History of atomic theory

The atomic theory is developed since 2000 years ago the Greek philosopher Democritus proposed that there was a limit to how small one could be divide matter, this smallest matter indivisible particle was called “ atom”. However this atomic theory of Democritus was criticized by Aristotle who proposed a model based on four basic “ elements” of earth, air, fire and water.

Aristotle’s view held for the next 2000 years as it better suited religious beliefs of the time. In 1801, an English teacher named John Dalton proposed his atomic theory which stated matter is composed of all small indivisible atoms, elements contain one type of atom; different elements contain different atoms, compounds contain more than one type of atom. In 1904, British physicist J. J Thomson and others demonstrated that cathode rays (electrons) were present in all matter.

Thomson proposed that the atom was a sphere of positive charge in which embedded were rings of negative charges (electrons) , like “ plums in a pudding”. Lather on, in 19 century, the discoveries of radiation leads to a new progress of Physics. We knew that certain elements emitting radiation, this suggested that atoms are no longer indivisible and not indestructible, as proposed in Dalton’s atomic model.

Radiation provides an important tool for the study of matter. In 1911, Rutherford first suggested the use of alpha particles to probe the internal structure of the atom. Finally, the nucleus and its protons were discovered.

Rutherford’s scattering experiment

Ernest Rutherford, directed an experiment to Hans Geiger and Ernest Marsden in 1909, in which the newly discovered alpha particles (Helium nuclei) were fired at a thin gold foil layer which only a few atom thick. At that time the atom was thought to be analogous “ plum pudding” by Thomson with a negative charge (the plum) throughout the positive sphere (the pudding).

Most of the alpha particle passed through with no or only very small deflections in a vacuum (see figure), as would been excepted on the Thomson model of the atom current at the time. About 1 in 8000 was deflected through angles greater than 90 degree. The result was so unexpected that Rutherford was very unexpected that Rutherford was promoted to write….” it was almost incredible as if you fired a 15 inch shell at a piece of tissue and it came back and hit you.”

Therefore Rutherford concluded that majority of the mass of the atom was concentrated in the nucleus. The small size of the nucleus explained the small number of the alpha particles that were repelled each other in this way. Rutherford showed that the size of the nucleus was less than about 10^{-14} m.

Bohr’s model, how it fit experimental observation

Accelerated electrons emitted and lose energy which predicted by Predicted by Maxwell and confirmed by Hertz, however the electron doesn’t spin into the nucleus because of energy losing.

Later on, In 1903. H. G. J Morseley found simple, regular relationship between the frequencies of X-ray emission line, thus it provided an evidence to support Bohr’s model. Scientist began to work on an alternating model to replace it.

Bohr’s postulates

In 1913 the Danish physicist Niels Bohr (1885~1992), put forward some radical propositions to account for the discrepancies between Rutherford’s model of the atom and the available experimental evidence. Bohr’s postulates are

1. Electron can remove in certain allowed orbits—stationary states (energy)—without radiating energy.

2. when an electron falls from a higher energy level to a lower energy level, it emits energy that is quantised by the plank relationship $E_2 – E_1 = hf$.

3. Augular momentum (mvr) is quantised and can only take values of the nh / 2π when n is the principal quantum number.

The first postulate account for the stability of the atom. However why the these stationary state excited was unknown. They exist was a fact.

The second postulate explains the line emission spectra. Emission (or absorption) of Energy is discontinuous and corresponds to a transition between two stationary states. Since the energy can be quantized, the emission, the frequency of the emitted (or absorbed) radiation is predetermined. A transition between different states will lead to difference frequencies or colours.

The third postulate effectively sets limits on the radius of the allowed orbits.

Bohr’s model

Bohr realised that if his model was correct, each atom would have a spectral fingerprint to the differences between electron energy level in that atom.

The Rydberg equation which also known as Balmer equation, has given him evidence for the quantised emission of the energy from hydrogen atom, leading him to going on to further his model and define his postulate.

So the hydrogen spectrum was very significant to the development of Bohr's model of the atom, because without the understanding of it, Bohr may not continue his work of the model.

Produced and observable when hydrogen gas was excited by the addition of energy. The equation in the original form was modified by Rhydberg until it worked and could be applied to explain the spectrum of hydrogen by using integer values of n, only as suggested by Bohr in his postulates.

Quantum number and quantum changes

It's possible to determine the energy of each orbit using Bohr's model and from this construct an energy level Figure for hydrogen. The figure below shows the energy these energy levels. Alternatively a transition between stationary states can be show in figure. B. The Balmer series of lines occurs when the electrons fall to the n= 2 level from n= 2 level n= 3, 4, 5 and 6 levels. This is illustrated differently in figure. B.

How Bohr describes the hydrogen spectra

Bohr's model of the atom wad quite similar to that of Rutherford's with two important differences firstly, it assisted positions to the electrons, but secondly the electron energy level s were quantised.

This was radically new, the idea that electrons had energy states and could absorb and emit energy to change states, and had no evidence. Bohr realised that if his model was correct, each atom would have a spectral " fingerprint to the differences between electron energy level in that atom.

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The energy levels describe by Bohr is clearly marked. According to Bohr, the Balmer series (shown on the bottom of the diagram as the hydrogen spectrum) was cause by changing energy levels, in the process releasing light. As shown, larger energy changes produce more energetic photons, as seen in Balmer's series, as further, this diagram shows how the Balmer series is formed by successive electron transition to the 2 nd shell (transition to other shell produce additional lines named after their discoveries.)

This is a great achievement that the Bohr ' s model is able to provide a physical basis for the Balmer series formula. From his second postulate E_f – E_i = hf. (i) states for initial energy level (f)states for final energy level

$E_i = 1/n_i^2 E_1$ and $E_f = 1/n_f^2$ E1

hence: $hf = 1/n_f^2 E_1 - 1/n_i^2 E_1 = (1/n_f^2 - 1/n_i^2) E_1$

And since c= fA=> A= c/f

the expression reduces to $1/A = E_1/hc X(1/n_f^2 - 1/n_i^2)$

where R states for Rydberg's constant, R_H (hydrogen)1. 097X 10^7 m^{-1} .

By using the mixture of classical physics and quantum physics, Bohr was able to define the equation for the spectral lines of hydrogen. He didn’t know why the electrons obeyed his rules. These were purely empirical results.

Problems with the model

For all the success, the Bohr model of the atom had serious limitations: It and ad hoc mixture of classical and quantum physics; it allows some laws of the classical physics held and others did not. Hydrogen has only one electron, and Bohr’s postulate are only able to explain it. It can’t work fo r multi-electron atoms.

It could not explain the relative intensities of the spectral line; some lines were more intense to the others and it was not known why this should occur. Certain spectral line were foun of a number of very fine and close lines and the cause of these hyperfine spectral lines could not be explained. The splitting of spectral lines when the sample was placed in a magnetic field (called the Zeeman effect, and discussed below) could also not be explained.

The postulates faced a problem that it is suited for larger atom. Hydrogen is the simplest atom containing only one electron. Similarity He^{+} and Li^{+} have one electron. Bohr’s model works with theses atom and ions. In all the other atoms however the electron interact with each other. In a larger atoms the outer electrons are shielded from the nucleus by the inner electrons. Interaction between electrons also result in different energy levels. The affect Bohr’s model to the extent that the spectra of multiple electrons could not be explain.

When the spectrum of the hydrogen was examined it was noted that the emission line varied in intensity. Some were quite intense and others were less intense; some were sharp and some were boarder. The following figure. illustrate these differences. Bohr's model could not explain these features but later it was explained that electron orbited in a ellipse and not in a circles. As the developing of the light spectroscopes improved it was found that some of the spectral " lines" were made up with hyperfine lines. This suggestion spitted Bohr's energy level theory; however there was no explanation for this.

The Zeeman Effect

Zeeman Effect occurred when a magnetic field us pass through a discharge tube. The magnetic field increased the hyperfine splitting of spectral lines, further breaking them up. As the limitation, Bohr's model was unable to explain the experimental evidence.

In 1896 a Dutch physics Pieter Zeeman (1865-1943) found that when he placed a source of sodium light between the poles of a strong magnet the lines split into three or more. This could not explain by Bohr's model, The spectral line of some elements can even split to 15 lines. This is called the anomalous Zeeman effect. It can't be explain by that time, and it leads to the new developing of the model to explain it. This begin with the work of de Broglie.

The following is the formal definition of Zeeman Effect: The splitting of single spectral lines of an emission or absorption spectrum of a substance into three or more components when the substance is placed in a magnetic field.

The effect occurs when several electron orbits in the same shell, which normally have the same energy level, have different energies due to their different orientations in the magnetic field. A *normal Zeeman Effect* is observed when a spectral line of an atom splits into three lines under a magnetic field. Astronomers can use the Zeeman Effect to measure magnetic fields of stars. The following diagrams shows the normal spectral line and the Zeeman effect.

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