General Chemistry (I)

For 1st year University Students.

Prepared by:

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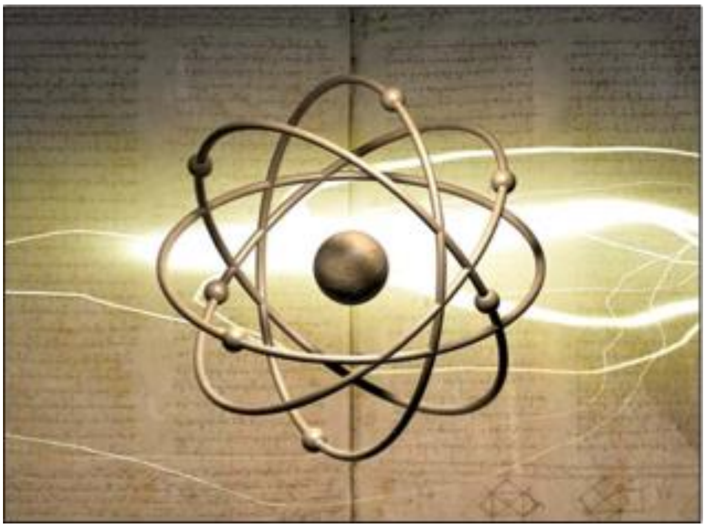
First Level (Freshman) - Course's Contents of Physical Science and Geological Science Programs.

Chm 101: General Chemistry (I) - 3 Credits (Lecture 2 Hrs. / W + Lab. 2 Hrs. / W).

Contents:

The history of atomic theory - Different Atomic models - Thomson Model of Atom - Bohr Model of the Atom - Rutherford's Nuclear Model of Atom - Electron Discovered - Protons & Nucleus Discovered - Electron's Charge Measured - Quantum Model of the Atom - Quantum number - Modern Atomic Theory - Schrodinger's atomic Model (The Cloud Model) - de Broglie (Dual Behavior of Matter) - Heisenberg's Uncertainty Principle - Quantum Mechanical model of atom (Schrodinger's equation) - Shapes of Atomic Orbitals - Filling of Orbitals in Atom - Aufbau Principle - Hond rules-Periodic table - Periodic table properties - ionization energy - electronegative force - electron affinity - oxidation and reduction state - types of bonds - hybridization.



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Atoms, Molecules, and Chemistry

What are atoms?

1. Atoms are the fundamental units of matter.

2. They are composed of smaller parts:

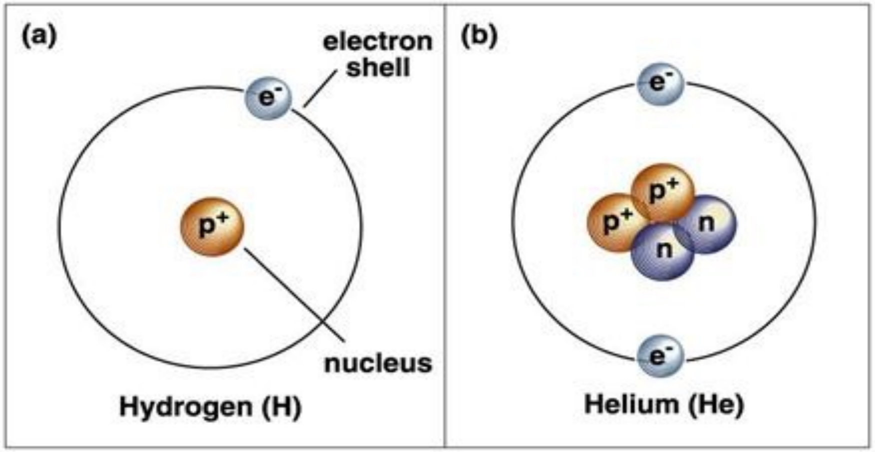
   a. Nucleus

      i. Protons (positively charged)

      ii. Neutrons (have no charge)

   b. Electrons

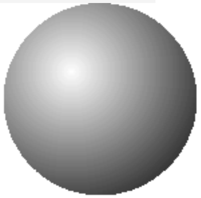
The different components of atoms determine the type and "behavior" of the elements…



The History of Atomic Structure

Democritus (460 - 370 BC)

- Was the first person to come up with the idea of atom
- Believed that all matter was composed of indivisible particles he called "**ATOMS**"
- He also believed that different atoms:
  - Are different sizes
  - Have different properties



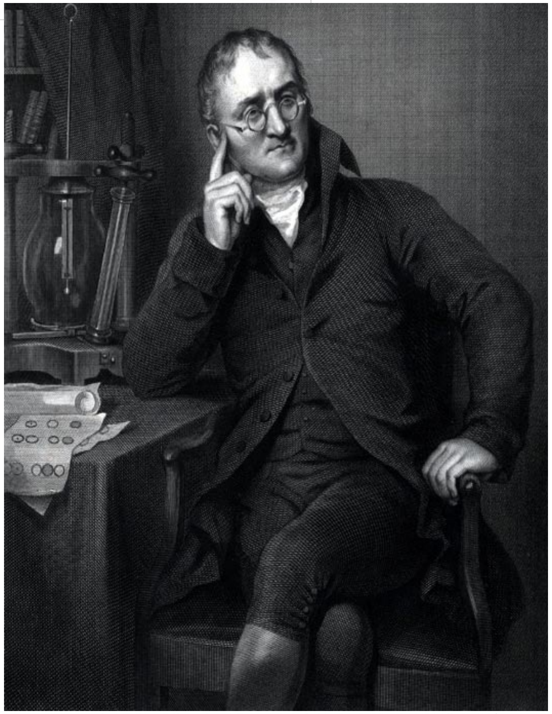
Democritus (400 B.C.)

Democritus (460 - 370 BC)

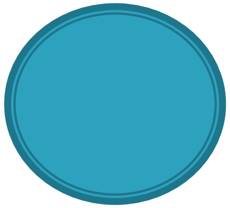
John Dalton (1766-1844)

- Dalton is the "Father of Atomic Theory"

- Dalton's ideas were so brilliant that they have remained essentially intact up to the present time and has only been slightly corrected.



John Dalton (1766-1844)

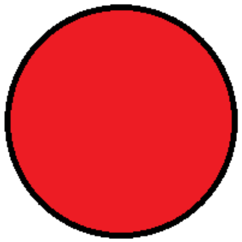


Dalton's Atom

Dalton's Atomic Theory (1803)

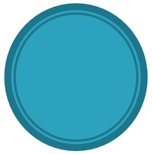
1. **All matter is composed of extremely small particles called atoms**.

2. **All atoms of a given element are identical, having the same**:
   - size
   - mass
   - chemical properties



Dalton's "Billiard Ball" Model of Atom

3. **All atoms of different elements are different**. (a Sodium atom is different than chlorine)



Dalton's Atom

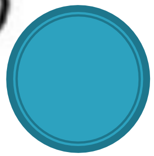
Dalton's Atomic Theory (1803)

4. **Atoms cannot be created, divided into smaller particles, or destroyed**.

In a chemical reaction, atoms of different elements are separated, joined or rearranged. They are never changed into the atoms of another element.

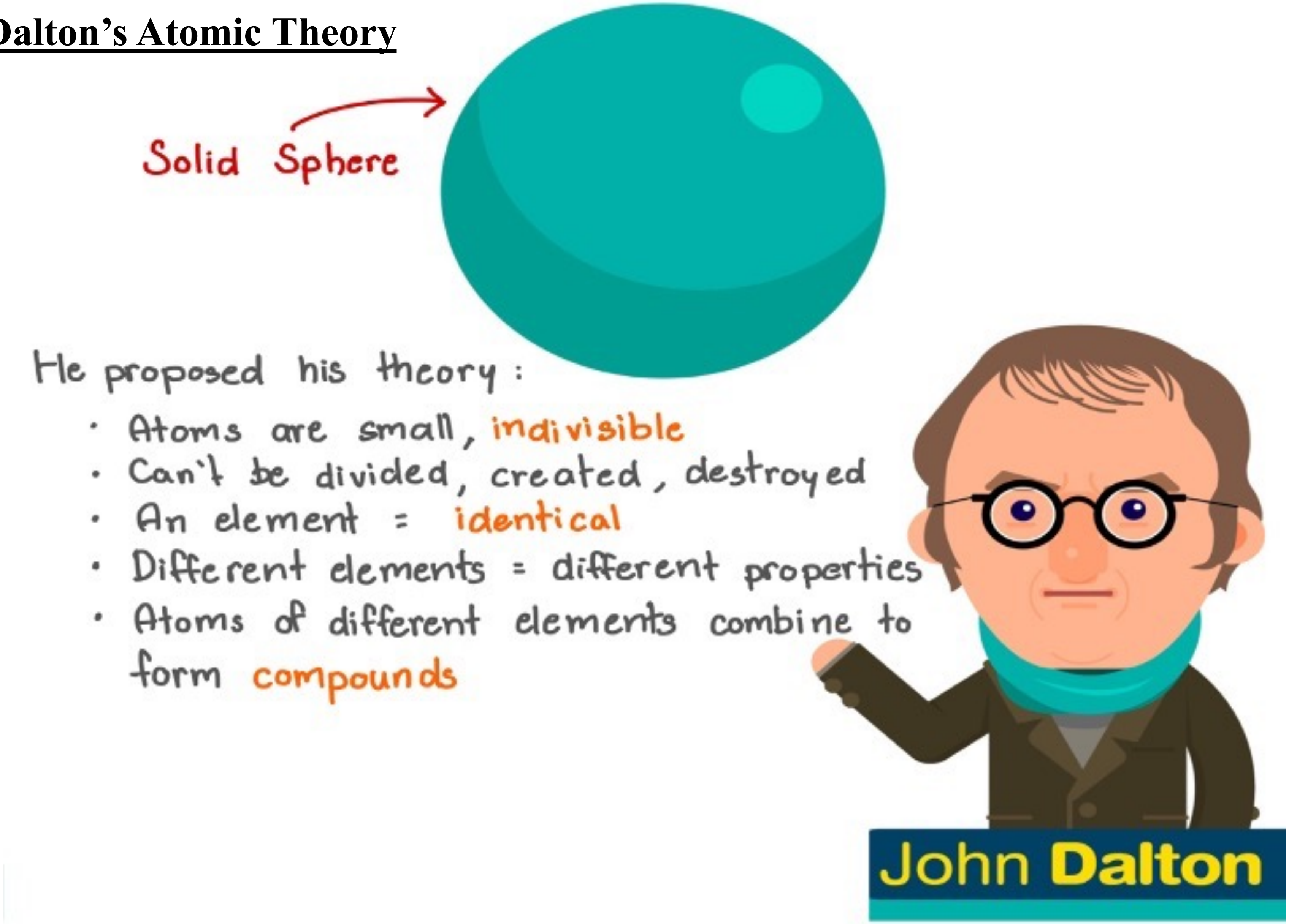
5. **Law of Definite Proportions: Different atoms combine in simple whole number ratios to form compounds**

(you can't have a ½ of a Carbon bonding with Oxygen; it's a whole atom or no atom)

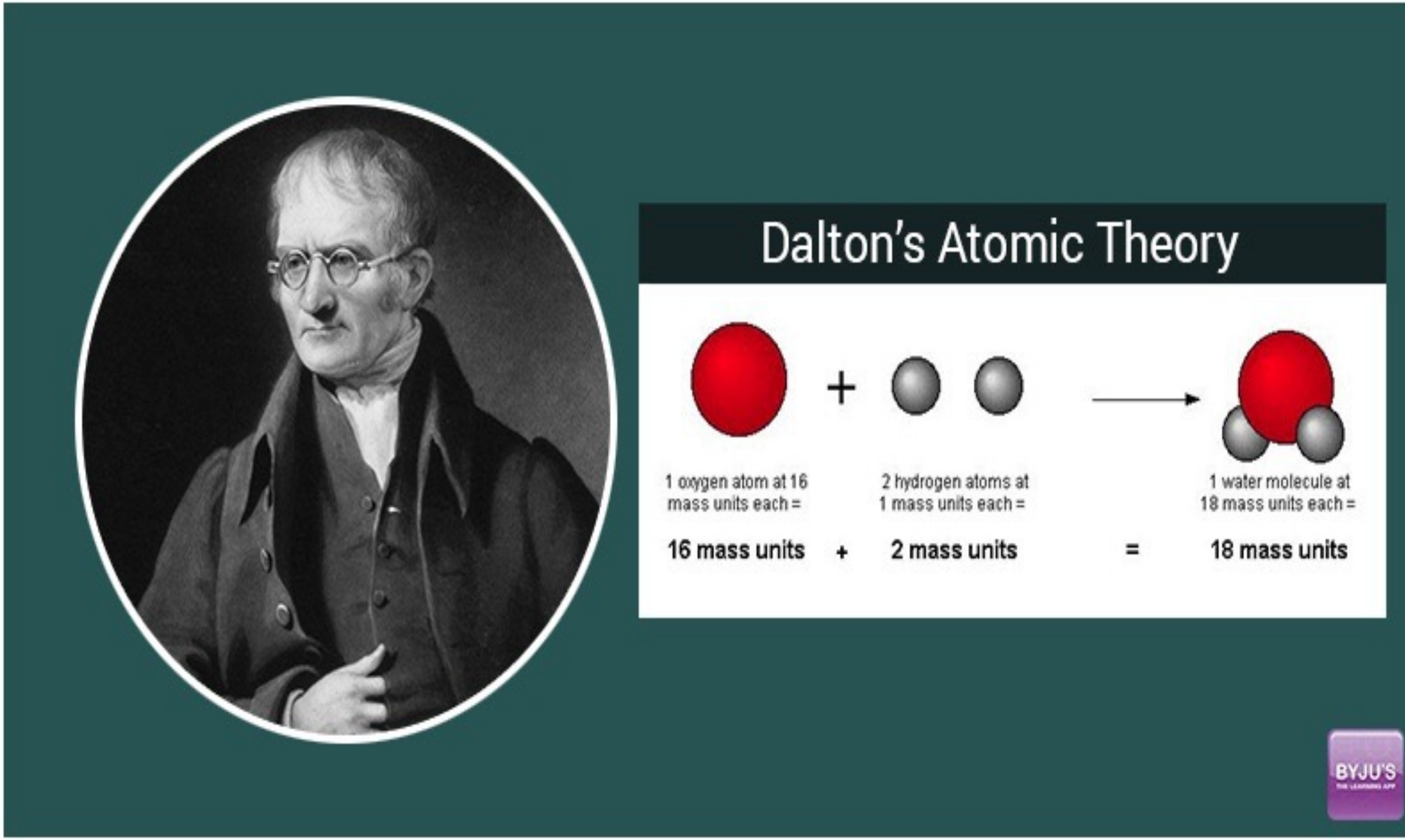


Dalton's Atom

Dalton's Atomic Theory



Dalton’s Atomic Theory



Cathode Rays

Cathode Rays (Discovering of Electrons) (Plücker 1858 and Thomson 1897)

Cathode Rays experiments

These experiments relied on the following basis:

1. All Gases under the normal conditions of pressure and temperature are insulator for electricity.
2. If a closed tube was evacuated of gases until the pressure becomes less than 0.01 to 0.001 mmHg, at these conditions the gas becomes conductor at a specific potential.

Cathode Rays (Discovering of Electrons) (Plücker 1858 and Thomson 1897)

Cathode Rays' experiments

3. If the potential between the two electrodes increase up to 10 KV, it can be noticed a current of rays from Cathode to Anode, causing flashing of light on the walls of the evacuated tube, these rays called Cathode rays, or Electrons.

Julius Plücker (16 June 1801 – 22 May 1868)

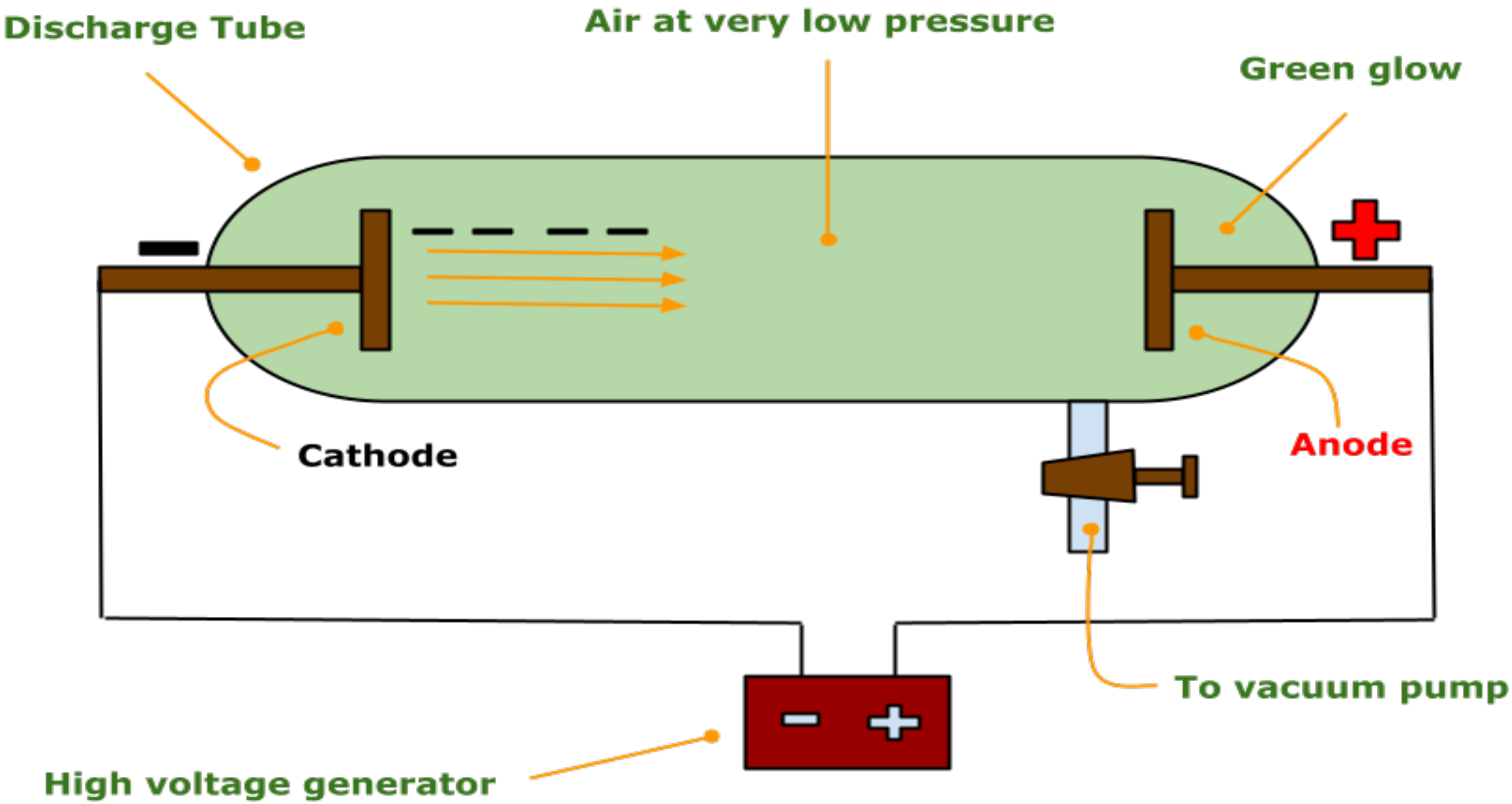
was a pioneer in the investigations of cathode rays (1858) that led eventually to the discovery of the electron (1897).

Joseph John Thomson

Born	**18 December 1856** **Cheetham Hill, Manchester, England**
Died	**30 August 1940 (aged 83)** **Cambridge, England**

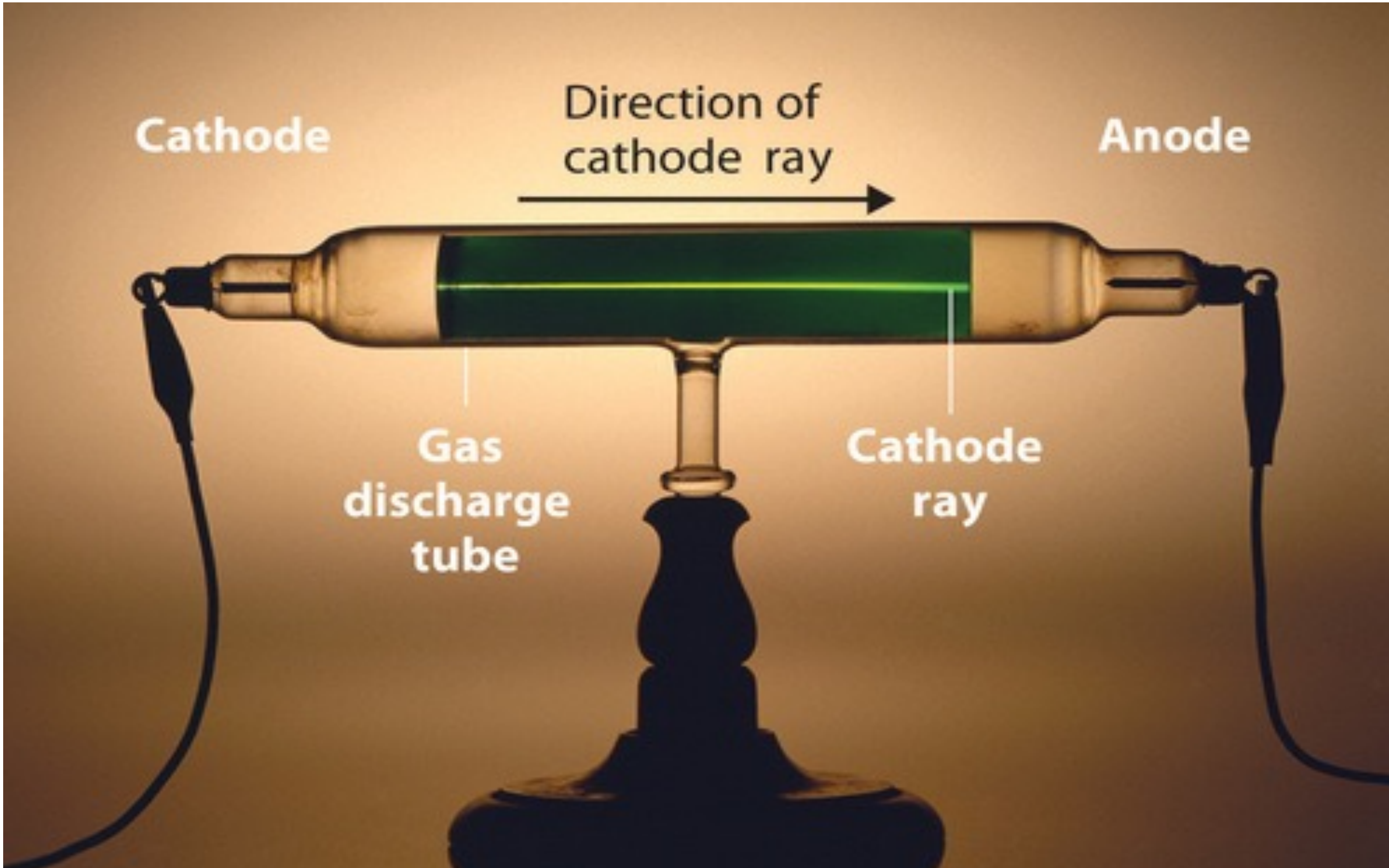
J. J. Thomson

Cathode Rays' experiments

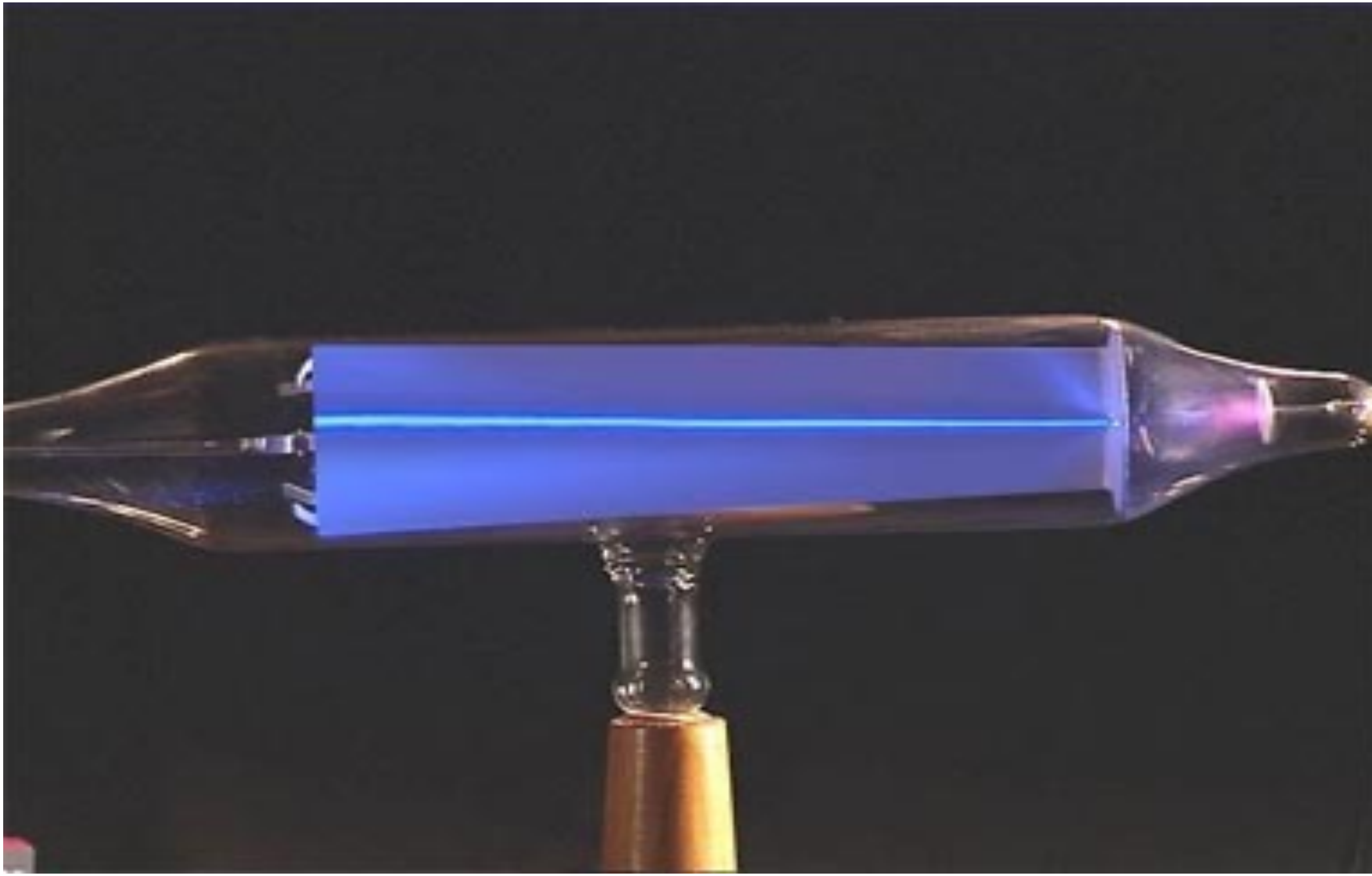


Production of cathode rays

Cathode Rays' experiments



Cathode Rays' experiments

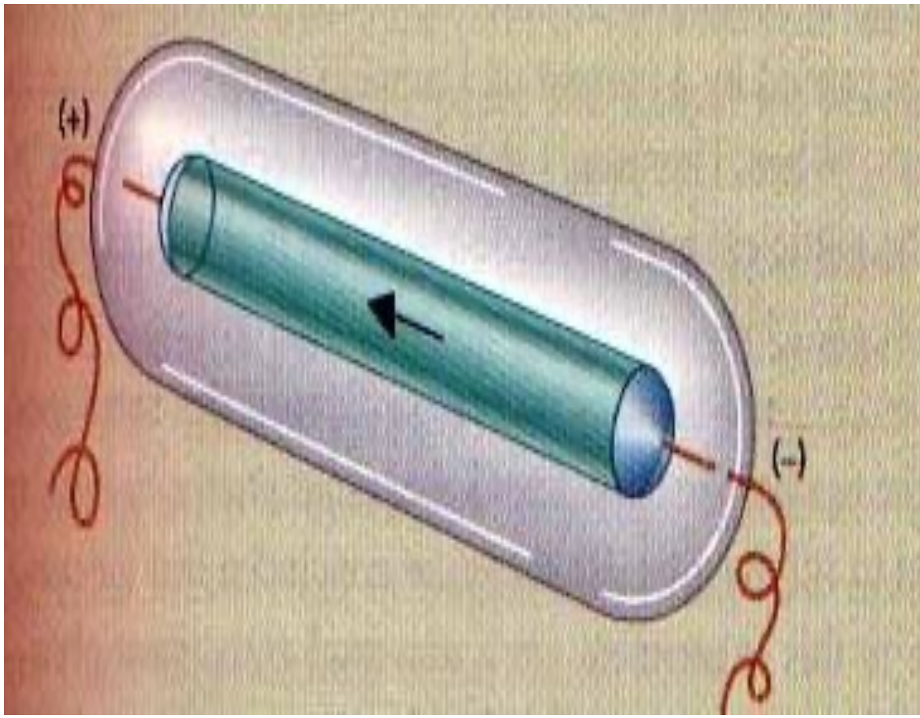


Characters of Cathode Rays

Characters of Cathode Rays

Cathode rays are current of unseen rays produced from cathode causing flashing light on evacuated tube's walls.

1. Consists of very small particles called electrons.
2. Move in straight lines.
3. Have a heat effect.
4. Affected by both of electric and magnetic fields.
5. Have a negative charge.
6. Don't change according to material types or gas type, which means that they are contained in all substances.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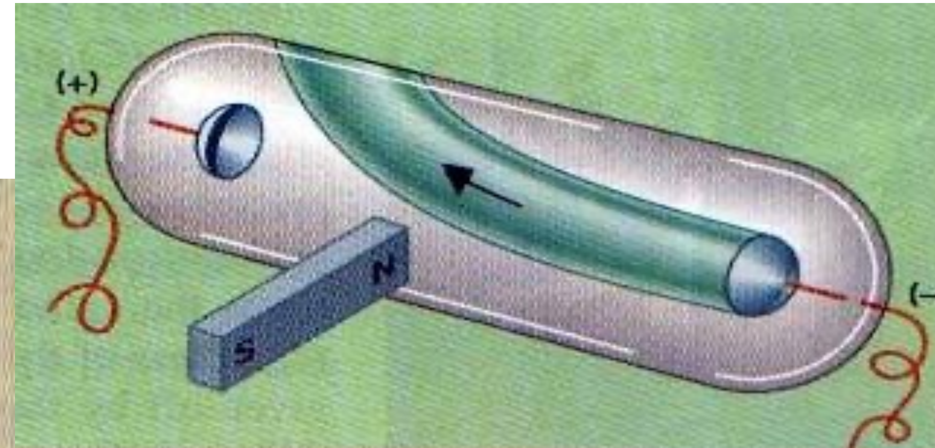
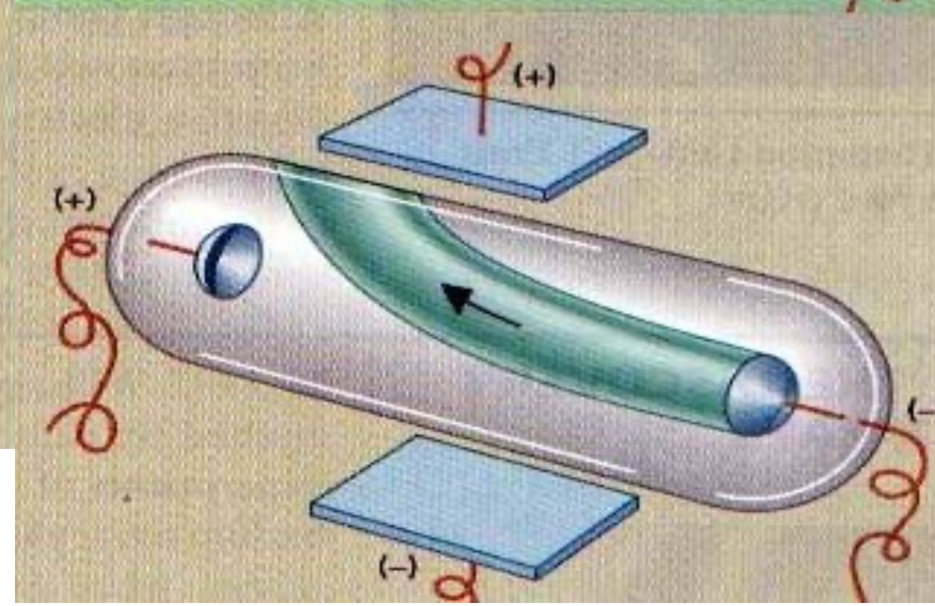
(–)

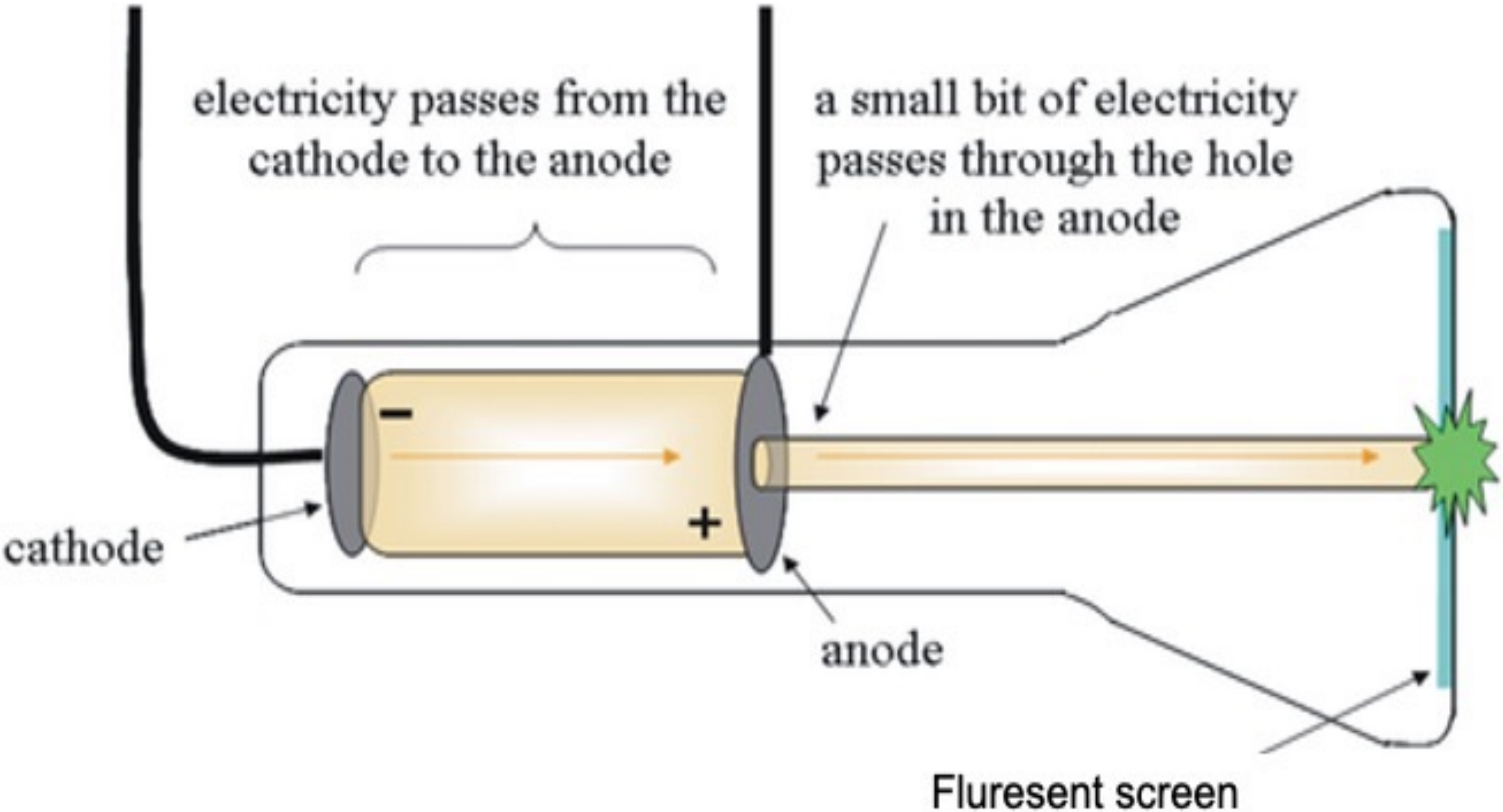

(+)
(–)
N
S


(+)
(+)
(–)
(–)

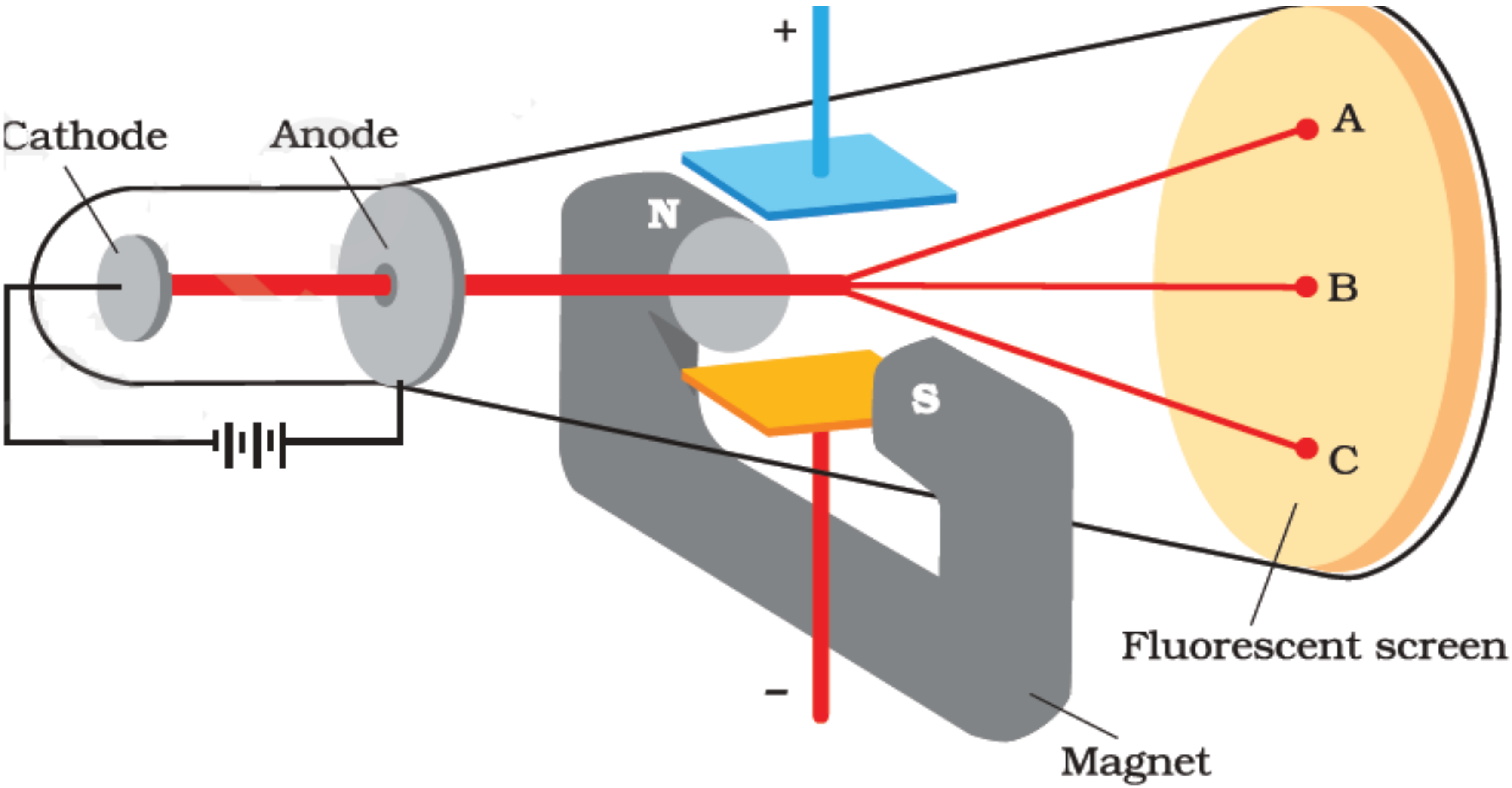
Cathode Rays' experiments

Applications of Cathode Rays:

Cathode Rays' experiments



Cathode Rays' experiments



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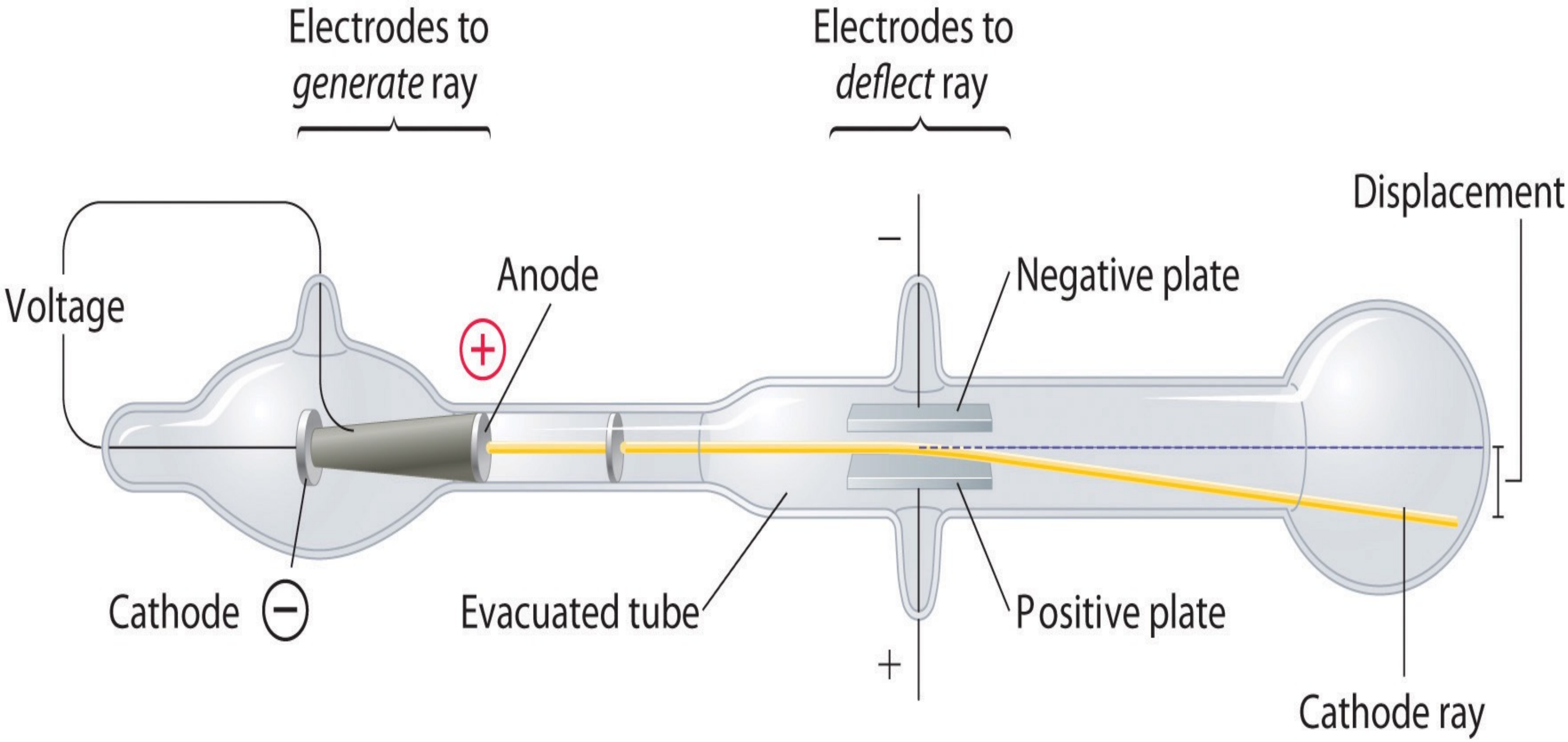

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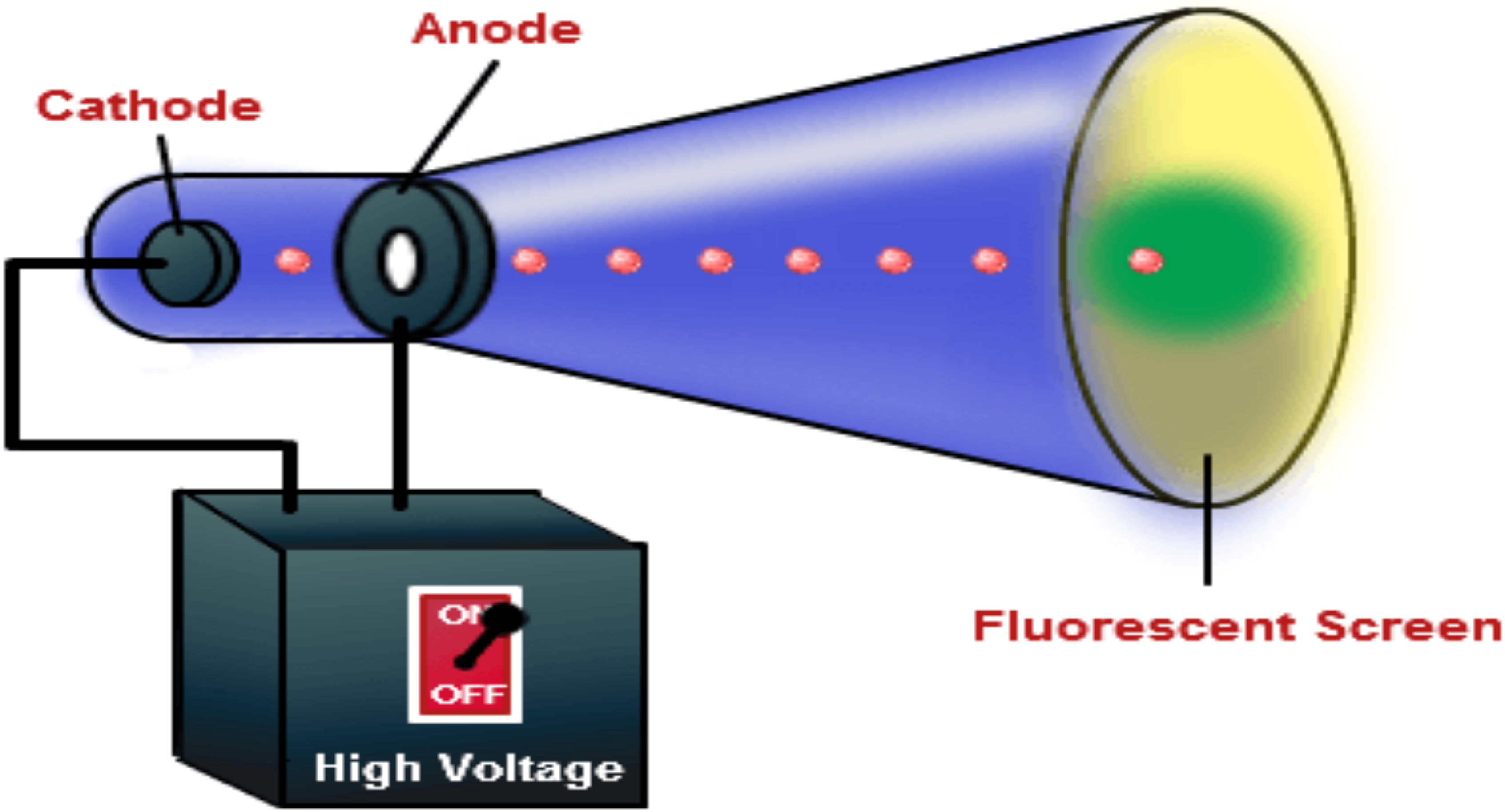
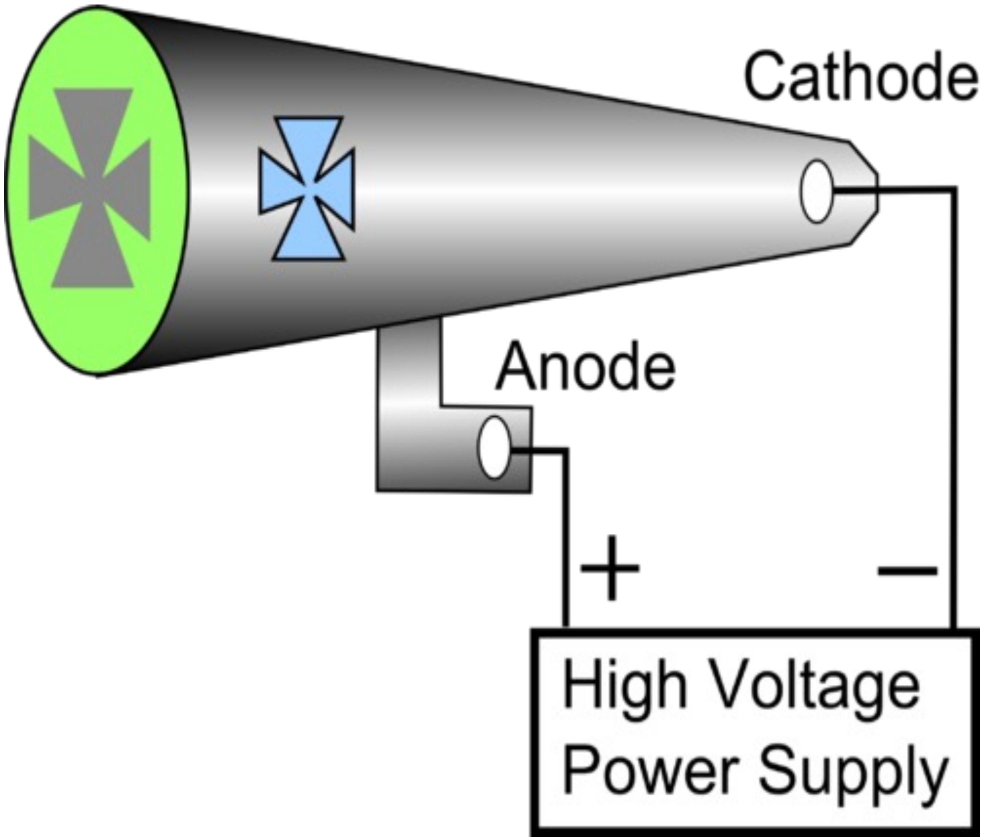


Cathode Rays' experiments

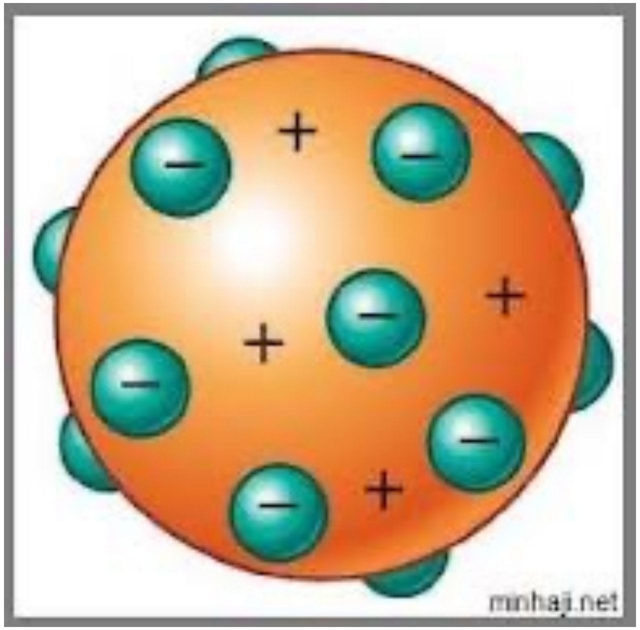


Atomic Structure

Thomson Model 1897

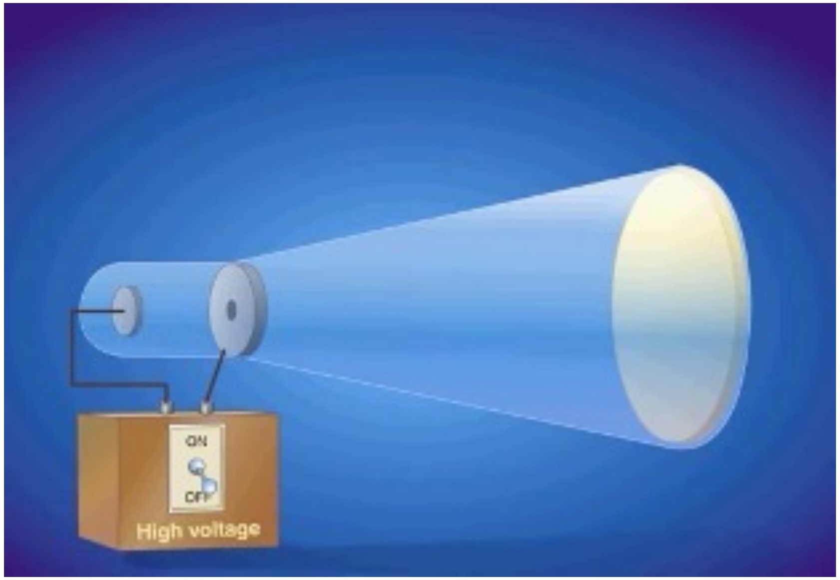
An atom is a spherical body of positive electricity that is permeated with negative electrons.

John Joseph Thomoson



Thomson Model

$$e/m = -1.76 \times 10^{8} \text{ Colum / gm}$$

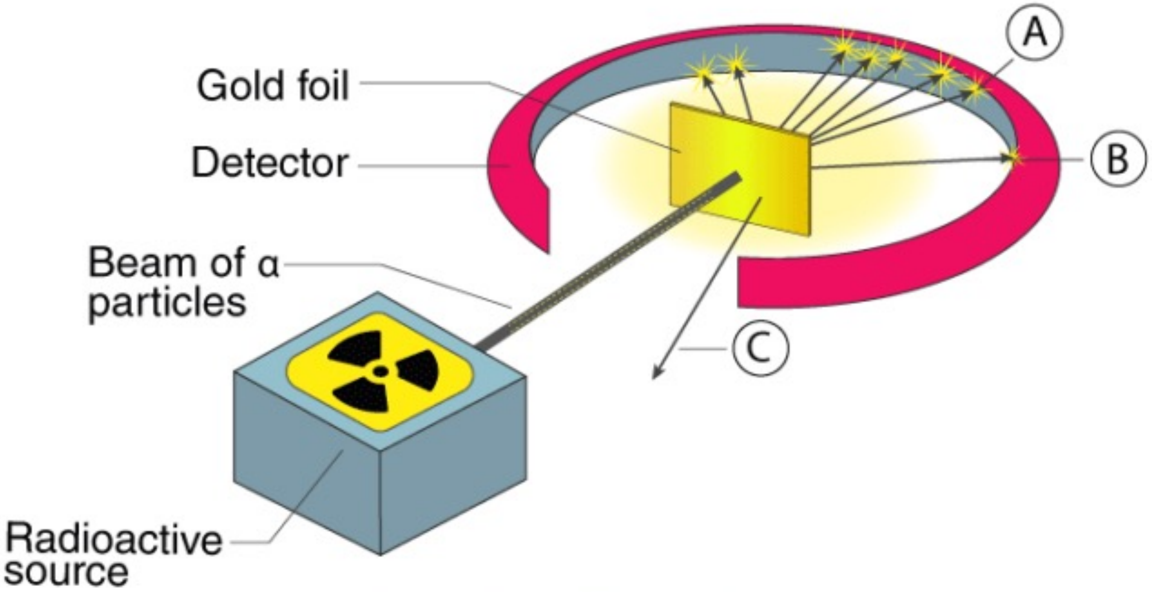


Millikan Experiment

$$e = -\ 1.602 \times 10^{-19}\ C$$

$$m = 9.1 \times 10^{-28}\ g$$

Rutherford Model for Atomic Structure



Rutherford's Experiment

Ernst Rutherford (1871-1937)

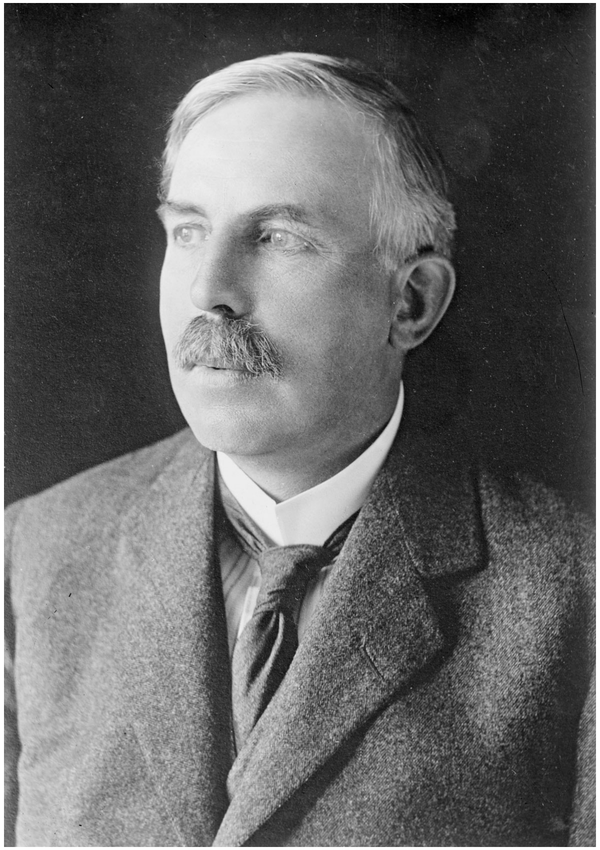
Born: August 30, 1871, Brightwater, New Zealand

Died: October 19, 1937, Cambridge, United Kingdom

Discoveries: Rutherford model, Atomic nucleus, Proton, and MORE

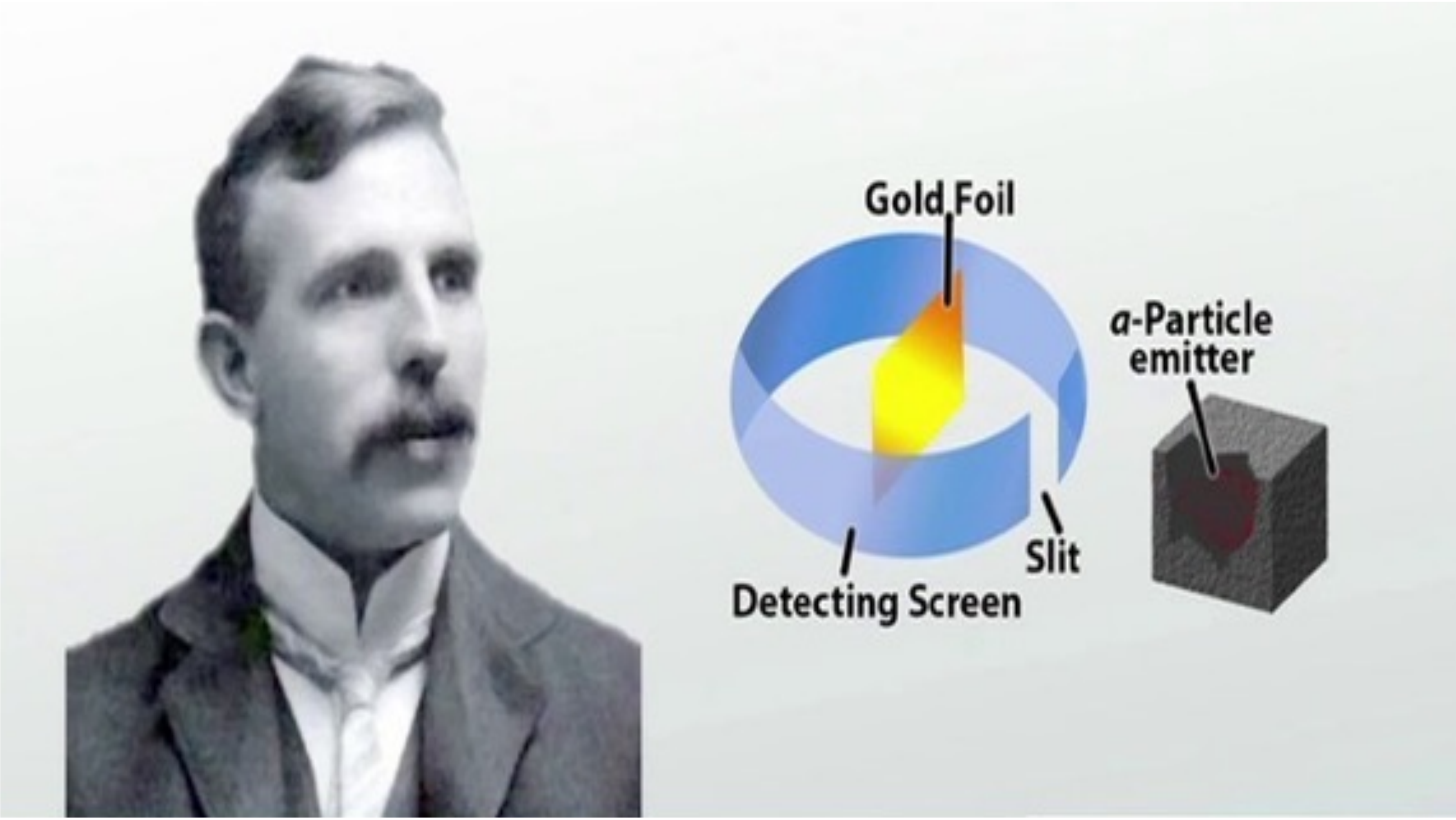
Awards: Nobel Prize in Chemistry, Copley Medal, Matteucci Medal, and MORE

Nationality: British, New Zealand



Ernst Rutherford (1871-1937)

Rutherford Model for Atomic Structure.



Rutherford exp.

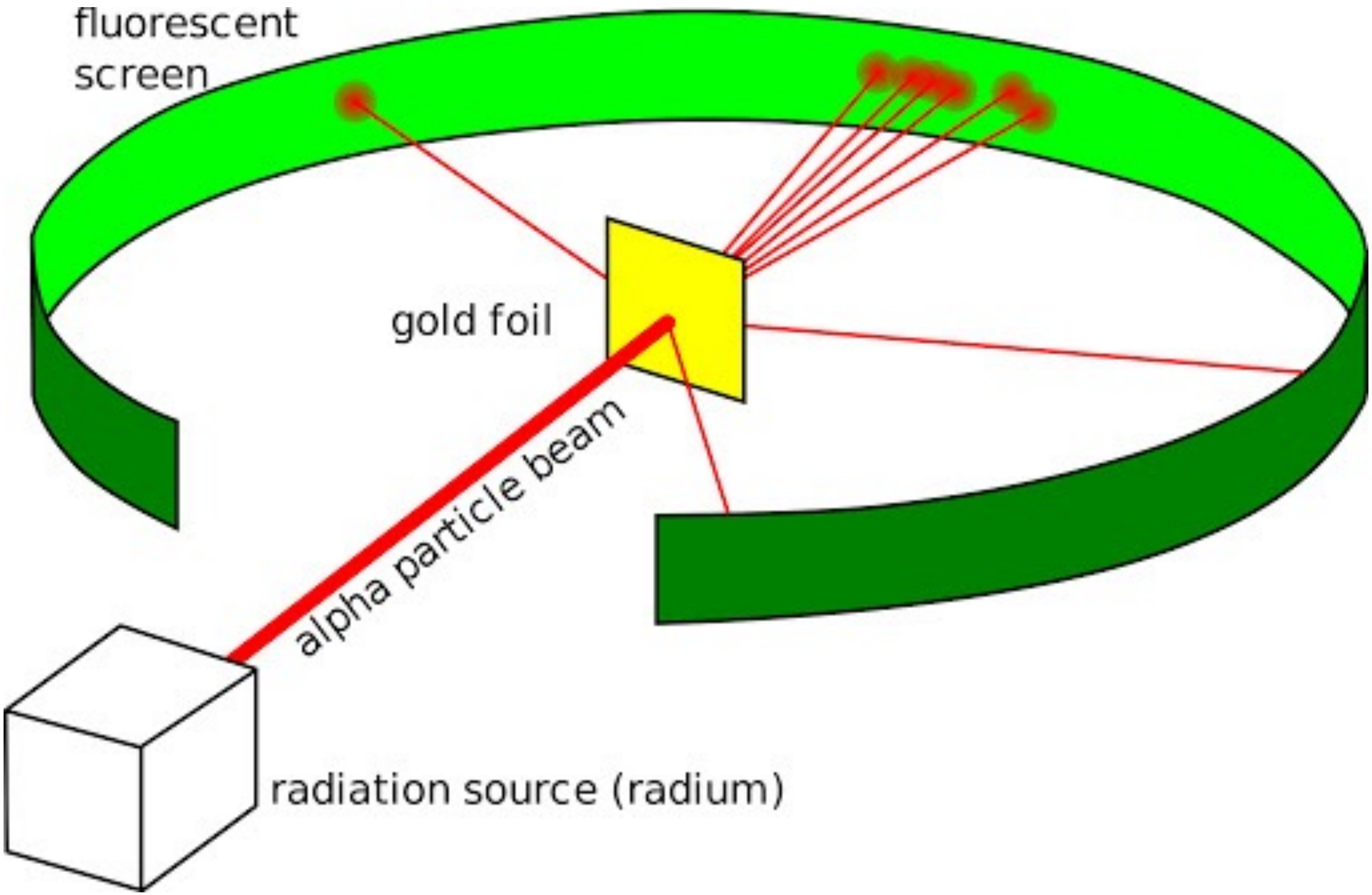
Equipment:

1. A metal plate covered with zinc sulfide (zinc sulfide gives a flash when alpha particles fall on it).
2. A source for alpha particles.
3. Thin slice of gold.

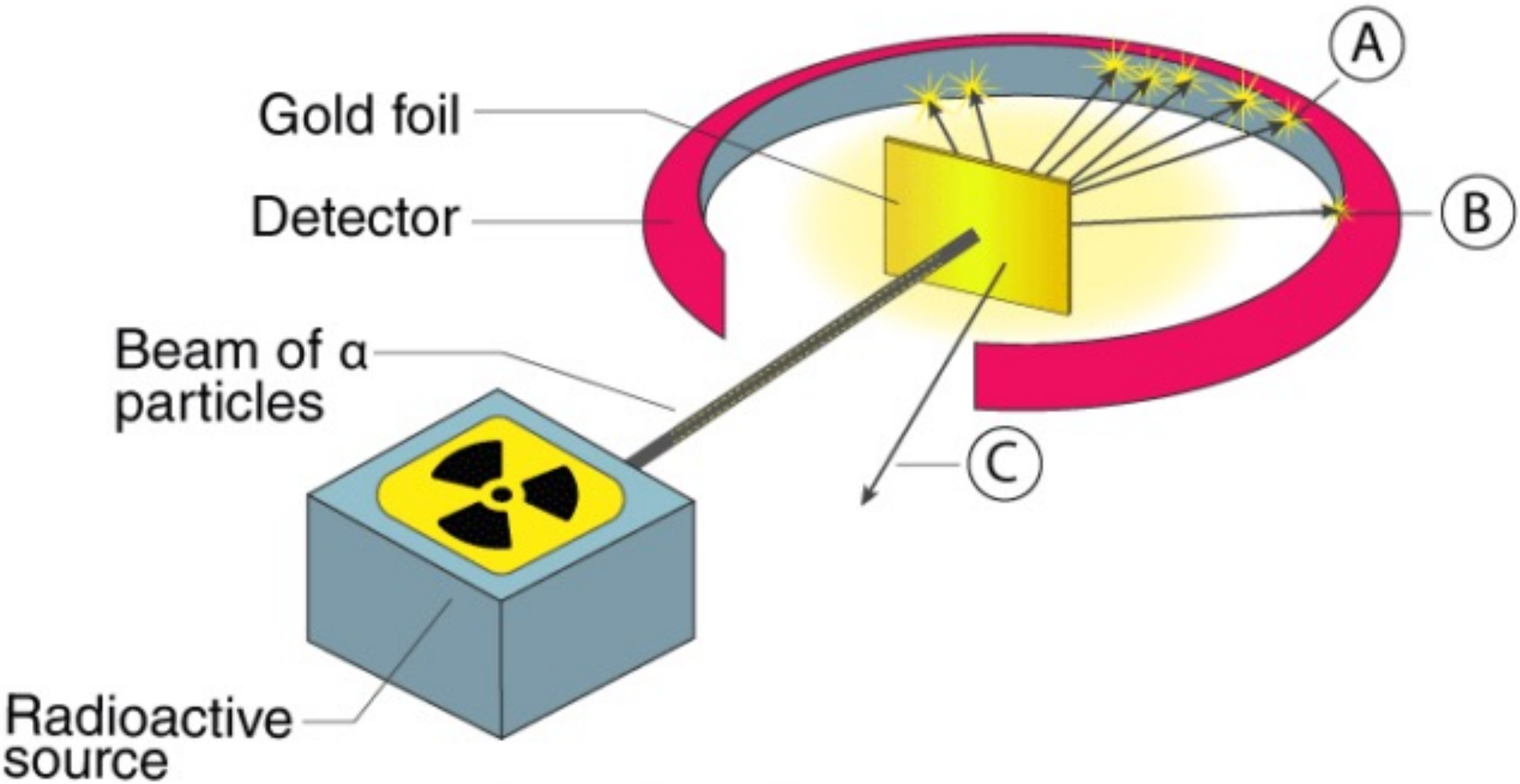
Procedures:

1. The alpha particles were allowed to collide with the metal plate lined with the zinc sulfide layer.
2. The location and number of alpha particles hitting the plate were determined from the flashes in the absence and presence of a thin plate of gold.

Rutherford Model for Atomic Structure.

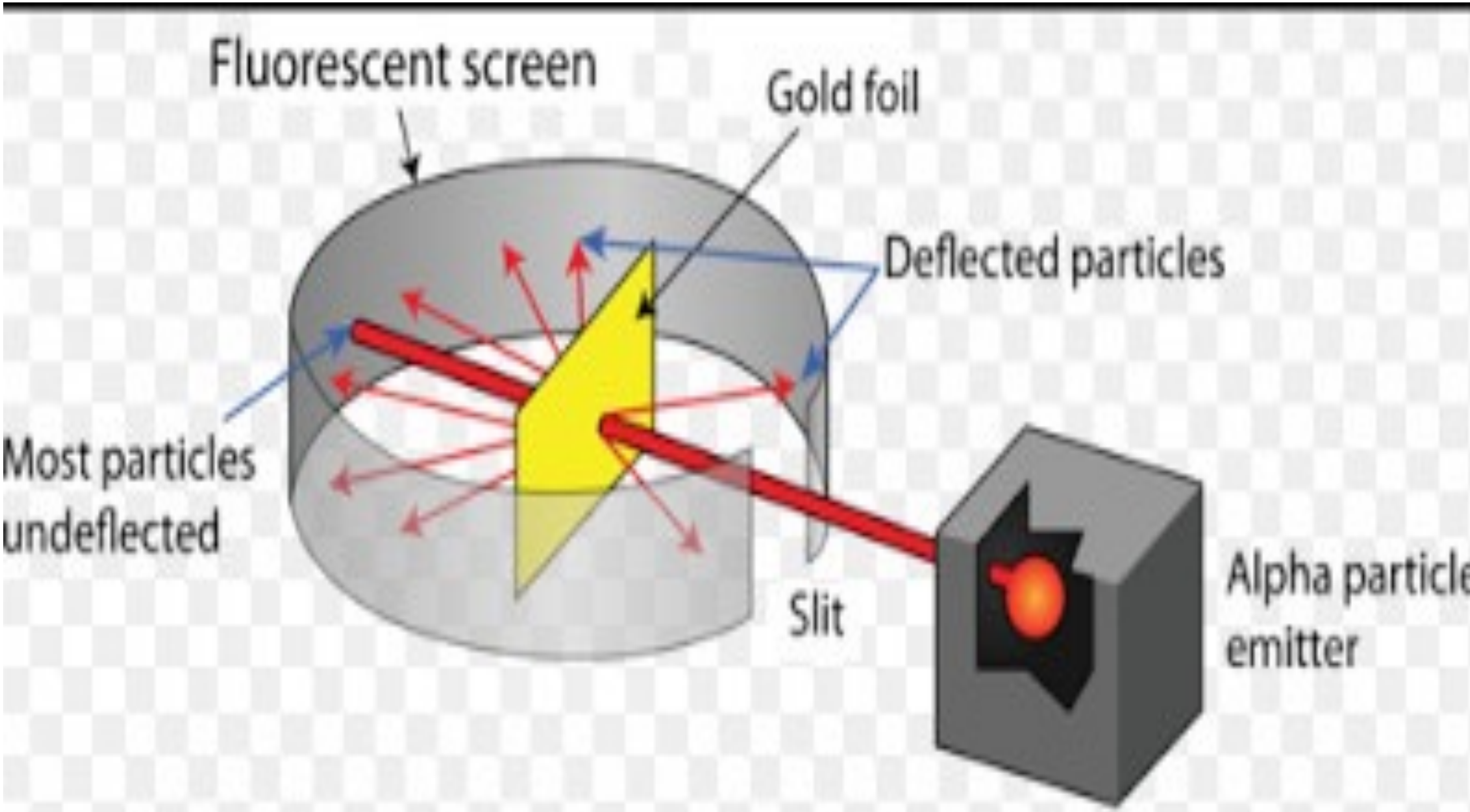


Rutherford Model for Atomic Structure.



Rutherford's Experiment

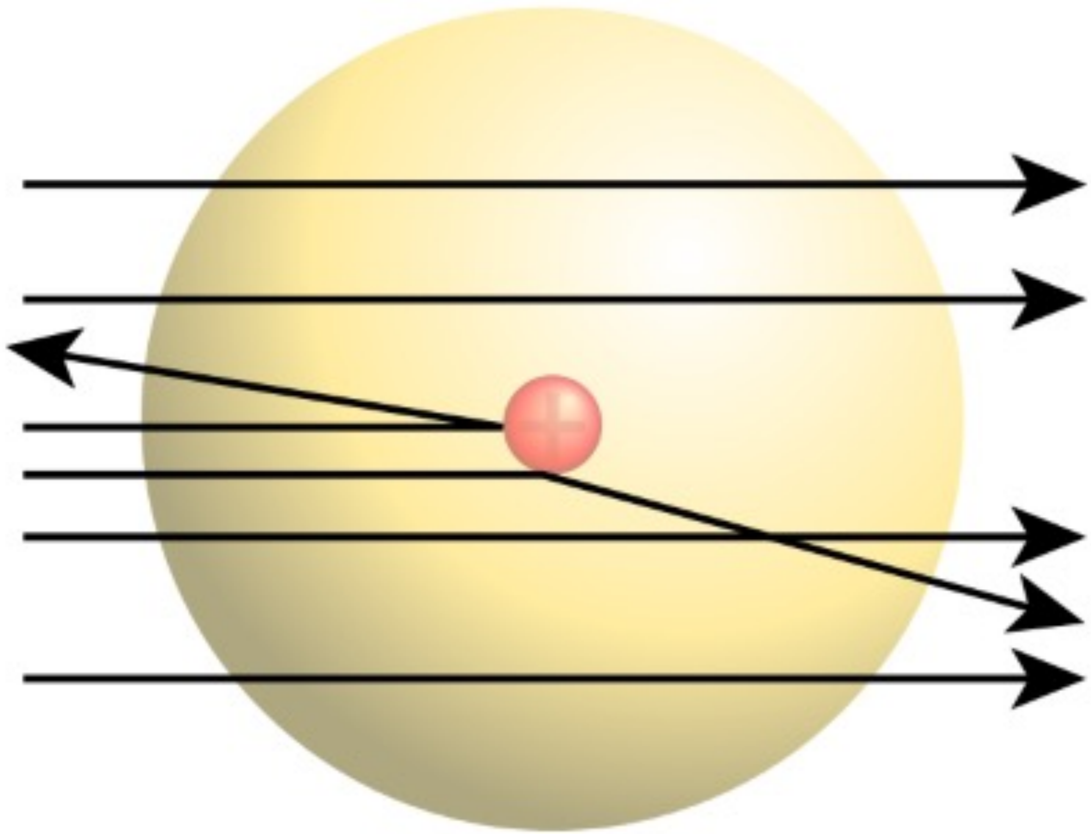
Rutherford Model for Atomic Structure.



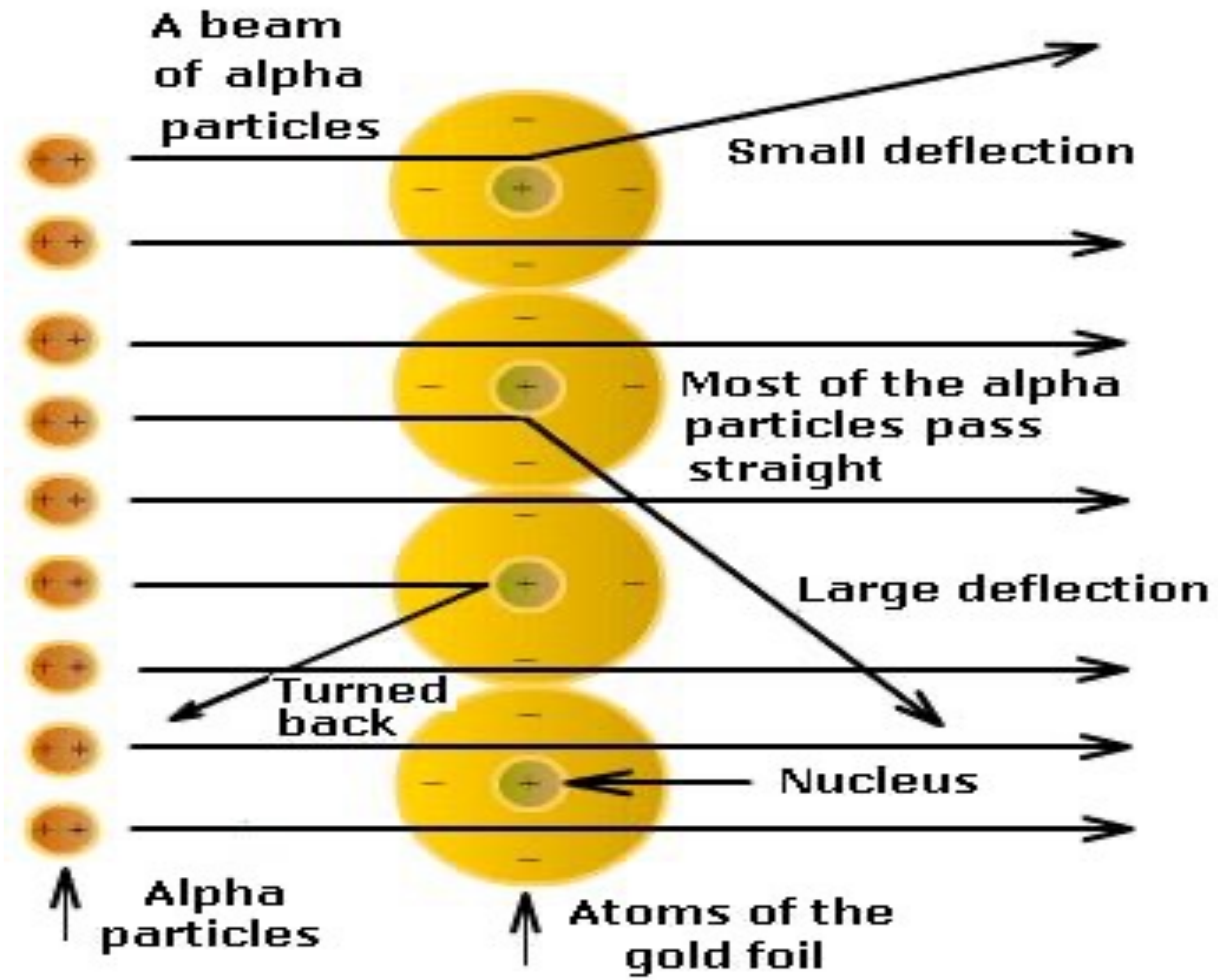
Notices of Rutherford Experiment:

1. Most of the alpha particles passed straight through the foil without any deflection from their path.
2. A small fraction of them was deflected from their original path by small angles.
3. Only a few particles bounced back.

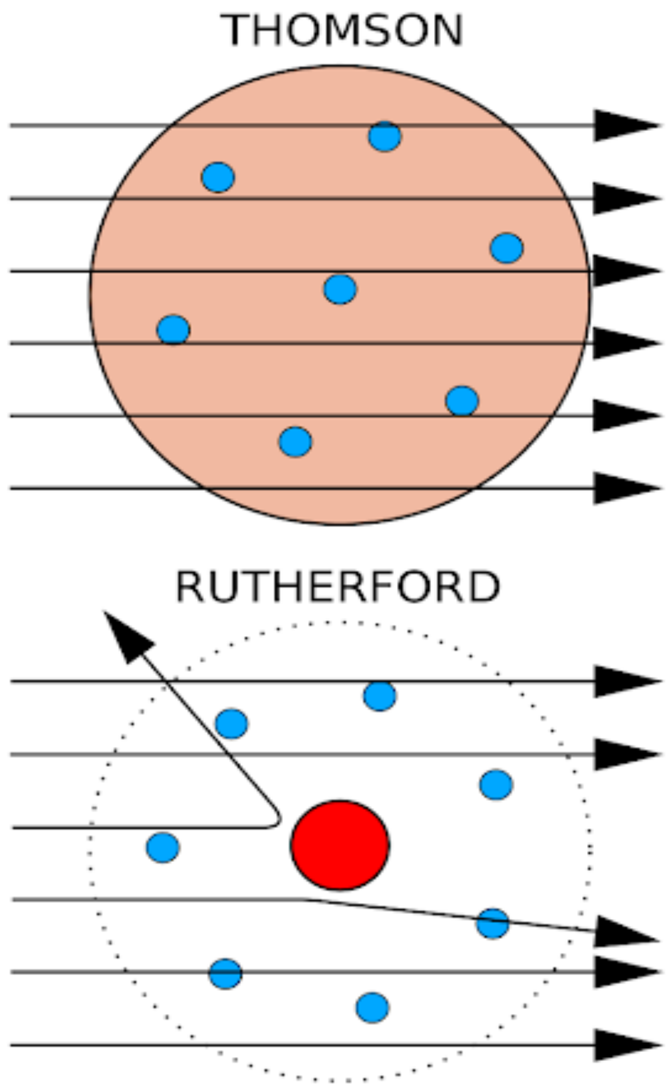
Rutherford Model for Atomic Structure.



Rutherford Model for Atomic Structure.



Rutherford Model for Atomic Structure.

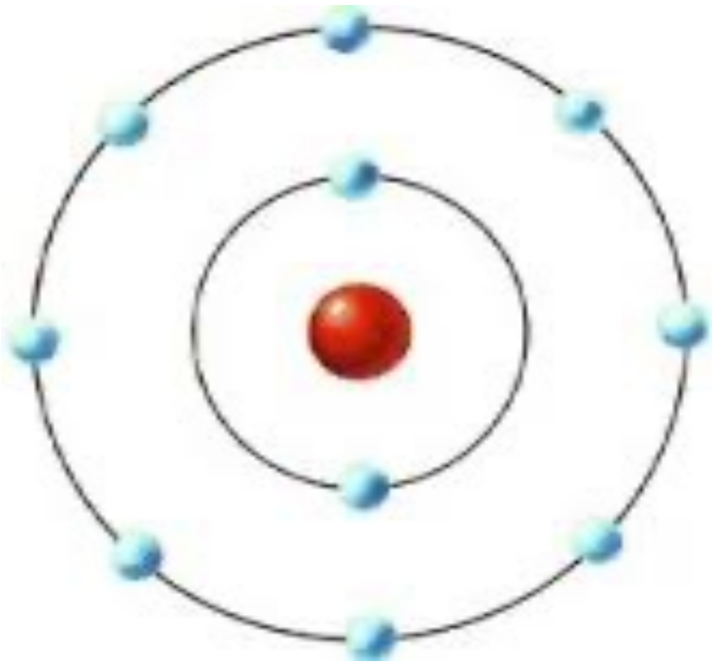


Rutherford Model:

1. Atom is a huge vacuum contains a central nucleus surrounded by negative electrons rotates around the nucleus and far away from it.
2. The atom mas is mainly contained at the nucleus which contains the positive charges as well.
3. Atom is electrically neutral, because it contains equal numbers of positive charges, at nucleus, and negative charges (electrons).

Rutherford Model:

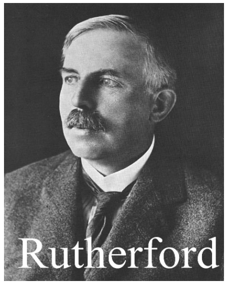
4. As an electron rotates around the nucleus, a centrifugal force is created that is equivalent to the electron's attraction to the nucleus.



Rutherford Atom

Discovering positive and neutral particles (proton, positron and neutron)

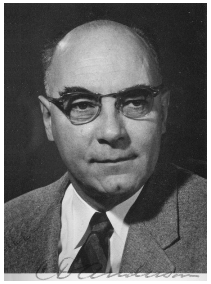
❖ Protons are positive and much heavier than electrons (Rutherford 1920).



❑ A neutron has a slightly heavier mass than a proton but is electrically neutral (Chadwick 1932).

Chadwick

✓ A positron has the same mass as an electron but is positively charged (Andersen 1932).



Andersen

Quantum theory and the atomic spectrum

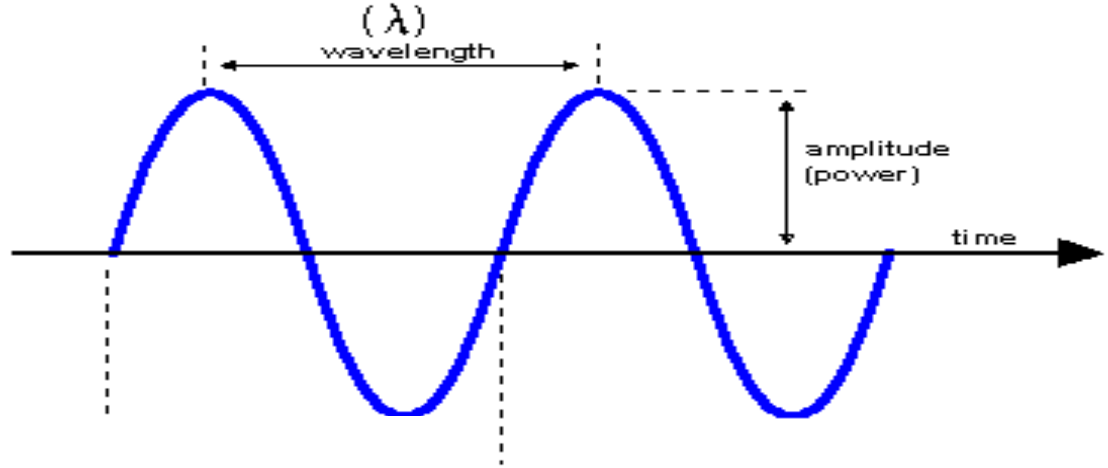
- Electromagnetic radiation consists of an electric and magnetic field that oscillate perpendicular to the direction of propagation of the radiation.
- The radiation travels in waves having a wavelength λ and a frequency γ.
- **Wavelength (λ) :**

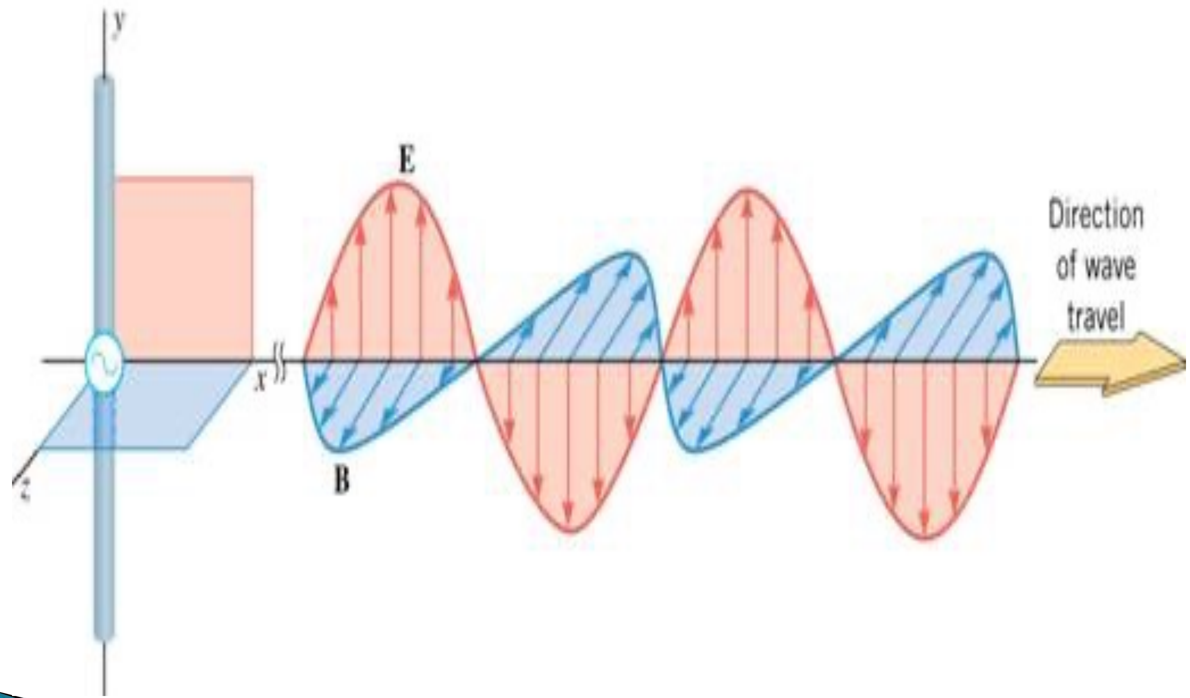
It is the distance between two consecutive peaks or troughs.

Frequency (γ):

It is the number of vibrations produced for any object per second.

Quantum theory and the atomic spectrum





Max Planck proved that an electromagnetic ray is not a flux or a constant current, but rather it is in the form of a quantum or quantum of energy. And it has an energy (E) estimated by Max Planck's law

$$E = h\gamma$$

γ = frequency **h = Max Plank's constant**

Einstein suggested that the beam consists of infinitesimal minutes that he called photons having energy:

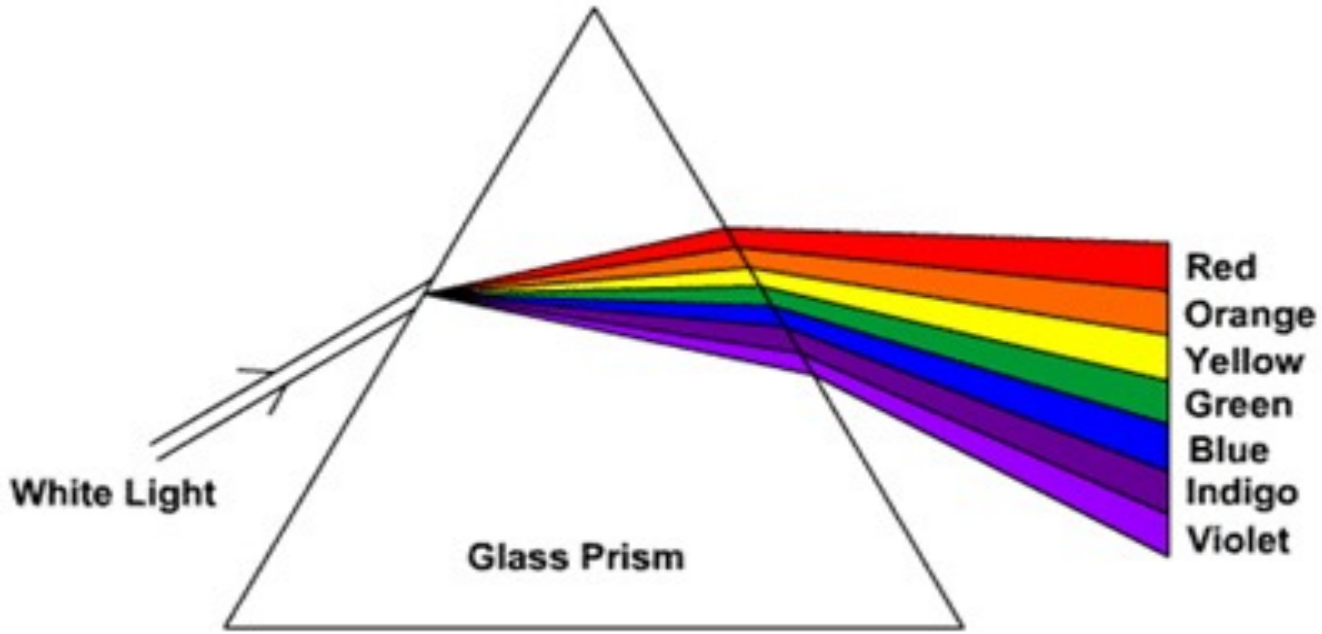
$$E = mc^2$$

E = Photon's Energy m = Photon's mass. c = Light speed.

- When gas molecules or atoms of the element are heated, it produces electromagnetic radiation, when analyzing it, we get a number of parallel spectrum lines in different regions (of different wavelengths).
- This spectrum is known as linear atomic spectrum and it varies from element to element.
- The complexity of the atomic spectrum increases with the number of electrons of the element.
- There is a relationship between the atomic spectrum of excited atoms and their internal structure.

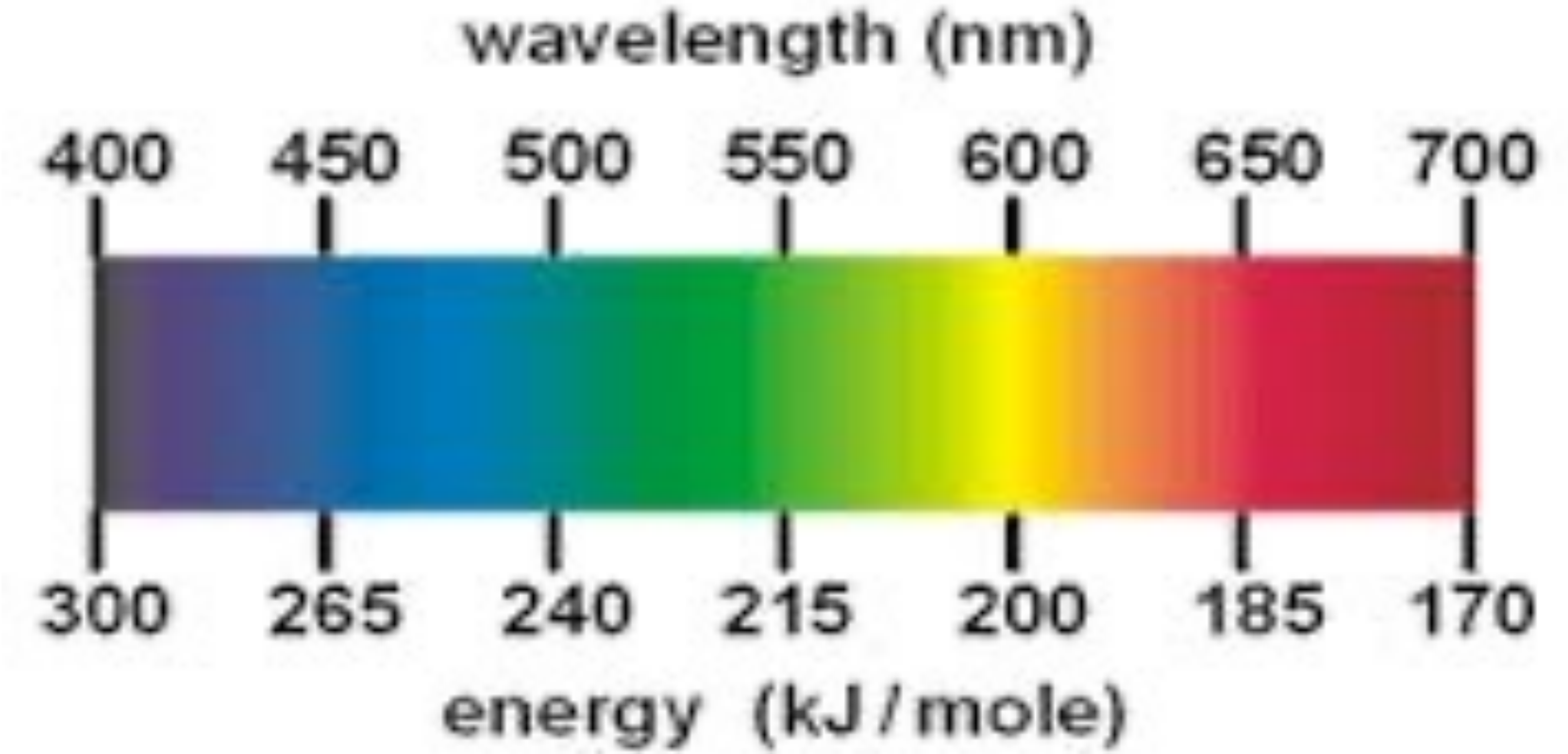
Spectra resulting from the scattering of the white ray through the glass prism:

Red - orange - yellow - green - blue - indigo - violet

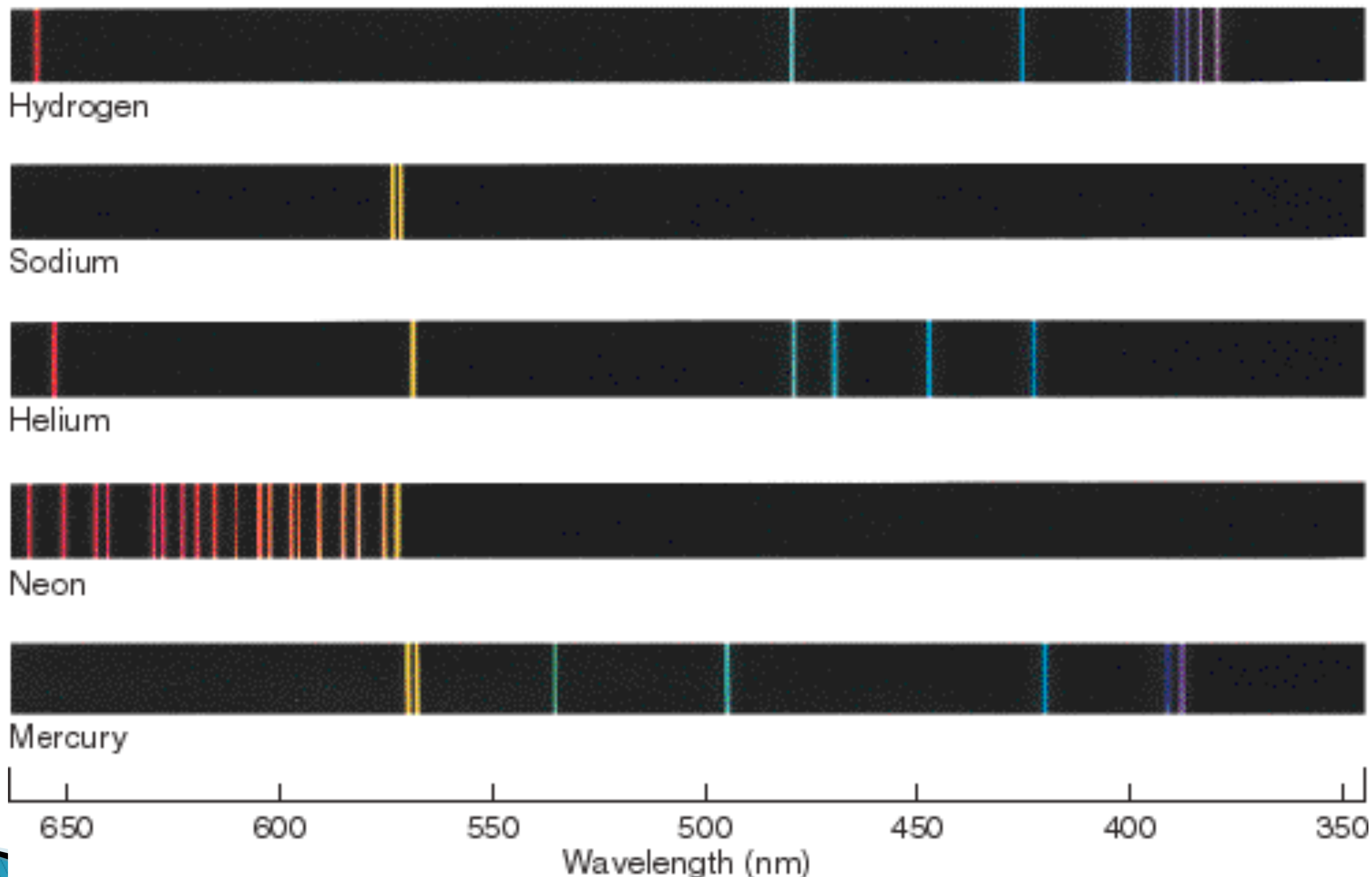


There is a relationship between the atomic spectrum of excited atoms and their internal structure.

Quantum theory and the atomic spectrum



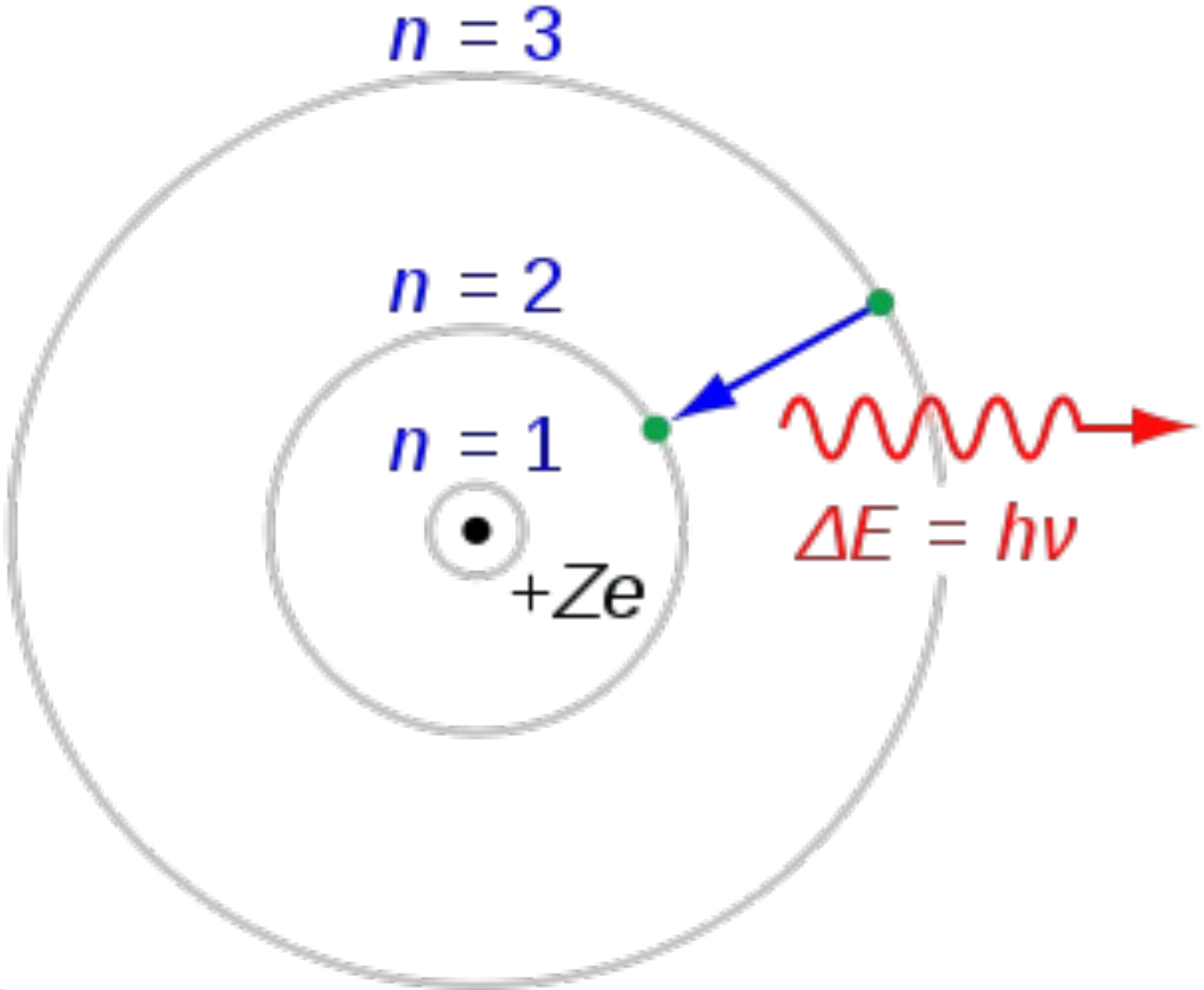
Benefits of the emission spectrum include identification of elements and compounds (Linear spectrum of elements such as a human fingerprint).



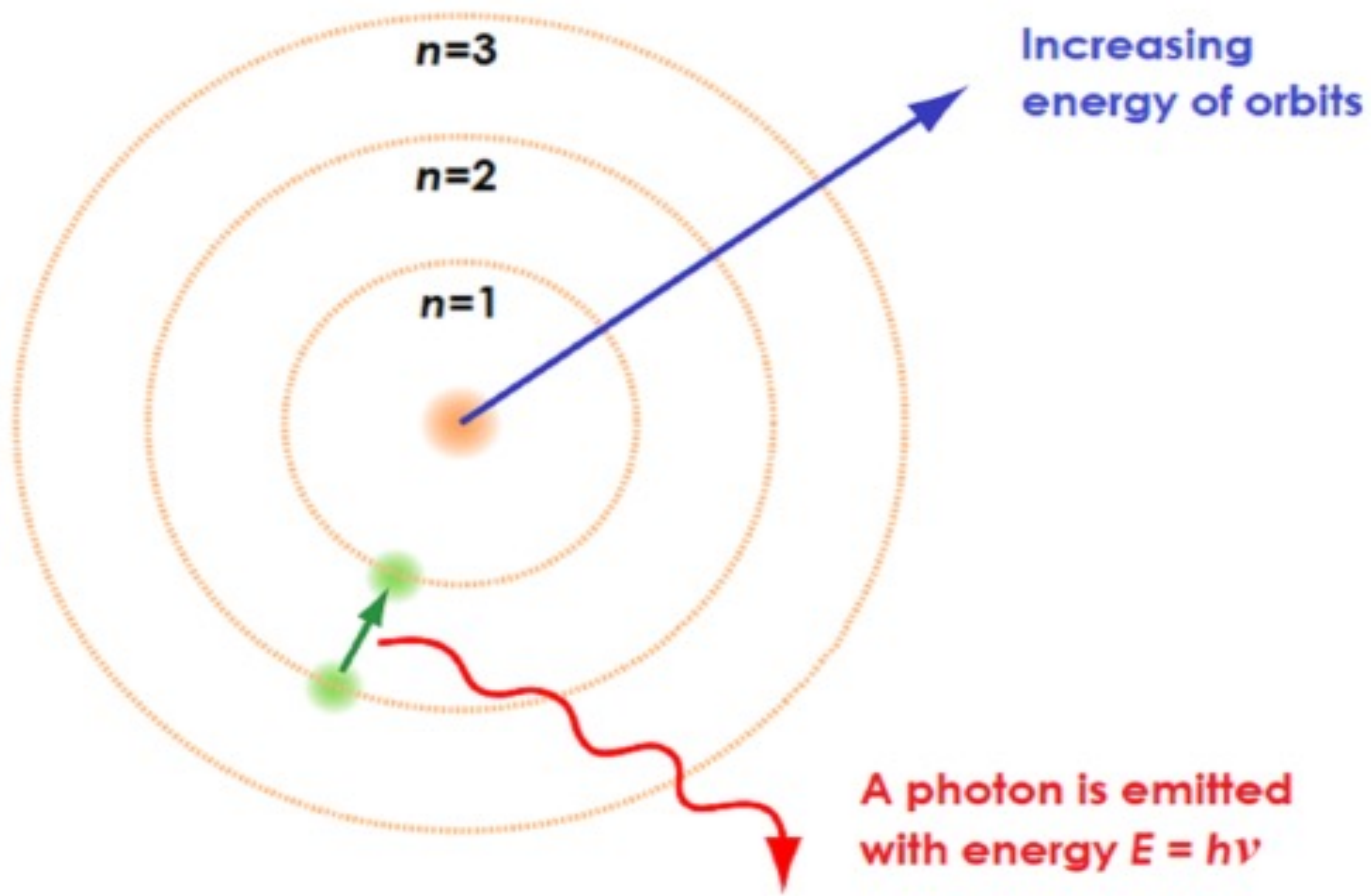
Bohr Atomic Model

- The electron orbits the nucleus at a finite number of constant and specific energy levels without losing or gaining energy in the normal state of an atom.
- Each electron during its rotation around the nucleus has a specific energy that depends on the distance of the energy level in which it rotates from the nucleus, as the energy level increases with the increase of its radius.
- The largest number of energy levels in the normal state of an atom, 7 levels, expresses the energy of each level with an integer number called the principal quantum number.

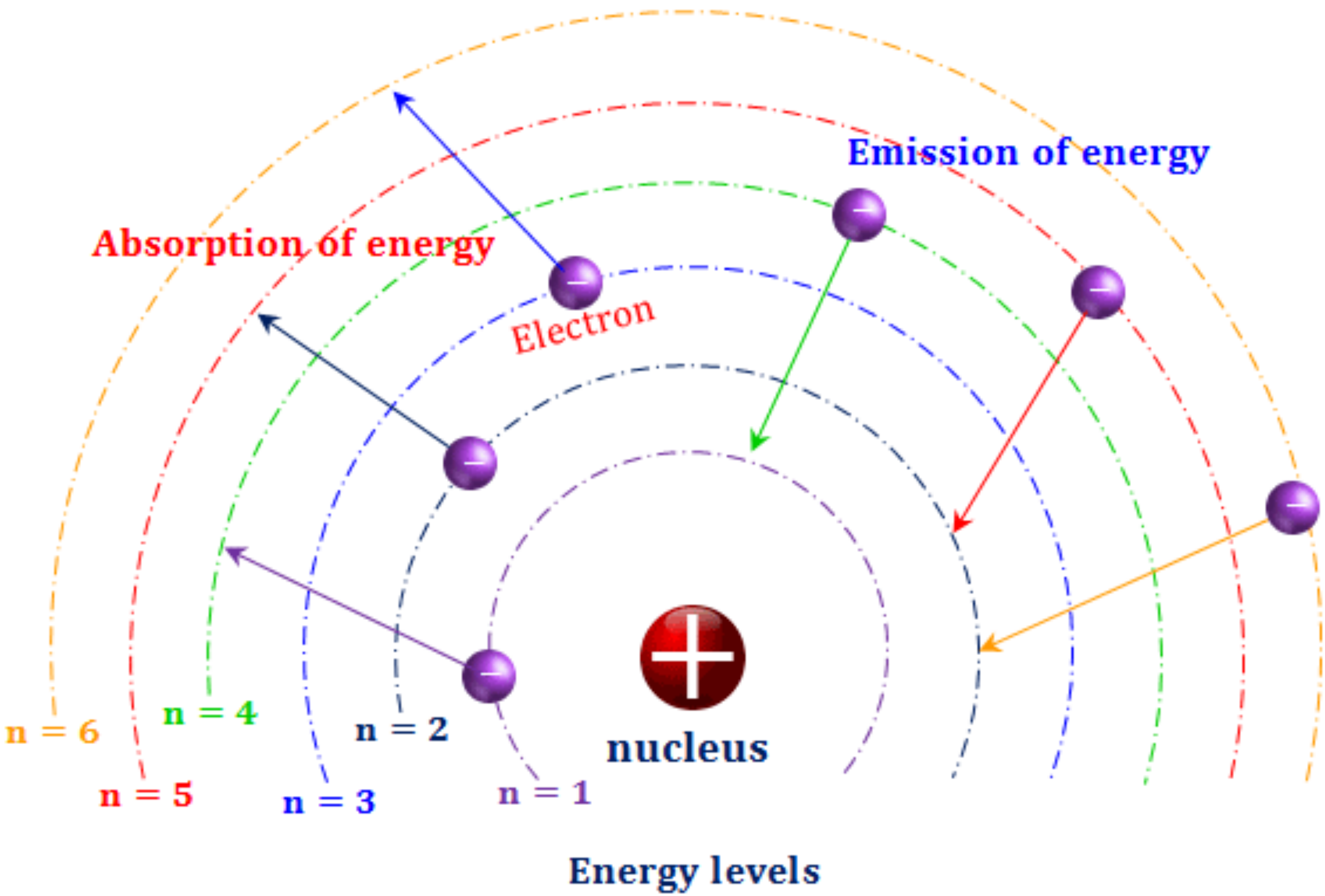
Bohr Atomic Model



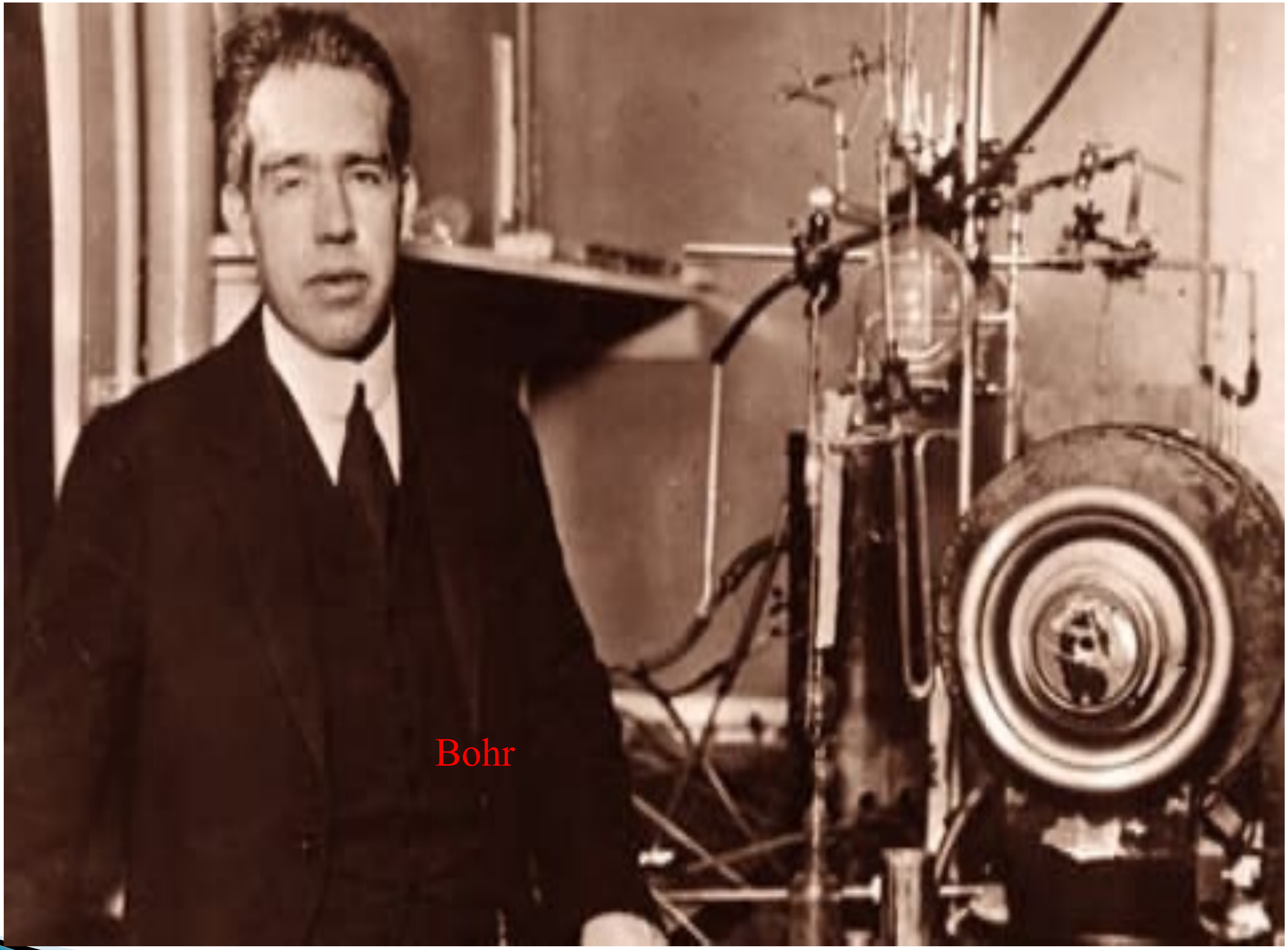
Bohr Atomic Model



Bohr Atomic Model



Bohr Atomic Model



Bohr Atomic Model

- In the stable state of the atom, the electron rotates at the energy level appropriate to its energy, and when the atom is excited, the electron temporarily jumps from its energy level E_a to a higher level E_b, and when it loses the excitation energy, the electron returns to its original location, giving radiation of a distinct wavelength and frequency for each element.
- The energy of the emitted photon is in the form of an electromagnetic spectrum equal to the energy difference between the two planes:

$$E_b - E_a = h\gamma$$

Thus, Bohr was able to explain the origin of the linear spectrum of the elements.

Bohr Atomic Model

- An electron does not move from one level to another unless the amount of energy gained or lost is equal to the energy difference between the two levels.
- As an electron rotates around the core, its angular momentum (mvr) is a multiple of the value (h / 2π)

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

Where: h=Planck's constant, m = the mass of the electron, v = the velocity of the electron, r = the radius of the orbital, and n = the number of the orbit.

- So, Bohr was able to calculate the radii of electronic orbits and the total energy Of the electron and the energy differences between the orbitals.

Disadvantages of the Bohr model

- The theory refers to knowing the exact position of the electron and its velocity, which is not possible and does not agree with Heisenberg's principle of uncertainty, which states that it is impossible to know at the same time the amount of motion and location of a moving particle.
- Bohr was not interested in the wave nature of the electron and considered it only a physical body, and it has been proven that electrons are similar to light in that they have a dual nature (wave and particle).
- Bohr used flat planes (that is, in one plane), which means that the hydrogen atom is flat. This is the opposite of the truth, as the atom has three directions in space.
- Bohr's theory did not succeed in explaining the more complex spectra of the hydrogen atom, i.e. the spectra of atoms with more than one electron.

The efforts that led to the development of Bohr's theory and the arrival of the modern atomic theory

1. The wave nature of the electron: (de-Brawley's equation):

- An electron, like light, diffuses and interferes, so it has a dual nature (particle and wave) and not just a particle.
- Scientist De Brawley was able to derive an equation to calculate the wavelength of an electron wave.

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

- Where; h = Planck's constant, m = the mass of an electron, and v = its velocity.

- mv expresses the particle of the electron, and λ expresses its wave nature.
- Electron diffraction has great uses in the study of crystal and molecular structure and imaging.

2. Heisenberg's Uncertainty Principle

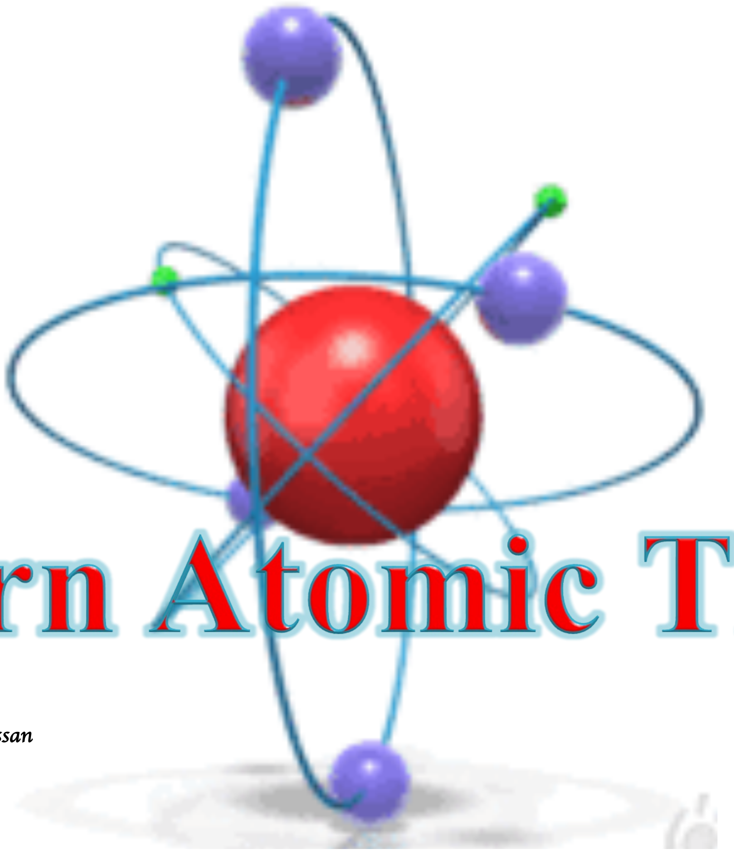
It is impossible to know at the same time the momentum and location of a moving particle.

$$(\Delta x)(\Delta\, mv) \geq \frac{h}{2\pi}$$

Δx = the amount of uncertainty with respect to the electron's position.

Δmv = the uncertainty of the amount of traffic.

That is, if the location of the electron is precisely determined, it is not possible to accurately determine the amount of its motion, and vice versa.

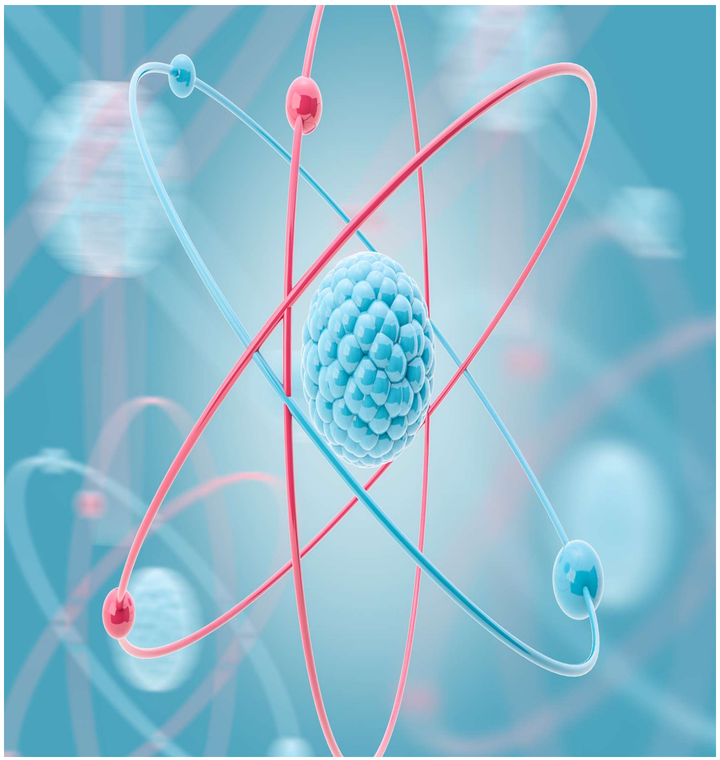


Modern Atomic Theory

Dr. Ibrahim A.I. Hassan

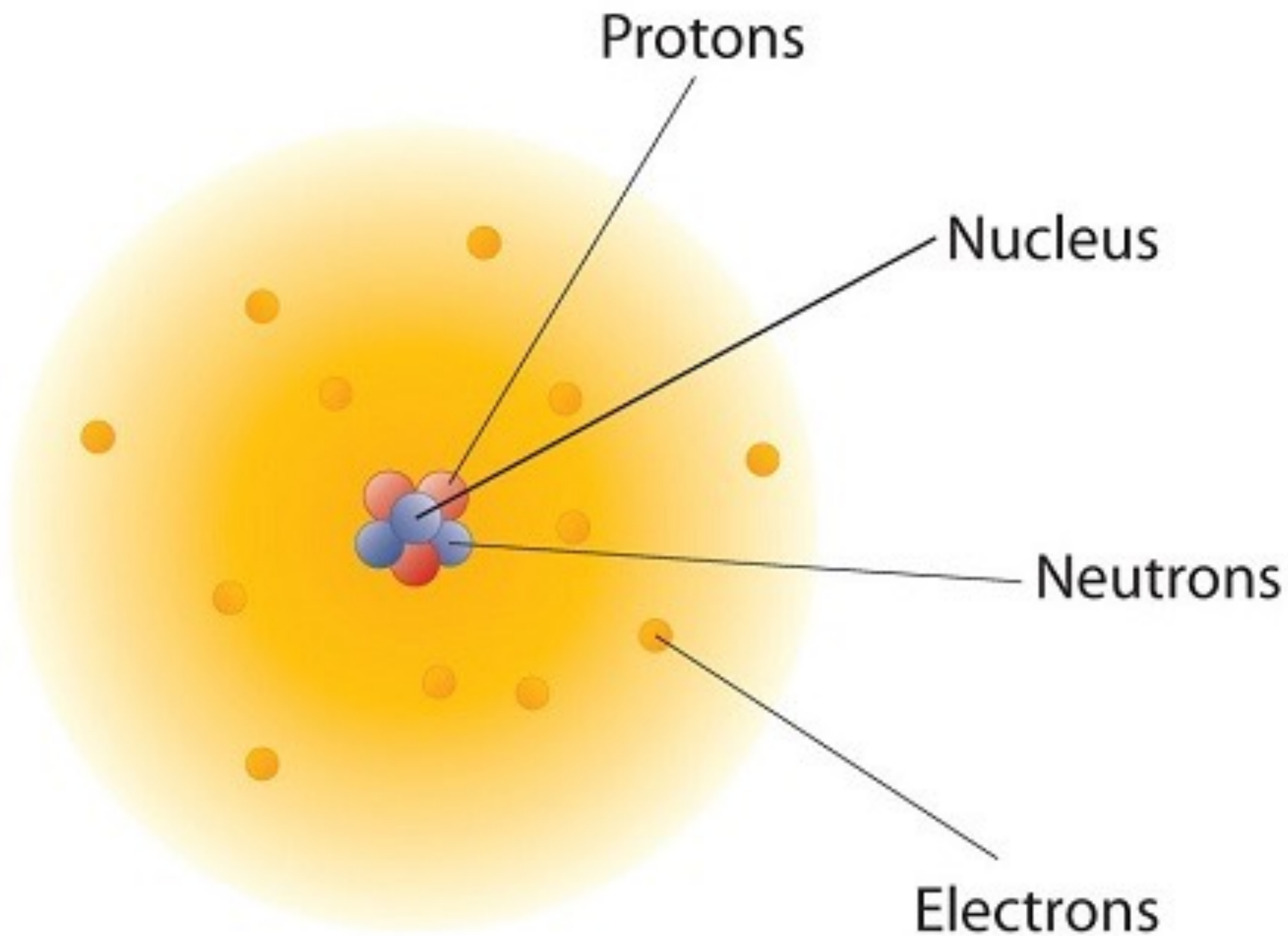
Modern Atomic Theory

- An atom consists of a nucleus containing a positive charge and most of the mass of an atom is concentrated in it.

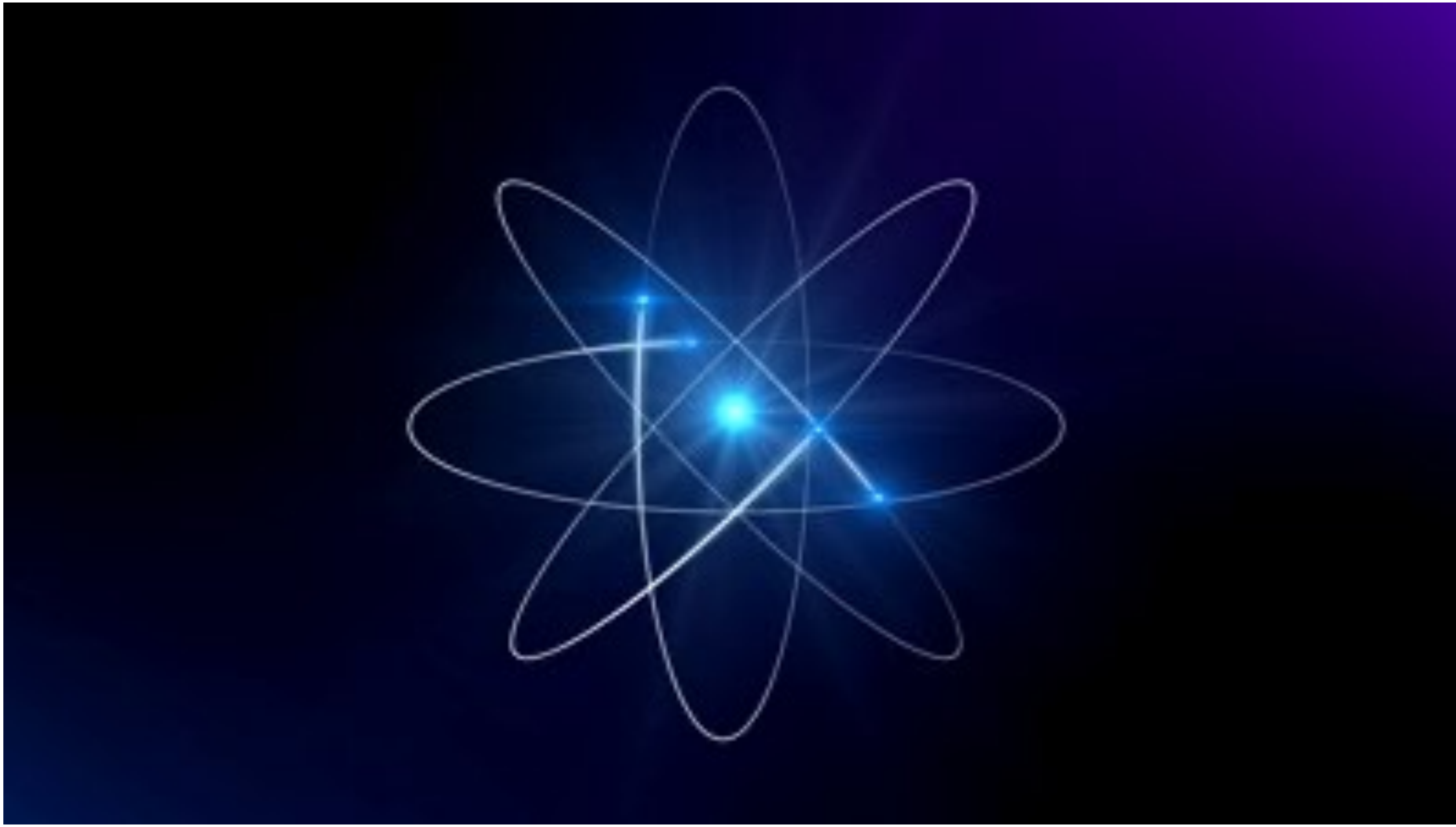


- The nucleus is surrounded by negatively charged electrons that move very quickly and have the properties of waves.
- Electrons occupy the void regions around the nucleus and have certain energies and are found in what is known as an electronic cloud.

Modern Atomic Theory



Modern Atomic Theory



Quantum numbers:

It describes the movement of electrons in an atom and determines its location.

1. **Principal quantum number (n):**

- It indicates the energy level.
- Shows the size of the electronic cloud.
- Its value is 1 - 7 and does not take the value zero.
- The seven energy levels take the symbols:

K, L, M, N, O, P, Q.

Quantum numbers and interstitial form of rotations

1. **Principal quantum number (n):**

- The number of electrons by which each principal energy level is saturated is twice the square of the shell number ($2n^2$).

e.g. The fourth level is saturated with 32 electrons ($2 \times 4^2 = 32$).

- This law does not apply to energy levels higher than the fourth because an atom becomes unstable if the number of electrons in any level exceeds 32 electrons.

Quantum numbers and interstitial form of rotations:.

2. Secondary quantum number (ℓ):

- Indicates the location of the sub-energy level (below the principal energy level).
- It is a number whose value determines the shape of the shell in which the electron is moving.
- Its number equals the number of its main level.
- The number of sub-energy levels does not exceed four levels.
- The four sub-energy levels have the symbols S, P, d, and F.
- They differ in energy: F> d> P> S.
- A positive number less than n and takes the values (0, 1, 2, 3)

$$\leq \ell \leq n-1$$

Quantum numbers and interstitial form of rotations:.

3. Magnetic quantum number (m):

- Determines the shape and orientation of the orbital in space.
- Number of directions per orbital = ($2\ell + 1$), where ℓ is the value of the secondary quantum number.
- The magnetic quantum number ranges between $+\ell$ & $-\ell$.
- In the case of the secondary S level it is:

$$m = 2 \times 0 + 1 = 1$$

Therefore, the S level has one-way direction in space and is the sphere around the nucleus.

Quantum numbers and interstitial form of rotations:.

3. Magnetic quantum number (m):

- In the case of the secondary level P:

$$m = 2 \times 1 + 1 = 3 \ \ (+1, 0, -1)$$

- Therefore, the P level has three directions in space.
- It consists of two connected bulges, each of them can accommodate an electron, which is difficult to be present when the bulges meet.

Quantum numbers and interstitial form of rotations:.

3. Magnetic quantum number (m):

- **In the case of the secondary level d:**

m = 2 x 2 + 1 = 5 (+2, +1, 0, -1, -2)

Therefore, the d level has five directions in space.

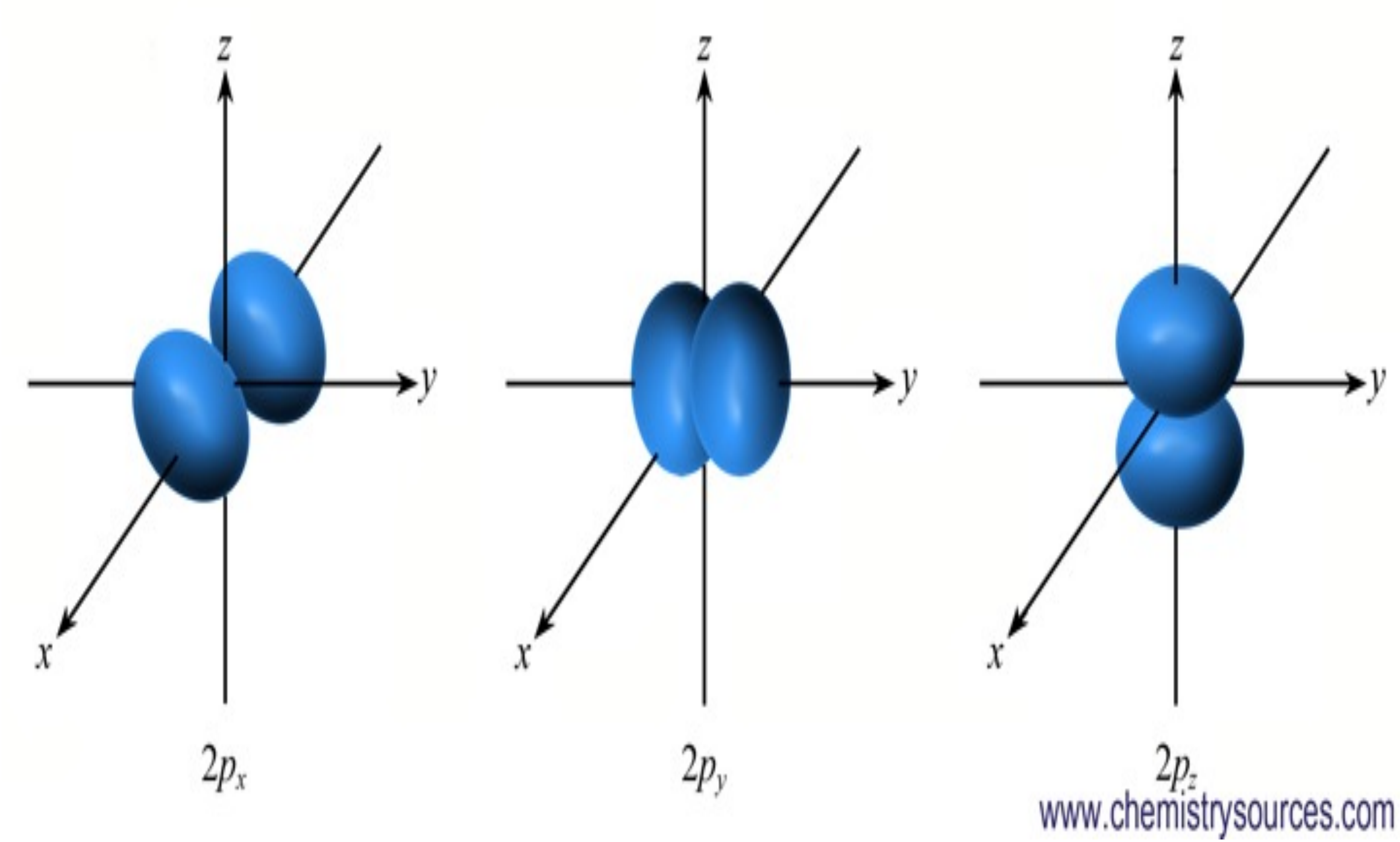
- **In the case of the secondary level f:**

m = 2 x 3 + 1 = 7 (+3, +2, +1, 0, -1, -2, -3)

Therefore, the f level has seven directions in space.

Quantum numbers and interstitial form of rotations:.

3. Magnetic quantum number (m)



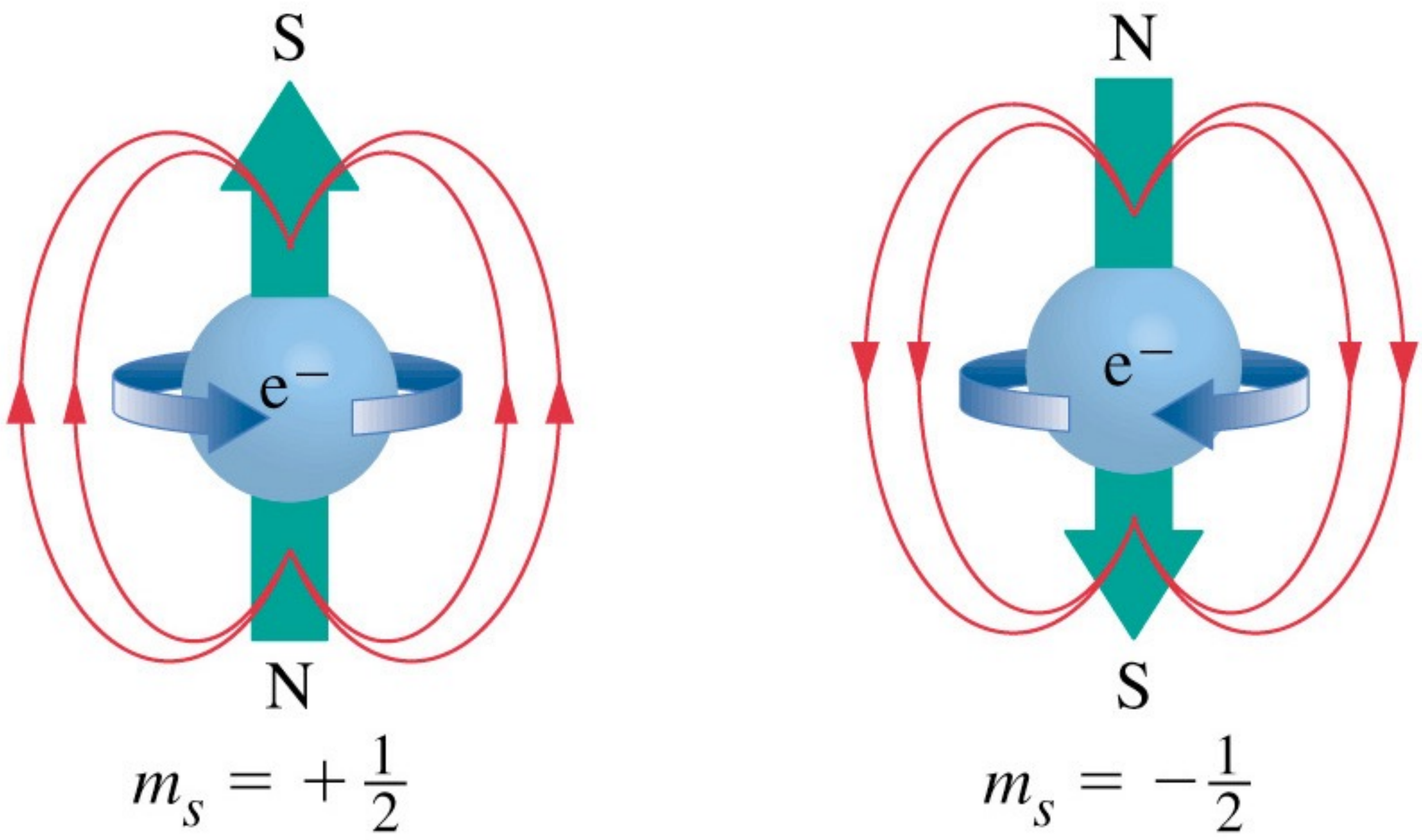
Quantum numbers and interstitial form of rotations:.

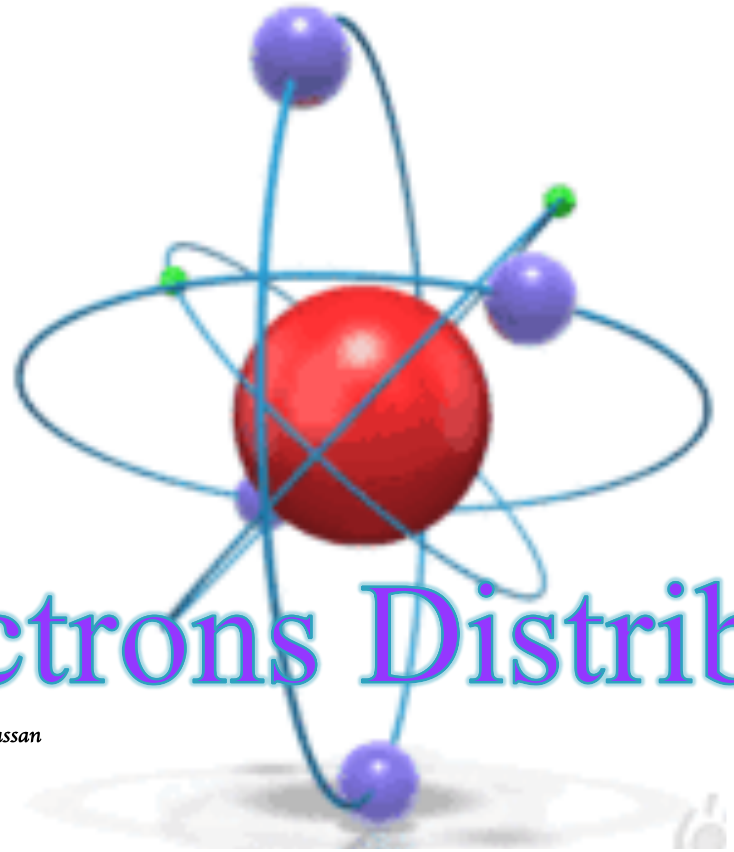
4. Spindle quantum number (s):

- ✓ It is a number that determines the type of spindle movement of an electron around its axis.
- ✓ Every electron has two movements: the first around the nucleus ... and the second around itself and is called the spindle movement.
- ✓ The spindle movement of one of the two electrons is counter to the movement of the other electron, which reduces the repulsive forces between them.
- ✓ The quantum number has two values: -1/2, +1/2

Quantum numbers and interstitial form of rotations:.

4. # Spindle quantum number (s)





Electrons Distribution

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The probability distribution of electrons in an atom:

1. **Principle of construction upward:**

Electrons fill the lower-energy levels first, then the higher-energy levels.

For example:

^{30}Zn

$$1S^2 \rightarrow 2S^2 \rightarrow 2p^6 \rightarrow 3S^2 \rightarrow 3p^6 \rightarrow 4S^2 \rightarrow 3d^{10}$$

The probability distribution of electrons in an atom:

2. **Hund's rule:**

At a certain sublevel, duplication does not occur between two electrons until their individual orbitals are operated first.

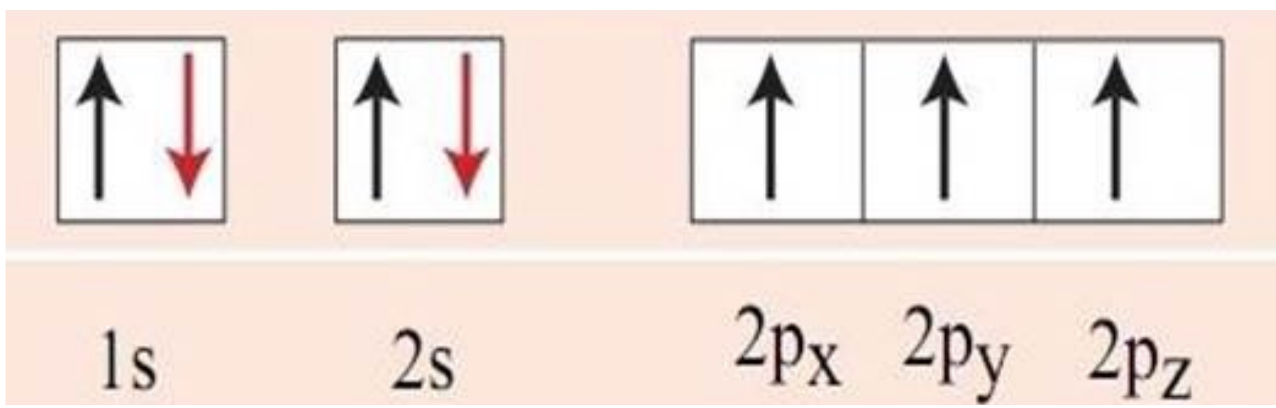
The probability distribution of electrons in an atom:

1. **Principle of construction upward:**
2. **Hund's rule:**

Example: ^{7}N

Principle of construction upward:- $1s^2\ 2s^2\ 2p^3$

Hund's rule:

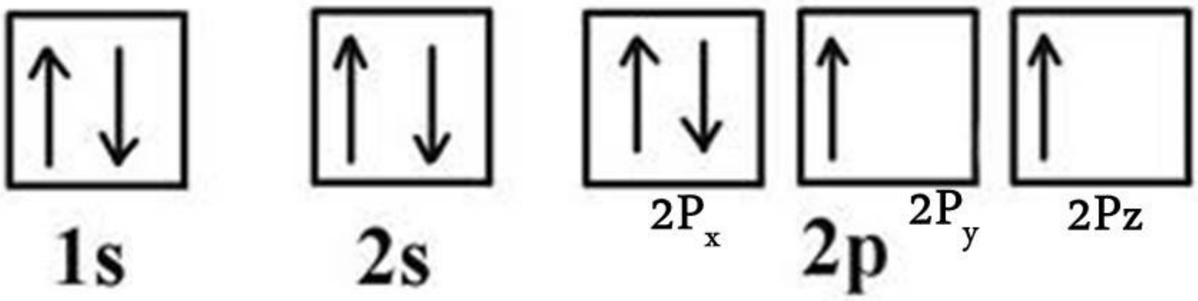


The probability distribution of electrons in an atom:

Hund's rule:

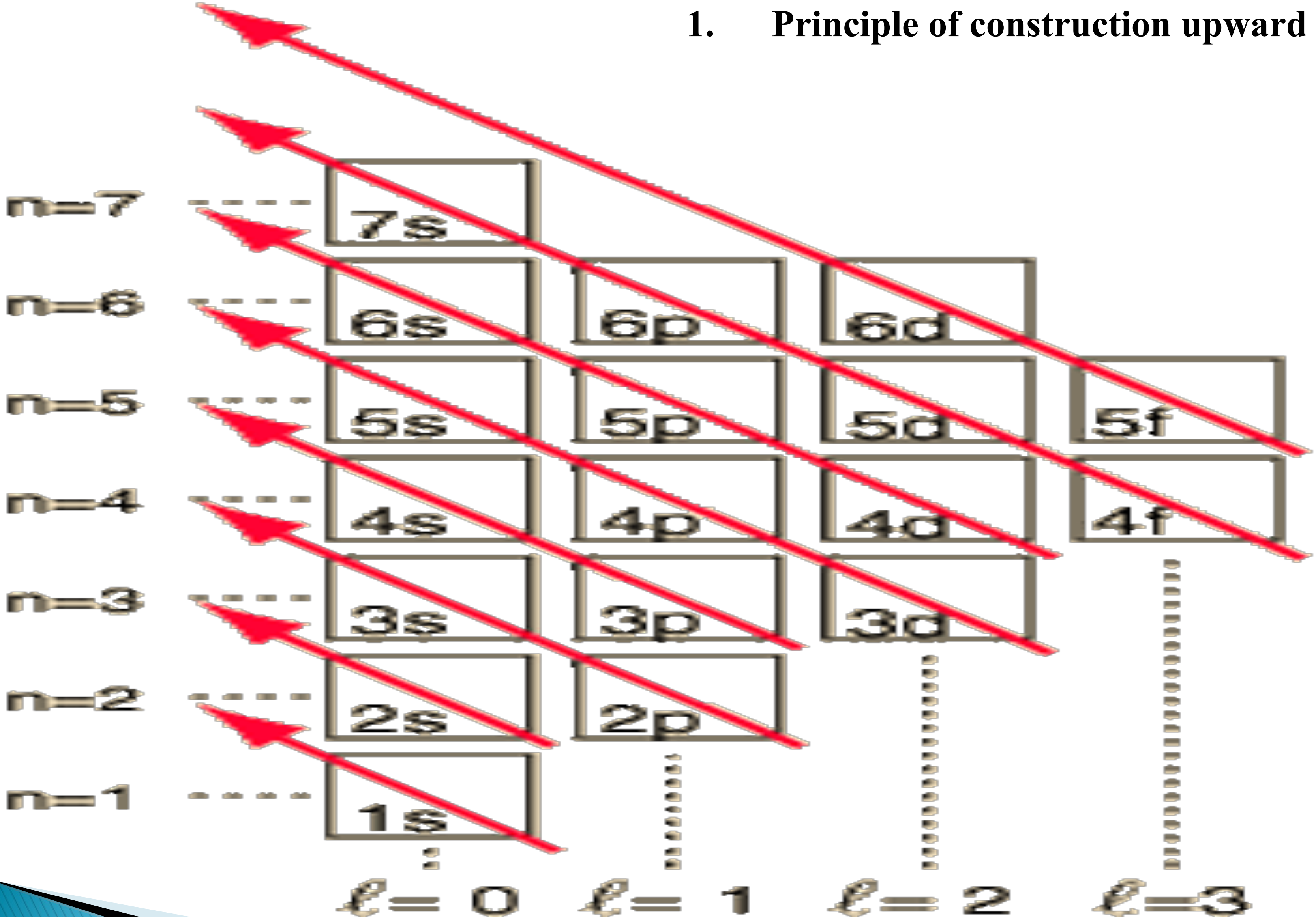
Example:

^{8}O Oxygen

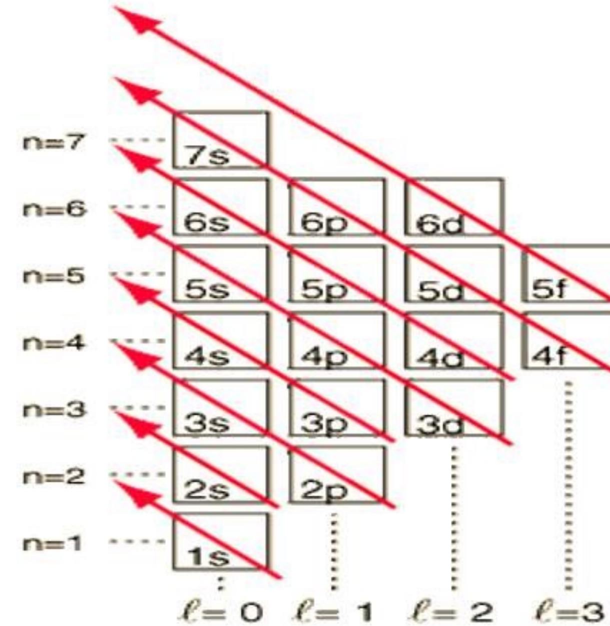


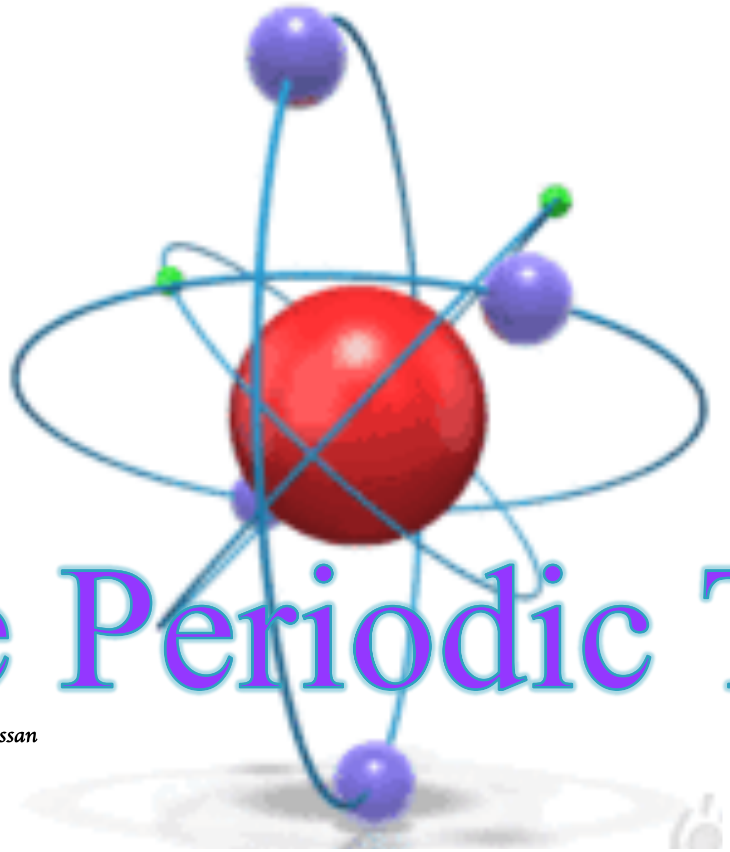
The probability distribution of electrons in an atom

1. **Principle of construction upward**



n =	1	2	3	4	5
^{2}He =	$1S^2$				
	2				
^{10}Ne =	$1S^2$	$2S^2$ $2P^6$			
	2	8			
^{18}Ar =	$1S^2$	$2S^2$ $2P^6$	$3S^2$ $3P^6$		
	2	8	8		
^{36}Kr =	$1S^2$	$2S^2$ $2P^6$	$3S^2$ $3P^6$	$4S^2$ $3d^{10}$ $4P^6$	
	2	8	18	8	
^{54}Xe =	$1S^2$	$2S^2$ $2P^6$	$3S^2$ $3P^6$	$4S^2$ $3d^{10}$ $4P^6$	$5S^2$ $4d^{10}$ $5P^6$
	2	8	18	18	8





The Periodic Table

Dr. Ibrahim A.I. Hassan

The Periodic Table

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57-71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

The Periodic Table and the Properties of the Elements

- **The elements are arranged within a table so that each element exceeds the one preceding it with a single electron.**
- **This table was called the Periodic Table of the Elements.**
- **In the Periodic Table of Elements, the chemical properties of the elements of one group are similar.**

According to the electronic distribution and the number of electrons in the outer orbit, known as the valence shell, the elements were divided into four groups:

1. **Group or block S:**

Includes hydrogen, alkaloid group, and alkaloid earths, where outer electron level begins with filling the S orbital and ends with its completion.

2. **Group or block p:**

In which the electron begins entering the p-level and ends with its completion in the inert gas.

3. **Group or block d:**

It is the set of elements in which the **d** orbital is filled, and it is called the transition elements, and there are three groups of them:

3. **Group or block d:**

A. **The first series of transition elements; in which the 3d full-level sublevel continues.**

B. **The second transitional chain of elements; in which the 4d sublevel is sequentially filled.**

C. **The third transition chain of elements; in which the 5d sublevel is sequentially filled.**

4. Group or f block:

It includes the elements in which the f block is filled, and these elements are known as rare-earth elements and have the characteristics of the transition elements, and the f-orbital in these elements is not completely filled.

a) **The series of Lanthanides with which the 4f block is filled.**

b) **The chain of Actinides in which the 5f block is filled.**

The Periodic Table

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57-71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

First Level (Freshman) - Course's Contents of Physical Science and Geoogical Science Programs.

Chm 101: General Chemistry (I) - 3 Credits (Lecture 2 Hrs. / W + Lab. 2 Hrs. / W).

Referances:

- **Paul Monk, Physical Chemistry: understanding our chemical world, John Wiley & Sons Ltd, 2004. Mcmurry Fay, Chemistry, 4th edition.**
- **General Chemistry: Principles and Modern Applications, Ralph Petrucci, F. Herring, Jeffry Madura, Carey Bissonnette, Pearson; 11th Edition (March 31, 2016).**