UNIT 3 - RADIATION and MATTER

OPTO-ELECTRONICS

3) EMISSION and ABSORPTION SPECTRA

You must be able to:

- Use the following terms correctly: photon, work function, photoelectric emission (photoemission) and threshold frequency.

- Describe experiments to display emission and absorption spectra.

- Use the following terms correctly - ground state, excited state, ionisation level, electron transition, emission spectrum and absorption spectrum.

- State that electrons in free atoms occupy discrete energy levels.

- Draw a diagram to represent the energy levels for a hydrogen atom.

- Describe spontaneous emission of radiation as a random process analogous to that of radioactive decay.

- Explain an emission line spectrum in terms of electron transitions from higher to lower energy levels and explain why some emission lines are brighter than others.

- State that a photon of light emitted from an atom has an energy (hf) equal to the difference in energy (ΔE) between the 2 energy levels involved in the electron transition responsible for creating the photon. $\Delta E = h f$.

- Explain an absorption line spectrum in terms of electrons in lower energy levels absorbing photons of correct energy (hf) and making transitions of energy difference (ΔE) to higher energy levels.

- Solve problems involving the equation $\Delta E = h f$.

- Explain the presence of absorption lines in the spectrum of sunlight.

RUTHERFORD-BOHR MODEL OF THE ATOM

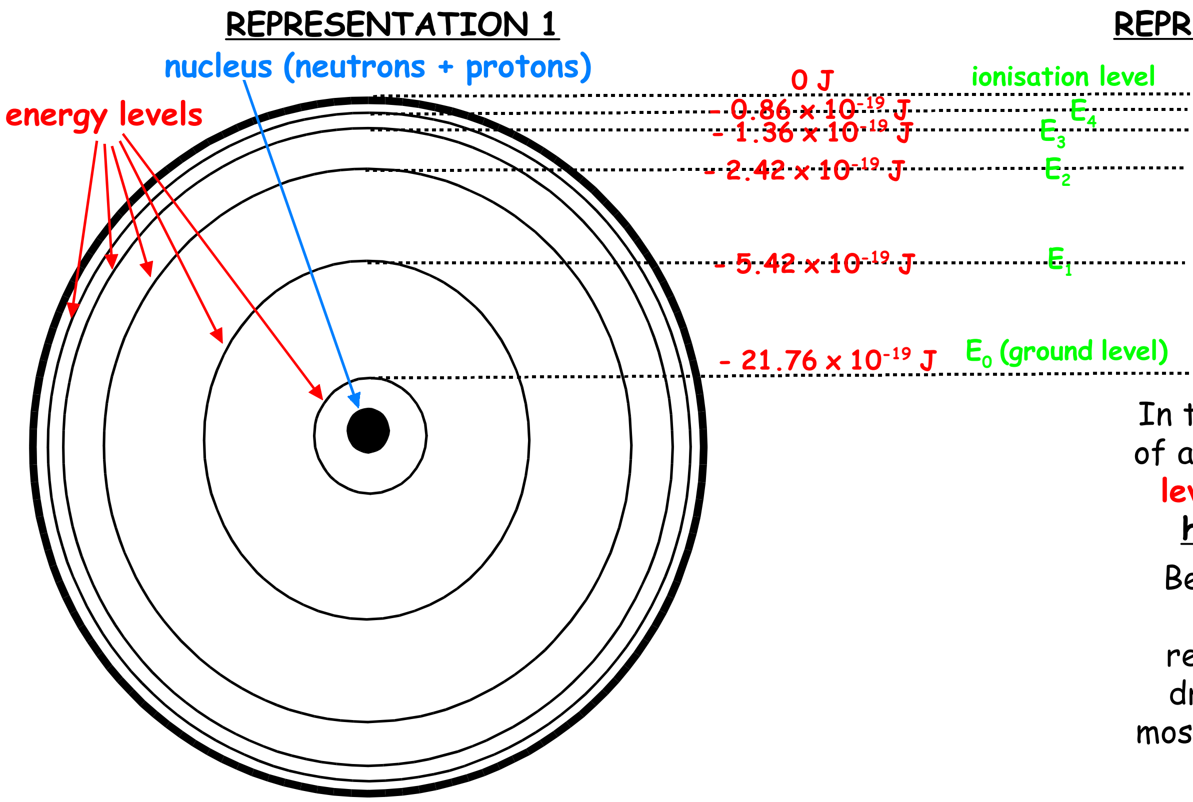
Free (unreacted) atoms consist of a tiny, central **nucleus** (containing particles called **neutrons** and **protons**) surrounded by particles called **electrons**.

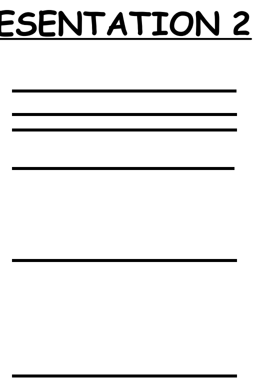
The **electrons** circle around the **nucleus** at fixed distances from it. Electrons at each distance have a **fixed energy value** - so each distance is known as an **energy level**.

Electrons can move from one **energy level** to another **energy level**, but cannot stop **between** the **energy levels**.

As an **electron** gets closer to the **nucleus**, the **electron** loses energy - so the **energy levels** closer to the nucleus have **more negative** energy values.

two representations of some of the energy levels in a hydrogen atom





In this representation of an atom, the **energy levels** are shown by **horizontal lines**.

Because this is the easier of the 2 representations to draw, it is the one most commonly used in tests/exams.

A hydrogen atom has only 1 electron, but this can move to any of the possible energy levels.

The **energy level** closest to the **nucleus** (the level with lowest energy) is called the **ground level** (E_o) - An **electron** in this **energy level** is said to be in its **ground state**.

The **energy levels** further from the **nucleus** (E_1, E_2, E_3, etc) are called **excited energy levels** - An **electron** in any of these **energy levels** is said to be in an **excited state**.

An **electron** can reach a distance so far away from the **nucleus** that the **electron** can escape from the atom - We say the **electron** has reached the **ionisation level** (where it has **0 Joules of energy**). When this happens, the **atom** is said to be in an **ionisation state**.

ATOMIC SPECTRA

Under certain circumstances, free (unreacted) atoms can **give out** (**emit**) or **take in** (**absorb**) **photons** of **electromagnetic energy**, including **photons** of **different coloured light**.
REMEMBER - The **colour** of light depends on its **frequency**/**wavelength**.

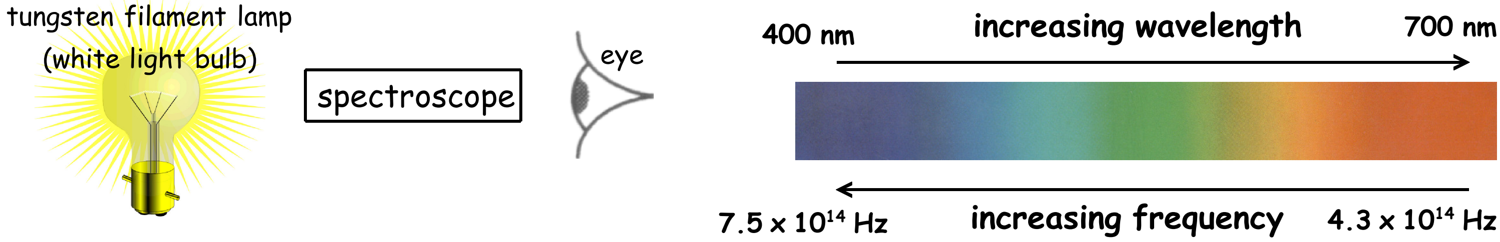
When the light is passed through a **prism**, **diffraction grating** or **spectroscope**, an **atomic spectrum** is produced.

Different atoms produce different atomic spectra (e.g., mercury atoms produce a different **spectrum** from sodium atoms.) As a result, an atom can be identified by observing its **spectrum**.

1) EMISSION SPECTRA

(a) Continuous Spectra

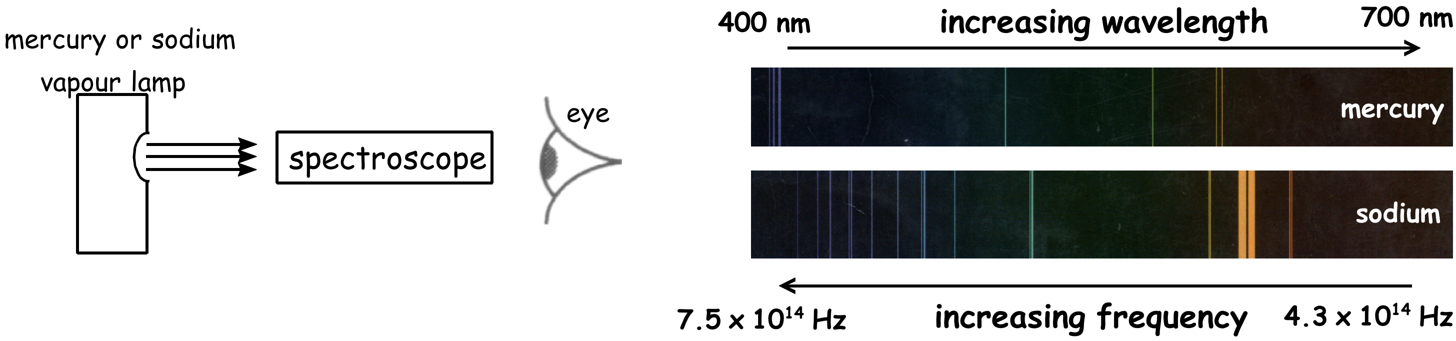
A tungsten filament lamp (a normal light bulb) emits **white light**. When the **white light** is passed through a spectroscope, a **continuous spectrum** is obtained. This contains all 7 colours of the **visible spectrum:**



(b) Line Spectra

A **mercury vapour lamp** or **sodium vapour lamp** emits different photons of **specific frequency/wavelength** (and hence **colour**). When the light is passed through a spectroscope, a series of **different coloured lines** on a **black background** is obtained. Each line occupies an exact position corresponding to its **exact frequency/wavelength**.

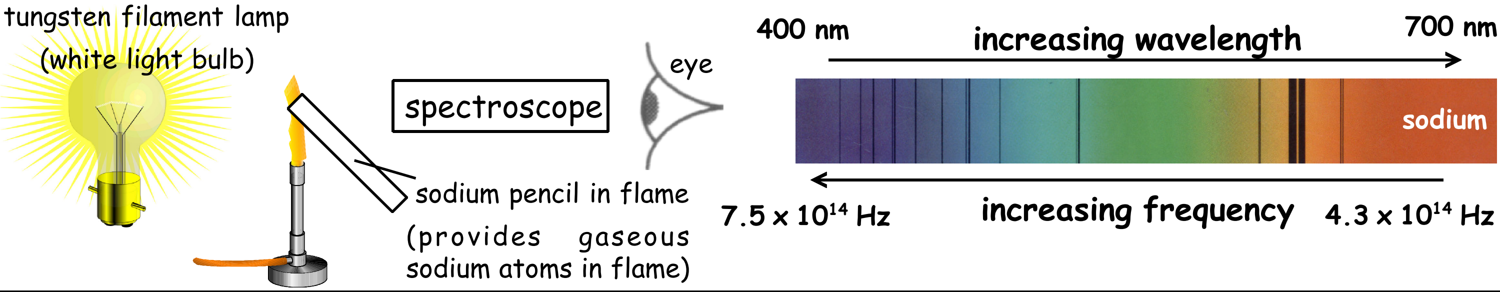
Notice the different colours and positions of the emission lines for mercury and sodium.



2) ABSORPTION SPECTRA

When **white light** (containing photons of all 7 different colours of the visible spectrum) is passed through atoms of an element like sodium which are in the **gaseous state**, the **gaseous atoms** absorb photons from the **white light** of **specific frequency/wavelength** (and hence **colour**). When the light is passed through a spectroscope, a **continuous spectrum** with a **series of black absorption lines** is obtained. Each **black absorption line** occupies an exact position corresponding to the exact **frequency/wavelength** of the photons from the white light that have been **absorbed** by the gaseous atoms.

[COMPARE THE LINE **EMISSION** AND **ABSORPTION** SPECTRA OF **SODIUM**].



HOW EMISSION LINE SPECTRA ARE CREATED

At **any time**, an **electron** in an **excited** (**higher**) **energy level** of an atom can make a **transition** (**jump**) to a **less excited** (**lower**) **energy level** in the same atom (including the **ground level**, E_o). This process is **random** - We cannot predict when it will happen (just like we cannot predict when the radioactive decay of an atomic nucleus will take place.)

When an **electron** makes such a **transition** (**jump**), **one photon of electromagnetic energy is emitted from the atom**. **The energy of this photon is exactly equal to the difference in energy between the 2 energy levels involved.**

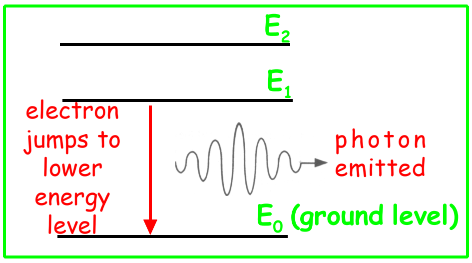
This equation applies:

$$\Delta E = h\,f$$

difference in energy between the 2 energy levels involved in electron transition

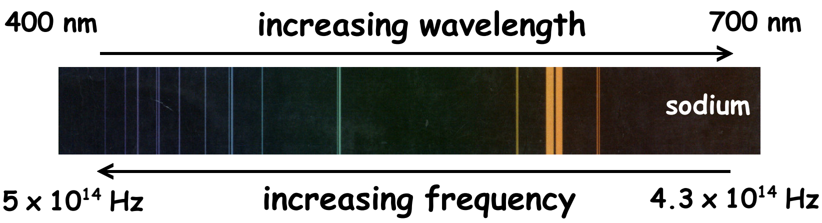
Planck's constant = 6.63×10^{-34} J s

frequency of photon emitted from atom



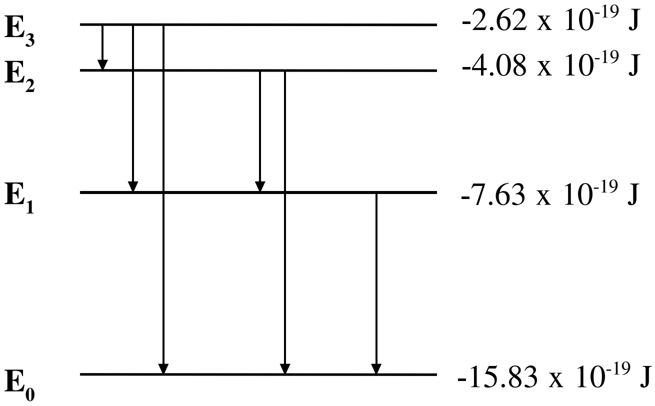
The emitted photon often has a frequency within the visible spectrum, so produces a coloured emission line in the atom's emission line spectrum. The photon may also have a frequency outwith the visible spectrum - in the infra-red or ultra-violet.

Various such electron transitions (jumps) of **different energy** (and hence different **frequency/wavelength**) are possible - so an emission line spectrum may consist of several emission lines of different **frequency/wavelength**, e.g., the **sodium line emission spectrum** shown below:



For example:

Atom **X** has 4 possible energy levels, as shown:



There are **6** possible downward electron transitions (jumps) - as shown by the **6** downward arrows. Each happens without outside influence - they are **spontaneous**.

Each downward electron transition (jump) will produce **one** emission line in the atom's emission spectrum (one photon being emitted per jump) - so the spectrum will have **6** emission lines.

The position of each emission line on the emission spectrum will depend on the **frequency/wavelength** of each emitted photon, which depends on the **difference in energy** (**ΔE**) between the 2 energy levels involved in the electron transition.

PROBLEM

Calculate the **energy** and **frequency** of the photon emitted when an electron jumps from energy level $\mathbf{E_2}$ to energy level $\mathbf{E_1}$.

$$\Delta E = (7.63 \times 10^{-19})\text{J} - (4.08 \times 10^{-19})\text{J}$$
$$= (3.55 \times 10^{-19})\text{J}$$

So, energy of emitted photon $= (3.55 \times 10^{-19})$J

Since we are calculating the change in energy, there is no need to use the - signs in front of the numbers.

$$f = \Delta E/h$$
$$= (3.55 \times 10^{-19})\text{J}/(6.63 \times 10^{-34})\text{J s}$$
$$= 5.35 \times 10^{14}\ \text{Hz}$$

Brightness of Emission Lines

Emission spectra are usually obtained by observing a **vapour lamp** through a **spectroscope**. The **vapour lamp** contains millions of atoms, each giving out photons - so many photons are emitted.

Some emission lines in an emission spectrum are brighter than others (see the 2 very bright orange lines in the sodium emission spectrum) - The brighter lines are caused by a larger number of electrons (from the same and other identical atoms) making the same energy transition (jump) - so more photons of light with the same frequency/wavelength are produced.

1) Explain, with the help of a labelled diagram, how an atom can emit a photon of electromagnetic energy:

2) All the possible energy levels of atom A are shown:

E_4	-2.50×10^{-19} J
E_3	-2.75×10^{-19} J
E_2	-3.25×10^{-19} J
E_1	-4.60×10^{-19} J
E_0	-6.95×10^{-19} J

(a) Show, using downward arrows, all the possible downward electron transitions

(b) All the photons emitted from the atom can be detected easily.

How many emission lines will be present in the atom's emission spectrum?__________

(c) Why will some emission lines be brighter than others?____________________________

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(d) Explain whether we can predict the exact moment when an electron will make a downward transition:____

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(e) Calculate the energy and frequency of the photon emitted from the atom when an electron makes a transition from energy level E_4 to E_0:

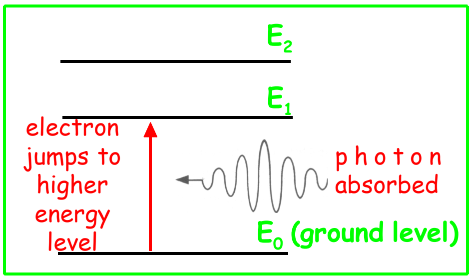
HOW ABSORPTION LINE SPECTRA ARE CREATED

An **atom** can **absorb** a photon of **electromagnetic energy**. **The atom can only do so if the energy of the photon is exactly equal to the difference in energy (ΔE) between any 2 energy levels in the atom.**

When a **photon** is absorbed, **one electron** makes a **transition** (**jump**) between **the 2 energy levels with exact energy difference ΔE**, from the **less excited** (**lower**) **energy level** to the **more excited** (**higher**) **energy level**.

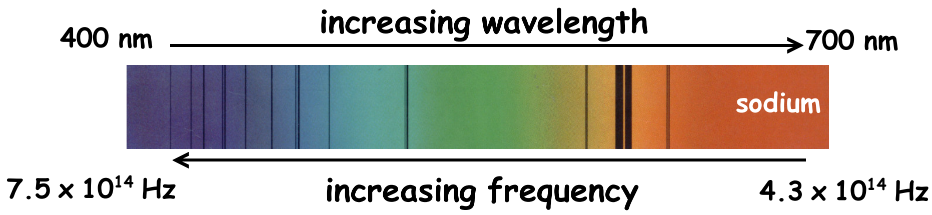
This equation applies: $\Delta E = h\,f$

- ΔE: difference in energy between the 2 energy levels involved in electron transition
- f: frequency of photon absorbed by atom
- h: Planck's constant = 6.63×10^{-34} J s



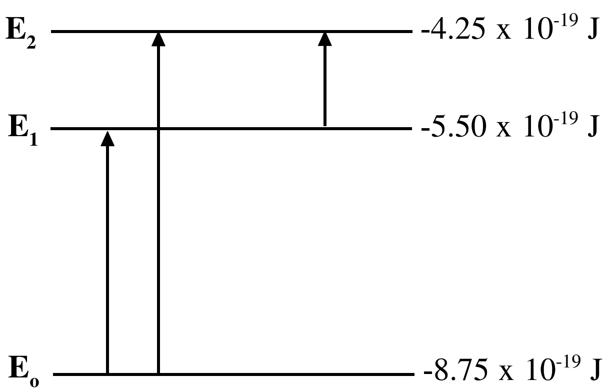
Various such electron transitions (jumps) of **different energy** are possible, provided **photons of suitable energy** are present to be **absorbed**.

The **absorbed photons** are **removed** from the incident electromagnetic radiation, so **black absorption lines** are produced on the atom's absorption line spectrum against a **coloured visible spectrum background** where **no photons** are being absorbed, e.g., the sodium line absorption spectrum shown below:



For example:

Atom **Y** has 3 possible energy levels, as shown:



There are **3** possible upward electron transitions (jumps) - as shown by the **3** upward arrows.

Each upward electron transition (jump) will produce **one** absorption line in the atom's absorption spectrum (one photon being absorbed per jump) - so the spectrum will have **3** absorption lines.

The position of each absorption line on the absorption spectrum will depend on the **frequency/wavelength** of each absorbed photon, which depends on the **difference in energy** (**ΔE**) between the 2 energy levels involved in the electron transition.

PROBLEM

Calculate the **energy** and **frequency** of the photon **absorbed** when an electron jumps from energy level $\mathbf{E_0}$ to energy level $\mathbf{E_2}$.

$$\Delta E = (8.75 \times 10^{-19})\text{J} - (4.25 \times 10^{-19})\text{J}$$
$$= (4.50 \times 10^{-19})\text{J}$$

So, energy of absorbed photon $= (4.50 \times 10^{-19})$J

Since we are calculating the **change in energy**, there is no need to use the - signs in front of the numbers.

$$f = \Delta E/h$$
$$= (4.50 \times 10^{-19})\text{J}/(6.63 \times 10^{-34})\text{J s}$$
$$= 6.79 \times 10^{14}\ \text{Hz}$$

Fraunhofer Lines - Absorption Lines in Sunlight

YOU MUST NEVER OBSERVE SUNLIGHT DIRECTLY

When **sunlight** is passed through a **spectroscope**, **black absorption lines** are observed in its **visible spectrum**.

These **absorption lines** are due to **photons** of **certain energies** from the sun's **hot core** being absorbed by **gaseous atoms** in the sun's **cooler outer layer**.

The **absorption lines** correspond to those produced by **hydrogen**, **helium**, **sodium** and **other atoms** - **So these must be present in the sun's atmosphere**.

1) Explain, with the help of a labelled diagram, how an atom can absorb a photon of electromagnetic energy:

2) All the possible energy levels of an atom are shown:

E_3 ———————— -2.56×10^{-19} J

E_2 ———————— -5.92×10^{-19} J

E_1 ———————— -8.80×10^{-19} J

E_0 ———————— -16.64×10^{-19} J

(a) Show, using upward arrows, all the possible upward electron transitions.

(b) All the photons absorbed by the atom can be detected easily.

How many absorption lines will be present in the atom's absorption spectrum?________

(c) An electron in this atom makes a transition from energy level E_0 to level E_1. What must be the (a) energy and (b) frequency of the photon absorbed?

(d) A photon of energy 6.24×10^{-19} J is absorbed by this atom, causing one electron to make a transition from a lower to a higher energy level.

(i) Between which 2 energy levels does the electron jump? __________________

(ii) Calculate the absorbed photon's frequency and wavelength.

3) Explain why the visible spectrum of sunlight contains black absorption lines.

4) What name is given to these absorption lines in the sun's visible spectrum?

5) What do these absorption lines tell us about the sun's atmosphere?

Notes