CHAPTER 1

ATOMIC THEORY AND NATURE OF ATOMS

Chemistry studies the composition of matter and the changes it undergoes. Matter has been defined as anything that has weight and can occupy space. All materials are made up of matter. And the three states of matter are solid, liquid and gas.

For centuries, philosopher, chemist and physicist tried to answer the question of what matter was made up of using a variety of experiments and observations.

1.1 THE EARLY IDEAS OF THE ATOM

In 440BC a Greek philosopher named Democritus came up with a conceived idea that matter was composed of very tiny and indestructible particles. He stated that these particles were the smallest unit of matter which he called **atomos** – meaning indivisible in Greek. The theory was not generally accepted because it was a mere philosophy (guess). In 1803, John Dalton proposed the atomic theory as follows:

1. All elements are composed of atoms and atoms are indivisible and indestructible particles
2. Atoms of the same elements are exactly alike (identical) in size, shape, mass and chemistry. While atoms of different elements are different.
3. Atoms combine with atoms of other elements in simple whole-number ratios to form chemical compounds.

Dalton‟s atomic theory became one of the foundations of chemistry and the dawn of quantitative analysis, because soon after the theory, scientific investigation became rapid. Curiosity and the urge to understand the world became increased.

1.2 JOSEPH JOHN THOMSON ATOMIC MODEL

Discovery of the electrons

In 1897, J.J. Thomson discovered the electrons while investigating the electrical conductivity of gases at very low pressure. In his experiment, a very high potential difference (pd) was placed across a glass tube containing gas at a very low pressure. At that experimental condition, a glow coming from the negative terminal was observed and it was attracted to the anode. A hole in the anode allowed the rays to pass through and hit the end of the glass tubes surface. The rays was deflected to the north pole by the magnetic field in the glass tube. The beam hit the fluorescent screen at the positively charged plate, when it was exposed to an electric field. However, in the absence of both the magnetic and electrical field or when both fields are on such that the fields cancel each other influence, the beam was not deflected, it travelled in a straight line and hit the fluorescent screen at point directly opposite the source.

Thomson concluded that the special ray must be negatively charged because it was attracted to the positively charged plate and moved away from the negatively charge plate. His conclusion was supported by the electromagnetic theory which state that a moving charged body behaves like a magnet which interact with the electric and magnetic field through which it passes. From his experiment, Thompson was also able to show that the ray was the same irrespective of the type of gas in the tube or the type of metal used as the electrodes, stating that the ray must be inside every atom. The glass tube Thompson used in his experiment is also called the cathode ray tube, while the negatively charged ray that he discovered are now known as electrons.

Properties of the cathode rays;

1. They are identical irrespective of cathodes (different metals) or gas used in the experiment, all produced same results.
2. The rays are deflected by magnetic and electric fields.
3. The rays produced some chemical reactions similar to those produced by light.
4. They travel in a straight line when it is not interfered with.

Also, from J.J. Thompson"s investigation he showed that at ordinary pressure gases are electrical insulators, but at pressure below 0.01amt, when subjected to a high voltage, gases break down and conduct electricity.

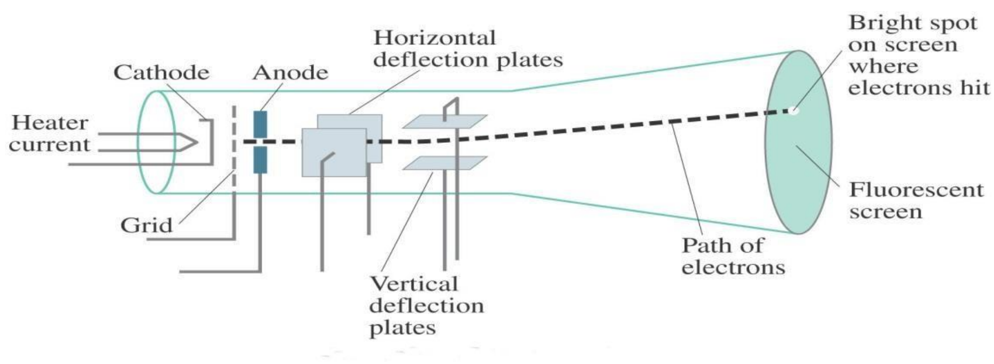
Based on the knowledge he obtained from the cathode ray experiment- that is the effect of electric and magnetic field on negatively charged particles, J.J. Thompson was able to derive the ratio of electric charge to mass of an electron as -1.76 X10^8 C/g. Where C = Coulomb. Also, a scientist called R.A. Millikan conducted a number of experiments in 1916 and found the charge of an electron to be -1.60 X 10^{-19}C. Finally, the mass of the electron was calculated from the data obtained by J.J. Thompson"s and R.A. Millikan.

$$\text{Mass of an electron} = \frac{\text{charge}}{\text{Charge/mass}}$$

$$= \frac{-1.60 \times 10^{-19}}{-1.76 \times 10^{8}\ \text{C/g.}}$$

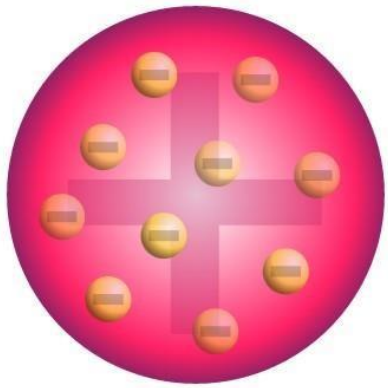
$$9.09\text{x } 10^{-28}\text{g}$$

FIGURE 1.1 THE CATHODE RAY TUBE



By the early 1900s,it became clear that cathode rays have identical properties regardless of the element used to produce them. Also, that all elements must contain negatively charged electrons and atoms are neutral. Therefore, there must be positive particles in the atom to balance the negative charge of the electrons. Thus J.J. Thompson later proposed a revised model for the atom called the plum pudding model. This is shown in figure 1.2. He describe an atom as a uniform positive shpere of matter (the dough) in which negatively charged electrons (raisins) were stocked.

FIGURE 1.2 THE PLUM – PUDDING MODEL



1.3 RUTHERFORD EXPERIMENT

Discovery of the nucleus

In 1910, Hans Gieger and Ernest Marsden under the supervision of Ernest (later Lord) Rutherford discovered the nucleus. Streams of positively charged particles were fired at a thin sheet of gold (gold foil). Rutherford"s team observed that most of the alpha particles passed through the gold foil without any deflection at all, while a few of the particles (1:800) were highly deflected through quite large angle. Sometimes also, but less frequently, α-particles would bounce back in the direction from which it came. This observation conflicted with Thompson"s model, because it was anticipated that since the positive charge in the Thomson"s atom was evenly spread out, the beam (α-particle) should have easily passed through the foil with little or no deflection. Based on this observation, Rutherford proposed that most of the particles passed through the foil without any deflection because the atom consisted of largely empty spaces and does not interact much with the α-particles. However, occasionally one of the α-particle interacted with the very dense positively charged center which he called the nucleus and the verylarge repulsion of this positive charged particle and the nucleus resulted in the large deviation in the path of the α-particles, even to the extent of deflecting the α-particle along the direction from which it came. This model was clearly different from Thompson"s.

Rutherford"s model though very useful, it could not adequately explain the arrangement of electrons round the nucleus nor did it answer the question of what prevented the electrons from falling into the nucleus. This is because unlike charges attract and according to classical physics a moving charged body continually loses energy while it spirals inwards and finally collapses into the middle. Rutherford model of an atom fail to explain why electron in a given atoms donot collapse into the nucleus when in motion and could not explain the position and arrangement of electrons around the nucleus.

FIGURE 1.3 RUTHERFORD EXPERIMENT (EXPECTED AND ACTUAL RESULT)

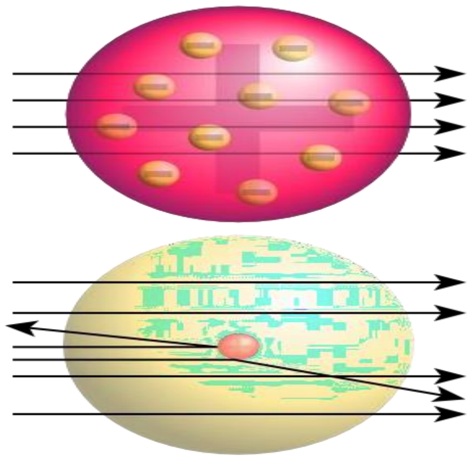
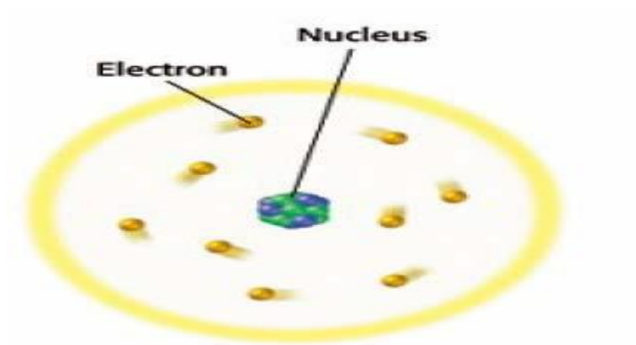


FIGURE 1.4 RUTHERFORD'S MODEL



1.4 CHADWICK EXPERIMENT

Discovery of the neutrons

After a series of experiments were conducted it was suspected that another type of subatomic particle exists. For example, experimental observations show that the mass of an oxygen atom was 16 times larger than the proton and oxygen has only eight positive charges, therefore, accounting for the remaining mass of the atom became a puzzle.

This puzzle was unraveled by James Chadwick in 1932. In his experiment, Chadwick bombarded a sheet of beryllium foil with alpha-particles and unknown radiation was produced. The result from the experiment showed that the particles emitted from beryllium consisted of electrically neutral particles with mass slightly higher than the proton.

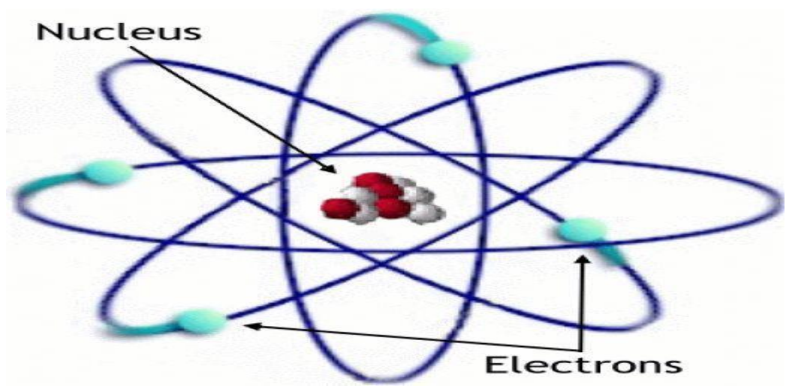
1.5 THE MODERN ATOMIC MODELS

Two atomic models are in use today, they are; the Bohr atomic model and the quantum mechanical model. Of these two models, the Bohr model is simpler and relatively easy to understand. However, because of some shortcomings of the Bohr model, the quantum mechanical model is more valid.

1.5.1 THE BOHR’S MODEL OF THE ATOM

In 1913, a Danish physicist Niels Bohr (1885 – 1962) proposed an improvement to Rutherford‟s model. Bohr explained the arrangement of electrons in an atom and provided answers to what prevented the electrons from falling into the nucleus. He proposed that although the electrons are negatively charged and the nucleus positively charged, the movement of the electrons in definite concentric circular path around nucleus prevented the electrons from collapsing into the nucleus. The concentric circular path (orbit) that the electron orbits is much like the planet orbit around the sun. The orbits are located at certain fixed distances from the nucleus. Bohr based isexplanation of the atomic model and the electronic energy levels of atoms on the experimental observation of the emission spectrum of light by atoms.

FIGURE 1.5 THE BOHR’S ATOMIC MODEL



1.5.2 ELECTROMAGNETIC SPECTRUM

Bohr based his explanations on the electromagnetic spectrum. When a white light source or emitted radiation is passed through a prism, the light splits from its source into series of colouredbands called spectrum. Atoms also give rise to spectrum called the atomic spectrum. Electromagnetic spectrum is the entire distribution of electromagnetic radiation according to frequency and wavelength. It is the entire range of lights that exist and they all travel at the same speed of light in vacuum (3 x 10^8 m/s).

Atoms of an element absorb energy when heated, by passing its gas or vapour through an electric discharge, the atoms become excited and the electron are promoted to a higher energy level. On falling to their lower energy states lose energy and emit light. The colours are characteristics of the element, for example hydrogen is red, and sodium is orange and so on.

Each line in an emission corresponds to a discrete (separate) wavelength, frequency and energy and thus the pattern of lines in the spectrum of each element is unique (like a finger print) to that element and thus can be used in the qualitative and quantitative identification of elements. Since the wavelength of the emission spectrum tells us what the sample is and the intensity of each wavelength can tell us how much of the element is present.

Since all electromagnetic radiation have the same speed (i.e. speed is constant), then frequency of radiation relating to spectral line of wavelength (λ) is giving by the equation;

$$f = \frac{c}{h}$$

$$f\lambda = c$$

As can be seen from the equation, the wavelength of light increases as the frequency decreases

f (or v) = frequency of radiation express as hertz (Hz) is equivalent to cycle "per second"
c = speed of light (3×10^8 m/s.)
λ = wavelength

The work of Max Planck has also showed that electromagnetic radiation may be regarded as stream of particles called photon. The energy carried by a photon is related to its frequency by the expression;

$E = hf$

but $f = \frac{c}{h}$

Therefore, $E = \frac{hc}{h}$

Where h = Planck"s constant with a value of 6.63×10^{-34} j-sec. A single photon carries one quantum of energy.

Example1.1 Calculate the frequency of radiation having a wave length of 6.5×10^{-7}meters.

Solution: $f = \frac{c}{h}$

$c = 3 \times 10^8$m/s

$\lambda = 6.5 \times 10^{-7}$m

$f = \frac{3 \times 108}{6.5 \times 10-7\text{m}} = 4.6 \times 10^{14}$ hertz

Example1.2 Calculate the frequency and energy of a photon having a wave length of 1. 7µm.

Solution: $f = \frac{c}{h}$

$c = 3 \times 10^8$m/s

1micrometer = 10^{-6}m

1.7micrometer = 1.7×10^{-6}m

$\lambda = 1.7 \times 10^{-6}$m

$$f = \frac{3X\,10^{8}}{1.7\,K\,10^{-6}} = 1.77 \text{ x } 10^{14} \text{ hertz}$$

$E = hf$

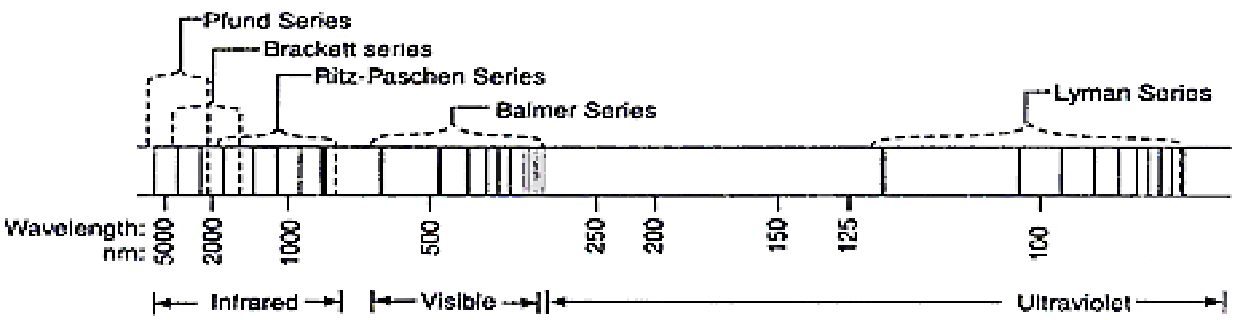
$f = 1.77 \text{ X } 10^{14}$ hertz
$h = 6.63 \text{ X } 10^{-34}$ j-sec

$E = 1.77 \text{ X } 10^{14} \text{ X } 6.63 \text{ X } 10^{-34}$
$E = 1.174 \text{ X} 10^{-19}$j

1.5.3 BOHR'S MODEL OF THE HYDROGEN ATOM

Bohr based his theory of the hydrogen atom on the electromagnetic spectrum that hydrogen produces. At room temperature, hydrogen gas does not emit light but when electricity is passed through a discharge tube containing hydrogen gas, the molecules break down into atoms and the tube glows with a reddish pink light. The emission spectrum reveals that the atom emits radiation in the form of lines. Several series of discrete lines with each corresponding to a different wavelength in the electromagnetic spectrum are observed. Figure 1.7 shows the emission spectrum of atomic hydrogen.

FIGURE 1.6: THE ATOMIC HYDROGEN SPECTRUM, SHOWING THE FIRST FIVE SERIES OF THE SPECTRAL LINES.



It can be observed from the spectrum that the intensity and distance between the lines decreases as the frequency increases after which a continuum is observed. That is the lines in each series become more closely spaced at increase frequency (decreasing wavelength)

The wave length (λ) of the radiation is related to the frequency (f) by the equation

$$\lambda = \frac{\,}{\int}$$

f (or v) = frequency of radiation express as hertz (Hz) is equivalent to cycle "per second"
c = speed of light ($3 \text{ x } 10^{8}$ m/s.)
λ = wavelength (m)

In spectroscopy, frequency is expressed as wave number (A) and the wave number is the reciprocal of wave length. A = $(\frac{1}{h})$ m^{-1}

Calculating the wavelength of the spectral lines of the hydrogen atom is done by using an expression giving by Rydberg and therefore known as Rydberg equation:

$$\text{Frequency (A)} = \frac{1}{(\frac{-}{h})} = R\left(\frac{1}{n1^2} - \frac{1}{n2^2}\right)$$

Where:

n_1= quantum number corresponding to the initial energy level (Start).
n_2=quantum number corresponding to the final energy level.
R is the Rydberg Ritz constants= 10967876m^{-1}

TABLE1.1 THE FIVE SERIES OF LINES IN THE SPECTRUM OF HYDROGEN ATOM

Order	Series	Converge toward	n_1	n_2
First series	Lyman series.	Ultraviolet region.	**1**	**2,3,4,5,------∞**
Second series	Balmer series	Visible region	**2**	**3,4,5,6,------∞**
Third series	Paschen series	Infrared region	**3**	**4,5,6,7,------∞**
Fourth series	Brackette.	Infrared region	**4**	**5,6,7,8,------∞**
Fifth series	Pfund series	Infrared region	**5**	**6,7,8,9,------∞**

It is required that $n_2 > n_1$ and *n* values are all integral up to infinity.

Example1.3 Calculate the wavelength of the first, second and third lines in the Lymer series

Solution: Frequency (A) = $\frac{1}{(\frac{-}{h})} = R\left(\frac{1}{n1^2} - \frac{1}{n2^2}\right)$

For the Lymer series, n_1 = 1 therefore, the first line have n_2 = 2, the second line has n_2 = 3 and the third line has n_2 = 4. Substituting these values into the Rydberg equation we have;

First line: $(\frac{1}{h}) = R\left(\frac{1}{n1^2} - \frac{1}{n2^2}\right)$

$$\frac{1}{h} = 10967876\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$

$$\frac{1}{h} = 10967876\left(\frac{1}{1} - \frac{1}{4}\right)$$

$$\frac{1}{h} = 10967876 \text{ X } 0.75$$

$$\frac{1}{h} = 8225907m^{-1}$$

$$\lambda = \left(\frac{1}{82259.07525}\right) = 1.216 \text{ X } 10^{-7}m$$

Second line: $(\frac{1}{h}) = R\left(\frac{1}{n1^2} - \frac{1}{n2^2}\right)$

$$\frac{1}{h} = 10967876\left(\frac{1}{1^2} - \frac{1}{3^2}\right)$$

$\frac{1}{h} = 10967876(\frac{1}{1} - \frac{1}{9})$

$\frac{1}{h} = 10967876X\ 0.89$

$\frac{1}{h} = 9761409.64m^{-1}$

$\lambda = (\frac{1}{9761409.64}) = 1.024\ X\ 10^{-7}m$

Third line: $(\frac{1}{h}) = R\ (\frac{1}{n1^2} - \frac{1}{n2^2})$

$\frac{1}{h} = 10967876(\frac{1}{1^2} - \frac{1}{4^2})$

$\frac{1}{h} = 10967876(\frac{1}{1} - \frac{1}{16})$

$\frac{1}{h} = 10967876(\frac{15}{16})$

$\frac{1}{h} = 10967876X\ 0.9375$

$\frac{1}{h} = 10282383.75\ m^{-1}$

$\lambda = (\frac{1}{10282383.75}) = 9.73\ X\ 10^{-8}m$

Example1.4 Calculate wavelengths and the energy of transition of the second and third lines in the Balmer series of a hydrogen atom. h = Planck‟s constant = 6.63 X 10^{-34}j/sec, c = speed of light 3 x 10^8 m/s

Solution: Frequency (A) = $\frac{1}{(\frac{}{h})} = R\ (\frac{1}{n1^2} - \frac{1}{n2^2})$

For the Balmer series, $n_1 = 2$ therefore, the second line has $n_2 = 4$ and the third line has $n_2 = 5$. Substituting these values into the Rydberg equation we have;

Second line: $(\frac{1}{h}) = R\ (\frac{1}{n1^2} - \frac{1}{n2^2})$

$\frac{1}{h} = 10967876(\frac{1}{2^2} - \frac{1}{4^2})$

$\frac{1}{h} = 10967876(\frac{1}{4} - \frac{1}{16})$

$\frac{1}{h} = 10967876(\frac{4-1}{16})$

$\frac{1}{h} = 10967876(\frac{3}{16})$

$\frac{1}{h} = 10967876X\ 0.1875$

$\frac{1}{h} = 2056476.75\ m^{-1}$

$\lambda = (\frac{1}{2056476.75}) = 4.863\ X\ 10^{-7}m$

Energy of transition $E = \frac{hC}{h}$

$E = (\frac{6.63\ X\ 10^{-34}K\ 3K10^{8}}{4.863\ X\ 10-7}) = 4.01\ x10^{-19}j$

Third line: $(\frac{1}{h}) = R\ (\frac{1}{n1^2} - \frac{1}{n2^2})$

$\frac{1}{h} = 10967876(\frac{1}{2^2} - \frac{1}{5^2})$

$\frac{1}{h} = 10967876(\frac{1}{4} - \frac{1}{25})$

$\frac{1}{h} = 10967876\ (\frac{25-4}{100})$

$\frac{1}{h} = 10967876(\frac{21}{100})$

$\frac{1}{h} = 10967876\ X\ 0.21$

$\frac{1}{h} = 2303253.96\ m^{-1}$

$\lambda = (\frac{1}{2303253.96}) = 4.34\ X\ 10^{-7}m$

Energy of transition $E = \frac{hC}{h}$

$E = (\frac{6.63\ X\ 10^{-34}K\ 3K10^{8}}{4.34\ X\ 10-7}) = 4.58\ x10^{-19}j$

1.5.4 BOHR'S POSTULATE

The summaries of Bohr"s postulates are as follow:-

1. The electron moves in circular orbits at different energies around the nucleus.
2. The radius of the orbit is quantized. That is orbits are at definite distances from the nucleus.
3. The electron revolve only in the orbit which have fixed value of energy, thus an electron in an atom can have only definite or discrete value of energy.
4. As long as the electron remains in a particular orbit, it neither loses nor gain energy. These orbits are known as stationary state or ground state.
5. If an electron in the ground state (stationary state) with energy E_i, is given sufficient energy and goes into another allowed orbit, with energy E_f, radiation will be emitted, with energy and frequency. $hf = E_i - E_f$

1.5.5 THE LIMITATIONS OF BOHR'S ATOMIC MODEL: -

1. It can only explain the line spectrum of an atom with one electron, such as, H atom, He^{+}, Li^{2+}, Be^{3+} etc. but unable to explain the line spectra of atoms with more than one electron.

2. It cannot explain the splitting of spectra lines in magnetic field known as ZEEMAN EFFECT.

3. Bohr"s model is unable to explain the splitting of spectra lines in an electric field and the phenomenon is known as STARK EFFECT.

4. Bohr"s model is also unable to explain Heisenberg"s uncertainty principles.

1.6 SUBATOMIC PARTICLES

Research studies conducted by Thompson, Chadwick, Rutherford and others reveals that the atom consist of smaller particles. Also, scientist have been able to discover hundreds of subatomic particles, however three of these particles; the Proton, Neutron and Electron are of interest to the chemist.

1.6.1 THE NUCLEUS

The nucleus is the center of the atom and although the size is small compared to the size of the electron, it accounts for 99.9% the mass of an atom because the proton and neutron are found in the nucleons. The proton and neutron are called nucleon because they are found in the nucleus of an atom.

1.6.2 PROTONS, NEUTRON AND ELECTRON

Protons are found in the nucleus of an atom, a proton has a charge of +1, and it is about 1840 times heavier than an electron. Every atom has a specific proton number which tells us the name of the element. An atom chemical identity is determined by its number of protons. A proton is what is left after hydrogen atom losses its single electron.

The neutrons are found in the nucleus. It carries no charge i.e. it is a neutral particle. A neutron like the proton is considered to have a mass of l amu. The neutrons do not take part in chemical reaction under normal condition.

Electrons are found outside the nucleus, an electron has a charge of -1. The mass of an electron is negligible. An electron is approximately $\frac{1}{1840}$ the mass of the proton.

TABLE 1.2: THE SUBATOMIC PARTICLES

Particle	Symbol	Mass (g)	Relative electrical charge	Approximate relative mass (a m u)
Electron	e^-	9.091×10^{-28}	1-	$\frac{1}{1840}$
Proton	p^+	1.67252×10^{-24}	1+	1
Neutron	n^0	1.67495×10^{-24}	0	1

Note: amu is atomic mass unit

Atoms of an element are electrically neutral, that is they have equal number of protons and electrons. However, an atom can lose or gain electrons to become an ion. If it gains electron, it becomes negatively charged and is called an anion. If it losses an electron it becomes positively charged and is called a cation.

1.7 ATOMIC NUMBER AND MASS NUMBER

The proton number is referred to as the atomic number. The atomic number is the whole number that increases across each row of the periodic table from left to right. The sum total of the number of protons and neutrons in the nucleus of an atom is the mass number. If an element X is represented thus: $^{A}X_{Z}$.

A = Mass number
Z = Atomic number. Therefore
A-Z = Neutron number

1.8 THE RELATIVE ATOMIC MASS

The mass number of an atom cannot be weighed directly because it is very small; however experimental methods are used to determine the mass of one atom relative to another.

The internationally agreed standard is an atom of carbon-12. This is an isotope of carbon with six protons and six neutrons, having a mass of 12 atomic mass unit (amu). Based on this standard, 1amu is exactly equal to one- twelfth the mass of one carbon-12 atom ($\frac{1}{12th}$, the mass of carbon 12). It is because the actual masses of the proton and neutron are very small that for practical purpose, the atomic mass unit (amu) is used.

Mass of one carbon-12 atom = 12 atomic mass unit therefore,

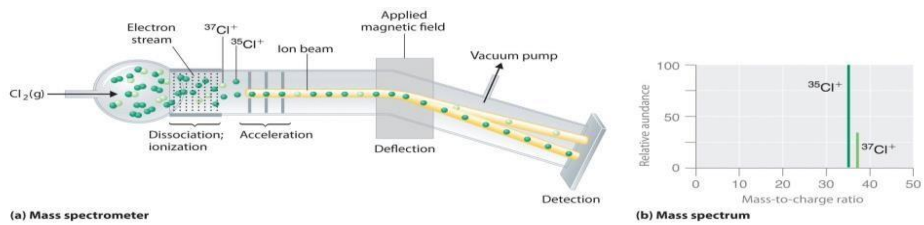
$$1 \text{ atomic mass unit} = \frac{\text{mass of one carbon}-12 \text{ atom}}{12}$$

The atomic mass unit is related to the number of protons and neutrons. It has nearly the same value as the mass number.

1.9 THE MASS SPECTROMETER

The process of producing and analysing spectra is called spectroscopy and the instrument for analysing the spectra is called a spectrometer. The mass spectrometer is an instrument designed to separate atoms of slightly different masses. First the sample to be separated is vapourized so as to allow it to move through the machine. The vapourized sample is then hit with high energy electrons, knocking off one or more electrons thus converting the atoms into positively charged ions. The ions are accelerated to a higher speed by an electric and magnetic field; also because ions are charged particles they can be deflected by an electric and magnetic field. The degree of deflection is thus dependent on the mass to charge ratio. The masses of each particle aredetermined by measuring the exact strength of the field and the degree of deflection of the particles in the mass spectrometer. Different elements have different mass which corresponds to a different deflection. Figure 1.7 shows the mass spectrometer

FIGURE 1.7: THE MASS SPECTROMETER



1.10 MASS SPECTRA AND ISOTOPES

The proton number (atomic number) tells us what an element is. Every element has its specific number of proton; however the number of neutrons and electrons may vary. For example carbon 12 has a proton number of 6 and a neutron number of 6 but carbon 14 has a proton number of 6 but a neutron number of 8 with a mass number of 14. This phenomenon is known as isotopy. Carbon 12 and carbon 14 are known as isotopes of carbon. Thus isotopy is defined as a phenomenon in which an element has the same number of proton but different number of neutrons. In other words isotopy is the existence of an element with the same atomic number but different mass number.

If an element X is represented thus: $^{A}_{Z}X$

A = the mass number (values are different for isotopes)
Z = the atomic number. (Values are same for isotopes) Therefore,
A-Z = Neutron number

Majority of elements found in nature are a mixture of isotopes. For example hydrogen has three Isotopes namely: hydrogen $^{1}_{1}H$, deuterium $^{2}_{1}H$ and tritium $^{3}_{1}H$, which is unstable and Disintegrates. Almost all hydrogen exists as hydrogen with no neutron. Also, Bromine exists in nature as $^{79}_{35}Br$ and bromine $^{81}_{35}Br$. A mass spectrometer is used to find out the masses of these Isotopes using the carbon -12 scale and the proportion of each isotope. Any naturally occurring Sample of bromine contains approximately 50.52% of $^{79}_{35}Br$, and 49.48% of $^{81}_{35}Br$

1.12 CALCULATING RELATIVE ATOMIC MASSES FROM SPECTRA

The mass spectrum of an element will show the presence of isotopes. From the relative heights of the peaks, we can work out the relative atomic mass of the element.

Example 4.3: Calculate the relative atomic mass of bromine from the mass spectra of bromine

Solution: Relative abundance of $^{79}_{35}Br$ = 50.52%
Relative abundance of $^{81}_{35}Br$ = 49.48%

$$\text{Relative atomic mass} = \frac{79\ x\ 50.52}{100} + \frac{81\ x\ 49.48}{100}$$

$$= \frac{3991.08}{100} + \frac{4007.88}{100}$$

$$= 39.9108 + 40.0788$$

$$= 79.9896\ 9 \approx 80$$

1.13 USES OF ISOTOPES

Isotopes have a wide range of uses and it has been used in various aspects of life and nature.

Medicine: Isotopes have been used in the field of medicine as tracers; these are substances containing a radioactive Isotope of an atom. These tracers are taken up by the organs to be monitored. For example, Iodine – 131 has been used to study the thyroid function.

Biology: Phosphorus – 32 has been used to study intake of nutrients in plants.

Power Generation: The fission of uranium 235 and the fusion of H – 3 have been used to create-explosives.

History and Archaeology: Carbon – 1.4 has been used by historians and archaeologist in determining the ages of ancient remains, such as rocks, bones, plants and so on. This technique is known as radiocarbon dating.

PRACTICE QUESTIONS

1. Refute Dalton"s Theory of "indivisible" atom using J.J. Thompson"s and Rutherford Model of the atom.

2. A line spectrum in the hydrogen atom has a wavelength corresponding to 4.57 x 10^{-7}m. Calculate the frequency and the energy associated with this transition.

3. Calculate the energy of a photon of radiation having a frequency of 3.54 x 10^{14} hertz.

4. Differentiate between the following;

   i. The spectra line of white light and the spectral lines of elements.
   ii. Ground state of an electron and the excited state

5. Give reasons for the following:
   (i) The nucleus accounts for the mass of an atom.
   (ii) The number of protons tells us the name of the element.
   (iii) Atomic masses unlike the atomic numbers are not whole numbers.

6. (a) Complete the table below:

Atom or ion of element	**A**	**B**	**C**	**D**	**E**
Mass number	10	-	39	-	19
Number of proton	5	7	-	20	9
Number of neutron	-	7	20	20	-

7. What is the atomic mass of chlorine, if $^{35}_{17}Cl$ has a relative abundance of 75% and $^{37}_{17}Cl$ has a relative abundance of 25%

8. Calculate the wavelength the frequency and energy of the lines in the Balmer series when n_2 = 3 and 5

9. The wave number of a line in the Lyman series is 10282383.75m^{-1}
i. Calculate the frequency and energy of the series
iii. Which line in the series is it?

Assume that the atomic mass of an element(X) is 25.77. X has three isotopes; 74.7% of $^{26}_{13}$X, 24.7% of $^{a}_{13}$X and 0.6% of $^{28}_{13}$X . Calculate for the value of a