

Find The Energy Required To Excite A Hydrogen Electron From N = 1 To N = 4.

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Understanding the energy transitions of electrons in hydrogen atoms is fundamental in quantum physics and spectroscopy. When an electron in a hydrogen atom moves from one energy level to another, it either absorbs or emits a specific amount of energy, corresponding to a photon of particular wavelength or frequency. This article explores how to calculate the energy required to excite a hydrogen electron from the ground state (N=1) to the fourth energy level (N=4). We will delve into the principles of the Bohr model, the formula for energy levels, step-by-step calculations, and the significance of these transitions in spectroscopic analysis.

Understanding Hydrogen Electron Energy Levels

Bohr Model of the Hydrogen Atom

The Bohr model, proposed by Niels Bohr in 1913, provides a simplified way to understand the quantized energy levels of electrons in hydrogen atoms. According to this model:

- The electron orbits the nucleus in specific, quantized circular paths called energy levels.
- Each energy level corresponds to a principal quantum number, denoted by N, where N=1, 2, 3, etc.
- The energy of each level is discrete, not continuous.

Energy Quantization in Hydrogen

In the Bohr model, the energy of an electron in the nth level of hydrogen is given by the formula:

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

Where:

- E_n is the energy of the level in electron volts (eV).
- 13.6 eV is the ionization energy of hydrogen (the energy needed to remove the electron from the ground state).

- n is the principal quantum number (N).

The negative sign indicates the bound state of the electron within the atom; the more negative the energy, the more tightly bound the electron.

Calculating the Energy Difference for N=1 to N=4 Transition

Step 1: Identify the Initial and Final Energy Levels

- Initial energy level: $n_i = 1$
- Final energy level: $n_f = 4$

Step 2: Write Down the Energy Levels

Using the formula:

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

Calculate:

- $E_1 = - \frac{13.6}{1^2} = -13.6 \text{ eV}$
- $E_4 = - \frac{13.6}{4^2} = - \frac{13.6}{16} = -0.85 \text{ eV}$

Step 3: Determine the Energy Required for Transition

The energy needed to excite the electron from N=1 to N=4 is the energy difference:

$$\Delta E = E_f - E_i$$

Since the electron absorbs energy, this difference will be positive:

$$\Delta E = (-0.85 \text{ eV}) - (-13.6 \text{ eV}) = 13.6 \text{ eV} - 0.85 \text{ eV} = 12.75 \text{ eV}$$

\]

Therefore, the energy required to excite a hydrogen electron from $N=1$ to $N=4$ is approximately 12.75 eV.

Additional Insights into Electron Transitions

Significance of the $N=1$ to $N=4$ Transition

- This transition corresponds to the emission or absorption of photons in the ultraviolet region of the electromagnetic spectrum.
- The exact wavelength can be calculated using the energy-wavelength relation, which is useful in spectroscopy.

Energy Wavelength Calculation

Using the relation:

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

Where:

- E is the energy in joules,
- $h = 6.626 \times 10^{-34} \text{ Js}$ (Planck's constant),
- $c = 3.00 \times 10^8 \text{ m/s}$,
- λ is the wavelength in meters.

First, convert the energy from eV to joules:

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

Thus,

$$E = 12.75 \text{ eV} \times 1.602 \times 10^{-19} \approx 2.043 \times 10^{-18} \text{ J}$$

Now, compute the wavelength:

\[
\\lambda = \\frac{hc}{E} = \\frac{6.626 \\times 10^{-34} \\times 3.00 \\times 10^8}{2.043 \\times 10^{-18}} \\approx 9.74 \\times 10^{-8} \\text{m}
\\]

or approximately 97.4 nanometers, which confirms the ultraviolet nature of this transition.

Summary of Key Concepts

- The energy levels of hydrogen are quantized and described by the Bohr model.
- Transition energies can be calculated using the energy level formula: $E_n = -13.6 \\text{eV} / n^2$.
- The energy required to excite an electron from $N=1$ to $N=4$ is 12.75 eV.
- The corresponding photon wavelength is approximately 97.4 nm, in the ultraviolet spectrum.
- These calculations are essential in understanding atomic spectra, astrophysical phenomena, and quantum mechanics.

Applications of Electron Transition Calculations

Spectroscopy and Identification of Elements

- Spectroscopic analysis relies on known transition energies to identify elements in stars, laboratories, and chemical samples.
- The hydrogen emission spectrum, including the Balmer series, is a fundamental reference in astrophysics.

Quantum Mechanics and Atomic Physics

- Calculations of energy differences validate quantum models of atomic structure.
- They help in developing laser technologies, where precise energy transitions are exploited.

Educational and Research Significance

- Understanding these calculations enhances comprehension of atomic behavior.
- They serve as foundational examples in quantum physics coursework.

Conclusion

Calculating the energy required to excite a hydrogen electron from $N=1$ to $N=4$ involves understanding the principles of the Bohr model and quantum energy levels. By applying the energy formula, we find that approximately 12.75 eV is needed for this transition, corresponding to ultraviolet photons around 97.4 nm. These calculations are vital in spectroscopy, astrophysics, and quantum physics, providing insights into atomic structure and electromagnetic interactions.

By mastering these calculations, scientists and students can better understand atomic behavior, spectral lines, and the fundamental principles governing matter at the quantum level.

Frequently Asked Questions

What is the energy required to excite a hydrogen electron from $N=1$ to $N=4$?

The energy required is approximately 10.2 eV, which is the difference between the energy levels at $N=4$ and $N=1$ in the hydrogen atom.

How do I calculate the energy difference between $N=1$ and $N=4$ in a hydrogen atom?

Use the formula $\Delta E = 13.6 \text{ eV} (1/n_1^2 - 1/n_2^2)$. For $N=1$ to $N=4$, $\Delta E = 13.6 \text{ eV} (1/1^2 - 1/4^2) = 13.6 \text{ eV} (1 - 1/16) = 13.6 \text{ eV} \cdot 15/16 \approx 12.75 \text{ eV}$.

What is the significance of the energy difference between these levels?

It represents the amount of energy the hydrogen atom must absorb to excite its electron from the ground state ($N=1$) to the fourth energy level ($N=4$).

Is the energy calculated in electron volts (eV) or joules?

The energy is commonly expressed in electron volts (eV), but it can be converted to joules by multiplying with 1.602×10^{-19} J/eV.

How does the energy required change if the electron is excited from N=1 to N=2 instead?

The energy difference from N=1 to N=2 is approximately 10.2 eV, which is less than the 12.75 eV needed for N=1 to N=4.

Can the energy required be obtained using the Bohr model of the hydrogen atom?

Yes, using the Bohr model, the energy levels are given by $E_n = -13.6 \text{ eV} / n^2$. The difference between levels $n=1$ and $n=4$ is calculated as shown, yielding approximately 12.75 eV.

What wavelength of light corresponds to this energy transition?

Using the relation $E = hc/\lambda$, where h is Planck's constant and c is the speed of light, the wavelength is approximately 97 nm, corresponding to ultraviolet light.

Why is the energy difference negative in the energy level formula?

The negative sign indicates that the energy levels are bound states, with the ground state being the most stable; the difference is taken as a positive value to signify energy absorption.

How is this excitation energy relevant in spectroscopy?

The energy difference corresponds to specific spectral lines in hydrogen's emission or absorption spectra, allowing identification of hydrogen in various astrophysical and laboratory contexts.

What assumptions are made in calculating the energy required for this transition?

Calculations assume the hydrogen atom is isolated, using the Bohr model with non-relativistic quantum mechanics, neglecting effects like fine structure or external fields.

Additional Resources

Find The Energy Required To Excite A Hydrogen Electron From $N = 1$ To $N = 4$

In the realm of atomic physics, understanding how electrons transition between energy levels is fundamental to grasping the nature of matter and light. Today, we delve into a classic yet intriguing question: What is the energy needed to excite a hydrogen atom's electron from its ground state ($N=1$) to a higher energy level ($N=4$)? This inquiry not only illustrates core concepts of quantum mechanics but also lays the groundwork for comprehending phenomena such as spectral lines, atomic interactions, and even the design of lasers and other photonic devices.

The Hydrogen Atom and Its Energy Levels: An Overview

Hydrogen, the simplest atom with one proton and one electron, serves as the cornerstone for atomic physics. Its energy levels are well understood and can be described precisely using quantum mechanics.

Quantum Numbers and Energy Levels

In the Bohr model and quantum mechanics, the energy levels of a hydrogen atom are characterized by the principal quantum number, N , which takes positive integer values: $N=1, 2, 3, \dots$. Each level corresponds to a specific energy state:

- $N=1$: Ground state (lowest energy)
- $N=2, 3, 4, \dots$: Excited states (higher energies)

The energy associated with each level can be calculated using the Bohr formula, which accurately reflects the energy of an electron in a hydrogen atom:

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

Here, 13.6 eV is the ionization energy of hydrogen in its ground state (the energy required to remove the electron entirely).

This negative sign indicates that the electron is bound to the nucleus; the energy becomes zero when the electron is completely free (ionized).

Calculating the Energy Difference: From $N=1$ to $N=4$

The fundamental principle for calculating the energy required to excite the electron from one level to another is subtracting the initial energy from the final energy:

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

where:

- E_{N_i} is the initial energy (N=1),
- E_{N_f} is the final energy (N=4).

Since the energies are negative, a positive ΔE indicates energy must be supplied to the atom to promote the electron from the lower to the higher level.

Step-by-Step Calculation

1. Identify the energies at N=1 and N=4:

$$E_1 = -\frac{13.6 \text{ eV}}{1^2} = -13.6 \text{ eV}$$

$$E_4 = -\frac{13.6 \text{ eV}}{4^2} = -\frac{13.6 \text{ eV}}{16} = -0.85 \text{ eV}$$

2. Calculate the energy difference:

$$\Delta E = E_4 - E_1 = (-0.85 \text{ eV}) - (-13.6 \text{ eV}) = 13.6 \text{ eV} - 0.85 \text{ eV} = 12.75 \text{ eV}$$

Thus, the energy required to excite the electron from N=1 to N=4 is approximately 12.75 eV.

Interpreting the Results: What Does 12.75 eV Mean?

The calculated value signifies that to elevate the hydrogen atom's electron from its ground state to the fourth energy level, an energy input of about 12.75 electronvolts is necessary. This is a measurable and significant quantity; for comparison, visible light photons have energies roughly between 1.6 eV (red) and 3.3 eV (violet). Therefore, the excitation involves ultraviolet or higher-energy photons.

This energy requirement has practical implications:

- Spectroscopy: When hydrogen absorbs photons with energies matching these transitions, it produces spectral lines characteristic of hydrogen, such as

those in the Balmer series.

- Astrophysics: These energy levels explain the emission and absorption lines observed in stellar spectra.
- Quantum Mechanics Validation: The precise calculation supports the quantum model over classical physics, confirming the quantized nature of atomic energy levels.

Broader Context: The Significance of Excitation Energies

Understanding excitation energies like the 12.75 eV for N=1 to N=4 transition is crucial for multiple scientific and technological domains.

Spectral Line Formation

When an electron drops from N=4 back down to N=1, it emits a photon with energy approximately 12.75 eV, corresponding to ultraviolet radiation. This process creates spectral lines observed in laboratory spectra and cosmic phenomena. The Rydberg formula, which predicts spectral lines, relies on these energy differences, making the precise calculation vital for accurate spectroscopic analysis.

Atomic and Molecular Physics

Knowledge of excitation energies aids in modeling atomic interactions, plasma physics, and chemical reactions involving hydrogen. It helps determine how atoms absorb energy, ionize, or emit light under various conditions.

Technological Applications

- Lighting and Laser Technologies: Understanding atomic transitions enables the design of lasers like the hydrogen laser, which relies on specific energy level transitions.
- Astrophysical Observations: Telescopes equipped with ultraviolet detectors analyze stellar and interstellar hydrogen, utilizing these energy differences to infer physical conditions.

Limitations and Assumptions in the Calculation

While the calculation provides a clear estimate, it's essential to recognize underlying assumptions:

- Hydrogenic Model: The calculation assumes a perfect hydrogen atom without external perturbations or relativistic effects.
- Neglecting Fine Structure: Spin-orbit coupling and relativistic corrections slightly modify energy levels but are negligible at this scale.
- No External Fields: Magnetic or electric fields can shift energy levels (Zeeman or Stark effects), affecting the actual energy needed.

Despite these simplifications, the calculation remains accurate for most practical purposes involving hydrogen excitation.

Final Thoughts: A Window Into Atomic Transitions

The calculation of 12.75 eV to excite a hydrogen atom from $N=1$ to $N=4$ exemplifies the quantized nature of atomic systems. It bridges fundamental quantum mechanics with observable phenomena, from spectral lines to astrophysical insights. As science progresses, understanding these basic principles continues to underpin advances in spectroscopy, quantum computing, and our comprehension of the universe.

In essence, the energy required to excite a hydrogen electron from the ground state to a higher level not only reveals the intricate structure of the atom but also illuminates the fundamental processes that govern the interaction of matter and light across the cosmos.

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