Atomic Structure and Chemical Bonding

Model of an Atom

Dalton's Atomic Theory

John Dalton (1808) presented his Dalton's atomic theory.

Comparison of Dalton's Atomic Theory with the Modern Atomic Theory

Dalton's Atomic Theory	Modern Atomic Theory
1. Atoms are indivisible particles. 2. Atoms can neither be created nor destroyed. 3. The atoms of an element are alike in all respects and differ from atoms of other elements.	1. Atoms are divisible into sub-atomic particles such as protons, neutrons and electrons. 2. Atoms can be created and destroyed by nuclear fusion and fission. 3. The atoms of an element may not be alike in all respects, as seen in the case of isotopes. Isotopes are atoms of the same element with the same atomic number but different mass numbers.

Discovery of Electrons

J. J. Thomson studied the characteristics and the constituents of cathode rays.

Properties of Cathode Rays

1. On application of an electric field to the cathode rays deflected towards the positively charged plate. This indicated that cathode rays must be **negatively charged**.
2. When a small wheel made of a light material was kept in the path of cathode rays, it rotated. This confirmed the particulate nature of cathode rays.
3. Thomson found that the charge to mass ratio of electrons denoted as **e/m** is **1.76×10^{11} C/kg**.

Conclusion: From his experiment, Thomson arrived at the conclusion that

1. Cathode rays are nothing but a stream of negatively charged particles called **electrons**.
2. These negatively charged particles are an integral part of all atoms.
3. Electrons have both definite mass and definite electric charge, both of which are independent of the nature of the gas in the discharge tube.

Discovery of Protons

A German scientist, E. Goldstein modified the discharge tube and passed an electric current through it.

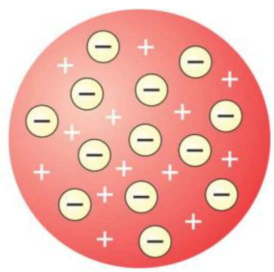
He found that positively charged rays were emitted from the anode in the discharge tube. These rays were called **canal rays.**

Properties of Anode Rays

1. They are made up of positively charged particles.
2. These positive rays are deflected by electric and magnetic fields but in a direction opposite to the cathode rays. Thus, these rays consist of positively charged particles called **protons**.
3. Charge to mass ratio e/m differs from gas to gas. Its value is maximum when hydrogen is taken in the discharge tube, and its value is much less than that of an electron.

Thomson's Theory of an Atom

J. J. Thomson proposed a model of an atom to be similar to a plum pudding. He proposed that an atom consists of a positively charged sphere where the electrons are embedded in it. The atom as a whole is electrically neutral.



Thomson's model of an Atom

Rutherford's Model of an Atom

Rutherford (1911) overturned Thomson's atomic model by his gold foil experiment. His experiment demonstrated that the atom has a tiny massive nucleus. From the experiment, he concluded that

- As most of the alpha particles passed through the gold foil without getting deflected, most of the space inside the atom is empty.
- Very few particles deflected from their path; this indicated that the positive charge of the atom occupies very little space.
- A small fraction of alpha particles bounced back by 180°, this indicated that the entire positive charge and mass of the atom was concentrated in a very small volume within the atom.

The main features of Rutherford's Theory of an Atom

- There is a positively charged centre in the atom called the **nucleus** in which nearly all the mass of the atom is concentrated.
- Negatively charged particles called **electrons** revolve around the nucleus in paths called **orbits**.

Drawbacks of Rutherford's Model of an Atom

- Rutherford's atomic model could not explain how the moving electrons remain in orbit.
- Any charged particle during acceleration would give out energy, and while revolving it, would lose energy and eventually fall into the nucleus. This means that the atom would be highly unstable.
- The major drawback of Rutherford's atomic model was that it could not explain the stability of atoms.

Niels Bohr's Atomic Model

- Electrons were confined to clearly defined orbits.
- Electrons do not radiate energy while revolving in the orbit.
- An electron must absorb or emit specific amount of energy to transition between these fixed orbits.

Discovery of the Neutron

- In 1932, James Chadwick discovered another sub-atomic particle inside present in the nucleus and named them neutron (n). Neutron had no charge and had a mass nearly equal to that of a proton.

Properties of Electron, Proton and Neutron

Sub-atomic particle	Symbol	Location in the atom	Relative Charge	Relative mass	Actual mass
Electron	e	Outside the nucleus	−1	1/1840 a.m.u.	9.1×10^{-31} kg
Proton	p	Inside the nucleus	+1	1 a.m.u.	1.673×10^{-27} kg
Neutron	n	Inside the nucleus	0	1 a.m.u.	1.675×10^{-27} kg

Parameters of an Atom

Atomic number (Z) = Number of protons (p)

Atomic mass number (A) = Number of protons (p) + Number of neutrons (n)

The atomic number, atomic mass number and symbol of an element are written as - $^{\text{Mass number}}_{\text{Atomic number}}X$ or $^{Z}_{A}X$

The mass of an atom is measured in a unit called **Dalton** and is expressed as **u**.
Particles which constitute the nucleus are called nucleons. Proton and neutrons are nucleons.

Distribution of Electrons in Orbits

Bohr-Bury Scheme

- According to Bohr's model, electrons occupy certain stable orbits or shells. Each shell has a definite energy. The maximum number of electrons in different shells is as follows:
  - The first orbit or K shell will have 2×1^2 = **2** electrons.
  - The second shell will have 2×2^2 = **8** electrons.
  - The third shell will have 2×3^2 = **18** electrons.
  - The fourth shell will have 2×4^2 = **32** electrons and so on.
- Electrons are not accommodated in a given shell unless the inner shells are filled.

Electronic Configuration of Elements

- Electrons in the K shell have minimum energy. Electrons in subsequent shells have higher energies.
- The arrangement of electrons of each element is called the electronic configuration of the element.
- It is represented by numbers which correspond to the electrons in the shells.

The electronic configurations of some elements:

<table>
<tr><th colspan="3">Electronic configuration</th></tr>
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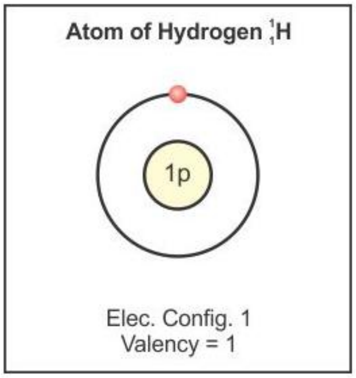
<td>
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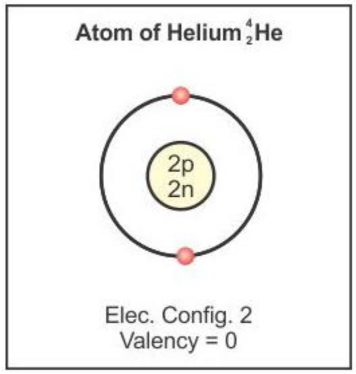
<td>
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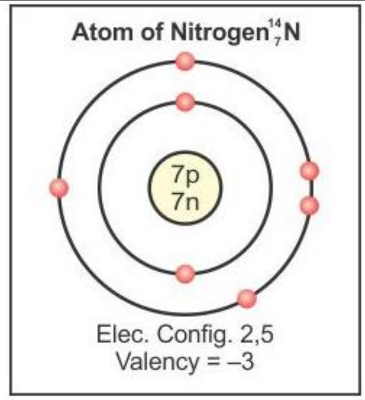
<td>
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</tr>
<tr><td>Hydrogen</td><td>Helium</td><td>Nitrogen</td></tr>
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- The symbol of hydrogen is H; the number of electrons is one. Therefore, the electronic configuration is also 1. Because it has only one electron, it will occupy the K shell.
- The symbol of helium is He. It has two electrons. The electronic configuration is also two. Both these electrons will occupy the K shell. This arrangement is known as a **Duplet**.
- The symbol of nitrogen is N. The number of electrons is 7. Therefore, the electronic configuration is 2, 5. This means five electrons are in the L shell.
- All Noble gases, except helium, have eight electrons in the outermost shell. This arrangement is known as an **Octet**.

Valence Electrons

The outermost shell of an atom is called its valence shell, and the number of electrons present in the valence shell is known as valence electrons.

- The number of valence electrons also indicates the metallic or non-metallic nature of the element. If an atom has three or lesser valence electrons, then it is a metal. On the other hand, an element with four or a higher number of valence electrons is a non-metal.
- The valence electrons participate in chemical bonding.
- Valency of an element represents the combining capacity of the element.
- Elements that have eight electrons in the valence shells, and Helium with two electrons are called **inert gases**.

Reasons for Chemical Activity of Atoms

The chemical activity of an element depends on the number of electrons in the valence shell of its atoms. Chemically active elements have an incomplete octet in the valence shell of their atoms.
Atoms complete their octets by sharing, accepting or donating electrons.

Chemical Bond

A chemical bond is defined as the force of attraction between any two atoms in a molecule to maintain stability.

Noble Gases

- Have stable electronic configuration, i.e. their outermost shell is complete hence chemically unreactive.
- They have 8 valence electrons except helium with 2 valence electrons.

Atoms of Elements – Other than Noble gases

- Have unstable electronic configuration, i.e. their outermost shell is incomplete.
- They can lose, gain or share electrons and are chemically reactive.

Reasons for Chemical Bonding

- The driving force for atoms to combine is related to the tendency of each atom to attain stable electronic configuration of the nearest inert noble gas.

Isotopes

- Atoms of the same element differing in the number of neutrons in their nuclei are known as **isotopes**. Thus, isotopes of an element have the same atomic number but different atomic mass numbers.

Examples of Isotopes		
Element	**Number of Isotopes**	
Hydrogen	Three	Protium (^1_1H) Deuterium (^2_1H) Tritium (^3_1H)

Subatomic particles in Isotopes				
Element	**Isotopes**	**Number of protons**	**Number of neutrons**	**Number of electrons**
Carbon	Carbon-12	6	6	6
	Carbon-13	6	7	6
	Carbon-14	6	8	6

Average atomic mass of chlorine

The isotopes of chlorine, found in nature are in the ratio 3 : 1.

So, in any sample of chlorine, $^{35}_{17}Cl$ will constitute 75% and $^{37}_{17}Cl$ will constitute 25%.

Therefore, in any sample of chlorine, the average atomic mass will be 35.5 u.

Radioactive Isotopes

- **The unstable isotopes which emit various types of radiations are known as radioactive isotopes.**
- A few commonly used radioactive isotopes are carbon-14, arsenic-74, sodium-24, iodine-131, cobalt-60 and uranium-235.

Applications of Isotopes

- Uranium-235 isotope is the fuel of choice for nuclear power plants.
- Cobalt-60 is the isotope of choice for radiotherapy.

Isobars

The atoms of different elements having different atomic numbers but the same mass number are known as isobars.
Examples of Isobars:

Isobars	Number of protons	Number of neutrons	Mass number
Chlorine-37	17	20	37
Argon-37	18	19	

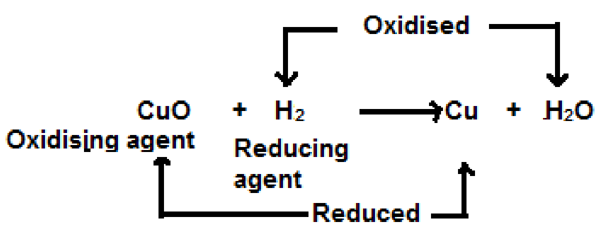
Methods for Achieving Chemical Bonding

There are three methods in which atoms can achieve a stable configuration.

- Transfer of one or more electrons from one atom to the other to form an **electrovalent bond**.
- Sharing of one, two or three pairs of electrons between two atoms to form a **covalent bond.**
- When the shared electron pairs are contributed by only one of the combining atoms, the bond formed is known as a **coordinate bond.**

Redox Reaction

A chemical reaction in which loss of electrons and the gain of electrons take place simultaneously is called a redox reaction.
Example:



In the reaction, hydrogen acts as a reducing agent and reduces copper oxide to copper. This is a reduction reaction.
Reduction: $Cu^{+2} + 2e^{-} \rightarrow Cu$
Simultaneously, copper oxide acts as a oxidizing agent and oxidizes hydrogen to water and this is an oxidation reaction.
Oxidation: $2H - 2e^{-} \rightarrow 2H^{+}$

Electrovalent (or Ionic) Bond

- The chemical bond formed between two atoms by transfer of one or more electrons from the atom of a metallic electropositive element to an atom of a non-metallic electronegative element.
- The compound formed by such electrovalent bonding is called ionic compound.
- The number of electrons donated or accepted by the valence shell of an atom of an element so as to achieve stable electronic configuration is called electrovalency.
- Since the electrostatic force of attraction between opposite charges is much higher, it makes the ionic compounds stable.

Formation of Electrovalent Compounds

A metallic atom loses electrons to attain a stable electronic configuration and become a cation.

Example:	Na	$\xrightarrow{-e^-}$	Na^+
	2, 8, 1		2, 8
	(Neutral)		(Cation)

A non-metallic atom gains electrons to attain a stable electronic configuration and become an anion.

Example:	Cl	$\xrightarrow{+e^-}$	Cl^-
	2, 8, 7		2, 8, 8
	(Neutral)		(Anion)

Cations and anions are oppositely charged particles which attract one another to form an electrovalent bond leading to the formation of an electrovalent compound.

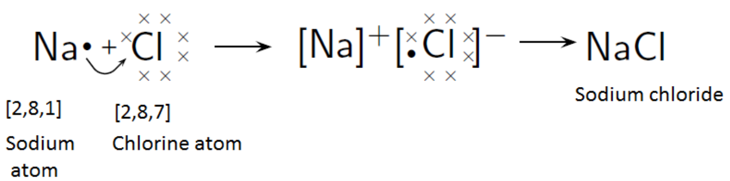
Ionic Equation

$$\underset{(2,\,8,\,1)}{Na} - 1e^- \rightarrow \underset{(2,\,8)}{Na^+}$$

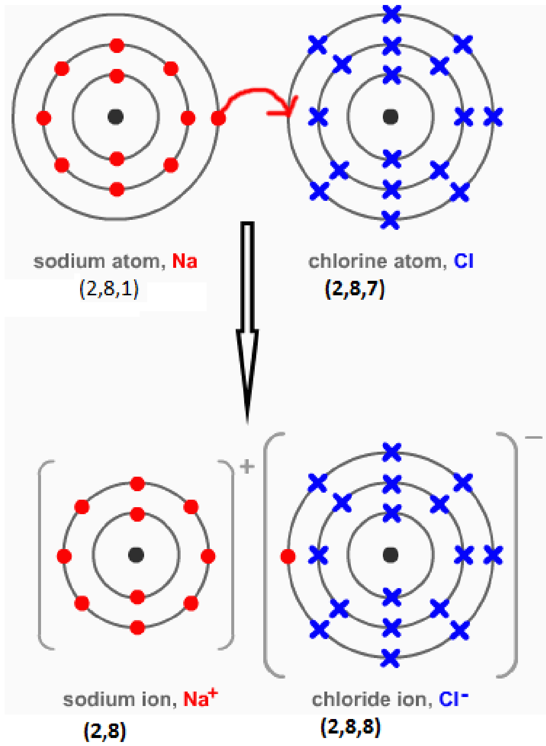
$$\underset{(2,\,8,\,7)}{Cl} + 1e^- \rightarrow \underset{(2,\,8,\,8)}{Cl^-}$$

$$Na + Cl \rightarrow Na^+\,Cl^- \rightarrow NaCl$$

Electron dot Structural Diagram



Atomic or Orbit Structural Diagram



Covalent Bond

- The chemical bond formed due to mutual sharing of electrons between the given pairs of atoms of non-metallic elements. The compound formed with such sharing of electrons is called covalent compound.
- The atoms of non-metals usually have 5, 6 or 7 electrons in their outermost shell (except carbon which has 4 and hydrogen which has just 1 electron in the outermost shell.)
- The atoms of such elements do not favour the loss of its electrons due to energy considerations and thus the transfer of electrons is not possible. Each atom contributes equal number of electron(s).

Types of Covalent bonds

- **Single Covalent Bond**
  It is formed by sharing one pair of electrons between the atoms, each contributing one electron.
  A single covalent bond is denoted by putting a short line (—) between the two atoms.
  Example: Hydrogen molecule can be written as H—H.

- **Double Covalent Bond**
  It is formed by sharing two pair of electrons between the atoms, each contributing two electrons.
  Example: Oxygen molecule can be written as O=O.

- **Triple Covalent Bond**
  It is formed by sharing three pair of electrons between the atoms, each contributing three electrons.
  Example: Nitrogen molecule can be written as N≡N.

Covalency

The number of electron pairs which an atom shares with one or more atoms of the same kind or different kind to achieve stable electronic configuration is called covalency.

Non-polar Covalent Compounds

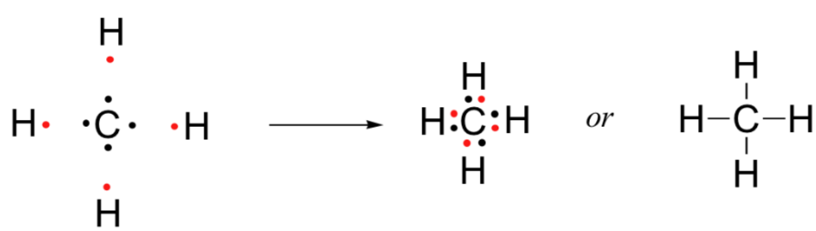
Covalent compounds are said to be non-polar when the shared pair of electrons are equally distributed between the two atoms. Examples: H_2, Cl_2, O_2.

Polar Covalent Compounds

Covalent compounds are said to be polar when shared pair of electrons are unequally distributed between the two atoms. Examples: H_2O, NH_3.

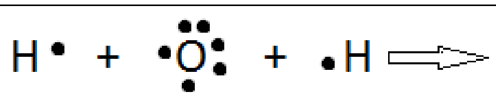
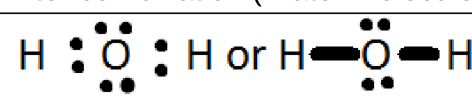
Formation of Methane Molecule – Non-Polar Covalent Compound

One atom of carbon shares four electron pairs, one with each of the four atoms of hydrogen.

	
Before combination (4 [H] and 1 [C] atom)	After combination (CH_4 – Methane molecule)

Formation of Water – Polar Covalent Compound

Atom	Electronic configuration	Nearest noble gas	To attain stable electronic configuration of nearest noble gas.
Hydrogen	${}^{1}_{1}H$ [1]	Helium [2]	Hydrogen needs one electron to complete the duplet.
Oxygen	${}^{16}_{8}O$ [2,6]	Neon [2,8]	Oxygen needs two electrons to complete the octet.

Before combination [2 H and 1 O atom]	After combination (Water molecule)
	 Two single covalent bond