Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

(https://intercom.help/kognity)

Teacher view

Index

Table of contents

Notebook

Glossary

Reading assistance

E. Nuclear and quantum physics / E.1 Structure of the atom

The big picture

? Guiding question(s)

- What is the current understanding of the nature of an atom?
- What is the role of evidence in the development of models of the atom?
- In what ways are previous models of the atom still valid despite recent advances in understanding?

Keep the guiding questions in mind as you learn the science in this subtopic. You will be ready to answer them at the end of this subtopic. The guiding questions require you to pull together your knowledge and skills from different sections, to see the bigger picture and to build your conceptual understanding.

The Sun is a star and emits electromagnetic radiation. We cannot travel to stars, like the Sun, so we can only deduce their properties by looking at the light they emit.

Figure 1 shows the spectrum of visible light emitted by the Sun. By studying the light, we know that the Sun is composed mainly of hydrogen and helium with some heavier elements, such as sodium.

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

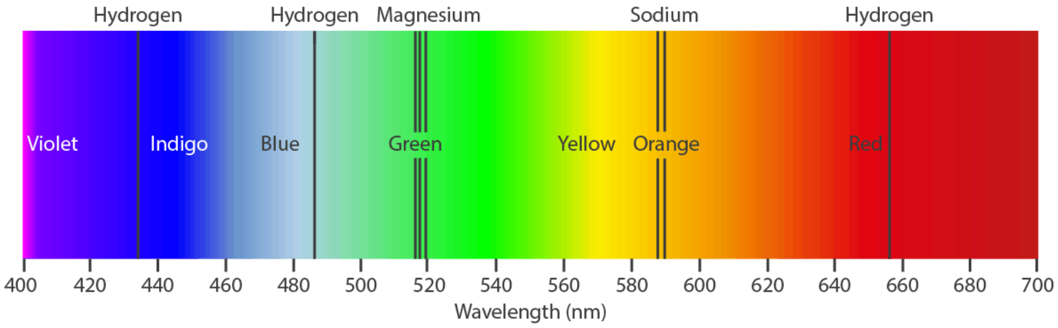


Figure 1. The visible light spectrum of the Sun.

More information for figure 1

The image depicts the visible light spectrum of the Sun, which ranges from 400 to 700 nanometers (nm) in wavelength. The spectrum displays distinct color bands from violet (400 nm) to red (700 nm), passing through indigo, blue, green, yellow, and orange. Superimposed on these bands are vertical lines, which are absorption lines indicating the presence of specific elements within the Sun's atmosphere. The absorption lines for hydrogen are visible at several points across the spectrum, specifically around the violet, blue, and red bands. Magnesium lines appear in the green region, while sodium lines are in the yellow area. This spectrum is used to deduce the composition of the Sun, showing it contains hydrogen, helium, and some heavier elements like sodium and magnesium. The absorption lines correspond to different elements present in the Sun, helping scientists understand the atomic structure and processes occurring within the star.

[Generated by AI]

Why are there dark lines in the visible light spectrum of the Sun? How do we know that these lines correspond to different elements in the Sun? How does this relate to the structure of the atom?

Prior learning

Before you study this subtopic make sure that you understand the following:

- Basic structure of the atom, including protons, neutrons and electrons.
- Electric force and Coulomb's law (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)).
- Gravitational potential energy (see subtopic A.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43083/)).
- Electric potential energy (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)).

Student
view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

E. Nuclear and quantum physics / E.1 Structure of the atom

The nucleus and atomic energy levels

Learning outcomes

By the end of this section you should be able to:

- Describe the Geiger—Marsden—Rutherford experiment and how it led to the discovery of the nucleus.
- Understand that photons are emitted and absorbed during atomic transitions and use the equation:

$$E = hf$$

- Understand that emission and absorption spectra provide evidence for discrete atomic energy levels and chemical composition.

Humans have always wondered what matter is made from. Democritus, an Ancient Greek philosopher, thought that all matter was made of solid hard lumps, which he called ‘atomos’.

In 1897, the English physicist J J Thomson discovered the negatively charged electron. This led to the **Thomson model** or ‘plum pudding’ model of the atom (**Figure 1**), where a dough (positive charge) has dried fruit (negatively charged electrons) embedded in it.

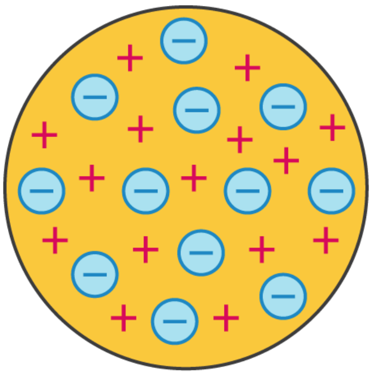


Figure 1. The Thomson model of the atom.

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Section Student... (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43191/print/) Assign

More information for figure 1

The image depicts the Thomson or 'plum pudding' model of the atom. It shows a large, yellow oval representing a positively charged substance. Embedded within the oval are multiple blue circles with negative signs, symbolizing negatively charged electrons. These electrons are scattered throughout the positively charged area, illustrating the concept that electrons are embedded in a positive matrix, similar to plums in a pudding. The positive and negative signs are evenly distributed to portray the uniform distribution of charge within the atom as per this model.

[Generated by AI]

We now know that atoms have complex structures and are made of protons, neutrons and electrons. How did we find out the structure of something so tiny as an atom?

Geiger—Marsden—Rutherford experiment

In the early 1900s, a series of experiments were performed by physicists Ernest Rutherford, Hans Geiger and Ernest Marsden. In these experiments, particles called alpha particles were fired at a thin sheet of gold foil. An alpha particle has a positive charge and is a helium nucleus, consisting of two protons and two neutrons (see subtopic E.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44319/)).

Figure 2 shows the gold foil experiment. The electric force is a field force and acts between charged particles (see subtopic A.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43136/)). It can be either attractive or repulsive. Rutherford expected that the positively charged alpha particles would interact with the positive and negative charges in the 'plum pudding' atoms in the gold foil. However, he calculated that the electric forces would be very weak and so the alpha particles would pass straight through the atoms of the gold foil or be very slightly deflected.

Most of the alpha particles did go straight through the gold foil with some minor deviations from their paths. However, some alpha particles deviated from their straight paths by greater angles. A few alpha particles were scattered at angles greater than 90°. This was very unexpected.

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

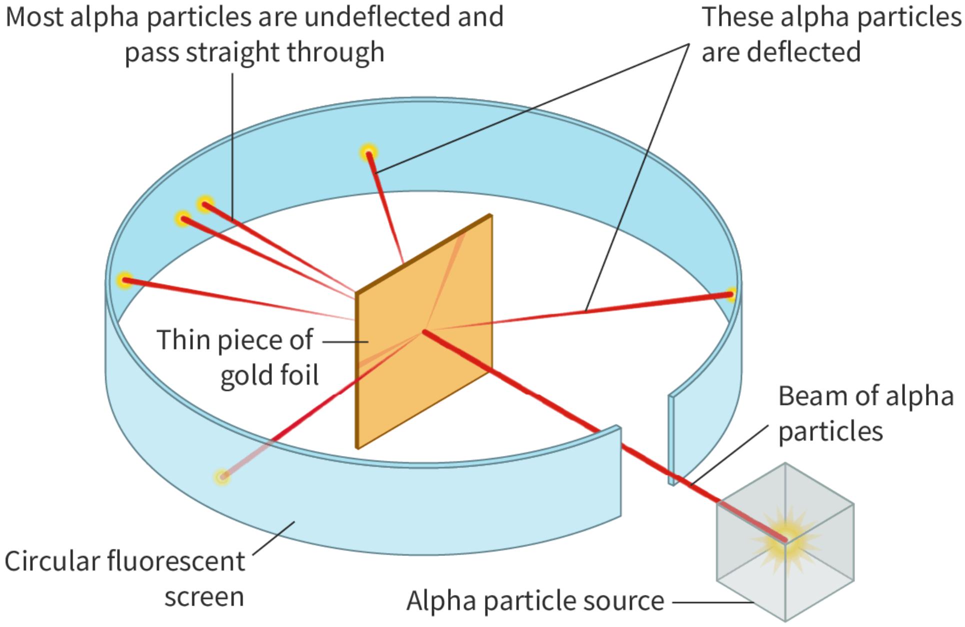


Figure 2. The gold foil experiment.

More information for figure 2

The image is a diagram illustrating the gold foil experiment conducted to observe alpha particle scattering. It features a central thin piece of gold foil, labeled accordingly, indicating the point where alpha particles interact. Surrounding the gold foil is a circular fluorescent screen, which is labeled as well. The diagram shows a beam of alpha particles being emitted from a source towards the gold foil. The paths of the alpha particles are depicted with red lines illustrating their trajectories. Most lines pass straight through the foil, indicating that most alpha particles are undeflected. A few lines demonstrate significant bending away from the direct path, showing how some alpha particles are deflected upon striking the gold foil. Text within the image helps to clarify that most particles pass through without deflection, while others deviate due to interaction with the foil. This experiment was crucial in understanding atomic structure and the presence of a dense, positively charged nucleus within atoms.

[Generated by AI]

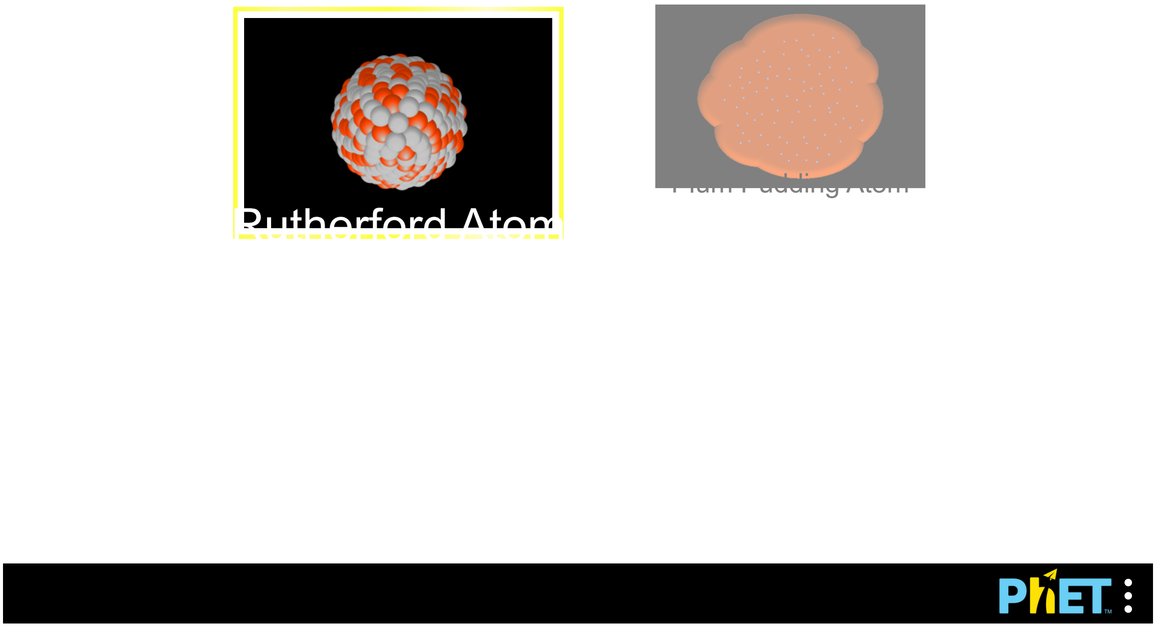
Use the simulation in **Interactive 1** to investigate Rutherford scattering.

1. Select the ‘Rutherford Atom’ tab.
2. Select ‘Traces’ in the ‘Alpha Particle’ box.
3. Click the ‘Alpha Particles’ blue button. What happens to the paths of the alpha particles?
4. At the top left are two buttons. Select the lower button, which shows the nuclear view. The upper button shows the atomic model, according to a model in which the circles represent possible ‘energy levels’ of electrons. Observe the paths of the alpha particles.

Student
view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

5. Now compare the scattering of the alpha particles with what would be expected if the ‘plum pudding’ model were true. Select the ‘Plum Pudding Atom’ tab and select ‘Traces’ in the ‘Alpha Particle’ box.
6. Click the ‘Alpha Particles’ blue button. What happens to the paths of the alpha particles?



Interactive 1. Rutherford scattering simulation.

More information for interactive 1

This interactive simulation provides a hands-on exploration of the Geiger—Marsden—Rutherford experiment, which was a breakthrough in our understanding of atomic structure. The experiment, conducted in the early 1900s, demonstrated how the scattering of alpha particles by gold foil led to the discovery of the atomic nucleus. Through this simulation, users can observe the paths of alpha particles as they interact with the atom’s structure, specifically its nucleus, allowing a deeper understanding of how Rutherford’s model of the atom was developed and how it rejected the earlier "plum pudding" model proposed by J.J. Thomson.

In the simulation, users are presented with the option to select between two atomic models: Rutherford's nuclear model and Thomson's plum pudding model. By switching between the "Rutherford Atom" and "Plum Pudding Atom" tabs, users can observe the differences in the scattering patterns.

Student
view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

In the “Rutherford Atom mode,” alpha particles are directed toward a gold foil-like structure, representing the experiment. Some particles pass through unimpeded, while others deflect at sharp angles, and a few even bounce straight back. The deflection occurs due to electrostatic repulsion from a small, dense, positively charged nucleus. The visualization demonstrates how most of an atom is empty space, with the nucleus accounting for most of the mass.

In the plum pudding mode, most particles pass straight through the foil with minimal deviation. Alpha particles are shown moving through an atom with evenly spread positive charge and embedded electrons. The simulation shows that all alpha particles pass through with slight deflections, supporting the idea that charge is distributed throughout the atom. This visualization reflects Thomson’s assumption that an atom is a soft, diffuse structure with no central concentration of charge.

A key aspect of this interactive experience is the "Nuclear View" versus "Atomic Model View," where users can visualize the differences between the two models of the atom. The "Nuclear View" shows the small, dense nucleus surrounded by electron orbits, while the "Atomic Model View" depicts energy levels and the distribution of electrons in the atom.

From this simulation, users can understand why the Plum Pudding Model failed to explain experimental results, how scientific models evolve based on experimental evidence, gain insight into atomic structure (specifically the nucleus and empty space), and recognize the importance of Rutherford’s experiment in shaping modern atomic theory.

This interactive simulation helps visualize these critical concepts, making it easier to grasp why the nuclear model replaced the Plum Pudding Model.

For the Rutherford atom, most of the alpha particles are not deflected. Some alpha particles are scattered at small angles and a very few are deflected at large angles. For the ‘plum pudding’ atom, the alpha particles are not noticeably scattered.

(https://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_all.html)These observations led to a new model of the atom – the Rutherford model. This model states that most of the atom is empty space. Almost all of the positively charged matter in the atom is concentrated in a very small region at its centre, called the nucleus. Electrons are in orbit around the nucleus.

Because most of the atom is empty space, most of the alpha particles do not deviate from their paths. The alpha particles that deviate at greater angles are the ones that get close enough to the nucleus to be deflected by the repulsive electric force. Since the nucleus is very small compared to the whole atom, such deflection events are rare.

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Figure 3 shows the expected paths of the alpha particles for the Thomson model and the Rutherford model and the observed results for the Rutherford model.

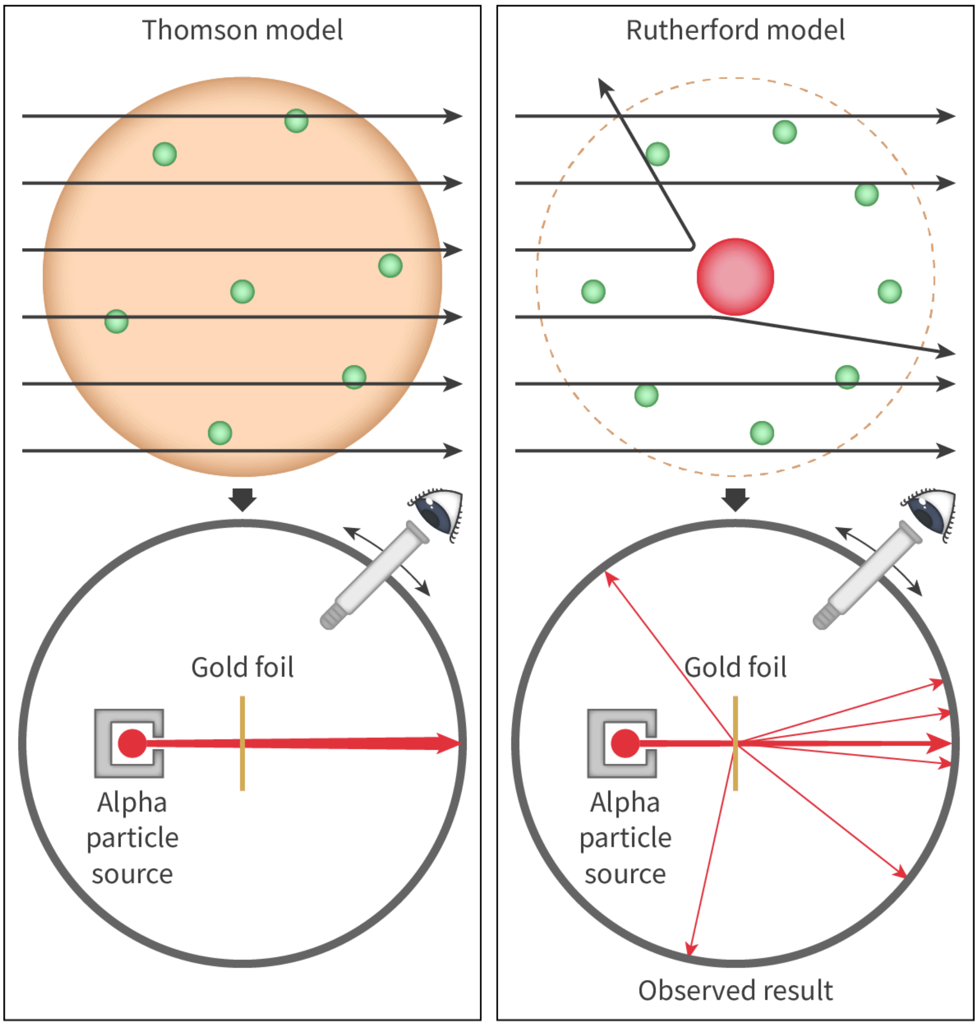


Figure 3. Comparing the paths of alpha particles for the Thomson model and the Rutherford model.

More information for figure 3

The image is a comparison diagram illustrating the paths of alpha particles in two atomic models: the Thomson model and the Rutherford model.

On the left side, the Thomson model consists of a large sphere containing evenly distributed small green circles representing positive charge. Black arrows pass straight through the sphere, indicating alpha particle paths.

Below the sphere is a secondary diagram showing an alpha particle source directed at a gold foil. A detector around the setup is positioned to observe the particles. According to the Thomson model, alpha particles pass straight through, highlighting how they interact uniformly with the distributed charge.

On the right side, the Rutherford model depicts a central red circle representing a dense nucleus, surrounded by smaller green circles within a dashed boundary. Black arrows show that alpha particles deflect at various angles, reflecting occasional direct hits with the nucleus.

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

Similarly, the diagram below the sphere shows the alpha particle source, gold foil, and detector. However, in the Rutherford model, paths of deflected particles vary, demonstrating how some particles reflect sharply when interacting with the dense nucleus.

The image contrasts the two models by showing different particle behaviors, supporting the observed results of the Rutherford experiment.

[Generated by AI]

Nature of Science

Aspect: Models

In science, a 'model' is a way of representing an aspect of the real world, such as atoms or gravitational forces. A model is not reality; it is a tool for making explanations and predictions. If new observations contradict an established model, then scientists will attempt to create a new model that fits the new observations as well as the old.

Different models for the same phenomenon may be used in different situations if that is helpful. For example, although the Rutherford model does not fit all of our observations and is not consistent with quantum theory, it can still be used successfully to predict some of the behaviours of atoms.

International Mindedness

Rutherford's ideas for the structure of the atom, along with the modifications introduced by Niels Bohr, shaped a model of the atom called the planetary model. In this model, the nucleus (Sun) is at the centre and the electrons (planets) move around it in different orbits (**Figure 4**).

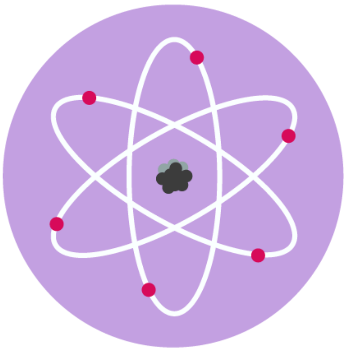


Figure 4. The planetary model of the atom.

More information for figure 4

Student
view

This is an illustration of the planetary model of the atom, with a central dark gray nucleus representing the atom's core. Surrounding the nucleus are several white curved lines depicting the orbital paths of electrons. At various points along these orbits, small red circles represent the electrons themselves. The background is a solid lavender color, providing contrast to the white orbits and red electrons, which highlights the atomic structure.

[Generated by AI]

Although this model has been superseded by more modern models, such as the electron cloud model, the image in **Figure 4** is often still used to represent an atom on banknotes, logos, maps, etc. It is a visualisation of a physics concept that people all over the world can understand. Can you think of any other images that communicate a physics idea in a universal way?

Nuclear notation

Today, we know that the nucleus of an atom is made up of protons and neutrons, while electrons orbit the nucleus. **Table 1** shows information about these particles.

Table 1. Position, mass and charge of protons, neutrons and electrons.

Particle	Where it is?	Mass (kg)	Charge (C)
Proton	In the nucleus	1.673×10^{-27}	1.60×10^{-19}
Neutron	In the nucleus	1.675×10^{-27}	0
Electron	In orbit around the nucleus	9.110×10^{-31}	-1.60×10^{-19}

The number of protons in the nucleus determines the chemical element of an atom. This is called the proton number. The number of protons and neutrons in the nucleus is known as the nucleon number.

We can describe an atom of an element using nuclear notation:

$$^{A}_{Z}\mathrm{X}$$

where:

- X is the chemical symbol of the element

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

- A is the nucleon number
- Z is the proton number.

For example, the nuclear notation for carbon-12, which has 6 protons and 6 neutrons in its nucleus, is:

$$^{12}_{6}\mathrm{C}$$

Exercise 1

Click a question to answer

Atomic energy levels

Imagine that you lift an object above the ground. When you lift the object a certain height, Δh, the object gains gravitational potential energy (see subtopic A.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43083/)):

$$\Delta E = mg\Delta h$$

You can hold the object at any height you want, giving the object any amount of gravitational potential energy.

In an atom, electrons orbit the nucleus. But electrons cannot be any distance from the nucleus. They can only orbit at certain discrete distances from the nucleus.

Nature of Science

Aspect: Measurement

Some quantities in nature can take any value within a range. An example is pouring water into a glass. You can pour in tiny drops or you can fill the glass up. These quantities are called **continuous**.

Other quantities cannot take any value within a range. They can only take certain values. For example, imagine that you put marbles in the glass. You can only put in one, two, three marbles, etc. These quantities are called **discrete**. The values are (integer)

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

multiples of a base value, which is the smallest value the quantity can take.

Can you think of other physical quantities that are continuous and other quantities that are discrete?

Electrons can only orbit the nucleus at discrete distances. This means that atoms have a set of discrete atomic energy levels that correspond to the allowed orbits of the electrons. **Figure 5** shows some of the atomic energy levels for hydrogen, which has one proton in the nucleus. We assign an integer value *n* to every level.

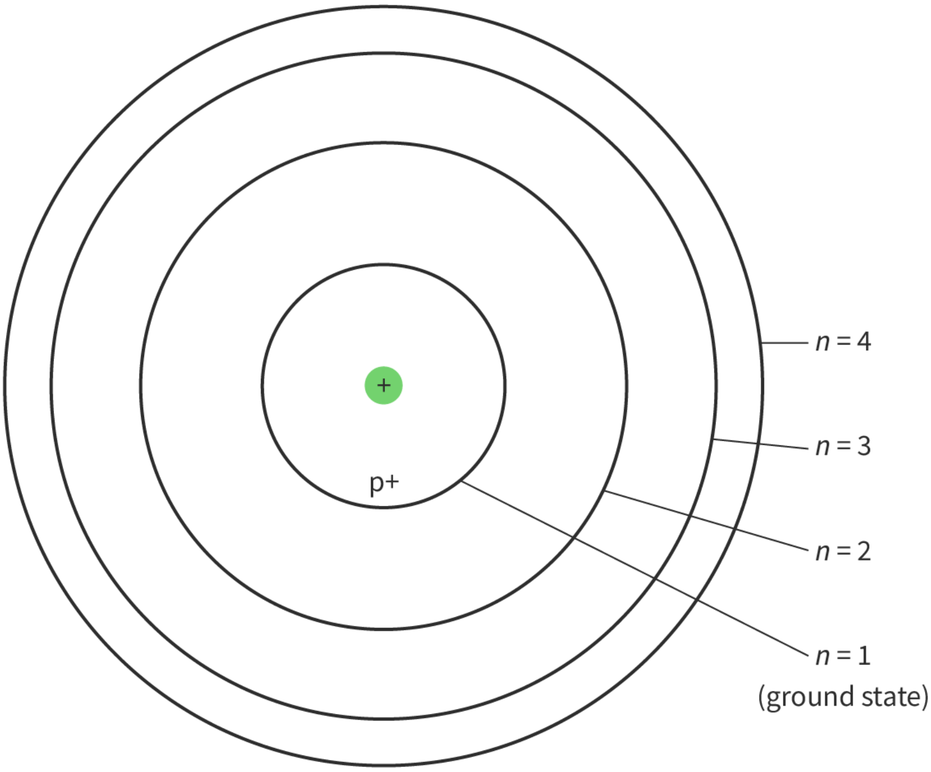


Figure 5. Some of the atomic energy levels for hydrogen.

More information for figure 5

The diagram shows the energy levels of a hydrogen atom. At the center is a nucleus marked with a green circle containing a plus sign labeled 'p+', representing a proton. Surrounding the nucleus are four concentric circles representing the orbitals where an electron can be present. Each orbital is labeled with an integer 'n' which represents the principal quantum number of the energy level. The closest circle to the nucleus is labeled 'n = 1 (ground state)', indicating the lowest energy level. The subsequent layers are labeled 'n = 2', 'n = 3', and 'n = 4'. These labels are written in line with the concentric circles from the smallest to the outermost orbit, representing increasing energy levels.

[Generated by AI]

Student
view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

Figure 6 shows an energy level diagram for hydrogen.

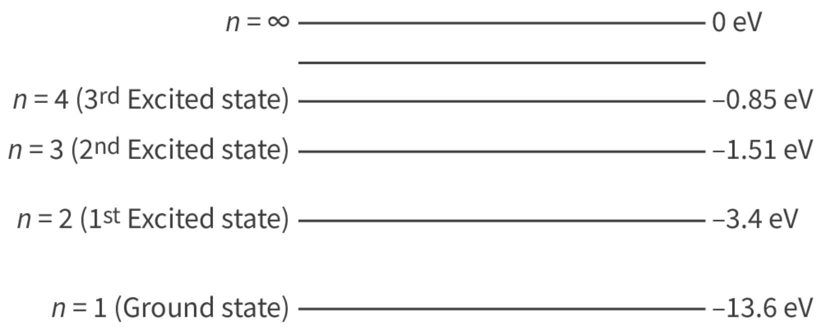


Figure 6. Energy level diagram for hydrogen.

More information for figure 6

The image is an energy level diagram for a hydrogen atom. It shows horizontal lines representing different energy states, labeled with their corresponding principal quantum numbers and energy values in electron volts (eV). The lines are as follows:

1. n = 1, Ground state at -13.6 eV
2. n = 2, 1st Excited state at -3.4 eV
3. n = 3, 2nd Excited state at -1.51 eV
4. n = 4, 3rd Excited state at -0.85 eV
5. n = ∞ at 0 eV, representing the ionized state.

These lines are parallel and increase in energy from bottom to top, indicating that they represent quantized energy levels that a hydrogen atom can occupy, with energy increasing as 'n' increases.

[Generated by AI]

The parallel lines in **Figure 6** show the energy levels of the atom. The first line ($n = 1$) corresponds to the lowest energy that an atom can have, called the ground state. The rest of the lines correspond to energies higher than the ground state, called excited states. The second line ($n = 2$) corresponds to the first excited state, and the third line ($n = 3$) corresponds to the second excited state, etc.

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

Study skills

The energy of an atom is usually expressed in electronvolts (eV), rather than joules (J). An electronvolt is defined as the (kinetic) energy that an electron gains when it is accelerated across a potential difference of 1 volt:

$$1 \text{ eV} = 1.6 \times 10^{-19} \text{ J}$$

Atomic transitions

An atom changes its energy level when electrons change their distance from the nucleus. This change in distance happens in a discrete way – the electrons 'jump' from one orbit to another.

This is called an atomic transition. Energy levels closer to the nucleus have lower energies than energy levels further from the nucleus.

The transition from a lower energy level to a higher energy level is called excitation. The transition from a higher energy level to a lower energy level is called de-excitation.

- When an atom gains energy, an electron may transition to a higher energy level.
- When an electron loses energy, it transitions to a lower energy level.

We can see from **Figure 6** that the ground state ($n = 1$) of an atom has the lowest energy (the most negative energy). This is because the negatively charged electrons are attracted to the positively charged nucleus so the energy is negative (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)).

The highest energy level ($n = \infty$) has zero energy. When an atom gains enough energy that an electron leaves the atom, the electron is no longer attracted to the nucleus so the highest energy level has zero energy.

In energy level diagrams, as the energy increases, it becomes less negative, with zero as the maximum.

Worked example 1

The diagram (not to scale) shows some of the energy levels of an atom. The arrows show two atomic transitions, from $n = 1$ to $n = 2$, and from $n = 3$ to $n = 1$.

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

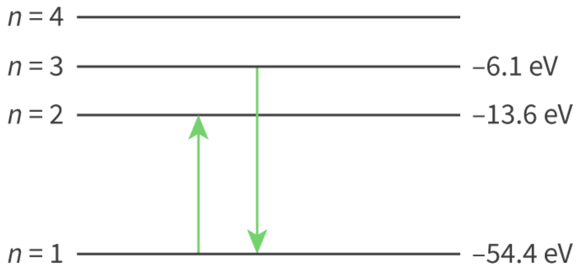


Figure 7. Energy levels of an atom.

More information for figure 7

The diagram illustrates the energy levels of an atom with four distinct levels. The top level, labeled "n = 4," has an associated energy of approximately -6.1 electron volts (eV). Below it is "n = 3," with an energy level of -13.6 eV, followed by "n = 2," and finally "n = 1" at the bottom, each separated by horizontal lines. Two vertical arrows indicate atomic transitions: one arrow goes upward from "n = 1" to "n = 2," and the other goes downward from "n = 3" to "n = 1." The energy levels suggest changes in potential energy as electrons move between these states.

[Generated by AI]

1. In which atomic transition does the electron gain energy and in which atomic transition does the electron lose energy?
2. How much energy is gained or lost in each atomic transition?
3. How many possible atomic transitions exist between the first three energy levels?

1. Deduce whether the electron gains or loses energy by comparing the energies of the energy levels involved in the transition.

   - $n = 1$ has a lower energy than $n = 2$, so the transition from $n = 1$ to $n = 2$ is an excitation, and the electron gains energy.
   - $n = 3$ has a higher energy than $n = 1$, so the transition from $n = 3$ to $n = 1$ is a de-excitation and the electron loses energy.

2. From $n = 1$ to $n = 2$:

$$(-54.4) - (-13.6) = -40.8 \text{ eV}$$

40.8 eV is gained.

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

From $n = 3$ to $n = 1$:

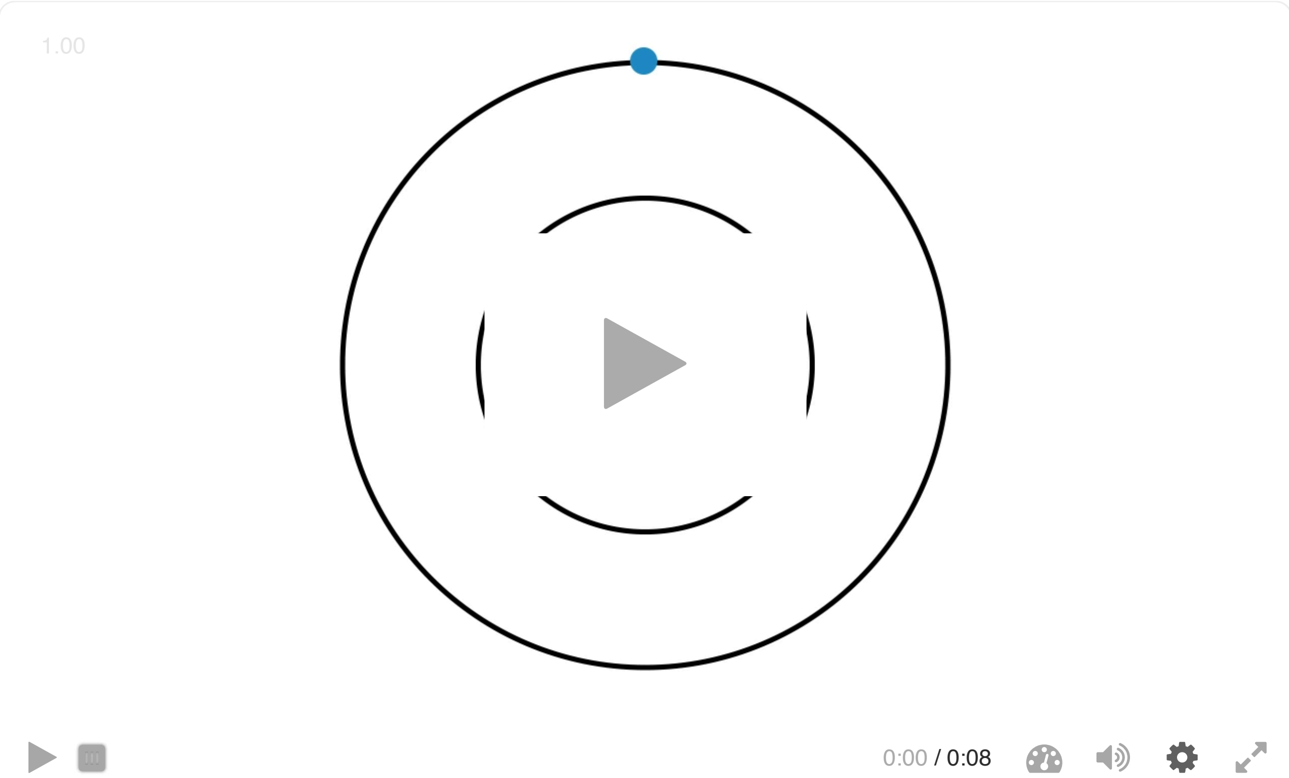
$$(-6.1) - (-54.4) = 48.3 \text{ eV}$$

48.3 eV is lost.

3. possible excitations are: $n = 1$ to $n = 2$, $n = 1$ to $n = 3$, and $n = 2$ to $n = 3$.
   possible de-excitations are: $n = 3$ to $n = 2$, $n = 3$ to $n = 1$, and $n = 2$ to $n = 1$.
   total number of possible transitions = 6

Since atomic energy levels have different energies, when electrons make the transition from one energy level to another, energy is either emitted or absorbed.

Interactive 2 shows a hydrogen atom. When the atom absorbs energy, the electron makes a transition to a higher energy level. When the energy is emitted, the electron makes a transition to a lower energy level.



Interactive 2. Two Atomic Transitions in a Hydrogen Atom.

More information for interactive 2

The animation video shows a simplified model of a hydrogen atom with a nucleus at the center and an electron orbiting around it. The nucleus is represented by a red-colored circle, while the electron is represented by a blue circle and it is smaller in size compared to the nucleus. There are two orbits (energy levels) around the nucleus.

The electron moves between the two energy levels, jumping to a higher orbit when it absorbs energy and dropping to a

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

lower orbit when it emits energy. The energy absorption is represented by a wavy arrow moving towards the orbit and the energy emission is represented by a wavy arrow moving away from the orbit. Each transition corresponds to the absorption or emission of a photon. The visualization highlights the concept that energy levels are discrete, meaning that only specific amounts of energy can be absorbed or emitted. This demonstrates the fundamental principle of atomic transitions and how photons are linked to changes in an electron's energy state.

Because the atomic energy levels are discrete, then the energy emitted or absorbed is also discrete. These 'packets' of electromagnetic energy are called photons.

When an atomic transition takes place, the photon that is emitted or absorbed has an energy equal to the difference in energy levels. In other words, every transition is associated with a photon with a certain energy, and a certain frequency (and wavelength).

Theory of Knowledge

If the natural sciences are trying to observe and describe physical reality, how can it be that scientific knowledge changes over time? Does reality change or our understanding of it?

Millikan's oil-drop experiment (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)) proved the existence of the electron and led to the Thomson model of the atom. Rutherford's gold foil experiment established a new model of the atom, where a nucleus exists and electrons orbit the nucleus.

But there were problems with the Rutherford model. According to electromagnetic theory, the electrons were accelerating around the nucleus, so they should be emitting radiation and losing energy, reducing their radius of motion until they fell into the nucleus and destroyed the atom. Niels Bohr introduced the idea that the electrons could only exist in certain orbits. Further discoveries and theories by Heisenberg and Schroedinger shaped new understandings of the structure of the atom, establishing that we cannot be certain about the position and orbit of an electron; we can only calculate probabilities about it.

These changes in thinking about the structure of the atom are known as **paradigm shifts**. How do changing paradigms in the natural sciences affect the validity of the knowledge claims being made?

The equation for the energy of a photon is shown in **Table 2**.

Student view

Table 2. Equation for the energy of a photon.

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

Equation	Symbols	Units
$E = hf$	E = energy of photon	joules (J)
	h = Planck constant (6.63×10^{-34} J Hz^{-1})	Given in section 1.6.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/fundamental-constants-id-45155/)of the DP physics data booklet
	f = frequency of photon	hertz (Hz)

The wave equation is (see subtopic C.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43778/)):

$$v = f\lambda$$

The speed of an electromagnetic wave is c (3.00×10^8 m s^{-1}):

$$c = f\lambda$$

$$f = \frac{c}{\lambda}$$

Substituting this into the equation for the energy of a photon gives:

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

Worked example 2

A laser pointer produces red light with a wavelength of 650 nm. What is the energy of each photon produced by the laser pointer?

$$\begin{aligned}\lambda &= 650 \text{ nm} \\ &= 650 \times 10^{-9} \text{ m}\end{aligned}$$

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

$$
\begin{aligned}
E &= \frac{hc}{\lambda} \\
&= \frac{(6.63\times10^{-34}\times3.00\times10^{8})}{650\times10^{-9}} \\
&= 3.06\times10^{-19}\ \text{J} \\
&= 3.1\times10^{-19}\ \text{J (2 s.f.)}
\end{aligned}
$$

Emission and absorption spectra

Light from the Sun is made up of lots of different colours. We can see these colours if we pass the light through a prism (**Figure 8**).

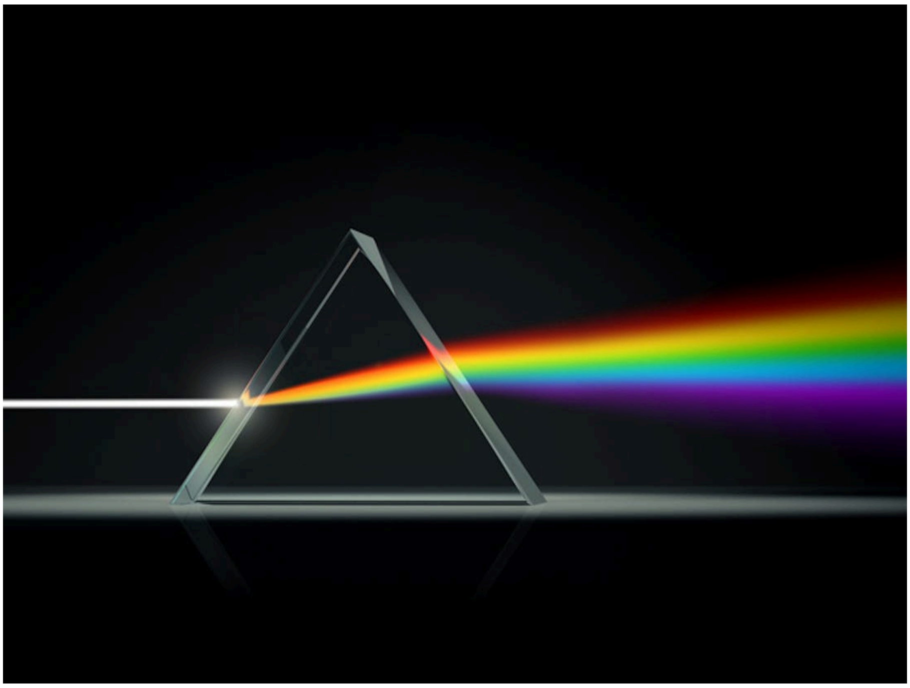


Figure 8. Passing light through a prism.

Credit: artpartner-images, Getty Images

This produces a rainbow-like distribution of light called a continuous spectrum. Each colour corresponds to a different band of wavelengths (and frequencies).

If hydrogen gas is excited (by passing a current through it or by heating it to a high temperature), it will emit photons. The hydrogen atoms only emit photons with specific energies (wavelengths). What would you expect to see in the emission spectrum for hydrogen? Click on 'Show or hide solution' to see the answer.

Student view

You would not see a continuous spectrum of light, consisting of all colours. You would see only specific wavelengths, and thus colours, emitted.

Figure 9 shows the emission spectrum of hydrogen. It consists of bright emission lines on a dark background.

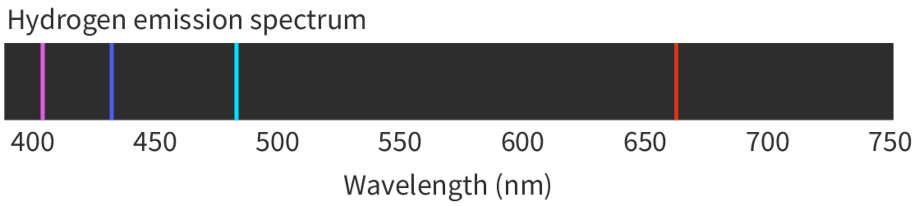


Figure 9. The emission spectrum for hydrogen.

More information for figure 9

The image represents the emission spectrum of hydrogen, displaying bright emission lines on a dark background. The spectrum is shown as a horizontal bar with a wavelength scale ranging from 400 nm to 750 nm. Bright lines appear at specific wavelengths: around 410 nm, 434 nm, 486 nm, and 656 nm, corresponding to the visible emission lines of hydrogen. The x-axis is labeled "Wavelength (nm)," showing the range numerically from 400 to 750 nm.

[Generated by AI]

If we shine white light on hydrogen gas, the hydrogen atoms only absorb the photons with energies that correspond to the possible transitions between energy levels. Therefore the spectrum of the light transmitted through the hydrogen gas will show dark lines on a bright background (**Figure 10**). This is called an absorption spectrum.

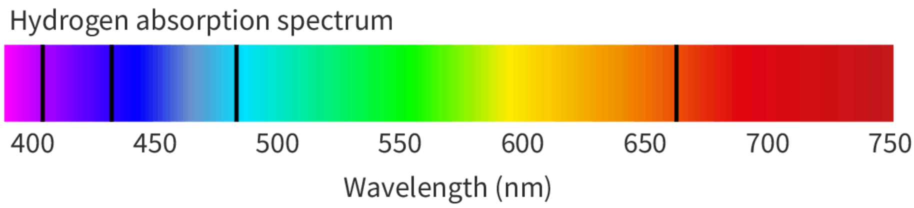


Figure 10. The absorption spectrum of hydrogen.

More information for figure 10

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

The image is a representation of the hydrogen absorption spectrum, displayed as a continuous range of colors from 400 nanometers (nm) to 750 nm, transitioning through purple, blue, cyan, green, yellow, and red. There are distinct dark vertical lines appearing at specific wavelengths where hydrogen absorbs light, interrupting the continuous spectrum. The labels below the spectrum show the corresponding wavelengths: 400, 450, 500, 550, 600, 650, 700, and 750 nm. This visual indicates the specific wavelengths absorbed by hydrogen atoms, correlating with possible transitions between energy levels as light passes through hydrogen gas.

[Generated by AI]

Figure 11 shows a light bulb and its continuous emission spectrum. When the light passes through the cloud of hydrogen gas, the hydrogen atoms get excited and absorb some of the photons, producing the absorption spectrum. The hydrogen atoms then undergo de-excitation and photons are emitted, producing the emission spectrum. Emission and absorption spectra are **complementary** to each other and are evidence for the discrete nature of atomic energy levels.

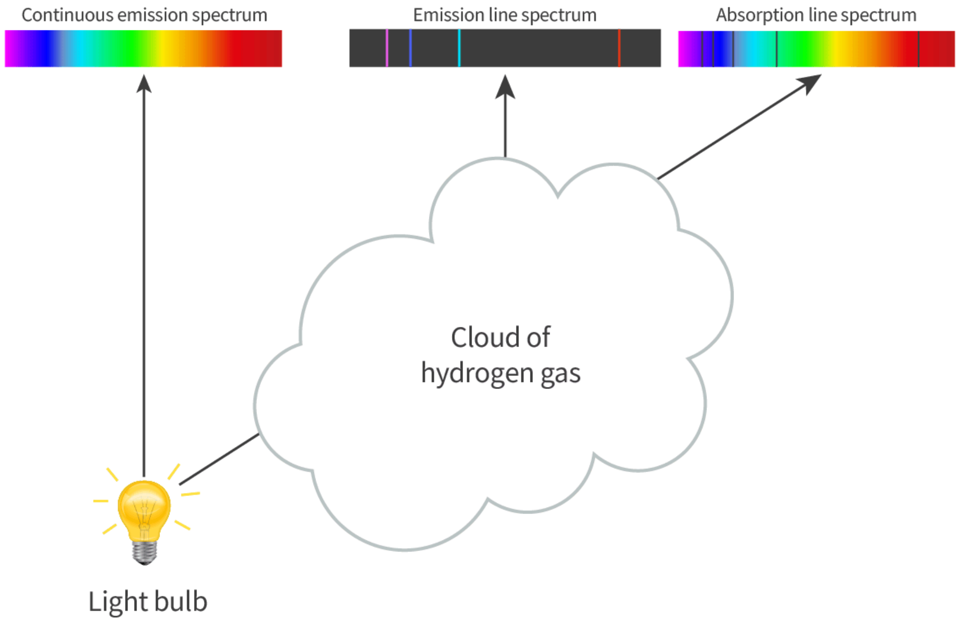


Figure 11. Continuous spectrum, emission spectrum and absorption spectrum.

More information for figure 11

This is a diagram illustrating the concepts of continuous, emission, and absorption spectra. At the bottom left, there is an image of a lit light bulb, symbolizing a source of continuous emission spectrum. This spectrum is depicted at the top left as a continuous band of colors ranging from violet to red. The center of the image shows a cloud labeled 'cloud of hydrogen gas.' An arrow extends from the light bulb to the cloud, indicating that the light passes through this gas. To the top right of the cloud, an 'absorption line spectrum' is shown, consisting of a band of colors similar to the continuous spectrum but

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

with dark lines indicating absorbed wavelengths. Above the cloud, an 'emission line spectrum' displays narrow colored lines on a dark background, representing the specific wavelengths emitted by excited hydrogen atoms. Arrows connect the cloud to both the absorption and emission spectra, illustrating the process of absorption and re-emission of photons by hydrogen gas.

[Generated by AI]

Atoms of different elements have different sets of energy levels, so they also have different emission and absorption spectra. **Figure 12** shows the emission spectra for three different elements: hydrogen, helium and mercury.

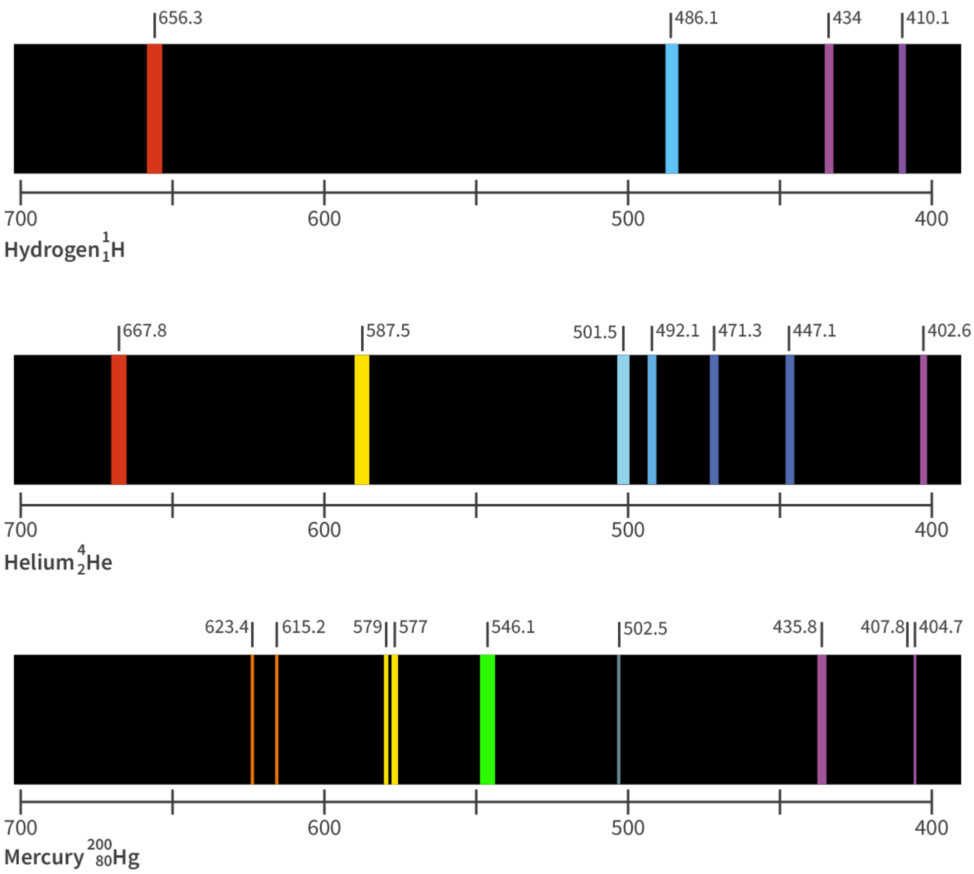


Figure 12. The emission spectra for hydrogen, helium and mercury.

More information for figure 12

The image depicts the emission spectra of hydrogen, helium, and mercury, each represented by different colored lines at specific wavelengths on a black background.

1. **Hydrogen Spectrum:**
2. A single red line at 656.3 nm.
3. A blue line at 486.1 nm.

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

4. Violet lines appear at 434 nm and 410.1 nm.
5. The X-axis ranges from 400 to 700 nm, with labels for every 100 nm.
6. **Helium Spectrum:**
7. A strong red line at 667.8 nm.
8. Yellow line at 587.5 nm.
9. Multiple blue lines at 501.5 nm, 492.1 nm, 471.3 nm, and 447.1 nm.
10. A purple line at 402.6 nm.
11. The X-axis is similarly labeled from 400 to 700 nm.
12. **Mercury Spectrum:**
13. Strong orange lines at 623.4 nm and 615.2 nm.
14. Multiple yellow lines at 579 nm and 577 nm.
15. A bright green line at 546.1 nm.
16. Blue and violet lines appear at 502.5 nm, 435.8 nm, 407.8 nm, and 404.7 nm.
17. The X-axis has the same range as the others.

Each spectrum is labeled with the respective element name and atomic symbol next to it. The overall image is organized with clear horizontal arrangements for easy comparison of elemental emission lines.

[Generated by AI]

We can identify an element by looking at its emission or absorption spectra. This is how scientists can determine the chemical composition of a star. They observe the light emitted or absorbed by the star and compare its spectrum to the known spectra of atoms produced in the laboratory.

Nature of Science

Aspect: Observations

In 1868, during a solar eclipse, astronomers observed an emission line originating from the Sun, with a wavelength that corresponded to a yellow colour. This emission line was evidence for the existence of an unknown element. Scientists called it helium, from the Greek ‘helios’, which means Sun. Helium was discovered on Earth, almost 30 years later.

What other examples can you think of where an observation is made before a theory is developed?

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Worked example 3

The diagram shows some of the energy levels of hydrogen and its emission spectrum. Identify lines 1 to 4 by matching them to atomic transitions A to D.

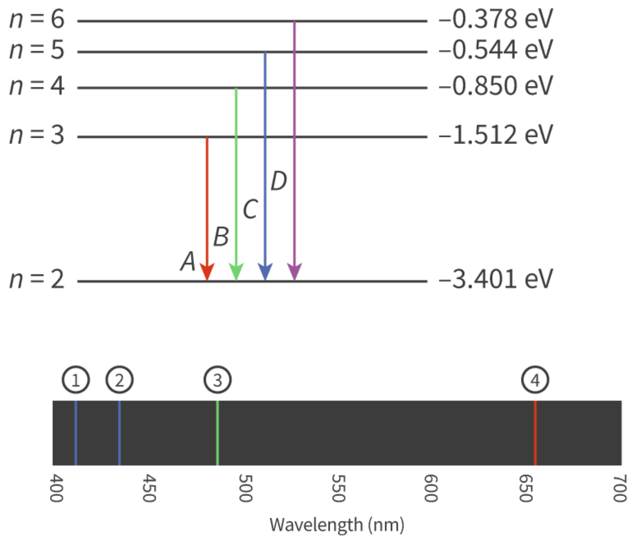


Figure 13. Energy level diagram for hydrogen.

More information for figure 13

The image is an energy level diagram for hydrogen, illustrating the transition of electrons between different energy levels. On the left side, energy levels from n=2 to n=6 are depicted with their corresponding energy values: -3.401 eV for n=2, -1.512 eV for n=3, -0.850 eV for n=4, -0.544 eV for n=5, and -0.378 eV for n=6.

Four transitions are shown with arrows labeled A, B, C, and D. Transition A is from n=3 to n=2, transition B is from n=4 to n=2, transition C is from n=5 to n=2, and transition D is from n=6 to n=2. Each arrow corresponds to a line in a spectrum box below the energy levels, with lines labeled 1 to 4.

The emission spectrum is shown at the bottom, with numbers 1 to 4 marked on specific positions along the wavelength axis, ranging from 400 nm to 700 nm. Lines corresponding to each transition are highlighted in matching colors on the wavelength spectrum: line 1 is around 450 nm, line 2 around 485 nm, line 3 near 500 nm, and line 4 at about 656 nm, matching the transitions A, B, C, and D respectively.

[Generated by AI]

Student view

Energy carried by photons for each atomic transition:

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

A: $-1.512 - (-3.401) = 1.889$ eV $= 3.0224 \times 10^{-19}$ J
B: $-0.850 - (-3.401) = 2.551$ eV $= 4.0816 \times 10^{-19}$ J
C: $-0.544 - (-3.401) = 2.857$ eV $= 4.5712 \times 10^{-19}$ J
D: $-0.378 - (-3.401) = 3.023$ eV $= 4.8368 \times 10^{-19}$ J

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

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

Wavelength of photons:

A: $\frac{(6.63\times10^{-34}\times3.00\times10^{8})}{3.0224\times10^{-19}} = 6.581\times10^{-7}\text{ m} = 658\text{ nm (3 s.f.)}$

B: $\frac{(6.63\times10^{-34}\times3.00\times10^{8})}{4.0816\times10^{-19}} = 4.873\times10^{-7}\text{ m} = 487\text{ nm (3 s.f.)}$

C: $\frac{(6.63\times10^{-34}\times3.00\times10^{8})}{4.5712\times10^{-19}} = 4.351\times10^{-7}\text{ m} = 435\text{ nm (3 s.f.)}$

D: $\frac{(6.63\times10^{-34}\times3.00\times10^{8})}{4.8368\times10^{-19}} = 4.112\times10^{-7}\text{ m} = 411\text{ nm (3 s.f.)}$

Transition A corresponds to line 4.
Transition B corresponds to line 3.
Transition C corresponds to line 2.
Transition D corresponds to line 1.

Work through the activity to check your understanding of atomic transitions.

Activity

- **IB learner profile attribute:** Knowledgeable
- **Approaches to learning:** Thinking skills — Applying key ideas and facts in new contexts
- **Time required to complete activity:** 25 minutes
- **Activity type:** Individual activity

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Look at this hydrogen atom simulation (https://foothill.edu/astronomy/astrosims/hydrogen-atom/). The long figure at the top shows a slice through a hydrogen atom, with the nucleus at the left. The atom is initially in the ground state, *n* = 1.

The simulation allows you to fire photons at the atom. You can drag the slider along the bar (near the bottom of the page) to choose a photon's characteristics. If you move the bar to the right, the photon's energy (shown in eV) increases, its frequency (in Hz) increases and its wavelength (in nm) decreases. (These three quantities are all related, so you cannot choose their values separately.)

Try firing a photon, by pressing the ‘Fire photon’ button at the bottom of the page. Try different photons — you will find that most photons pass straight through the atom.

If the energy of the photon corresponds to a possible atomic transition, then the atom will get excited, and the electron will make a transition to a higher energy level. The electron will then go back to the ground state, emitting the photon again. Your goal is to fire photons that will be absorbed by the atom and excite it from its ground state to *n* = 2, *n* = 3, *n* = 4, *n* = 5 and *n* = 6.

Figure 14 shows the energy level diagram for hydrogen for the first six energy levels.

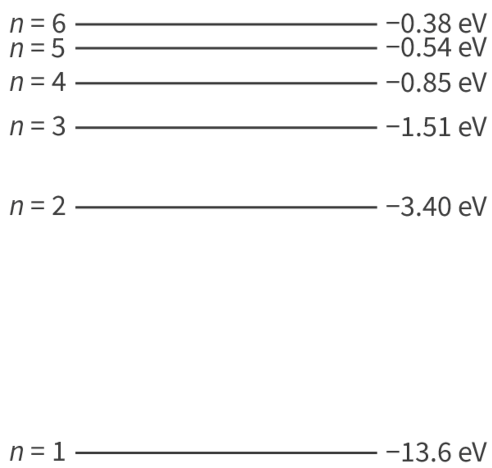


Figure 14. Energy level diagram for hydrogen showing the first six energy levels.

More information for figure 14

The image is an energy level diagram for hydrogen showcasing the first six energy levels. It displays horizontal lines, each representing an energy level of the hydrogen atom, labeled with the principal quantum number 'n'. The energy levels are labeled as follows: n = 6 at -0.38 eV, n = 5 at -0.54 eV, n = 4 at -0.85 eV, n = 3 at -1.51 eV, n = 2 at -3.40 eV, and n = 1 at -13.6 eV. The lines are displayed from top to bottom in descending order of energy values, with n=6 having the highest (least negative) energy and n=1 having the lowest (most negative) energy.

[Generated by AI]

Student view

Calculate the energy of the photon needed for each transition. Move the slider to select the energy (and frequency and wavelength) of the photon. Click ‘Fire Photon’. Is the photon absorbed by the hydrogen atom? Does it cause excitation?

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

What happens during de-excitation? Does it happen in one step or not? Calculate the energy of a photon emitted during de-excitation. Does your calculation match the event log?

Energies of photons that could be absorbed: 10.2 eV, 12.09 eV, 12.75 eV, 13.06 eV and 13.22 eV

Extension

An atom is said to be 'ionised' when an electron leaves the atom entirely. Try to work out how to ionise the atom in the simulation.

5 section questions

Question 1

SL HL Difficulty:

What did the Geiger—Marsden—Rutherford experiment show about the structure of the atom?

1. Most of the matter in an atom is concentrated in the centre ✓
2. Electrons orbit the nucleus at discrete distances
3. Positive charges are spread throughout the atom
4. The nucleus contains neutrons

Explanation

The Geiger—Marsden—Rutherford experiment showed that most of the mass of the atom is concentrated in the centre, in the nucleus.

Question 2

SL HL Difficulty:

A photon is a discrete 'packet' of 1 energy ✓ .

When an atom absorbs a photon, an electron may make a transition from a 2 lower ✓ energy level to a 3 higher ✓ energy level.

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

When an electron makes a transition from a 4 higher ✓ energy level to a 5 lower ✓ energy level, a photon is emitted.

Accepted answers and explanation

#1 energy

#2 lower

#3 higher

#4 higher

#5 lower

General explanation

A photon is a discrete ‘packet’ of energy.

When an atom absorbs a photon, an electron may make a transition from a lower energy level to a higher energy level.

When an electron makes a transition from a higher energy level to a lower energy level, a photon is emitted.

Question 3

SL HL Difficulty:

An atom absorbs a photon and an electron makes a transition between the n = 2 (−2.4 eV) and n = 3 (−1.5 eV) energy levels. What is the wavelength of the photon?

1 1.4 µm ✓

2 0.9 µm

3 0.9 nm

4 1.4 nm

Explanation

The energy of the photon corresponds to the difference between the energy levels:

$$\begin{aligned}(-2.4)–(-1.5) &= -0.9 \text{ eV} \\ &= -1.44\times10^{-19} \text{ J}\end{aligned}$$

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

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

$$
\begin{aligned}
\lambda &= \frac{hc}{E} \\
&= \frac{(6.63\times10^{-34}\times3.00\times10^{8})}{1.44\times10^{-19}} \\
&= 1.38\times10^{-6}\ \text{m} \\
&= 1.4\ \mu\text{m (2 s.f.)}
\end{aligned}
$$

Question 4

SL HL Difficulty:

The energy of a photon is 4.3 eV. What is its frequency?

1 1.0×10^{15} Hz ✓

2 6.9×10^{33} Hz

3 2.9×10^{-33} Hz

4 4.6×10^{-52} Hz

Explanation

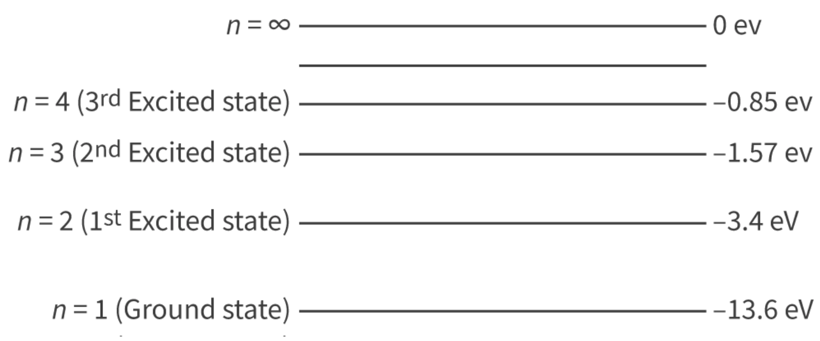
$$
\begin{aligned}
4.3\ \text{eV} &= 4.3\times1.6\times10^{-19} \\
&= 6.88\times10^{-19}\ \text{J}
\end{aligned}
$$

$$
\begin{aligned}
E &= hf \\
f &= \frac{E}{h} \\
&= \frac{6.88\times10^{-19}}{6.63\times10^{-34}} \\
&= 1.0377\times10^{15}\ \text{Hz} \\
&= 1.0\times10^{15}\ \text{Hz (2 s.f.)}
\end{aligned}
$$

Question 5

SL HL Difficulty:

The diagram shows the energy level diagram for hydrogen.



More information

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

When white light is shone through cold hydrogen gas, at which of the following frequencies will there be a dark line on the absorption spectrum?

1 3.3×10^{15} Hz ✓

2 2.4×10^{15} Hz

3 1.5×10^{15} Hz

4 2.1×10^{15} Hz

Explanation

$$E = hf$$

The largest amount of energy that can be absorbed by a single electron transition is from —13.6 eV to 0 eV.

Remembering to use energy in joules:

$$f = \frac{E}{h} = \frac{13.6 \times 1.6 \times 10^{-19}}{6.63 \times 10^{-34}} = 3.3 \times 10^{15} Hz$$

E. Nuclear and quantum physics / E.1 Structure of the atom

Deviations from Rutherford scattering and the Bohr model for hydrogen (HL)

E.1.7: Nuclear densities (HL) E.1.8: Deviations from Rutherford scattering (HL) E.1.9: Closest approach in head-on scattering experiments (HL)
E.1.10: Discrete energy levels in the Bohr model for hydrogen (HL) E.1.11: Quantization of angular momentum in the Bohr model (HL)
Section Student... (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/deviations-from-rutherford-scattering-hl-id-46599/print/) Assign

Higher level (HL)

Learning outcomes

By the end of this section you should be able to:

- Understand why deviations from Rutherford scattering occur and the concept of distance of closest approach.
- Understand the relationship between nucleon number and radius of a nucleus and use the equation:

Student view

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

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

- Understand that quantisation of angular momentum can explain discrete energy levels in the Bohr model for hydrogen and use the equations:

$$mvr = \frac{nh}{2\pi} \text{ and } E = -\frac{13.6}{n^2}\ \mathrm{eV}$$

Imagine you had a scientific theory about objects interacting. You throw balls through an open doorway and your theory says that they should travel straight through. You test it by throwing 1000 balls through a doorway. Most of the balls go through the door as expected, but a few of them bounce right back at you from thin air! What do you do?

You change your theory!

Deviations from Rutherford scattering

In Rutherford scattering (see section E.1.1 (/study/app/math-aa-hl/sid-423-cid-762593/book/atoms-and-photons-id-46593/)), the number of alpha particles detected after scattering depends on the angle at which they are deflected from their straight path (**Figure 1**). We call this angle the scattering angle.

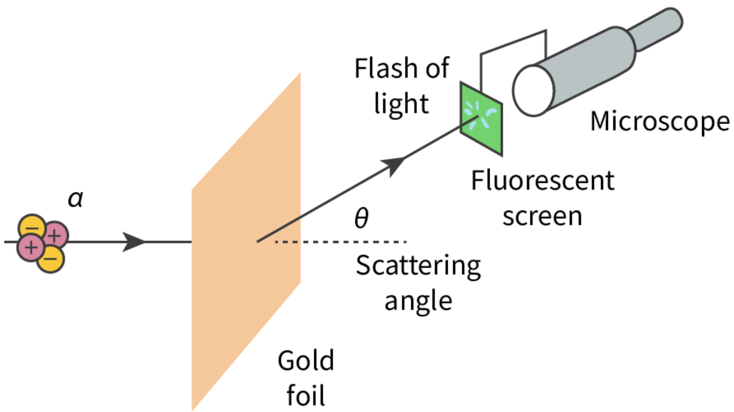


Figure 1. Detecting scattered alpha particles.

More information for figure 1

The diagram illustrates the process of alpha particles scattering off a gold foil, which is labeled in the image. The alpha particles, denoted by the symbol α, are shown as small clusters of circles with positive signs, indicating their positive charge. These particles travel towards a gold foil, represented by an orange rectangle. As the particles pass through or near the gold foil, they scatter at different angles, labeled as θ (theta), which represents the scattering angle. This angle is marked as a dashed line between the path of the particles and the perpendicular line to the gold foil.

Beyond the gold foil, there is a flashed light indicating the detection of the scattered particles on a fluorescent screen, which is represented as a green rectangle. The screen is positioned to detect the angle and intensity of scattering, working in conjunction with a microscope placed behind it to observe the results. The components

Student view

including the flash of light, the fluorescent screen, and the microscope are connected by lines showing how they align to measure and observe the particle scattering.

[Generated by AI]

The Geiger–Marsden–Rutherford experiment showed that the greater the scattering angle, the fewer the scattered alpha particles that were observed.

Also, the number of scattered alpha particles observed in a particular time interval at for a certain angle depends on the energy of the alpha particles. **Figure 2** shows what Rutherford predicted.

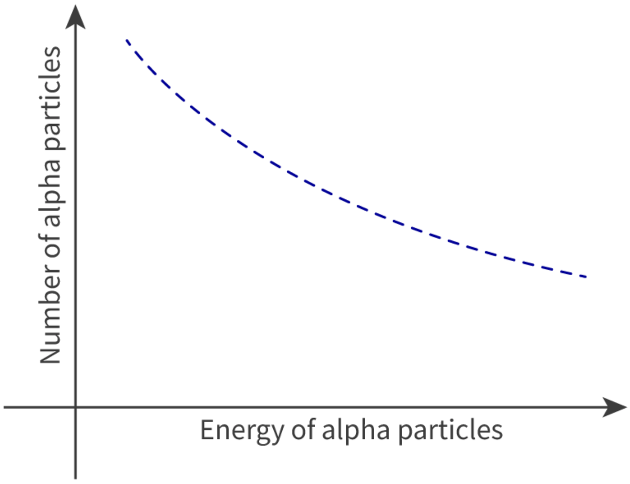


Figure 2. Predicted graph of number of alpha particles against energy of alpha particles.

More information for figure 2

The graph depicts a predicted relationship between the number of alpha particles and their energy. The X-axis represents the energy of alpha particles, labeled without specific units, while the Y-axis represents the number of alpha particles. The graph shows a downward sloping curve, indicating an inverse relationship. As the energy of alpha particles increases, the number of alpha particles decreases. This trend reflects Rutherford's prediction that higher energy particles have a greater chance of interacting with atomic nuclei, thereby reducing the number observed.

[Generated by AI]

We can see from the graph in **Figure 2** that Rutherford predicted that as the energy of the alpha particles increases, the number of scattered alpha particles observed should decrease. This is in part due to high energy particles having a lower interaction time with the atomic nucleus.

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

The alpha particles and the nucleus are both positively charged. Rutherford's model predicts that they will repel each other according to Coulomb's law (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)). Since the nucleus is much heavier than the alpha particle, it experiences negligible deflection, and only the alpha particle changes direction. If the alpha particle travels faster (has more energy), the change in direction will be less, so fewer alpha particles will be detected at a certain angle.

However, Rutherford's experiments did not verify his prediction. **Figure 3** shows his prediction (dashed line) and the actual results (solid line) for scattering of alpha particles by an angle of 60°, by atoms in a sheet of lead-208. How is the prediction different from the actual results? Over which range of energies are the curves the same? Over which range of energies do the results deviate from the prediction?

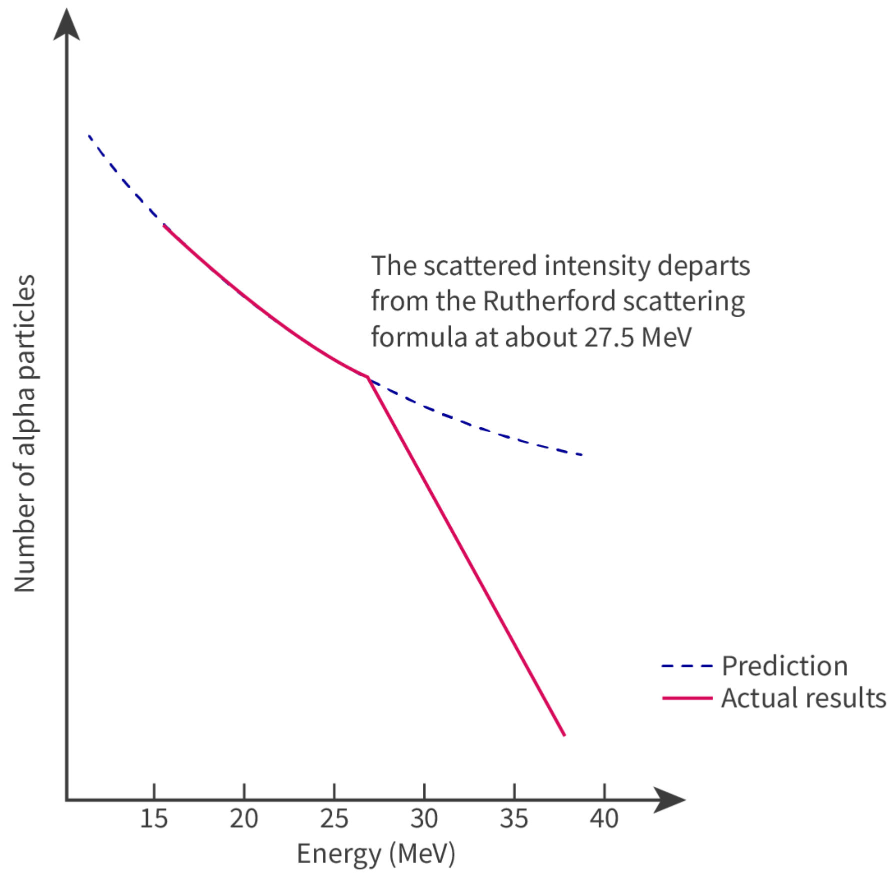


Figure 3. Rutherford's prediction and actual results.

More information for figure 3

The graph shows a comparison between Rutherford's predicted and actual results for the scattering of alpha particles.

- **X-axis:** Represents energy in MeV (Mega-electron Volts), ranging from 15 to 45.
- **Y-axis:** Represents the number of alpha particles.
- **Dashed Line (Prediction):** Starts at a higher number of alpha particles at 15 MeV and gradually decreases as energy increases.

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

- **Solid Line (Actual Results):** Starts similarly to the predicted line but diverges significantly after 27.5 MeV.

Observation: Both the prediction and actual results follow a similar trend up to approximately 27.5 MeV, where a noticeable deviation occurs in the actual results. The scattered intensity departs from the Rutherford scattering formula at about 27.5 MeV, as indicated by the annotation on the graph. The predicted line continues to decrease gradually, while the actual results show a steeper decline.

[Generated by AI]

Distance of closest approach

In Rutherford scattering, the only force between the alpha particle and the nucleus is the electric force. Could the deviation shown in the graph in **Figure 3** be explained by the existence of another force?

Imagine that an alpha particle is fired head-on at the nucleus of an atom with proton number Z. The alpha particle has kinetic energy E_k. The alpha particle and the nucleus are far enough apart that there is no electric force between them. As the alpha particle approaches the nucleus, it is repelled by the electric force and decelerates. At a certain point, the alpha particle will momentarily stop – it will have reached its minimum distance from the nucleus (**Figure 4**).

We call this distance the distance of closest approach. All of the kinetic energy of the alpha particle has transferred to the electric potential energy of the system (nucleus and alpha particle) (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)).

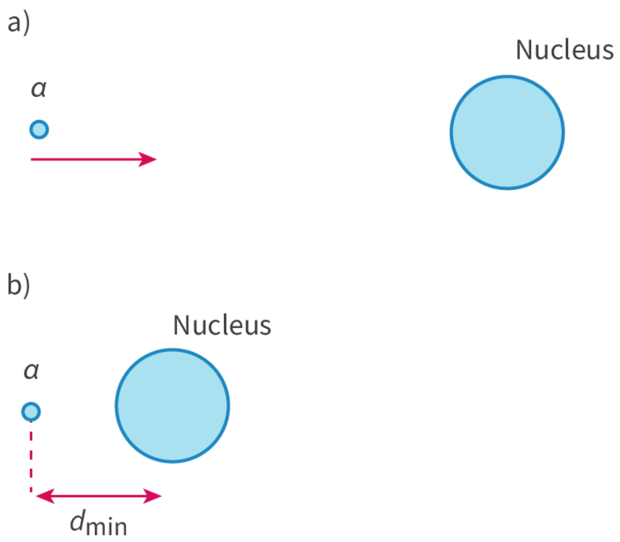


Figure 4. An alpha particle shot head-on at a nucleus (a) stops at its minimum distance from the nucleus (b).

More information for figure 4

Student
view

The image consists of two diagrams labeled (a) and (b). Diagram (a) shows an alpha particle labeled 'α' approaching a nucleus from the left side, depicted as a smaller circle moving towards a larger circle labeled 'Nucleus'. The arrow pointing right indicates its movement. Diagram (b) shows the same alpha particle halting at a minimum distance 'd_min' near the left side of the nucleus. The arrow now points left, indicating it has stopped and reversed direction, representing the transfer of kinetic energy to electric potential energy.

[Generated by AI]

Electric potential energy is given by:

$$E_{\mathrm{p}} = k\frac{q_1 q_2}{r}$$

where:

- q_1 is the alpha particle, which is positively charged: $q_1 = 2e$
- q_2 is the nucleus, which is positively charged: $q_2 = Ze$
- r is the distance of closest approach: $r = d_{\mathrm{min}}$

The system now has an electric potential energy of:

$$E = k\frac{2e \times Ze}{d_{\mathrm{min}}}$$

where E is also equal to the initial kinetic energy of the particle, thus substituting and rearranging for d_{min}:

$$d_{\mathrm{min}} = k\frac{2Ze^2}{E_K}$$

This equation shows that when the alpha particle has greater kinetic energy, the distance of closest approach decreases and the alpha particle will come closer to the nucleus.

Worked example 1

What is the distance of closest approach between an alpha particle and a silver nucleus $({}_{47}\mathrm{Ag})$ when the alpha particle is fired head-on at the nucleus with an energy of 5.0 MeV?

$$Z = 47$$

$$\begin{aligned} E &= 5.0\ \mathrm{MeV} \\ &= 8.0 \times 10^{-13}\ \mathrm{J} \end{aligned}$$

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

$$d_{\min} = k\frac{2Ze^2}{E}$$
$$= 8.99 \times 10^9 \times \frac{2 \times 47 \times (1.60 \times 10^{-19})^2}{8.0 \times 10^{-13}}$$
$$= 2.704 \times 10^{-14}\ \text{m}$$
$$= 2.7 \times 10^{-14}\ \text{m (2 s.f.)}$$

If the initial (kinetic) energy of the alpha particle is large enough, then it can come so close to the nucleus that a new force acts between the alpha particle and the nucleus. This is called the strong nuclear force (see subtopic E.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44319/)), and it acts on protons and neutrons only.

The strong nuclear force is approximately 100 times stronger than the electric force, and, except at exceptionally close range (smaller than the diameter of a nucleon) it is only attractive. However, its range is in the order of 10^{-15} m, so it only influences the results of Rutherford scattering when the alpha particle comes very close to the nucleus. This is why deviations occur only at relatively higher energies of the alpha particles.

Nuclear densities

Rutherford scattering and other experiments have shown that the size of the nucleus depends on the number of nucleons (proton and neutrons) it contains.

Assuming that the nucleus of the atom is a sphere, the radius of this sphere can be expressed as shown in the equation in **Table 1**.

Table 1. Equation for the radius of a nucleus.

Equation	Symbols	Units
$R = R_0 A^{\frac{1}{3}}$	R = radius of nucleus	metres (m)
	R_0 = Fermi radius (1.20×10^{-15} m) Given in section 1.6.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/fundamental-constants-id-45155/) of the DP physics data booklet	metres (m)
	A = nucleon number	unitless

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Worked example 2

Determine the radius of a nucleus of helium $\left({}^{4}_{2}\mathrm{He}\right)$. Give your answer to 2 significant figures.

$$A = 4$$

$$\begin{aligned} R &= R_0 A^{\frac{1}{3}} \\ &= 1.20\times10^{-15}\times4^{\frac{1}{3}} \\ &= 1.905\times10^{-15}\ \mathrm{m} \\ &= 1.9\times10^{-15}\ \mathrm{m}\ (2\ \mathrm{s.f.}) \end{aligned}$$

Since all nucleons within a nucleus are approximately the same distance from one another, all atomic nuclei have approximately the same density, regardless of which element they come from. We can demonstrate this as follows:

The equation for density (see subtopic B.1 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43777/)) is:

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

The mass of a neutron and the mass of a proton can be considered to be equal to each other.

The total mass of a nucleus can be expressed as:

$$m = Am_{\mathrm{nucleon}}$$

The volume of a sphere is given by:

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

Substituting for *R* gives:

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

$$V = \frac{4}{3}\pi {R_0}^3 A$$

So the density of a nucleus is:

$$\rho = \frac{Am_{\mathrm{nucleon}}}{\frac{4}{3}\pi {R_0}^3 A}$$

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

$$\rho = \frac{m_{\text{nucleon}}}{\frac{4}{3}\pi {R_0}^3} \simeq 2.3 \times 10^{17}\ \text{kg m}^{-3}$$

Bohr model and discrete energy levels

An atom has discrete atomic energy levels (see section E.1.1 (/study/app/math-aa-hl/sid-423-cid-762593/book/atoms-and-photons-id-46593/)). This is because the orbits of an electron around the nucleus are also discrete.

As well as behaving like particles, electrons also behave like standing waves that exist around the nucleus. This is known as wave–particle duality (see section E.2.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-wave-nature-of-matter-hl-id-46463/)).

The wavelength associated with the electron is:

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

where:

- h is the Planck constant (6.63×10^{-34} J s)
- m is the mass of the electron
- v is the linear speed of the electron.

To form a standing wave, the length of the space that the wave occupies must be an integer number of wavelengths (see subtopic C.4 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43788/)). In this case, the space is the length of the orbit:

$$\frac{2\pi r}{\lambda} = n$$

$$\frac{2\pi r}{\frac{h}{mv}} = n$$

$$mvr = \frac{nh}{2\pi}$$

Remember from previous study (Subtopic A.4 (/study/app/math-aa-hl/sid-423-cid-762593/book/angular-momentum-hl-id-43165/)) that mvr is also equal to the angular momentum of a particle.

The electron can be considered a point mass. The equation can therefore be written as shown in **Table 2**.

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Table 2. Equation for the quantisation of angular momentum (mvr) for hydrogen.

Equation	Symbols	Units
$mvr = \frac{nh}{2\pi}$	m = mass of electron	kilograms (kg)
	v = linear speed of electron	metres per second (m s^{-1})
	r = distance from nucleus (radius of motion)	metres (m)
	n = energy level (positive integer)	unitless
	h = Planck constant (6.63×10^{-34} J s) Given in section 1.6.3 (/study/app/math-aa-hl/sid-423-cid-762593/book/fundamental-constants-id-45155/) of the DP physics data booklet	joule second (J s)

Figure 5 shows a way of visualising electrons as standing waves, for the three lowest energy levels.

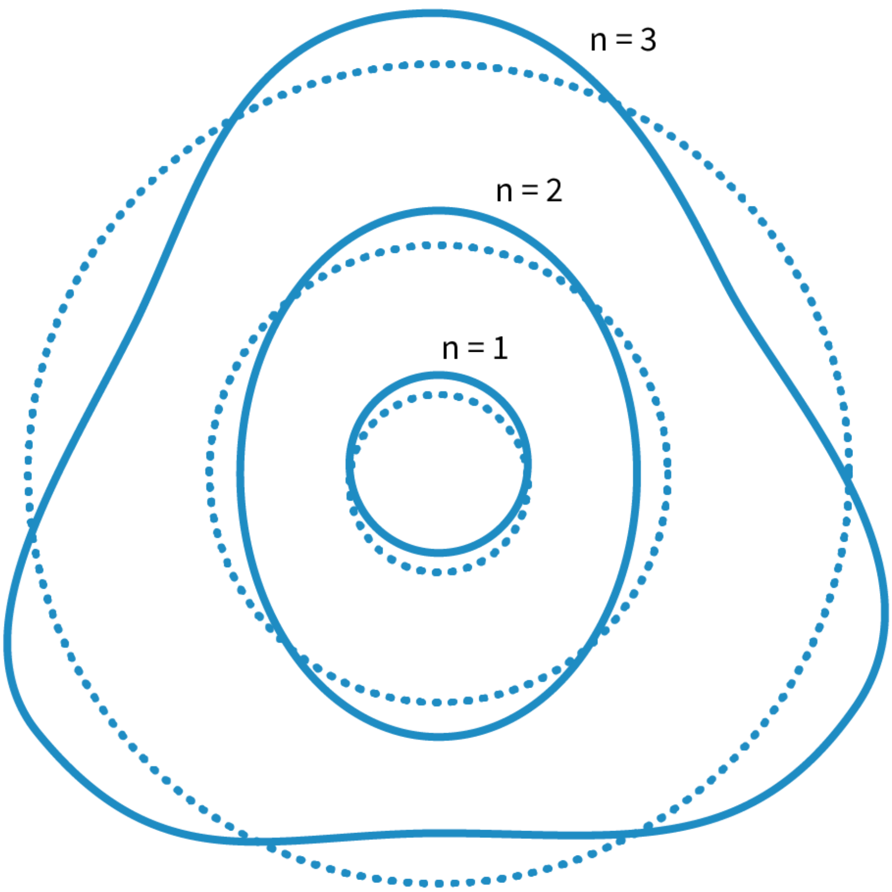


Figure 5. The electron standing waves for the three lowest energy levels in a hydrogen atom.

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

More information for figure 5

The image illustrates the electron standing waves for three different energy levels in a hydrogen atom, labeled as n=1, n=2, and n=3. Each energy level is represented by a circle with increasing size as the energy level increases. The smallest circle, labeled n=1, is within larger concentric circles labeled n=2 and n=3. These circles represent the areas of probability where the electron is most likely to be found, with each level having a specific pattern and radial distribution.

[Generated by AI]

Accepting that the electron in a hydrogen atom can only have discrete values for its energy, the Bohr model predicts the magnitude of the energy levels as given by the equation in **Table 3.**

Table 3. Equation for energy for the Bohr model for hydrogen.

Equation	Symbols	Units
$E = -\frac{13.6}{n^2}$ eV	E = energy	electronvolts (eV)
	n = energy level (positive integer)	unitless

There are two things to note about this equation:

- It only gives the energy for the hydrogen atom.
- The energy is negative because the electron is attracted to the nucleus (the proton).

Exercise 1

Click a question to answer

Higher level (HL)

Work through the activity to check your understanding of the Bohr model for hydrogen.

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

Activity

- **IB learner profile attribute:** Knowledgeable
- **Approaches to learning:** Thinking skills — Applying key ideas and facts in new contexts
- **Time required to complete activity:** 25 minutes
- **Activity type:** Individual activity

In a hydrogen atom, there is one proton in the nucleus and one electron orbiting the nucleus. You are going to use your understanding of circular motion (see subtopic A.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43136/)) and the electric force and electric potential energy (see subtopic D.2 (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-44743/)) to derive the expression for the energy of a hydrogen atom:

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

The electron is moving in circular motion. The electric force between the electron and the proton provides the centripetal force. The equation for centripetal force is:

$$F = \frac{mv^2}{r}$$

1. Write down the Coulomb's law equation for electric force.
2. Equate the equations for force and solve for the linear speed of the electron.
3. Write down the equation for angular momentum and solve for linear speed.
4. Combine the equations for linear speed and solve for radius of motion *r*.
5. Write down the equation for kinetic energy. Substitute for linear speed in the equation.
6. Write down the equation for electric potential energy.
7. Add the kinetic energy and the electric potential energy to find the total energy.
8. Substitute the expression for *r* in the equation for total energy and cancel expressions where possible.
9. Substitute the values of the constants and convert to eV.

1. $F = k\frac{q_1 q_2}{r^2}$

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

2. $\frac{mv^2}{r} = k\frac{q_1 q_2}{r^2}$

$$v^2 = k\frac{e^2}{mr}$$

3. $mvr = n\frac{h}{2\pi}$

$$v = n\frac{h}{2\pi mr}$$

4. $v^2 = k\frac{e^2}{mr}$

and

$$v = n\frac{h}{2\pi mr}$$

where

$$r = n^2\left(\frac{h^2}{4\pi^2 kme^2}\right)$$

5. $E_\mathrm{k} = \frac{1}{2}mv^2 \quad E_\mathrm{k} = k\frac{e^2}{2r}$

6. $E_p = k\frac{q_1 q_2}{r} \quad E_\mathrm{p} = -k\frac{e^2}{r}$

7. $E = E_\mathrm{k} + E_\mathrm{p} \quad E = k\frac{e^2}{2r} - k\frac{e^2}{r} \quad E = -k\frac{e^2}{2r}$

8. $E = -\frac{2\pi^2 me^4 k^2}{n^2 \times h^2}$

9. $E = -\frac{13.6}{n^2}\ \mathrm{eV}$

6 section questions

Question 1

HL Difficulty:

What are deviations from Rutherford scattering at high energies evidence for?

1 The existence of the strong nuclear force ✓

2 The atom consists of a nucleus and orbiting electrons

3 The Thomson model of the atom is wrong

Student view

Overview
(/study/ap
aa-
hl/sid-
423-
cid-
762593/c

4 Energy in atoms is discrete

Explanation

Rutherford's predictions were based on the assumption that the only force between the nucleus and the alpha particle was the electric force, which is repulsive. There is a decrease in the number of scattered alpha particles at high energies, which is explained by the strong nuclear force, which is attractive.

Question 2

HL Difficulty:

According to the Bohr model for hydrogen, which statement explains why atomic energy levels are discrete?

1 The angular momentum of the electron is quantised ✓

2 Photons are emitted when electrons change their orbits

3 All nuclei have the same density

4 The strong nuclear force exists

Explanation

In the Bohr model for hydrogen, in order to have discrete energy levels, the electron needs to have discrete values for angular momentum.

Question 3

HL Difficulty:

What is the energy of the first excited state of the hydrogen atom?

1 −3.4 eV ✓

2 −13.6 eV

3 13.6 eV

4 3.4 eV

Explanation

first excited state: $n = 2$

Student view

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

Question 4

HL Difficulty:

Why are deviations from Rutherford scattering mainly observed at high energies?

1 The alpha particles approach the nucleus close enough to feel the strong nuclear force ✓

2 The alpha particles do not interact with the nucleus

3 The alpha particles destroy the atoms

4 The nuclei are decaying

Explanation

The strong nuclear force has a very short range, so the alpha particles need high energies to approach the nucleus close enough to experience it.

Question 5

HL Difficulty:

An alpha particle with an energy of 25 MeV is fired directly towards a stationary atom of gold $\left({}^{197}_{79}\mathrm{Au}\right)$. Determine the distance of the closest approach? Give your answer to an appropriate number of significant figures.

The distance of closest approach is 1 9.1 ✓ $\times 10^{-15}$ m

Accepted answers and explanation

#1 9.1

General explanation

$$A = 197$$

$$Z = 79$$

$$E = 25 \text{ MeV}$$
$$= 4.0\times10^{-12} \text{ J}$$

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

$$
\begin{aligned}
d_{\min} &= k\frac{2Ze^2}{E} \\
&= 8.99 \times 10^9 \times \frac{2 \times 79 \times (1.6 \times 10^{-19})^2}{4.0 \times 10^{-12}} \\
&= 9.091 \times 10^{-15}\ \text{m} \\
&= 9.1 \times 10^{-15}\ \text{m (2 s.f.)}
\end{aligned}
$$

Question 6

HL Difficulty:

Calculate the radius of the orbit of an electron in a hydrogen atom in its second excited state (n = 3). Its velocity is 7.2×10^5 m s^{-1}.

1 4.8×10^{-10} m ✓

2 3.2×10^{-10} m

3 1.8×10^{-13} m

4 2.6×10^{-13} m

Explanation

$$mvr = \frac{nh}{2\pi}$$

$$
\begin{aligned}
r &= \frac{nh}{2\pi mv} = \frac{3 \times 6.63 \times 10^{-34}}{2\pi \times 9.1 \times 10^{-31} \times 7.2 \times 10^5} \\
&= 4.8 \times 10^{-10}\ \text{m}
\end{aligned}
$$

E. Nuclear and quantum physics / E.1 Structure of the atom

Summary and key terms

Section Student... (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/summary-and-key-terms-id-46600/print/) Assign

- The Geiger–Marsden–Rutherford experiment is evidence that the atom is mostly empty space, with a positive nucleus at the centre and electrons orbiting the nucleus.
- A nucleus can be described using nuclear notation, ${}^{A}_{Z}\text{X}$, where X is the chemical symbol, A is the nucleon number and Z is the proton number.
- An atom has discrete atomic energy levels because electrons have discrete orbits.
- When an atom absorbs a photon with energy equal to a difference in energy levels, an electron makes a transition to a higher energy level. When an electron makes a transition to a lower energy level, a photon is emitted with energy equal to a difference in energy levels.

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

- Emission and absorption spectra have lines that correspond to the electromagnetic radiation emitted or absorbed during atomic transitions. They are evidence for discrete atomic energy levels and can be used to determine chemical composition.

Higher level (HL)

- Deviation from Rutherford scattering at high energies is evidence for the existence of the strong nuclear force.
- The relationship between the nucleon number and the radius of a nucleus shows that all nuclei have the same density.
- According to the Bohr model for hydrogen, the electron only exists in certain orbits because its angular momentum is quantised.

Section Student... (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/atoms-and-photons-id-46593/print/) Assign

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

Key terms

Review these key terms. Do you know them all? Fill in as many gaps as you can using the terms in this list.

1. In the Rutherford model of the atom, positive charge is concentrated in the of the atom.
2. The number is the number of protons in the nucleus of the atom. The number is the number of protons and neutrons.
3. An atom has discrete atomic . The energy level with the lowest energy is known as the state.
4. During an atomic , energy is emitted or absorbed via a discrete 'packet' of energy called a .
5. An spectrum shows the wavelengths emitted by an atom. An spectrum shows the wavelengths absorbed by an atom.
6. [HL] Deviations from Rutherford scattering at energy provide evidence for the force.
7. [HL] In the Bohr model for , the of the electron gives rise to the discrete atomic energy levels.

high | nucleus | hydrogen | photon | emission | nucleon | absorption | energy levels | strong nuclear | angular momentum | ground | proton | transition

Check

Interactive 1. Atomic Models and Energy Levels.

E. Nuclear and quantum physics / E.1 Structure of the atom

Checklist

Section Student… (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/checklist-id-46601/print/) Assign

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

What you should know

At the end of this subtopic you should be able to:

- Describe the Geiger—Marsden—Rutherford experiment and how it led to the discovery of the nucleus.
- Understand that photons are emitted and absorbed during atomic transitions and use the equation:

$$E = hf$$

- Understand that emission and absorption spectra provide evidence for discrete atomic energy levels and chemical composition.

Higher level (HL)

- Understand why deviations from Rutherford scattering occur and the concept of distance of closest approach.
- Understand the relationship between nucleon number and radius of a nucleus and use the equation:

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

- Understand that quantisation of angular momentum can explain discrete energy levels in the Bohr model for hydrogen and use the equations:

$$mvr = \frac{nh}{2\pi} \text{ and } E = -\frac{13.6}{n^2} \text{ eV}$$

E. Nuclear and quantum physics / E.1 Structure of the atom

Investigation

Section Student... (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/investigation-id-46602/print/) Assign

- **IB learner profile attribute:** Inquirer
- **Approaches to learning:** Research skills – Using search engines and libraries effectively
- **Time required to complete activity:** 1 hour
- **Activity type:** Individual activity

Student view

Overview (/study/app/math-aa-hl/sid-423-cid-762593/c

Theories about the atom were introduced in the ancient world. However, the main development of the model of the atom has been based on experiments and theories during the last few centuries.

Your task

You are going to research the evolution of the model of the atom. Research the discovery of the atom, the electron, the proton and the neutron.

For each discovery, find out:

- who discovered it and when
- the experiments and evidence that led to the discovery
- how the discovery changed the model of the atom.

Create a timeline showing the development of the model of the atom.

Creativity, activity, service

- **Strand:** Creativity
- **Learning outcome:** Demonstrate the skills and recognise the benefits of working collaboratively

Often, theatre is a good way to communicate complex ideas to an amateur or young audience. Work with a group to produce a theatre performance or sketch to explain the development of the atomic model to younger students.

E. Nuclear and quantum physics / E.1 Structure of the atom

Reflection

Section Student... (0/0) Feedback Print (/study/app/math-aa-hl/sid-423-cid-762593/book/reflection-id-47886/print/) Assign

Student view

Overview
(/study/app/math-aa-hl/sid-423-cid-762593/c

Teacher instructions

The goal of this section is to encourage students to reflect on their learning and conceptual understanding of the subject at the end of this subtopic. It asks them to go back to the guiding questions posed at the start of the subtopic and assess how confident they now are in answering them. What have they learned, and what outstanding questions do they have? Are they able to see the bigger picture and the connections between the different topics?

Students can submit their reflections to you by clicking on 'Submit'. You will then see their answers in the 'Insights' part of the Kognity platform.

Reflection

Now that you've completed this subtopic, let's come back to the guiding questions introduced in The big picture (/study/app/math-aa-hl/sid-423-cid-762593/book/the-big-picture-id-43191/).

- What is the current understanding of the nature of an atom?
- What is the role of evidence in the development of models of the atom?
- In what ways are previous models of the atom still valid despite recent advances in understanding?

With these questions in mind, take a moment to reflect on your learning so far and type your reflections into the space provided.

You can use the following questions to guide you:

- What main points have you learned from this subtopic?
- Is anything unclear? What questions do you still have?
- How confident do you feel in answering the guiding questions?
- What connections do you see between this subtopic and other parts of the course?

⚠ Once you submit your response, you won't be able to edit it.

Student view

0/2000

Submit

Rate subtopic E.1 Structure of the atom

Help us improve the content and user experience.