HL IB Physics

Structure of the Atom

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Rutherford's Gold Foil Experiment

Rutherford's Gold Foil Experiment

- Evidence for the structure of the atom was discovered by Ernest Rutherford at the beginning of the 20th century from the study of **alpha particle scattering**
- The experiment consisted of beams of high-energy alpha particles fired at thin gold foil and a detector on the other side to determine
  - The different **angles of deflection** of the alpha particles
  - The **number of alpha particles** that were deflected at each angle

Apparatus for the Rutherford Scattering Experiment

- The setup for the scattering experiment consisted of:
  - A source of alpha particles in a lead container
  - A thin sheet of gold foil
  - A movable detector
  - An evacuated chamber

Purpose of the lead container

- Alpha particles are emitted in all directions, so the source was placed in a **lead container**
- This was to produce a collimated beam of alpha particles
- This is because alpha particles are absorbed by lead, so a long narrow hole at the front allowed a concentrated beam of alpha particles to escape and be directed as needed

Purpose of the thin sheet of gold foil

- The target material needed to be **extremely thin**, about 10^{-6} m thick
- This is because a thicker foil would stop the alpha particles completely
- Gold was chosen due to its malleability, meaning it was easy to hammer into thin sheets

Purpose of the evacuated chamber

- Alpha particles are highly ionising, meaning they only travel about 5 cm before interacting with molecules of air
- So, the apparatus was placed in an **evacuated chamber**
- This was to ensure that the alpha particles did not collide with any particles on their way to the foil target

Findings from the Rutherford Scattering Experiment

- An alpha (α) particle is the nucleus of a helium atom, so it has a **positive** charge
- The observations from Rutherford's experiment were:

A. The majority of α-particles passed straight through the foil undeflected

  - This suggests the atom is mostly empty space

B. Some α-particles deflected through small angles of <10°

  - This suggests there is a positive nucleus at the centre (since two positive charges would repel)

C. Only a small number of α-particles deflected straight back at angles of >90°

  - This suggests the nucleus is extremely small and is where most of the mass and charge of the atom are concentrated
  - This led to the conclusion that atoms consist of small, dense positively charged nuclei surrounded by negatively charged electrons

Emission & Absorption Spectrum

Spectra & Atomic Energy Levels

- Atomic spectra are observed when atoms **emit** or **absorb** light of certain wavelengths
  - These are known as **emission spectra** and **absorption spectra**
- Atomic spectra provide **evidence** that electrons in atoms can only transition between discrete atomic energy levels

Emission Spectra

- Emission spectra can be produced by heating a low-pressure gas
  - Heating provides energy to **excite** electrons to higher energy levels
  - When an electron transitions back to a **lower** energy level, it **emits** a photon
- Each transition corresponds to a **specific wavelength** of light which correlates to an observable spectral line
- The resulting **emission spectrum** contains a set of discrete wavelengths, represented by coloured lines on a black background

Absorption Spectra

- Absorption spectra can be produced by passing white light through a **cool, low-pressure gas**
  - Only photons with the exact energy required to excite electrons will be absorbed
- Each absorbed photon corresponds to a **specific wavelength** of light which correlates to an observable dark line in a continuous spectrum of wavelengths
- The resulting **absorption spectrum** contains a set of discrete wavelengths, represented by dark lines on a coloured background
  - These lines correspond to the same lines observed on an emission spectrum for the **same** element

Spectra & Chemical Composition

- The **chemical composition** of a substance can be investigated using emission and absorption spectra
- Each element produces a unique pattern of spectral lines
- No two elements produce the same set of spectral lines, therefore, elements can be identified by their atomic spectrum

- For example:
  - **Hydrogen** is known to produce strong spectral lines in the **red** portion of the visible spectrum, at **656 nm**
  - When sodium is burned, a characteristic **yellow** flame is observed due to it producing strong spectral lines in the yellow portion of the spectrum, at **589 nm**
  - When mercury is burned, most of the emission lines are below **450 nm**, which produces a characteristic **blue** light
- Elements such as **sodium** and **mercury** are known for their use in street lights, as well as **neon** for its use in colourful signs
- This can be achieved when
  - An electrical discharge is applied to the vapourised substance
  - The energy supplied excites orbital electrons within individual atoms to a higher energy state
  - When the electrons move back down to the ground state, a specific wavelength of light is emitted

Photon Energy

Photons & Atomic Transitions

The Photon Model

- Photons are fundamental particles that make up all forms of electromagnetic radiation
- A photon is defined as

A massless "packet" or a "quantum" of electromagnetic energy

- This means that the energy transferred by a photon is not continuous but as discrete packets of energy
  - In other words, each photon carries a specific amount of energy and transfers this energy all in one go
  - This is in contrast to waves which transfer energy continuously

Atomic Energy Levels

- Electrons in an atom occupy certain energy states called **energy levels**
  - Electrons will occupy the **lowest** possible energy level as this is the most **stable** configuration for the atom
  - When an electron **absorbs** or **emits** a photon, it can move between these energy levels, or be removed from the atom completely

Excitation

- When an electron moves to a higher energy level, the atom is said to be in an **excited state**
  - To **excite** an electron to a higher energy level, it must **absorb** a photon
- Electrons can also move back down to a lower energy level by **de-excitation**
  - To **de-excite** an electron to a lower energy level, it must **emit** a photon

Ionisation

- When an electron is removed from an atom, the atom becomes **ionised**
  - An electron can be removed from any energy level it occupies
  - However, the **ionisation energy** of an atom is the **minimum** energy required to remove an electron from the **ground state** of an atom

Representing Energy Levels

- Energy levels can be represented as a series of horizontal lines
  - The line at the bottom with the greatest negative energy represents the ground state
  - The lines above the ground state with decreasing energies represent excited states
  - The line at the top, usually 0 V or infinity ∞, represents the ionisation energy

Energy Levels in a Hydrogen Atom

Worked example

Explain how atomic spectra provide evidence for the quantisation of energy in atoms.

Answer:

Step 1: Outline the meaning of atomic spectra

- Atomic spectra show the spectrum of discrete wavelengths emitted or absorbed by a specific atom

Step 2: Describe the relationship between energy and wavelength

- Photon energy is related to frequency and wavelength
- Therefore, photons with discrete wavelengths have discrete energies equal to the difference between two energy levels

Step 3: Explain how atomic spectra give evidence for the quantisation of energy

- Photons arise from electron transitions between energy levels
- This happens when an electron is excited, or de-excited, from one energy level to another, by either emitting or absorbing light of a specific wavelength
- Since atomic spectra are made up of discrete wavelengths, this shows that atoms must contain discrete, or quantised, energy levels

Calculating Photon Energy

- Each line of the emission spectrum corresponds to a different **energy level transition** within the atom
  - Electrons can transition between energy levels absorbing or emitting a **discrete amount of energy**
  - An excited electron can transition down to the next energy level or move to a further level closer to the ground state
- For example, if an atom has **six** energy levels:
  - At low temperatures, most electrons will occupy the ground state n = 1
  - At high temperatures, electrons may be excited to the most excited state n = 6
- The energy of a photon can be calculated using the formula:

$$E = hf$$

- Using the wave equation, energy can also be equal to:

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

- Where:
  - E = energy of the photon (J)
  - h = Planck's constant (J s)
  - c = the speed of light (m s^{-1})
  - f = frequency (Hz)
  - λ = wavelength (m)
- This equation tells us:
  - The higher the frequency of EM radiation, the higher the energy of the photon
  - The energy of a photon is inversely proportional to the wavelength
  - A long-wavelength photon of light has a lower energy than a shorter-wavelength photon

Difference in discrete energy levels

- The difference between two energy levels is equal to a specific photon energy
- The energy of the photon is given by:

$$\Delta E = hf = E_2 - E_1$$

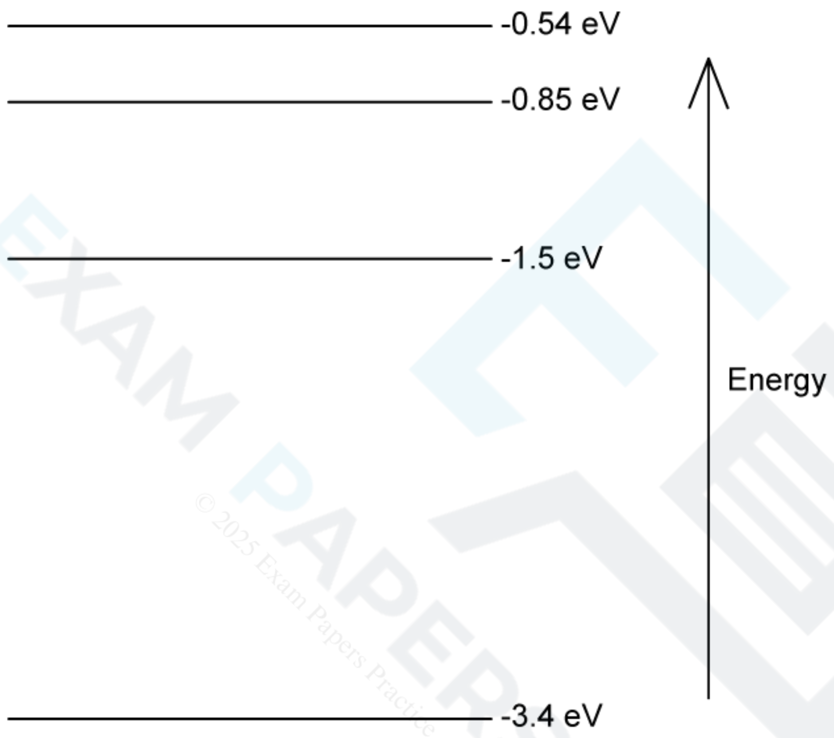
- Where:
  - E_1 = energy of the lower level (J)
  - E_2 = energy of the higher level (J)
- Using the wave equation, the wavelength of the emitted, or absorbed, radiation can be related to the energy difference by the equation:

$$\lambda = \frac{hc}{E_2 - E_1}$$

- This equation shows that:
  - The larger the difference in energy between two levels ΔE, the shorter the wavelength λ and vice versa

Worked example

Some electron energy levels in atomic hydrogen are shown below.



The longest wavelength produced as a result of electron transitions between two of the energy levels is 4.0×10^{-6} m.

(a) Draw an arrow to show the transition that would produce:

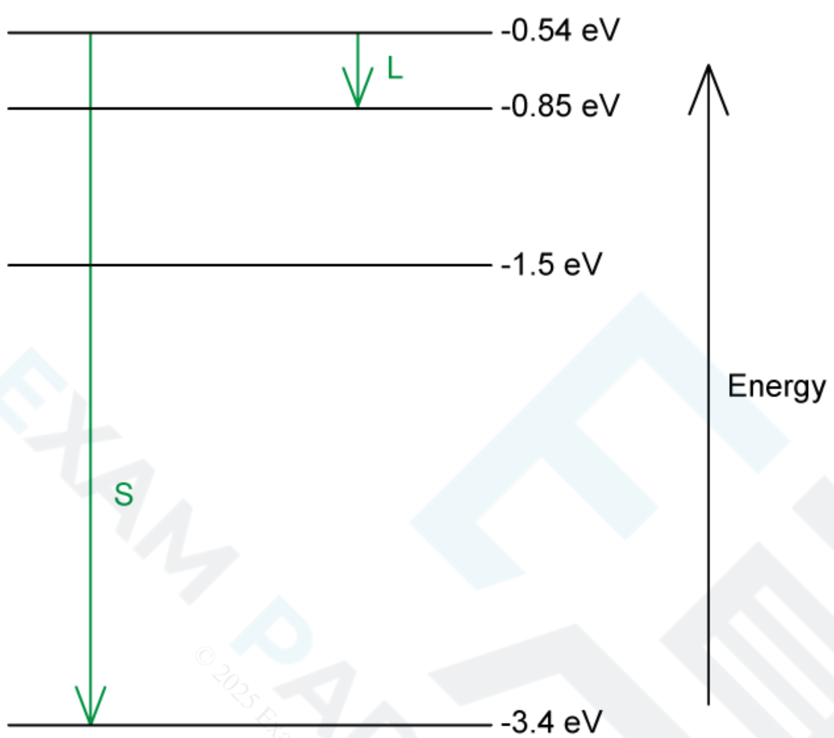
- A photon of wavelength 4.0×10^{-6} m. Mark with the letter **L**.
- The photon with the shortest wavelength. Mark with the letter **S**.

(b) Calculate the wavelength for the transition giving rise to the shortest wavelength.

Answer:

(a)

- Photon energy and wavelength are inversely proportional
- Therefore, the largest energy change corresponds to the shortest wavelength (line **S**)
- The smallest energy change corresponds to the longest wavelength (line **L**)



(b)

Step 1: Write down the equation linking wavelength and energy

$$\lambda = \frac{hc}{\Delta E} = \frac{hc}{E_2 - E_1}$$

Step 2: Identify the energy levels that give rise to the shortest wavelength

- The shortest wavelength photon will come from a transition between the energy levels that have the largest difference:
  - $E_2 = -0.54 \text{ eV}$
  - $E_1 = -3.4 \text{ eV}$
- Therefore, the greatest possible difference in energy is

$$\Delta E = E_2 - E_1 = -0.54 - (-3.4) = 2.86 \text{ eV}$$

Step 3: Calculate the wavelength

- To convert from eV to J: multiply by 1.6×10^{-19} J

$$\lambda = \frac{(6.63 \times 10^{-34})(3.0 \times 10^{8})}{2.86 \times (1.6 \times 10^{-19})}$$

$\lambda = 4.347 \times 10^{-7}$ m = 435 nm

Rutherford Scattering & Nuclear Radius (HL)

Nuclear Radius

- The radius of a nucleus depends on the nucleon number *A* of the atom
  - The greater the number of nucleons a nucleus has, the greater the space the nucleus occupies, hence giving it a larger radius
- The exact relationship between the **radius** and **nucleon number** can be determined from experimental data, such as Rutherford scattering
- By doing this, physicists were able to deduce the following relationship:

$$R = R_0 A^{\frac{1}{3}}$$

- Where:
  - R = nuclear radius (m)
  - A = nucleon / mass number
  - R_0 = Fermi radius
- The constant of proportionality $R_0 = 1.20 \times 10^{-15}$ m is known as the **Fermi radius**
- This is the radius of a hydrogen nucleus which contains only one proton (A = 1)

Nuclear Density

- Assuming that the nucleus is spherical, its volume is equal to:

$$V = \frac{4}{3}\pi R^3$$

- Combining this with the expression for nuclear radius gives:

$$V = \frac{4}{3}\pi\left(R_0 A^{\frac{1}{3}}\right)^3 = \frac{4}{3}\pi R_0^{\ 3} A$$

- This tells us that the nuclear volume *V* is proportional to the mass of the nucleus m, which is equal to

$$m = Au$$

- Where *u* = atomic mass unit (kg)
- Using the definition for density, nuclear density is equal to:

$$\rho = \frac{m}{V}$$

$$\rho = \frac{Au}{\frac{4}{3}\pi R_0^{\ 3} A} = \frac{3u}{4\pi R_0^{\ 3}}$$

- Since the mass number A cancels out, the remaining quantities in the equation are all **constants**
- Therefore, this shows the density of the nucleus is:
  - The **same** for all nuclei
  - Independent of the **radius**
- The fact that nuclear density is constant shows that nucleons are **evenly separated** throughout the nucleus regardless of their size
- The accuracy of nuclear density depends on the accuracy of the constant R_0
  - As a guide, nuclear density should always be of the order 10^{17} kg m^{-3}
- Nuclear density is significantly **larger** than atomic density which suggests:
  - The **majority** of the atom's mass is contained in the nucleus
  - The nucleus is very **small** compared to the atom
  - Atoms must be predominantly **empty space**

Worked example

Determine the value of nuclear density.

You may take the constant of proportionality R_0 to be 1.20 fm.

Answer:

Step 1: Derive an expression for nuclear density

- Using the equation derived above, the density of the nucleus is:

$$\rho = \frac{3u}{4\pi R_0^{\ 3}}$$

Step 2: List the known quantities

- Atomic mass unit, $u = 1.661 \times 10^{-27}$ kg
- Constant of proportionality, $R_0 = 1.20$ fm $= 1.20 \times 10^{-15}$ m

Step 3: Substitute the values to determine the nuclear density

$$\rho = \frac{3 \times (1.661 \times 10^{-27})}{4\pi (1.20 \times 10^{-15})^3} = 2.3 \times 10^{17} \text{ kg m}^{-3}$$

Rutherford Scattering Experiment

- In the Rutherford scattering experiment, alpha particles are fired at a thin gold foil
- Initially, before interacting with the foil, the particles have kinetic energy equal to

$$E_k = \frac{1}{2}mv^2$$

- Some of the alpha particles are found to come straight back from the gold foil
- This indicates that there is **electrostatic repulsion** between the alpha particles and the gold nucleus
- At the point of closest approach *d*, the repulsive force **reduces** the speed of the alpha particles to **zero** momentarily, before any change in direction
- At this point, the **initial kinetic energy** E_k of the alpha particle is equal to the **electric potential energy** E_p of the target nucleus:

$$E_k = E_p$$

- Where the electric potential energy is given by

$$E_p = k\frac{Qq}{d}$$

- Where:
    - Charge of an alpha particle, *Q* = 2e
    - Charge of a target nucleus, *q* = Ze
    - *Z* = proton (atomic) number
    - *e* = elementary charge (C)
    - k = Coulomb constant

- This gives an expression for the **potential energy** at the point of **repulsion**:

$$E_p = k\frac{(2e)(Ze)}{d} = k\frac{2Ze^2}{d}$$

- Which, due to the conservation of energy also gives the **initial kinetic energy** possessed by the alpha particle
- Rearranging for the distance of closest approach *d*

$$d = k\frac{2Ze^2}{E_p} = k\frac{2Ze^2}{E_k}$$

- This gives a value for the **radius** of the nucleus, assuming the alpha particle is fired at a high energy

Deviations from Rutherford Scattering (HL)

Deviations from Rutherford Scattering

- Rutherford's scattering experiment predicted that
  - As the scattering angle **increases**, the number of alpha particles scattered at that angle **decreases**
- This was found to be correct at low to moderately high energies
- However, at **very high energies** (>27.5 MeV) significant **deviations** from Rutherford's predictions were observed
- Instead of the number decreasing at the expected rate, the number of alpha particles being back-scattered **sharply decreases** to zero

- Rutherford's alpha scattering experiment originally assumed that the alpha particles **only** interact through **electrostatic repulsion**
- However, if the energy of the alpha particles exceeds 27.5 MeV, then they will be close enough to **interact** with the nucleus via the strong nuclear force
- Factoring in the interactions due to the strong nuclear force explains the scattering pattern observed in the experimental results
- Therefore, deviations from Rutherford scattering provide **evidence** for the **strong nuclear force**

Worked example

Alpha particles undergo scattering after being fired at a thin gold $^{197}_{79}\text{Au}$ foil. The gold is then replaced to make a comparison.

Describe the predicted difference in the scattering pattern when the foil is replaced with aluminium $^{30}_{13}\text{Al}$ foil of the same thickness.

Answer:

Step 1: Compare the relative charges of the nuclei

- The force between nuclei due to the electric repulsion is

$$F = k\frac{qQ}{r^2} \Rightarrow r^2 F = kqQ$$

- Therefore, the charge of a nucleus is proportional to the square of the distance between it and an alpha particle

$$Q \propto r^2$$

- Gold has 79 protons, so $Q_{gold} = +79e$
- Aluminium has 13 protons, so $Q_{aluminium} = +13e$
- Therefore, an alpha particle will get closer to the nucleus with less charge i.e. the aluminium nucleus than the gold nucleus

Step 2: Predict the patterns and deviations from Rutherford scattering

- Deviations from Rutherford scattering occur when alpha particles get close enough for the strong nuclear force to begin to become more significant than the electric force
- At very small separations (<1.5 fm) the effect of the strong nuclear force becomes significant
- Alpha particles will be able to get closer to aluminium nuclei at lower energies than the gold nuclei
- Therefore, alpha particles will be less affected by electric repulsion and able to get close enough for interactions with the strong nuclear force
- Hence, more deviation will be seen with aluminium foil than with gold foil

The Bohr Model of Hydrogen (HL)

The Bohr Model of Hydrogen

- Hydrogen is the simplest atom in existence, making it ideal for experiments investigating the nature of electron energy levels
- Line spectra produced by hydrogen atoms showed that
  - Electrons are able to jump, or **transition**, between specific energy levels producing specific energy photons
  - Different transitions can be categorised into **series**, or families, of lines
- The Lyman series converges on the **ground state** ($n = 1$) for electrons
  - The Balmer series converges on the **second** energy level ($n = 2$)
  - The Ritz-Paschen converges on the **third** energy level ($n = 3$) and so on

- The Lyman series photons will have the **highest** energies since they have the **shortest** wavelength
  - These transitions tend to produce **ultraviolet** photons
- The Pfund series photons will have the **lowest** energies since they have the **longest** wavelength
  - These transitions tend to produce **infrared** photons

- The finding of these electron transitions helped scientists to understand how electrons work to produce photons of specific wavelength and energy
- This led to the development of the **Bohr model of hydrogen**, which states that
  - Electrons can only move in fixed orbits
  - The orbital radius of electrons is restricted to certain values
- The **discrete energy** of the transitions in the Bohr model for hydrogen are described by the equation:

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

- Where
  - E = photon energy (J)
  - n = an integer 1, 2, 3 etc. to describe the energy level of an atom

Worked example

Determine the frequency of an emitted photon from a hydrogen atom when an electron makes a transition between levels n = 4 and n = 2.

Answer:

Step 1: List the known quantities

- Transition between n = 4 and n = 2
- Planck's constant, $h = 6.63 \times 10^{-34}$ J s
- Electronvolt, eV = 1.6×10^{-19} J

Step 2: Determine an equation for the change in energy Δ*E*

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

$$\Delta E = E_4 - E_2$$

$$\Delta E = \left(-\frac{13.6\ eV}{4^4}\right) - \left(-\frac{13.6\ eV}{2^4}\right)$$

Step 3: Calculate the change in energy, in eV, for the photon using the given equation

$$\Delta E = -13.6\left(\frac{1}{4^2} - \frac{1}{2^2}\right) = 2.55\ eV$$

Step 4: Rearrange the photon energy equation for frequency f

$$E = hf \quad \Rightarrow \quad f = \frac{E}{h}$$

Step 5: Substitute the known values into the equation for frequency

$$f = \frac{2.55 \times (1.6 \times 10^{-19})}{6.63 \times 10^{-34}} = 6.15 \times 10^{14}\ \text{Hz}$$

Quantisation of Angular Momentum

- Angular momentum is a property of any spinning or rotating body, very similar to linear momentum
  - In linear motion, momentum is the product of mass and velocity
  - In rotational motion, the momentum is the product of moment of inertia and angular speed
- Angular momentum is a vector, which means:
  - The **magnitude** is equal to the momentum of the particle times its radial distance from the centre of its circular orbit
  - The **direction** of the angular momentum vector is normal to the plane of its orbit with the direction being given by the corkscrew rule

- Niels Bohr proposed that the angular momentum *L* of an electron in an energy level is quantised in integer multiples of Planck's constant over 2π:

$$L = n\frac{h}{2\pi}$$

- Where:
  - *n* = an integer (n = 1, 2, 3...)
  - *h* = Planck's constant
- Hence the angular momentum for an electron in a circular orbit is **constant**

- De Broglie proposed that an electron with momentum *p* = *mv* has a wavelength λ given by

$$\lambda = \frac{h}{p}$$

- For an electron moving in a straight line, the matter wave takes a familiar wave shape consisting of peaks and troughs
  - Although the electron itself isn't oscillating up and down, only the matter wave is

- For the same electron moving in a circle, the matter wave still has a sinusoidal shape but is wrapped into a circle

Worked example

Determine the velocity of the electron in the first Bohr orbit of the hydrogen atom.

You may use the following values:

- Mass of an electron = 9.1×10^{-31} kg
- Radius of the orbit = 0.529×10^{-10} m
- Planck's constant = 6.63×10^{-34} kg m^2 s^{-1}

Answer:

Step 1: List the known quantities

- First orbital level, n = 1
- Mass of an electron, $m = 9.1 \times 10^{-31}$ kg
- Radius of the orbit, $r = 0.529 \times 10^{-10}$ m
- Planck's constant, $h = 6.63 \times 10^{-34}$ kg m^2 s^{-1}

Step 2: Write the Bohr Condition equation and rearrange for velocity, *v*

$$\frac{nh}{2\pi} = mvr \quad \Rightarrow \quad v = \frac{nh}{2\pi mr}$$

Step 3: Substitute the values in and calculate the velocity *v*

$$v = \frac{1 \times (6.63 \times 10^{-34})}{2\pi \times (9.1 \times 10^{-31})(0.529 \times 10^{-10})}$$

Step 4: Write the final answer

Velocity of an electron ($n = 1$): $v = 2.2 \times 10^6$ m s^{-1}