A Group Of Hydrogen Atoms In A Discharge Tube Emit Violet Light Of Wavelength 410 Nm.Determine The Quantum

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Understanding the behavior of hydrogen atoms and their emission spectra is fundamental in atomic physics and quantum mechanics. When hydrogen atoms are excited in a discharge tube, they emit light at specific wavelengths, revealing vital information about their energy levels and the quantized nature of atomic energy states. This article delves into the phenomenon of hydrogen emission lines, focusing on the violet light with a wavelength of 410 nm, and guides you through the process of determining the quantum involved in this emission.

Introduction to Atomic Spectra and Hydrogen Emission Lines

Hydrogen, the simplest atom consisting of one proton and one electron, exhibits a distinctive emission spectrum when its electrons transition between energy levels. These spectral lines are characteristic signatures that allow scientists to analyze atomic structures and validate quantum theories.

What Are Atomic Emission Spectra?

Atomic emission spectra arise when electrons in an atom transition from higher to lower energy states, releasing photons with energies corresponding to the difference between these levels. These photons manifest as specific wavelengths of light, forming a spectrum unique to each element.

Hydrogen's Spectral Series

Hydrogen's emission spectrum comprises several series, each corresponding to electron transitions ending at a particular energy level:

- **Lyman Series:** Transitions ending at the n=1 level (ultraviolet region)
- **Balmer Series:** Transitions ending at n=2 (visible region)

- **Paschen Series:** Transitions ending at n=3 (infrared region)
- And others such as Brackett and Pfund series

The violet light at 410 nm falls within the visible region, generally associated with the Balmer series, which involves transitions ending at n=2.

Understanding the Quantum Number and Transition in Hydrogen

In atomic physics, the energy levels of electrons in hydrogen are quantified by the principal quantum number, n, which takes positive integer values (n=1, 2, 3, ...). When an electron transitions between two energy levels, the change in the quantum number determines the emitted photon's energy and wavelength.

Energy Levels and Bohr Model of Hydrogen

The Bohr model provides a simplified way to calculate the energy of electrons in hydrogen:

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

where:
- E_n is the energy of level n
- 13.6 eV is the ionization energy of hydrogen

The negative sign indicates that the electron is bound within the atom. When an electron transitions from a higher level n_i to a lower level n_f, the energy of the emitted photon is:

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

which corresponds to the photon energy:

$$ E_{photon} = h \nu = \frac{hc}{\lambda} $$

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 of emitted light

Calculating the Quantum Transition for the 410 nm Violet Emission

Given the wavelength of the emitted violet light, the goal is to determine the quantum transition responsible for this emission—specifically, the initial and final energy levels involved.

Step 1: Convert Wavelength to Energy

Using the relation:

$$ E_{photon} = \frac{hc}{\lambda} $$

where:
- $\lambda = 410\, \text{nm} = 410 \times 10^{-9}\, \text{m}$

Calculating:

$$ E_{photon} = \frac{6.626 \times 10^{-34} \times 3.0 \times 10^{8}}{410 \times 10^{-9}} $$

$$ E_{photon} \approx \frac{1.9878 \times 10^{-25}}{4.10 \times 10^{-7}} $$

$$ E_{photon} \approx 4.846 \times 10^{-19}\, \text{J} $$

Converting Joules to electronvolts (1 eV = 1.602×10^{-19} J):

$$ E_{photon} \approx \frac{4.846 \times 10^{-19}}{1.602 \times 10^{-19}} \approx 3.027\, \text{eV} $$

Step 2: Determine the Transition Level

Since the emission occurs in the visible region and aligns with the Balmer series, which involves transitions ending at level $n=2$, the initial level n_i is higher than 2.

Using the energy level formula:

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

The energy difference:

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

Given $n_f = 2$:

$$E_2 = -\frac{13.6}{2^2} = -\frac{13.6}{4} = -3.4\, \text{eV}$$

So,

$$\Delta E = E_{n_i} - (-3.4\, \text{eV}) = E_{n_i} + 3.4\, \text{eV}$$

But since the photon energy is approximately 3.027 eV, and the transition is from n_i to 2:

$$E_{n_i} + 3.4\, \text{eV} \approx 3.027\, \text{eV}$$

Rearranged:

$$E_{n_i} \approx 3.027 - 3.4 = -0.373\, \text{eV}$$

Now, equate this to the energy level formula:

$$E_{n_i} = -\frac{13.6}{n_i^2}$$

which gives:

$$-\frac{13.6}{n_i^2} \approx -0.373$$

Simplify:

$$\frac{13.6}{n_i^2} \approx 0.373$$

$$n_i^2 \approx \frac{13.6}{0.373}$$

$$n_i^2 \approx 36.44$$

$$n_i \approx \sqrt{36.44} \approx 6.04$$

Since n_i must be an integer, the closest integer is 6.

Result: The transition responsible for the violet emission at 410 nm is from $n=6$ to $n=2$.

Summary of the Quantum Transition

Initial Level (n_i)	Final Level (n_f)	Wavelength (nm)	Transition Type
6	2	410	Balmer Series

This transition from $n=6$ to $n=2$ corresponds to the violet emission observed.

Significance of the Quantum Transition in Atomic Physics

Understanding such quantum transitions is pivotal in validating quantum theory and the Bohr model's accuracy. The quantized energy levels explain why hydrogen emits light at specific wavelengths, leading to the formation of the characteristic emission lines. Analyzing the transition from $n=6$ to $n=2$ confirms the energy quantization and supports the concept of discrete energy states.

Applications in Modern Science and Technology

- Spectroscopy: Used to identify elements in stars and distant celestial bodies.
- Laser Technology: Understanding atomic transitions underpins the development of laser sources.
- Quantum Mechanics: Empirical validation of the quantized energy levels in atoms.
- Astrophysics: Analyzing spectral lines to determine the composition and physical conditions of astronomical objects.

Conclusion

The violet light emitted by hydrogen atoms in a discharge tube with a wavelength of 410 nm results from a specific electronic transition between energy levels. Calculations based on quantum physics reveal that this emission corresponds to a transition from the $n=6$ level to the $n=2$ level. This process exemplifies the discrete nature of atomic energy levels and underscores the importance of quantum mechanics in explaining atomic spectra.

By understanding and calculating these quantum transitions, scientists can interpret spectral data, explore atomic structures, and develop technological applications that rely on precise control of atomic emissions. The study of hydrogen's emission lines continues to be a cornerstone in the field of atomic physics and quantum mechanics, illustrating the profound link between microscopic quantum phenomena and observable spectral features.

References

- Griffiths, D. (2017). Introduction to Quantum Mechanics. Cambridge University Press.
- Tipler, P. A., & Llewellyn, R. (2013).

Frequently Asked Questions

What is the significance of the wavelength 410 nm emitted by hydrogen atoms in a discharge tube?

The wavelength 410 nm corresponds to a specific electronic transition in the hydrogen atom, indicating the emission of violet light as electrons move between energy levels. It helps identify the energy difference between those levels.

How do you calculate the energy of the emitted photon when hydrogen atoms emit light at 410 nm?

The energy (E) of the photon can be calculated using the formula $E = hc/\lambda$, where h is Planck's constant (6.626×10^{-34} Js), c is the speed of light (3×10^{8} m/s), and λ is the wavelength (410 nm or 410×10^{-9} m). Substituting values gives $E \approx 4.85\times10^{-19}$ Joules.

What quantum number change corresponds to the emission of violet light at 410 nm in hydrogen?

The emission at 410 nm corresponds to an electronic transition where the principal quantum number decreases from n=6 to n=2, or a similar transition involving higher energy levels, indicating a specific quantum jump.

How can the Rydberg formula be used to determine the transition in hydrogen atoms emitting at 410 nm?

The Rydberg formula ($1/\lambda = R(1/n_1^2 - 1/n_2^2)$) relates the wavelength to the initial and final energy levels. By substituting λ=410 nm and R (Rydberg constant $\approx 1.097\times10^{7}$ m^{-1}), we can find the values of n_1 and n_2, identifying the electronic transition responsible for the emission.

What does the emission of violet light at 410 nm tell us about the energy

levels involved in hydrogen's atomic spectrum?

It indicates a transition involving relatively high energy levels, specifically between excited states and lower energy states, highlighting the quantized nature of hydrogen's atomic energy levels and the specific electron transitions that produce violet emission.

Additional Resources

Hydrogen Emission Spectrum: Unveiling the Quantum of Violet Light at 410 nm

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Introduction

In the fascinating world of atomic physics and spectroscopy, the emission spectra of elements serve as a window into the quantum realm. Hydrogen, the simplest and most abundant element in the universe, has been a cornerstone in understanding atomic structure. When a group of hydrogen atoms in a discharge tube emit violet light at a wavelength of 410 nm, it presents an intriguing opportunity to explore the quantum mechanics governing atomic transitions. This article aims to delve deeply into the process of determining the quantum involved in this emission, blending scientific rigor with an engaging analysis reminiscent of an expert review.

The Significance of Hydrogen's Emission Spectrum

Hydrogen's emission spectrum is characterized by discrete lines, each corresponding to specific electronic transitions within the atom. These spectral lines are fundamental evidence for the quantized nature of atomic energy levels, a core principle established by Niels Bohr and later refined by quantum mechanics.

In the case of the violet emission at 410 nm, the spectral line falls within the visible range, specifically in the violet region, and is part of the Balmer series—transitions where electrons fall to the second energy level (n=2). Understanding this emission involves:

- Recognizing the transition responsible for the 410 nm line
- Applying quantum theory to determine the quantum number change
- Calculating the energy and frequency associated with this transition

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Theoretical Foundations: Quantized Energy Levels in Hydrogen

Hydrogen Atom and Bohr Model

The Bohr model postulates that electrons orbit the nucleus in discrete energy levels, with energies given by:

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

where:
- E_n is the energy of the nth level
- n is the principal quantum number (n=1, 2, 3, ...)

Electrons transition between these energy levels, emitting or absorbing photons with energies equal to the difference between initial and final states.

Quantized Transition and Emission

When an electron transitions from a higher energy level n_i to a lower level n_f, the emitted photon has an energy:

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

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 of emitted light

The quantum number change ($\Delta n = n_i - n_f$) determines the spectral line.

Analyzing the Violet Line at 410 nm

Step 1: Convert Wavelength to Energy

Given:

$$\lambda = 410\, \text{nm} = 410 \times 10^{-9}\, \text{m}$$

Calculate the energy of the photon:

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

$$E = \frac{(6.626 \times 10^{-34})(3.0 \times 10^{8})}{410 \times 10^{-9}}$$

$$ E \approx \frac{1.9878 \times 10^{-25}}{4.10 \times 10^{-7}} $$

$$ E \approx 4.846 \times 10^{-19}\, \text{J} $$

Convert this energy to electron volts (eV):

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

$$ E \approx \frac{4.846 \times 10^{-19}}{1.602 \times 10^{-19}} \approx 3.027\, \text{eV} $$

Thus, the photon emitted has an energy of approximately 3.03 eV.

Step 2: Determine the Corresponding Transition

Since the emission occurs in the Balmer series (transitions ending at $ n_f=2 $), the initial level $ n_i $ can be deduced by solving:

$$ E_{n_i} - E_2 = E_{photon} $$

Using the energy levels:

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

we have:

$$ -\frac{13.6}{n_i^2} - \left(-\frac{13.6}{2^2}\right) = 3.03\, \text{eV} $$

Simplify:

$$ -\frac{13.6}{n_i^2} + \frac{13.6}{4} = 3.03 $$

$$ -\frac{13.6}{n_i^2} + 3.4 = 3.03 $$

$$ -\frac{13.6}{n_i^2} = -0.37 $$

$$ \frac{13.6}{n_i^2} = 0.37 $$

$$ n_i^2 = \frac{13.6}{0.37} \approx 36.76 $$

$$ n_i \approx \sqrt{36.76} \approx 6.06 $$

Since n_i must be an integer, the closest quantum level is n=6.

Therefore, the transition responsible for the 410 nm violet emission is from $n=6$ to $n=2$.

Quantum Number Change and Spectral Series

This transition ($6 \rightarrow 2$) is part of the Balmer series, which is characterized by visible lines resulting from transitions where electrons fall to $n=2$. The Balmer series formula for the wavelength of emitted photons is given by the Rydberg formula:

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

where R is the Rydberg constant ($1.097 \times 10^{7}\, \text{m}^{-1}$).

Calculating the Rydberg Constant and Confirming the Transition

Substituting the known values:

$$\frac{1}{\lambda} = 1.097 \times 10^{7} \left(\frac{1}{2^2} - \frac{1}{6^2} \right)$$

$$\frac{1}{\lambda} = 1.097 \times 10^{7} \left(\frac{1}{4} - \frac{1}{36} \right)$$

$$\frac{1}{\lambda} = 1.097 \times 10^{7} \left(\frac{9}{36} - \frac{1}{36} \right)$$

$$\frac{1}{\lambda} = 1.097 \times 10^{7} \times \frac{8}{36}$$

$$\frac{1}{\lambda} = 1.097 \times 10^{7} \times \frac{2}{9}$$

$$\frac{1}{\lambda} \approx 1.097 \times 10^{7} \times 0.2222$$

$$\frac{1}{\lambda} \approx 2.438 \times 10^{6}\, \text{m}^{-1}$$

Calculating wavelength:

$$\lambda = \frac{1}{2.438 \times 10^{6}} \approx 4.10 \times 10^{-7} \, \text{m} = 410\, \text{nm}$$

This confirms that the emission at 410 nm indeed corresponds to the $6 \rightarrow 2$ transition, with a quantum jump of $\Delta n = 4$.

Broader Implications and Quantum Insights

Quantum Number Change and Energy Quantization

The transition from $n=6$ to $n=2$ exemplifies the quantized energy jumps in hydrogen atoms. Each spectral line is a direct manifestation of electrons transitioning between discrete energy levels, reinforcing the quantum hypothesis.

Determining the Quantum

The key quantum here is the principal quantum number n of the initial state:

- Initial quantum level: n=6
- Final quantum level: n=2

The change in quantum number:

$$\Delta n = n_i - n_f = 6 - 2 = 4$$

This quantum jump corresponds to the emission of a photon with a wavelength of 410 nm, linked to an energy of approximately 3.03 eV.

Practical Applications and Significance

Understanding this emission process is not just academic; it has real-world implications:

- Spectroscopic Techniques: Used in astrophysics to analyze stellar compositions.
- Laser Development: Hydrogen spectral lines inform the design of laser systems.
- Quantum Mechanics Validation: Empirical confirmation of energy quantization principles.
- Educational Foundations: Serves as a classic example in physics curricula for quantum theory.

Conclusion

The violet light of wavelength 410 nm emitted by hydrogen atoms in a discharge tube encapsulates the profound principles of quantum mechanics. Through meticulous calculations rooted in the Bohr model and Ryd

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