Department of Mathematics and Natural Sciences

CHE101: Lecture 02

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- Subatomic particles
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Lec 01 Recap….

- The philosophers of ancient Greece believed that everything was made of one or, at most, a few elemental substances (elements), whose properties gave rise to the properties of everything. But…..

Democritus (460–370 bc):

- the father of atomism,
- took a different approach.
- reasoned that if you cut a piece of, say, aluminum foil smaller and smaller, you reach a particle of aluminum too small to cut
- So matter must be ultimately composed of indivisible particles with nothing but empty space between them.
- He called the particles atoms (Greek atomos, "uncuttable")

Elements, Compounds, and Mixtures: An Atomic Overview

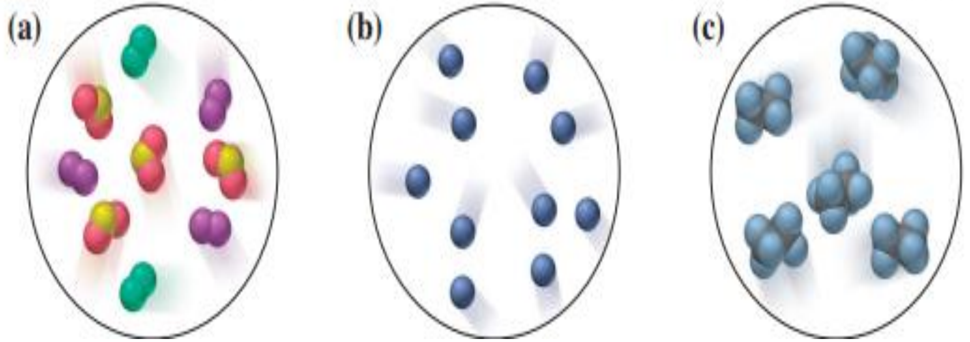
Substance:

- contains only one type of particle, an element or a compound;

- Particles of an element have only one kind of atom (one color), and

- particles of a compound have two or more kinds of atoms always in fixed ratio.

Problem The scenes below represent atomic-scale views of three samples of matter:



Mixture:

- If it contains more than one type, it is a mixture.
- Variable composition and proportion

The Observations That Led to an Atomic View of Matter

- Any model of the composition of matter had to explain two widespread observations known as

(i) the law of mass conservation
(ii)the law of definite (or constant) composition.

As you'll see, an atomic theory developed in the early 19th century explained these mass laws and another now known as the law of multiple proportions

The Observations That Led to an Atomic View of Matter

Mass Conservation

- The most fundamental chemical observation of the 18th century was the law of mass conservation……………

" the total mass of substances does not change during a chemical reaction. The number of substances may change and, by definition, their properties must, but the total amount of matter remains constant"………

- The great French chemist and statesman Lavoisier first stated this law on the basis of experiments

- Mass conservation means that, based on all chemical experience, matter cannot be created or destroyed.

The Observations That Led to an Atomic View of Matter

Definite Composition:

- states that no matter what its source, a particular compound is composed of the same elements in the same parts (fractions) by mass.
- The fraction by mass (mass fraction) is the part of the compound's mass that each element contributes.
- It is obtained by dividing the mass of each element by the mass of the compound.
- The percent by mass (mass percent, mass %) is the fraction by mass expressed as a percentage (multiplied by 100)

Consider calcium carbonate, the major compound in seashells, marble, and coral. It is composed of three elements—calcium, carbon, and oxygen. The following results are obtained from a mass analysis of 20.0 g of calcium carbonate:

Analysis by Mass (grams/20.0 g)	Mass Fraction (parts/1.00 part)	Percent by Mass (parts/100 parts)
8.0 g calcium	0.40 calcium 8/20=0.4	40% calcium
2.4 g carbon	0.12 carbon 2.4/20=0.12	12% carbon
9.6 g oxygen	0.48 oxygen	48% oxygen

$$\text{Mass of element in sample} = \text{mass of compound in sample} \times \frac{\text{mass of element in compound}}{\text{mass of compound}}$$

Mass fraction * mass of sample total

eg= calcium carbonate 20 * mass frac of calcium 0.4= 8(mass of calcium in the sample)

The Observations That Led to an Atomic View of Matter

Multiple Proportions

- Dalton and others made an observation that applies when two elements form more than one compound, now called the law of multiple proportions:

- if elements A and B react to form two compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers.

- The law of multiple proportions tells us that "in two compounds of the same elements, the mass fraction of one element relative to the other element changes in increments based on ratios of small whole numbers"

The Observations That Led to an Atomic View of Matter

Multiple Proportions

- Consider two compounds, let's call them I and II, that carbon and oxygen form.
- These compounds have very different properties:
- the density of carbon oxide I is 1.25 g/L, whereas that of II is 1.98 g/L;
- I is poisonous and flammable, but II is not.
- Mass analysis shows that Carbon oxide I is 57.1 mass % oxygen and 42.9 mass % carbon Carbon oxide II is 72.7 mass % oxygen and 27.3 mass % carbon

To see the phenomenon of multiple proportions, we use the mass percents of oxygen and of carbon to find their masses in a given mass, say 100 g, of each compound. Then we divide the mass of oxygen by the mass of carbon in each compound to obtain the mass of oxygen that combines with a fixed mass of carbon:

	Carbon Oxide I	Carbon Oxide II
g oxygen/100 g compound	57.1	72.7
g carbon/100 g compound	42.9	27.3
g oxygen/g carbon	$\frac{57.1}{42.9} = 1.33$	$\frac{72.7}{27.3} = 2.66$

If we then divide the grams of oxygen per gram of carbon in II by that in I, we obtain a ratio of small whole numbers:

$$\frac{2.66 \text{ g oxygen/g carbon in II}}{1.33 \text{ g oxygen/g carbon in I}} = \frac{2}{1}$$

Dalton's Atomic Theory

- With over 200 years of hindsight, it's easy to see how the mass laws could be explained by an atomic model—matter existing in indestructible units, each with a particular mass—
- but it was a major breakthrough in 1808 when John Dalton (1766–1844) presented his atomic theory of matter in *A New System of Chemical Philosophy*
- Dalton expressed his theory in a series of postulates

The Observations That Led to an Atomic View of Matter

- All matter consists of atoms, tiny indivisible particles of an element that cannot be created or destroyed. (derives from the "eternal, indestructible atoms" proposed by Democritus more than 2000 years earlier and reflects mass conservation)

- Atoms of one element cannot be converted into atoms of another element. In chemical reactions, the atoms of the original substances recombine to form different substances.

- Atoms of an element are identical in mass and other properties and are different from atoms of any other element. (This contains Dalton's major new ideas: unique mass and properties for the atoms of a given element.)

- Compounds result from the chemical combination of a specific ratio of atoms of different elements. (from the law of definite composition)

How the Theory Explains the Mass Laws

Mass conservation:

- Atoms cannot be created or destroyed (postulate 1) or converted into other types of atoms
- Therefore, a chemical reaction, in which atoms are combined differently, cannot possibly result in a mass change.

Definite composition:

- A compound is a combination of a specific ratio of different atoms each of which has a particular mass
- Thus, each element in a compound constitutes a fixed fraction of the total mass.

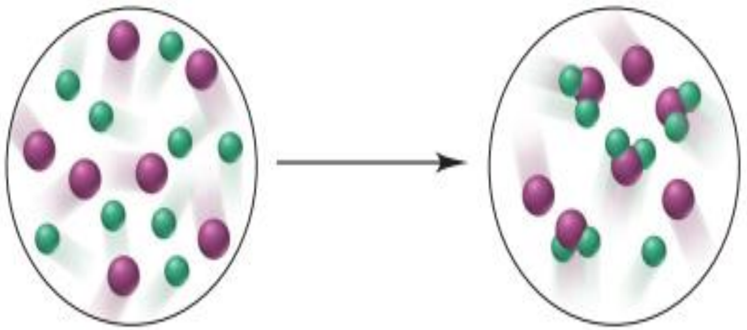
Multiple proportions:

- Atoms of an element have the same mass (postulate 3) and are indivisible
- The masses of element B that combine with a fixed mass of element A give a small, whole-number ratio because different numbers of B atoms combine with each A atom in different compounds.

How the Theory Explains the Mass Laws

Problem 1:

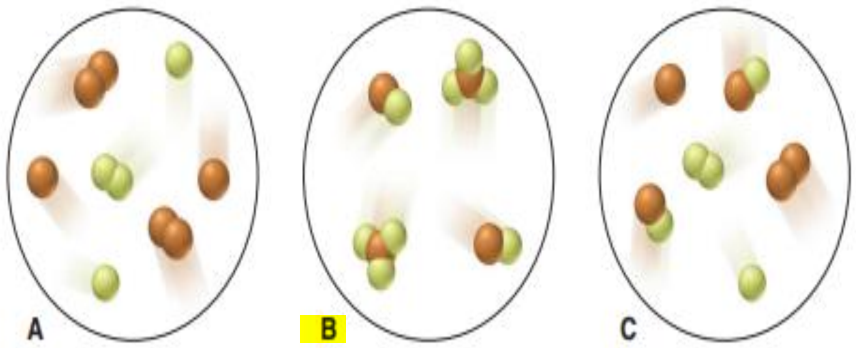
The scenes represent an atomic-scale view of a chemical reaction:



- Which of the mass laws—mass conservation, definite composition, or multiple proportions—is (are) illustrated?

Problem 2:

Which sample(s) best display(s) the fact that compounds of bromine (orange) and fluorine (yellow) exhibit the law of multiple proportions?



The Observations that Led to the Nuclear Atom Model

- Dalton's model established that masses of reacting elements could be explained in terms of atoms

- But not why atoms bond as they do

- Why, for example, do two, and not three, hydrogen atoms bond with one oxygen atom in a water molecule?

- Moreover, Dalton's model did not predict the existence of subatomic charged particles.

- These were observed in later experiments that led to the discovery of electrons and the atomic nucleus.

Key Notes (Lec 01)

- A substance is matter with a fixed composition.

- The two types of substances are elements and compounds: an element consists of a single type of atom, which may occur separately or as molecules; a compound consists of molecules (or formula units) made up of two or more different atoms combined in a fixed ratio.

- A mixture consists of two or more substances intermingled physically and, thus, has a variable composition.

Key Notes (Lec 01)

- A compound's properties differ from those of its components, but a mixture's properties do not.

- Three mass laws led to an atomic theory of matter: DALTON

(i) mass is conserved during a chemical change;

(ii) any sample of a compound has its elements in the same proportions by mass;

(iii) in different compounds made of the same two elements, the masses of one element that combine with a given mass of the other can be expressed as a ratio of small integers.

Key Notes (Lec 01)

- All the atoms of an element have the same number of protons (atomic number, Z) and thus the same chemical behavior.
- Isotopes of an element have atoms with different masses because they have different numbers of neutrons.
- The atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.
- Unlike compounds, mixtures can be separated by physical means into their components.
- A heterogeneous mixture has a nonuniform composition with visible boundaries between the components.

element wise mix

atom wise mix

- A homogeneous mixture (solution) has a uniform composition because the components (elements and/or compounds) are mixed as individual atoms, ions, or molecules.

Studies whole Dalton's model full night believing it as the only atomic model

Next lecture :-

Discovery of the Electron

- In 1897 the British physicist J. J. Thomson conducted a series of experiments that showed the atoms were not indivisible particles.

- Two electrodes from a high-voltage source are sealed into a glass tube from which the air has been evacuated.

- The negative electrode is called cathode and the positive one is anode.

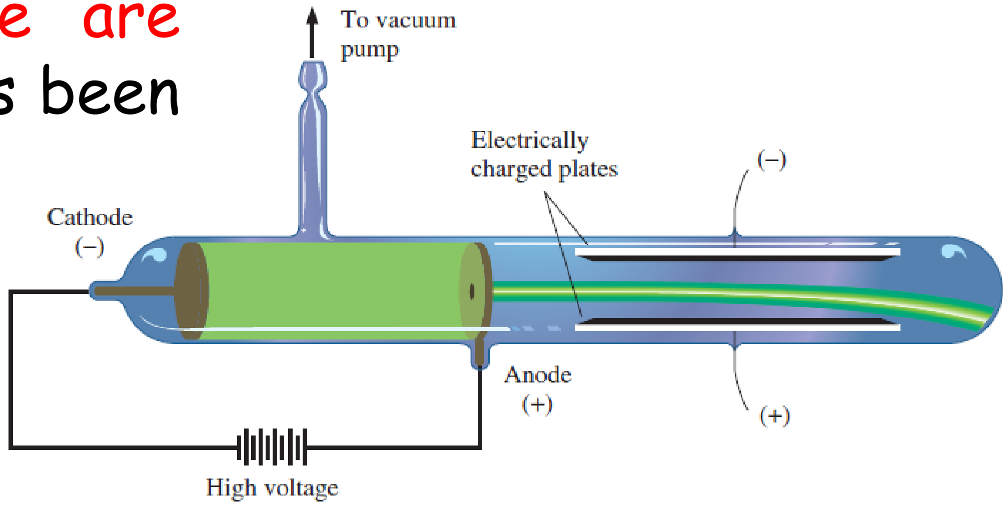


Fig: Formation of cathode rays

- When high-voltage current is turned on, the glass tube emits a greenish light caused by interaction of the glass with cathode rays.

Determination of the Charge of an Electron

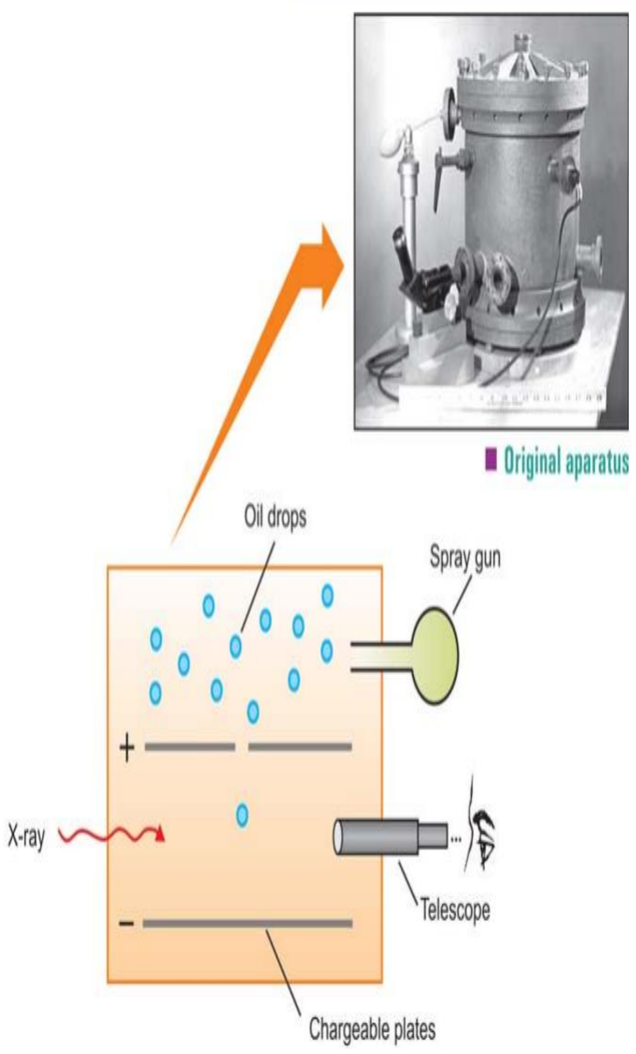
- The absolute value of the charge on an electron was measured by R.A. Milikan (1908) by what is known as the Milikan's Oil-drop Experiment

- He sprayed oil droplets from an atomizer into the apparatus. An oil droplet falls through a hole in the upper plate.

- The air between the plates is then exposed to X-rays which eject electrons from air molecules. Some of these electrons are captured by the oil droplet and it acquires a negative charge. When the plates are earthed, the droplet falls under the influence of gravity

- The drop captures one or more electrons and gets a negative charge, Q.

- Thus, Q = ne where n = number of electrons and e = charge of the electron. From measurement with different drops, Milikan established that electron has the charge – 1.60 × 10^{-19} coulombs



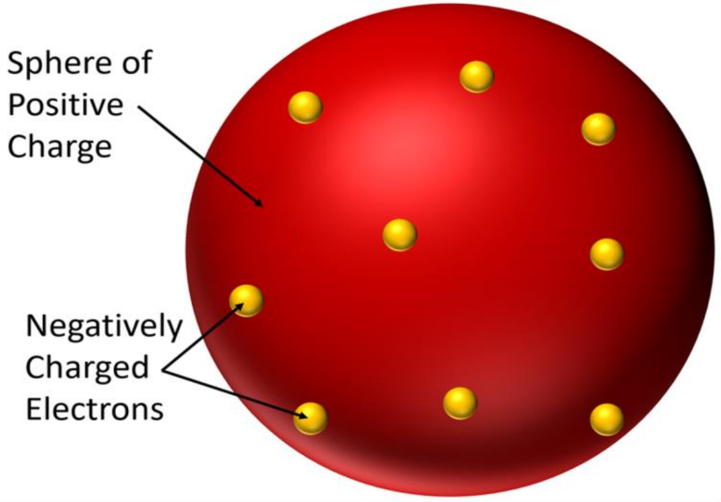
The Plum Pudding Model

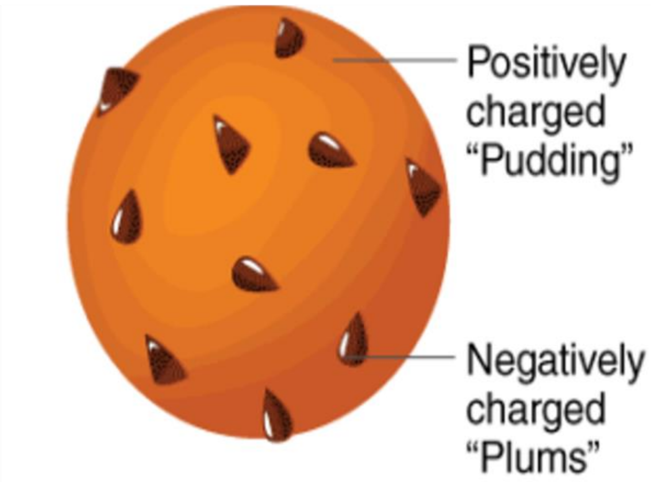
- Presence of electrons in all matter posed some major questions about the structure of atoms.
- Matter is electrically neutral, atoms must be also, what positive charges balance them?
- If an electron has such a tiny mass, what accounts for an atom's much larger mass?

The Plum Pudding Model

- To address these issues, Thomson proposed his "plum-pudding" model—" a spherical atom composed of diffuse, positively charged matter with electrons embedded like "raisins in a plum pudding."

- named after a desert made from sponge with plums stuck inside.





Positive Rays

- In 1886 Eugen Goldstein used a discharge tube with a hole in the cathode.

- He observed that while cathode rays were streaming away from the cathode, there were coloured rays produced simultaneously which passed through the perforated cathode and caused a glow on the wall opposite to the anode

- Thomson studied these rays and showed that they consisted of particles carrying a positive charge. He called them Positive rays.

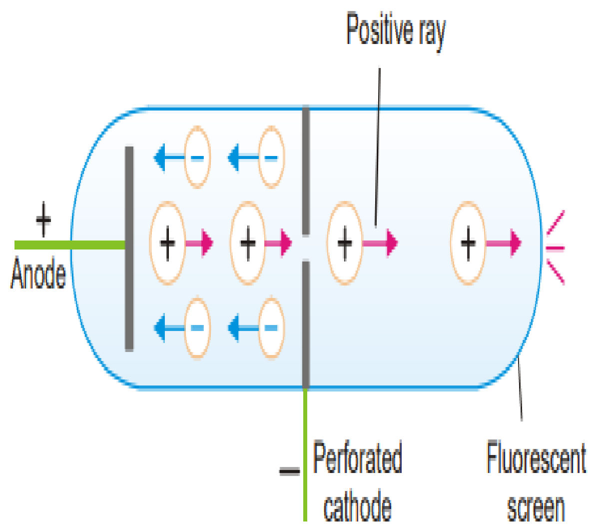


Figure 1.4
Production of Positive rays.

Idea of nucleus:
The gold foil experiment or α- scattering experiment

- Having known that atom contains electrons and a positive ion, Rutherford proceeded to perform experiments to know as to how and where these were located in the atom.

- In 1909 Rutherford and Marsden performed their historic Alpha Particle-Scattering Experiment

Idea of nucleus:

The gold foil experiment or α- scattering experiment

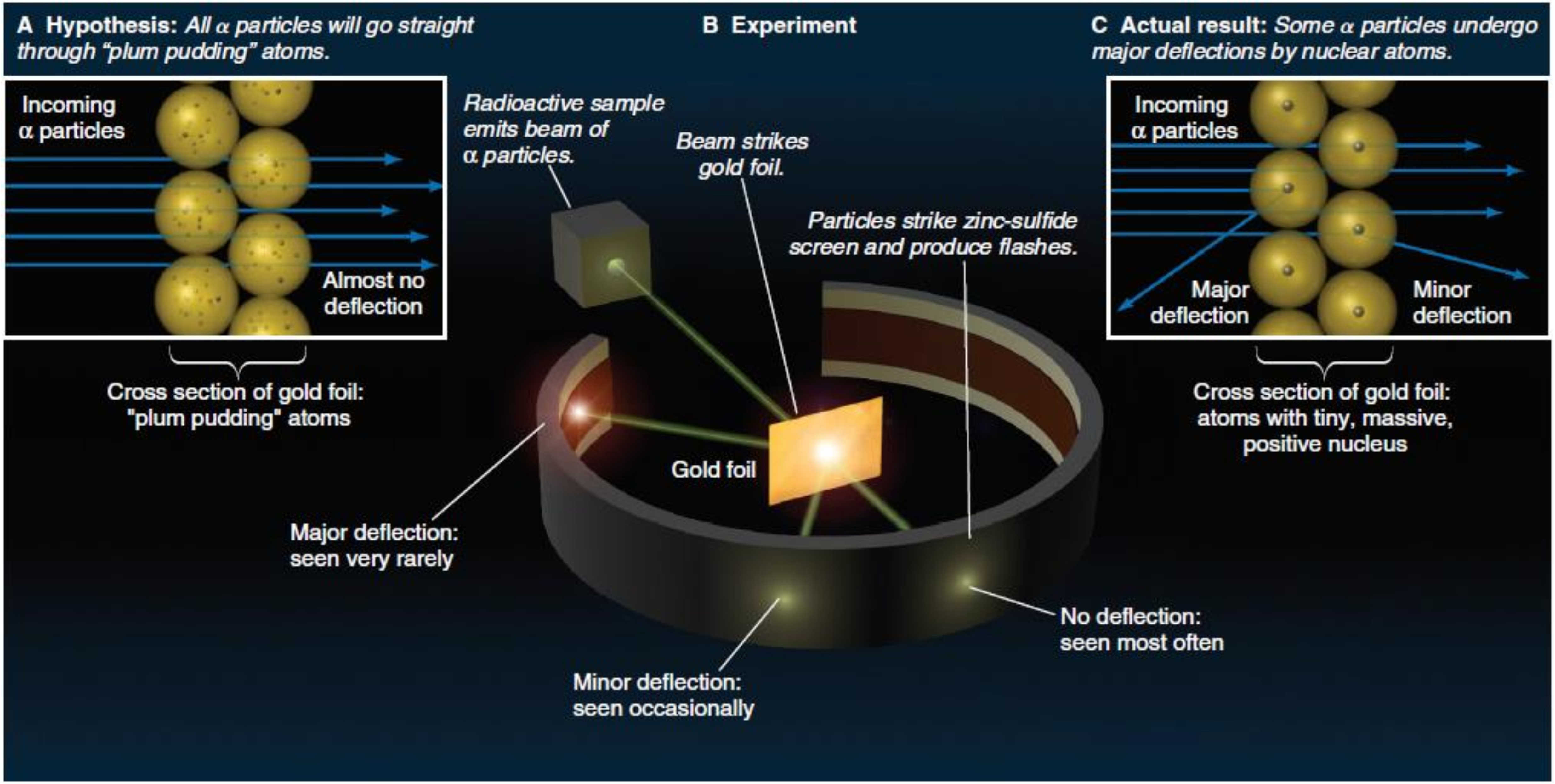
- **Rutherford** projected a beam of alpha (α) particles from a radioactive source upon a very thin gold foil.
- The α- particles emitted from radioactive elements with great velocities, on the average about 180,000 miles per second.

Observations:

1. Most alpha particles were observed to pass straight through the gold foil without deflection

2. A few were scattered at large angles

3. Some even bounced back toward the source as if the α-particles have met with some obstacles in their onward journey


A Hypothesis: All α particles will go straight through "plum pudding" atoms.
Incoming α particles
Almost no deflection
Cross section of gold foil: "plum pudding" atoms
B Experiment
Radioactive sample emits beam of α particles.
Beam strikes gold foil.
Particles strike zinc-sulfide screen and produce flashes.
Gold foil
Major deflection: seen very rarely
Minor deflection: seen occasionally
No deflection: seen most often
C Actual result: Some α particles undergo major deflections by nuclear atoms.
Incoming α particles
Major deflection
Minor deflection
Cross section of gold foil: atoms with tiny, massive, positive nucleus

Inferences/ Rutherford's Atom Model

(1) Most of the space of an atom is empty.

(2) Most of the mass of an atom is concentrated at the centre (99.95 % or more) – called the nucleus which is positively charged and exceedingly small as compared to the total size of atom.

(3) Electrons move around the nucleus, almost like the solar system in which the planets move around the sun.

(4) The number of electrons must be equal to the number of positively charged particles in the nucleus so that the atom as a whole is neutral.

(5) Due to rapid rotation of electron around the nucleus, the inward force of electrostatic attraction (centripetal force) between electron and nucleus is exactly counterbalanced by the outward centrifugal force.

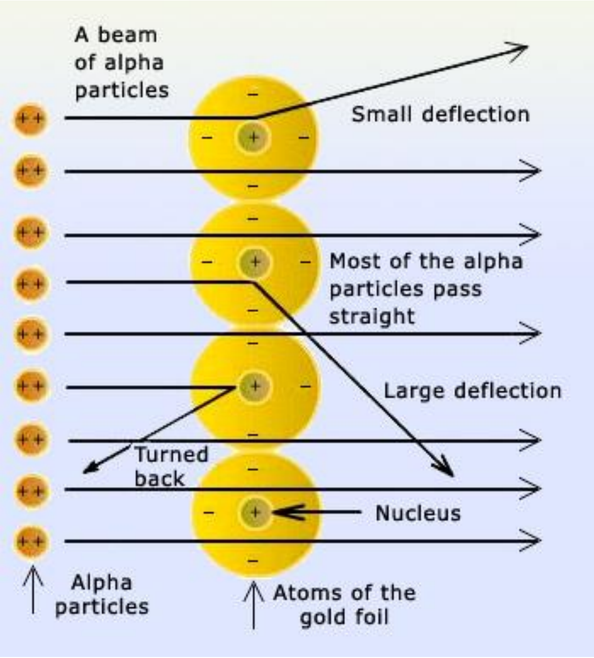


Fig: Deflection of α- particles by nuclei in a metal foil

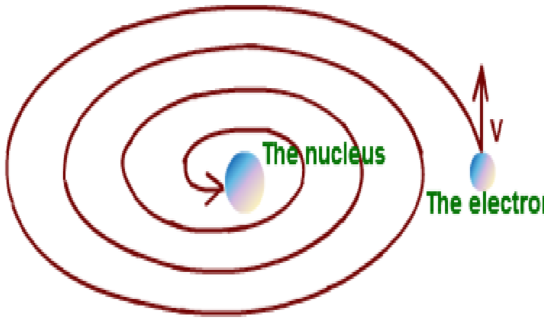
- The Rutherford atomic model has been alternatively called the nuclear atom, or the planetary model of the atom or Solar system model.

- Although most of the mass of an atom is in its nucleus, the nucleus occupies only a very small portion of the space of the atom. **Nuclei** have **diameters** of about **10^{-15} m** whereas **atomic diameter** is about **10^{-10} m**, a **hundred thousand times larger**.

Limitations of Rutherford Atom Model

- This model is based upon Newton's laws of motion and gravitation. But Newton's laws of motion and gravitation can only be applied to neutral bodies such as planets and not to charged bodies such as tiny electrons moving round a positive nucleus.

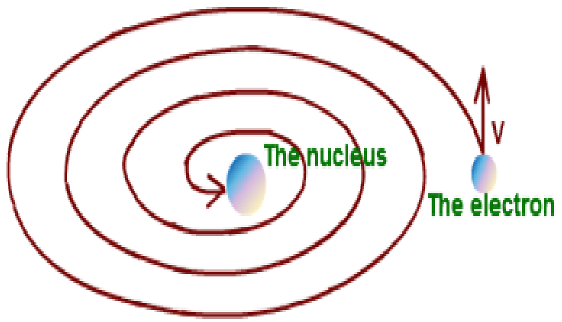
- According to Maxwell's electromagnetic theory, any charged body such as electrons rotating in an orbit must radiate energy continuously thereby losing kinetic energy. Thus rotating electron will lose energy and its orbit will become smaller and smaller and it will ultimately fall into the nucleus following a spiral path, annihilating the atom model.



In the planetary model of atom, the electron should emit energy and spirally fall on the nucleus.

Limitations of Rutherford Atom Model

- If there is continuous emission of radiation, the spectra of an atom will be a band or continuous spectra. But an atom gives discontinuous or line spectra.
- Rutherford did not give any idea about the shape of the orbits.
- There was no explanation about the rotation of electrons in an atom with many electrons.



In the planetary model of atom, the electron should emit energy and spirally fall on the nucleus.

Quantum Theory And Bohr Atom

- Rutherford model laid the foundation of the model picture of the atom.
- However, it did not tell anything as to the position of the electrons and how they were arranged around the nucleus.
- Rutherford recognised that electrons were orbiting around the nucleus.
- But according to the classical laws of Physics an electron moving in a field of force like that of nucleus, would give off radiations and gradually collapse into the nucleus.
- Thus Rutherford model failed to explain why electrons did not do so

Quantum Theory And Bohr Atom

- Neils Bohr, a brilliant Danish Physicist, pointed out that the old laws of physics just did not work in the submicroscopic world of the atom.

- He closely studied the behaviour of electrons, radiations and atomic spectra.

- In 1913 Bohr proposed a new model of the atom based on the modern Quantum theory of energy.

- he was able to explain as to why an orbiting electron did not collapse into the nucleus and how the atomic spectra were caused by the radiations emitted when electrons moved from one orbit to the other

- it is first necessary to acquaint ourselves with the nature of electromagnetic radiations and the atomic spectra as also the Quantum theory of energy

Electromagnetic Radiations

- Energy can be transmitted through space by electromagnetic radiations.
- Some forms of radiant energy are radio waves, visible light, infrared light, ultraviolet light, X-rays and γ-radiations
- are so named because they consist of waves which have electrical and magnetic properties
- electromagnetic waves are produced by a periodic motion of charged particles
- vibratory motion of electrons would cause a wave train of oscillating electric field and another of oscillating magnetic field.
- These electromagnetic waves travel through empty space with the speed or velocity of light

Characteristics of Waves

Waves are characterised by the following properties :

Wavelength:

- The wavelength is defined as the distance between two successive crests or troughs of a wave.
- Wavelength is denoted by the Greek letter λ (lambda).
- It is expressed in centimetres or metres or in angstrom units. One angstrom, Å, is equal to 10–8 cm. It is also expressed in nanometers (1nm = 10– 9 m)

Frequency

- The frequency is the number of waves which pass a given point in one second.
- Frequency is denoted by the letter v (nu) and is expressed in hertz (hz). 1/T
- It is noteworthy that a wave of high frequency (b) has a shorter wavelength, while a wave of low frequency (a) has a longer wavelength

Characteristics of Waves

Speed:

- Speed The speed (or velocity) of a wave is the distance through which a particular wave travels in one second.
- Speed is denoted by c and it is expressed in cm per second. If the speed of a wave is c cm/sec, it means that the distance travelled by the wave in one second is c cm.
- Speed is related to frequency and wavelength by the expression $c = v\lambda$ or
- Speed = Frequency × Wavelength

Wave number

- Another quantity used to characterise radiation is the wave number. This is reciprocal of the wavelength and is given the symbol v (nu bar).
- That is, $v = 1/\lambda$
- The wave number is the number of wavelengths per unit of length covered. Its units are cm^{-1}

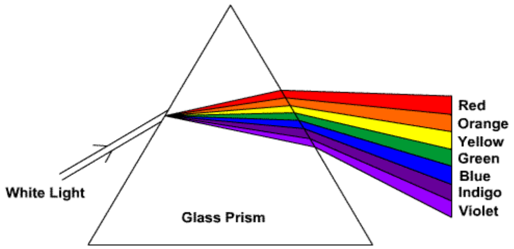
- The wavelength of a violet l
wave number.
- The frequency of strong yel
1014 sec– 1. Calculate the

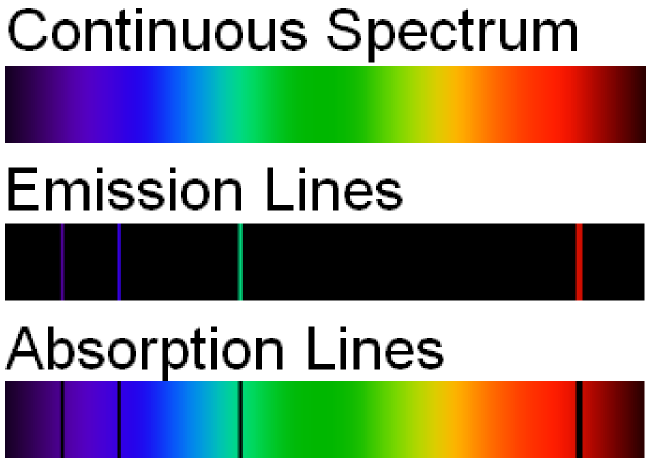
Mathematical Problems

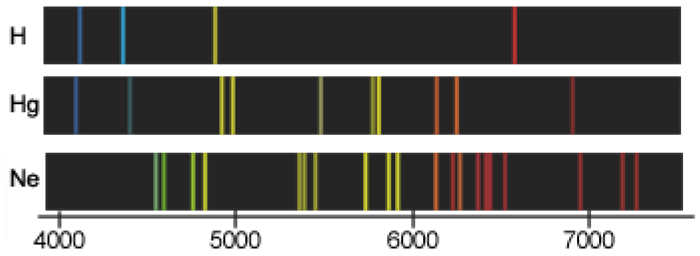
- The wavelength of a violet light is 400 nm. Calculate its frequency and wave number.

- The frequency of strong yellow line in the spectrum of sodium is 5.09 × 1014 sec– 1. Calculate the wavelength of the light in nanometers

Spectrum





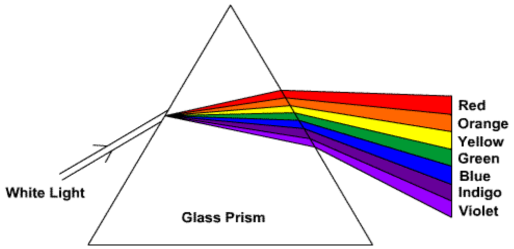


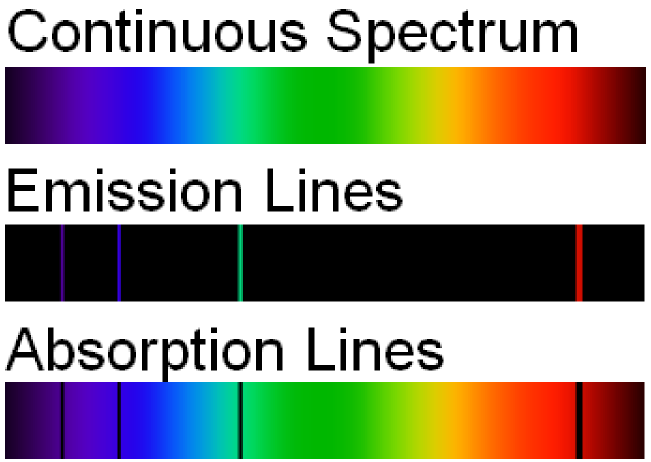
When white light is passed through a prism it is resolved into various color components of different wave length, and what is obtained is called a spectrum. In the seventeenth century Newton showed that sun light is composed of various color components.

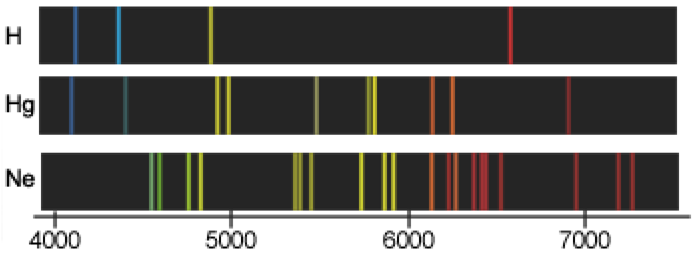
Line Spectrum: A spectrum showing only certain colors or specific wavelengths of light.

Continuous spectrum: A spectrum containing light of all wavelengths.

Spectrum







Emission spectrum: If atoms or molecules are heated to sufficiently high temperature, they emit light of certain wavelengths. The pattern of frequencies emitted by the substance is called emission spectrum. From the emission spectrum we get bright lines.

Absorption spectrum: When white light is passed through a substance, black lines appear in the spectrum where light of some wavelengths have been absorbed by the substance. The pattern of frequencies absorbed by the substance is called absorption spectrum. From the absorption spectrum we get dark lines.

Atomic Spectrum Of Hydrogen

The emission line spectrum of hydrogen can be obtained by passing electric discharge through the gas contained in a discharge tube at low pressure. The light radiation emitted is then examined with the help of a **spectroscope.** The bright lines recorded on the photographic plate constitute the atomic spectrum of hydrogen (Fig. 1.18).

In 1884 J.J. Balmer observed that there were four prominent coloured lines in the visible hydrogen spectrum :

(1) a *red line* with a wavelength of 6563 Å.
(2) a *blue-green line* with a wavelength 4861 Å.
(3) a *blue line* with a wavelength 4340 Å.
(4) a *violet line* with a wavelength 4102 Å.

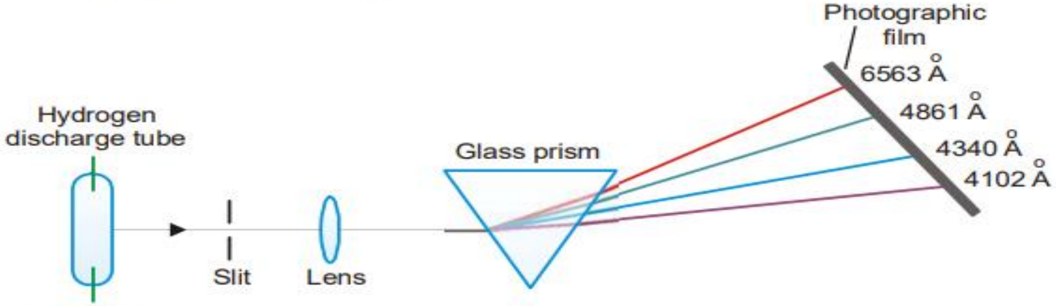


Figure 1.18
The examination of the atomic spectrum of hydrogen with a spectroscope.

The above series of four lines in the visible spectrum of hydrogen was named as the **Balmer Series.** By carefully studying the wavelengths of the observed lines, Balmer was able empirically to give an equation which related the wavelengths (λ) of the observed lines. The **Balmer Equation** is

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

where R is a constant called the **Rydberg Constant** which has the value 109, 677 cm^{-1} and $n = 3, 4, 5, 6$ etc. That is, if we substitute the values of 3, 4, 5 and 6 for n, we get, respectively, the wavelength of the four lines of the hydrogen spectrum.

Name	Region where located
(1) Lyman Series	Ultraviolet
(2) Balmer Series	Visible
(3) Paschen Series	Infrared
(4) Brackett Series	Infrared
(5) Pfund Series	Infrared

Balmer equation had no theoretical basis at all. Nobody had any idea how it worked so accurately in finding the wavelengths of the spectral lines of hydrogen atom. However, in 1913 Bohr put forward his theory which immediately explained the observed hydrogen atom spectrum. Before we can understand Bohr theory of the atomic structure, it is necessary to acquaint ourselves with the quantum theory of energy.

Thank you