

BIOPHYSICS LECTURE

CHAPTER 2

LESSON 1: ATOMIC MODELS

J.J Thompson's "Plum Pudding Model" of the atom failed to explain experimental results associated with the atomic structure of elements.

Rutherford's atomic model is described by the positively charged particles where most of the mass of an atom was concentrated in the nucleus. Electrons surround the nucleus of an atom and revolve around it with at high speeds in circular paths called orbits. This model did not explain how electrons were distributed in the space surrounding the nucleus.

There is a need for a better model to explain why the electrons are not attached to the protons. And this is when the Bohr's model of the hydrogen atom came along.

HISTORY OF THE ATOM

- Atomic Theory – everything is made up of tiny particles that can be break up, separated by each other in an empty space
- 1800's – John Dalton described atoms as 'solid spheres'
- 1897 – J.J Thompson Plum Pudding Model; couldn't be solid spheres
- 1909 – Ernest Rutherford; Alpha particles, thin sheet of gold; proves JJ's theory wrong; Nuclear Model; flaw is that they didn't see anything stopping the atoms
- 1913 – Niels Bohr; suggested that electrons orbited the electrons; they were held in shells; this idea was important bcs the shells prevent the collapse
- Ernest Rutherford – Protons
- James Chadwick – Neutrons

BOHR ATOMIC MODEL

- Described the atoms as positively charged nucleus comprised of protons and neutrons surrounded by negatively charged electron orbiting the nucleus in atomic shells

$$K_n = \frac{1}{2} \cdot \frac{mv_n^2}{4\pi\epsilon_0 r^2} = \frac{22.069eV}{n^2} \quad \text{(Kinetic energies in Bohr model)}$$

$$U_n = -\frac{1}{4\pi\epsilon_0} \cdot \frac{mq^2}{4\pi\epsilon_0 r^2} = -\frac{44.138eV}{n^2} \quad \text{(Potential energies in Bohr model)}$$

The sum of these energies is:

$$E_n = -\frac{1}{4\pi\epsilon_0} \cdot \frac{mq^2}{8\pi\epsilon_0 r^2} \quad \text{(Total energies in Bohr model) or}$$

Where,

- m = mass of an electron ($m = 1.019 \times 10^{-31}$ kg)
- e = charge of the electron (1.602×10^{-19} Coulombs)
- h = Planck's constant (6.626×10^{-34} J.s = 4.136×10^{-15} eV.s)
- ϵ_0 = permittivity constant (8.854×10^{-12} F/m, F/m)

In this model, the electrons in an atom can only be found in certain discrete energy states. These states are associated with the orbits or shells of electrons in an atom. Only certain photon energies are allowed when electrons jump down from higher levels to lower levels, producing the hydrogen spectrum.



When an electron moves from a higher energy level to a lower one, a photon is emitted. The energy of the emitted photon is expressed by the equation below.

$$E = hf = \frac{hc}{\lambda}$$

where,

$f = \frac{c}{\lambda}$ is the frequency in Hertz

c = 3.00×10^8 m/s is the speed of light in vacuum

λ is the wavelength of light in meter

The Bohr model worked for explaining the hydrogen atom and other single electron system such as He^+ but...

- is unable to explain spectra of atoms with more than one electron
- there is no way of explaining why some lines is more intense than the other or

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1. Understanding the eBook Biophysics Lecture Notes Study And Exam Help
 - The Rise of Digital Reading Biophysics Lecture Notes Study And Exam Help
 - Advantages of eBooks Over Traditional Books
2. Identifying Biophysics Lecture Notes Study And Exam Help
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Biophysics Lecture Notes Study And Exam Help
 - User-Friendly Interface
4. Exploring eBook Recommendations from Biophysics Lecture Notes Study And Exam Help
 - Personalized Recommendations
 - Biophysics Lecture Notes Study And Exam Help User Reviews and Ratings
 - Biophysics Lecture Notes Study And Exam Help and Bestseller Lists
5. Accessing Biophysics Lecture Notes Study And Exam Help Free and Paid eBooks
 - Biophysics Lecture Notes Study And Exam Help Public Domain eBooks
 - Biophysics Lecture Notes Study And Exam Help eBook Subscription Services
 - Biophysics Lecture Notes Study And Exam Help Budget-Friendly Options

6. Navigating Biophysics Lecture Notes Study And Exam Help eBook Formats
 - ePub, PDF, MOBI, and More
 - Biophysics Lecture Notes Study And Exam Help Compatibility with Devices
 - Biophysics Lecture Notes Study And Exam Help Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Biophysics Lecture Notes Study And Exam Help
 - Highlighting and Note-Taking Biophysics Lecture Notes Study And Exam Help
 - Interactive Elements Biophysics Lecture Notes Study And Exam Help
8. Staying Engaged with Biophysics Lecture Notes Study And Exam Help
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Biophysics Lecture Notes Study And Exam Help
9. Balancing eBooks and Physical Books Biophysics Lecture Notes Study And Exam Help
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Biophysics Lecture Notes Study And Exam Help
10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Biophysics Lecture Notes Study And Exam Help
 - Setting Reading Goals Biophysics Lecture Notes Study And Exam Help
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Biophysics Lecture Notes Study And Exam Help
 - Fact-Checking eBook Content of Biophysics Lecture Notes Study And Exam Help
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
 - Integration of Multimedia Elements

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