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The Development of Atomic Structures by Dalton, Thomson Rutherford and Bohr, and their Mathematical Equations

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Abstract

Thomson's atom is a solid ball or billiard ball with a positive charge that contains several negatively charged particles or electrons. These electrons will be spread on the ball like raisins on bread. The main difference between Thomson's and Rutherford's atomic models is that Thomson's model does not contain information about the atomic nucleus, while Rutherford's model does. The theory of atomic structure helps scientists understand why elements behave in certain ways in chemical reactions. For example, electron configuration determines how elements bond and form compounds. In this paper, a literature review was conducted on the development of Thomson's atomic structure model. The study method was carried out to identify elements based on their atomic number, determine their reactivity based on the number of valence electrons, and understand how atoms unite to form molecules through chemical bonds. The results of the study, by studying atomic theory, can find out about the chemical and physical properties, as well as the uses of particles or substances that exist around the universe.

Keywords: Atoms, Thomson, Rutherford, chemical properties, physical properties, substances around nature.

1. Introduction

Atoms are the smallest particles that are the basic components of all objects in the universe (Wendt and Wendt, 2010). Atoms cannot be divided into smaller parts without losing their chemical properties, so they are often considered the fundamental units that make up matter. In the context of physics and chemistry, atoms are structural components that form objects with certain structures and properties. This understanding has undergone significant development along with the advancement of science, especially through the theories of atomic models put forward by leading scientists such as Dalton, Thomson, and Rutherford (Bailey and Bailey, 2022).

In the early 19th century, John Dalton introduced the first modern atomic theory which stated that atoms are basic particles that cannot be divided further. According to Dalton, atoms are the smallest entities of an element that retain the identity and chemical properties of the element (Kragh, 2019). Dalton also asserted that atoms of different elements have different sizes and masses, and that these atoms can combine to form chemical compounds.

Along with the development of technology and experimental methods, J.J. Thomson in the late 19th century discovered that atoms are not completely indivisible (Gooday, 2020), but consist of negatively charged particles called electrons. Thomson described the atom as a positively charged sphere with electrons scattered around it, like raisins in a bun. This model is known as the "raisin bun model". This theory paved the way for a deeper understanding of atomic structure, but still left many unanswered questions about the arrangement of electrons and the distribution of positive charge in the atom.

Ernest Rutherford, through his alpha particle scattering experiment, made a breakthrough in understanding atomic structure (Leone et al., 2018). Rutherford discovered that most of the mass of an atom is concentrated in a small, positively charged nucleus at the center of the atom, while electrons move around the nucleus in empty space. Rutherford's atomic model replaced Thomson's model and became the basis for the development of modern atomic theory (Barrette, 2021). This finding confirmed that atoms are not solid particles, but are mostly composed of empty space with a small, very massive nucleus.

In addition to basic theories about atomic structure, the development of mathematical formulas in chemistry and physics also played an important role in understanding atomic behavior. Atomic formulas are used to explain the relationships between atoms in chemical reactions, as well as to describe the mathematical processes that occur in the reaction. For example, chemical reaction equations must satisfy the principle of equivalence, namely the number of reactant atoms on both sides of the equation must be the same. This is achieved by adding coefficients to the substances involved without changing their atomic indices.

This study aims to review in more depth the development of atomic theory from Dalton to Rutherford, and analyze the structure of the atom and the associated mathematical equations. By understanding these theories in detail, we can gain a better understanding of how the atom, as the basic unit of matter, has been the focus of scientific study for centuries and continues to make important contributions to the development of science.

2. Method

The method used in this study involves reviewing various literature, including books and scientific papers that have been published in academic journals. The purpose of this study is to identify elements based on their atomic number, analyze the level of reactivity of the element by referring to the number of valence electrons, and understand the mechanism of molecule formation through the chemical bonding process between atoms. In addition, this study also seeks to dig deeper into information about the effect of electron configuration on the physical and chemical properties of atoms, and how interactions between atoms form stable molecular structures. Through this approach, it is hoped that a comprehensive understanding of the behavior of atoms in various chemical reaction conditions can be obtained.

3. Results and Discussion

3.1. Definition of Atom

Atom is the smallest particle of matter that cannot be divided into smaller parts. The idea of atoms has actually existed since before Christ. One of the early figures who put forward the theory of atoms was Democritus. According to Democritus' thinking, a substance can be divided into smaller and smaller parts until it reaches the smallest part that cannot be divided anymore, which is called an atom. The term "atom" itself comes from the Greek word "atomos," which literally means "uncuttable" or "undivided." This concept is the basis for the understanding that atoms are fundamental, inseparable units of matter, although this view later developed and underwent various revisions along with the advancement of science in the following period.

3.2. Atomic Structure

Atoms consist of three main subatomic particles such as protons, neutrons, and electrons. Protons and neutrons are in the nucleus of an atom, while electrons continue to move around the nucleus of an atom due to the influence of their electric charge. Protons are positively charged (+), neutrons have no charge (neutral), while electrons are negatively charged (-). Negatively charged electrons are attracted by positively charged protons in the nucleus of an atom, creating a balance that keeps the electrons spinning in their orbits (Fantz and Lettry, 2018).

Although electrons and protons have opposite charges, under normal conditions, the number of protons and electrons in an atom is the same, so the atom is neutral, neither positively nor negatively charged. However, if an atom has an excess or deficiency of electrons, then the atom becomes an ion. Ions are positively charged if they lose electrons, and negatively charged if they receive additional electrons. This phenomenon causes electrical interactions between atoms and affects how atoms bond together in the formation of molecules (Vallabhajosula, 2023).

Although protons are positively charged, it does not mean that all atoms in the universe are positively charged. Atoms remain neutral as long as the number of protons and electrons is balanced. However, if an atom becomes an ion, there is an imbalance of charge that causes atoms to attract or repel each other depending on the type of ion.

3.3. Development of Atomic Theory

The theory of atoms began to emerge long before Christ. One of the early figures who discussed the concept of atoms was Democritus. He stated that a substance can continue to be divided into smaller parts until it reaches the smallest particle that cannot be divided anymore, and this particle is called an atom. The word "atom" comes from the Greek "atomos," which means "uncuttable."

The scientific development of the atomic concept began in 1805 by John Dalton, followed by important discoveries from Thomson in 1897, Rutherford in 1911, and finally refined by Bohr in 1914. The results of experiments conducted by these scientists provide a clearer picture of the arrangement of particles in atoms.

3.4. Dalton's Atomic Model

In 1808, John Dalton, a British chemist, introduced an atomic theory stating that atoms are the smallest particles that make up matter (Bailey and Bailey, 2022). His ideas about atoms include several important points:

a). Atoms are the smallest particles that cannot be broken down further.
b). Atoms of the same element have identical properties, while atoms of different elements will have different masses and properties.
c). Compounds are formed when atoms of different elements combine.
d). Chemical reactions involve only the rearrangement of atoms without changing the atoms themselves.

Dalton's atomic model described matter as a collection of tiny atoms that could not be broken down any further. This theory eliminated the idea that matter could be divided indefinitely (continuously). Although we now know that atoms are made up of subatomic particles such as protons, neutrons, and electrons, atoms are still considered the smallest units that retain the properties of an element. Atoms are the smallest units involved in chemical reactions. If atoms were broken down into subatomic particles, then the element would no longer have the same identity as we know it today, and the process of breaking atoms into subatomic particles is very difficult to do.

Dalton's atomic model was the first scientific atomic model developed based on experimental results, such as the law of conservation of mass and the law of definite proportions. Although the idea of atoms as the smallest particles had been put forward long before Dalton, for example by the Greek philosopher Democritus around 400 BC, this view was not considered scientific because it was not supported by empirical evidence.

Figure 1: Dalton's atomic model

3.4.1. Advantages of Dalton's Atomic Theory

John Dalton was one of the first scientists to conduct experiments and research related to the concept of atoms. Although before Dalton, there were scientists such as Democritus who also put forward ideas about atoms, Democritus' thoughts were not based on the results of scientific experiments. Dalton, on the other hand, succeeded in introducing the concept of atoms with a more scientific basis of experiments and observations, making him a pioneer in the development of modern atomic theory. Dalton's theory states that two or more atoms, either from the same or different elements, can combine to form molecules. This is in accordance with the fact that molecules, both single elements and compounds, are formed through the combination of atoms. For example:

$$O_2, H_2, N_2, O_3, \text{and others}$$

For different elements

$$H_2\mathrm{O}, H_2SO_4, \mathrm{HCL}, \text{and others}$$

Dalton's theory provides a simple explanation of the structure of matter, stating that all matter is made up of tiny particles called atoms. Each element is made up of atoms that are identical in mass and properties, which distinguishes it from other elements. Dalton's theory supports and explains the law of conservation of mass, which states that the mass of a substance does not change during a chemical reaction. In addition, his theory explains the law of definite proportions, where compounds are always formed from elements in fixed mass ratios.

3.4.2. Thomson's Atomic Model

The weaknesses in Dalton's atomic theory were corrected by J.J. Thomson through experiments he conducted using a cathode ray tube. This experiment revealed the existence of negatively charged particles smaller than atoms, which were later known as electrons. This discovery changed the previous understanding that atoms were the smallest particles that could not be divided anymore (Nair, 2019).

According to Thomson's atomic model, atoms are not the smallest indivisible particles, but solid balls consisting of positive charges with negatively charged electrons spread throughout them. This positive charge is evenly distributed throughout the ball, while electrons are spread across its surface, similar to raisins in bread. Because of the balance between positive and negative charges, the atom as a whole is neutral.

This model later became known as the raisin bread model or plum pudding model, which describes how electrons are randomly distributed in a positively charged ball. Thomson's experiments succeeded in showing that atoms have a more complex internal structure than Dalton had thought, and that atoms are composed of different subatomic particles.

Although this model could explain some aspects of the atom, such as the presence of electrons and the neutral nature of the atom, Thomson's atomic model was unable to explain how the particles were organized in more detail or how the atom maintained its stability. This model was later refined by Ernest Rutherford who discovered the structure of the atomic nucleus and explained that positive charge is concentrated in the center of the atom, namely in the nucleus.

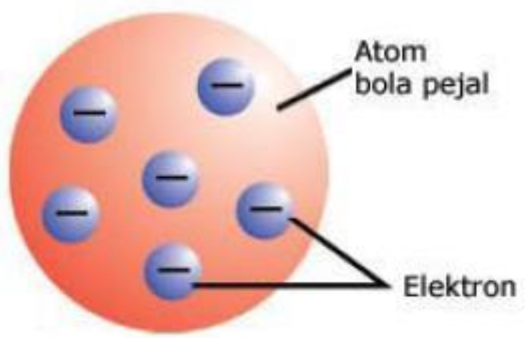

Figure 2: Thomson's atomic model

Based on the theory of other scientists such as Sir Humphry Davy, who showed that gases become better conductors of electricity at low pressure, J.J. Thomson conducted an experiment using a cathode ray tube. Sir Humphry Davy had proven that gases at low pressure were able to conduct electricity more efficiently, and this became the basis for Thomson's experiment.

In his experiment, Thomson used a vacuum tube, where the gas in the tube was at very low pressure. This tube is equipped with two electrodes, namely the cathode (negative electrode) and the anode (positive electrode), which produce a flow of charged particles when an electric current is passed through the gas. The cathode ray tube used by Thomson allowed him to observe negatively charged particles moving from the cathode to the anode, which were later known as electrons.

As seen in Figure 3, this tube is designed in such a way that cathode rays (electron flow) can be detected at the end of the tube, indicating that these negatively charged particles are actually the basic components of atoms. This experiment helped Thomson develop his atomic model and prove that atoms are not indivisible entities, but are composed of smaller particles such as electrons.

Figure 3: Cathode Ray Tube

When a high-voltage electric current is passed through a gas in a tube at low pressure and high temperature, the gas will glow, and the color of the glow depends on the type of gas in the tube. However, if the gas pressure is further reduced, the area in front of the cathode will begin to darken, and this dark area will expand as the gas pressure decreases. Eventually, the entire tube will be dark, but the part of the tube in front of the cathode will begin to glow greenish. Through experiments, it was proven that this greenish light came from particle radiation emitted by the cathode. Because this radiation comes from the cathode, it is called cathode rays. Further experiments showed that cathode rays actually consist of negatively charged particles called electrons. Thomson proved that cathode rays are not ordinary light, but a stream of negatively charged particles, and this is the first evidence of the existence of electrons. The various important events that occurred during the cathode ray experiment are as follows:

a). The cathode rays move straight from the cathode to the anode.

b). The cathode rays can turn a pinwheel, indicating that the rays have mass.
c). Cathode rays are deflected towards the positive pole of an electric or magnetic field, indicating that these rays are negatively charged.

From the results of his experiments, Thomson concluded that atoms consist of positively charged spheres in which negatively charged particles, namely electrons, are scattered within them. This is known as the Thomson atomic model or often called the plum pudding model, where the positive charge is likened to bread, and electrons are raisins scattered within it. In short, Thomson's atomic model explains:

a). Atoms consist of positive charges that are evenly distributed throughout the volume of the atom.
b). Electrons are attached to the surface of the positive charge at certain positions.
c). Electrons do not move around the nucleus, but remain in their positions.
d). Electrons vibrate at certain frequencies in their positions.
e). The mass of the atom is evenly distributed throughout the volume of the sphere.

Proof of the existence of negatively charged particles in atoms: Thomson succeeded in proving the existence of electrons, which means that atoms are no longer considered the smallest part of an element. This was an important advance in understanding atomic structure because it showed that atoms have subatomic components. No Thomson's model was unable to explain how positive and negative charges are arranged in atoms, so its structure is still unclear, does not explain chemical reactions between atoms: This model fails to explain how atoms interact chemically, especially how atoms form chemical bonds. when alpha particles were fired at a thin gold layer in Rutherford's experiment, many alpha particles penetrated the layer, showing that most of the space in the atom is empty. This contradicts Thomson's model which considers atoms to be solid balls that are full. This model is also unable to explain the complex emission spectrum lines of hydrogen atoms, even though hydrogen only has one electron.

3.4.3. Rutherford's Atomic Model

In 1911, Hans William Geiger and Ernest Marsden, under the supervision of Ernest Rutherford, conducted an alpha ray scattering experiment to test the truth of Thomson's atomic model. In their experiment, they used an alpha particle emitter directed through a small hole in a lead screen, producing a sharp beam of alpha particles. This beam was then directed at a thin gold foil. On the other side of the gold foil was a zinc sulfide (ZnS) screen that fluoresces when struck by alpha particles. The experiment produced several results that were inconsistent with Thomson's model: most of the alpha particles passed through the gold foil without deflection. However, some alpha particles were significantly deflected, and some even bounced back to the source of the alpha particles.

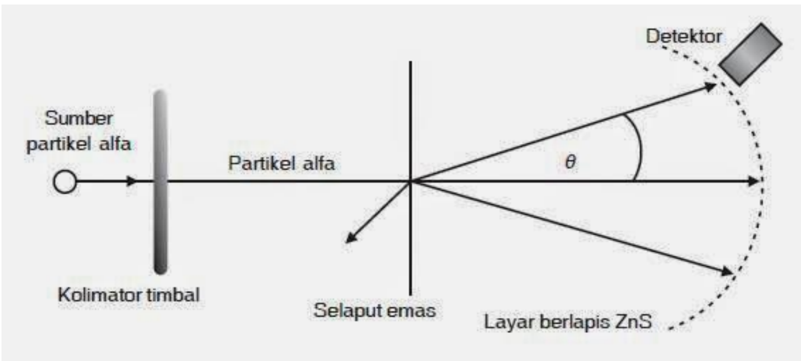

Figure 4: Source of the Alpha Particles

From their observations, it was found that when alpha particles were fired at a very thin gold foil, most of the alpha particles continued to travel with only a slight deviation (less than 1°). However, Marsden found that about one in 20,000 alpha particles experienced significant deflection, with an angle of up to 90° or even reflected back. This clearly contradicts Thomson's atomic theory, which describes atoms as solid and homogeneous structures in all parts. If the atom is really like the "raisin bun" model proposed by Thomson, the alpha particles should be evenly distributed, not reflected. Therefore, Rutherford's experimental results showed that Thomson's atomic model was unacceptable, because it could not explain this phenomenon. Rutherford's interpretation of this experiment was that atoms must have a positively charged nucleus with a large mass.

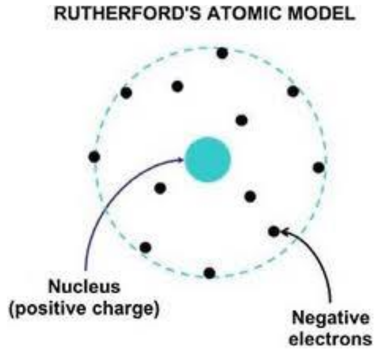

Figure 5: Rutherford's Atom Model

Rutherford's Hypothesis:

1) Atoms are mostly empty space with a positively charged nucleus at the center and one or more electrons orbiting it, much like planets orbit the sun in the solar system. Most of the mass of an atom is concentrated in the nucleus.
2) Overall, the atom is neutral. The positive charge in the nucleus is proportional to the number of electrons orbiting it, so the positive nuclear charge is equal to the atomic number multiplied by the elementary charge.
3) The nucleus and electrons attract each other. This attraction produces a centripetal force that controls the motion of electrons in their orbits, much like gravity controls the motion of planets in the solar system.
4) In chemical reactions, the nucleus of an atom is unchanged, only the electrons in the outermost shell of the atom change.

Based on the results of his extraordinary experiments, Rutherford put forward the concept of the atomic nucleus. According to him, the majority of the mass and positive charge of an atom are concentrated in the center of the atom, called the atomic nucleus. Electrons move around the nucleus at a considerable distance, and the distance between the nucleus and the atomic shell is known as the atomic radius.

With this model, the experiment of scattering alpha rays on a thin gold plate can be explained as follows:

a). Most alpha particles pass through the plate without obstruction because they pass through the empty area in the atom.

b). Alpha particles approaching the atomic nucleus are deflected due to the repulsive force of the positive charge of the nucleus.

c). Alpha particles moving directly towards the nucleus are reflected because the nucleus is very dense and positively charged. Rutherford's model is able to hypothesize that atoms consist of atomic nuclei and electrons surrounding them. While the weakness of this model is that it cannot explain the stability of the atom as a whole. Because electrons that move faster will emit electromagnetic radiation continuously, the energy of the electrons will decrease, causing the electrons to spiral towards the nucleus until they are destroyed in a very short time (10^{-8} seconds). However, in reality, the electrons remain stable in their orbits, toms should emit a continuous spectrum, but the observed spectrum of hydrogen atoms shows a line spectrum (Balmer series) and this model is unable to explain how electrons are distributed outside the atomic nucleus.

3.4.4. Bohr Atomic Model

To overcome the weaknesses of the Rutherford model, Bohr introduced four basic postulates in his atomic model:

a). The hydrogen atom consists of one electron moving in a circular orbit around the nucleus, and its motion is influenced by the Coulomb force according to the principles of classical mechanics.

b). A stable orbit for an electron in a hydrogen atom only exists if its angular momentum (L) is a multiple of Planck's constant divided by 2π

$$L = n\,.\,\hbar = n\,.\frac{h}{2\pi} \tag{1}$$

where n = 1,2,3,… and are called the principal quantum numbers, and h is Planck's constant.

c). In a steady orbit, an electron around the nucleus of an atom does not emit electromagnetic energy, in this case its total energy E does not change.

d). If an atom makes a transition from a high energy state $E_U$to a lower energy state E_I, a photon with energy $h\upsilon = E_U - E_I$ is emitted. If a photon is absorbed, the atom will transition from a low energy state to a high energy state.

”Bohr stated that electrons only occupy certain orbits around the atomic nucleus, each of which is associated with a multiple of some fundamental quantum value. (John Gribbin, 2002)”

Bohr's model of the hydrogen atom depicts negatively charged electrons orbiting the atomic shell in a certain path around the positively charged atomic nucleus. When electrons jump from one orbit to another, they are always accompanied by the emission or absorption of a certain amount of electromagnetic energy hf.

According to Bohr:

"There are quantum physics rules that allow only a certain number of electrons in each orbit. There is only room for two electrons in the orbit closest to the nucleus. (John Gribbin, 2005)"

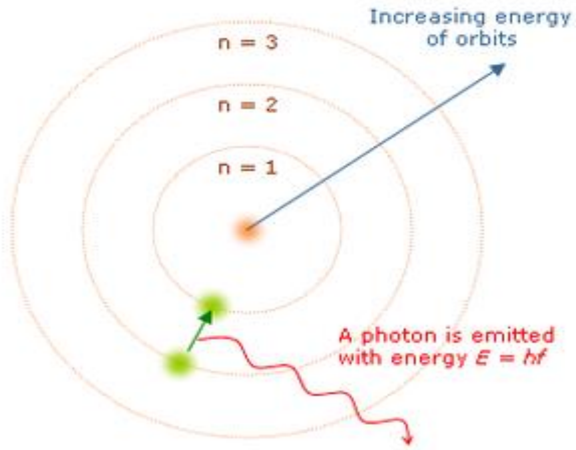

Figure 5: Bohr atomic model

This model is a development of several previous atomic models, such as the plum pudding model (1904), the Saturnian model (1904), and the Rutherford model (1911). Because the Bohr model is a refinement of the Rutherford model, many sources refer to it as the Rutherford-Bohr model. The main success of the Bohr model lies in its ability to explain the Rydberg formula for the spectral emission lines of the hydrogen atom. Although the Rydberg formula had been known through previous experiments, there was no theoretical basis that could explain it until the emergence of the Bohr model.

This model not only explains the structure of the Rydberg formula, but also provides justification for the experimental results in terms of fundamental physical constants. The Bohr model is essentially the basic model of the hydrogen atom. As a theory, this model can be considered the first approximation to the hydrogen atom using the more modern and accurate principles of quantum mechanics. Although the Bohr model is now considered obsolete in the context of the development of quantum mechanics, its simplicity and precise results for certain systems make it still taught as an introduction to understanding quantum mechanics.

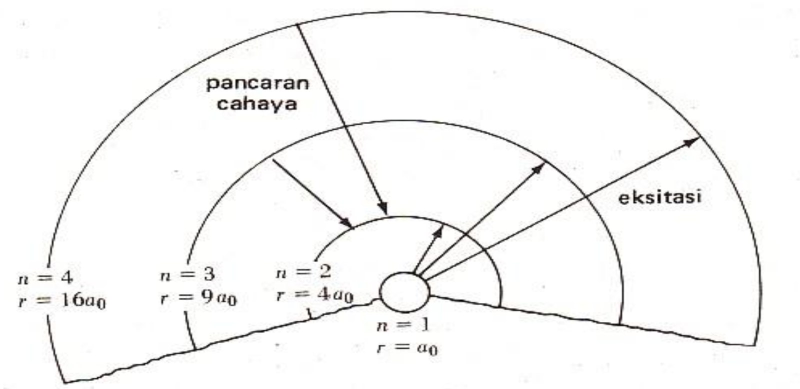

Figure 6: Bohr model for hydrogen atom

- The allowed orbits for electrons are numbered n = 1, n = 2, n = 3 etc. These numbers are called quantum numbers, the letters K, L, M, N are also used to name the orbits
- The orbital radius is expressed as 1^2, 2^2, 3^2, 4^2, …n^2. For a particular orbit with a minimum radius of a_0 = 0.53 Å

$$a_0 = \frac{4\pi\varepsilon_0\hbar^2}{me^2} \quad (2)$$

- If an electron is attracted to the nucleus and is possessed by an n orbit, energy is emitted and the electron's energy becomes lower by

$$E_n \frac{-B}{n^2} \tag{3}$$

Numeric constants with values $2.179 \times 10^{-18}\, J = -13.6 eV$

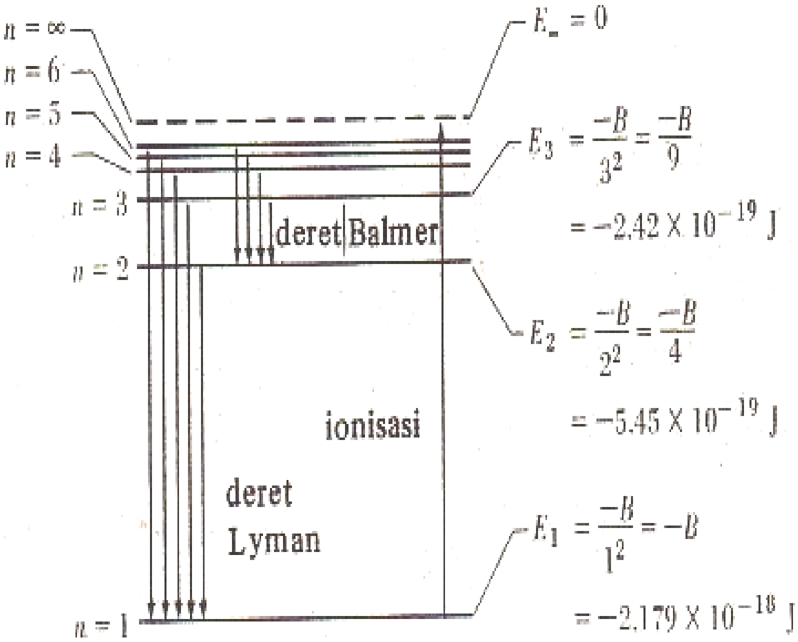

Figure 7: Energy levels of hydrogen atoms

3.5. Energy levels of electrons in hydrogen atoms

The Bohr model is only accurate for one-electron systems such as hydrogen or helium atoms that are ionized once. The derivation of the formula for the energy levels of the hydrogen atom uses the Bohr model.

The derivation of the formula is based on three simple assumptions:

1) The energy of an electron in orbit is the sum of its kinetic energy and potential energy:
$$E = E_{Kinetics} + E_{Potential} \tag{4}$$
$$=\frac{1}{2} m_e v^2 \frac{kq_e^2}{r}$$
with $k = 1/(4\pi\varepsilon 0)$, and q_e is the electron charge
2) The angular momentum of an electron can only have certain discrete values:
$$L = m_e v r = n\frac{h}{2\pi} = n\hbar \tag{5}$$
with n = 1,2,3… and is called the principal quantum number, h is Planck's constant, and $\hbar = h/(2\pi)$
3) Electrons are in orbits governed by the Coulomb force. This means that the Coulomb force is equal to the centripetal force:
$$\frac{kq_e^2}{r^2} = \frac{m_e v^2}{r} \tag{6}$$

By multiplying both sides of equation (3) by r we get:

$$\frac{kq_e^2}{r} = m_e v^2 \tag{7}$$

The term on the left hand side represents potential energy, so the equation for energy becomes:

$$E = \frac{1}{2} m_e v^2 - \frac{kq_e^2}{r} = -\frac{1}{2} m_e v^2 \tag{8}$$

By solving equation (2) for r, we obtain the permissible radius value:

$$r = \frac{n\hbar}{m_e v} \tag{9}$$

By plugging equation (6) into equation (4), we obtain:

$$kq_e^2 \frac{m_e v}{n\hbar} = m_e v^2 \tag{10}$$

By dividing both sides of equation (7) by mev we obtain

$$\frac{kq_e^2}{r^2} = \frac{m_e v^2}{r} \tag{11}$$

By inserting the value of v into the energy equation (equation (5), and then substituting the values for k and, the energy at different orbital levels of the hydrogen atom can be determined as follows:

$$\begin{aligned} E_n &= \frac{-1}{2} m_e \left(\frac{kq_e^2}{n\hbar}\right)^2 \\ &= \frac{-1}{2} m_e \left(\frac{1}{m_{\pi\epsilon 0}} \frac{kq_e^2}{n\hbar}\right)^2 \\ &= \frac{-m_e q_e^4}{8h^2 \epsilon_0^2} n^2 \end{aligned} \tag{12}$$

By entering the prices of all constants, we get,

$$E_n = (-13.6eV) \frac{1}{n^2} \tag{13}$$

Thus, the lowest energy level for a hydrogen atom (n = 1) is -13.6 eV. The next energy level (n = 2) is -3.4 eV. The third energy level (n = 3) is -1.51 eV, and so on. These energy values are negative, indicating that the electron is in a bound state with a proton. Positive energy values are associated with atoms that are in an ionized state, namely when the electron is no longer bound, but in a scattered state. With quantum theory, Bohr also found a mathematical formula that could be used to calculate the wavelengths of all lines that appear in the spectrum of a hydrogen atom.

The calculated values turned out to be very suitable with those obtained from direct experiments. However, for elements that are more complicated than hydrogen, Bohr's theory turned out to be unsuitable in predicting the wavelengths of spectrum lines. Nevertheless, this theory is recognized as a step forward in explaining physical phenomena that occur at the atomic level. Planck's quantum theory is recognized as true because it can be used to explain various physical phenomena that at that time could not be explained by classical theory.

a). Strengths and Weaknesses of Bohr's Theory
- The success of Bohr's theory lies in its ability to predict lines in the hydrogen atom spectrum
- One of the discoveries is a set of fine lines, especially if the excited atoms are placed in a magnetic field

b). Weaknesses
- This fine line structure is explained through a modification of Bohr's theory but this theory has never succeeded in describing the spectrum other than the hydrogen atom
- Not yet able to explain the presence of fine structure in the spectrum, namely 2 or more lines that are very close together
- Not yet able to explain the spectrum of complex atoms
- The relative intensity of each emission spectrum line.
- The Zeeman effect, namely the splitting of the spectrum line when the atom is in a magnetic field.

4. Conclusion

Atoms are the smallest species composed of protons (positive charge), neutrons, and electrons (negative charge). In its development, several atomic models and related theories were found. The atomic models include: John Dalton's atomic model, Thomson's atomic model, Rutherford's atomic model, Bohr's atomic model and modern atomic theory. According to Democritus, atoms are the smallest particles that cannot be divided again. According to Dalton, atoms are solid balls with a neutral charge. According to JJ. Thomson, atoms are positively charged balls around which electrons are scattered (raisin bread). According to Rutherford's theory, atoms are positively charged particles surrounded by electrons on certain and non-fixed trajectories. According to Neils Bohr's theory, atoms have stationary trajectories, where electrons surround protons on certain trajectories and energy levels. And finally, according to modern atomic theory, atoms are particles that have wave-like properties, which can change at certain energy levels and electrons surround protons on trajectories and sub-energy.

References

Bailey, J., & Bailey, J. (2022). History of the Atom, 1803–1932. *Inventive Geniuses Who Changed the World: Fifty-Three Great British Scientists and Engineers and Five Centuries of Innovation,* 173-201.

Bailey, J., & Bailey, J. (2022). History of the Atom, 1803–1932. *Inventive Geniuses Who Changed the World: Fifty-Three Great British Scientists and Engineers and Five Centuries of Innovation,* 173-201.

Barrette, J. (2021). Nucleus-nucleus scattering and the Rutherford experiment. *Journal of the Royal Society of New Zealand, 51*(3-4), 434-443.

Fantz, U., & Lettry, J. (2018). Focus on sources of negatively charged ions. *New Journal of Physics, 20*(6), 060201.

Gooday, G. (2020). Periodical physics in Britain: Institutional and industrial contexts, 1870–1900. *Science periodicals in nineteenth-century Britain: Constructing scientific communities,* 238-73.

Kragh, H. (2019). The periodic system and the idea of a chemical element: From Mendeleev to superheavy elements. *Centaurus, 61*(4), 329-344.

Kurniawan, A., & Firman, H. (2018). The Scientific Revolution in the Development of Atomic Physics: Is it Continuous or Discontinuous? *International Journal of Scientific & Engineering Research, 7*(6).

Leone, M., Robotti, N., & Verna, G. (2018). 'Rutherford's experiment'on alpha particles scattering: the experiment that never was. *Physics Education, 53*(3), 035003.

Nair, K. R. (2019). *Atomic Spectroscopy.* MJP Publisher.

Ngee, G. H. (2008). Using Open Ended Questions to Engage and Challange Mathematically Gifted Students.

Salazar R, E., Obaya V, A. E., Giammatteo, L., & Vargas-Rodríguez, Y. (2019). Evaluating a Didactic Strategy to Promote Atomic Models Learning in High School Students through Hake's Method. *Online Submission, 7*(5), 293-312.

Suryelita, S., Guspatni, G., & Defriati, P. (2019, October). Description of learning difficulties on atomic structure and periodic table topics of tenth grade students in SMAN 7 Padang. In *Journal of Physics: Conference Series (Vol. 1317, No. 1, p. 012147).* IOP Publishing.

Tro, N. J., Fridgen, T. D., Shaw, L., & Boikess, R. S. (2017). *Chemistry: A molecular approach (Vol. 5).* Boston: Pearson.

Vallabhajosula, S. (2023). Atoms and Radiation. In *Molecular Imaging and Targeted Therapy: Radiopharmaceuticals and Clinical Applications* (pp. 37-47). Cham: Springer International Publishing.

Wendt, S., & Wendt, S. (2010). Small, Smaller, Smallest–How the Components of Matter Were Found. *Roots of Modern Technology: An Elegant Survey of the Basic Mathematical and Scientific Concepts,* 241-265.