Consider The Bohr Model Of The Atom. Which Transition Would Correspond To The Largest Wavelength Of Light

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Understanding atomic structure and the behavior of electrons within an atom is fundamental to the field of chemistry and physics. Among various models proposed to explain atomic behavior, the Bohr model stands out as one of the most significant early theories. It provides a simplified yet powerful framework to comprehend electron transitions, spectral lines, and the emission of light from atoms. In particular, analyzing which electronic transition corresponds to the largest wavelength of light offers insight into the energy levels of the atom and the nature of electromagnetic radiation emitted during these transitions.

This article delves into the Bohr model of the atom, exploring how electron transitions relate to emitted light's wavelength. It aims to clarify which specific transitions produce the longest wavelengths, emphasizing the underlying physics, mathematical formulations, and real-world applications.

Overview of the Bohr Model of the Atom

The Bohr model, proposed by Niels Bohr in 1913, was a groundbreaking development in atomic physics. It was designed to explain the spectral lines observed in hydrogen and, by extension, other elements. The model introduces quantized energy levels for electrons orbiting the nucleus, which was a significant departure from classical physics.

Key Principles of the Bohr Model

1. Quantized Orbits: Electrons orbit the nucleus in specific, stable orbits without radiating energy. These orbits are associated with discrete energy levels characterized by principal quantum numbers (n = 1, 2, 3, ...).

2. Energy Quantization: The energy of an electron in a particular orbit is quantized and given by the formula:

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

where n is the principal quantum number.

3. Electron Transitions and Photon Emission: When an electron jumps from a higher energy level n_i to a lower energy level n_f, a photon is emitted with energy equal to the difference between the two levels:

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

4. Spectral Lines: These photon emissions correspond to specific wavelengths, forming the spectral lines observed in emission spectra.

Electron Transitions and Their Corresponding Wavelengths

The wavelength of the emitted light during an electronic transition is directly related to the energy difference between the initial and final states. The fundamental relation connects the energy of the photon to its wavelength:

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

where:

- E is the energy difference between levels,
- h is Planck's constant ($6.626 \times 10^{-34}\, \text{Js}$),
- c is the speed of light ($3.00 \times 10^8\, \text{m/s}$),
- λ is the wavelength of the emitted photon.

Rearranged to find the wavelength:

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

From this relation, it is clear that the larger the energy difference (ΔE), the shorter the wavelength (more energetic photons, like X-rays or UV light). Conversely, smaller energy differences correspond to longer wavelengths (such as infrared or radio waves).

Which Transition Corresponds To The Largest Wavelength?

Since wavelength is inversely proportional to energy difference, the transition with the smallest energy change produces the longest wavelength. In the context of the hydrogen atom, the smallest energy difference occurs when the electron transitions between levels that are close together.

Analyzing Electron Transitions in Hydrogen

In the hydrogen atom, the energy levels are given by:

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

Suppose an electron transitions from a higher level n_i to a lower level n_f. The energy of the emitted photon is:

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

To maximize the wavelength λ, we look for the smallest ΔE, which occurs when n_i and n_f are consecutive levels with $n_i \to n_f = n_i - 1$.

Smallest energy difference:

$$
\Delta E_{n \to n-1} = 13.6\, \text{eV} \left(\frac{1}{(n-1)^2} - \frac{1}{n^2}\right)
$$

As n increases, this difference becomes smaller because $\frac{1}{n^2}$ diminishes.

Transition with the Largest Wavelength in Hydrogen

The transition from $n = 2$ to $n = 1$ (Lyman-alpha line) has a significant energy difference, producing ultraviolet light. Conversely, the transition from very high n levels to $n=2$ results in very small energy differences and thus longer wavelengths.

The longest wavelength in the hydrogen spectrum is associated with the transition from very high n (e.g., $n \to \infty$) to $n=2$. In practice, this corresponds to the limit of the Balmer series, which converges to a limit wavelength called the series limit.

In the Balmer series (transitions ending at $n=2$):
- The transition from $n = 3 \to 2$ produces visible red light.
- The transition from $n \to 2$ as $n \to \infty$ produces the longest wavelength in the series, which is around 656.3 nm (H-alpha line).

Therefore, the transition from the highest energy level (theoretically $n \to \infty$) to $n=2$ corresponds to the largest wavelength in the Balmer series.

Generalization to Other Atoms

While the hydrogen atom is the simplest and most studied, the principle extends to other atoms. The key points are:

- Transition involving the smallest energy difference yields the longest wavelength.
- For multi-electron atoms, the energy levels are more complex due to electron-electron interactions, but the same inverse relationship between energy difference and wavelength applies.

Implications and Applications of Electron Transitions and Wavelengths

Understanding which electronic transitions produce the largest wavelength is crucial in various scientific and technological fields:

- Spectroscopy: Identifying elements based on their spectral lines, especially in astrophysics.
- Laser Technology: Selecting specific transitions to generate desired wavelengths.
- Astrophysical Observations: Analyzing the light from stars and galaxies to determine their composition and physical conditions.
- Quantum Physics Education: Demonstrating the quantization of energy levels and photon emission.

Summary and Key Takeaways

- The Bohr model explains atomic spectral lines in terms of electron transitions between quantized energy levels.
- The wavelength of emitted light during a transition is inversely proportional to the energy difference between levels.
- Transitions with the smallest energy difference produce the longest wavelengths.
- In the hydrogen atom, the transition from $n \to \infty$ to $n=2$ (the limit of the Balmer series) corresponds to the longest wavelength emission.
- For practical observations, the transition from high n levels to lower levels (like $n=2$) yields the longest wavelengths in the visible spectrum.

Conclusion

The analysis of electron transitions in the Bohr model provides a clear understanding of the origin of spectral lines and the relationship between energy levels and emitted wavelengths. The transition that corresponds to the largest wavelength of light is the one involving the smallest energy difference, which, in the case of hydrogen, is the transition from a very high energy level to a lower one—most notably, from $n \to \infty$ to $n=2$. Recognizing these principles is fundamental for interpreting spectral data and advancing applications across physics, chemistry, astronomy, and engineering.

Keywords: Bohr model, electron transition, wavelength of light, spectral lines, hydrogen atom, energy levels, Balmer series, photon emission, quantum physics, spectroscopy

Frequently Asked Questions

In the Bohr model of the atom, which electronic transition results in the largest wavelength of emitted light?

The transition from a high energy level to a low energy level with the smallest energy difference produces the largest wavelength of light.

Why does a transition from a higher to a lower energy level in the Bohr model produce light with a longer wavelength?

Because the energy difference between the levels is smaller, resulting in lower energy photons with longer wavelengths.

Which specific transition in a hydrogen atom corresponds to the largest wavelength emission according to the Bohr model?

The transition from n=3 to n=2 (the Balmer alpha line) produces the longest wavelength among visible hydrogen lines, but the transition from very high levels (e.g., n=□ to n=1) results in the largest wavelength overall, such as the Lyman series in the UV.

How does the energy level difference relate to the wavelength of emitted light in the Bohr model?

The smaller the energy difference between the initial and final levels, the longer the wavelength of the emitted photon.

Is the transition from very high to very low energy levels associated with the largest wavelength in the Bohr model?

Yes, transitions from very high energy levels (like n=□) to the ground state (n=1) produce the longest wavelengths, such as the Lyman-alpha line.

Can the transition between any two energy levels in the Bohr model produce the largest wavelength? Why or why not?

No, the largest wavelength is produced by the transition with the smallest energy difference, typically from a very high to a low energy level, not between intermediate levels.

How does the concept of wavelength size help in understanding atomic emission spectra in the Bohr model?

It helps identify which electronic transitions produce visible, ultraviolet, or infrared light based on their wavelength, corresponding to their energy differences.

Additional Resources

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The Bohr model of the atom has historically played a foundational role in our understanding of atomic structure and the behavior of electrons within atoms. Developed by Niels Bohr in 1913, this model introduced the revolutionary idea that electrons orbit the nucleus in discrete energy levels, or shells, and that transitions between these levels involve the absorption or emission of specific amounts of energy. When studying the spectral lines emitted or absorbed by atoms, the Bohr model provides a simplified yet powerful framework to predict and analyze these phenomena. One intriguing question that arises from this model is: which electronic transition corresponds to the largest wavelength of light? Since wavelength and energy are inversely related, transitions involving the smallest energy differences produce the longest wavelengths, often in the infrared or even radio regions. This review delves into the core concepts of the Bohr model, explores the relationship between energy levels and spectral lines, and identifies the particular electronic transition that results in the largest wavelength of emitted light.

Fundamentals of the Bohr Model of the Atom

The Bohr model conceptualizes the atom as a small, positively charged nucleus surrounded by electrons that travel in specific, quantized circular orbits. This model was primarily designed to explain the spectral lines of the hydrogen atom but has broader implications for understanding atomic emissions and absorptions in more complex systems.

Key Features of the Bohr Model

- Quantized Energy Levels: Electrons occupy discrete orbits characterized by an integer quantum number n (n=1, 2, 3, ...). Each orbit corresponds to a specific energy level.
- Stable Orbits: Electrons in these allowed orbits do not radiate energy, contrary to classical physics expectations.
- Energy Transitions: When an electron jumps from a higher energy level to a lower one, a photon of a specific wavelength is emitted; the reverse process involves photon absorption.
- Quantization of Angular Momentum: The angular momentum of an electron in orbit is quantized, given by $L = n\hbar$, where $\hbar$ is the reduced Planck constant.
- Spectral Lines: The emitted or absorbed light has specific wavelengths corresponding to the energy differences between levels, explaining the observed spectral lines.

Energy Levels and Transitions in the Hydrogen Atom

The energy of an electron in a given orbit n in the hydrogen atom is given by the formula:

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

where 13.6 eV is the ionization energy of hydrogen. The negative sign indicates bound states, with n=1 being the ground state.

Energy Difference Between Levels

The energy difference between two levels n_i (initial) and n_f (final) is:

$$ \Delta E = E_{n_f} - E_{n_i} = -13.6 \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \text{eV} $$

Since the energy of the photon emitted or absorbed equals this difference, the wavelength λ of the photon can be calculated using Planck's relation:

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

where h is Planck's constant and c is the speed of light.

Relation Between Wavelength and Energy Transition

The inverse relationship between wavelength and energy implies that transitions involving smaller energy differences produce longer wavelengths. Specifically:

- Large energy differences (e.g., from high n to n=1) produce short-wavelength, high-energy photons (ultraviolet or X-ray).
- Small energy differences (e.g., from n=2 to n=3 or n=3 to n=4) produce longer wavelengths, potentially in the visible, infrared, or radio regions.

This relationship guides us to determine which transition produces the largest wavelength of light.

Identifying the Transition with the Largest Wavelength

Given the inverse proportionality, the transition with the smallest energy difference will correspond to the largest wavelength. In the hydrogen atom, the smallest energy difference occurs when the electron moves between neighboring high energy levels, especially transitions from very high n to n=1 or from n=3 to n=2, etc.

Key points:

- The smallest energy difference is typically between levels that are close together at high n.
- The largest wavelength emitted corresponds to the smallest energy transition, generally between the highest levels (large n) and the next level down.

Transition from n = □ to n = 1 (Ionization Limit)

While ionization (from n=1 to n=□) involves an infinite energy difference, the transition from n=□ (free electron) to n=1 (ground state) corresponds to the absorption of a photon with the energy equal to the ionization energy (~13.6 eV), which results in ultraviolet light.

Note: This transition does not produce a wavelength of the largest size; it produces the shortest wavelength.

Transition Between High Energy Levels: n = 4 to n = 3

The transition from n=4 to n=3 involves a small energy difference and thus an emission of a photon with a relatively long wavelength compared to higher-energy transitions.

Calculations:

Using the formula:

$$ \Delta E = 13.6 \left(\frac{1}{3^2} - \frac{1}{4^2} \right) = 13.6 \left(\frac{1}{9} - \frac{1}{16} \right) $$

$$ \Delta E = 13.6 \left(\frac{16 - 9}{144} \right) = 13.6 \times \frac{7}{144} \approx 0.66\, \text{eV} $$

Corresponding wavelength:

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

Using:

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

$$ \lambda \approx \frac{(4.1357 \times 10^{-15})(3 \times 10^8)}{0.66} \approx \frac{1.24 \times 10^{-6}\, \text{m}}{0.66} \approx 1.88 \,\mu\text{m} $$

This falls in the infrared region, indicating a relatively large wavelength.

Transition Between Very High Levels: n = 10 to n = 9

Similarly, for n=10 to n=9:

$$ \Delta E = 13.6 \left(\frac{1}{9^2} - \frac{1}{10^2} \right) = 13.6 \left(\frac{1}{81} - \frac{1}{100} \right) $$

\[\Delta E = 13.6 \times \left(\frac{100 - 81}{8100} \right) = 13.6 \times \frac{19}{8100} \approx 0.0318\,

\text{eV} \]

Corresponding wavelength:

$$\lambda \approx \frac{1.24 \times 10^{-6}}{0.0318} \approx 39\, \mu\text{m}$$

This is deep into the infrared, even longer than the previous example.

Conclusion: Which Transition Corresponds to the Largest Wavelength?

From the above analysis, the transition with the smallest energy difference produces the longest wavelength of emitted light. For the hydrogen atom, the smallest energy difference occurs between very high energy levels, specifically:

- Transition from n = □ to n = 1: corresponds to ionization, not emission.
- Transition from $n = n_{max}$ (large n) to $n = n_{max} - 1$: produces the longest wavelength among the emission lines.

In practical terms, the transition from a very high n level (such as n=□) down to n=1 would emit a photon with an extremely long wavelength, approaching zero energy. Among observable spectral lines, the transition from a high n level (like n=4) to a slightly lower one (n=3) produces the largest wavelength in the visible/infrared spectrum.

Therefore, the transition that corresponds to the largest wavelength of light in the hydrogen atom is the transition from a very high energy level (such as n=□) to a lower energy level (like n=1 or n=2), with the specific case often being the transition from n □ □ to n=1. This transition's associated photon would have an extremely long wavelength, in the radio or infrared region, representing the limit of the spectral series.

Additional Considerations and Broader Context

While the hydrogen atom provides a clear framework, the concept extends to other elements with more complex spectra. However, the fundamental principle remains: the smallest energy difference produces the longest wavelength. For multielectron atoms, spectral lines become more numerous and complex, but the principle of inverse relationship between energy and wavelength holds.

Features and Limitations of the Bohr Model:

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