Introduction to Structure of Atom

An atom is present at the most basic level in everything we see around us. In fact, every living organism is composed of atoms. Every non-living thing around you such as tables, chairs, water, etc is made up of matter. But the building blocks of matter are atoms. Therefore, living or non-living, everything is composed of atoms. Let us take a look at the structure of atom.

Atoms

Atom is a Greek word which means "indivisible." The Greeks believed that matter can be broken down into very small invisible particles called atoms. Greek philosophers such as Democritus and John Dalton put forward the concept of the atom.

Democritus explained the nature of matter. He also proposed that all substances are made up of matter. He stated atoms are constantly moving, invisible, minuscule particles that are different in shape, size, temperature and cannot be destroyed.

Learn the concept of the Atomic number here in detail.

Later in the year 1808, John Dalton proposed the atomic theory and explained the law of chemical combination. By the end of 18th and the early 20th centuries, many scientists such as J.J Thomson, Gold stein, Rutherford, Bohr among others developed and proposed several concepts on "atom."

Atom is the smallest unit of matter that is composed of a *positively charged centre termed as "nucleus"* and the central nucleus is *surrounded by negatively charged electrons*. Even though an atom is the smallest unit of matter but it retains all the chemical properties of an element. For example, a silver spoon is made up of silver atoms with few other constituents. A silver atom obtains its properties from tiny subatomic particles that it is composed of.

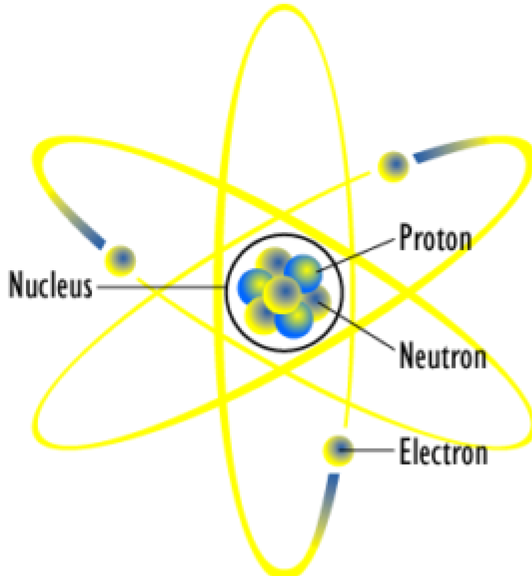
Atoms are further arranged and organized to form larger structures known as molecules. Atoms, molecules adhere to the general chemistry and physics rule even when they are part of living breathing human body. Let us study atoms and molecules to further understand how the atoms react, behave and interact in cells form an essential part of the living and nonliving world.

Learn more about Thomson's Model of Atom in detail.

Structure of Atom

Structure of an atom can be basically divided into two parts:

- an atomic nucleus
- extra nucleus part



The tiny atomic nucleus is the centre of an atom constituting positively charged particles **"protons"** and uncharged particles **"neutrons."** On the other hand, the extra nucleus part is a much larger region which is composed of a cloud of negatively charged particles called an **electron.** Electrons revolve around the orbit or centre of the nucleus. The attraction between the protons and electrons holds the structure of an atom together.

Learn how electrons are distributed in different shells in detail here.

Generally, all atoms are composed of these three subatomic particles except hydrogen. *Hydrogen is an exception to all atoms as it just contains one proton and one electron but lacks neutrons.* The number of protons indicates what element an atom is whereas the number of electrons indicates the type of reactions will happen in an atom.

The atomic nucleus in the structure of the atom is composed of a fixed number of protons and the proton attracts the same number of electrons thereby making an atom electrically neutral. Ions are formed by addition or removal of electrons from an atom.

Learn about the Disadvantages of Rutherford's Atomic Model here.

Discovery of an Electron

A British Physicist named J.J Thompson in the year 1897 proposed that an atom constitutes of at least one negatively charged particle. He named it "corpuscles" which was later called "electron."

An electron is represented "e" and it is known to contribute to the negative charge of an atom. The absolute charge of an electron is the negative charge of 1.6×10^{-19} coulombs. The relative mass of an electron is 1/1836, thus the mass of an electron is very small and is considered as 0.

Discovery of Proton

Proton was discovered by Rutherford when he conducted the famous gold foil experiment. Initially, in 1886 Goldstein discovered the presence of positively charged rays while conducting an experiment in the discharged tube by using perforated cathode. The rays were named as anode rays or canal rays. A series of experiments led to the discovery of protons. Protons are known as the particles that contribute to the positive charge of the atom.

Proton is represented by "p". The absolute charge of a proton is the positive charge of 1.6×10^{-19} coulomb. The mass of a proton is 1.6×10^{-24} g and is considered 1 that is mass of a hydrogen atom.

Discovery of Neutron

Neutron was not discovered until the year 1932. It was discovered by James Chadwick by using scattered particle to calculate the mass of the neutral particle. The subatomic particle "neutron" is present in an atom's nucleus. Neutron is represented by "n" and is considered a neutral particle. The mass of a neutron is measured to be 1.6×10^{-24} g.

Gram is not a very appropriate unit for the calculation of such tiny subatomic particles. Therefore they are alternatively calculated in Dalton or amu (atomic mass unit). Neutron and a proton have a mass that is nearly 1 amu.

Learn more about Bohr's Atomic Model here.

Solved Question for You

Question: What is the net charge of an atom?

Ans: There is no net charge of an atom. Electrons are the negatively charged particle whereas protons are the positively charged particles. The equal positive charge of the proton and the negative charge of the electron cancel each other. Therefore, the atom has no net charge. In an atom that is neutral, the number of electrons revolving around the nucleus and the number of protons inside the nucleus are equal in number.

Atomic Number

Elements are the building blocks for all the matter you see in the world. Now, the question is how can you distinguish these elements? The answer is "Atomic Number." Every element has their unique atomic number that helps to distinguish between two different elements. Let us study further to know about its significance.

History of Atomic Number

We all have heard about periodic tables and how the elements are arranged in a periodic table in a very informative order. The history of periodic table dates back to late 1860's when Dmitri Mendeleev first discovered the periodic law. This soon became a major organizing concept of chemical sciences. However, the law had some drawbacks. Some of the elements when arranged according to Mendeleev's law were out of sequence.

It was not until Wilhelm Röntgen and his discovery of x rays in 1895 that helped other scientists to further research on Mendeleev's discovery. English physicist H.G.J. Moseley in 1913 studied wavelengths of X-rays emitted by different chemical elements. Moseley hypothesized the modern periodic table given by Mendeleev was on the basis of the number of protons present in the atomic nucleus.

This hypothesis forms the basis of atomic number. Each and every element has a unique number that represents the number of protons present within an atomic nucleus.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers

Atomic Number

We know that an atom consists of electrons, protons, and neutrons. Atomic number is one of the fundamental properties of an atom. Each atom can be characterized by a unique atomic number. It is represented by the letter "Z."

The total number of protons present in the nucleus of an atom represents the atomic number of a particular atom. Every atom of a particular element is composed of the same number of protons and therefore have the same atomic number. However, atoms of different elements have unique atomic numbers that vary from one element to the other.

An atom does not have any net charge and is thus electrically neutral. This means that the number of electrons will be equal to the number of protons present in an atom thereby making an atom electrically neutral.

"Atomic Number = No. of Protons = No. of Electrons"

For example, each atom of oxygen has 8 protons in their nucleus so the atomic number is 8. Similarly, the atomic number of carbon is 6 because all atoms of carbon have 6 protons in their nucleus.

Atomic numbers are whole numbers because it is the total number of protons and protons are generally units of matter. It ranges from 1 to 118. It starts with hydrogen and ends with the heaviest known element Oganesson (Og).

Theoretically, the atomic numbers can be increased if more elements are discovered. However, with the addition of more number of protons and neutrons, the elements become prone to radioactive decay.

Importance of Atomic Numbers

- Helps in the identification of a particular element of an atom.
- Forms the basis of the arrangement of the elements. The elements are arranged in increasing order of the atomic numbers of the elements.
- Helps in the determination of the properties of any element. However, valence electron determines the chemical bondingbehaviourr of an element.

Examples of Atomic Numbers

It does not matter how many electrons or neutrons present in an atom. Atomic number is always determined by the number of protons. For example

- If an atom has one proton then the atomic number is 1 and the element is hydrogen.
- Every carbon atom will have 6 protons and the atomic number of carbon is 6.
- Every cesium atom has 55 protons and hence the atomic number is 55.

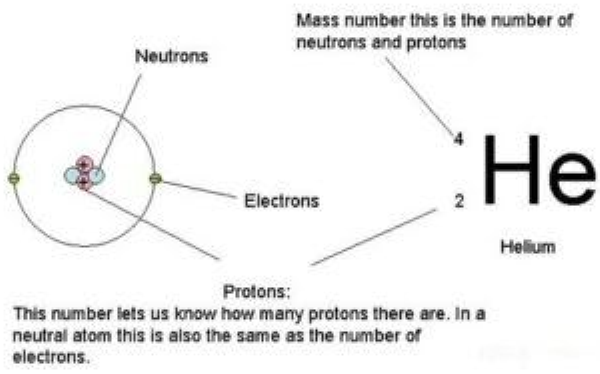
Atomic numbers of few elements:

- Hydrogen= 1
- Helium=2
- Lithium=3
- Beryllium=4
- Boron=5
- Carbon=6
- Nitogen=7
- Oxygen=8

Finding the Atomic Number

It can be calculated on the basis of information given. Firstly, an atomic number can be calculated by the number of protons in an atom of an element.

Secondly, if an element name or symbol is given then it can be identified by using a periodic table chart. It is a simple positive and whole number and is denoted by the letter z. For example, if the symbol Al (aluminum) is given. If you search the periodic table you can see the atomic number of Al is 13. It is easy to find out in a periodic table because the elements are present in increasing order.



(Source Credit: Wikipedia)

Thirdly it can be identified by the isotope symbol. Isotope symbol can be written in many ways but the symbol is always given while writing an isotope. For example, if an isotope of carbon is given as ^{14}C or $^{14}_{6}C$.

- ^{14}C- If the isotope is written this way we know that the symbol represents carbon atom so the atomic number of carbon atom is 6.
- $^{14}_{6}C$- If the isotope is written this way the atomic number will be the whole number and smaller of the two numbers which is 6. It is usually the subscript of the element symbol.

Solved Examples for You

Question 1: If the atomic number of sodium is 11. Find out how many electrons and protons are present in a calcium atom.

Solution: We know that "Atomic Number = No. of Protons = No. of Electrons"

Thus, number of electrons=11 and number of protons =11

Question 2: What is the atomic number of chlorine?

a. 18
b. 19
c. 17
d. 16

Solution: The correct answer is "C". The atomic number of chlorine is 17 because the number of protons in a chlorine atom is 17.

Bohr's Model of Atom

Quantum mechanics, Quantum physics, the theory of relativity, etc are the modern subjects that interests, astound, and confuse almost everybody. These topics form the basis of modern physics. However, the very first-time quantum theory was incorporated in Bohr's Model of an atom or Bohr atomic model. Later this model became the predecessor of complete quantum mechanical models.

The physicist Niels Bohr said, **"Anyone who is not shocked by quantum theory has not understood it."** He also said, **"We must be clear that when it comes to atoms, language can only be used as in poetry."** So what exactly is this Bohr atomic model? Let us find out!

Bohr atomic model and the models after that explain the properties of atomic electrons on the basis of certain allowed possible values. The model explained how an atom absorb or emit radiation when electrons on subatomic level jump between the allowed and stationary states. German-born physicists James Franck and Gustav Hertz obtained the experimental evidence of the presence of these states.

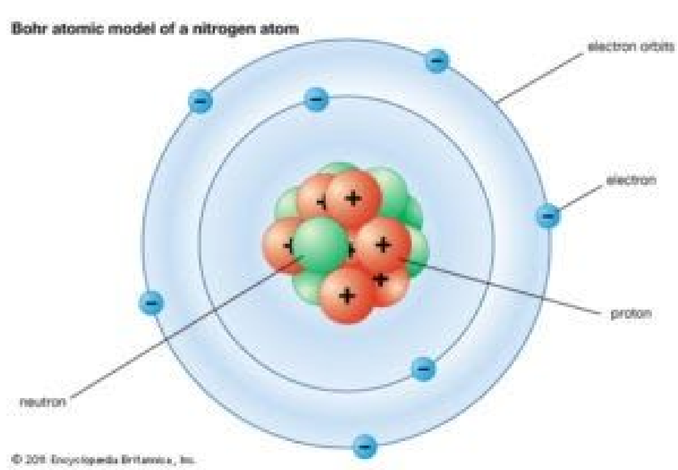
Bohr Atomic Model

A Danish physicist named Neil Bohr in 1913 proposed the Bohr atomic model. He modified the problems and limitations associated with Rutherford's model of an atom. Earlier in Rutherford Model, Rutherford explained in an atom a nucleus is positively charged and is surrounded by electrons (negatively charged particles).

Learn about Rutherford's Atomic Model here in detail.

The electrons move around in a predictable path called **orbits**. Bohr modified Rutherford's model where he explained that electrons move around in fixed orbital shells. Furthermore, he explained that each orbital shell has fixed energy levels. Therefore, Rutherford basically explained a nucleus of an atom whereas Bohr took the model one step ahead. He explained about electrons and the different energy levels associated with it.

According to Bohr Atomic model, a small positively charged nucleus is surrounded by revolving negatively charged electrons in fixed orbits. He concluded that electron will have more energy if it is located away from the nucleus whereas the electrons will have less energy if it located near the nucleus.



Bohr's Model of an Atom (Source Credit: Britannica)

Postulates of the Bohr Atomic Model

- Electrons revolve around the nucleus in a fixed circular path termed "orbits" or "shells" or "energy level."
- The orbits are termed as "stationary orbit."

- Every circular orbit will have a certain amount of fixed energy and these circular orbits were termed orbital shells. The electrons will not radiate energy as long as they continue to revolve around the nucleus in the fixed orbital shells.
- The different energy levels are denoted by integers such as n=1 or n=2 or n=3 and so on. These are called as quantum numbers. The range of quantum number may vary and begin from the lowest energy level (nucleus side n=1) to highest energy level. Learn the concept of an Atomic number here.
- The different energy levels or orbits are represented in two ways such as 1, 2, 3, 4… or K, L, M, N….. shells. The lowest energy level of the electron is called the ground state. Learn the concept of Valency here in detail here.
- The change in energy occurs when the electrons jump from one energy level to other. In an atom, the electrons move from lower to higher energy level by acquiring the required energy. However, when an electron loses energy it moves from higher to lower energy level.

Therefore,

- 1^{st} orbit (energy level) is represented as K shell and it can hold up to 2 electrons.
- 2^{nd} orbit (energy level) is represented as L shell and it can hold up to 8 electrons.
- 3^{rd} orbit (energy level) is represented as M shell and it can contain up to 18 electrons.
- 4^{th} orbit (energy level) is represented as N Shell and it can contain maximum 32 electrons.

The orbits continue to increase in a similar manner.

Watch and Learn more about Modern Atomic Theory

Distribution of Electrons in Orbits or Shells:

Electronic distribution of various orbits or energy levels can be calculated by the formula **$2n^2$.** Here, **'n'** denotes the number of orbits.

- The number of electrons in K shell (1st orbit) can be calculated by $2n^2 = 2 \times 1^2 = 2$. Thus, maximum number of electrons in 1st orbit = 2
- Similarly, The number of electrons in L shell (2nd orbit)= $2 \times 2^2 = 8$. Thus, maximum number of electrons in 2nd orbit = 8

We can determine the maximum number of electrons in a similar way.

Read about Thomson's Model of an Atom, the very first model of an Atom by J.J. Thomsons.

Limitations of Bohr's Model of an Atom:

Bohr atomic model had few limitations. They are:

- Failure to explain Zeeman Effect (how atomic spectra are affected by magnetic fields).
- It contradicts Heisenberg Uncertainty Principle.
- Unable to explain how to determine the spectra of larger atoms.

What are Isotopes? Learn the concept of Isotopes and Isobars.

Solved Examples for You

Example 1: Calculate the maximum number of electrons an o shell can hold.

Solution: We know that O shell means 5th shell.

Therefore, n=5. Applying the formula $2n^2 = 2 \times 5^2 = 50$

Thus, the maximum number of electrons O shell can hold is 50.

Example 2: What happens when an electron changes its orbit from outer to inner energy? Energy remains constant

1. absorbed
2. no change
3. released

Solution: The answer is 4**.** Energy is released when an electron jumps from higher to lower energy level.

Charged Particles in Matter

What happens when you comb your dry hair and then bring the comb near few small pieces of paper? Combing your dry hair will move some of the subatomic particles called electrons from hair to the comb. The charge of an electron is negative and thus the comb will also gain the charge of an electron (negative charge). Finally, when the comb is brought near the bits of paper, the bits of paper will get attracted to the comb.

Similarly, what happens when you rub a silk cloth on a glass rod and then the glass rod is brought near an inflated balloon? The electrons from the glass rod are transferred to the silk cloth. Hence, the glass rod becomes positively charged. When the positively charged glass rod is brought near the inflated balloon, it will be attracted to the glass rod.

These activities prove two objects when rubbed together, the objects become electrically charged. However, how do the objects become electrically charged? All matter is composed of divisible particles called atoms. The atoms are composed of tiny charged subatomic particles named electrons, protons, and neutrons.

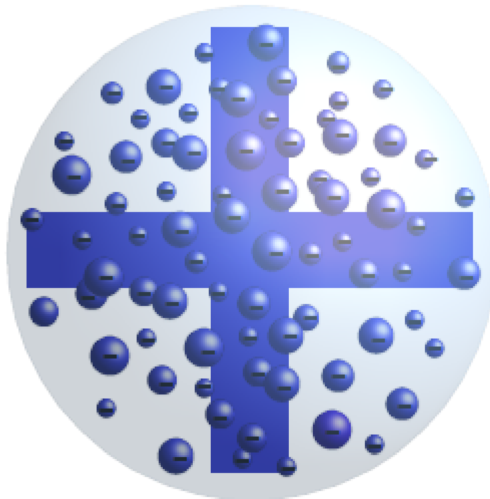
Earlier it was postulated by John Dalton that atoms are indivisible particles. However, the contribution of many scientists by the very end of 1800's or the start of 1900's gradually started revealing that atoms are divisible into tiny subatomic particles.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

Discovery of Subatomic particles

Discovery of Existence and Charge of an Electron



Thomson's Atomic Model (Source Credit: Wikipedia)

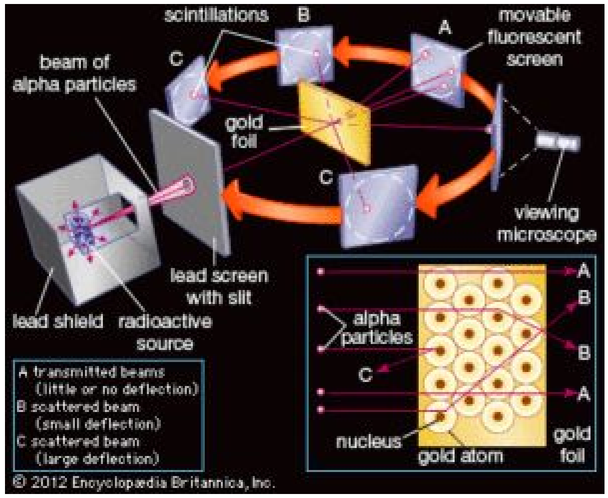
A British Physicist named **J.J Thompson** in the year 1897 proposed that an atom constitutes of at least one negatively charged particle. He discovered electrons while working on an experiment. In the experiment, gas was passed at low pressure through a glass discharged tube.

Two electrodes were placed at the two opposite ends of the glass discharge tube which was further connected to high voltage supply (battery). The electrode joined to the negative end was called cathode and the electrode joined to the positive end was called anode.

Thomson observed a stream of negatively charged particles emitted from the cathode. Furthermore, the stream of particles moved towards the anode. The streams of negatively charged particles discharged from the cathode end of the discharge tube are called cathode rays. He named it as "corpuscles" which was later called "electron."

An electron is represented by "e" and it is known to contribute to the negative charge of an atom. The absolute charge of an electron is the negative charge of 1.6×10^{-19} coulombs. The relative mass of an electron is 1/1836 (1/2000), thus the mass of an electron is very small and is considered as 0. Therefore, the mass of an electron is negligible and the charge of an electron is -1.

Discovery of Proton



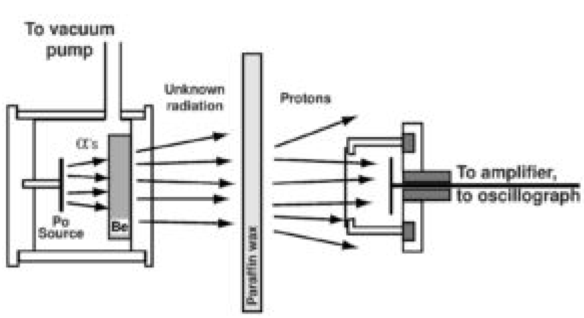
Rutherford's Gold Foil Experiment (Source Credit: Britannica)

Proton was discovered by **Rutherford** when he conducted the famous gold foil experiment. Initially, in 1886 Goldstein discovered the presence of positively charged rays while conducting an experiment in the discharged tube by using perforated cathode. The experiment was similar to J.J. Thomson but with different situations. This led to the emission of charged particles from the anode.

The positively charged radiations produced from the anode end of the discharge tube are called canal rays. The rays were named as anode rays or canal rays. A series of experiments led to the discovery of protons. Protons are known as the particles that contribute to the positive charge of the atom.

The mass of the particles was found to be 2000 times more than that of an electron. Protons were represented by the letter "p". The absolute charge of a proton is the positive charge of 1.6×10^{-19} coulomb. The mass of a proton is 1.6×10^{-24} g and is considered 1 (which is mass of a hydrogen atom). The mass of a proton is considered one unit and it's charge is considered to be +1.

Discovery of Neutron



Chadwick Experiment (Source Credit: Wikimedia)

Neutron was not discovered until the year 1932. It was discovered by **James Chadwick** by using scattered particle to calculate the mass of the neutral particle. The subatomic particle "neutron" is present in an atom's nucleus. Neutron is represented by the letter "n" and is considered a neutral particle.The mass of a neutron is measured as 1.6 x 10^{-24}g.

Gram is not a very appropriate unit for the calculation of such tiny subatomic particles. Therefore, they are alternatively calculated in Dalton or amu (atomic mass unit). A Neutron and a proton have a similar mass that is nearly 1 amu.

Summary of Subatomic Particles:

Name	Representation	Absolute Charge	Relative Charge	Mass (Kg)	Relative mass
Electron	e^-	-1.60×10^{-19} C	-1	9.1×10^{-31} kg	1/1840 (negligible)
Proton	P^+	1.60×10^{-19} C	+1	1.672×10^{-27} kg	1u
Neutron	n	0.0 C	0	1.674×10^{-27} kg	1u

Conclusion

The discovery of the subatomic particles raised many questions. One of them was how these subatomic particles are arranged in an atom. This led to the proposal of different models of these charged particles. These form the basis of the vast subject of "structure of an atom."

Solved Questions For You

Q1: Which of the options represent the properties of electrons?

1. the relative charge of -1
2. negligible mass
3. the charge is opposite and equal to a proton
4. all of these

Ans: The answer is 4 (all of these), as per the defined nature of electrons.

Isobars

Do you know isobars were originally called as “isobares”? Isobars are atoms or nuclides of separate chemical elements having the same number of nucleons (protons+ neutrons). The name was given by Alfred Walter Stewart in 1918. It is originally taken from the combination of Greek words- isos means equal and bar means weight.

More About Isobars

Atoms of chemical elements having same atomic mass but a different atomic number are called Isobars. The sum of the number of protons and neutrons together form the atomic mass. Therefore, we can also say the number of nucleons present in the nucleus is equal to the atomic mass of an atom. It will have the same number of nucleons.

The number of protons and neutrons alone will vary but the number of nucleons or the sum of protons and neutrons in isobars will always be same. Isobars always have different atomic structure because of the difference in atomic numbers. The number of neutrons makes up the difference in the number of nucleons. Therefore, they are always different chemical elements having same atomic masses. Thus, isobar has different chemical properties.

Learn the concept of an Atomic number here.

By now we know that the atomic masses in isobars are same. Therefore, they have similar physical properties. For example, the isobars iron and nickel. Iron and nickel have atomic number 26 and 28 respectively. However, the mass number is 58.

Another example is argon and calcium with atomic number 18 and 20, respectively. However, they have the same atomic mass 40. Furthermore, sometimes there are series of an element that has same atomic mass. For example, Cobalt (Co), Nickel (Ni), Copper (Cu), and Iron (Fe) have same atomic mass 64 but the atomic number varies.

Examples

- ${}_{18}Ar^{40}$ ${}_{19}K^{40}$ ${}_{20}Ca^{40}$
- ${}_{32}$ Ce^{76}, ${}_{34}Se^{76}$
- ${}_{11}Na^{24}$ ${}_{12}Mg^{24}$
- ${}_{26}Fe^{58}$, ${}_{27}$ Ni^{58}
- ${}_{27}Co^{64}$ ${}_{28}Ni^{64}$

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr’s Model of Atom
- Charged Particles in Matter

- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

Difference between Isotopes and Isobars

Isotope	Isobar
Isotopes are atomic structures of same elements having a different mass number/atomic mass	Isobars are different chemical elements having same atomic mass.
Atomic numbers of all isotopic forms of a single element are equal.	Atomic numbers of isobars vary from each other
They are the same chemical element but their forms are different	They are different elements altogether.
All isotopic forms of single elements have different physical properties.	Physical properties can be similar to each other

Conclusion

Isotope and isobar indicate the relationships between chemical elements. Therefore, it is necessary to understand the difference between them. Isotopes provide the relationship of different forms of the single chemical element. Isobars indicate the relationship between different chemical elements but having the same atomic mass and thus the same physical properties.

Solved Questions for You

Q1: Isobars have the same number of ______?

1. Protons
2. Electrons
3. Neutrons
4. Nucleons

Ans: The answer is option 4 (Nucleons). The number of protons and neutrons alone will vary in isobars but the number of nucleons or the sum of protons and neutrons will always be same.

Q2. Which of the following statements about Isobars is false?

1. Their atomic numbers vary from each other.
2. Chemically they are the same element but their forms are different.
3. Physical properties can be similar to each other.
4. They are different chemical elements having same atomic mass.

Ans: The statement 2 is false. Isobars are different elements altogether.

Isotopes

We all have come across the word "radioactivity" somewhere or the other. Do you know an isotope form the basis of radioactivity? Many times terms like nuclear energy, nuclear reactors, and nuclear weapons have popped up in news and textbooks. We have read stories and watched movies about superheroes born out of some sort of radiation exposure. What is radioactivity and how it is associated with isotopes?

Suggested Videos

Radioactivity is one of the properties of atoms. Every radioactive atom will have an unstable nucleus. Thus it has a tendency to release subatomic particles to obtain stability thereby releasing radiation or energy in the process. Elements or atoms are present both in radioactive and non-radioactive variety.

The number of neutrons in both the varieties will be different. These different varieties of the same element are termed isotope. Read through this article to find out more about isotopes.

Browse more Topics Under Structure Of Atom

- ntroduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom

- Charged Particles in Matter
- Isobars
- Mass Number
- Neutrons
- Rutherford’s Model of an Atom
- Thomson’s Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr’s Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

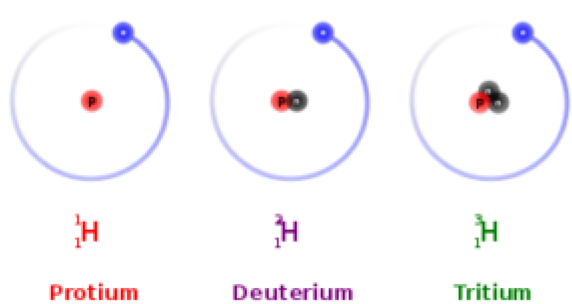
Isotopes

Elements with the same atomic number but a different mass number are defined as “Isotopes”. The number of protons and neutrons combined together is called atomic mass or mass number of an element, whereas the total number of protons gives the atomic number of an element.

In a certain element, the number of protons will always remain constant. However, the number of neutrons can change. The number of neutrons varies but the number of protons always remains same in an isotope of a single element. Thus, the definition “Elements with the same atomic number but a different mass number are termed as Isotopes.”

For Example

Hydrogen has three most stable isotopic forms- protium, deuterium, as well as tritium. All the three isotopic forms of hydrogen have the same number of protons but vary in the number of neutrons. The number of neutrons in protium is zero, the number of neutrons in deuterium is one and the number of neutrons in tritium is two.



Isotopes of Hydrogen (Source: Wikipedia)

Carbon has three isotopic forms- Carbon-12, Carbon-13, as well as Carbon-14. The numbers 12, 13, and 14 represents the atomic masses of different isotopic forms of carbon. As discussed, atomic number is the unique property by which we can determine the element. Therefore, the atomic number 6 of carbon in all the forms is constant. Carbon-12 is the stable isotope of the carbon element whereas carbon-14 is the radioactive isotope.

Examples of Isotopes

- Isotopic forms of Oxygen – Oxygen -16, Oxygen -17, Oxygen -18
- Isotopic forms of Uranium- U-235, U-238
- Chlorine- 35, Chlorine – 37 are the isotopic forms of chlorine
- Isotopic forms of Fluorine – Fluorine 17, Fluorine 18, Fluorine 19
- Hydrogen – 1, Hydrogen – 2, Hydrogen – 3 are the isotopic forms of Hydrogen
- Isotopic forms of Carbon-Carbon– 12, Carbon – 13, Carbon- 14

There are approximately 275 different isotopes of 81 stable elements. There are more than 800 natural and synthetic radioactive isotopes present. A single element present in the periodic table can have multiple isotopic forms.

Chemical properties

Generally, the chemical properties of isotopes of any element are almost identical. The exception to this case is the isotopes of hydrogen because the numbers of neutrons have a major effect on the size of the nucleus of a hydrogen atom.

Physical properties

The physical properties of isotopes in a particular element vary from each other. This is because the physical properties of any isotope depend on the mass. The mass of each isotope of a single element varies from one another. Processes such as fractional distillation and diffusion are used to separate isotopes from one another. We make use of the fact that isotopes have different physical properties.

Uses of Isotopes

- Carbon dating makes use of Carbon-14, an isotope of Carbon. This isotope of carbon is present in the atmosphere as radioactive carbon. The amount of carbon-14 obtained in fossils help palaeontologists to calculate the age of the fossils.
- Uranium Isotopes are popular for its use in nuclear reactors. U-235 is used as a fuel in nuclear reactors.

- Radioactive Isotopes are used for medicinal purposes. They are used for detection of tumours, blood clots, etc. Arsenic-74, An isotope of arsenic, is used for determining the presence of a tumour. Similarly, sodium-24 is used for the detection of blood clots.
- Cobalt (cobalt-60) isotope of carbon is applied in cancer treatments.
- Iodine (Iodine-131) isotope of carbon helps in the treatment of goitre.

Learn more about Isobars here.

Solved Question for You

Question: Isotopes of a single element vary in the number of

1. Neutrons
2. Protons
3. Electrons
4. All of the above

Solution: The answer is 1 (Neutrons). The number of protons will always constant in a single element. However, the number of neutrons can change. The number of neutrons varies in isotopes of an element but the number of protons always remains same.

Mass Number

Did you know that protons and neutrons together are termed as nucleons? Nucleons reside inside the nucleus and are denoted by the mass number is denoted as "A." All atoms of a particular chemical element will have the same atomic number, number of protons present in the nucleus of an atom, but different mass numbers. But what exactly are these mass numbers? Let's know more about it.

Mass Number

It is defined as the sum of protons and neutrons. The mass number is almost equal to the atomic mass of a particular atom. Therefore it can be written as

Mass no. of an atom = No. of protons + No. of neutrons

Thus, it represents the total number of neutrons present in the nucleus of an atom.

Use of Mass Number

Mass number helps to give an idea of the isotopic mass. Isotopic mass is measured in atomic mass units or “u.” An isotope of an element will have the same atomic number but a different mass number. Isotope mainly differs in the number of neutrons. Different isotopes of the same element will have a different mass number. However, isotopes of different elements can have the same mass number such as carbon-14 (6 protons + 8 neutrons) and nitrogen-14 (7 protons + 7 neutrons).

Properties of Mass Number

The various properties of mass number are enumerated here. Let us have a look,

- Sum of protons and neutrons provide this number of a certain element.
- It is represented by the letter A.
- Protons and Neutrons are together termed as nucleons.
- Example: Atoms of a carbon consist of 6 protons and 6 neutrons. Therefore, the mass number of Carbon is 12.
- The number of neutrons may vary in an element. However, the total number of protons is same in all atoms of an element. Therefore, the atoms of the same element with a same atomic number but a different mass number are termed as isotopes.
- Generally, atomic mass and mass numbers are two different terms and may vary slightly. In most cases, they are not the same. However, the weight of an electron is almost negligible so we can consider the atomic mass of an atom to be almost equal to its mass number.

X= Chemical Symbol of a particular element

N= Neutron number

Z= Atomic Number= Number of protons

A= Mass number= Z+N

Solved Examples for You

Q: The atomic numb er of aluminum is 13 and atomic mass is 27u. Calculate the number of protons and neutrons.

Sol: Given, Atomic number= 13

We know that, Z= Atomic Number= Number of protons

Therefore, number of proton=13

We know that Mass no. of an atom = No. of protons + No. of neutrons

Mass no= atomic mass= 27 u

27 u = 13 + n Or, n = 27-13 = 14

Thus, Number of proton=13 and number of neutron= 14

Q: Calculate the number of neutrons for the symbol $^{35}Cl_{17}$

Sol: We know that, Z= Atomic Number= Number of protons

Atomic number=Z= 17

Given, Mass no= A= 35

We know that Mass no. of an atom = No. of protons + No. of neutrons

35= 17 + no of neutrons

Or, No of neutrons= 35-17= 18

Therefore, Number of neutrons=18

Neutrons

Do you know what type of charge a neutron has? Positive, negative or is it neutral? Well, neutrons play a crucial role in nuclear power production. Atom, the smallest unit of matter, is composed of a centre – the nucleus which is surrounded by the subatomic particles called electrons. A nucleus, where all the mass of an atom is concentrated, in turn, is composed of protons and neutrons. Let’s find out more about these neutral subatomic particles below.

Who Discovered Neutron

It was observed after the discovery of electrons and protons that all mass of an atom cannot be accounted on the basis of just electrons and protons. Rutherford was unable to explain any other atoms of different elements except for hydrogen on the basis of protons and electrons.

Scientists realized that the model given by Rutherford was incomplete. For example, the atomic mass of helium should be double of the hydrogen atom. However, on calculation, the mass of helium atom was found to be four times more than that of hydrogen. Therefore, Rutherford postulated the presence of neutral particles in an atom. He believed that the neutral particle had the same mass as that of a proton. However, experimental evidence was lacking.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

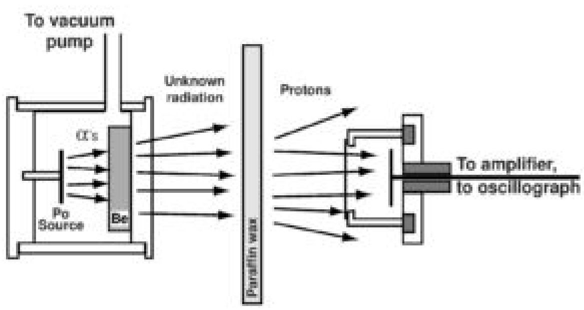
Experiments Conducted and their Observations

Later on, various experiments were conducted and several theories and observations resulted in the discovery of neutrons. Few noteworthy experiments and their observations are listed below.

- Bothe and H. Becker while bombarding beryllium with alpha particles observed electrically neutral radiation in the year 1930. The scientists believed it to be photons with high energies, also known as gamma rays.
- In the year 1932, Irène and Frédéric Joliot-Curie demonstrated that this ray has the potential to eject protons when it strikes paraffin or any H-containing compounds.
- This however raised many questions such as how a photon that does not have mass can release a particle 1836 times heavier than an electron (protons). Therefore, it was concluded that the ejected rays cannot be photons.

In the year 1932, James Chadwick conducted the same experiment but he used different bombardment targets other than paraffin. After research and analysis of the various targets, he discovered the presence of a new particle. This new particle has the mass similar to that of the proton but no charge. He named these particles neutrons.

Furthermore, he derived the nearly accurate mass of this new particle by applying the conservation of energy and momentum. He measured the mass of a neutron to be nearly same as that of a proton. $\mathbf{M_N = 1.00866\ u = 1.6749 \times 10^{-27}\ kg}$



Chadwick Experiment (Source Credit: Wikimedia)

Function of Neutrons

- James Chadwick concluded the nucleus consists of proton and neutron. Together these two subatomic particles are termed as nucleon.
- The discovery helped to understand the atomic mass and the atomic number more clearly. It also helped in understanding the concept of isotopes which forms the basis of radioactivity.

The beryllium reaction when it is bombarded with the alpha particle is:

$$\mathbf{{}^{9}_{4}Be + {}^{4}_{2}\alpha \rightarrow [{}^{13}_{6}C] \rightarrow [{}^{12}_{6}C] + {}^{1}_{0}n}$$

X= Chemical Symbol of a particular element

N= Neutron number

Z= Atomic Number= Number of protons

A= Atomic Mass number= Z+N

After the discovery of the neutron, every chemical element present in the periodic table was modified and written accordingly. For instance, uranium was written as $^{235}_{92}U$. This means that a single atom of uranium contains 235 nucleons (protons + neutrons). Out of 235 nucleons, 92 are protons and 143 neutrons.

Properties of Neutrons

- Neutrons reside inside the nucleus with protons.
- Hydrogen atom lack nucleus.
- Mass of a neutron is equivalent to the mass of a proton. Therefore, the mass of neutron= **1.676×10^{-27} kg = 1.676×10^{-24} g.** However, the mass of an electron is negligible and was found to be 1/1840 times that of the proton.
- Neutrons do not have charge and thus they are electrically neutral.

Conclusion

Subatomic Particle	**Representation**	**Mass**	**Charge**
Electron	e^-	9.1×10^{-31} kg	-1.60×10^{-19} C
Proton	p^+ (H^+)	1.672×10^{-27} kg	1.60×10^{-19} C
Neutron	n^0	1.674×10^{-27} kg	0.0 C

Solved Questions for You

Question: Neutron was emitted by bombardment of beryllium and _______ ?

1. alpha particles

2. gamma particles
3. beta particles
4. none

Solution: The answer is 1 (alpha particles).

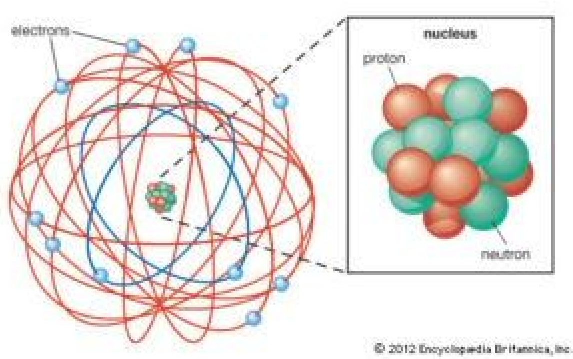
Rutherford's Model of an Atom

We know a structure of an atom consists of electrons, protons, and neutrons. This was accurately presented after several scientists came up with different models. The classic model of an atom was given by Ernest Rutherford called the Rutherford atomic model or Rutherford model of the atom. However, it is not considered the accurate representation of an atom anymore. Let us know more about this model.

Watch Video Lecture on Rutherford Atomic Model

Rutherford Atomic Model

Rutherford proposed that an atom is composed of empty space mostly with electrons orbiting in a set, predictable paths around fixed, positively charged nucleus.



Rutherford's Atomic Model (Source Credit: Britannica)

History

The concept of atom dates back to 400 BCE when Greek philosopher Democritus first conceived the idea. However, it was not until 1803 John Dalton proposed again the idea of the atom. But at that point of time, atoms were considered indivisible. This idea of an atom as indivisible particles continued until the year 1897 when British Physicist J.J. Thomson discovered negatively charged particles which were later named electrons.

To understand this concept in more detail you should read the Structure of Atom

He proposed a model on the basis of that where he explained electrons were embedded uniformly in a positively charged matrix. The model was named plum pudding model. However, J.J. Thomson's plum pudding model had some limitations. It failed to explain certain experimental results related to the atomic structure of elements.

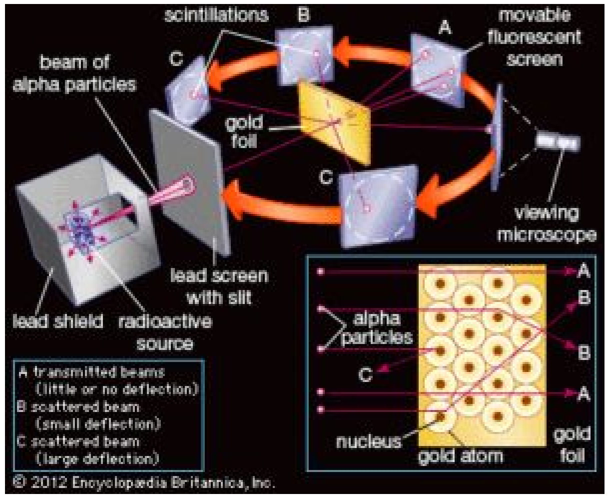
A British Physicist "Ernest Rutherford" proposed a model of the atomic structure known as Rutherford's Model of Atoms. He conducted an experiment where he bombarded α-particles in a thin sheet of gold. In this experiment, he studied the trajectory of the α-particles after interaction with the thin sheet of gold.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

Rutherford Atomic Model Experiment

In Rutherford's experiment, he bombarded high energy streams of α-particles on a thin gold foil of 100 nm thickness. The streams of α-particles were directed from a radioactive source. He conducted the experiment to study the deflection produced in the trajectory of α-particles after interaction with the thin sheet of gold. To study the deflection, he placed a screen made up of zinc sulfide around the gold foil. The observations made by Rutherford contradicted the plum pudding model given by J.J. Thomson.



Rutherford's Gold Foil Experiment (Source Credit: Britannica)

Observations of Rutherford Model Experiment

On the basis of the observations made during the experiment, Rutherford concluded that

- Major space in an atom is empty **–** A large fraction of α-particles passed through the gold sheet without getting deflected. Therefore, the major part of an atom must be empty.
- The positive charge in an atom is not distributed uniformly and it is concentrated in a very small volume **–** Few α-particles when bombarded were deflected by the gold sheet. They were deflected minutely and at very small angles. Therefore he made the above conclusion.
- Very few α-particles had deflected at large angles or deflected back. Moreover, very few particles had deflected at 180°. Therefore, he concluded that the positively charged particles covered a small volume of an atom in comparison to the total volume of an atom.

[What is the difference between the Thomsons and Rutherford Atomic model?]

Postulates of Rutherford atomic model based on observations and conclusions

- An atom is composed of positively charged particles. Majority of the mass of an atom was concentrated in a very small region. This region of the atom was called as the **nucleus** of an atom. It was found out later that the very small and dense nucleus of an atom is composed of neutrons and protons.
- Atoms nucleus is surrounded by negatively charged particles called **electrons**. The electrons revolve around the nucleus in a fixed circular path at very high speed. These fixed circular paths were termed as **"orbits."**

- An atom has no net charge or they are **electrically neutral** because electrons are negatively charged and the densely concentrated nucleus is positively charged. A strong electrostatic force of attractions holds together the nucleus and electrons.
- The size of the nucleus of an atom is very small in comparison to the total size of an atom.

Learn more about Thomsons Model of an Atom, by J.J. Thomsons, which was the first model of Atom.

Drawbacks of Rutherford Model

Limitations of Rutherford Atomic Model

Rutherford's experiment was unable to explain certain things. They are:

- Rutherford's model was unable to explain the stability of an atom. According to Rutherford's postulate, electrons revolve at a very high speed around a nucleus of an atom in a fixed orbit. However, Maxwell explained accelerated charged particles release electromagnetic radiations. Therefore, electrons revolving around the nucleus will release electromagnetic radiation.
- The electromagnetic radiation will have energy from the electronic motion as a result of which the orbits will gradually shrink. Finally, the orbits will shrink and collapse in the nucleus of an atom. According to the calculations, if Maxwell's explanation is followed Rutherford's model will collapse with 10^{-8} seconds. Therefore, Rutherford atomic model was not following Maxwell's theory and it was unable to explain an atom's stability.
- Rutherford's theory was incomplete because it did not mention anything about the arrangement of electrons in the orbit. This was one of the major drawbacks of Rutherford atomic model.

Conclusion

Even though the early atomic models were inaccurate and could not explain the structure of atom and experimental results properly. But it formed the basis of the quantum mechanics and helped the future development of quantum mechanics.

Solved Questions for You

Question: Name the part of an atom discovered by Rutherford α-particles scattering experiment

1. Electrons
2. Protons
3. Neutrons

4. Nucleus

Answer**:** The answer is 4. Rutherford α-particles scattering experiment led to the discovery of nucleus.

Thomson's Model of an Atom

All of us have seen a plum pudding and a watermelon in our daily life. Do you know one of the early models of an atom has been compared to a plum pudding, raisin pudding, and even a watermelon? The model we are talking about is the Thomson's atomic model. Any idea why it has been given such names? Let us find out more about Thomson's atomic model or Thomson's Model of an Atom.

History

Before the discovery of subatomic particles, John Dalton came up with Dalton's atomic theory where he suggested that atoms are indivisible particles. It explained atoms cannot be broken down into further smaller particles. However, the discovery of subatomic particles disapproved the postulates proposed in Dalton Atomic Theory.

The discovery of subatomic particles led to the search how the subatomic particles are arranged in an atom. J.J. Thomson was the first and one of the many scientists who proposed models for the structure of an atom. J.J. Thomson discovered negatively charged particles by cathode ray tube experiment in the year 1897.

The particles were named electrons. J.J Thomson believed electrons to be two thousand times lighter than a proton. He assumed that an atom is composed of a cloud of negative charge in a sphere of positive charges. J.J Thomson and Rutherford first demonstrated the ionization of air in x rays.

Learn about the concept of an Atomic number here in detail.

Thomson Model of an atom

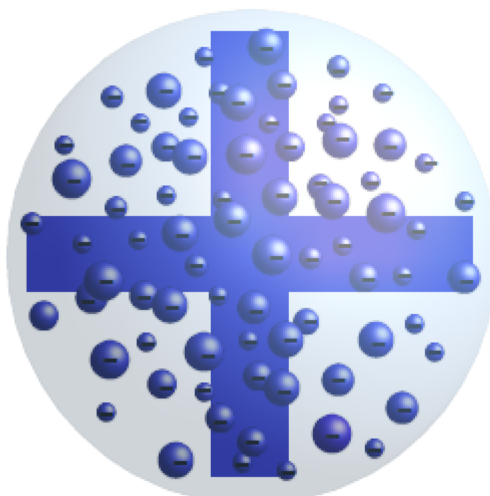
The description of Thomson's atomic model is one of the many scientific models of the atom. It was proposed by J.J Thomson in the year 1904 just after the discovery of electrons. However, at that time the atomic nucleus was yet to be discovered. So, he proposed a model on the basis of known properties available at that time. The known properties are:

- Atoms are neutrally charged

- Negatively charged particles called electrons are present in an atom. Learn about Charged particles in Matter in more detail here.

Thomson's Atomic Model- Postulates

- According to the postulates of Thomson's atomic model, an atom resembles a sphere of positive charge with electrons (negatively charged particles) present inside the sphere.
- The positive and negative charge is equal in magnitude and therefore an atom has no charge as a whole and is electrically neutral.
- Thomson's atomic model resembles a spherical plum pudding as well as a watermelon. It resembles a plum pudding because the electrons in the model look like the dry fruits embedded in a sphere of positive charge just like a spherical plum pudding. The model has also been compared to a watermelon because the red edible part of a watermelon was compared to the sphere having a positive charge and the black seeds filling the watermelon looked similar to the electrons inside the sphere.



Thomson's Atomic Model

(Source Credit: Wikipedia)

Limitations of Thomson's Atomic Model

- Thomson's atomic model failed to explain how the positive charge holds on the electrons inside the atom. It also failed to explain an atom's stability.
- The theory did not mention anything about the nucleus of an atom.
- It was unable to explain the scattering experiment of Rutherford.

Learn about the Rutherford's Atomic Model in detail here.

Conclusion

Even though Thomson's atomic model was inaccurate and had a few drawbacks, it provided the base for several other atomic structure models afterward. It is one of the foundation models that led to significant and revolutionary inventions later.

Solved Questions for You

Question: Who gave the first model of an atom?

1. Ernest Rutherford
2. J.J Thomson
3. Eugen Goldstein
4. Neils Bohr

Solution: The answer is 2 (J.J. Thomson)

Valency

We all know that the chemical formula for water is H_2O. How many times have we wondered why it is written as "H_2O" and not something else? What is the reason for this particular formula? The answer to the above question is "Valency". Let us know more about Valency and how it helps in determining a formula!

What is Valency?

Valency is the measure of the combining capacity of atoms or molecules. Therefore, it is the capacity of an atom of a single element to react and combine with particular numbers of atoms of another element.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons

- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

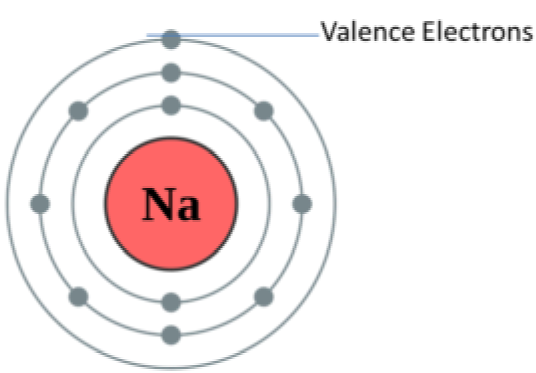
The Concept of Valency Explained

Electrons in an atom are arranged in different orbitals (shells) represented as K, L, M, N, and so on. The electrons present in the outermost shell/orbit of an atom are called valence electrons. The valence electrons take part in any chemical reaction because the outermost orbit usually contains more energy than the electrons present in other orbits.

According to the Bohr-bury scheme, the outermost orbit of an atom will have a maximum of 8 electrons. However, if the outermost orbit is completely filled then very little to no chemical activity is observed in the particular element. Their combining capacity becomes negligible or zero.

Understand the concept of Sub-atomic Particles here in detail.

This is why noble gases are least reactive because their outermost orbit is completely filled. However, the reactivity of other elements depends upon their capacity to gain noble gas configuration. It will also help to determine the valency of an atom.



Valence Electron

Achieving Complete Octet

If the outermost shell of an atom has a total of 8 electrons then the atom is said to have attained a complete octet. An atom has to gain, lose or share a particular number of electrons from its outermost orbit to obtain complete octet. Therefore, a capacity of an atom is the total number of electrons gained, lost, or shared to complete its octet arrangement in the outermost atom. This capacity of an atom will also determine the valency of an atom.

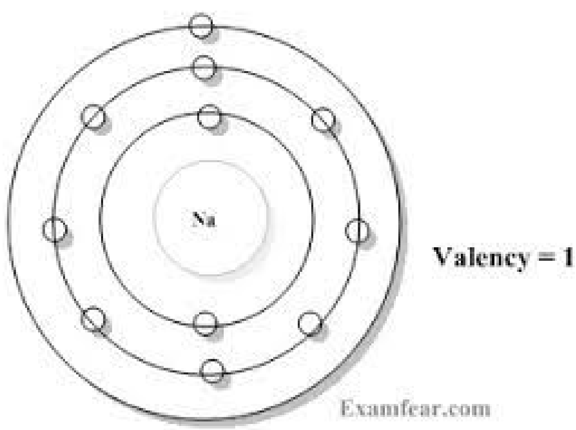
For instance, hydrogen has 1 electron in its outermost orbit so it needs to lose 1 electron to attain stability or octet. Thus, the valency of hydrogen is 1. Similarly, magnesium has 2 electrons in its outermost orbit and it needs to lose them to attain octet and obtain stability. Therefore, the valency of magnesium is 2.

Stability is also determined by the ability of atoms to gain electrons. For instance, Fluorine has 7 electrons in its outermost orbit. It is difficult to lose 7 electrons but it is easy to gain one electron. Thus, it will gain one electron to obtain octet so the valency of fluorine is 1.

Examples of Valency

Valency of Sodium

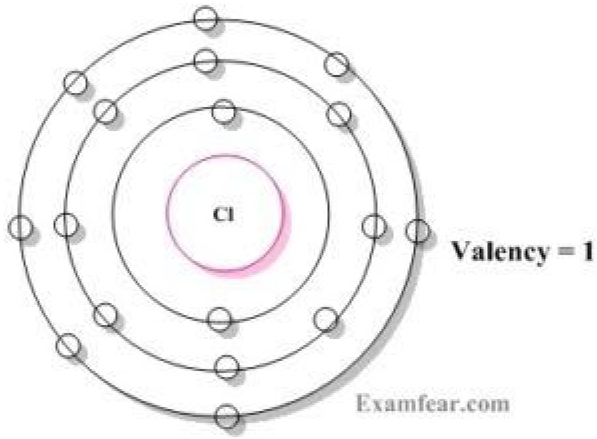
The atomic number of sodium is 11 (Z=11). The electronic configuration of sodium can be written as 2, 8, 1. 2, 8, 1 electrons are distributed in the shells K, L, M respectively. Therefore, valence electron in sodium is 1 and it needs to lose 1 electron from the outermost orbit to attain octet. Hence, the valency of sodium is 1.



(Source: examfear)

Valency of Chlorine

The atomic number of chlorine is 17 (Z=17). The electronic configuration of chlorine can be written as 2, 8, 7. 2, 8, 7 electrons are distributed in the shells K, L, M respectively. Therefore, valence electron in chlorine is 7 and it needs to gain 1 electron from the outermost orbit to attain octet. Hence, the valency of chlorine is 1.



(Source: examfear)

Examples of Valency on the basis of Chemical Formula

Ammonia (NH_3)

We know valency is the capacity of an atom to combine with a particular number of atoms of another element. In the case of ammonia, one nitrogen atom combines with 3 hydrogen atoms. The atomic number of hydrogen is 7. The electronic configuration is 2, 5. 2, 5 electrons are distributed in the orbits K, L. Therefore, a nitrogen atom needs to gain 3 electrons in its outermost orbit to complete octet.

The nitrogen atom combines with 3 hydrogen atom in the case of NH3. Therefore, its valency is 3. However, the hydrogen atoms present in ammonia combines with one nitrogen atom. Therefore, the valency of hydrogen is one. This is how chemical formulae of compounds are formed by swapping the valencies.

Uses of Valency

- It helps to determine a chemical formula.
- It helps to determine how many atoms of an element will combine with another element to form any chemical formula.

Methods of Determining Valency

The valency of the same group of the element present in the periodic table is the same. If we consider group 8 in the periodic table, all the elements of group 8 have completely filled outermost orbit and have attained octet arrangement. So, the elements of group 8 have zero valencies. The valency of any element can be determined primarily by 3 different methods:

1) The Octet Rule

If we cannot use the periodic table to determine valency then the octet rule is followed. This rule states that atoms of an element or chemicals have a tendency to obtain 8 electrons in their outermost orbit either by gaining or losing electrons in whatever form of compound it is present. An atom can have a maximum of 8 electrons in its outermost orbit. The presence of 8 electrons in the outermost shell indicates stability of an atom.

An atom tends to lose electron if it has one to four electrons in its outermost orbit. When an atom donates these free electrons it has positive valency. An atom will gain electrons if it has four to seven electrons in its outermost orbit. In such cases, it is easier to accept electron rather than donating it. Therefore, we determine the valency by subtracting the numbers of electrons from 8. All noble gases have 8 electrons in its outermost orbit except helium. Helium has 2 electrons in its outermost orbit.

Read How Electronics are distributed in Different Orbits

2) Using the Periodic Table

In this method, valency is calculated by referring to the periodic table chart. For example, all the metals, be it hydrogen, lithium, sodium and so on, present in column 1 have valency +1. Similarly, all the elements present in column 17 have valency -1 such as fluorine, chlorine, and so on. All the noble gases are arranged in column 18. These elements are inert and have valency 0.

However, there is an exception to this method of valency determination. Certain elements like copper, iron, and gold have multiple active shells. This exception is usually noticed in transitional metals from column 3 through 10. It is also observed in heavier elements from column 11 through 14, lanthanides (57-71), and actinides (89-103).

3) On the Basis of the Chemical Formulae

This method is based on the octet rule. The valencies of many transitional elements or radicals can be determined in a particular compound by observing how it chemically unites with elements of known valency. In this case, the octet rule is followed where the elements and radicals combine and try to attain eight electrons in the outermost shell in order to become stable.

For instance, consider the compound NaCl. We know that the valency of sodium (Na) is +1 and Chlorine (Cl) is -1. Both sodium and chlorine have to gain one electron and lose one electron respectively to achieve stable outermost orbit. Therefore, sodium donates an electron and chlorine accept the same electron. This is how the valency is determined. It is the classic example of ionic reaction as well.

Learn about Thomson's Atomic Model here in detail.

Difference between Valency and Oxidation Number

The Combining capacity of an atom is called valency. Thus, it is the number of valence electron an atom has to gain or lose from its outermost orbit. The oxidation number is the charge an atom can carry.

For instance, nitrogen has valency 3 but its oxidation number can range from -3 to +5. Oxidation number is an assumed charge of a particular atom in a molecule or ion. It helps to determine the capacity of an atom to gain or lose electrons within a particular species.

The Valency of First 20 Elements

Element	Symbol	Atomic Number	Valency
Hydrogen	H	1	1
Helium	He	2	0
Lithium	Li	3	1
Beryllium	Be	4	2
Boron	B	5	3
Carbon	C	6	4
Nitrogen	N	7	3
Oxygen	O	8	2
Fluorine	F	9	1
Neon	Ne	10	0
Sodium	Na	11	1

Magnesium	Mg	12	2
Aluminum	Al	13	3
Silicon	Si	14	4
Phosphorus	P	15	3
Sulphur	S	16	2
Chlorine	Cl	17	1
Argon	Ar	18	0
Potassium	K	19	1
Calcium	Ca	20	2

Learn about the concept of Isotopes and Isobars.

A Solved Question for You

Q: Determine the valencies of neon, phosphorus, sulfur.

Answer: Neon=0, Phosphorus= 3, Sulphur= 2. Explanation:

- Atomic Number of Neon=10
  Electronic configuration of Neon= 2, 8
  Therefore, Valency =0 (It is already in its octate arrangement or stable state)
- Atomic Number of Phosphorus =15
  Electronic configuration of phosphorus= 2, 8, 5
  Therefore, Valency= 8-5=3
- Atomic Number of Sulphur =16
  Electronic configuration of Sulphur = 2, 8, 6
  Therefore, Valency= 8-6=2

How are Electrons Distributed in Different Orbits (Shells)?

By now, we know that, in an atom, electrons revolve around the nucleus while protons and neutrons are inside the nucleus. The question is, how do electrons revolve? Do they move randomly? Or do they follow a specific route? How are they arranged? The answer to these questions is Electronic Configuration. Let's learn about the arrangement of electrons around the nucleus.

Distribution of Electrons in Different Orbits

Neils Bohr gave the planetary model of an atom. He was the first person to suggest the periodicity in the properties of the elements. "Bohr atomic model" forms the basis of the electronic structure of an atom. He was the person to describe the arrangement of electrons (electronic configuration) in different orbits/shells.

He proposed that electrons are distributed in circular electronic shells (orbits). These electrons revolve in the orbits around the nucleus from a fixed distance. In this topic, we will learn more about the electronic configuration of different elements.

Bohr Bury Schemes

The distribution of electrons in an atom is called as Electronic Configuration. Formula $2n^2$ helps in the determination of the maximum number of electrons present in an orbit, here n= orbit number. The formula helps in determination of arrangement of electrons and is known as "Bohr Bury Schemes."

Read more about Atomic models and Atomic numbers.

Electrons are negatively charged subatomic particles arranged like a cloud of negative charges outside the nucleus of an atom. The arrangement depends upon of their potential energies in different orbits. The different energy levels are known as 1, 2, 3, 4….. and the corresponding shells are known as K, L, M, N and so on.

For instance,

- 1st energy level- K shell/orbit
- 2nd energy level- L shell/orbit
- 3rd energy level- M shell/orbit and so on.

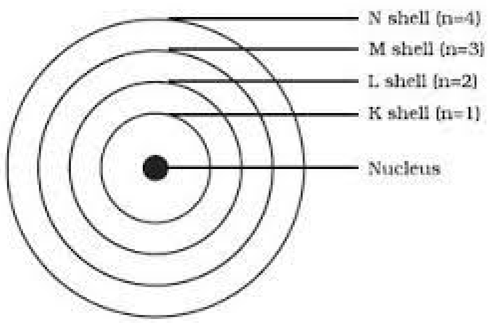


Fig. 4.3: A few energy levels in an atom

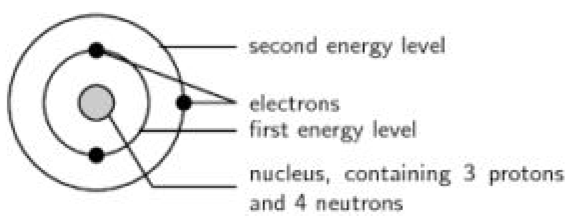
(Source: schools.aglasem)

Learn about Isobars and Isotopes here.

The Arrangement of Electrons in Different Orbits

The shells begin from the centre and gradually move outwards. So K shell will always have minimum energy. Similarly, L shell is a little away from nucleus so it will have higher energy than K shell. The outermost shell will have maximum energy. Now it is important to understand the distribution and arrangement of electrons in the atoms of any elements in the different energy levels.

An atom of any element is most stable when it has minimum energy. An atom will first fill the lowest energy level so as to attain the state of minimum energy. Gradually, the electrons will fill the higher energy levels. Therefore, electrons will first fill K shell, then L shell, M shell, N shell, and so on.



(Source: siyavula)

Electronic Configuration of Elements

According to the postulate of Neils Bohr, “electrons revolve around the centre of an atom (nucleus) in a predictable pathway named orbits”. The representation of the orbits is done by letters and numbers such as K, L, M, N, O…. and 1,2,3,4…. respectively. The arrangement and distribution of electrons in different orbits was given by Bohr and Bury.

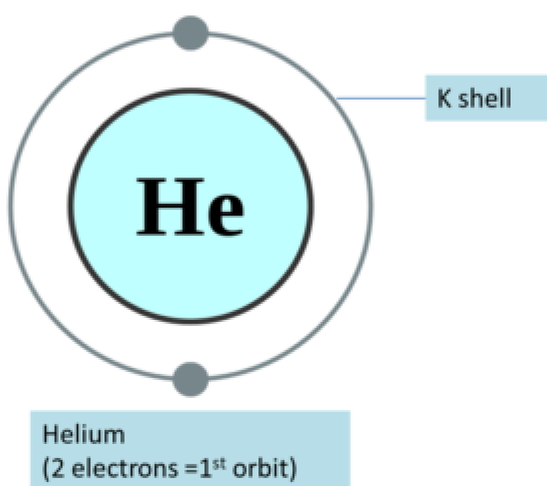
The arrangement of electrons in different shells and sub-shells is known as the electronic configuration of a particular element. The electronic configuration diagram represents an element in its ground state or stable state. There are a set of rules to remember while distribution off electrons in different orbits.

- **Rule 1**: The maximum number of electrons present in a particular shell is calculated by the formula $2n^2$, where "n" represents the shell number. For instance, K shell is the first shell and it can hold up to $2(1)^2 = 2$ electrons. Similarly, L shell is the second shell and it can hold up to $2(2)^2 = 8$ electrons. This formula helps to calculate the maximum number of electrons that an orbit can accommodate.
- **Rule 2**: The maximum capacity to hold electrons in the outermost shell is 8.
- **Rule 3**: The electrons will fill the inner shells before the outer shells. First electrons will fill the K-shell and then L shell and so on. Thus, electronic configuration of elements follows an ascending order.

Examples of the Electronic Configuration

1) Helium

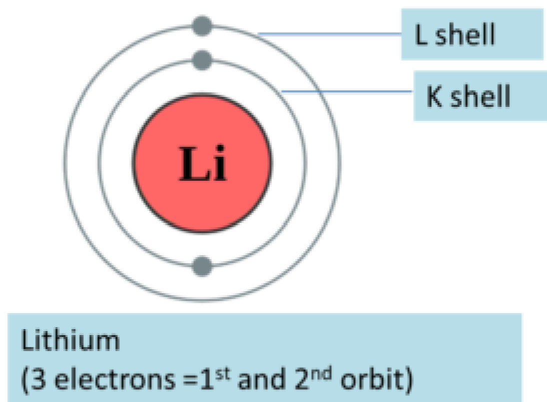
The atomic number of the element = 2. The total number of electrons present in Helium = 2. The maximum number of electrons in K shell (1st orbit) = 2. Therefore, shells needed = 1.



Electronic Configuration of Helium

2) Lithium

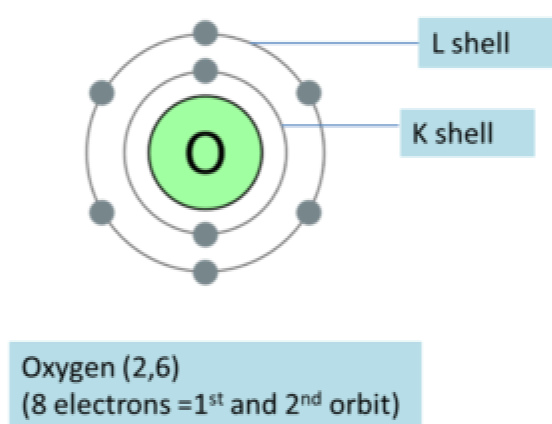
The atomic number of the element = 3. Lithium has 3 electrons. We can apply rule number 3 to fill the electrons in different in different orbits. The maximum number of electrons accommodated in K shell (1^{st} orbit) will be 2. The second orbit will accommodate rest of the electrons. Electronic configuration of Lithium= 2, 1. Therefore, the total number of shells required = 2.



Electronic Configuration of Lithium

3) Oxygen

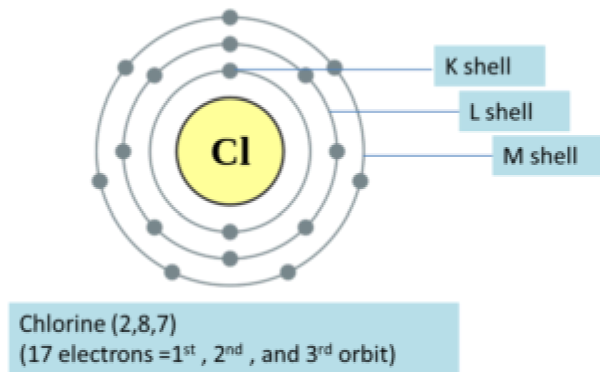
The atomic number of the element = 8. Oxygen has 8 electrons. The maximum number of electrons accommodated in the K shell (1^{st} orbit) will be 2. The second orbit will accommodate rest of the electrons left (6 electrons). Electronic configuration of Oxygen = 2, 6. Therefore, the total number of shells required = 2 (1st and 2^{nd} shell/orbit).



Electronic Configuration of Oxygen

4) Chlorine

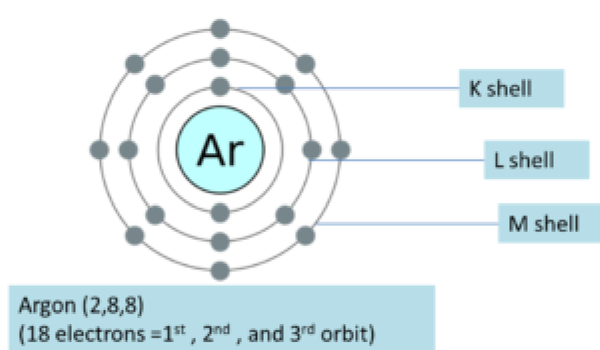
The atomic number of the element= 17. Chlorine has 17 electrons. The maximum number of electrons accommodated in the K shell (1^{st} orbit) will be 2. The second orbit will fill up to 8 electrons. Finally, The third orbit will accommodate rest of the electrons left. Electronic configuration of Chlorine = 2, 8, 7. Therefore, the total number of shells required = 3 (1^{st}, 2^{nd}, and 3^{rd} shell).



Electronic Configuration of Chlorine

5) Argon

The atomic number of the element= 18. Argon has 18 electrons. The maximum number of electrons accommodated in K shell is 2. The second orbit will fill up to 8 electrons. The third orbit can fill up to 18 electrons and it will accommodate left electrons of the element. Electronic configuration of Argon = 2, 8, 8. Therefore, the total number of shells = 3.



Electronic Configuration of Argon

Uses of the Electronic Configuration

- Electronic Configuration helps to understand the structure of periodic table with respect to each element.
- It also helps in understanding and explanation of the chemical bonds between the atoms.
- It explains the different properties and peculiar properties of certain elements. For example, electronic configuration explains the reason for the unique properties of lasers and semiconductors.

Importance of the Electronic Configuration

The electronic configuration is very important and basic part of understanding chemistry. It forms the basis of the periodic table. Additionally, the stability of any orbital will depend upon the electronic configuration of that element. It also helps us to understand the arrangement of elements in different periods and groups.

Electronic Distribution of First 18 Elements

Element	**Symbol**	**Atomic Number**	**No. of Electrons**	**Electronic configuration**				**Valency**
				Shells				
				K	**L**	**M**	**N**	
Hydrogen	H	1	1	1				1
Helium	He	2	2	2				0

Lithium	Li	3	3	2	1		1
Beryllium	Be	4	4	2	2		2
Boron	B	5	5	2	3		3
Carbon	C	6	6	2	4		4
Nitrogen	N	7	7	2	5		3
Oxygen	O	8	8	2	6		2
Fluorine	F	9	9	2	7		1
Neon	Ne	10	10	2	8		0
Sodium	Na	11	11	2	8	1	1
Magnesium	Mg	12	12	2	8	2	2
Aluminum	Al	13	13	2	8	3	3

Silicon	Si	14	14	2	8	4	4
Phosphorus	P	15	15	2	8	5	3
Sulphur	S	16	16	2	8	6	2
Chlorine	Cl	17	17	2	8	7	1
Argon	Ar	18	18	2	8	8	0

Learn about Charged Particles in Matter here.

A Solved Question for You

Q: Find the Electronic Configuration of Potassium (K).

1. 2,8,8,4
2. 2,8,8,1
3. 2, 8, 7
4. 2,8,8,3

Ans: The correct answer is option 2 (2, 8, 8, 1).

Solution: Atomic number of the potassium = 19. Potassium has 18 electrons. Applying rule number 3, the maximum number of electrons accommodated in K shell is 2. After filling the first orbit, the second orbit will consist of 8 electrons. Although it can fill up to 18 electrons the 3rd orbit will fill up to 8 electrons.

The reason behind the arrangement of electrons in such a manner is due to the presence of subshells. An atom always tends to remain in its stable state. Furthermore, it is necessary to arrange the electrons in the sub-shells in such a way that the element gains stability to attain the lowest energy level.

There are separate principles to fill the electrons in its subshells. Hence, the fourth orbit will accommodate the one electron left. Thus, the Electronic configuration of potassium = 2, 8, 8, 1

Sub-Atomic Particles

The word 'atom' comes from the Greek word 'a-tomio' which means 'uncuttable' or 'non-divisible'. Scientists believed that atoms were indivisible for the longest time. However, in the early 20th century, some scientists showed that atoms can be further divided into smaller parts such as electrons, protons, and neutrons. These are called sub-atomic particles. Want to know how these sub-atomic particles were discovered? Let's dive in!

Suggested Videos

Discovery Of Electron

Dalton's atomic theory successfully explained the following laws – conservation of mass, constant composition and multiple proportions. However, it failed to explain certain other observations like the generation of electricity on rubbing glass or ebonite with silk or fur. These observations propelled the discovery of sub-atomic particles in the 20th century. Let's learn about the discovery of the first sub-atomic particle – Electron.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Atomic Models

- [Shapes of Atomic Orbitals]
- [Energies of Orbitals]
- [Quantum Numbers]
- [Development Leading to Bohr's Model of Atom]
- [Emission and Absorption Spectra]
- [Towards Quantum Mechanical Model of Atom]

Cathode Ray Tube Experiment

Initial experiments by Michael Faraday suggested the particulate nature of electricity. His later experiments also gave an insight into the structure of the atom. Before we learn the results of these experiments, we need to remember a basic rule – 'Like charges repel each other, while unlike charges attract each other'.

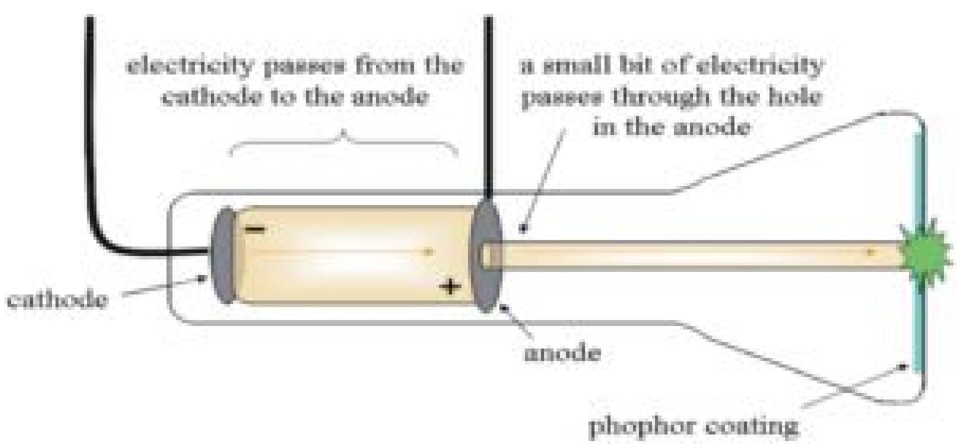
Many scientists, especially Faraday began to study electrical discharge through gases in special tubes called cathode ray discharge tubes. These tubes are made of glass and contain two thin pieces of metal called electrodes. The electrical discharge through gases is visible only at very low temperatures or very low pressures. Alteration of the pressures of different gases is possible through evacuation.

Experiment

On applying a high voltage through the electrodes, current begins to flow from the negative (cathode) to the positive (anode) electrode through a stream of particles. These are cathode rays or cathode ray particles.

Further, they made a hole in the anode and coated the tube behind with phosphorescent material like zinc sulfide. Now, when the rays pass through the anode and strike the coating, a bright spot is seen on it.

Cathode Ray Tube Experiment



[Source: Wikimedia Commons]

Results Of The Experiment

- The cathode rays start at the cathode and travel towards the anode.
- These rays are not visible as they are, but are visible with phosphorescent materials which glow when the rays hit them.
- Cathode rays travel in straight lines in the absence of electrical or magnetic fields.
- These rays behave like negatively charged particles in the presence of electrical or magnetic fields. This suggests that cathode rays consist of negatively charged, sub-atomic particles – electrons.
- The characteristics of the cathode rays or electrons are not dependent on the nature of the gas in the tube or the material of the electrodes.

These results ultimately led to the conclusion that electrons are the basic constituent of all atoms. In this manner, one of the first sub-atomic particles of an atom came to light.

Did you know that televisions also works like cathode ray tubes? The picture that you see is due to fluorescence on the television screen that is coated with fluorescent or phosphorescent materials!

Charge To Mass Ratio Of an Electron

British physicist J. J. Thomson, conducted experiments to determine the ratio of electrical charge (e) to the mass of the electron (m_e). He did this using cathode ray tubes and applying electrical and magnetic fields perpendicular to each other and also to the path of the electrons.

Thomson argued that in the presence of an electrical or magnetic field, the amount of deviation of the particles from their path depends on:

- The magnitude of the negative charge on the particle. Greater the magnitude of charge on the particle → greater the interaction with the electrical or magnetic field → greater the deflection of the particle.
- The mass of the particle. Lighter the particle → greater the deflection.
- The strength of the magnetic or electrical field. Increase in voltage across the electrodes → greater the deflection of electrons from its original path.

On application of only electric field or only magnetic field, the electrons deviate from their path and hit opposite points in the tube. By balancing the electrical and magnetic field, one can bring the electron back to the path it would follow in the absence of the two fields. Thomson determined the charge to mass ratio by calculating the amount of deflection of electrons under both fields. It is given as follows:

$$e / m_e = 1.758820 \times 10^{11} \text{ C kg}^{-1}$$

Here, m_e – the mass of electron in kg, e – the magnitude of charge on the electron in Coulomb (C). Since electrons are negatively charged, the charge on them is $-e$.

Charge On The Electron

To determine the charge on electrons, scientist R. A. Millikan conducted the oil drop experiment. Through this experiment, he found that the charge on the electron is 1.6×10^{-19} C. Today, the accepted value of electrical charge is -1.6022×10^{-19} C. The mass of the electron is obtained using this value and Thomson's value of e/m_e ratio:

$$m_e = e / (e/m_e) = (1.6022 \times 10^{-19} \text{ C}) / (1.758820 \times 10^{11} \text{ C kg}^{-1}) = 9.1094 \times 10^{-31} \text{ kg}$$

Discovery Of Protons And Neutrons

Proton

The cathode ray tube experiment also led to the discovery of positively charged particles or canal rays. Their characteristics are as follows:

- Unlike cathode rays, they depend on the nature of the gas in the cathode ray tube. These are the positively charged gaseous ions.
- The charge to mass ratio of these particles depends on the gas from which they originate.
- Some positively charged particles carry a multiple of the fundamental unit of electrical charge.
- These particles behave opposite to cathode rays in an electrical or magnetic field.

The lightest and smallest positively charged particle was obtained from hydrogen and called proton. This was characterized in 1919.

Neutron

Later, scientists also discovered an electrically neutral constituent of the atom. In 1932, the scientist Chadwick conducted an experiment where he bombarded a thin sheet of beryllium with α particles. He observed the emission of electrically neutral particles with a mass slightly greater than that of protons. As a result, these were named neutrons.

Important Properties of Sub-atomic Particles

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

Solved Examples For You

Question: Calculate the charge of one mole of electrons.

Solution: 1 mole = 6.022×10^{23} atoms, charge of one electron = 1.6022×10^{-19} C

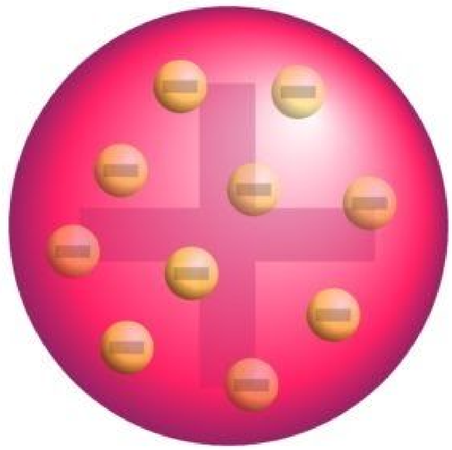
Therefore, charge of one mole of electrons = $1.6022 \times 10^{-19} \times 6.022 \times 10^{23} = 9.65 \times 10^{4}$ C.

Atomic Models

Soon after the discovery of the sub-atomic particles of an atom, scientists were eager to figure out the distribution of these particles within the atom. Several atomic models were proposed to explain the structure of the atom. However, a lot of them could not explain the stability of the atom. Let's learn about two of these atomic models that have led to our current concept of the atom.

Thomson's Atomic Model

In 1898, J. J. Thomson proposed the first of many atomic models to come. He proposed that an atom is shaped like a sphere with a radius of approximately 10^{-10}m, where the positive charge is uniformly distributed. The electrons are embedded in this sphere so as to give the most stable electrostatic arrangement.



Thomsons' atomic model [Source: Wikimedia Commons]

Doesn't the figure above remind you of a cut watermelon with seeds inside? Or, you can also think of it as a pudding with the electrons being the plum or the raisins in the pudding. Therefore, this model is also referred to as the **watermelon model**, **the plum pudding model** or **the raisin pudding model**.

An important aspect of this model is that it assumes that the mass of the atom is uniformly distributed over the atom. Thomson's atomic model was successful in explaining the overall neutrality of the atom. However, its propositions were not consistent with the results of later experiments. In 1906, J. J. Thomson was awarded the Nobel Prize in physics for his theories and experiments on electricity conduction by gases.

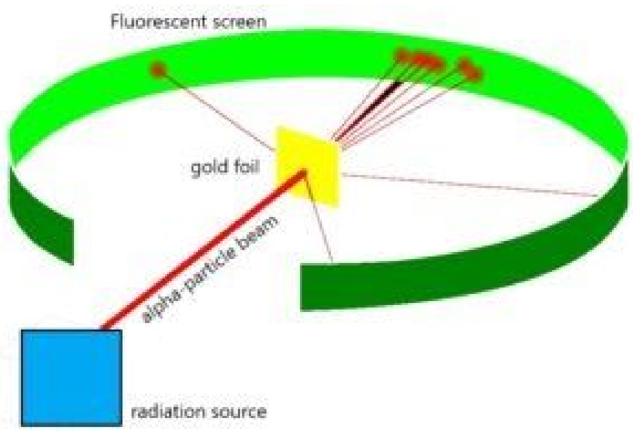
Rutherford's Atomic Model

The second of the atomic models was the contribution of Ernest Rutherford. To come up with their model, Rutherford and his students – Hans Geiger and Ernest Marsden performed an experiment where they bombarded very thin gold foil with α-particles. Let's understand this experiment.

α-Particle Scattering Experiment

Experiment

In this experiment, high energy α-particles from a radioactive source were directed at a thin foil (about 100nm thickness) of gold. A circular, fluorescent zinc sulfide screen was present around the thin gold foil. A tiny flash of light was produced at a point on the screen whenever α-particles struck it.

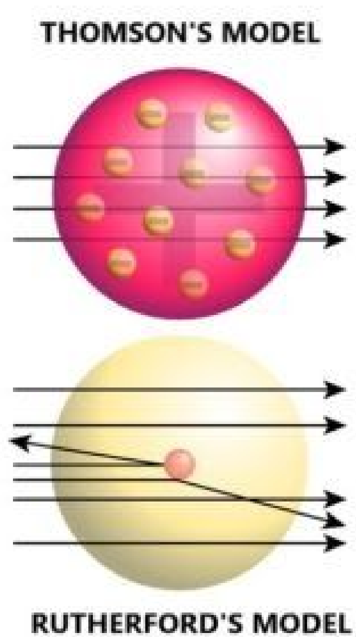


Rutherford's alpha-particle scattering experiment [Source: Wikimedia Commons]

Results

Based on Thomson's model, the mass of every atom in the gold foil should be evenly spread over the entire atom. Therefore, when α-particles hit the foil, it is expected that they would slow down and change directions only by small angles as they pass through the foil. However, the results from Rutherford's experiment were unexpected –

- Most of the α-particles passed undeflected through the foil.
- A small number of α-particles were deflected by small angles.
- Very few α-particles (about 1 in 20,000) bounced back.



Thomson's model versus Rutherford's model [Source: Wikimedia Commons]

Conclusions of the α-scattering experiment

Based on the above results, Rutherford made the following conclusions about the structure of the atom:

- Since most of the α-particles passed through the foil undeflected, most of the space in the atom is empty.
- The deflection of a few positively charged α-particles must be due to the enormous repulsive force. This suggests that the positive charge is not uniformly spread throughout the atom as Thomson had proposed. The positive charge has to be concentrated in a very small volume to deflect the positively charged α-particles.
- Rutherford's calculations show that the volume of the nucleus is very small compared to the total volume of the atom and the radius of an atom is about 10^{-10}m, while that of the nucleus is 10^{-15}m.

Nuclear Model Of The Atom

Based on his observations and conclusions, Rutherford proposed his model of the structure of the atom. According to this model –

- Most of the mass of the atom and the positive charge is densely concentrated in a very small region in the atom. Rutherford called this region the **nucleus**.
- Electrons surround the nucleus and move around it at very high speeds in circular paths called **orbits**. This arrangement also resembles the solar system, where the nucleus forms the sun and the electrons are the revolving planets. Therefore, it is also referred to as the **Planetary Model**.
- Electrostatic forces of attraction hold the nucleus and electrons together.

Drawbacks Of Rutherford's Atomic Model

- According to Rutherford's atomic model, the electrons (planets) move around the nucleus (sun) in well-defined orbits. Since a body that moves in an orbit must undergo acceleration, the electrons, in this case, must be under acceleration. According to Maxwell's electromagnetic theory, charged particles when accelerated must emit electromagnetic radiation. Therefore, an electron in an orbit will emit radiation and eventually the orbit will shrink. If this is true, then the electron will spiral into the nucleus. But this does not happen. Thus, Rutherford's model does not explain the stability of the atom.

- Contrarily, let's consider that the electrons do not move and are stationary. Then the electrostatic attraction between the electrons and the dense nucleus will pull the electrons into the nucleus to form a miniature version of Thomson's model.
- Rutherford's model also does not state anything about the distribution of the electrons around the nucleus and the energies of these electrons.

Thus, Thomson and Rutherford's atomic models revealed key aspects of the structure of the atom but failed to address some critical points. Now that we know the two atomic models, let's try to understand a few concepts.

Atomic Number And Mass Number

As we know now, a positive charge on the nucleus is due to the protons. Also, the charge on the proton is equal but opposite to that of the electron. **Atomic Number (Z)** is the number of protons present in the nucleus. For example, the number of protons in sodium is 11 whereas it is 1 in hydrogen, Therefore, the atomic numbers of sodium and hydrogen are 11 and 1, respectively.

Also, to maintain electrical neutrality, the number of electrons in an atom is equal to the number of protons (atomic number, Z). Therefore, the number of electrons in sodium and hydrogen is 11 and 1, respectively.

Atomic number = the number of protons in the nucleus of an atom

= the number of electrons in a neutral atom

The positive charge on the nucleus is due to protons, but the mass of the atom is due to protons and neutrons. They are collectively known as **nucleons**. **Mass number (A)** of the atom is the total number of nucleons.

Mass number (A) = the number of protons (Z) + the number of neutrons (n)

Therefore, the composition of an atom is represented using the element symbol (X) with the mass number (A) as super-script on the left and atomic number (Z) as sub-script on the left – ${}^{A}_{Z}X$.

Learn about Atomic number here in more detail here.

Isobars and Isotopes

Isobars are atoms with the same mass number but a different atomic number. For example, ${}^{14}_{6}C$ and ${}^{14}_{7}N$.

Learn about Isobars here in more detail here.

Isotopes, on the other hand, are atoms with the same atomic number but a different mass number. This means that the difference in the isotopes is due to the presence of a different number of neutrons in the nucleus. Let's understand this using hydrogen as an example –

- 99.985% of hydrogen atoms contain only one proton. This isotope is **protium ($^{1}_{1}H$)**.
- The isotope containing one proton and one neutron is **deuterium ($^{2}_{1}D$)**.
- The isotope with one proton and two neutrons is **tritium ($^{3}_{1}T$)**. This isotope exists in trace amounts on earth.

Other common isotopes are – carbon atoms with 6 protons and 6, 7, or 8 neutrons ($^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$) and chlorine atoms with 17 protons and 18 or 20 neutrons ($^{35}_{17}Cl$, $^{37}_{17}Cl$).

Note: Chemical properties of atoms are under the influence of the number of electrons, which are dependent on the number of protons in the nucleus. The number of neutrons has a very little effect on the chemical properties of an element. Therefore, all isotopes of an element show same chemical behaviour.

Learn about Isotopes here in more detail here.

Solved Examples For You

Question 1: Match the columns:

1. Mass number	a. Nuclear Model of Atom
2. J.J. Thomson	b. Number of protons
3. Rutherford	c. Number of nucleons
4. Atomic number	d. Plum pudding model

Solution: 1 → c, 2 → d, 3 → a, 4 → b.

Question 2: Calculate the number of protons, neutrons, and electrons in $^{56}_{26}Fe$.

Solution: In $^{56}_{26}Fe$, atomic number (Z) = 26, mass number (A) = 56.

Number of protons = number of electrons = Z = 26.

Number of neutrons = A – Z = 56 – 26 = 30.

Shapes of Atomic Orbitals

In lower classes, we were taught about the solar system model of an atom (Bohr's Atomic Model). It made us think that the electrons revolved around the nucleus in elliptical orbits. But this is not the case. In this topic, we will learn about the Shapes of Atomic Orbitals.

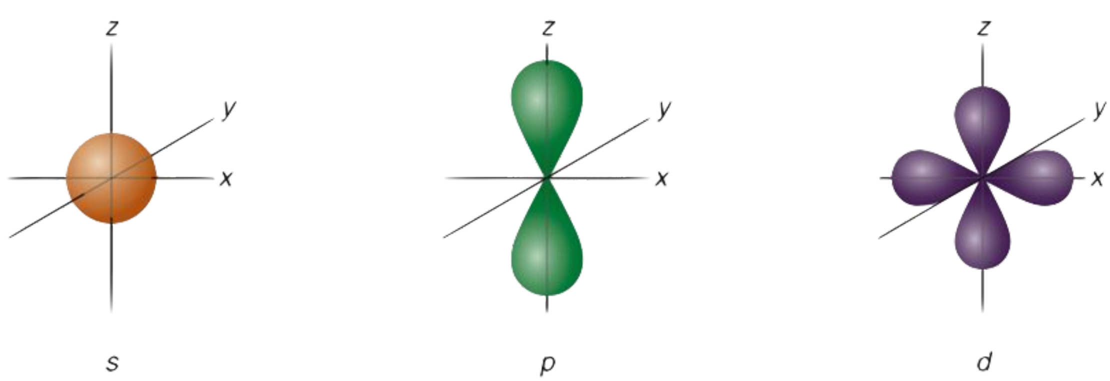
Shapes of Atomic Orbitals

In an atom, there are a large number of orbitals. There are various shapes of atomic orbitals. An orbital which is of small size states that there is more chance of finding the electron near the nucleus. The shapes of atomic orbitals and the orientation define that there is no probability of finding the electron along some certain directions than among others.

In other words, orbitals are the regions of space in which electrons are usually to be found. Each of the orbitals is denoted by a number and a letter. Herein, the number is denoted by the level of energy of the electron in the orbital.

1 refers to the energy level that is closest to the nucleus, however, 2 refers to the next level of energy further out. The symbols s, p, d, f originally comes from the words meaning sharp, principal, diffuse and fundamental respectively. This has been used to ascertain the spectral lines in the atomic spectra of different atoms.

The total permitted values of m for a given value of l to give the number of orbitals of one type within a subshell. For example – there is one orbital of s – a type of the value of l is 0, 3 p – orbitals of the value of l is one, 5 d – orbitals if the value of l is 2 and so on.



Browse more Topics Under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom

Determination of Shapes of Atomic Orbitals

S – Orbitals

For the s orbital, the boundary surface diagram looks like a sphere having the nucleus as its centre which in 2 dimensions can be seen as a circle. The s – orbitals are spherically symmetric having the probability of finding the electron at a given distance equal in all the directions. With the increase in the value of a principal quantum number, say n, the size of the s orbital will also increase.

P – Orbitals

Each p orbitals comprises of two sections referred to as lobes that lie on either side of the plane passing through the nucleus. All the three p orbitals vary in the way of the orientation of the lobes. However, they are very identical in terms of the size, shape, and energy. These lobes are placed along one of the x, y, z-axis and are given description 2px, 2py, and 2pz.

Therefore, we can say that there are about 3 p orbitals whose axes are mutually perpendicular. Just like the s – orbitals, with an increase in size and energy of p orbitals quantum number (4p > 3p > 2p), the size and energy of p orbitals also increase.

D – Orbitals

Magnetic orbital quantum number for d orbitals is given as (-2, -1, 0, 1, 2). For which we can say that there are 5 d – orbitals which are designated as d_{xy}, d_{yz}, d_{zx}, d

x

2

–

y

2

and d

z

2

. Out of these 5 d orbitals, the shapes of the first 4 d – orbitals are similar to each other which are comparatively different from the dz2 orbital whereas the energy of all 5 d – orbitals is same.

Solved Examples for You

Question: Which of the following d orbitals has a diagonal nodal plane?

A. d_{xy}
B. d_{yz}
C. d
D. *x*
E. 2
F. –
G. *y*
H. 2
I.
J. d_{zx}

Solution: Option C. d

x

2

–

y

2

Question: The electron density of $3d_{xy}$ orbital in YZ plane is:

A. 50%
B. 95%
C. 33.33%
D. 0%

Solution: Option D) 0%. YZ plane is a nodal plane for d_{xy} orbital. The electron density is only in XY plane.

Energies of Orbitals

Electrons are negatively charged and have their own energy. The energy of an electron defines which orbit it will be in. In this topic, we will learn about the Energies of Orbitals, factors affecting them and lot more.

Browse more Topics Under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

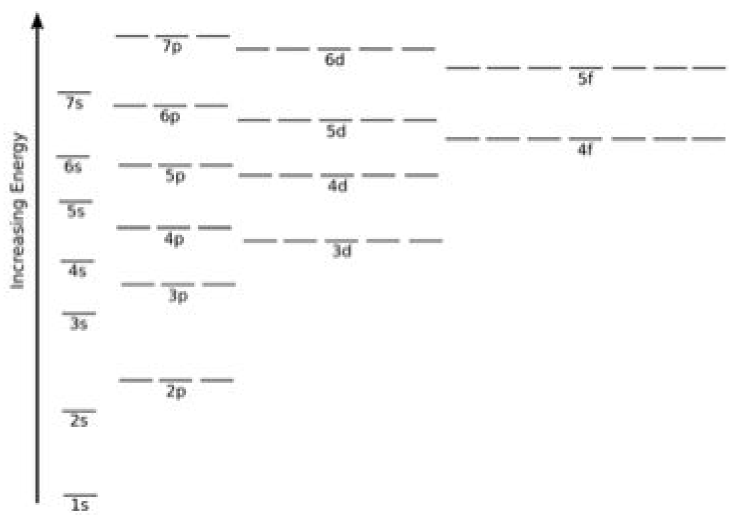
Energies of Orbitals

The energy which is essential to take an electron present in that orbital to infinity or the release of energy when an electron from an infinity it is added to that orbital, it is referred to as the energy of orbitals.

This orbital energy is dependent upon the principle of the quantum number (n) as well as the azimuthal quantum number (l) which is that it depends on the shell and subshells. For all those orbitals that belong to the same subshell, it is the same and those orbitals that are with the same energy are stated as degenerate orbitals.

The order of the increase in energy along the various orbitals is stated as –

$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f$$



Factors affecting the Orbital Energy

- The s orbital electron will be more tightly bound to the nucleus as compared to the p orbital electron, which is more tightly bound in regard to a d orbital electron for a given value of the principal quantum number.
- As compared to p orbital electrons, s orbital electrons will have more negative or lesser amount of energy. Here, the p orbital electrons will have lesser energy than that of d orbital electrons.
- As the extent of shielding from the nucleus is different for the electrons in different orbitals, it leads to the splitting of energy levels that have the same principal quantum number. Therefore, the orbital energy would depend on the values of both the principal quantum number and azimuthal quantum number, symbolized as n and l respectively. Hence, the lower value of (n + 1) for an orbital, the lower is its energy.

- With the increase in the atomic number (Zeff), the orbital energy decreases in the same subshell.

The energies of orbitals of hydrogen and hydrogen-like particles are dependent upon the value of the principal quantum (n) numbers only as well as those of multi-electron atoms that depend upon the principal quantum number (n) as well as the azimuthal quantum number (l). Energy level diagrams are the diagram that represents the orbitals arrangement in order of their increasing energies.

Some Important Observations

The important observations derived from the energy level diagrams of multi-electron atoms are

- The subshell of a particular shell does not have equal energies. For example – 2s and 2p have different energies.
- In a particular shell, the subshell that holders the lower value of l has the lower energy. In the 2nd shell, 2s (l = 0) has a lower energy than 2p (l = 1). However, in the shell 3, energy is placed in order as – 3s < 3p < 3d
- For the same value of n, the differences between the energies of s and p subshell are small whereas, between p and d subshell, it is large and so on.
- With the increase in the value of n, the subshell of the lower shell may have higher energy than that of a higher shell which means 3d has higher energy than 4s.

Solved Question for You

Question: What is the lowest value of n that allows g orbitals to exist?

Solution: For g-orbitals, l = 4.

As for any value ‘n‘ of principal quantum number, the Azimuthal quantum number (l) can have a value from zero to (n – 1).

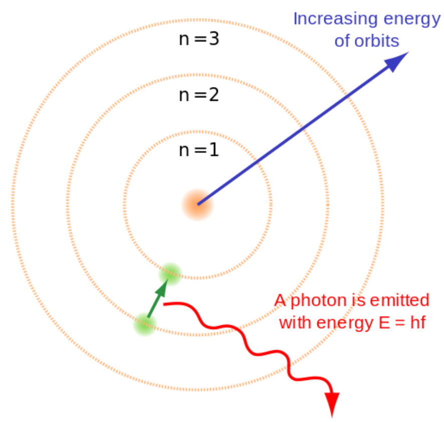
For l = 4, minimum value of n = 5.

Quantum Numbers

Quantum numbers are numbers assigned to all the electrons in an atom and they describe certain characteristics of the electron. It is very important to understand Quantum Numbers in order to understand the Structure of Atom. In this chapter, we’ll learn everything about Quantum Numbers.

Quantum Number

An [atom](#) consists of a large number of orbitals which are distinguished from each other on the basis of their [shape](#), [size](#) and orientation in space. The [orbital](#) characteristics are used to define the state of an electron completely and are expressed in terms of three numbers as stated, Principal quantum number, Azimuthal quantum number and Magnetic quantum number and Spin Quantum number.



Quantum number are those numbers that designate and distinguish various atomic orbitals and electrons present in an atom. A set of four numbers through which we can get the complete information about all the electrons in an atom, be it energy, location, space, type of orbital occupied, and even the orientation of that orbital is called Quantum Numbers.

Browse more Topics under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom
- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals

- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

Principal Quantum Number

The Principal Quantum Number represents the principal energy level or shell in which an electron revolves around the nucleus. It is denoted by the letter n and can have any integral value except the 0 i.e. n = 1, 2, 3, 4 … ….etc. The energies of the various principal shells will follow the sequence as :

$$K < L < M < N < O.....$$

$$1 < 2 < 3 < 4 < 5.......$$

Azimuthal Quantum Number

Azimuthal quantum number, also known as orbital quantum number determines the subshell to which an electron belongs. As a matter of result, the number of electronic jump increases and the number of lines at the same time.

- For a given value of n, it can have any integral value ranging from 0 to n – 1.
- For the 1st Shell, say K, n =1, you can have only one value i.e. l = 0
- For the 2nd Shell, say L, n = 2, you can have two values i.e. l = 0 and 1
- For the 3rd Shell, say M, n = 3, you can have three values i.e. l = 0, 1 and 2
- For the 4th shells, say N, n = 4, you can have 4 values i.e. l = 0, 1, 2 and 3

Magnetic Quantum Number

Magnetic Quantum Number denoted by the symbol m is what represents the orientation of atomic orbital in space. The value of the Magnetic Quantum Number, m, depends on the value of l. Magnetic Quantum Number can have a total number of (2l + 1).

Sublevel	l	m_l
s	0	0
p	1	-1, 0, +1
d	2	-2, -1, 0, +1, +2
f	3	-3, -2, -1, 0, +1, +2, +3

Spin Quantum Number

Spin Quantum Number represents the direction of the spin of the electrons. This can either be in the direction of clockwise or even anti-clockwise. Spin Quantum Number is denoted by the symbol s. It can have about only two values i.e. +1/2 or -1/2

Solved Examples for You

Question: An electron is in one of the 3d orbitals. Give the possible values of n, l and ml for this electron.

Solution: For the 3d orbital,

- Principal quantum number (n) = 3
- Azimuthal quantum number (l) = 2
- Magnetic quantum number (ml) = – 2 , – 1 , 0 , 1, 2

Question: Calculate the minimum and maximum number of electrons which have magnetic quantum number m= 1 and spin quantum number s= 1/2 in chromium

Solution: Spin can either be +1/2 or -1/2. So it is 2 electrons. If one has spin +1/2, other will have spin -1/2. So it is 2 electrons.

Question: Give the physical significance of Principal Quantum Number.

Solution: The principal quantum number (n) signifies the size of the electron cloud.

Development Leading to Bohr's Model of Atom

Through his experiments, the scientist Neils Bohr improved upon Rutherford's model of the atom. One of the concepts that played a major role in the formulation of Bohr's model is the dual nature of electromagnetic radiation. This means that radiations possess both wave-like and particle-like characteristics. Let's understand this concept in more detail.

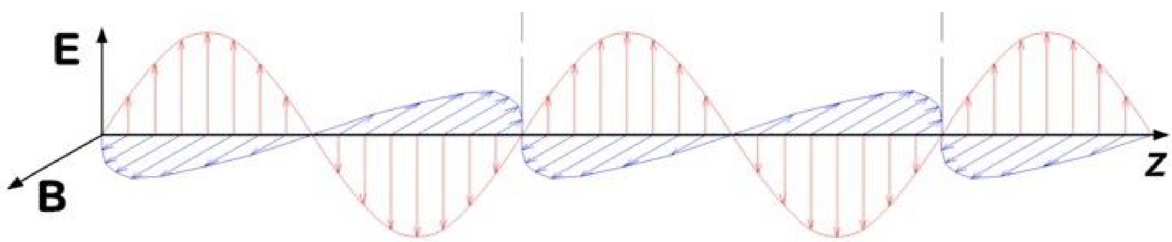
Wave Nature of Electromagnetic Radiation

In 1870, James Maxwell proposed that when electrically charged particles move under acceleration, alternating magnetic and electrical fields are generated and transmitted. These fields are transmitted as waves and are called **electromagnetic waves** or **electromagnetic radiation**.

For years scientists have speculated about the nature of light as a form of radiation. In the early years, scientists believed that light comprised of particles or corpuscles. The wave nature of light was established only in the early 19th century. Through the concept of electromagnetic radiation, Maxwell was the first to show that electricity, magnetism, and light are different manifestations of the same phenomenon. Let's understand a few simple properties of electromagnetic wave motion.

Properties of Electromagnetic Wave Motion

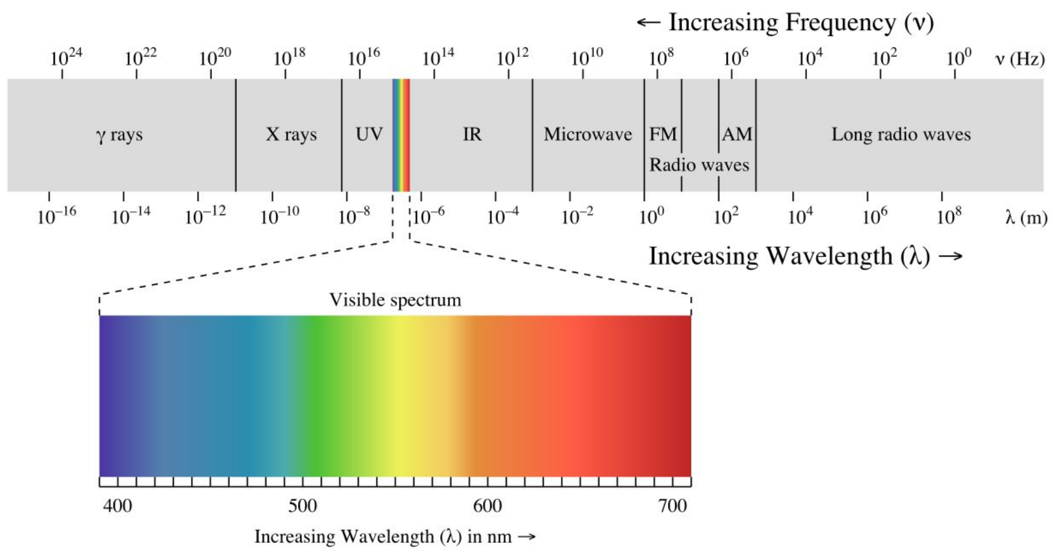
- Oscillating charged particles produce oscillating electric and magnetic fields that are perpendicular to each other. These fields are also perpendicular to the direction of propagation of the wave.



The electric and magnetic field of the electromagnetic wave. The electric field in red (E axis), the magnetic field in blue (B axis), the wave propagates along the Z-axis. [Source: Wikimedia Commons]

- Electromagnetic waves don't need a medium for propagation like sound waves or water waves. They can travel through a vacuum.
- Today, there are many different types of electromagnetic radiations that differ from each other in wavelength or frequency. They all constitute an **electromagnetic spectrum**. Different regions of this spectrum have different names and use. For example, radio frequency region around 10^6 Hz is used for broadcasting, microwave region around 10^{10} Hz is used for radar, the infrared region around 10^{13} Hz is used for heating and 10^{16} Hz is the UV component of the sun's rays. **Visible light** is the small portion around 10^{15} Hz and it is only this part that our eyes can see. Special instruments are needed to detect non-visible light.
- Electromagnetic radiation has different properties. Let's learn a few of them.
    - **Frequency (ν)** – It is the number of waves that pass a given point in one second. The SI unit is Hertz (Hz, s^{-1}), named after Heinrich Hertz.

- **Wavelength (λ)** – Wavelength has the same units as the length which is the meter (m). But since many waves of small wavelength make up the electromagnetic radiation, we use smaller units.



Source: Wikimedia

- **Wavenumber** – It is the number of wavelengths per unit length. Its units are the reverse of wavelength – m^{-1} or cm^{-1}.
- **The speed of light (c)** – This is the speed at which all types of electromagnetic radiations, regardless of wavelength travel in a vacuum (3.0×10^8 ms^{-1}). The wavelength, frequency, and speed of light are related by the equation:

$$c = \nu \lambda$$

Particle Nature of Electromagnetic Radiation

Although the wave nature of electromagnetic radiation explains phenomena like ‘diffraction‘ and ‘interference’, some other important features remain unexplained. The unexplained observations are:

- Black-body radiation i.e. the nature of emission of radiation from hot bodies.
- Photoelectric effect i.e. the ejection of electrons from a metal surface when radiation strikes it.
- Variation of the heat capacity of solids.
- Line spectra of atoms with reference to hydrogen.

Before we go any further, let’s understand the phenomena of black-body radiation and the photoelectric effect.

Suggested Videos on Bohr Atomic Model

Black-body Radiation

In this phenomenon, solids when heated emit radiations over a wide range of wavelengths. The best example of this is the heating of an iron rod in a furnace or over a flame. Have you ever observed the different colours as the iron rod becomes hotter and hotter? It starts off as a dull red colour which then becomes redder as the temperature increases.

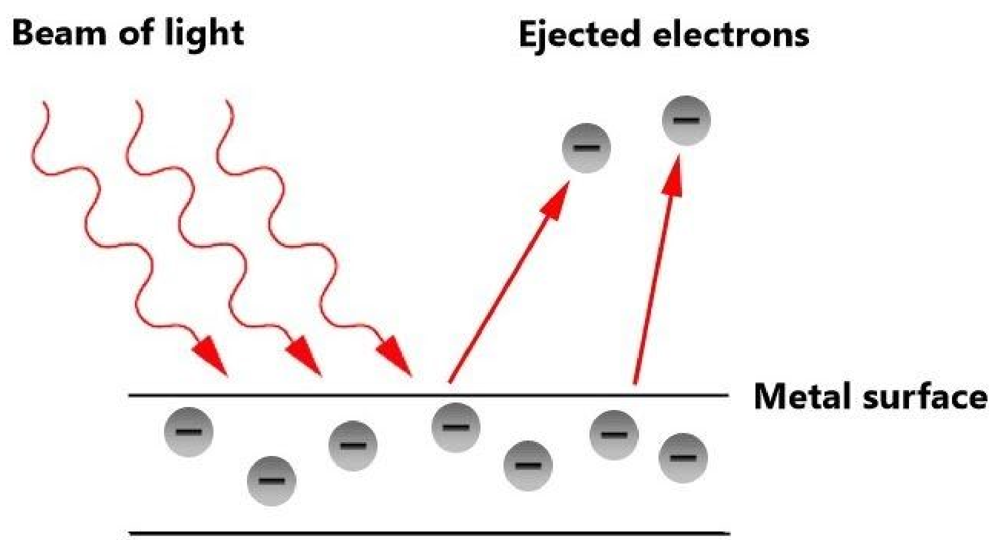
As the temperature rises further, it turns white and then blue. This simply means that the frequency of the emitted radiation goes from a lower frequency to a higher frequency as the temperature increases. The red colour lies in the lower frequency region while blue colour lies in the higher frequency region of the spectrum.

A **black-body** is an ideal body that emits and absorbs radiations of all frequencies. The radiation emitted by such a body is **black-body radiation**. The frequency distribution of the emitted radiation from a black body depends only on its temperature. The radiation intensity at a given temperature increases with the decrease of wavelength, it reaches a maximum and then starts decreasing with a further decrease in wavelength.

Photoelectric Effect

H. Hertz performed a very interesting experiment in 1887. Electrons were ejected when he exposed certain metals to a beam of light. We call this phenomenon as the **Photoelectric effect.** His observations were as follows:

- There is no time lag between the striking of the beam of light on the metal and the ejection of electrons from its surface.
- The number of electrons ejected is proportional to the brightness or intensity of light.
- There is a characteristic minimum frequency or **threshold frequency** for each metal, below which photoelectric effect is not observed. Above this threshold frequency, the electrons are ejected with a certain kinetic energy which increases with an increase in the frequency of the light used.



Source: Wikimedia

Planck’s Quantum Theory

Classical physics or the wave theory of light do not satisfactorily explain the phenomena of black-body radiation or photoelectric effect.

Explanation for Black-Body Radiation

In 1900, Max Planck gave the first concrete explanation for the phenomenon of black-body radiation. He suggested that atoms or molecules emit or absorb energy only in discrete amounts called **quantum** and not in a continuous manner. Quantum is the smallest amount of energy that is emitted or absorbed in the form of electromagnetic radiation. The energy of the quantum is proportional to its frequency. It is as follows –

$$E = hv$$

where ‘E’ is the energy of the quantum, ‘v’ is the frequency and ‘h’ is the proportionality constant or the Planck’s constant and has a value of 6.626×10^{-34} Js. Using this theory, Planck was able to explain that the intensity distribution of the radiation from a black-body is a function of frequency or wavelength at different temperatures.

Explanation for Photoelectric Effect

In 1905, Einstein used Planck’s quantum theory to explain the photoelectric effect. According to Planck’s quantum theory, shining a beam of light on a metal surface can be viewed as shooting the metal with a beam of particles or photons.

In this case, when a photon of sufficient energy strikes an electron in the metal, it transfers its energy to the electron immediately and the electron gets ejected without any time lag. A more intense beam of light has a larger number of photons and therefore, ejects a larger number of electrons.

Finally, greater the energy carried by a photon, greater is the kinetic energy of the ejected electron. This means that the kinetic energy of the ejected electron is proportional to the frequency of the electromagnetic radiation. The following equation gives the kinetic energy of the ejected electron –

$$hv = hv_0 + 1/2\ m_e v^2$$

where, m_e – the mass of the electron, v – velocity associated with the ejected electron, hv – energy of the striking photon, hv_0 – the minimum energy required to eject an electron or work function (W_o).

Dual Behavior Of Electromagnetic Radiation

The particle nature of light explains the phenomena of black-body radiation and the photoelectric effect. The wave nature of light, on the other hand, explains interference and diffraction. This contrast posed a dilemma for scientists.

Finally, they accepted the idea that light possesses both wave-like and particle-like properties i.e. light has dual behaviour. Light has wave-like properties when it propagates whereas, on interaction with matter, it shows particle-like properties.

Solved Examples for You

Question: A sodium lamp emits yellow light of wavelength (λ) 580nm. What are the frequency (v) and wavenumber of this light?

Solution: We know that c = v λ and that regardless of wavelength, all electromagnetic radiations travel at a speed (c) of $3.0 \times 10^8\ ms^{-1}$. Therefore, v = c/λ

$= 3.0 \times 10^8\ ms^{-1} / 580 \times 10^{-9}m$

$= 5.172 \times 10^{14}$ per second

Wave number = 1/ wavelength = 1/λ = $1/\ 580 \times 10^{-9}\ m$

$= 1.724 \times 10^6\ m^{-1}$

Question: If a photon of wavelength 4×10^{-7} m strikes a metal surface and the work function (hv_0) of the metal is 2.13eV, then what is the kinetic energy of the emission?

Solution: We know that $E = hv = hc/\lambda$

where h = Planck's constant = 6.626×10^{-34}Js, c =speed of light = 3.0×10^8 m/s.

Therefore, $E = hc/\lambda = (6.626 \times 10^{-34} \times 3.0 \times 10^8)/ 4 \times 10^{-7}$

$= 4.97 \times 10^{-19}$J = 3.102 eV (Since, 1J = 6.24×10^{18} eV)

Now, kinetic energy = $hv - hv_0 = E - hv_0$ = 3.102 – 2.13 eV = 0.972 eV.

Emission and Absorption Spectra

Aren't you fascinated when you see a rainbow? The colours we see in it never fail to captivate us! Did you know that even though we identify the distinct colours of a rainbow, it is actually a continuous range of colours? A similar range of colours appears when white light passes through a prism. This range of colours is a spectrum. Let's learn about the different forms of spectra in more detail.

Spectrum

Ordinary white light consists of waves of all wavelengths in the visible range. This is why, when white light passes through a prism, a series of coloured bands are seen called **spectrum**. This spectrum of white light ranges from violet at 7.5×10^{14} Hz to red at 4×10^{14} Hz.

Since the colours merge into each other i.e. violet merges into blue, blue into green and so on, we call it a **continuous spectrum**. When this light passes through an object or medium, the wave with the shortest wavelength (violet) deviates the most than the one with the longest wavelength (red).

The interaction of electromagnetic radiation with matter causes the atoms and molecules to absorb energy and go to a higher energy state. Since this state is unstable, they need to emit radiations to return to their normal states. This gives rise to emission and absorption spectra. Let's learn about them in detail.

Browse more Topics Under Structure Of Atom

- Introduction: Structure of Atom
- Atomic Number
- Bohr's Model of Atom

- Charged Particles in Matter
- Isobars
- Isotopes
- Mass Number
- Neutrons
- Rutherford's Model of an Atom
- Thomson's Model of an Atom
- Valency
- How are Electrons Distributed in Different Orbits (Shells)?
- Sub-Atomic Particles
- Atomic Models
- Shapes of Atomic Orbitals
- Energies of Orbitals
- Quantum Numbers
- Development Leading to Bohr's Model of Atom
- Emission and Absorption Spectra
- Towards Quantum Mechanical Model of Atom

Emission and Absorption Spectra

The **emission spectrum** is the spectrum of radiation emitted by a substance that has absorbed energy. Atoms, molecules, and ions that have absorbed radiation are called '**excited**'. The **absorption spectrum** is the opposite of the emission spectrum.

It is the spectrum formed by electromagnetic radiation that has passed through a medium, in which radiation of some frequencies is absorbed. **Spectroscopy** is the study of emission and absorption spectra.

Line Spectrum

Unlike visible light which shows a continuous spectrum of all wavelengths, the emission spectra of atoms in the gas phase emit light only at specific wavelengths with dark spaces between them. This is called **line spectra** or **atomic spectra** since the emitted radiation is identified by bright lines in the spectra.

Learn about Atomic Spectra here in detail.

Each element has its own unique **line emission spectrum**. Did you know that just the way fingerprints are used to identify people, the characteristic lines in an atomic spectrum are used to identify unknown atoms!

Robert Bunsen, a German chemist was the first to identify elements using line spectra. Using spectroscopic methods, we discovered elements such as helium, rubidium, scandium, gallium, thallium, etc.

Learn more about Group 18 Elements here in detail.

Line Spectrum of Hydrogen

Hydrogen molecules dissociate when we pass electric discharge through gaseous hydrogen. Subsequently, the energetically excited H_2 atoms emit electromagnetic radiation of discrete frequencies giving rise to a spectrum. The hydrogen spectrum has many series of lines. These lines are named after their discoverers.

In 1885, the scientist Balmer showed that if spectral lines are expressed as wavenumber, then the visible lines of the hydrogen spectrum obey the following formula –

$$\text{Wavenumber} = 109{,}677\ (1/2^2 - 1/n^2)\ \text{cm}^{-1}$$

where $n \geq 3$.

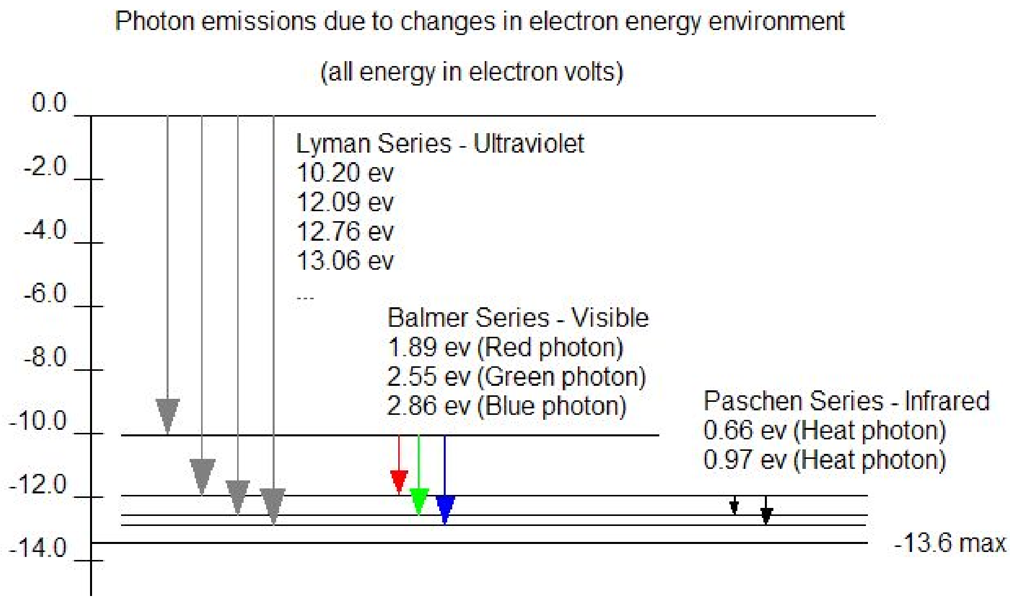
We call this series of lines, **Balmer series**. These lines are the only lines in the hydrogen spectrum that appear in the visible region of electromagnetic radiation. Johannes Rydberg, a Swedish spectroscopist, showed that all series of lines in the hydrogen spectrum can be described by the formula –

$$\text{Wavenumber} = 109{,}677\ (1/n_1^2 - 1/n_2^2)\ \text{cm}^{-1}$$

where $n_1 = 1, 2....$ and $n_2 = n_1 + 1, n_1 + 2......$

The value 109,677 is the **Rydberg constant** for hydrogen. The lines that correspond to $n_1 = 1, 2, 3, 4, 5$ are called Lyman, Balmer, Paschen, Brackett and Pfund series, respectively.

Series	n_1	n_2	**Spectral Region**
Lyman	1	2,3…	Ultraviolet
Balmer	2	3,4…	Visible
Paschen	3	4,5…	Infrared
Brackett	4	5,6…	Infrared
Pfund	5	6,7…	Infrared



Source: Wikimedia

The hydrogen atom has the simplest line spectrum of all elements. For heavier atoms, the line spectrum becomes more and more complex. However, there are certain features that are common to all line spectra –

- Line spectrum of every element is unique.
- There is regularity in the line spectrum of each element.

Now, that we understand the line spectrum of hydrogen, let's understand the features of the hydrogen atom, its structure, and its spectrum.

Bohr's Model For Hydrogen Atom

In 1913, Neils Bohr was the first to explain the features of the hydrogen atom and its spectrum. His model is based on the following postulates –

- Electrons in a hydrogen atom can move around the nucleus in a circular path that has a fixed radius and energy. These paths, arranged concentrically around the nucleus, are **orbits**.
- The energy of an electron in orbit does not change with time. However, when the electron absorbs a required amount of energy, it moves from a lower stationary state to a higher stationary state. On the other hand, when the electron emits energy, it moves from a higher stationary state to a lower stationary state. This energy change is not continuous.

- The following equation gives the frequency of radiation absorbed or emitted when an electron transitions between two stationary states that differ in energy by ΔE –

$$v = \Delta E/h = (E_2 - E_1)/h$$

where, E_1 and E_2 are the energies of the lower and higher energy states, respectively. This equation is commonly called **Bohr's frequency rule.**

- In a given stationary state, the angular momentum of an electron is given as –

$$m_e\, v\, r = n.(h/2\pi) \quad \text{where } n = 1,2,3\ldots\ldots$$

Thus, an electron can only move in those orbits whose angular momentum is an integral multiple of $h/2\pi$. This is why there are only certain orbits.

Hydrogen Atom described by Bohr's Model

For the hydrogen atom, Bohr's theory describes that –

- The stationary states for the electron are numbered n = 1,2,3…These are **Principal Quantum numbers**.
- The radii of the stationary states are given as $r_n = n^2a_0$, where a_0 = 52,9pm. Therefore, the radius of the first stationary state i.e. the **Bohr orbit** is 52.9pm. The electron in the hydrogen atom is in this orbit i.e. n=1. As *n* increases, *r* will increase and the electron will be found away from the nucleus.
- The energy of the stationary state is given by the equation –

$$E_n = -R_H\,(1/n^2) \text{ where } n=1,2,3\ldots\ldots$$

R_H is the **Rydberg constant** and the value is 2.18×10^{-18} J. Using this equation, you can calculate the energy of any state. For example, the energy of the stationary state for n=2 will be:

$$E_2 = -2.18\times10^{-18}\text{J}\ (1/2^2) = -0.545\times10^{-18}\ \text{J}$$

The energy of an electron is taken as zero when it is not under the influence of the nucleus. In this situation, n=∞ and the atom are called an ionized hydrogen atom.

- Bohr's theory is also applicable to ions with only one electron, for example, the hydrogen atom. For these kinds of ions, the energies of stationary states are –

$$E_n = -2.18 \times 10^{-18}\,(Z^2/n^2)\text{J}$$

The radii are given by $r_n = 52.9(n^2)/Z$ pm, where Z is the atomic mass number.

- With Bohr's theory, it is also possible to calculate the velocities of electrons moving in orbits. The magnitude of the velocity of the electron increases with an increase in positive charge on the nucleus and decreases with an increase in principal quantum number.

Learn more about Bohr's Model of Hydrogen Atom here in detail.

Explanation of Line Spectrum of Hydrogen

Bohr's model can explain the line spectrum of the hydrogen atom. According to assumption 2, radiation is absorbed when an electron goes from orbit of lower energy to higher energy; whereas radiation is emitted when it moves from higher to lower orbit. The energy gap between the two orbits is –

$$\Delta E = E_f - E_i \text{ where f is the final orbit, i is the initial orbit}$$

Since, $E_n = -R_H (1/n^2)$, we can now say $\Delta E = (-R_H /n^2_f) - (-R_H /n^2_i) = R_H (1/n^2_i - 1/n^2_f)$

$= 2.18 \times 10^{-18}$ J $(1/n^2_i - 1/n^2_f)$

The frequency and wavenumber associated with the absorption and emission of the photon can also be calculated –

$$v = \Delta E/h = R_H(1/n^2_i - 1/n^2_f) = 2.18 \times 10^{-18} \text{ J}/6.626 \times 10^{-34} \text{ J s } (1/n^2_i - 1/n^2_f)$$

$$= 3.29 \times 10^{15} (1/n^2_i - 1/n^2_f) \text{ Hz}$$

$$\text{Wavenumber} = v/c = R_H/hc\ (1/n^2_i - 1/n^2_f) = 3.29 \times 10^{15} \text{ s}^{-1}/ 3 \times 10^8 \text{ ms}^{-s} (1/n^2_i - 1/n^2_f)$$

$$= 1.09677 \times 10^7 (1/n^2_i - 1/n^2_f) \text{ m}^{-1}$$

- When $n_f > n_i$, the term in brackets is positive and energy is absorbed.
- When $n_i > n_f$, ΔE is negative and energy is released.

Limitations of Bohr's Model

Bohr's model, although an improvement over Rutherford's model, failed to account for the following points –

- It fails to explain the finer details of the hydrogen atom spectrum observed by spectroscopic techniques. Bohr's model is unable to explain the spectrum of atoms other than hydrogen. It also fails to explain the concepts of Zeeman and the Stark effect.
- It does not explain the ability of atoms to form molecules with chemical bonds.

Solved Example For You

Question: An electron in a hydrogen atom transitions from energy level with n=4 to n=2. What is the wavelength of the emitted light?

Solution: Wavenumber = R $(1/n_1^2 - 1/n_2^2)$ cm^{-1}

Here, R = 109,677, $n_1 = 4$, $n_2 = 2$

Therefore, wavenumber = 109,677 $(1/4^2 – 1/2^2)$ cm^{-1} = 20564.4 cm^{-1}

Now, λ = 1/wavenumber = 1/20564.4 = 486 x 10^{-7}cm = 486 nm.

Towards Quantum Mechanical Model of Atom

Although Bohr's model of the atom is the most commonly used model, scientists continued to develop new and improved models for the atom. Two concepts that contributed significantly to the development of improved models were – the Dual behaviour of matter and the Heisenberg uncertainty principle. Let's learn about these concepts in more detail.

Dual Behavior of Matter

In 1924, de Broglie, the French physicist proposed that matter, like radiation, should show dual behaviour i.e. both wave-like and particle-like properties. This means that just like photons, electrons should also have a wavelength as well as momentum. From this, de Broglie gave an equation that relates wavelength and momentum of a material particle –

$$\lambda = h/mv = h/p$$

where h is the Planck's constant (6.626 x 10^{-34}), m is the mass of the particle, v is the velocity and p is the momentum of the particle.

This prediction of de Broglie's was experimentally proven when it was found that an electron undergoes diffraction, a property characteristic of waves. This knowledge has been used to construct an electron microscope. Just as an ordinary microscope uses the wave nature of light, an electron microscope uses the wave-like behaviour of electrons.

An electron microscope achieves a magnification of about 15 million times and therefore is a powerful tool in scientific research. According to de Broglie, every object in motion has a wave-like character. The wave properties of ordinary objects are hard to detect because their wavelengths are so short.

However, experimental detection of the wavelengths associated with electrons and other sub-atomic particles is possible. For example, to find out the wavelength of a ball of mass 0.1 kg moving with a velocity of $10ms^{-1}$, we will use de Broglie's equation –

$$\lambda = h/mv = (6.626 \times 10^{-34})/(0.1 \times 10) = 6.626 \times 10^{-34} \text{ m}$$

Heisenberg's Uncertainty Principle

In 1927, the German physicist, Werner Heisenberg came up with the uncertainty principle as a result of concepts like the dual behaviour of matter and radiation.

Source: Wikimedia

His principle states that – **It is impossible to determine, simultaneously, the exact position and exact momentum (or velocity) of an electron.** He gave the following equation –

$$\Delta x \times \Delta p_x \geq h/4\pi \quad \text{OR}$$

$$\Delta x \times \Delta(mv_x) \geq h/4\pi \quad \text{OR}$$

$$\Delta x \times \Delta v_x \geq h/4\pi m$$

where Δx is the uncertainty in position and Δp_x or Δv_x is the uncertainty in momentum (or velocity) of the particle. If Δx is small, then Δv_x is large, while if Δv_x is small then Δx will be large. Thus, is we physically measure the electron's position or velocity, the outcome is always fuzzy or a blur.

We can best explain the uncertainty principle with an example. Suppose you are asked to measure the thickness of a sheet of paper with an unmarked ruler. Obviously, the results will be inaccurate and meaningless. To do this correctly, you will have to use a graduated or marked ruler with units smaller than the thickness of the sheet of paper.

Similarly, in order to measure the position of an electron, we need to measure with a meter stick with units smaller than the dimensions of an electron. We can observe an electron by illuminating it with “light” or electromagnetic radiation. This “light” should have a wavelength smaller than the dimensions of an electron.

The collision of the high momentum photons of light ($p=h/\lambda$) with the electrons, change the energy of the electrons. This way, we will be able to calculate the position of the electron but we will still know very little about the velocity of the electron after a collision.

The Significance of Uncertainty Principle

- It rules out the existence of definite paths or trajectories of electrons and other similar particles. The trajectory of an object involves the location and velocity of the object. If we know the location of a body at a particular instant and we know its velocity and the forces acting on it, we can determine where the body will be sometime later.
 Since it is not possible to determine the position and velocity of an electron at any given instant to an arbitrary degree of precision, it is not possible to know the trajectory of an electron.
- The effect of Heisenberg Uncertainty Principle is significant only for the motion of microscopic objects and is negligible for that of macroscopic objects. For example, if we apply the uncertainty principle to an object of mass 1 milligram (10^{-6} kg), then

$$\Delta v.\Delta x = h/4\pi m = 6.626 \times 10^{-34}\ \mathrm{Js}/(4 \times 3.1416 \times 10^{-6}\mathrm{kg}) \approx 10^{-28}\ \mathrm{m^2s^{-1}}$$

- For milligram-sized or heavier objects, the associated uncertainties are of hardly any real consequence. The reason for this is that the value of $\Delta v \Delta x$ obtained is small and insignificant.
- The precise statements of the position and momentum of electrons have to be replaced by the statements of probability, that the electron has at a given position and momentum.

For a microscopic object like an electron, $\Delta v.\Delta x$ is much larger and the uncertainties are of real consequence. For example, for an electron with a mass of 9.11×10^{-31} kg, according to the Heisenberg uncertainty principle –

$$\Delta v.\Delta x = h/4\pi m = 6.626 \times 10^{-34}\ \mathrm{Js}/(4 \times 3.1416 \times 9.11 \times 10^{-31}\mathrm{kg}) = 10^{-4}\ \mathrm{m^2s^{-1}}$$

This means that if you try to find the location of an electron to an uncertainty of about 10^{-8}m, then the uncertainty Δv in velocity would be –

$$10^{-4} m^2 s^{-1} / 10^{-8} m \approx 10^4\ ms^{-1}$$

This value is so large that the idea of electrons moving in Bohr's orbits does not hold good. This is why the precise statements of the position and momentum of electrons have to be replaced by the statements of probability, that the electron has at a given position and momentum. This is what happens in the quantum mechanical model of the atom.

Reasons for Failure of Bohr Model

Let's understand the reasons for the failure of Bohr's model –

- Bohr's model describes an electron as a charged particle that moves around the nucleus in well-defined circular orbits. The model fails to consider the wave character of the electron.
- An orbit is a clearly defined path which is defined completely only if the position and velocity of the electron are known exactly at the same time. According to Heisenberg's uncertainty principle, this is not possible. Therefore, Bohr's model of the hydrogen atom not only ignores the dual behaviour of matter but also contradicts Heisenberg's uncertainty principle.

Thus, we did not extend Bohr's model to other atoms because of these limitations. New models were needed that could account for the wave-particle duality of matter and be consistent with Heisenberg's principle. This led to the advent of quantum mechanics.

Solved Example For You

Question: An electron is moving with a velocity of $2.05 \times 10^7\ ms^{-1}$. What is its wavelength?

Solution: According to de Broglie's equation, we know that –

$$\lambda = h/mv$$

where λ is the wavelength, h is the Planck's constant, m is the mass of the particle and, v is the velocity.

We also know that $v = 2.05 \times 10^7\ ms^{-1}$, $h = 6.626 \times 10^{-34}$ Js, $m = 9.10939 \times 10^{-31}$ kg

So substituting the appropriate values, we get –

$$\lambda = h/mv = 6.626 \times 10^{-34}\ Js/(9.10939 \times 10^{-31}\ kg)\ (2.05 \times 10^7\ ms^{-1})$$

$$\lambda = 3.548 \times 10^{-11}\ m$$