The Nuclear Model - At the end of the 1800s, J.J. Thomson's discovery of the electron led to the ## understanding of atoms as containing subatomic particles.- *Example*: This model is often likened to a "plum pudding" model. ! [plum pudding.jpg](https://prod-files-secure.s3.us-west-2.amazonaws.com/8f3c221c-927a-482b-9747-5ece94d3f9dc/66a12b95-eaee-4abd-89db-a458b1fe3ad8/plum_pudding.jpg) ## Subtopic 2: **Rutherford's Experiment** - **Key Concept**: Rutherford's experiment involved bombarding a thin gold foil with alpha particles to investigate the distribution of mass and charge in the atom. - *Example*: Most of the atom's mass was found to be concentrated in a small volume at the atom's center, leading to the proposal of a nuclear model of the atom. ! [Rutherfords-Experiment-700x437.jpg](https://prod-files-secure.s3.us-west-2.amazonaws.com/8f3c221c-927a-482b-9747-5ece94d3f9dc/c519c0fc-0676-4725-bd7d-46b49d44067c/Rutherfords-Experiment-700x437.jpg) ## Subtopic 3: **Alpha Particles** - **Key Concept**: Alpha particles were found to be massive, positively charged particles emitted during radioactive decay. - *Example*: These particles were detected by the small flashes of light when they collided with a zinc-sulfide-coated screen. ## Subtopic 4: **Rutherford's Conclusion** - **Key Concept**: Rutherford's experiment led to the conclusion that the atom had a small, dense, positively charged nucleus # Rutherford's Gold Foil Experiment ## **Surprising Results** - Rutherford expected *minor deflections* of alpha particles passing through gold foil. - *Example*: Like expecting a light breeze to move a beach ball. ## **Unexpected Observations** 1. **Major Deflections**: Some particles were scattered at large angles, even >90?. 2. **Rebounds**: A few particles rebounded at large angles. - *Reflection*: Rutherford likened this to firing a cannon at tissue paper and having it bounce back. ## **Revised Model** - **Nuclear Model**: Rutherford proposed a tiny, massive *nucleus* with concentrated positive charge. - *Example*: Similar to a cannonball hitting tissue paper and bouncing back. - *Insight*: This led to the understanding that most of an atom's mass and positive charge is in the nucleus. ## **Emission Spectra: Understanding Atomic Arrangement** - **Key Concept**: Emission spectra are unique patterns of light emitted by atoms, used to identify elements. - *Example*: Gas-discharge tubes demonstrate this with different gases emitting unique colors. ## **Gas-Discharge Tubes and Emission Spectra** 1. **Principle**: Gas-discharge tubes contain low-pressure gas and electrodes, emitting light when a potential difference is applied. 2. **Unique Colors**: Different gases emit unique colors, forming a characteristic emission spectrum. - *Example*: Passing emitted light through a diffraction grating or prism produces an emission spectrum, unique to each element. - The spectrum is a series of distinct lines of different colors. Each line corresponds to a particular wavelength of light emitted by the atoms of that gas ## **Spectroscopy and Emission Spectrum Analysis** - **Instrument**: A spectroscope is used to study emission spectra in greater detail. - **Process**: Light passes through a slit, diffracts through a grating, and forms an image at different positions for each wavelength. # NOTE : - **Emission Spectrum**: - This term refers to the specific pattern of light (wavelengths) emitted by atoms or molecules of an element when they are excited. Each element has its unique emission spectrum, which acts like a "fingerprint."**Energy Levels of the Hydrogen Atom**: - Substituting the expression for rnr_nrn into the equation for energy, we get : En=-Ke^4me/2h^2 x 1/n^2 - Substituting numerical values: En=-2.17x10-18 Jx1/n2 - Converting the energy from joules to electron volts (1 eV=1.6x10−19 J): En=−13.6eVxn21 ### Summary – **Bohr Radius** (n=1n): r1=0.053 nm – **Energy Levels**: En=−13.6 eVx1n2 – The energy is **quantized** and is inversely proportional to the square of the principal quantum number n. – The lowest energy state (n=1) has an energy of −13.6eV, and as n increases, the energy levels get closer together.– When an electron transitions between these energy levels, it either **absorbs** or **emits** a photon with an energy equal to the difference between the initial and final states: ?E=Efinal−Einitial ### Visible, Ultraviolet, and Infrared Spectral Lines: – The hydrogen atom can emit electromagnetic radiation in different regions of the spectrum: – **Ultraviolet light** is emitted when the electron falls to the ground state (n=1) from a higher energy level.– When an atom absorbs a photon, it gains energy equal to that of the photon, and when it returns to a lower energy state, it emits a photon with a characteristic energy **Energy Change in Atomic Transitions:** – When an atom transitions from an initial energy level *Ei* to a final energy level *Ef*, the change in energy (?Eatom) is calculated by: ?E*atom*=*Ef−Ei* – If the energy decreases (a transition to a lower energy level), the energy difference corresponds to the energy of the emitted photon.**Deriving the Orbital Radius**: – Using Bohr's quantization of angular momentum and the above relationship, we can solve for the orbital radius rn=n2h24?2Kmee2 – Substituting known values for constants and n=1n = 1n=1 (for the ground state) r1=(6.626x10−34 J?s)2x124?2(9.0x109 N?m2/C2)(9.11x10−31 kg)(1.60x10−19 C)2r_1 = \frac{(6.626 \times 10^{-34} \, \text{J} \cdot \text{s})^2 \times 1^2}{4 \pi^2 (9.0 \times 10^9 \, \text{N} \cdot \text{m}^2/\text{C}^2) (9.11 \times 10^{-31} \, \text{kg}) (1.60 \times 10^{-19} \, \text{C})^2} r1=5.3x10−11 m or 0.053 nmr – This is known as the **Bohr radius**.### Final Summary: – **Key Takeaways**: Emission spectra show emitted light wavelengths, aiding in identification, while absorption spectra reveal absorbed light, enabling analysis of element composition.– **Relationship for Orbital Radius:** – By combining the quantization of angular momentum and Coulomb's force, Bohr derived an expression for the radius of the electron's orbit for different energy levels in a hydrogen atom.**Photon Energy Equation:** – The energy of the emitted photon *Ephoton* is equal to the negative of the change in the atom's energy: *Ephoton*=*−?Eatom* – Alternatively, it can be expressed as: *Ephoton=Ei−Ef* **Planck's Equation for Photon Energy:** – The energy of the emitted photon is also given by Planck's equation: *Ephoton*=hf – Here, h is Planck's constant, and f is the frequency of the emitted photon.# NOTE – **Summary Insight**: Both emission and absorption spectra are crucial tools for identifying and analyzing elements and compounds in various fields.– **Application to Hydrogen Spectrum:** – Bohr's model accurately predicted the wavelengths of light emitted by a hydrogen atom, matching experimental measurements by other scientists.## **Determining Star Composition** – **Key Concept**: Scientists determine the composition of stars by comparing the missing lines in the observed spectrum with the known emission spectra of various elements.– Aimed to resolve the issues in the nuclear model by incorporating Planck's concept of quantized energy levels and Einstein's theory of light ## **Stable Condition and Stationary States** – Bohr proposed that the laws of electromagnetism do not apply inside the atom.**Quantization of Angular Momentum**: – According to Bohr's model, the angular momentum of an electron orbiting the nucleus is quantized and is given by the equation mevr=nh/2?## Subtopic 1: **Thomson's Model** – **Key Concept**: J.J. Thomson proposed that the atom contained a massive, positively charged substance with negatively charged electrons distributed throughout, similar to raisins in a muffin.# Absorption Spectra ## **Key Concept**: - **Absorption Spectra** are produced when white light passes through a gas sample and then through a prism, resulting in a continuous spectrum with dark lines corresponding to absorbed wavelengths.**Analyzing Unknown Materials**: Spectroscopy allows scientists to analyze and identify unknown materials by observing their emission and absorption spectra.**Industrial Applications**: Industries like steel mills use spectroscopy to determine the composition of materials, such as scrap iron, for commercial purposes.**Coulomb's Law (Force between Electron and Nucleus)**: - The centripetal force keeping the electron in orbit is provided by the electrostatic attraction between the positively charged nucleus and the negatively charged electron: - F = Ke^2/r^2 where: - K = Coulomb's constant (9.0x109N?m2/C2) - e = charge of the electron (1.60x10-19C) 3.### The Significance of the Bohr Model: - Even though the Bohr model could not accurately predict the spectral lines of atoms with more than one electron, it was crucial in advancing the understanding of atomic structure.- **Contributions of the Bohr Model**: - It accurately predicted the **emission spectrum of hydrogen**, including the precise wavelengths of its spectral lines.- The emission spectrum shows *which wavelengths of light* an atom or molecule emits when its electrons transition from higher to lower energy levels.## Emission Spectra - Spectrum of an incandescent solid is a continuous band of colors from red through violet - **Key Concept**: Emission spectra show the wavelengths of light emitted by atoms of a gas.- Example: Mercury and neon emission spectra display distinct colored lines corresponding to specific wavelengths.- When the emission spectrum of a combination of elements is photographed, an analysis of the lines on the photograph can indicate the identities and the relative concentrations of the elements present.- Using Einstein's photoelectric theory, Bohr determined that the energy of each emitted photon is *EPhoton=hf*, where h is Planck's constant and f is the frequency of the photon.- **Limitations of the Bohr Model:** - The model only worked for hydrogen and couldn't predict the spectrum of helium, the next simplest element.where: - me = mass of the electron (9.11x10-31kg) 9.11x10-31 kg9.11 \times 10^{-31} \, \text{kg} - v = velocity of the electron - r = radius of the orbit - n = principal quantum number (integer values: 1, 2, 3, ...) - h = Planck's constant (6.626x10-34J?s) 2.**Calculating the Total Energy of the Electron**: - The total energy of the electron in orbit is the sum of its kinetic energy (KE) and potential energy (PE) E=KE+PE - The kinetic energy is given by KE= 1/2mv^2 - The potential energy due to electrostatic attraction is: PE=-Ke^2/r - Using KE=-1/2PE the total energy E is : E=-Ke^2/2r 6.- It provided a method to calculate the **ionization energy** of a hydrogen atom, which matched experimental data closely.- Bohr's insights into electron arrangements also influenced our understanding of **chemical bonding** and the **periodic table**, emphasizing that each element has a unique electron configuration.### Historical Impact: - Niels Bohr received the **Nobel Prize in Physics in 1922** for his work on the structure of atoms and their radiation, marking a major step in the development of modern atomic theory.## **Observation of Fraunhofer Lines** - **Key Concept**: Fraunhofer lines are dark lines observed in the spectrum of sunlight, caused by gases in the Sun's atmosphere absorbing specific wavelengths of light.- *Example*: When sunlight passes through the Sun's atmosphere, certain gases absorb characteristic wavelengths, creating dark lines in the spectrum.- Moving an electron farther from the nucleus requires work, so atoms with electrons in distant orbits have higher energy.ELECTRON ORBITAL IN HYDROGEN : rn= h^2(n^2)/4pi^2Kme(e^2) ### Bohr's Model of the Hydrogen Atom 1.rn?n2r_n \propto n^2 - **Energy** decreases in magnitude as En?-1/n2 En?-1/n2E_n \propto -1/n^2 ### Why is the Energy Negative in the Bohr Model?### Energy Levels and Spectral Lines: - The **energy levels** in a hydrogen atom are quantized and given by the formula: En=-13.6eVx1/n^2 where n is the principal quantum number.- Had a major flaw: according to electromagnetic theory, an accelerating electron (moving in a curved path) should lose energy by emitting radiation.- Bohr postulated that the energy levels in an atom are quantized, meaning they exist in specific, discrete amounts.## **Quantization of Energies** - The quantization of energies of electrons in atoms can be likened to a flight of stairs with decreasing-height steps.- This concept forms the basis of Bohr's atomic model, where electrons orbit the nucleus in fixed, quantized energy levels.- **Bohr's Explanation of Atomic Emission Spectra:** - Bohr proposed that stable atoms do not radiate energy while in a fixed orbit.### Predictions of the Bohr Model - **Predictive Power of a Theory:** - Bohr's model provided predictions that could be tested experimentally, a key aspect of a strong scientific theory.- Coulomb's law for the force F between an electron (charge -e) and a proton (charge +e) separated by distance r is given by: F=-Ke^2/r^2 - The centripetal acceleration of the electron is a=r/v^2, where v is the electron's velocity.- **Quantization of Angular Momentum:** - Bohr proposed that the electron's angular momentum is quantized, meaning it can only have certain discrete values.- The **four visible lines** in the hydrogen spectrum (red, blue-green, blue-violet, and violet) correspond to transitions from n=3,4,5,and6->n=2, respectively.For instance, astronomers use the emission spectra of distant stars to determine their chemical composition.- The practical application of emission spectra lies in their uniqueness, making it possible to differentiate and identify elements based on their emitted light.**Corresponding Bright and Dark Lines**: For a gas, the bright lines of the emission spectrum and the dark lines of the absorption spectrum occur at the same wavelengths.## Revising Rutherford's Nuclear model - **19th Century Studies of Atomic Spectra:** - Physicists focused on atomic spectra to understand atomic structure.- Hydrogen, the simplest element with a simple spectrum, was studied extensively.- This energy loss would cause the electron to spiral into the nucleus within a fraction of a second, making atoms unstable.- **Problems with the Planetary Model:** - Contradicted the stability of atoms, as it implied atoms would collapse.- However, atoms emit light only at specific wavelengths (seen in atomic spectra), which the planetary model could not explain.- **Quantization of Energy and Orbits:** - Because energy is quantized, the electron orbits (distances from the nucleus) are also quantized.- Therefore, the energy of an emitted photon is both the product of Planck's constant and frequency, as well as the decrease in the atom's energy.- Here, me is the electron's mass, r is the radius of the orbit, h is Planck's constant, and n is an integer (the principal quantum number).- For a hydrogen atom, Bohr's model calculates the ionization energy as the energy difference between the ground state (n=1) and the ionized state (n->?).- The model introduced the concept of **quantized electron orbits**, which laid the groundwork for quantum mechanics.- **Use of Emission Spectra**: - Emission spectra are used as a tool to *identify elements*, because each element's spectrum is unique.These spectra are used to identify unknown samples or gases by comparing their emitted wavelengths with those of known gases.# Unknown Samples - An unknown gas can be identified by comparing its wavelengths with the wavelengths measured in the spectra of known gas samples.- The elements that are present in the greatest concentrations produce emission lines of greater relative intensities.**Gas Identification**: The set of wavelengths absorbed by a gas forms its absorption spectrum.- *Example*: Gaseous elements absorb the same wavelengths that they emit when excited.**Composition Determination**: The composition of a gas can be determined from the wavelengths of the dark lines in its absorption spectrum.- *Example*: By analyzing the dark lines in the spectrum, scientists can identify the gases present in the star's atmosphere.- **Rutherford's Nuclear Model:** - Proposed that electrons orbit the nucleus like planets orbit the Sun.- Predicted that electrons would emit energy continuously at all wavelengths.- An electron in a stable orbit in an atom does not radiate energy, despite accelerating.## **Energy of an Atom:** - The total energy of an atom is the sum of the kinetic energy of the electrons and the potential energy due to the attraction between electrons and the nucleus.- This agreement between prediction and experiment led to widespread acceptance of Bohr's model for hydrogen.### Development of Bohr's Model - **Application of Newton's Second Law and Coulomb's Law:** - Bohr used Newton's second law of motion (F=ma) and Coulomb's law to describe the electron's motion around the nucleus.**Relationship Between Centripetal Force and Coulomb's Force**: - Equating the centripetal force to the Coulomb force: mev2r=Ke2r2 4.### Visual Representation - For n=1,2,3,...: - **Radius** increases as rn?n2.This negative sign indicates that the electron is in a **bound state**, meaning it is attracted to the nucleus and requires energy to be freed from this attraction.### Understanding Zero Energy: - The **zero energy level** is defined as the point where the electron is **infinitely far** from the nucleus, having no kinetic energy (i.e., it is at rest).In this state, the atom is ionized, and there is no interaction between the nucleus and the electron.The energy is negative because it represents the energy required to pull the electron from its orbit (bound state) to a state of zero interaction (ionization).E1=-13.6eV Thus, to ionize a hydrogen atom from its ground state, **13.6 eV** of energy is required.- Through a comparison of the line intensities, the percentage composition of the gaseous material can be determined.- The visible hydrogen spectrum has four main lines: red, green, blue, and violet.- **Niels Bohr's Contribution:** - Joined Rutherford's group in England in 1911.- Electromagnetic energy is emitted only when an electron transitions from a higher energy level (excited state) to a lower energy level.- A successful theory should not only explain observed phenomena but also make accurate predictions.In the Bohr model, the **total energy** of an electron in orbit around a nucleus is negative.- **Infrared light** is emitted when the electron falls to the n=3 or higher levels (e.g., n=4,5,...).For example, the Sun was found to be mostly hydrogen and helium using this technique.- Electrons have quantized amounts of energy, each called an *energy level*.- It also lacked a strong explanation for why electromagnetic laws should work everywhere except within the atom itself.- The quantization condition is: mevr=nh/2?- Therefore, any electron bound to the nucleus (i.e., in any orbit) has a total energy **less than zero**.### Ionization Energy: - **Ionization energy** is the amount of energy needed to completely remove the electron from the atom, taking it from a bound state to a free state.- **Visible light** is emitted when the electron falls to the n=2 level from a higher level (n>=3).- Scientists had questions about the location and nature of the atom's mass, leading to the development of the nuclear model.## **Importance of Spectroscopy** 1.- Any theory on atomic structure needed to explain these specific wavelengths and support the nuclear model.- An electron's energy cannot have a value between allowed energy levels.- The smallest allowable amount of energy is the *ground state*, while any level above it is an *excited state*.- Electrons closer to the nucleus have lower energy, while electrons further away have higher energy (excited states).This set of transitions is known as the **Lyman series**.This set of transitions is known as the **Balmer series**.This set of transitions is known as the **Paschen series** (for n=3) and others.- This stable condition is called a *stationary state*, with each state having .its own specific amount of energy.## **Practice & Application** 1./2c 2.3.2.5