

# ENGINEERING CHEMISTRY-I

AS PER SBTE, JHARKHAND SYLLABUS OF  
FIRST SEMESTER POLYTECHNIC STUDENTS



**KHANNA PUBLISHERS**

# **ENGINEERING CHEMISTRY-I**

**As Per SBTE, Jharkhand Syllabus**  
*of*  
First Semester Polytechnic Students

*Compiled by*  
**Khanna Editorial Team**



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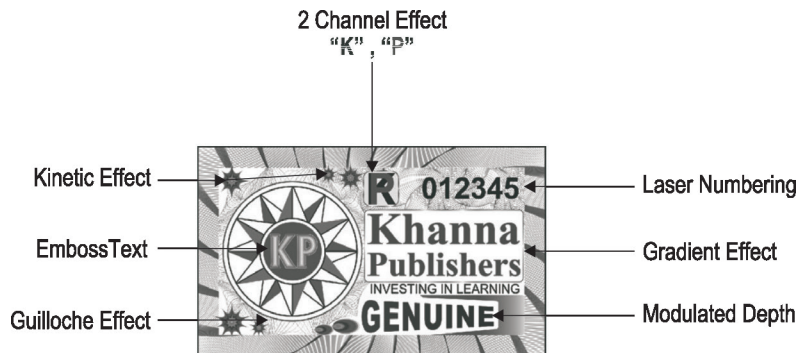
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# Preface

It gives us immense pleasure in presenting this revised edition to the diploma students of SBTE, Jharkhand for their first semester course as per the E-scheme. The book 'Engineering Chemistry-1', written in simple and lucid language, provides an important opportunity for diploma students to learn the core concepts of chemistry and understand how those concepts apply to their lives and the world around them. The text has been developed to meet the scope and sequence of most general chemistry courses. We tried to add a number of innovative features designed to enhance student learning.

## **Salient Features**

- Written as per SBTE, Jharkhand E-Scheme
- It provides the basic concepts of chemistry in simple and lucid language
- The orbital configuration of different elements have been drawn
- Important definitions have been given in each chapter.

Although every care has been taken to make this edition error-free, some mistake might be crept in. If you bring to our notice such mistakes, we would be thankful to you, and we will try to rectify them in the next edition.

Also any suggestions and comments for the improvement of the book will be gratefully acknowledged.

***–Publishers***

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# Atomic Structure

## Objectives

After reading this chapter, students should be able to :

- Define Atom, Atomic number and Atomic Mass number
- Recognize the fundamental particles of atom
- Understand Bohr's Theory, Hund's Rule and Aufbau's Principle
- Know the formation of Electrovalent & Covalent Compounds
- Identify the difference between electrovalent & covalent compounds.

## 1.1 INTRODUCTION

Chemistry has been defined as the study of matter in terms of its structure, composition and the properties. As you are aware, matter is made up of atoms, and therefore an understanding of the structure of atom is very important. You have studied in your earlier classes that the earliest concept of atom ( smallest indivisible part of matter ) was given by ancient (600-400 BC) Indian and Greek philosophers. At that time there were no experimental evidence. The origin of the concept of atom was based on their thoughts on 'What would happen if we continuously keep dividing matter'. John Dalton revived the concept of atom in the beginning of nineteenth century in terms of his atomic theory which successfully explained the laws of chemical combination. Later experiments showed that the atom is not indivisible but has an internal structure.

In this chapter, you will learn about the internal structure of an atom which will help you to understand the correlations between its structure and properties.

## 1.2 DEFINITION OF ATOM

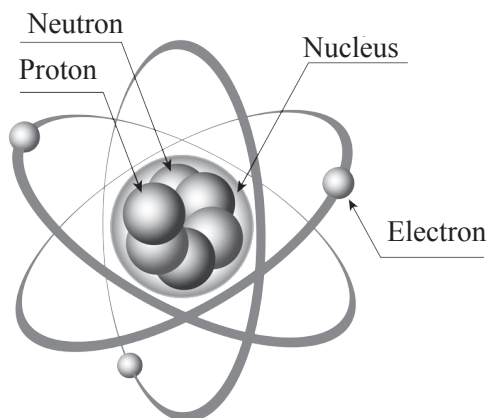
An atom is the smallest constituent unit of ordinary matter that has the properties of a chemical element. Every solid, liquid, gas, and plasma is composed of neutral or ionized atoms. Atoms are very small; typical sizes are around 100 picometers (a ten-billionth of a meter, in the short scale).

Atoms are small enough that attempting to predict their behavior using classical physics – as if they were billiard balls, for example – gives noticeably incorrect predictions due to quantum effects. Through the development of physics, atomic models have incorporated quantum principles to better explain and predict the behavior.

Every atom is composed of a nucleus and one or more electrons bound to the nucleus. The nucleus is made of one or more protons and typically a similar number of neutrons. Protons

and neutrons are called nucleons. More than 99.94% of an atom's mass is in the nucleus. The protons have a positive electric charge, the electrons have a negative electric charge, and the neutrons have no electric charge. If the number of protons and electrons are equal, that atom is electrically neutral. If an atom has more or fewer electrons than protons, then it has an overall negative or positive charge, respectively, and it is called an ion.

A matter is anything that has mass and occupies space. Pen, paper, clips, sand, air, ice, etc. are different forms of matter. Every matter is made up of tiny particles. These particles are so tiny that they can't be seen with naked eyes. Let's learn about different characteristics of particles of matter.



**Figure 1:** Atomic structure.

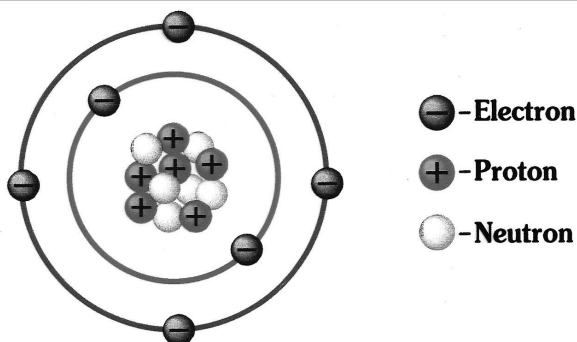
### 1.3 FUNDAMENTAL PARTICLES OF ATOM

A fundamental particle - also known as an elementary particle - is one that cannot be broken into smaller pieces. An atom is not a fundamental particle, because it consists of protons and neutrons in its nucleus and has a cloud of electrons orbiting the nucleus. Protons and neutrons are also not fundamental particles, because they consist of quarks. However, quarks and electrons cannot be broken apart and are fundamental particles.

In 1897 J.J. Thomson discovered electron as a constituent of atom. He determined that an electron had a negative charge and had very little mass as compared to that of the atom. Since an atom was found to be electrically neutral it was inferred that some source of positive charge must be present in the atom. This soon led to the experimental discovery of the proton, which is a positively charged subatomic particle. Proton was found approximately 1840 times heavier than an electron. Further experiments revealed that the atomic masses were more than that expected from the presence of just protons and electrons in the atom. For example, the mass of helium atom was expected to be double that of hydrogen atom but was actually found to be almost four times the mass of hydrogen atom. This suggested the presence of neutral particles with mass comparable to that of protons in the atom. Sir James Chadwick discovered this neutral particle and called it neutron subsequently in 1932. Thus we may conclude that atoms are not indivisible but are made up of three fundamental particles whose characteristics are given in Table 1.1.

**Table 1.1** Fundamental particles of atom and their characteristics.

Particle	Symbol	Mass/ kg	Actual Charge / C	Relative charge
Electron	<i>e</i>	$9.109\ 389 \times 10^{-31}$	$- 1.602\ 177 \times 10^{-19}$	-1
Proton	<i>p</i>	$1.672\ 623 \times 10^{-27}$	$1.602\ 177 \times 10^{-19}$	+1
Neutron	<i>n</i>	$1.674\ 928 \times 10^{-27}$	0	0

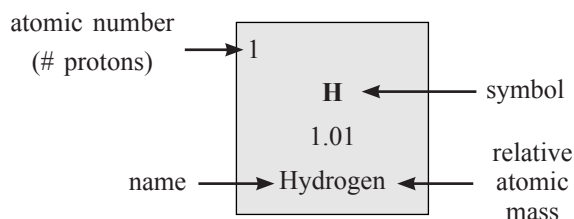
**Figure 2:** Fundamental particles of atomic.

## 1.4 DEFINITION OF ATOMIC NUMBER AND ATOMIC MASS

Atoms of each element contain a characteristic number of protons. In fact, the number of protons determines what atom we are looking at (e.g., all atoms with six protons are carbon atoms); the number of protons in an atom is called the atomic number. In contrast, the number of neutrons for a given element can vary. Forms of the same atom that differ only in their number of neutrons are called isotopes. Together, the number of protons and the number of neutrons determine an element's mass number: mass number = protons + neutrons. If you want to calculate how many neutrons an atom has, you can simply subtract the number of protons, or atomic number, from the mass number.

A property closely related to an atom's mass number is its atomic mass. The atomic mass of a single atom is simply its total mass and is typically expressed in atomic mass units or amu. By definition, an atom of carbon with six neutrons, carbon-12, has an atomic mass of 12 amu. Other atoms don't generally have round-number atomic masses for reasons that are a little beyond the scope of this article. In general, though, an atom's atomic mass will be very close to its mass number, but will have some deviation in the decimal places.

Since an element's isotopes have different atomic masses, scientists may also determine the relative atomic mass—sometimes called the atomic weight—for an element. The relative atomic mass is an average of the atomic masses of all the different isotopes in a sample, with each isotope's contribution to the average determined by how big a fraction of the sample it makes up. The relative atomic masses given in periodic table entries—like the one for hydrogen, below—are calculated for all the naturally occurring isotopes of each element, weighted by the abundance of those isotopes on earth. Extraterrestrial objects, like asteroids or meteors, might have very different isotope abundances.



### 1.4.1 What is an atom's atomic number?

The number of protons in the nucleus of an atom determines an element's atomic number. In other words, each element has a unique number that identifies how many protons are in one atom of that element. For example, all hydrogen atoms, and only hydrogen atoms, contain one proton and have an atomic number of 1. All carbon atoms, and only carbon atoms, contain six protons and have an atomic number of 6. Oxygen atoms contain 8 protons and have an atomic number of 8. The atomic number of an element never changes, meaning that the number of protons in the nucleus of every atom in an element is always the same.

### 1.4.2 What is an atom's mass number?

The mass number (symbol  $A$ ), also called atomic mass number or nucleon number, is the total number of protons and neutrons (together known as nucleons) in an atomic nucleus.

The mass number is written either after the element name or as a superscript to the left of an element's symbol. For example, the most common isotope of carbon is carbon-12, or  $^{12}\text{C}$ , which has 6 protons and 6 neutrons.

All atoms have a mass number which is derived as follows.

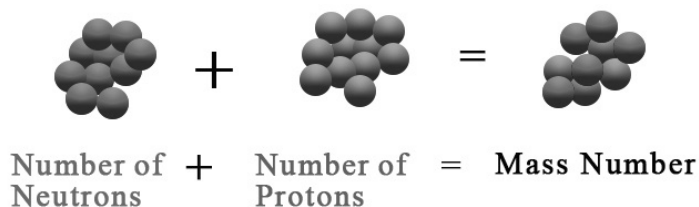


Figure 3: Atomic mass number.

## 1.5 ISOTOPES AND ISOBARS

### 1.5.1 Isotopes

Those elements that have the same atomic number but a different mass number are referred to as isotopes. Isotopes occur due to the presence of a different number of neutrons in elements having the same atomic number as mass number is the sum of the number of neutrons and protons. Many but not all elements have isotopes. The isotopes of hydrogen are protium (has one proton and no neutrons), deuterium (has one proton and one neutron) and tritium (has one

proton and two neutrons). The chemical properties of isotopes are same owing to the fact that they have the same number of protons and hence same number of electrons which determines the chemical properties of an element.

### Isotopes of hydrogen

Hydrogen has three main isotopes; protium ( ${}^1\text{H}$ ), deuterium ( ${}^2\text{H}$ ) and tritium ( ${}^3\text{H}$ ). These isotopes form naturally in nature. Protium and deuterium are stable. Tritium is radioactive and has a half-life of about 12 years.

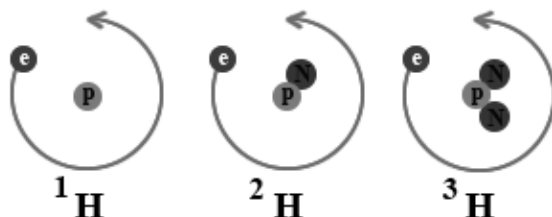


Figure 4: Isotopes of hydrogen.

### Isotopes of oxygen

There are three stable isotopes of oxygen ( ${}_{8}\text{O}$ ):  ${}^{16}\text{O}$ ,  ${}^{17}\text{O}$ , and  ${}^{18}\text{O}$ . Radioactive isotopes with mass numbers from  ${}^{12}\text{O}$  to  ${}^{24}\text{O}$  have also been characterized, all short-lived, with the longest-lived being  ${}^{15}\text{O}$  with a half-life of 122.24 seconds. The shortest-lived is  ${}^{12}\text{O}$  with a half-life of  $580(30)\times 10^{-24}$  second.

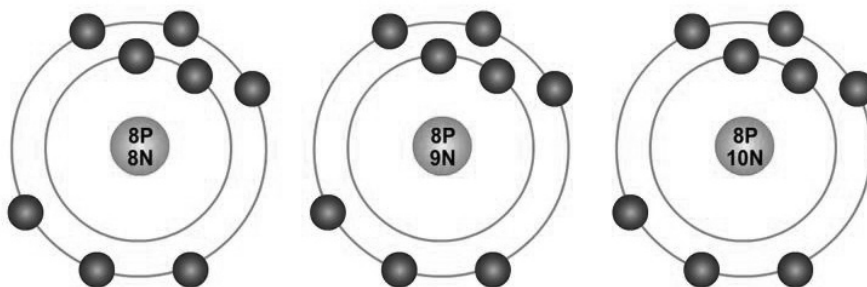


Figure 5: Isotopes of oxygen.

## 1.5.2 Isobars

Isobars, on the other hand, are the atoms having the same mass number but a different atomic number. For example, the atomic number of carbon and nitrogen is 6 and 7 respectively. Carbon-14 an isotope of carbon has a mass number of 14 which is same as that of nitrogen and hence carbon-14 and nitrogen are Isobars

## Isotopes and radioactive decay

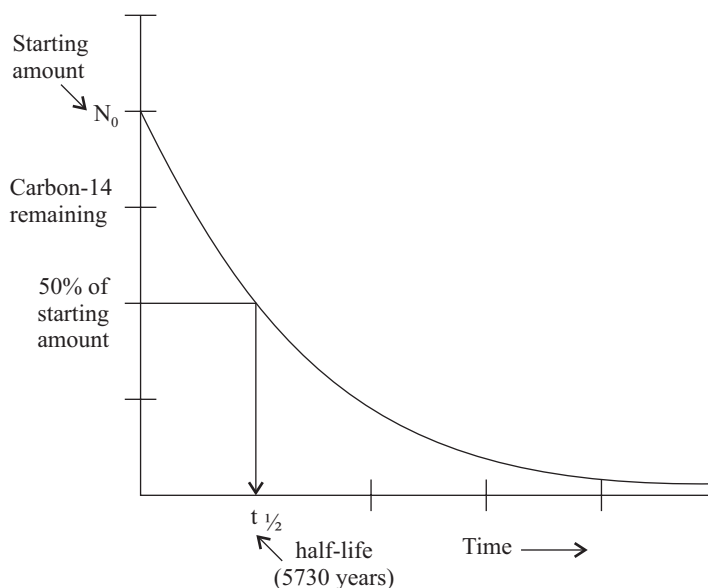
As mentioned above, isotopes are different forms of an element that have the same number of protons but different numbers of neutrons. Many elements—such as carbon, potassium, and uranium—have multiple naturally occurring isotopes. Carbon-12 contains six protons, six neutrons, and six electrons; therefore, it has a mass number of 12 (six protons plus six neutrons). Carbon-14 contains six protons, eight neutrons, and six electrons; its mass number is 14 (six protons plus eight neutrons). These two alternate forms of carbon are isotopes.

Some isotopes are stable, but others can emit, or kick out, subatomic particles to reach a more stable, lower-energy, configuration. Such isotopes are called radioisotopes, and the process in which they release particles and energy is known as decay. Radioactive decay can cause a change in the number of protons in the nucleus; when this happens, the identity of the atom changes (e.g., carbon-14 decaying to nitrogen-14).

Radioactive decay is a random but exponential process, and an isotope's half-life is the period over which half of the material will decay to a different, relatively stable product. The ratio of the original isotope to its decay product and to stable isotopes changes in a predictable way; this predictability allows the relative abundance of the isotope to be used as a clock that measures the time from the incorporation of the isotope (e.g., into a fossil) to the present.

Graph of radioactive decay of carbon-14. The amount of carbon-14 decreases exponentially with time. The time at which half of the original carbon-14 has decayed—and half still remains—is designated as  $t_{1/2}$ . This time is also known as the half-life of the radioisotope and, for carbon-14, is equal to 5730 years.

For example, carbon is normally present in the atmosphere in the form of gases like carbon dioxide, and it exists in three isotopic forms: carbon-12 and carbon-13, which are stable, and carbon-14, which is radioactive. These forms of carbon are found in the atmosphere in relatively constant proportions, with carbon-12 as the major form at about 99%, carbon-13 as a minor form at about 1%, and carbon-14 present only in tiny amounts. As plants pull carbon dioxide from the air to make sugars, the relative amount of carbon-14 in their tissues will be equal to the concentration of carbon-14 in the atmosphere. As animals eat the plants, or eat other animals that ate plants, the concentrations of carbon-14 in their bodies will also match the



atmospheric concentration. When an organism dies, it stops taking in carbon-14, so the ratio of carbon-14 to carbon-12 in its remains, such as fossilized bones, will decline as carbon-14 decays gradually to nitrogen-14.

After a half-life of approximately 5,730 years, half of the carbon-14 that was initially present will have been converted to nitrogen-14. This property can be used to date formerly living objects such as old bones or wood. By comparing the ratio of carbon-14 to carbon-12 concentrations in an object to the same ratio in the atmosphere, equivalent to the starting concentration for the object, the fraction of the isotope that has not yet decayed can be determined. On the basis of this fraction, the age of the material can be calculated with accuracy if it is not much older than about 50,000 years. Other elements have isotopes with different half lives, and can thus be used to measure age on different timescales. For example, potassium-40 has a half-life of 1.25 billion years, and uranium-235 has a half-life of about 700 million years and has been used to measure the age of moon rocks.

### 1.5.3 All the Isotopes of an Element Have Identical Chemical Properties

The chemical properties of an atom of the element depend on the number of protons and electrons, not on the number of neutrons. Since all the isotopes of an element contain the same number of protons and electrons, therefore, the chemical properties of all the isotopes of an element are identical (or same). We can explain the identical chemical properties of all the isotopes of an element on the basis of their electronic configurations as follows.

All the isotopes of an element contain the same number of electrons because of which they have identical electronic configurations having the same number of valence electrons. Since all the isotopes of an element have identical electronic configurations containing the same number of valence electrons, therefore, all the isotopes of an element show identical chemical properties. For example, the two isotopes of chlorine,  $^{35}\text{Cl}$  and  $^{37}\text{Cl}$ , both have the same number of 17 electrons in them due to which both of them have the same electronic configuration of 2, 8, 7. Since both the isotopes of chlorine,  $^{35}\text{Cl}$  and  $^{37}\text{Cl}$ , have identical electronic configurations (having the same number of 7 valence electrons), they show identical chemical properties.

## 1.6 DALTON'S ATOMIC THEORY

Various atoms and molecules as depicted in John Dalton's "A New System of Chemical Philosophy (1808)".

In the early 1800s, John Dalton used the concept of atoms to explain why elements always react in ratios of small whole numbers (the law of multiple proportions). For instance, there are two types of tin oxide: one is 88.1% tin and 11.9% oxygen and the other is 78.7% tin and 21.3% oxygen (tin(II) oxide and tin dioxide respectively). This means that 100g of tin will combine either with 13.5g or 27g of oxygen. 13.5 and 27 form a ratio of 1:2, a ratio of small whole numbers. This common pattern in chemistry suggested to Dalton that elements react in

whole number multiples of discrete units—in other words, atoms. In the case of tin oxides, one tin atom will combine with either one or two oxygen atoms.

Dalton also believed atomic theory could explain why water absorbs different gases in different proportions. For example, he found that water absorbs carbon dioxide far better than it absorbs nitrogen. Dalton hypothesized this was due to the differences between the masses and configurations of the gases' respective particles, and carbon dioxide molecules ( $\text{CO}_2$ ) are heavier and larger than nitrogen molecules ( $\text{N}_2$ ).

## 1.7 NIELS BOHR ATOMIC THEORY

Niels Henrik David Bohr (1885-1962) was published his Niels Bohr atomic theory in 1913. He was a Danish physicist and he got Nobel Prize in 1922 in physics for atomic structure and quantum mechanics. He was one of the most influential scientists in 20th century.

Niels Bohr model of the atom is very small, positively charged nucleus which contains proton and neutron surrounded by negatively charged electrons. This negatively charged electrons travel in circular orbits around the nucleolus. Niels Bohr atomic model is more similar to the solar system. But in Niels Bohr atomic model, electro static force providing attraction force rather than gravity force.

- Niels Bohr atomic model was an improved model from an earlier cubic model, the plum-pudding model, the saturation model, and the Rutherford model.
- Niels Bohr atomic model not only explained the structure of the Rydberg formula and its reasoning but also gave a clear acceptability for its disambiguated result in terms of fundamental physical constant.
- Niels Bohr applies quantum Rutherford's atomic structure and assuming that the electrons travel in stationary orbits defined by their angular momentum. This led to calculation of possible energy level for these orbits and the postulations that the emission of light occurs when an electron transfers in to the lower energy level of orbit.
- Niels Bohr atomic model is basically considered as a primitive hydrogen atom model.
- As a Niels Bohr atomic theory, it could be accepted as the derivation of a first order approximation of an atom of hydrogen using broader and much more accurate quantum mechanics.
- Niels Bohr atomic theory could be utilized and accepted to be an absolute scientific theory.

The Niels Bohr atomic model is basically introduced to impart students to the knowledge of quantum mechanics because of its simplicity and credibility for a particular selected system, before going to models which are more accurate and complex valance atom.

### 1.7.1 Bohr Model

In 1913 the physicist Niels Bohr proposed a model in which the electrons of an atom were assumed to orbit the nucleus but could only do so in a finite set of orbits, and could jump between these orbits only in discrete changes of energy corresponding to absorption or radiation of a photon. This quantization was used to explain why the electrons orbits are stable (given that normally, charges in acceleration, including circular motion, lose kinetic energy which is

emitted as electromagnetic radiation, see synchrotron radiation) and why elements absorb and emit electromagnetic radiation in discrete spectra.

Later in the same year Henry Moseley provided additional experimental evidence in favor of Niels Bohr's theory. These results refined Ernest Rutherford's and Antonius Van den Broek's model, which proposed that the atom contains in its nucleus a number of positive nuclear charges that is equal to its (atomic) number in the periodic table. Until these experiments, atomic number was not known to be a physical and experimental quantity. That it is equal to the atomic nuclear charge remains the accepted atomic model today.

### 1.7.2 Niels Bohr Atomic Theory Experiment

Niels Bohr atomic theory can be used to explain the structure of all atoms. Niels Bohr atomic theory was a combination of Rutherford's model and an idea of atom, with quantum theory, which is introduced by Max Planck.

He expanded the Rutherford model upon this theory by proposing that electrons travel only on certain successively larger orbits. Niels Bohr suggested that the outer orbits could not hold much of electrons compared with the inner orbits. The atom's chemical property depends upon the outer orbits.

Niels Bohr atomic theory experiment explains the following important things that are

- Electrons would have only certain chemical motions.
- Electrons in atoms orbit the nucleus in orbits that have a set size and energy.
- The lowest energy is found in the smallest orbit, since the energy of the orbit is related to its size.
- When electrons move or jump from one orbit to another one, electrons absorb or emit the radiation with a frequency  $\nu$  determined by the energy difference of the energy levels according to Planck's relations.

$$\Delta E = E_2 - E_1 = h\nu$$

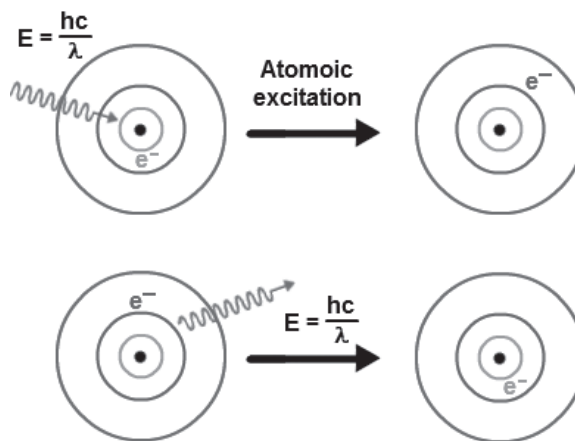
$$\nu = 1/T$$

Where,

$h$  = Planck's constant

$T$  = Frequency of the radiation emitted at an orbit of period

Electrons could orbit the nucleus, but only in discrete orbits at fixed distance from the nucleus, which didn't emit radiation. Atoms can move between one orbit to another orbit by absorbing or emitting energies by quantum mechanisms. When electrons move to a higher orbit, with a large nucleus, the absorption of a photon causing an atomic excitation and when



**Figure 6:** Niels Bohr atomic theory.

electrons moves for lower orbit, with a smaller radius, the emission of photon causing an atomic de-excitation.

**Note:** In each case the wavelength of absorbed or emitted lights is exactly such that the photon carries the energy difference between two orbits. The energy can be calculated by using following equation.

$$E = hc/\lambda$$

Where,

$h$  = Planck's constant,

$c$  = Speed of light, and

$\lambda$  = Wavelength of the light

An atom can absorb or emit only certain discrete wavelength or energies. Niels Bohr atomic theory experiment explains how the Bohr model is works.

### **Niels Bohr Atomic Theory Contribution**

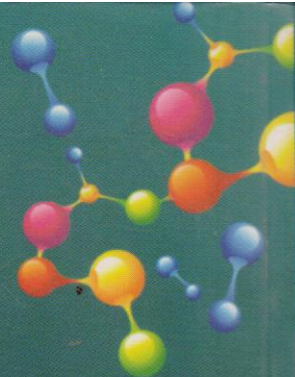
- Niels Bohr postulated atomic theory based on quantum theory that the negatively charged electrons travel around the positively charged nucleus in a stationary orbit.
- This also led to the theory of the different energy levels in atoms, that is if an electron transmits from higher orbit level to an lower level orbit it must release energy or electromagnetic radiation.
- The Niels Bohr contributed the quantum theory of atomic structure and today's research of quantum physics which is our future.

## **1.8 SHAPES OF ORBITALS – ATOMIC ORBITALS**

According to quantum atomic model, an atom can have many possible numbers of orbitals. These orbitals can be categorized on the basis of their size, shape or orientation. A smaller sized orbital means there is a greater chance of getting an electron near the nucleus. The orbital wave function or  $\psi$  is a mathematical function used for representing the coordinates of an electron. The square of the orbital wave function or  $\psi^2$  represents the probability of finding an electron. This wave function also helps us in drawing boundary surface diagrams. Boundary surface diagrams of constant probability density for different orbitals help us understand the shape of orbitals.

### **1.8.1 Determination of shapes of orbitals**

**s-orbital:** Boundary surface diagram for s-orbital looks like a sphere having nucleus as its center which in two dimensions can be seen as a circle. Hence, we can say that s-orbitals are spherically symmetric having probability of finding the electron at a given distance equal in all the directions. The size of the s-orbital is also found to increase with increase in the value of principle quantum number ( $n$ ), thus,  $4s > 3s > 2s > 1s$ .



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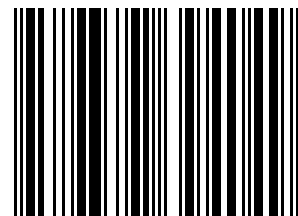
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