A Ground State H Atom Absorbs A Photon And Makes A Transition To The N = 4 Energy Level. It Then Emits

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Understanding the behavior of hydrogen atoms during photon absorption and emission processes is fundamental to quantum physics, spectroscopy, and many applications in astrophysics and chemistry. When a hydrogen atom in its ground state absorbs a photon, it undergoes an energy transition to a higher energy level — in this case, the N=4 energy level. Subsequently, the atom can emit photons as it returns to lower energy states, producing characteristic spectral lines. This article provides an in-depth exploration of this process, including the underlying quantum mechanics, the significance of energy levels, and the spectral implications of such transitions.

Hydrogen Atom Energy Levels: An Overview

The hydrogen atom, the simplest atom consisting of one proton and one electron, serves as a fundamental model in quantum mechanics. Its energy levels are quantized, meaning the electron can only occupy specific discrete energy states, labeled by the principal quantum number, N.

Principal Quantum Number (N)

- Defines the energy level of the electron.
- N = 1 corresponds to the ground state.
- Higher N values indicate excited states with higher energy and larger average distance from the nucleus.

Energy Level Formula

The energy of an electron in the nth energy level of a hydrogen atom is given by:

$$ E_n = -13.6\, \text{eV} \times \frac{1}{n^2} $$

Where:
- E_n is the energy of the nth level.
- 13.6 eV is the ionization energy of hydrogen.

For example:
- N=1: $E_1 = -13.6\, \text{eV}$
- N=4: $E_4 = -13.6\, \text{eV} \times \frac{1}{16} = -0.85\, \text{eV}$

This energy difference between levels is crucial for understanding photon absorption and emission.

Photon Absorption: Transition from N=1 to N=4

When a hydrogen atom in its ground state (N=1) absorbs a photon with the appropriate energy, it transitions to a higher energy state—in this case, N=4.

The Absorption Process

- Photon Energy Requirement: For absorption, the photon must have energy equal to the difference between the initial and final energy levels:

$$ \Delta E = E_{N=4} - E_{N=1} $$

- Calculating the Energy Difference:

$$
\begin{aligned}
\Delta E &= (-0.85\, \text{eV}) - (-13.6\, \text{eV}) \\
&= 13.6\, \text{eV} - 0.85\, \text{eV} \\
&= 12.75\, \text{eV}
\end{aligned}
$$

- Photon Wavelength:

Using the relationship between energy and wavelength:

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

Where:
- h is Planck's constant ($6.626 \times 10^{-34}\, \text{Js}$)
- c is the speed of light ($3.0 \times 10^{8}\, \text{m/s}$)
- λ is the wavelength in meters

Calculating:

$$
\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34})(3.0 \times 10^{8})}{12.75 \times 1.602 \times 10^{-19}} \approx 97.2\, \text{nm}
$$

This wavelength falls within the ultraviolet (UV) spectrum.

Implications of Absorption

- The UV photon excites the electron from N=1 to N=4.
- This excitation process is fundamental in astrophysics, where stellar atmospheres absorb specific UV wavelengths.
- Such transitions are also used in laboratory spectroscopy to analyze hydrogen.

Electron Transition from N=4 to Lower Energy Levels and Emission

After excitation, the hydrogen atom is in an unstable, higher energy state. It will eventually return to a lower energy state, emitting photons characteristic of the energy difference between the levels.

Possible Transition Pathways from N=4

- The atom can directly transition from N=4 to N=1, emitting a photon with energy:

$$
\Delta E_{4 \to 1} = E_1 - E_4 = -13.6\, \text{eV} - (-0.85\, \text{eV}) = -12.75\, \text{eV}
$$

- Alternatively, it can cascade through intermediate levels:

1. N=4 to N=3
2. N=3 to N=2
3. N=2 to N=1

This cascade results in the emission of multiple photons at different wavelengths.

Calculating Transition Energies and Wavelengths

Transition	Energy Difference (ΔE)	Wavelength (λ)	Spectral Region
N=4 to N=3	$E_3 - E_4 = -1.51\, \text{eV} - (-0.85\, \text{eV}) = -0.66\, \text{eV}$	$\approx 1870\, \text{nm}$	Infrared (IR)
N=3 to N=2	$-3.4\, \text{eV} - (-1.51\, \text{eV}) = -1.89\, \text{eV}$	$\approx 656\, \text{nm}$	Red (Balmer series)
N=2 to N=1	$-13.6\, \text{eV} - (-3.4\, \text{eV}) = -10.2\, \text{eV}$	$\approx 121.6\, \text{nm}$	Ultraviolet (Lyman series)

Note: The transitions from higher levels produce a series of spectral lines, notably the

Balmer series (visible), Lyman series (UV), and Paschen series (IR).

Significance of Emission Lines

- These emission lines serve as fingerprints for identifying hydrogen in various astrophysical environments.
- The Lyman-alpha line at 121.6 nm (N=2 to N=1) is one of the most prominent in cosmic spectra.
- Balmer lines in the visible spectrum are key in stellar classification and nebular analysis.

Quantum Mechanical Perspective on Transitions

The absorption and emission processes are governed by quantum mechanics, specifically by the selection rules that dictate the allowed transitions.

Selection Rules for Electric Dipole Transitions

- Change in principal quantum number: Δn can be any integer.
- Change in orbital angular momentum quantum number: $\Delta l = \pm 1$.
- Change in magnetic quantum number: $\Delta m_l = 0, \pm 1$.

In hydrogen, transitions like N=1 to N=4 are allowed because they satisfy these rules.

Transition Probability and Lifetimes

- The probability of spontaneous emission is characterized by Einstein coefficients.
- The lifetime of the excited state influences the intensity and width of spectral lines.
- For the N=4 state, the lifetime is on the order of nanoseconds, allowing for observable emission lines.

Astrophysical and Practical Applications

Understanding these transitions has broad implications:

- Astronomy: Analyzing stellar spectra, nebulae, and interstellar medium to determine composition and physical conditions.
- Spectroscopy: Identifying elements and measuring physical parameters in laboratories.
- Quantum Mechanics: Testing and validating quantum theories of atomic structure.
- Laser Technology: Utilizing hydrogen emission lines in certain laser applications.

Summary

In summary, when a ground state hydrogen atom absorbs a photon with an energy of approximately 12.75 eV (wavelength ~97.2 nm), it makes a transition from N=1 to N=4. This excitation prepares the atom for subsequent emission events, where it can either directly transition back to the ground state or cascade through intermediate levels, emitting photons characteristic of each transition. These emissions produce spectral lines across ultraviolet, visible, and infrared regions, forming the basis of many spectroscopic techniques and astrophysical observations.

Understanding these processes not only enriches our knowledge of atomic physics but also provides practical tools for exploring the universe and advancing technological innovations. The detailed study of hydrogen atom transitions exemplifies the elegant interplay between quantum mechanics and observable phenomena, highlighting the importance of spectral analysis in science.

Keywords: hydrogen atom, photon absorption, energy levels, N=4 transition, emission spectrum, spectral lines, quantum mechanics, spectroscopy, astrophysics, Lyman series, Balmer series, atomic transitions

Frequently Asked Questions

What is the initial energy state of the hydrogen atom before absorbing the photon?

The hydrogen atom is in its ground state, with the principal quantum number n = 1.

What happens when the hydrogen atom absorbs a photon in the ground state?

The atom absorbs the photon and transitions from n = 1 to a higher energy level, specifically n = 4 in this case.

Which type of transition occurs when the hydrogen atom moves from n = 1 to n = 4?

It undergoes an electronic excitation transition from the ground state (n = 1) to a higher energy state (n = 4), absorbing energy equal to the difference between these levels.

What is the process called when the hydrogen atom emits a photon after being excited to n = 4?

It is called spontaneous emission, where the atom releases a photon as it transitions from a higher to a lower energy level.

What are the possible emission transitions after the hydrogen atom reaches the n = 4 energy level?

The atom can emit photons as it transitions from n = 4 to lower energy levels such as n = 3, n = 2, or directly back to n = 1, resulting in various spectral lines.

How can the emitted photon's wavelength be determined after the transition from n = 4 to n = 1?

Using the Rydberg formula for hydrogen, the wavelength can be calculated based on the energy difference between n = 4 and n = 1 energy levels.

Additional Resources

Hydrogen Atom Transitions: From Ground State Absorption to Emission at N=4

Introduction: The Beauty of Atomic Transitions

In the realm of atomic physics, the hydrogen atom stands as the quintessential model, offering profound insights into the fundamental processes governing matter at the quantum level. Its simplicity—featuring a single electron orbiting a proton—makes it an ideal candidate to explore atomic energy levels, photon interactions, and spectral emissions. This article delves into a specific yet illustrative process: a hydrogen atom initially in its ground state absorbs a photon, transitions to a higher energy state (N=4), and subsequently emits radiation as it returns to lower energy states. This not only exemplifies the quantum mechanics of atomic transitions but also underpins many practical applications, from astrophysics to spectroscopy.

The Ground State of Hydrogen: The Baseline

Understanding the Ground State

The ground state of the hydrogen atom, denoted as N=1, is its lowest energy configuration. In quantum mechanics, the principal quantum number N characterizes the energy level of the electron:

- N=1: Ground state, the lowest energy level.
- N=2, 3, 4, ...: Excited states, with increasing energy.

In the ground state, the hydrogen atom’s electron occupies the most stable and lowest energy orbital, with quantum numbers:

- Principal quantum number: N=1
- Orbital angular momentum quantum number: l=0

- Magnetic quantum number: m=0
- Spin quantum number: s=1/2

This state is stable under normal conditions, but it can be excited by absorbing energy in the form of a photon.

Photon Absorption: Transition from N=1 to N=4

The Mechanics of Absorption

When a hydrogen atom in its ground state encounters a photon with sufficient energy, it can absorb that photon, causing the electron to jump to a higher energy level—in this case, N=4. This process is governed by the principles of quantum electrodynamics and selection rules that dictate whether a transition is allowed.

Energy of the Absorbing Photon

The energy of the photon must match the energy difference between the initial and final states:

$$
E_{photon} = E_{N=4} - E_{N=1}
$$

The energy levels of hydrogen are given by the Rydberg formula:

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

Calculating the energies:

- $E_1 = -13.6 \, \text{eV}$
- $E_4 = -13.6 \, \text{eV} \times \frac{1}{16} = -0.85 \, \text{eV}$

Thus,

$$
E_{photon} = (-0.85) - (-13.6) = 12.75 \, \text{eV}
$$

This photon must have an energy of approximately 12.75 eV to induce the transition.

Wavelength of the Incident Photon

Using the relation $E = hc / \lambda$:

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

\]

where:

- $h = 4.1357 \times 10^{-15} \, \text{eV·s}$,
- $c = 3 \times 10^8 \, \text{m/s}$,

we find:

$$
\lambda \approx \frac{(4.1357 \times 10^{-15})(3 \times 10^8)}{12.75} \approx 97 \, \text{nm}
$$

This wavelength (~97 nm) lies in the ultraviolet region, indicating that ultraviolet photons are necessary for this excitation.

Selection Rules for Absorption

Quantum mechanics imposes selection rules for electric dipole transitions:

- $\Delta N \neq 0$, but specific Δl values are critical.
- $\Delta l = \pm 1$: change in the orbital angular momentum quantum number.
- $\Delta m = 0, \pm 1$: change in magnetic quantum number.

Since the initial state has $l=0$, the final state at $N=4$ with $l=1$ (p orbital) is allowed. The transition involves a change from an s orbital to a p orbital, consistent with these rules.

The Excited State: N=4

Characteristics of the N=4 Level

The N=4 energy level is an excited state with multiple sublevels distinguished by quantum numbers:

- Orbital angular momentum quantum number $l=0,1,2,3$ (s, p, d, f orbitals).
- Magnetic quantum number m varies from $-l$ to $+l$.

Within N=4, the electron could occupy various sub-orbitals:

- 4s ($l=0$)
- 4p ($l=1$)
- 4d ($l=2$)
- 4f ($l=3$)

Typically, the transition from ground to N=4 involves excitation to the 4p orbital because of the selection rules.

Stability and Lifetime

Once in the N=4 state, the hydrogen atom is in a non-equilibrium state with excess energy. It is metastable, meaning it can remain in this state for a variable amount of time depending on the transition probabilities and available pathways.

Emission: The Atom's Return to Lower Energy Levels

De-excitation Pathways

After excitation, the hydrogen atom will eventually return to a lower energy state by emitting photons. These emission processes can follow several paths:

- Direct Transition (N=4 → N=1): Emission of a high-energy photon (~97 nm), known as a Lyman-alpha transition.
- Sequential Cascades: The atom may transition through intermediate levels:

- N=4 → N=3
- N=3 → N=2
- N=2 → N=1

Each step emits a photon with characteristic wavelength, producing a cascade of emissions that shape the spectral lines observed.

Spectral Lines and Their Significance

The emitted photons correspond to specific spectral lines:

- Lyman series: Transitions ending at N=1 (ultraviolet region).
- Balmer series: Transitions ending at N=2 (visible region).
- Paschen series: Transitions ending at N=3 (infrared region).

For the N=4 to N=1 transition, the photon emitted is in the ultraviolet Lyman-alpha wavelength (~121.6 nm), which is one of the most prominent features in astrophysical spectra.

Detailed Breakdown of the Emission Cascade

Sequence of Emissions

The atom's return to the ground state can involve multiple steps, releasing photons of different energies:

1. N=4 → N=3: Emission at ~100.5 nm (Lyman-beta).
2. N=3 → N=2: Emission at ~102.6 nm (Lyman-alpha).
3. N=2 → N=1: Emission at ~121.6 nm (Lyman-alpha).

Alternatively, direct N=4 → N=1 is also possible but less probable due to transition rates.

Implications for Spectroscopy

This cascade produces a distinctive spectral signature, which astronomers utilize to analyze cosmic hydrogen clouds, star-forming regions, and the intergalactic medium. The relative intensities of these lines inform us about physical conditions like temperature, density, and radiation fields.

Practical Applications and Significance

Astrophysics and Cosmology

- Mapping Hydrogen: The spectral lines generated by these transitions are fundamental in understanding the distribution and dynamics of hydrogen in space.
- Studying Early Universe: The Lyman series helps probe the epoch of reionization and the formation of the first galaxies.
- Diagnostics: Emission line ratios provide diagnostics for physical conditions in astrophysical plasmas.

Laboratory Spectroscopy

- Calibration Standards: Hydrogen emission lines serve as standards in spectroscopic calibration.
- Quantum Mechanics Validation: Observations of these transitions validate quantum models of atomic structure.

Technological Applications

- UV Detectors: Devices designed to detect ultraviolet photons rely on understanding such atomic transitions.
- Laser Physics: Stimulated emission involving hydrogen energy levels informs laser development in the UV spectrum.

Conclusion: The Quantum Dance of Hydrogen

The journey of a hydrogen atom from the ground state to an excited N=4 state and back again exemplifies the elegance of quantum mechanics. The photon-induced excitation and subsequent emission not only underscore fundamental physical principles but also empower diverse scientific fields. From the ultraviolet glow of distant nebulae to the calibration of advanced spectroscopic instruments, these atomic transitions are at the heart of our understanding of the universe. Their study continues to illuminate the microscopic and cosmic realms, demonstrating the profound interconnectedness of quantum phenomena and observational astronomy.

In summary, the process involves precise energy considerations, strict selection rules, and a cascade of emissions that produce characteristic spectral lines. Understanding these transitions enables scientists to decode the universe's secrets, making hydrogen’s simple yet profound quantum dance a cornerstone of modern science.

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