Ratna Sagar

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9

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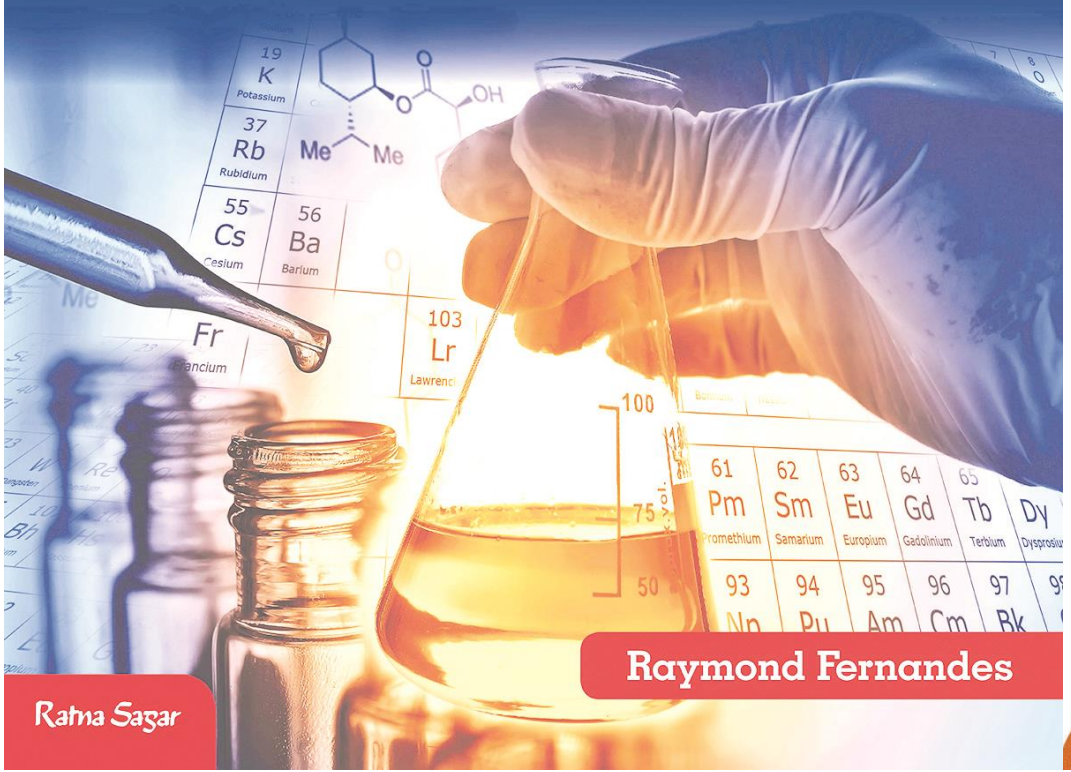
Living Science CHEMISTRY



Raymond Fernandes

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ICSE
Living Science
Chemistry

Class 9

Chapter 4 Atomic Structure and Chemical Bonding

LEARNING OBJECTIVES

Dalton's Atomic Theory

- Atom and subatomic particles

Discovery of Electron

- Characteristics of electrons

Discovery of Proton

- Characteristics of protons
- Structure of the atom

Discovery of the Atomic Nucleus

- Bohr's atomic model

The Neutron

- Bohr's atomic model
- Present concept of the atom

Atomic Number and Mass Number

- Bohr-Bury Scheme
- Valence electrons
- Inertness of noble gases
- Isotopes and Isobars

Chemical Bonding

Ionic or Electrovalent Bond

- Formation of ionic compounds

Covalent Bonds

- Some covalent molecules and their structures

Dalton's Atomic Theory

The postulates of Dalton's atomic theory are:

1. All matter is made of tiny particles called **atoms**.

2. Atoms are indivisible particles. They cannot be created or destroyed in a chemical reaction.

3. All atoms of a given element are identical in mass and chemical properties.

4. Atoms of one element differ in mass and chemical properties from those of other elements.

5. Compounds are formed when two or more different kinds of atoms combine in definite proportions.

6. The relative number and kinds of atoms are constant in a given compound.

Atom and subatomic particles

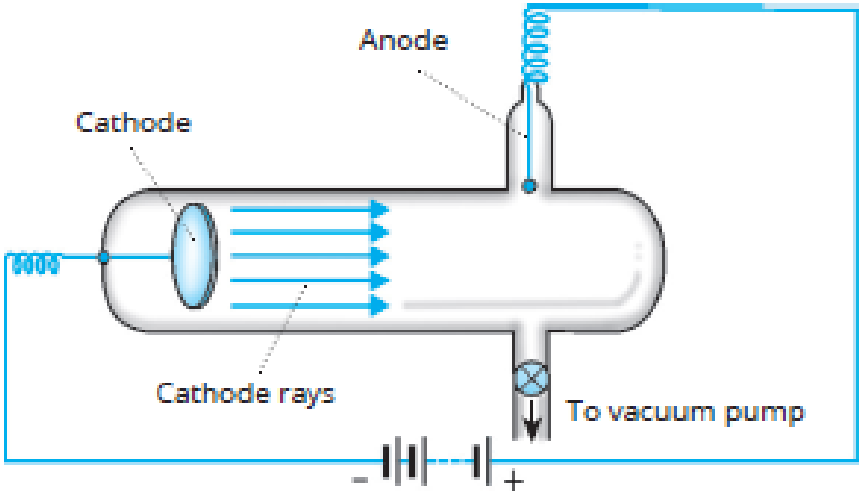
An **atom** is the smallest particle of an element which can take part in a chemical reaction. Atoms are building blocks of all matter.
Under normal conditions, the atoms of elements except those of noble gases like helium and argon are not capable of independent existence

The subdivision of the atom led to the introduction of subatomic particles namely electrons, protons and neutrons. These particles are called **elementary fundamental particles** because these are essential constituents of atoms.

Discovery of Electron

In 1878, cathode rays were discovered by **William Crookes**. He found that when an electric discharge is passed through a tube containing a small quantity of gas at very low pressure, rays are emitted from the negative plate called the cathode. Since the rays originated from the cathode they were called **cathode rays**.

In 1897, Sir **J J Thomson** found that these cathode rays consist of negatively charged particles, which were named as **electrons**. Electrons are essential constituents of all matters.



Emission of cathode rays in the discharge tube opposite to the cathode

Characteristics of electrons

1. The charge on an individual electron is -1.602×10^{-19} coulomb. Thus, electron has 1 unit of negative charge (–1). Thus, the relative charge of an electron is –1.

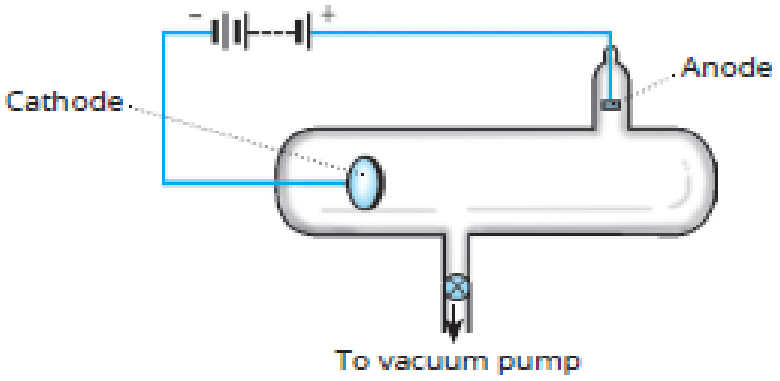
2. The mass of an individual electron is 9.107×10^{-28} g. This weight of an electron expressed in atomic mass unit is only 0.0005486 u which is about 1/1837 of the weight of a hydrogen atom.

3. Electrons have a particle nature and travel in a straight line. This was deduced because a stream of electrons could cast a shadow if an object was placed in their path.

4. Electrons deflect towards a positive field and away from a negative field.

An electron is defined as a subatomic fundamental particle having mass equal to 1/1837 of that of a hydrogen atom and carrying one unit of negative charge.

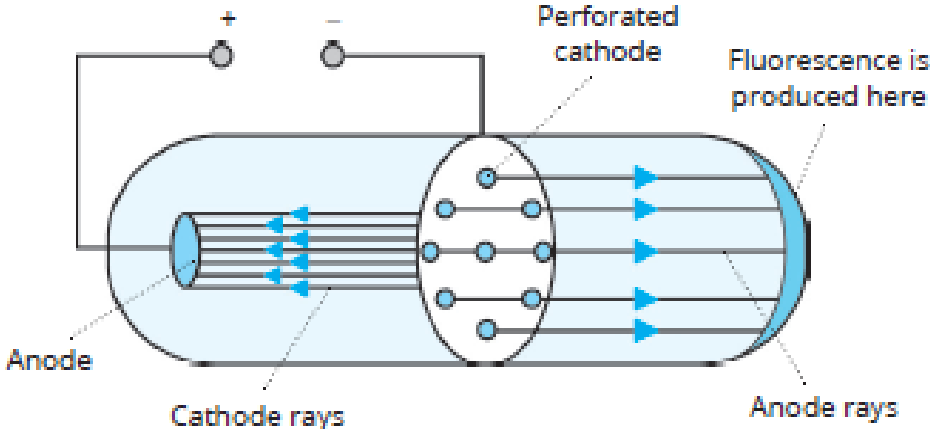
As electron has charge –1 and negligible mass, it is represented as $^{-1}_{0}e$.



Emission of pink glow by air inside the discharge tube

Discovery of Proton

If negatively charged particles are present in an atom, positively charged particles must also be present, because all atoms are electrically neutral. The search for such a type of particle led to the discovery of the subatomic positive particle called the **proton**.



Goldstein's discharge tube experiment producing anode rays

Characteristics of protons

1. The charge on an individual proton is $+1.602 \times 10^{-19}$ coulomb. Therefore, the relative charge of a proton is +1.

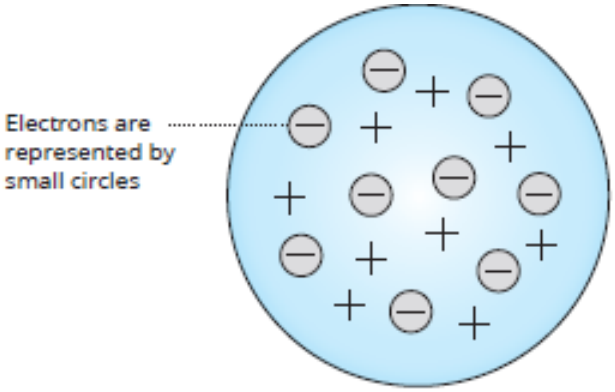
2. The mass of an individual proton is 1.672×10^{-24} g. This mass of a proton expressed in atomic mass unit is 1.0073 u which is nearly the mass of a hydrogen atom.

3. Protons deflect towards a negative field and away from a positive field.

The protons were believed to be situated in the central region of the atom called the **nucleus**. A proton is defined as a subatomic particle having mass equal to that of a hydrogen atom (1 u) and carrying one unit of positive charge.

Discovery of Proton

J J Thomson's model of an atom: In 1898 J J Thomson proposed a model for the structure of an atom which is called Thomson's **plum-pudding model** of atom. Thomson proposed the following:



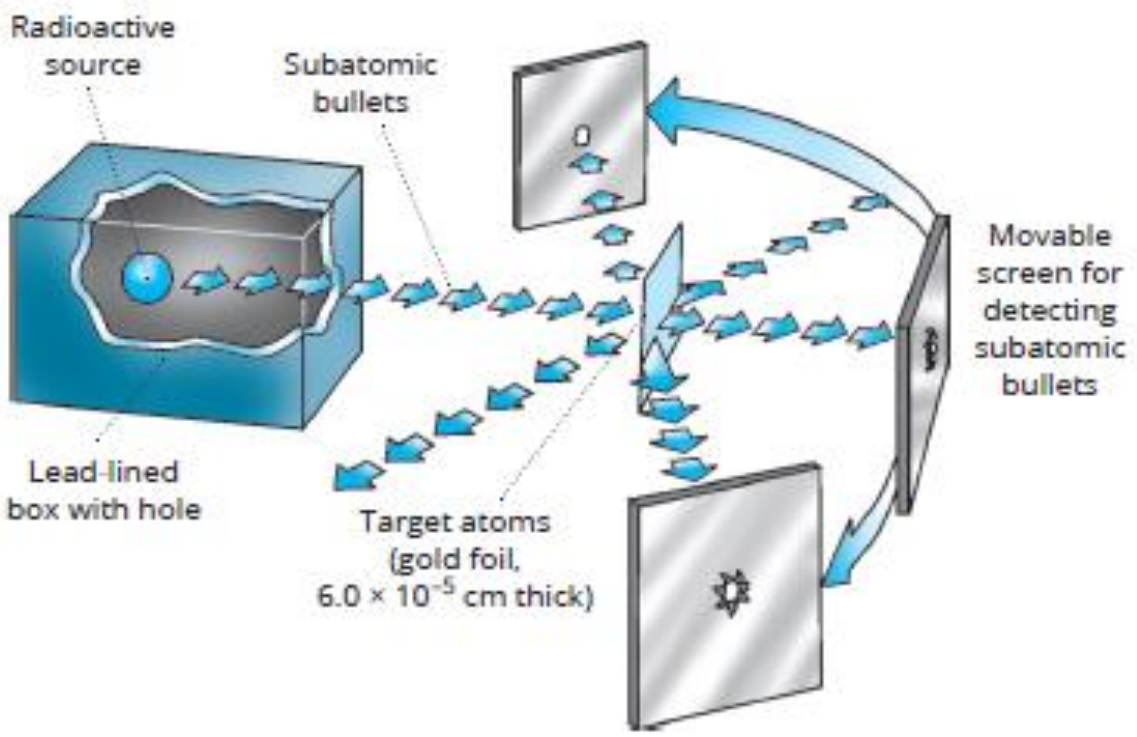
1. An atom is considered to be a sphere of uniform positive charge into which the negatively charged electrons are embedded.
2. The negative and positive charges are equal in magnitude. Thus, the atom as a whole is electrically neutral.
3. The mass of the atom is considered to be evenly spread over the atom.

Discovery of the Atomic Nucleus

In 1911, **Lord Rutherford** performed a unique experiment on the scattering of α-particles, which are positively charged helium atoms, projected against sheets of metals. He used a piece of radium as the source of α -particles.

This experiment is now called **Rutherford's alpha-particle scattering experiment**. In this experiment, fast moving a-particles were allowed to fall on a thin gold foil. The following observations were made from this experiment:

Rutherford's gold foil experiment



1. Most of the a-particles passed through the foil without any deflection.

2. A small fraction were deflected through small angles.

3. Only a very few a-particles (1 in 12000) suffered a deflection of 180° .

From this experiment, it was concluded that

1. Very little space in an atom is occupied by matter.

2. Most of the matter in an atom is located at the centre.

3. At this centre positive charge is concentrated.

4. Only such an atom could account for the behaviour of α -particles striking a metal sheet. This line of reasoning led Rutherford to suggest the **nuclear model** of the atom.

Rutherford's findings

1. Most of the matter in an atom is located at the centre in a very little space called **nucleus**. The volume occupied by the nucleus is very small as compared to the total volume of the atom.

2. The nucleus of an atom is associated with all the positive charges of the atom and hence contains all the protons.

3. Since the atom is neutral as a whole, an equal number of electrons must surround the nucleus to counterbalance the number of protons in the nucleus.

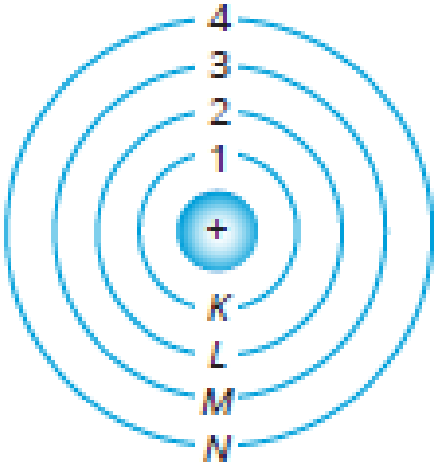
4. In order to confer stability on the nuclear model of the atom, electrons were supposed to be revolving around the nucleus.

Bohr's model of atom

Niels Bohr, in 1913, gave the following postulates about the model of an atom.

1. In an atom, the electrons revolve around the nucleus in certain circular paths called **orbits**.

2. Each of these orbits is associated with certain amount of energy. These orbits are known as **energy levels** or **energy shells**.



Energy shells

3. As long as an electron revolves in a particular orbit, it neither loses nor absorbs energy.

4. When energy is initially supplied to the atom, an electron moves to an orbit of higher energy. When this electron drops back to the original orbit, the absorbed energy is radiated by the atom.

Neutron

Neutron is a subatomic fundamental particle having mass equal to that of a hydrogen atom (1 u) and carrying zero charge. It is represented as ${}_0^1n$.

Characteristics of neutrons

1. A neutron has no charge. It is electrically neutral.

2. The mass of an individual neutron is 1.672×10^{-24} g. This mass of a neutron expressed in atomic mass unit is 1.0073 u.

3. Neutrons do not get deflected under the influence of an electric or magnetic field.

Present concept of the atom

The atom is now known to be composed of two parts:

Nucleus

The nucleus contains all the protons and neutrons of the atom. They are also known as **nucleons**.

Extranuclear region

The extranuclear region of an atom consists of electrons revolving around the nucleus in definite orbits. These orbits are called **shells** and are labelled as *K*, *L*, *M*, etc., starting from the innermost shell that is closest to the nucleus.

Note: **Refer to Table 4.1 for Comparison among proton, neutron and electron**

Atomic number of an element (Z)

= Number of protons in the nucleus of its atom
= Number of positive charges on the nucleus of its atom
= Number of electrons in the neutral atom of the element

Mass Number (*A*)

The sum total of the neutrons and protons inside the nucleus of an atom is called the **mass number**.

Bohr-Bury Scheme

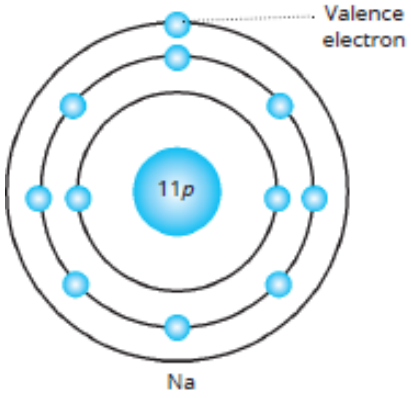
The distribution of electrons in different energy shells (orbits) of an atom is governed by Bohr–Bury Scheme. According to this scheme:

1. The maximum number of electrons which can be accommodated in an energy shell (orbit) is given by $2n^2$, where n is the number of the shell (or orbit).

2. The outermost shell of an atom cannot accommodate more than 8 atoms, even it has the capacity to accommodate more electrons. If, however the outermost shell is *K* shell or first shell, then it can hold a maximum of 2 electrons only.

3. Electrons in an atom do not occupy a new shell or orbit unless the inner shells are filled.

Valence electrons



The electrons present in the outermost shell of an atom are called **valence electrons.** For example, the electronic configuration of sodium atom is 2, 8, 1.

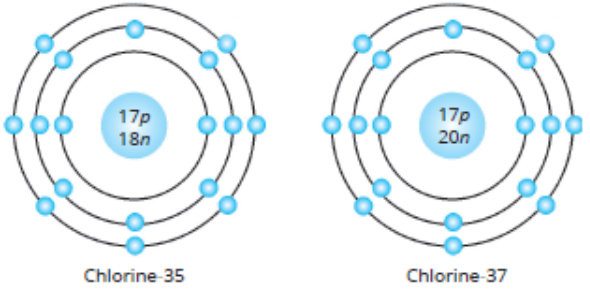
Inertness of noble gases

Atoms having 8 electrons or octet of electrons in their outermost shell are chemically inert or unreactive. When the atom has just one shell, i.e. *K* shell, then 2 electrons or duplet of electrons is considered a stable arrangement of electrons. Helium is the only noble gas having duplet of electrons in the outermost shell.

Note: Refer to Table 4.5 for Electronic configuration of noble gases

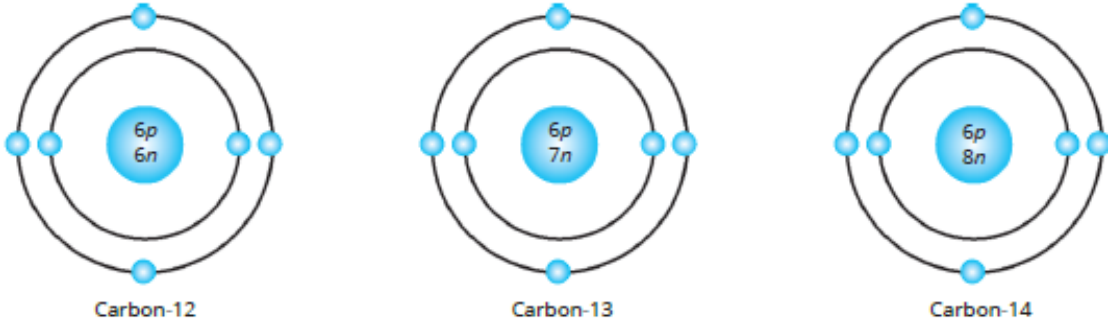
Isotopes

Atoms of the same element which have same number of protons but different number of neutrons inside their nuclei are called **isotopes**. The different mass numbers are due to the atoms having different number of neutrons in the nucleus, number of protons remaining constant for atoms of a particular element.



Isotopes of Chlorine

Isotopes of Carbon



Characteristics of isotopes

The isotopes of an element have the following characteristics:

1. The isotopes of an element have the same number of protons and electrons but different number of neutrons.
2. The isotopes of an element have different mass numbers and hence, different masses.
3. The isotopes of an element possess the same electronic configuration. Therefore, they have the same number of valence electrons and exhibit the same chemical properties.
4. The isotopes of an element have different physical properties such as density, melting point, boiling point, etc.

Applications of isotopes

Isotopes find important applications in a variety of areas such as in the field of medicine, radiocarbon dating, tracer technique, production of cheaper electricity, industry, agriculture, discovery of new elements and new isotopes.

Isobars

Elements having the same atomic mass numbers but different atomic numbers are called **isobars**. Potassium and argon are isobars.
Isobars have different physical and chemical properties

Chemical Bonding

In a compound, atoms of different elements are held together by bonds. These are chemical bonds. The types of bonds that are present in a compound are largely responsible for its physical and chemical properties. A chemical bond is said to form between two atoms when they are held together in a molecule.

Chemical bonds are of three types:

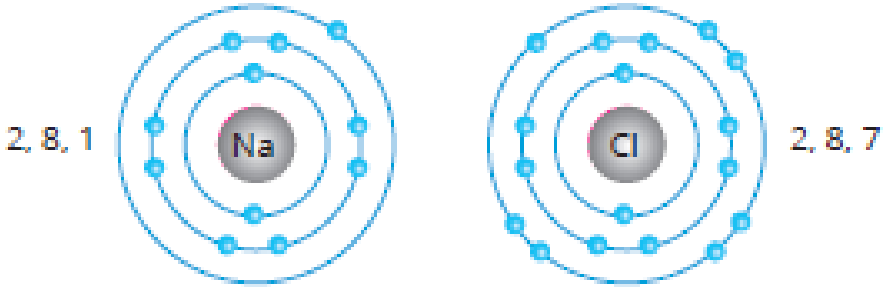
a. Ionic or electrovalent bond **b.** Covalent bond **c.** Coordinate or dative bond

Ionic or electrovalent bond

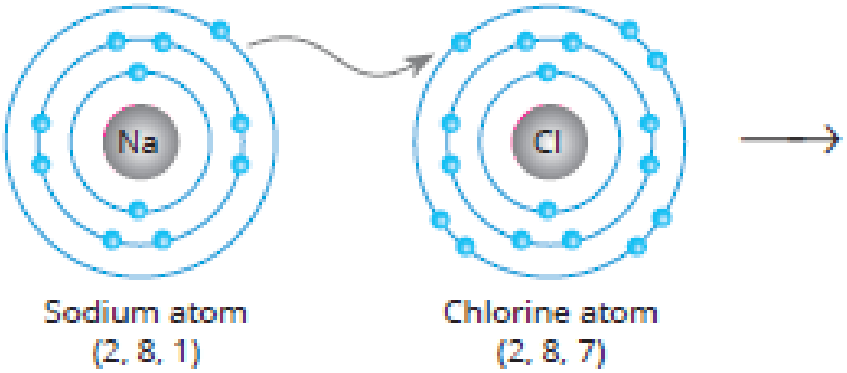
The chemical bond formed between two atoms by transfer of electrons from one atom to another is called **ionic** or **electrovalent bond**. The number of electrons lost or gained by the atoms is called its **electrovalency**. The compounds thus formed are called **ionic compounds** or **electrovalent compounds**.

Example: When sodium reacts with chlorine, it transfers an electron to chlorine atom. By losing one electron, sodium atom forms sodium ion (Na^+) and chlorine atom forms chloride ion (Cl^-). Due to the opposite charges generated on the initially neutral atoms they are attracted to each other by strong electrostatic forces forming an **ionic** or **electrovalent bond**.

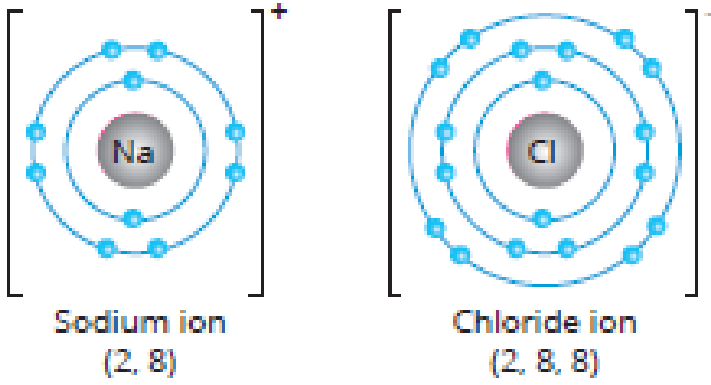
Electronic configurations of atoms of sodium and chlorine



Ionic bond formation in sodium chloride

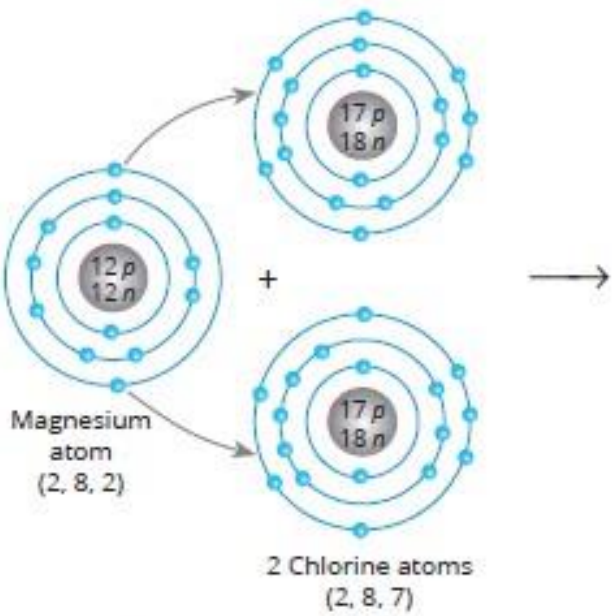


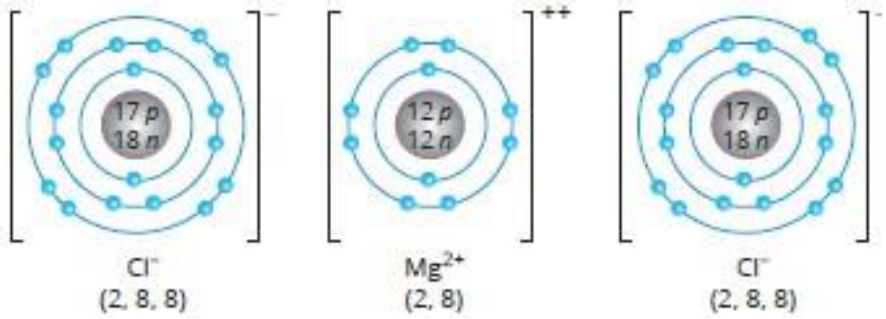
Thus, 1 atom of sodium combines with 1 atom of chlorine. Therefore, formula of sodium chloride is NaCl.



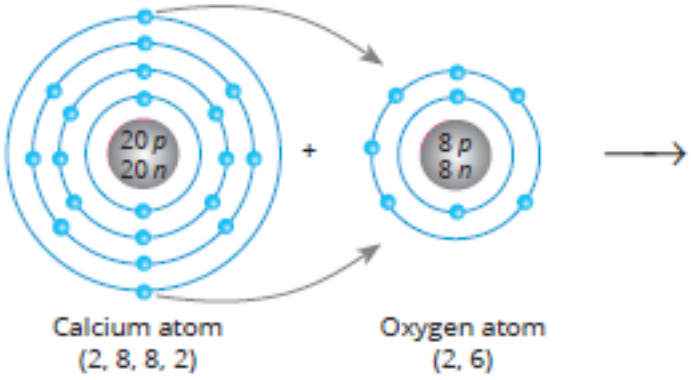
Formation of some Ionic compounds

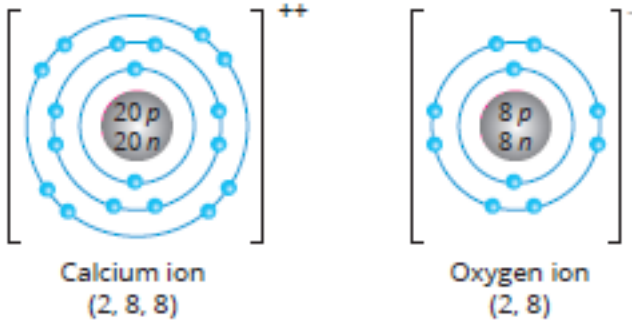
Formation of magnesium chloride ($MgCl_2$)





Formation of calcium oxide





Covalent bond

The chemical bond formed due to mutual sharing of two or more electrons between two atoms is called a **covalent bond**.

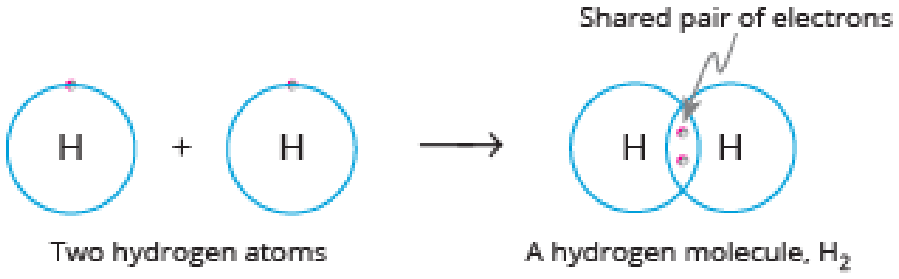
A covalent bond formed by sharing of one pair of electrons is called **single covalent bond**. It is denoted by a single dash (—) between the two atoms. For example, the two hydrogen atoms in H_2 molecule.

The covalent bond formed by sharing of two pairs of electrons between the two atoms is called **double covalent bond**. It is denoted by a set of two dashes . For example, oxygen molecule O_2 where the two oxygen atoms share two pairs of electrons.

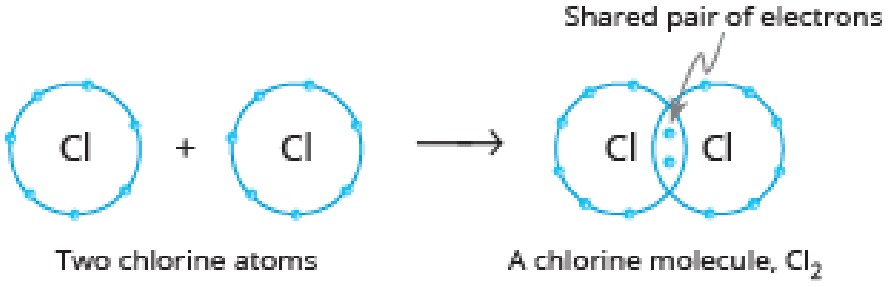
The **triple covalent bond** is formed by sharing of three pairs of electrons. It is denoted by a set of three dashes . For example, two nitrogen atoms share three pairs of electrons to form a molecule of nitrogen.

Formation of hydrogen molecule, H_2

A pair of electrons in the overlapped region constitutes a bond and therefore the hydrogen molecule has a **single bond**.

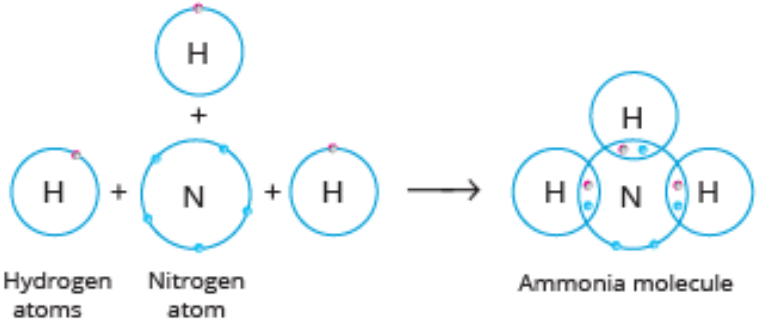


Formation of chlorine molecule, Cl_2



Formation of nitrogen molecule, N_2

Formation of ammonia molecule, NH_3



Note: Refer to P-55-57 of the book for Some covalent molecules and their structures

SUMMARY...

1. Dalton's atomic theory states that all matter is composed of tiny, indivisible and indestructible particles called atoms that are identical in mass and in chemical properties.

2. The modern atomic theory differs from the conventional Dalton's theory because it states that atoms can be further subdivided into electrons, protons and neutrons.

3. An atom is the smallest particle of an element which can take part in a chemical reaction.

4. An electron is defined as a subatomic fundamental particle having mass equal to 1/1837 of that of a hydrogen atom and carrying one unit of negative charge. It is represented as $_{-1}{}^{0}e$.

5. A proton is defined as a subatomic particle having mass equal to that of a hydrogen atom and carrying one unit of positive charge. It is represented by the symbol $_{1}{}^{1}p$.

6. An atom consists of two regions namely the nucleus and the extranuclear region.

7. The nucleus is a positively charged core consisting of protons and neutrons.

8. The extranuclear region consists of electrons revolving around the nucleus in different but specific orbits.

9. Neutron is a subatomic fundamental particle having mass equal to that of a hydrogen atom and carrying zero charge. It is represented as $_0{}^1n$.

10. The sum total of the protons and neutrons present in the nucleus of an atom is called the atomic mass number. It is represented by '*A*'.

11. The number of electrons revolving around the nucleus of an atom or the number of protons in the nucleus of an atom is called the atomic number and is represented by '*Z* '

12. Atoms take part in chemical reactions because they want to complete their outer octet by either sharing or transferring electrons.

13. Isotopes are atoms of the same elements that have the same atomic number but different atomic mass numbers because they have different number of neutrons in their nuclei.

14. Isobars are elements having the same atomic mass numbers but different atomic numbers.

15. The chemical bond formed between two atoms by transfer of electrons from one atom to another is called ionic or electrovalent bond.

16. The chemical bond formed due to mutual sharing of two or more electrons between two atoms is called a covalent bond.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