2.5 Energy Levels & Photon Emission

PHYSICS

AQA A Level Revision Notes

EXAM PAPERS PRACTICE

A Level Physics AQA

2.5 Energy Levels & Photon Emission

CONTENTS

Ionisation & Excitation

- **Ionisation** of an atom is:

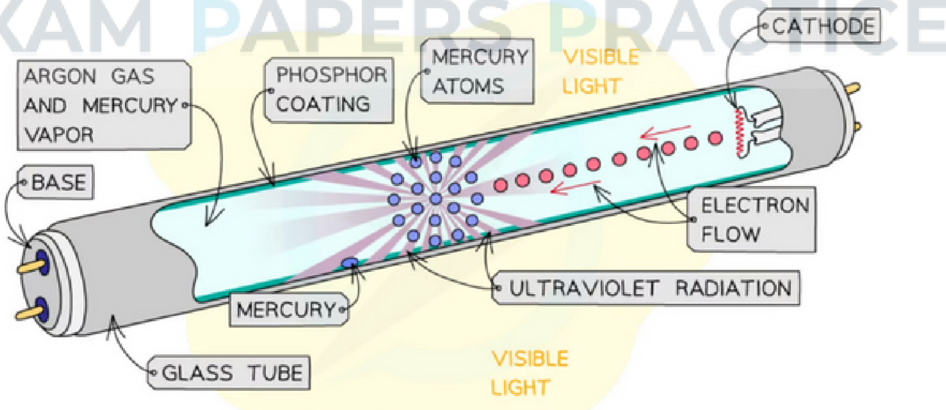
  The removal, or addition, of an electron from, or to, an atom when given sufficient energy

- **Excitation** of electrons is:

  When an electron is given enough energy to move up an energy level, but not enough to leave the atom

Fluorescent Tube

- Fluorescence occurs when an electron in an atomic orbital absorbs energy from an interaction with a photon or a collision with another atom
- Fluorescent tubes are partially evacuated glass tubes filled with low-pressure mercury vapour with a phosphor coating on the glass
- When an electric current is passed through the vapour, the electrons in mercury are **excited** and move to a **higher energy level**
- This high energy level state is unstable and so the electron moves back to its original state, or **de-excites**
- As it de-excites, the electron releases some of that energy in the form of a **UV photon**
- This UV light then excites the electrons in the phosphor coating
- As a result, visible light photons are released when the electrons return to their original energy state, which provides the fluorescent glow



Fluorescent tubes operate on the basis of excitation and de-excitation of electrons leading to the emission of visible light

The Photon Model

- Photons are fundamental particles which make up all forms of electromagnetic radiation
- A photon is a massless "packet" or a "quantum" of electromagnetic energy
- What this means is that the energy is not transferred continuously, 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, rather than supplying a consistent amount of energy

Calculating Photon Energy

- The energy of a photon can be calculated using the formula:

$$E = hf$$

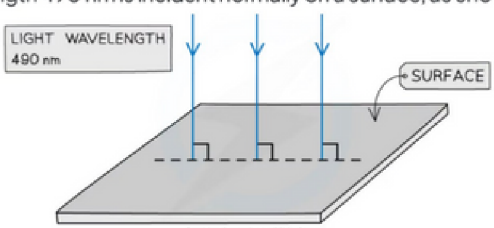
- Using the wave equation, energy can also be equal to:

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

- 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

Worked Example

Light of wavelength 490 nm is incident normally on a surface, as shown in the diagram.



The power of the light is 3.6 mW. The light is completely absorbed by the surface. Calculate the number of photons incident on the surface in 2.0 s.

Step 1: Write the known quantities

Wavelength, $\lambda = 490 \text{ nm} = 490 \times 10^{-9}$ m

Power, $P = 3.6 \text{ mW} = 3.6 \times 10^{-3}$ W

Time, t = 2.0 s

Step 2: Write the equations for wave speed and photon energy

$$\text{Wave speed: } c = f\lambda \rightarrow f = \frac{c}{\lambda}$$

$$\text{Photon energy: } E = hf \rightarrow E = \frac{hc}{\lambda}$$

Step 3: Calculate the energy of one photon

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34}) \times (3.0 \times 10^{8})}{490 \times 10^{-9}} = 4.06 \times 10^{-19} \text{ J}$$

Step 4: Calculate the number of photons hitting the surface every second

$$\frac{\text{Power of light source}}{\text{Energy of one photon}} = \frac{3.6 \times 10^{-3}}{4.06 \times 10^{-19}} = 8.9 \times 10^{15} \text{ s}^{-1}$$

Step 5: Calculate the number of photons that hit the surface in 2 s

$$(8.9 \times 10^{15}) \times 2 = 1.8 \times 10^{16}$$

Exam Tip

Make sure you learn the definition for a photon: *discrete quantity / packet / quantum of electromagnetic energy* are all acceptable definitions. The values of Planck's constant and the speed of light will always be available on the datasheet, however, it helps to memorise them to speed up calculation questions!

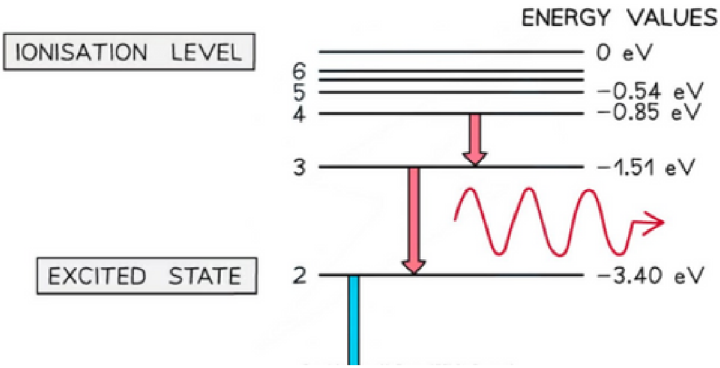
Line Spectra & Energy Levels

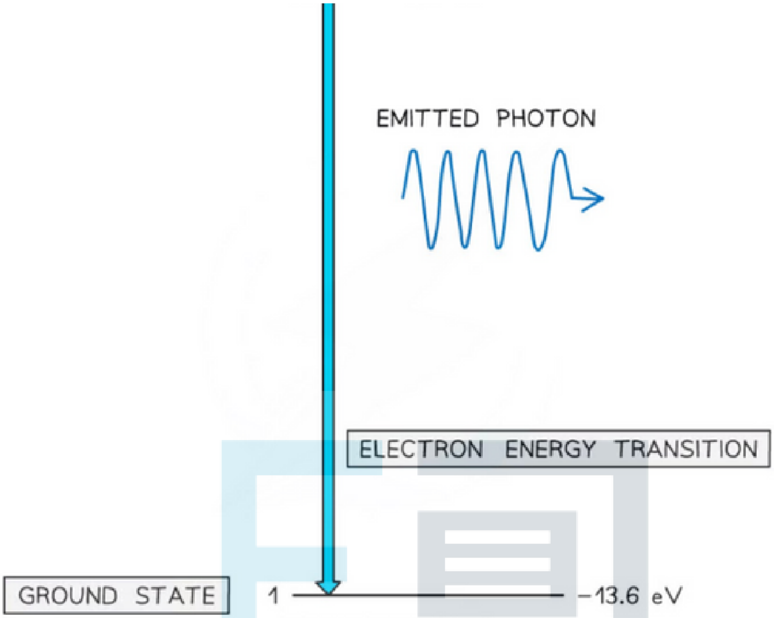
Atomic Energy Levels

- Electrons in an atom can have only certain specific energies
    - These energies are called **electron energy levels**

- They can be represented as a series of stacked horizontal lines increasing in energy
- Normally, electrons occupy the lowest energy level available, this is known as the **ground state**
- Electrons can gain energy and move up the energy levels if it absorbs energy either by:
    - Collisions with other atoms or electrons
    - Absorbing a photon
    - A physical source, such as heat

- This is known as **excitation**, and when electrons move up an energy level, they are said to be in an **excited state**
- If the electron gains enough energy to be removed from the atom entirely, this is known as **ionisation**
- When an electron returns to a lower energy state from a higher excited state, it releases energy in the form of a photon





Electron energy levels in atomic hydrogen. Photons are emitted when an electron moves from a higher energy state to a lower energy state

Line Spectra

- Line spectra is a phenomenon which occurs when excited atoms emit light of certain wavelengths which correspond to different colours
- The emitted light can be observed as a series of coloured lines with dark spaces in between
  - These series of coloured lines are called **line or atomic spectra**

- Each element produces a unique set of spectral lines
- No two elements emit the same set of spectral lines, therefore, elements can be identified by their line spectrum
- There are two types of line spectra: **emission spectra** and **absorption spectra**

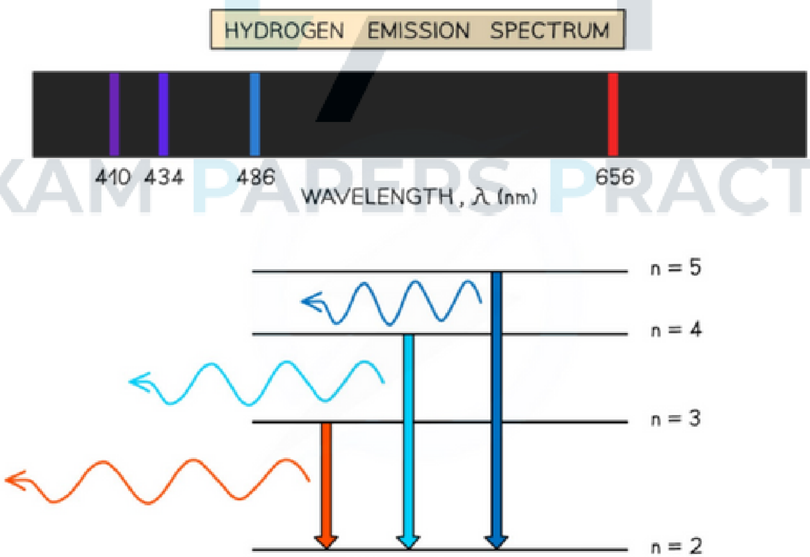
Emission Spectra

- When an electron transitions from a higher energy level to a lower energy level, this results in the **emission** of a photon
- Each transition corresponds to a different wavelength of light and this corresponds to a line in the spectrum
- The resulting emission spectrum contains a set of discrete wavelengths, represented by coloured lines on a black background
- Each emitted photon has a wavelength which is associated with a discrete change in energy, according to the equation:

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

- Where:
  - ΔE = change in energy level (J)
  - h = Planck's constant (J s)
  - f = frequency of photon (Hz)
  - c = the speed of light (m s^{-1})
  - λ = wavelength of the photon (m)

- Therefore, this is evidence to show that electrons in atoms can only transition between discrete energy levels



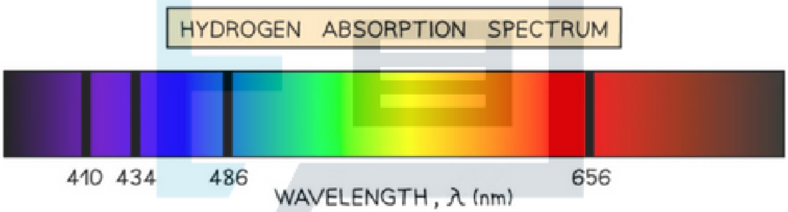
Emission spectrum of Hydrogen gas

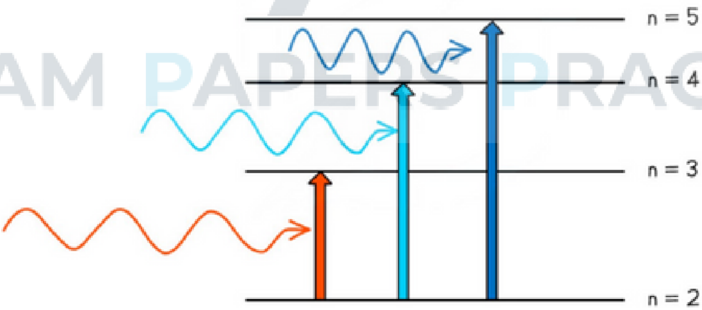
Absorption Spectra

- An atom can be raised to an excited state by the absorption of a photon
- When white light passes through a **cool, low pressure gas** it is found that light of certain wavelengths are missing
    - This type of spectrum is called an absorption spectrum

- An absorption spectrum consists of a continuous spectrum containing all the colours with dark lines at certain wavelengths
- These dark lines correspond exactly to the differences in energy levels in an atom
- When these electrons return to lower levels, the photons are emitted in all directions, rather than in the original direction of the white light
    - Therefore, some wavelengths appear to be missing

- The wavelengths missing from an absorption spectrum are the same as their corresponding emission spectra of the same element





Absorption spectrum of Hydrogen gas

Difference in Discrete Energy Levels

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

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

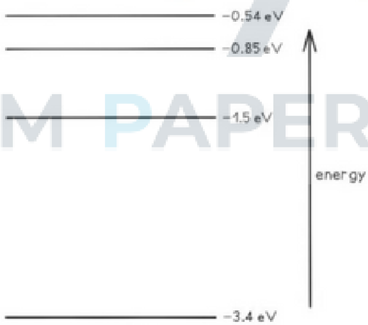
- Where:
  - E_1 = Energy of the higher level (J)
  - E_2 = Energy of the lower level (J)
  - h = Planck's constant (J s)
  - f = Frequency of photon (Hz)
- Using the wave equation, the wavelength of the emitted, or absorbed, radiation can be related to the energy difference by the equation:

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

- This equation shows that the larger the difference in energy of two levels ΔE ($E_2 - E_1$) the shorter the wavelength λ and vice versa

Worked Example

Some electron energy levels in atomic hydrogen are shown below.

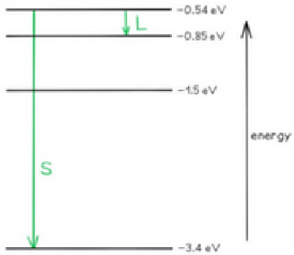


The longest wavelength produced as a result of electron transitions between two of the energy levels is 4.0×10^{-6} m.a) Draw and mark:

- The transition giving rise to the wavelength of 4.0×10^{-6} m with letter **L**.
- The transition giving rise to the shortest wavelength with letter **S**.

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

Part (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**)

Part (b)

Step 1: Write down the equation linking the wavelength and the energy levels

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

Step 2: Identify the energy levels giving rise to the shortest wavelength

This will be the energy levels which have the largest difference

E_1 = 0.54 eV

E_2 = 3.4 eV

We can ignore the negative sign we just want the difference between the levels and the wavelength should be positive

Step 3: Calculate the wavelength

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

$$\lambda = \frac{(6.63 \times 10^{-34}) \times (3 \times 10^{8})}{(3.4 - 0.54) \times (1.6 \times 10^{-19})} = 4.347 \times 10^{-7}\ \text{m} = 435\ \text{nm}$$

2.5.3 Wave-Particle Duality

Wave-Particle Duality

- Light can behave as a particle (i.e. photons) **and** a wave
- This phenomenon is called the wave-particle nature of light or **wave-particle duality**
- Light interacts with matter, such as electrons, **as a particle**
    - The evidence for this is provided by the photoelectric effect
- Light propagates through space **as a wave**
    - The evidence for this comes from the diffraction and interference of light in Young's Double Slit experiment

Light as a Particle

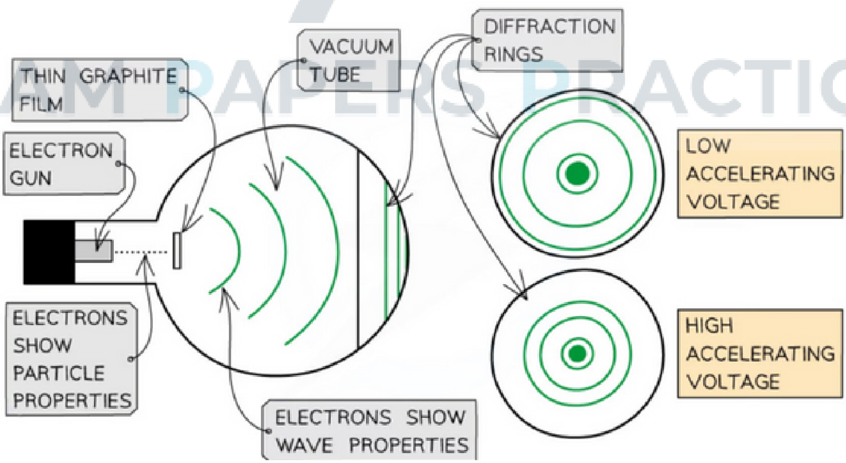
- Einstein proposed that light can be described as a quanta of energy that behave as particles, called photons
- The photon model of light explains that:
    - Electromagnetic waves carry energy in discrete packets called photons
    - The energy of the photons are quantised according to the equation $E = hf$
    - In the photoelectric effect, each electron can absorb only a single photon - this means only the frequencies of light above the threshold frequency will emit a photoelectron
- The wave theory of light does **not** support the idea of a threshold frequency
    - The wave theory suggests any frequency of light can give rise to photoelectric emission if the exposure time is long enough
    - This is because the wave theory suggests the energy absorbed by each electron will increase gradually with each wave
    - Furthermore, the kinetic energy of the emitted electrons should increase with radiation intensity
    - However, in the photoelectric effect, this is not what is observed
- If the frequency of the incident light is above the threshold and the intensity of the light is increased, **more photoelectrons are emitted per second**
- Although the wave theory provides good explanations for phenomena such as interference and diffraction, it fails to explain the photoelectric effect

Compare wave theory and particulate nature of light

The wave theory of light suggests...	This is wrong because...
Any frequency of light can give rise to photoelectric emission if the exposure time is long enough	Photoelectrons will be released immediately if the frequency is above the threshold for that metal
The energy absorbed by each electron will increase gradually with each wave	Energy is absorbed instantaneously – photoelectrons are either emitted or not emitted after exposure to light
The kinetic energy of the emitted electrons should increase with radiation intensity	If the intensity of the light is increased, more photoelectrons are emitted per second

Electron Diffraction

- Louis de Broglie discovered that matter, such as electrons, can behave as a wave
- He showed a diffraction pattern is produced when a beam of electrons is directed at a thin graphite film
- Diffraction is a property of waves, and cannot be explained by describing electrons as particles

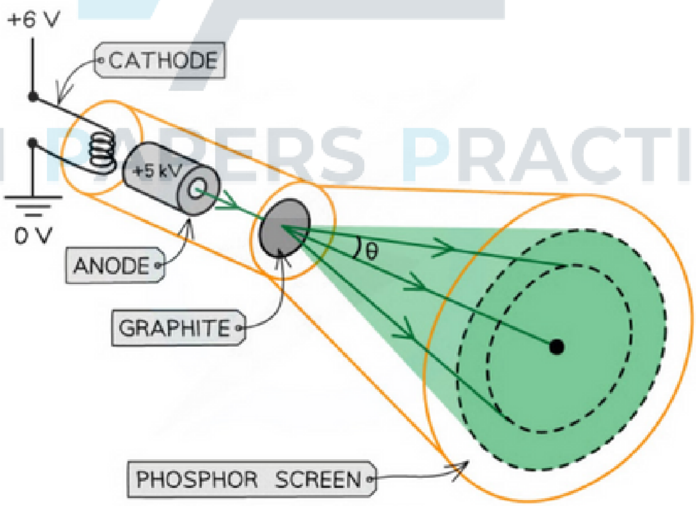


Electrons accelerated through a high potential difference demonstrate wave-particle duality

- In order to observe the diffraction of electrons, they must be focused through a gap similar to their size, such as an atomic lattice
- Graphite film is ideal for this purpose because of its crystalline structure
    - The gaps between neighbouring planes of the atoms in the crystals act as slits, allowing the electron waves to spread out and create a diffraction pattern
- The diffraction pattern is observed on the screen as a series of concentric rings
    - This phenomenon is similar to the diffraction pattern produced when light passes through a diffraction grating
    - If the electrons acted as particles, a pattern would not be observed, instead, the particles would be distributed uniformly across the screen
- It is observed that a larger accelerating voltage reduces the diameter of a given ring, while a lower accelerating voltage increases the diameter of the rings

Investigating Electron Diffraction

- Electron diffraction tubes can be used to investigate the wave properties of electrons
- The electrons are accelerated in an electron gun to a high potential, such as 5000 V, and are then directed through a thin film of graphite
- The electrons diffract from the gaps between carbon atoms and produce a circular pattern on a fluorescent screen made from phosphor



Experimental setup to demonstrate electron diffraction

- Increasing the voltage between the anode and the cathode causes the energy, and hence speed, of the electrons to increase
- The kinetic energy of the electrons is proportional to the voltage across the anode-cathode:

2.5.4 The de Broglie Wavelength

The de Broglie Wavelength

- Using ideas based upon the quantum theory and Einstein's theory of relativity, de Broglie suggested that the momentum (p) of a particle and its associated wavelength (λ) are related by the equation:

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

- Since momentum p = mv, the de Broglie wavelength can be related to the speed of a moving particle (v) by the equation:

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

- Since kinetic energy $E = ½mv^2$
- Momentum and kinetic energy can be related by:

$$E = \frac{p^2}{2m} \quad \text{or} \quad p = \sqrt{2mE}$$

- Combining this with the de Broglie equation gives a form which relates the de Broglie wavelength of a particle to its kinetic energy:

$$\lambda = \frac{h}{\sqrt{2mE}}$$

- Where:
  - λ = the de Broglie wavelength (m)
  - h = Planck's constant (J s)
  - p = momentum of the particle (kg m s^{-1})
  - E = kinetic energy of the particle (J)
  - m = mass of the particle (kg)
  - v = speed of the particle (m s^{-1})

Worked Example

A proton and an electron are each accelerated from rest through the same potential difference.

Determine the ratio: $\dfrac{\textit{de Broglie wavelength of the proton}}{\textit{de Broglie wavelength of the electron}}$

- Mass of a proton = 1.67×10^{-27} kg
- Mass of an electron = 9.11×10^{-31} kg

Step 1: Consider how the proton and electron can be related via their masses

The proton and electron are accelerated through the same p.d., therefore, they both have the same kinetic energy

Step 2: Write the equation relating the de Broglie wavelength of a particle to its kinetic energy

$$\lambda = \frac{h}{p} = \frac{h}{\sqrt{2mE}}$$

$$\lambda \propto \frac{1}{\sqrt{m}}$$

Step 3: Calculate the ratio

$$\frac{\text{de Broglie wavelength of the proton}}{\text{de Broglie wavelength of the electron}} = \frac{1}{\sqrt{m_p}} \div \frac{1}{\sqrt{m_e}}$$

$$\sqrt{\frac{m_e}{m_p}} = \sqrt{\frac{9.11 \times 10^{-31}}{1.67 \times 10^{-27}}} = 2.3 \times 10^{-2}$$

This means the de Broglie wavelength of the proton is 0.023 times smaller than that of the electron **OR** the de Broglie wavelength of the electron is about 40 times larger than that of the proton

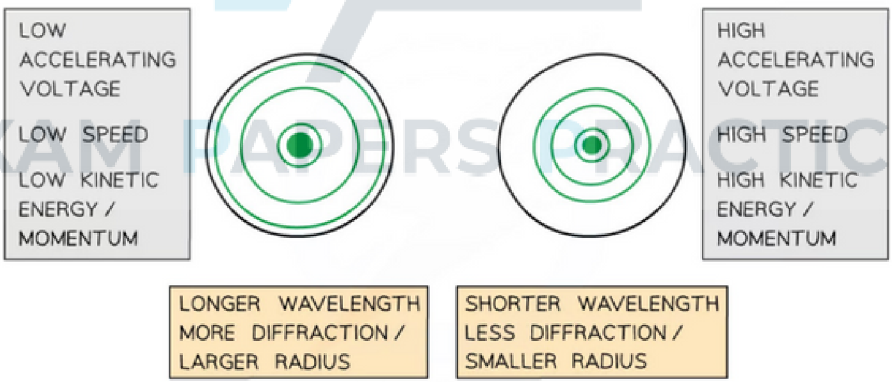
2.5.5 Diffraction Effects of Momentum

Diffraction Effects of Momentum

- When electrons pass through a slit similar in size to themselves, they exhibit a wavelike property(Diffraction), meaning they spread out like a wave passing through a narrow gap
- The regular spacing of atoms in a crystalline solid act as a diffraction grating, scattering the electrons in a predictable manner
- The observed diffraction pattern can be used to deduce the structure of the crystal producing that pattern
- High energy electrons have a **shorter wavelength** and can therefore be used to look at the **size of the nucleus** of an atom as opposed to the arrangement of atoms in a crystal
- The de Broglie wavelength tells us about the wave-particle relationship:

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

- Where:
  - λ = the de Broglie wavelength (m)
  - h = Planck's Constant (J s)
  - m = mass of the electron (kg)
  - v = velocity of the electron (m s^{-1})



Comparison of electron diffraction patterns at different values of momentum

Momentum

- Momentum is equal to $p = mv$, so, from de Broglie's equation:
  - A smaller momentum will result in a **longer** wavelength
  - A larger momentum will result in a **shorter** wavelength

Kinetic Energy

- If the electron speed / kinetic energy is increased, by increasing the accelerating voltage, then:
  - The wavelength of the wave will decrease
  - The diffraction rings will appear closer together
- The higher the kinetic energy of the electron, the higher its momentum hence the smaller its wavelength

Radius of the Diffraction Pattern

- The radius of the diffraction pattern depends on the wavelength:
  - The longer the wavelength, the more the light spreads out hence a **larger radius is produced**
  - The shorter the wavelength, the **smaller the radius produced**
- Therefore, electrons with smaller momentum will produce a more diffuse diffraction pattern

2.5.6 Analysis of the Nature of Matter

Analysis of the Nature of Matter

The Evolving Field of Quantum Mechanics

- The field of quantum mechanics is a relatively new field of research, compared to fields in classical mechanics (Newton's laws, wave theory etc)
- Around 1900, discoveries, such as the electron and the gamma photon, began to conflict with the existing models scientists held about the nature of matter
- Soon after, new theories about the nature of matter began to emerge from **Max Planck**, **Niels Bohr** and **Albert Einstein**, who are seen as the pioneers of **Quantum Theory**
- Unsurprisingly, the understanding of quantum mechanics and the nature of matter has changed over time
  - Scientists use the existing models to make predictions eg. **wave theory**
  - Sometimes the results of their experiments were not as predicted and did not fit with the existing model eg. **the photoelectric effect**
  - Scientists then have to change the model so that it can explain the new evidence eg. **wave–particle duality**

Development of Scientific Theories

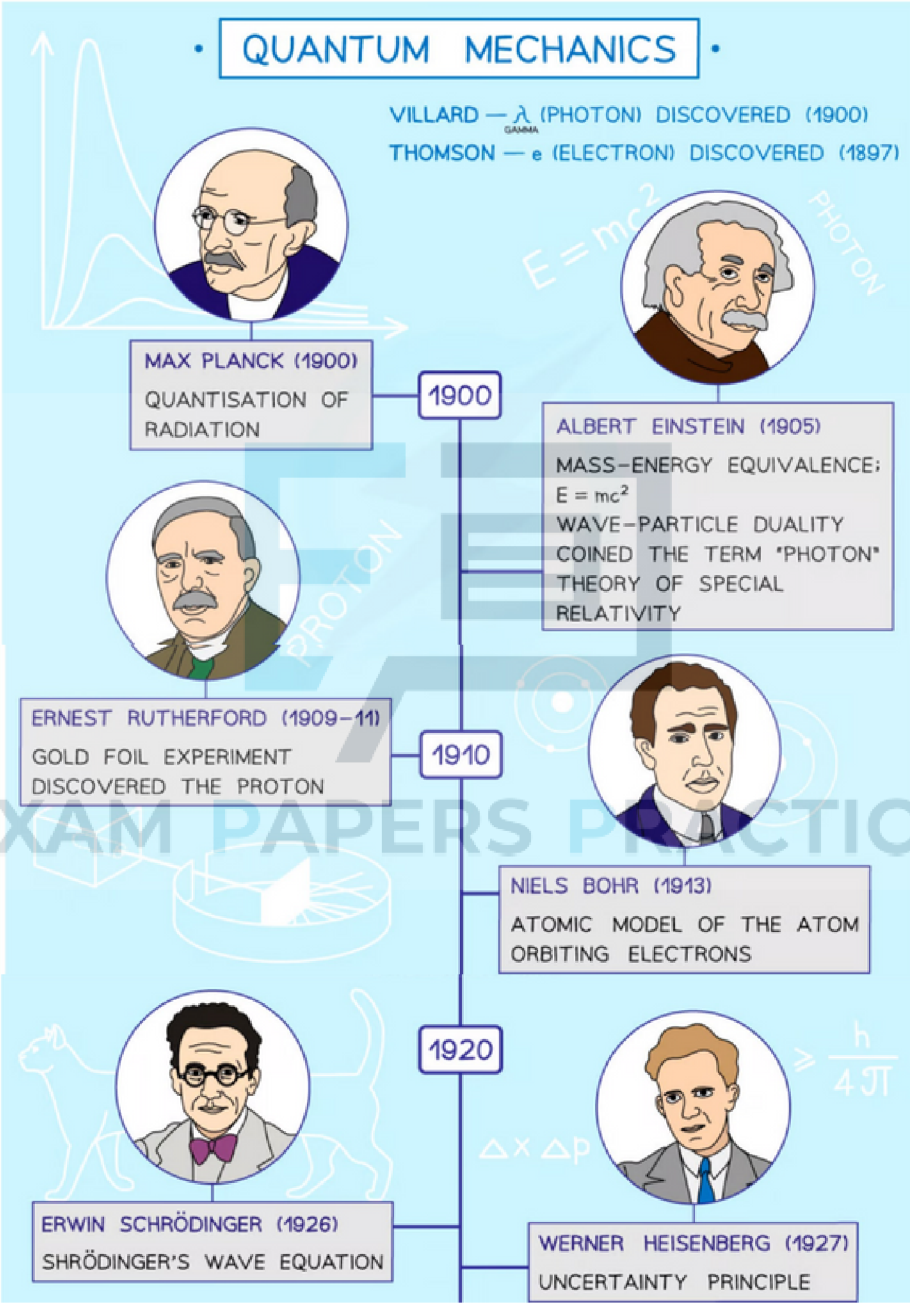
- When a new theory is suggested, such as de Broglie's idea of particles acting as waves, the theory needs to be **evaluated**
- Other scientists evaluate the theory in a process called **peer review**
- When enough evidence is found to support the theory, it is **validated**
- De Broglie's theory is currently accepted to be correct until any conflicting evidence is found

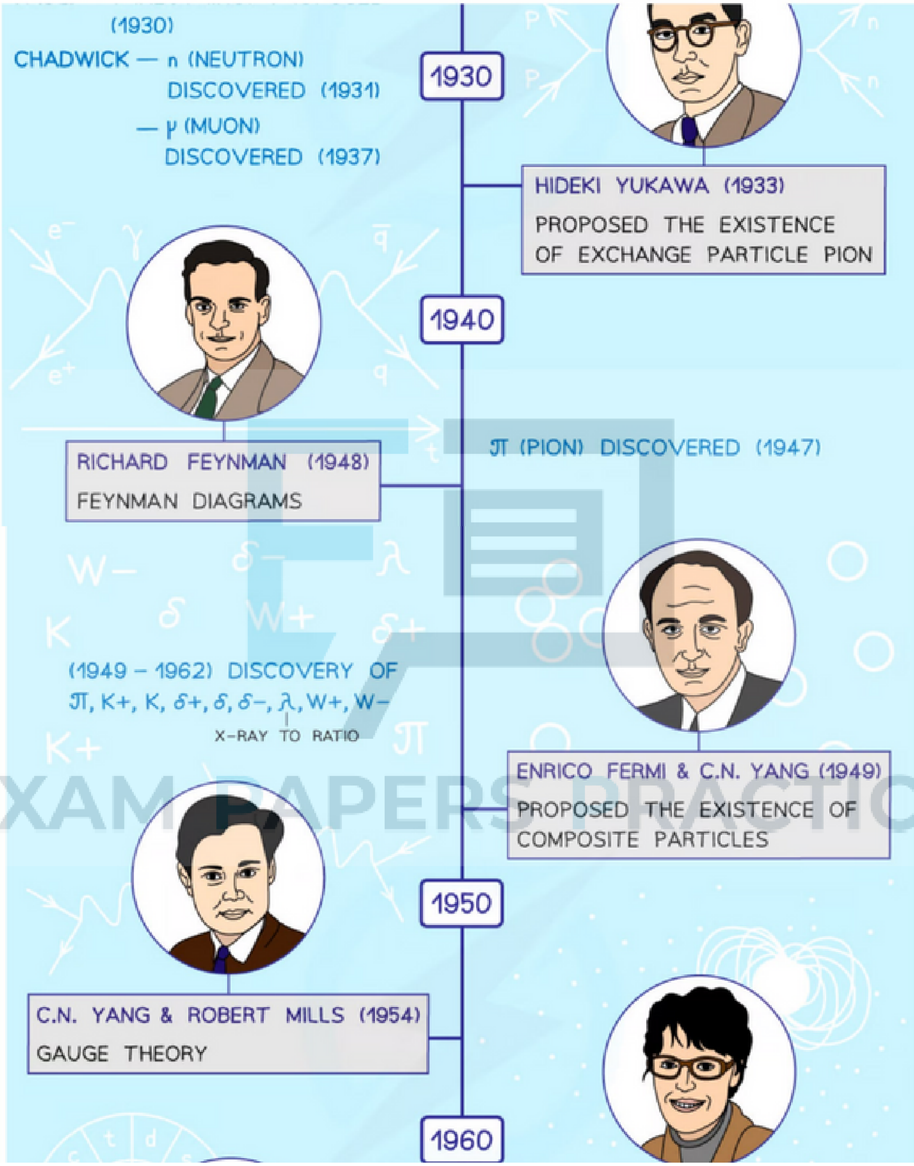
Peer Review

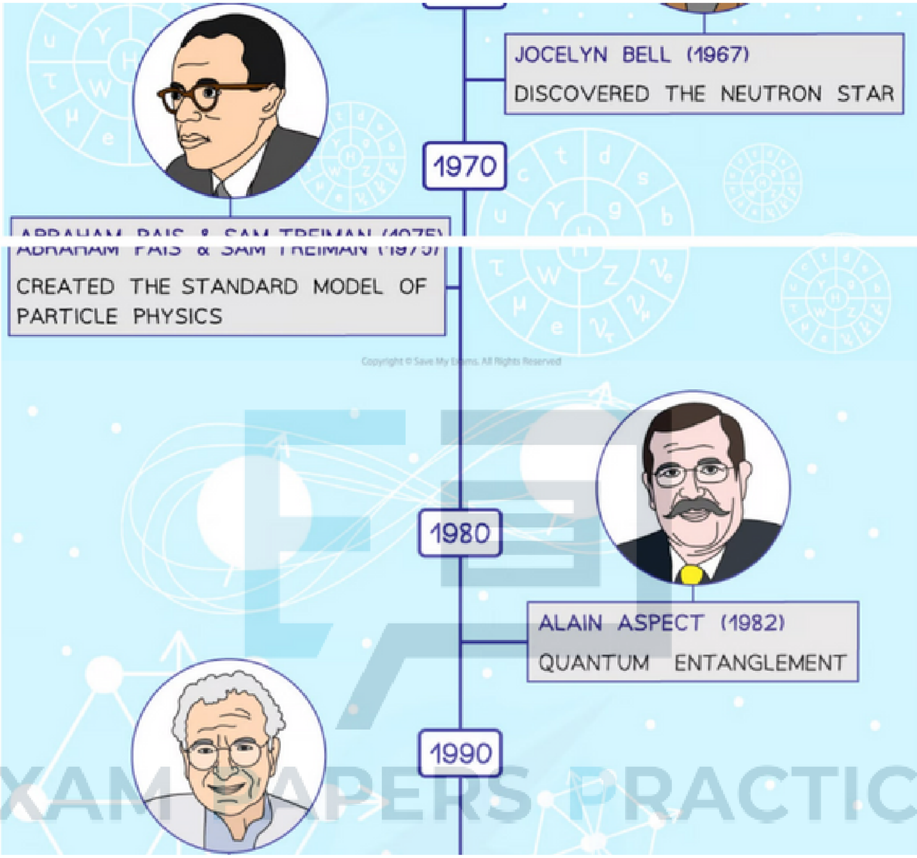
- Scientists across the world form a **scientific community**
- Scientific claims from new research that is published in journals must be peer-reviewed or evaluated by other scientists who are experts in that field of science
- During the peer review process, scientists must check:
  - **Validity** – does the research achieve what it says it does? Is the method appropriate and how have any errors been addressed?
  - **Originality** – are the results new or has anyone else already carried out similar research? If so, has their work been credited?
  - **Significance** – are the findings of the research important or ground-breaking?

Evaluating Scientific Claims

- Claims that are from research that is not peer-reviewed should always be questioned
- There could be a multitude of problems. These could be the method used, the accuracy of results or the conclusions drawn
- Research should also be checked if it is:
  - **Repeatable** by the same scientists who carried it out
  - **Reproducible** by other teams of scientists


QUANTUM MECHANICS
VILLARD — λ (PHOTON) DISCOVERED (1900)
GAMMA
THOMSON — e (ELECTRON) DISCOVERED (1897)
E = mc²
PHOTON
MAX PLANCK (1900)
QUANTISATION OF RADIATION
1900
ALBERT EINSTEIN (1905)
MASS–ENERGY EQUIVALENCE; E = mc²
WAVE–PARTICLE DUALITY
COINED THE TERM "PHOTON"
THEORY OF SPECIAL RELATIVITY
PROTON
ERNEST RUTHERFORD (1909–11)
GOLD FOIL EXPERIMENT
DISCOVERED THE PROTON
1910
NIELS BOHR (1913)
ATOMIC MODEL OF THE ATOM
ORBITING ELECTRONS
1920
Δx Δp ≥ h/4π
ERWIN SCHRÖDINGER (1926)
SHRÖDINGER'S WAVE EQUATION
WERNER HEISENBERG (1927)
UNCERTAINTY PRINCIPLE


(1930)
CHADWICK — n (NEUTRON) DISCOVERED (1931)
— μ (MUON) DISCOVERED (1937)
1930
HIDEKI YUKAWA (1933)
PROPOSED THE EXISTENCE OF EXCHANGE PARTICLE PION
1940
RICHARD FEYNMAN (1948)
FEYNMAN DIAGRAMS
π (PION) DISCOVERED (1947)
(1949 – 1962) DISCOVERY OF
π, K+, K, δ+, δ, δ−, λ, W+, W−
X-RAY TO RATIO
ENRICO FERMI & C.N. YANG (1949)
PROPOSED THE EXISTENCE OF COMPOSITE PARTICLES
1950
C.N. YANG & ROBERT MILLS (1954)
GAUGE THEORY
1960


JOCELYN BELL (1967)
DISCOVERED THE NEUTRON STAR
1970
ABRAHAM PAIS & SAM TREIMAN (1975)
CREATED THE STANDARD MODEL OF PARTICLE PHYSICS
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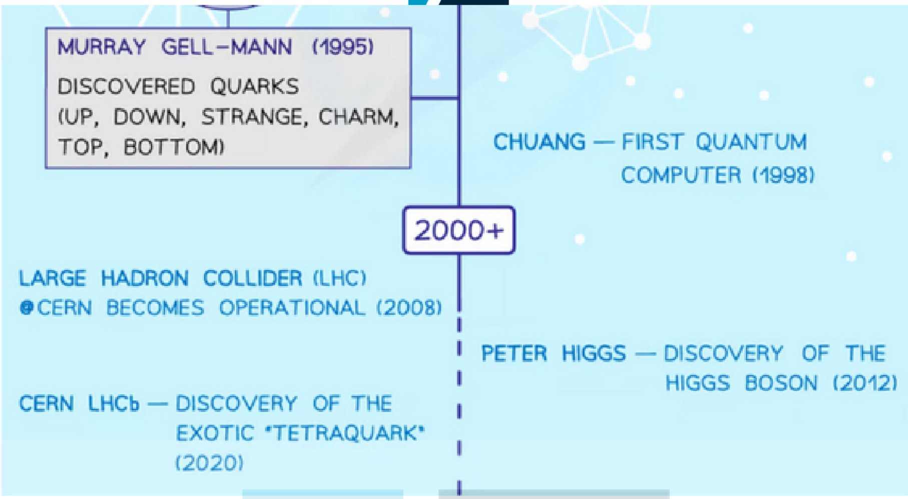
1980
ALAIN ASPECT (1982)
QUANTUM ENTANGLEMENT
1990



Timeline of the great advancements in quantum theory since 1900