

What Wavelength Photon Would Be Required To Ionize A Hydrogen Atom In The Ground State And Give The Ejected

What Wavelength Photon Would Be Required To Ionize A Hydrogen Atom In The Ground State And Give The Ejected electron? This question delves into fundamental concepts of atomic physics, specifically the interaction of photons with hydrogen atoms. Understanding the energy requirements for ionization helps us comprehend processes ranging from stellar radiation effects to applications in spectroscopy and quantum mechanics. In this article, we will explore the principles underlying photon-induced ionization of hydrogen, how to calculate the necessary photon wavelength, and the implications of these calculations in various scientific contexts.

Understanding Hydrogen Atom Energy Levels

The Ground State of Hydrogen

The hydrogen atom, the simplest of atoms, consists of a single proton in the nucleus surrounded by a single electron. Its energy levels are quantized, meaning the electron can only occupy specific energy states. The lowest energy state, known as the ground state, has an energy of approximately -13.6 eV relative to the ionization threshold. This energy level is denoted as $n=1$ in the Bohr model.

Energy Levels and Transition Dynamics

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

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

where n is the principal quantum number ($n=1, 2, 3, \dots$). The negative sign indicates that the electron is bound to the nucleus, and energy must be supplied to free it.

Ionization of Hydrogen: Basic Principles

What is Ionization?

Ionization occurs when an electron gains enough energy to escape the attractive pull of the nucleus, resulting in a free electron and a proton. For hydrogen in the ground state, this process requires supplying energy equal to the ionization energy, which is the energy difference between the ground state and the free electron (at zero kinetic energy).

Ionization Energy of Hydrogen

The ionization energy (IE) for hydrogen in the ground state is:

$E = 13.6 \text{ eV}$

This is the minimum amount of energy that must be delivered to the atom to eject the electron.

Photon Energy and Wavelength Relationship

Energy of a Photon

Photons, the quanta of electromagnetic radiation, carry energy given by:

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

where:

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

By rearranging, the wavelength is:

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

Calculating the Wavelength for Ionization

Since ionization requires at least 13.6 eV, the photon must have an energy equal to or greater than this value:

$E_{\text{photon}} \geq 13.6 \text{ eV}$

Converting electron volts to joules:

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

Thus:

$E_{\text{photon}} = 13.6 \times 1.602 \times 10^{-19} \text{ J} \approx 2.179 \times 10^{-18} \text{ J}$

The corresponding wavelength:

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

$\lambda = \frac{6.626 \times 10^{-34} \times 3.00 \times 10^8}{2.179 \times 10^{-18}}$

$\lambda \approx \frac{1.9878 \times 10^{-25}}{2.179 \times 10^{-18}}$

$\lambda \approx 9.11 \times 10^{-8} \text{ m}$

Converting to nanometers:

$\lambda \approx 91.1 \text{ nm}$

Therefore, a photon with a wavelength of approximately 91.1 nm or shorter (higher energy) will be capable of ionizing a hydrogen atom in its ground state.

Implications of the Wavelength Calculation

Photon Wavelength for Ionization

- Minimum Wavelength: About 91.1 nm
- Type of Radiation: Ultraviolet (UV) region of the electromagnetic spectrum

- Significance: Photons with wavelengths shorter than 91.1 nm possess enough energy to ionize hydrogen.

Photon Wavelengths Longer Than 91.1 nm

Photons with longer wavelengths (lower energy) are insufficient to ionize hydrogen in the ground state. They may cause excitation to higher energy levels but cannot free the electron.

Additional Considerations in Ionization Processes

Photons with Excess Energy

When a photon with energy greater than 13.6 eV interacts with hydrogen, the excess energy is converted into the kinetic energy of the ejected electron:

$$KE_{\text{electron}} = E_{\text{photon}} - 13.6 \text{ eV}$$

This kinetic energy influences the speed and behavior of the ejected electron.

Cross Section and Probability

The probability of ionization depends on the photon energy and the atomic cross section. The cross section peaks near the ionization threshold and decreases at higher energies.

Applications and Observations

Astronomical Significance

Ultraviolet radiation with wavelengths shorter than 91.1 nm plays a critical role in ionizing hydrogen in interstellar space, impacting star formation and nebular emission spectra.

Laboratory and Technological Applications

- Spectroscopy: Understanding ionization thresholds helps in designing experiments involving UV photons.

- Photonic Devices: UV lasers can be used to ionize hydrogen in controlled experiments.

Summary and Conclusion

To ionize a hydrogen atom in its ground state, a photon must carry an energy of at least 13.6 eV. The corresponding wavelength for such a photon is approximately 91.1 nm, placing it within the ultraviolet spectrum. Photons with shorter wavelengths (higher energy) can ionize hydrogen more

efficiently, imparting additional kinetic energy to the ejected electron. This fundamental relationship between photon energy, wavelength, and atomic ionization is pivotal in understanding various phenomena across astrophysics, spectroscopy, and quantum physics.

In essence:

- Wavelength required: ~91.1 nm
- Photon energy: 13.6 eV
- Significance: Defines the UV threshold for hydrogen ionization, critical for understanding cosmic and laboratory processes.

Frequently Asked Questions

What wavelength photon is required to ionize a hydrogen atom in its ground state?

A photon with a wavelength shorter than approximately 91.2 nanometers is required to ionize a hydrogen atom in the ground state.

How is the ionization energy of a hydrogen atom related to the photon wavelength needed for ionization?

The ionization energy corresponds to the energy of the photon, which can be calculated from its wavelength using the relation $E = hc/\lambda$; for hydrogen, this energy is about 13.6 eV, corresponding to a wavelength of roughly 91.2 nm.

What is the energy of the photon needed to ionize hydrogen from its ground state?

Approximately 13.6 electron volts (eV).

How do you calculate the wavelength of the photon required to ionize a hydrogen atom?

Use the equation $\lambda = hc / E$, where h is Planck's constant, c is the speed of light, and E is the ionization energy (13.6 eV for hydrogen).

What happens to the electron after a photon of sufficient energy ionizes a hydrogen atom?

The electron is ejected from the atom with kinetic energy equal to the photon energy minus the ionization energy, leaving behind a proton.

Is the wavelength required to ionize hydrogen longer or shorter than visible light?

It is shorter; approximately 91.2 nm, which falls in the ultraviolet (UV) region of the electromagnetic spectrum.

Can visible light ionize a hydrogen atom in the ground state?

No, visible light photons have longer wavelengths and thus lower energy than required; they cannot ionize hydrogen in the ground state.

What is the significance of the ionization wavelength in astrophysics?

It helps determine the radiation environment and the presence of UV radiation capable of ionizing hydrogen in space, influencing star formation and interstellar medium studies.

How does the energy of the ejected electron relate to the photon energy?

The kinetic energy of the ejected electron equals the photon energy minus the ionization energy of hydrogen (13.6 eV).

Why is understanding the ionization wavelength of hydrogen important in spectroscopy?

Because it allows scientists to identify UV radiation sources and study the physical conditions of astrophysical and laboratory plasmas containing hydrogen.

Additional Resources

Photon Wavelength for Hydrogen Ionization: An In-Depth Exploration

Introduction

In the realm of atomic physics, understanding the interaction between photons and atoms is fundamental to grasping the behavior of matter at the quantum level. Among the simplest and most studied atoms is hydrogen, whose single electron orbiting a proton forms the basis for much of modern physics. A key question often posed by students, educators, and researchers alike is: What wavelength of photon is required to ionize a hydrogen atom in its ground state, and what are the characteristics of the ejected electron?

This query opens a window into the elegant interplay between quantum energy levels, photon properties, and atomic structure. In this comprehensive review, we will dissect this problem step-by-step, covering the essential concepts, calculations, and implications involved.

Understanding the Ground State of Hydrogen

Before diving into the specifics of photon energy, it's crucial to understand what the "ground state" of hydrogen signifies.

The Hydrogen Atom and Its Energy Levels

Hydrogen's electron energy levels are quantized, meaning the electron can only occupy specific, discrete energy states. These are given by the Bohr model and refined by quantum mechanics, with energies expressed as:

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

where:

- E_n is the energy of the level corresponding to principal quantum number n ,
- $n = 1, 2, 3, \dots$.

Ground state ($n=1$):

$$E_1 = -13.6 \text{ eV}$$

This negative value indicates the electron is bound within the atom, with the magnitude representing the energy required to free it entirely from the atom (ionization energy).

Ionization Energy of Hydrogen

To free the electron from the ground state, a photon must supply at least the energy equivalent to the ionization energy, which is the energy difference between the ground state and the free state (which is considered zero energy).

Ionization energy of hydrogen:

$$\Delta E_{\text{ionization}} = 0 \text{ eV} - (-13.6 \text{ eV}) = 13.6 \text{ eV}$$

This value sets the minimum energy threshold a photon must have to ionize hydrogen in its ground state.

Connecting Photon Energy to Wavelength

The energy of a photon (E) relates to its wavelength (λ) via the fundamental relation:

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

where:

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

Converting the ionization energy from electronvolts to joules:

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

Thus:

$$E = 13.6 \text{ eV} \times 1.602 \times 10^{-19} \text{ J/eV} = 2.179 \times 10^{-18} \text{ J}$$

This is the minimum photon energy required for ionization.

Calculating the Wavelength Needed for Ionization

Using the relation:

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

Substituting the known values:

$$\lambda = \frac{(6.626 \times 10^{-34} \text{ Js}) \times (3.00 \times 10^8 \text{ m/s})}{2.179 \times 10^{-18} \text{ J}} \approx 91.2 \text{ nm}$$

Result:

> To ionize a hydrogen atom in its ground state, a photon with a wavelength of approximately 91.2 nanometers (nm) or shorter is required.

Significance of Wavelength and Photon Energy

This wavelength falls within the ultraviolet (UV) range of the electromagnetic spectrum. Specifically, the UV spectrum spans from about 10 nm to 400 nm, making 91.2 nm a high-energy UV photon. Such photons are capable of penetrating the outer layers of Earth's atmosphere but are blocked by ozone, which is why UV-based ionization experiments are typically conducted in space or specialized laboratories.

What Happens During Ionization?

When a hydrogen atom absorbs a photon with energy equal to or greater than 13.6 eV, two primary outcomes are possible:

- Photon energy exactly matches the ionization threshold:
The electron is ejected with virtually zero kinetic energy, just enough to escape the Coulomb attraction of the proton.
- Photon energy exceeds the threshold:
The excess energy becomes the kinetic energy ($K.E.$) of the ejected electron, resulting in a higher velocity.

This process is described by the photoelectric effect, which is fundamental to quantum physics.

The Ejected Electron: Kinetic Energy and Speed

Given an incident photon of energy (E_{photon}), the kinetic energy ($K.E.$) of the ejected electron is:

$$K.E. = E_{\text{photon}} - E_{\text{ionization}}$$

For the minimum energy photon (at threshold):

$$\text{K.E.} = 0$$

For photons with energy greater than 13.6 eV:

$$\text{K.E.} = (E_{\text{photon}} - 13.6 \text{ eV})$$

Example:

Suppose a photon has energy of 15 eV:

$$\text{K.E.} = 15 \text{ eV} - 13.6 \text{ eV} = 1.4 \text{ eV}$$

Converting to joules:

$$1.4 \text{ eV} \times 1.602 \times 10^{-19} \text{ J/eV} = 2.2428 \times 10^{-19} \text{ J}$$

The electron's velocity (v) can be found from:

$$\text{K.E.} = \frac{1}{2} m_e v^2$$

where:

$$m_e = 9.109 \times 10^{-31} \text{ kg} \text{ (electron mass)}$$

Solving for v :

$$v = \sqrt{\frac{2 \text{ K.E.}}{m_e}} = \sqrt{\frac{2 \times 2.2428 \times 10^{-19} \text{ J}}{9.109 \times 10^{-31} \text{ kg}}} \approx 7.02 \times 10^5 \text{ m/s}$$

This speed is roughly 2.3 times the speed of sound in air, illustrating the high-energy nature of photo-ejected electrons.

Broader Implications and Applications

Understanding the photon wavelength necessary for hydrogen ionization isn't just an academic exercise—it underpins numerous technologies and scientific pursuits:

- Spectroscopy:

Determines elemental composition based on absorption and emission lines.

- Astrophysics:

Explains phenomena such as stellar atmospheres and interstellar medium ionization.

- Photovoltaics and Photoelectric Devices:

Harnesses photon energies to generate electricity.

- Quantum Mechanics and Particle Physics:

Provides foundational insight into wave-particle duality and quantum energy quantization.

Summary: Key Takeaways

Aspect	Details
Ground state energy of hydrogen	-13.6 eV
Minimum photon energy for ionization	13.6 eV
Corresponding wavelength	~91.2 nm
Photon range	Ultraviolet (UV) spectrum
Ejected electron's kinetic energy	Zero at threshold, increases with photon energy
Electron speed (for excess energy)	Up to $\sim (7 \times 10^5) \text{ m/s}$

Final Thoughts

The question of the wavelength needed to ionize hydrogen in its ground state encapsulates fundamental principles of quantum physics, wave-particle duality, and electromagnetic radiation. The precise calculation of approximately 91.2 nm underscores the quantized energy structure of atoms and highlights how specific photon energies dictate atomic interactions.

Whether you're a student grappling with quantum concepts, a researcher probing the cosmos, or an engineer designing UV-based technologies, understanding these relationships forms the cornerstone of modern scientific inquiry and technological innovation.

In essence, to liberate an electron from a hydrogen atom in its lowest energy configuration, a photon in the ultraviolet spectrum—specifically, with a wavelength shorter than about 91.2 nanometers—must be absorbed, providing the necessary energy to overcome the atom's ionization barrier and set the electron free with a spectrum of kinetic energies depending on the excess photon energy.

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