# PKP - *CHEM*- 11 <Neet+Jee>

# { Structure *Of Atoms* }

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*** What is the structure of atoms ?**

_ Atomic structure refers to the structure of an atom comprising a **nucleus** (centre) in which the **protons** (positively charged) and **neutrons** (neutral) are present. The negatively charged particles called **electrons** revolve around the **centre of the nucleus**

*** Who discovered atom ?**

The first scientific theory of atomic structure was proposed by John Dalton in the 1800s.

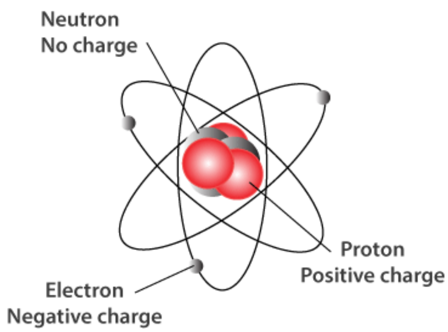
The history of atomic structure and quantum mechanics dates back to the times of Democritus, the man who first proposed that matter is composed of atoms. The study about the structure of an atom gives a great insight into the entire class of chemical reactions, bonds and their physical properties

*** Discovery of SUB – ATOMIC PARTICAL**

**SIZE-** A Subatomic particle is nothing but a particle which is smaller than an atom in size. Typically, an atom can be broken down into three subatomic particles, namely: protons, electrons, and neutrons.

**PARTICAL -** For a long time, it was believed that atoms are the ultimate particles that matter is made up of and that these atoms cannot be divided further. The experiments conducted during the latter half of the nineteenth century and early years of the twentieth century revealed that the atom is not the ultimate particle. The continued efforts of the scientists led to the discovery of subatomic particles.

The three primary subatomic particles that constitute an atom are illustrated below



**Limitation :** The limitations of Dalton's atomic theory to explain certain observations formed the basis for the discovery of electrons and protons. Further investigations revealed the existence of neutrons. The components of atoms are called **subatomic particles** and generally include the proton, electron, and the neutron.

## Discovery and Features of Subatomic Particles

The discovery of the three basic subatomic particles and some of their important features are discussed in this subsection.

### Protons

Protons and Neutrons together make up the nucleus of an atom and are hence called nucleons. Some important points regarding the discovery and properties of protons are listed below.

- Protons are positively charged subatomic particles.
- The number of protons in an atom is equal to the number of electrons in it.
- The discovery of protons is credited to Ernest Rutherford.
- Protons can be produced via the removal of an electron from a hydrogen atom.
- The mass of a proton is 1.676 * 10-24 grams.
- The charge of a proton is +1.602 * 10-19 Coulombs.

### Electrons

Electrons are the subatomic particles that revolve around the nucleus of an atom. These electrons may be removed from or gained by an atom to form ions. Electrons of different atoms come together to participate in chemical bonding. A few points detailing the discovery and the properties of electrons are listed below.

- Electrons are negatively charged subatomic particles
- An equal number of electrons and protons are found in the atoms of all elements.
- J. Thompson is credited with the discovery of electrons since he was the first person to accurately calculate the mass and the charge on an electron.
- The mass of an electron is negligible when compared to the mass of a proton. It is found to have a mass equal to (1/1837) times the mass of a proton.
- The charge of an electron is equal to -1.602 * 10-19 Coulombs.

## IMPORTANT QUESTION

### What are the three types of subatomic particles?

Subatomic particles include electrons, negatively charged, nearly massless particles that account for much of the atom's bulk, that include the stronger building blocks of the atom's compact yet very dense nucleus, the protons that are positively charged, and the strong neutrons that are electrically neutral.

### What is the smallest subatomic particle?

Quarks represent the smallest subatomic particles that are known. The modern elementary particles are thought to be certain building blocks of matter, substituting protons, neutrons and electrons as the fundamental particles of the universe.

### Is anything smaller than an atom?

Subatomic particles are lighter than atoms in the physical sciences. They may be artificial particles, such as neutrons and protons, or elementary particles that are not constructed of such particles in compliance with the standard model.

### Is a photon smaller than an atom?

The Quantum of Electromagnetic Radiation is a photon, while an atom is the central part of all matter. Its size can be close to the size of Electron-like subatomic particles, but it is smaller than an atom.

### Does a photon have size?

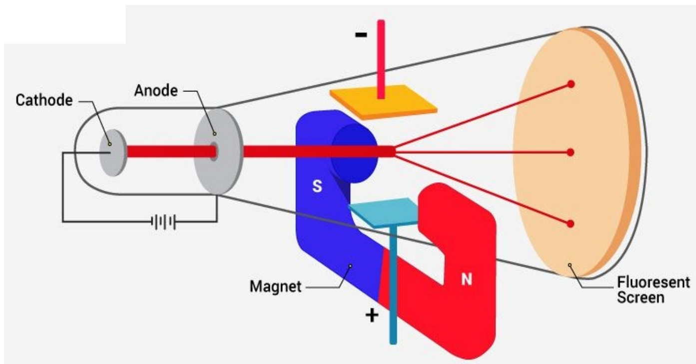
The particle aspect is often helpful, depending on the case, and sometimes the wave aspect is. Although photons do not have a physical diameter and may be viewed as point particles, they are given a probabilistic size by their quantum behaviour. Under this description, a photon has no absolute "scale."

# CHARGE TO MASS RATIO OF ELECTRON

**e/m = 1.758820 × 1011 C/kg**

Where,

- **m** = mass of an electron in kg = **9.10938356 × 10-31 kilograms**.
- **E** = magnitude of the charge of an electron in coulombs =**1.602 x 10-19 coulombs**.




While carrying out the discharge tube experiment, Thomson observed that the particles of the cathode deviate from their path. He noticed the amount of deviation in the presence of an electrical or magnetic field depends on various related parameters. They are:

1.Particles with a greater magnitude of the charge experienced greater interaction with the electric or magnetic field. Thus, they exhibited greater deflection.

2.Lighter particle experienced greater deflection. Thus, deflection is inversely proportional to the mass of the particle.

3.Deflection of the particle from their path is directly proportional to the strength of the electrical and the magnetic field present.

# IMPORTANT QUESTION

### What is the mass of one proton?

The proton is a stable subatomic particle with a positive charge equal to that of an electron and a rest mass of 1.67262 1027 kg, or 1,836 times the mass of an electron.

### Do Protons have mass?

Protons, neutrons, and electrons: The nucleus contains protons and neutrons, each of which has a mass of one amu. Protons, on the other hand, have a charge of +1, whereas neutrons are uncharged. Electrons have a charge of -1 and have a mass of around 0 amu. They circle the nucleus and have a mass of roughly 0 amu.

### What is an electron?

An electron is a subatomic particle with a negative charge. It can be either free (not bound to any atom) or tied to an atom's nucleus. The energy levels of electrons in atoms are represented by spherical shells of varied radii. The unit electrical charge is defined as the charge on a single electron.

### Where do protons get their mass?

These particles are made up of three quarks that are bonded together by gluons, the particles that convey the strong force, and move at dizzying speeds. The mass of protons and neutrons is determined by the energy of this interaction between quarks and gluons.

### Who named Proton?

Ernest Rutherford discovered the proton in the early 1900s. During this time, his study led to the first splitting of the atom, when he found protons through a nuclear reaction. His finding was given the term "protons" from the Greek word "protos," which means "first."

## Atomic Models

In the 18th and 19th centuries, many scientists attempted to explain the structure of the atom with the help of atomic models. Each of these models had their own merits and demerits and were pivotal to the development of the **modern atomic model**. The most notable contributions to the field were by the scientists such as



John Dalton, J.J. Thomson, Ernest Rutherford and Niels Bohr. Their ideas on the structure of the atom are discussed in this subsection

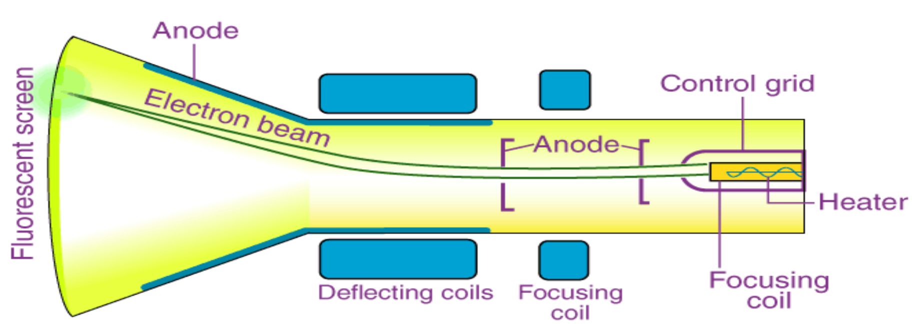
## Thomson Atomic Model

The English chemist Sir Joseph John Thomson put forth his model describing the atomic structure in the early 1900s.

He was later awarded the Nobel prize for the **discovery of “electrons”**. His work is based on an experiment called cathode ray experiment.

## Cathode Ray Experiment

It has a tube made of glass which has two openings, one for the vacuum pump and the other for the inlet through which a gas is pumped in.

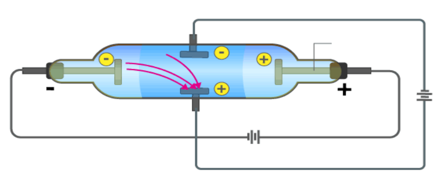



**Cathode Ray Tube Experiment**

The role of the vacuum pump is to maintain “partial vacuum” inside the glass chamber. A high voltage power supply is connected using electrodes i.e. cathode and Anode is fitted inside the glass tube.

### Observations:

- When a high voltage power supply is switched on, there were rays emerging from the cathode towards the anode. This was confirmed by the ‘Fluorescent spots’ on the ZnS screen used. These rays were called “Cathode Rays”.
- When an external electric field is applied, the cathode rays get deflected towards the positive electrode, but in the absence of electric field, they travel in a straight line.
- When rotor Blades are placed in the path of the cathode rays, they seem to rotate. This proves that the cathode rays are made up of particles of a certain mass, so that they have some energy.




**The direction of Cathode rays in the absence of an Electric Field**

- With all this evidence, Thompson concluded that cathode rays are made of negatively charged particles called “electrons”.
- On applying the electric and magnetic field upon the cathode rays (electrons), Thomson found the charge to mass ratio (e/m) of electrons. (e/m) for electron: 17588 × 1011 e/bg.

From this ratio, the charge of the electron was found by Mullikin through oil drop experiment. [Charge of e– = 1.6 × 10-16 C and Mass of e– = 9.1093 × 10-31 kg]

Conclusions:

Based on conclusions from his cathode ray experiment, Thomson described the atomic structure as a positively charged sphere into which negatively charged electrons were embedded.

It is commonly referred to as the “**plum pudding model”** because it can be visualized as a plum pudding dish where the pudding describes the positively charged atom and the plum pieces describe the electrons.

Thomson’s atomic structure described atoms as electrically neutral, i.e. the positive and the negative charges were of equal magnitude.

**Limitations of Thomson’s Atomic Structure:** Thomson’s atomic model does not clearly explain the stability of an atom. Also, further discoveries of other subatomic particles, couldn’t be placed inside his atomic model.

.Rutherford Atomic Theory

Rutherford, a student of J. J. Thomson modified the atomic structure with the discovery of another **subatomic particle called “Nucleus”**. His atomic model is based on the Alpha ray scattering experiment.

Alpha Ray Scattering Experiment

Construction:

- A very thin gold foil of 1000 atoms thick is taken.
- Alpha rays (doubly charged Helium He2+) were made to bombard the gold foil.
- Zn S screen is placed behind the gold foil.

Observations:

- Most of the rays just went through the gold foil making scintillations (bright



spots) in the ZnS screen.
- A few rays got reflected after hitting the gold foil.
- One in 1000 rays got reflected by an angle of 180° (retraced path) after hitting the gold foil.

Conclusions:

- Since most rays passed through, Rutherford concluded that most of the space inside the atom is empty.

- Few rays got reflected because of the repulsion of its positive with some other positive charge inside the atom.
- 1/1000th of rays got strongly deflected because of a very strong positive charge in the center of the atom. He called this strong positive charge as “nucleus”.
- He said most of the charge and mass of the atom resides in the Nucleus

## Rutherford’s Structure of Atom

Based on the above observations and conclusions, Rutherford proposed his own atomic structure which is as follows.

- The nucleus is at the center of an atom, where most of the charge and mass are concentrated.
- Atomic structure is spherical.
- Electrons revolve around the nucleus in a circular orbit, similar to the way planets orbit the sun.

## Limitations of Rutherford Atomic Model

- If electrons have to revolve around the nucleus, they will spend energy and that too against the strong force of attraction from the nucleus, a lot of energy will be spent by the electrons and eventually, they will lose all their energy and will fall into the nucleus so the stability of atom is not explained.
- If electrons continuously revolve around the ‘nucleus, the type of spectrum expected is a continuous spectrum. But in reality, what we see is a line spectrum

## Subatomic Particles

### Protons

- Protons are positively charged subatomic particles. The charge of a proton is 1e, which corresponds to approximately 1.602 × 10-19
- The mass of a proton is approximately 1.672 × 10-24
- Protons are over 1800 times heavier than electrons.
- The total number of protons in the atoms of an element is always equal to the atomic number of the element.

### Neutrons

- The mass of a neutron is almost the same as that of a proton i.e. 1.674×10-24
- Neutrons are electrically neutral particles and carry no charge.
- Different isotopes of an element have the same number of protons but vary in the number of neutrons present in their respective nuclei.

### Electrons

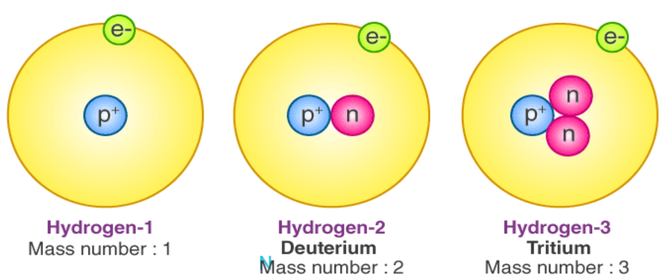
- The charge of an electron is -1e, which approximates to -1.602 × 10-19
- The mass of an electron is approximately 9.1 × 10-31.
- Due to the relatively negligible mass of electrons, they are ignored when calculating the mass of an atom.

### Atomic Structure of Isotopes

Nucleons are the components of the nucleus of an atom. A nucleon can either be a proton or a neutron. Each element has a unique number of protons in it, which is described by its unique atomic number. However, several atomic structures of an element can exist, which differ in the total number of nucleons.

These variants of elements having a different nucleon number (also known as the mass number) are called isotopes of the element. Therefore, the isotopes of an element have the same number of protons but differ in the number of neutrons.

The atomic structure of an isotope is described with the help of the chemical symbol of the element, the atomic number of the element, and the mass number of the isotope. For example, there exist three known naturally occurring isotopes of hydrogen, namely, protium, deuterium, and tritium. The atomic structures of these hydrogen isotopes are illustrated below



## Dalton’s Atomic Theory

The English chemist **John Dalton** suggested that all matter is made up of atoms, which were indivisible and indestructible. He also stated that all the atoms of an element were exactly the same, but the atoms of different elements differ in size and mass.

Chemical reactions, according to Dalton’s atomic theory, involve a rearrangement of atoms to form products. According to the postulates proposed by Dalton, the atomic structure comprised atoms, the smallest particle responsible for the chemical reactions to occur.

**The following are the postulates of his theory:**

- Every matter is made up of atoms.
- Atoms are indivisible.
- Specific elements have only one type of atoms in them.
- Each atom has its own constant mass that varies from element to element.
- Atoms undergo rearrangement during a chemical reaction.
- Atoms can neither be created nor be destroyed but can be transformed from one form to another.

Dalton’s atomic theory successfully explained the Laws of chemical reactions, namely, the Law of conservation of mass, Law of constant properties, Law of multiple proportions and Law of reciprocal proportions.

## Demerits of Dalton’s Atomic Theory

- The theory was unable to explain the existence of isotopes.
- Nothing about the structure of atom was appropriately explained.
- Later, the scientists discovered particles inside the atom that proved, the atoms are divisible.

## Atomic Structures of Some Elements

The structure of atom of an element can be simply represented via the total number of protons, electrons, and neutrons present in it. The atomic structures of a few elements are illustrated below.

### Hydrogen

The most abundant isotope of hydrogen on the planet Earth is protium. The atomic



number and the mass number of this isotope are 1 and 1, respectively.

**Structure of Hydrogen atom:** This implies that it contains one proton, one electron, and no neutrons ( total number of neutrons = mass number – atomic number)

## Carbon

Carbon has two stable isotopes – 12C and 13C. Of these isotopes, 12C has an abundance of 98.9%. It contains 6 protons, 6 electrons, and 6 neutrons.

**Structure of Carbon atom:** The electrons are distributed into two shells and the outermost shell (valence shell) has four electrons. The tetravalency of carbon enables it to form a variety of chemical bonds with various elements.

## Oxygen

There exist three stable isotopes of oxygen – 18O, 17O, and 16O. However, oxygen-16 is the most abundant isotope.

**Structure of Oxygen atom:** Since the atomic number of this isotope is 8 and the mass number is 16, it consists of 8 protons and 8 neutrons. 6 out of the 8 electrons in an oxygen atom lie in the valence shell.

## Bohr's Atomic Theory

Neils Bohr put forth his model of the atom in the year 1915. This is the most widely used atomic model to describe the atomic structure of an element which is based on Planck's theory of quantization.

## Postulates:

- The electrons inside atoms are placed in discrete orbits called "stationery orbits".
- The energy levels of these shells can be represented via quantum numbers.
- Electrons can jump to higher levels by absorbing energy and move to lower energy levels by losing or emitting its energy.
- As longs as, an electron stays in its own stationery, there will be no absorption or emission of energy.
- Electrons revolve around the nucleus in these stationery orbits only.
- The energy of the stationary orbits is quantized.

### Limitations of Bohr's Atomic Theory:

- Bohr's atomic structure works only for single electron species such as H, He+, Li2+, Be3+, ….



- When the emission spectrum of hydrogen was observed under a more accurate spectrometer, each line spectrum was seen to be a combination of no of smaller discrete lines.
- Both Stark and Zeeman effects couldn't be explain using Bohr's theory.

**Heisenberg's uncertainty principle:** Heisenberg stated that no two conjugate physical quantities can be measured simultaneously with 100% accuracy. These will always be some error or uncertainty in the measurement.

**Drawback:** Position and momentum are two such conjugate quantities that were measured accurately by Bohr (theoretically).

**Stark effect:** Phenomenon of deflection of electrons in the presence of an electric field.

**Zeeman effect:** Phenomenon of deflection of electrons in the presence of a magnetic field.

## Dual Nature of Matter

The electrons which were treated to be particles, the evidence of photoelectric effect shows they also have wave nature. This was proved by Thomas young with the help of his double slit experiment.

De-Broglie concluded that since nature is symmetrical, so should be light or any other matter wave.

## Quantum Numbers

- **Principal Quantum number (n):** t denotes the orbital number or shell number of electron.
- **Azimuthal Quantum numbers (l):** It denotes the orbital (sub-orbit) of the electron.
- **Magnetic Quantum number:** itnotes the number of energy states in each orbit.
- **Spin Quantum number(s)** It denotes the direction of spin, S = -½ = Anticlockwise and ½ = Clockwise.

## Electronic Configuration of an Atom

The electrons have to be filled in the s, p, d, f in accordance with the following rule.

**1. Aufbau's principle:** The filling of electrons should take place in accordance with the ascending order of energy of orbitals:

- Lower energy orbital should be filled first and higher energy levels.
- The energy of orbital **(p + l)** value it two orbitals have same (n + l) value, E α n
- Ascending order of energy 1s, 2s, 2p, 3s, 3p, 4s, 3d, . . .

**2. Pauli's exclusion principle:** No two electrons can have all the four quantum numbers to be the same or if two electrons have to be placed in an energy state they should be placed with opposite spies.

**3. Hund's rule of maximum multiplicity:** In the case of filling degenerate (same energy) orbitals, all the degenerate orbitals have to be singly filled first and then only pairing has to happen.

## IMPORTANT QUESTION

### What are subatomic particles?

Subatomic particles are the particles that constitute an atom. Generally, this term refers to protons, electrons, and neutrons.

### How do the atomic structures of isotopes vary?

They vary in terms of the total number of neutrons present in the nucleus of the atom, which is described by their nucleon numbers.

### What are the shortcomings of Bohr's atomic model?

According to this atomic model, the structure of an atom offers poor spectral predictions for larger atoms. It also failed to explain the Zeeman effect. It could only successfully explain the hydrogen spectrum.

### How can the total number of neutrons in the nucleus of a given isotope be determined?

The mass number of an isotope is given by the sum of the total number of protons and neutrons in it. The atomic number describes the total number of protons in the nucleus. Therefore, the number of neutrons can be determined by subtracting the atomic number from the mass number.

The **electron configuration** of an element describes how electrons are distributed in its atomic orbitals. Electron configurations of atoms follow a standard notation in which all electron-containing atomic subshells (with the number of electrons they hold written in superscript) are placed in a sequence. For example, the electron

configuration of sodium is 1s22s22p63s1.

However, the standard notation often yields lengthy electron configurations (especially for elements having a relatively large atomic number). In such cases, an abbreviated or condensed notation may be used instead of the standard notation. In the abbreviated notation, the sequence of completely filled subshells that correspond to the electronic configuration of a noble gas is replaced with the symbol of that noble gas in square brackets. Therefore, the abbreviated electron configuration of sodium is [Ne]3s1 (the electron configuration of neon is 1s22s22p6, which can be abbreviated to [He]2s22p6).

| | | | |
|---|---|---|---|
| Neon | $1s^2 2s^2 2p^6$ | | |
| Aluminum | $1s^2 2s^2 2p^6 3s^2 3p^1$ | Becomes | [Ne] $3s^2 3p^1$ |
| Ar gon | $1s^2 2s^2 2p^6 3s^2 3p^6$ | | |
| Calcium | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ | Becomes | [Ar] $4s^2$ |

Electron Configurations are useful for:

- •Determining the valency of an element.
- •Predicting the properties of a group of elements (elements with similar electron configurations tend to exhibit similar properties).
- •Interpreting atomic spectra.

This notation for the distribution of electrons in the atomic orbitals of atoms came into practice shortly after the Bohr model of the atom was presented by Ernest Rutherford and Niels Bohr in the year 1913.

## Writing Electron Configurations

### Shells

The maximum number of electrons that can be accommodated in a shell is based on the principal quantum number (n). It is represented by the formula 2n2, where ‘n’ is the shell number. The shells, values of n, and the total number of electrons that can be accommodated are tabulated below.

| | |
|---|---|
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| Shell and ‘n’ value | Maximum electrons present in the shell |
|---|---|
| K shell, n=1 | 2*12 = 2 |
| L shell, n=2 | 2*22 = 8 |
| M shell, n=3 | 2*32 = 18 |
| N shell, n=4 | 2*42 = 32 |

## Subshells

- •The subshells into which electrons are distributed are based on the azimuthal quantum number (denoted by ‘l’).
- •This quantum number is dependent on the value of the principal quantum number, n. Therefore, when n has a value of 4, four different subshells are possible.
- •When n=4. The subshells correspond to l=0, l=1, l=2, and l=3 and are named the s, p, d, and f subshells, respectively.
- •The maximum number of electrons that can be accommodated by a subshell is given by the formula 2*(2l + 1).
- •Therefore, the s, p, d, and f subshells can accommodate a maximum of 2, 6, 10, and 14 electrons, respectively.

All the possible subshells for values of n up to 4 are tabulated below.

| Principle Quantum Number Value | Value of Azimuthal Quantum Number | Resulting Subshell in the Electron Configuration |
|---|---|---|
| n=1 | l=0 | 1s |
| n=2 | l=0 | 2s |
| | l=1 | 2p |
| n=3 | l=0 | 3s |
| | l=1 | 3p |
| | l=2 | 3d |
| n=4 | l=0 | 4s |
| | l=1 | 4p |
| | l=2 | 4d |
| | l=3 | 4f |

Thus, it can be understood that the 1p, 2d, and 3f orbitals do not exist because the value of the azimuthal quantum number is always less than that of the principal quantum number.

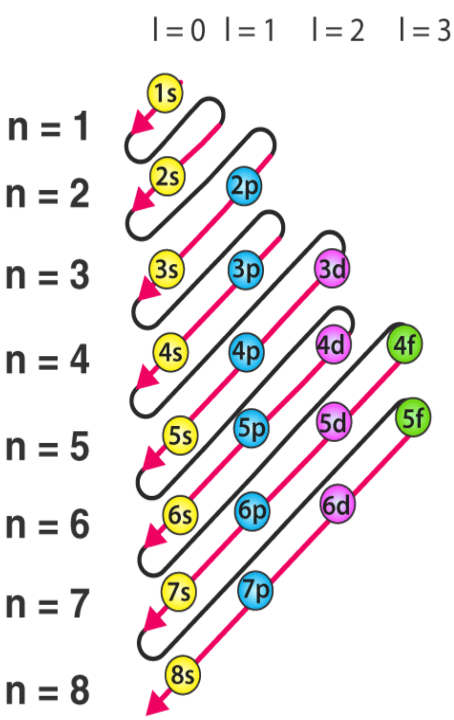
## Notation

- The electron configuration of an atom is written with the help of subshell labels.
- These labels contain the shell number (given by the principal quantum number), the subshell name (given by the azimuthal quantum number) and the total number of electrons in the subshell in superscript.
- For example, if two electrons are filled in the 's' subshell of the first shell, the resulting notation is '1s2'.
- With the help of these subshell labels, the electron configuration of magnesium (atomic number 12) can be written as 1s2 2s2 2p6 3s2.

## Filling of Atomic Orbitals

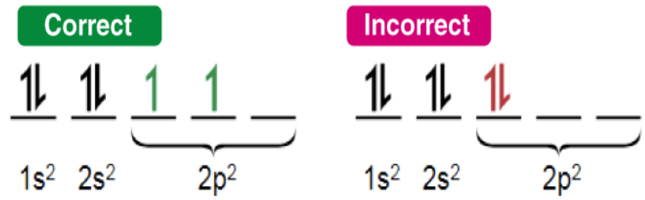
## Aufbau Principle

- This principle is named after the German word 'Aufbeen' which means 'build up'.
- The Aufbau principle dictates that electrons will occupy the orbitals having lower energies before occupying higher energy orbitals.
- The energy of an orbital is calculated by the sum of the principal and the azimuthal quantum numbers.
- According to this principle, electrons are filled in the following order: 1s, 2s, 2p, 3s,




3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p…

The order in which electrons are filled in atomic orbitals as per the Aufbau principle is illustrated uppar




## Representation of electronic Configuration of Atom

The electron configurations of a few elements are provided with illustrations in this subsection.

## Electron Configuration of Hydrogen

The atomic number of hydrogen is 1. Therefore, a hydrogen atom contains 1 electron, which will be placed in the s subshell of the first shell/orbit. The electron configuration of hydrogen is **1s1**, as illustrated below.

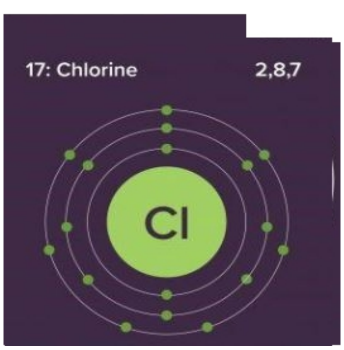
Electron Configuration of Oxygen

The atomic number of oxygen is 8, implying that an oxygen atom holds 8 electrons. Its electrons are filled in the following order:

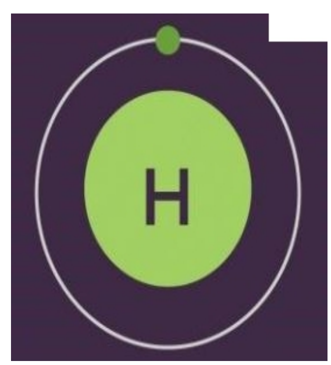
K shell – 2 electrons

L shell – 6 electrons

Therefore, the electron configuration of oxygen is **1s2 2s2 2p4**, as shown in the illustration provided below.





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## Chlorine Electronic Configuration

Chlorine has an atomic number of 17. Therefore, its 17 electrons are distributed in the following manner:

K shell – 2 electrons

L shell – 8 electrons

M shell – 7 electrons

The electron configuration of chlorine is illustrated below. It can be written as **1s22s22p63s23p5** or as $[Ne]3s23p^5$

# IMPORTANT QUESTION

### What is meant by the electronic configuration of an element?

The electronic configuration of an element is a symbolic notation of the manner in which the electrons of its atoms are distributed over different atomic orbitals. While writing electron configurations, a standardized notation is followed in which the energy level and the type of orbital are written first, followed by the number of electrons present in the orbital written in superscript. For example, the electronic configuration of carbon (atomic number: 6) is $1s_2 2s_2 2p_2$.

### What are the three rules that must be followed while writing the electronic configuration of elements?

The three rules that dictate the manner in which electrons are filled in atomic orbitals are:

- **The Aufbau principle:** electrons must completely fill the atomic orbitals of a given energy level before occupying an orbital associated with a higher energy level. Electrons occupy orbitals in the increasing order of orbital energy level.
- **Pauli's exclusion principle: states that no two electrons can have equal values for all four quantum numbers. Consequently, each subshell of an orbital can accommodate a maximum of 2 electrons and both these electrons MUST have opposite spins.**
- Hund's rule of maximum multiplicity: All the subshells in an orbital must be singly occupied before any subshell is doubly occupied. Furthermore, the spin of all the electrons in the singly occupied subshells must be the same (in order to maximize the overall spin).

### Why are electronic configurations important?

Electron configurations provide insight into the chemical behaviour of elements by helping determine the valence electrons of an atom. It also helps classify elements into different blocks (such as the s-block elements, the p-block elements, the d-block elements, and the f-block elements).



This makes it easier to collectively study the properties of the elements.

### List the electron configurations of all the noble gases.

The electronic configurations of the noble gases are listed below.

- Helium (He) – $1s_2$
- Neon (Ne) – $[He]2s_2 2p_6$
- Argon (Ar) – $[Ne]3s_2 3p_6$

- Krypton (Kr) – [Ar]3d104s24p6
- Xenon (Xe) – [Kr]4d105s25p6
- Radon (Rn) – [Xe]4f145d106s26p6

### What is the electronic configuration of copper?

The electronic configuration of copper is [Ar]3d104s1. This configuration disobeys the aufbau principle due to the relatively small energy gap between the 3d and the 4s orbitals. The completely filled d-orbital offers more stability than the partially filled configuration

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