For A Hydrogen Atom, What Is The Excited State (n2) If A Wavelength Of 93.8 Nm Is Emitted When N1=1?

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Understanding the excited states of the hydrogen atom is fundamental in atomic physics and quantum mechanics. When a hydrogen atom transitions from a higher energy level (excited state, n_2) to a lower energy level (n_1), it emits a photon whose wavelength corresponds to the energy difference between these levels. In this article, we explore how to determine the specific excited state (n_2) in hydrogen when a photon with wavelength 93.8 nm is emitted during a transition to the ground state (n_1=1). We will delve into the principles behind atomic emission spectra, the Bohr model, the relevant formulas, and detailed calculations to answer this question comprehensively.

Fundamentals of Hydrogen Atom Energy Levels

The hydrogen atom is the simplest atom, consisting of one proton and one electron. Its energy levels are quantized, meaning the electron can only occupy specific energy states characterized by the principal quantum number n (n=1, 2, 3, ...). Each energy level corresponds to a specific energy, with higher n representing higher energy states.

When an electron transitions from a higher energy level (n_2) to a lower one (n_1), the energy difference (ΔE) is released as a photon. The wavelength (λ) of this photon is related to this energy difference via the Planck-Einstein relation:

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

where:
- h is Planck's constant ($6.626 \times 10^{-34} \, \mathrm{Js}$)
- c is the speed of light ($3.00 \times 10^8 \, \mathrm{m/s}$)
- λ is the wavelength of emitted photon

The energy levels of the hydrogen atom are given by the Bohr formula:

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

The energy difference between levels n_2 and n_1 is:

$$\Delta E = E_{n_1} - E_{n_2} = 13.6 \, \mathrm{eV} \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

Since the transition is to the ground state (n_1=1), this simplifies to:

$$ \Delta E = 13.6 \, \mathrm{eV} \left(1 - \frac{1}{n_2^2} \right) $$

Calculating the Excited State (n_2) from Wavelength Data

Given the emitted wavelength λ = 93.8 nm, which corresponds to the transition from n_2 to $n_1=1$, our goal is to find n_2.

Step 1: Convert Wavelength to Energy

First, convert the wavelength into energy in electronvolts (eV):

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

Where:
- $h = 6.626 \times 10^{-34} \, \mathrm{Js}$
- $c = 3.00 \times 10^{8} \, \mathrm{m/s}$
- $\lambda = 93.8 \, \mathrm{nm} = 93.8 \times 10^{-9} \, \mathrm{m}$

Calculating:

$$
E = \frac{(6.626 \times 10^{-34})(3.00 \times 10^{8})}{93.8 \times 10^{-9}} \approx \frac{1.9878 \times 10^{-25}}{9.38 \times 10^{-8}} \approx 2.12 \times 10^{-18} \, \mathrm{J}
$$

Converting Joules to eV:

$$
1 \, \mathrm{eV} = 1.602 \times 10^{-19} \, \mathrm{J}
$$

$$
E (\mathrm{eV}) = \frac{2.12 \times 10^{-18}}{1.602 \times 10^{-19}} \approx 13.24 \, \mathrm{eV}
$$

So, the energy of the photon emitted is approximately 13.24 eV.

Step 2: Use the Energy Difference Formula

Recall:

$$
\Delta E = 13.6 \, \mathrm{eV} \left(1 - \frac{1}{n_2^2} \right)
$$

Set ΔE to 13.24 eV:

$$
13.24 = 13.6 \left(1 - \frac{1}{n_2^2} \right)
$$

Divide both sides by 13.6:

$$
\frac{13.24}{13.6} = 1 - \frac{1}{n_2^2}
$$

$$
0.974 = 1 - \frac{1}{n_2^2}
$$

Rearranged:

$$
\frac{1}{n_2^2} = 1 - 0.974 = 0.026
$$

Calculate n_2^2:

$$
n_2^2 = \frac{1}{0.026} \approx 38.46
$$

Finally, find n_2:

$$
n_2 \approx \sqrt{38.46} \approx 6.20
$$

Since quantum numbers are integers, the closest integer value for n_2 is 6.

Result: The excited state n_2 is 6.

Interpretation and Confirmation of Results

The calculation indicates that the electron transitions from the energy level $n_2 = 6$ to the ground state $n_1=1$, emitting a photon with a wavelength of approximately 93.8 nm.

Why n=6?
The transition from $n=6$ to $n=1$ produces the Balmer series in the ultraviolet region, consistent with the observed wavelength.

Additional notes:

- The transition from $n=6$ to $n=1$ is a highly energetic ultraviolet emission.
- The wavelength of 93.8 nm falls within the extreme ultraviolet (EUV) spectrum, typical for high-level electronic transitions in hydrogen.

Additional Transitions in Hydrogen Spectrum

Hydrogen emission lines are classified into series based on the lower energy level:

- Lyman series: Transition to $n_1=1$, ultraviolet region.
- Balmer series: Transition to $n_1=2$, visible region.
- Paschen series: Transition to $n_1=3$, infrared region.
- Brackett series: Transition to $n_1=4$, infrared.
- Pfund series: Transition to $n_1=5$, infrared.

The transition from $n=6$ to $n=1$ is part of the Lyman series, emphasizing its ultraviolet nature.

Significance of Determining n_2 in Atomic Physics

Understanding the excited states and spectral lines of hydrogen is crucial for various scientific fields:

- Astrophysics: Identifying spectral lines in stars and interstellar medium.
- Quantum Mechanics: Validating the Bohr model and quantum theory.
- Spectroscopy: Analyzing atomic and molecular structures.
- Plasma Physics: Understanding emission spectra in plasma environments.

Accurate calculations of energy levels and transitions facilitate the development of models for atomic behavior and contribute to technologies such as lasers, atomic clocks, and spectroscopy instruments.

Summary of Key Steps to Find n_2

1. Convert wavelength to photon energy using $E = \frac{hc}{\lambda}$.
2. Relate the photon energy to the difference in energy levels using the Bohr model formula.
3. Solve for n_2 by algebraic manipulation and approximation.
4. Interpret the result within the context of hydrogen spectral series.

Final answer: The excited state n_2 for the hydrogen atom, emitting a photon of wavelength 93.8 nm when transitioning to $n_1=1$, is approximately 6.

Conclusion

By applying fundamental principles of atomic physics and spectral analysis, we have determined that the hydrogen atom's electron is excited to the $n=6$ energy level before transitioning down to the ground state and emitting a photon with a wavelength of 93.8 nm. This process exemplifies the quantized nature of atomic energy levels and highlights the importance of spectral lines in understanding atomic structure. Whether you are a student, researcher, or enthusiast, mastering these calculations enhances comprehension of atomic spectra and the underlying quantum mechanics governing atomic behavior.

Keywords for SEO:

- Hydrogen atom excited state
- Hydrogen spectral lines
- Transition n=6 to n=1 hydrogen
- Wavelength emission hydrogen atom
- Bohr model hydrogen spectrum

Frequently Asked Questions

What is the excited state (n_2) of a hydrogen atom if a wavelength of 93.8 nm is emitted during the transition from n_2 to $n_1=1$?

The excited state n_2 is approximately 4, as the transition from n=4 to n=1 emits a photon with a wavelength of about 93.8 nm.

How can we determine the initial energy level (n_2) of a hydrogen atom given the emitted wavelength of 93.8 nm for the n_2 to $n_1=1$ transition?

By using the Rydberg formula and the wavelength, we solve for n_2, which results in $n_2 \approx 4$ for the given wavelength.

Why does a transition from n=4 to n=1 in hydrogen

emit a photon with a wavelength of 93.8 nm?

Because the energy difference between levels n=4 and n=1 corresponds to the photon energy associated with 93.8 nm, which lies in the ultraviolet region.

What is the significance of the wavelength 93.8 nm in hydrogen spectral lines?

It corresponds to the Lyman series transition from n=4 to n=1, indicating the electron drops from the fourth energy level to the ground state.

How does the wavelength of 93.8 nm relate to the energy difference between the excited and ground states in hydrogen?

The wavelength of 93.8 nm indicates a large energy difference, specifically between the n=4 and n=1 energy levels, calculated using the Rydberg formula.

Additional Resources

Understanding the Excited State of a Hydrogen Atom from Emission Wavelengths: An In-Depth Analysis

When studying the spectral lines of the hydrogen atom, one of the fundamental questions involves determining the energy levels involved when the atom transitions between states. Specifically, if a photon with a wavelength of 93.8 nm is emitted during a transition from an excited state \(n_2 \) to a lower state \(n_1 \), and \(n_1 = 1 \), what is the value of \(n_2 \)? This question not only probes the core principles of quantum mechanics and atomic physics but also offers insight into the spectral characteristics of hydrogen, which serves as the archetype for understanding atomic structure.

Fundamentals of Hydrogen Atom Spectroscopy

Before delving into the specifics of the problem, it’s crucial to establish a solid foundation regarding the spectral behavior of the hydrogen atom and the principles that govern its emission lines.

The Bohr Model and Quantized Energy Levels

- Quantization of Energy: The hydrogen atom’s energy levels are discrete, meaning electrons can only occupy certain allowable energy states characterized by the principal quantum number \(n \).

- Energy Level Formula: According to Bohr’s model, the energy of the nth level in a hydrogen atom is given by:

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

where 13.6 eV is the ionization energy of hydrogen.

- Transition of Electrons: When an electron transitions from a higher energy level n_2 to a lower energy level n_1, it emits a photon with energy corresponding to the difference between these levels:

$$
\Delta E = E_{n_2} - E_{n_1}
$$

Spectral Series of Hydrogen

Hydrogen's emission lines are categorized into series based on the lower energy level n_1:

- Lyman Series: Transitions ending at $n_1 = 1$, lying in the ultraviolet region.
- Balmer Series: Transitions ending at $n_1 = 2$, in the visible region.
- Paschen Series: Transitions ending at $n_1 = 3$, in the infrared region.
- And others (Brackett, Pfund, etc.) for higher n_1.

Since the emission wavelength in question is 93.8 nm, which falls within the ultraviolet spectrum, the transition likely belongs to the Lyman series.

Connecting Wavelength and Energy: The Fundamental Relationship

The key to solving the problem lies in understanding the relationship between the emitted photon’s wavelength and the energy difference between the initial and final states.

The Equation: $E = h \nu = \frac{hc}{\lambda}$

- Planck’s constant (h): $6.626 \times 10^{-34} \, \text{Js}$
- Speed of light (c): $3.00 \times 10^{8} \, \text{m/s}$
- Wavelength (λ): Given as 93.8 nm, which is $93.8 \times 10^{-9} \, \text{m}$

The energy of the photon emitted during the transition:

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

\]

Calculating this:

$$
\Delta E = \frac{(6.626 \times 10^{-34})(3.00 \times 10^{8})}{93.8 \times 10^{-9}}
$$

$$
\Delta E \approx \frac{1.9878 \times 10^{-25}}{9.38 \times 10^{-8}} \approx 2.12 \times 10^{-18} \, \text{J}
$$

Converting to electronvolts:

$$
1\, \text{eV} = 1.602 \times 10^{-19} \, \text{J}
$$

$$
\Delta E \approx \frac{2.12 \times 10^{-18}}{1.602 \times 10^{-19}} \approx 13.24\, \text{eV}
$$

Thus, the energy difference between the two levels involved in the transition is approximately 13.24 eV.

Determining the Transition: From Energy Difference to Quantum Numbers

Given that the final state $n_1 = 1$, and the energy difference is approximately 13.24 eV, we can now determine the initial excited level n_2.

Step 1: Express Energy Difference in Terms of n_2 and n_1

Using the Bohr energy formula:

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

The energy difference:

$$
\Delta E = E_{n_2} - E_{n_1} = - \frac{13.6}{n_2^2} - \left(- \frac{13.6}{1^2} \right) = 13.6 \left(1 - \frac{1}{n_2^2} \right)
$$

\]

Set this equal to the calculated photon energy:

$$
13.6 \left(1 - \frac{1}{n_2^2} \right) = 13.24
$$

Step 2: Solve for n_2

$$
1 - \frac{1}{n_2^2} = \frac{13.24}{13.6} \approx 0.974
$$

$$
\frac{1}{n_2^2} = 1 - 0.974 = 0.026
$$

$$
n_2^2 = \frac{1}{0.026} \approx 38.46
$$

$$
n_2 \approx \sqrt{38.46} \approx 6.2
$$

Since n_2 must be an integer quantum number, and the calculated value is close to 6, the most plausible value is:

$$
\boxed{n_2 = 6}
$$

Final Result:

The excited state n_2 is 6, corresponding to a transition from $n=6$ to $n=1$, with the emission of a photon of wavelength approximately 93.8 nm.

Implications and Validations of the Result

Having deduced that $n_2 = 6$, it’s vital to validate this conclusion against known spectral lines and understand its physical significance.

Known Spectral Data

- The Lyman-alpha line (transition $n=2 \to 1$) corresponds to a wavelength of about 121.6 nm.

- The Lyman-beta line ($3 \to 1$) is around 102.6 nm.
- The Lyman-gamma ($4 \to 1$): approximately 97.3 nm.
- The Lyman-delta ($5 \to 1$): roughly 93.8 nm.
- The Lyman-epsilon ($6 \to 1$): about 93.2 nm.

The calculated wavelength (93.8 nm) aligns with the Lyman-delta line, which involves the transition $n=5 \to 1$. However, our calculation suggested $n_2 \approx 6$, so we need to re-examine the calculations for precision.

Re-evaluation for Exactness

Using the precise energies:

$$
\Delta E = 13.6 \left(1 - \frac{1}{n_2^2} \right) = 13.24\, \text{eV}
$$

$$
1 - \frac{1}{n_2^2} = \frac{13.24}{13.6} \approx 0.974
$$

$$
\frac{1}{n_2^2} = 0.026
$$

$$
n_2^2 \approx 38.46
$$

$$
n_2 \approx 6.2
$$

This suggests that the transition is most consistent with $n=6 \to 1$. Since the known wavelength for $6 \to 1$ is around 93.2 nm (Lyman-epsilon), the slight discrepancy could be attributed to rounding errors.

Therefore, the most accurate conclusion is:

$$
\boxed{\text{The excited state } n_2 = 6}
$$

corresponding to the emission line near 93.8 nm, matching the spectral line for $6 \to 1$.

Additional Considerations and Contextual Insights

Significance of the Wavelength and Transition

- The ultraviolet wavelength of approximately 93.8 nm indicates a high-energy transition.
- Such lines are crucial in astrophysics, as hydrogen is the most abundant element in the universe, and its spectral lines are used to infer conditions in stars and interstellar media

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