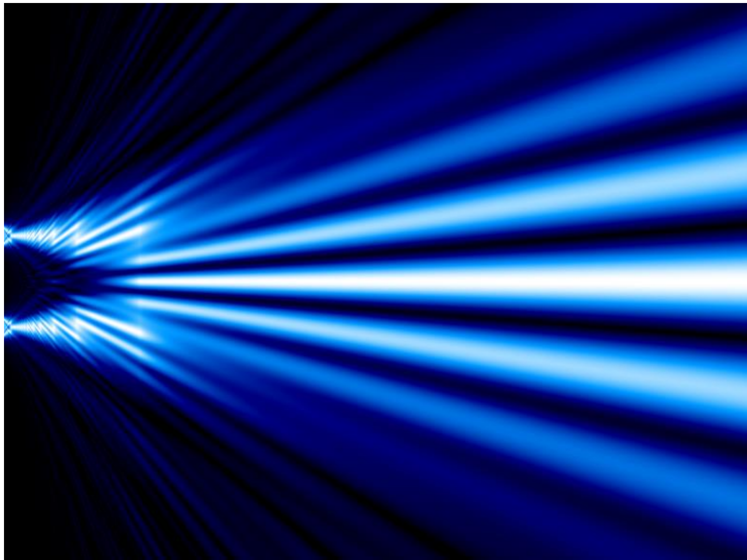
Waves & The Particle Nature of Light

2015 EdExcel A Level Physics
Topic 5

The Particle Nature of Light



Physics Tutor Online

Particles or Waves?

I won the Nobel Prize for discovering the electron in 1897.

This experiment shows what happens when electrons are aimed at a graphite crystal:

What’s this? Is that a diffraction pattern I see? What does this mean?

Electron Diffraction

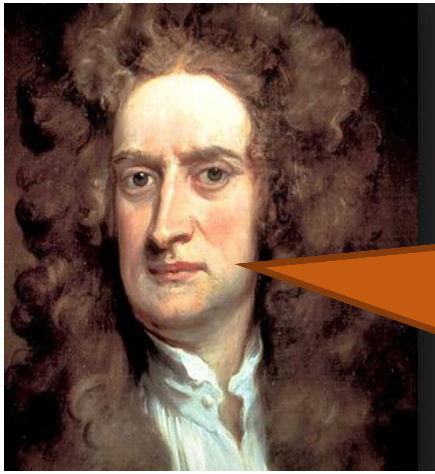
Electron diffraction patterns are seen when electrons are passed through graphite crystal. Diffraction is seen because the distance between the atoms is of the same order as the de Broglie wavelength of the electrons.

de Broglie wavelength $\lambda = \frac{h}{p}$

1) What is the de Broglie wavelength of electrons travelling at around $2x10^7ms^{-1}$ (electron mass = $9.1x10^{-31}kg$)?

2) What would happen to the diffraction pattern if the voltage to the electrons (and therefore their speed) was increased?

Particle Theory and Wave Theory



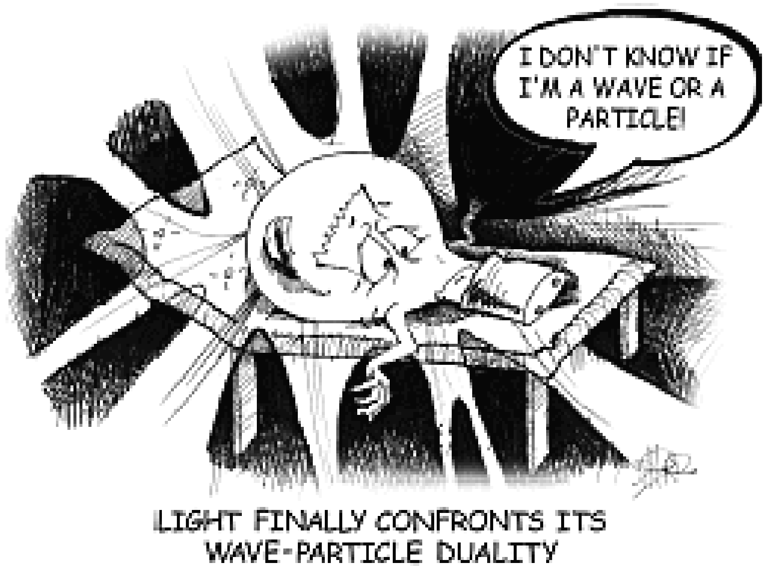
Light can be reflected, refracted and diffracted. These three things are called "wave behaviour" so light must travel as a wave. It makes sense!

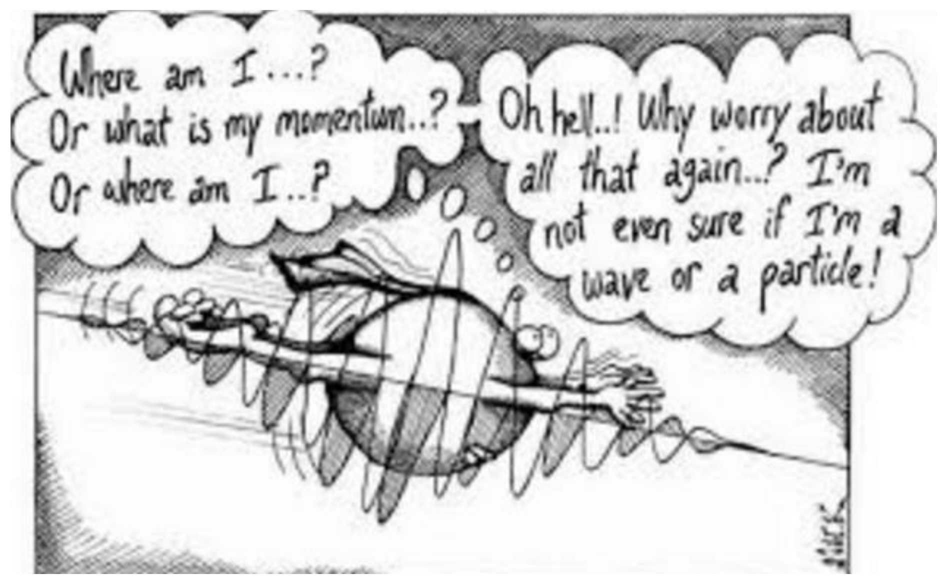
Isaac Newton

Ah yes, but light and other EM radiations also demonstrate some "particle behaviour" so many scientists say that it travels as a set of particles called photons. Here's the evidence…

Max Planck

Particle Theory and Wave Theory cont.





Introduction to Wave Particle Duality

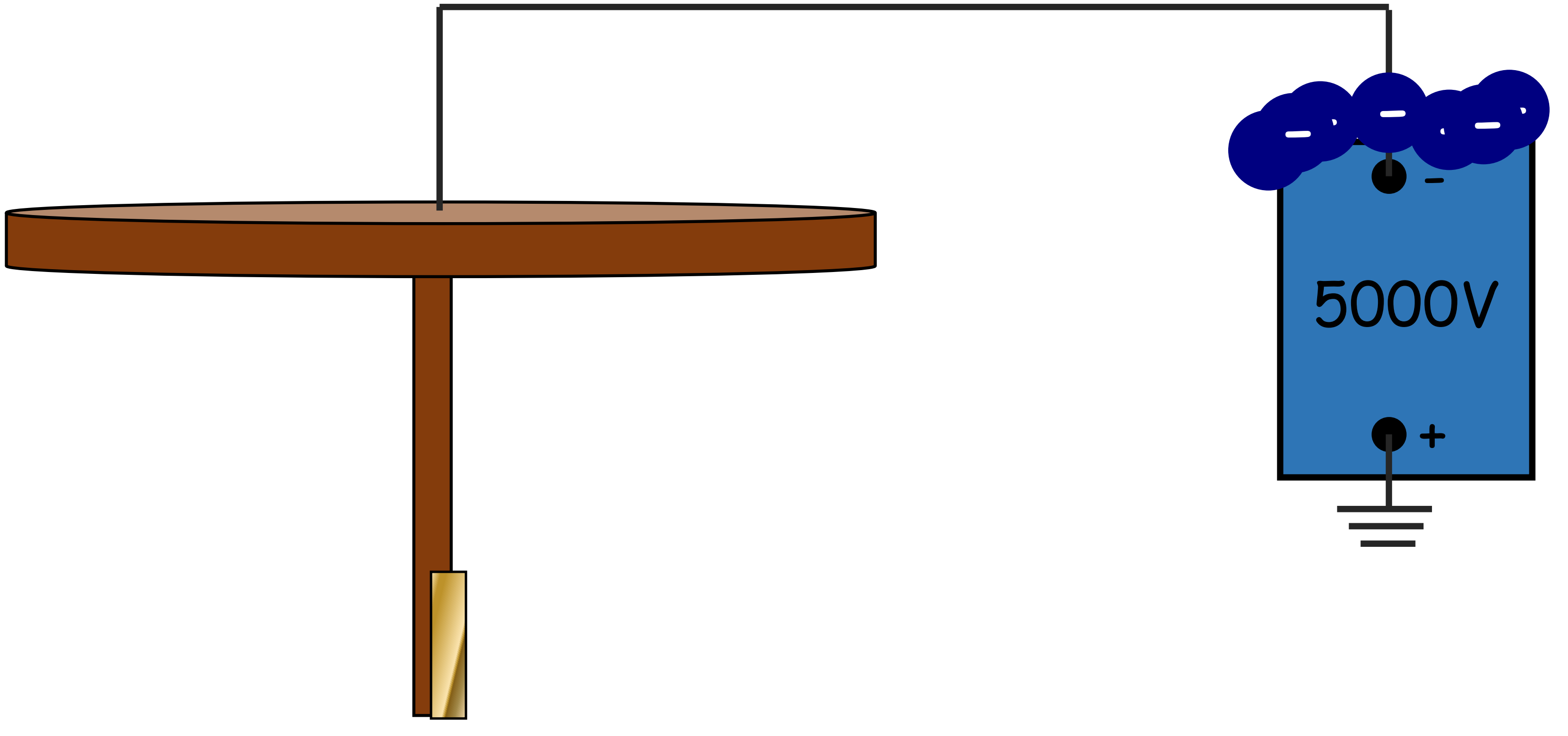
Some basic principles:

1) The wavelength of blue light is around 400nm ($4x10^{-7}m$)
2) The wavelength of red light is around 650nm ($6.5x10^{-7}m$)
3) Therefore blue light is higher frequency than red light
4) Light is treated as being a wave. Therefore the amount of energy a light wave contains should depend on its intensity or brightness.

Photoelectric Emission

Photoelectric Effect refers to the emission, or ejection, of electrons from the surface of, generally, a metal in response to incident light.

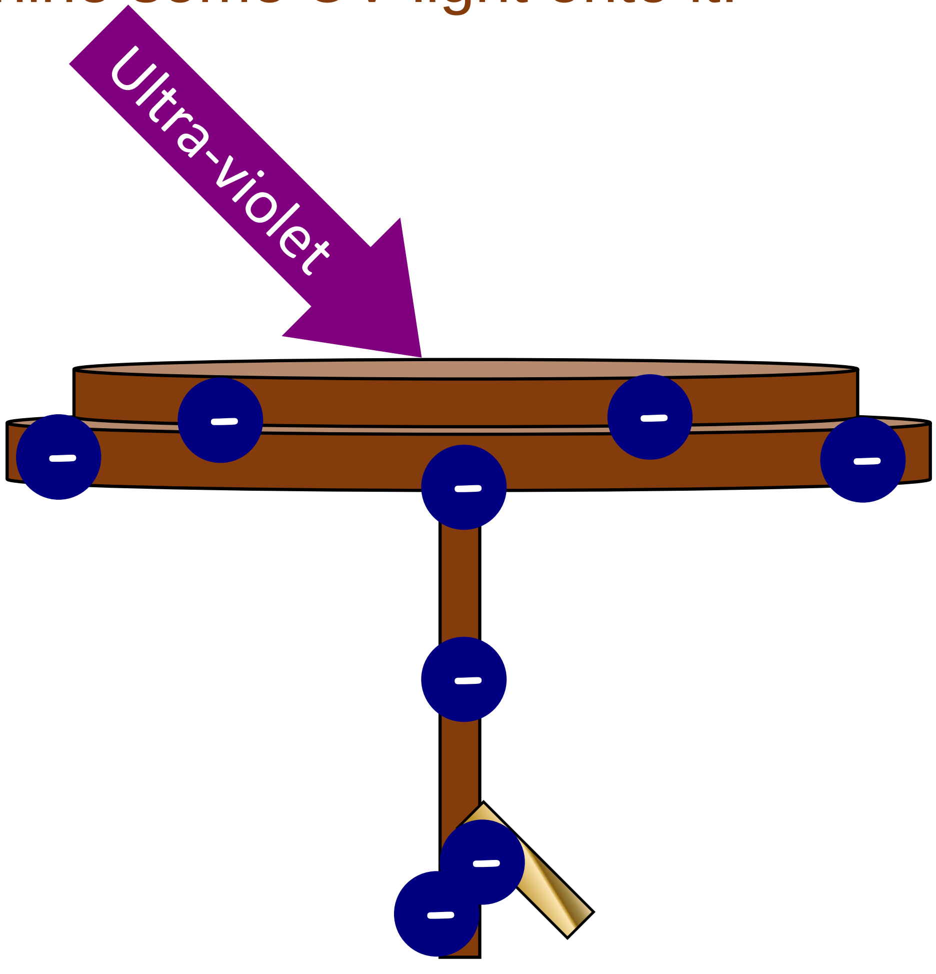
Consider a gold-leaf electroscope…



Photoelectric Emission

Let's put a piece of zinc on top:

Now shine some UV light onto it:

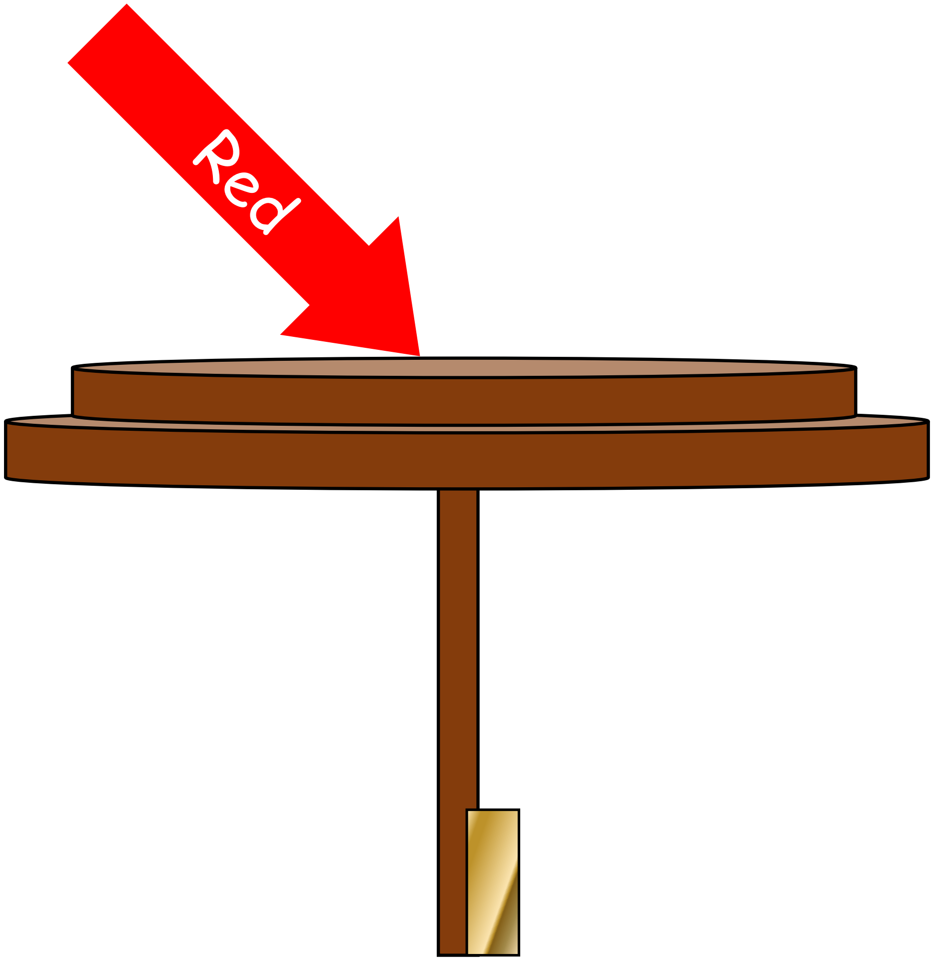


Ultra-violet light is causing the zinc to emit electrons – this is "Photoelectric Emission".

This causes the gold leaf to rise!

Photoelectric Emission cont.

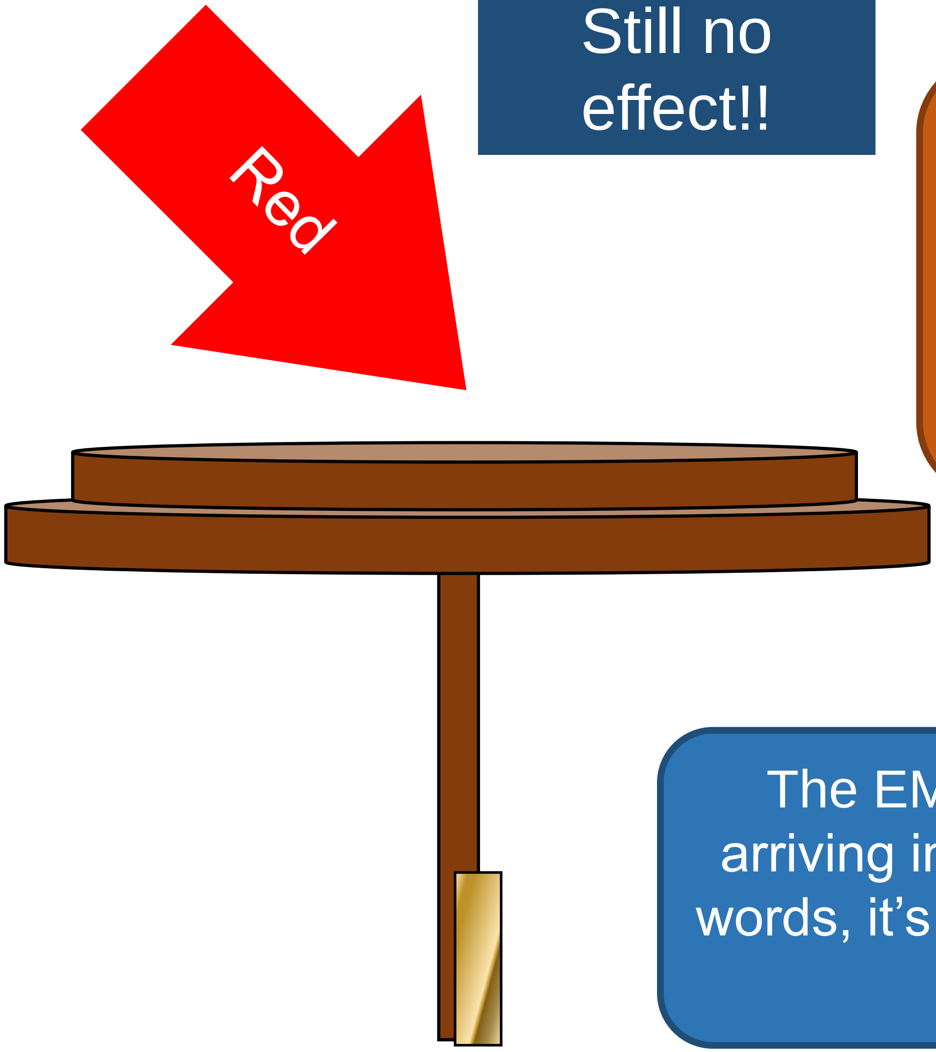
What if red light was used instead?



No effect (although the electrons would eventually be naturally emitted)

Photoelectric Emission cont

What if BRIGHTER red light was used instead?

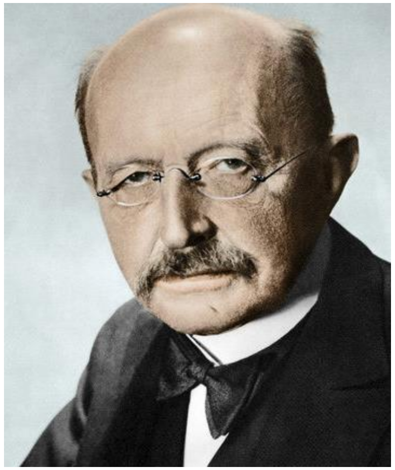


Still no effect!!

Weird. The electrons only had enough energy to leave the surface of the metal when UV light was used. Red light did nothing, even when it was made really bright. What's the solution?

The EM radiation must be arriving in "packets" – in other words, it's behaving like a set of particles!

Some conclusions and definitions



For zinc, this effect is only seen when UV light is used, i.e. when the light has a frequency of 1×10^{15}Hz or higher. This is called the "**Threshold Frequency**" and is generally lower for more reactive metals.

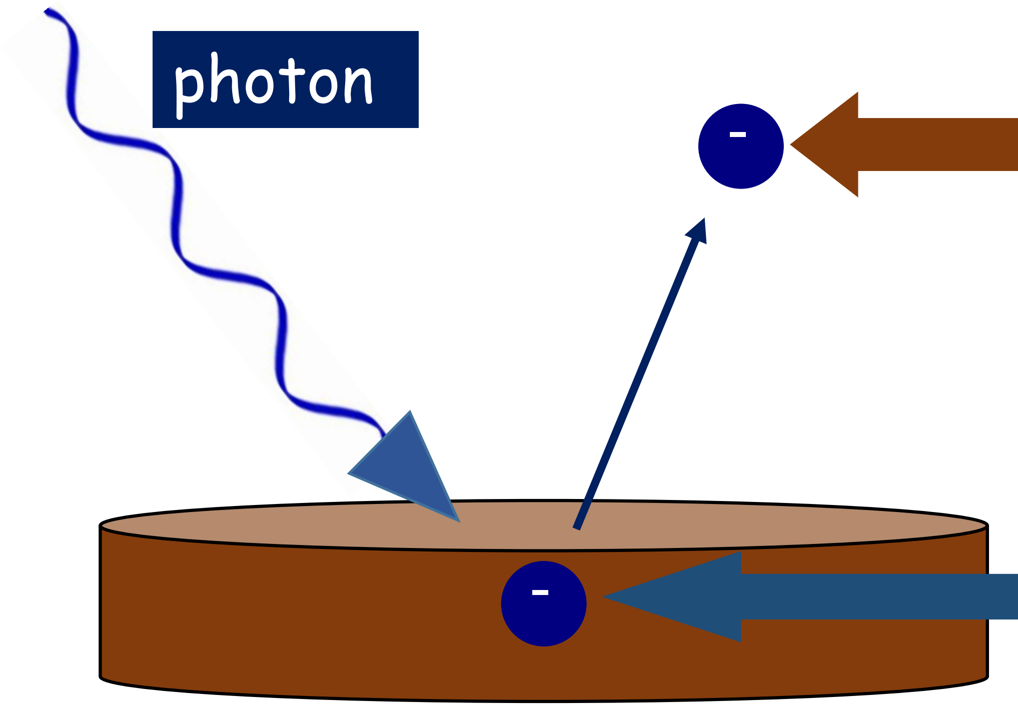
Max Planck (1858-1947) proposed that electromagnetic radiation, like light, comes in small packets. The general name for these packets is "**quanta**".

In the specific case of electromagnetic radiation, a quanta is called a "**photon**" and its energy depends on its frequency, not how bright it is. In other words, photon energy is proportional to frequency:

$E=hf$ where h = Planck's Constant (6.63×10^{-34} Js)

The amount of energy needed to release an electron from a metal is called the "**work function**" and is given the symbol φ. Generally, work functions are lower for more reactive metals.

Photoelectron Energy



…and some energy is given to the electron as kinetic energy.

Some energy is needed to release the electron (the work function φ)…

Photon Energy = work function + kinetic energy of electron

Calculating Photon Energy

Remember that the energy of a photon is proportional to its frequency, so E=hf, where h = Planck's Constant = 6.63×10^{-34}Js.

On the previous slide we said that…

Photon energy = work function + kinetic energy of electron

Therefore this equation becomes:

$$hf = \varphi + 1/2mv^2$$

The Electronvolt

Recall the equation W=QV…

This equation states that the work done on an electron (of charge 1.6×10^{-19}C) as it moves through a potential difference of 1V is given by:

$$W = QV = 1.6\times10^{-19} \times 1 = 1.6\times10^{-19}J$$

This is called "the electronvolt", i.e. the work done on one electron as it moves through 1 volt

$$1\ eV = 1.6\times10^{-19}J$$

Convert the following work functions into electronvolts:

1) Caesium – 3.36×10^{-19}J
2) Sodium – 3.78×10^{-19}J
3) Zinc – 5.81×10^{-19}J

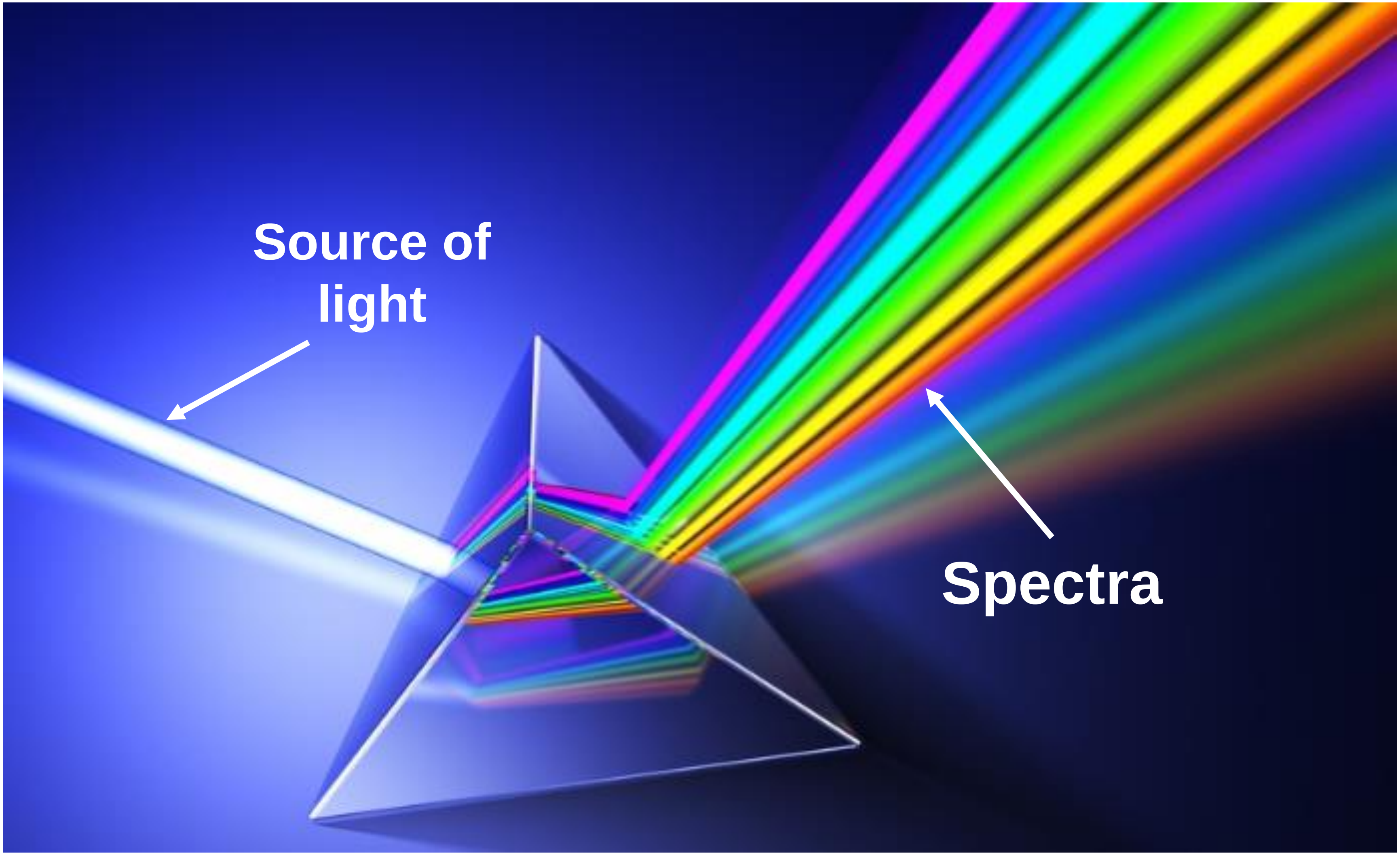
Practice Questions

UV light is shone onto a caesium plate and electrons are emitted from the surface. If the UV light had a frequency of 350nm and the work function of caesium is 1.1eV calculate:

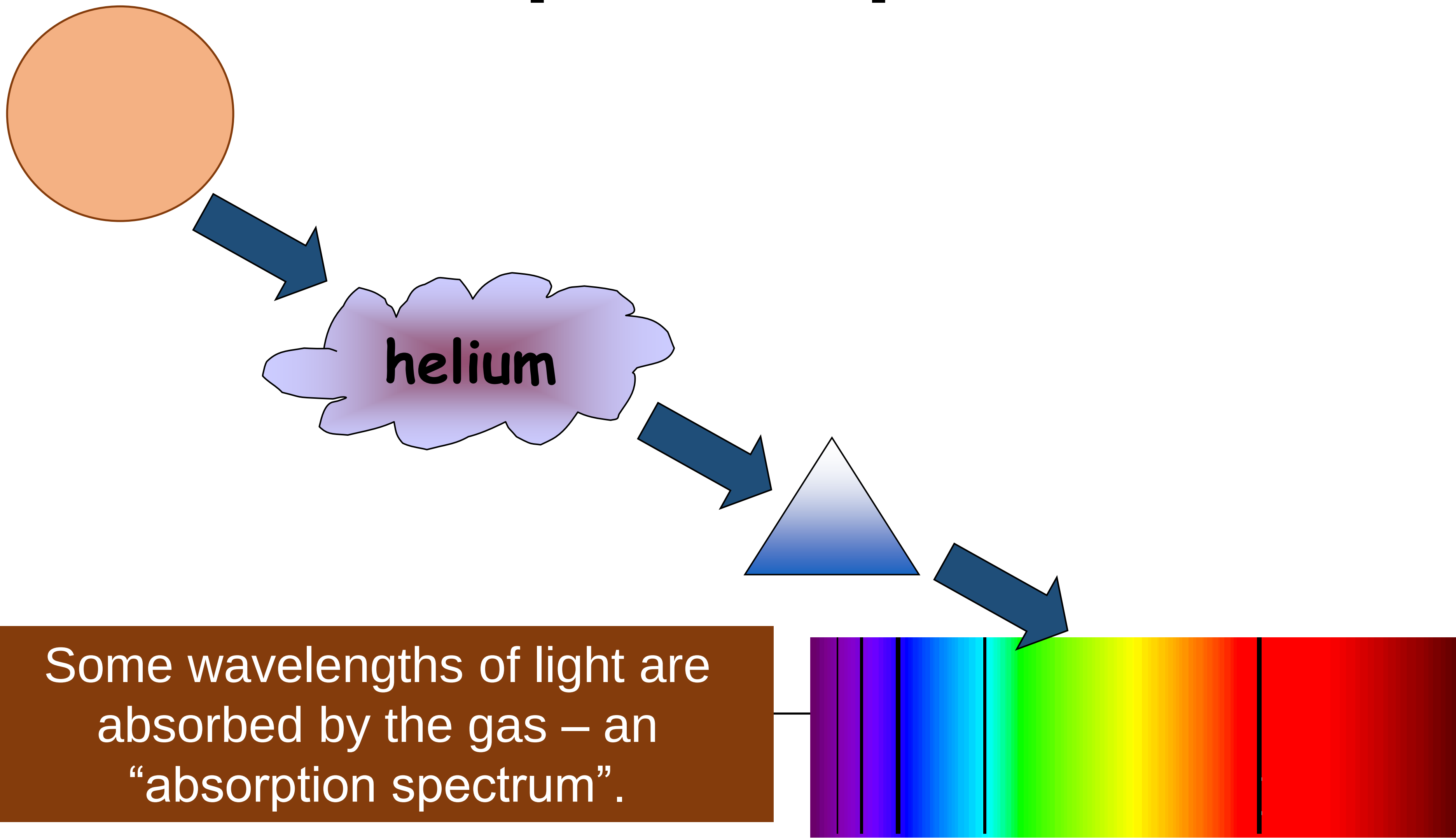
I. The kinetic energy of the electrons emitted in J
II. The kinetic energy in eV
III. The average speed of the electrons

Spectra – introduction

Spectra

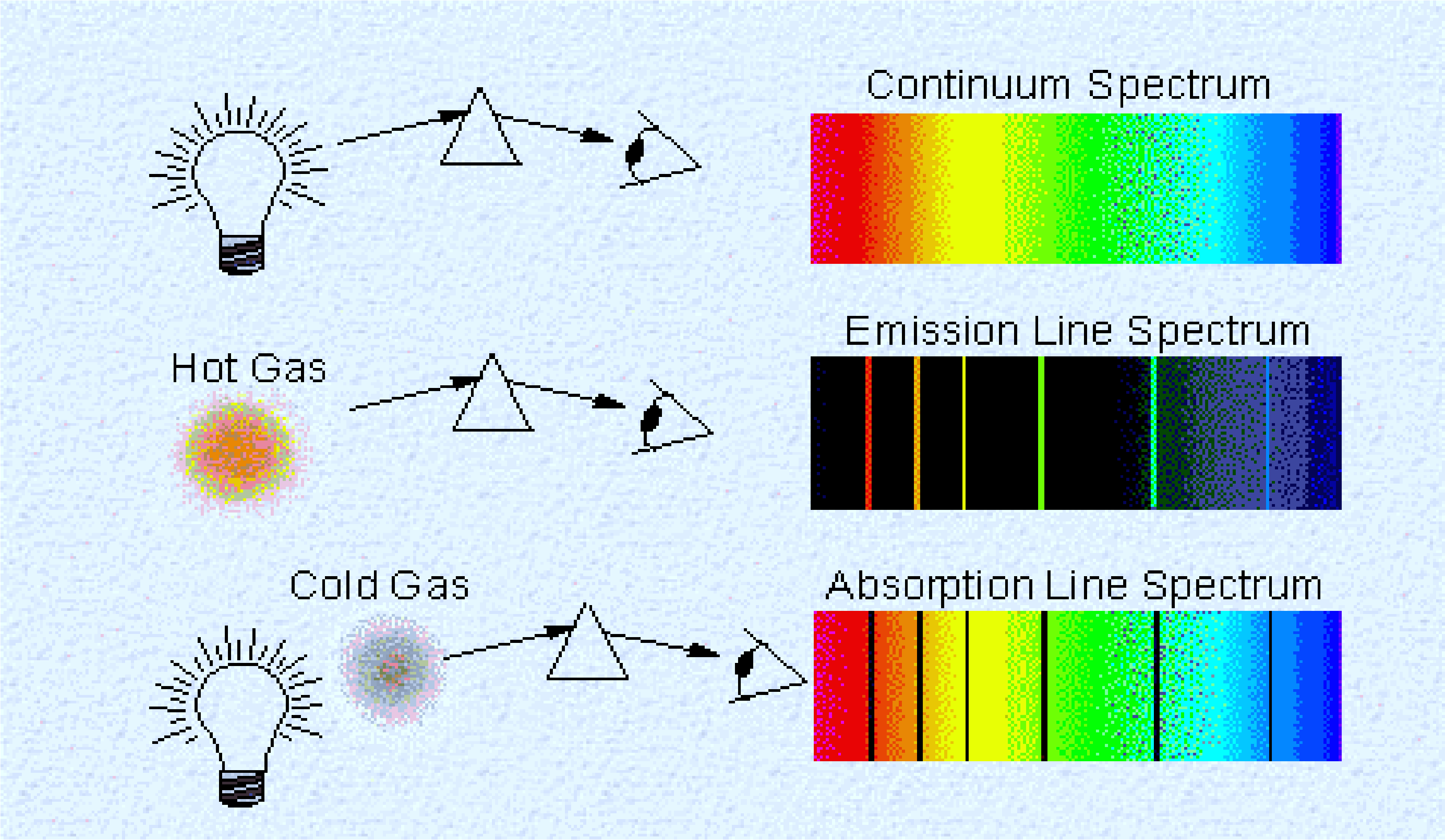


Absorption Spectra



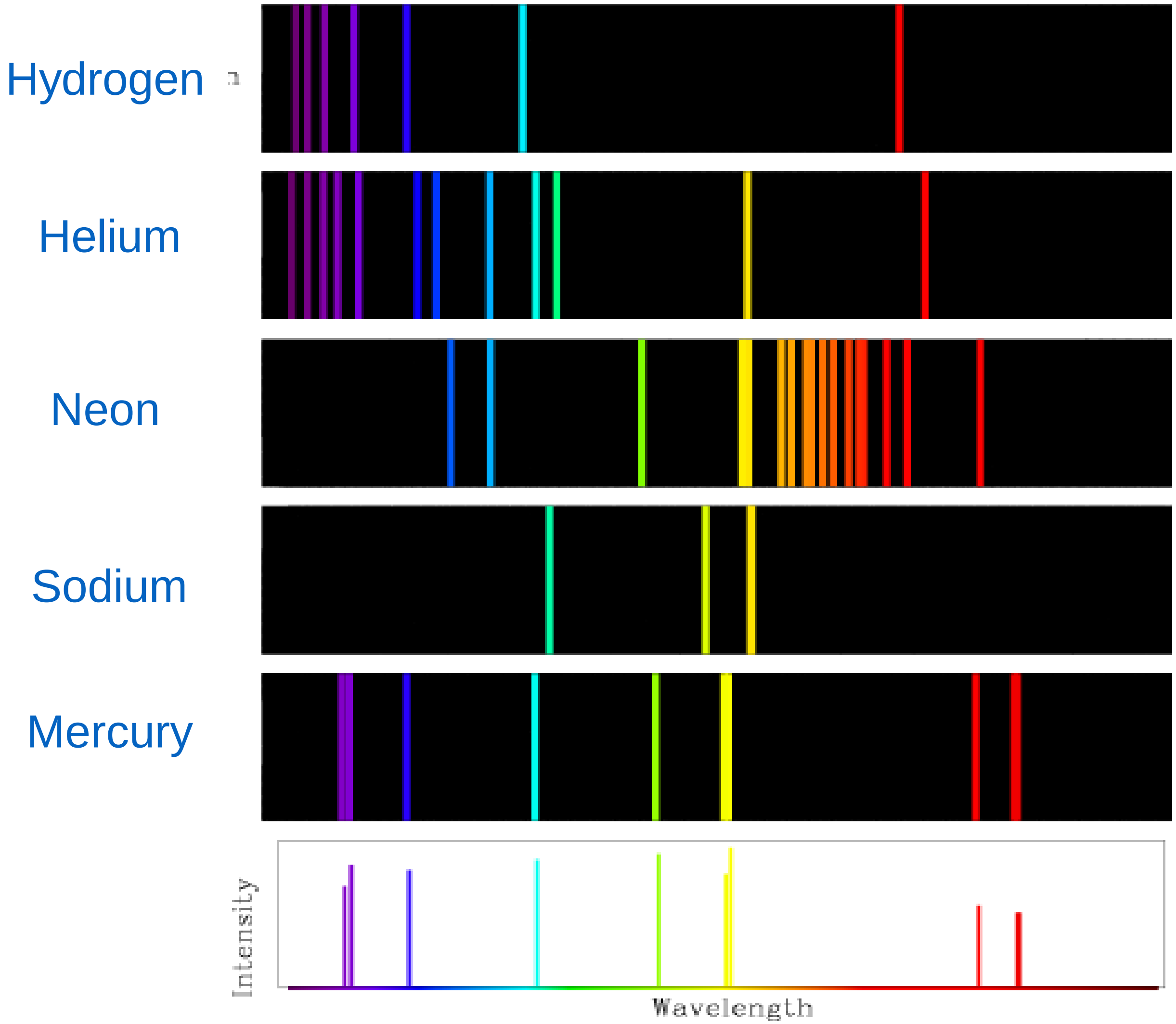
Some wavelengths of light are absorbed by the gas – an “absorption spectrum”.

Spectra



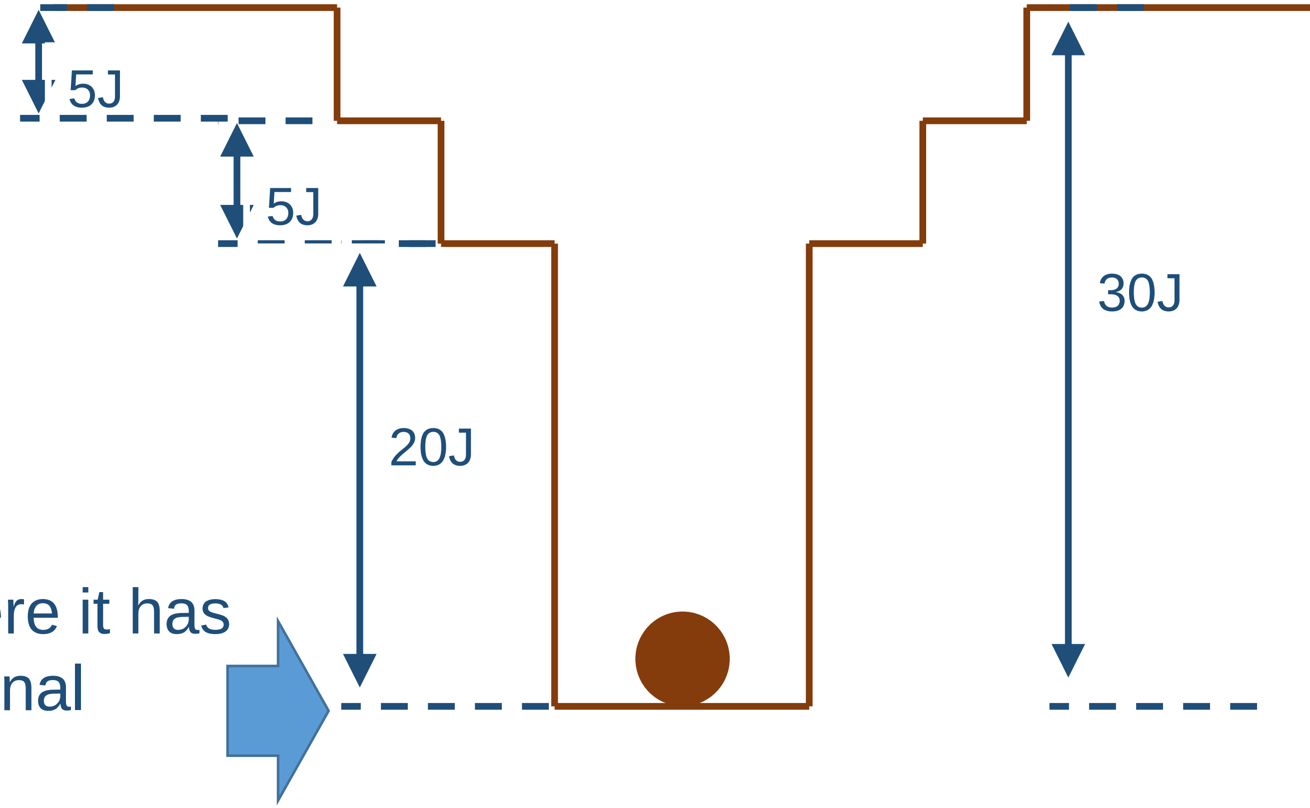
Continuum Spectrum
Hot Gas
Emission Line Spectrum
Cold Gas
Absorption Line Spectrum

Emission Spectra



Spectra

Consider a ball in a hole:

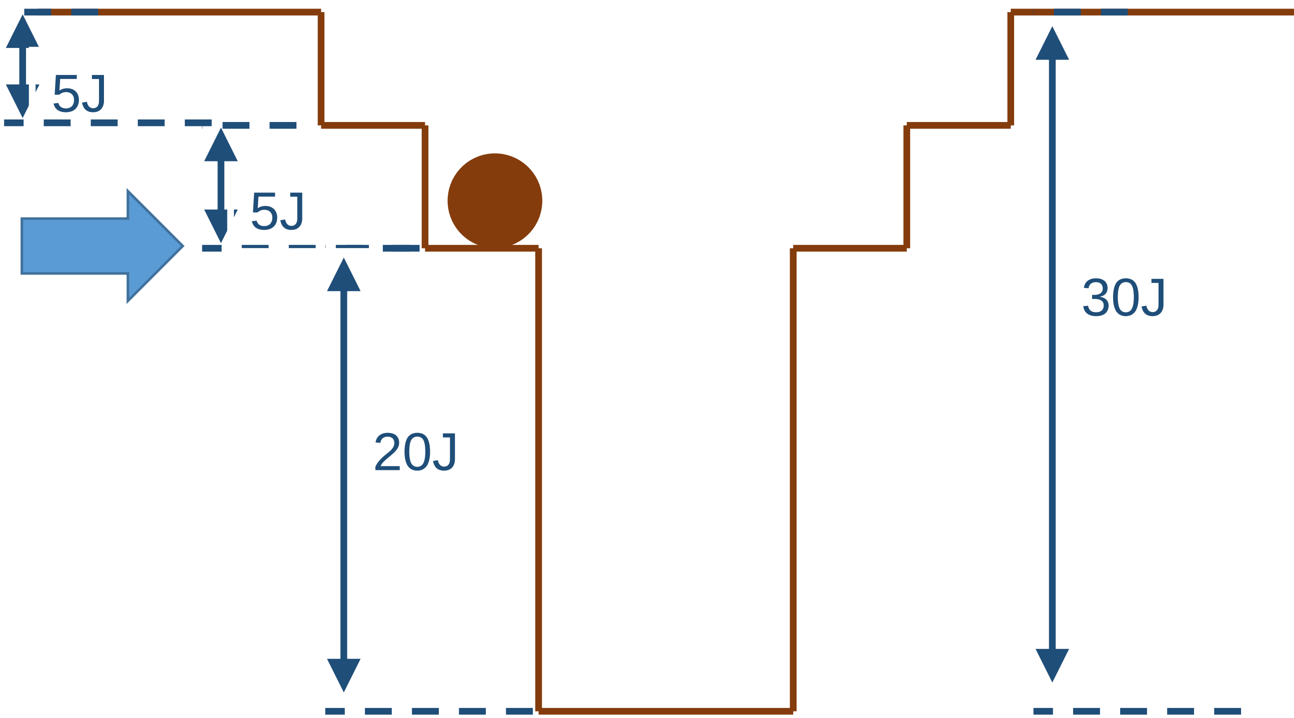


When the ball is here it has its lowest gravitational potential energy.

Spectra

Consider a ball in a hole:

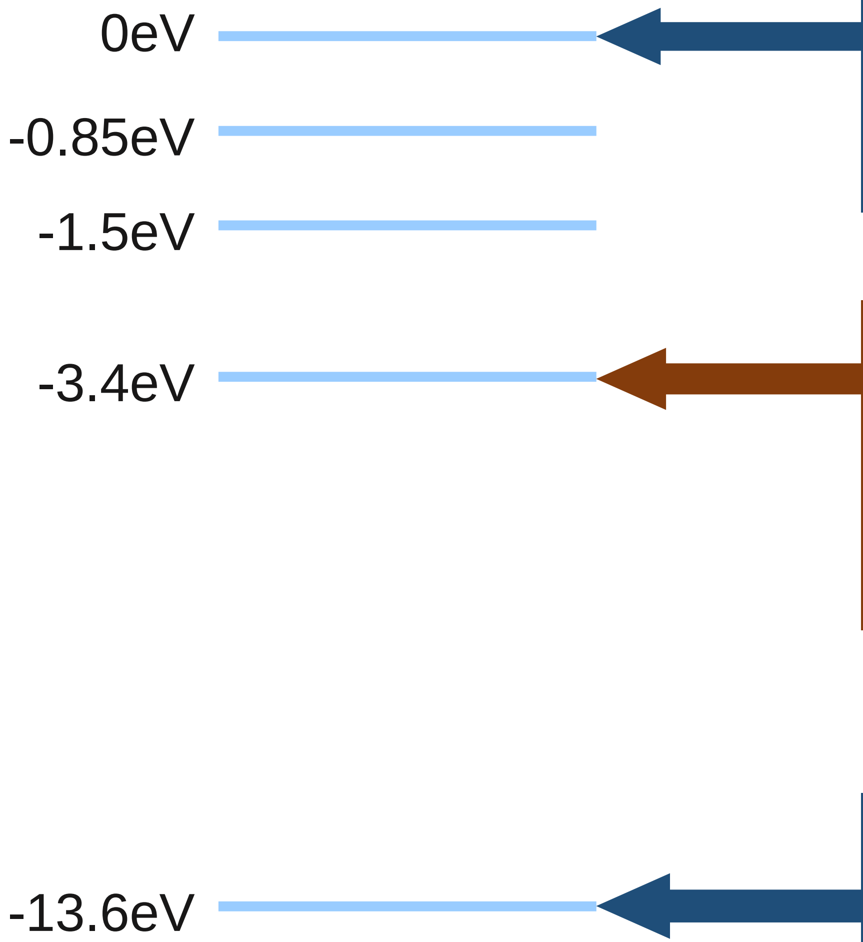
We can give it potential energy by lifting it up:



If it falls down again it will lose this gpe:

Spectra

A similar thing happens to electrons. We can "excite" them and raise their energy level:



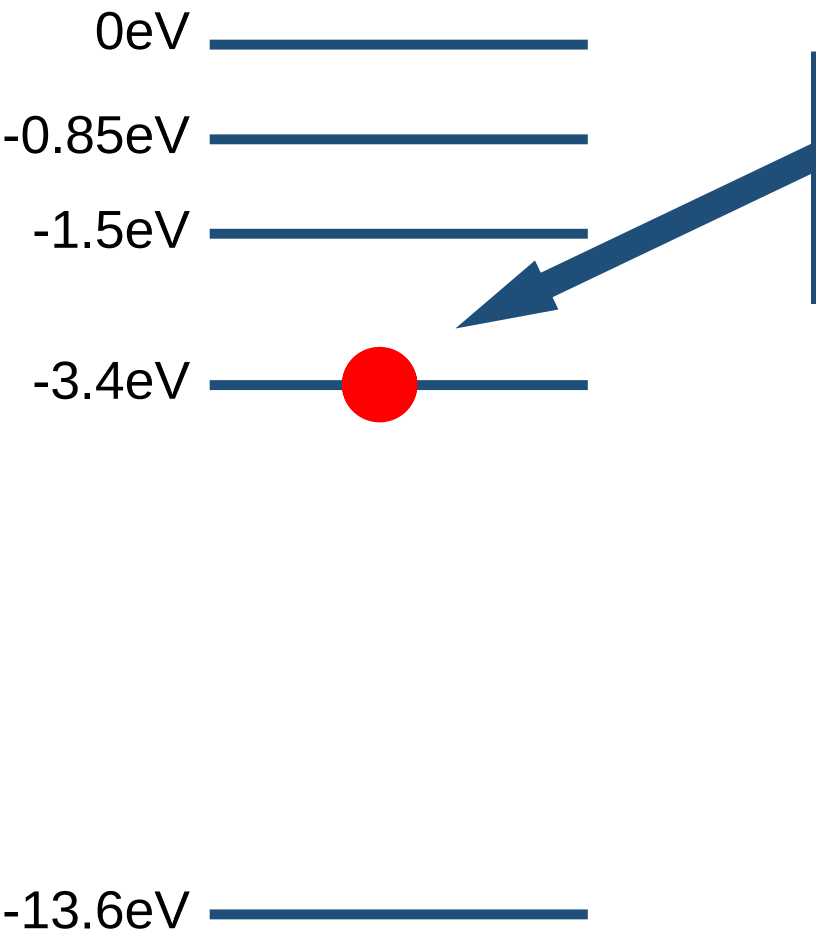
An electron at this energy level would be "free" – it's been "ionised".

These energy levels are negative because an electron here would have less energy than if its ionised.

This is called "The ground state"

Spectra

If we illuminate the atom we can excite the electron:



Q. What photon wavelength would be needed to excite this electron to ionise it?

Spectra

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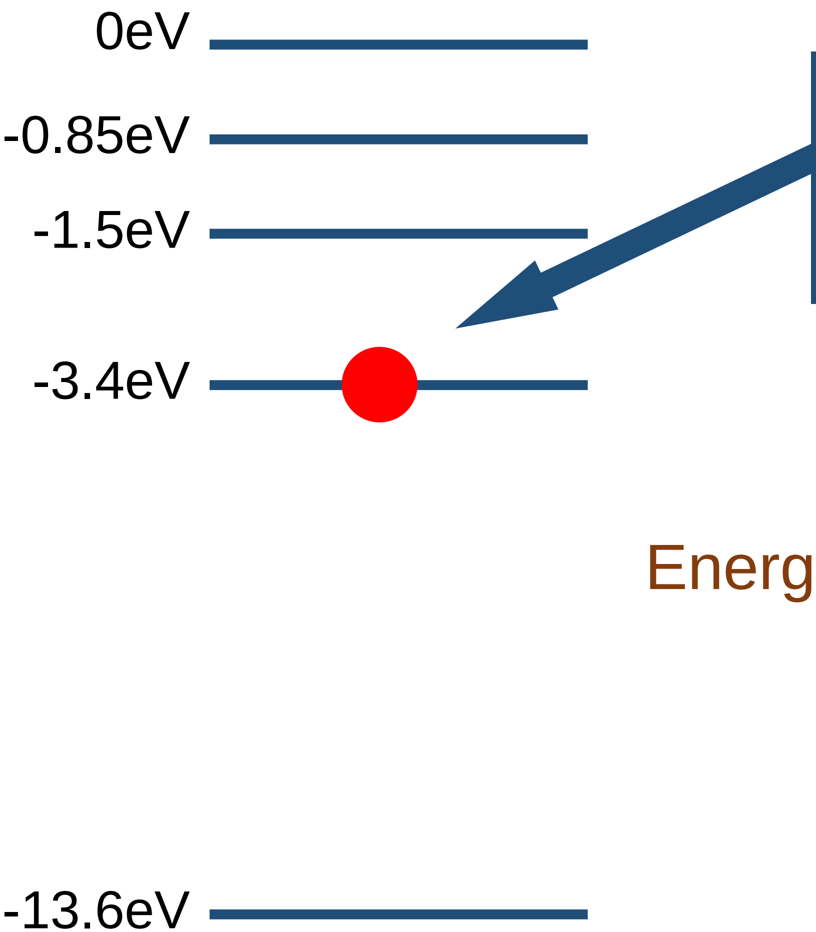

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Energy change = 3.4eV = 5.44×10^{-19} J.

Spectra

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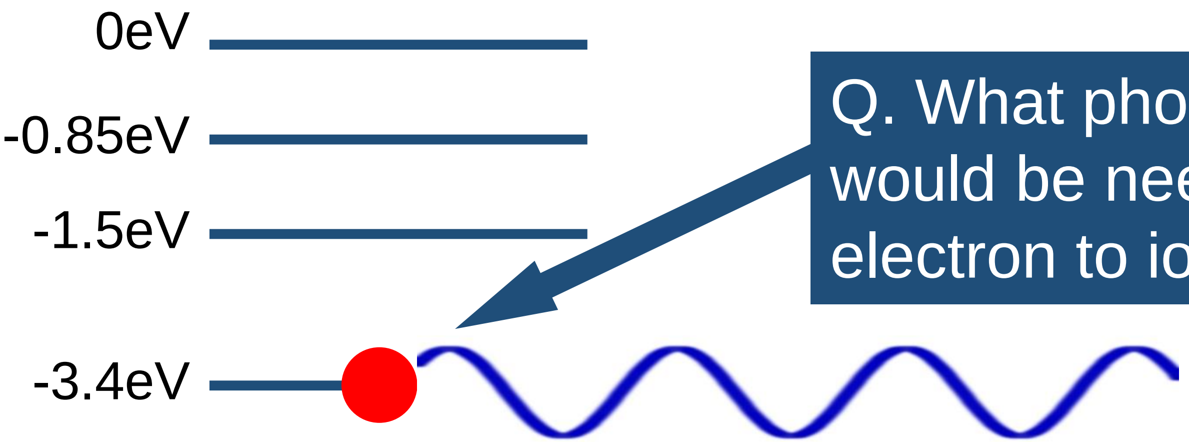

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Energy change = 3.4eV = 5.44×10^{-19}J.

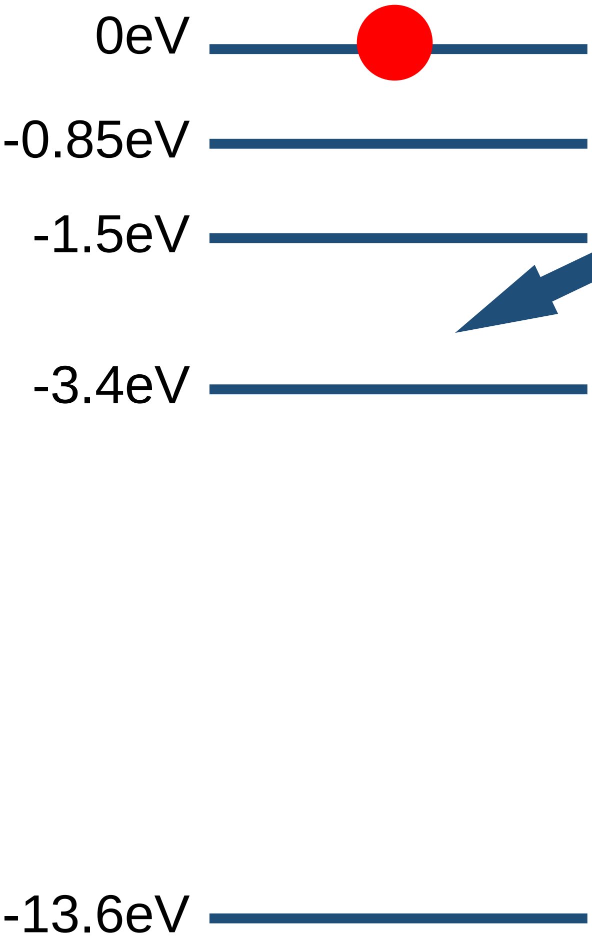
Using $E=hc/\lambda$ wavelength = 3.66×10^{-7}m

(In other words, ultra violet light)

-13.6eV

Spectra

If we illuminate the atom we can excite the electron:



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Using $E = hc/\lambda$ wavelength = 3.66×10^{-7}m

(In other words, ultra violet light)

Practice questions

1) State the ionisation energy of this atom in eV.

2) Calculate this ionisation energy in joules.

3) Calculate the wavelength of light needed to ionise the atom.

4) An electron falls from the -1.5eV to the -3.4eV level. What wavelength of light does it emit and what is the colour?

5) Light of frequency 1×10^{14}Hz is incident upon the atom. Will it be able to ionise the atom?