Atomic Structure & Chemical Bonding

Thomson's Atomic Model

Atom Is Divisible

Do you recall **Dalton's atomic theory**? Dalton postulated in his theory that an **atom is indivisible**. However, the later discoveries of **protons** and **electrons** proved this to be erroneous.

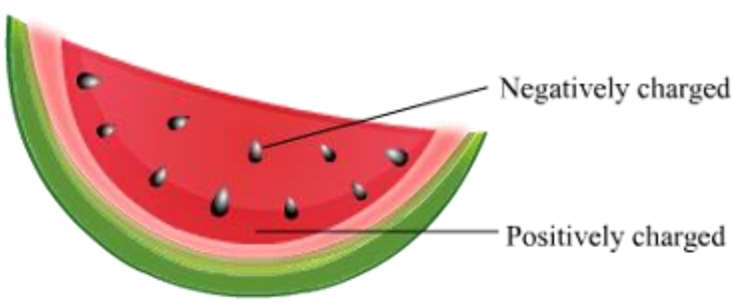
In 1886, while carrying out an experiment in a gas discharge tube, E. Goldstein discovered positively charged radiations which led to the discovery of the subatomic particles called protons. Later, in 1897, J. J. Thomson discovered another type of subatomic particle—the negatively charged electron. Consequent to these discoveries, an atom was no longer indivisible; rather, it became a sum total of differently charged subatomic particles.

We know that an atom is neutral. It is made up of an equal number of oppositely charged particles—protons and electrons. Now, the question that arises is this:

How are the subatomic particles arranged inside an atom?

Many scientists performed varied experiments to develop different models for the structure of an atom. The first such model was proposed by J. J. Thomson. His atomic model is compared to a plum pudding and a watermelon; hence, it is known by the names 'the plum-pudding model'.

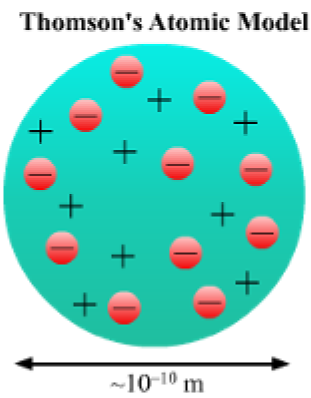
The Plum-Pudding Model of an Atom



Let us understand Thomson's atomic model with the help of a slice of a watermelon. The slice consists of a red edible portion with embedded black seeds. Now, if we liken this watermelon to an atom, then (as per Thomson's model) the positive charge in the atom is spread all over the red edible part; and the negatively charged particles, like the seeds, are embedded in this positively charged space.

In the same way, we can liken an atom to a plum pudding. In this case, the positive charge is spread all over the pudding, while the negatively charged particles are embedded like plums in this positively charged space.

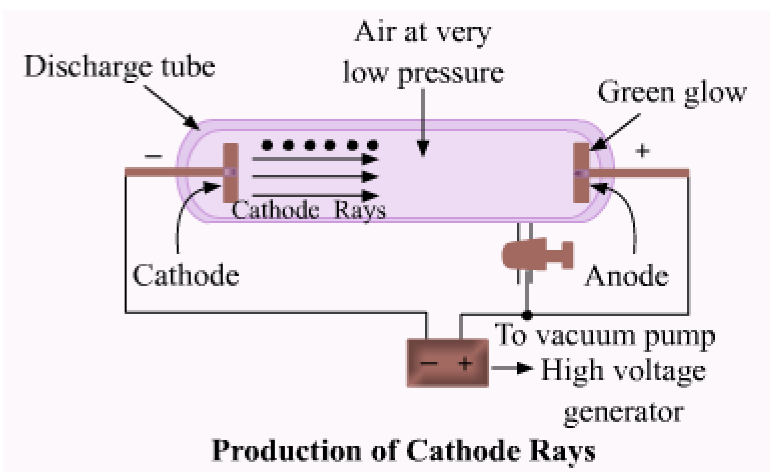
According to Thomson's atomic model:



1. An atom consists of a positively charged sphere with electrons embedded in it.

2. The negative and positive charges present inside an atom are equal in magnitude. Therefore, an atom as a whole is electrically neutral.

Cathode Rays



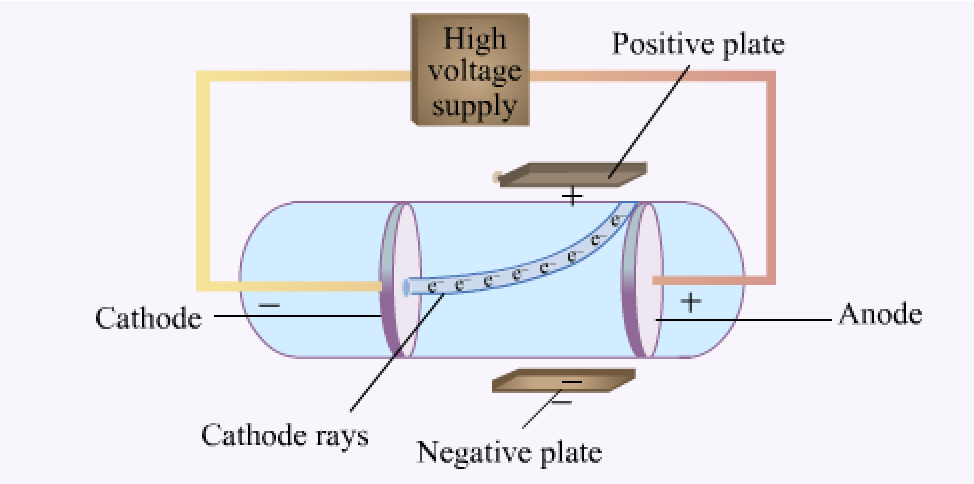
J.J Thomson discovered that there are small particle present in the atom and that atom is divisible. J.J Thomson and his colleagues conducted experiments using discharge tube apparatus.

A discharge tube apparatus consists of a glass tube of about 15 cm length and 3 cm in diameter, filled with gas at low pressure. The tube is connected with the vacuum pump and two metal electrodes are fitted to the ends of the tube.

Low pressure was created inside the tube and high voltage was applied to the electrodes of the tube. This produced greenish glow at the anode end of the tube. The greenish glow at anode was produced due to the emission of the streams of rays from the cathode. These rays are known as cathode rays. Cathode rays will emit with blue glow.

When J.J Thomson placed a light paddle wheel inside the tube in the path of the cathode rays, the wheel started rotating. This led him to conclude that cathode rays are particulate in nature.

Properties of Cathode Rays



Experiment to demonstrate properties of cathode rays

When J.J Thomson applied an electric field in the direction parallel to the path of cathode rays, he observed that the rays were deflected towards the anode.

This observation led to the conclusion that cathode rays are negatively charged.

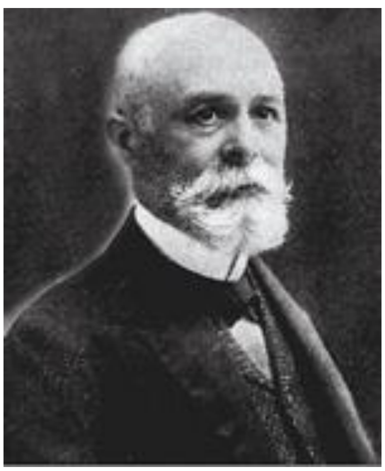
When the above experiment was conducted with different gases, same observation were made and he named these negatively charged particles as electrons.

An electron is lighter than hydrogen atom and has very small mass in comparison to the mass of an atom.

Thus, J.J Thomson's experiment and discovery of electron proved that atom is divisible and is made up of sub – atomic particles.

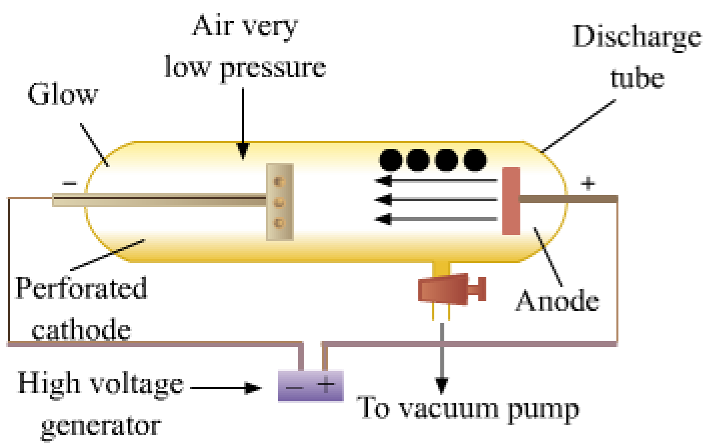
Know Your Scientist

Sir Joseph John Thomson (1856−1940) was a British physicist. He is known for the discovery of electrons and for his model of an atom, popularly known as ‘the plum-pudding model’. He received the Nobel Prize in Physics in 1906 for discovering electrons and for his research on conduction in gases. In 1912, while working on the composition of canal rays, he and his colleague (F. W. Aston) found the first evidence for isotopes of neon.



Eugen Goldstein (1850−1930) was a German physicist. He is known for the discovery of canal rays which led to the discovery of protons. He also investigated comets using gas discharge tubes. His experiments established that a small object (like a ball) placed in the path of cathode rays produces emissions, flaring outward just like in case of a comet’s tail.

Canal Rays



Production of anode rays

After J.J Thomson's discovery of atom another question arose that: if electrons are present inside the atom, then how is atom electrically neutral? Does this mean that there are positively charged particle also present inside the atom?

To find out the answers to such questions, Goldstein conducted an experiment similar to that of J.J Thomson's but with some modifications, for example he used perforated cathode in the discharge tube.

It was observed during the experiment that some rays were travelling in the direction opposite to that of cathode rays. Goldstein named these rays as anode rays.

When he applied an electric field in the direction parallel to that of the rays he observed that rays deflected towards cathode, thereby he concluded that anode rays are positively charged.

However, the deflection of anode rays in the discharge tube was found to be very less than that of cathode rays, because the emission of cathode rays was not dependent on the nature of the gas taken in the discharge tube. The deflection was seen highest for the hydrogen gas, when taken in the discharge tube.

The positive particles of hydrogen were found to be lightest and were named protons. Their mass is approximately equal to 1840 times that of electron. This mass is assumed as **1 atomic mass unit**. The charge on a proton (+1) is equal to charge on an electron in magnitude (-1).

Rutherford's Atomic Model

Why the Plum-Pudding Model Failed

The plum-pudding model of an atom was unable to explain the findings of Rutherford's experiment while studying radioactivity.

In an experiment with gold foil, Rutherford bombarded the gold foil with **alpha particles**. With Thomson's model as the basis, Rutherford expected small deviations; however, his findings were different from what was expected.

As we go further into this lesson, we will learn more about Rutherford's gold-foil experiment, his observations and his conclusions. We will also learn about the atomic model that he came up with on the basis of his conclusions.

Set up for Rutherford's experiment:

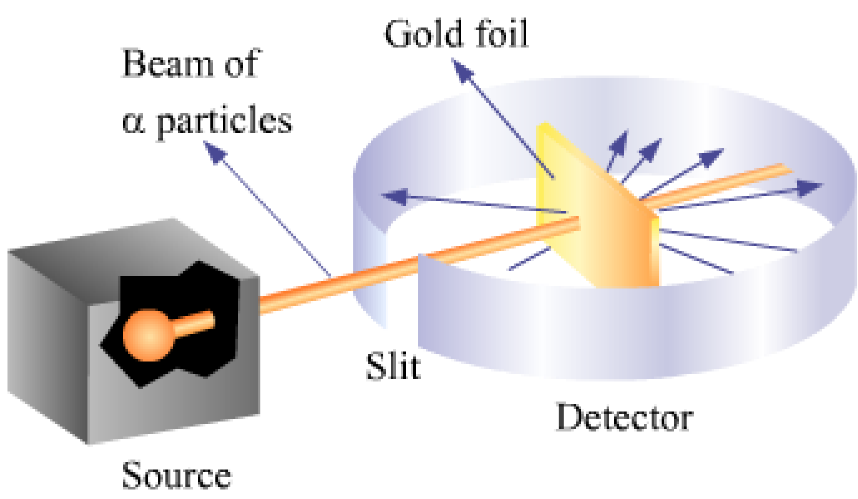
1. A thin gold foil, approximately 1000 atoms thick, was taken. Gold was chosen for its high malleability.

2. A detector screen with a small slit (for emission of radiation from the atom) was placed around the foil.

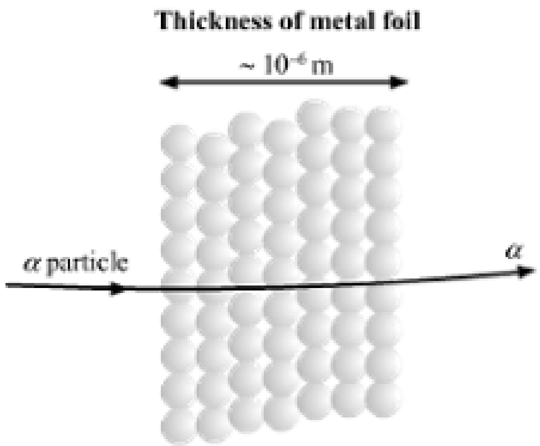
3. A source of alpha particles was kept in front of the foil.

4. The foil was bombarded with fast-moving alpha particles.

The set-up for Rutherford's gold-foil experiment is shown in the figure.



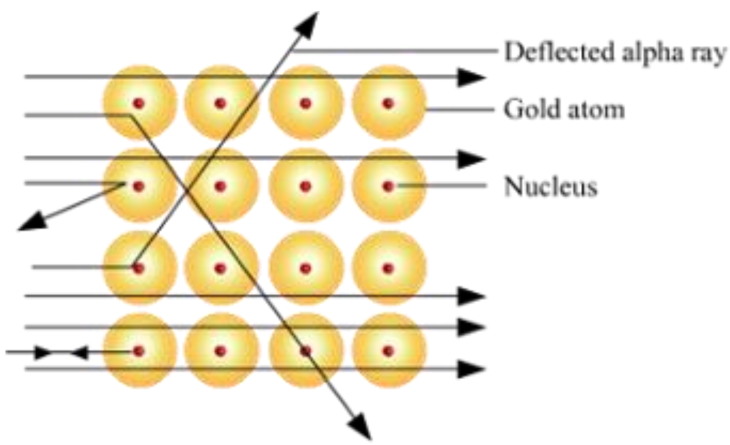
Rutherford's Expectations and Observations



What Rutherford expected?

Rutherford expected that the alpha particles would pass straight through the foil and only a small fraction of alpha particles would be deflected. This expectation was in compliance with Thomson's atomic model.

What Rutherford observed?



Rutherford's findings were contrary to his expectation. He observed that:

1. Most of the fast-moving alpha particles passed straight through the gold foil.

2. Some particles were deflected through the foil by small angles.

3. One out of every 12000 particles rebounded, i.e., they got deflected by an angle of 180°.

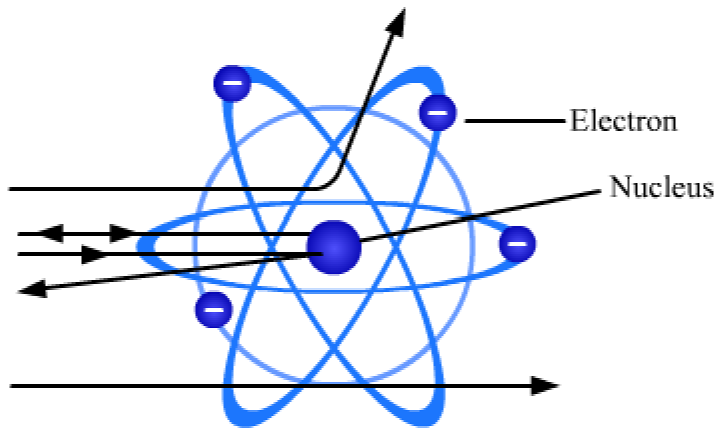
What Rutherford Concluded from His Observations

Rutherford then carefully studied his observations and made the following conclusions.

1. Most alpha particles passed through the gold foil without any deflection. This indicates that most of the space inside an atom is empty.

2. Very few particles suffered a deflection from their path. This means that positive charge occupies very little space inside an atom.

3. Only a small fraction of particles underwent a 180° deflection. This shows that the entire positive charge and mass of an atom are present within a very small volume inside the atom.



Rutherford's Atomic Model

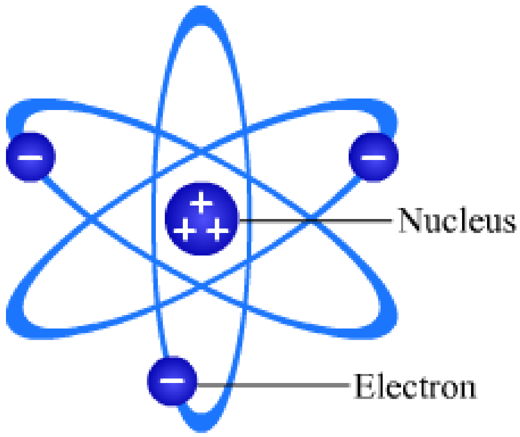
Based on his conclusions in the gold-foil experiment, Rutherford devised his own atomic model. The major features of **Rutherford atomic model** or **the nuclear model of an atom** are as follows:

1. An atom consists of a nucleus at its centre and all the protons are present inside this nucleus.

2. Electrons reside outside the nucleus and revolve around the nucleus in well-defined orbits.

3. The size of the nucleus is very small as compared to the size of the atom. As per Rutherford's calculations, the nucleus is 10^5 times smaller than the atom.

4. Since the mass of the electrons is negligible as compared to the mass of the protons, almost all the mass of the atom is concentrated in its nucleus.



Know Your Scientist

Ernest Rutherford (1871−1937) was a British chemist and physicist. He is known as 'the father of nuclear physics'. He discovered radioactive half-life. He proved that alpha radiations are nothing but helium ions. He was awarded the Nobel Prize in Chemistry in 1908 for his work on 'the disintegration of elements' and 'the chemistry of radioactive substances'. He was the first scientist to split an atom in a nuclear reaction. The element 'rutherfordium' (atomic number 104) is named after him.

Solved Examples

Hard

Example 1:

What would have been observed if neutrons had been used to bombard the gold foil?

1. **The observations of the experiment would have remained the same in spite of the change in the nature of the bombarding particles.**
2. **The neutrons would have suffered no deflection from the subatomic particles.**
3. **All the neutrons would have been absorbed by the gold atoms.**
4. **All the neutrons would have rebounded.**

Solution:

The correct answer is B.

Neutrons do not carry any charge; so, they do not suffer any repulsion. Hence, if neutrons had been used to bombard the gold foil, no deflection would have occurred. It is also possible that some neutrons would have been absorbed by the nucleus.

Medium

Example 2:

State whether the following statements are true (T) or false (F).

1. **Increasing the energy of the alpha particles will lead to more deflection.____**
2. **Speed of the alpha particles can be increased by increasing their energy.____**
3. **Use of aluminium sheet will lead to the same result as in case of gold foil.____**

Solution:

1. **T**: Increasing the energy of the alpha particles will cause them to strike closer to the nucleus. Consequently, they will suffer greater deflection.

2. **T**: The kinetic energy of the alpha particles is directly related to their velocity. So, increasing their energy will result in an increase in the speed of the particles.

3. **F**: The positive charge on the nucleus in case of an aluminium foil is much smaller as compared to that on the gold nucleus. So, the result will vary.

Easy

Example 3:

One of the postulates of Rutherford's atomic model is that

1. **an atom consists of a positively charged sphere.**
2. **an atom has its mass concentrated in its nucleus.**
3. **the nucleus of an atom is composed of electrons and protons.**
4. **the mass of an atom is the sum of the masses of all electrons and protons.**

Solution:

The correct answer is B.

The postulates of Rutherford's atomic model are as follows:

1. An atom consists of a nucleus at its centre and all the protons are present inside this nucleus.

2. Electrons reside outside the nucleus and revolve around the nucleus in well-defined orbits.

3. The size of the nucleus is very small as compared to the size of the atom. As per Rutherford's calculations, the nucleus is 10^5 times smaller than the atom.

4. Since the mass of the electrons is negligible as compared to the mass of the protons, almost all the mass of the atom is concentrated in its nucleus.

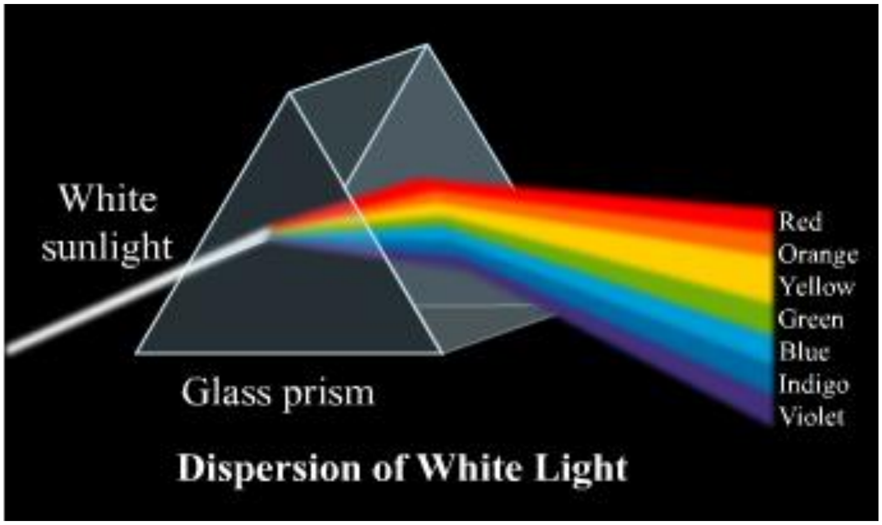
   Rutherford also noticed that the actual mass of the nucleus was much more higher than the sum of the masses of protons and electrons. This lead him to predict that nucleus contains some kind of neutral particle whose mass must be equal to that of proton.

   This was experimentally proved by James Chadwick in the year 1932. He proved that nucleus of atom contains an additional neutral particle and called them **neutrons.** The mass of these neutrons is equal to that of protons.

Bohr's Atomic Model

The Shell Model of an Atom

In the late nineteenth century, scientists researched on the dispersal of white light into its constituent seven colours—known as the spectrum.

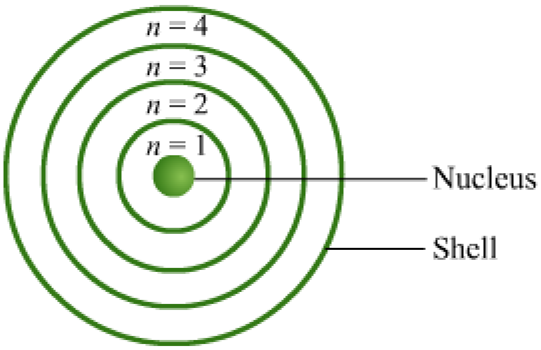


The spectrum was used for analyzing newly discovered elements. It was observed that this spectrum was different for different elements.

Now, the above observation could not be explained using Rutherford's model of an atom. Thus, Niels Bohr made some modifications to Rutherford's model. The modified atomic model of Niels Bohr is also known as **the shell model of an atom**.

Let us go through this lesson to learn more about this model.

Bohr's Model of an Atom



The postulates of Bohr's model of an atom are as follows:

1. Only certain special **orbits**, known as discrete orbits of electrons, are allowed inside the atom.

2. While revolving in the discrete orbits, the electrons do not radiate energy.

3. An electron can jump from one orbit to another by absorbing or emitting a fixed amount of energy in the form of radiation.

Bohr named these orbits as **energy levels**. These orbits (or shells) are represented by the letters **K, L, M, N**…, or the numbers n = 1, 2, 3, 4….

Bohr's model of an atom explains how:

- the electrons are arranged and distributed in the extra-nuclear space of the atom.
- the atom attains stability due to the presence of energy levels around the nucleus, in which the electrons revolve without radiating energy.
- each energy shell can accommodate only a fixed number of electrons.
- the filling of a shell begins only when the preceding shell has been completely filled.
- this gives the atom a shell-like structure.

Distribution of Electrons or Electronic Configuration

Let us see how the electrons are distributed in different orbits in an atom. Bohr, along with Charles R. Bury, suggested certain rules to show this electronic distribution. These rules (known as the Bohr−Bury scheme of electronic configuration) have to be followed while writing the **electronic configuration** of an atom.

1. The maximum number of electrons in a shell is given by the formula $2n^2$, where n is the orbit number or the energy level index (i.e., 1, 2, 3…).

Orbit numbers	Names of the shells	Numbers of electrons in the shells
1	K-shell	$2 \times (1)^2 = 2$
2	L-shell	$2 \times (2)^2 = 8$
3	M-shell	$2 \times (3)^2 = 18$
4	N-shell	$2 \times (4)^2 = 32$

2. The maximum number of electrons that can be accommodated in the outermost shell is 8.

3. The filling of the shells takes place in a stepwise manner. First, one shell is filled completely, then the next shell, and so on.

Solved Examples

Medium

Example 1:

Which of the following statements is true for Bohr's model of an atom?

1. **Electrons go around the nucleus based on the strength of the force of attraction extended by the nucleus.**
2. **Electrons lose energy while travelling in orbits around the nucleus.**
3. **The energy shells can accommodate only a fixed number of electrons.**
4. **The energy of the shells is continuous.**

Solution:

The correct answer is C.

According to Bohr's model of an atom:

- Only certain special orbits, known as discrete orbits of electrons, are allowed inside the atom.
- While revolving in the discrete orbits, the electrons do not radiate energy.
- The energy shells can accommodate only a fixed number of electrons.
- The energy of the shells is discrete, and not continuous.

Easy

Example 2:

An element has 12 electrons. How many energy shells does this element possess?

1. **1**
2. **2**
3. **3**
4. **4**

Solution:

The correct answer is C.

According to the Bohr−Bury scheme, each energy shell can accommodate $2n^2$ electrons, where n is the energy number. In case of any element, the electrons are distributed as shown in the table.

Shells	Energy numbers	Numbers of electrons in the shells
K	1	$2 \times (1)^2 = 2$
L	2	$2 \times (2)^2 = 8$
M	3	$2 \times (3)^2 = 18$

Now, the given element has 12 electrons. So, the K-shell fills first with 2 electrons. The remaining 10 electrons are divided among the L-shell and the M-shell as follows: 8 in the L-shell and 2 in the M-shell. Thus, the element has a total of 3 energy shells.

Hard

Example 3:

What is the electronic configuration for an element having 17 electrons?

1. **2, 8, 7**
2. **2, 6, 9**
3. **2, 10, 5**
4. **2, 7, 8**

Solution:

The correct answer is A.

According to the Bohr−Bury scheme, each energy shell can accommodate $2n^2$ electrons, where n is the energy number. In case of any element, the electrons are distributed as shown in the table.

Shells	Energy numbers	Numbers of electrons in the shells
K	1	$2 \times (1)^2 = 2$
L	2	$2 \times (2)^2 = 8$
M	3	$2 \times (3)^2 = 18$

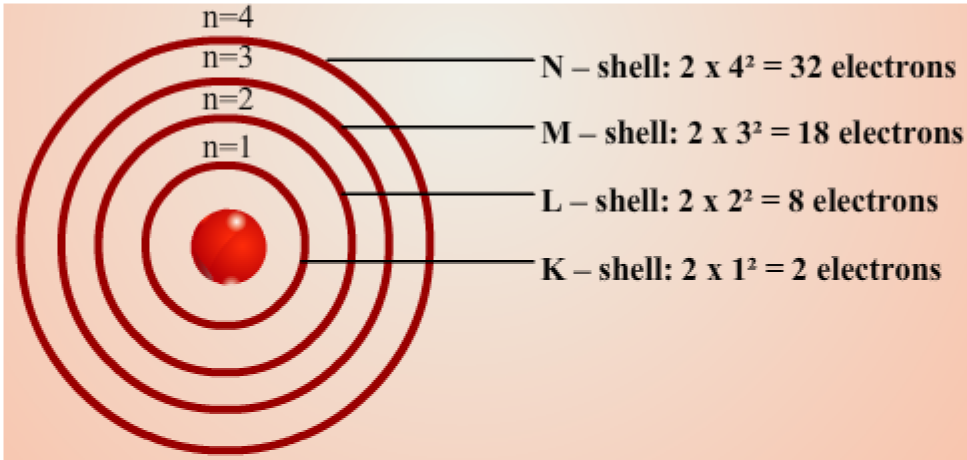
Now, the given element has 17 electrons. So, the K-shell fills first with 2 electrons. The remaining 15 electrons are divided among the L-shell and the M-shell as follows: 8 in the L-shell and 7 in the M-shell. Thus, the electronic configuration for the element is 2, 8, 7.

Electronic Configuration of Atoms

Bohr postulated that electrons can move from one shell to another by absorbing or emitting energy in the form of radiation. Let us understand this phenomenon.

As you know, in Bohr’s atomic model, electrons are arranged in energy shells and each of these shells has a fixed amount of energy. The electrons residing in a particular shell possess the characteristic energy of the shell. The energy of an electron remains constant as long as it remains in a particular energy level.

In an atom, shells are arranged in order of increasing energy.



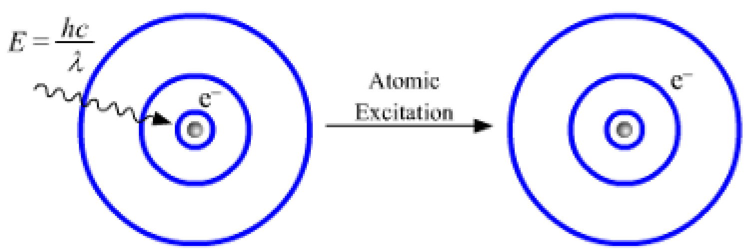
So, the shells in increasing order of energy are as follows:

K < L < M < N

Excitation and De-Excitation of Electrons

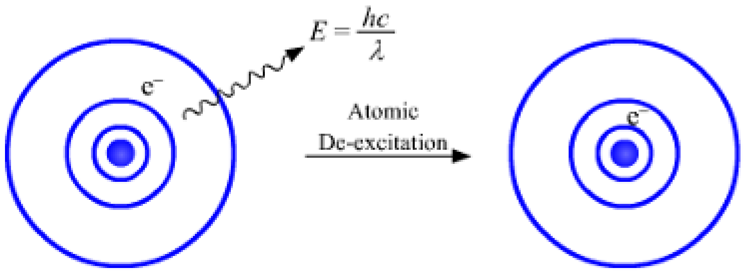
Excitation of electrons:

An electron lying in a lower energy shell absorbs energy from incident light to gain energy corresponding to that of the adjacent higher energy level. This is shown in the figure.



De-excitation of electrons:

An electron lying in a higher energy shell emits energy in the form of electromagnetic radiation to lose energy corresponding to that of the adjacent lower energy level. This is shown in the figure.



The emitted radiation may lie in any region of the electromagnetic spectrum, depending upon the frequency of the radiation.

Sub-Shells in an Atom and Discovery of Neutron

Sub-Shells in an Atom:

Further research on atomic structure indicated that the energy shells K, L, M, N, etc., are further divided into sub-shells of different shapes and energies. The division of shells into sub-shells is as follows:

Main energy levels	**Sub-shell(s) in the main levels**
K (1^{st})	s
L (2^{nd})	s, p
M (3^{rd})	s, p, d
N (4^{th})	s, p, d, f

In an atom, these sub-shells are arranged in order of increasing energy and are filled successively

Discovery of Neutrons:

- In 1932, J. Chadwick discovered the sub-atomic particles known as **neutrons**.
- Neutrons are present in all atomic nuclei, except that of hydrogen.
- The mass of a neutron was found to be 1.6749×10^{-27} kg, which is slightly more than that of a proton.
- Mass of an atom = Mass of all neutrons + Mass of all protons

Did You Know?

A hydrogen atom contains only one electron and one proton. It does not have any neutron in its nucleus. The removal of the electron from the atom leaves behind the single-proton-containing nucleus; so, H^+ is sometimes referred to as simply 'proton'.

Know Your Scientist

Neils Bohr (1885−1962) was a Danish physicist. He made major contributions to the understanding of the atomic structure and quantum mechanics. He was awarded the Nobel Prize in Physics in 1922. He is credited with developing the planetary model of an atom. While working on quantum mechanics, he postulated that electrons jump from one energy level to another by absorbing or emitting discrete amounts of energy. He also identified the uranium isotope U-238.

James Chadwick (1891−1974) was a British physicist. He is credited with the discovery of neutrons, for which he received the Nobel Prize in Physics in 1935. He was part of the Manhattan Project in the US and helped in the development of the atomic bombs that were dropped on Hiroshima and Nagasaki during the Second World War.

Atomic Number and Mass Number

In the 1830s, representation of elements and compounds was a major concern for chemists.

Many symbolic notations for elements were devised during this period. Gradually, the representations became standardized. Currently, the general symbolic notation for an element is:

$^{A}_{Z}E$. Now, take for example the specific symbolic notations for oxygen and nitrogen.

Element	Symbolic notation
Oxygen	${}^{16}_{8}O$
Nitrogen	${}^{14}_{7}N$

Wondering what these symbolic notations represent? Go through this lesson to find out.

You know that the symbolic notation of oxygen is ${}^{16}_{8}O$. In this notation, the letter 'O' symbolises the element 'oxygen'; the number '16' represents the **mass number** of oxygen; and the number '8' indicates the **atomic number** of oxygen.

Thus, in the general symbolic notation of an element $\left(\text{i.e., } {}^{A}_{Z}E\right)$, the letter 'E' is the symbol of the element, the letter 'A' is its mass number, and the letter 'Z' is its atomic number.

The **atomic number** is the number of protons present in the nucleus of an atom. It is denoted by **Z**.

The total number of the protons and the neutrons present in the nucleus of an atom is known as **mass number**. It is denoted by **A**.

Atomic Number and Mass Number

Symbolic Notations of Some Elements

Elements	Symbolic notations	Symbols	Atomic numbers	Mass numbers
Hydrogen	${}^{1}_{1}H$	H	1	1
Helium	${}^{4}_{2}He$	He	2	4
Lithium	${}^{7}_{3}Li$	Li	3	7
Beryllium	${}^{9}_{4}Be$	Be	4	9
Boron	${}^{11}_{5}B$	B	5	11
Carbon	${}^{12}_{6}C$	C	6	12
Nitrogen	${}^{14}_{7}N$	N	7	14
Oxygen	${}^{16}_{8}O$	O	8	16
Fluorine	${}^{19}_{9}F$	F	9	19

Neon	$^{20}_{10}Ne$	Ne	10	20

Symbolic Notations of Some Elements

Elements	Symbolic notations	Symbols	Atomic numbers	Mass numbers
Sodium	$^{23}_{11}Na$	Na	11	23
Magnesium	$^{24}_{12}Mg$	Mg	12	24
Aluminium	$^{27}_{13}Al$	Al	13	27
Silicon	$^{28}_{14}Si$	Si	14	28
Phosphorus	$^{31}_{15}P$	P	15	31
Sulphur	$^{32}_{16}S$	S	16	32
Chlorine	$^{35}_{17}Cl$	Cl	17	35
Argon	$^{40}_{18}Ar$	Ar	18	40
Potassium	$^{39}_{19}K$	K	19	39
Calcium	$^{40}_{20}Ca$	Ca	20	40

Relation between Atomic Number and Mass Number

Mass number **(A)** of an atom = Number of protons + Number of neutrons

Therefore, Mass number **(A)** = Atomic number **(Z)** + Number of neutrons

Therefore, Number of neutrons = **A - Z**

Hence, the number of neutrons can be calculated if the atomic number and mass number of an element are known.

An atom of sodium contains 11 protons and 12 neutrons. **Can you calculate the mass number of a sodium atom?**

Now, mass number **(A)** = number of protons + number of neutrons

Therefore, mass number of sodium atom = 11 + 12 = 23

Hence, the mass number of sodium is 23.

An atom of carbon is represented as ${}^{12}_{6}C$. **Can you tell the number of neutrons and protons present in carbon atom?**

It is seen from the symbolic notation of carbon that the atomic number and mass number of carbon atom is 6 and 12 respectively.

Now, number of neutrons = mass number − atomic number = 12 − 6 = 6

Since the number of protons is equal to the atomic number of that element. Thus, the number of protons present in a carbon atom is 6.

Solved Examples

Easy

Example 1:

What is the symbol of the element sodium?

1. **Na**
2. **N**
3. **So**
4. **S**

Solution:

The correct answer is A.

The symbol of sodium is Na. It is derived from the Latin name for the element, i.e., ‘natrium’.

Example 2:

What is the atomic number of an element having five protons and six neutrons?

1. **11**
2. **9**
3. **6**
4. **5**

Solution:

The correct answer is D.

The atomic number of an element is the number of protons or electrons present in an atom of the element. Since an atom of the given element has five protons, its atomic number is 5.

Medium

Example 3:

What is the number of neutrons in an element having 39 protons and 89 as its mass number?

1. **45**
2. **50**
3. **55**
4. **60**

Solution:

The correct answer is B.

We know that:

Mass number = Number of protons + Number of neutrons

In case of the given element:

Mass number = 89

Number of protons = 39

So,

89 = 39 + Number of neutrons

=> Number of neutrons = 89 -39 = 50

Hard

Example 4:

What is the symbol of the element having 22 neutrons and 40 as its mass number?

1. **Al**
2. **Mg**
3. **Ar**
4. **Ca**

Solution:

The correct answer is C.

The given element has:

Mass number = 40

Number of neutrons = 22

We know that:

Mass number = Number of protons + Number of neutrons

So,

40 = Number of protons + 22

=> Number of protons = 40 - 22 = 18

Also,

Atomic number = Number of protons = 18

Argon is the element having 18 as its atomic number and 40 as its mass number. The symbol of argon is Ar.

Did You Know?

- Water is the major constituent of the human body. It is made up of two elements: hydrogen and oxygen.

- Almost all the mass of our body is made up of the following six elements.

1. Oxygen (65%)
2. Carbon (18%)
3. Hydrogen (10%)
4. Nitrogen (3%)
5. Calcium (1.5%)
6. Phosphorus (1%)

- Some of the other elements found in our body are:
- Sulphur (0.25%)
- Sodium (0.15%)
- Magnesium (0.05%)
- Zinc (0.7%)

Whiz Kid

The Periodic Table

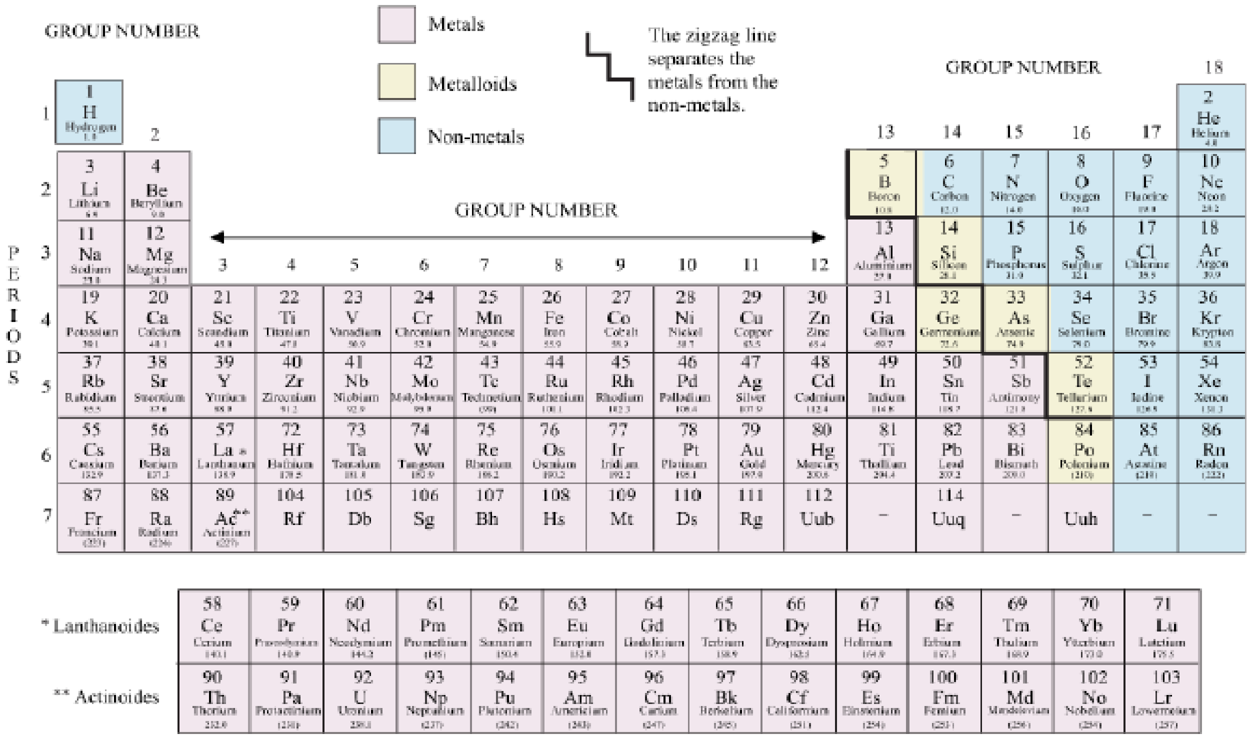
The periodic table is a table classifying all the known elements.

It is divided into 18 columns (called groups) and 7 rows (called periods).

The elements are arranged in the rows or periods by order of increasing atomic number.

The elements in the columns or groups display similar chemical and physical properties. This feature of the periodic table makes it easy to study the vast number of elements.

The periodic table is shown in the figure.



Isotopes

In 1910, Frederick Soddy recorded the existence of elements having different atomic masses to show similar properties. These elements came to be known as isotopes. Isotopes are defined as atoms of same element having the same atomic number, but different mass numbers. These atoms contain an equal number of protons and electrons, but a different number of neutrons.

For example, in nature, hydrogen is found in three forms with different mass numbers, namely protium $({}^1_1H)$, deuterium $({}^2_1H)$, and tritium $({}^3_1H)$. These are the three naturally occurring isotopes of hydrogen. The atomic number of each isotope is 1, but the mass number varies i.e. it is 1, 2, and 3 respectively. Some other examples of isotopes include C- 12 and C-14, which are isotopes of carbon, and Cl-35 and Cl-37, which are isotopes of chlorine.

However, it was F. W. Aston who, in 1919, discovered various stable isotopes for a number of elements. In that year, he was successful in proving the existence of two isotopes of neon:

neon-20 and neon-22.

The isotopes of an element are species with different numbers of **neutrons** but the same numbers of protons and electrons.

Isotopes are pure substances just like elements and compounds.

We know that the atomic mass of an element is the sum total of its protons and neutrons. Now, in case of an element having isotopes, the number of neutrons does not remain fixed. The atomic mass of such an element includes the atomic masses of its isotopes. This gives rise to the term ‘average atomic mass’.

The average atomic mass of an element with isotopes takes into account the atomic masses of all its isotopes with respect to their abundance in nature.

Applications of Isotopes

1. An isotope of uranium exhibits nuclear fission properties. It is used in nuclear reactions as a fuel.
2. An isotope of cobalt is used for treating cancer.
3. An isotope of iodine is used for treating goitre.
4. An isotope of carbon is used in radiocarbon dating to determine the age of an organic sample.
5. An isotope of calcium is used in biomedical research on cellular functions and bone formation in mammals.
6. An isotope of iron is used for detecting sulphur in air.
7. An isotope of hydrogen (tritium) is used in estimating the age of water bodies and the rate of their replenishment through precipitation.

Isobars and Isotones

Isobars

Isobars are elements having the same mass number but different atomic numbers. Take, for example, ${}^{13}_{6}C$ and ${}^{13}_{7}N$. They have the same mass number (i.e., 13) but different atomic numbers (i.e., 6 and 7).

One can also say that isobars have an equal number of nucleons but different numbers of protons, neutrons and electrons.

Here are two more examples of isobars:

1. ${}^{14}_{6}C$ and ${}^{14}_{7}N$
2. ${}^{40}_{18}Ar$, ${}^{40}_{19}K$ and ${}^{40}_{20}Ca$

Isotones

Isotones are atoms of different elements having the same number of neutrons.

Here are some examples isotones:

1. $^{3}_{1}H$ and $^{4}_{2}He$
2. $^{32}_{16}S$ and $^{31}_{15}P$
3. $^{14}_{6}C$ and $^{15}_{7}N$

Solved Examples

Easy

Example 1:

Which of the following atomic pairs are isotopes of each other?

A. $^{40}_{18}\mathbf{Ar}$ **and** $^{40}_{20}\mathbf{Ca}$

B. $^{40}_{18}\mathbf{Ar}$ **and** $^{36}_{18}\mathbf{Ar}$

C. $^{40}_{18}\mathbf{Ar}$ **and** $^{40}_{19}\mathbf{K}$

D. $^{40}_{16}\mathbf{S}$ **and** $^{40}_{17}\mathbf{Cl}$

Solution:

The correct answer is B.

Isotopes are those species of the same element which possess the same number of protons but different numbers of neutrons. In other words, isotopes have the same atomic number but different atomic masses. Among the given pairs, only $^{40}_{18}Ar$ and $^{36}_{18}Ar$ is an atomic pair with the same number of protons

Medium

Example 2:

$^{24}_{12}\mathbf{Mg}$, $^{25}_{12}\mathbf{Mg}$ **and** $^{26}_{12}\mathbf{Mg}$ **are the three stable isotopes of magnesium. The average atomic mass of magnesium is 24.3. The relative abundance of** $^{24}_{12}\mathbf{Mg}$ **and** $^{26}_{12}\mathbf{Mg}$ **are**

78.7% and 11.2% respectively. What percentage of magnesium existing in nature is $^{25}_{12}Mg$?

A. 10%
B. 12%
C. 14%
D. 16%

Solution:

The correct answer is A.

Let us take:

Average atomic mass of magnesium = M

Atomic mass and relative abundance of $^{24}_{12}Mg$ **=** m_1 and p

Atomic mass and relative abundance of $^{25}_{12}Mg$ **=** m_2 and q

Atomic mass and relative abundance of $^{26}_{12}Mg$ **=** m_3 and r

Now, we know that:

M = 24.3

m_1 = 24 and p = 78.7%

m_2 = 25

m_3 = 26 and r = 11.2%

The average atomic mass of magnesium can be found as:

$$M = \frac{(m_1 \times p) + (m_2 \times q) + (m_3 \times r)}{100}$$
$$\Rightarrow 24.3 = \frac{(24 \times 78.7) + (25 \times q) + (26 \times 11.2)}{100}$$
$$\Rightarrow 24.3 = \frac{1888.8 + 25q + 291.2}{100}$$
$$\Rightarrow 2430 = 2180 + 25q$$
$$\Rightarrow 25q = 2430 - 2180$$
$$\Rightarrow 25q = 250$$
$$\Rightarrow \therefore q = \frac{250}{25} = 10\%$$

Hence, 10% of the magnesium existing in nature is $^{25}_{12}Mg$.

Radioactivity

Radioactivity is the phenomenon wherein the nucleus of an unstable atom loses energy after emitting ionized particles or radiations.

In this process, the original atom transforms into a new stable atom.

In nature, there are many unstable isotopes that emit radiations. The common types of radiations are:

1. **Alpha radiation**

1. It consists of a stream of positively charged particles that are generally helium ions, having atomic mass as 4 and carrying a +2 charge.
2. It results in a decrease in atomic number by 2 and a decrease in mass number by 4.

2. **Beta radiation**

1. It consists of a stream of electrons. So, it is negatively charged.
2. It results in an increase in atomic number by 1, but mass number remains unchanged.

3. **Gamma radiation**

1. It is a photon (packet of light) with very high energy.
2. It does not result in a change in either atomic number or mass number.

Elementary Idea of Chemical Bonding

Chemical Bonding

Elements are rarely capable of free existence. In a compound, atoms of different elements are held together by bonds. The types of bonds present in a compound are largely responsible for its physical and chemical properties. The different bonds can be classified as **strong** and **weak**.

Why do elements undergo bond formation?

Elements are made of atoms, which comprise of protons, electrons, and neutrons. The protons and the neutrons reside in the nucleus and the electrons revolve around in definite paths called **orbits**. The electrons present in the last shell are called valence electrons. These electrons are responsible for all the chemical reactions of that element.

Every element has a tendency to attain a stable outer octet. To do so, it either gains or loses or shares its electrons; and in this process, it forms the bonds.

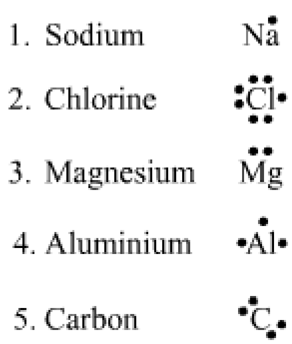
Types of strong bonds:

- Ionic or electrovalent bond
- Covalent bond
- Metallic bond

Types of weak bonds:

- Bonds formed due to van der Waal’s interaction
- Hydrogen bond

This representation of elements with valence electrons as dots around elements is referred to as **Electron Dot structures** for elements. The electron dot structure of some of the elements are:



Chemical Bonding

A chemical bond is an attractive force which holds various constituents (such as atoms, ions) together in different chemical species.

Kossel-Lewis Approach to Chemical Bonding

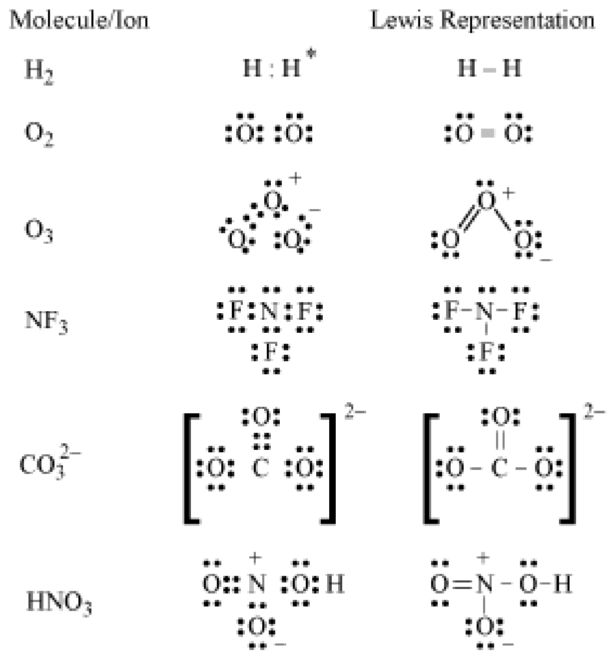
- Lewis postulated that atoms attain the stable octet when they are chemically bonded.
- **Lewis symbols**
- Notations to represent valence electrons in an atom
- Example:

- Significance of Lewis symbols − The number of dots represents the number of valence electrons.
- Octet rule- Atoms tend to gain, lose or share electrons so as to have eight electrons in their valence shells.
- **Lewis dot structure**

Representation of molecules and ions in terms of the shared pairs of electrons and the octet rule

Steps to writing Lewis dot structure:

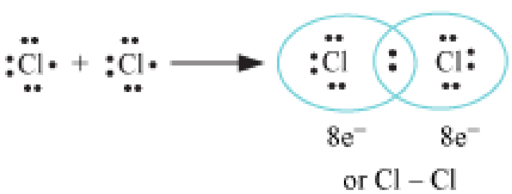
- Add the valence electrons of the combining atoms. This will give the total number of electrons required to write the structure.
- One negative charge means the addition of an electron. Similarly, one positive charge implies the removal of an electron from the total number of electrons.
- The chemical symbol of the atoms and the skeletal structure of the compound should be known. Then, distribute the total number of electrons as bonding shared pairs between the atoms in proportion to the total bonds.
- The least electronegative atom occupies the central position of the molecule/ion. For example in NF_3, nitrogen occupies the central position whereas the three fluorine atoms occupy the terminal positions.
- When the shared pairs of electrons have been accounted for single bonds, utilise the remaining electron pairs for either multiple bonding or count them as lone pairs. Here, the basic requirement is that each bonded atom gets an octet of electrons.
- Lewis representation of some molecules



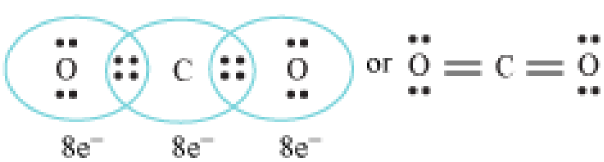
(*- Each hydrogen atom attains the electronic configuration of helium i.e. a duplet of electrons)

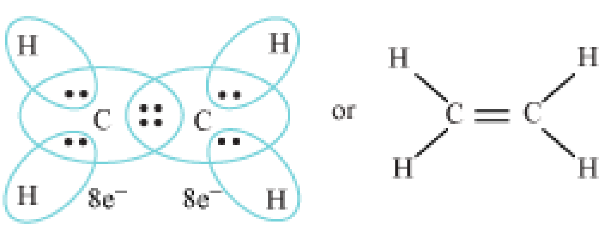
- **Covalent bond**

- Single covalent bond − Sharing of one electron pair

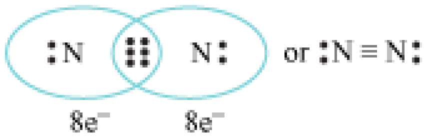


- Double bond − Sharing of two electron pairs





- Triple bond − Sharing of three electron pairs



N_2 molecule

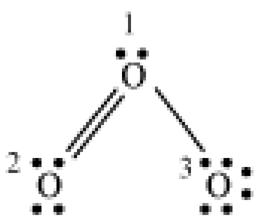
C_2H_2 molecule

- **Formal Charge**

$$\begin{bmatrix}\text{Formal charge (F.C)}\\ \text{on an atom in a}\\ \text{Lewis structure}\end{bmatrix} = \begin{bmatrix}\text{Total number of}\\ \text{valence electrons}\\ \text{in the free atom}\end{bmatrix} - \begin{bmatrix}\text{Total number of}\\ \text{nonbonding (lone}\\ \text{pair electrons)}\end{bmatrix} - \frac{1}{2}\begin{bmatrix}\text{Total number of}\\ \text{bonding (shared}\\ \text{electrons)}\end{bmatrix}$$

- Example:

Lewis structure of O_3 is



F.C. on the O-1 atom $= 6 - 2 - \frac{1}{2}(6) = +1$

F.C. on the O-2 atom $= 6 - 4 - \frac{1}{2}(4) = 0$

F.C. on the O-3 atom $= 6 - 6 - \frac{1}{2}(2) = -1$

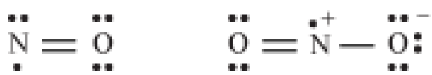
- Smaller the formal charge on the atoms, lower is the energy of the structure.
- The concept of formal charge is based on covalent bonding in which electron pairs are equally shared by neighbouring atoms.

- Limitations of the octet rule:
- Incomplete octet of the central atom

  Examples: LiCl, BeH_2, BCl_3

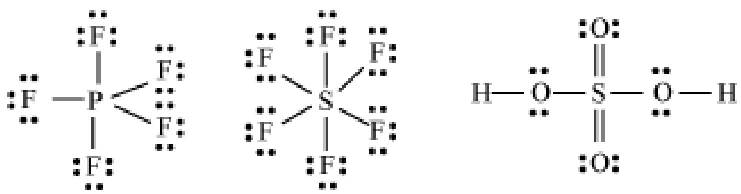
- Odd electron molecules

  Examples: NO, NO_2



- Expanded octet

  Examples: PF_5, SF_6, H_2SO_4



PF_5 — 10 electrons around the P atom

SF_6 — 12 electrons around the S atom

H_2SO_4 — 12 electrons around the S atom

Some other drawbacks of octet rule:

- It is based upon chemical inertness of noble gases. However, some noble gases can combine to form compounds such as XeF_2, KrF_2, $XeOF_2$, etc.
- It does not account for shape of molecules
- It does not explain the relative stability of molecules

Conditions for Formation of Covalent Bond

- Presence of four or more electrons in the outermost shell of an atom (exception H, Be, B and Al)
- High electronegativity of both the atoms
- High electron affinity for both the atoms
- High ionisation energy of both the atoms

- Electronegativity difference between combining atoms should be zero or very low

Formation of Some Covalently Bonded Molecules

Compound	Molecule	Type and Number of Covalent Bonds
Hydrogen (H_2)	H−H	One single bond
Chlorine (Cl_2)	Cl−Cl	One single bond
Nitrogen (N_2)	N≡N	One triple bond
Water (H_2O)	H−O−H	Two single bonds between O and H
Ammonia (NH_3)	H −N− H | H	Three single bonds between N and H
Carbon tetrachloride (CCl_4)	Cl | Cl −C− Cl | Cl	Four single bonds between C and Cl
Methane (CH_4)	H | H −C− H | H	Four single bonds between C and H

Difference between Properties of Ionic and Covalent Compounds

Ionic Compounds	Covalent Compounds
The constituent particles are ions.	The constituent particles are molecules.
They exist as hard solids.	They exist as gases, liquids or soft solids.
They have high melting and boiling points	They have low melting and boiling points.

They are good conductors of electricity in the aqueous or molten state.	They do not conduct electricity.
They ionise in solution and behave as electrolytes.	Only polar compounds form ions in aqueous solutions.
They undergo dissociation.	They do not undergo dissociation.
They are soluble in water.	They are soluble only in organic solvents.
They undergo fast chemical reactions.	They undergo slow chemical reactions.

Formation of Ionic Compounds and Their Properties

We know that common salt is an important dietary mineral essential for animal life. Common salt is chemically known as sodium chloride. The chemical formula of sodium chloride is NaCl. It suggests that it is made up of sodium, which is a reactive metal, and chlorine, which is a non-metal.

Do you know that sodium chloride does not exist as molecules, but aggregates as oppositely charged ions?

An ion is a charged species, which can be negatively charged or positively charged. A negatively charged species is called an 'anion' and a positively charged species is called a 'cation'.

Sodium chloride (NaCl) is formed by the combination of sodium (Na^+) and chloride (Cl^-) ions. Sodium and chloride ions are oppositely charged. Hence, they are held by a strong electrostatic force of attraction in sodium chloride compound. **But why do they react or combine with each other?** This can be explained by considering the formation of sodium chloride.

This representation of elements with valence electrons as dots around elements is referred to as **Electron Dot structures** for elements.

Do you know what type of a compound sodium chloride is? Sodium chloride is an ionic compound.

Ionic compounds:

These are compounds that are formed by the transfer of electrons. In other words, these are compounds that are made up of ions.

The bonding in such compounds is called **ionic bonding or electrovalent bonding**. This type of bonding is also known as **electrostatic bonding** as the forces that hold the ions together are electrostatic in nature. The transfer of electrons always takes place from a metal to a non-metal. Thus, metals and non-metals combine with each other to attain a noble gas configuration.

We know that inert (noble) gases are very stable and almost unreactive. This is because of their stable electronic configuration in which their valence shell is complete. Hence, they do not take part in the formation of ionic compounds. The given table lists some elements with their electronic configurations.

S.No.	**Type of element**	**Element**		**Symbol**	**Atomic number**	**Electronic configuration** **K L M N**	**Number of valence electrons**
1.	**Noble gases**	1.	Helium	He	2	2	2
		2.	Neon	Ne	10	2, 8	8
		3.	Argon	Ar	18	2, 8, 8	8
2.	**Metals**	1.	Sodium	Na	11	2, 8, 1	1
		2.	Potassium	K	19	2, 8, 8, 1	1
		3.	Magnesium	Mg	12	2, 8, 2	2
		4.	Calcium	Ca	20	2, 8, 8, 2	2
		5.	Aluminium	Al	13	2, 8, 3	3

3.	**Non-metals**	1.	Nitrogen	N	7	2, 5	5
		2.	Phosphorus	P	15	2, 8, 5	5
		3.	Oxygen	O	8	2, 6	6
		4.	Sulphur	S	16	2, 8, 6	6
		5.	Fluorine	F	9	2, 7	7
		6.	Chlorine	Cl	17	2, 8, 7	7

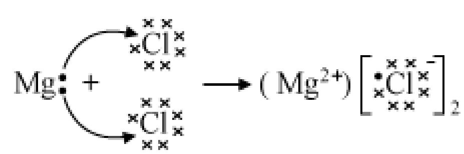
Let us now see the formation of magnesium chloride, which is also an ionic compound.

The atomic number of magnesium is 12. Thus, its electronic configuration is 2, 8, 2. Since it contains two more electrons than a stable noble gas configuration, it loses these two electrons to form Mg^{2+}. On the other hand, the atomic number of chlorine is 17. Thus, its electronic configuration is 2, 8, 7. It requires one more electron to complete its octet. For this, two chlorine atoms accept two electrons that were lost by Mg atom to form two chloride (Cl^-) ions. The chemical equations involved in the process are given below:

$$\begin{array}{ccccc} Mg & \rightarrow & Mg^{2+} & + & 2e^- \\ \text{Magnesium} & & \text{Magnesium ion} & & \text{Electron} \\ (2,8,2) & & (2,8) & & \end{array}$$

$$\begin{array}{ccccc} 2Cl & + & 2e^- & \rightarrow & 2Cl^- \\ \text{Chlorine} & & \text{Electron} & & \text{Chloride ion} \\ (2,8,7) & & & & (2,8,8) \end{array}$$

The reaction between magnesium and chlorine can be represented as follows:



On the similar basis, the formation of sodium chloride (NaCl) and calcium oxide (CaO) is depicted in the table below:

Compound	Formation
Sodium chloride (NaCl)	$\underset{(2,8,1)}{Na} \longrightarrow \underset{(2,8)}{Na^+} + e^-$ $\underset{(2,8,7)}{Cl} + e^- \longrightarrow \underset{(2,8,8)}{Cl^-}$ $Na^+ + Cl^- \longrightarrow [Na^+][Cl^-]$
Calcium oxide (CaO)	$\underset{(2,8,8,2)}{Ca} \longrightarrow \underset{(2,8,8)}{Ca^{2+}} + 2\,e^-$ $\underset{(2,6)}{O} + 2\,e^- \longrightarrow \underset{(2,8)}{O^{2-}}$ $Ca^{2+} + O^{2-} \longrightarrow [Ca^{2+}][O^{2-}]$

Potassium oxide (K_2O) is also an ionic compound. It is made of two potassium atoms and one oxygen atom.

Can you draw the Electron Dot structure of potassium and oxygen atoms? Can you show the formation of potassium oxide?

Let us now try to find out the properties of ionic compounds by performing the following activities.

1) Take samples of sodium chloride, potassium iodide, and barium chloride and observe their physical state.

2) After that, take a small amount of a sample on a metal spatula and heat it directly on a flame. Observe what happens to the sample.

3) Now, try to dissolve each sample in water, petrol, and kerosene and observe the solubility of compounds.

4) Now, take a container and fill it with distilled water. Take two electrodes and place them in water. Then, connect the electrodes to a bulb and a battery through electric wires (as shown in **figure 1)**. When the switch is closed, the bulb will not glow as distilled water does not conduct electricity. Now, instead of distilled water, take a solution of an ionic compound and observe.

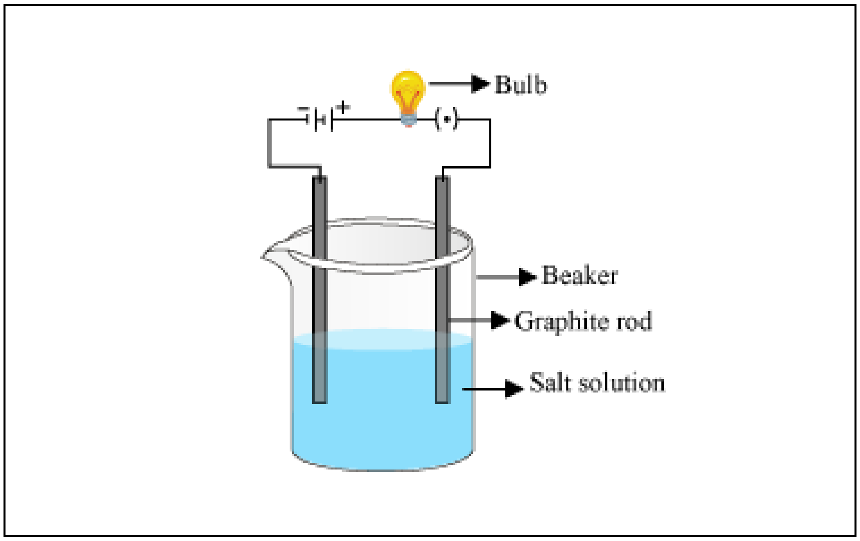


Figure 1: Conductivity of salt

We will observe that

- all compounds are solids
- all have high melting and boiling points
- all samples are soluble in water but insoluble in kerosene and petrol
- the solution of all samples can conduct electricity

When the switch is closed, the bulb starts glowing. This shows that solutions of ionic compounds conduct electricity.

Hence, we can summarize the properties of ionic compounds as follows:

Ionic compounds are hard and brittle crystalline solids: The electrostatic force holding the ions present in ionic compounds are very strong. Therefore, these compounds are quite hard, as they are made up of small crystals.

Ionic compounds have high melting and boiling points: A lot of energy is required to overcome the strong electrostatic force of attraction, which holds the ions present in ionic compounds together. Thus, these compounds have high melting and boiling points.

Table 1: Melting and boiling points of some ionic compounds

Ionic compound	Melting point (K)	Boiling point (K)
NaCl	1074	1686
LiCl	887	1600

$CaCl_2$	1045	1900
CaO	2850	3120
$MgCl_2$	981	1685

Ionic compounds dissolve only in polar solvents: Ionic compounds are polar in nature due to the presence of opposite charges. Therefore, these compounds dissolve only in polar solvents such as water. These compounds are insoluble in organic solvents such as kerosene, alcohol, and petrol.

Ionic compounds conduct electricity in a solution or molten state: Ionic compounds consist of small ions, which can conduct electricity.

Covalent and Coordinate Bond

We know that a majority of substances used by us daily, from paper and plastics to coal and petrol, are all made up of carbon. Food grains, pulses, medicines, cotton, synthetic fibres, wood, etc. are all made up of carbon.

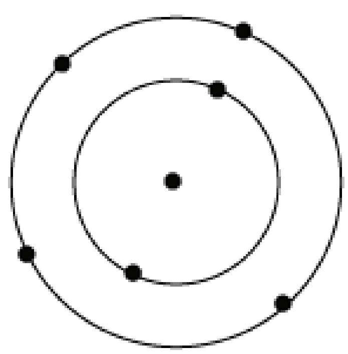
Carbon is also a major part of all living things. In air, it is present as carbon dioxide and comprises around 0.03% of the total atmosphere.

Let us study about carbon and its bonding in its compound in more detail.

Carbon is a non-metal having the symbol '**C**' and atomic number **six**. Since the atomic number of carbon is six, its electronic configuration is 2, 4.

This means that carbon contains two electrons in K shell and 4 electrons in L shell (outermost shell). Hence, it has four electrons in its valence shell.

Since carbon has four electrons in its valence shell, it requires four more electrons to complete its octet. Therefore, it is a tetravalent element.



In order to complete its octet i.e., to attain its noble gas configuration and to stabilise itself, carbon can:

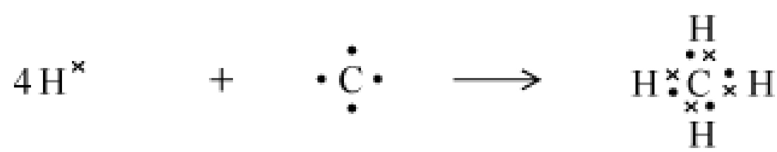
- Either lose four electrons to form $\mathbf{C^{4+}}$ or gain four electrons to form $\mathbf{C^{4-}}$. This, however, requires a lot of energy and would make the system unstable.

- Therefore, carbon completes its octet by sharing its four electrons with the other carbon atoms or with atoms of other elements.

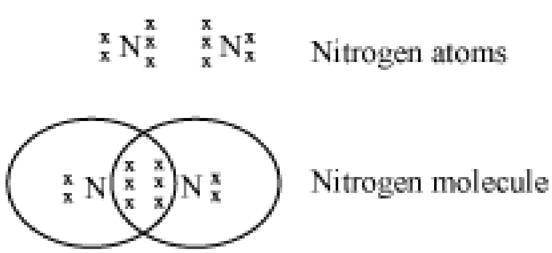
The bonds that are formed by sharing electrons are known as covalent bonds. Covalently bonded molecules have strong intermolecular forces, but intramolecular forces are weak.

Carbon has four valence electrons and requires four more electrons to complete its octet. Therefore, it is capable of bonding with four other atoms of carbon or atoms of other elements having a valency of 1.

For example, the simplest molecule, methane, can be formed with hydrogen (H) atoms that have only one electron in its K shell. To attain the noble gas configuration, it combines with four hydrogen atoms as shown in the figure.



Nitrogen has an atomic number of 7. In order to attain an octet, each nitrogen atom in a molecule of nitrogen contributes three electrons, thereby giving rise to forming three shared pairs of electrons. This is said to constitute a triple bond between the two atoms. The electron dot structure of N_2 and its triple bond can be depicted as follows.:



Now, let us study the properties of carbon covalent compounds.

- Covalent bonds are made by atoms by sharing their electrons. Formation of ions does not take place in this process.

In addition, these compounds do not have any extra electrons. Hence, covalent compounds are non-conductors of electricity.

- As all organic compounds contain covalent bonds, they also have low melting and boiling points. This becomes evident from the following data.

Compound	Melting point (K)	Boiling point (K)
Acetic acid	290	391
Chloroform	209	334
Ethanol	156	351
Methane	90	111

Also, from the above data, it can be inferred that the forces of attraction between the carbon molecules in carbon compounds is not very strong.

- Because of their low melting and boiling points, these compounds mostly exist as liquids or gases at room temperature.

A covalent bond formed between two different atoms, with different electronegativities is known as **polar covalent bond.**

For example when a Covalent bond is formed between H and Cl, it is polar in nature because Cl is more electronegative than H atom.

Therefore, electron cloud is shifted towards Cl atom. As a result, a partial negative charge appears on Cl atom and, an equal positive charge on H atom.

$\mathbf{H^{\delta+}{----}Cl^{\delta-}}$

A covalent bond formed between two like atoms, is known as **Non-polar bond**. Since difference of electronegativity is zero, therefore, both atoms attract electron pair equally and no charge appears on any atom, and the whole molecule becomes neutral.
For e.g. **H------H**

Non-polar covalent bond:

Let us now compare the properties of electrovalent compounds with those of covalent compounds.

Electrovalent compounds	Covalent compounds

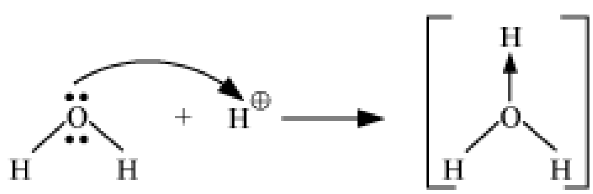
These types of compounds are formed when the atoms containing one to three valence electrons donate electrons. On the other hand, the atoms containing five to seven electrons accept thosethe donated electrons. Thus, both the participating atoms acquire the nearest noble gas configuration.	These types of compounds are formed when the participating atoms mutually contribute one electron each to form an electron pair which that is shared by them in such a manner that both of them attain the nearest noble gas configuration.
They are usually crystalline solids.	They are usually gases or liquids or soft solids.
They are good conductors of electricity.	They are poor conductors of electricity.
They usually have high melting and boiling points.	They usually have low melting and boiling points.
They are soluble in polar solvents but insoluble in organic solvents.	They are soluble in organic solvents but insoluble in polar solvents.

Coordinate Bond

It is formed when the shared pair of electrons is provided by one of the two atoms and shared by both. Examples of coordinate-bonded compounds are:

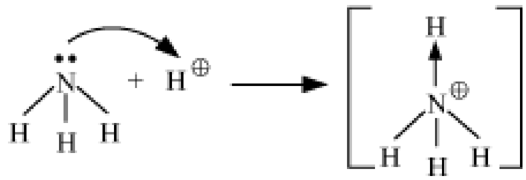
- Hydronium ion
- Ammonium ion

Hydronium ion: The compound giving hydronium ion is water. One lone pair on the oxygen atom of water molecule is shared with the hydrogen ion. In this combination, the hydrogen ion accepts a pair of electrons and distributes the charge over the entire hydronium radical.



Coordinate bond formation in hydronium ion

Ammonium ion: One lone pair on the nitrogen atom of ammonia molecule is shared with the hydrogen ion. In this combination, the hydrogen ion accepts a pair of electrons and distributes the charge across the entire ammonium radical.



Coordinate bond formation in ammonium ion

Conditions for formation of coordinate bond

- Presence of at least one lone pair of electrons on any of the two atoms. This atom acts like a donor.
- Shortage of a lone pair of electron on the second atom. This atom acts like an acceptor.

Formation of Hydronium (H_3O^+) Ion

- Based on ionisation of water
- One water molecule contains two hydrogen atoms covalently bonded with one oxygen atom.
- Oxygen atom has two lone pairs of electrons for donation.

- Addition of an acid to water, results in dissociation of the water molecules as shown below:
  - $H_2O \xrightarrow{HA} H^+ + OH^-$

- H^+ ions accepts an electron pair from oxygen atom of another water molecule forming coodinately bonded hydronium ion:

  - $H^+ + H-\ddot{\underset{\cdot\cdot}{O}}-H \longrightarrow \left[\begin{array}{ccccc} H & - & \ddot{O} & - & H \\ & & \downarrow & & \\ & & H & & \end{array}\right]^+$

- Hydronium ion is also referred to as hydrated proton.
- Similarly, water molecules undergo self-ionisation where one molecule acts as an acid:

  - $H_2O \xrightarrow{HA} H^+ + OH^-$