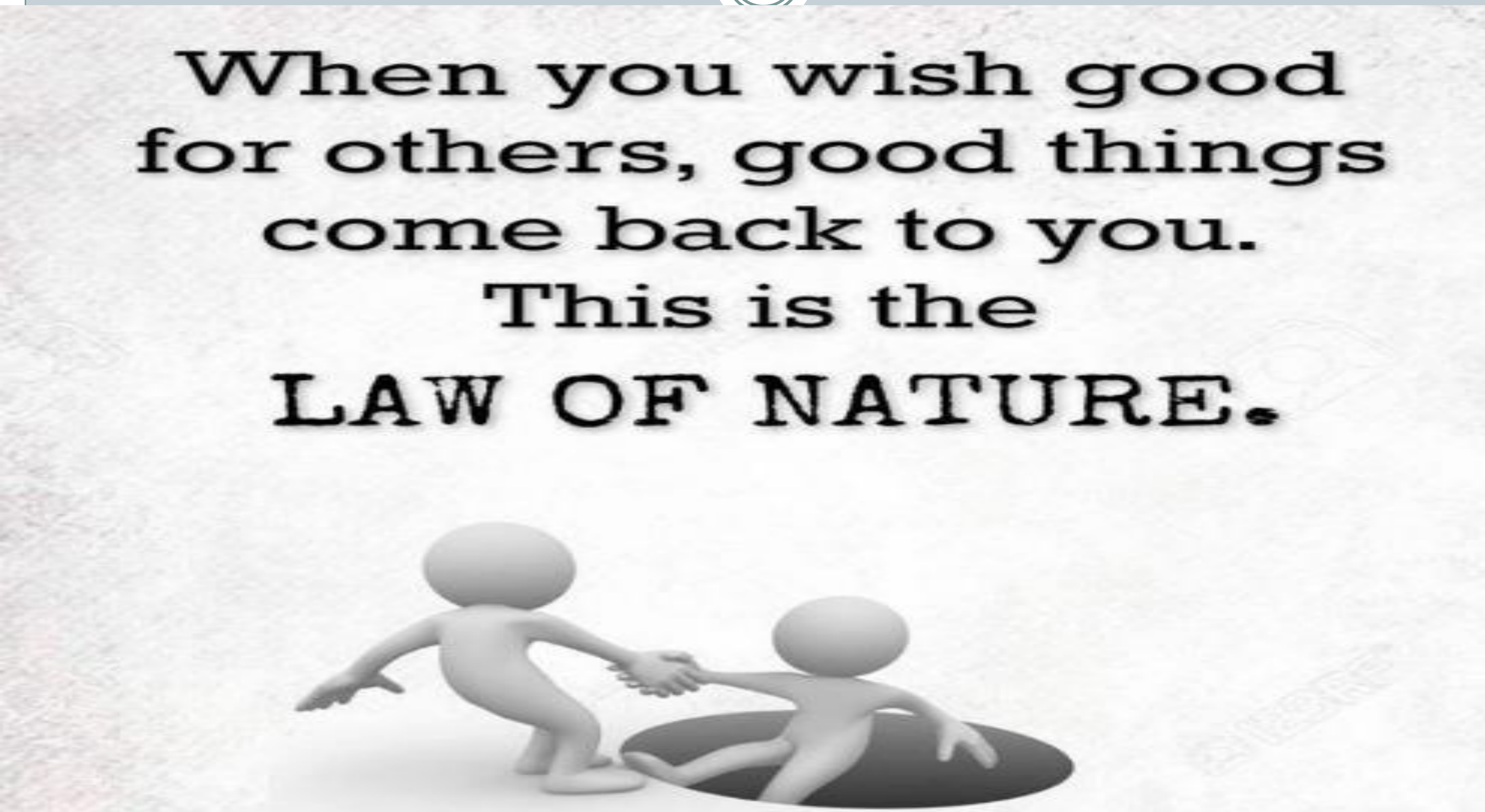

Thank
You

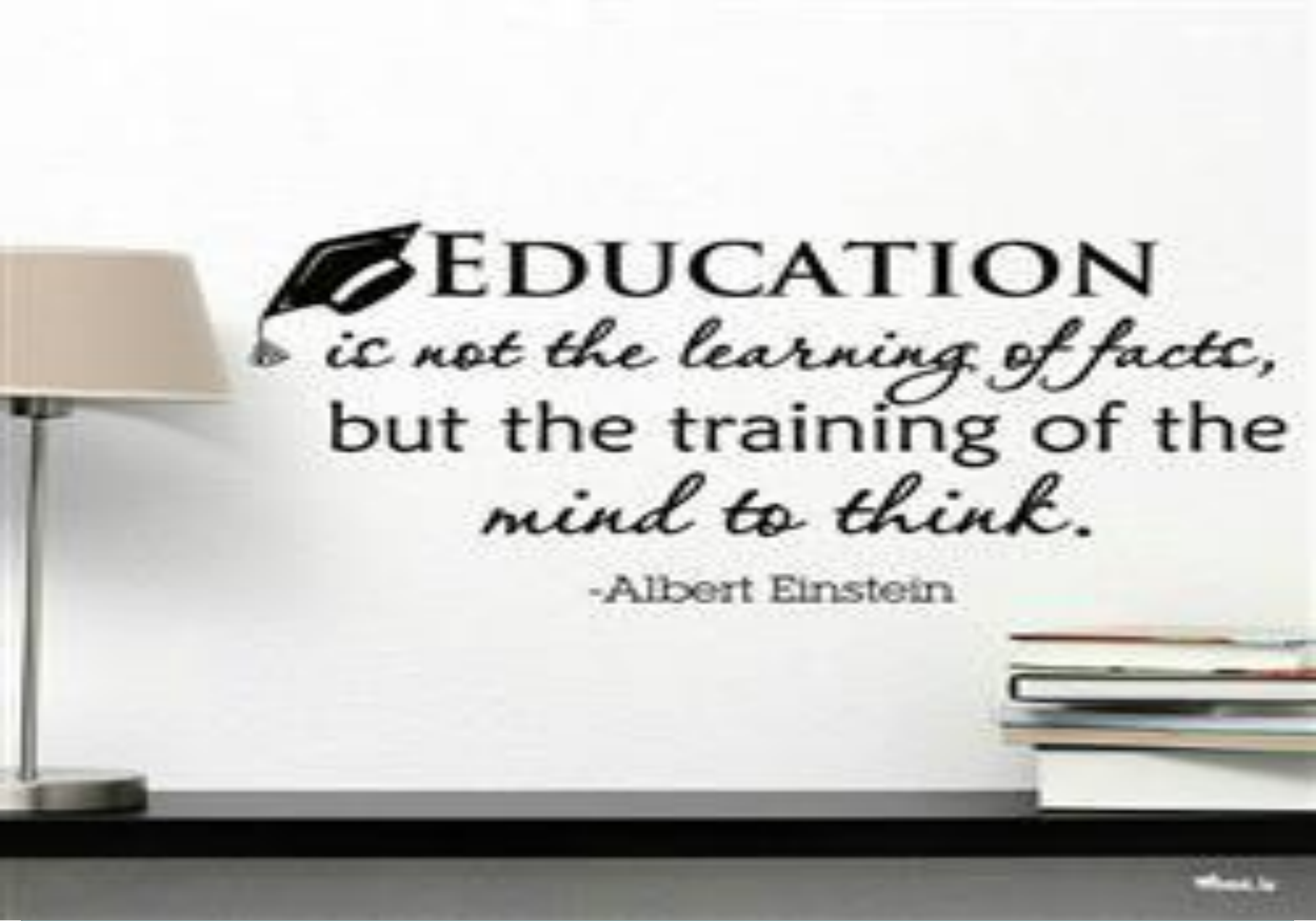

TAKE CARE

LECTURE NO:9



MESSAGE OF THE DAY:




EDUCATION
is not the learning of facts,
but the training of the
mind to think.
-Albert Einstein


WELCOME
BACK
2+2


HAHA! SO WE MEET AGAIN!!


>> Ready for Anything
WELCOME BACK TO
VIRTUAL CLASSES!
LESSON 1
A
B
C
#WEAREALWAYSTOGETHER
AMERICAN

A WARM WELCOME TO ALL THE STUDENTS IN THE ONLINE CLASSES.THIS IS D.CHARLES

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.

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POINTS TO PONDER:

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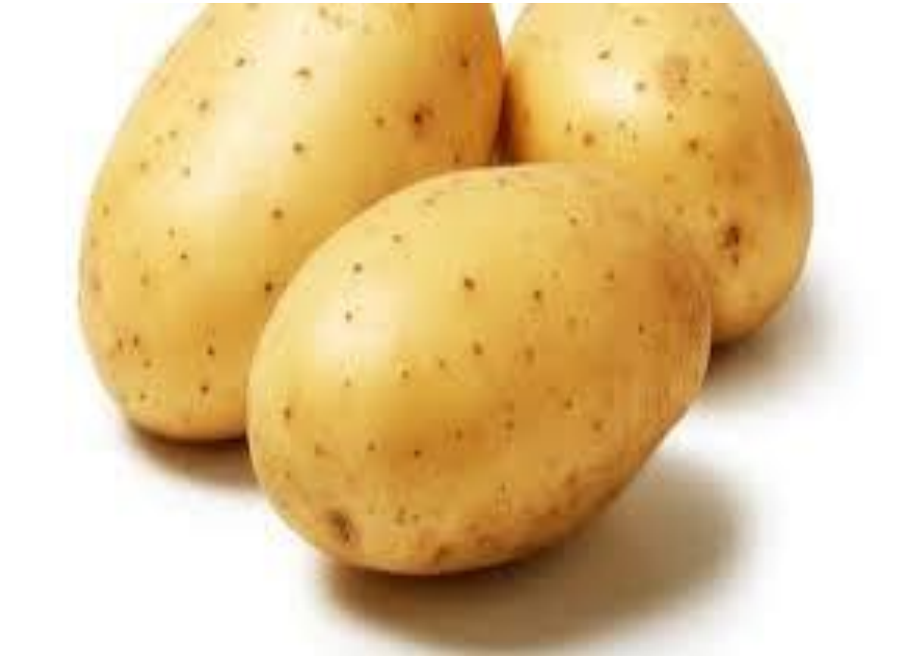

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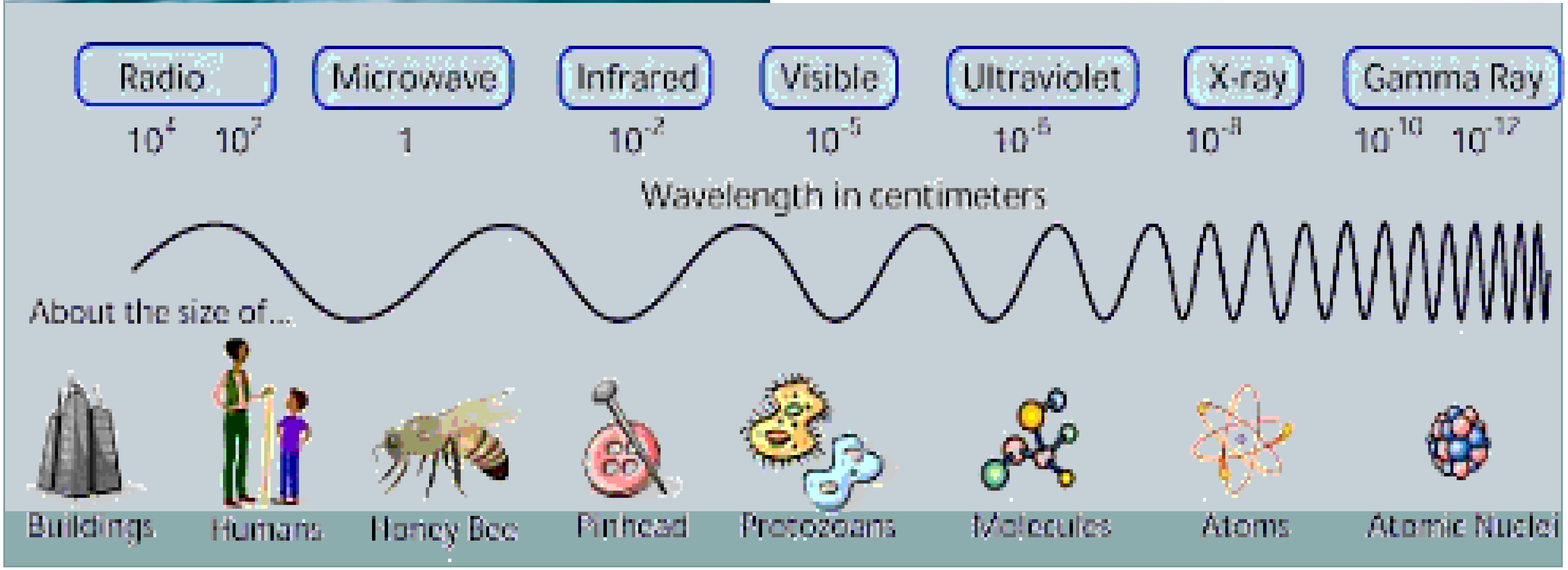


LESSON OBJECTIVES:9

- BY THE END OF THIS PART OF LESSON,STUDENTS WILL BE ABLE TO:
- Explain PLANCK'S THEORY ?
- What were the assumptions of Planck's quantum theory?
- Define frequency, wavelength and wavenumber?

Planck's Quantum Theory

a) Energy is not emitted or absorbed continuously. It is emitted or absorbed in the form of wave packets or quanta.

In case of light the quantum of energy is often called Photon.

Planck's Quantum Theory

b) The amount of energy associated with quantum of radiation is directly proportional to the frequency (v) of radiation i.e.

$$E \propto \nu$$

$$E = h\nu$$

Where, h =Planck's Constant =6.626×10^{-34} J.s

Planck's Quantum Theory

c) A body can emit or absorb energy only in terms of integral multiple of a quantum/photon.

$$E = nh\nu$$

Where n= 1, 2, 3,....

According to Planck's Quantum Theory:

$$E = h\nu$$

Therefore,

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

Thus greater the Wavelength of radiation, lower will be the energy

Planck's Quantum Theory.

Planck's Quantum Theory

- In 1901, Max Planck, suggested that the radiation of energy was not a continuous process. Planck made the following assumptions:

1. Energy is radiated in 'packages' or quanta.
2. Each quantum consists of a specific amount of energy, E, which is directly proportional to the frequency of the radiation: $E = hf$
3. A fraction of a quantum can never be radiated nor absorbed, only whole numbers of quanta.

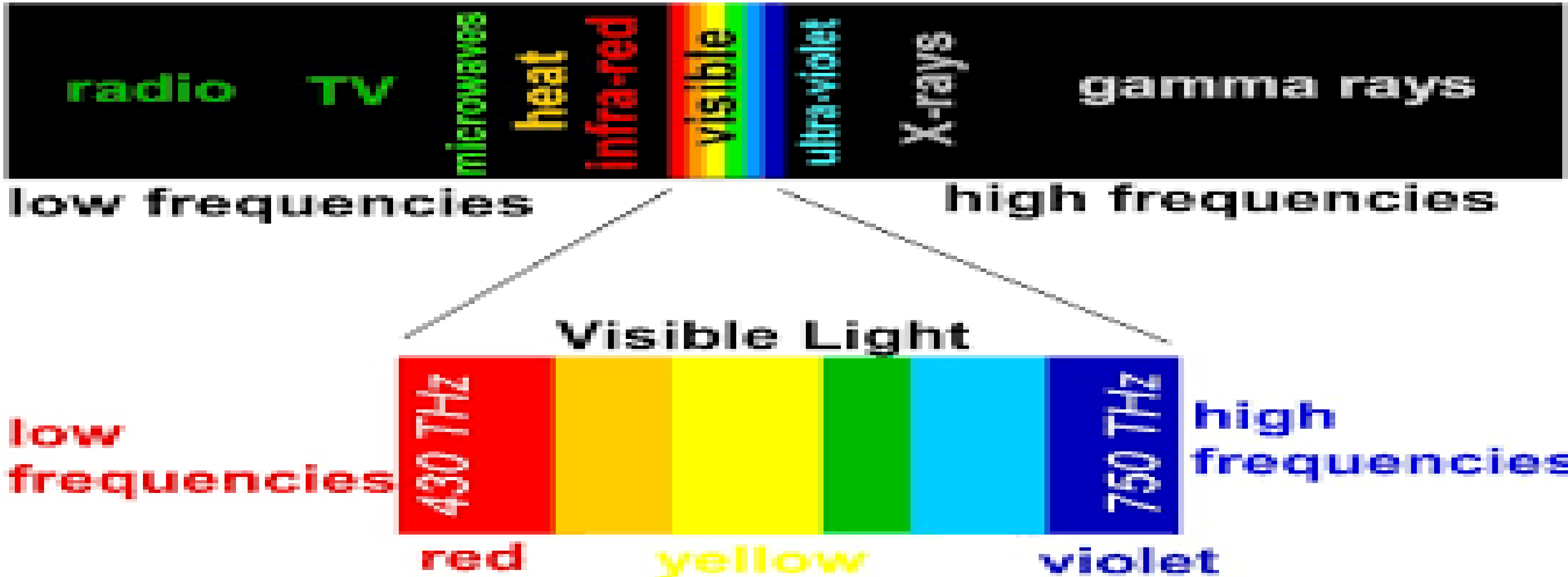
$E_{quanta} \propto$

According to Planck's Quantum Theory:

$$E = h\nu$$

Therefore,

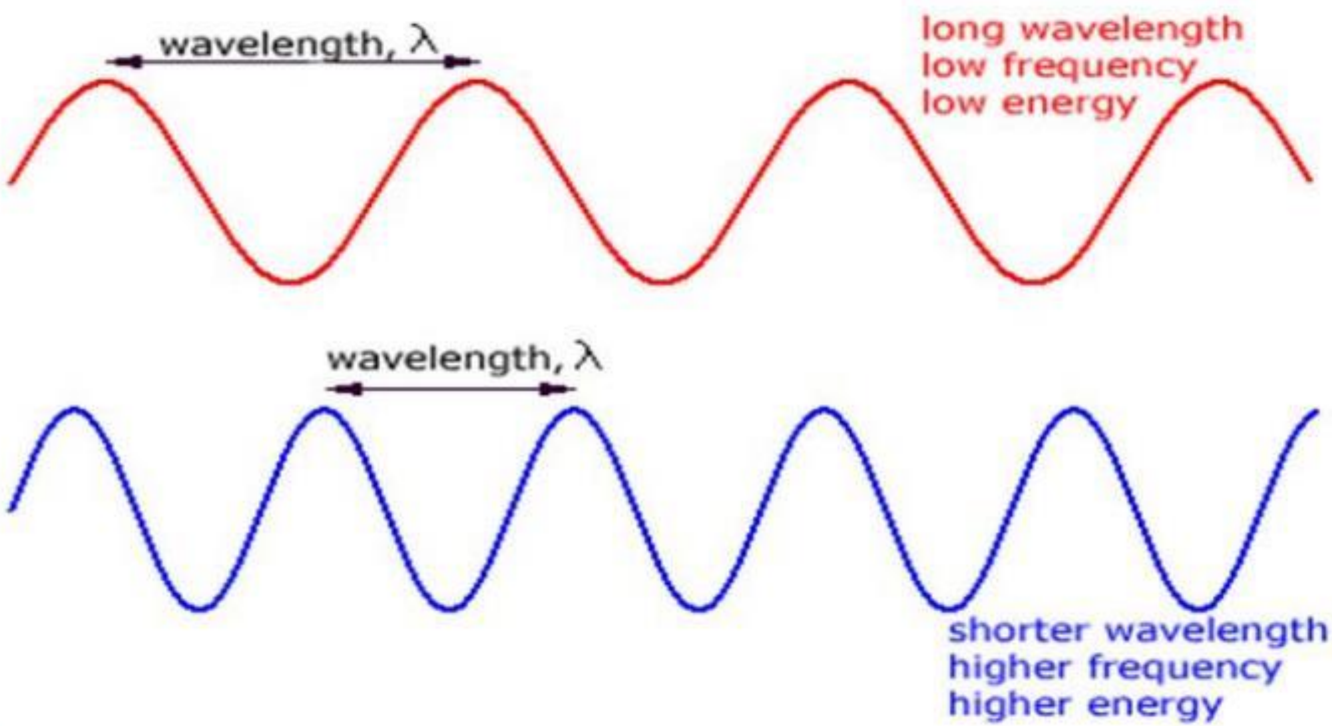
$$E = h\frac{c}{}$$



The **frequency** is the number of waves that pass a point in space during any time interval, usually one second. We measure it in units of cycles (waves) per second, or hertz. The **frequency** of visible **light** is referred to as color, and ranges from 430 trillion hertz, seen as red, to 750 trillion hertz, seen as violet.

What is wavelength?

Wavelength is a measure of distance, so the units for wavelength are always distance units, such as meter, centimeters, millimeters, etc.



WAVE NUMBER:

- **Wave number** is the number of waves per unit distance.
  Wave number is defined as the number of wavelengths per unit distance the spacial wave frequency and is known as spatial frequency. It is a scalar quantity represented by k and the mathematical representation is given as follows: k=1λ Where,
- k is the wave number λ is the wavelength.

PLANK'S QUANTUM THEORY

In 1900, Max Planck proposed quantum theory of radiation. The main points of his theory are

(i) Energy is not emitted or absorbed continuously. It is emitted or absorbed in a discontinuous manner in the form of wave packets called quanta. In case of light, the quantum of energy is often called photon.

(ii) Each wave packet or quantum has a definite amount of energy.

(iii) The amount of energy of the quantum is directly proportional to the frequency of radiation by the equation

$$E \propto \nu$$

or $E = h\nu$ _______ (1)

where E = energy of the quantum ν = frequency of radiation

h = Planck's constant = 6.625×10^{-34} Js

(iv) A body can emit or absorb energy only in terms of integral multiple of a quantum.

i.e. $E = nh\nu$ where n = 1, 2, 3 ...

Frequency

It is the number of waves, which passes through a given point in one second.

It is denoted by ν. Its units are cycles s^{-1} or s^{-1} or Hertz (Hz).

1 Hz = 1 cycles s^{-1}

It is related to wavelength as

$c = \nu\lambda$

or $\nu = c/\lambda$ _______ (2)

where c = velocity of light = $3 \times 10^8\ ms^{-1}$

λ = wavelength of any light radiation

Wave Length

It is the distance between two adjacent crests or troughs in a beam of radiation.

It is denote by λ.

Its units are picometer, nanometer or angstrom etc.

$1\ Å = 10^{-10}$ m, 1 nm = 10^{-9} m, 1 pm = 10^{-12} m

Put eq. (2) in eq. (1)

$E = hc/\lambda$ _______ (3)

λ is related to the wave number i.e.; $\bar{\nu}$ as

$\lambda = 1/\bar{\nu}$ _______ (4)

Wave Number

It is the number of waves per unit distance.

It is denoted by $\bar{\nu}$. Its units are m^{-1} or cm^{-1} etc.

Put eq. (4) in eq. (3)

$E = hc\bar{\nu}$ _______ (5)

...ations (1), (3) and (5) shows that energy of light is directly proportional to its frequency and wave number ...ly proportional to its wavelength.

...er the wave number of photons, greater is the energy associated with them. The relationships of energy, ..., wavelength, wave number about the photon of light are accepted by scientists and used by Bohr in his ...odel.

...2.7:

...of light with energy 10^{-19} J is emitted by a source of light.

...ert this light into the wave length, frequency and wave number of the photon in terms of meters, Hertz ... respectively.

...

$E = 10^{-19}$ J

$h = 6.625 \times 10^{-34}$ Js

$c = 3 \times 10^8$ m/s

$\nu = ?$

$\lambda = ?$

$\bar{\nu} = ?$

Since $E = h\nu$

or $\nu = \dfrac{E}{h} = \dfrac{10^{-19}}{6.625 \times 10^{-34}} = 1.509 \times 10^{14}\ s^{-1}$

Since $\lambda = \dfrac{c}{\nu} = \dfrac{3 \times 10^8}{1.509 \times 10^{14}} = 1.988 \times 10^{-6}$ m

Since $\bar{\nu} = \dfrac{1}{\lambda} = \dfrac{1}{1.988 \times 10^{-6}} = 5.030 \times 10^5\ m^{-1}$

...Convert this energy of photon into ergs and calculate the wavelength in cm, frequency in Hz and wave ...ber in cm^{-1}.

Since 1J = 10^7 erg

$E = 10^{-19}\ J = 10^{-19} \times 10^7 = 10^{-12}$ erg

$h = 6.625 \times 10^{-34}\ Js = 6.625 \times 10^{-34} \times 10^7$ erg s $= 6.625 \times 10^{-27}$ erg s

$c = 3 \times 10^8$ m/s $= 3 \times 10^{10}$ cm/s

$\nu = ?$

$\lambda = ?$

$\bar{\nu} = ?$

> 1 J = 10^7 erg
> 1m = 100 cm
> erg is the unit of energy in c.g.s. system.

$\nu = \dfrac{E}{h} = \dfrac{10^{-12}}{6.625 \times 10^{-27}} = 1.509 \times 10^{14}\ s^{-1}$

$\lambda = \dfrac{c}{\nu} = \dfrac{3 \times 10^{10}}{1.509 \times 10^{14}} = 1.988 \times 10^{-4}$ cm

$\bar{\nu} = \dfrac{1}{\lambda} = \dfrac{1}{1.988 \times 10^{-4}} = 5.030 \times 10^3\ cm^{-1}$

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or $$E = h\nu \quad \text{——— (1)}$$

where E = energy of the quantum ν = frequency of radiation

h = Planck's constant $= 6.625 \times 10^{-34}$ Js

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Put eq. (4) in eq. (3)

$$E = hc\bar{\nu} \quad \text{——— (5)}$$

PLANCK'S QUANTUM THEORY:

- PLANCK'S QUANTUM THEORY : In 1900,Max Planck proposed quantum theory of radiation. The main points of his theory are:
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- Each wave packets or quantum has a definite amount of energy.
- The amount of energy of the quantum is directly proportional to the frequency of radiation by the equation; E α V or E =h v------(1). where E=energy of the quantum v=frequency of radiation, h=Planck's constant =6.625x10^{-34}Js.
- A body can emit or absorb energy only in the terms of integral multiple of a quantum .i.e. E=nhv. where n=123.......

Today's Learning

Don't forget to put your name on!

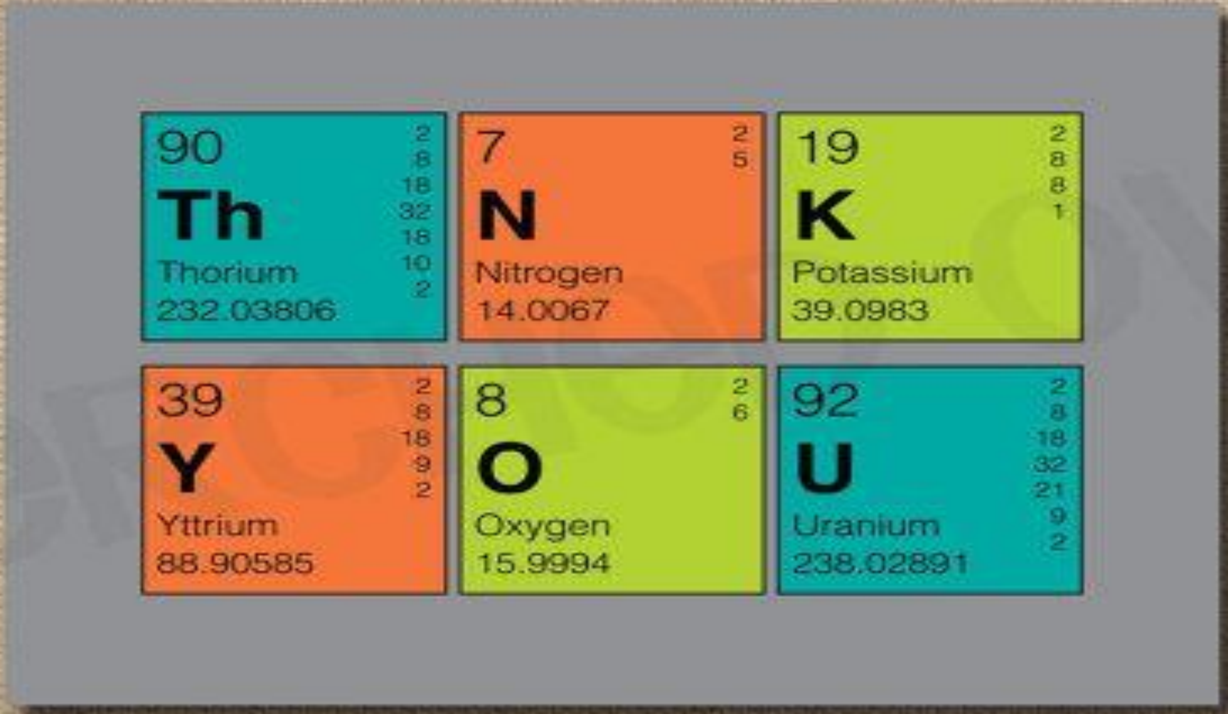
Something I can do now that I couldn't do before the lesson is...	A question I would like to know the answer to is...	I need to improve on...

HOME-WORK:

- NOTE:ATTEMPT ANY ONE QUESTION:
- Explain Planck's quantum theory?
- Define frequency, wavelength and wavenumber and show the relationship between them?

LESSON CLOSURE:

good bye
good bye
SCHOOL BUS



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Uranium
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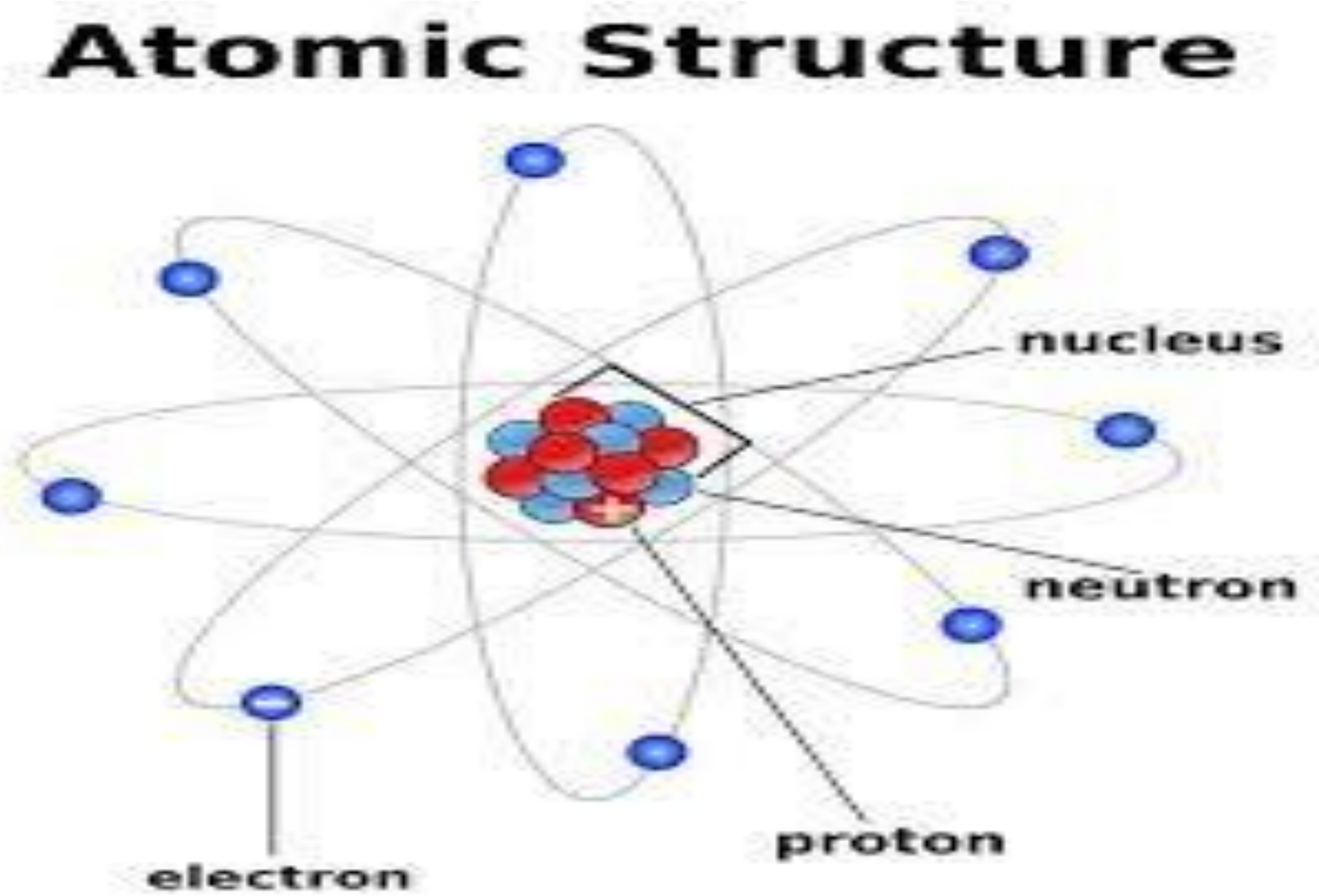
Thank
You


TAKE CARE

LECTURE NO:10



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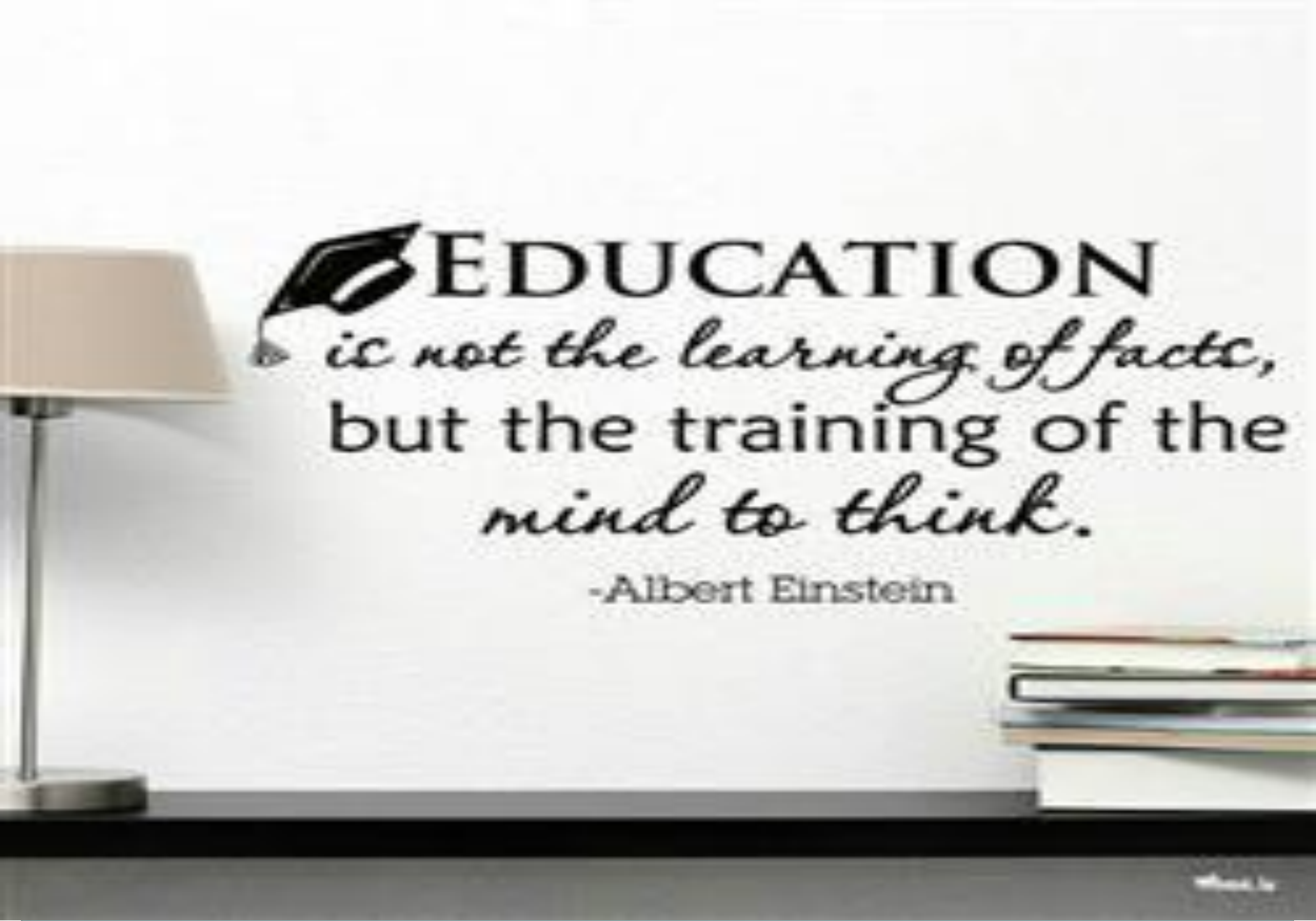
EDUCATION
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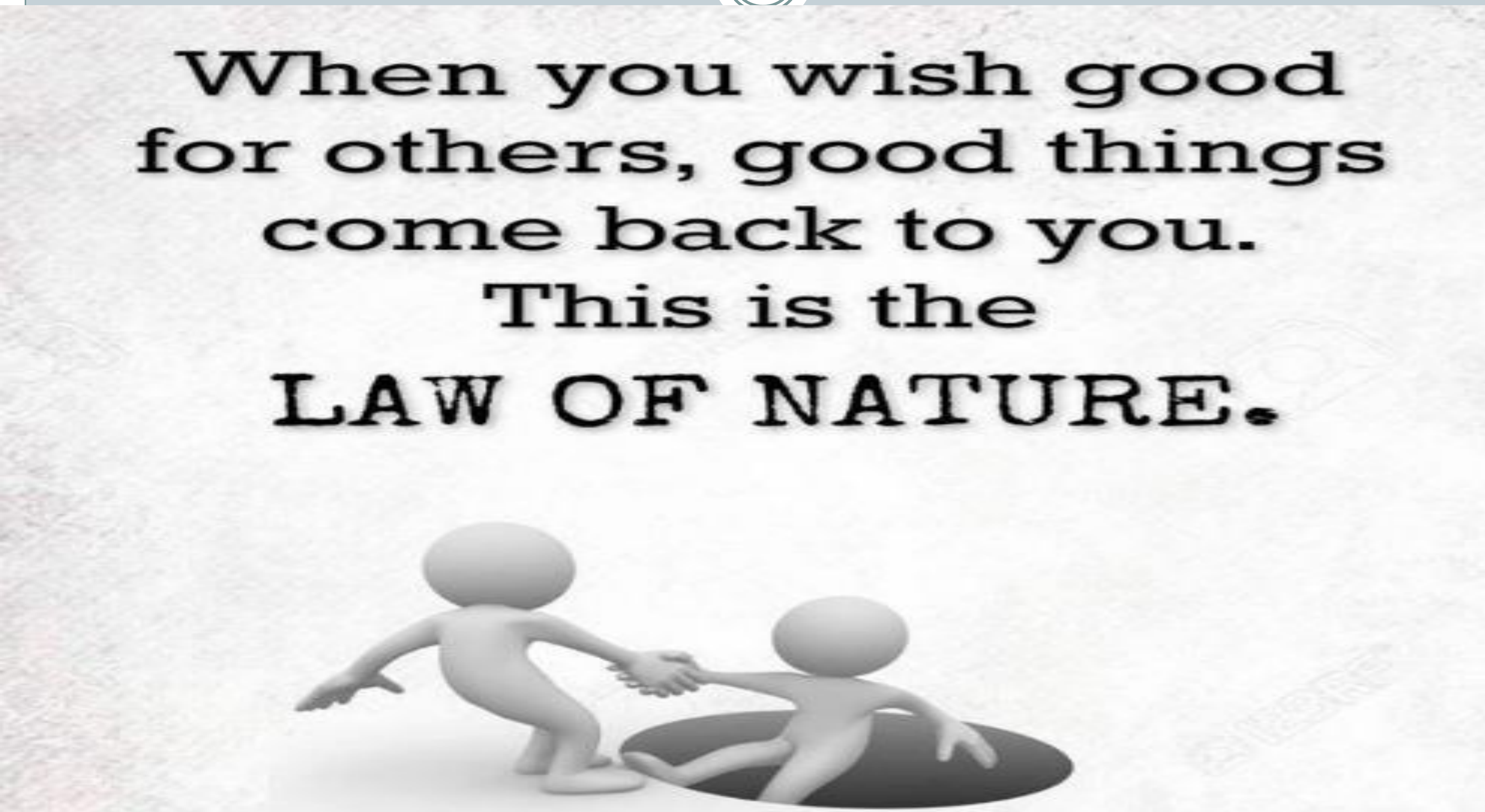
-Albert Einstein

MESSAGE OF THE DAY:




WELCOME
BACK
2+2


HAHA! SO WE MEET AGAIN!!


>> Ready for Anything
WELCOME BACK TO
VIRTUAL CLASSES!
#WEAREALWAYSTOGETHER
AMERICAN

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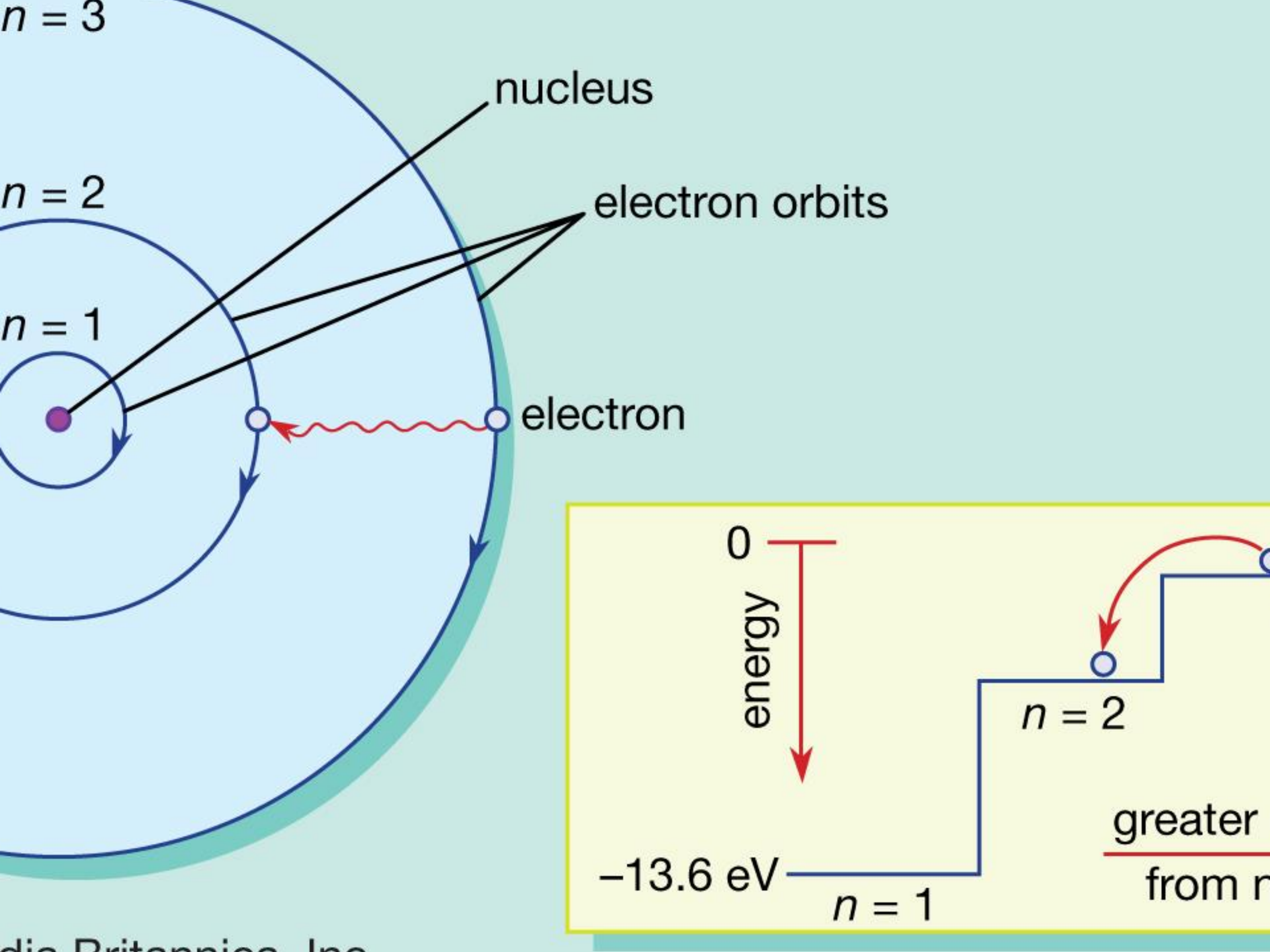
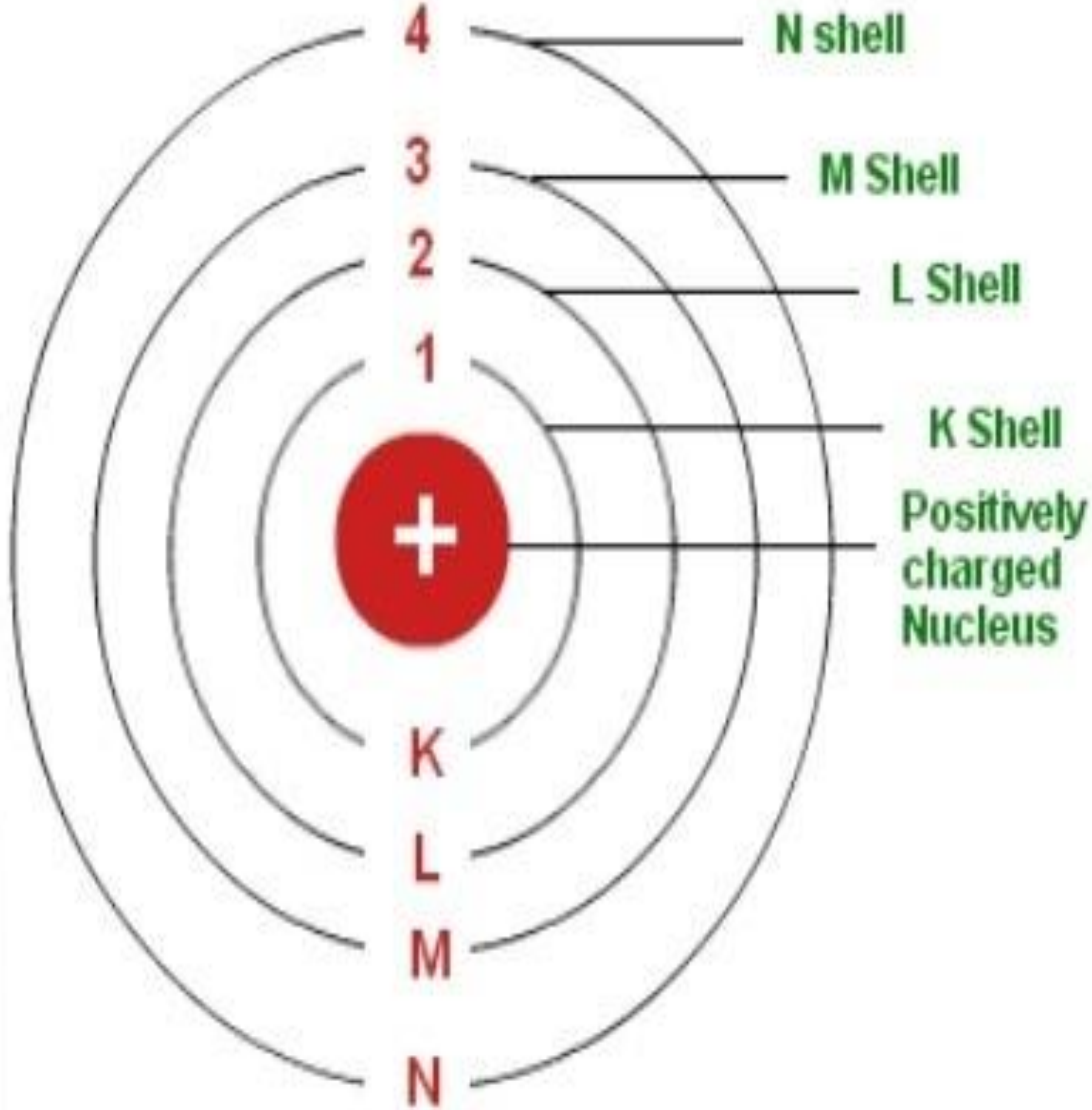


n = 3
nucleus
n = 2
electron orbits
n = 1
electron
0
energy
n = 2
greater
–13.6 eV
from
n = 1

Bohr's Model of the Atom



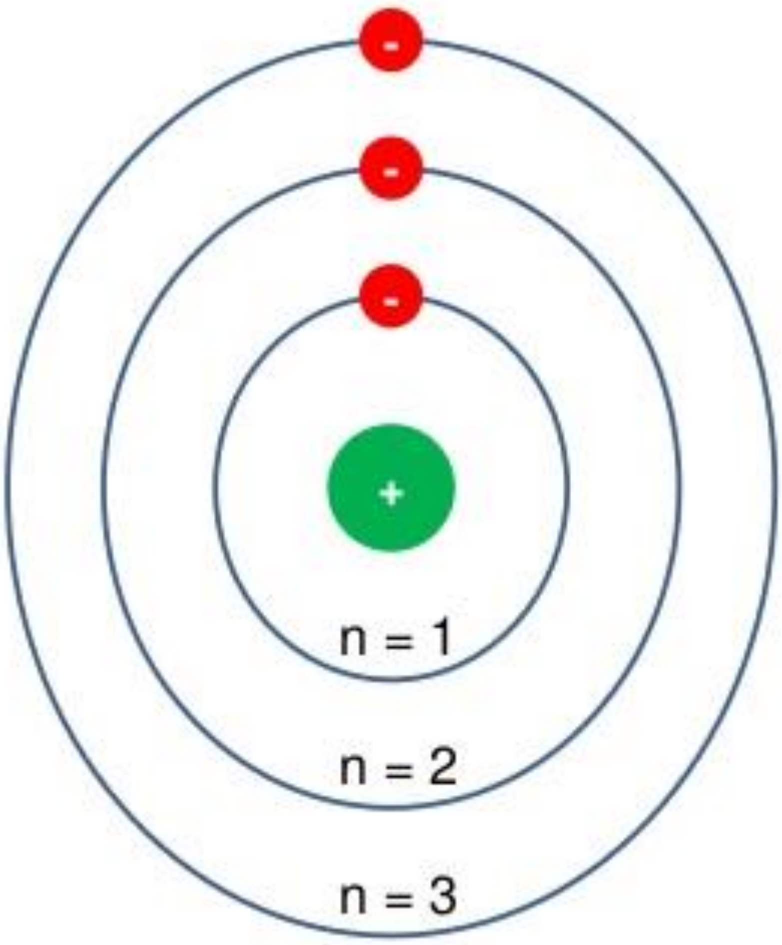
LESSON OBJECTIVES:10

- BY THE END OF THIS PART OF LESSON,STUDENTS WILL BE ABLE TO:
- EXPLAIN BOHR'S ATOMIC MODEL?
- WRITE THE POSTULATES OF BOHR'S ATOMIC MODEL?
- WHAT ARE THE LIMITATIONS OF BOHR'S ATOMIC MODEL?
- WHAT IS MEANT BY ZEEMAN'S AND STARK EFFECT?
- DERIVATION OF RADIUS OF ORBIT OF AN ATOM ?

Bohr's Model of Atom

1- Electrons revolves around the nucleus in definite energy levels called orbits or shells in an atom without radiating energy.

2- As long as an electron remain in a shell it never gains or losses energy.



Postulates of Bohr's theory

- An electron revolves in a circular orbit about the nucleus and its motion is governed by the ordinary laws of mechanics and electrostatics, with the restriction that its angular momentum is quantized (can only have certain discrete values)

$$\text{angular momentum} = m \cdot v \cdot r = nh/2\pi$$

m = mass of electron

v = velocity of electron

r = radius of orbit

n = 1,2,3,4,...(energy levels)

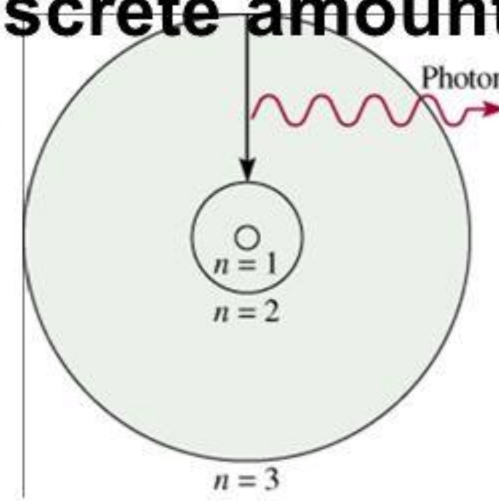
h = Planck's constant

Postulates of Bohr's Atomic Model

- Electrons revolve round the nucleus with definite velocities in concentric circular orbits situated at definite distances from the nucleus. The energy of an electron in a certain orbit remains constant. As long as it remains in that orbit, it neither emits nor absorbs energy. These are termed stationary states or main energy states.

Bohr's Postulates:

- When electron absorbs energy, it will move from ground state to higher energy level that is called excited stated
  - The excited state is unstable, then electron will lose energy and fall back to the lower energy levels
  - These energy will be radiated as light.
- The light emitted with a discrete amount of energy is called photon :



Postulates Neil Bohr model of an atom

- Bohr's first postulate was that *an electron in an atom could revolve in certain stable orbits without the emission of radiant energy*,
- contrary to the predictions of electromagnetic theory. According to this postulate, each atom has certain definite stable states in which it can exist, and each possible state has definite total energy. These are called the stationary states of the atom.
- Bohr's second postulate defines these stable orbits. This postulate states that the *electron* revolves around the nucleus *only in those orbits for which the angular momentum is some integral multiple of h/*2π where *h* is the Planck's constant (= 6.6 × 10–34 J s).
- Thus the angular momentum (*L*) of the orbiting electron is quantised. That is $L = nh/2\pi$
- Bohr's third postulate incorporated into atomic theory the early quantum concepts that had been developed by Planck and Einstein.
- It states that *an electron might make a transition from one of its specified non-radiating orbits to another of lower energy. When it does so, a photon is emitted having energy equal to the energy difference between the initial and final states. The frequency of the emitted photon is then given by*
- $h\nu = E_i - E_f$
- where E_i and E_f are the energies of the initial and final states and $E_i > E_f$.

THE BOHR MODEL OF THE ATOM

- 1. Electrons in the atom can only have fixed amounts of energy.
- The electrons revolve around the nucleus only in certain allowed orbits called stationary states.
- When in a stationary state, the electron cannot radiate any of its energy.
- The electron is only found in the stationary state so the energy is quantized inside the atom.

Limitations of Bohr's Atomic Model :

Bohr's model of the hydrogen atom is considered as an improvement over Rutherford's nuclear model because it accounts for the stability of atom or ion. Limitations of Bohr's atomic model are as follows :

(i) Bohr's atomic model failed to explain the atomic spectrum of atoms other than hydrogen atom.

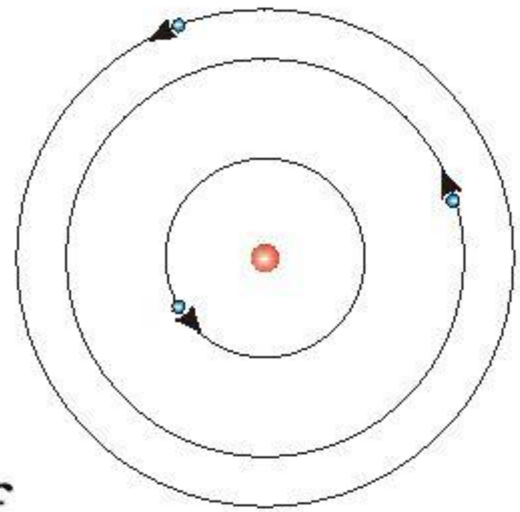
(ii) Bohr's model could not give better explanation of hydrogen spectrum because it could not explain the spectrum when two spectral lines are very close to each other (i.e. doublet)

(iii) It failed to explain Zeeman effect i.e. splitting of spectral lines under the influence of magnetic field.

(iv) It could not explain the ability of atoms to form molecule by chemical bonds.

Limitations of the Bohr Model

a. Model could not calculate the wavelengths of observed spectra of multi-electron atoms.
b. Model could not explain the chemical behavior of atoms.
c. Bohr used classical mechanics to understand the behaviors of small particles.
d. The Bohr model is also known as the planetary, solar system, or satellite model.



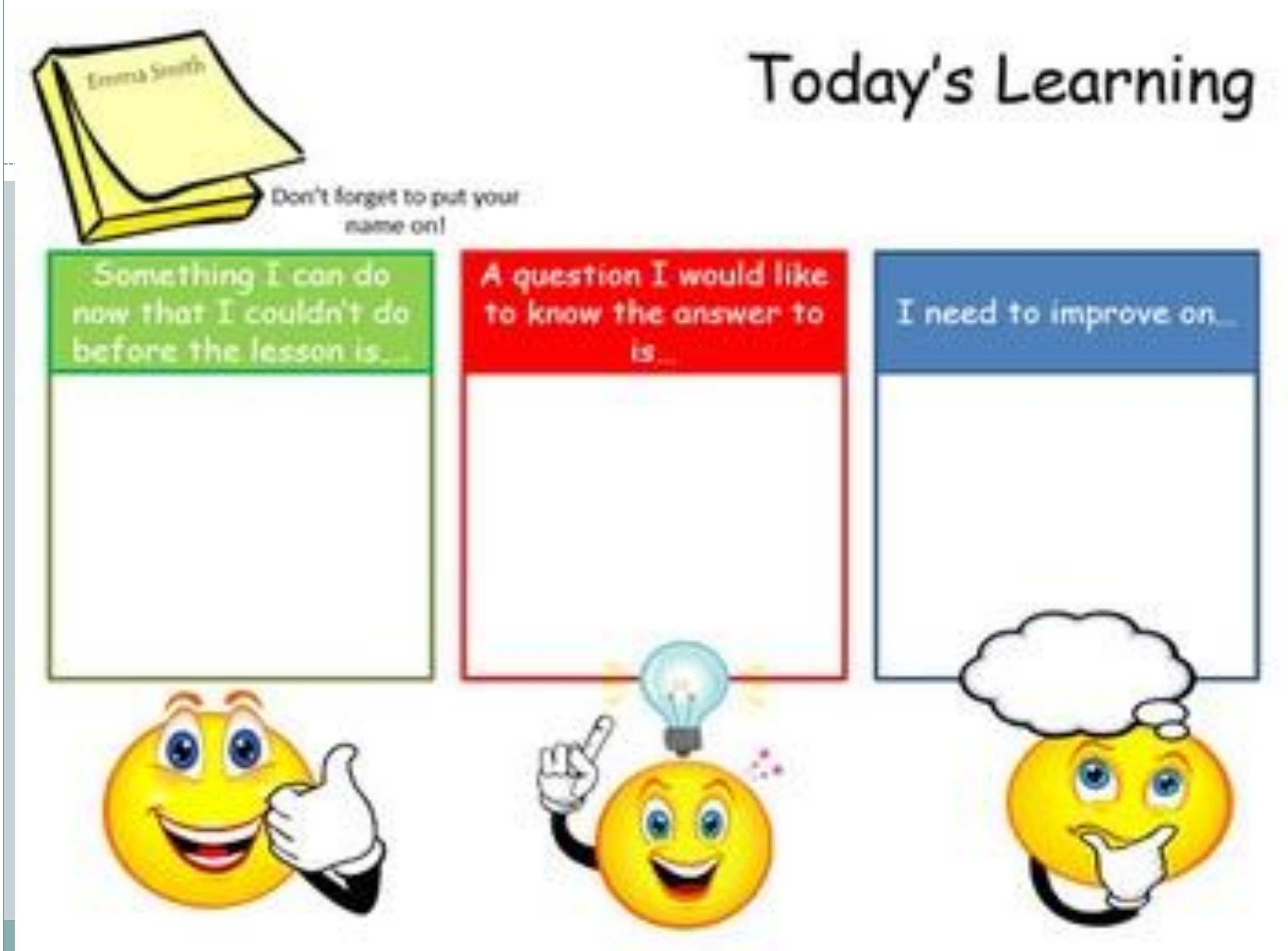
SHORTCOMINGS OF BOHR ATOM MODEL

- According to Bohr, the radiation results when an electron jumps from one orbit to another it emits radiation . But how it happens it was not explained by him.
- Bohr successfully explained the observed spectra of atoms such as He,Li,Be etc. but he couldn't explain atoms with large number of electrons.
- Bohr atom model doesn't obey Heisenberg's Uncertainty principle
- There is no explanation for zeeman and stark effect.

5.5.7 Defects of Bohr's Atomic Model

1. Bohr's theory can successfully explain the origin of the spectrum of H-atom and ions like He^{+1}, Li^{+2} and Be^{+3}, etc. These are all one electron systems. But this theory is not able to explain the origin of the spectrum of multi-electrons or poly-electrons system like He, Li and Be, etc.
2. When the spectrum of hydrogen gas is observed by means of a high resolving power spectrometer, the individual spectral lines are replaced by several very fine lines, i.e. original lines are seen divided into other lines. The H_{α}- line in the Balmer series is found to consist of five - component lines. This is called fine structure or multiple structure. Actually, the appearance of several lines in a single line suggests that only one quantum number is not sufficient to explain the origin of various spectral lines.
3. Bohr suggested circular orbits of electrons around the nucleus of hydrogen atom, but researches have shown that the motion of electron is not in a single plane, but takes place in three dimensional space. Actually, the atomic model is not flat.
4. When the excited atoms of hydrogen (which give an emission line spectrum) are placed in a magnetic field, its spectral lines are further split up into closely spaced lines. This type of splitting of spectral lines is called Zeeman effect. So, if the source which is producing the Na - spectrum is placed in a weak magnetic field, it causes the splitting of two lines of Na into component lines. Similarly, when the excited hydrogen atoms are placed in an electrical field, then similar splitting of spectral lines takes place which is called "Stark effect". Bohr's theory does not explain either Zeeman or Stark effect.

However, in 1915, Sommerfeld suggested the moving electrons might describe in addition to the circular orbits elliptic orbits as well wherein the nucleus lies at one of the foci of the ellipse.

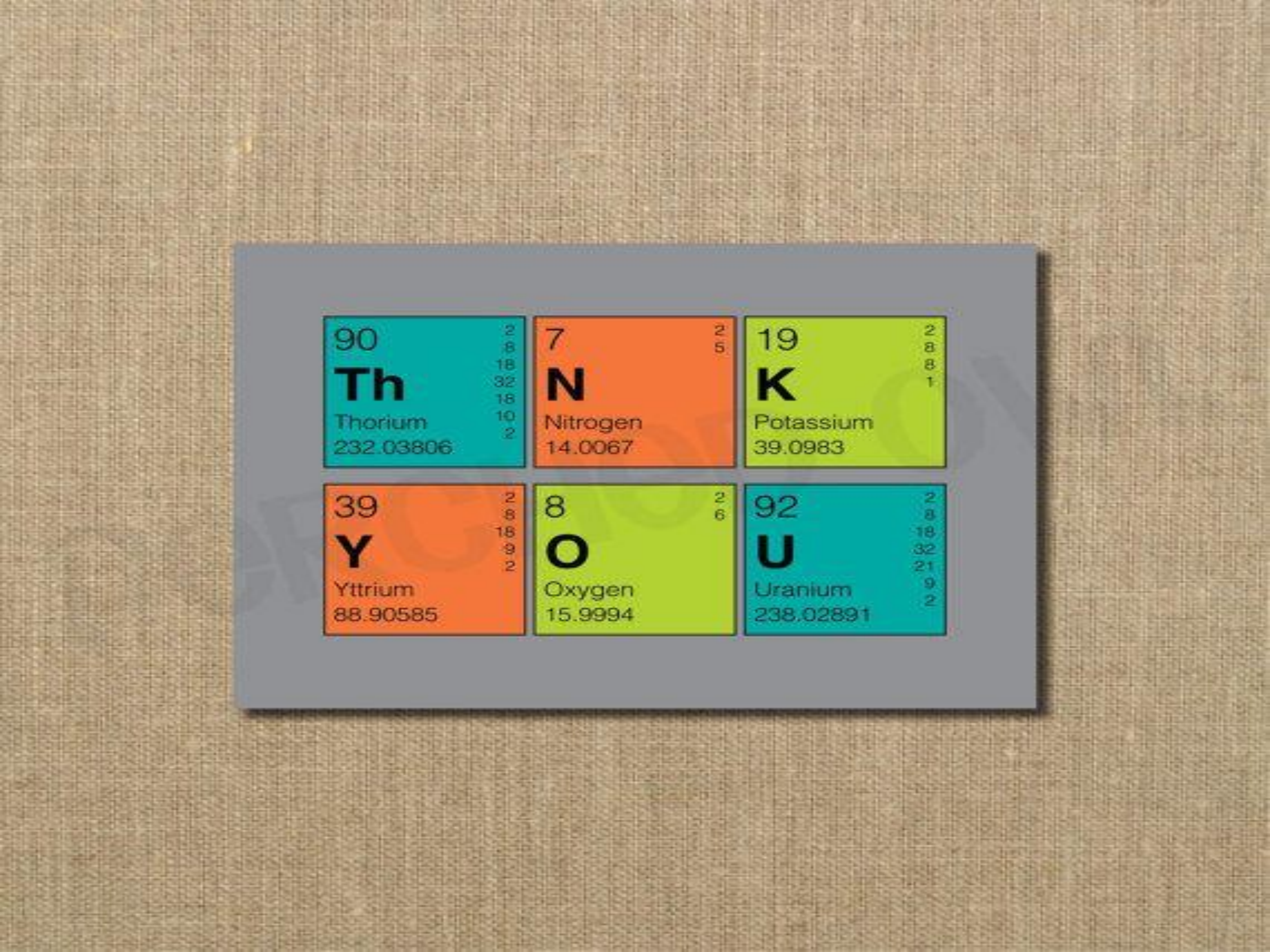

Today's Learning
Emma Smith
Don't forget to put your name on!
Something I can do now that I couldn't do before the lesson is...
A question I would like to know the answer to is...
I need to improve on...

HOME-WORK:

- NOTE:ATTEMPT ANY ONE QUESTION:
- Write the postulates of Bohr's atomic model?
- What are limitation of Bohr's atomic model?
- Explain Zeeman and Stark's effect ?

LESSON CLOSURE:


good bye
good bye
SCHOOL BUS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Th
Thorium
232.03806
7
2 5
N
Nitrogen
14.0067
19
2 8 8 1
K
Potassium
39.0983
39
2 8 18 9 2
Y
Yttrium
88.90585
8
2 6
O
Oxygen
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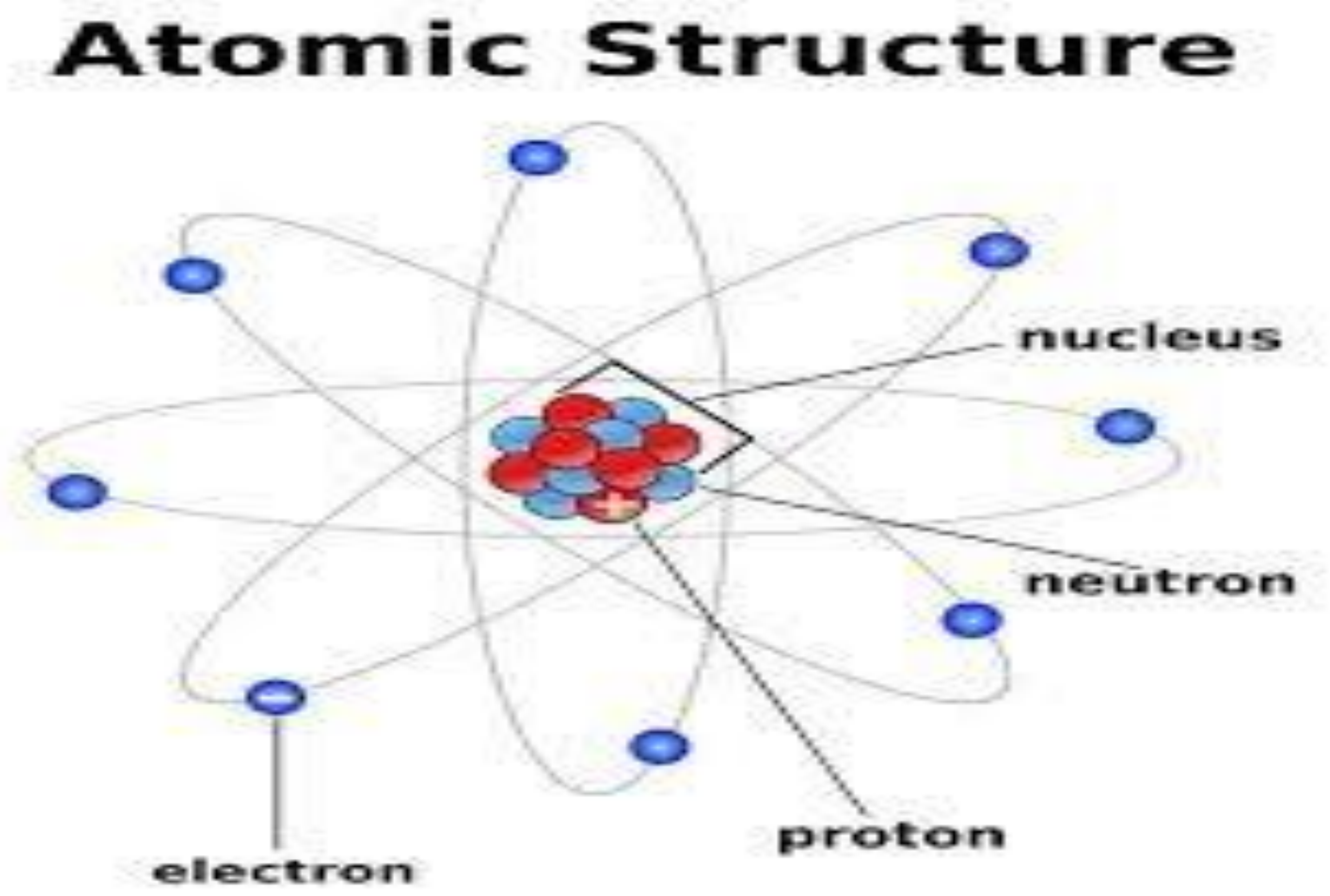
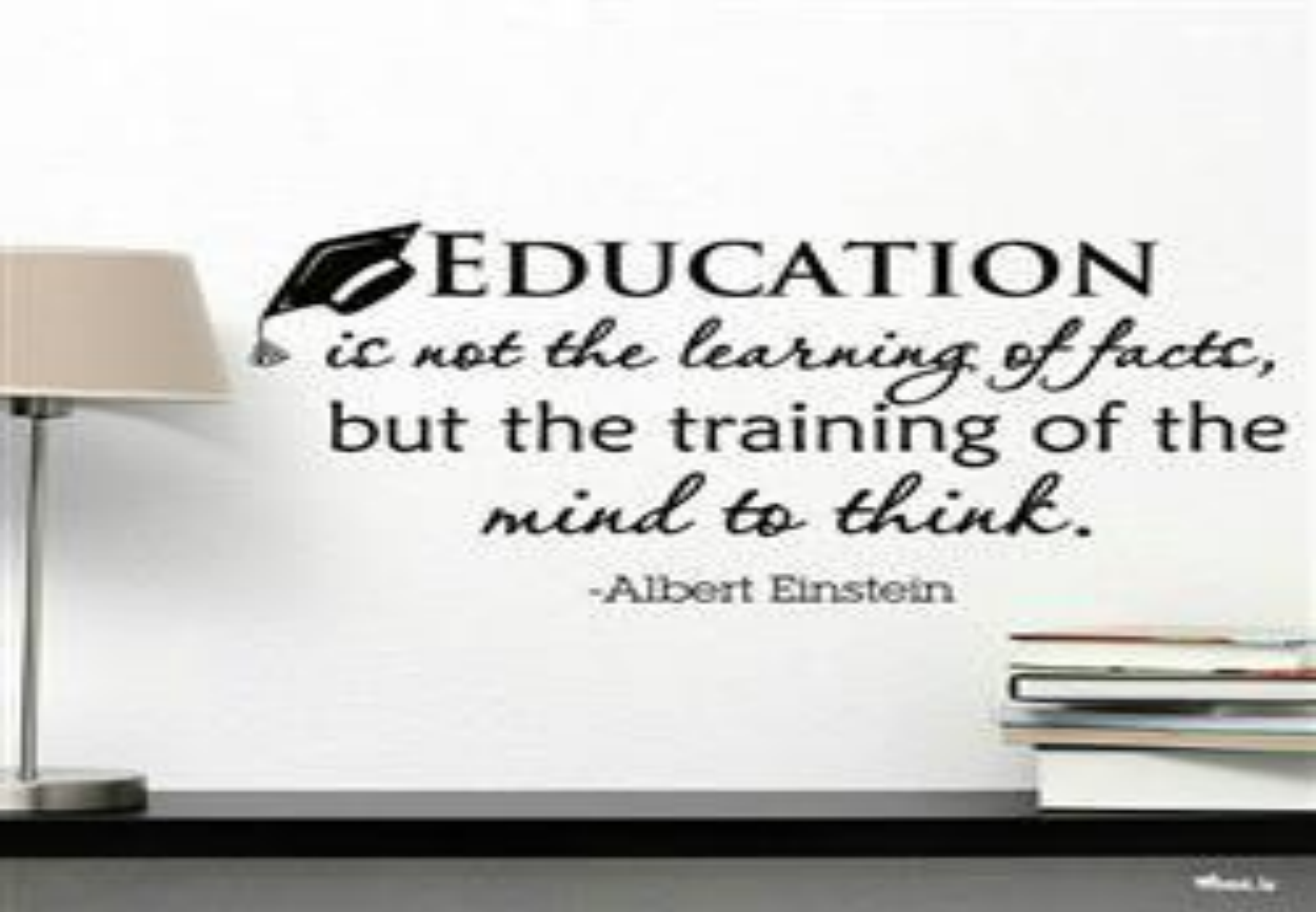
2 8 18 32 21 9 2
U
Uranium
238.02891


Thank
You

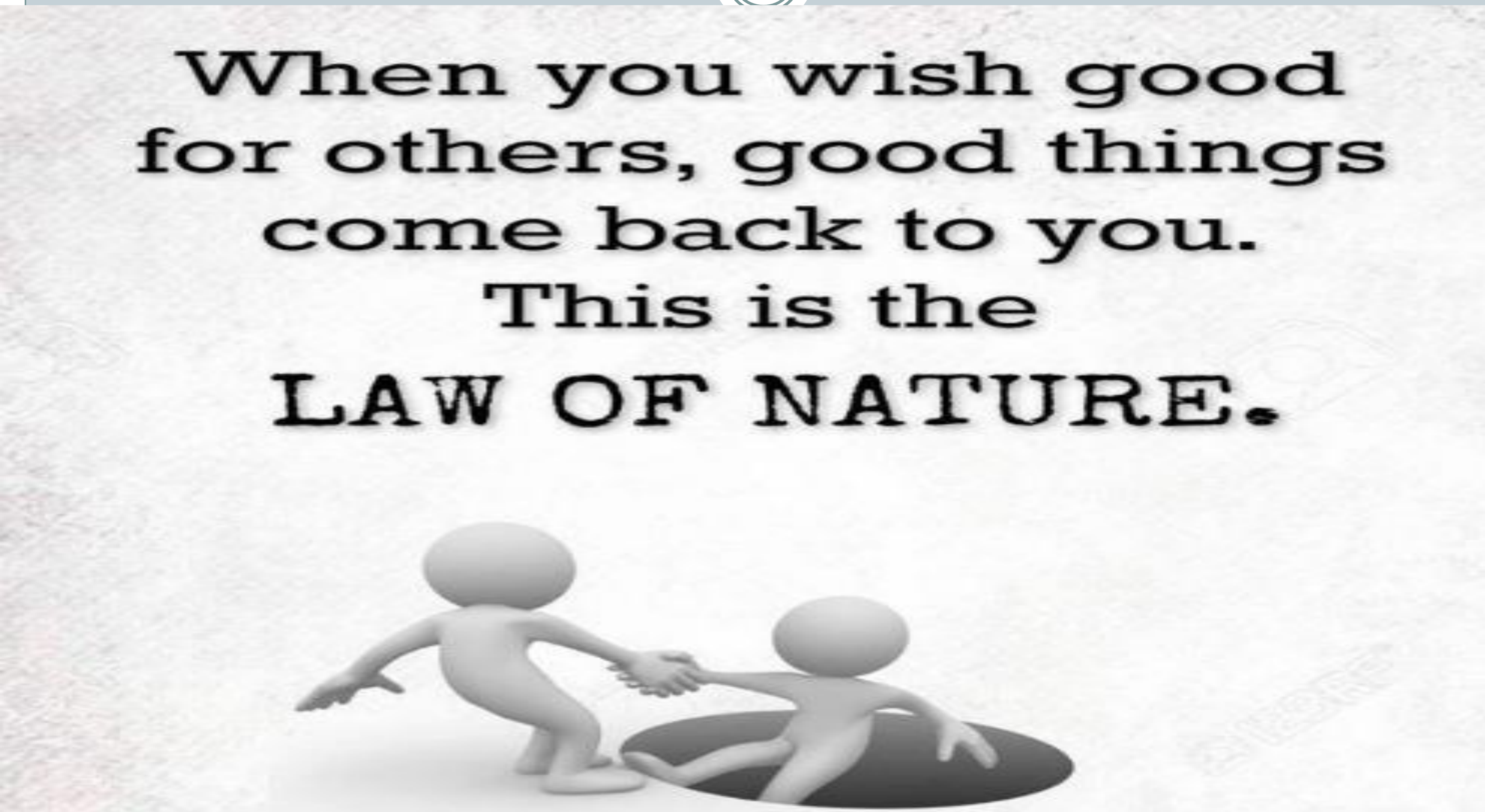

TAKE CARE

LECTURE NO:11




EDUCATION
is not the learning of facts,
but the training of the
mind to think.
-Albert Einstein

MESSAGE OF THE DAY:

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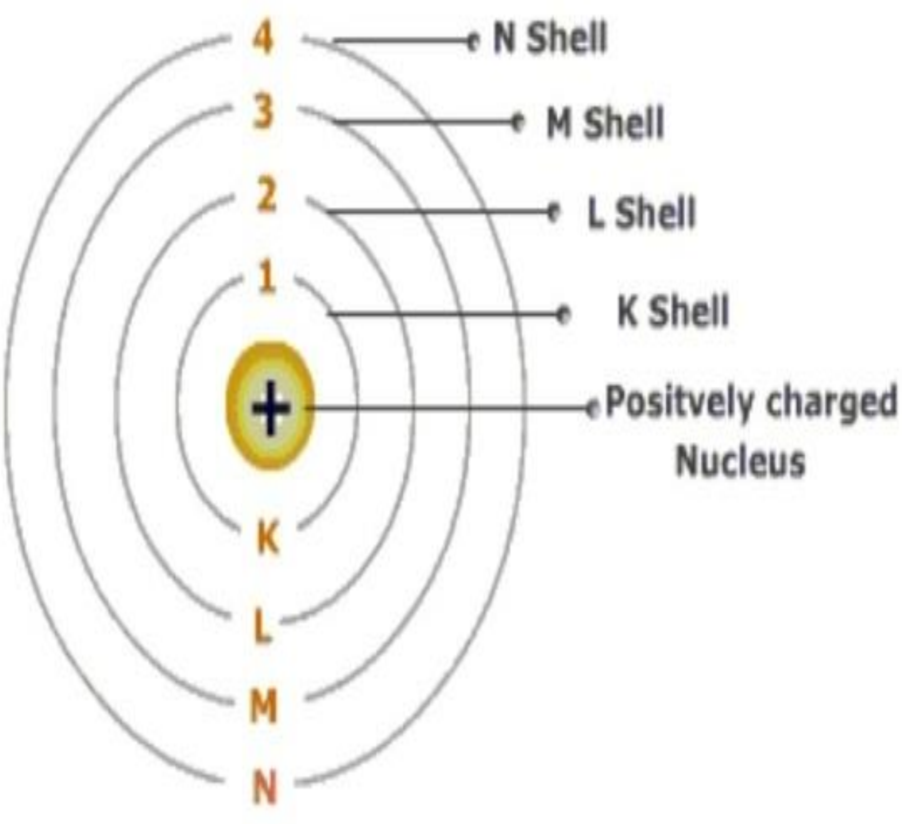
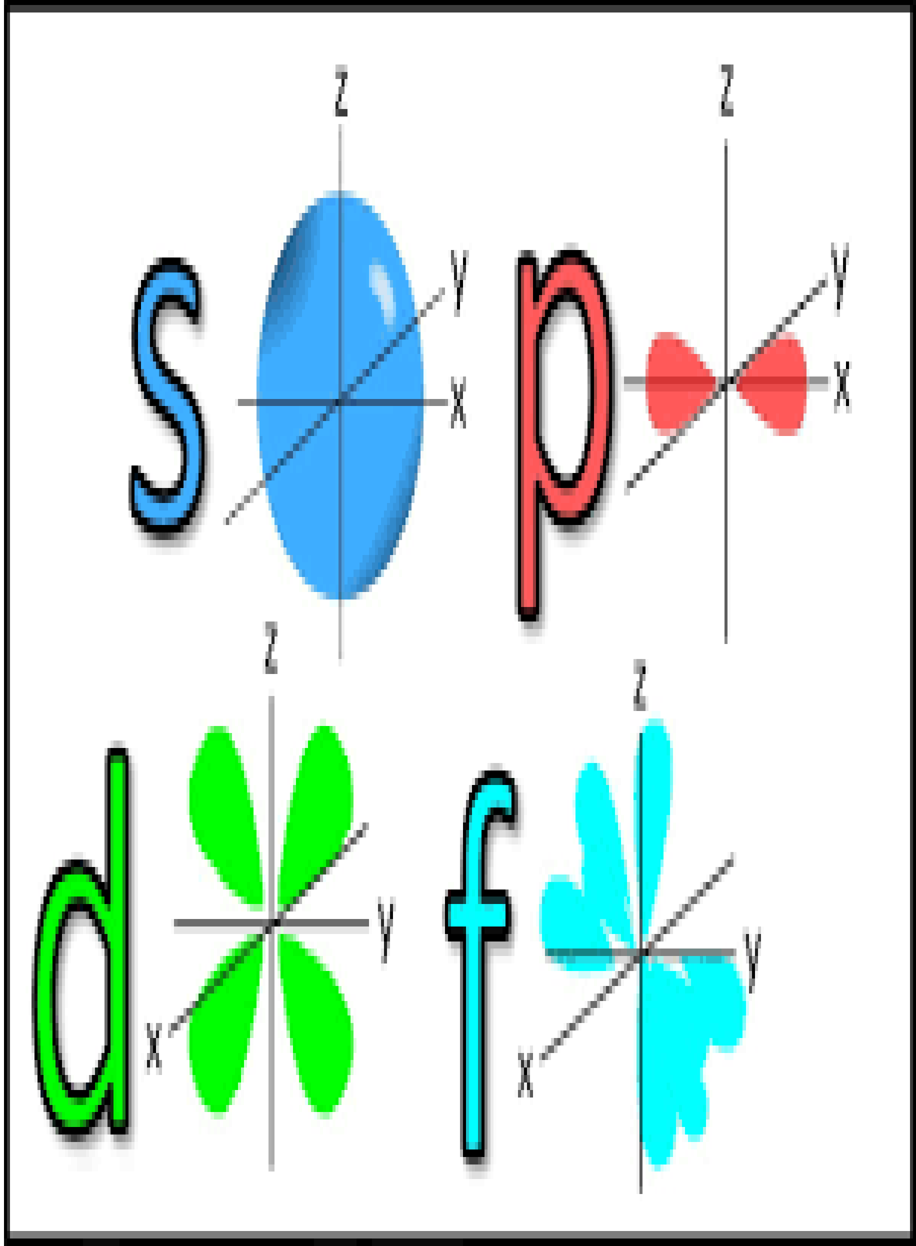
POINTS TO PONDER:

- These orbits are represented by the symbols K, L, M, N etc....or numbers 1,2,3,4 respectively.

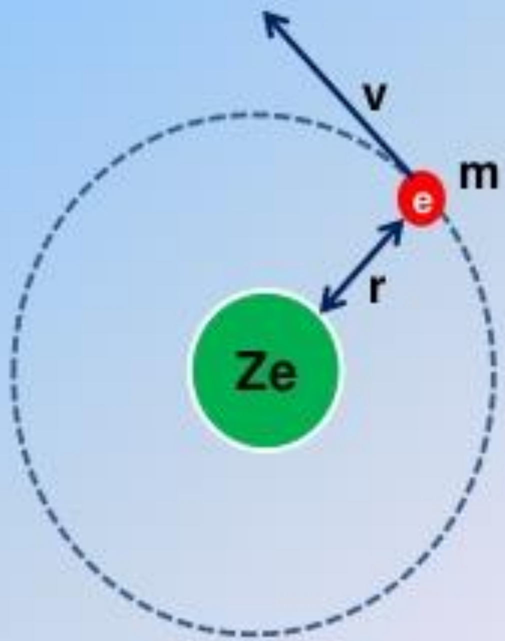




Derivation of Radius of an Orbit of an Atom

Consider an atom having an electron e^- moving around the nucleus having charge Ze where Z is the atomic number.

Let m be the mass, r the radius of the orbit and v, the velocity of the revolving electron.



LESSON OBJECTIVES:11

- BY THE END OF THIS PART OF LESSON,STUDENTS WILL BE ABLE TO:
- DIFFERENTIATE BETWEEN ORBIT AND ORBITAL?
- DERIVATION OF RADIUS OF ORBIT OF AN ATOM ?
- DERIVATION OF ENERGY OF AN ELECTRON OF AN ATOM ?

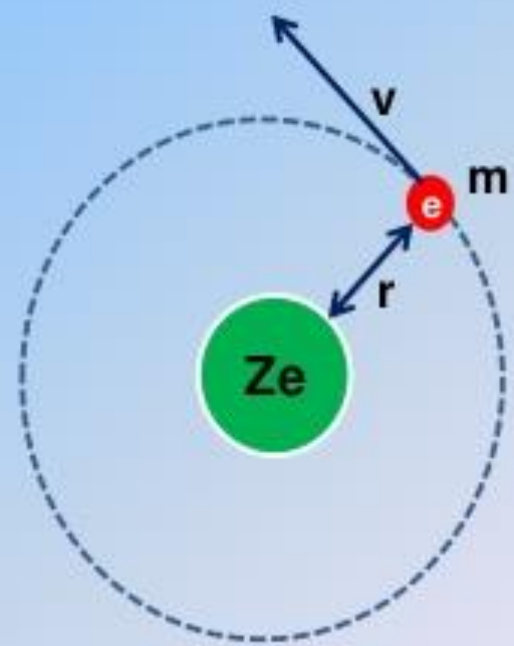
	Orbit	Orbital
1.	An orbit is a well defined circular path around the nucleus in which the electron revolves.	An orbital is the three dimensional space around the nucleus within which the probability of finding an electron is maximum.
2	The maximum number of electrons in any orbit is given by $2n^2$ where n is the number of the orbit.	The maximum number of electrons present in any orbital is two.

Difference Between Orbit and Orbital	
Orbit	Orbital
Orbits represent the planner motion of the electron.	Orbitals represent the three-dimensional motion of the electron around the nucleus.
They are Circular in shape.	They are Different in shapes from each other.
It is a well-defined path that revolving electrons follow around a nucleus.	It is simply a region of space around a nucleus where the electron is present.
The orbits theory doesn't satisfy the Heisenberg uncertainty principle.	Orbitals Concept completely complies with the Heisenberg uncertainty principle.
Can accommodate 2n2electron in the shells where n is the number of orbits.	Only two electrons can be present in an orbital.
the planner motion of electron can be designated by L, M, N, etc.	The 3D motion of electron by s, p, d, f orbitals.

Derivation of Radius of an Orbit of an Atom

Consider an atom having an electron e^- moving around the nucleus having charge Ze where Z is the atomic number.

Let m be the mass, r the radius of the orbit and v, the velocity of the revolving electron.



2.2 BOHR'S ATOMIC MODEL AND ITS APPLICATIONS

Defects of the Rutherford Atomic model

According to the Rutherford's atomic model the structure of atom was like a planet planetry system revolving around the sun. It was defective because according to the classical electromagnetic theory the revolving electron arround nucleus should lose energy continuously. As a result the electron would be accelerated to the nucleus and the radius of orbit become smaller and smaller and ultimately it should fall in to the nucleus. Therefore the atomic structure would collapse and atom will cease to exist. If electron emits energy continuously, it should form continuous spectrum. But actually line spectrum is obtained. Neil Bohr (1913), an English scientist removed these defects and proposed another possible structure of atom called as the Bohr's atomic model. According to this model,

1. Electrons revolve around the nucleus in definite energy levels called orbits or shells.
2. As long as an electron remain in a shell it never gains or losses energy.
3. The gain or loss of energy occurs within orbits only due to jumping of electrons from one energy level to another energy level.
4. Angular momentum (mvr) of an electron is equal to $nh/2\pi$.
   The angular momentum of an orbit depends upon its quantum number and it is an integral multiple of the factor $h/2\pi$

   i.e. $mvr = nh/2\pi$

   Where n = 1, 2, 3, ...

2.2.1 Derivation of Radius of an Orbit of an Atom

Bohr derived expressions for the calculations of radius of nth orbit of an atom of hydrogen or ions like He^{+1}, Li^{+2} etc.

Figure 2.9 Electron revolving in an atom

Let us consider an atom having an electron $\bar{e}$ moving around the nucleus having charge Z e, where Z is the atomic number. Let m be the mass, r the radius of the orbit and v, the velocity of the revolving electron.

According to coulomb's law, the electrostatic force

$$\text{Coulomb's force of attraction} = \frac{Ze \times \bar{e}}{4\pi \epsilon_o r^2} = \frac{Ze^2}{4\pi \epsilon_o r^2}$$

Where ϵ_o is the vacuum permittivity constant with a value $8.84 \times 10^{-12}\ C^2 J^{-1} m^{-1}$.

centrifugal force acting on the moving electron $= \frac{mv^2}{r}$

These two forces are equal and opposite and balance each other. So,

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi \epsilon_o r^2} \quad \text{.............................. (1)}$$

$$mv^2 = \frac{Ze^2}{4\pi \epsilon_o r}$$

$$r = \frac{Ze^2}{4\pi \epsilon_o mv^2} \quad \text{.............................. (2)}$$

Thus we conclude that the radius of a moving electron is inversely proportional to the square of its velocity.

Now we consider angular momentum. According to Neil Bohr,

$$mvr = \frac{nh}{2\pi} \quad \text{.............................. (3)}$$

$$v = \frac{nh}{2\pi mr}$$

Taking square on both sides,

$$v^2 = \frac{n^2h^2}{4\pi^2m^2r^2} \quad \text{.............................. (4)}$$

Putting this value of v^2 in (2),

$$r = \frac{Ze^2}{4\pi \epsilon_o m} \frac{4\pi^2m^2r^2}{n^2h^2}$$

$$\frac{1}{1} = \frac{Ze^2\pi mr}{\epsilon_o n^2h^2}$$

or $\quad Ze^2\pi mr = \epsilon_o n^2h^2$

$$r = \epsilon_o n^2h^2 / Ze^2m\pi \quad \text{.............................. (5)}$$

For hydrogen, $Z = 1$

So, $$r = \frac{\epsilon_o n^2h^2}{\pi me^2} \quad \text{.............................. (6)}$$

or $\quad r = n^2 a^o$

where $a^o = \frac{\epsilon_o h^2}{\pi me^2}$, a constant quantity having a value of

$0.529 \times 10^{-10}m \quad = 0.529A^o = 10^{-10}m)$

so, $\quad r = n^2 \times 0.529$ Å

Therefore radius of orbits having n $= 1, 2,$------ are as follows.

When n = 1, $r = 1^2 \times 0.529$Å $= 0.529$ Å

When n = 2, $r = 2^2 \times 0.529$Å $= 4 \times 0.529$Å $= 2.116$Å

2.2.2 Derivation of Energy of an Orbit

The energy of an electron in an orbit is the sum of its potential energy and kinetic energy.

$\therefore$ $E_{Total} = K.E + P.E.$

$$= \left(\frac{1}{2}mv^2\right) + \left(-\frac{Ze^2}{4\pi\epsilon_o r}\right) \quad \text{.......................... (7)}$$

$$E_{Total} = \frac{1}{2}mv^2 - \frac{Ze^2}{4\pi\epsilon_o r}$$

This potential energy is governed by the coulomb's Law of Electrostatic force.

Putting the value of mv^2 from eq. (1) into eq. (7)

$$E_n = \frac{1}{2}\left(\frac{Ze^2}{4\pi\epsilon_o r}\right) - \frac{Ze^2}{4\pi\epsilon_o r}$$

$$= \frac{Ze^2}{4\pi\epsilon_o r}\left(\frac{1}{2} - \frac{1}{1}\right) = \frac{Ze^2}{4\pi\epsilon_o r}\left(-\frac{1}{2}\right)$$

$$= -\frac{Ze^2}{8\pi\epsilon_o r} \quad \text{.......................... (8)}$$

Now putting the value of r from eq. (5) in to eq. (8)

$$E_n = -\frac{Ze^2}{8\pi\epsilon_o} \times \left[\frac{\pi mZe^2}{\epsilon_o n^2h^2}\right]$$

$$= -\frac{mZ^2e^4}{8\epsilon_o^2 n^2h^2} \quad \text{.......................... (9)}$$

For Hydrogen atom; $z = 1$

$$\therefore \quad E_n = -\frac{me^4}{8\epsilon_o^2 n^2h^2}$$

$$= -\frac{me^4}{8\epsilon_o^2 h^2}\left[\frac{1}{n^2}\right]$$

But $\frac{me^4}{8\epsilon_o^2 h^2} = 2.178 \times 10^{-18}$ J

This value is obtained by putting the values of constants.

$$\therefore \quad E_n = -2.178 \times 10^{-18}\left[\frac{1}{n^2}\right] \text{J} \quad \text{.......................... (10)}$$

$$= -\frac{k}{n^2} \text{ where } k = 2.178 \times 10^{-18} \text{ J}$$

The negative sign indicates decrease in energy of the electron.

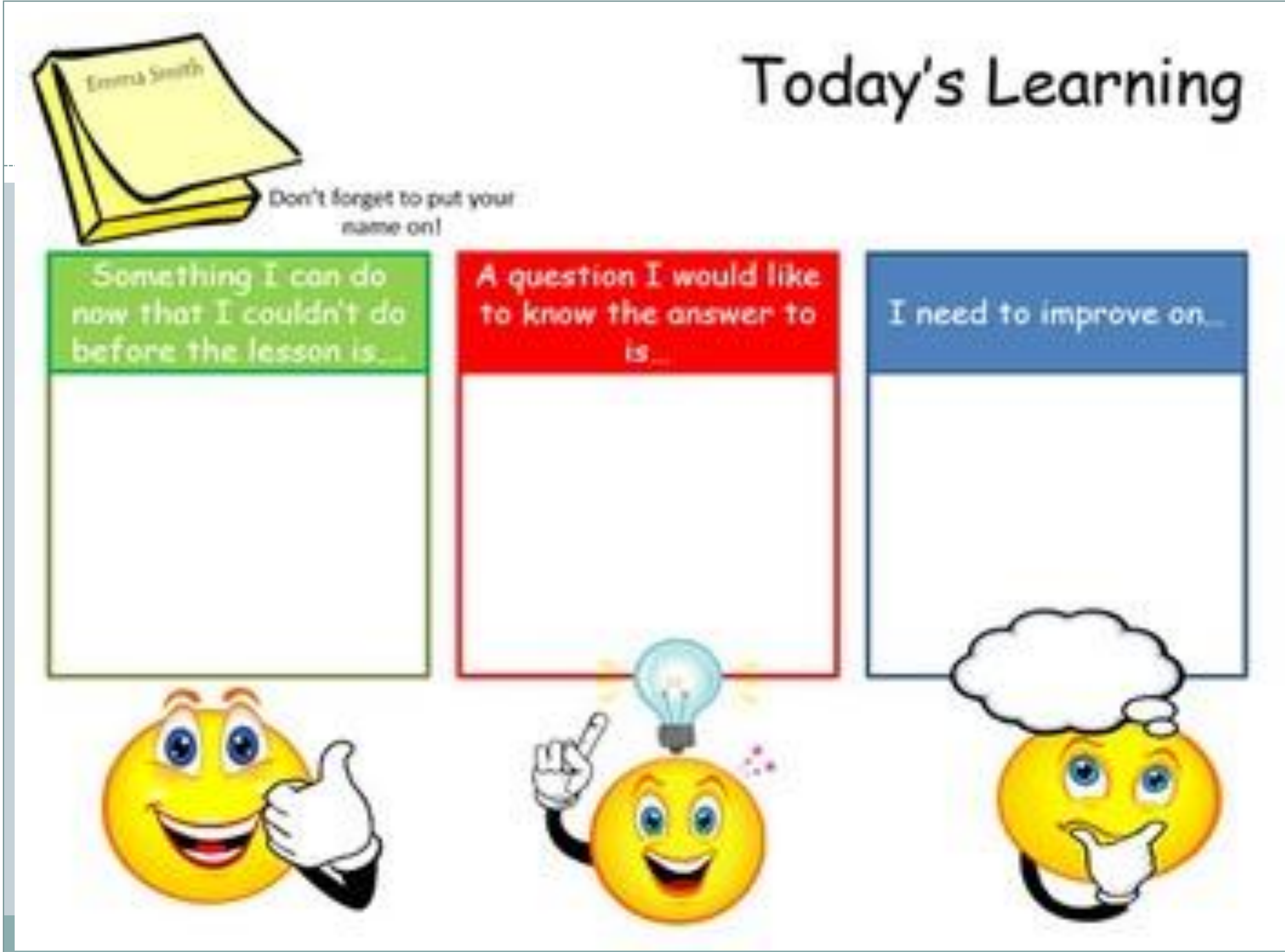
The value of energy obtained is in Joules/atom. If this quantity is multiplied by Avogad No. and divided by 1000, the value of E_n becomes.

$$E_n = -\frac{1313.35}{n^2} \text{ kJ / mole.}$$

This energy is associated with 1.008 gram-atoms of hydrogen.

If n = 1, 2, 3 4, 5......

then $E_1 = -\frac{1313.35}{1^2} = -1331.35\text{kJ mole}^{-1}$


Emma Smith
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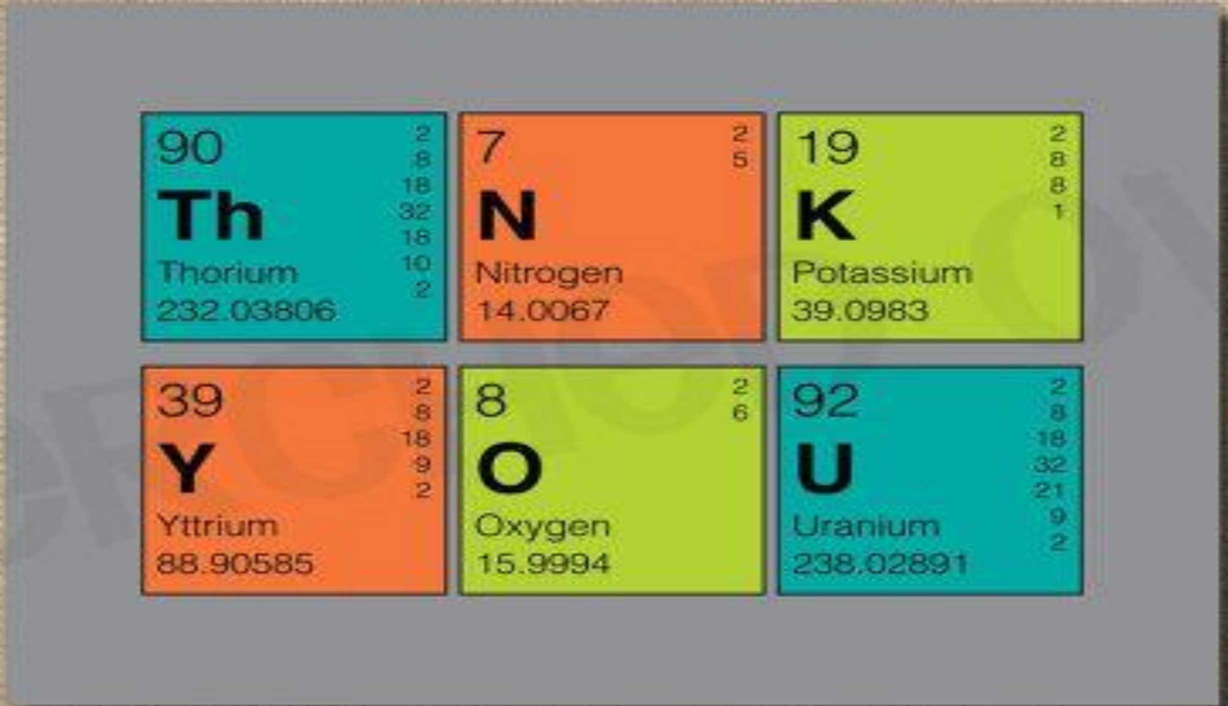
Don't forget to put your name on!
Something I can do now that I couldn't do before the lesson is...
A question I would like to know the answer to is...
I need to improve on...

HOME-WORK:

- NOTE:ATTEMPT ANY ONE QUESTION:
- Do derivation of radius of orbit of an atom?
- Do derivation of energy of an electron of an atom?
- Differentiate between orbit and orbital?

LESSON CLOSURE:


good bye
good bye
SCHOOL BUS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Thorium
232.03806
2 8 18 32 18 10 2
7
N
Nitrogen
14.0067
2 5
19
K
Potassium
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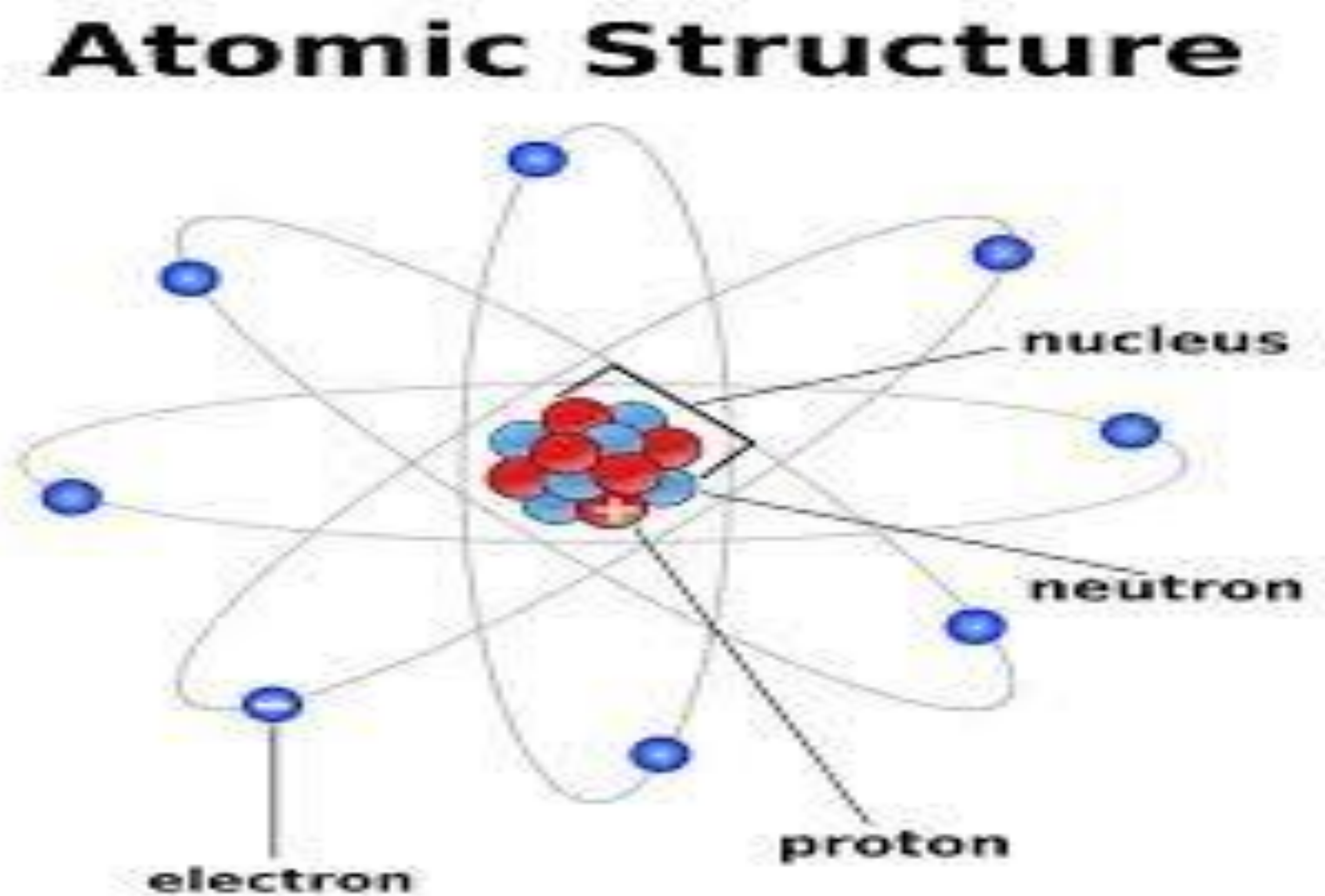
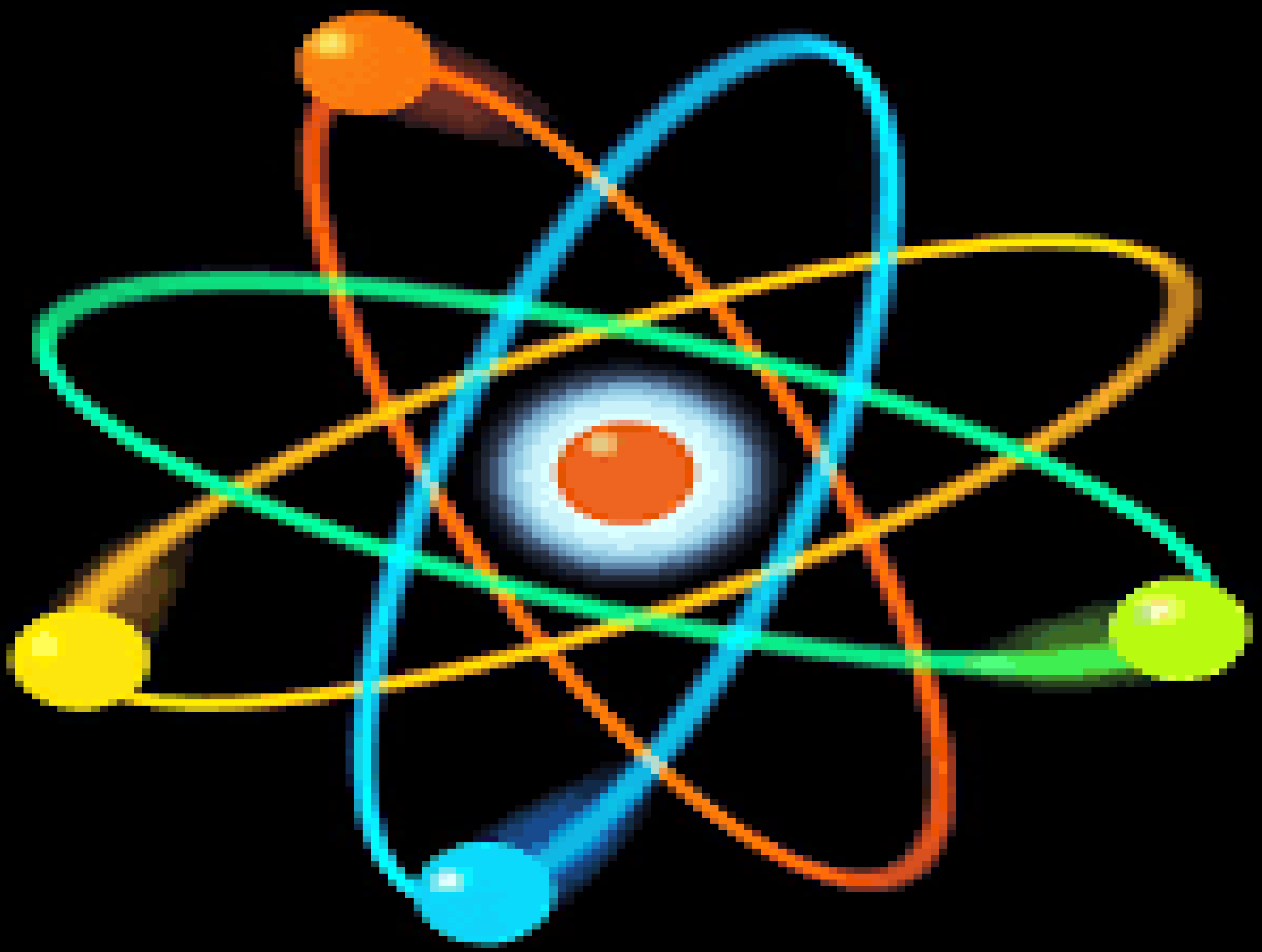
Y
Yttrium
88.90585
2 8 18 9 2
8
O
Oxygen
15.9994
2 6
92
U
Uranium
238.02891
2 8 18 32 21 9 2


Thank
You

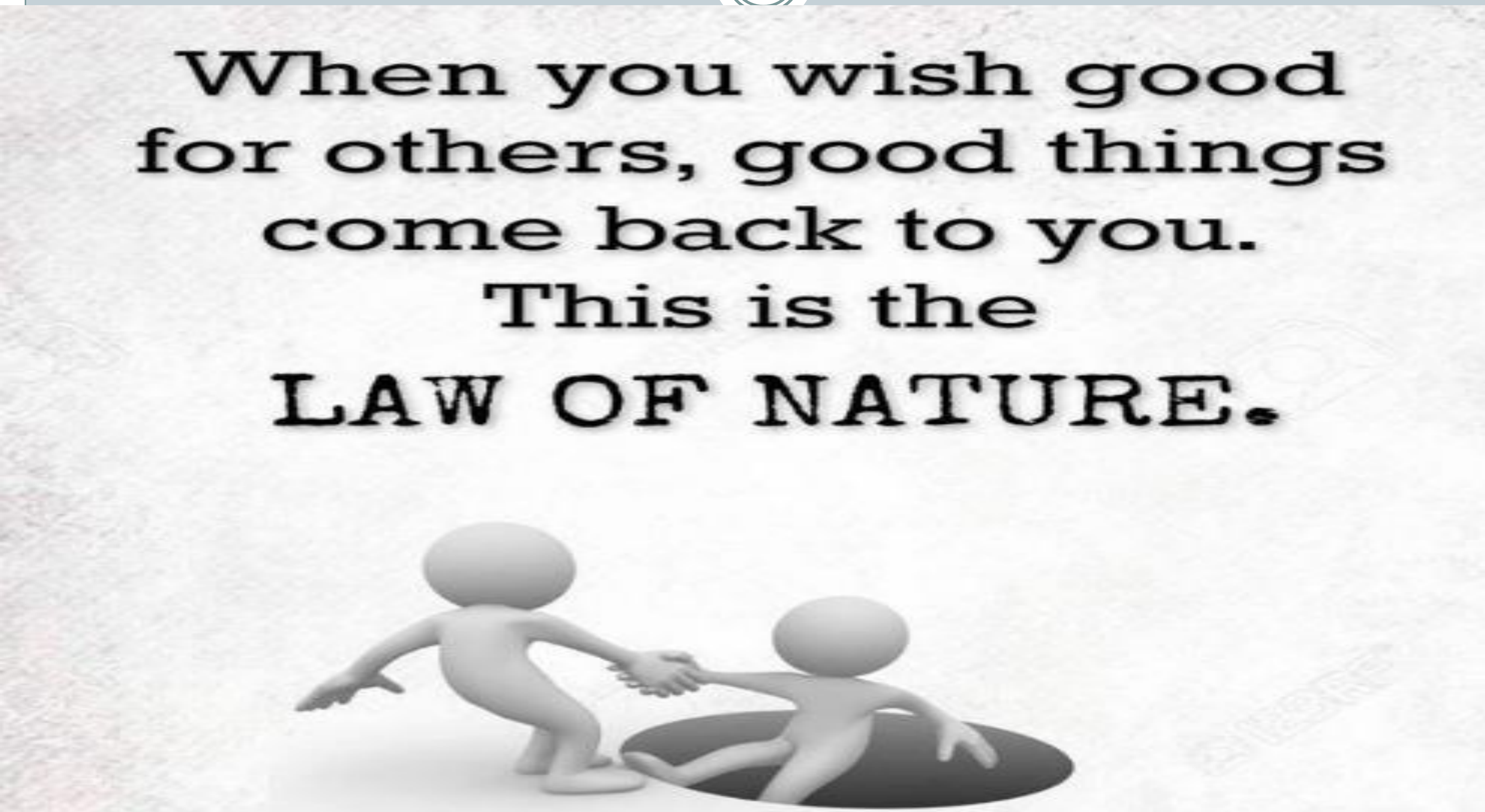

TAKE CARE

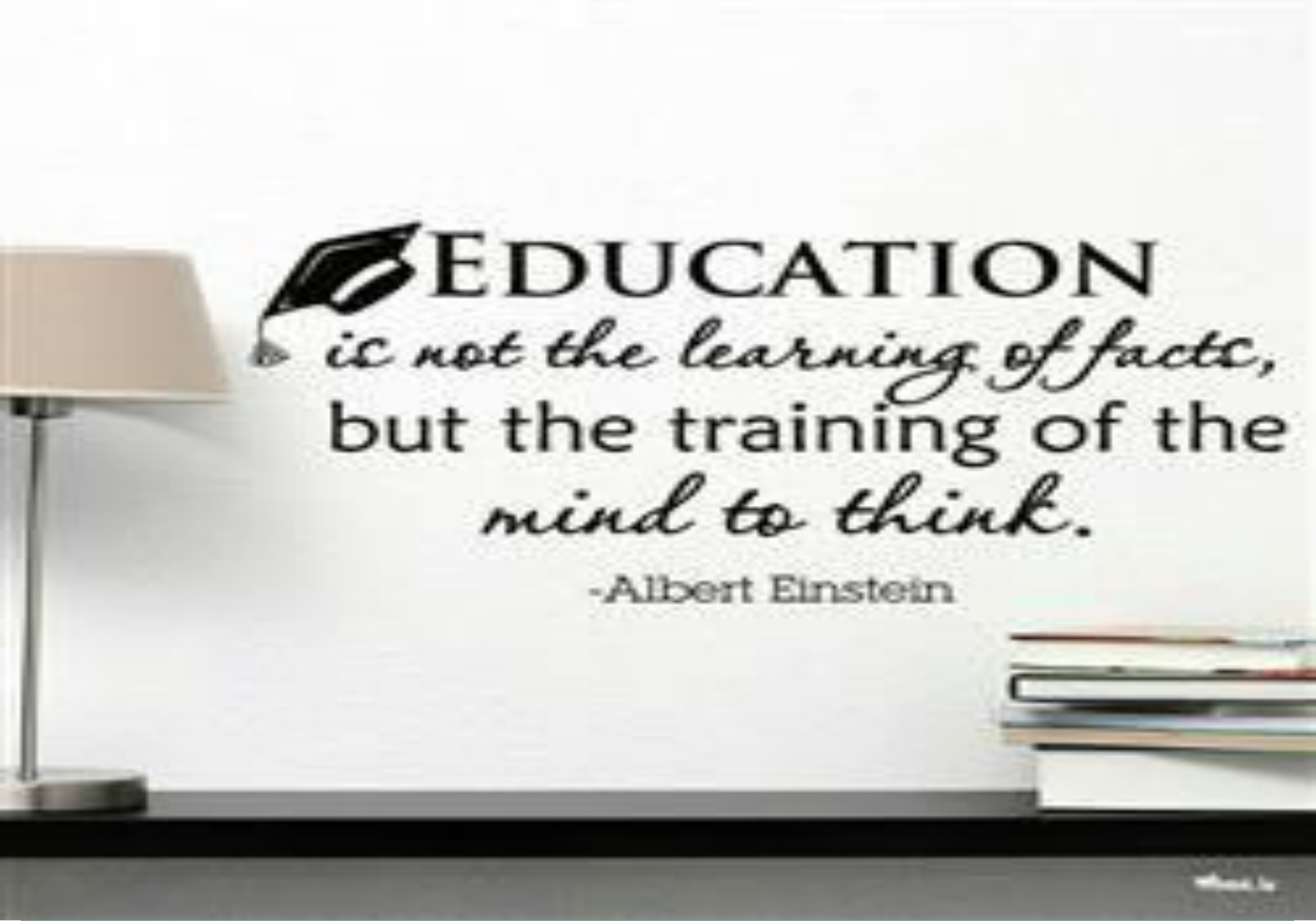
LECTURE NO:12





MESSAGE OF THE DAY:




EDUCATION
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HAHA! SO WE MEET AGAIN!!


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LESSON OBJECTIVES:12

- BY THE END OF THIS PART OF LESSON,STUDENTS WILL BE ABLE TO:
- DERIVATION OF ENERGY DIFFERENCE OF AN ELECTRON OF AN ATOM ?
- DERIVATION OF WAVE NUMBER OF AN ELECTRON OF AN ATOM?
- DEFINE HYDROGEN SPECTRUM ?HOW IT WAS DISCOVERED ?

$$E_2 = -\frac{1313.35}{2^2} = -328.32 \text{ kJ mole}^{-1}$$

$$E_3 = -\frac{1313.35}{3^2} = -145.92 \text{ kJ mole}^{-1}$$

$$E_4 = -\frac{1313.35}{4^2} = -82.08 \text{ kJ mole}^{-1}$$

$$E_5 = -\frac{1313.35}{5^2} = -52.53 \text{ kJ mole}^{-1}$$

The first energy level when n = 1 is known as the ground state of the hydrogen atom. All other energy levels are known as excited states.

According to Eq. (9) $\qquad E = -\dfrac{mz^2e^4}{8\epsilon_o^2 n^2h^2}$

Let E_1 be the energy of the orbit n_1 and E_2 that of n_2,

Then $\qquad \Delta E = E_2 - E_1$

$$= \frac{mz^2e^4}{8\epsilon_o^2 n_2^2h^2} - \frac{mz^2e^4}{8\epsilon_o^2 n_1^2h^2}$$

$$= \frac{mz^2e^4}{8\epsilon_o^2 n_1^2h^2} - \frac{mz^2e^4}{8\epsilon_o^2 n_2^2h^2}$$

$$= \frac{mz^2e^4}{8\epsilon_o^2 h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

For Hydrogen, Z = 1

$$\therefore \Delta E = \frac{me^4}{8\epsilon_o^2 h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{.............. (11)}$$

Here, $\qquad \dfrac{me^4}{8\epsilon_o^2 h^2} = 2.18 \times 10^{-18}$ J.

$$\therefore \qquad \Delta E = 2.18 \times 10^{-18} \text{ J}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{.............. (12)}$$

but according to Plank's Quantum theory,

$$\Delta E = h\nu$$

$$\therefore \qquad h\nu = 2.18 \times 10^{-18} \text{ J} \times \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right].$$

again $\quad h\nu \quad = \dfrac{me^4}{8\epsilon_o^2 h^2}\left[\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right]$

or $\quad \nu \quad = \dfrac{me^4}{8\epsilon_o^2 h^3}\left[\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right]$ Hz or cycles sec^{-1} $\qquad$ .... (13)

2.2.3 Derivation of Wave Number ($\bar{\upsilon}$)

The relationship between frequency (ν) and wave number ($\bar{\upsilon}$) is

$$\nu = \bar{\upsilon} c \qquad \text{.........................} (14)$$

Where c is the velocity of light.

Putting the value of ν from eq. (13) in eq. (14)

$$\bar{\upsilon} c = \frac{Z^2 m e^4}{8 \epsilon_0^2 h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

or

$$\bar{\upsilon} = \frac{Z^2 m e^4}{8 \epsilon_0^2 h^3 c}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{.........................} (15)$$

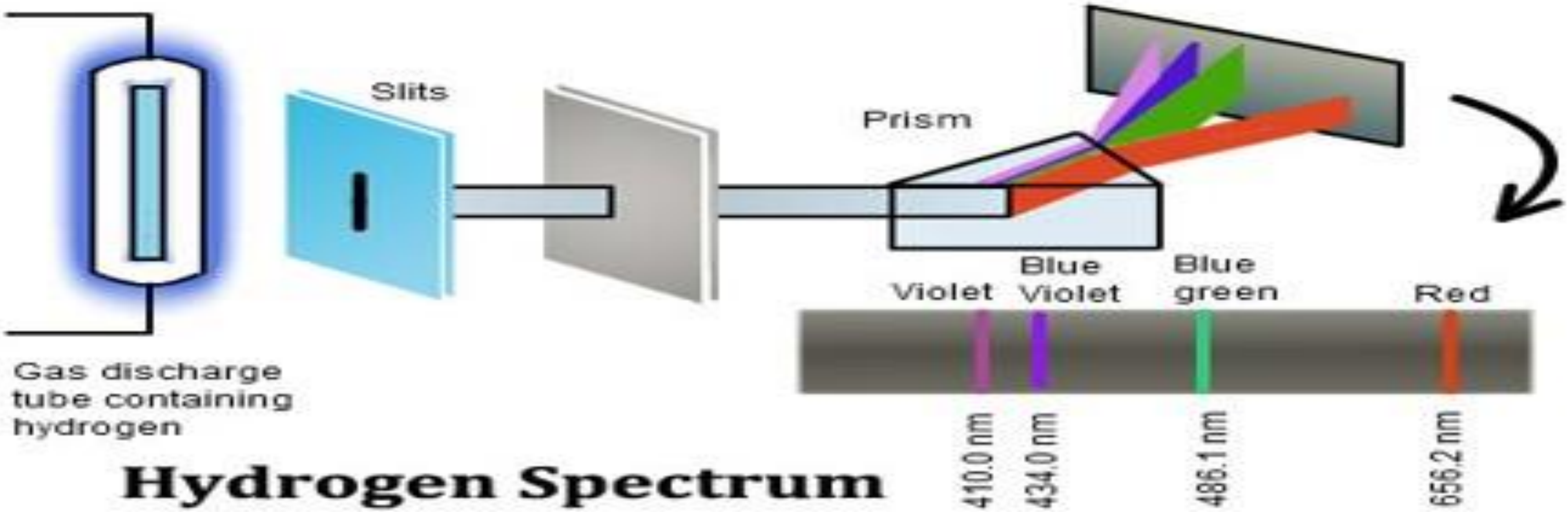
Putting values of constants, $\dfrac{me^4}{8 \epsilon_0^2 h^3 c} = R = 1.09678 \times 10^7 \text{ m}^{-1}$,

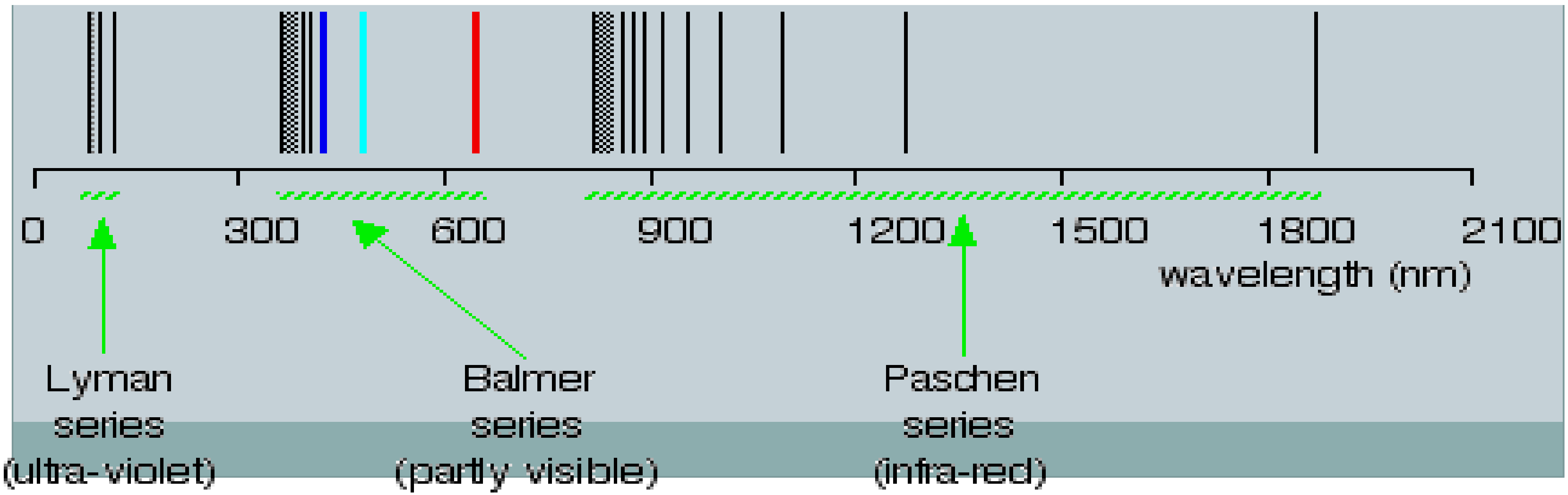
R called Rydberg's constant?

$$\bar{\upsilon} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \qquad \text{.........................} (16)$$

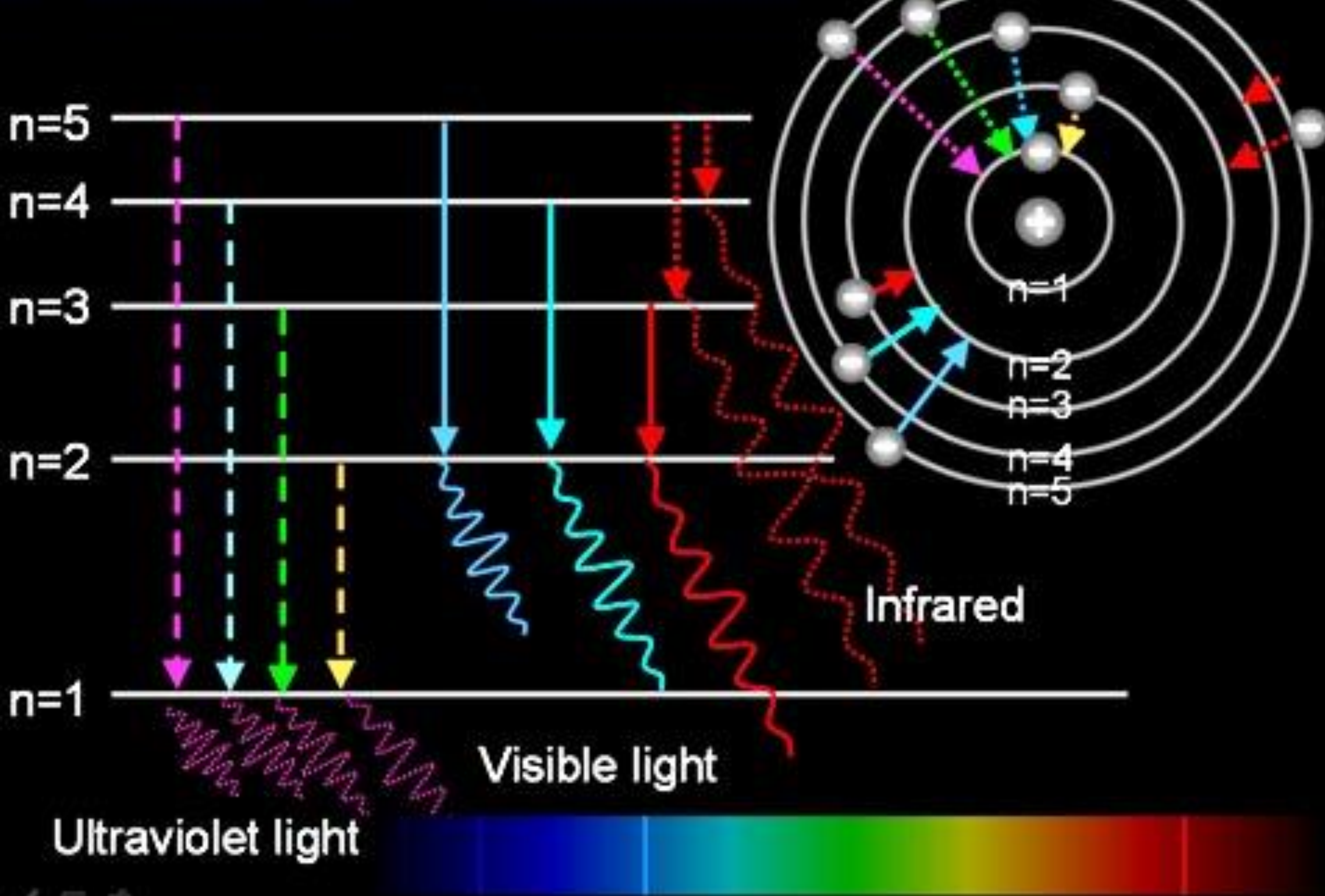
$$\therefore \bar{\upsilon} = 1.09678 \times 10^7\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \text{m}^{-1}$$

POINTS TO PONDER:





SPECTRA OF HYDROGEN



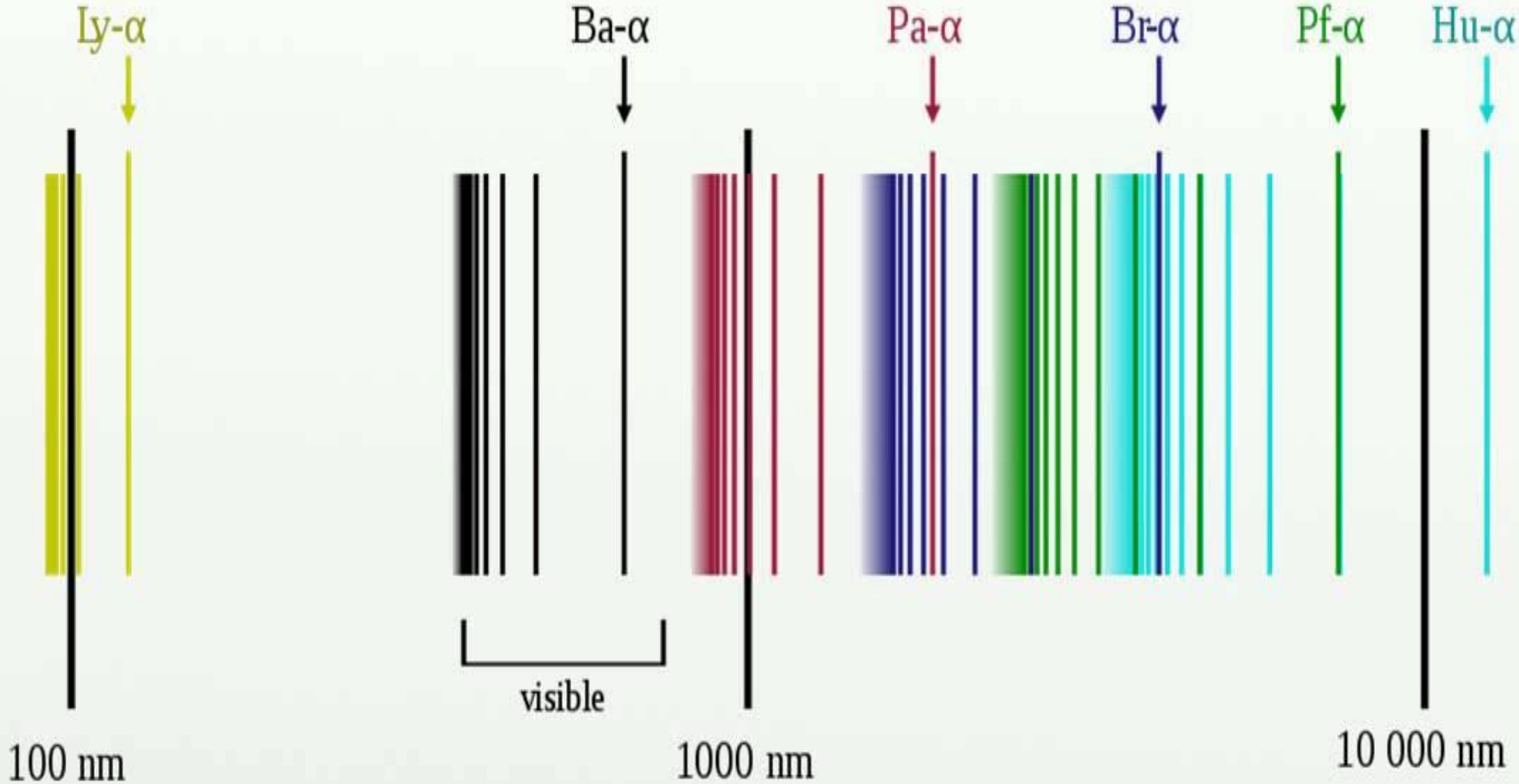
Bohr's Explanation for Hydrogen Spectrum

When current is passed through Hydrogen gas in the discharge tube at low pressure, the molecules of Hydrogen break in to atoms.

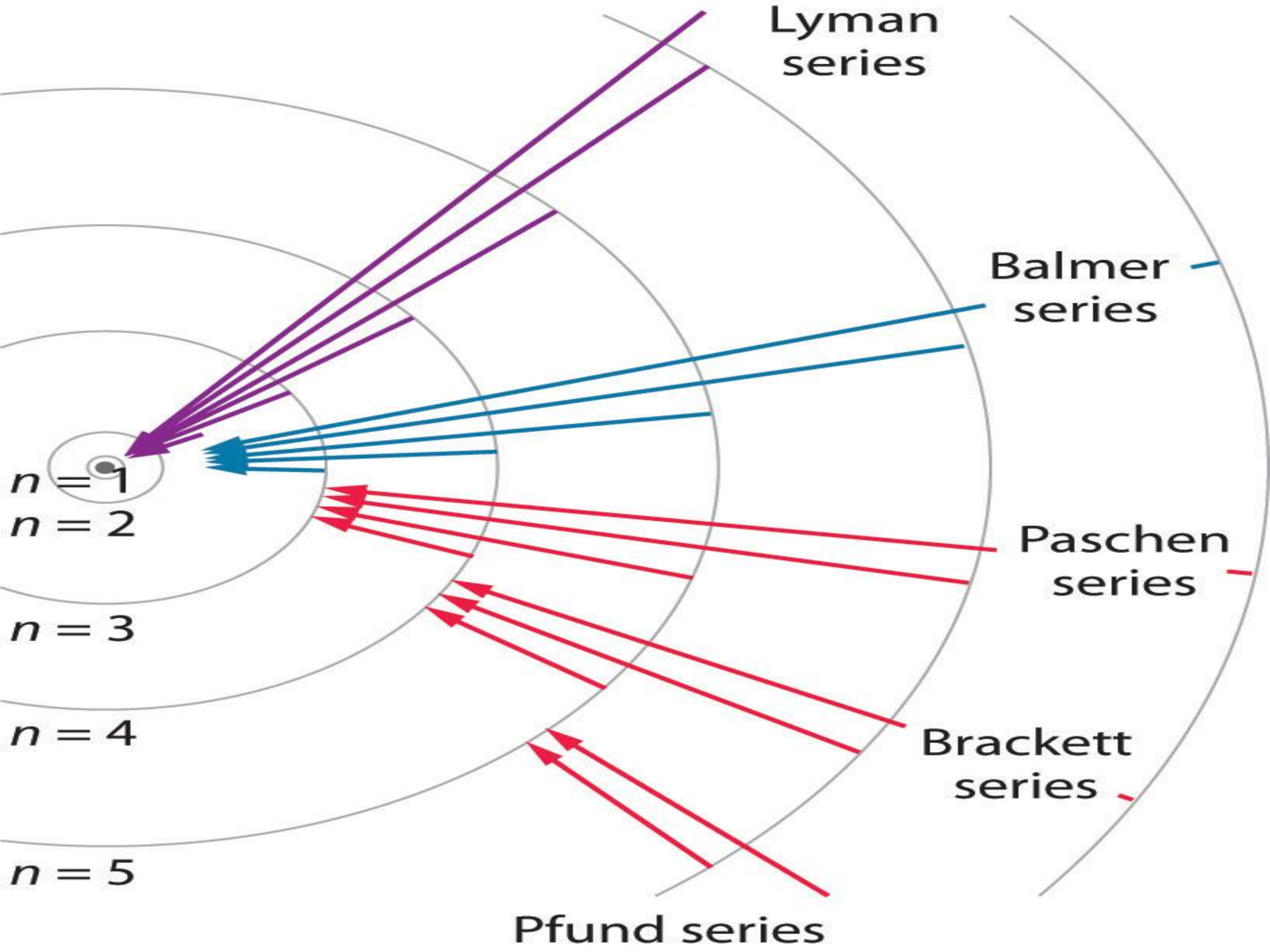
$$\underset{\text{(gas)}}{H_2} \xrightarrow[\text{Low Pressure}]{\text{Electrical Discharge}} \underset{\text{(atoms)}}{2\,H}$$

These atoms absorbs energy from electric spark.

Hydrogen spectral series



https://en.wikipedia.org/wiki/File:Hydrogen_spectrum.svg

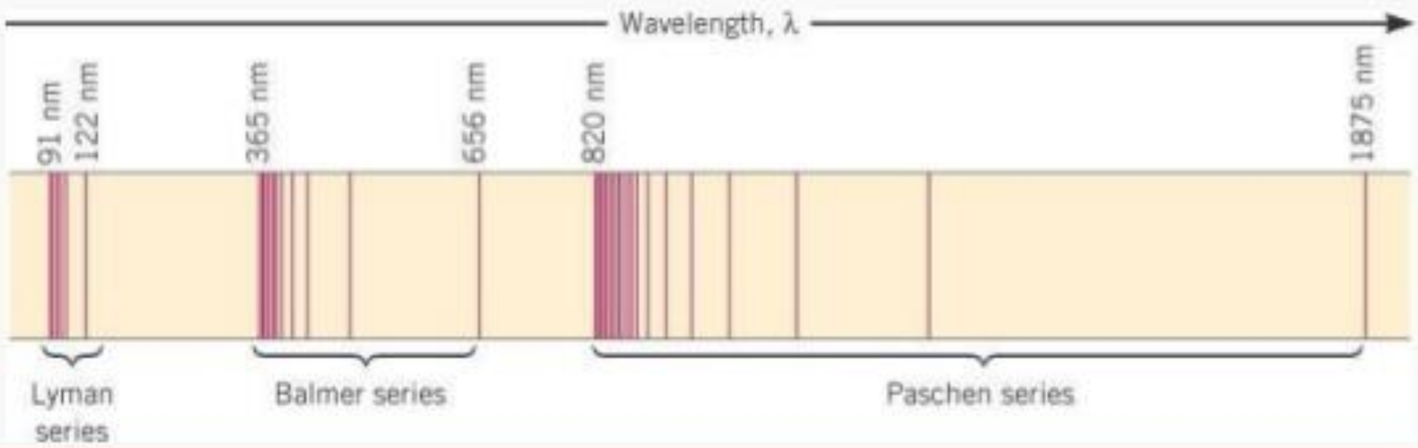

Lyman series
Balmer series
Paschen series
Brackett series
Pfund series
$n = 1$
$n = 2$
$n = 3$
$n = 4$
$n = 5$

The Line Spectrum of Hydrogen

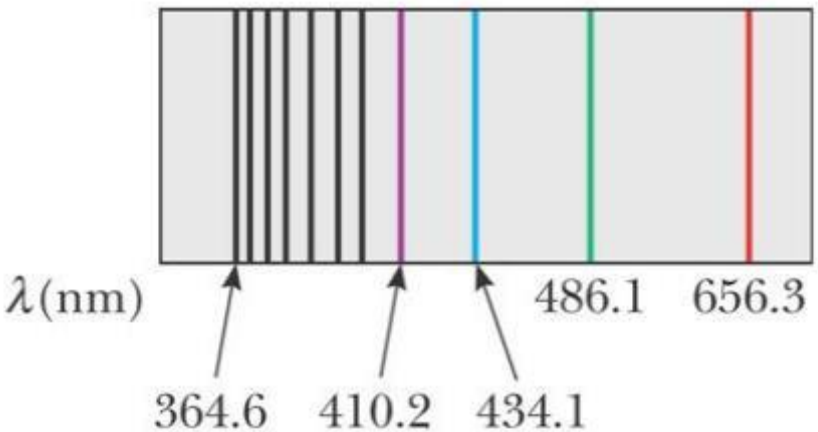
Lyman series $$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right) \qquad n = 2, 3, 4, \ldots$$

Balmer series $$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad n = 3, 4, 5, \ldots$$

Paschen series $$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right) \qquad n = 4, 5, 6, \ldots$$



Balmer series of hydrogen



The Balmer series of spectral lines for atomic hydrogen.

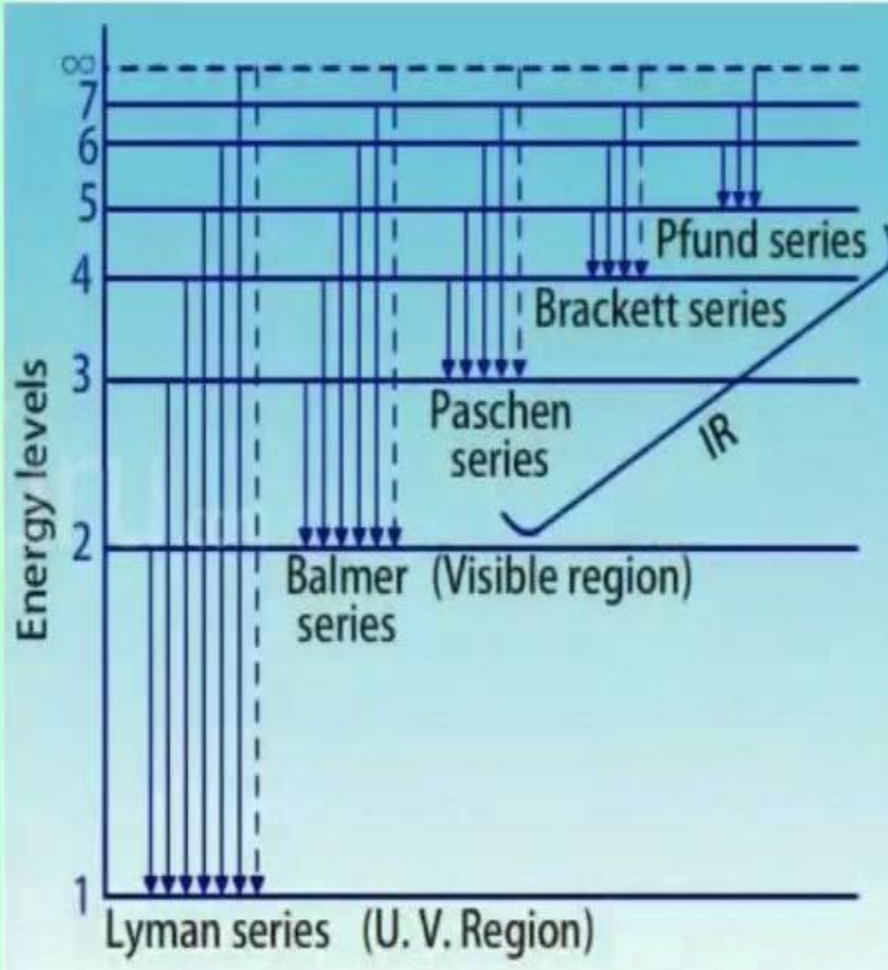
- Johann Jocob Balmer (1825-1898)
- The empirical equation by Johannes Rydberg (1854-1919):

$$\frac{1}{\lambda} = R_H \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \qquad n = 3,4,5,\ldots$$

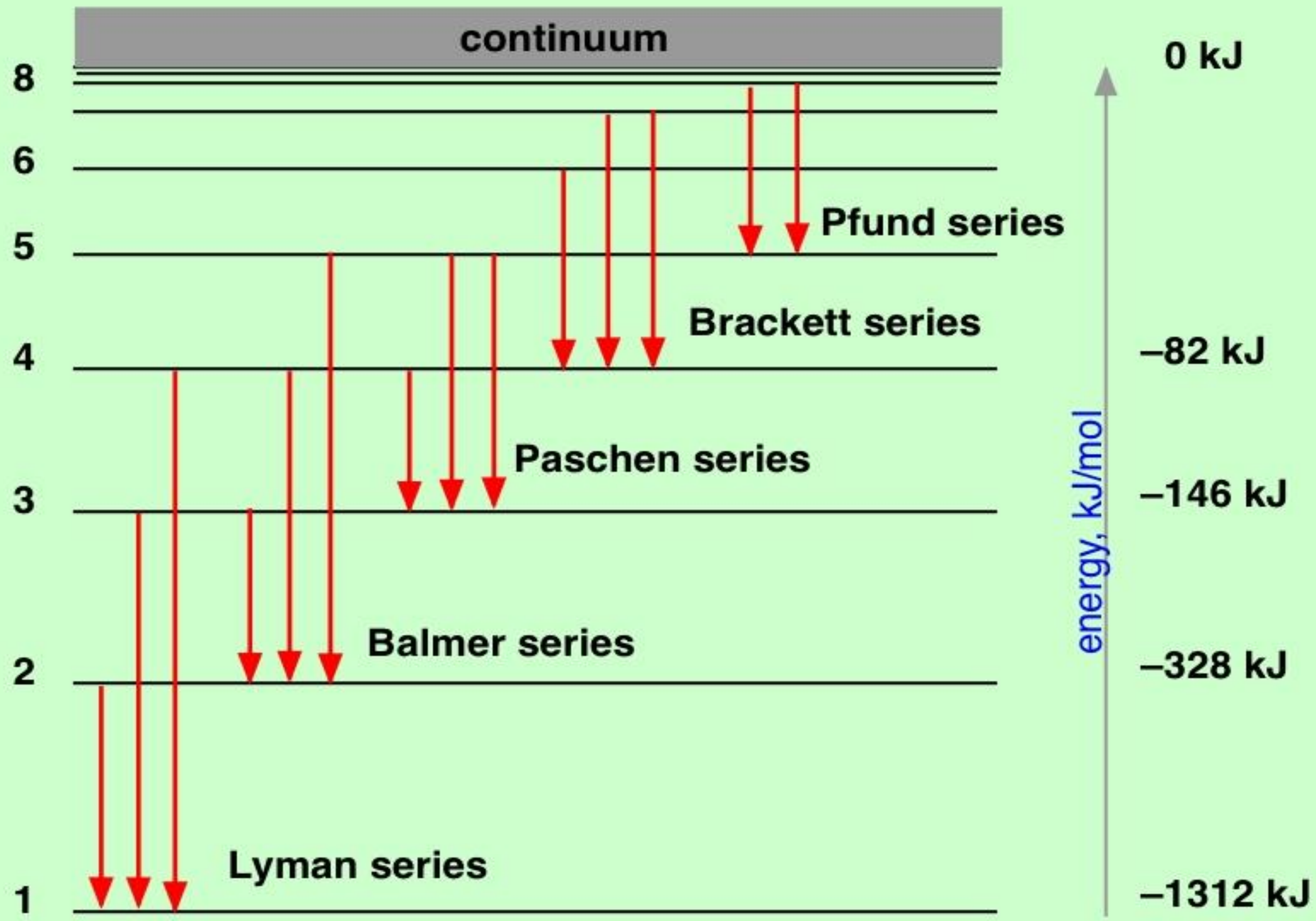
  - R_H: Rydberg constant = 1.0973732 x 10^7 m^{-1}.
  - The series limit
  - The measured spectral lines agree with the empirical equation to within 0.1%.

Explanation of Line Spectrum of Hydrogen

The lines which arise due to the transitions from higher energy levels to first energy level are grouped as Lyman Series. Similarly, the lines obtained as a result of transitions of electrons from higher energy levels to second, third, fourth and fifth energy levels give rise to Balmer, Paschen, Brackett and Pfund series respectively.

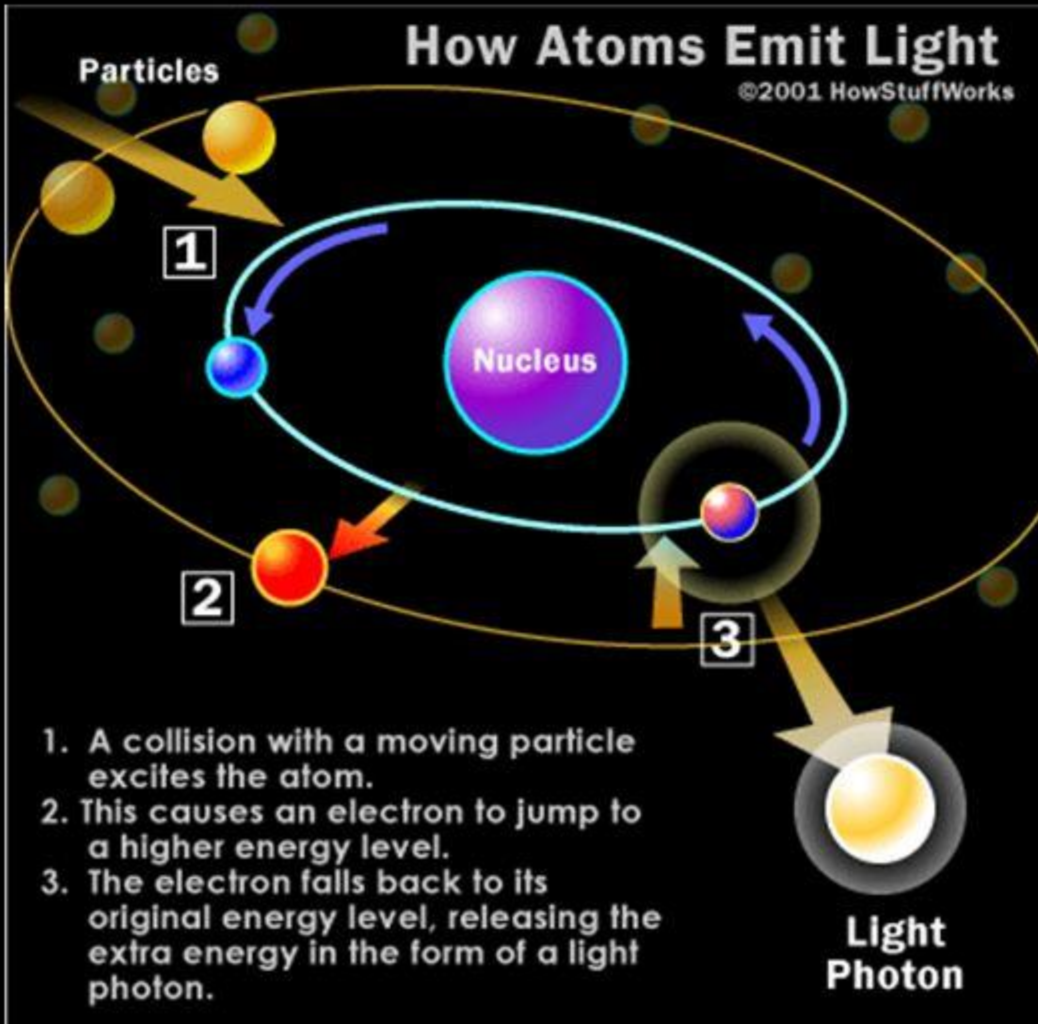


Generation of various spectral series in hydrogen spectrum



continuum
0 kJ
8
6
5
Pfund series
Brackett series
4
–82 kJ
Paschen series
3
–146 kJ
energy, kJ/mol
Balmer series
2
–328 kJ
Lyman series
1
–1312 kJ

Photon Model



Light is made of particles or bundles of energy that travel in a beam or stream.

2.5 Hydrogen Spectrum

When Hydrogen is enclosed in a container and heated, it emits radiation. These radiation e actually emitted due to excitation and de-excitation of electron of hydrogen.

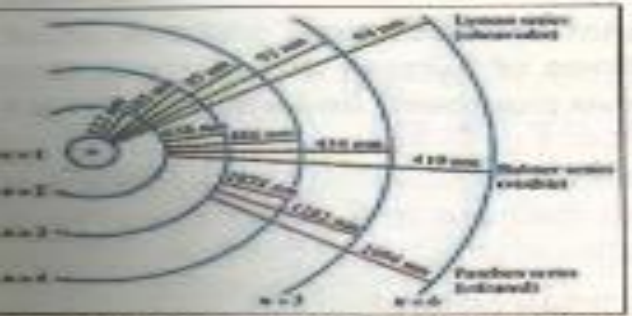
According to equation (16).

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

With this equation, Bohr was able to predict the wave length in the hydrogen emission ectrum and the electron transition (changes of energy levels) that occur in hydrogen atom.

The wave no of different spectral lines can be calculated corresponding to the values of , and n_2. In the hydrogen spectrum, different series of lines have been identified for n_1 and n_2 alues. These series are,

yman series	n_1	=	1	n_2 = 2,3,4,5 ...
almer series	n_1	=	2	n_2 = 3,4,5,6 ...
aschen series	n_1	=	3	n_2 = 4,5,6,7 ...
rackett series	n_1	=	4	n_2 = 5,6,7,8 ...
fund series	n_1	=	5	n_2 = 6,7,8,9 ...



ig. 2.10: Spectrum of hydrogen atom showing electronic transition to give ifferent series

Only the Balmer series was observed in the visible part of the spectrum. Lyman series lie in the in the ultraviolet region while the Paschen and Bracket series have been observed in the infrared region.

Origin of Hydrogen Spectrum on the Basis of Bohr's Model.

The first spectral lines were discovered in 1887 by Lymen and Balmer. No satisfactory reason became available till 1913. Neil Bohr presented his explanation of line spectra in 1913.

According to Bohr when current is passed through the ydrogen gas in the discharge tube at low pressure, the molecules of hydrogen break in to toms. These atoms absorb energy from the electric spark. The electrons of hydrogen atoms re excited to high energy levels .The higher energy orbits to which the electron migrate depend pon the amount of energy absorbed by the electron. Above diagram shows the possibilities of novement of electron from lower to higher levels.

These excited electrons being unstable come back to one of the lower energy levels. The lectrons may come to the lowest energy levels. In this way they emit energy, they had absorbed. ymen series is produced when the electrons jump from n=2,3,4,5,6...etc to n=1.In balmer series he electrons from n=3,4,5,6....come back to n=2.

Let us caculate the various lines of Lymen series, Balmer series, Paschen series, Brackett series and Pfund series from Bohr's equation of wave number.

$$\bar{\nu} = 1.09678\times10^7 [\frac{1}{n_1^2} - \frac{1}{n_2^2}] m^{-1}$$

Lymen Series

The various lines in Lymen series got their explanation by considering that the electronos of hydrogen atom fall back to n=1 from higher levels. The higer levels occupied by the electrons due to the electric spark.

First line: $n_1 = 1$ $n_2 = 2$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{1^2} - \frac{1}{2^2}) m^{-1} = 82.26\times10^5 m^{-1}$$

Second line: $n_1 = 1$ $n_2 = 3$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{1^2} - \frac{1}{3^2}) m^{-1} = 97.60\times10^5 m^{-1}$$

Third line: $n_1 = 1$ $n_2 = 4$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{1^2} - \frac{1}{4^2}) m^{-1} = 102.70\times10^5 m^{-1}$$

Limiting line: $n_1 = 1$ $n_2 = \infty$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{1^2} - \frac{1}{\infty^2}) m^{-1} = 109.678\times10^5 m^{-1}$$

This limiting line shows that the energy difference beween the first level and the infinity level is the ionozation energy of the hydrogen atom. All these lines of Lymen series have close values. They appear in the form of a group. These values of wave numbers lie in the UV region of the spectrum.

Balmer series:

In this series the electrons fall back to n=2.

First line (Hα line): $n_1 = 2$ $n_2 = 3$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{2^2} - \frac{1}{3^2}) m^{-1} = 15.234\times10^5 m^{-1}$$

Second line (Hβ line): $n_1 = 2$ $n_2 = 4$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{2^2} - \frac{1}{4^2}) m^{-1} = 20.566\times10^5 m^{-1}$$

Third line (Hγ line): $n_1 = 2$ $n_2 = 5$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{2^2} - \frac{1}{5^2}) m^{-1} = 23.05\times10^5 m^{-1}$$

Limiting line: $n_1 = 2$ $n_2 = \infty$

$$\bar{\nu} = 1.09678\times10^7 (\frac{1}{2^2} - \frac{1}{\infty^2}) m^{-1} = 27.421\times10^5 m^{-1}$$

All these lines of balmer series are very close to each other and appear in the form of group of lines. These lines lie in the visible region of the spectrum.

Paschen Series:

The electrons from higher levels fall back to n = 3.

First line: $n_1 = 3$ $n_2 = 4$

$\bar{\nu} = 1.09687 \times 10^7 (\frac{1}{3^2} - \frac{1}{4^2}) m^{-1} = 5.3310 \times 10^5 m^{-1}$

Second line: $n_1 = 3$ $n_2 = 5$

$\bar{\nu} = 1.09678 \times 10^7 (\frac{1}{3^2} - \frac{1}{5^2}) m^{-1} = 7.799 \times 10^5 m^{-1}$

Limiting line: $n_1 = 3$ $n_2 = \infty$

$\bar{\nu} = 1.09678 \times 10^7 (\frac{1}{3^2} - \frac{1}{\infty^2}) m^{-1} = 12.187 \times 10^5 m^{-1}$

These are the again the groups of lines close to each other and appear in IR region.

Brackett series:

The electrons from higher levels fall back to n = 4.

First line:	$n_1 = 4$	$n_2 = 5$	$\bar{\nu} = 2.45 \times 10^5 m^{-1}$
Second line:	$n_1 = 4$	$n2 = 6$	$\bar{\nu} = 3.808 \times 10^5 m^{-1}$
Limiting line:	$n_1 = 4$	$n_2 = \infty$	$\bar{\nu} = 6.855 \times 10^5 m^{-1}$

Pfund series:

The electrons from higer energy levels fall back to n = 5.

First line:	$n_1 = 5$	$n_2 = 6$	$\bar{\nu} = 1.340 \times 10^5 m^{-1}$
Second line:	$n_1 = 5$	$n_2 = 7$	$\bar{\nu} = 2.148 \times 10^5 m^{-1}$
Limiting line:	$n_1 = 5$	$n_2 = \infty$	$\bar{\nu} = 4.387 \times 10^5 m^{-1}$

2.3 PLANK'S QUANTUM THEORY

Max Plank (1900) proposed a theory about nature of light. According to this theory,

(a) Energy is not emitted or absorbed continuously. It is emited or absorbed in the form of wave packets or quanta. In case of light the quantum of energy is often called photon.

(b) The amount of energy associated with quantum of radiation is directly proportional to the frequency (υ) of radiation. i.e.

$E \propto \nu$ or $E = h\upsilon$ .... (1)

where h = Plank's constant and has a value of 6.625×10^{-34} Joules sec.

(c) A body can emit or absorb energy only in terms of integral multiple of a quantum.

$E = nh\upsilon$ where n = 1, 2, 3

Now $\upsilon \propto 1/\lambda$

or $\upsilon = c/\lambda$

Where "λ" is wave length and "c" is the velocity of light, a constant quantity.

Putting the value of ν in eq. (1), we get

$E = hc/\lambda$ .... (2)

Thus greater the value of λ, smaller will be the energy.

Now $\bar{\upsilon} = 1/\lambda$ .... (3)

Where $\bar{\upsilon}$ = wave number.

Putting the value of $1/\lambda$ in eq. (2), we get.

$E = hc\bar{\upsilon}$

Today's Learning

Don't forget to put your name on!

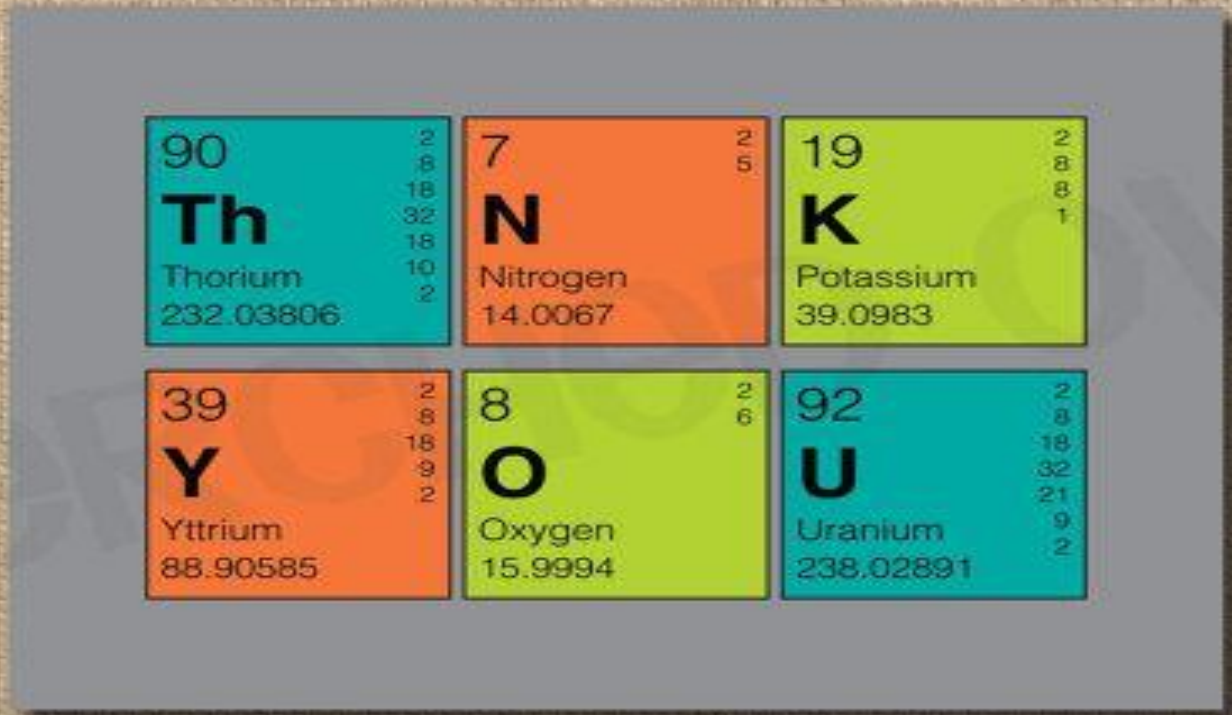
Something I can do now that I couldn't do before the lesson is…	A question I would like to know the answer to is…	I need to improve on…

HOME-WORK:

- NOTE:ATTEMPT ANY ONE QUESTION:
- HOW TO DETERMINE THE ENERGY DIFFERENCE OF AN ELECTRON IN AN ORBIT OF AN ATOM?
- HOW TO DETERMINE THE WAVE NUMBER OF AN ELECTRON OF AN ATOM?
- EXPLANATION OF HYDROGEN SPECTRUM OF AN ATOM?
- EXPLANATION OF DIFFERENT SPECTRAL LINES?

LESSON CLOSURE:


good bye
good bye
SCHOOL BUS

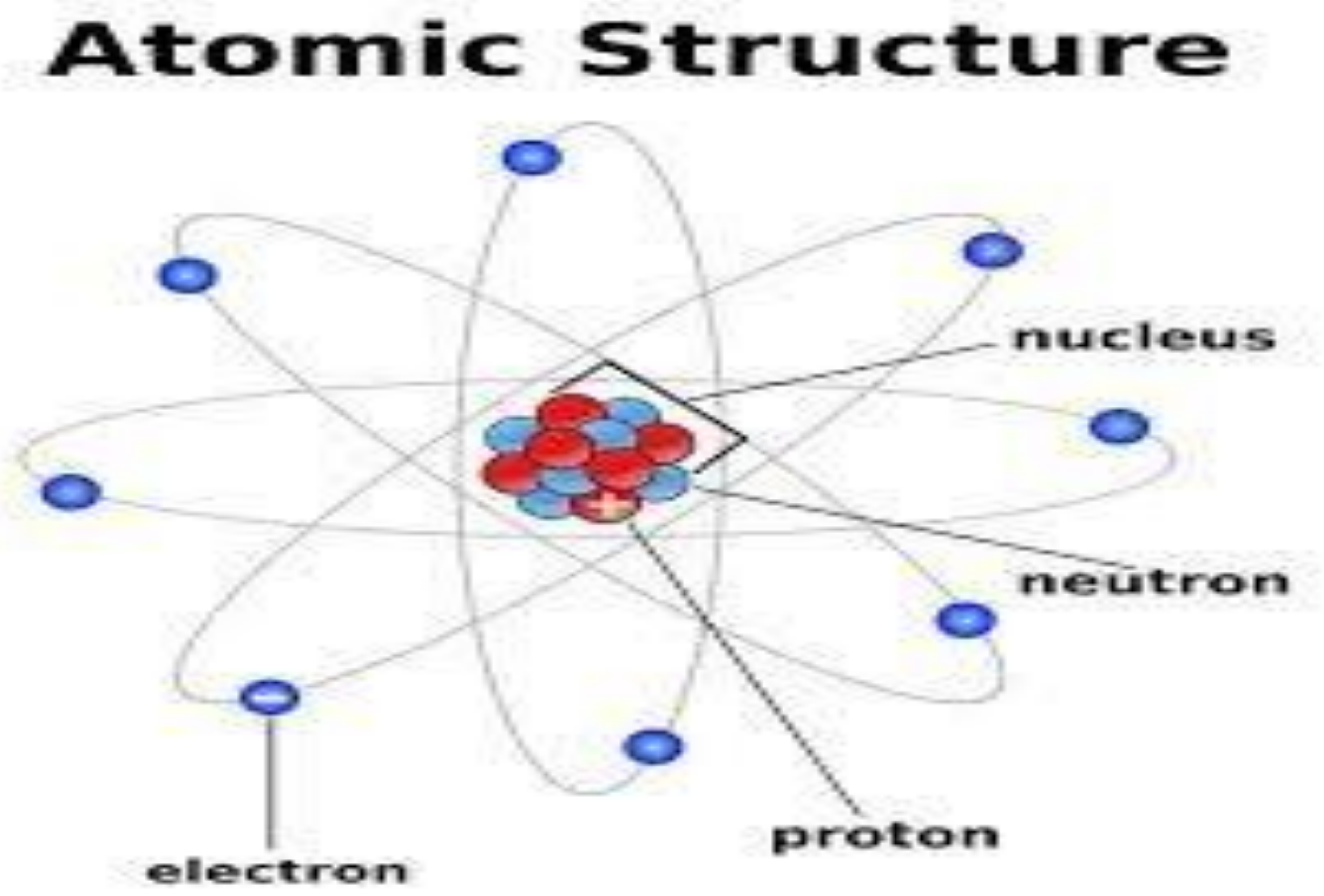


90
Th
Thorium
232.03806
2 8 18 32 18 10 2
7
N
Nitrogen
14.0067
2 5
19
K
Potassium
39.0983
2 8 8 1
39
Y
Yttrium
88.90585
2 8 18 9 2
8
O
Oxygen
15.9994
2 6
92
U
Uranium
238.02891
2 8 18 32 21 9 2

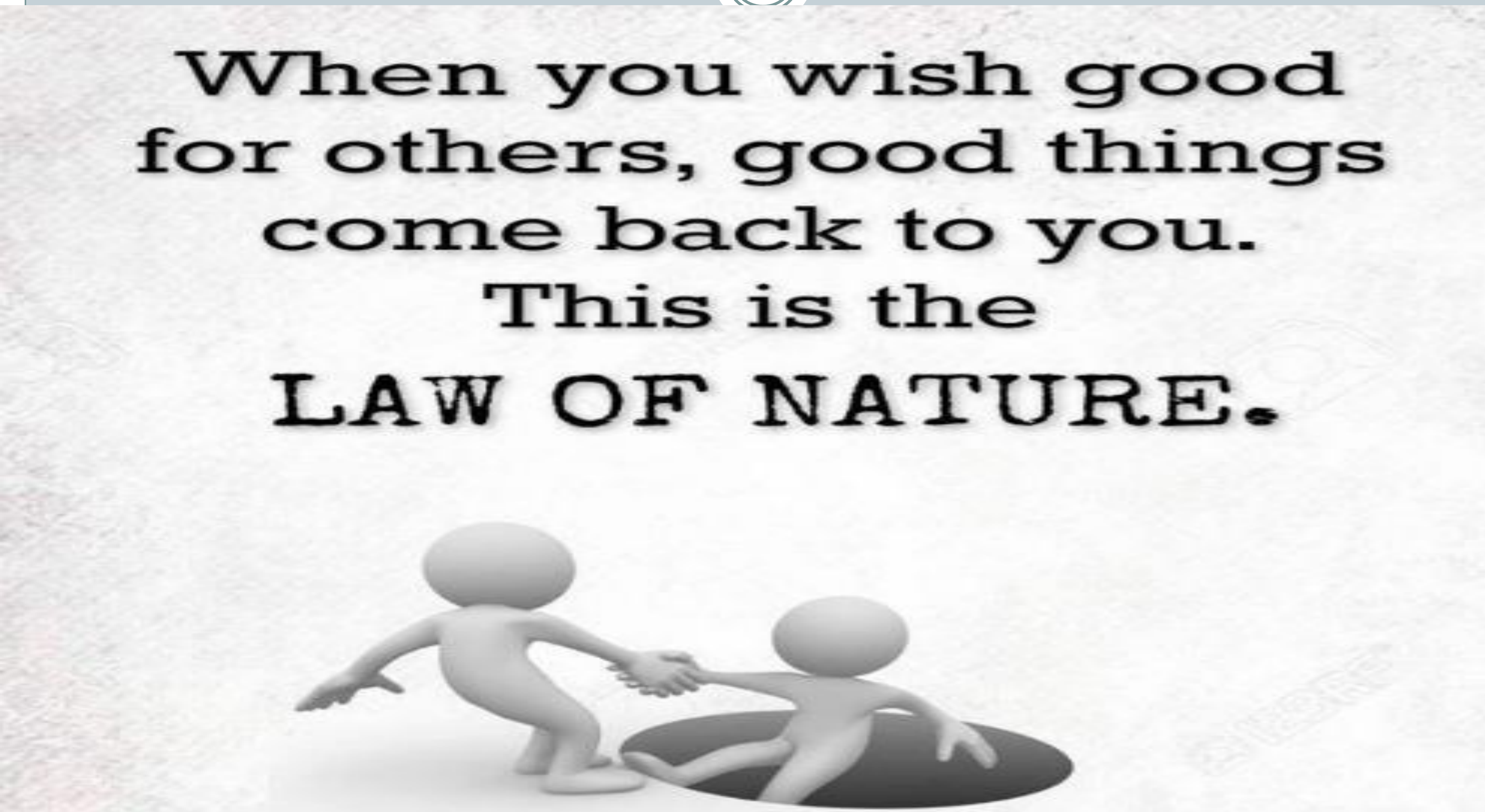

Thank
You

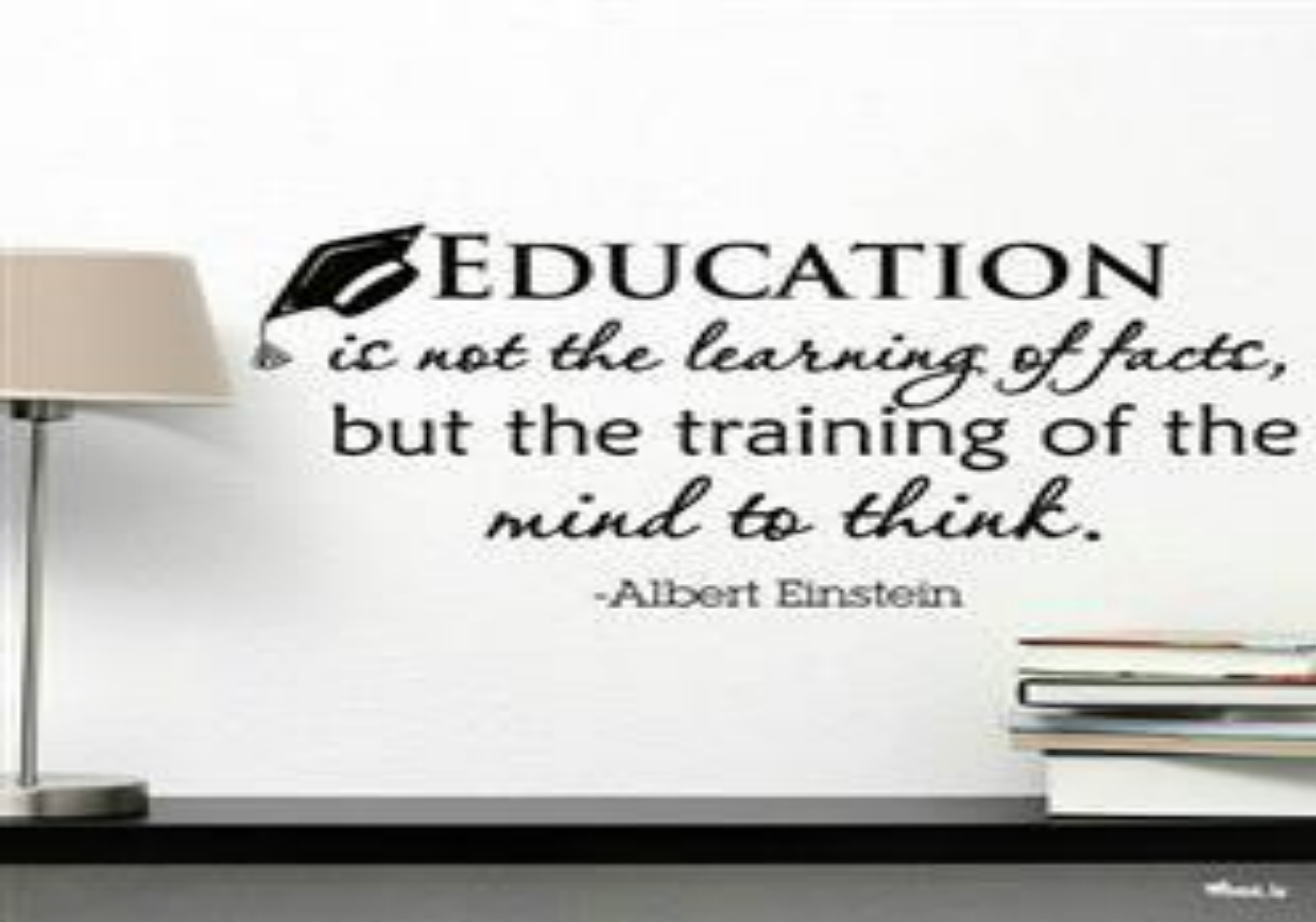

TAKE CARE

LECTURE NO:13



MESSAGE OF THE DAY:




EDUCATION
is not the learning of facts,
but the training of the
mind to think.
-Albert Einstein


HAHA! SO WE MEET AGAIN!!


>> Ready for Anything
WELCOME BACK TO VIRTUAL CLASSES!
LESSON 1
A
B
C
#WEAREALWAYSTOGETHER
AMERICAN
AS

A WARM WELCOME TO ALL THE STUDENTS IN THE ONLINE CLASSES.THIS IS D.CHARLES

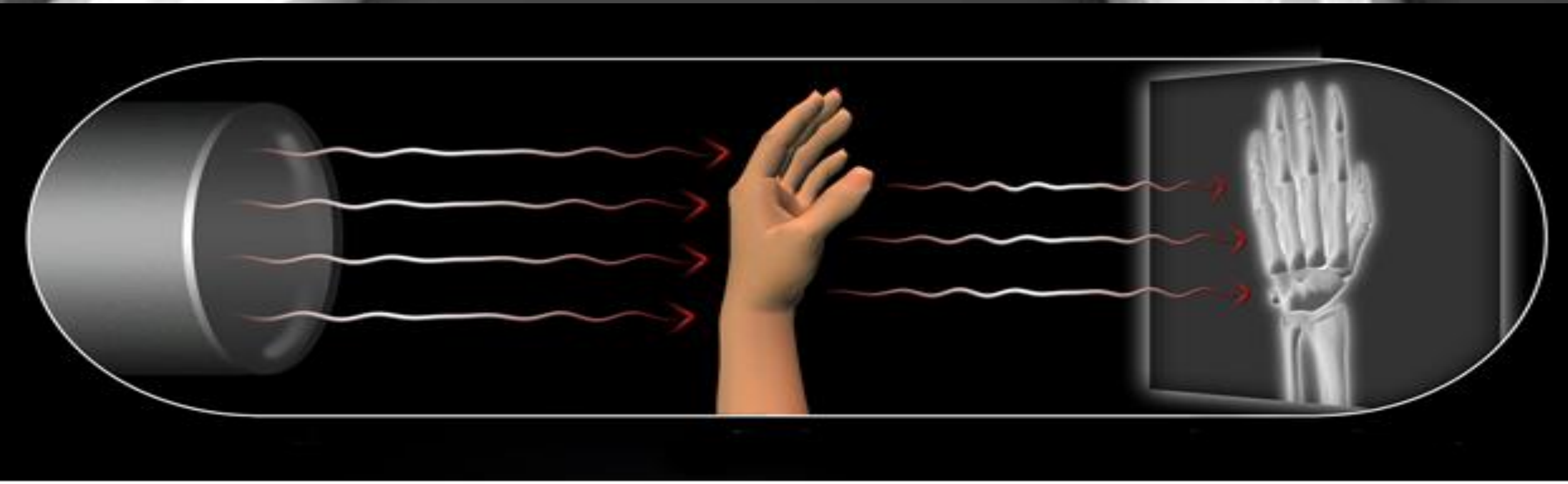
WE ARE GOING TO START OUR ONLINE CHEMISRTY LESSON TODAY.I HOPE YOU ALL WILL ENJOY AND LEARN.

.

RULES OF THE CLASS:

1)BE ON TIME FOR ALL YOUR CLASSES.
2)RESPECT ALL PARTICIPANTS OF THE CLASS.
3)DO NOT CREATE ANY DISTURBANCE.
4)RAISE HAND IF YOU HAVE A QUESTION.
5)PAY ATTENTION TO YOUR TEACHER.

POINTS TO PONDER:



LESSON OBJECTIVES:13

- BY THE END OF THIS PART OF LESSON,STUDENTS WILL BE ABLE TO:
- Define X-rays ? Who discovered X-rays ? How X-rays were discovered?
- What is meant by X rays -diffraction ? Write its applications?
- State Moseley law ? Write its importance ?

1. Introduction:

- **X-Ray Definition**:X-Rays are short wavelength electromagnetic radiation between UV & gamma ray, which consist of wavelength in the region about **0.1Å to 100Å**
- For analytical purpose, the range of 0.7-2.0 A^{o} is the most useful region.
- A German professor Rontgen in 1895 discovered X-ray while working with a discharge tube
- Barium platinocyanide screen placed near discharge tube began to glow. The glow continued even when a wooden screen was placed between them
- These x-rays could pass through bodies, which are opaque to ordinary light

Definition of X-ray

- It is a type of electromagnetic radiation characterized by wavelengths between approximately 1 A and 10^{-4} A.

- They are invisible, penetrative especially at higher photon energies, and travel with the same speed as visible light.

- They are usually produced by bombarding a target of high atomic number with fast electrons in a high vacuum

What is an x ray?

An X-ray machine is basically like a camera. It uses X-rays to expose the film, instead of visible light.

X-rays are similar to light in that they are electromagnetic waves, but they are more energetic so they can penetrate many materials to varying degrees. When the X-rays hit the film, they expose it just as light would. Various structures such as bone, fat, muscle, tumors and all other masses absorb X-rays at different levels *(they do not let the x ray energy pass through).* The image on the film lets you view distinct structures inside the body because of the different levels of exposure on the film.

Characteristic x rays

- There must be a vacancy in a subshell of an atom to produce characteristic x-rays.
- The bombarding electrons can eject electrons from the inner shells of the atoms of the metal target.
- Those vacancies will be quickly filled by electrons dropping down from higher levels, emitting x-rays with sharply defined frequencies associated with the difference between the atomic energy levels of the target atoms, called characteristic x-rays.

X-ray diffraction

- Definition
  - The scattering of x-rays by crystal atoms that produces a pattern that yields information about the structure of the crystal. The wavelengths of x-rays are comparable in size to the distances between atoms in most crystals. X-ray diffraction is the basis of x-ray crystallography

http://science.education.nih.gov/supplements/nih4/technologyother/glossary.htm

X – Ray Diffraction

- " Every crystalline substance gives a pattern; the same substance always gives the same pattern; & in a mixture of substances each produces its pattern independently of the others"
- The **X-ray diffraction pattern** of a pure substance is, therefore, like a **fingerprint of the substance**. It is based on the scattering of x-rays by crystals.
- Definition : The atomic planes of a crystal cause an incident beam of X-rays to interfere with one another as they leave the crystal. The phenomenon is called X-ray diffraction

2.4 X-RAYS

Wilhelm Roentgen (1895) accidentally discovered that if cathode rays are pointed to fall on a heavy metal target, there are produced some penetrating short wave length rays. He called them the X-rays. The X-rays are electromagnetic radiations of very high frequency depending upon the nature of anode. Oftenly a tungsten filament is used for this purpose.

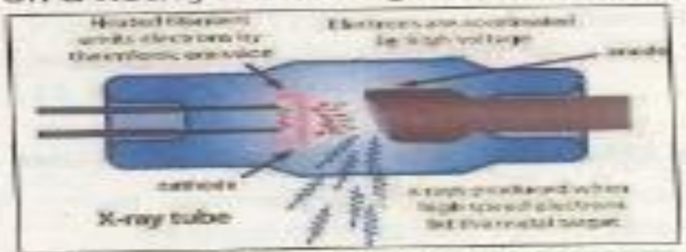


Figure 2.11: Cathode rays pointed at heavy metal (W, Cu etc.)

X-rays are emitted from the target in all directions. A small portion of them is used for useful purpose through the windows. The wavelength of X-rays produced depends upon the nature of target metal. Every metal has its own characteristic X-rays.

2.4.1 Atomic Number and X-rays

Moseley undertook a systematic and comprehensive study of X-rays in 1913. His researches covered a range of wavelengths 0.04 – 0.08 Å.

Moseley proved that the frequencies of X-rays increase in a regular manner from one element to the other in the Periodic Table. He further suggested that the frequencies of these rays are directly proportional to the no of protons in the nucleus. The no of protons in the nucleus is called "Atomic Number".

Moreover Moseley drew the following conclusions from the detailed analysis of spectra lines which he obtained from 38 different elements. (from Al to Au) as targets in X-rays tube.

(a) The spectral lines could be classified into two distinct groups. One, which belong to shorter wavelength, called K-series and the other with longer wavelength called as L-series.

(b) If the target element is of higher at no., the wave-length of X-rays becomes shorter

(c) A relationship between frequency (υ) and atomic number (Z) of the elements given as

$$\sqrt{\upsilon} = a(Z-b)$$

This is called Moseley Law. Where a and b are called constant quantities. This law states that the frequency of a spectral line in x-ray spectrum varies as the square of atomic number of an element emitting it.

2.4.2 X-Rays, Atomic Numbers and Orbital Structure

In 1913 Henry G. J. Moseley, a student of Rutherford, used the technique of X-ray spectroscopy (just discovered by Max von Laue) to determine the atomic numbers of the elements. X-rays are produced in a cathode-ray tube when the electron beam (cathode ray) strikes on a metal target. The explanation for the production of X-rays is as follows: When an electron in the cathode ray hits a metal atom in the target, it can (if it has sufficient energy) knock out electron from an inner shell of the atom. This produces a metal ion with an electron missing from an inner orbital. The electronic configuration is unstable, and an electron from an orbital of higher energy drops in to the half-filled orbital and a photon is emitted. The photon corresponds to electromagnetic radiations in the x-ray region.

2.4.3 Uses of X-Rays

Practical applications followed closely on the heels of discovery of X-rays. Because X-rays have different penetrating powers for different types of matter, they can be used to photograph interior of objects. Roentgen's original announcement of X-rays was made on December 28, 1895. On January 20^{th}, 1896, in Dartmouth, New Hampshire, X-rays were used to assist in setting a person's broken arm. This was something of a record time for turning a scientific discovery in to practical application.

The layers of the closely packed particles in a crystal constitute planes. In 1912 Max von Laue in Germany suggested that the particles and the planes of a crystal might be separated by distances that are of the same order of magnitude as the wavelength of X-rays. Therefore a beam of X-rays should be diffracted by a crystal. The suggestion was quickly verified experimentally and the highly symmetrical diffraction pattern that appears on a photographic film is known as the Laue pattern of a substance. In England in 1913 William Bragg and Lawrence Bragg devised a simpler apparatus to determine the internal structure of a crystal, which is called X-ray diffraction technique.

2.5 THE QUANTUM NUMBERS AND ORBITALS

Schrodinger in 1926 gave an equation in which electrons are treated as moving with wave like motion in the three dimensional space around the nucleus. It differs from Bohr's atomic model in the sense that the electrons move in orbits. It also specifies the distance between the electron and the nucleus.

The solution of Schrodinger's wave equation gives certain mathematical integers. These sets of numerical values, which give the acceptable picture of an atom, are called Quantum Numbers. There are four quantum numbers which can describe the electron completely.

1. **Principal Quantum No. (n)**

It determines the size of the orbit and the distance from the nucleus. Greater the distance from the nucleus, larger will be the size of the orbit. The shells are named as,

If $n = 1$ — K shell.
$n = 2$ — L shell.
$n = 3$ — M shell.
$n = 4$ — N shell.

The no of electrons accommodated in various orbits are as follows.

K = 2, L = 8, M = 18, N = 32.

Today's Learning

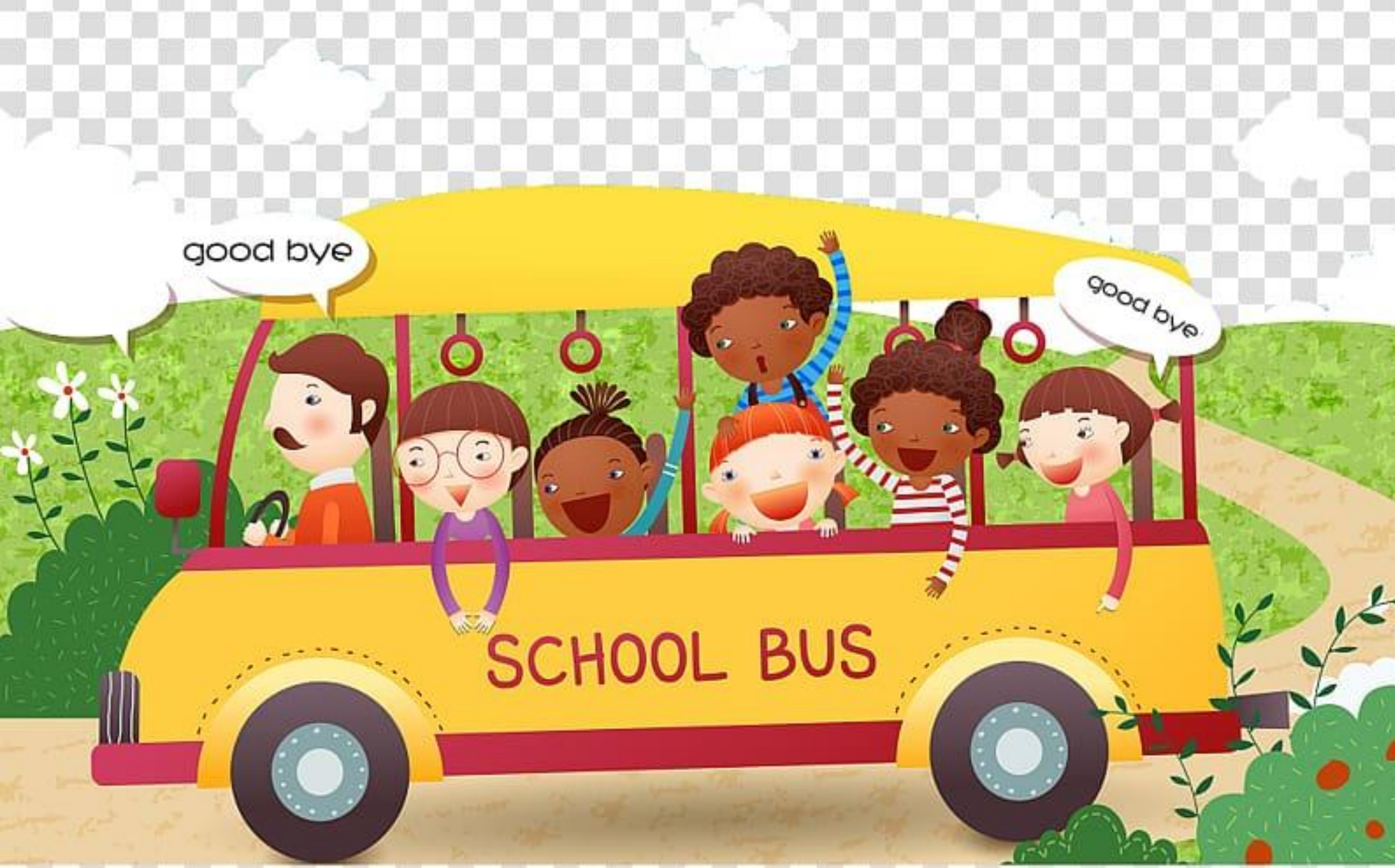
Don't forget to put your name on!

Something I can do now that I couldn't do before the lesson is…	A question I would like to know the answer to is…	I need to improve on…

HOME-WORK:

- NOTE:ATTEMPT ANY ONE QUESTION:
- How X-rays were discovered ?
- Write the uses of X-rays ?
- State Moseleys law ?
- Explain X-diffraction technique in detail?

LESSON CLOSURE:


good bye
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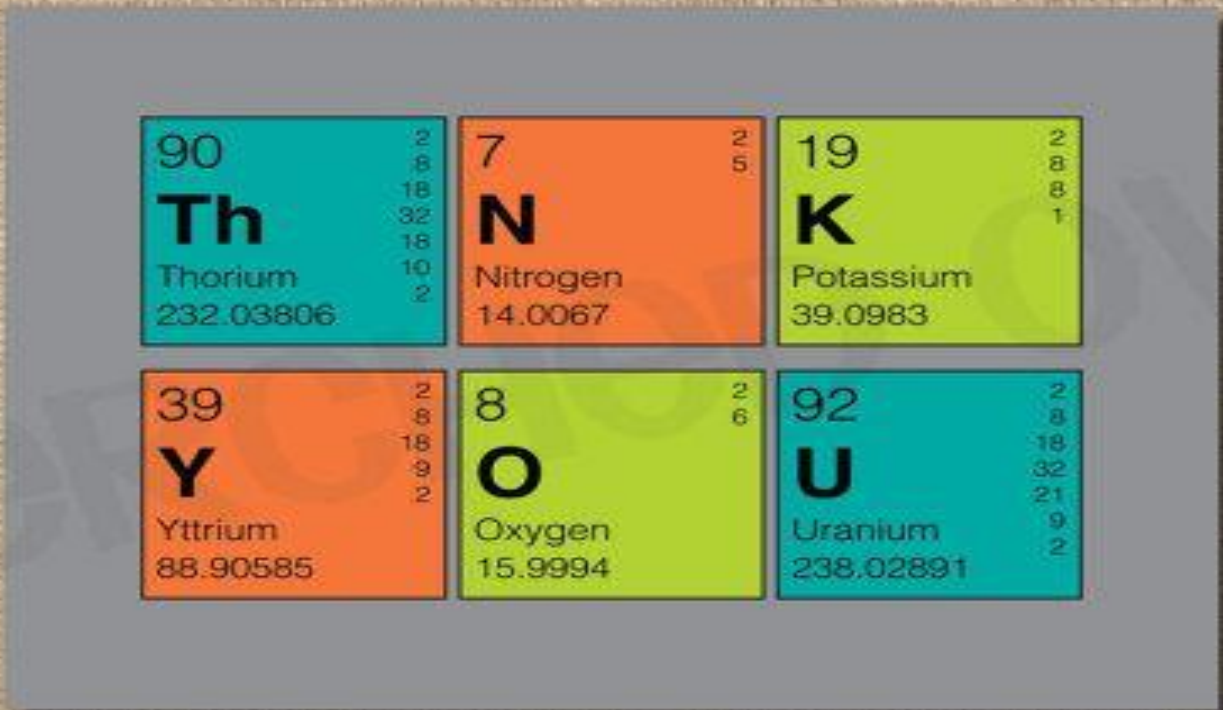
SCHOOL BUS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Yttrium
88.90585
2 8 18 9 2
8
O
Oxygen
15.9994
2 6
92
U
Uranium
238.02891
2 8 18 32 21 9 2


Thank
You


TAKE CARE