MODULE 1 BASIC STRUCTURE OF AN ATOM AND ATOMIC MODELS

UNIT 1 THE ATOM, ITS STRUCTURE AND CHARGE QUANTIZATION

CONTENTS

1.0 INTRODUCTION

This unit defines an atom, discusses its constituents and the size of an atom. The structure of the atom is discussed. The charge of an atom and charge quantization is also discussed.

2.0 A OBJECTIVES

At the end of this unit you will be able to:

- explain what is meant by an atom;
- describe the constituents of an atom;
- state the size of an atom; and
- explain charge of an atom and charge quantization.

2.0 B How to Study this Unit

1. You are expected to read carefully through this unit twice before attempting to answer the activity questions. Do not look at the solution or guides provided at the end of the unit until you are satisfied that you have done your best to get all the answers.
2. Share your difficulties in understanding the unit with your mates, facilitators and by consulting other relevant materials or internet.
3. Ensure that you only check correct answers to the activities as a way of confirming what you have done.
4. Note that if you follow these instructions strictly, you will feel fulfilled at the end that you have achieved your aim and could stimulate you to do more.

3.0 MAIN CONTENT

3.1 An Atom

Is the Smallest particle of an element that can take part in a chemical reaction and retains all chemical properties of that element.

3.2 Atomic Structure

The simplest of the atom is spherical in shape. Note however that some atoms, especially the heavy ones do not exhibit spherical shape. It has a nucleus that is positively charged and located at the centre of the atom. The nucleus is made up of protons and neutrons. The negative electrons move round the nucleus in energy shells. The electrons are held together by strong Columbic forces. The positively charged nucleus and the negatively charged electrons combined to give an atom a neutral charge. Atoms have a radius of 10^{-10} m and the radius of the nucleus is 10^{-16} m.

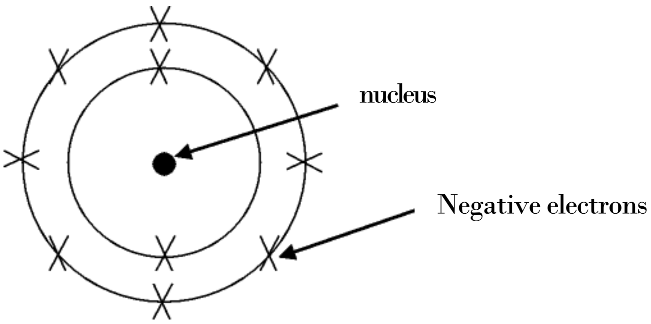

Fig. 1 An Atom

The Structure of the Atom

- charge carried by electrons
- e- has very small mass
- bulk of the atom is +ve charge
- most mass resides in +ve charge

Question: what is the spatial distribution of charge inside an atom?

3.3 Charge Quantization

Electricity consists of charges carried by electrons, protons, neutrons etc. Electric charges are of two forms: Positive and Negative. Negative charges are called electrons and positive charges are called protons.

The charge carried by both electron and proton are exactly equal but opposite.

The minimum electric charge is denoted by a symbol *e*, the electronic charge has a magnitude of 1.6×10^{-19} Coulombs.

That is $e = 1.6 \times 10^{-19}$ C.

Any physical existing charge in the universe is an integral multiple of *e*,
i.e multiple integral of $e = ne$, where n is any number. So charge (e) exists in discrete form and **not** in continuous amount. This is referred to as **charge quantization.**

4.0 CONCLUSION

a) In this unit you have learnt that an atom has a radius of 10^{-10} m.
b) The atom has a nucleus of radius 10^{-16} m
c) The nucleus contains neutrons, protons, and electrons travel round the nucleus.
d) Electrons carry a charge, -1.6×10^{-19} C.
e) Electric charge is quantized.

5.0 SUMMARY

You have learnt in this unit:

- What an atom is.
- The structure of an atom.
- The constituents' particles of the atom.
- Charge and charge quantization.

6.0 TUTORS MARKED ASSIGNMENT

1. Define an atom
2. Briefly describe the structure of an atom.

7.0 REFERENCES/FURTHER READING

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UNIT 2 MASS SPECTRA

CONTENTS

1.0 INTRODUCTION

In unit we would explain mass spectra and discuss its principle of operation. Isotopes and the relative abundance of isotopes will be treated. We would also calculate the relative atomic masses of some elements.

2.0 A Objectives

At the end of this unit, you will be able to:

- explain what a mass spectrum is;
- describe the operation of a mass spectrum;
- define isotopes;
- calculate the relative atomic masses of elements; and
- explain types of radio isotopes by origin

2.0 B How to Study this Unit:

1. You are expected to read carefully through this unit twice before attempting to answer the activity questions. Do not look at the solution or guides provided at the end of the unit until you are satisfied that you have done your best to get all the answers.
2. Share your difficulties in understanding the unit with your mates, facilitators and by consulting other relevant materials or internet.
3. Ensure that you only check correct answers to the activities as a way of confirming what you have done.
4. Note that if you follow these instructions strictly, you will feel fulfilled at the end that you have achieved your aim and could stimulate you to do more.

3.0 MAIN CONTENT

3.1 Mass Spectra

(Singular, mass spectrum): Is a pattern of the chemical constituents of a substance separated according to their mass and presented as a spectrum as measured using a mass spectrometer. One version of this device is known as the **Bainbridge mass spectrometer**. It operates on the principle that a beam of ions is made to pass through a velocity selector. In the velocity selector the charged particles all move with the same velocity. This is achieved by a combination of a vertically downward electric field E that is perpendicular to a magnetic field B. In the velocity selector the magnetic force is $q\mathbf{v}\times\mathbf{B}$ and the electric field is q**E.** When the charged particles move in the straight horizontal line through the fields, the magnetic force is equal to the electric field i.e qvB = qE

$$v = \frac{E}{B}$$

The charged particles from the velocity selector enter the second magnetic field B_0 that has the same direction as in the velocity selector. Upon entering the second magnetic field the ions move in semicircles of radius r before striking a photographic plate. If the ions are positively charged, the beam deflects upward, and if the ions are negatively charged, the beam deflects upward. While in circular motion, the centripetal $(F = \frac{mv^2}{r})$ force is equal to the magnetic force (qvB_0)

$$i.e \quad \frac{mv^2}{r} = B_0qv$$

$$\frac{mv}{r} = B_0q$$

$$\frac{m}{q} = \frac{B_0r}{v}$$

$$but \ v = \frac{E}{B}$$

$$\frac{m}{q} = \frac{B_0rB}{E}$$

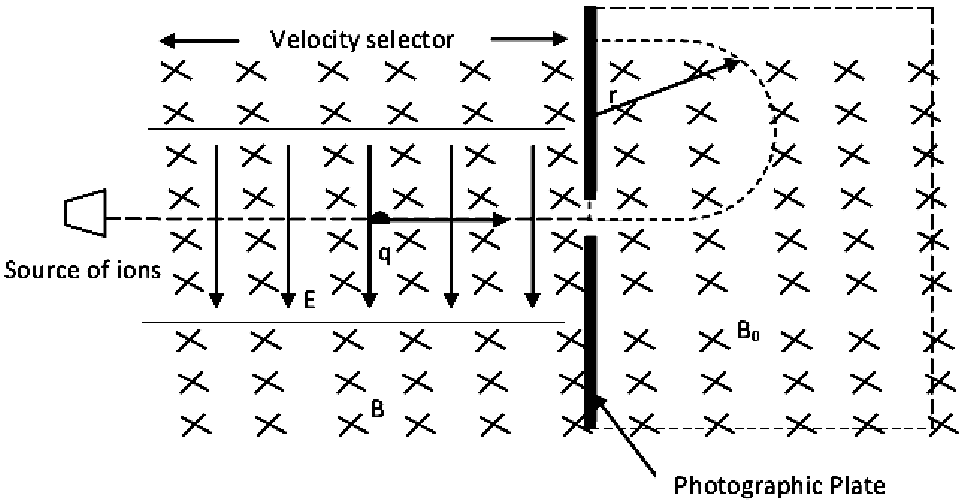

Fig. 2 Mass Spectrometer

Therefore, we can determine $\frac{m}{q}$ by measuring the radius r of the curvature and knowing the fields B, B_0, and E. In practice the mass spectrometer measures the various isotopes of a given ion, since the ions all carry the same charge q. In this way, the mass ratios can be determined even if q is unknown.

Types of Isotopes: Radioactive

Types of radioactive decay

- alpha
- beta positive
- beta negative
- electron capture

Types of radioactive isotopes (by origin)

- long-lived radioactive nuclides
- cosmogenic
- anthropogenic
- radiogenic

Radioactive isotopes are nuclides (isotope-specific atoms) that have unstable nuclei that decay, emitting alpha, beta, and sometimes gamma rays. Such isotopes eventually reach stability in the form of nonradioactive isotopes of other chemical elements, their "radiogenic daughters." Decay of a radionuclide to a stable radiogenic daughter is a function of time measured in units of half-lives.

Types of radioactive decay (return to top)

1) ***alpha (a) decay*** results from an excess of mass. In this type of decay, alpha particles (consisting of two protons and two neutrons) are emitted from the nucleus. Both the atomic number and neutron number of the daughter are reduced by two, so the mass number decreases by four. An example is the decay of ^{238}U:

$$^{238}_{92}U \rightarrow ^{234}_{90}Th + ^{4}_{2}He + \gamma$$

2) ***ß+ - or "positron decay"*** results from an excess of protons. In this type of decay, a positively charged beta particle and a neutrino are emitted from the nucleus. The atomic number decreases by one and the neutron number is increased by one. An example is the decay of radioactive ^{18}F to stable^{18}O:

$$^{18}_{9}F \rightarrow ^{18}_{8}O + \beta^{+} + \nu + Q$$

where *ß+* is the positron, *v* is the neutrino, and *Q* is the total energy given off.

3) ***ß- - or "negatron decay"*** results from an excess of neutrons. In this type of decay, a negatively charged beta particle and a neutrino are emitted from the nucleus. The atomic number increases by one and the neutron number is reduced by one. An example is the decay of radioactive ^{14}C to stable ^{14}N:

$$^{14}_{6}C \rightarrow ^{14}_{7}N + \beta^{-} + \bar{\nu} + Q$$

where *ß-* is the beta particle, *v* is the antineutrino, and *Q* is the end point energy (0.156 MeV).

4) ***electron capture*** also results from an excess of protons. In this type of decay, an electron is spontaneously incorporated into the nucleus and a neutrino is emitted from the nucleus. The atomic number decreases by one and the neutron number increases by one. Electron capture may be followed by the emission of a gamma ray. An example is the decay of ^{123}I to ^{123}Te:

$$^{123}_{53}I + e^{-} \rightarrow ^{123}_{52}Te + \nu + \gamma$$

Types of radioactive isotopes by origin (return to top)

1) ***Long-lived radioactive nuclides***

Some radioactive nuclides that have very long half lives were created during the formation of the solar system (~4.6 billion years ago) and are still present in the earth. These include ^{40}K (t½ = 1.28 billion years), ^{87}Rb (t½ = 48.8 billion years), ^{238}U (t½ = 447 billion years), and ^{186}Os (t½ = 2 x 106 billion years, or 2 million billion years).

2) ***Cosmogenic***

Cosmogenic isotopes are a result of cosmic ray activity in the atmosphere. Cosmic rays are atomic particles that are ejected from stars at a rate of speed sufficient to shatter other atoms when they collide. This process of transformation is called spallation. Some of the resulting fragments produced are unstable atoms having a different atomic structure (and atomic number), and so are isotopes of another element. The resulting atoms are considered to have cosmogenic radioactivity. Cosmogenic isotopes are also produced at the surface of the earth by direct cosmic ray irradiation of atoms in solid geologic materials.

Examples of cosmogenic nuclides include ^{14}C, ^{36}Cl, ^{3}H, ^{32}Si, and ^{10}Be. Cosmogenic nuclides, since they are produced in the atmosphere or on the surface of the earth and have relatively short half-lives (10 to 30,000 years), are often used for age dating of waters.

3) *Anthropogenic*

Anthropogenic isotopes result from human activities, such as the processing of nuclear fuels, reactor accidents, and nuclear weapons testing. Such testing in the 1950s and 1960s greatly increased the amounts of tritium (^{3}H) and ^{14}C in the atmosphere; tracking these isotopes in the deep ocean, for instance, allows oceanographers to study ocean flow, currents, and rates of sedimentation. Likewise, in hydrology it allows for the tracking of recent groundwater recharge and flow rates in the vadose zone. Examples of hydrologically useful anthropogenic isotopes include many of the cosmogenic isotopes mentioned above: ^{3}H, ^{14}C, ^{36}Cl, and ^{85}Kr.

4) *Radiogenic*

Radiogenic isotopes are typically stable daughter isotopes produced from radioactive decay. In the geosciences, radiogenic isotopes help to determine the nature and timing of geological events and processes. Isotopic systems useful in this research are primarily K-Ar, Rb-Sr, Re-Os, Sm-Nd, U-Th-Pb, and the noble gases (^{4}H, ^{3}H-^{3}He, ^{40}Ar).

Because of their stable evolution in groundwater, such naturally occurring isotopes are useful hydrologic tracers, allowing evaluation of large geographic areas to determine flow paths and flow rates. Consequently, they are helpful in building models that predict fracturing, aquifer thickness, and other subterranean features.

3.2 Isotopes

Are elements with unequal mass number but equal atomic number (or same atomic number but different neutrons number)

E.g 1. $^{2}_{1}H$, $^{3}_{1}H$

2. $^{35}_{17}Cl$, $^{37}_{17}Cl$ etc.

Example 1

Any naturally occurring sample of chlorine contains ^{35}Cl and ^{37}Cl in the proportion 75.53% : 24.47%. For a sample of 100 atoms of 100 atoms of chlorine, determine the average mass of chlorine.

Solution

Average mass = $35X\frac{75.53}{100} + 37X\frac{24.47}{100}$

$= 35.50$

$\therefore$ Relative atomic mass of Cl = 35.57

4.0 CONCLUSION

In this unit you have known what is a mass spectrometer and its operation. You have also known what is meant by isotopes. You have known how to calculate the relative atomic masses of elements.

5.0 SUMMARY

Mass spectra records the chemical constituents of a substance separated according to their mass and presented as a spectrum as measured by a mass spectrometer.

Mass spectrometer measures the charge to mass ratios.

Isotopes are elements with same atomic number but different neutrons number.

6.0 TUTORS MARKED ASSIGNMENT

1. Naturally occurring Strontium contains ^{84}Sr, ^{86}Sr, ^{87}Sr and ^{88}Sr in relative abundance of 0.56%, 9.86%, 7.02%, 82.56% respectively. Calculate the relative atomic mass of Strontium.(**Ans. 87.71)**
2. The isotopes of iron are ^{54}Fe, ^{56}Fe, ^{57}Fe and ^{58}Fe in relative abundance of 5.84%, 91.68%, 2.17% and 0.31% respectively. Calculate the relative atomic mass of iron.(**Ans. 55.91)**
3. Explain any 4 types of radio isotopes by origin

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UNIT 3 ATOMIC MODELS

CONTENTS

1.0 INTRODUCTION

In this unit we treat the various models of an atom which were presented by different scientists after its discovery. An atomic model is a way the scientist gives its own picture of how an atom looks like. The various models of an atom J. J Thomson model or Plum pudding model, Rutherford's model, Electron Cloud model, Bohr's model etc

2.0 A OBJECTIVE

At the end of this unit you are expected to describe the various atomic models.

2.0 B How to Study this Unit

1. You are expected to read carefully through this unit twice before attempting to answer the activity questions. Do not look at the solution or guides provided at the end of the unit until you are satisfied that you have done your best to get all the answers.
2. Share your difficulties in understanding the unit with your mates, facilitators and by consulting other relevant materials or internet.
3. Ensure that you only check correct answers to the activities as a way of confirming what you have done.
4. Note that if you follow these instructions strictly, you will feel fulfilled at the end that you have achieved your aim and could stimulate you to do more.

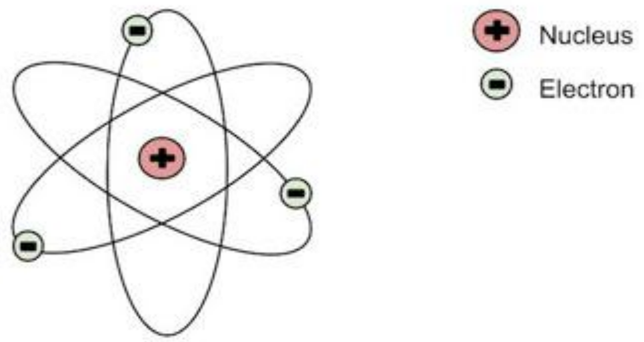
3.0 MAIN CONTENT

3.1 Atomic models

Atomic models are various pictures of atom which scientists presented how an atom looks like. Some of these models include:

Rutherford Model

In many ways, the **Rutherford model of the atom** is the classic model of the atom, even though it's no longer considered an accurate representation. Rutherford's model shows that an atom is mostly empty space, with electrons orbiting a fixed, positively charged nucleus in set, predictable paths.



Rutherford Model of the Atom

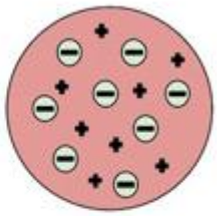
This model of an atom was developed by **Ernest Rutherford**, a New Zealand native working at the University of Manchester in England in the early 1900s. Rutherford spent most of his academic career researching aspects of radioactivity, and in 1908, won the Nobel Prize for his discoveries related to radioactivity. It was after this that Rutherford began developing his model of the atom.

Discovery of the Atom

The atom was first conceived of by the Greek philosopher **Democritus** in approximately 400 BCE. The concept was lost during the Dark Ages of Europe, until 1803, when the British scientist **John Dalton** speculated that everything was composed of very tiny indivisible particles called atoms.

Dalton's simple model of an atom persisted until 1897, when another British physicist, **J.J. Thomson**, discovered that atoms contained tiny negatively charged particles called electrons. From 1897 to 1909, scientists thought that atoms were composed of electrons spread uniformly throughout a positively charged matrix. J.J. Thompson's model was known as the **plum pudding model**.

Dalton's Model of the Atom
(1803)

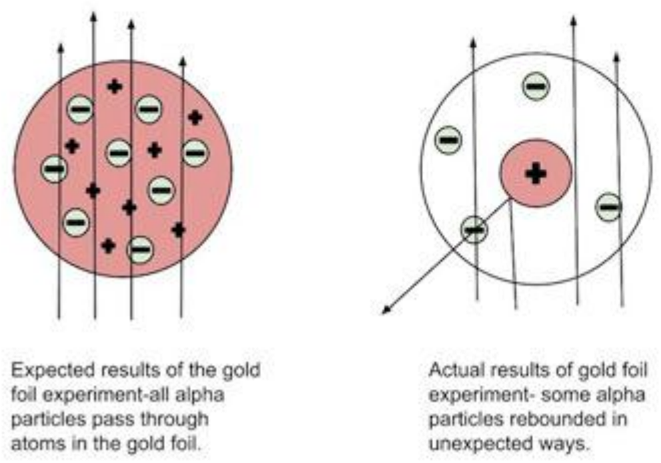


Plum Pudding Model of the Atom
(1897)

Dalton's model of the atom depicted a tiny, solid, indivisible sphere. Thompson's plum pudding model shows electrons (the green circles) distributed in a positively charged matrix.

Development of the Rutherford Model

In 1909, Rutherford conducted his famous **gold foil experiment**. In the experiment, Rutherford and his colleague Hans Geiger bombarded a piece of gold foil with positively charged alpha particles, expecting them to travel straight through the foil. Instead, many alpha particles ricocheted off of the foil, suggesting that there was something positive these particles were colliding with. They named this positive force the **nucleus**. The Rutherford Model was created based on this new data.

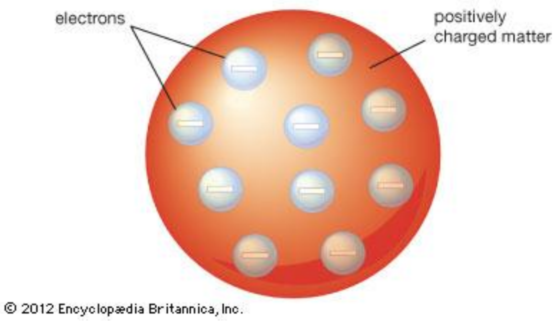


This diagram depicts the expected and the actual results of the gold foil experiment. The diagram on the left shows particles passing through the positively charged matrix of the plum pudding model. The diagram on the right shows particles ricocheting off of the nucleus in the center of the atom.

Problems with the Rutherford Model

In the years after Rutherford discovered the nucleus, chemists and particle physicists discovered that electron behavior was much more complicated than depicted in the Rutherford model. Electrons did not travel in set paths, their speeds were inconsistent, and their location around the nucleus could change based on how much energy they had. It was no longer accurate to depict electrons as traveling in straight paths. Instead, physicists began to represent them by an electron cloud that could suggest where electrons might be at any given time. The electron cloud model is the current model of the atom.

Thomson atomic model, earliest theoretical description of the inner structure of atoms, proposed about 1900 by Lord Kelvin and strongly supported by Sir Joseph John Thomson, who had discovered (1897) the electron, a negatively charged part of every atom. Though several alternative models were advanced in the 1900s by Lord Kelvin and others, Thomson held that atoms are uniform spheres of positively charged matter in which electrons are embedded. Popularly known as the plum-pudding model, it had to be abandoned (1911) on both theoretical and experimental grounds in favour of the Rutherford atomic model, in which the electrons describe orbits about a tiny positive nucleus.



Other models include **Electron cloud** and **Bohr's models** etc.

CONCLUSION

We have in this unit learnt various atomic models as viewed by different scientists.

5.0 SUMMARY

The various atomic models are J. J Thomson model or Plum pudding model, Rutherford's model, Electron Cloud model, Bohr's model etc.

6.0 TUTORS MARKED ASSIGNMENT

1. Briefly describe Rutherford's model of an atom.
2. Explain the plum pudding model of an atom.

7.0 REFERENCES/FURTHER READING

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UNIT 4 BOHR'S MODEL OF AN ATOM

CONTENTS

1.0 INTRODUCTION

Of all the atomic models, Bohr's model was the widely accepted model. In this unit, we will treat Bohr's model briefly by outlining the main features of his model.

2.0 A OBJECTIVES

At the end of this unit, you should be able to:

- outline bohr's theory of an atom; and
- define orbital and angular momentum of an atom

2.0 B How to Study this Unit:

1. You are expected to read carefully through this unit twice before attempting to answer the activity questions. Do not look at the solution or guides provided at the end of the unit until you are satisfied that you have done your best to get all the answers.
2. Share your difficulties in understanding the unit with your mates, facilitators and by consulting other relevant materials or internet.
3. Ensure that you only check correct answers to the activities as a way of confirming what you have done.
4. Note that if you follow these instructions strictly, you will feel fulfilled at the end that you have achieved your aim and could stimulate you to do more.

3.0 MAIN CONTENT

3.1 Bohr's Model

Bohr's model of an atom was widely accepted as the correct model of an atom. Some of its basic revolutionary suggestions of an atom are:

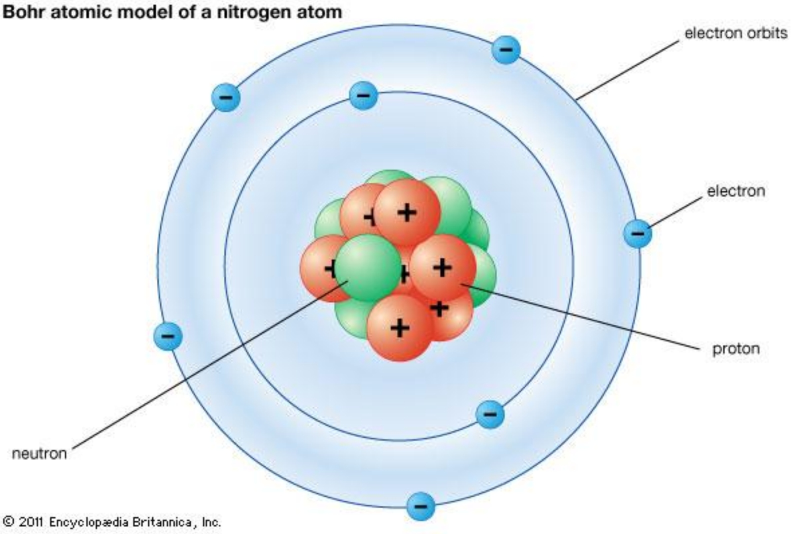
- Electrons can revolve round the nucleus in only in certain **allowed orbits** and while they are in these orbits they do not emit radiations (energy). Each orbit has a fixed amount of energy (has kinetic energy due to motion and potential energy due to attraction of the nucleus.)

- The **allowed orbits** are those for which the orbital angular momentum is equal to an integral multiple of $\hbar$ ($\hbar = \frac{h}{2\pi}$) and h is the Planck's constant. L = mvr = n$\hbar$ where n = 1, 2, 3, 4, and n is called **principal quantum number.**
- An electron can **jump** from the orbit
- of energy E_2 to another of energy E_1 and the energy difference is emitted as one quantum of radiation of frequency f given by Planck's equation:

$$E_2 - E_1 = hf$$

Bohr atomic model, description of the structure of atoms, especially that of hydrogen, proposed (1913) by the Danish physicist Niels Bohr. The Bohr model of the atom, a radical departure from earlier, classical descriptions, was the first that incorporated quantum theory and was the predecessor of wholly quantum-mechanical models. The Bohr model and all of its successors describe the properties of atomic electrons in terms of a set of allowed (possible) values. Atoms absorb or emit radiation only when the electrons abruptly jump between allowed, or stationary, states. Direct experimental evidence for the existence of such discrete states was obtained (1914).

Bohr atomic model of a nitrogen atom



4.0 CONCLUSION

In this unit you have learnt Bohr's revolutionary suggestions of the hydrogen model of an atom. The angular momentum of an allowed orbit, and Planck's energy equation.

5.0 SUMMARY

The Bohr's theory of an atom are:

- Electrons can revolve round the nucleus in only in certain **allowed orbits** and while they are in these orbits they do not emit radiations (energy). Each orbit has a fixed amount of energy (has kinetic energy due to motion and potential energy due to attraction of the nucleus.)
- The **allowed orbits** are those for which the orbital angular momentum is equal to an integral multiple of $\hbar\ (\hbar = \frac{h}{2\pi})$ and h is the Planck's constant.

  $L = mvr = n\hbar$ where n = 1, 2, 3, 4, and n is called **principal quantum number.**
- An electron can **jump** from the orbit of energy E_2 to another of energy E_1 and the energy difference is emitted as one quantum of radiation of frequency f given by Planck's equation:

$$E_2 – E_1 = hf$$

6.0 TUTORS MARKED ASSIGNMENT

1. Describe Bohr's theory of the hydrogen atom.
2. Define orbital angular momentum of an atom.

7.0 REFERENCES/FURTHER READING

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UNIT 5 HYDROGEN SPECTRA

CONTENTS

1.0 INTRODUCTION

In this unit you will learn how to use Bohr's energy equation for an orbit to calculate the energy of any obit and the transition energy of an exited particle.

2.0 A OBJECTIVE

At the end of this unit you will be able to calculate the energy of any given orbit and the transition energy.

2.0 B How to Study this Unit

1. You are expected to read carefully through this unit twice before attempting to answer the activity questions. Do not look at the solution or guides provided at the end of the unit until you are satisfied that you have done your best to get all the answers.
2. Share your difficulties in understanding the unit with your mates, facilitators and by consulting other relevant materials or internet.
3. Ensure that you only check correct answers to the activities as a way of confirming what you have done.
4. Note that if you follow these instructions strictly, you will feel fulfilled at the end that you have achieved your aim and could stimulate you to do more.

3.0 MAIN CONTENT

3.1 Hydrogen Spectra

These are horizontal lines drawn one above the other to represent the energy transition (or energy levels) in increasing order in a hydrogen atom.

A transition is a jump from one energy level to another.
Bohr derived a formula for energy of an electron in any energy level as:

$$E_n = -\frac{me^4}{8\varepsilon_0^2 h^2 n^2}$$

Where m = mass of an electron
e = charge of an electron
ε_0 = Permittivity of free space or vacuum
h = Planck's constant

Recasting the above formula i.e $E_n = -\frac{me^4}{8\varepsilon_0^2 h^2} X \frac{1}{n^2}$
If m = 9.1 X 10^{-31} Kg
e = 1.6 X 10^{-19} Coloumb
ε_0 = 8.85 X 10^{-12} farad per meter
h = 6.6 X 10^{-34} Js

Then $E_n = -\frac{9.1X10^{-31}(1.6X10^{-19})^4}{8(8.85X10^{-12})^2(6.6X10^{-34})} X \frac{1}{n^2}$

$$E_n = -\frac{2.179X10^{-18}}{n^2} Joules$$

But one electron volt (1eV) = 1.6 X 10^{-19} joule

Therefore the energy of an electron in any energy level is:

$$E_n = -\frac{2.179X10^{-18}}{1.6X10^{-19}} X \frac{1}{n^2} eV$$

$$E_n = -\frac{13.61875}{n^2} eV$$

For n = 1, E1 = -13.62 eV
For n = 2, E2 = -3.41 eV
For n = 3, E3 = -1.51 eV and so on.

The energy levels in the spectral of hydrogen atom as presented by Bohr are as follows:

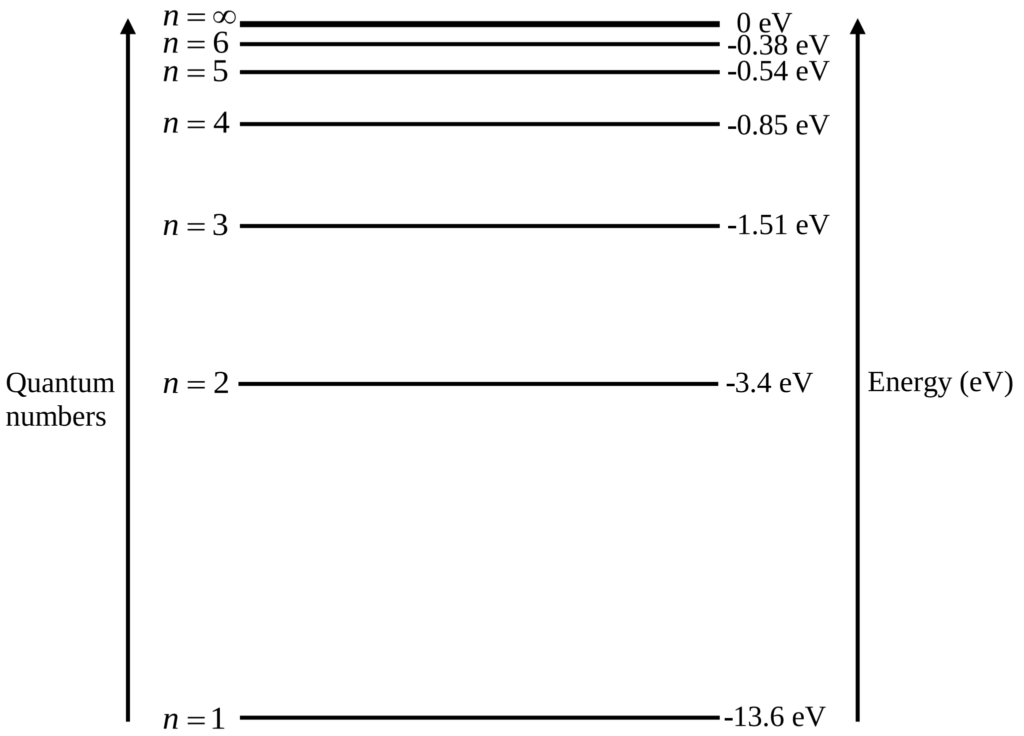

Fig.3 Hydrogen Spectra

3.2 Solved examples

1. Calculate the frequency of electromagnetic radiation emitted by a hydrogen atom which undergoes a transition between energy levels of -1.36 X 10^{-19} J and -5.45 X 10^{-19} J. (Take Planck's constant h = 6.6 X 10^{-34} Js)

Solution

Using $E_2 - E_1 = hf$

$1.36 \times 10^{-19} - (-5.45 \times 10^{-19}) = 6.6 \times 10^{-34} \times f$

$$f = \frac{-1.36X10^{-19} + 5.45X10^{-19}}{6.6X10^{-34}} Hz$$

$$f = 6.2X10^{14} Hz$$

2. The three lowest energy levels of a fictitious atom are shown in fig. 4 below

3 ______________________________ 1.8 eV
2 ______________________________ 4.0 eV
1 ______________________________ 16.0 eV

Fig. 4: Lowest energy levels of a fictitious atom

i. Determine the minimum energy required in joules to eject an electron in the lowest state from the atom
ii. Assuming the energy level n has energy k / n^2 , determine the energy of level n = 4 in electronvolts
iii Determine the wavelength of the radiation associated with a transition from level n = 2 to n = 3.

Solution

i. Using $E_\infty - E_0 = hf$

$$0-(-16.0eV) = hf$$

$$16eV = hf = E$$

But 1 eV = 1.6 X 10^{-19} J

$$E = 16X1.6X10^{-19} j$$
$$E = 2.56X10^{-18} j$$

ii
$$E = -\frac{K}{n^2}$$
$$E = \frac{-16.0}{n^2}eV \quad where\ k = -16.0eV$$
$$but \quad n = 4$$
$$E = \frac{-16.0}{4^2}$$
$$E = \frac{-16.0}{16}eV$$
$$E = -1.0eV$$

iii.
$$E_2 - E_3 = hf = \frac{hc}{\lambda} \qquad since \quad c = f\lambda$$
$$-1.8eV - (-4.0eV) = \frac{6.6X10^{-34}X3X10^8}{\lambda}$$
$$-1.8eV + 4.0eV = \frac{1.98X10^{-25}}{\lambda}$$
$$2.2eV = \frac{1.98X10^{-25}}{\lambda}$$
$$\lambda = \frac{1.98X10^{-25}}{2.2eV}$$
$$\lambda = \frac{1.98X10^{-25}}{2.2X1.6X10^{-16}}$$
$$\lambda = 5.625X10^{-7}m$$

4.0 CONCLUSION

In this unit you have learnt how to:

a) Calculate the energy level of any orbit if the quantum number n is known
b) Calculate the energy of transition.

5.0 SUMMARY

Hydrogen spectra are horizontal lines drawn one above the other to represent the energy levels in increasing order in a hydrogen atom.

Transition energy is the energy an electron gives out or absorbs when it jumps to another energy level.

6.0 TUTORS MARKED ASSIGNMENT

1. The longest wavelength in a Lyman series of hydrogen spectra is due to an electron transition from the first excited state -3eV to the ground state -13.6 eV. Calculate the wavelength. (h = 6.6 X 10^{-34} Js , c = 3 X 10^{8} m/s, 1eV = 1.6 X 10^{-19} J).
2. In Balma series of hydrogen, the longest wavelength, 6.6 X 10^{-5} cm, is due to an electron transition from the second excited (n = 3) to the first excited state (n = 2), -3.4 eV. Calculate the energy of the second excited state. (h = 6.6 X 10^{-34} Js, c = 3 X 10^{8} m/s, 1 eV = 1.6 X 10^{-19} J).
3. The fig. 5 below shows some of the energy levels in the hydrogen atom. E_0 is the ground state.

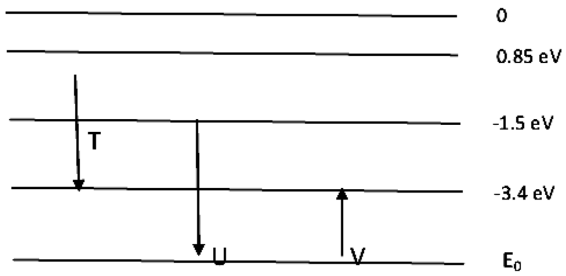

Fig. 5

(a) Calculate the wavelength of the photon emitted when an electron falls from energy level 4 to 2 as shown by T.
(b) When an electron makes the energy U, it emits a photon of wavelength 1.02 X 10^{-7} m. Calculate the energy E_0.

7.0 REFERENCES/FURTHER READING

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