

The Atomic Spectrum of Hydrogen

When atoms are excited by heat or by electric discharge, they tend to give off electromagnetic energy. These photons of energy are emitted only at certain wavelengths that are characteristic of the specific atoms involved in the sample. These wavelengths make up the atomic spectrum of the element and tell us much about the electronic structure of the atom. Quantum theory tells us that atoms can only exist in certain states, each of which has a fixed amount of energy. When an atom changes energy states, it must absorb or emit an amount of energy that is just equal to the difference between the energies of the initial and the final states. When this energy is emitted, it is done so in the form of electromagnetic energy. This energy may fall within the visible region of the spectrum, or it may lie outside the visible in the lower energy infrared or the higher energy ultraviolet. In some cases the energy is even given off in the form of X-rays, but that takes specialized instruments to measure.

The emission spectrum of an element is obtained when excited electrons fall from a high-energy state to a lower state and emit a photon of "light". Since there are many possible transitions that an electron could make within an atom, the spectra of most atoms are extremely complex. Before you begin this lab, be sure that you have viewed the spectra of several elements through a diffraction rating to see the uniqueness of each pattern of colored lines.

Light is absorbed or emitted by atoms in the form of photons, each of which has a specific amount of energy, E . This energy is related to the wavelength of light by the equation

$$E_{\text{photon}} = hc/\lambda$$

where h is Planck's constant, 6.626×10^{-34} joule seconds, c is the speed of light in a vacuum, 2.997925×10^8 meters/second, and λ is the wavelength in meters. The energy, E_{photon} is in joules and is the energy given off by one electron when it jumps from a higher to a lower energy level. Since total energy is conserved, the change in energy of the atom, ΔE_{atom} , must equal the energy of the photon emitted.

$$\Delta E_{\text{atom}} = \Delta E_{\text{photon}}$$

where ΔE_{atom} is equal to the energy in the upper level minus the energy in the lower one. Thus, the energy transition from a higher energy level to a lower one is equal to the energy of the photon released.

$$E_{\text{higher}} - E_{\text{lower}} = hc/\lambda$$

The simplest of all atomic spectra is that of the hydrogen atom. In 1886 Balmer showed that the lines in the spectrum of the hydrogen atom had wavelengths that could be expressed by a rather simple equation. Bohr, in 1913, explained the spectrum on a theoretical basis with his famous model of the hydrogen atom. According to Bohr's theory, the energies allowed to a hydrogen atom are given by the equation

$$E_n = -b/n^2$$

Where b is a constant predicted by the theory and n is the Principal quantum number (1, 2, 3; ...etc.). It has been found that all the lines in the atomic spectrum of hydrogen can be associated with energy levels in the atom that are predicted with great accuracy by Bohr's equation. When we write the equation above in terms of a mole of H atoms we get the following equation:

$$N(E_{\text{higher}} - E_{\text{lower}}) = Nhc/\lambda$$

Multiplying both sides of the equation by N , the number of atoms/mol (6.02205×10^{23}), converting J to kJ , and substituting the numerical value for b , we obtain

$$E_n = -1312.04 \text{ kJ}/n^2 \text{ mol}$$

Atomic Spectrum Of Hydrogen Pre Lab Answers

B Lingard



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Table of Contents Atomic Spectrum Of Hydrogen Pre Lab Answers

1. Understanding the eBook Atomic Spectrum Of Hydrogen Pre Lab Answers
 - The Rise of Digital Reading Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Atomic Spectrum Of Hydrogen Pre Lab Answers
 - 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 Atomic Spectrum Of Hydrogen Pre Lab Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Personalized Recommendations
 - Atomic Spectrum Of Hydrogen Pre Lab Answers User Reviews and Ratings
 - Atomic Spectrum Of Hydrogen Pre Lab Answers and Bestseller Lists
5. Accessing Atomic Spectrum Of Hydrogen Pre Lab Answers Free and Paid eBooks

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 - Atomic Spectrum Of Hydrogen Pre Lab Answers Budget-Friendly Options
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 - Atomic Spectrum Of Hydrogen Pre Lab Answers Compatibility with Devices
 - Atomic Spectrum Of Hydrogen Pre Lab Answers Enhanced eBook Features
 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Highlighting and Note-Taking Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Interactive Elements Atomic Spectrum Of Hydrogen Pre Lab Answers
 8. Staying Engaged with Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Atomic Spectrum Of Hydrogen Pre Lab Answers
 9. Balancing eBooks and Physical Books Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Atomic Spectrum Of Hydrogen Pre Lab Answers
 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
 11. Cultivating a Reading Routine Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Setting Reading Goals Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Carving Out Dedicated Reading Time
 12. Sourcing Reliable Information of Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Fact-Checking eBook Content of Atomic Spectrum Of Hydrogen Pre Lab Answers
 - Distinguishing Credible Sources
 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development

- Exploring Educational eBooks

14. Embracing eBook Trends

- Integration of Multimedia Elements
- Interactive and Gamified eBooks

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