The Hydrogen Emission Spectrum Exhibits A Line At 1.093010-19 J. If The Final Quantum State Is N = 4,

The Hydrogen Emission Spectrum Exhibits A Line At 1.093010-19 J. If The Final Quantum State Is N = 4, understanding the underlying principles of atomic spectra, quantum mechanics, and hydrogen's energy levels is essential for grasping the significance of this emission line. This article delves into the scientific concepts behind hydrogen's spectral lines, the quantum theory that explains them, and the specific details pertaining to the emission at the energy value of 1.093010×10^{-19} Joules with a final state of N=4.

Introduction to Hydrogen Emission Spectrum

Hydrogen, the simplest and most abundant element in the universe, plays a pivotal role in spectroscopy and quantum physics. Its spectral lines serve as fundamental evidence for the quantization of atomic energy levels. When an electron in a hydrogen atom transitions from a higher to a lower energy state, it emits a photon with an energy equal to the difference between these states. The collection of such emitted photons forms the hydrogen emission spectrum, characterized by discrete lines at specific wavelengths and energies.

Understanding Quantum States in Hydrogen

Quantum Numbers and Energy Levels

In the Bohr model and quantum mechanics, the energy of an electron in a hydrogen atom is described by the principal quantum number, N, which takes positive integer values: N=1, 2, 3, and so forth.

- N=1: Ground state (lowest energy)
- N=2, 3, 4, ...: Excited states

The energy associated with each level is given by the formula:

$$E_N = -13.6\, \text{eV} \times \frac{1}{N^2}$$

where 13.6 eV is the ionization energy of hydrogen.

Energy Transitions and Spectral Lines

When an electron drops from a higher N to a lower N, a photon is emitted with energy equal to:

$$ \Delta E = E_{initial} - E_{final} $$

These transitions result in spectral lines, which are observed in the emission spectrum.

Calculating the Transition Energy for N=4

Given the emission line at 1.093010×10^{-19} Joules, and the final quantum state N=4, we aim to determine the initial state N_i and understand the transition process.

Step 1: Convert the Known Energy to Electronvolts (eV)

Since 1 eV = 1.602×10^{-19} Joules,

$$ E = 1.093010 \times 10^{-19}\ \text{J} \approx \frac{1.093010 \times 10^{-19}}{1.602 \times 10^{-19}}\ \text{eV} \approx 0.682\ \text{eV} $$

This is the energy of the photon emitted during the transition.

Step 2: Determine the Initial Energy Level

Using the energy formula:

$$ E_N = -13.6\, \text{eV} \times \frac{1}{N^2} $$

The energy difference:

$$ \Delta E = E_{N_i} - E_{4} $$

$$ 0.682\, \text{eV} = \left(-13.6 \times \frac{1}{N_i^2} \right) - \left(-13.6 \times \frac{1}{16} \right) $$

$$ 0.682 = -13.6 \left(\frac{1}{N_i^2} - \frac{1}{16} \right) $$

Rearranged:

$$ -\frac{0.682}{13.6} = \frac{1}{N_i^2} - \frac{1}{16} $$

Calculating:

$$ -0.05015 = \frac{1}{N_i^2} - 0.0625 $$

$$ \frac{1}{N_i^2} = 0.0625 - 0.05015 = 0.01235 $$

$$ N_i^2 = \frac{1}{0.01235} \approx 81 $$

$$ N_i \approx \sqrt{81} = 9 $$

Therefore, the electron transitions from N=9 to N=4, emitting a photon with energy approximately 0.682 eV, which matches the given energy.

Significance of the N=9 to N=4 Transition

This transition corresponds to the H-atom spectral line known as the Balmer series, specifically the H-8 line (or sometimes called the Paschen series when involving transitions to N=3, but here, N=4 final state indicates a Balmer series transition).

Key Characteristics of the Transition

- Type: Balmer series transition ($n \geq 3$ to n=2), but since the final state is N=4, it indicates a different series, possibly the Brackett series.
- Energy change: The emitted photon energy is 1.093010×10^{-19} Joules.
- Wavelength: The wavelength can be calculated using:

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

where:

- $h = 6.626 \times 10^{-34}\ \text{Js}$
- $c = 3.00 \times 10^{8}\ \text{m/s}$

Calculating:

$$ \lambda = \frac{6.626 \times 10^{-34} \times 3.00 \times 10^{8}}{1.093010 \times 10^{-19}} $$

$$ \lambda \approx 1.82 \times 10^{-6}\ \text{m} = 1820\ \text{nm} $$

which lies in the infrared region of the spectrum, indicating this is an IR transition.

Quantum Mechanical Explanation of the Spectrum

Bohr Model vs. Quantum Mechanics

While the Bohr model provides a simplified picture using quantized orbits, quantum mechanics offers a more accurate description using wavefunctions and probability distributions. The spectral lines result from electronic transitions between discrete energy levels, governed by selection rules.

Selection Rules for Electronic Transitions

Transitions are allowed when:

- The change in the orbital angular momentum quantum number $\Delta l = \pm 1$.
- The change in magnetic quantum number $\Delta m = 0, \pm 1$.
- The principal quantum number change ΔN can be any value, but the transition's intensity depends on the matrix elements of the dipole operator.

In hydrogen:

- Transitions from N=9 to N=4 are allowed since $\Delta N=5$.
- The transition's probability and intensity depend on the transition dipole moment.

The Role of Spectroscopy in Astrophysics and Quantum Physics

Spectral lines such as the one at energy 1.093010×10^{-19} Joules are foundational in astrophysics for identifying elements in stars and interstellar matter. They help determine physical conditions like temperature, density, and motion through Doppler shifts.

Applications of Hydrogen Spectral Lines

- Astronomical observations: Inferring the composition and dynamics of celestial bodies.
- Quantum mechanics validation: Confirming quantum theory predictions.
- Laser development: Utilizing specific spectral lines for laser action.

Summary and Conclusion

The emission line at 1.093010×10^{-19} Joules in the hydrogen spectrum, with the final quantum state N=4, is a transition from an initial state approximately N=9. This transition emits infrared radiation at approximately 1820 nm, characteristic of hydrogen's spectral series involving higher energy levels.

Understanding this spectral line involves applying quantum mechanics principles, calculating energy differences, and appreciating how atomic structure manifests in observable spectra. The hydrogen emission spectrum remains a cornerstone of atomic physics, providing insights into the fundamental nature of matter and the universe.

Key Takeaways

- The energy of the emitted photon corresponds to a transition from N=9 to N=4.
- The spectral line's wavelength is approximately 1820 nm, in the infrared region.
- The transition obeys the quantum mechanical selection rules for hydrogen.
- Spectroscopic analysis of such lines informs both laboratory physics and astrophysical research.
- The fundamental principles underpinning hydrogen's spectrum continue to influence modern physics, from quantum theory to cosmology.

By understanding the detailed calculations and physical principles behind this emission line, scientists can better interpret spectroscopic data and deepen our knowledge of atomic and cosmic phenomena.

Frequently Asked Questions

What does the hydrogen emission line at 1.093010−19 J indicate about the atom's energy transition?

It indicates a specific electronic transition in hydrogen where an electron moves between energy levels, emitting a photon with energy 1.093010−19 J corresponding to a particular wavelength or spectral line.

Given the emission energy and the final quantum state N=4, how can we determine the initial quantum number N for this transition?

Using the Rydberg formula, we can calculate the initial quantum number N by solving for N in the energy difference equation, considering the known energy of the final state N=4 and the emitted photon energy.

What is the significance of the quantum number N=4 in the context of the hydrogen emission spectrum?

N=4 represents the final energy level of the electron after the transition, indicating that the electron has fallen to the fourth energy level from a higher level, emitting a photon in the process.

How does the wavelength of the emitted photon relate to its energy in the hydrogen emission spectrum?

The wavelength is inversely proportional to the photon energy, calculated using the relation λ = hc / E, where h is Planck's constant and c is the speed of light. A higher energy photon corresponds to a shorter wavelength.

Can the observed spectral line at 1.093010−19 J be associated with any specific transition in the hydrogen atom? If so, which one?

Yes, it corresponds to a transition where the electron drops from a higher energy level to N=4, likely from N=5 or higher, depending on the calculated initial state, producing the observed spectral line.

How does the Bohr model help in understanding the emission spectrum of hydrogen, particularly the line at 1.093010−19 J?

The Bohr model explains the emission spectrum by quantized energy levels and electron transitions between them, allowing calculation of the photon energy emitted during a transition, such as the one at 1.093010−19 J when moving to N=4.

Additional Resources

Hydrogen Emission Spectrum

Introduction: Unlocking the Mysteries of Hydrogen’s Spectral Lines

The hydrogen emission spectrum is one of the most fundamental and well-studied phenomena in atomic physics. It serves as a cornerstone for understanding quantum mechanics, atomic structure, and the nature of light itself. When hydrogen atoms are energized—whether through electrical discharge, heating, or other excitation methods—they emit light at specific wavelengths corresponding to energy transitions within the atom. These emissions produce a distinct spectral pattern, which has been meticulously analyzed since the late 19th century.

In this article, we delve into a specific spectral line of hydrogen, characterized by an energy of 1.093010 × 10^(-19) joules. This line correlates to an electron transition that terminates at the quantum number N = 4. We will explore what this means in terms of atomic energy levels, the quantum mechanical principles involved, and the significance of this spectral line in both historical and modern contexts.

Understanding the Hydrogen Emission Spectrum

The Basics of Spectral Lines

Hydrogen's emission spectrum is primarily characterized by the Balmer series, which includes visible lines, but also extends into ultraviolet and infrared regions. These spectral lines are produced when electrons transition between discrete energy levels within the atom, emitting photons with specific energies corresponding to the difference between initial and final states.

Each spectral line can be described by the Rydberg formula:

$$
\frac{1}{\lambda} = R \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)
$$

where:
- λ is the wavelength of emitted light,
- R is the Rydberg constant ($\sim 1.097373 \times 10^7 \text{ m}^{-1}$),
- n_i is the initial (higher) quantum number,
- n_f is the final (lower) quantum number.

This formula allows us to predict the wavelengths (and energies) of spectral lines for various electron transitions.

The Significance of Energy in Spectral Lines

The energy of a photon emitted during a transition can be calculated using Planck’s relation:

$$
E = h \nu = \frac{hc}{\lambda}
$$

where:
- E is the photon energy,
- h is Planck’s constant ($\sim 6.626 \times 10^{-34}$ Js),
- ν is the frequency,
- c is the speed of light ($\sim 3.0 \times 10^8$ m/s),
- λ is the wavelength.

In our specific case, the spectral line has an energy of 1.093010×10^{-19} Joules, which provides insight into the transition involved.

The Spectral Line at 1.093010×10^{-19} Joules: A Deep Dive

Linking Energy to Quantum Numbers

Given the energy of the emitted photon, we can determine the initial and final energy levels involved in this transition. For hydrogen, the energy levels are described by the Bohr model:

$$
E_n = - \frac{13.6\, \text{eV}}{n^2}
$$

where n is the principal quantum number (1, 2, 3, ...). To convert the given energy from joules to electronvolts (eV):

$$
E (\text{eV}) = \frac{E (\text{J})}{1.602 \times 10^{-19}}
$$

Applying this:

$$
E = \frac{1.093010 \times 10^{-19}}{1.602 \times 10^{-19}} \approx 0.682 \, \text{eV}
$$

This photon energy (~0.68 eV) suggests a transition involving relatively high energy levels, since the energy difference between levels decreases as n increases.

Determining the Final State: N = 4

The problem states that the final quantum state is N = 4, indicating the electron transitions down to the $n=4$ level. To identify the initial state n_i, we use the energy difference:

$$
\Delta E = E_{n_i} - E_{4}
$$

Given the energy levels:

$$
E_n = -13.6\, \text{eV} / n^2
$$

then:

$$
\Delta E = -\frac{13.6}{n_i^2} + \frac{13.6}{16}
$$

(since $E_4 = -13.6/16 \approx -0.85\, \text{eV}$).

Setting $\Delta E \approx 0.68\, \text{eV}$:

$$
0.68 = -\frac{13.6}{n_i^2} + 0.85
$$

$$
-\frac{13.6}{n_i^2} = 0.68 - 0.85 = -0.17
$$

$$
\frac{13.6}{n_i^2} = 0.17
$$

$$
n_i^2 = \frac{13.6}{0.17} \approx 80
$$

$$
n_i \approx \sqrt{80} \approx 8.94
$$

Since n_i must be an integer, the transition likely occurs from $n_i=9$ to $n=4$, which is consistent with the known hydrogen spectral lines.

The Transition: $n=9 \rightarrow n=4$

This transition corresponds to the upper level $n=9$ down to $n=4$, emitting a photon with energy approximately 0.68 eV, aligning with the given spectral line energy of 1.093010×10^{-19} Joules.

Implications of the Spectral Line at N=4

The Balmer Series and Its Significance

Transitions ending at $n=4$ are part of the Paschen series, which lies in the infrared region, but given the energy, this particular line is more consistent with the Humphreys series (transitions ending at $n=4$) or a higher-order Balmer line, depending on the actual wavelength.

In our case, since the energy corresponds to visible or near-infrared light, it highlights the importance of hydrogen's spectral lines for various scientific and technological applications:

- Astrophysics: Identifying hydrogen lines in stellar spectra helps determine composition, temperature, and movement.
- Quantum Mechanics: Validating models of atomic structure and energy quantization.
- Spectroscopy: Developing precise measurement techniques for atomic transitions.

Significance in Modern Science

Understanding these spectral lines enables scientists to:

- Refine atomic models: Confirm predictions of quantum mechanics.
- Measure cosmic distances: Using spectral lines as standard candles.
- Develop laser technologies: Utilizing specific transitions within hydrogen for laser emissions.

Broader Context: Quantum Transitions and Energy Levels

Quantum Mechanical Perspective

The transition from a higher level (n_i) to a lower level $(n_f=4)$ involves the quantized energy difference, dictated by the Schrödinger equation. The energy levels of hydrogen are well approximated by the Bohr model but are more accurately described by quantum mechanics, which accounts for fine structure, Zeeman effects, and other phenomena.

Selection Rules for Transitions

Not all transitions are allowed; quantum mechanics imposes selection rules:

- Change in orbital angular momentum quantum number: $\Delta l = \pm 1$
- Change in magnetic quantum number: $\Delta m = 0, \pm 1$

For electric dipole transitions, the most common and intense spectral lines occur when these rules are satisfied.

The Role of Quantum Numbers

- Principal quantum number n: Determines energy level.
- Angular momentum quantum number l: Defines orbital shape.
- Magnetic quantum number m: Orientation of the orbital.
- Spin quantum number s: Electron spin.

The specific transition at $n=9 \rightarrow 4$ involves changes in these quantum numbers, consistent with the selection rules and resulting in the observed spectral line.

Practical Applications and Future Directions

Spectroscopy and Astrophysics

The precise measurement of spectral lines like the one at 1.093010 × 10^(-19) Joules assists in:

- Mapping stellar atmospheres.
- Determining chemical abundances in galaxies.
- Studying cosmic phenomena like quasars and nebulae.

Advances in Quantum Technologies

Understanding hydrogen spectral lines facilitates:

- Development of high-precision atomic clocks.
- Quantum computing and information processing, where atomic states serve as qubits.
- Laser development based on specific atomic transitions.

Ongoing Research

Current research aims to:

- Refine measurements of spectral lines to test fundamental physics.
- Explore hydrogen-like ions for broader applications.
- Investigate effects of external fields (magnetic, electric) on spectral lines to understand atomic interactions better.

Conclusion: The Enduring Significance of Hydrogen's Spectral Lines

The spectral line at 1.093010 × 10^(-19) Joules, corresponding to a transition ending at \(N=4\),

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