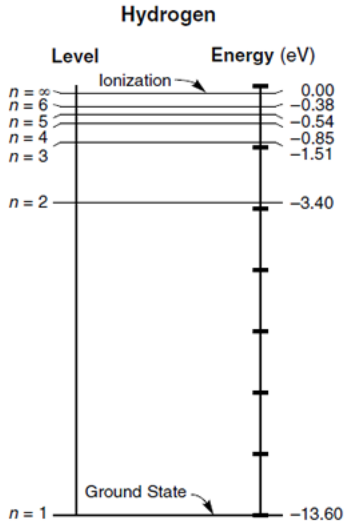
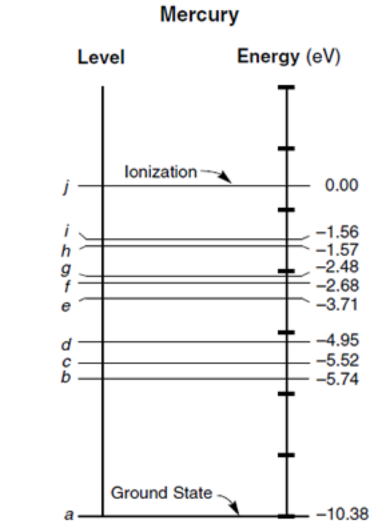
Physics with Synno – Light-Matter – Lesson 3

L&M.5 Emission Spectra

Previously, it was stated that electrons could move between energy levels in an atom. Each atom has its own series of energy levels.



Energy Levels for the Hydrogen Atom



A Few Energy Levels for the Mercury Atom

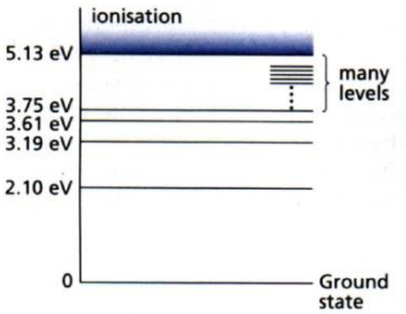
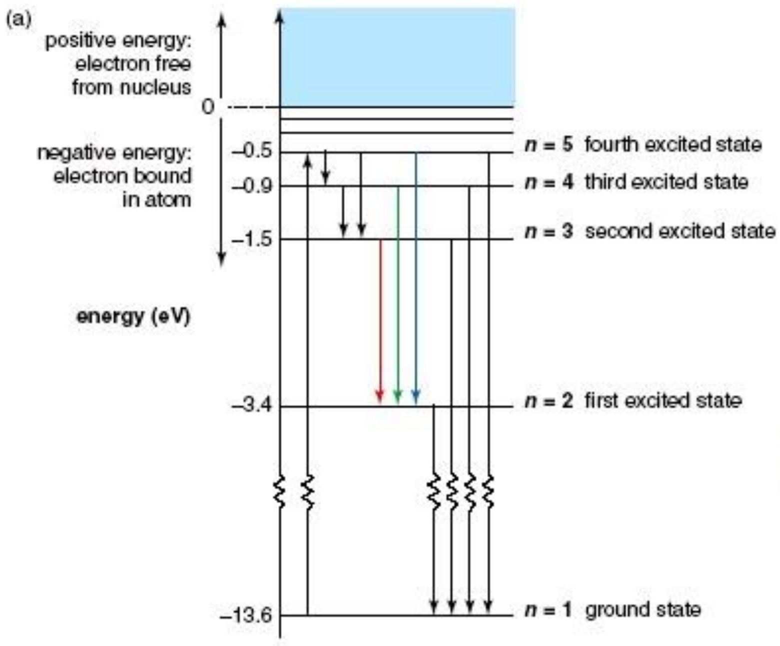


Figure 21.15
Energy levels for sodium.

If an electron moves from a higher energy level to a lower one the energy will be given out in the form of a photon. The energy of the photon given out is given by:

$$E_{photon} = E_{higher\ energy\ level} - E_{lower\ energy\ level} = h\ f$$



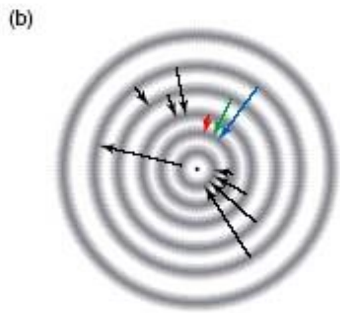
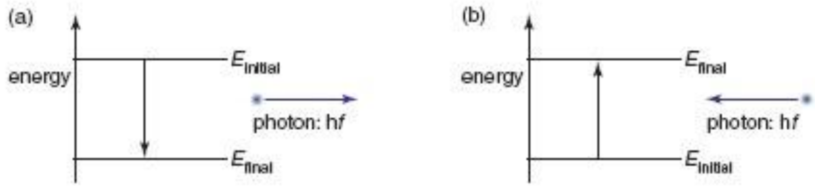


Figure 11.9 (a) Atomic energy level view of the spectral series of hydrogen, and (b) electron orbit view of the spectral series of hydrogen as illustrated in (a).

Figure 11.10
(a) Emission of light:
$E_{photon} = hf = E_{initial} - E_{final}$
(b) Absorption of light:
$E_{photon} = hf = E_{final} - E_{initial}$



Example 1: Some energy levels for sodium are shown above (fig 21.15).
What is the energy of the photon emitted when an electron in an atom of sodium falls from:
i) The 2.10 eV level to the ground state?

$$E_{photon} = E_{higher} - E_{lower}$$
$$E_{photon} = 5.13 - 0$$
$$E_{photon} = 5.13\ eV$$

ii) The 3.75 eV level to the 2.10 eV level?

$$E_{photon} = E_{higher} - E_{lower}$$
$$E_{photon} = 3.75 - 2.10$$
$$E_{photon} = 1.65\ eV$$

Example 2: What is the minimum energy required to ionise a hydrogen atom?

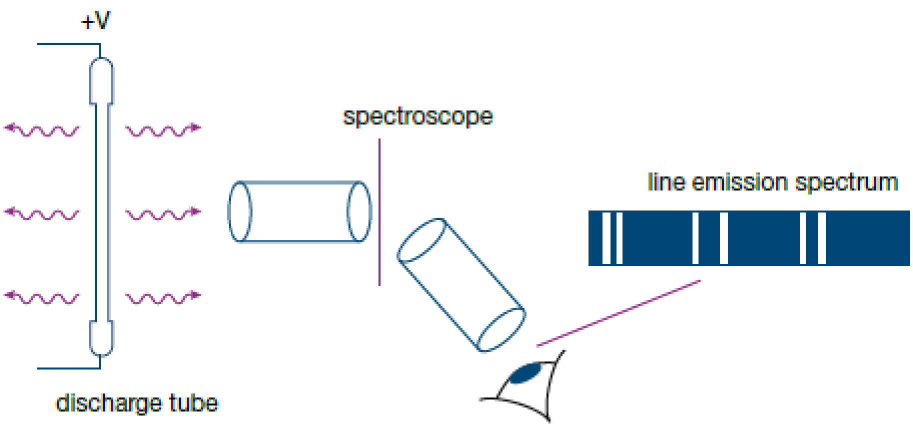
Using figure 11.9 a above
Ground state is -13.6 eV
Thus 13.6 eV of energy required to ionise Hydrogen

What is the frequency of the photon associated with this energy?

$$E_{photon} = h\,f$$
$$13.6 = 4.1 \times 10^{-15} \times f$$
$$f = 3.32 \times 10^{15}\ Hz$$

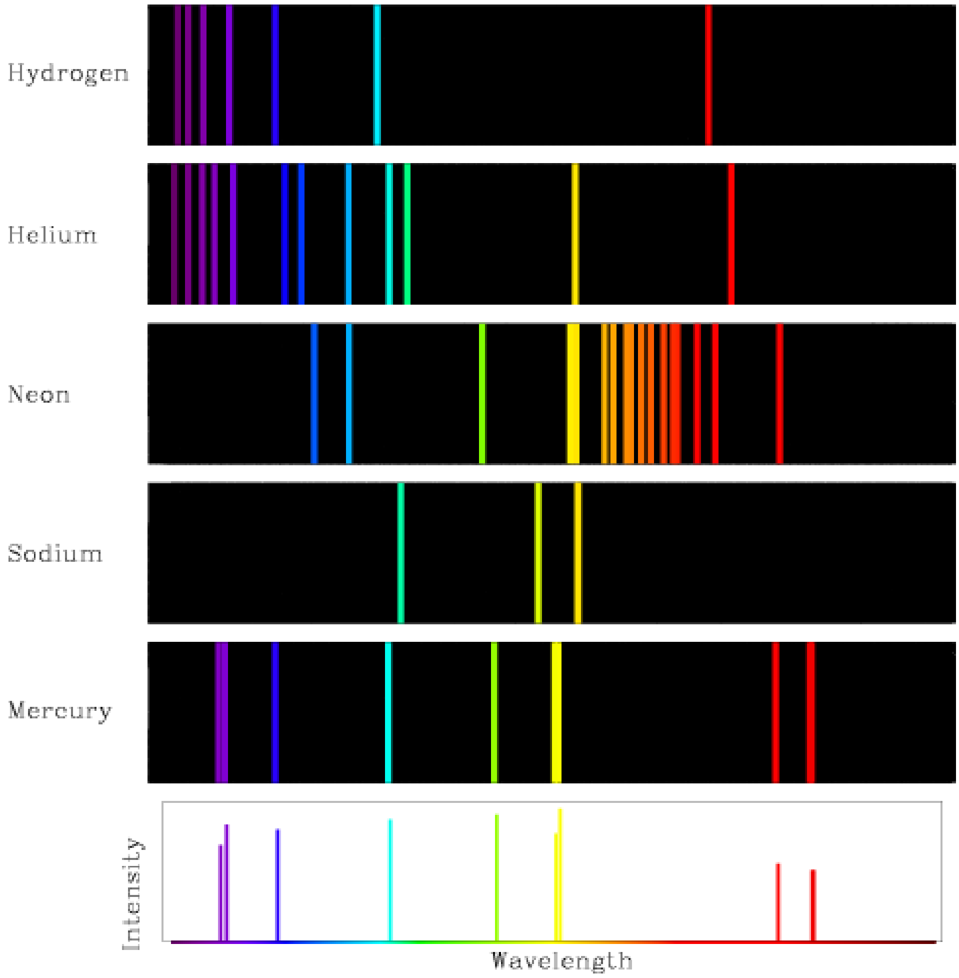
L&M.5.1 Emission Spectra

An emission spectra can be produced by applying a voltage to a gas. This will cause the gas to emit light. The frequencies of the light emitted depends on the energy levels of the electrons, thus each element has its own emission spectra.



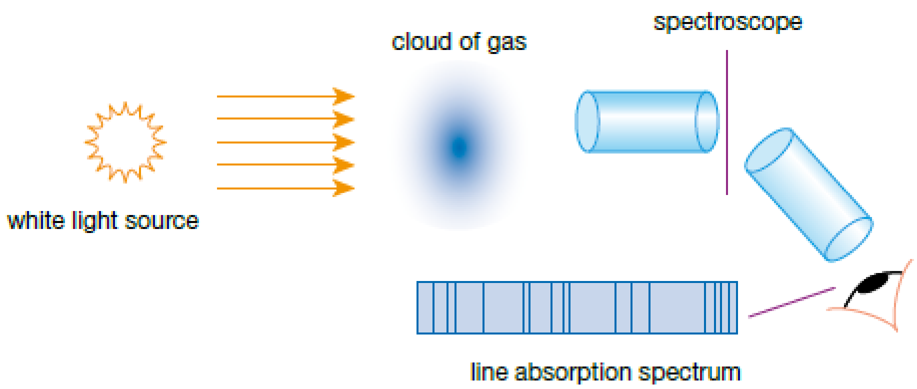
The frequency of the light emitted corresponds to the difference in energy levels in the atom.

Applying a voltage to the gas excites electrons into a higher energy level and when they drop back to the lower level the energy is released as a photon of light. Metal vapour lamps produce this type of spectrum.



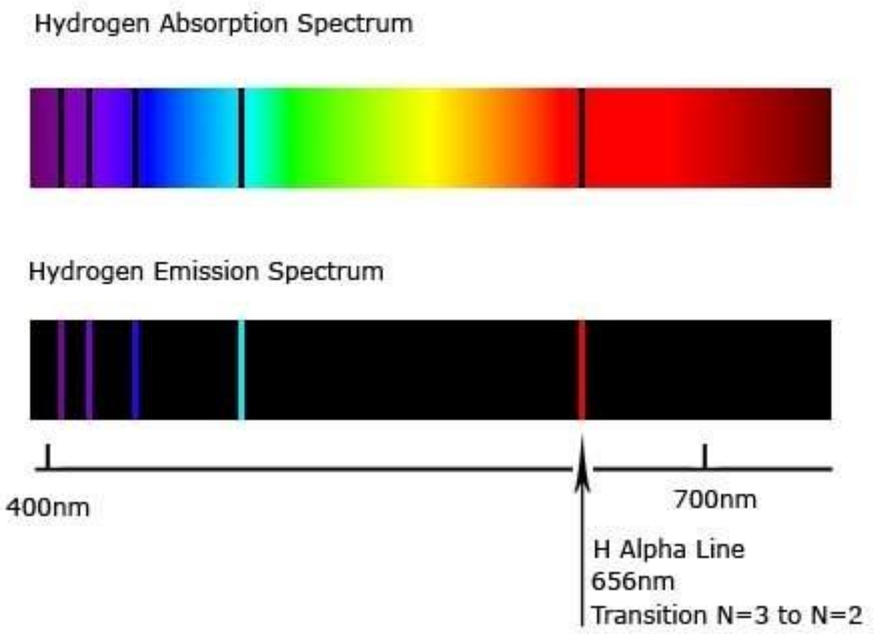
L&M.5.2 Absorption Spectra

When light is passed through a cool gas an absorption spectrum is created. Some of the frequencies of light are absorbed by the gas.



The absorption spectrum is unique for each type of gas. The missing frequencies correspond to the line in the emission spectrum of that gas.
Astronomers use absorption spectra to determine the chemical composition of stars.

Read: Text Page 350 & 351 - Bohr Model



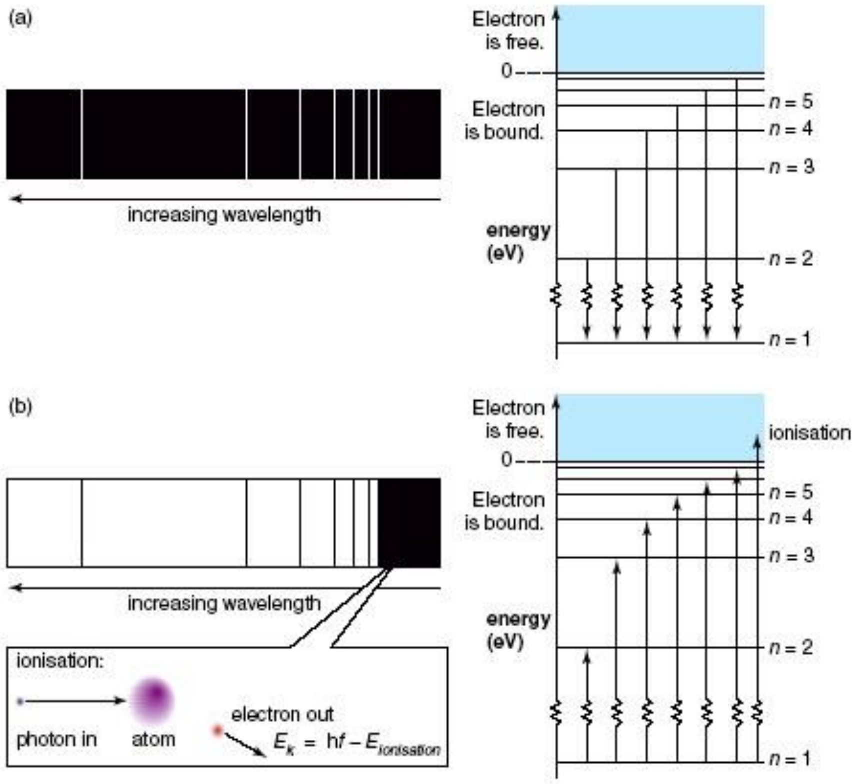
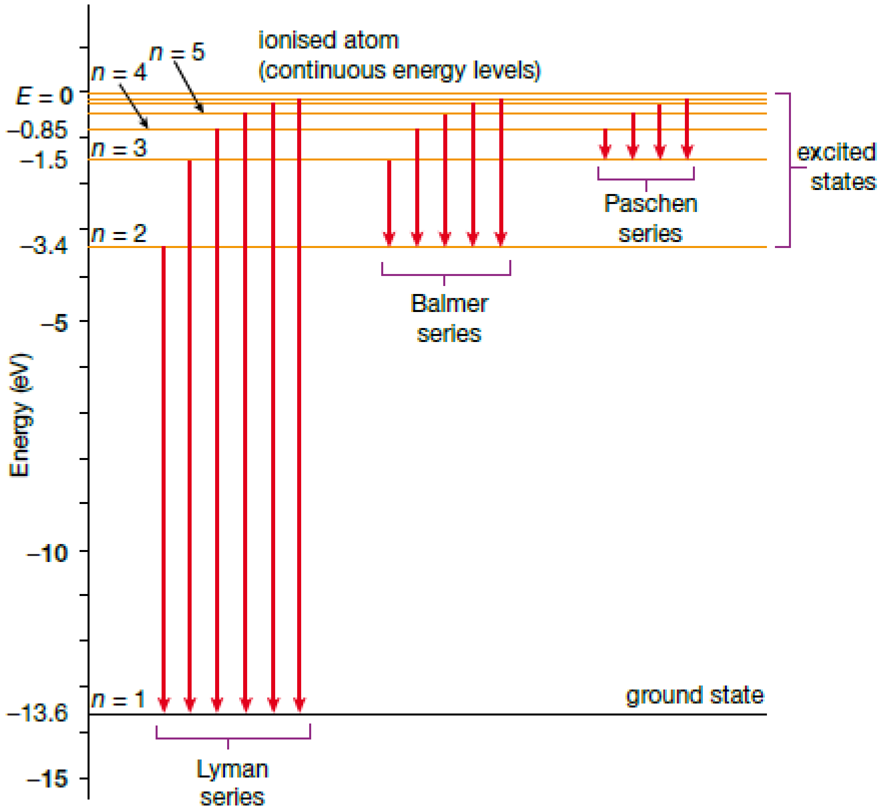


Figure 11.12
(a) Emission spectrum and (b) absorption spectrum of hydrogen atoms at room temperature. In the absorption spectrum, transfer of photon energy to eject an electron is labelled 'ionisation'. All wavelengths less than the limiting value (that is, photons with energy greater than the limiting value) can be absorbed.

L & M 5.3 Bohr and Energy Levels

The Bohr model for the atom having discrete energy levels worked well for hydrogen.



Where the energy of each level is $E_n = \frac{-13.6}{n^2}$

This model could not explain the energy levels of more complex atoms. It was, however, significant in suggesting discrete allowable energy levels.

In the development of his theory Bohr made links to the quantum nature and provided a conceptual framework for future developments. Particularly those of Heisenberg and Schrödinger.

Example

Calculate the wavelength (in nm) of the photon produced when an electron drops from the n = 4 energy level of the hydrogen atom to the n = 2 energy level.

Solution

E_4 = -0.85 eV E_2 = -3.4 eV
ΔE = -0.85-(-3.4)
= 2.55 eV

$E = \frac{h\,c}{\lambda} \rightarrow \lambda = \frac{h\,c}{E}$
$\lambda = 4.87 \times 10^{-7} m$
= 487 nm

L & M 6 Electrons, Waves and Energy Levels

Louis De Broglie pictured the electron in a hydrogen atom travelling along one of the allowed orbits around the nucleus, together with its associated wave. In de Broglie's mind the circumference of each allowed orbit contained a *whole number* of wavelengths of the electron-wave so that it formed a standing wave around the orbit.

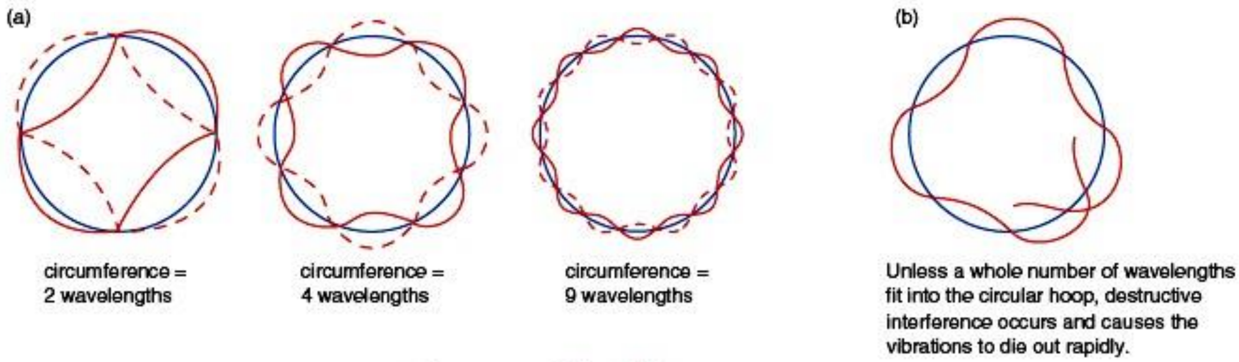


Figure 11.23 A model of the atom showing the electron as a standing wave

An electron-wave whose wavelength was slightly longer, or shorter, would not join onto itself smoothly. It would quickly collapse due to destructive interference. Only orbits corresponding to standing waves would survive.

i.e. The stable orbits of the Hydrogen atom are those where the circumference is exactly equal to a whole number of wavelengths.

L & M 7 Light Sources

Source	Method for producing Light	Light produced
Incandescent	Heating a metal filament to a very high temperature	A continuous spectrum. Some of the wavelengths produced are not visible to humans.
LEDs	A small amount of electricity is used to excite electrons to a higher energy level. When the electrons drop down they release a photon. Made of mainly silicon with the addition of another material, impurity.	Specific colour. Depend on the type and amount of impurity.
Lasers	Raising the atoms in a gas to an excited state by heating or an electric current.	Polarised. Monochromatic – one colour or limited wavelengths. Coherent – waves are in phase.
Synchrotron	Accelerating electrons to very high speeds.	Extremely bright. Highly polarised. Emitted in short pulses. Can produce a wide range of wavelengths.

Text Questions: Page 359 Exercise 10.3 Questions 1 → 10